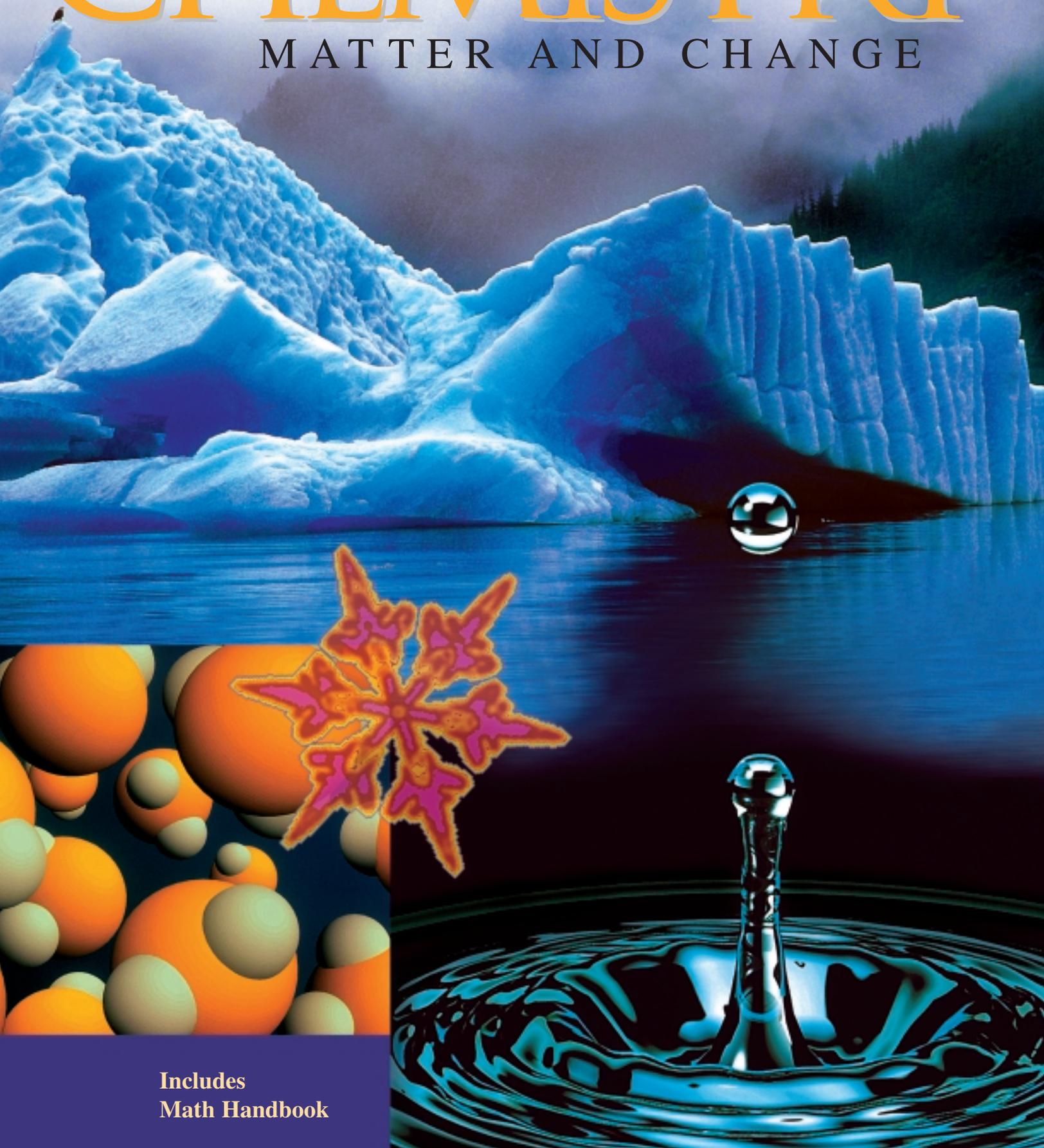


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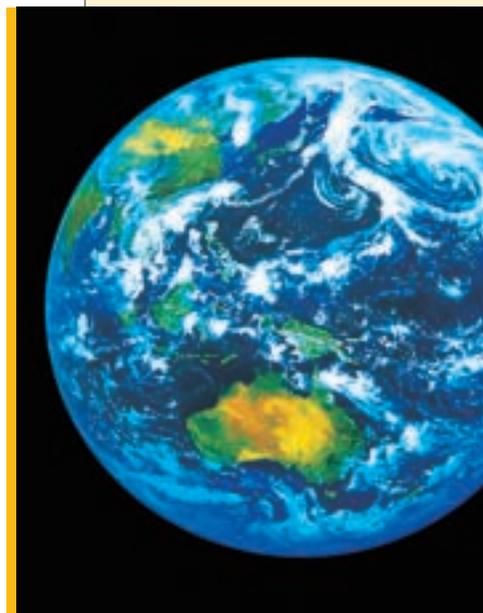
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Contents In Brief

Chapter 1	Introduction to Chemistry	2
Chapter 2	Data Analysis	24
Chapter 3	Matter—Properties and Changes	54
Chapter 4	The Structure of the Atom	86
Chapter 5	Electrons in Atoms.	116
Chapter 6	The Periodic Table and Periodic Law	150
Chapter 7	The Elements.	178
Chapter 8	Ionic Compounds.	210
Chapter 9	Covalent Bonding	240
Chapter 10	Chemical Reactions	276
Chapter 11	The Mole.	308
Chapter 12	Stoichiometry	352
Chapter 13	States of Matter	384
Chapter 14	Gases	418
Chapter 15	Solutions.	452
Chapter 16	Energy and Chemical Change.	488
Chapter 17	Reaction Rates	528
Chapter 18	Chemical Equilibrium.	558
Chapter 19	Acids and Bases	594
Chapter 20	Redox Reactions	634
Chapter 21	Electrochemistry	662
Chapter 22	Hydrocarbons	696
Chapter 23	Substituted Hydrocarbons and Their Reactions	736
Chapter 24	The Chemistry of Life	774
Chapter 25	Nuclear Chemistry	804
Chapter 26	Chemistry in the Environment	840



Introduction to Chemistry

What You'll Learn

- ▶ You will describe the relationship between chemistry and matter.
- ▶ You will recognize how scientific methods can be used to solve problems.
- ▶ You will distinguish between scientific research and technology.

Why It's Important

- ▶ You, and all the objects around you, are composed of matter. By studying matter and the way it changes, you will gain an understanding of your body and all the "stuff" you see and interact with in your everyday life.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about chemistry and matter.

The four nebulae shown here contain a stew of elements. The red color in two of the nebulae is emitted by hydrogen atoms. The Horsehead Nebula can be seen on the right. The fourth nebula is the bluish structure below the horse's head. The round, bright object on the left is the star, Zeta Orionis.



DISCOVERY LAB



Where is it?

When an object burns, the quantity of ashes that remain is smaller than the original object that was burned. What happened to the rest of the object?

Safety Precautions



Do not place matches in the sink. Use caution around flames.

Procedure

1. Measure the mass of a large kitchen match. Record this measurement and detailed observations about the match.
2. Carefully strike the match and allow it to burn for five seconds. Then, blow it out. **CAUTION:** *Keep hair and loose clothing away from the flame.* Record observations about the match as it burns and after the flame is extinguished.
3. Allow the match to cool. Measure and record the mass of the burned match.
4. Place the burned match in a container designated by your instructor.
5. Repeat this procedure. Compare your data from the two trials.

Analysis

How do you account for the change in mass? Where is the matter that appears to have been lost?

Materials

large kitchen matches
laboratory balance
lab notebook
pen
stopwatch or clock

Section

1.1

The Stories of Two Chemicals

Objectives

- **Explain** the formation and importance of ozone.
- **Describe** the development of chlorofluorocarbons.

Take a moment to look around you. Where did all the “stuff” you see come from? All the stuff in the universe is made from building blocks formed in stars such as the ones shown in the photo on the opposite page. And, as you learned in the **DISCOVERY LAB**, this stuff changes form.

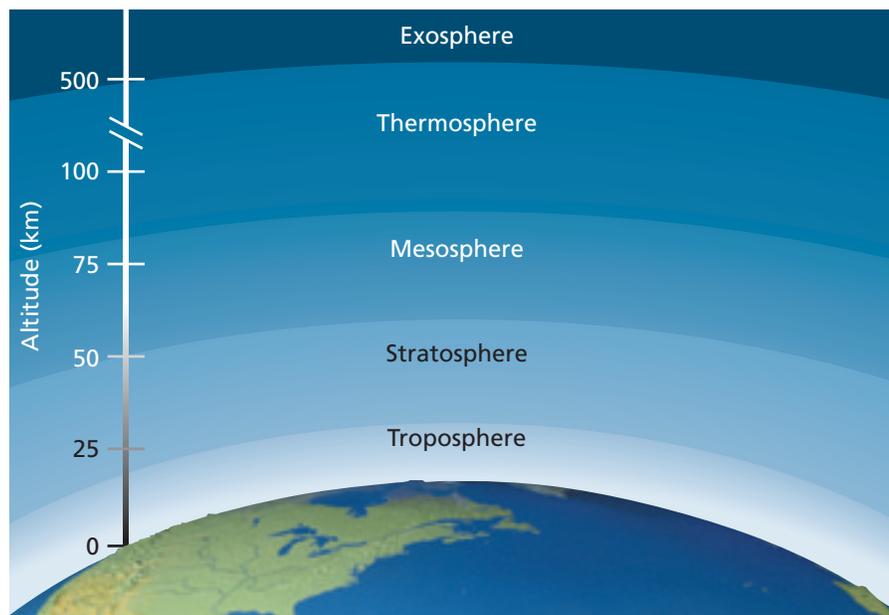
Scientists are naturally curious. They continually ask questions about and seek answers to all that they observe in the universe. One of the areas in which scientists work is the branch of science called chemistry. Your introduction to chemistry will begin with two unrelated discoveries that now form the basis of one of the most important environmental issues of our time.

The Ozone Layer

You are probably aware of some of the damaging effects of ultraviolet radiation from the Sun if you have ever suffered from a sunburn. Overexposure to ultraviolet radiation also is harmful to plants and animals, lowering crop yields and disrupting food chains. Living things can exist on Earth because ozone, a chemical in Earth’s atmosphere, absorbs most of this radiation before it reaches Earth’s surface. A chemical is any substance that has a definite composition. Ozone is a substance that consists of three particles of oxygen.

Figure 1-1

Earth's atmosphere consists of several layers. The layer nearest Earth is the troposphere. The stratosphere is above the troposphere.



Earth's atmosphere As you can see in **Figure 1-1**, Earth's atmosphere consists of layers. The lowest layer is called the troposphere and contains the air we breathe. The troposphere is where the clouds shown in **Figure 1-2** occur and where airplanes fly. All of Earth's weather occurs in the troposphere.

The stratosphere is the layer above the troposphere. It extends from about 15 to 50 kilometers (km) above Earth's surface. The ozone that protects Earth is located in the stratosphere. About 90% of Earth's ozone is spread out in a layer that surrounds and protects our planet.

Ozone formation How does ozone enter the stratosphere? Ozone is formed when oxygen gas is exposed to ultraviolet radiation in the upper regions of the stratosphere. Particles of oxygen gas are made of two smaller oxygen particles. The energy of the radiation breaks the gas particles into oxygen particles, which then interact with oxygen gas to form ozone. **Figure 1-3** illustrates this process. Ozone also can absorb radiation and break apart to reform oxygen gas. Thus, there tends to be a balance between oxygen gas and ozone levels in the stratosphere.

Ozone was first identified and measured in the late 1800s, so its presence has been studied for a long time. It was of interest to scientists because air currents in the stratosphere move ozone around Earth. Ozone forms over the equator where the rays of sunlight are the strongest and then flows toward the poles. Thus, ozone makes a convenient marker to follow the flow of air in the stratosphere.

In the 1920s, G.M.B. Dobson began measuring the amount of ozone in the atmosphere. Although ozone is formed in the higher regions of the stratosphere, most of it is stored in the lower stratosphere, where it can be measured by instruments on the ground or in balloons, satellites, and rockets. Dobson measured levels of stratospheric ozone of more than 300 Dobson units (DU). His measurements serve as a basis for comparison with recent measurements.

During 1981–1983, a research group from the British Antarctic Survey was monitoring the atmosphere above Antarctica. They measured surprisingly low levels of ozone, readings as low as 160 DU, especially during the Antarctic spring in October. They checked their instruments and repeated their measurements. In October 1985, they reported a confirmed decrease in the amount of ozone in the stratosphere and concluded that the ozone layer was thinning.



Figure 1-2

The troposphere extends to a height of about 15 km. Cumulonimbus clouds, or thunderheads, produce thunder, lightning, and rain.

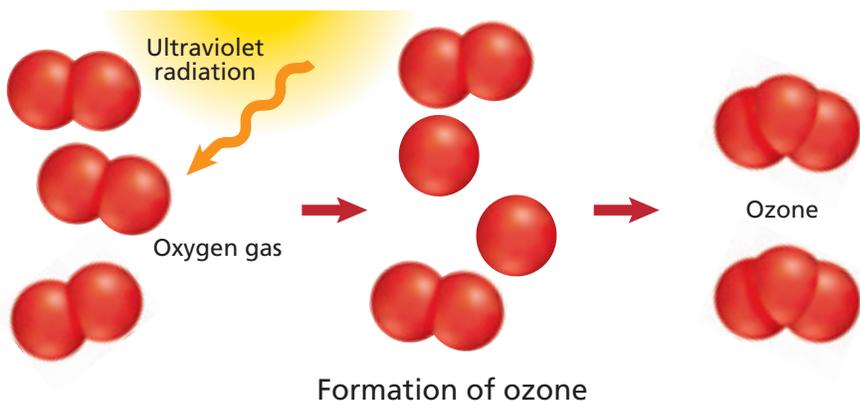


Figure 1-3

This model of the formation of ozone shows that ultraviolet radiation from the Sun causes oxygen gas to break down into two individual particles of oxygen. These individual oxygen particles combine with oxygen gas to form ozone, which consists of three oxygen particles.

Although the thinning of the ozone layer is often called the ozone hole, it is not actually a hole. You can think of it as being similar to the old sock in **Figure 1-4a** in which the material of the heel is wearing thin. You might be able to see your skin through the thinning sock. So although the ozone is still present in the atmosphere, the protective layer is much thinner than normal. This fact has alarmed scientists who never expected to find such low levels. Measurements made from balloons, high-altitude planes, and satellites have supported the measurements made from the ground, as the satellite map in **Figure 1-4b** shows. What could be causing the ozone hole?

Chlorofluorocarbons

The story of the second chemical in this chapter begins in the 1920s. Refrigerators, which used toxic gases such as ammonia as coolants, were just beginning to be produced large scale. Because ammonia fumes could escape from the refrigerator and harm the members of a household, chemists began to search for safer coolants. Thomas Midgley, Jr. synthesized the first chlorofluorocarbons in 1928. A chlorofluorocarbon (CFC) is a chemical that consists of chlorine, fluorine, and carbon. There are several different chemicals that are classified as CFCs. They are all made in the laboratory and do not occur naturally. CFCs are nontoxic and stable. They do not readily react with other chemicals. At the time, they seemed to be ideal coolants for refrigerators. By 1935, the first self-contained home air-conditioning units and eight million new refrigerators in the United States used CFCs as coolants. In addition to their use as refrigerants, CFCs also were used in plastic foams and as propellants in spray cans.

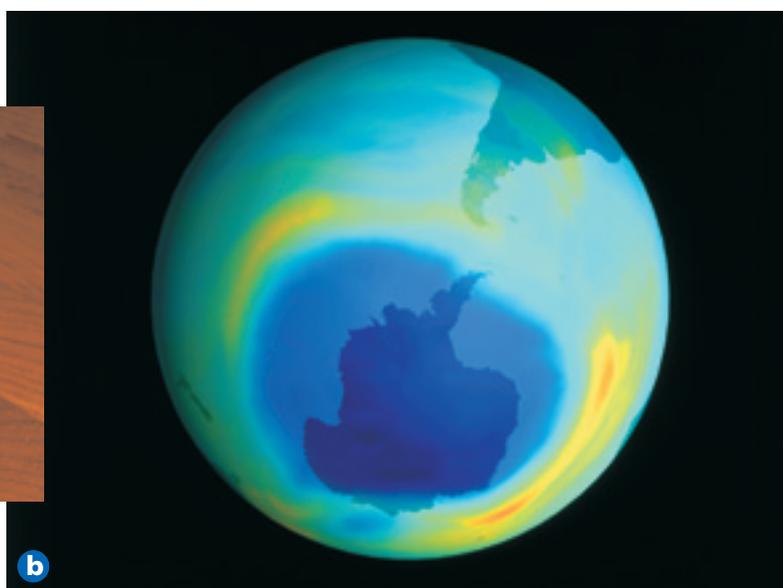


Figure 1-4

- a** The thinning heel of this sock models the thinning of the ozone layer in the stratosphere.
- b** This colored satellite map of stratospheric ozone over Antarctica was taken on September 15, 1999. The lowest amount of ozone (light purple) appears over Antarctica (dark purple). Blue, green, orange, and yellow show increasing amounts of ozone.

Concentration of CFCs in the Atmosphere

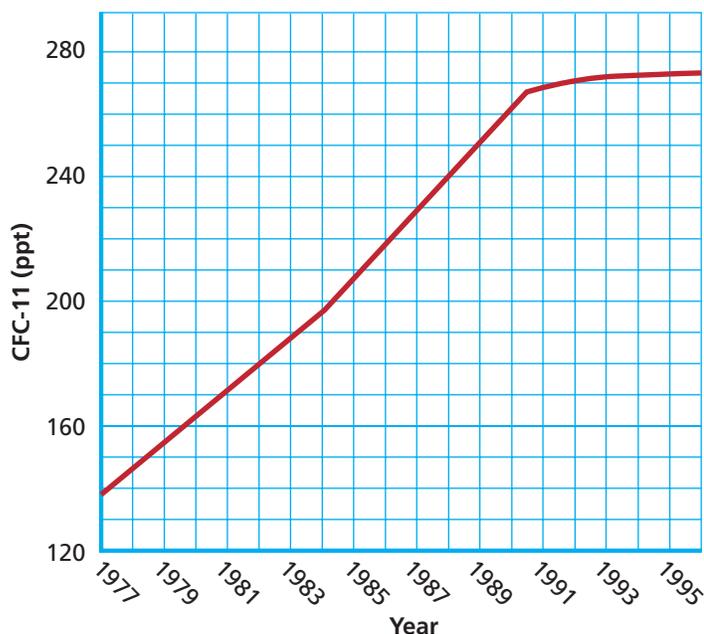


Figure 1-5

Quantities of CFCs in the atmosphere continued to rise until a ban on products containing them went into effect in many countries.

and measured: the protective ozone layer in the atmosphere was thinning, and increasingly large quantities of useful CFCs were drifting into the atmosphere. Could there be a connection between the two occurrences? Before you learn the answer to this question, you need to understand some of the basic ideas of chemistry and know how chemists—and most scientists, for that matter—solve problems.

Now think of all the refrigerators in your neighborhood, in your city, across the country, and around the world. Think of the air conditioners in homes, schools, office buildings, and cars that also used CFCs. Add to your mental list all of the aerosol cans and plastic foam cups and food containers used each day throughout the world. If all of these products contained or were made with CFCs, imagine the quantities of these chemicals that could be released into the environment in a single day.

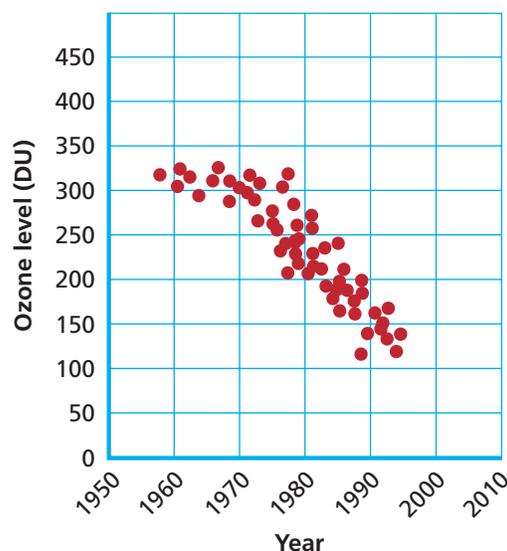
Scientists first began to notice the presence of CFCs in the atmosphere in the 1970s. They decided to measure the amount of CFCs in the stratosphere and found that quantities in the stratosphere increased year after year. This increase is shown in **Figure 1-5**. But, it was thought that CFCs did not pose a threat to the environment because they are so stable.

Two separate occurrences had been noticed

Section 1.1 Assessment

1. Why is ozone important in the atmosphere?
2. Where is ozone formed and stored?
3. What are CFCs? How are they used?
4. **Thinking Critically** Why do you think ozone is formed over the equator? What is the connection between sunlight and ozone formation?
5. **Comparing and Contrasting** What general trend in ozone concentration is shown in the graph at the right? How does the data for the years 1977–1987 on this graph compare to the same time span on the graph in **Figure 1-5**? What do you notice?

October Ozone Concentrations



Matter, the stuff of the universe, has many different forms. You are made of matter. There is matter in the bed, blankets, and sheets on which you sleep as well as in the clothes you wear. There is matter in the food you eat, and in medications and vitamins you may take. You have learned that ozone is a chemical that occurs naturally in the environment, whereas CFCs do not. Although both chemicals are invisible gases, they, too, are matter.

Chemistry: The Central Science

Chemistry is the study of matter and the changes that it undergoes. A basic understanding of chemistry is central to all sciences—biology, physics, Earth science, ecology, and others. Chemistry also is central to our everyday lives, as **Figure 1-6** illustrates. It will continue to be central to discoveries made in science and technology in the twenty-first century.

Objectives

- **Define** chemistry and matter.
- **Compare** and **contrast** mass and weight.
- **Explain** why chemists are interested in a submicroscopic description of matter.

Vocabulary

chemistry
matter
mass
weight



Figure 1-6

High-tech fabrics that don't hinder athletic performance, water, fertilizers, pesticides, food, grocery items, clothing, building materials, hair care products, plastics, and even the human body are made of chemicals.

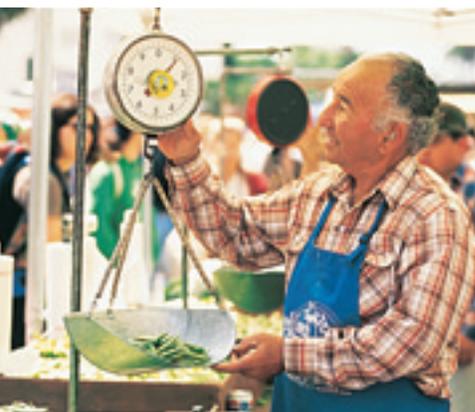


Figure 1-7

A scale measures the downward pull of gravity on an object. If this scale were used on the Moon, the reading would be less than on Earth.

Matter and its Characteristics

You recognize matter in the everyday objects you are familiar with, such as those shown in **Figure 1-6**. But, how do you define matter? **Matter** is anything that has mass and takes up space. **Mass** is a measurement that reflects the amount of matter. It is easy to see that your textbook has mass and takes up space. Is air matter? You can't see it and you can't always feel it. However, when you inflate a balloon, it expands to make room for the air. The balloon gets heavier. Thus, air must be matter. Is everything made of matter? The thoughts and ideas that "fill" your head are not matter; neither are heat, light, radio waves, nor magnetic fields. What else can you name that is not matter?

Mass and weight When you go to the supermarket to buy a pound of vegetables, you place them on a scale like the one shown in **Figure 1-7** to find their weight. **Weight** is a measure not only of the amount of matter but also of the effect of Earth's gravitational pull on that matter. This force is not exactly the same everywhere on Earth and actually becomes less as you move away from Earth's surface at sea level. You may not notice a difference in the weight of a pound of vegetables from one place to another, but subtle differences do exist.

It might seem more convenient for scientists to simply use weight instead of mass. You might wonder why it is so important to think of matter in terms of mass. Scientists need to be able to compare the measurements that they make in different parts of the world. They could identify the gravitational force every time they weigh something but that is not practical or convenient. This is why they use mass as a way to measure matter independent of gravitational force.

What you see and what you don't What can you observe about the outside of your school building or a skyscraper downtown? You know that there is more than meets the eye to such a building. There are beams inside the walls that give the building structure, stability, and function. Consider another

problem-solving LAB

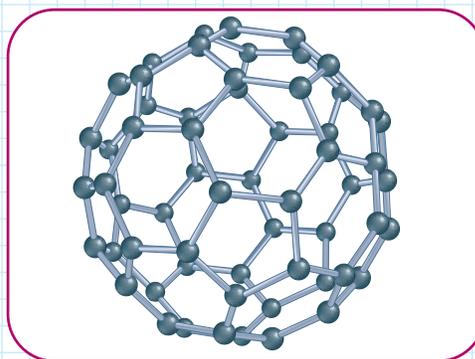
Chemical Models

Making Models Until the mid-1980s, scientists thought there were only two forms of carbon, each with a unique structure: diamond and graphite. As with many scientific discoveries, another form of carbon came as a surprise.

Buckminsterfullerene, also called the buckyball, was found while researching interstellar matter. Scientists worked with various models of carbon structures until they determined that 60 carbons were most stable when joined together in a shape that resembles a soccer ball.

Analysis

Examine the structure in the diagram. Where are the carbon atoms? How many carbons is each carbon connected to? Identify the pentagons and hexagons on the faces on the buckyball.



Thinking Critically

Go back to **Figure 1-3** to see an example of another model that chemists use. What process does the figure show? Explain why this information could not be shown in a photograph. What do the colored particles represent? Use **Table C-1** in Appendix C to help you in your identification.

Table 1-1

Branches of Chemistry		
Branch	Area of emphasis	Examples
Organic chemistry	Most carbon-containing chemicals	Pharmaceuticals, plastics
Inorganic chemistry	In general, matter that does not contain carbon	Minerals, metals and nonmetals, semi-conductors
Physical chemistry	The behavior and changes of matter and the related energy changes	Reaction rates, reaction mechanisms
Analytical chemistry	Components and composition of substances	Food nutrients, quality control
Biochemistry	Matter and processes of living organisms	Metabolism, fermentation

example. When you bend your arm at the elbow, you observe that your hand comes toward your shoulder. Muscles that you cannot see under the skin contract and relax to move your arm.

Much of matter and its behavior is macroscopic; that is, you do not need a microscope to see it. You will learn in Chapter 3 that the tremendous variety of stuff around you can be broken down into more than 100 types of matter called elements, and that elements are made up of particles called atoms. Atoms are so tiny that they cannot be seen even with optical microscopes. Thus, atoms are *submicroscopic*. They are so small that 100 million million atoms could fit onto the period at the end of this sentence.

The structure, composition, and behavior of all matter can be explained on a submicroscopic level. All that we observe about matter depends on atoms and the changes they undergo. Chemistry seeks to explain the submicroscopic events that lead to macroscopic observations. One way this can be done is by making a model, a visual representation of a submicroscopic event. **Figure 1-3** is such a model. The **problem-solving Lab** on the opposite page gives you practice interpreting a simple chemical model.

Branches in the field of chemistry Because there are so many types of matter, there are many areas of study in the field of chemistry. Chemistry is traditionally broken down into the five branches listed in **Table 1-1**.

Additional areas of chemistry include theoretical chemistry, which focuses on why and how chemicals interact, and environmental chemistry, which deals with the role chemicals play in the environment.

Careers Using Chemistry

Chemistry Teacher

Do you love chemistry? Would you like to help others love it—or at least understand it? If so, you might make an excellent chemistry teacher.

Chemistry teachers work in high schools and two-year and four-year colleges. They lecture, guide discussions, conduct experiments, supervise lab work, and lead field trips. High school teachers might also be asked to monitor study halls and serve on committees. College instructors might be required to do research and publish their findings.

Section 1.2 Assessment

- Define matter.
- Compare and contrast mass and weight.
- Why does chemistry involve the study of the changes in the world at a submicroscopic level?
- Thinking Critically** Explain why a scientist must be cautious when a new chemical that has many potential uses is synthesized.
- Using Numbers** If your weight is 120 pounds and your mass is 54 kilograms, how would those values change if you were on the moon? The gravitational force on the moon is 1/6 the gravitational force on Earth.

Objectives

- **Identify** the common steps of scientific methods.
- **Compare** and **contrast** types of data.
- **Compare** and **contrast** types of variables.
- **Describe** the difference between a theory and a scientific law.

Vocabulary

scientific method
qualitative data
quantitative data
hypothesis
experiment
independent variable
dependent variable
control
conclusion
model
theory
scientific law

Figure 1-8 shows students working together on an experiment in the laboratory. You know that each person in the group probably has a different idea about how to do the project and a different part of the project that interests him or her the most. Having many different ideas about how to solve a problem is one of the benefits of many people working together. Communicating ideas effectively to one another and combining individual contributions to form a solution can be difficulties encountered in group work.

A Systematic Approach

Scientists approach their work in a similar way. Each scientist tries to understand his or her world based on a personal point of view and individual creativity. Often, the work of many scientists is combined in order to gain new insight. It is helpful if all scientists use common procedures as they conduct their experiments.

A **scientific method** is a systematic approach used in scientific study, whether it is chemistry, biology, physics, or other sciences. It is an organized process used by scientists to do research, and it provides a method for scientists to verify the work of others. An overview of the typical steps of a scientific method is shown in **Figure 1-9**. The steps are not used as a checklist to be done in the same order each time. Therefore, all scientists must describe their methods when they publish their results. If other scientists cannot confirm the results after repeating the method, then doubt arises over the validity of the reported results.

Observation You make observations throughout your day in order to make decisions. Scientific study usually begins with simple observation. An observation is the act of gathering information. Quite often, the types of observations scientists first make are **qualitative data**—information that describes color, odor, shape, or some other physical characteristic. In general, anything that relates to the five senses is qualitative: how something looks, feels, sounds, tastes, or smells.

Figure 1-8

During your chemistry course, you will have opportunities to use scientific methods to perform investigations and solve problems.



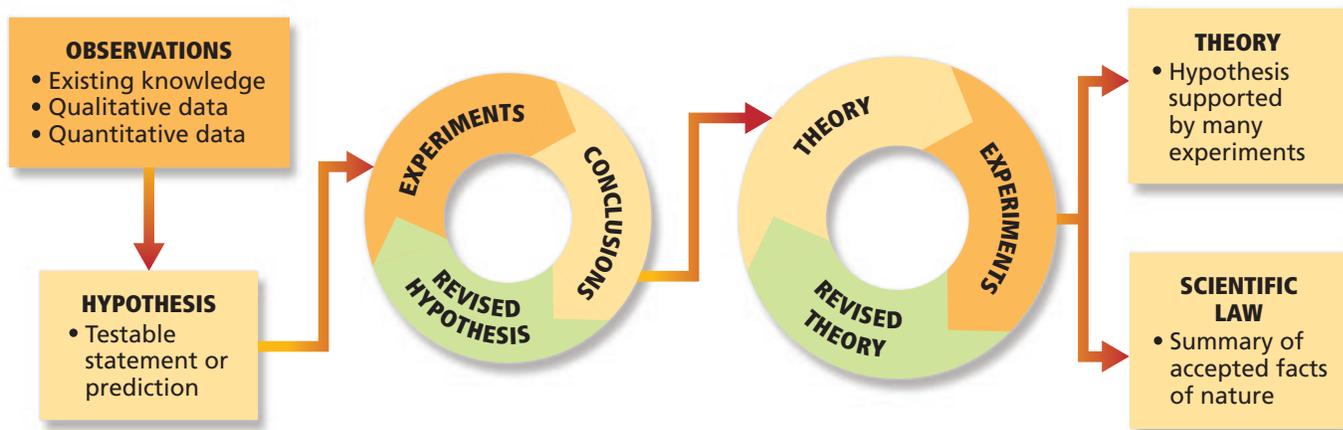


Figure 1-9

The steps in a scientific method are repeated until a hypothesis has been supported or discarded.

Chemists frequently gather another type of data. For example, they can measure temperature, pressure, volume, the quantity of a chemical formed, or how much of a chemical is used up in a reaction. This numerical information is called **quantitative data**. It tells you how much, how little, how big, how tall, or how fast. What kind of qualitative and quantitative data can you gather from **Figure 1-10**?

Hypothesis Let's return to the stories of two chemicals that you read about earlier. Even before quantitative data showed that ozone levels were decreasing in the stratosphere, scientists observed that CFCs were found there. Chemists Mario Molina and F. Sherwood Rowland were curious about how long CFCs could exist in the atmosphere.

Molina and Rowland examined the interactions that can occur among various chemicals in the troposphere. They determined that CFCs were stable there for long periods of time. But they also knew that CFCs drift up into the stratosphere. They formed a hypothesis that CFCs break down in the stratosphere due to interactions with ultraviolet light from the Sun. In addition, the calculations they made led them to hypothesize that a chlorine particle produced by this interaction would break down ozone.

A **hypothesis** is a tentative explanation for what has been observed. Molina and Rowland's hypothesis stated what they believed to be happening, even though there was no formal evidence at that point to support the statement.

Experiments A hypothesis means nothing unless there are data to support it. Thus, forming a hypothesis helps the scientist focus on the next step in a scientific method, the experiment. An **experiment** is a set of controlled observations that test the hypothesis. The scientist must carefully plan and set up one or more laboratory experiments in order to change and test one variable at a time. A variable is a quantity or condition that can have more than one value.

Suppose your chemistry teacher asks your class to use the materials shown in **Figure 1-11** to design an experiment in which to test the hypothesis that table salt dissolves faster in hot water than in water at room temperature (20°C).

Figure 1-10

Compare the qualitative and quantitative observations you can make from this photo with your classmates. Did you observe the same things?



Figure 1-11

These materials can be used to determine the effect of temperature on the rate at which table salt dissolves.



Because temperature is the variable that you plan to change, it is an **independent variable**. Your group determines that a given quantity of salt completely dissolves within 1 minute at 40°C but that the same quantity of salt dissolves only after 3 minutes at 20°C. Thus, temperature affects the rate at which the salt dissolves. This rate is called a **dependent variable** because its value changes in response to a change in the independent variable. Although your group can determine the way the independent variable changes, it has no control over the way the dependent variable changes.

What other factors could you vary in your experiment? Would the amount of salt you try to dissolve make a difference? The amount of water you use? Would stirring the mixture affect your results? The answer can be yes to all of these questions. You must plan your experiment so that these variables are the same at each temperature, or you will not be able to tell clearly what caused your results. In a well-planned experiment, the independent variable should be the only condition that affects the experiment's outcome. A constant is a factor that is not allowed to change during the experiment. The amount of salt, water, and stirring must be constant at each temperature.

In many experiments, it is valuable to have a **control**, that is, a standard for comparison. In the above experiment, the room-temperature water is the control. The rate of dissolving at 40°C is compared to the rate at 20°C. **Figure 1-12** shows a different type of control. A chemical indicator has been added to each of three test tubes. Acid is added to the middle test tube, and the indicator turns yellow. This test tube can be used as a control. Compare the other two test tubes with the control. Does either one contain acid?

The interactions described between CFCs and ozone in Molina and Rowland's hypothesis take place high overhead. Many variables are involved. For example, there are different gases present in the stratosphere. Thus, it would be difficult to determine which gases, or possibly if all gases, are decreasing ozone levels. Winds, variations in ultraviolet light, and other factors could change the outcome of any experiment on any given day making comparisons difficult. Sometimes it is easier to simulate conditions in a laboratory where the variables can be more easily controlled.

An experiment may generate a large amount of data. These data must be carefully and systematically analyzed. Because the concept of data analysis is so important, you will learn more about it in the next chapter.

Conclusion Scientists take the data that have been analyzed and apply them to the hypothesis to form a conclusion. A **conclusion** is a judgment based on the information obtained. A hypothesis can never be proven. Therefore, to say

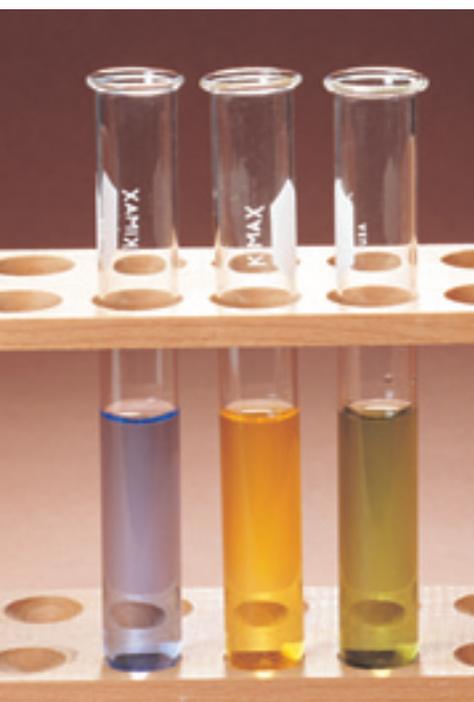


Figure 1-12

The control test tube in the middle lets you make a visual comparison.

that the data support a hypothesis is to give only a tentative “thumbs up” to the idea that the hypothesis may be true. If further evidence does not support it, then the hypothesis must be discarded or modified. The majority of hypotheses are not supported, but the data may still yield new information.

Molina and Rowland formed a hypothesis about the stability of CFCs in the stratosphere. They gathered data that supported their hypothesis and developed a model in which the chlorine formed by the breakdown of CFCs would react over and over again with ozone. You must realize by now a **model** is a visual, verbal, and/or mathematical explanation of experimental data.

A model can be tested and used to make predictions. Molina and Rowland’s model predicted the formation of chlorine and the depletion of ozone, as shown in **Figure 1-13**. Another research group found evidence of interactions between ozone and chlorine when taking measurements in the stratosphere, but they did not know the source of the chlorine. Molina and Rowland’s model predicted a source of the chlorine. They came to the conclusion that ozone in the stratosphere could be destroyed by CFCs and that they had enough support for their hypothesis to publish their discovery.

Theory A **theory** is an explanation that has been supported by many, many experiments. You may have heard of Einstein’s theory of relativity or the atomic theory. A theory states a broad principle of nature that has been supported over time. All theories are still subject to new experimental data and can be modified. Also, theories often lead to new conclusions.

A theory is considered successful if it can be used to make predictions that are true. In 1985, the announcement by the British Antarctic Survey that the amount of ozone in the stratosphere was decreasing lent Molina and Rowland’s hypothesis—that chlorine from CFCs could destroy ozone—further support.

Scientific law Sometimes, many scientists come over and over again to the same conclusion about certain relationships in nature. They find no exceptions. For example, you know that no matter how many times skydivers leap from a plane, they always wind up back on Earth’s surface. Sir Isaac Newton was so certain that an attractive force exists between all objects that he proposed his law of universal gravitation.

Newton’s law is a **scientific law** and, as such, a relationship in nature that is supported by many experiments. It is up to scientists to develop further hypotheses and experimentation to explain why these relationships exist.

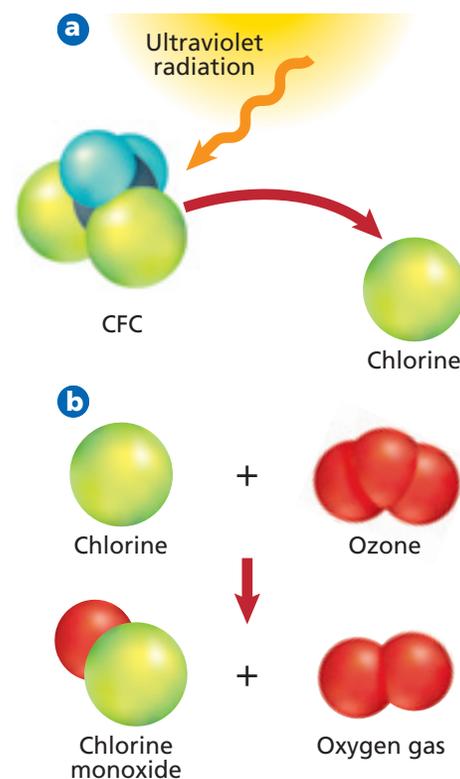


Figure 1-13

- a** Molina and Rowland’s model predicted that ultraviolet radiation causes a chlorine particle to split off from a CFC.
- b** The chlorine particle then destroys the ozone by combining with it to form oxygen gas and chlorine monoxide.

Section 1.3 Assessment

11. What is a scientific method? What are its steps?
12. You are asked to study the effect of temperature on the volume of a balloon. The balloon’s size increases as it is warmed. What is the independent variable? Dependent variable? What factor is held constant? How would you construct a control?
13. Critique Molina and Rowland’s hypothesis of ozone depletion as to its strengths and weaknesses.
14. Jacques Charles described the direct relationship between temperature and volume of all gases at constant pressure. Should this be called Charles’s law or Charles’s theory? Explain.
15. **Thinking Critically** Why must Molina and Rowland’s data in the laboratory be supported by measurements taken in the stratosphere?
16. **Interpreting Data** A report in the media states that a specific diet will protect individuals from cancer. However, no data are reported to support this statement. Is this statement a hypothesis or a conclusion?

Objectives

- **Compare** and **contrast** pure research, applied research, and technology.
- **Apply** knowledge of laboratory safety.

Vocabulary

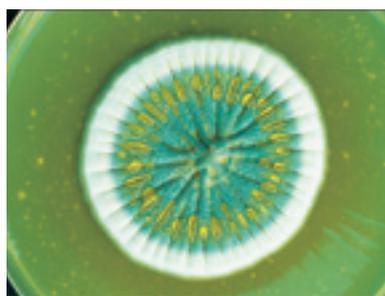
pure research
applied research
technology

Biology

CONNECTION

Many discoveries in science are made quite unexpectedly. Alexander Fleming is famous for making two such discoveries. The first occurred in 1922, when nasal mucus accidentally dripped onto bacteria that he had been growing for research. Rather than throwing the culture plate away, he decided to observe it over the next several days. He found that the bacteria died. As a result, he discovered that lysozyme, a chemical in mucus and tears, helps protect the body from bacteria.

Six years later, in 1928, Fleming found that one of his plates of *Staphylococcus* bacteria had been contaminated by a greenish mold, later identified as *Penicillium*. He observed it carefully and saw a clear area around the mold where the bacteria had died. In this case, a chemical in the mold—penicillin—was responsible for killing the bacteria.



Every day in the media, whether it's TV, newspapers, magazines, or the Internet, you are bombarded with the results of scientific investigations. Many deal with the environment, medicine, or health. As a consumer, you are asked to evaluate the results of scientific research and development. How do scientists use qualitative and quantitative data to solve different types of scientific problems?

Types of Scientific Investigations

Pure research seeks to gain knowledge for the sake of knowledge itself. Molina and Rowland conducted research on CFCs and their interactions with ozone as pure research, motivated by curiosity. No environmental evidence at the time indicated that there was a correlation to their model in the stratosphere. Their research only showed that CFCs could speed the breakdown of ozone in a laboratory setting.

When the ozone hole was reported in 1985, scientists had made measurements of CFC levels in the stratosphere that supported the hypothesis that CFCs could be responsible for the depletion of ozone. The pure research done only for the sake of knowledge became applied research. **Applied research** is research undertaken to solve a specific problem. Scientists continue to monitor the amount of CFCs in the atmosphere and the annual changes in the amount of ozone in the stratosphere. Applied research also is being done to find replacement chemicals for the CFCs that are now banned. Read the **Chemistry and Society** feature at the end of this chapter to learn about research into the human genome. What type of research does it describe?

Chance discoveries Sometimes, when a scientist plans research with a specific goal in mind, he or she will conduct experiments and reach a conclusion that is expected. Sometimes, however, the conclusion reached is far different from what was expected. Some truly wonderful discoveries in science have been made unexpectedly.

The discovery of nylon is one example. In 1928, E.I. DuPont de Nemours and Company appointed a young, 32-year-old chemist from Harvard, Wallace Carothers, as the director of its new research center. The goal was to create artificial fibers similar to cellulose and silk. In 1930, Julian Hill, a member of Carothers' team, dipped a hot glass rod in a mixture of solutions and unexpectedly pulled out long fibers such as the one shown in **Figure 1-14**. Carothers pursued the development of these fibers as a synthetic silk that could withstand high temperatures and eventually developed nylon in 1934. Nylon's first use was in a toothbrush with nylon bristles. During World War II, nylon was used as a replacement for silk in parachutes. Nylon is used extensively today in textiles and some kinds of plastics.

Students in the Laboratory

In your study of chemistry, you will learn many facts about matter. You also will do experiments in which you will be able to form and test hypotheses, gather and analyze data, and draw conclusions.

When you work in the chemistry laboratory, you are responsible for your safety and the safety of the people working nearby. **Table 1-2** on page 16 lists some safety rules that you must use as a guide each time you enter the lab.

Chemists and all other scientists use these safety rules. Before you do the **miniLAB** at the bottom of this page or the **CHEMLAB** at the end of this chapter, read the procedures carefully. Which safety rules in **Table 1-2** apply?

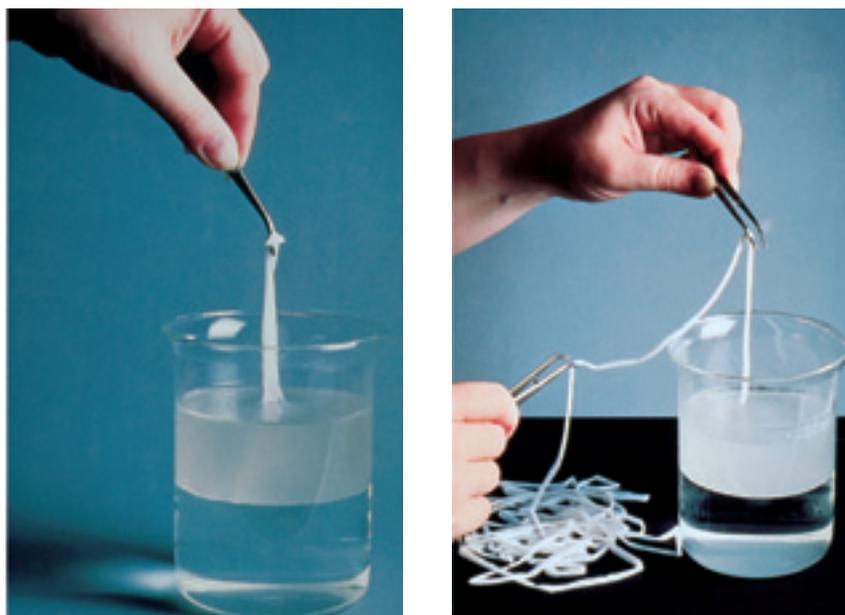


Figure 1-14

Strands of nylon can be pulled from the top layer of solutions. After its discovery, nylon was used mainly for war materials and was unavailable for home use until after World War II.

miniLAB

Developing Observation Skills

Observing and Inferring A chemist's ability to make careful and accurate observations is developed early. The observations often are used to make inferences. An inference is an explanation or interpretation of observations.

Materials petri dish (2), graduated cylinder, whole milk, water, vegetable oil, four different food colorings, toothpick (2), dishwashing detergent

Procedure

1. Add water to a petri dish to a height of 0.5 cm. Add 1 mL of vegetable oil.
2. Dip the end of a toothpick in liquid dishwashing detergent.
3. Touch the tip of the toothpick to the water at the center of the petri dish. Record your detailed observations.
4. Add whole milk to a second petri dish to a height of 0.5 cm.
5. Place one drop each of four different food colorings in four different locations on the surface of the milk, as shown in the photo. Do not put a drop of food coloring in the center.



6. Repeat steps 2 and 3.

Analysis

1. What did you observe in step 3?
2. What did you observe in step 6?
3. Oil, the fat in milk, and grease belong to a class of chemicals called lipids. What can you infer about the addition of detergent to dishwater?

Table 1-2

Safety in the Laboratory

- | | |
|--|---|
| <p>1. Study your lab assignment before you come to the lab. If you have any questions, be sure to ask your teacher for help.</p> | <p>9. Report any accident, injury, incorrect procedure, or damaged equipment to your teacher.</p> |
| <p>2. Do not perform experiments without your teacher's permission. Never work alone in the laboratory.</p> | <p>10. If chemicals come in contact with your eyes or skin, flush the area immediately with large quantities of water. Immediately inform your teacher of the nature of the spill.</p> |
| <p>3. Use the table on the inside front cover of this textbook to understand the safety symbols. Read all CAUTION statements.</p> | <p>11. Handle all chemicals carefully. Check the labels of all bottles before removing the contents. Read the label three times:</p> <ul style="list-style-type: none"> • Before you pick up the container. • When the container is in your hand. • When you put the bottle back. |
| <p>4. Safety goggles and a laboratory apron must be worn whenever you are in the lab. Gloves should be worn whenever you use chemicals that cause irritations or can be absorbed through the skin. Long hair must be tied back. See the photo below.</p> | <p>12. Do not take reagent bottles to your work area unless instructed to do so. Use test tubes, paper, or beakers to obtain your chemicals. Take only small amounts. It is easier to get more than to dispose of excess.</p> |
| <p>5. Do not wear contact lenses in the lab, even under goggles. Lenses can absorb vapors and are difficult to remove in case of an emergency.</p> | <p>13. Do not return unused chemicals to the stock bottle.</p> |
| <p>6. Avoid wearing loose, draping clothing and dangling jewelry. Bare feet and sandals are not permitted in the lab.</p> | <p>14. Do not insert droppers into reagent bottles. Pour a small amount of the chemical into a beaker.</p> |
| <p>7. Eating, drinking, and chewing gum are not allowed in the lab.</p> | <p>15. Never taste any chemicals. Never draw any chemicals into a pipette with your mouth.</p> |
| <p>8. Know where to find and how to use the fire extinguisher, safety shower, fire blanket, and first-aid kit.</p> | <p>16. Keep combustible materials away from open flames.</p> |
|  | <p>17. Handle toxic and combustible gases only under the direction of your teacher. Use the fume hood when such materials are present.</p> |
| | <p>18. When heating a substance in a test tube, be careful not to point the mouth of the test tube at another person or yourself. Never look down the mouth of a test tube.</p> |
| | <p>19. Do not heat graduated cylinders, burettes, or pipettes with a laboratory burner.</p> |
| | <p>20. Use caution and proper equipment when handling hot apparatus or glassware. Hot glass looks the same as cool glass.</p> |
| | <p>21. Dispose of broken glass, unused chemicals, and products of reactions only as directed by your teacher.</p> |
| | <p>22. Know the correct procedure for preparing acid solutions. Always add the acid slowly to the water.</p> |
| | <p>23. Keep the balance area clean. Never place chemicals directly on the pan of a balance.</p> |
| | <p>24. After completing an experiment, clean and put away your equipment. Clean your work area. Make sure the gas and water are turned off. Wash your hands with soap and water before you leave the lab.</p> |



Figure 1-15

Nuclear power and artificial limbs and joints are just a few of the technological advances that have improved human life.

Benefits of Chemistry

It is easy to understand the purpose of applied research because it addresses a specific problem. It also is easier to see its immediate benefit. Yet, when a sudden, unexpected event occurs in the world, whether it's the ozone hole or the AIDS epidemic, the first line of defense is to look at the pure research that has already been conducted.

The products that we use to make our lives easier and more comfortable are the result of technological applications of pure and applied research. **Technology** is the practical use of scientific information. It is concerned with making improvements in human life and the world around us. As **Figure 1-15** shows, advances in technology can benefit us in many ways.

Section 1.4 Assessment

17. Compare and contrast pure research and applied research.
18. What is technology? Is technology a product of pure research or applied research? Explain.
19. Explain why it is important to read each **CHEM-LAB** and **miniLAB** before you come to class.
20. **Thinking Critically** Explain the reason behind each of the following.
 - a. Wear goggles and an apron in the lab even if you are only an observer.
 - b. Report all accidents to your teacher.
 - c. Do not return unused chemicals to the stock bottle.
21. **Interpreting Scientific Diagrams** What safety precautions should you take when you see the following safety symbols?


The Rubber Band Stretch

Galileo Galilei (1564–1642) was an Italian philosopher, astronomer, and mathematician. Galileo pioneered the use of a systematic method of observation, experimentation, and analysis as a way to discover facts about nature. Modern science has its roots in Galileo's 17th-century work on the art of experimentation. This chapter introduced you to how scientists approach their work. In this **CHEMLAB**, you will have a chance to design a scientific method to study something you have observed many times before—the stretching of a rubber band.

Problem

What happens when you heat a stretched rubber band?

Objectives

- **Observe** the properties of a stretched and a relaxed rubber band.
- **Form a hypothesis** about the effect of heat on a stretched rubber band.
- **Design** an experiment to test your hypothesis.
- **Collect and analyze** data.
- **Draw conclusions** based on your analysis.

Materials

large rubber band
500-g mass
ring stand
clamp
hair dryer
meter stick or ruler

Safety Precautions



- Frequently observe the rubber band for any splits. Discard if rubber band is defective.
- The hair dryer can become hot, so handle it with care.

Pre-Lab

- Heat is the transfer of energy from a warmer object to a cooler object. If an object feels warm to your finger, your finger is cooler than the object and energy is being transferred from the object to your finger. In what direction does the energy flow if an object feels cooler to you?
- Your forehead is very sensitive to hot and cold. How can you use this fact to detect whether an object is giving off or absorbing heat?
- Read the entire **CHEMLAB**. It is important to know exactly what you are going to do during all chemistry experiments so you can use your laboratory time efficiently and safely. What is the problem that this experiment is going to explore?
- What typical steps in a scientific method will you use to explore the problem? Write down the procedure that you will use in each experiment that you design. Be sure to include all safety precautions.
- You will need to record the data that you collect during each experiment. Prepare data tables that are similar to the one below.

Rubber Band Data	
Experiment #	Observations
Trial 1	
Trial 2	
Trial 3	
Trial 4	

Procedure

1. Obtain one large rubber band. Examine the rubber band for any splits or cracks. If you find any defects, discard it and obtain a new one.
2. Record detailed observations of the unstretched rubber band.
3. Design your first experiment to observe whether heat is given off or absorbed by a rubber band as it is stretched. Have your instructor approve your plan.
4. Do repeated trials of your experiment until you are sure of the results. **CAUTION:** *Do not bring the rubber band near your face unless you are wearing goggles.*
5. Design a second experiment to observe whether heat is given off or absorbed by a rubber band as it contracts after being stretched. Have your instructor approve your plan.
6. Do repeated trials of your experiment until you are sure of the results.
7. Use your observations in steps 2, 4, and 6 to form a hypothesis and make a prediction about what will happen to a stretched rubber band when it is heated.
8. Use the remaining items in the list of materials to design a third experiment to test what happens to a stretched rubber band as it is heated. Have your



instructor approve your plan. Be sure to record all observations before, during, and after heating.

Cleanup and Disposal

1. Return the rubber band to your instructor to be reused by other classes.
2. Allow the hair dryer to cool before putting it away.

Analyze and Conclude

1. **Observing and Inferring** What results did you observe in step 4 of the procedure? Was energy gained or lost by the rubber band? By your forehead? Explain.
2. **Observing and Inferring** What results did you observe in step 6 of the procedure? Was energy gained or lost by the rubber band? By your forehead? Explain.
3. **Applying** Many substances expand when they are heated. Did the rubber band behave in the same way? How do you know?
4. **Drawing a Conclusion** Did the result of heating the stretched rubber band in step 8 confirm or refute your hypothesis? Explain.
5. **Making Predictions** What would happen if you applied ice to the stretched rubber band?
6. **Error Analysis** Compare your results and conclusion with those of your classmates. What were your independent and dependent variables? Did you use a control? Did all of the lab teams measure the same variables? Were the data that you collected qualitative or quantitative? Does this make a difference when reporting your data to others? Do your results agree? Why or why not?

Real-World Chemistry

1. When you put ice in a glass so that the ice rises higher than the rim, water does not overflow the glass when the ice melts. Explain.
2. Why do you think temperature extremes must be taken into account when bridges and highways are designed?

CHEMISTRY and Society

The Human Genome Project

From eye and skin color to the potential for developing disease, humans display remarkable and endless variety. Much of this variety is controlled by the human genome: the complete “instruction manual” found in the nucleus of all cells that is used to define a specific organism. Decoding and understanding these instructions is the goal of the Human Genome Project (HGP). The United States Department of Energy and the National Institutes of Health coordinated the project that began in 1990. Private industry also is involved. The HGP fosters cooperation as well as competition among researchers in a project that could hasten advances in treatment of human genetic conditions such as cancer and heart disease.

A common goal, a common approach

The year 2003 marks the fiftieth anniversary of Watson and Crick’s discovery of the structure of DNA. Chemicals called nitrogen bases connect the twisted strands that make up DNA. There are roughly three billion pairs of these bases in the human genome. Determining the sequence, or order, in which these base pairs occur was one common goal of researchers working on the HGP. Biologists, chemists, physicists, computer specialists, and engineers across the country approach the task from different angles. Use of scientific methods was the unifying theme in all of this work.

Prior to the HGP, researchers around the world used scientific methods to work independently on the mammoth task of decoding the human genome. Without coordination, however, the data they collected were analyzed and stored using different databases and were shared only through scientific journals and conferences. Those working on the HGP used common methods for gathering and analyzing data. Results were shared through databases that are rapidly available through the World Wide Web, enhancing communication, cooperation, and the flow of information. Because of this cooperation, a rough map of the human genome was completed several years ahead of schedule. Much work still remains to be done, however, and results will



have to be shared by researchers working on the project.

Looking toward the future

The Human Genome Project is spurring medical advances. But this powerful knowledge also raises important issues. For example, if a person’s genome map shows a predisposition for a certain illness, should an employer or insurance company be informed? While the HGP is opening exciting doors, the associated ethical, legal, and societal issues must be acknowledged and addressed.

Investigating the Issue

- 1. Communicating Ideas** Research some of the ethical, legal, and societal issues being raised by genome research. Write a brief essay giving your opinion about how one or more of these issues could be addressed.
- 2. Debating the Issue** The race to sequence the human genome took place in both the public and private sectors. Find information about several companies that worked independently on genome research. Is it ethical for them to sell this information, or should it be freely shared with others?

CHEMISTRY
Online



CLICK HERE

Visit the Chemistry Web site at science.glencoe.com to find links to more information about the Human Genome Project.

Summary

1.1 The Stories of Two Chemicals

- The building blocks of the matter in the universe formed in the stars.
- A chemical is any substance that has a definite composition.
- Ozone is a chemical that forms a protective layer in Earth's atmosphere.
- Ozone is formed in the stratosphere when ultraviolet radiation from the Sun strikes oxygen gas.
- Thinning of the ozone layer over Antarctica is called the ozone hole.
- CFCs are synthetic chemicals made of chlorine, fluorine, and carbon.
- CFCs were used as refrigerants and as propellants in aerosol cans.
- CFCs can drift into the stratosphere.

1.2 Chemistry and Matter

- Chemistry is the study of matter and the changes that it undergoes.
- Matter is anything that has mass and takes up space.
- Mass is a measure of the amount of matter.
- Weight is a measure not only of an amount of matter but also the effect of Earth's gravitational pull on that matter.
- There are five traditional branches of chemistry: organic chemistry, inorganic chemistry, physical chemistry, analytical chemistry, and biochemistry.

- Macroscopic observations of matter reflect the actions of atoms on a submicroscopic scale.

1.3 Scientific Methods

- Typical steps of a scientific method include observation, hypothesis, experiments, data analysis, and conclusion.
- Qualitative data describe an observation; quantitative data use numbers.
- An independent variable is a variable that you change in an experiment.
- A dependent variable changes in response to a change in the independent variable.
- A theory is a hypothesis that has been supported by many experiments.
- A scientific law describes relationships in nature.

1.4 Scientific Research

- Scientific methods can be used in pure research for the sake of knowledge, or in applied research to solve a specific problem.
- Laboratory safety is the responsibility of anyone who conducts an experiment.
- Many of the conveniences we enjoy today are technological applications of chemistry.

Vocabulary

- | | | |
|------------------------------|--------------------------------|-----------------------------|
| • applied research (p. 14) | • independent variable (p. 12) | • scientific law (p. 13) |
| • chemistry (p. 7) | • mass (p. 8) | • scientific method (p. 10) |
| • conclusion (p. 12) | • matter (p. 8) | • technology (p. 17) |
| • control (p. 12) | • model (p. 13) | • theory (p. 13) |
| • dependent variable (p. 12) | • pure research (p. 14) | • weight (p. 8) |
| • experiment (p. 11) | • qualitative data (p. 10) | |
| • hypothesis (p. 11) | • quantitative data (p. 11) | |

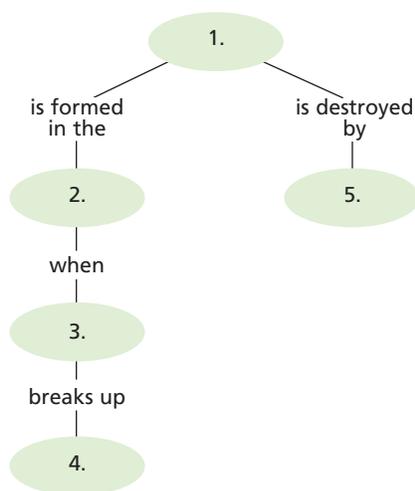


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 1 Assessment.

Concept Mapping

22. Complete the concept map using the following terms: stratosphere, oxygen gas, CFCs, ozone, ultraviolet radiation.



Mastering Concepts

23. What is a chemical? (1.1)
24. Where is ozone located in Earth's atmosphere? (1.1)
25. Explain the balance between oxygen and ozone in the stratosphere. Why is it important? (1.1)
26. What were common uses of CFCs? (1.1)
27. What is chemistry? (1.2)
28. Why is chemistry called the central science? (1.2)
29. Which measurement depends on gravitational force—mass or weight? Explain. (1.2)
30. Which branch of chemistry studies the composition of substances? Environmental impact of chemicals? (1.2)
31. How does qualitative data differ from quantitative data? Give examples of each. (1.3)
32. What is the function of a control in an experiment? (1.3)
33. What is the difference between a hypothesis, a theory, and a law? (1.3)
34. In the study of water, what questions might be asked in pure research? Applied research? Technology? (1.4)

Thinking Critically

35. **Compare and Contrast** Why is CFC depletion of the ozone layer a theory and not a scientific law?
36. **Classifying** CFCs break down to form chemicals that react with ozone. Is this a macroscopic or a microscopic observation?
37. **Communicating Ideas** Scientists often learn as much from an incorrect hypothesis as they do from one that is correct. Explain.
38. **Designing an Experiment** How would you design an experiment to evaluate the effectiveness of a "new and improved" chemical fertilizer on bean plants? Be sure to describe your hypothesis, procedure, variables, and control.
39. **Inferring** A newscaster reports, "The air quality today is poor. Visibility is only a quarter mile. Pollutants in the air are expected to rise above 0.085 parts per million (ppm) in the next eight hour average. Spend as little time outside today as possible if you suffer from asthma or other breathing problems." Which of these statements are qualitative and which are quantitative?
40. **Comparing and Contrasting** Match each of the following research topics with the branch of chemistry that would study it: water pollution, the digestion of food in the human body, the composition of a new textile fiber, metals to make new coins, a treatment for AIDS.

Writing in Chemistry

41. Based on your beginning knowledge of chemistry, describe the research into depletion of the ozone layer by CFCs in a timeline.
42. Learn about the most recent measures taken by countries around the world to reduce CFCs in the atmosphere since the Montreal Protocol. Write a short report describing the Montreal Protocol and more recent environmental measures to reduce CFCs.
43. Name a technological application of chemistry that you use everyday. Prepare a booklet about its discovery and development.

Cumulative Review

In chapters 2 through 26, this heading will be followed by questions that review your understanding of previous chapters.

Use these questions and the test-taking tip to prepare for your standardized test.

- Matter is defined as anything that _____.
 - exists in nature
 - is solid to the touch
 - is found in the universe
 - has mass and takes up space
- Mass is preferred as a measurement over weight for all of the following reasons EXCEPT _____.
 - it has the same value everywhere on Earth
 - it is independent of gravitational forces
 - it becomes less in outer space, farther from Earth
 - it is a constant measure of the amount of matter
- Which of the following is an example of pure research?
 - creating synthetic elements to study their properties
 - producing heat-resistant plastics for use in household ovens
 - finding ways to slow down the rusting of iron ships
 - searching for fuels other than gasoline to power cars
- When working with chemicals in the laboratory, which of the following is something you should NOT do?
 - Read the label of chemical bottles before using their contents.
 - Pour any unused chemicals back into their original bottles.
 - Use lots of water to wash skin that has been splashed with chemicals.
 - Take only as much as you need of shared chemicals.

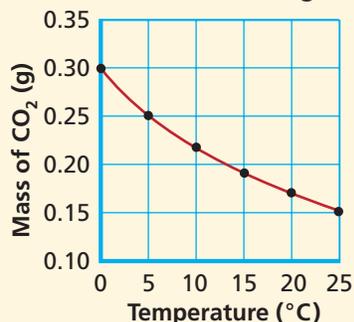
Interpreting Tables and Graphs Use the table and graph to answer questions 5–7.

- What must be a constant during the experiment?
 - temperature
 - mass of CO_2 dissolved in each sample
 - amount of beverage in each sample
 - independent variable
- Assuming that all of the experimental data are correct, what is a reasonable conclusion for this experiment?
 - Greater amounts of CO_2 dissolve in a liquid at lower temperatures.
 - The different samples of beverage contained the same amount of CO_2 at each temperature.
 - The relationship between temperature and solubility seen with solids is the same as the one seen with CO_2 .
 - CO_2 dissolves better in a liquid at higher temperatures.

Page From a Student's Laboratory Notebook

Step	Notes
Observation	Carbonated beverages taste fizzier (more gassy) when they are warm than when they are cold. (Carbonated beverages are fizzy because they contain dissolved carbon dioxide gas.)
Hypothesis	At higher temperatures, greater amounts of carbon dioxide gas will dissolve in a liquid. This is the same relationship between temperature and solubility seen with solids.
Experiment	Measure the mass of carbon dioxide (CO_2) in different samples of the same carbonated beverage at different temperatures.
Data Analysis	See graph below.
Conclusion	?

Mass of CO_2 Dissolved in a Carbonated Beverage



- The scientific method used by this student showed that _____.
 - the hypothesis is supported by the experimental data
 - the observation accurately describes what occurs in nature
 - the experiment is poorly planned
 - the hypothesis should be thrown out

TEST-TAKING TIP

More Than One Graphic If a test question has more than one table, graph, diagram, or drawing with it, use them all. If you answer based on just one graphic, you've probably missed an important piece of information. For questions 5-7 above, make sure that you accurately analyzed both graphics before answering the questions.

Data Analysis

What You'll Learn

- ▶ You will recognize SI units of measurement.
- ▶ You will convert data into scientific notation and from one unit to another.
- ▶ You will round off answers to the correct degree of certainty.
- ▶ You will use graphs to organize data.

Why It's Important

What do planting a garden, painting a room, and planning a party have in common? For each task, you need to gather and analyze data.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about measurement and data analysis.

Carpenters learn from their mistakes to “measure twice and cut once.”



DISCOVERY LAB



Materials

5 mL each in small, plastic containers:
alcohol
corn oil

glycerol
water
graduated cylinder

Layers of Liquids

How many layers will four different liquids form when you add them to a graduated cylinder?

Safety Precautions



Keep alcohol away from open flames.

Procedure

1. Pour the blue-dyed glycerol from its small container into the graduated cylinder. Allow all of the glycerol to settle to the bottom.
2. Slowly add the water by pouring it down the inside of the cylinder as shown in the photograph.
3. Repeat step 2 with the corn oil.
4. Repeat step 2 with the red-dyed alcohol.

Analysis

How are the liquids arranged in the cylinder? Hypothesize about what property of the liquids is responsible for this arrangement.

Section

2.1

Units of Measurement

Objectives

- **Define** SI base units for time, length, mass, and temperature.
- **Explain** how adding a prefix changes a unit.
- **Compare** the derived units for volume and density.

Vocabulary

base unit
second
meter
kilogram
derived unit
liter
density
kelvin

Suppose you get an e-mail from a friend who lives in Canada. Your friend complains that it has been too hot lately to play soccer or ride a bike. The high temperature each day has been about 35. You think that this figure must be wrong because a temperature of 35 is cold, not hot. Actually, 35 can be both hot and cold depending on which temperature units are used. For a measurement to be useful, it must include both a number and a unit.

SI Units

Measurement is a part of daily activities. Hospitals record the weight and length of each baby. Meters on gasoline pumps measure the volume of gasoline sold. The highway sign in **Figure 2-1** gives the measured distance from the sign's location to two different destinations. In the United States, these distances are shown in both kilometers and miles.

For people born in the United States, the mile is a familiar unit. People in most other countries measure distances in kilometers. Kilometers and miles are units of length in different measurement systems. The system that includes kilometers is the system used by scientists worldwide.

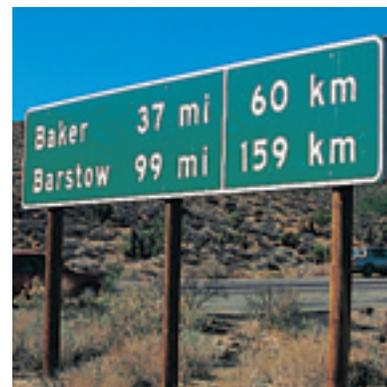


Figure 2-1

How many miles apart are Baker and Barstow? Which is longer, a mile or a kilometer?

Table 2-1

SI Base Units	
Quantity	Base unit
Time	second (s)
Length	meter (m)
Mass	kilogram (kg)
Temperature	kelvin (K)
Amount of a substance	mole (mol)
Electric current	ampere (A)
Luminous intensity	candela (cd)

For centuries, units of measurement were fairly inexact. A person might mark off the boundaries of a property by walking and counting the number of steps. The passage of time could be estimated with a sundial or an hourglass filled with sand. Such estimates worked for ordinary tasks. Scientists, however, need to report data that can be reproduced by other scientists. They need standard units of measurement. In 1795, French scientists adopted a system of standard units called the metric system. In 1960, an international committee of scientists met to update the metric system. The revised system is called the *Système Internationale d'Unités*, which is abbreviated SI.

Base Units

There are seven base units in SI. A **base unit** is a defined unit in a system of measurement that is based on an object or event in the physical world. A base unit is independent of other units. **Table 2-1** lists the seven SI base units, the quantities they measure, and their abbreviations. Some familiar quantities that are expressed in base units are time, length, mass, and temperature.

Time The SI base unit for time is the **second** (s). The frequency of microwave radiation given off by a cesium-133 atom is the physical standard used to establish the length of a second. Cesium clocks are more reliable than the clocks and stopwatches that you use to measure time. For ordinary tasks, a second is a short amount of time. Many chemical reactions take place in less than a second. To better describe the range of possible measurements, scientists add prefixes to the base units. This task is made easier because the metric system is a decimal system. The prefixes in **Table 2-2** are based on multiples, or factors, of ten. These prefixes can be used with all SI units. In Section 2.2, you will learn to express quantities such as 0.000 000 015 s in scientific notation, which also is based on multiples of ten.

Length The SI base unit for length is the **meter** (m). A meter is the distance that light travels through a vacuum in 1/299 792 458 of a second. A vacuum is a space containing no matter. A meter, which is close in length to a yard, is useful for measuring the length and width of a room. For distances between cities, you would use kilometers. The diameter of a drill bit might be reported in millimeters. Use **Table 2-2** to figure out how many millimeters are in a meter and how many meters are in a kilometer.

Table 2-2

Prefixes Used with SI Units				
Prefix	Symbol	Factor	Scientific notation	Example
giga	G	1 000 000 000	10^9	gigameter (Gm)
mega	M	1 000 000	10^6	megagram (Mg)
kilo	k	1000	10^3	kilometer (km)
deci	d	1/10	10^{-1}	deciliter (dL)
centi	c	1/100	10^{-2}	centimeter (cm)
milli	m	1/1000	10^{-3}	milligram (mg)
micro	μ	1/1 000 000	10^{-6}	microgram (μg)
nano	n	1/1 000 000 000	10^{-9}	nanometer (nm)
pico	p	1/1 000 000 000 000	10^{-12}	picometer (pm)

Astronomy

CONNECTION

A star's temperature and size determine its brightness, or luminous intensity. The SI base unit for luminous intensity is the candela. The more massive a star and the hotter its temperature, the brighter the star will be. How bright a star appears from Earth can be misleading because stars are at different distances from Earth. Light spreads out as it travels from its source. Thus, distant stars will appear less bright than stars of equal intensity that are closer to Earth.



Mass Recall that mass is a measure of the amount of matter. The SI base unit for mass is the **kilogram** (kg). A kilogram is about 2.2 pounds. The kilogram is defined by the platinum-iridium metal cylinder shown in **Figure 2-2**. The cylinder is stored in a triple bell jar to keep air away from the metal. The masses measured in most laboratories are much smaller than a kilogram. For such masses, scientists use grams (g) or milligrams (mg). There are 1000 grams in a kilogram. How many milligrams are in a gram?

Derived Units

Not all quantities can be measured with base units. For example, the SI unit for speed is meters per second (m/s). Notice that meters per second includes two SI base units—the meter and the second. A unit that is defined by a combination of base units is called a **derived unit**. Two other quantities that are measured in derived units are volume and density.

Volume Volume is the space occupied by an object. The derived unit for volume is the cubic meter, which is represented by a cube whose sides are all one meter in length. For measurements that you are likely to make, the more useful derived unit for volume is the cubic centimeter (cm^3). The cubic centimeter works well for solid objects with regular dimensions, but not as well for liquids or for solids with irregular shapes. In the **miniLAB** on the next page, you will learn how to determine the volume of irregular solids.

Figure 2-3 shows the relationship between different SI volume units, including the cubic decimeter (dm^3). The metric unit for volume equal to one cubic decimeter is a **liter** (L). Liters are used to measure the amount of liquid in a container of bottled water or a carbonated beverage. One liter has about the same volume as one quart. For the smaller quantities of liquids that you will work with in the laboratory, volume is measured in milliliters (mL). A milliliter is equal in volume to one cubic centimeter. Recall that *milli* means one-thousandth. Therefore, one liter is equal to 1000 milliliters.

Density Why is it easier to lift a grocery bag full of paper goods than it is to lift a grocery bag full of soup cans? The volumes of the grocery bags are identical. Therefore, the difference in effort must be related to how much mass is packed into the same volume. **Density** is a ratio that compares the mass of an object to its volume. The units for density are often grams per cubic centimeter (g/cm^3).



Figure 2-2

The kilogram is the only base unit whose standard is a physical object. The standard kilogram is kept in Sèvres, France. The kilogram in this photo is a copy kept at the National Institute of Standards and Technology in Gaithersburg, Maryland.

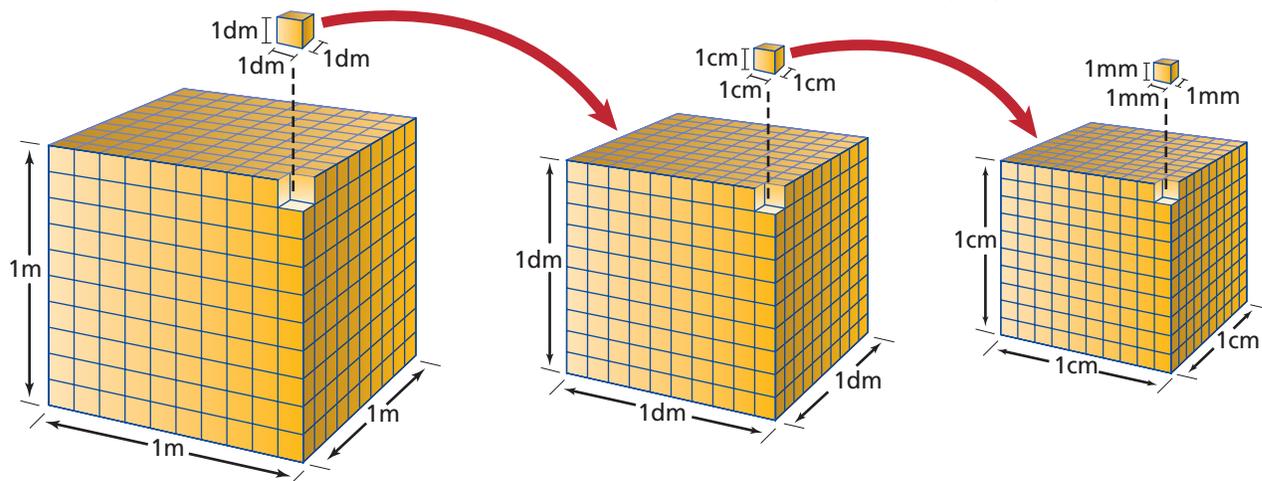


Figure 2-3

How many cubic centimeters (cm^3) are in one liter?

Figure 2-4

The piece of foam has the same mass as the quarter. Compare the densities of the quarter and the foam.



Consider the quarter and the piece of foam in **Figure 2-4**. In this case, the objects have the same mass. Because the density of the quarter is much greater than the density of the foam, the quarter occupies a much smaller space. How does density explain what you observed in the **DISCOVERY LAB**?

You can calculate density using this equation:

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

If a sample of aluminum has a mass of 13.5 g and a volume of 5.0 cm³, what is its density? Insert the known quantities for mass and volume into the density equation.

$$\text{density} = \frac{13.5 \text{ g}}{5.0 \text{ cm}^3}$$

The density of aluminum is 2.7 g/cm³. Density is a property that can be used to identify an unknown sample of matter. Every sample of pure aluminum has the same density. **How It Works** at the end of the chapter explains how ultrasound testing relies on the variation in density among materials.

miniLAB

Density of an Irregular Solid

Measuring To calculate density, you need to know both the mass and volume of an object. You can find the volume of an irregular solid by displacing water.

Materials balance, graduated cylinder, water, washer or other small object

Procedure

1. Find and record the mass of the washer.
2. Add about 15 mL of water to your graduated cylinder. Measure and record the volume. Because the surface of the water in the cylinder is curved, make volume readings at eye

level and at the lowest point on the curve. The curved surface is called a *meniscus*.

3. Carefully add the washer to the cylinder. Then measure and record the new volume.

Analysis

1. Use the initial and final volume readings to calculate the volume of the washer.
2. Use the calculated volume and the measured mass to find the density of the washer.
3. Explain why you cannot use displacement of water to find the volume of a sugar cube.
4. The washer is a short cylinder with a hole in the middle. Describe another way to find its volume.

Your textbook includes example problems that explain how to solve word problems related to concepts such as density. Each example problem uses a three-part process for problem solving: analyze, solve, and evaluate. When you analyze a problem, you first separate what is known from what is unknown. Then you decide on a strategy that uses the known data to solve for the unknown. After you solve a problem, you need to evaluate your answer to decide if it makes sense.

EXAMPLE PROBLEM 2-1

Using Density and Volume to Find Mass

Suppose a sample of aluminum is placed in a 25-mL graduated cylinder containing 10.5 mL of water. The level of the water rises to 13.5 mL. What is the mass of the aluminum sample?

1. Analyze the Problem

The unknown is the mass of aluminum. You know that mass, volume, and density are related. The volume of aluminum equals the volume of water displaced in the graduated cylinder. You know that the density of aluminum is 2.7 g/cm³, or 2.7 g/mL, because 1 cm³ equals 1 mL.

Known

density = 2.7 g/mL

volume = 13.5 mL – 10.5 mL = 3.0 mL

Unknown

mass = ? g

2. Solve for the Unknown

Rearrange the density equation to solve for mass.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{mass} = \text{volume} \times \text{density}$$

Substitute the known values for volume and density into the equation.

$$\text{mass} = 3.0 \text{ mL} \times 2.7 \text{ g/mL}$$

Multiply the values and the units. The mL units will cancel out.

$$\text{mass} = 3.0 \cancel{\text{ mL}} \times \frac{2.7 \text{ g}}{\cancel{\text{ mL}}} = 8.1 \text{ g}$$

3. Evaluate the Answer

You can check your answer by using it with the known data in the equation for density. The two sides of the equation should be equal.

$$\text{density} = \text{mass}/\text{volume}$$

$$2.7 \text{ g/mL} = 8.1 \text{ g}/3.0 \text{ mL}$$

If you divide 8.1 g by 3.0 mL, you get 2.7 g/mL.

Problem-Solving Process

THE PROBLEM

1. Read the problem carefully.
2. Be sure that you understand what it is asking you.

ANALYZE THE PROBLEM

1. Read the problem again.
2. Identify what you are given and list the known data.
3. Identify and list the unknowns.
4. Gather information you need from graphs, tables, or figures.
5. Plan the steps you will follow to find the answer.

SOLVE FOR THE UNKNOWN

1. Determine whether you need a sketch to solve the problem.
2. If the solution is mathematical, write the equation and isolate the unknown factor.
3. Substitute the known quantities into the equation.
4. Solve the equation.
5. Continue the solution process until you solve the problem.

EVALUATE THE ANSWER

1. Re-read the problem. Is the answer reasonable?
2. Check your math. Are the units and the significant figures correct? (See Section 2.3.)

PRACTICE PROBLEMS

1. A piece of metal with a mass of 147 g is placed in a 50-mL graduated cylinder. The water level rises from 20 mL to 41 mL. What is the density of the metal?
2. What is the volume of a sample that has a mass of 20 g and a density of 4 g/mL?
3. A metal cube has a mass of 20 g and a volume of 5 cm³. Is the cube made of pure aluminum? Explain your answer.



For more practice with density problems, go to **Supplemental Practice Problems** in Appendix A.

Temperature

Suppose you run water into a bathtub. You control the temperature of the water by adjusting the flow from the hot and cold water pipes. When the streams mix, heat flows from the hot water to the cold water. You classify an object as hot or cold by whether heat flows from you to the object or from the object to you. The temperature of an object is a measure of how hot or cold the object is relative to other objects.

Temperature scales Hot and cold are qualitative terms. For quantitative descriptions of temperature, you need measuring devices such as thermometers. In a thermometer, a liquid expands when heated and contracts when cooled. The tube that contains the liquid is narrow so that small changes in temperature can be detected. Scientists use two temperature scales. The Celsius scale was devised by Anders Celsius, a Swedish astronomer. He used the temperatures at which water freezes and boils to establish his scale because these temperatures are easy to reproduce. He defined the freezing point as 0 and the boiling point as 100. Then he divided the distance between these points into 100 equal units, or degrees Celsius.

The Kelvin scale was devised by a Scottish physicist and mathematician, William Thomson, who was known as Lord Kelvin. A **kelvin** (K) is the SI base unit of temperature. On the Kelvin scale, water freezes at about 273 K and boils at about 373 K. **Figure 2-5** compares the two scales. You will use the Celsius scale for your experiments. In Chapter 14, you will learn why scientists use the Kelvin scale to describe the behavior of gases.

It is easy to convert from the Celsius scale to the Kelvin scale. For example, the element mercury melts at -39°C and boils at 357°C . To convert temperatures reported in degrees Celsius into kelvins, you just add 273.

$$-39^{\circ}\text{C} + 273 = 234 \text{ K}$$

$$357^{\circ}\text{C} + 273 = 630 \text{ K}$$

It is equally easy to convert from the Kelvin scale to the Celsius scale. For example, the element bromine melts at 266 K and boils at 332 K. To convert temperatures reported in kelvins into degrees Celsius, you subtract 273.

$$266 \text{ K} - 273 = -7^{\circ}\text{C}$$

$$332 \text{ K} - 273 = 59^{\circ}\text{C}$$

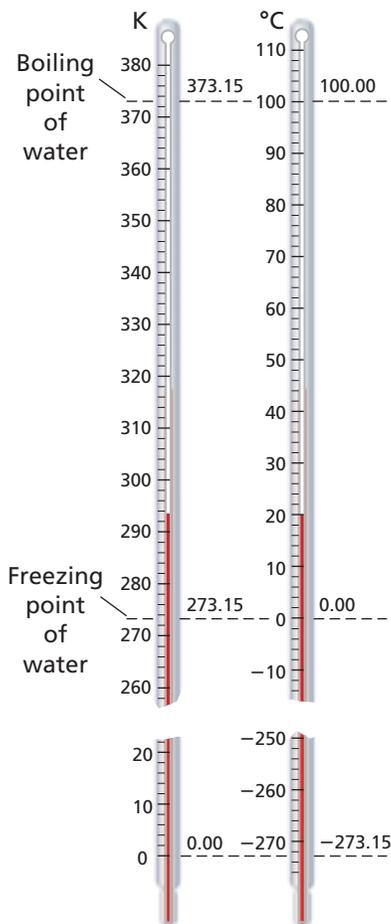


Figure 2-5

One kelvin is equal in size to one degree on the Celsius scale. The degree sign $^{\circ}$ is not used with temperatures on the Kelvin scale.

Section 2.1 Assessment

- List SI units of measurement for length, mass, time, and temperature.
- Describe the relationship between the mass, volume, and density of a material.
- Which of these samples have the same density?

Density Data		
Sample	Mass	Volume
A	80 g	20 mL
B	12 g	4 cm^3
C	33 g	11 mL

- What is the difference between a base unit and a derived unit?
- How does adding the prefix *mega-* to a unit affect the quantity being described?
- How many milliseconds are in a second? How many centigrams are in a gram?
- Thinking Critically** Why does oil float on water?
- Using Numbers** You measure a piece of wood with a meterstick and it is exactly one meter long. How many centimeters long is it?

Scientific Notation and Dimensional Analysis

A proton's mass is 0.000 000 000 000 000 000 000 001 672 62 kg. An electron's mass is 0.000 000 000 000 000 000 000 000 000 000 910 939 kg. If you try to compare the mass of a proton with the mass of an electron, the zeros get in the way. Numbers that are extremely small or large are hard to handle. You can convert such numbers into a form called scientific notation.

Scientific Notation

Scientific notation expresses numbers as a multiple of two factors: a number between 1 and 10; and ten raised to a power, or exponent. The exponent tells you how many times the first factor must be multiplied by ten. The mass of a proton is $1.627\ 62 \times 10^{-27}$ kg in scientific notation. The mass of an electron is $9.109\ 39 \times 10^{-31}$ kg. When numbers larger than 1 are expressed in scientific notation, the power of ten is positive. When numbers smaller than 1 are expressed in scientific notation, the power of ten is negative.

EXAMPLE PROBLEM 2-2

Convert Data into Scientific Notation

Change the following data into scientific notation.

- The diameter of the Sun is 1 392 000 km.
- The density of the Sun's lower atmosphere is 0.000 000 028 g/cm³.

1. Analyze the Problem

You are given two measurements. One measurement is much larger than 10. The other is much smaller than 10. In both cases, the answers will be factors between 1 and 10 that are multiplied by a power of ten.

2. Solve for the Unknown

Move the decimal point to produce a factor between 1 and 10. Count the number of places the decimal point moved and the direction.

1 392 000.

The decimal point
moved
6 places
to the left.

0.000 000 028

The decimal point
moved
8 places
to the right.

Remove the extra zeros at the end or beginning of the factor. Multiply the result by 10^n where n equals the number of places moved. When the decimal point moves to the *left*, n is a *positive* number. When the decimal point moves to the *right*, n is a *negative* number. Remember to add units to the answers.

- $1\ 392\ 000 = 1.392 \times 10^6$ km
- $0.000\ 000\ 028 = 2.8 \times 10^{-8}$ g/cm³

3. Evaluate the Answer

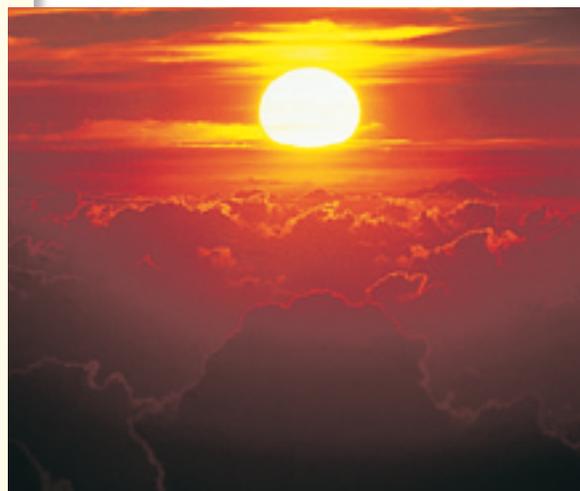
The answers have two factors. The first factor is a number between 1 and 10. Because the diameter of the Sun is a large number, 10 has a positive exponent. Because the density of the Sun's lower atmosphere is a small number, 10 has a negative exponent.

Objectives

- Express numbers in scientific notation.
- Use dimensional analysis to convert between units.

Vocabulary

scientific notation
conversion factor
dimensional analysis



The density of the Sun's lower atmosphere is similar to the density of Earth's outermost atmosphere.



For more practice converting to scientific notation, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

12. Express the following quantities in scientific notation.

- | | |
|----------------------|-----------------------|
| a. 700 m | e. 0.0054 kg |
| b. 38 000 m | f. 0.000 006 87 kg |
| c. 4 500 000 m | g. 0.000 000 076 kg |
| d. 685 000 000 000 m | h. 0.000 000 000 8 kg |

13. Express the following quantities in scientific notation.

- a. 360 000 s
 b. 0.000 054 s
 c. 5060 s
 d. 89 000 000 000 s

Adding and subtracting using scientific notation When adding or subtracting numbers written in scientific notation, you must be sure that the exponents are the same before doing the arithmetic. Suppose you need to add 7.35×10^2 m + 2.43×10^2 m. You note that the quantities are expressed to the same power of ten. You can add 7.35 and 2.43 to get 9.78×10^2 m. What can you do if the quantities are not expressed to the same power of ten?

As shown in **Figure 2-6**, some of the world's cities are extremely crowded. In 1995, the population figures for three of the four largest cities in the world were: 2.70×10^7 for Tokyo, Japan; 15.6×10^6 for Mexico City, Mexico; and 0.165×10^8 for São Paulo, Brazil. To find the total population for these three cities in 1995, you first need to change the data so that all three quantities are expressed to the same power of ten. Because the first factor in the data for Tokyo is a number between 1 and 10, leave that quantity as is: 2.70×10^7 . Change the other two quantities so that the exponent is 7. For the Mexico City data, you need to increase the power of ten from 10^6 to 10^7 . You must move the decimal point one place to the left.

$$15.6 \times 10^6 = 1.56 \times 10^7$$

For the São Paulo data, you need to decrease the power of ten from 10^8 to 10^7 . You must move the decimal point one place to the right.

$$0.165 \times 10^8 = 1.65 \times 10^7$$

Now you can add the quantities.

$$2.70 \times 10^7 + 1.56 \times 10^7 + 1.65 \times 10^7 = 5.91 \times 10^7$$

You can test this procedure by writing the original data in ordinary notation. When you add 27 000 000 + 15 600 000 + 16 500 000, you get 59 100 000. When you convert the answer back to scientific notation, you get 5.91×10^7 .



Figure 2-6

Population density is high in a city such as São Paulo, Brazil.

PRACTICE PROBLEMS

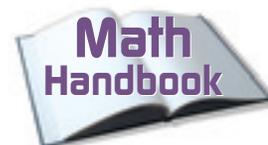
Solve the following addition and subtraction problems. Express your answers in scientific notation.

- | | |
|---|--|
| 14. a. 5×10^{-5} m + 2×10^{-5} m | e. 1.26×10^4 kg + 2.5×10^3 kg |
| b. 7×10^8 m - 4×10^8 m | f. 7.06×10^{-3} kg + 1.2×10^{-4} kg |
| c. 9×10^2 m - 7×10^2 m | g. 4.39×10^5 kg - 2.8×10^4 kg |
| d. 4×10^{-12} m + 1×10^{-12} m | h. 5.36×10^{-1} kg - 7.40×10^{-2} kg |



Multiplying and dividing using scientific notation Multiplying and dividing also involve two steps, but in these cases the quantities being multiplied or divided do not have to have the same exponent. For multiplication, you multiply the first factors. Then, you add the exponents. For division, you divide the first factors. Then, you subtract the exponent of the divisor from the exponent of the dividend.

Take care when determining the sign of the exponent in an answer. Adding +3 to +4 yields +7, but adding +3 to -4 yields -1. Subtracting -6 from -4 yields +2, but subtracting -4 from -6 yields -2.



Review arithmetic operations with positive and negative numbers in the **Math Handbook** on pages 887 to 889 of this text.

EXAMPLE PROBLEM 2-3

Multiplying and Dividing Numbers in Scientific Notation

Suppose you are asked to solve the following problems.

- a.** $(2 \times 10^3) \times (3 \times 10^2)$
b. $(9 \times 10^8) \div (3 \times 10^{-4})$

1. Analyze the Problem

You are given values to multiply and divide. For the multiplication problem, you multiply the first factors. Then you add the exponents. For the division problem, you divide the first factors. Then you subtract the exponent of the divisor from the exponent of the dividend.

$$\text{Quotient} = \frac{9 \times 10^8}{3 \times 10^{-4}}$$

↖ Dividend
↘ Divisor

2. Solve for the Unknown

a. $(2 \times 10^3) \times (3 \times 10^2)$

Multiply the first factors.

$$2 \times 3 = 6$$

Add the exponents.

$$3 + 2 = 5$$

Combine the factors.

$$6 \times 10^5$$

b. $(9 \times 10^8) \div (3 \times 10^{-4})$

Divide the first factors.

$$9 \div 3 = 3$$

Subtract the exponents.

$$8 - (-4) = 8 + 4 = 12$$

Combine the factors.

$$3 \times 10^{12}$$

3. Evaluate the Answer

You can test these procedures by writing the original data in ordinary notation. For example, problem a becomes 2000×300 . An answer of 600 000 seems reasonable.

PRACTICE PROBLEMS

Solve the following multiplication and division problems. Express your answers in scientific notation.

15. Calculate the following areas. Report the answers in square centimeters, cm^2 .

- a.** $(4 \times 10^2 \text{ cm}) \times (1 \times 10^8 \text{ cm})$
b. $(2 \times 10^{-4} \text{ cm}) \times (3 \times 10^2 \text{ cm})$
c. $(3 \times 10^1 \text{ cm}) \times (3 \times 10^{-2} \text{ cm})$
d. $(1 \times 10^3 \text{ cm}) \times (5 \times 10^{-1} \text{ cm})$

16. Calculate the following densities. Report the answers in g/cm^3 .

- a.** $(6 \times 10^2 \text{ g}) \div (2 \times 10^1 \text{ cm}^3)$
b. $(8 \times 10^4 \text{ g}) \div (4 \times 10^1 \text{ cm}^3)$
c. $(9 \times 10^5 \text{ g}) \div (3 \times 10^{-1} \text{ cm}^3)$
d. $(4 \times 10^{-3} \text{ g}) \div (2 \times 10^{-2} \text{ cm}^3)$



For more practice doing arithmetic operations using scientific notation, go to **Supplemental Practice Problems** in Appendix A.

Figure 2-7

Twelve teaspoons equal four tablespoons; four tablespoons equal 1/4 of a cup. How many teaspoons are equivalent to two 1/4 measuring cups?



Dimensional Analysis

Suppose you have a salad dressing recipe that calls for 2 teaspoons of vinegar. You plan to make 6 times as much salad dressing for a party. That means you need 12 teaspoons of vinegar. You could measure out 12 teaspoons or you could use a larger unit. According to **Figure 2-7**, 3 teaspoons are equivalent to 1 tablespoon and 4 tablespoons are equivalent to 1/4 of a cup. The relationship between teaspoons and tablespoons can be expressed as a pair of ratios. These ratios are conversion factors.

$$\frac{3 \text{ teaspoons}}{1 \text{ tablespoon}} = \frac{1 \text{ tablespoon}}{3 \text{ teaspoons}} = 1$$

A **conversion factor** is a ratio of equivalent values used to express the same quantity in different units. A conversion factor is always equal to 1. Because a quantity does not change when it is multiplied or divided by 1, conversion factors change the units of a quantity without changing its value. If you measure out 12 teaspoons of vinegar, 4 tablespoons of vinegar, or 1/4 of a cup of vinegar, you will get the same volume of vinegar.

Dimensional analysis is a method of problem-solving that focuses on the units used to describe matter. For example, if you want to convert a temperature in degrees Celsius to a temperature in kelvins, you focus on the relationship between the units in the two temperature scales. Scale drawings such as maps and the blueprint in **Figure 2-8** are based on the relationship between different units of length.

Dimensional analysis often uses conversion factors. Suppose you want to know how many meters are in 48 km. You need a conversion factor that relates kilometers to meters. You know that 1 km is equal to 1000 m. Because you are going to multiply 48 km by the conversion factor, you want to set up the conversion factor so the kilometer units will cancel out.

$$48 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = 48\,000 \text{ m}$$

When you convert from a large unit to a small unit, the number of units must increase. A meter is a much smaller unit than a kilometer, one one-thousandth smaller to be exact. Thus, it is reasonable to find that there are 48 000 meters in 48 kilometers.



Figure 2-8

Blueprints are scale drawings. On a blueprint, objects and distances appear smaller than their actual sizes but the relative sizes of objects remain the same. How can conversion factors be used to make a scale drawing?

PRACTICE PROBLEMS

Refer to **Table 2-2** to figure out the relationship between units.

- | | |
|------------------------------------|------------------------------------|
| 17. a. Convert 360 s to ms. | 18. a. Convert 245 ms to s. |
| b. Convert 4800 g to kg. | b. Convert 5 m to cm. |
| c. Convert 5600 dm to m. | c. Convert 6800 cm to m. |
| d. Convert 72 g to mg. | d. Convert 25 kg to Mg. |

EXAMPLE PROBLEM 2-4

Using Multiple Conversion Factors

What is a speed of 550 meters per second in kilometers per minute?

1. Analyze the Problem

You are given a speed in meters per second. You want to know the equivalent speed in kilometers per minute. You need conversion factors that relate kilometers to meters and seconds to minutes.

2. Solve for the Unknown

First convert meters to kilometers. Set up the conversion factor so that the meter units will cancel out.

$$\frac{550 \cancel{\text{m}}}{\text{s}} \times \frac{1 \text{ km}}{1000 \cancel{\text{m}}} = \frac{0.55 \text{ km}}{\text{s}}$$

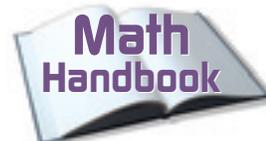
Next convert seconds to minutes. Set up the conversion factor so that the seconds cancel out.

$$\frac{0.55 \text{ km}}{\cancel{\text{s}}} \times \frac{60 \cancel{\text{s}}}{1 \text{ min}} = \frac{33 \text{ km}}{\text{min}}$$

3. Evaluate the Answer

To check your answer, you can do the steps in reverse order.

$$\frac{550 \text{ m}}{\cancel{\text{s}}} \times \frac{60 \cancel{\text{s}}}{1 \text{ min}} = \frac{33\,000 \cancel{\text{m}}}{\text{min}} \times \frac{1 \text{ km}}{1000 \cancel{\text{m}}} = \frac{33 \text{ km}}{\text{min}}$$



Review dimensional analysis and unit conversions in the **Math Handbook** on pages 900 and 901 of this text.

PRACTICE PROBLEMS

- How many seconds are there in 24 hours?
- The density of gold is 19.3 g/mL. What is gold's density in decigrams per liter?
- A car is traveling 90.0 kilometers per hour. What is its speed in miles per minute? One kilometer = 0.62 miles.



For more practice using conversion factors, go to **Supplemental Practice Problems** in Appendix A.

Using the wrong units to solve a problem can be a costly error. In 1999, the Mars Climate Orbiter crashed into the atmosphere of Mars instead of flying closely by as planned. The probe was destroyed before it could collect any data. Two teams of engineers working on the probe had used different sets of units—English and metric—and no one had caught the error in time.

Section 2.2 Assessment

- Is the number 5×10^{-4} greater or less than 1.0? Explain your answer.
- When multiplying numbers in scientific notation, what do you do with the exponents?
- Write the quantities 3×10^{-4} cm and 3×10^4 km in ordinary notation.
- Write a conversion factor for cubic centimeters and milliliters.
- What is dimensional analysis?
- Thinking Critically** When subtracting or adding two numbers in scientific notation, why do the exponents need to be the same?
- Applying Concepts** You are converting 68 km to meters. Your answer is 0.068 m. Explain why this answer is incorrect and the likely source of the error.

Objectives

- **Define** and **compare** accuracy and precision.
- **Use** significant figures and rounding to reflect the certainty of data.
- **Use** percent error to **describe** the accuracy of experimental data.

Vocabulary

accuracy
precision
percent error
significant figure

Suppose someone is planning a bicycle trip from Baltimore, Maryland to Washington D.C. The actual mileage will be determined by where the rider starts and ends the trip, and the route taken. While planning the trip, the rider does not need to know the actual mileage. All the rider needs is an estimate, which in this case would be about 39 miles. People need to know when an estimate is acceptable and when it is not. For example, you could use an estimate when buying material to sew curtains for a window. You would need more exact measurements when ordering custom shades for the same window.

Accuracy and Precision

When scientists make measurements, they evaluate both the accuracy and the precision of the measurements. **Accuracy** refers to how close a measured value is to an accepted value. **Precision** refers to how close a series of measurements are to one another. The archery target in **Figure 2-9** illustrates the difference between accuracy and precision. For this example, the center of the target is the accepted value.

In **Figure 2-9a**, the location of the arrow is accurate because the arrow is in the center. In **Figure 2-9b**, the arrows are close together but not near the center. They have a precise location but not an accurate one. In **Figure 2-9c**, the arrows are closely grouped in the center. Their locations are both accurate and precise. In **Figure 2-9d**, the arrows are scattered at a distance from the center. Their locations are neither accurate nor precise. Why does it make no sense to discuss the precision of the arrow location in **Figure 2-9a**?

Consider the data in **Table 2-3**. Students were asked to find the density of an unknown white powder. Each student measured the volume and mass of three separate samples. They reported calculated densities for each trial and an average of the three calculations. The powder was sucrose, also called table sugar, which has a density of 1.59 g/cm^3 . Who collected the most accurate data? Student A's measurements are the most accurate because they are closest to the accepted value of 1.59 g/cm^3 . Which student collected the most precise data? Student C's measurements are the most precise because they are the closest to one another.

Figure 2-9

An archery target illustrates the difference between accuracy and precision.

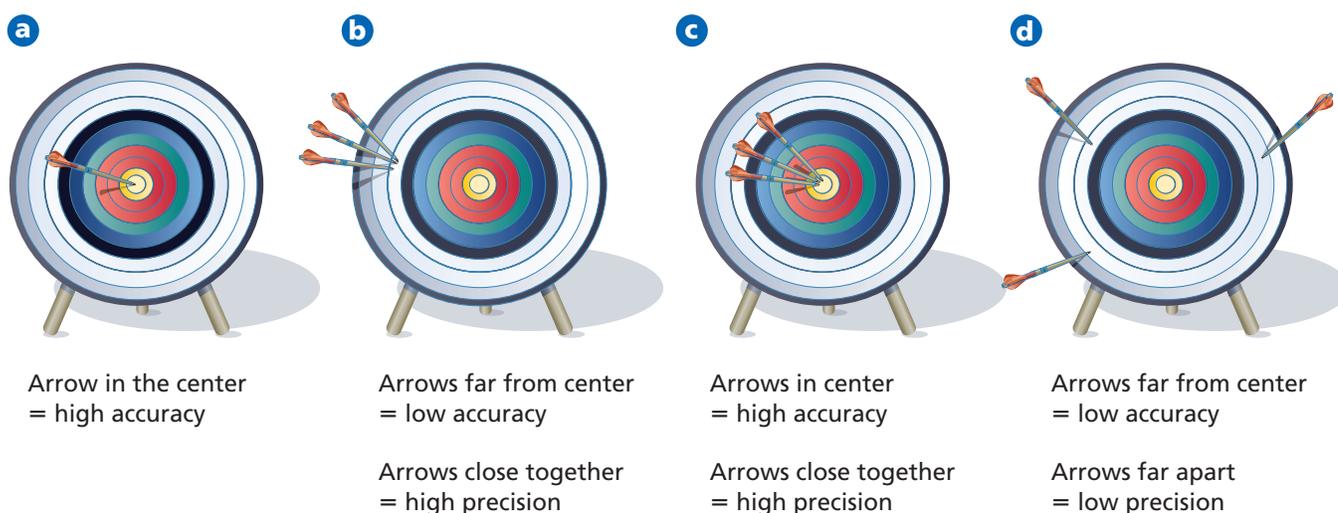


Table 2-3

Density Data Collected by Three Different Students			
	Student A	Student B	Student C
Trial 1	1.54 g/cm ³	1.40 g/cm ³	1.70 g/cm ³
Trial 2	1.60 g/cm ³	1.68 g/cm ³	1.69 g/cm ³
Trial 3	1.57 g/cm ³	1.45 g/cm ³	1.71 g/cm ³
Average	1.57 g/cm ³	1.51 g/cm ³	1.70 g/cm ³

Recall that precise measurements may not be accurate. Looking at just the average of the densities can be misleading. Based solely on the average, Student B appears to have collected fairly reliable data. However, on closer inspection, Student B's data are neither accurate nor precise. The data are not close to the accepted value and they are not close to one another.

What factors could account for inaccurate or imprecise data? Perhaps Student A did not follow the procedure with consistency. He or she might not have read the graduated cylinder at eye level for each trial. Student C may have made the same slight error with each trial. Perhaps he or she included the mass of the filter paper used to protect the balance pan. Student B may have recorded the wrong data or made a mistake when dividing the mass by the volume. External conditions such as temperature and humidity also can affect the collection of data.

Percent error The density values reported in **Table 2-3** are experimental values, which are values measured during an experiment. The density of sucrose is an accepted value, which is a value that is considered true. To evaluate the accuracy of experimental data, you can calculate the difference between an experimental value and an accepted value. The difference is called an error. The errors for the data in **Table 2-3** are listed in **Table 2-4**.

Scientists want to know what percent of the accepted value an error represents. **Percent error** is the ratio of an error to an accepted value.

$$\text{Percent error} = \frac{\text{error}}{\text{accepted value}} \times 100$$

For this calculation, it does not matter whether the experimental value is larger or smaller than the accepted value. Only the size of the error matters. When you calculate percent error, you ignore plus and minus signs. Percent error is an important concept for the person assembling bicycle gears in **Figure 2-10**. The dimensions of a part may vary within narrow ranges of error called tolerances. Some of the manufactured parts are tested to see if they meet engineering standards. If one dimension of a part exceeds its tolerance, the item will be discarded or, if possible, retooled.

**Figure 2-10**

The dimensions for each part used to build a bicycle gear have accepted values.

Table 2-4

Errors for Data in Table 2-3			
	Student A	Student B	Student C
Trial 1	-0.05 g/cm ³	-0.19 g/cm ³	+0.11 g/cm ³
Trial 2	+0.01 g/cm ³	+0.09 g/cm ³	+0.10 g/cm ³
Trial 3	-0.02 g/cm ³	-0.14 g/cm ³	+0.12 g/cm ³

Table 2-5

Student A's Data		
Trial	Density (g/cm ³)	Error (g/cm ³)
1	1.54	-0.05
2	1.60	+0.01
3	1.57	-0.02

EXAMPLE PROBLEM 2-5**Calculating Percent Error**

Calculate the percent errors. Report your answers to two places after the decimal point. Table 2-5 summarizes Student A's data.

1. Analyze the Problem

You are given the errors for a set of density measurements. To calculate percent error, you need to know the accepted value for density, the errors, and the equation for percent error.

Known

accepted value for density = 1.59 g/cm³
 errors: -0.05 g/cm³; 0.01 g/cm³; -0.02 g/cm³

Unknown

percent errors = ?

2. Solve for the Unknown

Substitute each error into the percent error equation. Ignore the plus and minus signs. Note that the units for density cancel out.

$$\text{percent error} = \frac{\text{error}}{\text{accepted value}} \times 100$$

$$\text{percent error} = \frac{0.05 \text{ g/cm}^3}{1.59 \text{ g/cm}^3} \times 100 = 3.14\%$$

$$\text{percent error} = \frac{0.01 \text{ g/cm}^3}{1.59 \text{ g/cm}^3} \times 100 = 0.63\%$$

$$\text{percent error} = \frac{0.02 \text{ g/cm}^3}{1.59 \text{ g/cm}^3} \times 100 = 1.26\%$$

3. Evaluate the Answer

The percent error is greatest for trial 1, which had the largest error, and smallest for trial 2, which was closest to the accepted value.

PRACTICE PROBLEMS

Use data from Table 2-4. Remember to ignore plus and minus signs.

29. Calculate the percent errors for Students B's trials.
30. Calculate the percent errors for Student C's trials.

Practice!

For more practice with percent error, go to **Supplemental Practice Problems** in Appendix A.

Significant Figures

Often, precision is limited by the available tools. If you have a digital clock that displays the time as 12:47 or 12:48, you can record the time only to the nearest minute. If you have a clock with a sweep hand, you can record the time to the nearest second. With a stopwatch, you might record time elapsed to the nearest hundredth of a second. As scientists have developed better measuring devices, they have been able to make more precise measurements. Of course, the measuring devices must be in good working order. For example, a balance must read zero when no object is resting on it. The process for assuring the accuracy of a measuring device is called calibration. The person using the instrument must be trained and use accepted techniques.

Scientists indicate the precision of measurements by the number of digits they report. A value of 3.52 g is more precise than a value of 3.5 g. The digits that are reported are called significant figures. **Significant figures**



include all known digits plus one estimated digit. Consider the rod in **Figure 2-11**. The end of the rod falls somewhere between 5 and 6 cm. Counting over from the 5-cm mark, you can count 2 millimeter tick marks. Thus, the rod's length is between 5.2 cm and 5.3 cm. The 5 and 2 are known digits that correspond to marks on the ruler. You can add one estimated digit to reflect the rod's location relative to the 2 and 3 millimeter marks. The third digit is an estimate because the person reading the ruler must make a judgment call. One person may report the answer as 5.23 cm. Another may report it as 5.22 cm. Either way, the answer has three significant figures—two known and one estimated.

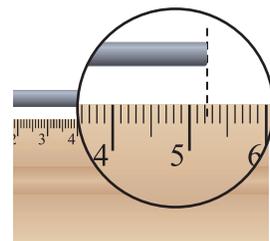


Figure 2-11

What determines whether a figure is known or estimated?

Rules for recognizing significant figures

1. *Non-zero numbers are always significant.* **72.3** g has three
2. *Zeros between non-zero numbers are always significant.* **60.5** g has three
3. *All final zeros to the right of the decimal place are significant.* **6.20** g has three
4. *Zeros that act as placeholders are not significant. Convert quantities to scientific notation to remove the placeholder zeros.* **0.0253** g and **4320** g each have three
5. *Counting numbers and defined constants have an infinite number of significant figures.* 6 molecules
60 s = 1 min

EXAMPLE PROBLEM 2-6

Applying Significant Figure Rules

Determine the number of significant figures in the following masses.

- 0.000 402 30 g
- 405 000 kg

1. Analyze the Problem

You are given two measurements of mass. Choose the rules that are appropriate to the problem.

2. Solve for the Unknown

Count all non-zero numbers (rule 1), zeros between non-zero numbers (rule 2), and final zeros to the right of the decimal place (rule 3). Ignore zeros that act as placeholders (rule 4).

- 0.000 402 30** g has five significant figures.
- 405 000** kg has three significant figures.

3. Evaluate the Answer

Write the data in scientific notation: 4.0230×10^{-4} g and 4.05×10^5 kg. Without the placeholder zeros, it is clear that 0.000 402 30 g has five significant figures and 405 000 kg has three significant figures.



Review significant figures in the **Math Handbook** on page 893 of this text.

PRACTICE PROBLEMS

Determine the number of significant figures in each measurement.

- | | |
|-----------------------------------|-------------------------------------|
| 31. a. 508.0 L | 32. a. 0.049 450 s |
| b. 820 400.0 L | b. 0.000 482 mL |
| c. 1.0200×10^5 kg | c. 3.1587×10^{-8} g |
| d. 807 000 kg | d. 0.0084 mL |



For more practice with significant figures, go to **Supplemental Practice Problems** in Appendix A.



Figure 2-12

The calculator often provides more significant figures than are appropriate for a given calculation.

Rounding Off Numbers

Suppose you are asked to find the density of an object whose mass is 22.44 g and whose volume is 14.2 cm³. When you use your calculator, you get a density of 1.580 281 7 g/cm³, as shown in **Figure 2-12**. A calculated density with eight significant figures is not appropriate if the mass has only four significant figures and the volume has only three. The answer should have no more significant figures than the data with the fewest significant figures. The density must be rounded off to three significant figures, or 1.58 g/cm³.

Rules for rounding numbers

In the example for each rule, there are three significant figures.

1. If the digit to the immediate right of the last significant figure is less than five, do not change the last significant figure. $2.532 \rightarrow 2.53$
2. If the digit to the immediate right of the last significant figure is greater than five, round up the last significant figure. $2.536 \rightarrow 2.54$
3. If the digit to the immediate right of the last significant figure is equal to five and is followed by a nonzero digit, round up the last significant figure. $2.5351 \rightarrow 2.54$
4. If the digit to the immediate right of the last significant figure is equal to five and is not followed by a nonzero digit, look at the last significant figure. If it is an odd digit, round it up. If it is an even digit, do not round up. $2.5350 \rightarrow 2.54$ but $2.5250 \rightarrow 2.52$

EXAMPLE PROBLEM 2-7

Applying the Rounding Rules

Round 3.515 014 to (a) five significant figures, then to (b) three significant figures, and finally to (c) one significant figure.

1. Analyze the Problem

You are given a number that has seven significant figures. You will remove two figures with each step. You will need to choose the rule that is appropriate for each step.

2. Solve for the Unknown

a. Round 3.515 014 to five significant figures.

Rule 1 applies. The last significant digit is 0. The number to its immediate right is 1, which is less than 5. The zero does not change. The answer is 3.5150.

b. Round 3.5150 to three significant figures.

Rule 4 applies. The last significant digit is 1. The number to its immediate right is a 5 that is not followed by a nonzero digit. Because the 1 is an odd number it is rounded up to 2. The answer is 3.52.

c. Round 3.52 to one significant figure.

Rule 3 applies. The last significant digit is 3. The number to its immediate right is a 5 that is followed by a nonzero digit. Thus, the 3 is rounded up to 4. The answer is 4.

3. Evaluate the Answer

The final answer, 4, makes sense because 3.515 013 7 is greater than the halfway point between 3 and 4, which is 3.5.

PRACTICE PROBLEMS

Round all numbers to four significant figures. Write the answers to problem 34 in scientific notation.

- 33. a.** 84 791 kg
b. 38.5432 g
c. 256.75 cm
d. 4.9356 m
- 34. a.** 0.000 548 18 g
b. 136 758 kg
c. 308 659 000 mm
d. 2.0145 mL

Addition and subtraction When you add or subtract measurements, your answer must have the same number of digits to the right of the decimal point as the value with the fewest digits to the right of the decimal point. For example, the measurement 1.24 mL has two digits to the right of the decimal point. The measurement 12.4 mL has one digit to the right of the decimal point. The measurement 124 mL has zero digits to the right of the decimal point, which is understood to be to the right of the 4. The easiest way to solve addition and subtraction problems is to arrange the values so that the decimal points line up. Then do the sum or subtraction. Identify the value with the fewest places after the decimal point. Round the answer to the same number of places.

EXAMPLE PROBLEM 2-8

Applying Rounding Rules to Addition

Add the following measurements: 28.0 cm, 23.538 cm, and 25.68 cm.

1. Analyze the Problem

There are three measurements that need to be aligned on their decimal points and added. The measurement with the fewest digits after the decimal point is 28.0 cm, with one digit. Thus, the answer must be rounded to only one digit after the decimal point.

2. Solve for the Unknown

Line up the measurements.

$$\begin{array}{r} 28.0 \text{ cm} \\ 23.538 \text{ cm} \\ + 25.68 \text{ cm} \\ \hline 77.218 \text{ cm} \end{array}$$

Because the digit immediately to the right of the last significant digit is less than 5, rule 1 applies. The answer is 77.2 cm.

3. Evaluate the Answer

The answer, 77.2 cm, has the same precision as the least precise measurement, 28.0 cm.

PRACTICE PROBLEMS

Complete the following addition and subtraction problems. Round off the answers when necessary.

- 35. a.** 43.2 cm + 51.0 cm + 48.7 cm
b. 258.3 kg + 257.11 kg + 253 kg
c. 0.0487 mg + 0.058 34 mg + 0.004 83 mg
- 36. a.** 93.26 cm – 81.14 cm
b. 5.236 cm – 3.14 cm
c. $4.32 \times 10^3 \text{ cm} - 1.6 \times 10^3 \text{ cm}$

Careers Using Chemistry

Scientific Illustrator

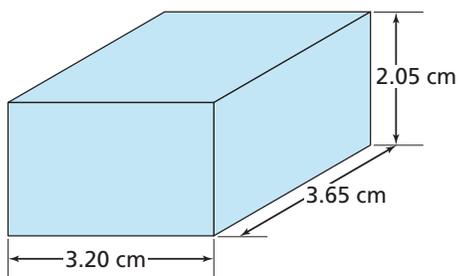
Imagine the expertise that went into illustrating this book. You can combine a science background with your artistic ability in the technically demanding career of scientific illustrator.

Scientific illustrations are a form of art required for textbooks, museum exhibits, Web sites, and publications of scientific research. These figures often show what photographs cannot—the reconstruction of an object from fragments, comparisons among objects, or the demonstration of a process or idea. Scientific illustrators use everything from paper and pencil to computer software to provide this vital link in the scientific process.



For more practice with rounding after addition or subtraction, go to **Supplemental Practice Problems** in Appendix A.

Multiplication and division When you multiply or divide numbers, your answer must have the same number of significant figures as the measurement with the fewest significant figures.



To find the volume of a rectangular object, multiply the length of the base times the width times the height.

EXAMPLE PROBLEM 2-9

Applying Rounding Rules to Multiplication

Calculate the volume of a rectangular object with the following dimensions: length = 3.65 cm; width = 3.20 cm; height = 2.05 cm.

1. Analyze the Problem

You are given measurements for the length, width, and height of a rectangular object. Because all three measurements have three significant figures, the answer will have three significant figures. Note that the units must be multiplied too.

2. Solve for the Unknown

To find the volume of a rectangular object, multiply the length times the width times the height.

$$3.20 \text{ cm} \times 3.65 \text{ cm} \times 2.05 \text{ cm} = 23.944 \text{ cm}^3$$

Because the data have only three significant figures, the answer can have only three significant figures.

The answer is 23.9 cm^3 .

3. Evaluate the Answer

To test if your answer is reasonable, round the data to one significant figure. Multiply $3 \text{ cm} \times 4 \text{ cm} \times 2 \text{ cm}$ to get 24 cm^3 . Your answer, 23.9 cm^3 , has the same number of significant figures as the data. All three measurements should have the same number of significant figures because the same ruler or tape measure was used to collect the data.



For more practice with rounding after multiplication, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Complete the following calculations. Round off the answers to the correct number of significant figures.

- | | |
|--|--|
| 37. a. $24 \text{ m} \times 3.26 \text{ m}$ | 38. a. $4.84 \text{ m}/2.4 \text{ s}$ |
| b. $120 \text{ m} \times 0.10 \text{ m}$ | b. $60.2 \text{ m}/20.1 \text{ s}$ |
| c. $1.23 \text{ m} \times 2.0 \text{ m}$ | c. $102.4 \text{ m}/51.2 \text{ s}$ |
| d. $53.0 \text{ m} \times 1.53 \text{ m}$ | d. $168 \text{ m}/58 \text{ s}$ |

Section 2.3 Assessment

- 39.** A piece of wood has a labeled length value of 76.49 cm. You measure its length three times and record the following data: 76.48 cm, 76.47 cm, and 76.59 cm. How many significant figures do these measurements have?
- 40.** Are the measurements in problem 39 accurate? Are they precise? Explain your answers.
- 41.** Calculate the percent error for each measurement in problem 39.
- 42.** Round 76.51 cm to two significant figures. Then round your answer to one significant figure.
- 43. Thinking Critically** Which of these measurements was made with the most precise measuring device: 8.1956 m, 8.20 m, or 8.196 m? Explain your answer.
- 44. Using Numbers** Write an expression for the quantity 506 000 cm in which it is clear that all the zeros are significant.

When you analyze data, you may set up an equation and solve for an unknown, but this is not the only method scientists have for analyzing data. A goal of many experiments is to discover whether a pattern exists in a certain situation. Does raising the temperature change the rate of a reaction? Does a change in diet affect a rat's ability to solve a maze? When data are listed in a table such as **Table 2-6**, a pattern may not be obvious.

Graphing

Using data to create a graph can help to reveal a pattern if one exists. A **graph** is a visual display of data. Have you ever heard the saying, "A picture is worth a thousand words?"

Circle graphs If you read a paper or a news magazine, you will find many graphs. The circle graph in **Figure 2-13a** is sometimes called a pie chart because it is divided into wedges like a pie or pizza. A circle graph is useful for showing parts of a fixed whole. The parts are usually labeled as percents with the circle as a whole representing 100%. The graph in **Figure 2-13a** is based on the data in **Table 2-6**. What percent of the chlorine sources are natural? What percent are manufactured compounds? Which source supplies the most chlorine to the stratosphere?

Bar graph A bar graph often is used to show how a quantity varies with factors such as time, location, or temperature. In those cases, the quantity being measured appears on the vertical axis (*y*-axis). The independent variable appears on the horizontal axis (*x*-axis). The relative heights of the bars show how the quantity varies. A bar graph can be used to compare population figures for a country by decade. It can compare annual precipitation for different cities or average monthly precipitation for a single location, as in **Figure 2-13b**. The precipitation data was collected over a 30-year period (1961–1990). During which four months does Jacksonville receive about half of its annual precipitation?

Objectives

- **Create** graphs to reveal patterns in data.
- **Interpret** graphs.

Vocabulary

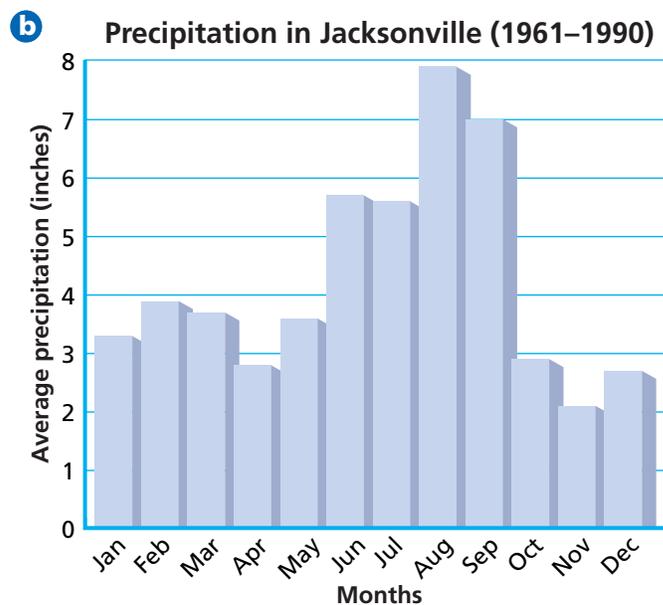
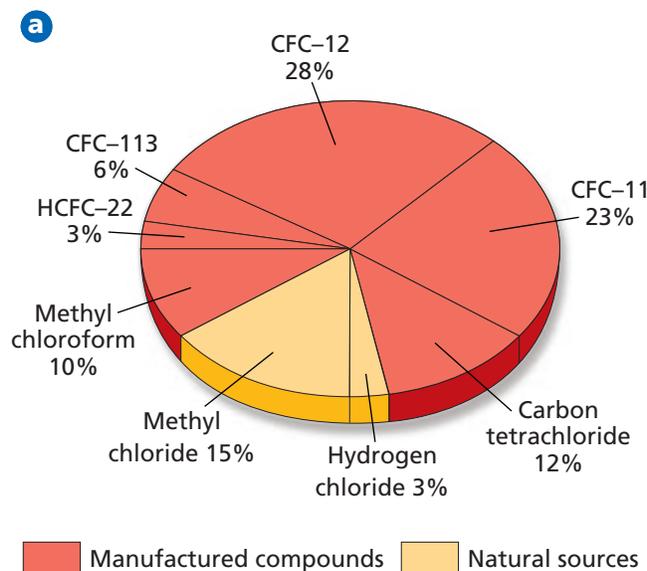
graph

Table 2-6

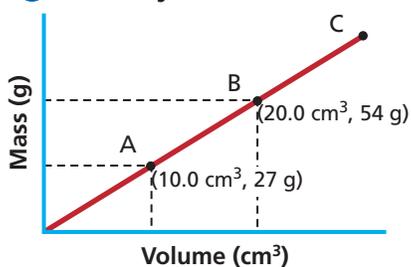
Sources of Chlorine in the Stratosphere	
Source	Percent
Hydrogen chloride	3
Methyl chloride	15
Carbon tetrachloride	12
Methyl chloroform	10
CFC-11	23
CFC-12	28
CFC-113	6
HCFC-22	3

Figure 2-13

What do circle graphs and bar graphs have in common?



a Density of Aluminum



b Temperature Versus Elevation

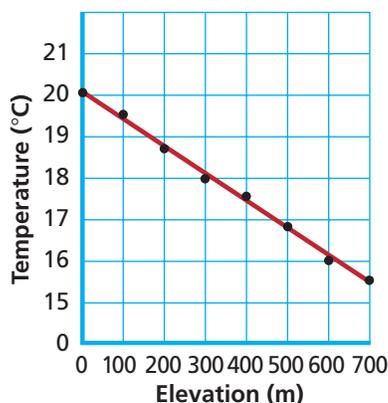


Figure 2-14

Compare the slopes of these two graphs.

Line Graphs

In chemistry, most graphs that you create and interpret will be line graphs. The points on a line graph represent the intersection of data for two variables. The independent variable is plotted on the x -axis. The dependent variable is plotted on the y -axis. Remember that the independent variable is the variable that a scientist deliberately changes during an experiment. In **Figure 2-14a**, the independent variable is volume and the dependent variable is mass. What are the values for the independent variable and the dependent variable at point B? **Figure 2-14b** is a graph of elevation versus temperature. Because the points are scattered, the line cannot pass through all the data points. The line must be drawn so that about as many points fall above the line as fall below it. This line is called a best fit line.

If the best fit line is straight, there is a linear relationship between the variables and the variables are *directly* related. This relationship can be further described by the steepness, or slope, of the line. If the line rises to the right, the slope is *positive*. A positive slope indicates that the dependent variable *increases* as the independent variable increases. If the line sinks to the right, the slope is *negative*. A negative slope indicates that the dependent variable *decreases* as the independent variable increases. Either way, the slope of the graph is constant. You can use the data points to calculate the slope of the line. The slope is the change in y divided by the change in x .

$$\text{slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{\Delta y}{\Delta x}$$

Calculate the slope for the line in **Figure 2-14a** using data points A and B.

$$\text{slope} = \frac{54 \text{ g} - 27 \text{ g}}{20.0 \text{ cm}^3 - 10.0 \text{ cm}^3} = \frac{27 \text{ g}}{10.0 \text{ cm}^3} = 2.7 \text{ g/cm}^3$$

When the mass of a material is plotted against volume, the slope of the line is the density of the material. Do the **CHEMLAB** at the end of the chapter to learn more about using a graph to find density.

problem-solving LAB

How does speed affect stopping distance?

Making and Using Graphs Use the steps below and the data to make a line graph.

Speed (m/s)	11	16	20	25	29
Stopping distance (m)	18	32	49	68	92

1. Identify the independent and dependent variables.
2. Determine the range of data that needs to be plotted for each axis. Choose intervals for the axes that spread out the data. Make each square on the graph a multiple of 1, 2, or 5.
3. Number and label each axis.

4. Plot the data points.
5. Draw a best fit line for the data. The line may be straight or it may be curved. Not all points may fall on the line.
6. Give the graph a title.

Analysis

What does the graph tell you about the relationship between speed and stopping distance?

Thinking Critically

There are two components to stopping distance: reaction distance (distance traveled before the driver applies the brake) and braking distance (distance traveled after the brake is applied). Predict which component will increase more rapidly as the speed increases. Explain your choice.

When the best fit line is curved, the relationship between the variables is nonlinear. In chemistry, you will study nonlinear relationships called inverse relationships. See pages 903–907 in the **Math Handbook** for more discussion of graphs. Do the **problem-solving LAB** to practice making line graphs.

Interpreting Graphs

An organized approach can help you understand the information on a graph. First, identify the independent and dependent variables. Look at the ranges of the data and consider what measurements were taken. Decide if the relationship between the variables is linear or nonlinear. If the relationship is linear, is the slope positive or negative? If a graph has multiple lines or regions, study one area at a time.

When points on a line graph are connected, the data is considered continuous. You can read data from a graph that falls between measured points. This process is called interpolation. You can extend the line beyond the plotted points and estimate values for the variables. This process is called extrapolation. Why might extrapolation be less reliable than interpolation?

Interpreting ozone data **Figure 2-15** is a graph of ozone measurements taken at a scientific settlement in Antarctica called Halley. The independent variable is months of the year. The dependent variable is total ozone measured in Dobson units (DU). The graph shows how ozone levels vary from August to April. There are two lines on the graph. Multiple lines allow scientists to introduce a third variable, in this case different periods of time. Having two lines on the same graph allows scientists to compare data gathered before the ozone hole developed with data from a recent season. They can identify a significant trend in ozone levels and verify the depletion in ozone levels over time.

The top line represents average ozone levels for the period 1957–1972. Follow the line from left to right. Ozone levels were about 300 DU in early October. By November, they rose to about 360 DU. Ozone levels slowly dropped back to around 290 DU by April. The bottom line shows the ozone levels from the 1999–2000 survey. The ozone levels were around 200 DU in August, dipped to about 150 DU during October, and slowly rose to a maximum of about 280 DU in January. At no point during this 9-month period were the ozone levels as high as they were at the corresponding points during 1957–1972. The “ozone hole” is represented by the dip in the bottom line. Based on the graph, was there an ozone hole in the 1957–1972 era?

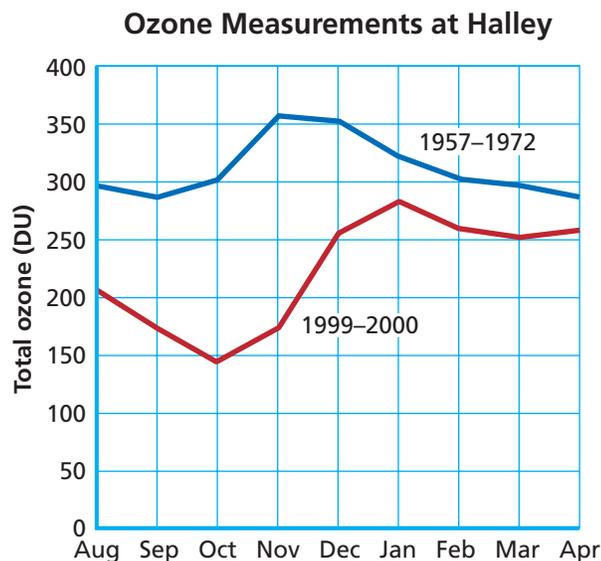


Figure 2-15

In Antarctica, spring begins in October and winter begins in April. Why are there no measurements from May to July?

Section 2.4 Assessment

45. Explain why graphing can be an important tool for analyzing data.
46. What type of data can be displayed on a circle graph? On a bar graph?
47. If a linear graph has a negative slope, what can you say about the dependent variable?
48. When can the slope of a graph represent density?
49. **Thinking Critically** Why does it make sense for the line in **Figure 2-14a** to extend to 0, 0 even though this point was not measured?
50. **Interpreting Graphs** Using **Figure 2-15**, determine how many months the ozone hole lasts.

Using Density to Find the Thickness of a Wire

The thickness of wire often is measured using a system called the American Wire Gauge (AWG) standard. The smaller the gauge number, the larger the diameter of the wire. For example, 18-gauge copper wire has a diameter of about 0.102 cm; 12-gauge copper wire has a diameter of about 0.205 cm. Such small diameters are difficult to measure accurately with a metric ruler. In this experiment, you will plot measurements of mass and volume to find the density of copper. Then, you will use the density of copper to confirm the gauge of copper wire.

Problem

How can density be used to verify the diameter of copper wire?

Objectives

- **Collect** and **graph** mass and volume data to find the density of copper.
- **Measure** the length and volume of a copper wire, and **calculate** its diameter.
- **Calculate** percent errors for the results.

Materials

tap water
100-mL graduated cylinder
small cup, plastic
balance
copper shot
copper wire (12-gauge, 18-gauge)

metric ruler
pencil
graph paper
graphing calculator (optional)

Safety Precautions



- Always wear safety goggles and a lab apron.

Pre-Lab

1. Read the entire CHEMLAB.
2. What is the equation used to calculate density?
3. How can you find the volume of a solid that has an irregular shape?
4. What is a meniscus and how does it affect volume readings?
5. If you plot mass versus volume, what property of matter will the slope of the graph represent?
6. How do you find the slope of a graph?
7. A piece of copper wire is a narrow cylinder. The equation for the volume of a cylinder is

$$V = \pi r^2 h$$

where V is the volume, r is the radius, h is the height, and π (pi) is a constant with a value of 3.14. Rearrange the equation to solve for r .

8. What is the relationship between the diameter and the radius of a cylinder?
9. Prepare two data tables.

Density of Copper

Trial	Mass of copper added	Total mass of copper	Total volume of water displaced
1			
2			
3			
4			

Diameter of Copper Wire

	12-gauge	18-gauge
Length		
Mass		
Measured diameter		
Calculated diameter		

Procedure

Record all measurements in your data tables.

1. Pour about 20 mL of water into a 100-mL graduated cylinder. Read the actual volume.
2. Find the mass of the plastic cup.
3. Add about 10 g of copper shot to the cup and find the mass again.
4. Pour the copper shot into the graduated cylinder and read the new volume.
5. Repeat steps 3 and 4 three times. By the end of the four trials, you will have about 40 g of copper in the graduated cylinder.
6. Obtain a piece of 12-gauge copper wire and a piece of 18-gauge copper wire. Use a metric ruler to measure the length and diameter of each wire.
7. Wrap each wire around a pencil to form a coil. Remove the coils from the pencil. Find the mass of each coil.



Cleanup and Disposal

1. Carefully drain off most of the water from the graduated cylinder. Make sure all of the copper shot remains in the cylinder.
2. Pour the copper shot onto a paper towel to dry. Both the copper shot and wire can be reused.

Analyze and Conclude

1. **Using Numbers** Complete the table for the density of copper by calculating the total mass of copper and the total water displaced for each trial.
2. **Making and Using Graphs** Graph total mass versus total volume of copper. Draw a line that best fits the points. Then use two points on your line to find the slope of your graph. Because density equals mass divided by volume, the slope will give you the density of copper. If you are using a graphing calculator, select the 5:FIT CURVE option from the MAIN MENU of the ChemBio program. Choose 1:LINEAR L1,L2 from the REGRESSION/LIST to help you plot and calculate the slope of the graph.
3. **Using Numbers** Calculate the percent error for your value of density.
4. **Using Numbers** To complete the second data table, you must calculate the diameter for each wire. Use the accepted value for the density of copper and the mass of each wire to calculate volume. Then use the equation for the volume of a cylinder to solve for the radius. Double the radius to find the diameter.
5. **Comparing and Contrasting** How do your calculated values for the diameter compare to your measured values and to the AWG values listed in the introduction?
6. **Error Analysis** How could you change the procedure to reduce the percent error for density?

Real-World Chemistry

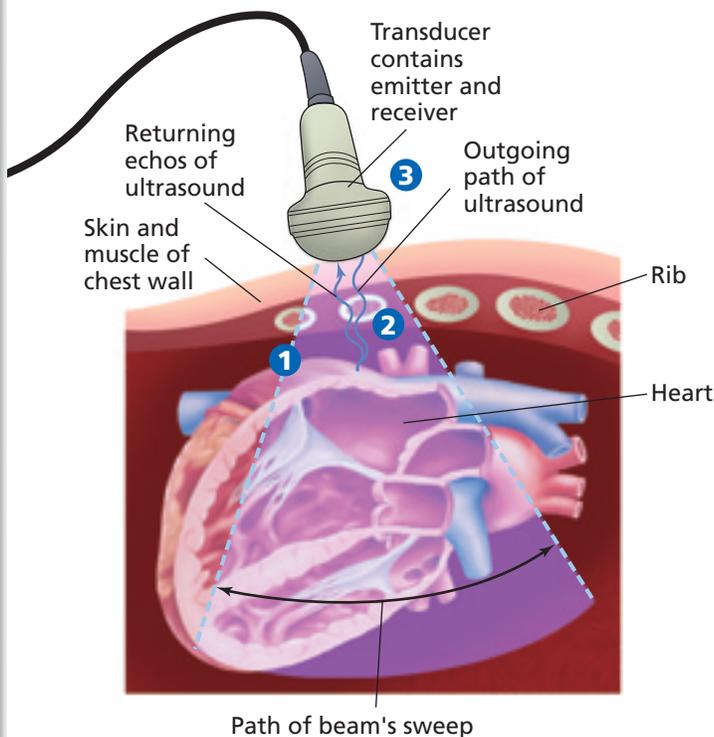
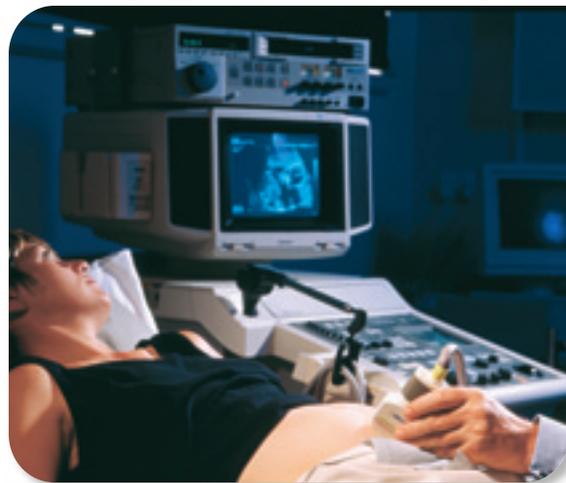
1. There is a standard called the British Imperial Standard Wire Gauge (SWG) that is used in England and Canada. Research the SWG standard to find out how it differs from the AWG standard. Are they the only standards used for wire gauge?
2. Interview an electrician or a building inspector who reviews the wiring in new or remodeled buildings. Ask what the codes are for the wires used and how the diameter of a wire affects its ability to safely conduct electricity. Ask to see a wiring diagram.

How It Works

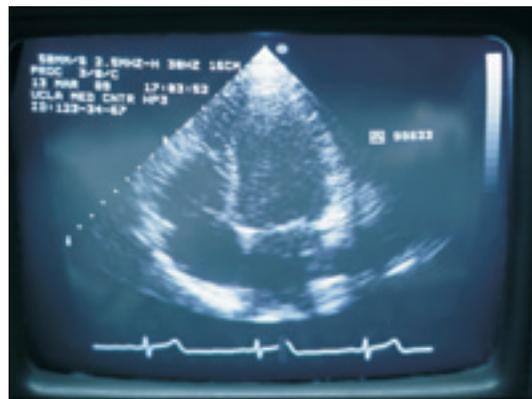
Ultrasound Devices

An ultrasound device is a diagnostic tool that allows doctors to see inside the human body without having to perform surgery. With an ultrasound device, doctors can detect abnormal growths, follow the development of a fetus in the uterus, or study the action of heart valves.

An ultrasound device emits high-frequency sound waves that can pass through a material, be absorbed, or reflect off the surface of a material. Waves are reflected at the border between tissues with different densities, such as an organ and a tumor. The larger the difference in density, the greater the reflection.



- 1 In the transducer, electrical pulses are changed into sound waves, which are aimed at a specific part of the body.
- 2 As the transducer is moved across the body, some sound is reflected back as echoes.
- 3 A receiver detects the reflected waves and converts the sound back into electrical pulses.



- 4 A computer analyzes the data and creates an image of an internal organ, such as the heart.

Thinking Critically

1. **Predicting** If all parts of the heart had the same density, would doctors be able to use ultrasound to detect heart defects? Explain.
2. **Inferring** Why is it considered safe to use ultrasound but not X rays during pregnancy?

Summary

2.1 Units of Measurement

- SI measurement units allow scientists to report data that can be reproduced by other scientists.
- Adding prefixes to SI units extends the range of possible measurements.
- SI base units include the meter for length, the second for time, the kilogram for mass, and the kelvin for temperature.
- Volume and density have derived units. Density is the ratio of mass to volume. Density can be used to identify a sample of matter.

2.2 Scientific Notation and Dimensional Analysis

- Scientific notation makes it easier to handle extremely large or small measurements.
- Numbers expressed in scientific notation are a product of two factors: (1) a number between 1 and 10 and (2) ten raised to a power.
- Numbers added or subtracted in scientific notation must be expressed to the same power of ten.
- When measurements are multiplied or divided in scientific notation, their exponents are added or subtracted, respectively.
- Dimensional analysis often uses conversion factors to solve problems that involve units. A conversion factor is a ratio of equivalent values.

2.3 How reliable are measurements?

- An accurate measurement is close to the accepted value. Precise measurements show little variation over a series of trials.
- The type of measurement instrument determines the degree of precision possible.
- Percent error compares the size of an error in experimental data to the size of the accepted value.
- The number of significant figures reflects the precision of reported data. Answers to calculations are rounded off to maintain the correct number of significant figures.

2.4 Representing Data

- Graphs are visual representations of data. Graphs can reveal patterns in data.
- Circle graphs show parts of a whole. Bar graphs can show how a factor varies with time, location, or temperature.
- The relationship between the independent and dependent variables on a line graph can be linear or nonlinear.
- Because the data on a line graph are considered continuous, you can interpolate or extrapolate data from a line graph.

Key Equations and Relationships

• density: $\text{density} = \frac{\text{mass}}{\text{volume}}$
(p. 28)

• conversion between temperature scales: $^{\circ}\text{C} + 273 = \text{K}$
 $\text{K} - 273 = ^{\circ}\text{C}$
(p. 30)

• percent error:
 $\text{percent error} = \frac{\text{error}}{\text{accepted value}} \times 100$
(p. 37)

• slope of graph: $\text{slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{\Delta y}{\Delta x}$
(p. 44)

Vocabulary

- accuracy (p. 36)
- base unit (p. 26)
- conversion factor (p. 34)
- density (p. 27)
- derived unit (p. 27)
- dimensional analysis (p. 34)
- graph (p. 43)
- kelvin (p. 30)
- kilogram (p. 27)
- liter (p. 27)
- meter (p. 26)
- percent error (p. 37)
- precision (p. 36)
- scientific notation (p. 31)
- second (p. 26)
- significant figure (p. 38)

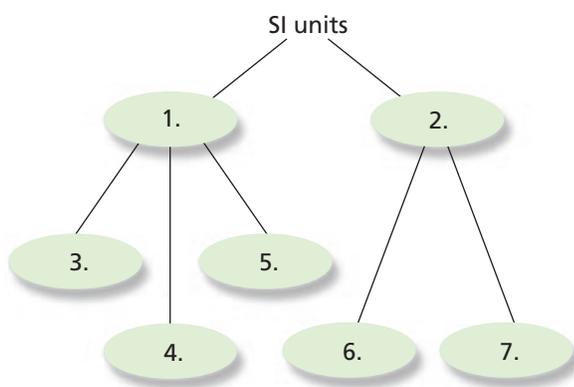


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 2 Assessment.

Concept Mapping

51. Use the following terms to complete the concept map: volume, derived unit, mass, density, base unit, time, length.



Mastering Concepts

52. Why must a measurement include both a number and a unit? (2.1)
53. Explain why scientists, in particular, need standard units of measurement. (2.1)
54. What role do prefixes play in the metric system? (2.1)
55. How many meters are there in one kilometer? In one decimeter? (2.1)
56. What is the relationship between the SI unit for volume and the SI unit for length? (2.1)
57. Explain how temperatures on the Celsius and Kelvin scales are related. (2.1)
58. How does scientific notation differ from ordinary notation? (2.2)
59. If you move the decimal place to the left to convert a number into scientific notation, will the power of ten be positive or negative? (2.2)
60. When dividing numbers in scientific notation, what must you do with the exponents? (2.2)
61. When you convert from a small unit to a large unit, what happens to the number of units? (2.2)

62. If you report two measurements of mass, 7.42 g and 7.56 g, are the measurements accurate? Are they precise? Explain your answers. (2.3)
63. When converting from meters to centimeters, how do you decide which values to place in the numerator and denominator of the conversion factor? (2.2)
64. Why are plus and minus signs ignored in percent error calculations? (2.3)
65. In 50 540, which zero is significant? What is the other zero called? (2.3)
66. Which of the following three numbers will produce the same number when rounded to three significant figures: 3.456, 3.450, or 3.448? (2.3)
67. When subtracting 61.45 g from 242.6 g, which factor determines the number of significant figures in the answer? Explain. (2.3)
68. When multiplying 602.4 m by 3.72 m, which factor determines the number of significant figures in the answer? Explain. (2.3)
69. Which type of graph would you choose to depict data on how many households heat with gas, oil, or electricity? Explain. (2.4)
70. Which type of graph would you choose to depict changes in gasoline consumption over a period of ten years? Explain. (2.4)
71. How can you find the slope of a line graph? (2.4)

Mastering Problems

Density (2.1)

72. A 5-mL sample of water has a mass of 5 g. What is the density of water?
73. An object with a mass of 7.5 g raises the level of water in a graduated cylinder from 25.1 mL to 30.1 mL. What is the density of the object?
74. The density of aluminum is 2.7 g/mL. What is the volume of 8.1 g?

Scientific Notation (2.2)

75. Write the following numbers in scientific notation.
- 0.004 583 4 mm
 - 0.03054 g
 - 438 904 s
 - 7 004 300 000 g
76. Write the following numbers in ordinary notation.
- 8.348×10^6 km
 - 3.402×10^3 g
 - 7.6352×10^{-3} kg
 - 3.02×10^{-5} s

77. Complete the following addition and subtraction problems in scientific notation.

- a. $6.23 \times 10^6 \text{ kL} + 5.34 \times 10^6 \text{ kL}$
- b. $3.1 \times 10^4 \text{ mm} + 4.87 \times 10^5 \text{ mm}$
- c. $7.21 \times 10^3 \text{ mg} + 43.8 \times 10^2 \text{ mg}$
- d. $9.15 \times 10^{-4} \text{ cm} + 3.48 \times 10^{-4} \text{ cm}$
- e. $4.68 \times 10^{-5} \text{ cg} + 3.5 \times 10^{-6} \text{ cg}$
- f. $3.57 \times 10^2 \text{ mL} - 1.43 \times 10^2 \text{ mL}$
- g. $9.87 \times 10^4 \text{ g} - 6.2 \times 10^3 \text{ g}$
- h. $7.52 \times 10^5 \text{ kg} - 5.43 \times 10^5 \text{ kg}$
- i. $6.48 \times 10^{-3} \text{ mm} - 2.81 \times 10^{-3} \text{ mm}$
- j. $5.72 \times 10^{-4} \text{ dg} - 2.3 \times 10^{-5} \text{ dg}$

78. Complete the following multiplication and division problems in scientific notation.

- a. $(4.8 \times 10^5 \text{ km}) \times (2.0 \times 10^3 \text{ km})$
- b. $(3.33 \times 10^{-4} \text{ m}) \times (3.00 \times 10^{-5} \text{ m})$
- c. $(1.2 \times 10^6 \text{ m}) \times (1.5 \times 10^{-7} \text{ m})$
- d. $(8.42 \times 10^8 \text{ kL}) \div (4.21 \times 10^3 \text{ kL})$
- e. $(8.4 \times 10^6 \text{ L}) \div (2.4 \times 10^{-3} \text{ L})$
- f. $(3.3 \times 10^{-4} \text{ mL}) \div (1.1 \times 10^{-6} \text{ mL})$

Conversion Factors (2.2)

79. Write the conversion factor that converts

- a. grams to kilograms
- b. kilograms to grams
- c. millimeters to meters
- d. meters to millimeters
- e. milliliters to liters
- f. centimeters to meters

80. Convert the following measurements.

- a. 5.70 g to milligrams
- b. 4.37 cm to meters
- c. 783 kg to grams
- d. 45.3 mm to meters
- e. 10 m to centimeters
- f. 37.5 g/mL to kg/L

Percent Error (2.3)

81. The accepted length of a steel pipe is 5.5 m. Calculate the percent error for each of these measurements.

- a. 5.2 m
- b. 5.5 m
- c. 5.7 m
- d. 5.1 m

82. The accepted density for copper is 8.96 g/mL. Calculate the percent error for each of these measurements.

- a. 8.86 g/mL
- b. 8.92 g/mL
- c. 9.00 g/mL
- d. 8.98 g/mL

Significant Figures (2.3)

83. Round each number to four significant figures.

- a. 431 801 kg
- b. 10 235.0 mg
- c. 1.0348 m
- d. 0.004 384 010 cm
- e. 0.000 781 00 mL
- f. 0.009 864 1 cg

84. Round each figure to three significant figures.

- a. 0.003 210 g
- b. 3.8754 kg
- c. 219 034 m
- d. 25.38 L
- e. 0.087 63 cm
- f. 0.003 109 mg

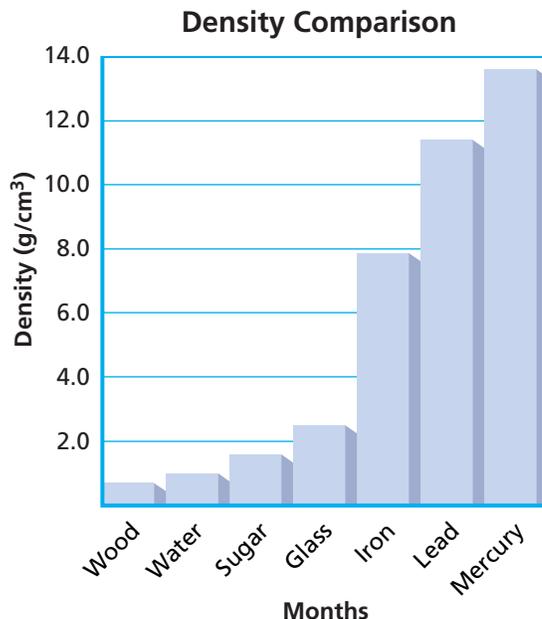
85. Round the answers to each of the following problems to the correct number of significant figures.

- a. $7.31 \times 10^4 + 3.23 \times 10^3$
- b. $8.54 \times 10^{-3} - 3.41 \times 10^{-4}$
- c. $4.35 \text{ dm} \times 2.34 \text{ dm} \times 7.35 \text{ dm}$
- d. $4.78 \text{ cm} + 3.218 \text{ cm} + 5.82 \text{ cm}$
- e. $3.40 \text{ mg} + 7.34 \text{ mg} + 6.45 \text{ mg}$
- f. $45 \text{ m} \times 72 \text{ m} \times 132 \text{ m}$
- g. $38736 \text{ km}/4784 \text{ km}$

Representing Data (2.4)

86. Use the accompanying bar graph to answer the following questions.

- a. Which substance has the greatest density?
- b. Which substance has the least density?
- c. Which substance has a density of 7.87 g/cm^3 ?
- d. Which substance has a density of 11.4 g/cm^3 ?





- 87.** Graph the following data with the volume on the x-axis and the mass on the y-axis. Then calculate the slope of the line.

Table 2-7

Density Data	
Volume (mL)	Mass (g)
2.0 mL	5.4
4.0 mL	10.8
6.0 mL	16.2
8.0 mL	21.6
10.0 mL	27.0

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 88.** You have a 23-g sample of ethanol with a density of 0.7893 g/mL. What volume of ethanol do you have?
- 89.** Complete the following problems in scientific notation. Round off to the correct number of significant figures.
- $(5.31 \times 10^{-2} \text{ cm}) \times (2.46 \times 10^5 \text{ cm})$
 - $(3.78 \times 10^3 \text{ m}) \times (7.21 \times 10^2 \text{ m})$
 - $(8.12 \times 10^{-3} \text{ m}) \times (1.14 \times 10^{-5} \text{ m})$
 - $(5.53 \times 10^{-6} \text{ km}) \times (7.64 \times 10^3 \text{ km})$
 - $(9.33 \times 10^4 \text{ mm}) \div (3.0 \times 10^2 \text{ mm})$
 - $(4.42 \times 10^{-3} \text{ kg}) \div (2.0 \times 10^2 \text{ kg})$
 - $(6.42 \times 10^{-2} \text{ g}) \div (3.21 \times 10^{-3} \text{ g})$
- 90.** Evaluate the following conversion. Will the answer be correct? Explain.
- $$\text{rate} = \frac{75 \text{ m}}{1 \text{ s}} \times \frac{60 \text{ s}}{1 \text{ min}} = \frac{1 \text{ h}}{60 \text{ min}}$$
- 91.** What mass of lead (density 11.4 g/cm³) would have an identical volume to 15.0 g of mercury (density 13.6 g/cm³)?
- 92.** Three students use a meterstick to measure a length of wire. One student records a measurement of 3 cm. The second records 3.3 cm. The third records 2.87 cm. Explain which answer was recorded correctly.
- 93.** Express each quantity in the unit listed to its right.
- | | | | |
|----------------------------|----|---------------|-----------------|
| a. 3.01 g | cg | d. 0.2 L | dm ³ |
| b. 6200 m | km | e. 0.13 cal/g | kcal/g |
| c. 6.24×10^{-7} g | μg | f. 3.21 mL | L |

- 94.** The black hole in the galaxy M82 has a mass about 500 times the mass of our Sun. It has about the same volume as Earth's moon. What is the density of this black hole?

$$\text{mass}_{\text{sun}} = 1.9891 \times 10^{30} \text{ kg}$$

$$\text{volume}_{\text{moon}} = 2.1968 \times 10^{10} \text{ km}^3$$

- 95.** The density of water is 1 g/cm³. Use your answer to question 94 to compare the densities of water and a black hole.

Thinking Critically

- 96. Comparing and Contrasting** What advantages do SI units have over the units in common use in the United States? Is there any disadvantage to using SI units?
- 97. Forming a Hypothesis** Why do you think the SI standard for time was based on the distance light travels through a vacuum?
- 98. Inferring** Explain why the mass of an object cannot help you identify what material the object is made from.
- 99. Drawing Conclusions** Why might property owners hire a surveyor to determine property boundaries rather than measure the boundaries themselves?

Writing in Chemistry

- 100.** Although the standard kilogram is stored at constant temperature and humidity, unwanted matter can build up on its surface. Scientists have been looking for a more reliable standard for mass. Research and describe alternate standards that have been proposed. Find out why no alternate standard has been chosen.
- 101.** Research and report on some unusual units of measurement such as bushels, pecks, firkins, and frails.
- 102.** Research the range of volumes used for packaging liquids sold in supermarkets.
- 103.** Find out what the acceptable limits of error are for some manufactured products or for the doses of medicine given at a hospital.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 104.** You record the following in your lab book: A liquid is thick and has a density of 4.58 g/mL. Which data is qualitative? Which is quantitative? (Chapter 1)



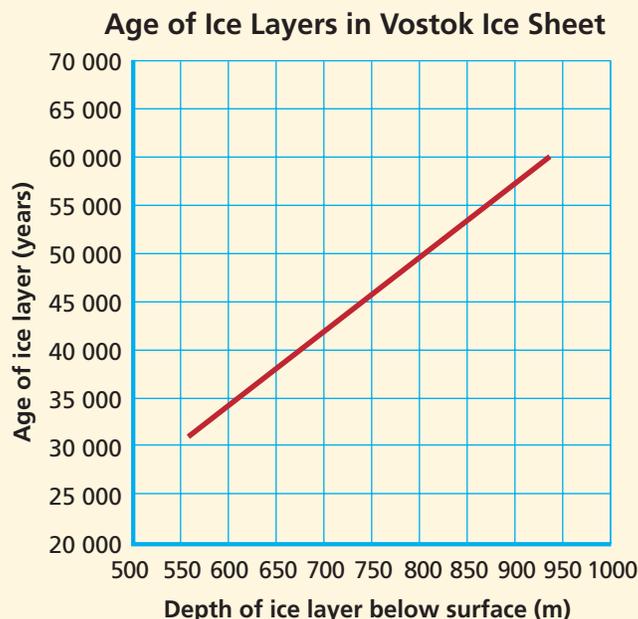
Use the questions and the test-taking tip to prepare for your standardized test.

- Which of the following is *not* an SI base unit?
 - second (s)
 - kilogram (kg)
 - degrees Celsius (°C)
 - meter (m)
- Which of the following values is NOT equivalent to the others?
 - 500 meters
 - 0.5 kilometers
 - 5000 centimeters
 - 5×10^{11} nanometers
- The correct representation of 702.0 g using scientific notation is _____.
 - 7.02×10^3 g
 - 70.20×10^1 g
 - 7.020×10^2 g
 - 70.20×10^2 g
- Three students measured the length of a stamp whose accepted length is 2.71 cm. Based on the table, which statement is true?
 - Student 2 is both precise and accurate.
 - Student 1 is more accurate than Student 3.
 - Student 2 is less precise than Student 1.
 - Student 3 is both precise and accurate.

Measured Values for a Stamp's Length			
	Student 1	Student 2	Student 3
Trial 1	2.60 cm	2.70 cm	2.75 cm
Trial 2	2.72 cm	2.69 cm	2.74 cm
Trial 3	2.65 cm	2.71 cm	2.64 cm
Average	2.66 cm	2.70 cm	2.71 cm

- Chemists found that a complex reaction occurred in three steps. The first step takes 2.5731×10^2 s to complete, the second step takes 3.60×10^{-1} s, and the third step takes 7.482×10^1 s. The total amount of time elapsed during the reaction is _____.
 - 3.68×10^1 s
 - 7.78×10^1 s
 - 1.37×10^1 s
 - 3.3249×10^2 s
- How many significant figures are there in a distance measurement of 20.070 km?
 - 2
 - 3
 - 4
 - 5

Interpreting Graphs Use the graph to answer the following questions.



- Using the graph, a student reported the age of an ice layer at 705 m as 4.250×10^4 years. The accepted value for the age of this ice layer is 4.268×10^4 years. The percent error of the student's value is _____.
 - 0.4217%
 - 99.58%
 - 0.4235%
 - 1.800%
- The slope of the graph is about _____.
 - 80 years/m
 - 80 m/year
 - 0.015 years/m
 - 1500 m/year
- What age is an ice layer found at a depth of 1000 m?
 - 6.75×10^4 years
 - 7.00×10^4 years
 - 6.25×10^4 years
 - 6.5×10^4 years



TEST-TAKING TIP

Practice Under Test-Like Conditions

Ask your teacher to set a time limit. Then do all of the questions in the time provided without referring to your book. Did you complete the test? Could you have made better use of your time? What topics do you need to review? Show your test to your teacher for an objective assessment of your performance.

Matter—Properties and Changes

What You'll Learn

- ▶ You will distinguish between physical and chemical properties.
- ▶ You will classify matter by composition: element, compound, or mixture.
- ▶ You will identify observable characteristics of chemical reactions.
- ▶ You will explain the fundamental law of conservation of mass.

Why It's Important

You are completely surrounded by matter. To better understand this matter—how it affects you, how you affect it, and how it can be manipulated for the benefit of society—you need to build a basic understanding of the types and properties of matter.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about matter.

Chemistry is the study of matter and its properties. Every aspect of these divers' environment, under water and on land, is some form of matter.



DISCOVERY LAB



Materials

large test tube
test-tube holder or rack
10 mL HCl
zinc metal
wood splint
match or burner

Observing Chemical Change

Consider the metal objects that are part of the everyday world. A mailbox, for example, stands outside day in and day out, without seeming to change. Under what conditions does metal exhibit chemical change?

Safety Precautions



Always wear eye goggles, gloves, and an apron when experimenting with chemicals. Use caution when handling an open flame.

Procedure

1. Place a piece of zinc metal in a large test tube.
2. Add approximately 10 mL of 3M hydrochloric acid (HCl) to the test tube. Record your observations.
CAUTION: HCl causes burns and hazardous fumes.
3. When the zinc and HCl have reacted for approximately 1 min, bring a lighted, glowing wood splint to the mouth of the test tube. **CAUTION:** Be sure the test tube is facing away from your face when the splint is brought near. Again record your observations.

Analysis

What may have caused the dynamic reaction you observed in step 3? Did you expect this reaction? Explain.

Section

3.1

Properties of Matter

Objectives

- **Identify** the characteristics of a substance.
- **Distinguish** between physical and chemical properties.
- **Differentiate** among the physical states of matter.

Vocabulary

substance
physical property
extensive property
intensive property
chemical property
states of matter
solid
liquid
gas
vapor

Imagine yourself scuba diving through a complex biological ecosystem such as a coral reef. What kinds of things fill your imagination? Regardless of what you envision, there is only one answer—you see matter. The diversity of matter in the world and in the universe is astounding. From pepperoni pizzas to supernovas, it's all matter. If we are to understand this diversity, we must start with a way of organizing and describing matter.

Substances

Recall from Chapter 1 that chemistry is the study of matter, and matter is anything that has mass and takes up space. Everything around you is matter; including things such as air and microbes, which you cannot see. For example, table salt is a simple type of matter that you are probably familiar with. Table salt has a unique and unchanging chemical composition. It is always 100% sodium chloride and its composition does not change from one sample to another. Matter that has a uniform and unchanging composition is called a **substance**, also known as a pure substance. Table salt is a substance. Another example of a pure substance is water. Water is always composed of hydrogen and oxygen. Seawater, on the other hand, is not a substance because samples taken from different locations will probably have

Careers Using Chemistry

Science Writer

Do you get excited about news in science and technology? Do you like to explain information in a way that others find interesting and understandable? Then consider a career as a science writer.

Science writers keep up-to-date on what is happening in the world of science and translate that news so nonscientists can understand it. These writers work for newspapers, magazines, scientific publications, television stations, and Internet news services. Lots of curiosity, as well as a degree in a science and/or journalism, is essential.

differing compositions. That is, they will contain differing amounts of water, salts, and other dissolved substances. Given this definition, what other pure substances are you familiar with? Substances are important; much of your chemistry course will be focused on the processes by which substances are changed into different substances.

Physical Properties of Matter

You are used to identifying objects by their properties—their characteristics and behavior. For example, you can easily identify a pencil in your backpack because you recognize its shape, color, weight, or some other property. These characteristics are all physical properties of the pencil. A **physical property** is a characteristic that can be observed or measured without changing the sample's composition. Physical properties describe pure substances, too. Because substances have uniform and unchanging compositions, they have consistent and unchanging physical properties as well. Density, color, odor, taste, hardness, melting point, and boiling point are common physical properties that scientists record as identifying characteristics of a substance. Sodium chloride forms solid, white crystals at room temperature, all having the same unique salty taste. **Table 3-1** lists several common substances and their physical properties.

Table 3-1

Physical Properties of Common Substances					
Substance	Color	State at 25°C	Melting point (°C)	Boiling point (°C)	Density (g/cm ³)
Oxygen	Colorless	Gas	-218	-183	0.0014
Mercury	Silver	Liquid	-39	357	13.5
Water	Colorless	Liquid	0	100	1.00
Sucrose	White	Solid	185	Decomposes	1.59
Sodium chloride	White	Solid	801	1413	2.17

Figure 3-1

Miners relied on the physical property of density to distinguish gold (19 g/cm³) from the worthless minerals in their sluice pans. The density of pyrite, a worthless mineral often mistaken for gold, is 5 g/cm³.



Extensive and intensive properties Physical properties can be further described as being one of two types. **Extensive properties** are dependent upon the amount of substance present. For example, mass, which depends on the amount of substance there is, is an extensive property. Length and volume are also extensive properties. Density, on the other hand, is an example of an intensive property of matter. **Intensive properties** are independent of the amount of substance present. For example, density of a substance (at constant temperature and pressure) is the same no matter how much substance is present.

A substance can often be identified by its intensive properties. In some cases, a single intensive property is unique enough for identification. During the California gold rush, miners relied on gold's characteristic density (19 g/cm³) to separate valuable gold-containing flakes from riverbed sand. The process used by the miners is shown in **Figure 3-1**. Another intensive property of gold is its distinctive appearance. Unfortunately, miners often learned that identification of gold based on appearance alone was misleading. **Figure 3-2** shows a nugget of the relatively worthless

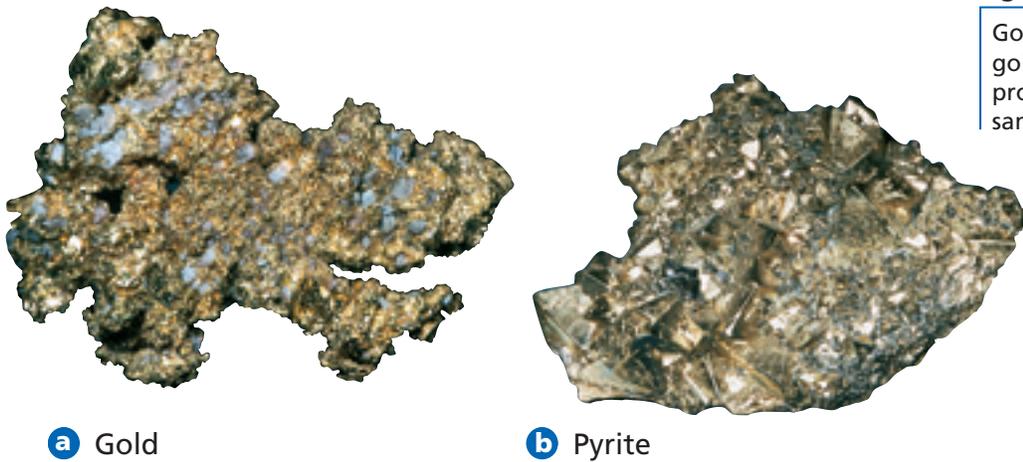


Figure 3-2

Gold **a** and pyrite, or "fool's gold" **b**, have similar physical properties but are different samples of matter.

mineral pyrite, often called "fool's gold," which looks very similar to actual gold nuggets. Such errors in identification based on the intensive property of appearance fooled many miners into falsely thinking they had struck it rich!

Chemical Properties of Matter

Some properties of a substance are not obvious unless the substance has changed composition as a result of its contact with other substances or the application of thermal or electrical energy. The ability of a substance to combine with or change into one or more other substances is called a **chemical property**. The ability of iron to form rust when combined with air is an example of a chemical property of iron. Similarly, the inability of a substance to change into another substance is also a chemical property. For example, when iron is placed in nitrogen gas at room temperature, no chemical change occurs. The fact that iron does not undergo a change in the presence of nitrogen is another chemical property of iron.

Observing Properties of Matter

Every substance has its own unique set of physical and chemical properties. **Table 3-2** lists several of these properties of copper. **Figure 3-3** shows physical and chemical properties of copper. What physical and chemical properties are evident in these photos?

Figure 3-3

These photos illustrate some of the physical and chemical properties of copper as it exists in the form of hardware **a** and the Statue of Liberty **b**.



Table 3-2

Properties of Copper	
Physical properties	Chemical properties
<ul style="list-style-type: none"> • Reddish brown, shiny • Easily shaped into sheets (malleable) and drawn into wires (ductile) • Good conductor of heat and electricity • Density = 8.92 g/cm³ • Melting point = 1085°C • Boiling point = 2570°C 	<ul style="list-style-type: none"> • Forms green copper carbonate compound when in contact with moist air • Forms new substances when combined with nitric acid and sulfuric acid • Forms a deep blue solution when in contact with ammonia

Observations of properties may vary depending on the conditions of the immediate environment. It is important to state the specific conditions in which observations are made because both chemical and physical properties depend on temperature and pressure. Consider the properties of water, for example. You may think of water as a liquid (physical property) that is not particularly chemically reactive (chemical property). You may also know that water has a density of 1.00 g/cm^3 (physical property). These properties, however, apply only to water at standard “room” temperature and pressure. At temperatures greater than 100°C , water is a gas (physical property) with a density of about 0.0006 g/cm^3 (physical property) that reacts rapidly with many different substances (chemical property). As you can see, the properties of water are dramatically different under different conditions.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

States of Matter

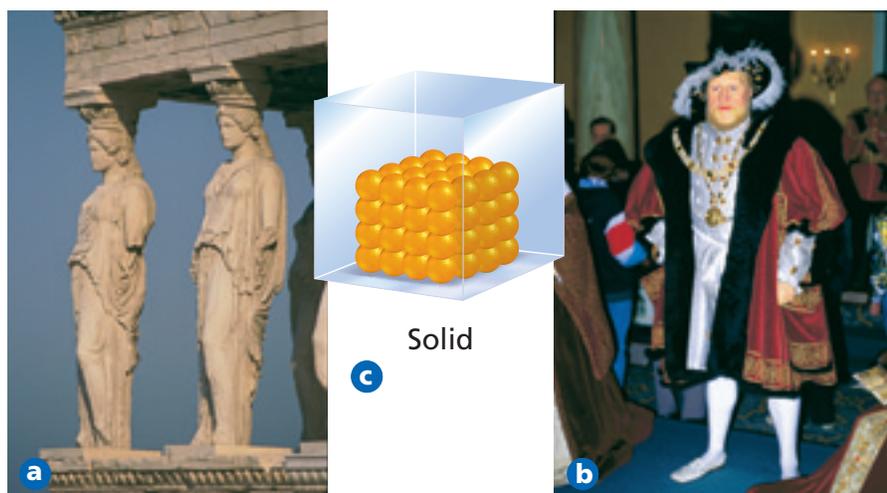
Imagine you are sitting on a bench, breathing heavily and drinking water after a tiring game of soccer. In this scenario, you are in contact with three different forms of matter; the bench is a solid, the water is a liquid, and the air you breathe is a gas. In fact, all matter that exists on Earth can be classified as one of these physical forms called **states of matter**. Scientists recognize a fourth state of matter called plasma, but it does not occur naturally on Earth except in the form of lightning bolts. The physical state of a substance is a physical property of that substance. Each of the three common states of matter can be distinguished by the way it fills a container.

Solids A **solid** is a form of matter that has its own definite shape and volume. Wood, iron, paper, and sugar are examples of solids. The particles of matter in a solid are very tightly packed; when heated, a solid expands, but only slightly. Because its shape is definite, a solid may not conform to the shape of the container in which it is placed. The tight packing of particles in a solid makes it incompressible; that is, it cannot be pressed into a smaller volume. It is important to understand that a solid is not defined by its rigidity or hardness; the marble statue in **Figure 3-4** is rigid whereas wax sculpture is soft, yet both are solids.

Liquids A **liquid** is a form of matter that flows, has constant volume, and takes the shape of its container. Common examples of liquids include water, blood, and mercury. The particles in a liquid are not rigidly held in place and are less closely packed than are the particles in a solid: liquid particles

Figure 3-4

The properties of the solid materials marble **a** and wax **b** make these sculptures possible. Particles in a solid are tightly packed **c**, giving definite shape and volume to the solid.



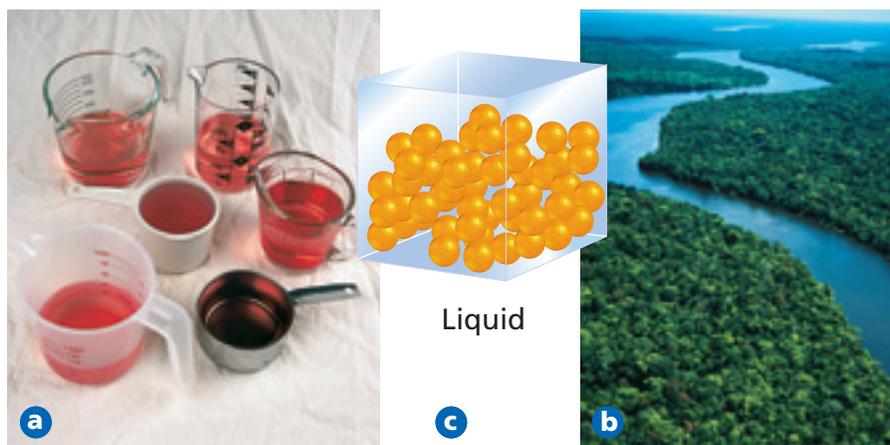


Figure 3-5

- a** Despite having different shapes, each of these measuring cups holds the same volume of liquid.
- b** River water flows to fit within the boundaries of its banks, regardless of the curves along its path.
- c** Molecules in a liquid are closely packed but can still move relatively freely.

are able to move past each other. This allows a liquid to flow and take the shape of its container, although it may not completely fill the container. A liquid's volume is constant: regardless of the size and shape of the container in which the liquid is held, the volume of the liquid remains the same. This is why measuring cups used in cooking, such as those pictured in **Figure 3-5**, can be made in a variety of shapes yet still measure the same volume. Because of the way the particles of a liquid are packed, liquids are virtually incompressible. Like solids, liquids tend to expand when heated.

Gases A **gas** is a form of matter that flows to conform to the shape of its container and fills the entire volume of its container. Examples of gases include neon, which is used in the lighted artwork in **Figure 3-6**; methane, which is used in cooking; and air, which is a mixture of gases. Compared to solids and liquids, the particles of gases are very far apart. Because of the significant amount of space between particles, gases are easily compressed. The **problem-solving LAB** in this section poses several important questions about the practical use of compressed gas.

It is likely that you are familiar with the word vapor as it relates to the word gas. The words gas and vapor, while similar, do not mean the same thing and should not be used interchangeably. The word gas refers to a substance that is naturally in the gaseous state at room temperature. The word **vapor** refers to the gaseous state of a substance that is a solid or a liquid at room temperature. For example, steam is a vapor because at room temperature water exists as a liquid.

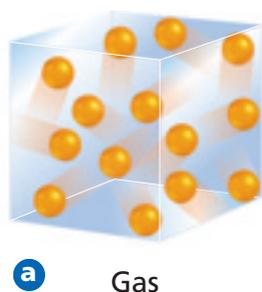


Figure 3-6

- a** Molecules in a gas are far apart and freely moving.
- b** Neon gas completely fills the tubes of the electric artwork.

How is compressed gas released?

Recognizing Cause and Effect Tanks of compressed gases are a common sight in a chemistry laboratory. For example, nitrogen is often flowed over a reaction in progress to displace other gases that might interfere with the experiment. Given what you know about the properties of gases, explain how compressed nitrogen is released.

Analysis

By definition, the particles of gases are far apart and gases tend to fill their container, even if the container is a laboratory room. Tanks of compressed gas come from the supplier capped to prevent the gas from escaping. In the lab a chemist or technician attaches a regulator to the tank and secures the tank to a stable fixture.



Thinking Critically

1. Explain why the flow of compressed gas must be controlled for practical use.
2. Predict what would happen if the valve on a full tank of compressed gas were suddenly opened all the way or if the full tank were punctured.

The fact that substances can change form, as in the example of water changing to steam, is another important concept in chemistry. If you review what you just learned about physical properties of substances, you can see that because the particular form of a substance is a physical property, changing the form introduces or adds another physical property to its list of characteristics. In fact, resources that provide tables of physical and chemical properties of substances, such as the *CRC Handbook of Chemistry and Physics*, generally include the physical properties of substances in all of the states in which they can exist.

Section 3.1 Assessment

1. Describe the characteristics that identify a sample of matter as being a substance.
2. Classify each of the following as a physical or chemical property.
 - a. Iron and oxygen form rust.
 - b. Iron is more dense than aluminum.
 - c. Magnesium burns brightly when ignited.
 - d. Oil and water do not mix.
 - e. Mercury melts at -39°C .
3. Create a table that describes the three common states of matter in terms of their shape, volume, and compressibility.
4. **Thinking Critically** Using what you know about the compressibility of gases, explain why the oxygen in a SCUBA tank is compressed.
5. **Interpreting Data** Bromine is a reddish-brown liquid that boils at 59°C . Bromine is highly reactive with many metals. For example, it reacts with sodium to form a white solid. Classify each of these properties of bromine as either a physical or a chemical property.

You learned in Section 3.1 that scientists can describe matter in terms of physical and chemical properties. For example, a physical property of copper allows it to be drawn into copper wire, and a chemical property of copper accounts for the fact that when a solution of copper ions is combined with ammonia, the copper solution changes to a deep blue color. The key concept in both of these examples is that the substance copper changed in some way. In this section, you'll explore how matter changes as a result of its physical and chemical properties.

Physical Changes

A substance often undergoes changes that result in a dramatically different appearance yet leave the composition of the substance unchanged. An example is the crumpling of a sheet of aluminum foil. While the foil goes from a smooth, flat, mirrorlike sheet to a round, compact ball, the actual composition of the foil is unchanged—it is still aluminum. Changes such as this, which alter a substance without changing its composition, are known as **physical changes**. Cutting a sheet of paper and breaking a crystal are other examples of physical changes in matter. Can you name some other physical changes? Your list might include verbs such as bend, grind, crumple, split, and crush, all of which indicate physical change.

As with other physical properties, the state of matter depends on the temperature and pressure of the surroundings. As temperature and pressure change, most substances undergo a change from one state (or phase) to another. For example, at atmospheric pressure and at temperatures below 0°C , water is in its solid state, which is known as ice. As heat is added to the ice, it melts and becomes liquid water. This change of state is a physical change because even though ice and water have very different appearances, their composition is the same. If the temperature of the water increases to 100°C , the water begins to boil and liquid water is converted to steam. Melting and formation of a gas are both physical changes and phase changes. **Figure 3-7** shows condensation, another common phase change. When you encounter terms such as boil, freeze, condense, vaporize, or melt in your study of chemistry, the meaning generally refers to a phase change in matter.

Objectives

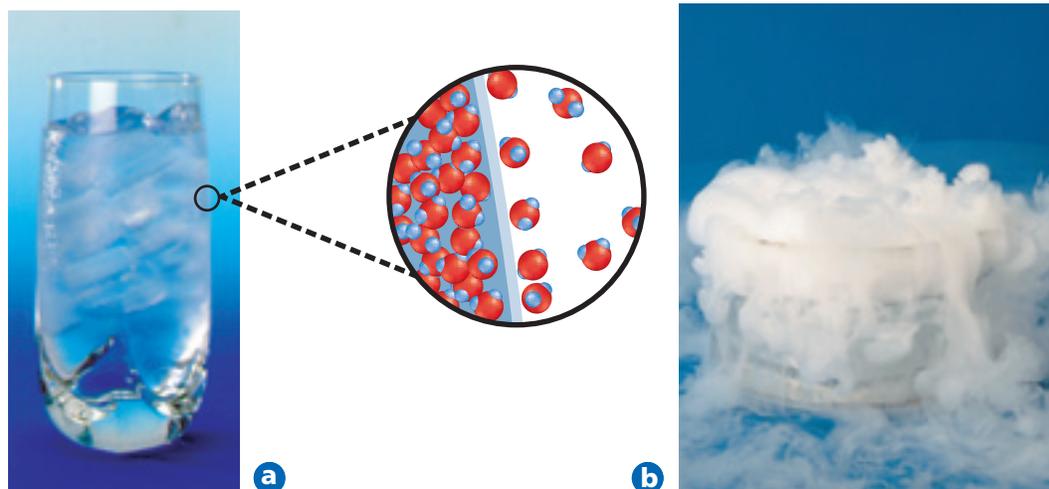
- **Define** physical change and list several common physical changes.
- **Define** chemical change and list several indications that a chemical change has taken place.
- **Apply** the law of conservation of mass to chemical reactions.

Vocabulary

physical change
chemical change
law of conservation of mass

Figure 3-7

- Condensation on an icy beverage glass is the result of the phase change of water in a gaseous state to water in a liquid state.
- The characteristic “fog” of dry ice is actually fine water droplets formed by condensation of water vapor from the air surrounding the very cold dry ice. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



The temperature and pressure at which a substance undergoes a phase change are important physical properties. These properties are listed as the melting and boiling points of the substance. **Table 3-1** on page 56 provides this information for several common substances. Like density, the melting point and boiling point are intensive physical properties that may be used to identify unknown substances. For example, if an unknown solid melts at 801°C and boils at 1413°C —very high temperatures—it is most probably sodium chloride, or common table salt. Tables of intensive properties, such as those given in the *CRC Handbook of Chemistry and Physics*, are indispensable tools in identifying unknown substances from experimental data.

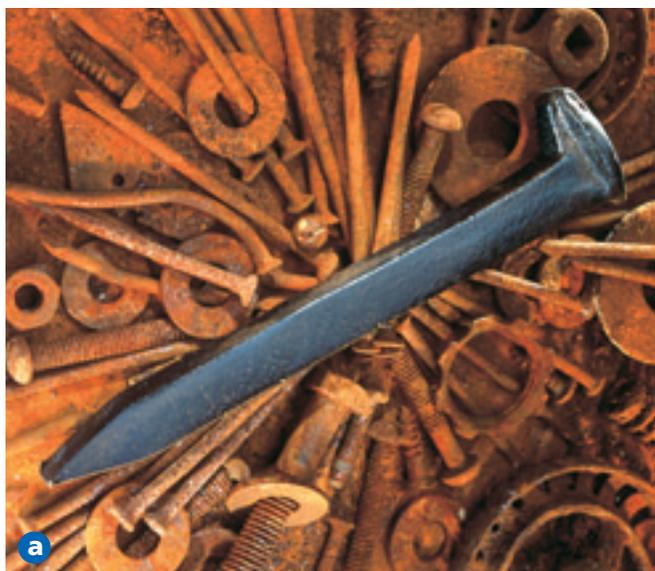
Chemical Changes

As you learned earlier, chemical properties relate to the ability of a substance to combine with or change into one or more substances. A process that involves one or more substances changing into new substances is called a **chemical change**, which is commonly referred to as a chemical reaction. The new substances formed in the reaction have different compositions and different properties from the substances present before the reaction occurred. For example, the crushing of grapes that is part of the wine-making process is a physical change, but the fermentation of the juice, sugars, and other ingredients to wine is a chemical change. The **Chemistry and Society** feature at the end of the chapter describes some interesting consequences of physical and chemical changes in the production of concrete.

Let's consider again the rusting of iron. When a freshly exposed iron surface is left in contact with air, it slowly changes into a new substance, namely, the rust shown in **Figure 3-8a**. The iron reacts with oxygen in the air to form a new substance, rust. Rust is a chemical combination of iron and oxygen. In chemical reactions, the starting substances are called reactants and the new substances that are formed are called products. Thus iron and oxygen are reactants and rust is a product. When you encounter terms such as explode, rust, oxidize, corrode, tarnish, ferment, burn, or rot, the meaning generally refers to a chemical reaction in which reactant substances produce different product substances.

Figure 3-8

- a** The formation of a gas or solid when reactants mix often indicates that a chemical reaction has taken place. Rust is the result of a chemical reaction.
- b** Color changes generally indicate that a chemical reaction has taken place. One example is the color change of tree leaves in the fall.





Evidence of a chemical reaction As **Figure 3-8a** shows, rust is a brownish-orange powdery substance that looks very different from iron and oxygen. Rust is not attracted to a magnet, whereas iron is. The observation that the product (rust) has different properties than the reactants (iron and oxygen) is evidence that a chemical reaction has taken place. A chemical reaction always produces a change in properties. **Figures 3-8 and 3-9** illustrate several common indicators of chemical change. The **CHEMLAB** at the end of the chapter provides a practical laboratory experience with chemical reactions.

Conservation of Mass

Although chemical reactions have been observed over the course of human history, it was only in the late eighteenth century that scientists began to use quantitative tools to monitor chemical changes. The revolutionary quantitative tool developed at this time was the analytical balance, which was capable of measuring very small changes in mass.

By carefully measuring mass before and after many chemical reactions, it was observed that, although chemical changes occurred, the total mass involved in the reaction remained constant. The constancy of mass in chemical reactions was observed so often that scientists assumed the phenomenon must be true for all reactions. They summarized this observation in a scientific law. The **law of conservation of mass** states that mass is neither created nor destroyed during a chemical reaction—it is conserved. This law was one of the great achievements of eighteenth-century science. The equation form of the law of conservation of mass is

$$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$$

The French scientist Antoine Lavoisier (1743–1794) was one of the first to use an analytical balance like the one shown in **Figure 3-10** to monitor chemical reactions. He studied the thermal decomposition of mercury(II) oxide, known then as calx of mercury. Mercury(II) oxide is a powdery red solid. When it is heated, the red solid reacts to form silvery liquid mercury and colorless oxygen gas as shown in **Figure 3-11** on the next page. The color change and production of a gas are indicators of a

Figure 3-9

- a** Energy changes indicate chemical reactions. For example, the burning of wood is a common example of a reaction that releases heat.
- b** The change in the smell of a substance or the production of an odor may be an indication of a chemical reaction.

Figure 3-10

The development of scientific tools such as this analytical balance gave a degree of precision to measurements that greatly improved general scientific understanding.

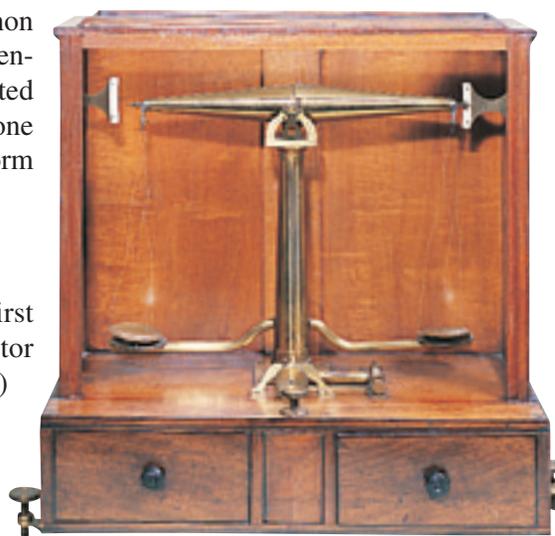
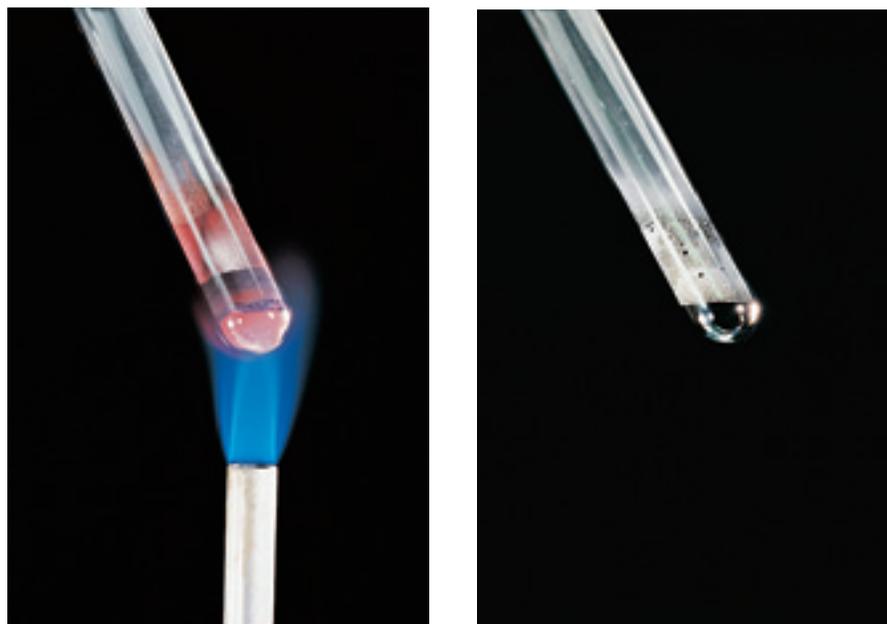
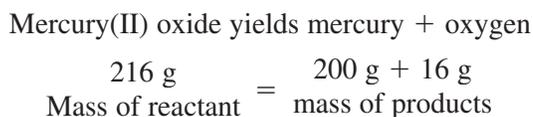


Figure 3-11

Lavoisier's experimental decomposition of mercury(II) oxide is one proof of the law of conservation of mass. Although a chemical reaction is obvious (powder to liquid mercury), matter was neither created nor destroyed.

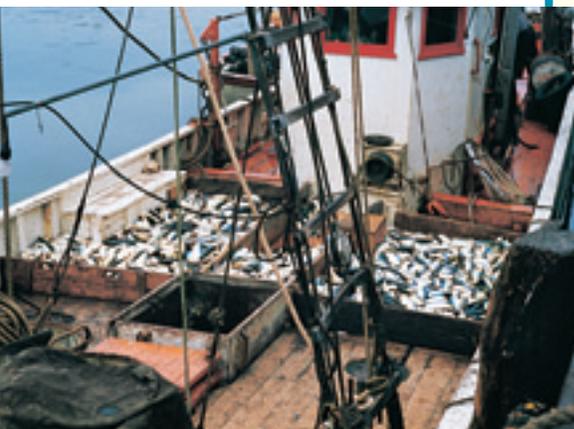


chemical reaction. When the reaction is performed in a closed container, the oxygen gas cannot escape and the mass before and after the reaction can be measured. The masses will be the same.



A more modern digital analytical balance can be used to prove the conservation of mass of this example. The law of conservation of mass is one of the most fundamental concepts of chemistry. Let's examine more closely some situations that illustrate the concept. Example Problem 3-1 leads you through a sample calculation. The practice problems also illustrate the law of conservation of mass.

Mercury occurs naturally in air, water, soil, and living organisms. Seafood that is intended for human consumption is monitored to ensure that the products do not contain levels of mercury exceeding the established limits for public safety.



EXAMPLE PROBLEM 3-1

Conservation of Mass

In an experiment, 10.00 g of red mercury(II) oxide powder is placed in an open flask and heated until it is converted to liquid mercury and oxygen gas. The liquid mercury has a mass of 9.26 g. What is the mass of oxygen formed in the reaction?

1. Analyze the Problem

You are given the mass of a reactant and the mass of one of the products in a chemical reaction. Applying the law of conservation of mass, the total mass of the products must equal the total mass of the reactants. This means that the mass of the liquid mercury plus the mass of the oxygen gas must equal the mass of the mercury(II) oxide powder.

Known

Mass of mercury(II) oxide =
10.00 g

Mass of liquid mercury = 9.26 g

Unknown

Mass of oxygen formed =
? g

2. Solve for the Unknown

Write an equation showing conservation of mass of reactants and products.

$$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$$

$$\text{Mass}_{\text{mercury(II) oxide}} = \text{Mass}_{\text{mercury}} + \text{Mass}_{\text{oxygen}}$$

Solve the equation for $\text{Mass}_{\text{oxygen}}$.

$$\text{Mass}_{\text{oxygen}} = \text{Mass}_{\text{mercury(II) oxide}} - \text{Mass}_{\text{mercury}}$$

Substitute known values and solve.

$$\text{Mass}_{\text{oxygen}} = 10.00 \text{ g} - 9.26 \text{ g}$$

$$\text{Mass}_{\text{oxygen}} = 0.74 \text{ g}$$

3. Evaluate the Answer

The sum of the masses of the two products equals the mass of the reactant, verifying that mass has been conserved. The answer is correctly expressed to the hundredths place.

PRACTICE PROBLEMS

- From a laboratory process designed to separate water into hydrogen and oxygen gas, a student collected 10.0 g of hydrogen and 79.4 g of oxygen. How much water was originally involved in the process?
- A student carefully placed 15.6 g of sodium in a reactor supplied with an excess quantity of chlorine gas. When the reaction was complete, the student obtained 39.7 g of sodium chloride. How many grams of chlorine gas reacted? How many grams of sodium reacted?
- In a flask, 10.3 g of aluminum reacted with 100.0 g of liquid bromine to form aluminum bromide. After the reaction, no aluminum remained, and 8.5 grams of bromine remained unreacted. How many grams of bromine reacted? How many grams of compound were formed?
- A 10.0-g sample of magnesium reacts with oxygen to form 16.6 g of magnesium oxide. How many grams of oxygen reacted?



For more practice with conservation of mass, go to **Supplemental Practice Problems** in Appendix A.

Section 3.2 Assessment

- Describe the results of a physical change and list three examples of physical change.
- Describe the results of a chemical change. List four indicators of chemical change.
- Solve each of the following.
 - In the complete reaction of 22.99 g of sodium with 35.45 g of chlorine, what mass of sodium chloride is formed?
 - A 12.2-g sample of X reacts with a sample of Y to form 78.9 g of XY. What is the mass of Y that reacted?
- Thinking Critically** A friend tells you, “Because composition does not change during a physical change, the appearance of a substance does not change.” Is your friend correct? Explain why.
- Classifying** Classify each of the following examples as a physical change or a chemical change.
 - crushing an aluminum can
 - recycling used aluminum cans to make new aluminum cans
 - aluminum combining with oxygen to form aluminum oxide

Objectives

- **Contrast** mixtures and substances.
- **Classify** mixtures as homogeneous or heterogeneous.
- **List** and **describe** several techniques used to separate mixtures.

Vocabulary

mixture
heterogeneous mixture
homogeneous mixture
solution
filtration
distillation
crystallization
chromatography

When scientists speak of the composition of matter, they are referring to the kinds and amounts of components of which the matter is made. On the basis of composition alone, all matter can be classified into two broad categories: substances or mixtures. You have already learned that a pure substance is a form of matter with a uniform and unchanging composition. You also know that the intensive properties of pure substances do not change, regardless of the physical state or amount of the substance. But what is the result when two or more substances are combined?

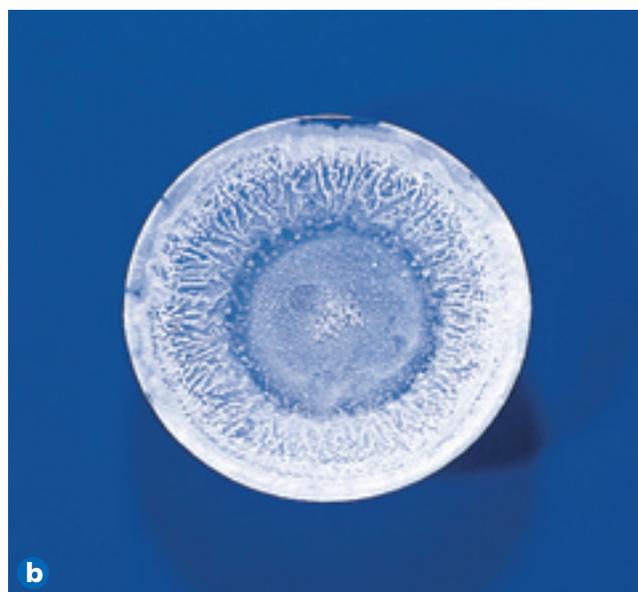
Mixtures

A **mixture** is a combination of two or more pure substances in which each pure substance retains its individual chemical properties. The composition of mixtures is variable, and the number of mixtures that can be created by combining substances is infinite. Although much of the focus of chemistry is the behavior of substances, it is important to remember that most everyday matter occurs as mixtures. Substances tend naturally to mix; it is difficult to keep things pure.

Two mixtures, sand and water, and table salt and water, are shown in **Figure 3-12a**. You know water to be a colorless liquid. Sand is a grainy solid that does not dissolve in water. When sand and water are mixed, the two substances are in contact, yet each substance retains its properties. The sand and water have not reacted. Just by looking at the sand–water mixture in beaker A, it is easy to see each separate substance. Some mixtures, however, may not look like mixtures at all. The mixture of table salt and water in the beaker labeled B is colorless and appears the same as pure water. How can you determine if it is a mixture? If you were to boil away the water, you would see a white residue. That residue, shown in **Figure 13-12b**, is the salt. Thus, the colorless mixture actually contained two separate substances. The salt and the water physically mixed but did not react and were separated by the physical method of boiling.

Figure 3-12

- a** The components of the sand and water mixture (left) are obvious, whereas the components of the table salt and water mixture (right) are not.
- b** The salt component becomes obvious when the mixture is boiled.



Types of mixtures The combinations of pure substances shown in **Figure 3-12** are indeed both mixtures, despite their obvious visual differences. Can you think of some way to further define mixtures? Mixtures themselves are classified as either heterogeneous or homogeneous. A **heterogeneous mixture** is one that does not blend smoothly throughout and in which the individual substances remain distinct. The sand and water mixture is an example of a heterogeneous mixture. Suppose you draw a drop from the top of the mixture using an eyedropper. The drop would be almost completely water. If you draw a second drop from the bottom of the mixture, that drop would contain mostly sand. Thus the composition of the sand–water mixture is not uniform—the substances have not blended smoothly and the two substances of the mixture (sand on the bottom and water on the top) remain distinct. In another example, fresh-squeezed orange juice is a mixture of juice and pulp. The pulp component floats on top of the juice component. Is your favorite pizza a mixture? The answer is yes when you consider that the pizza is a combination of distinct areas of dough, sauce, cheese, and toppings. We can therefore say that the existence of two or more distinct areas indicates a heterogeneous mixture.

A **homogeneous mixture** has constant composition throughout; it always has a single phase. Let’s examine the salt–water mixture using the eyedropper. A drop of the mixture from the top of the beaker has the same composition as a drop from the bottom of the beaker. In fact, every drop of the mixture contains the same relative amounts of salt and water.

Homogeneous mixtures are also referred to as **solutions**. You are probably most familiar with solutions in a liquid form, such as cough suppressant medicine and lemonade, but solutions may contain solids, liquids, or gases. **Table 3-3** lists the various types of solution systems and gives an example of each. Solutions are very important in chemistry, and, in fact, this textbook devotes an entire chapter to the study of solutions.

The solid–solid solution known as steel is called an alloy. An alloy is a homogeneous mixture of metals, or a mixture of a metal and a nonmetal in which the metal substance is the major component. The U.S. Mint’s golden dollar coin, shown in **Figure 3-13**, uses a metal alloy composed of 77% copper, 12% zinc, 7% manganese, and 4% nickel surrounding a copper core. Alloys are also used in spacecraft and automobiles. What might be the benefit of using alloys for these applications? Manufacturers combine the properties of various metals in an alloy to achieve greater strength and durability of their products.



Figure 3-13

Coins issued by the U.S. Mint are metal alloys. The combination of multiple metals gives the coins specific properties such as color, weight, and durability.

Table 3-3

Types of Solution Systems	
System	Example
Gas–gas	Air is primarily a mixture of nitrogen, oxygen, and argon gases.
Gas–liquid	Carbonated beverages contain carbon dioxide gas in solution.
Liquid–gas	Moist air contains water droplets in air (which is a mixture of gases).
Liquid–liquid	Vinegar contains acetic acid in water.
Solid–liquid	Sweetened powder drink contains sugar and other solid ingredients in water.
Solid–solid	Steel is an alloy of iron containing carbon.



Figure 3-14

The physical properties of the iron filings on the plate allow them to be easily separated from the sand using a magnet.

Separating Mixtures

Most matter exists naturally as mixtures. For students and scientists to gain a thorough understanding of matter, it is very important to be able to do the reverse of mixing, that is, to separate mixtures into their component substances. Because the substances in a mixture are physically combined, the processes used to separate a mixture are physical processes that are based on the difference in physical properties of the substances. Sometimes it is very easy to separate a mixture; separating a mixture of pennies and nickels is not a difficult task. More difficult would be separating a mixture of sand and iron filings. Or would it be? The demonstration illustrated in **Figure 3-14** shows how the sand–iron mixture is easily separated on the basis of the unique physical properties of the substances involved. Numerous techniques have been developed that take advantage of different physical properties in order to separate mixtures.

Heterogeneous mixtures composed of solids and liquids are easily separated by filtration. **Filtration** is a technique that uses a porous barrier to separate a solid from a liquid. As **Figure 3-15** shows, the mixture is poured through a piece of filter paper that has been folded into a cone shape. The liquid passes through, leaving the solids trapped in the filter paper.

miniLAB

Separating Ink Dyes

Applying Concepts Chromatography is an important diagnostic tool for chemists. Many types of substances can be separated and analyzed using this technique. In this experiment, you will use paper chromatography to separate the dyes in water-soluble black ink.

Materials 9-oz wide-mouth plastic cups (2); round filter paper; $\frac{1}{4}$ piece of 11-cm round filter paper; scissors; pointed object, approximately 3–4 mm diameter; water-soluble black felt pen or marker

Procedure



1. Fill one of the wide-mouth plastic cups with water to about 2 cm from the top. Wipe off any water drops on the lip of the cup.
2. Place the round filter paper on a clean, dry surface. Make a concentrated ink spot in the center of the paper by firmly pressing the tip of the pen or marker onto the paper.
3. Use a sharp object to create a small hole, approximately 3–4 mm or about the diameter of a pen tip, in the center of the ink spot.
4. Roll the $\frac{1}{4}$ piece of filter paper into a tight cone. This will act as a wick to draw the ink. Work the pointed end of the wick into the hole in the center of the round filter paper.
5. Place the paper/wick apparatus on top of the cup of water, with the wick in the water. The



water will move up the wick and outward through the round paper.

6. When the water has moved to within about 1 cm of the edge of the paper (about 20 minutes), carefully remove the paper from the water-filled cup and put it on the empty cup.

Analysis

1. Make a drawing of the round filter paper and label the color bands. How many distinct dyes can you identify?
2. Why do you see different colors at different locations on the filter paper?
3. How does your chromatogram compare with those of your classmates who used other types of black felt pens or markers? Explain the differences.



Figure 3-15

Filtration is a common technique used to remove impurities from drinking water. Clean water passes through the porous filter **a**, leaving behind the impurities that can be easily discarded **b**.

Most homogeneous mixtures can be separated by distillation. **Distillation** is a separation technique that is based on differences in the boiling points of the substances involved. In distillation, a mixture is heated until the substance with the lowest boiling point boils to a vapor that can then be condensed into a liquid and collected. When precisely controlled, distillation can separate substances having boiling points that differ by only a few degrees.

Did you ever make rock candy as a child? Making rock candy from a sugar solution is an example of separation by crystallization. **Crystallization** is a separation technique that results in the formation of pure solid particles of a substance from a solution containing the dissolved substance. When the solution contains as much dissolved substance as it can possibly hold, the addition of a tiny amount more often causes the dissolved substance to come out of solution and collect as crystals on some available surface. In the rock candy example, as water evaporates from the sugar–water solution, the sugar is left behind as a solid crystal on the string. Crystallization produces highly pure solids.

Chromatography is a technique that separates the components of a mixture (called the mobile phase) on the basis of the tendency of each to travel or be drawn across the surface of another material (called the stationary phase). The **miniLAB** in this section describes how you can separate a solution such as ink into its components as it spreads across a stationary piece of paper. The separation occurs because the various components of the ink spread through the paper at different rates.

Section 3.3 Assessment

15. How do mixtures and substances differ?
16. Consider a mixture of water, sand, and oil. How many phases are present? How could you separate this mixture into individual substances?
17. Classify each of the following as either a heterogeneous or homogeneous mixture.
 - a. orange juice
 - b. tap water
 - c. steel (a blend of iron and carbon)
 - d. air
 - e. raisin muffin
18. **Thinking Critically** When 50 mL of ethanol is mixed with 50 mL of water, a solution forms. The volume of the final solution is less than 100 mL. Propose an explanation for this phenomenon. (*Hint*: Consider what you know about the space between particles in liquids.)
19. **Applying Concepts** Describe the separation technique that could be used to separate each of the following mixtures.
 - a. two colorless liquids
 - b. a nondissolving solid mixed with a liquid
 - c. red and blue marbles of same size and mass

Objectives

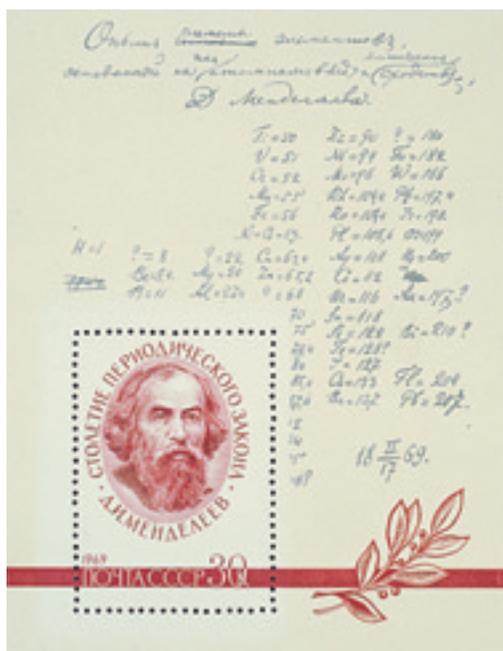
- **Distinguish** between elements and compounds.
- **Describe** the organization of elements on the periodic table.
- **Explain** how all compounds obey the laws of definite and multiple proportions.

Vocabulary

element
 periodic table
 compound
 law of definite proportions
 percent by mass
 law of multiple proportions

Figure 3-16

Although many early scientists have contributed to the modern organization of the elements, Mendeleev's system of rows and columns was a revolutionary advancement.



To this point you've examined many of the properties of matter. You've also learned how scientists have organized, classified, and described matter by arranging it into various subcategories of components. But there remains another fundamental level of classification of matter: the classification of pure substances as elements or compounds.

Elements

Recall that earlier in this chapter you considered the diversity of your surroundings in terms of matter. Although the diversity is astounding, in reality all matter can be broken down into a relatively small number of basic building blocks called elements. An **element** is a pure substance that cannot be separated into simpler substances by physical or chemical means. On Earth, 91 elements occur naturally. Copper, oxygen, and gold are examples of naturally occurring elements. There are also several elements that do not exist naturally but have been developed by scientists.

Each element has a unique chemical name and symbol. The chemical symbol consists of one, two, or three letters; the first letter is always capitalized and the remaining letter(s) are always lowercase. Why has so much effort been given to naming the elements? The names and symbols of the elements are universally accepted by scientists in order to make the communication of chemical information possible.

The 91 naturally occurring elements are not equally abundant. For example, hydrogen is estimated to make up approximately 75% of the mass of the universe. Oxygen and silicon together comprise almost 75% of the mass of Earth's crust, while oxygen, carbon, and hydrogen account for more than 90% of the human body. Francium, on the other hand, is one of the least abundant naturally occurring elements. It is estimated that there is probably less than 20 grams of francium dispersed throughout Earth's crust. To put that into perspective, the total mass of francium is approximately equal to the mass of your pencil or pen.

A first look at the periodic table As many new elements were being discovered in the early nineteenth century, chemists began to see patterns of similarities in the chemical and physical properties of particular sets of elements. Several schemes for organizing the elements on the basis of these similarities were proposed, with varying degrees of success. In 1869, the Russian chemist Dmitri Mendeleev made a significant contribution to the effort. Mendeleev devised the chart shown in **Figure 3-16**, which organized all of the elements that were known at the time into rows and columns based on their similarities and their masses. Mendeleev's organizational table was the first version of what has been further developed into the **periodic table** of elements. The periodic table organizes the elements into a grid of horizontal rows called periods and vertical columns called groups or families. Elements in the same group have similar chemical and physical properties. The table is called "periodic" because the pattern of similar properties repeats as you move from period to period.

One of the brilliant aspects of Mendeleev's original table was that its structure could accommodate elements that did not even exist at

the time. Notice the blank spots in Mendeleev's table. By analyzing the similarities among the elements and their pattern of repetition, Mendeleev was able to predict the properties of elements that were yet to be discovered.

In most cases, Mendeleev's predictions (and the blanks in the table) closely matched the characteristics of new elements as they were discovered. **Figure 3-18** on pages 72–73 shows samples of the elements in their arrangement in the periodic table. The standard modern version of the periodic table includes more than 100 elements. You'll study the periodic table in greater detail later in this textbook. In fact, the periodic table remains a dynamic tool as scientists continue to discover new elements.

Compounds

Take a moment to recall what you have learned about the organization of matter, using **Figure 3-17** as a guide. You know that matter is classified as pure substances and mixtures. As you learned in the previous section, mixtures can be homogeneous or heterogeneous. You also know that elements are pure substances that cannot be separated into simpler substances. There is yet another classification of pure substances—compounds. A **compound** is a combination of two or more different elements that are combined chemically. Most of the substances that you are familiar with and, in fact, much of the matter of the universe are compounds. Water, table salt, table sugar, and aspirin are examples of common compounds.

Today, there are approximately 10 million known compounds, and new compounds continue to be developed and discovered at the rate of about 100 000 per year. Can you recall some of the medicinal compounds that have made headlines in recent years? There appears to be no limit to the number of compounds that can be made or that will be discovered. Considering this virtually limitless potential, several organizations have assumed the task of collecting data and indexing the known chemical compounds. These organizations maintain huge databases that allow researchers to access information on existing compounds. The databases and retrieval tools enable scientists to build the body of chemical knowledge in an efficient manner.

The chemical symbols of the periodic table make it easy to write the formulas for chemical compounds. For example, table salt, or sodium chloride, is composed of one part sodium (Na) and one part chlorine (Cl), and its chemical formula is NaCl. Water is composed of two parts hydrogen (H) to one part oxygen (O), and its formula is H₂O.

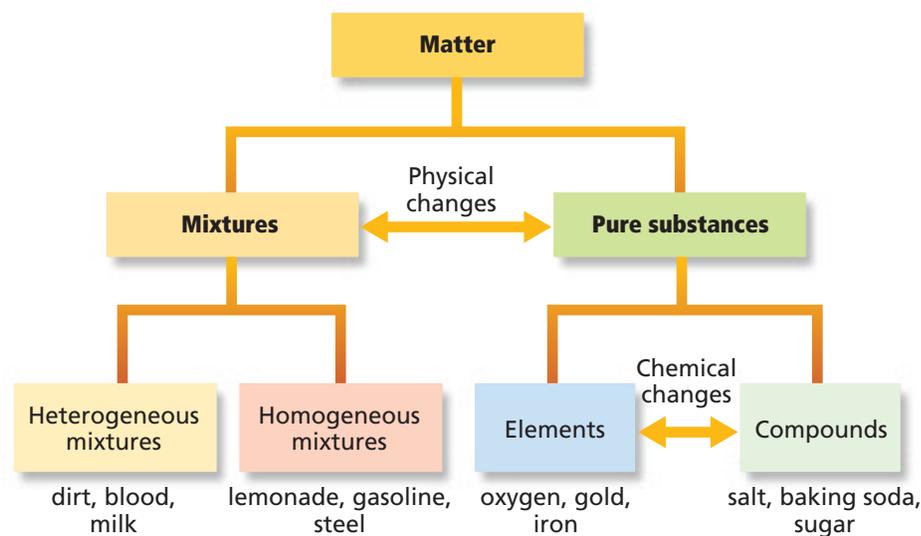


Figure 3-17

The concept of matter is far-reaching and can be overwhelming. But, when broken down as shown here, it becomes clear how elements, compounds, substances, and mixtures define all matter.

Periodic Table

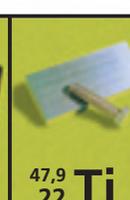
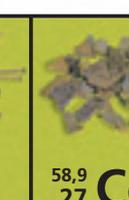
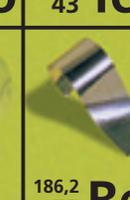
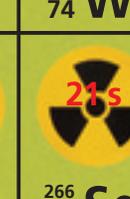
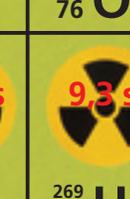
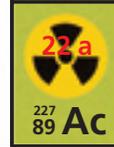
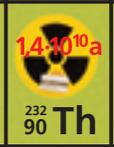
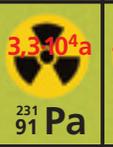
	IA 1	IIA 2	IIIB 3	IVB 4	VB 5	VIB 6	VIIB 7	8	VIII 9
1	 1,0 1 H								
2	 6,9 3 Li	 9,0 4 Be							
3	 23,0 11 Na	 24,3 12 Mg							
4	 39,1 19 K	 40,1 20 Ca	 45,0 21 Sc	 47,9 22 Ti	 50,9 23 V	 52,0 24 Cr	 54,9 25 Mn	 55,8 26 Fe	 58,9 27 Co
5	 85,5 37 Rb	 87,6 38 Sr	 88,9 39 Y	 91,2 40 Zr	 92,9 41 Nb	 95,9 42 Mo	 4,2 · 10 ⁶ a 98 43 Tc	 101,1 44 Ru	 102,9 45 Rh
6	 132,9 55 Cs	 137,3 56 Ba	 138,9 57 La	 178,5 72 Hf	 180,9 73 Ta	 183,8 74 W	 186,2 75 Re	 190,2 76 Os	 192,2 77 Ir
7	 22 min 223 87 Fr	 1600 a 226 88 Ra	 22 a 227 89 Ac	 65 s 261 104 Rf	 34 s 262 105 Db	 21 s 266 106 Sg	 440 ms 264 107 Bh	 9,3 s 269 108 Hs	 70 ms 268 109 Mt

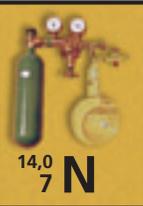
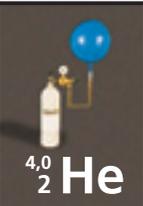
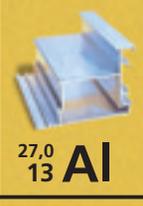
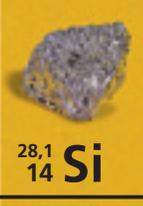
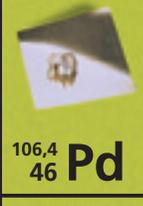
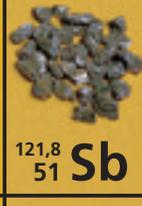
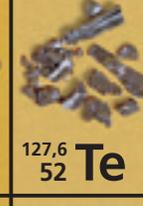
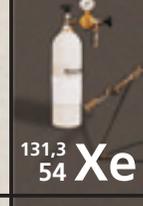
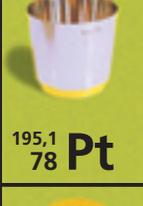
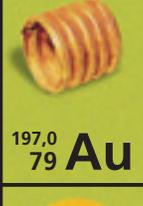
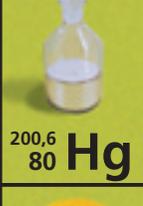
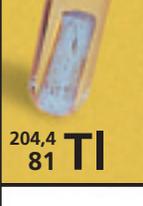
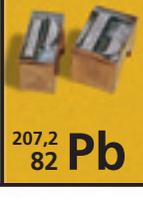
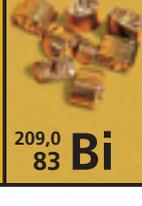
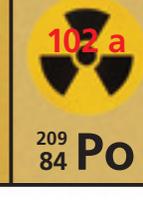
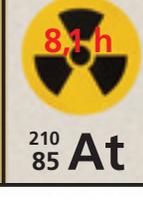
Figure 3-18

The periodic table shown above illustrates samples of many of the elements. Be sure to use the periodic table on pages 156-157 for reference throughout your chemistry course.

 138,9 57 La	 140,1 58 Ce	 140,9 59 Pr	 144,2 60 Nd	 18 a 145 61 Pm
 22 a 227 89 Ac	 1,4 · 10 ¹⁰ a 232 90 Th	 3,3 · 10 ⁴ a 231 91 Pa	 4,5 · 10 ⁹ a 238 92 U	 2,1 · 10 ⁶ a 237 93 Np

of the Elements

VIIIA
18

			IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	VIIIA 18			
			 10,8 5 B	 12,0 6 C	 14,0 7 N	 16,0 8 O	 19,0 9 F	 4,0 2 He			
			 27,0 13 Al	 28,1 14 Si	 31,0 15 P	 32,1 16 S	 35,5 17 Cl	 39,9 18 Ar			
10	IB 11	IIB 12	 58,7 28 Ni	 63,5 29 Cu	 65,4 30 Zn	 69,7 31 Ga	 72,6 32 Ge	 74,9 33 As	 79,0 34 Se	 79,9 35 Br	 83,8 36 Kr
			 106,4 46 Pd	 107,9 47 Ag	 112,4 48 Cd	 114,8 49 In	 118,7 50 Sn	 121,8 51 Sb	 127,6 52 Te	 126,9 53 I	 131,3 54 Xe
			 195,1 78 Pt	 197,0 79 Au	 200,6 80 Hg	 204,4 81 Tl	 207,2 82 Pb	 209,0 83 Bi	 209 84 Po 102 a	 210 85 At 81 h	 222 86 Rn 3,8 d
			 273 110 118 ms	 272 111 1,5 ms	 277 112 0,24 ms						

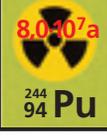
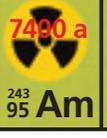
 150,4 62 Sm	 152,0 63 Eu	 157,3 64 Gd	 158,9 65 Tb	 162,5 66 Dy	 164,9 67 Ho	 167,3 68 Er	 168,9 69 Tm	 173,0 70 Yb	 175,0 71 Lu
 244 94 Pu 8,0 · 10 ⁷ a	 243 95 Am 7490 a	 247 96 Cm 1,6 · 10 ⁷ a	 247 97 Bk 1490 a	 251 98 Cf 900 a	 252 99 Es 472 d	 257 100 Fm 101 d	 258 101 Md 52 d	 259 102 No 58 min	 262 103 Lr 3,6 h

Figure 3-19

This classic apparatus, called a Hoffman apparatus, and other similar designs are used to separate water into its components.



Unlike elements, compounds can be broken down into simpler substances by chemical means. In general, compounds that naturally occur are more stable than the individual component elements. To separate a compound into its elements often requires external energy such as heat or electricity. **Figure 3-19** shows the apparatus used to produce the chemical change of water into its component elements of hydrogen and oxygen through a process called electrolysis. Here, one end of a long platinum electrode is exposed to the water in the tube and the other end is attached to a power source. An electric current splits water into hydrogen gas in the compartment on the right and oxygen gas in the compartment on the left. Because water is composed of two parts hydrogen and one part oxygen, there is twice as much hydrogen gas than oxygen gas.

Figure 3-20

Compounds such as sodium chloride (table salt) are often remarkably different from the components that comprise them.



The properties of a compound are different from those of its component elements. The example of water in **Figure 3-19** illustrates this fact. Water is a stable compound that is liquid at room temperature. When water is broken down into its components, it is obvious that hydrogen and oxygen are dramatically different than the liquid they form when combined. Oxygen and hydrogen are tasteless, odorless gases that vigorously undergo chemical reactions with many elements. This difference in properties is a result of a chemical reaction between the elements. **Figure 3-20** shows the component elements (sodium and chlorine) of the compound commonly called table salt (sodium chloride). When sodium and chlorine react with each other, the compound sodium chloride is formed. Note how different the properties of sodium chloride are from its component elements. Sodium is a highly reactive element that fizzes and burns when added to water. Chlorine is a poisonous, pale green gas. Sodium chloride, however, is a white, unreactive solid that flavors many of the foods you eat.

Law of Definite Proportions

An important characteristic of compounds is that the elements comprising them combine in definite proportions by mass. This observation is so fundamental that it is summarized as the **law of definite proportions**. This law states that, regardless of the amount, a compound is always composed of the same elements in the same proportion by mass. For example, consider the compound table sugar (sucrose), which is composed of carbon, hydrogen, and oxygen. The analysis of 20.00 g of sucrose from a bag of sugar is given in **Table 3-4**. Note that in Column 1 the sum of the individual masses of the elements equals 20.00 g, the amount of sucrose that was analyzed. This demonstrates the law of conservation of mass as applied to compounds: The mass of the compound is equal to the sum of the masses of the elements that make up the compound. Column 2 shows the ratio of the mass of each element to the total mass of the compound as a percentage called the **percent by mass**.

$$\text{percent by mass (\%)} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

Table 3-4

Sucrose Analysis from Bag Sugar		
	Column 1	Column 2
Element	Analysis by mass (g)	Percent by mass (%)
Carbon	8.44 g carbon	$\frac{8.44 \text{ g C}}{20.00 \text{ g sucrose}} \times 100 = 42.2\% \text{ carbon}$
Hydrogen	1.30 g hydrogen	$\frac{1.30 \text{ g H}}{20.00 \text{ g sucrose}} \times 100 = 6.50\% \text{ hydrogen}$
Oxygen	10.26 g oxygen	$\frac{10.26 \text{ g O}}{20.00 \text{ g sucrose}} \times 100 = 51.30\% \text{ oxygen}$
Total	20.00 g sucrose	= 100.0%

Now let's suppose you analyzed 500.0 g of sucrose isolated from a sample of sugar cane. The analysis is shown in **Table 3-5**. Note in Column 2 that the percent by mass values equal those in Column 2 in **Table 3-4**. Compounds with the same mass proportions must be the same compound; conversely, compounds with different mass proportions must be different compounds. Thus, you can conclude that, although the two samples of sucrose are from different sources, they have the same composition and they must be the same compound.

Table 3-5

Sucrose Analysis from Sugar Cane		
	Column 1	Column 2
Element	Analysis by mass (g)	Percent by mass (%)
Carbon	211.0 g carbon	$\frac{211.0 \text{ g C}}{500.0 \text{ g sucrose}} \times 100 = 42.20\% \text{ carbon}$
Hydrogen	32.5 g hydrogen	$\frac{32.5 \text{ g H}}{500.0 \text{ g sucrose}} \times 100 = 6.50\% \text{ hydrogen}$
Oxygen	256.5 g oxygen	$\frac{256.5 \text{ g O}}{500.0 \text{ g sucrose}} \times 100 = 51.30\% \text{ oxygen}$
Total	500.0 g sucrose	= 100.00%

History

CONNECTION

Antoine-Laurent Lavoisier (1743–1794) is recognized as the father of modern chemistry. While his fellow scientists tried to explain matter based on the elements fire, earth, air, and water, Lavoisier performed some of the first quantitative chemical experiments. His data and observations led to the statement of the law of conservation of mass. He also studied the nature of combustion and devised a system of naming elements.

Lavoisier held many public offices in France in which he attempted to reform the French monetary and taxation system and farming methods. He also supervised the French government's manufacture of gunpowder. During the Reign of Terror that followed the French Revolution, Lavoisier and other members of Fermés Generalé were arrested, tried, and condemned to the guillotine.





For more practice with percent by mass and law of definite proportions, go to **Supplemental Practice Problems** in Appendix A.

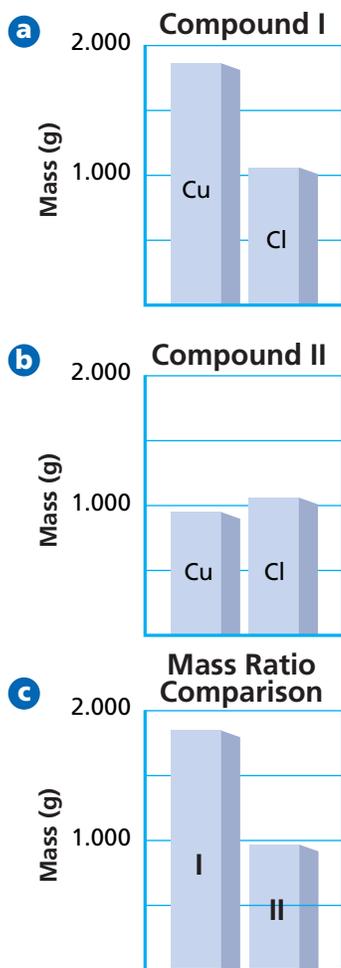


Figure 3-21

Bar graph **a** compares the relative masses of copper and chlorine in Compound I and bar graph **b** compares the relative masses of copper and chlorine in Compound II. **c** A comparison between the relative masses of copper in both compounds shows a 2:1 ratio.

PRACTICE PROBLEMS

- A 78.0-g sample of an unknown compound contains 12.4 g of hydrogen. What is the percent by mass of hydrogen in the compound?
- If 1.0 g of hydrogen reacts completely with 19.0 g of fluorine, what is the percent by mass of hydrogen in the compound that is formed?
- If 3.5 g of X reacts with 10.5 g of Y to form the compound XY, what is the percent by mass of X in the compound? The percent by mass of Y?
- Two unknown compounds are tested. Compound I contains 15.0 g of hydrogen and 120.0 g of oxygen. Compound II contains 2.0 g of hydrogen and 32.0 g of oxygen. Are the compounds the same?
- All you know about two unknown compounds is that they have the same percent by mass of carbon. With only this information, can you be sure the two compounds are the same?

Law of Multiple Proportions

Compounds composed of different elements are obviously different compounds. Can compounds that are composed of the same elements differ from each other? The answer is yes because those different compounds have different mass compositions. The **law of multiple proportions** states that when different compounds are formed by a combination of the same elements, different masses of one element combine with the same relative mass of the other element in a ratio of small whole numbers. Ratios compare the relative amounts of any items or substances. The comparison can be expressed using numbers separated by a colon or as a fraction. With regard to the law of multiple proportions, ratios express the relationship of elements in a compound.

The two distinct compounds water (H_2O) and hydrogen peroxide (H_2O_2) illustrate the law of multiple proportions. Each compound contains the same elements (hydrogen and oxygen). Water is composed of two parts hydrogen (the element that is present in the same amount in both compounds) to one part oxygen (the element that is present in different amounts in both compounds). Hydrogen peroxide is composed of two parts hydrogen and two parts oxygen. Hydrogen peroxide differs from water in that it has twice as much oxygen. When we compare the mass of oxygen in hydrogen peroxide to the mass of oxygen in water, we get the ratio 2:1.

In another example, copper (Cu) reacts with chlorine (Cl) under different sets of conditions to form two different compounds. **Table 3-6** provides an analysis of their composition. Note that the two copper compounds must be different because they have different percents by mass. Compound I contains 64.20% copper; compound II contains 47.27% copper. Compound I contains 35.80% chlorine; compound II contains 52.73% chlorine.

Table 3-6

Analysis Data of Two Copper Compounds					
Compound	% Cu	% Cl	Mass copper (g) in 100.0 g of compound	Mass chlorine (g) in 100.0 g of compound	Mass ratio $\left(\frac{\text{mass Cu}}{\text{mass Cl}}\right)$
I	64.20	35.80	64.20	35.80	1.793 g Cu/1 g Cl
II	47.27	52.73	47.27	52.73	0.8964 g Cu/1 g Cl



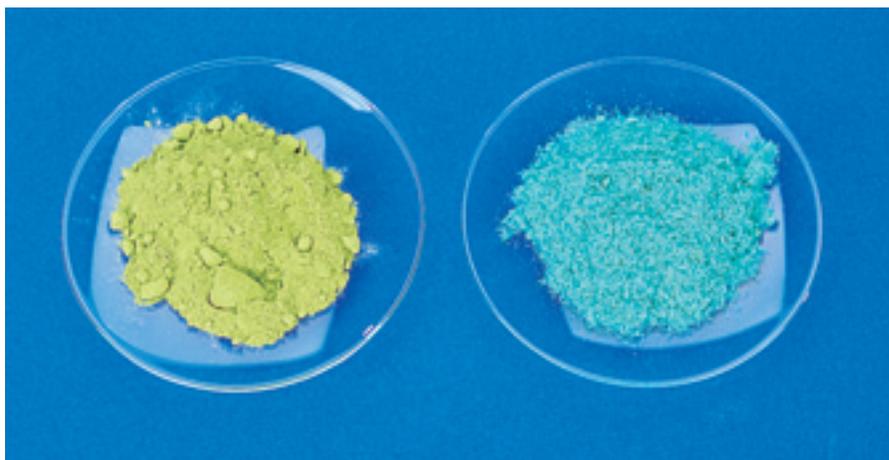


Figure 3-22

Analyses of the mass ratios of the two copper chloride compounds shown here indicate that they are indeed different compounds. The calculated mass ratio of compound I to compound II is 2.000 and fits the definition of the law of multiple proportions.

Compare the ratio of the mass of copper to the mass of chlorine for each compound (see the last column of **Table 3-6** and **Figure 3-21**). You'll notice that the mass ratio of copper to chlorine in compound I (1.793) is two times the mass ratio of copper to chlorine in compound II (0.896).

$$\frac{\text{mass ratio compound I}}{\text{mass ratio compound II}} = \frac{1.793 \text{ g Cu/g Cl}}{0.8964 \text{ g Cu/g Cl}} = 2.000$$

As the law of multiple proportions states, the different masses of copper that combine with a fixed mass of chlorine in the two different copper compounds, shown in **Figure 3-22**, can be expressed as a small whole-number ratio, in this case 2:1.

Considering that there is a finite number of elements that exist today and an exponentially greater number of compounds that are composed of these elements under various conditions, it becomes clear how important the law of multiple proportions is in chemistry.

Section 3.4 Assessment

25. How are elements and compounds similar? How are they different?
26. What is the basic organizing feature of the periodic table of elements?
27. Explain how the law of definite proportions applies to compounds.
28. What type of compounds are compared in the law of multiple proportions?
29. **Thinking Critically** Name two elements that have properties similar to those of element potassium (K). To those of krypton (Kr).
30. **Interpreting Data** Complete the following table and then analyze the data to determine if compounds I and II are the same compound. If the compounds are different, use the law of multiple proportions to show the relationship between them.

Analysis Data of Two Iron Compounds					
Compound	Total mass (g)	Mass Fe (g)	Mass O (g)	Mass % Fe	Mass % O
I	75.00	52.46	22.54		
II	56.00	43.53	12.47		

Matter and Chemical Reactions

One of the most interesting characteristics of matter, and one that drives the study and exploration of chemistry, is the fact that matter changes. By examining a dramatic chemical reaction, such as the reaction of the element copper and the compound silver nitrate in a water solution, you can readily observe chemical change. Drawing on one of the fundamental laboratory techniques introduced in this chapter, you can separate the products. Then, you will use a flame test to confirm the identity of the products.

Problem

Is there evidence of a chemical reaction between copper and silver nitrate? If so, which elements reacted and what is the name of the compound they formed?

Objectives

- **Observe** the reactants as they change into product.
- **Separate** a mixture by filtration.
- **Predict** the names of the products.

Materials

copper wire
 AgNO₃ solution
 sandpaper
 stirring rod
 50-mL graduated cylinder
 50-mL beaker
 funnel

filter paper
 250-mL Erlenmeyer flask
 ring stand
 small iron ring
 plastic petri dish
 paper clip
 Bunsen burner
 tongs

Safety Precautions



- Always wear safety goggles, gloves, and lab apron.
- Silver nitrate is toxic and will harm skin and clothing.
- Use caution around a flame.

Pre-Lab

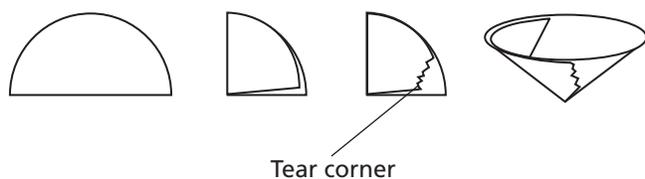
1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your observations.
3. Define the terms physical property and chemical property. Give an example of each.
4. Form a hypothesis regarding what you might observe if
 - a. a chemical change occurs.
 - b. a physical change occurs.
5. Distinguish between a homogeneous mixture and a heterogeneous mixture.

Reaction Observations	
Time (min)	Observations

Procedure

1. Obtain 8 cm of copper wire. Rub the copper wire with the sandpaper until it is shiny.
2. Measure approximately 25 mL AgNO₃ (silver nitrate) solution into a 50-mL beaker. **CAUTION:** Do not allow to contact skin or clothing.
3. Make and record an observation of the physical properties of the copper wire and AgNO₃ solution.

- Coil the piece of copper wire to a length that will fit into the beaker. Make a hook on the end of the coil to allow the coil to be suspended from the stirring rod.
- Hook the coil onto the middle of the stirring rod. Place the stirring rod across the top of the beaker immersing some of the coil in the AgNO_3 solution.
- Make and record observations of the wire and the solution every five minutes for 20 minutes.
- Use the ring stand, small iron ring, funnel, Erlenmeyer flask, and filter paper to set up a filtration apparatus. Attach the iron ring to the ring stand. Adjust the height of the ring so the end of the funnel is inside the neck of the Erlenmeyer flask.
- To fold the filter paper, examine the diagram below. Begin by folding the circle in half, then fold in half again. Tear off the lower right corner of the flap that is facing you. This will help the filter paper stick better to the funnel. Open the folded paper into a cone. Place the filter paper cone in the funnel.



- Remove the coil from the beaker and dispose of it as directed by your teacher. Some of the solid product may form a mixture with the liquid in the beaker. Decant the liquid by slowly pouring it down the stirring rod into the funnel. Solid product will be caught in the filter paper. Collect the filtrate—the liquid that runs through the filter paper—in the Erlenmeyer flask.
- Transfer the clear filtrate to a petri dish.
- Adjust a Bunsen burner flame until it is blue. Hold the paper clip with tongs in the flame until no additional color is observed. **CAUTION:** *The paper clip will be very hot.*
- Using tongs, dip the hot paper clip in the filtrate. Then, hold the paper clip in the flame. Record the color you observe.

Cleanup and Disposal

- Dispose of materials as directed by your teacher.
- Clean and return all lab equipment to its proper place.
- Wash hands thoroughly.

Analyze and Conclude

- Classifying** Which type of mixture is silver nitrate in water? Which type of mixture is formed in step 6? Explain.
- Observing and Inferring** Describe the changes you observed in step 6. Is there evidence a chemical change occurred? Why?
- Predicting** Predict the products formed in step 6. You may not know the exact chemical name, but you should be able to make an intuitive prediction.
- Using Resources** Use resources such as the *CRC Handbook of Chemistry and Physics*, the *Merck Index*, or the Internet to determine the colors of silver metal and copper nitrate in water. Compare this information with your observations of the reactants and products in step 6.
- Identifying** Metals emit characteristic colors in flame tests. Copper emits blue-green light. Do your observations in step 12 confirm the presence of copper in the filtrate collected in step 9?
- Communicating** Express in words the chemical equation that represents the reaction that occurred in step 6.
- Error Analysis** Compare your recorded observations with those of several other lab teams. Explain any differences.

Real-World Chemistry

- Analytical chemists determine the chemical composition of matter. Two major branches of analytical chemistry are qualitative analysis—determining what is in a substance—and quantitative analysis—measuring how much substance. Research and report on a career as an analytical chemist in the food industry.

CHEMISTRY and Society

Green Buildings

Until the Industrial Revolution, the amount of carbon dioxide (CO_2) in the atmosphere was fairly constant. Since the Industrial Revolution, however, the burning of fossil fuels has contributed to a significant increase in the amount of carbon dioxide in the atmosphere. As the level of carbon dioxide increases, Earth gradually warms up. Too much CO_2 in the atmosphere can change the conditions on Earth.

Another major source of carbon dioxide may be in the foundation of your building or on the sidewalks near your school. The production of cement, the key ingredient in concrete, releases tremendous amounts of carbon dioxide into the atmosphere. Chemistry may allow engineers to build “green buildings,” that are still practical yet have less of an impact on the environment.

Producing Cement

Cement generally begins with a mixture of limestone and sand placed in a kiln, which heats it to about 1480°C . As the mixture is heated, its chemical and physical properties change. After heating, the solid that remains is ground into a fine powder. This is cement. To make concrete, the cement is mixed with fine particles, such as sand, coarse particles, such as crushed stone, and water.

During the production of cement, carbon dioxide is released in two ways. First, when the limestone is heated it changes into lime and carbon dioxide. Second, the electrical energy used to heat the kiln is usually supplied by a power plant that burns fossil fuels, such as coal. Fossil fuels also release carbon dioxide and other substances.

Using Flyash

One way to reduce the amount of carbon dioxide released into the atmosphere is to find a replacement for cement in concrete. One such replacement is a substance known as flyash. Flyash is a waste product that accumulates in the smokestacks of power plants when ground coal is burned. It is a fine gray powder that consists of tiny glass beads.

Using flyash offers several advantages. Flyash ordinarily is dumped in landfills. Replacing cement with flyash can reduce CO_2 emissions and prevent



tons of waste from piling up in landfills. Flyash also produces better concrete. Traditional concrete has weak zones where tiny cracks allow water to flow through. Flyash contains fine particles that fill spaces and keep moisture out. Flyash also protects the steel surrounding the concrete, makes the concrete easier to work with, and extends the life of the concrete structure. In fact, flyash is so reliable the Romans used natural materials similar to flyash to build the concrete dome of the Pantheon.

Solutions to environmental problems require a willing commitment from scientists, architects, builders, and owners to look for ways to build durable structures and protect the environment.

Investigating the Issue

- 1. Communicating Ideas** Write a pamphlet for people who are building new homes telling them about the importance of green buildings.
- 2. Using the Internet** Investigate issues that influence the decision to use flyash. Discuss the advantages and disadvantages of flyash.



CLICK HERE

Visit the Chemistry Web site at science.glencoe.com to find links to more information about flyash and green buildings.

Summary

3.1 Properties of Matter

- A substance is a form of matter with a uniform and unchanging composition.
- Physical properties can be observed without altering a substance's composition. Chemical properties describe a substance's ability to combine with or change into one or more new substances.
- Both physical and chemical properties are affected by external conditions such as temperature and pressure.
- The three common states of matter are solid, liquid, and gas.

3.2 Changes in Matter

- A physical change alters the physical properties of a substance without changing its composition.
- A chemical change, also known as a chemical reaction, involves a change in a substance's composition.
- In a chemical reaction, reactants form products.
- The law of conservation of mass states that mass is neither created nor destroyed during a chemical reaction; it is conserved.

3.3 Mixtures of Matter

- A mixture is a physical blend of two or more pure substances in any proportion.

- Solutions are homogeneous mixtures.
- Mixtures can be separated by physical means. Common separation techniques include filtration, distillation, crystallization, and chromatography.

3.4 Elements and Compounds

- Elements are substances that cannot be broken down into simpler substances by chemical or physical means.
- The elements are organized in the periodic table of elements.
- A compound is a chemical combination of two or more elements. Properties of compounds differ from the properties of their component elements.
- The law of definite proportions states that a compound is always composed of the same elements in the same proportions.
- The law of multiple proportions states that if elements form more than one compound, those compounds will have compositions that are small, whole-number multiples of each other.

Key Equations and Relationships

- law of conservation of mass (p. 63)

$$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$$

- percent by mass = $\frac{\text{Mass}_{\text{element}}}{\text{Mass}_{\text{compound}}} \times 100$
 (p. 75)

Vocabulary

- chemical change (p. 62)
- chemical property (p. 57)
- chromatography (p. 69)
- compound (p. 71)
- crystallization (p. 69)
- distillation (p. 69)
- element (p. 70)
- extensive properties (p. 56)
- filtration (p. 68)
- gas (p. 59)
- heterogeneous mixture (p. 67)
- homogeneous mixture (p. 67)
- intensive properties (p. 56)
- law of conservation of mass (p. 63)
- law of definite proportions (p. 75)
- law of multiple proportions (p. 76)
- liquid (p. 58)
- mixture (p. 66)
- percent by mass (p. 75)
- periodic table (p. 70)
- physical changes (p. 61)
- physical property (p. 56)
- solid (p. 58)
- solution (p. 67)
- states of matter (p. 58)
- substance (p. 55)
- vapor (p. 59)

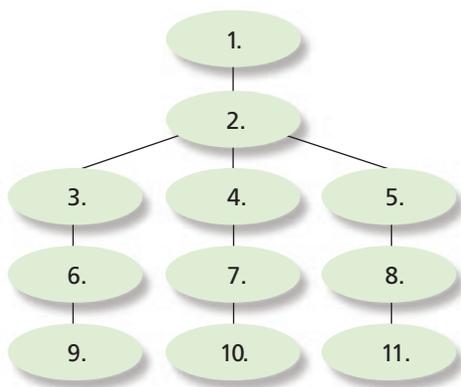


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 3 Assessment.

Concept Mapping

31. Organize the following terms into a logical concept map: state, physical properties, virtually incompressible, solid, gas, liquid, tightly packed particles, compressible, incompressible, particles far apart, loosely packed particles.



Mastering Concepts

32. List three examples of substances. Explain why each is a substance. (3.1)
33. List at least three physical properties of tap water. (3.1)
34. Identify each of the following as an extensive or intensive physical property. (3.1)
- melting point
 - mass
 - density
 - length
35. “Properties are not affected by changes in temperature and pressure.” Is this statement true or false? Explain. (3.1)
36. Classify each of the following as either solid, liquid, or gas at room temperature. (3.1)
- milk
 - air
 - copper
 - helium
 - diamond
 - candle wax

37. Classify each of the following as a physical property or a chemical property. (3.1)
- aluminum has a silvery color
 - gold has a density of 19 g/cm^3
 - sodium ignites when dropped in water
 - water boils at 100°C
 - silver tarnishes
 - mercury is a liquid at room temperature

38. A carton of milk is poured into a bowl. Describe the changes that occur in the milk’s shape and volume. (3.1)



39. Classify each of the following as a physical change or a chemical change. (3.2)
- breaking a pencil in two
 - water freezing and forming ice
 - frying an egg
 - burning wood
 - leaves turning color in the fall
40. Is a change in phase a physical change or a chemical change? Explain. (3.2)
41. List four indicators that a chemical change has probably taken place. (3.2)
42. Iron and oxygen combine to form iron oxide (rust). List the reactants and products of this reaction. (3.2)
43. Use **Table 3-1** to identify a substance that undergoes a phase change as its temperature increases from -250°C to -210°C . What phase change takes place? (3.2)

44. After burning for three hours, a candle has lost half of its mass. Explain why this example does not violate the law of conservation of mass. (3.2)



45. Describe the difference between a physical change and a chemical change. (3.2)
46. Describe the characteristics of a mixture. (3.3)
47. Describe a method that could be used to separate each of the following mixtures. (3.3)
- iron filings and sand
 - sand and salt
 - the components of ink
 - helium and oxygen gases
48. “A mixture is the chemical bonding of two or more substances in any proportion.” Is this statement true or false. Explain.

49. Which of the following are the same and which are different? (3.3)
- a substance and a pure substance
 - a heterogeneous mixture and a solution
 - a substance and a mixture
 - a homogeneous mixture and a solution
50. Describe how a homogeneous mixture differs from a heterogeneous mixture. (3.3)
51. A chemistry professor has developed a laboratory task to give her students practical experience using basic separation techniques. She prepares a liquid solution of water and another compound. Assuming you are a student in the class, name the technique you would use to separate and identify the components. Give specific details of the method.
52. State the definition of an element. (3.4)
53. Name the elements contained in the following compounds. (3.4)
- | | |
|-------------------------------|--|
| a. sodium chloride (NaCl) | c. ethanol (C ₂ H ₆ O) |
| b. ammonia (NH ₃) | d. bromine (Br ₂) |
54. How many naturally occurring elements are found on Earth? Approximately how many synthetic elements have been identified? (3.4)
55. What was Dmitri Mendeleev's major contribution to the field of chemistry? (3.4)
56. Is it possible to distinguish between an element and a compound? Explain. (3.4)
57. How are the properties of a compound related to those of the elements that comprise it? (3.4)
58. How are the elements contained within a group on the periodic table related? (3.4)
59. Which law states that a compound always contains the same elements in the same proportion by mass? (3.4)

Mastering Problems

Properties of Matter (3.1)

60. A scientist is given the task of identifying an unknown compound on the basis of its physical properties. The substance is a white solid at room temperature. Attempts to determine its boiling point were unsuccessful. Using **Table 3-1**, name the unknown compound.

Conservation of Mass (3.2)

61. A 28.0-g sample of nitrogen gas combines completely with 6.0 g of hydrogen gas to form ammonia. What is the mass of ammonia formed?

62. A substance breaks down into its component elements when it is heated. If 68.0 grams of the substance is present before it is heated, what is the combined mass of the component elements after heating?
63. A 13.0-g sample of X combines with a 34.0-g sample of Y to form the compound XY₂. What is the mass of the reactants?
64. Sodium chloride can be formed by the reaction of sodium metal and chlorine gas. If 45.98 g of sodium combines with an excess of chlorine gas to form 116.89 g sodium chloride, what mass of chlorine gas is used in the reaction?
65. Copper sulfide is formed when copper and sulfur are heated together. In this reaction, 127 g of copper reacts with 41 g of sulfur. After the reaction is complete, 9 g of sulfur remains unreacted. What is the mass of copper sulfide formed?

Law of Definite Proportions (3.4)

66. A 25.3-g sample of an unknown compound contains 0.8 g of oxygen. What is the percent by mass of oxygen in the compound?
67. Magnesium combines with oxygen to form magnesium oxide. If 18.06 g of magnesium reacts completely with 6.96 g of oxygen, what is the percent by mass of oxygen in magnesium oxide?
68. When mercury oxide is heated, it decomposes into mercury and oxygen. If 28.4 g of mercury oxide decomposes, producing 2.0 g oxygen, what is the percent by mass of mercury in mercury oxide?

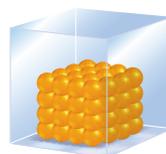
Law of Multiple Proportions (3.4)

69. Carbon reacts with oxygen to form two different compounds. Compound I contains 4.82 g carbon for every 6.44 g of oxygen. Compound II contains 20.13 g carbon for every 53.7 g of oxygen. What is the ratio of carbon to a fixed mass of oxygen for the two compounds?

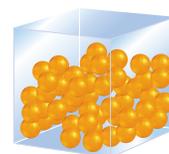
Mixed Review

Sharpen your problem-solving skills by answering the following.

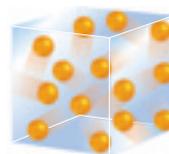
70. Which state of matter is the most compressible? The least? Explain why.



Solid



Liquid



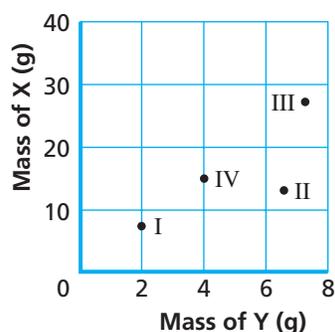
Gas



- 71.** Classify each of the following as a homogeneous mixture or a heterogeneous mixture. (3.3)
- brass (an alloy of zinc and copper)
 - a salad
 - blood
 - powder drink mix dissolved in water
- 72.** Phosphorus combines with hydrogen to form phosphine. In this reaction, 123.9 g of phosphorus combines with excess hydrogen to produce 129.9 g of phosphine. After the reaction, 310 g of hydrogen remains unreacted. What mass of hydrogen is used in the reaction? What was the initial mass of hydrogen before the reaction?
- 73.** A sample of a certain lead compound contains 6.46 grams of lead for each gram of oxygen. A second sample has a mass of 68.54 g and contains 28.76 g of oxygen. Are the two samples the same?

Thinking Critically

- 74. Applying Concepts** Air is a mixture of many gases, primarily nitrogen, oxygen, and argon. Could distillation be used to separate air into its component gases? Explain.
- 75. Interpreting Data** A compound contains elements X and Y. Four samples with different masses were analyzed, and the masses of X and Y in each sample were plotted on a graph. The samples are labeled I, II, III, and IV.



- Which samples are from the same compound? How do you know?
- What is the approximate ratio of mass X to mass Y in the samples that are from the same compound?
- What is the approximate ratio of mass X to mass Y in the sample(s) that are not from the same compound?

Writing in Chemistry

- 76.** Select a synthetic element and prepare a short written report on its development. Be sure to cover recent discoveries, list major research centers that conduct this type of research, and describe the properties of the synthesized element.
- 77.** Research the life of a scientist, other than Mendeleev, who contributed to the development of the modern periodic table of elements. Write a brief biography of this person and detail his or her scientific accomplishments.
- 78.** The results and interpretations of chemistry experiments and studies are recorded and published in literally hundreds of scientific journals around the world. Visit the local library and look at several of the articles in a chemistry journal such as *The Journal of the American Chemical Society*. Write a brief summary of your observations regarding the format and style of writing in chemistry.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 79.** What is chemistry? (Chapter 1)
- 80.** What is mass? Weight? (Chapter 1)
- 81.** Express the following in scientific notation. (Chapter 2)
- 34 500
 - 2665
 - 0.9640
 - 789
 - 75 600
 - 0.002 189
- 82.** Perform the following operations. (Chapter 2)
- $10^7 \times 10^3$
 - $(1.4 \times 10^{-3}) \times (5.1 \times 10^{-5})$
 - $(2 \times 10^{-3}) \times (4 \times 10^5)$
- 83.** Convert 65°C to Kelvins. (Chapter 2)
- 84.** Graph the following data. What is the slope of the line? (Chapter 2)



Energy Released by Carbon

Mass (g)	Energy released (kJ)
1.00	33
2.00	66
3.00	99
4.00	132

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Tables Use the table to answer questions 1 and 2.

Mass Analysis of Two Chlorine–Fluorine Compound Samples				
Sample	Mass chlorine (g)	Mass fluorine (g)	%Cl	%F
I	13.022	6.978	65.11	34.89
II	5.753	9.248	?	?

- What are the values for %Cl and %F, respectively, for Sample II?
 - 0.622 and 61.65
 - 61.65 and 38.35
 - 38.35 and 0.622
 - 38.35 and 61.65
- Which of the following statements best describes the relationship between the two samples?
 - The compound in Sample I is the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of definite proportions.
 - The compound in Sample I is the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of multiple proportions.
 - The compound in Sample I is not the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of proportions.
 - The compound in Sample I is not the same as in Sample II. Therefore, the mass ratio of Cl to F in both samples will obey the law of multiple proportions.
- After elements A and B react to completion in a closed container, the ratio of masses of A and B in the container will be the same as before the reaction. This is true because of the law of _____.
 - definite proportions
 - multiple proportions
 - conservation of mass
 - conservation of energy
- All of the following are physical properties of table sugar (sucrose) EXCEPT _____.
 - forms solid crystals at room temperature
 - appears as crystals white in color
 - breaks down into carbon and water vapor when heated
 - tastes sweet
- A substance is said to be in the solid state if _____.
 - it is hard and rigid
 - it can be compressed into a smaller volume
 - it takes the shape of its container
 - its matter particles are close together
- Na, K, Li, and Cs all share very similar chemical properties. In the periodic table of elements, they most likely belong to the same _____.
 - row
 - period
 - group
 - element
- A heterogeneous mixture _____.
 - cannot be separated by physical means.
 - is composed of distinct areas of composition.
 - is also called a solution.
 - has the same composition throughout.
- The percent by mass of sulfur in sulfuric acid, H_2SO_4 , is _____.
 - 32.69%
 - 64.13%
 - 16.31%
 - 48.57%
- Magnesium reacts explosively with oxygen to form magnesium oxide. All of the following are true of this reaction EXCEPT _____.
 - The mass of magnesium oxide produced equals the mass of magnesium consumed plus the mass of oxygen consumed.
 - The reaction describes the formation of a new substance.
 - The product of the reaction, magnesium oxide, is a chemical compound.
 - Magnesium oxide has physical and chemical properties similar to both oxygen and magnesium.
- Which of the following is NOT a chemical reaction?
 - dissolution of sodium chloride in water
 - combustion of gasoline
 - fading of wallpaper by sunlight
 - curdling of milk



TEST-TAKING TIP

When Eliminating, Cross It Out Consider each answer choice individually and cross out choices you've eliminated. If you can't write in the test booklet, use the scratch paper. List the answer choice letters on the scratch paper and cross them out there. You'll save time and stop yourself from choosing an answer you've mentally eliminated.

The Structure of the Atom

What You'll Learn

- ▶ You will identify the experiments that led to the development of the nuclear model of atomic structure.
- ▶ You will describe the structure of the atom and differentiate among the subatomic particles that comprise it.
- ▶ You will explain the relationship between nuclear stability and radioactivity.

Why It's Important

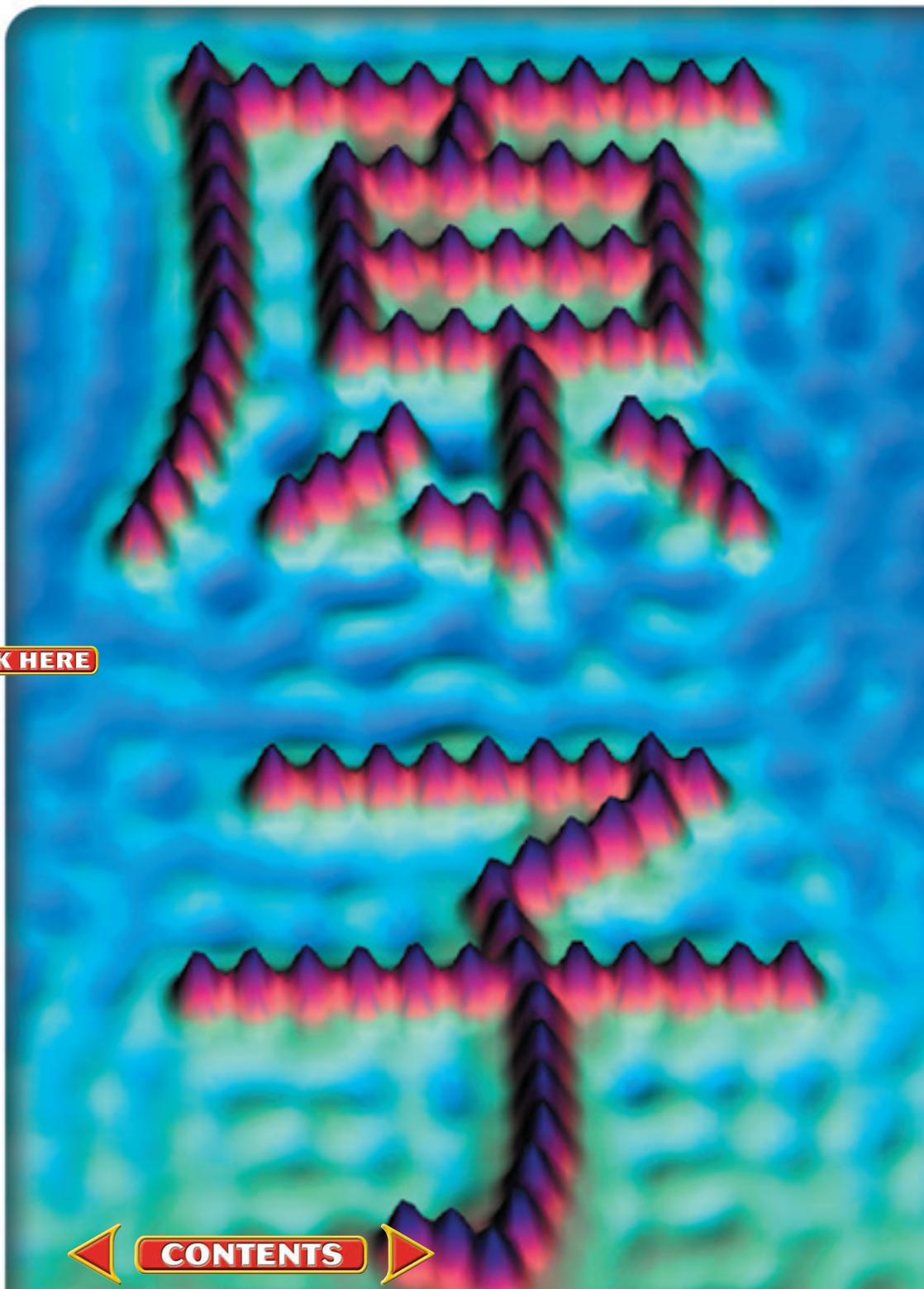
The world you know is made of matter, and all matter is composed of atoms. Understanding the structure of the atom is fundamental to understanding why matter behaves the way it does.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about atomic structure.

Not only can individual atoms be seen, but scientists now have the ability to arrange them into patterns and simple devices. The atoms shown here have been arranged to form the Japanese kanji characters for atom.



DISCOVERY LAB



Observing Electrical Charge

Electrical charge plays an important role in atomic structure and throughout chemistry. How can you observe the behavior of electrical charge using common objects?

Procedure

1. Cut out small round pieces of paper using the hole punch and spread them out on a table. Run a plastic comb through your hair. Bring the comb close to the pieces of paper. Record your observations.
2. Fold a 1-cm long portion of each piece of tape back on itself to form a handle. Stick two pieces of tape firmly to your desktop. Quickly pull both pieces of tape off of the desktop and bring them close together so that their non-sticky sides face each other. Record your observations.
3. Firmly stick one of the remaining pieces of tape to your desktop. Firmly stick the last piece of tape on top of the first. Quickly pull the pieces of tape as one from the desktop and then pull them apart. Bring the two tape pieces close together so that their non-sticky sides face each other. Record your observations.

Materials

metric ruler	10-cm long
plastic comb	piece of
hole punch	clear plastic
paper	tape (4)

Analysis

Use your knowledge of electrical charge to explain your observations. Which charges are similar? Which are different? How do you know?

Section

4.1

Early Theories of Matter

Objectives

- **Compare** and **contrast** the atomic models of Democritus and Dalton.
- **Define** an atom.

Vocabulary

Dalton's atomic theory
atom

Perhaps you have never seen a photo of individual atoms as shown on the previous page, but chances are you've heard of atoms ever since you were in elementary school. From atom smashers and atomic power to the reality of the atomic bomb, you are already familiar with many modern atom-based processes. Surprisingly, the idea that matter is composed of tiny particles (which we now call atoms) did not even exist a few thousand years ago. In fact, for more than a thousand years, great thinkers of their day argued against the idea that atoms existed. They were wrong. As you will see, the development of the concept of the atom and our understanding of atomic structure are fascinating stories involving scores of great thinkers and scientists.

The Philosophers

Science as we know it today did not exist several thousand years ago. No one knew what a controlled experiment was, and there were few tools for scientific exploration. In this setting, the power of mind and intellectual thought were considered the primary avenues to the truth. Curiosity sparked the interest of scholarly thinkers known as philosophers who considered the many mysteries of life. As they speculated about the nature of matter, many of the philosophers formulated explanations based on their own life experiences.

Figure 4-1

Many Greek philosophers thought matter was formed of air, earth, fire, and water. They also associated properties with each of the four basic components of matter. The pairings of opposite properties, such as hot and cold, and wet and dry, mirrored the symmetry and balance the philosophers observed in nature. These early nonscientific and incorrect beliefs were not completely dispelled until the 1800s.



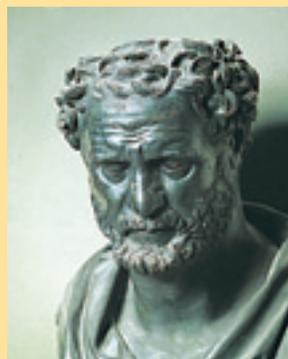
It wasn't surprising then, that many of them concluded that matter was composed of things such as earth, water, air, and fire. See **Figure 4-1**. It was also commonly accepted that matter could be endlessly divided into smaller and smaller pieces. While these early ideas were creative, there was no method for testing their validity.

The Greek philosopher Democritus (460–370 B.C.) was the first person to propose the idea that matter was not infinitely divisible. He believed matter was made up of tiny individual particles called *atomos*, from which the English word atom is derived. Democritus believed that atoms could not be created, destroyed, or further divided. Democritus and a summary of his ideas are shown in **Figure 4-2**.

While a fair amount of Democritus's ideas do not agree with modern atomic theory, his belief in the existence of atoms was amazingly ahead of his time. Despite this, his ideas did not turn out to be a major step toward our current understanding of matter. Over time, Democritus's ideas were met with criticism from other philosophers. "What holds the atoms together?" they asked. Democritus could not answer the question. Other criticisms came from Aristotle (384–322 B.C.), one of the most influential Greek philosophers. Aristotle is shown in **Figure 4-3**. He rejected the atomic "theory" entirely

Figure 4-2

The Greek philosopher Democritus (460–370 B.C.) proposed the concept of the atom more than two thousand years ago.



Democritus's Ideas

- Matter is composed of empty space through which atoms move.
- Atoms are solid, homogeneous, indestructible, and indivisible.
- Different kinds of atoms have different sizes and shapes.
- The differing properties of matter are due to the size, shape, and movement of atoms.
- Apparent changes in matter result from changes in the groupings of atoms and not from changes in the atoms themselves.



Aristotle

- One of the most influential philosophers.
- Wrote extensively on many subjects, including politics, ethics, nature, physics, and astronomy.
- Most of his writings have been lost through the ages.

because it did not agree with his own ideas on nature. One of Aristotle's major criticisms concerned the idea that atoms moved through empty space. He did not believe that the "nothingness" of empty space could exist. Unable to answer the challenges to his ideas, Democritus's atomic theory was eventually rejected.

In fairness to Democritus, it was impossible for him or anyone else of his time to determine what held the atoms together. More than two thousand years would pass before the answer was known. However, it is important to realize that Democritus's ideas were just that—ideas and not science. Without the benefit of being able to conduct controlled experiments, Democritus could not test to see if his ideas were valid.

Unfortunately for the advancement of science, Aristotle was able to gain wide acceptance for his ideas on nature—ideas that denied the existence of atoms. Incredibly, the influence of Aristotle was so great and the development of science so primitive that his denial of the existence of atoms went largely unchallenged for two thousand years!

John Dalton

Although the concept of the atom was revived in the 18th century, it took the passing of another hundred years before significant progress was made. The work done in the 19th century by John Dalton (1766–1844), a schoolteacher in England, marks the beginning of the development of modern atomic theory. Dalton revived and revised Democritus's ideas based upon the results of scientific research he conducted. The main points of **Dalton's atomic theory** are shown in **Figure 4-4**.



Dalton's Atomic Theory

- All matter is composed of extremely small particles called atoms.
- All atoms of a given element are identical, having the same size, mass, and chemical properties. Atoms of a specific element are different from those of any other element.
- Atoms cannot be created, divided into smaller particles, or destroyed.
- Different atoms combine in simple whole-number ratios to form compounds.
- In a chemical reaction, atoms are separated, combined, or rearranged.

Figure 4-3

The Greek philosopher Aristotle (384–322 B.C.) was influential in the rejection of the concept of the atom.



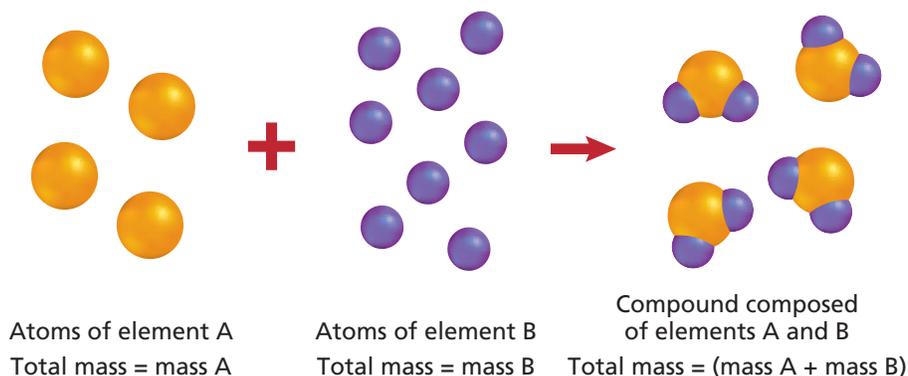
Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Figure 4-4

John Dalton's (1766–1844) atomic theory was a breakthrough in our understanding of matter.

Figure 4-5

Dalton's atomic theory explains the conservation of mass when a compound forms from its component elements. Atoms of elements A and B combine in a simple whole-number ratio, in this case two B atoms for each A atom, to form a compound. Because the atoms are only rearranged in the chemical reactions, their masses are conserved.



History

CONNECTION

Alchemy, a popular pursuit during the Middle Ages, was the search for a way to transform common metals into gold. “Scientists” who studied alchemy did not succeed in producing gold, but their experiments revealed the properties of various metals and helped advance our understanding of matter. During the seventeenth century some alchemists began focusing on identifying new compounds and reactions—these alchemists were probably the first true chemists. By the twentieth century, scientists could transform atoms of some elements into atoms of other elements by bombarding them with neutrons. Even this process, however, has not succeeded in producing gold.



The advancements in science since Democritus's day served Dalton well, as he was able to perform experiments that allowed him to refine and verify his theories. Dalton studied numerous chemical reactions, making careful observations and measurements along the way. He was able to accurately determine the mass ratios of the elements involved in the reactions. Based on this research, he proposed his atomic theory in 1803. In many ways Democritus's and Dalton's theories are similar. What similarities and differences can you find between the two theories?

Recall from Chapter 3 that the law of conservation of mass states that mass is conserved in any process, such as a chemical reaction. Dalton's atomic theory easily explains the conservation of mass in chemical reactions as being the result of the separation, combination, or rearrangement of atoms—atoms that are not created, destroyed, or divided in the process. The formation of a compound from the combining of elements and the conservation of mass during the process are shown in **Figure 4-5**. Dalton's convincing experimental evidence and clear explanation of the composition of compounds and conservation of mass led to the general acceptance of his atomic theory.

Was Dalton's atomic theory a huge step toward our current atomic model of matter? Yes. Was all of Dalton's theory accurate? No. As is often the case in science, Dalton's theory had to be revised as additional information was learned that could not be explained by the theory. As you will soon learn, Dalton was wrong about atoms being indivisible (they are divisible into several subatomic particles) and about all atoms of a given element having identical properties (atoms of an element may have slightly different masses).

Defining the Atom

Many experiments since Dalton's time have proven that atoms do actually exist. So what exactly then is the definition of an atom? To answer this question, consider a gold ring. Suppose you decide to grind the ring down into a pile of gold dust. Each fragment of gold dust still retains all of the properties of gold. If it were possible—which it is not without special equipment—you could continue to divide the gold dust particles into still smaller particles. Eventually you would encounter a particle that could not be divided any further and still retain the properties of gold. This smallest particle of an element that retains the properties of the element is called an **atom**.

Just how small is a typical atom? To get some idea of its size, consider the population of the world. In the year 2000, the world population was approximately 6 000 000 000 (six billion) people. By comparison, a typical solid copper penny contains almost five billion times as many atoms of copper!

World population	6 000 000 000
Atoms in a penny	29 000 000 000 000 000 000

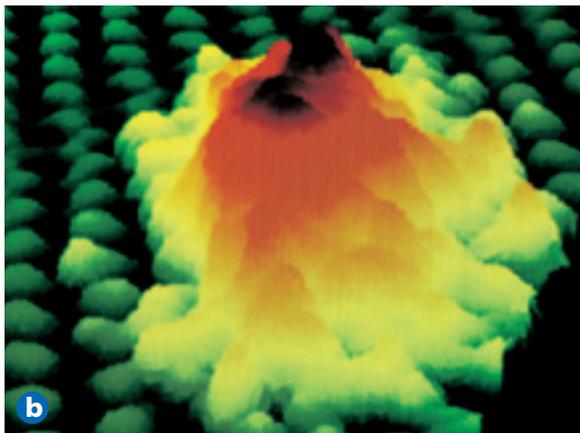
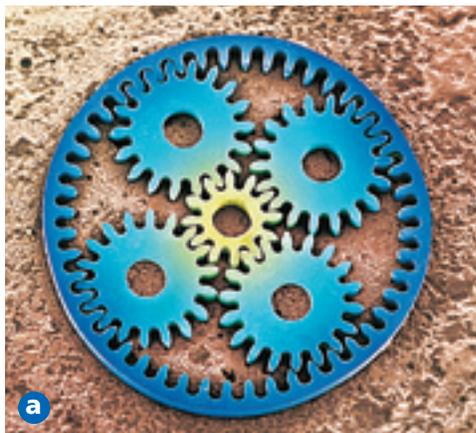


Figure 4-6

- a** This colorized scanning electron micrograph shows a microgear mechanism built by the precise arrangement of individual atoms.
- b** A mound of gold atoms (yellow, red, and brown) is easily discerned from the graphite substrate (green) it rests on.

The diameter of a single copper atom is 1.28×10^{-10} m. Placing six billion copper atoms (equal in number to the world's population) side by side would result in a line of copper atoms less than one meter long.

You might think that because atoms are so small there would be no way to actually see them. However, an instrument called the scanning tunneling microscope allows individual atoms to be seen. Do the **problem-solving LAB** on page 96 to analyze scanning tunneling microscope images and gain a better understanding of atomic size. As **Figures 4-6a** and **4-6b** illustrate, not only can individual atoms be seen, scientists are now able to move individual atoms around to form shapes, patterns, and even simple machines. This capability has led to the exciting new field of nanotechnology. The promise of nanotechnology is molecular manufacturing—the atom-by-atom building of machines the size of molecules. As you'll learn in later chapters, a molecule is a group of atoms that are bonded together and act as a unit. While this technology is not yet feasible for the production of consumer products, progress toward that goal has been made. To learn more about nanotechnology, read the **Chemistry and Society** at the end of this chapter.

The acceptance of atomic theory was only the beginning of our understanding of matter. Once scientists were fairly convinced of the existence of atoms, the next set of questions to be answered emerged. What is an atom like? How are atoms shaped? Is the composition of an atom uniform throughout, or is it composed of still smaller particles? While many scientists researched the atom in the 1800s, it was not until almost 1900 that answers to some of these questions were found. The next section explores the discovery of subatomic particles and the further evolution of atomic theory.

Section 4.1 Assessment

1. Why were Democritus's ideas rejected by other philosophers of his time?
2. Define an atom using your own words.
3. Which statements in Dalton's original atomic theory are now considered to be incorrect? Describe how modern atomic theory differs from these statements.
4. **Thinking Critically** Democritus and Dalton both proposed the concept of atoms. Describe the method each of them used to reach the conclusion that atoms existed. How did Democritus's method hamper the acceptance of his ideas?
5. **Comparing and Contrasting** Compare and contrast the atomic theories proposed by Democritus and John Dalton.

Subatomic Particles and the Nuclear Atom

Objectives

- **Distinguish** between the subatomic particles in terms of relative charge and mass.
- **Describe** the structure of the nuclear atom, including the locations of the subatomic particles.

Vocabulary

cathode ray
electron
nucleus
proton
neutron

In 1839, American inventor Charles Goodyear accidentally heated a mixture of natural rubber and sulfur. The resulting reaction greatly strengthened the rubber. This new rubber compound revolutionized the rubber industry and was eventually used in the manufacturing of automobile tires. Accidental discoveries such as this have occurred throughout the history of science. Such is the case with the discovery of subatomic particles, the particles that make up atoms.

Discovering the Electron

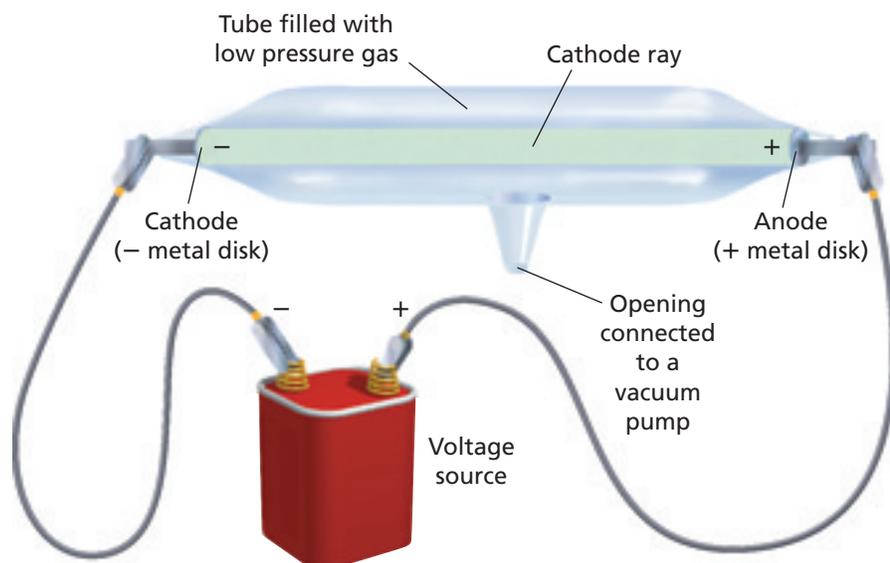
Has your hair ever clung to your comb? Have you ever received a shock from a metal doorknob after walking across a carpeted floor? Observations such as these led scientists in the 1800s to look for some sort of relationship between matter and electric charge. To explore the connection, some scientists wondered how electricity might behave in the absence of matter. With the help of the recently invented vacuum pump, they passed electricity through glass tubes from which most of the air (and most of the matter) had been removed.

A typical tube used by researchers studying the relationship between mass and charge is illustrated in **Figure 4-7**. Note that metal electrodes are located on opposite ends of the tube. The electrode connected to the negative terminal of the battery is called the cathode, and the electrode connected to the positive terminal is called the anode.

One day while working in a darkened laboratory, English physicist Sir William Crookes noticed a flash of light within one of the tubes. The flash was produced by some form of radiation striking a light-producing coating that had been applied to the end of the tube. Further work showed there were rays (radiation) traveling from the cathode to the anode within the tube. Because the ray of radiation originated from the cathode end of the tube, it became known as a **cathode ray**. The accidental discovery of the cathode ray led to the invention of one of the most important technological and social developments of the 20th century—the television. Television and computer monitor images are formed as radiation from the cathode strikes light-producing chemicals that coat the backside of the screen.

Figure 4-7

Examine the parts of a typical cathode ray tube. Note that the electrodes take on the charge of the battery terminal to which they are connected. The cathode ray travels from the cathode to the anode.



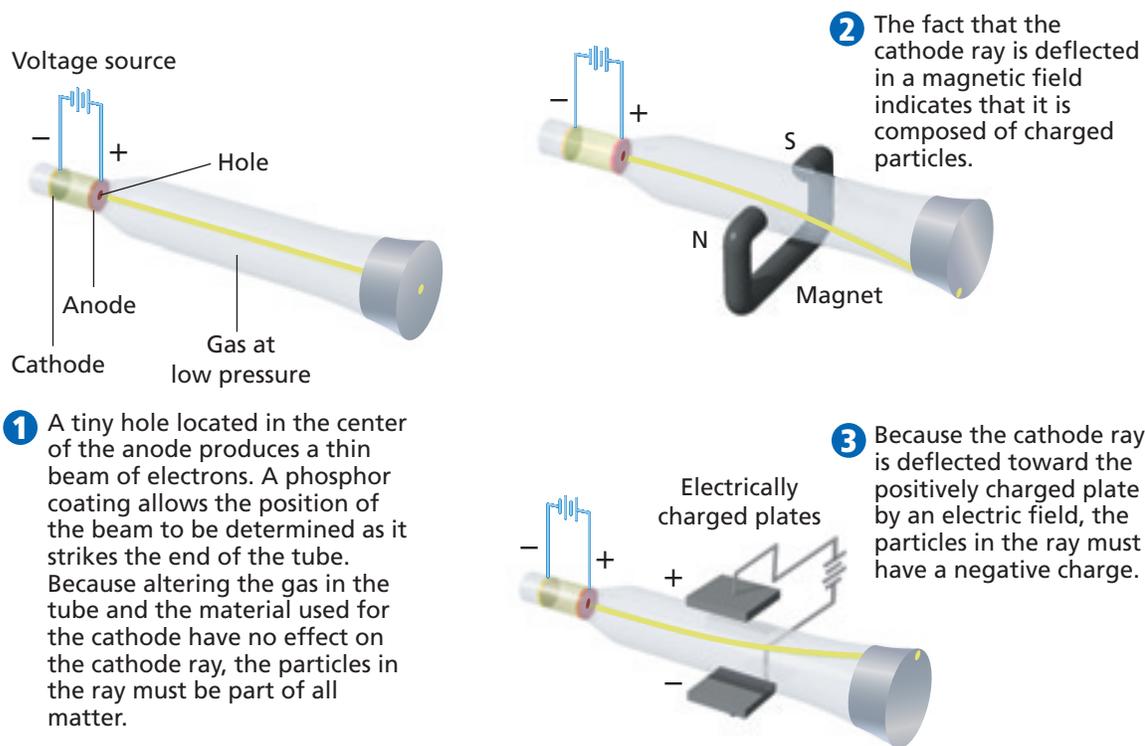


Figure 4-8

Multiple experiments helped determine the properties of cathode rays.

Scientists continued their research using cathode ray tubes, and by the end of the 1800s they were fairly convinced of the following:

- Cathode rays were actually a stream of charged particles.
- The particles carried a negative charge. (The exact value of the negative charge was not known, however.)

Because changing the type of electrode or varying the gas (at very low pressure) in the cathode ray tube did not affect the cathode ray produced, it was concluded that the ray's negative particles were found in all forms of matter. These negatively charged particles that are part of all forms of matter are now called **electrons**. The range of experiments used to determine the properties of the cathode ray are shown in **Figure 4-8**.

In spite of the progress made from all of the cathode ray tube experiments, no one had succeeded in determining the mass of a single cathode ray particle. Unable to measure the particle's mass directly, English physicist J.J. Thomson (1856–1940) began a series of cathode ray tube experiments in the late 1890s to determine the ratio of its charge to its mass. By carefully measuring the effect of both magnetic and electric fields on a cathode ray, Thomson was able to determine the charge-to-mass ratio of the charged particle. He then compared that ratio to other known ratios. Thomson concluded that the mass of the charged particle was much less than that of a hydrogen atom, the lightest known atom. The conclusion was shocking because it meant there were particles smaller than the atom. In other words, Dalton was wrong: Atoms were divisible into smaller subatomic particles. Because Dalton's atomic theory had become so widely accepted, and because Thomson's conclusion was so revolutionary, many fellow scientists found it hard to believe this new discovery. But Thomson was correct. He had identified the first subatomic particle—the electron.

The next significant development came in 1909, when an American physicist named Robert Millikan (1868–1953) determined the charge of an electron.

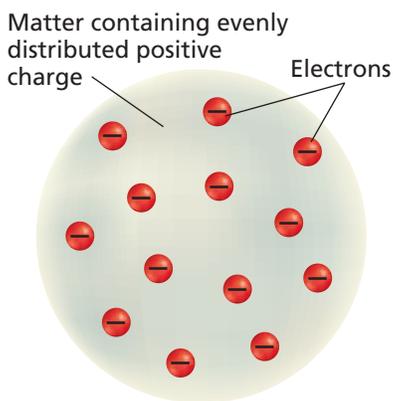


Figure 4-9

J.J. Thomson's plum pudding atomic model proposed that negatively charged electrons were distributed throughout a uniform positive charge.

So good was Millikan's experimental setup and technique that the charge he measured almost one hundred years ago is within 1% of the currently accepted value. This charge has since been equated to a single unit of negative charge; in other words, a single electron carries a charge of $1-$. Knowing the electron's charge and using the known charge-to-mass ratio, Millikan calculated the mass of a single electron.

$$\text{Mass of an electron} = 9.1 \times 10^{-28} \text{ g} = \frac{1}{1840} \text{ mass of a hydrogen atom}$$

As you can see, the mass of an electron is extremely small.

The existence of the electron and the knowledge of some of its properties raised some interesting new questions about the nature of atoms. It was known that matter is neutral. You know matter is neutral from everyday experience; you do not receive an electrical shock (except under certain conditions) when you touch an object. If electrons are part of all matter and they possess a negative charge, how is it that all matter is neutral? Also, if the mass of an electron is so extremely small, what accounts for the rest of the mass in a typical atom?

In an attempt to answer these questions, J.J. Thomson proposed a model of the atom that became known as the plum pudding model. As you can see in **Figure 4-9**, Thomson's model consisted of a spherically shaped atom composed of a uniformly distributed positive charge within which the individual negatively charged electrons resided. A more modern name for this model might be the chocolate-chip cookie dough model, where the chocolate chips are the electrons and the dough is the uniformly distributed positive charge. As you are about to learn, the plum pudding model of the atom did not last for very long.

The Nuclear Atom

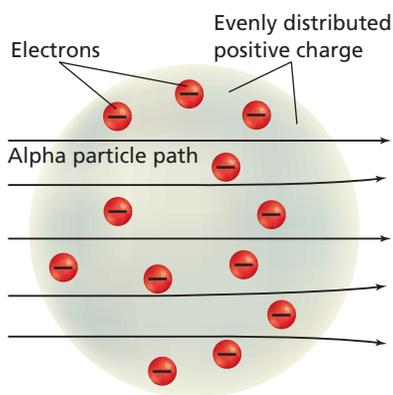
The story of the atom continues with the role played by Ernest Rutherford (1871–1937). As a youth, Rutherford, who was born in New Zealand, placed second in a scholarship competition to attend the prestigious Cambridge University in England. He received a fortunate break when the winner of the competition decided not to attend. By 1908, Rutherford won the Nobel Prize in chemistry and had many significant discoveries to his credit.

In 1911 Rutherford became interested in studying how positively charged alpha particles (radioactive particles you will learn more about later in this chapter) interacted with solid matter. A small group of scientists that included Rutherford designed and conducted an experiment to see if alpha particles would be deflected as they passed through a thin foil of gold. In the experiment, a narrow beam of alpha particles was aimed at a thin sheet of gold foil. A zinc sulfide coated screen surrounding the gold foil produced a flash of light whenever it was struck by an alpha particle. By noting where the flashes occurred, the scientists could determine if the atoms in the gold foil deflected the alpha particles.

Rutherford was aware of Thomson's plum pudding model of the atom and expected only minor deflections of the alpha particles. He thought the paths of the massive (relative to electrons) and fast-moving alpha particles would be only slightly altered by a nearby encounter or collision with an electron. And because the positive charge within the gold atoms was thought to be uniformly distributed, he thought it would not alter the paths of the alpha particles either. **Figure 4-10** shows the results Rutherford anticipated from the experiment. After a few days of testing, Rutherford and his fellow scientists were amazed to discover that a few of the alpha particles were deflected at

Figure 4-10

Ernest Rutherford expected most of the fast-moving and relatively massive alpha particles to pass straight through the gold atoms. He also expected a few of the alpha particles to be slightly deflected by the electrons in the gold atoms.



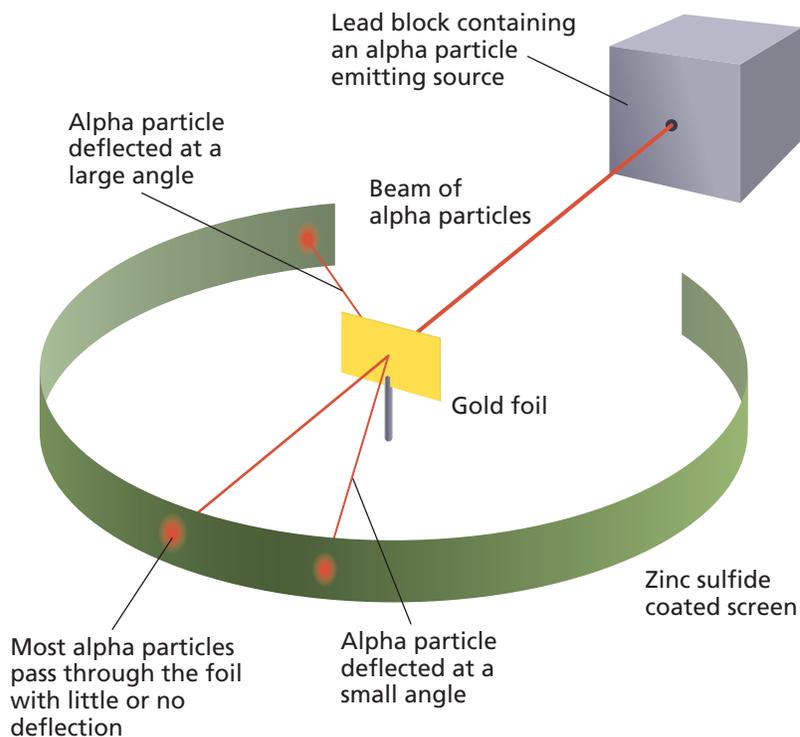


Figure 4-11

As Rutherford expected, most all of the alpha particles passed straight through the gold foil, without deflection. Surprisingly, however, some alpha particles were scattered at small angles, and on a few occasions they were deflected at very large angles.

very large angles. Several particles were even deflected straight back toward the source of the alpha particles. Rutherford likened the results to firing a large artillery shell at a sheet of paper and having the shell come back and hit you! These results, shown in **Figure 4-11**, were truly astounding.

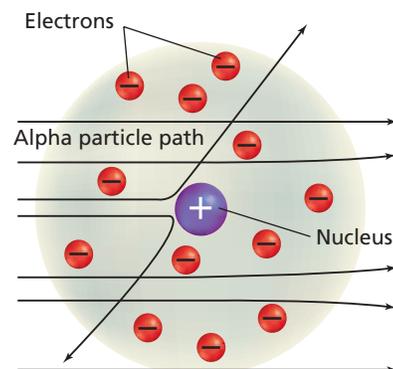
Rutherford concluded that the plum pudding model was incorrect because it could not explain the results of the gold foil experiment. He set out to develop a new atomic model based upon his findings. Considering the properties of the alpha particles and the electrons, and the frequency of the deflections, he calculated that an atom consisted mostly of empty space through which the electrons move. He also concluded that there was a tiny, dense region, which he called the **nucleus**, centrally located within the atom that contained all of an atom's positive charge and virtually all of its mass. Because the nucleus occupies such a small space and contains most of an atom's mass, it is incredibly dense. Just how dense? If a nucleus were the size of the dot in the exclamation point at the end of this sentence, its mass would be approximately as much as that of 70 automobiles!

According to Rutherford's new nuclear atomic model, most of an atom consists of electrons moving rapidly through empty space. The electrons move through the available space surrounding the nucleus and are held within the atom by their attraction to the positively charged nucleus. The volume of space through which the electrons move is huge compared to the volume of the nucleus. A typical atom's diameter, which is defined by the volume of space through which the electrons move, is approximately 10 000 times the diameter of the nucleus. To put this in perspective, if an atom had a diameter of two football fields, the nucleus would be the size of a nickel!

The concentrated positive charge in the nucleus explains the deflection of the alpha particles—the repulsive force produced between the positive nucleus and the positive alpha particles causes the deflections. Alpha particles closely approaching the nucleus were deflected at small angles, while alpha particles directly approaching the nucleus were deflected at very large angles. You can see in **Figure 4-12** how Rutherford's nuclear atomic model explained the

Figure 4-12

Rutherford's nuclear model of the atom explains the results of the gold foil experiment. Most alpha particles pass straight through, being only slightly deflected by electrons, if at all. The strong force of repulsion between the positive nucleus and the positive alpha particles causes the large deflections.



results of the gold foil experiment. The nuclear model also explains the neutral nature of matter: the positive charge of the nucleus balancing the negative charge of the electrons. However, the model still could not account for all of the atom's mass. Another 20 years would pass before this mystery was solved.

Completing the Atom—The Discovery of Protons and Neutrons

By 1920, eight years after his revolutionary gold foil experiment, Rutherford had refined the concept of the nucleus. He concluded that the nucleus contained positively charged particles called protons. A **proton** is a subatomic particle carrying a charge equal to but opposite that of an electron; that is, a proton has a positive charge of $1+$.

In 1932, Rutherford's coworker, English physicist James Chadwick (1891–1974), showed that the nucleus also contained another subatomic particle, a neutral particle called the neutron. A **neutron** has a mass nearly equal to that of a proton, but it carries no electrical charge. Thus, three subatomic particles are the fundamental building blocks from which all atoms are

problem-solving LAB

Interpreting STM Images

Measuring The invention of the scanning tunneling microscope (STM) in 1981 gave scientists the ability to visualize individual atoms, and also led to their being able to manipulate the positions of individual atoms. Use the information shown in the STM images to interpret sizes and make measurements.

Analysis

Figure A is an STM image of silicon atoms that have been bonded together in a hexagonal pattern. The image is of an area 18.1 nm wide by 19.0 nm high ($1 \text{ nm} = 1 \times 10^{-9} \text{ m}$).

Figure B is an STM image of 48 iron atoms that have been arranged into a circular "corral." The corral has a diameter of 1426 nm. There is a single electron trapped inside the "corral."

Thinking Critically

- Using a metric ruler and the dimensions of **Figure A** given above, develop a scale for making measurements off of the image. Use your scale to estimate the distance between adjacent silicon nuclei forming a hexagon.
- What evidence is there that an electron is trapped inside the "corral" of iron atoms in **Figure B**? Estimate the distance between adjacent iron atoms. (Hint: Use the number of atoms and the formula $\text{circumference} = \pi \times \text{diameter}$.)

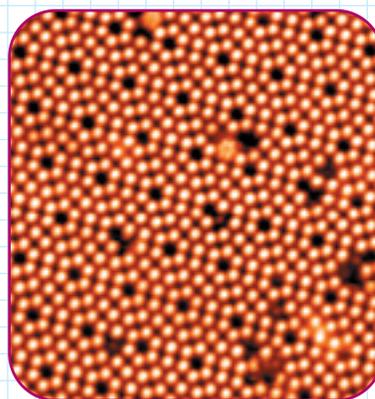


Figure A

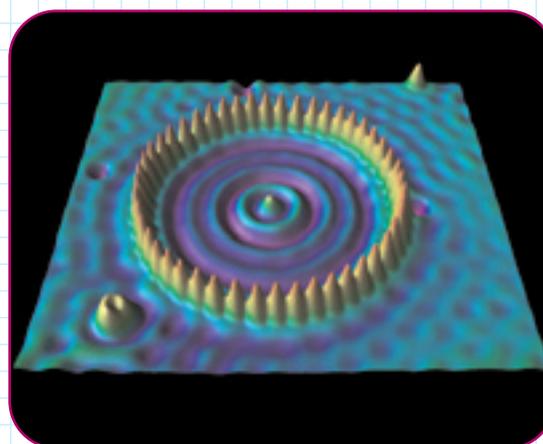


Figure B

Table 4-1

Properties of Subatomic Particles					
Particle	Symbol	Location	Relative electrical charge	Relative mass	Actual mass (g)
Electron	e^-	In the space surrounding the nucleus	1-	$\frac{1}{1840}$	9.11×10^{-28}
Proton	p^+	In the nucleus	1+	1	1.673×10^{-24}
Neutron	n^0	In the nucleus	0	1	1.675×10^{-24}

made—the electron, the proton, and the neutron. Together, electrons, protons, and neutrons account for all of the mass of an atom. The properties of electrons, protons, and neutrons are summarized in **Table 4-1**.

You know an atom is an electrically neutral particle composed of electrons, protons, and neutrons. Atoms are spherically shaped, with a tiny, dense nucleus of positive charge surrounded by one or more negatively charged electrons. Most of an atom consists of fast-moving electrons traveling through the empty space surrounding the nucleus. The electrons are held within the atom by their attraction to the positively charged nucleus. The nucleus, which is composed of neutral neutrons (hydrogen's single-proton nucleus is an exception) and positively charged protons, contains all of an atom's positive charge and 99.97% of its mass. Since an atom is electrically neutral, the number of protons in the nucleus equals the number of electrons surrounding the nucleus. The features of a typical atom are shown in **Figure 4-13**. To gain more perspective on the size of typical atoms, do the **CHEMLAB** at the end of this chapter.

Subatomic particle research is still a major interest of modern scientists. In fact, the three subatomic particles you have just learned about have since been found to have their own structures. That is, they contain sub-subatomic particles. These particles will not be covered in this textbook because it is not understood if or how they affect chemical behavior. As you will learn in coming chapters, behavior can be explained by considering only an atom's electrons, protons, and neutrons.

You should now have a solid understanding of the structure of a typical atom. But what makes an atom of one element different from an atom of another element? In the next section, you'll find out.

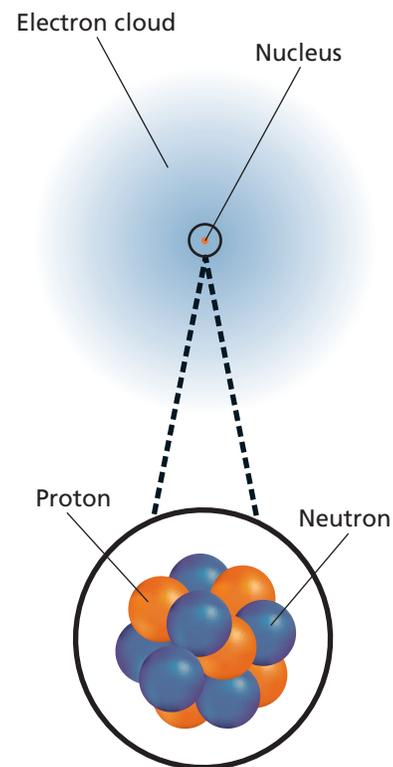


Figure 4-13

Atoms consist of a "cloud" of fast moving, negatively charged electrons surrounding a tiny, extremely dense nucleus containing positively charged protons and neutral neutrons. The nucleus contains virtually all of the atom's mass, but occupies only about one ten-thousandth the volume of the atom.

Section 4.2 Assessment

- Briefly evaluate the experiments that led to the conclusion that electrons were negatively charged particles found in all matter.
- Describe the structure of a typical atom. Be sure to identify where each subatomic particle is located.
- Make a table comparing the relative charge and mass of each of the subatomic particles.
- Thinking Critically** Compare and contrast Thomson's plum pudding atomic model with Rutherford's nuclear atomic model.
- Graphing** Make a timeline graph of the development of modern atomic theory. Be sure to include the discovery of each subatomic particle.

Objectives

- **Explain** the role of atomic number in determining the identity of an atom.
- **Define** an isotope and **explain** why atomic masses are not whole numbers.
- **Calculate** the number of electrons, protons, and neutrons in an atom given its mass number and atomic number.

Vocabulary

atomic number
isotope
mass number
atomic mass unit (amu)
atomic mass

Look at the periodic table on the inside back cover of this textbook. As you can see, there are more than 110 different elements. This means that there are more than 110 different kinds of atoms. What makes an atom of one element different from an atom of another element? You know that all atoms are made up of electrons, protons, and neutrons. Thus, you might suspect that atoms somehow differ in the number of these particles. If so, you are correct.

Atomic Number

Not long after Rutherford's gold foil experiment, the English scientist Henry Moseley (1887–1915) discovered that atoms of each element contain a unique positive charge in their nuclei. Thus, the number of protons in an atom identifies it as an atom of a particular element. The number of protons in an atom is referred to as the element's **atomic number**. Look again at the periodic table and you will see that the atomic number determines the element's position in the table. Consider hydrogen, located at the top left of the table. The information provided by the periodic table for hydrogen is shown in **Figure 4-14**. Note that above the symbol for hydrogen (H), you see the number 1. This number, which corresponds to the number of protons in a hydrogen atom, is the atomic number of hydrogen. Hydrogen atoms always contain a single proton. Moving across the periodic table to the right, you'll next come to helium (He). Helium has an atomic number of 2, and thus has two protons in its nucleus. The next row begins with lithium (Li), atomic number 3, followed by beryllium (Be), atomic number 4, and so on. As you can see, the periodic table is organized left-to-right and top-to-bottom by increasing atomic number. How many protons does a gold atom contain? A silver atom?

Remember that because all atoms are neutral, the number of protons and electrons in an atom must be equal. Thus, once you know the atomic number of an element, you know both the number of protons and the number of electrons an atom of that element contains.

$$\text{Atomic number} = \text{number of protons} = \text{number of electrons}$$

For instance, an atom of lithium, atomic number of 3, contains three protons and three electrons. How many electrons does an atom of element 97 contain?

Figure 4-14

The atomic number of an element equals the positive charge contained in its nucleus.

Hydrogen	Chemical name
1	Atomic number
H	Chemical symbol
1.008	Average atomic mass

Hydrogen, with an atomic number of 1, is the first element in the periodic table. A hydrogen atom has one proton and a charge of 1+ in its nucleus.

EXAMPLE PROBLEM 4-1

Using Atomic Number

Complete the following table.

Composition of Several Elements				
	Element	Atomic number	Protons	Electrons
a.	Pb	82	—	—
b.	—	—	8	—
c.	—	—	—	30

1. Analyze the Problem

You are given the information in the table. Apply the relationship among atomic number, number of protons, and number of electrons to complete most of the table. Once the atomic number is known, use the periodic table to identify the element.

2. Solve for the Unknown

Apply the atomic number relationship and then consult the periodic table to identify the element.

- Atomic number = number of protons = number of electrons
 $82 = \text{number of protons} = \text{number of electrons}$
Element 82 is lead (Pb).
- Atomic number = number of protons = number of electrons
Atomic number = $8 = \text{number of electrons}$
Element 8 is oxygen (O).
- Atomic number = number of protons = number of electrons
Atomic number = number of protons = 30
Element 30 is zinc (Zn).

The completed table is shown below.

Composition of Several Elements				
	Element	Atomic number	Protons	Electrons
a.	Pb	82	82	82
b.	O	8	8	8
c.	Zn	30	30	30

3. Evaluate the Answer

The answers agree with atomic numbers and element symbols given in the periodic table.

PRACTICE PROBLEMS

- How many protons and electrons are in each of the following atoms?
 - boron
 - radon
 - platinum
 - magnesium
- An atom of an element contains 66 electrons. What element is it?
- An atom of an element contains 14 protons. What element is it?



For more practice with problems using atomic numbers, go to **Supplemental Practice Problems** in Appendix A.

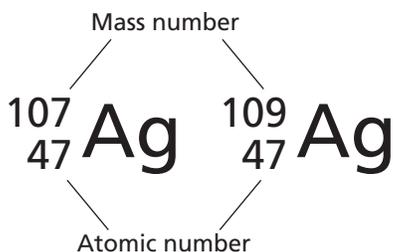


Figure 4-15

Ag is the chemical symbol for the silver used in these coins. The silver in each coin is comprised of 51.84% silver-107 ($^{107}_{47}\text{Ag}$) isotopes and 48.16% silver-109 ($^{109}_{47}\text{Ag}$) isotopes.

Isotopes and Mass Number

Earlier you learned that Dalton's atomic theory was wrong about atoms being indivisible. It was also incorrect in stating that all atoms of a particular element are identical. While it is true that all atoms of a particular element have the same number of protons and electrons, the number of neutrons on their nuclei may differ. For example, there are three different types of potassium atoms. All three types contain 19 protons (and thus 19 electrons). However, one type of potassium atom contains 20 neutrons, another contains 21 neutrons, and still another 22 neutrons. Atoms such as these, with the same number of protons but different numbers of neutrons, are called **isotopes**.

In nature most elements are found as a mixture of isotopes. Usually, no matter where a sample of an element is obtained, the relative abundance of each isotope is constant. For example, in a banana, which is a rich source of potassium, 93.25% of the potassium atoms have 20 neutrons, 6.7302% will have 22 neutrons, and a scant 0.0117% will have 21 neutrons. In another banana, or in a totally different source of potassium, the percentage composition of the potassium isotopes will still be the same.

As you might expect, the isotopes do differ in mass. Isotopes containing more neutrons have a greater mass. In spite of differences in mass and the number of neutrons, isotopes of an atom have essentially the same chemical behavior. Why? Because, as you'll learn in greater detail later in this textbook, chemical behavior is determined by the number of electrons an atom has, not by its number of neutrons and protons. To make it easy to identify each of the various isotopes of an element, chemists add a number after the element's name. The number that is added is called the **mass number**, and it represents the sum of the number of protons and neutrons in the nucleus. For example, the potassium isotope with 19 protons and 20 neutrons has a mass number of 39 ($19 + 20 = 39$), and the isotope is called potassium-39. The potassium isotope with 19 protons and 21 neutrons has a mass number of 40 ($19 + 21 = 40$), and is called potassium-40. What is the mass number and name of the potassium isotope with 19 protons and 22 neutrons?

Chemists often write out isotopes using a shortened type of notation involving the chemical symbol, atomic number, and mass number, as shown in **Figure 4-15**. Note that the mass number is written as a superscript to the left of the chemical symbol, and the atomic number is written as a subscript to the left of the chemical symbol. The three potassium isotopes you have just learned

Figure 4-16

The three naturally occurring potassium isotopes are potassium-39, potassium-40, and potassium-41. How do their masses compare? Their chemical properties?

	Potassium-39	Potassium-40	Potassium-41
Protons	19	19	19
Neutrons	20	21	22
Electrons	19	19	19

about are summarized in **Figure 4-16**. The number of neutrons in an isotope can be calculated from the atomic number and mass number.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

EXAMPLE PROBLEM 4-2

Using Atomic Number and Mass Number

A chemistry laboratory has analyzed the composition of isotopes of several elements. The composition data is given in the table at the right. Data for one of neon's three isotopes is given in the table. Determine the number of protons, electrons, and neutrons in the isotope of neon. Name the isotope and give its symbol.

1. Analyze the Problem

You are given some data for neon in the table. The symbol for neon can be found from the periodic table. From the atomic number, the number of protons and electrons in the isotope are known. The number of neutrons in the isotope can be found by subtracting the atomic number from the mass number.

Known

Element: neon
Atomic number = 10
Mass number = 22

Unknown

Number of protons,
electrons, and neutrons = ?
Name of isotope = ?
Symbol for isotope = ?

2. Solve for the Unknown

The number of protons equals the number of electrons which equals the atomic number.

Number of protons = number of electrons = atomic number = 10

Use the atomic number and the mass number to calculate the number of neutrons.

Number of neutrons = mass number - atomic number

Number of neutrons = 22 - 10 = 12

Use the element name and mass number to write the isotope's name.

neon-22

Use the chemical symbol, mass number, and atomic number to write out the isotope in symbolic notation form.

${}_{10}^{22}\text{Ne}$

3. Evaluate the Answer

The relationships among number of electrons, protons, and neutrons have been applied correctly. The isotope's name and symbol are in the correct format.

Isotope Composition Data

Element	Atomic number	Mass number
a. Neon	10	22
b. Calcium	20	46
c. Oxygen	8	17
d. Iron	26	57
e. Zinc	30	64
f. Mercury	80	204

PRACTICE PROBLEM

14. Determine the number of protons, electrons, and neutrons for isotopes b. through f. in the table above. Name each isotope, and write its symbol.



For more practice with problems using atomic number and mass number, go to **Supplemental Practice Problems** in Appendix A.

Mass of Individual Atoms

Recall from **Table 4-1** that the masses of both protons and neutrons are approximately 1.67×10^{-24} g. While this is a very small mass, the mass of an electron is even smaller—only about $\frac{1}{1840}$ that of a proton or neutron. Because these extremely small masses expressed in scientific notation are difficult to work with, chemists have developed a method of measuring the mass of an atom relative to the mass of a specifically chosen atomic standard. That standard is the carbon-12 atom. Scientists assigned the carbon-12 atom a mass of exactly 12 atomic mass units. Thus, one **atomic mass unit (amu)** is defined as $\frac{1}{12}$ the mass of a carbon-12 atom. Although a mass of 1 amu is very nearly equal to the mass of a single proton or a single neutron, it is important to realize that the values are slightly different. As a result, the mass of silicon-30, for example, is 29.974 amu, and not 30 amu. **Table 4-2** gives the masses of the subatomic particles in terms of amu.

Table 4-2

Masses of Subatomic Particles	
Particle	Mass (amu)
Electron	0.000 549
Proton	1.007 276
Neutron	1.008 665

Because an atom's mass depends mainly on the number of protons and neutrons it contains, and because protons and neutrons have masses close to 1 amu, you might expect the atomic mass of an element to always be very near a whole number. This, however, is often not the case. The explanation involves how atomic mass is defined. The **atomic mass** of an element is the weighted

miniLAB

Modeling Isotopes

Formulating Models Because they have different compositions, pre- and post-1982 pennies can be used to model an element with two naturally occurring isotopes. From the penny "isotope" data, the mass of each penny isotope and the average mass of a penny can be determined.

Materials bag of pre- and post-1982 pennies, balance

Procedure

1. Get a bag of pennies from your teacher, and sort the pennies by date into two groups: pre-1982 pennies and post-1982 pennies. Count and record the total number of pennies and the number of pennies in each group.
2. Use the balance to determine the mass of ten pennies from each group. Record each mass to the nearest 0.01 g. Divide the total mass of each group by ten to get the average mass of a pre- and post-1982 penny "isotope."

Analysis

1. Using data from step 1, calculate the percentage abundance of each group. To do this, divide the number of pennies in each group by the total number of pennies.
2. Using the percentage abundance of each "isotope" and data from step 2, calculate the



atomic mass of a penny. To do this, use the following equation for each "isotope."

$$\text{mass contribution} = (\% \text{ abundance})(\text{mass})$$

Sum the mass contributions to determine the atomic mass.

3. Would the atomic mass be different if you received another bag of pennies containing a different mixture of pre- and post-1982 pennies? Explain.
4. In step 2, instead of measuring and using the mass of a single penny of each group, the average mass of each type of penny was determined. Explain why.

Calculating the Weighted Average Atomic Mass of Chlorine

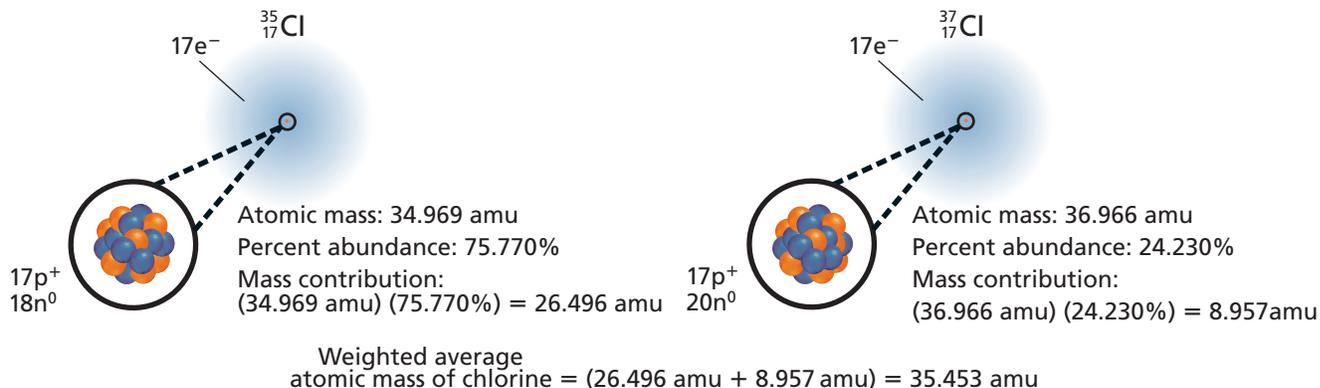


Figure 4-17

To determine the weighted average atomic mass of chlorine, the mass contribution of each of the two isotopes is calculated, and then those two values are added together.

average mass of the isotopes of that element. For example, the atomic mass of chlorine is 35.453 amu. Chlorine exists naturally as a mixture of about 75% chlorine-35 and 25% chlorine-37. Because atomic mass is a weighted average, the chlorine-35 atoms, which exist in greater abundance than the chlorine-37 atoms, have a greater effect in determining the atomic mass. The atomic mass of chlorine is calculated by summing the products of each isotope's percent abundance times its atomic mass. See **Figure 4-17**. For hands-on practice in calculating atomic mass, do the **miniLAB** on the previous page.

You can calculate the atomic mass of any element if you know its number of naturally occurring isotopes, their masses, and their percent abundances. The following Example Problem and Practice Problems will provide practice in calculating atomic mass.

EXAMPLE PROBLEM 4-3

Calculating Atomic Mass

Given the data in the table at the right, calculate the atomic mass of unknown element X. Then, identify the unknown element, which is used medically to treat some mental disorders.

1. Analyze the Problem

You are given the data in the table. Calculate the atomic mass by multiplying the mass of each isotope by its percent abundance and summing the results. Use the periodic table to confirm the calculation and identify the element.

Known

For isotope ^6X :
 mass = 6.015 amu
 abundance = 7.50% = 0.0750

For isotope ^7X :
 mass = 7.016 amu
 abundance = 92.5% = 0.925

Unknown

atomic mass of X = ? amu
 name of element X = ?

2. Solve for the Unknown

Calculate each isotope's contribution to the atomic mass.

For ^6X : Mass contribution = (mass)(percent abundance)
 mass contribution = (6.015 amu)(0.0750) = 0.451 amu

Isotope Abundance for Element X

Isotope	Mass (amu)	Percent abundance
^6X	6.015	7.5%
^7X	7.016	92.5%

Continued on next page

For ${}^7\text{X}$: Mass contribution = (mass)(percent abundance)

$$\text{mass contribution} = (7.016 \text{ amu})(0.925) = 6.490 \text{ amu}$$

Sum the mass contributions to find the atomic mass.

$$\text{Atomic mass of X} = (0.451 \text{ amu} + 6.490 \text{ amu}) = 6.941 \text{ amu}$$

Use the periodic table to identify the element.

The element with a mass of 6.941 amu is lithium (Li).

3. Evaluate the Answer

The result of the calculation agrees with the atomic mass given in the periodic table. The masses of the isotopes have four significant figures, so the atomic mass is also expressed with four significant figures.



For more practice with atomic mass problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

15. Boron has two naturally occurring isotopes: boron-10 (abundance = 19.8%, mass = 10.013 amu), boron-11 (abundance = 80.2%, mass = 11.009 amu). Calculate the atomic mass of boron.
16. Helium has two naturally occurring isotopes, helium-3 and helium-4. The atomic mass of helium is 4.003 amu. Which isotope is more abundant in nature? Explain.
17. Calculate the atomic mass of magnesium. The three magnesium isotopes have atomic masses and relative abundances of 23.985 amu (78.99%), 24.986 amu (10.00%), and 25.982 amu (11.01%).

Analyzing an element's mass can give you insight into what the most abundant isotope for the element may be. For example, note that fluorine (F) has an atomic mass that is extremely close to a value of 19 amu. If fluorine had several fairly abundant isotopes, it would be unlikely that its atomic mass would be so close to a whole number. Thus, you might conclude that virtually all naturally occurring fluorine is probably in the form of fluorine-19 (${}^{19}_9\text{F}$). You would be correct, as 100% of naturally occurring fluorine is in the form of fluorine-19. While this type of reasoning generally works well, it is not foolproof. Consider bromine (Br), with an atomic mass of 79.904 amu. With a mass so close to 80 amu, it seems likely that the most common bromine isotope would be bromine-80 (${}^{80}_{35}\text{Br}$). This is not the case, however. Bromine's two isotopes, bromine-79 (78.918 amu, 50.69%) and bromine-81 (80.917 amu, 49.31%), have a weighted average atomic mass of approximately 80 amu, but there is no bromine-80 isotope.

Section 4.3 Assessment

18. Which subatomic particle identifies an atom as that of a particular element? How is this particle related to the atom's atomic number?
19. What is an isotope? Give an example of an element with isotopes.
20. Explain how the existence of isotopes is related to atomic masses not being whole numbers.
21. **Thinking Critically** Nitrogen has two naturally occurring isotopes, N-14 and N-15. The atomic mass of nitrogen is 14.007 amu. Which isotope is more abundant in nature? Explain.
22. **Communicating** List the steps in the process of calculating average atomic mass given data about the isotopes of an element.

Unstable Nuclei and Radioactive Decay

You now have a good understanding of the basic structure of matter and how matter interacts and changes through processes called chemical reactions. With the information you have just learned about the atom's nuclear nature, you are ready to learn about a very different type of reaction—the nuclear reaction. This section introduces you to some of the changes that can take place in a nucleus; you will revisit and further explore this topic in Chapter 25 when you study nuclear chemistry.

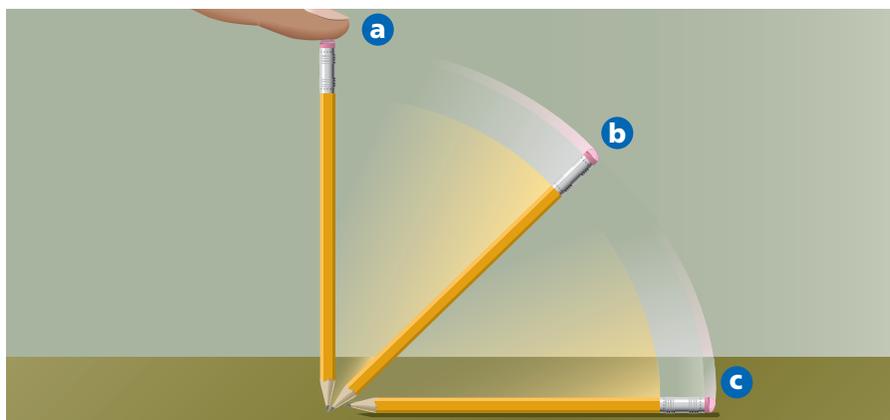
Radioactivity

Recall from Chapter 3 that a chemical reaction involves the change of one or more substances into new substances. Although atoms may be rearranged, their identities do not change during the reaction. You may be wondering why atoms of one element do not change into atoms of another element during a chemical reaction. The reason has to do with the fact that chemical reactions involve only an atom's electrons—the nucleus remains unchanged.

As you learned in the previous section, the number of protons in the nucleus determines the identity of an atom. Thus, because there are no changes in the nuclei during a chemical reaction, the identities of the atoms do not change. There are, however, reactions that do involve an atom of one element changing into an atom of another element. These reactions, which involve a change in an atom's nucleus, are called **nuclear reactions**.

In the late 1890s, scientists noticed that some substances spontaneously emitted radiation in a process they called **radioactivity**. The rays and particles emitted by the radioactive material were called **radiation**. Scientists studying radioactivity soon made an important discovery—radioactive atoms undergo significant changes that can alter their identities. In other words, by emitting radiation, atoms of one element can change into atoms of another element. This discovery was a major breakthrough, as no chemical reaction had ever resulted in the formation of new kinds of atoms.

Radioactive atoms emit radiation because their nuclei are unstable. Unstable systems, whether they're atoms or the pencil standing on its sharpened tip shown in **Figure 4-18a**, gain stability by losing energy. As you can see in **Figure 4-18b** and **Figure 4-18c**, the pencil gains stability (and loses energy) by toppling over. When resting flat on the table top, the pencil has



Objectives

- **Explain** the relationship between unstable nuclei and radioactive decay.
- **Characterize** alpha, beta, and gamma radiation in terms of mass and charge.

Vocabulary

nuclear reaction
radioactivity
radiation
radioactive decay
alpha radiation
alpha particle
nuclear equation
beta radiation
beta particle
gamma ray

Figure 4-18

Unstable systems, such as this pencil momentarily standing on its tip, gain stability by losing energy. In this case, the pencil loses gravitational potential energy as it topples over. Unstable atoms also gain stability by losing energy—they lose energy by emitting radiation.

Careers Using Chemistry

Radiation Protection Technician

Do you like the idea of protecting the health of others? Would you enjoy the challenge of removing contaminated materials? If so, consider a career as a radiation protection technician.

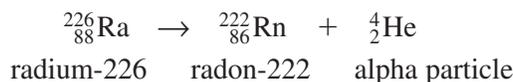
Radiation protection technicians use radiation measuring instruments to locate and assess the risk posed by contaminated materials. Then they decontaminate the area with high-pressure cleaning equipment and remove the radioactive materials. The work often requires the technicians to wear protective suits. The physically demanding work is carefully planned and carried out, with an emphasis on safety.

less gravitational potential energy than it did in its upright position, and thus, it is more stable. Unstable nuclei lose energy by emitting radiation in a spontaneous process (a process that does not require energy) called **radioactive decay**. Unstable radioactive atoms undergo radioactive decay until they form stable nonradioactive atoms, often of a different element. Several types of radiation are commonly emitted during radioactive decay.

Types of Radiation

Scientists began researching radioactivity in the late 1800s. By directing radiation from a radioactive source between two electrically charged plates, scientists were able to identify three different types of radiation. As you can see in **Figure 4-19**, some of the radiation was deflected toward the negatively charged plate, some was deflected toward the positively charged plate, and some was not deflected at all.

Alpha radiation Scientists named the radiation that was deflected toward the negatively charged plate **alpha radiation**. This radiation is made up of **alpha particles**. Each alpha particle contains two protons and two neutrons, and thus has a 2+ charge. As you know, opposite electrical charges attract. So the 2+ charge explains why alpha particles are attracted to the negatively charged plate shown in **Figure 4-19**. An alpha particle is equivalent to a helium-4 nucleus and is represented by ${}^4_2\text{He}$ or α . The alpha decay of radioactive radium-226 into radon-222 is shown below.



Note also that a new element, radon (Rn), is created as a result of the alpha decay of the unstable radium-226 nucleus. The type of equation shown above is known as a **nuclear equation** because it shows the atomic number and mass number of the particles involved. It is important to note that both mass number and atomic number are conserved in nuclear equations. The accounting of atomic numbers and mass numbers below shows that they are conserved.

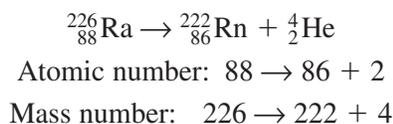
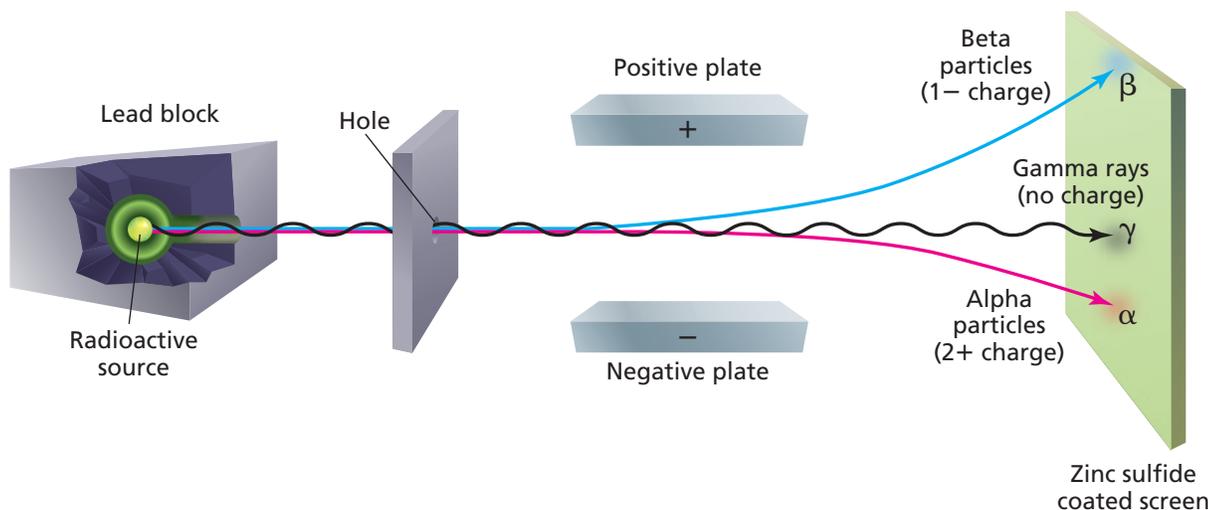
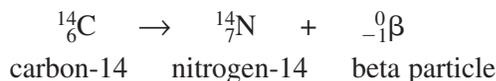


Figure 4-19

Because alpha, beta, and gamma radiation possess different amounts of electrical charge, they are affected differently by an electric field. Gamma rays, which carry no charge, are not deflected by the electric field.

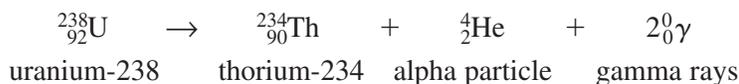


Beta radiation Scientists named the radiation that was deflected toward the positively charged plate **beta radiation**. This radiation consists of fast moving electrons called **beta particles**. Each beta particle is an electron with a 1− charge. The negative charge of the beta particle explains why it is attracted to the positively charged plate shown in **Figure 4-19**. Beta particles are represented by the symbol ${}_{-1}^0\beta$. The beta decay of radioactive carbon-14 into nitrogen-14 is shown below.



The beta decay of unstable carbon-14 results in the creation of the new atom nitrogen (N).

Gamma radiation The third common type of radiation is called gamma radiation, or gamma rays. **Gamma rays** are high-energy radiation that possess no mass and are denoted by the symbol ${}^0_0\gamma$. Because they possess no electrical charge, gamma rays are not deflected by electric or magnetic fields. Gamma rays usually accompany alpha and beta radiation, and they account for most of the energy lost during the radioactive decay process. For example, gamma rays accompany the alpha decay of uranium-238.



Because gamma rays are massless, the emission of gamma rays by themselves cannot result in the formation of a new atom. **Table 4-3** summarizes the basic characteristics of alpha, beta, and gamma radiation.

Nuclear stability You may be wondering why some atoms are stable while others are not. The primary factor in determining an atom's stability is its ratio of neutrons to protons. The details of how the neutron-to-proton ratio determines stability will be covered in Chapter 25. For now, it is enough that you know that atoms containing either too many or too few neutrons are unstable. Unstable nuclei lose energy through radioactive decay in order to form a nucleus with a stable composition of neutrons and protons. Alpha and beta particles are emitted during radioactive decay, and these emissions affect the neutron-to-proton ratio of the newly created nucleus. Eventually, radioactive atoms undergo enough radioactive decay to form stable, nonradioactive atoms. This explains why there are so few radioactive atoms found in nature—most of them have already decayed into stable atoms.

Table 4-3

Characteristics of Alpha, Beta, and Gamma Radiation			
Radiation type	Symbol	Mass (amu)	Charge
Alpha	${}^4_2\text{He}$	4	2+
Beta	${}^0_{-1}\beta$	$\frac{1}{1840}$	1−
Gamma	${}^0_0\gamma$	0	0

Section 4.4 Assessment

23. Explain how unstable atoms gain stability. What determines whether or not an atom is stable?
24. Create a table comparing the mass and charge of alpha, beta, and gamma radiation.
25. In writing a balanced nuclear equation, what must be conserved?
26. **Thinking Critically** Explain how a nuclear reaction differs from a chemical reaction.
27. **Classifying** Classify each of the following as a chemical reaction, a nuclear reaction, or neither.
 - a. Thorium emits a beta particle.
 - b. Two atoms share electrons to form a bond.
 - c. A sample of pure sulfur emits heat energy as it slowly cools.
 - d. A piece of iron rusts.

Very Small Particles

This laboratory investigation will help you conceptualize the size of an atom. You will experiment with a latex balloon containing a vanilla bean extract. Latex is a polymer, meaning that it is a large molecule (a group of atoms that act as a unit) that is made up of a repeating pattern of smaller molecules. The scent of the vanilla extract will allow you to trace the movement of its molecules through the walls of the solid latex balloon.

Problem

How small are the atoms that make up the molecules of the balloon and the vanilla extract? How can you conclude the vanilla molecules are in motion?

Objectives

- **Observe** the movement of vanilla molecules based on detecting their scent.
- **Infer** what the presence of the vanilla scent means in terms of the size and movement of its molecules.
- **Formulate models** that explain how small molecules in motion can pass through an apparent solid.
- **Hypothesize** about the size of atoms that make up matter.

Materials

vanilla extract or flavoring
9-inch latex balloon (2)
dropper

Safety Precautions



- Always wear safety goggles and a lab apron.
- Be careful not to cut yourself when using a sharp object to deflate the balloon.

Pre-Lab

1. Read the entire CHEMLAB.
2. Describe a polymer and give an example.
3. Identify constants in the experiment.
4. What is the purpose of the vanilla extract?
5. As a liquid evaporates, predict what you think will happen to the temperature of the remaining liquid.
6. When you smell an aroma, is your nose detecting a particle in the solid, liquid, or gas phase?
7. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your data and observations.

Data Table			
Observations		Initial	Final
Balloon 1 with vanilla	Relative size		
	Relative temperature		
Balloon 2 without vanilla	Relative size		
	Relative temperature		

Procedure

- Using the medicine dropper, add 25 to 30 drops of vanilla extract to the first balloon.



- Inflate the balloon so its walls are tightly stretched, but not stretched so tightly that the balloon is in danger of bursting. Try to keep the vanilla in one location as the balloon is inflated. Tie the balloon closed.
- Feel the outside of the balloon where the vanilla is located and note the temperature of this area relative to the rest of the balloon. Record your observations in the data table.
- Use only air to inflate a second balloon to approximately the same size as that of the first, and tie it closed. Feel the outside of the second balloon. Make a relative temperature comparison to that of the first balloon. Record your initial observations.
- Place the inflated balloons in a small, enclosed area such as a closet or student locker.
- The next day, repeat the observations in steps 3 and 4 after the vanilla has dried inside the balloon. Record these final observations.
- To avoid splattering your clothes with dark brown vanilla, do not deflate the balloon until the vanilla has dried inside.

Cleanup and Disposal

- After the vanilla has dried, deflate the balloon by puncturing it with a sharp object.
- Dispose of the pieces of the balloon as directed by your teacher.

Analyze and Conclude

- Observing and Inferring** How did the relative volumes of balloons 1 and 2 change after 24 hours? Explain.
- Observing and Inferring** By comparing the relative temperatures of balloons 1 and 2, what can you conclude about the temperature change as the vanilla evaporated? Explain.
- Observing and Inferring** Did the vanilla's odor get outside the balloon and fill the enclosed space? Explain.
- Predicting** Do you think vanilla will leak more rapidly from a fully inflated balloon or from a half-inflated balloon? Explain.
- Hypothesizing** Write a hypothesis that explains your observations.
- Comparing and Contrasting** Compare your hypothesis to Dalton's atomic theory. In what ways is it similar? How is it different?
- Error Analysis** What factors might affect the results of different groups that performed the experiment? What types of errors might have occurred during the procedure?

Real-World Chemistry

- Explain why helium-filled, Mylar-foil balloons can float freely for several weeks, but latex balloons for less than 24 hours.
- How are high-pressure gases stored for laboratory and industrial use to prevent loss?

CHEMISTRY and Society

Nanotechnology

Imagine a technology able to produce roads that repave themselves, greenhouses productive enough to end starvation, and computers the size of cells. Sound like science fiction? It may not be.

Starting at the bottom

This yet-to-be realized technology is generally known as nanotechnology. The prefix *nano-* means one billionth. A nanometer is roughly the size of several atoms put together. The goal of nanotechnology is to manipulate individual atoms in order to create a wide variety of products.

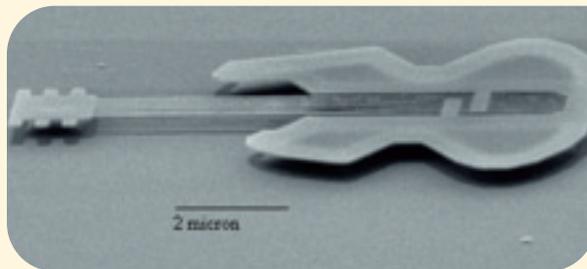
In order to manipulate the immense number of atoms required to make a product, scientists plan on constructing tiny robots called nanorobots. Nanorobots would have two objectives—to manipulate atoms and to copy themselves (self-replication). Through self-replication, countless nanorobots could be created. This work force of nanorobots would then work together to quickly and efficiently assemble new products.

Possibilities and effects

The benefits of nanotechnology could potentially go beyond those of all other existing technologies. The quality and reliability of manufactured products could improve dramatically. For example, a brick could repair itself after cracks form, and a damaged road could repave itself. Furthermore, even as the quality and capability of products increase, their prices would decrease. With the use of nanorobot workforces and a readily available supply of atoms, the cost of atom-assembled products would be low.

Unlike many technological advances of the past, nanotechnology could also benefit the environment. The need for traditional raw materials, such as trees and coal, would be greatly reduced. Also, because the atoms are arranged individually, the amount of waste could be carefully controlled and limited. Thin materials called nanomembranes may even be able to filter existing pollutants out of air and water.

Advances in medicine could also be amazing. Nanodevices smaller than human cells could be used to detect health problems, repair cells, and carry medicines to specific sites in the body.



When will it happen?

The promise of nanotechnology is still years away. Some think nanodevices will be available in the next decade, whereas others expect the technology to take much longer to develop. So far, researchers have accomplished rearranging atoms into specific shapes such as letters and symbols, and have also succeeded in developing several very simple nanodevices. One such device, a nanoguitar, made of crystalline silicon, is about 1/20th the width of a single human hair.

Investigating the Issue

- 1. Communicating Ideas** Read about the Industrial Revolution and write a brief essay describing how technological advances affected society.
- 2. Debating the Issue** Nanotechnology supporters argue that dangers posed by the technology will be addressed as nanotechnology is developed. Should researchers be prevented from developing nanotechnology?



Visit the Chemistry Web site at science.glencoe.com to find links to more information about nanotechnology and the issues that surround it.

Summary

4.1 Early Theories of Matter

- The Greek philosopher Democritus was the first person to propose the existence of atoms.
- In 1808, Dalton proposed his atomic theory, which was based on numerous scientific experiments.
- All matter is composed of atoms. An atom is the smallest particle of an element that maintains the properties of that element. Atoms of one element are different from atoms of other elements.

4.2 Subatomic Particles and the Nuclear Atom

- Atoms are composed of negatively charged electrons, neutral neutrons, and positively charged protons. Electrons have a $1-$ charge, protons have a $1+$ charge, and neutrons have no charge. Both protons and electrons have masses approximately 1840 times that of an electron.
- The nucleus of an atom contains all of its positive charge and nearly all of its mass.
- The nucleus occupies an extremely small volume of space at the center of an atom. Most of an atom consists of empty space surrounding the nucleus through which the electrons move.

4.3 How Atoms Differ

- The number of protons in an atom uniquely identifies an atom. This number of protons is the atomic number of the atom.

- Atoms have equal numbers of protons and electrons, and thus, no overall electrical charge.
- An atom's mass number is equal to its total number of protons and neutrons.
- Atoms of the same element with different numbers of neutrons and different masses are called isotopes.
- The atomic mass of an element is a weighted average of the masses of all the naturally occurring isotopes of that element.

4.4 Unstable Nuclei and Radioactive Decay

- Chemical reactions involve changes in the electrons surrounding an atom. Nuclear reactions involve changes in the nucleus of an atom.
- The neutron-to-proton ratio of an atom's nucleus determines its stability. Unstable nuclei undergo radioactive decay, emitting radiation in the process.
- Alpha particles are equivalent to the nuclei of helium atoms, and are represented by ${}^4_2\text{He}$ or α . Alpha particles have a charge of $2+$.
- Beta particles are high-speed electrons and are represented by ${}^0_{-1}\beta$. Beta particles have a $1-$ charge.
- Gamma rays are high-energy radiation and are represented by the symbol ${}^0_0\gamma$. Gamma rays have no electrical charge and no mass.

Key Equations and Relationships

- Determining the number of protons and electrons

$$\text{Atomic number} = \text{number of protons} = \text{number of electrons} \quad (\text{p. 98})$$

- Determining the number of neutrons

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number} \quad (\text{p. 101})$$

Vocabulary

- alpha particle (p. 106)
- alpha radiation (p. 106)
- atom (p. 90)
- atomic mass (p. 102)
- atomic mass unit (amu) (p. 102)
- atomic number (p. 98)
- beta particle (p. 107)
- beta radiation (p. 107)
- cathode ray (p. 92)
- Dalton's atomic theory (p. 89)
- electron (p. 93)
- gamma ray (p. 107)
- isotope (p. 100)
- mass number (p. 100)
- neutron (p. 96)
- nuclear equation (p. 106)
- nuclear reaction (p. 105)
- nucleus (p. 95)
- proton (p. 96)
- radiation (p. 105)
- radioactive decay (p. 106)
- radioactivity (p. 105)

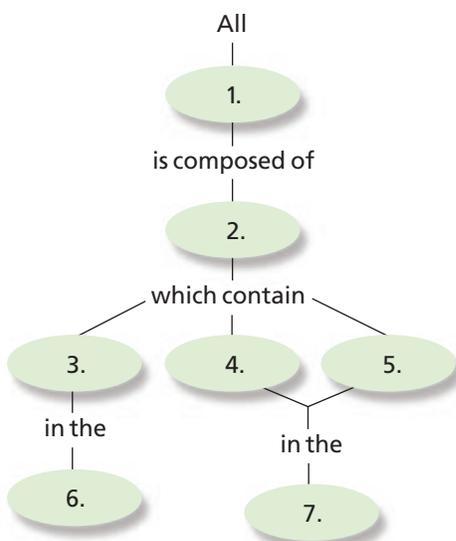


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 4 Assessment.

Concept Mapping

- 28.** Complete the concept map using the following terms: electrons, matter, neutrons, nucleus, empty space around nucleus, protons, and atoms.



Mastering Concepts

- 29.** Who originally proposed the concept that matter was composed of tiny indivisible particles? (4.1)
- 30.** Whose work is credited with being the beginning of modern atomic theory? (4.1)
- 31.** Explain why Democritus was unable to experimentally verify his ideas. (4.1)
- 32.** State the main points of Dalton's atomic theory using your own words. Which parts of Dalton's theory were later found to be in error? Explain why. (4.1)
- 33.** Explain how Dalton's atomic theory offered a convincing explanation of the observation that mass is conserved in chemical reactions. (4.1)
- 34.** Which subatomic particle was discovered by researchers working with cathode ray tubes? (4.2)
- 35.** What experimental results led to the conclusion that electrons were part of all forms of matter? (4.2)
- 36.** What is the charge and mass of a single electron? (4.2)
- 37.** List the strengths and weaknesses of Rutherford's nuclear model of the atom. (4.2)
- 38.** What particles are found in the nucleus of an atom? What is the net charge of the nucleus? (4.2)
- 39.** Explain what keeps the electrons confined in the space surrounding the nucleus. (4.2)
- 40.** Describe the flow of a cathode ray inside a cathode ray tube. (4.2)
- 41.** Which outdated atomic model could be likened to chocolate chip cookie dough? (4.2)
- 42.** What caused the deflection of the alpha particles in Rutherford's gold foil experiment? (4.2)
- 43.** Which subatomic particles account for most all of an atom's mass? (4.2)
- 44.** How is an atom's atomic number related to its number of protons? To its number of electrons? (4.2)
- 45.** What is the charge of the nucleus of element 89? (4.2)
- 46.** Explain why atoms are electrically neutral. (4.2)
- 47.** Does the existence of isotopes contradict part of Dalton's original atomic theory? Explain. (4.3)
- 48.** How do isotopes of a given element differ? How are they similar? (4.3)
- 49.** How is the mass number related to the number of protons and neutrons an atom has? (4.3)
- 50.** What do the superscript and subscript in the notation ${}_{19}^{40}\text{K}$ represent? (4.3)
- 51.** Explain how to determine the number of neutrons an atom contains if you know its mass number and its atomic number. (4.3)
- 52.** Define the atomic mass unit. What were the benefits of developing the atomic mass unit as a standard unit of mass? (4.3)
- 53.** What type of reaction involves changes in the nucleus of an atom? (4.4)
- 54.** Explain how energy loss and nuclear stability are related to radioactive decay. (4.4)
- 55.** Explain what must occur before a radioactive atom ceases to undergo further radioactive decay. (4.4)
- 56.** Write the symbols used to denote alpha, beta, and gamma radiation and give their mass and charge. (4.4)
- 57.** What change in mass number occurs when a radioactive atom emits an alpha particle? A beta particle? A gamma particle? (4.4)
- 58.** What is the primary factor determining whether or not an atom is stable or unstable? (4.4)

Mastering Problems

Atomic Number and Mass Number (4.3)

59. How many protons and electrons are contained in an atom of element 44?
60. For each of the following chemical symbols, determine the element name and the number of protons and electrons an atom contains.
- a. V c. Ir
b. Mn d. S
61. A carbon atom has a mass number of 12 and an atomic number of 6. How many neutrons does it have?
62. An isotope of mercury has 80 protons and 120 neutrons. What is the mass number of this isotope?
63. An isotope of xenon has an atomic number of 54 and contains 77 neutrons. What is the xenon isotope's mass number?
64. How many electrons, protons, and neutrons are contained in each of the following atoms?
- a. $^{132}_{55}\text{Cs}$ c. $^{163}_{69}\text{Tm}$
b. $^{59}_{27}\text{Co}$ d. $^{70}_{30}\text{Zn}$
65. How many electrons, protons, and neutrons are contained in each of the following atoms?
- a. gallium-64 c. titanium-48
b. fluorine-23 d. helium-8

Atomic Mass (4.3)

66. Chlorine, which has an atomic mass of 35.453 amu, has two naturally occurring isotopes, Cl-35 and Cl-37. Which isotope occurs in greater abundance? Explain.
67. Silver has two isotopes, $^{107}_{47}\text{Ag}$ has a mass of 106.905 amu (52.00%), and $^{109}_{47}\text{Ag}$ has a mass of 108.905 amu (48.00%). What is the atomic mass of silver?
68. Data for chromium's four naturally occurring isotopes is provided in the table below. Calculate chromium's atomic mass.

Chromium Isotope Data		
Isotope	Percent abundance	Mass (amu)
Cr-50	4.35%	49.946
Cr-52	83.79%	51.941
Cr-53	9.50%	52.941
Cr-54	2.36%	53.939

Mixed Review

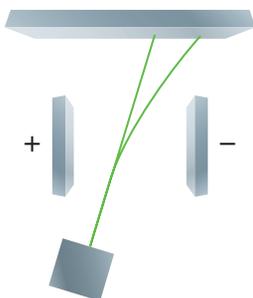
Sharpen your problem-solving skills by answering the following.

69. Describe a cathode ray tube and how it operates.
70. Explain how J. J. Thomson's determination of the charge-to-mass ratio of the electron led to the conclusion that atoms were composed of subatomic particles.
71. How did the actual results of Rutherford's gold foil experiment differ from the results he expected?
72. Complete the table below.

Composition of Various Isotopes					
Isotope	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
		32	16		
				24	20
Zn-64					
	9			10	
	11	23			

73. Approximately how many times greater is the diameter of an atom than the diameter of its nucleus? Knowing that most of an atom's mass is contained in the nucleus, what can you conclude about the density of the nucleus?
74. Is the charge of a nucleus positive, negative, or zero? The charge of an atom?
75. Why are electrons in a cathode ray tube deflected by magnetic and electric fields?
76. What was Henry Moseley's contribution to our understanding of the atom?
77. What is the mass number of potassium-39? What is the isotope's charge?
78. Boron-10 and boron-11 are the naturally occurring isotopes of elemental boron. If boron has an atomic mass of 10.81 amu, which isotope occurs in greater abundance?
79. Calculate the atomic mass of titanium. The five titanium isotopes have atomic masses and relative abundances of 45.953 amu (8.00%), 46.952 amu (7.30%), 47.948 amu (73.80%), 48.948 amu (5.50%), and 49.945 amu (5.40%).

- 80.** Identify the two types of radiation shown in the figure below. Explain your reasoning.



- 81.** Describe how each type of radiation affects an atom's atomic number and mass number.
- 82.** Silicon is very important to the semiconductor manufacturing industry. The three naturally occurring isotopes of silicon are silicon-28, silicon-29, and silicon-30. Write the symbol for each.

Thinking Critically

- 83. Applying Concepts** Which is greater, the number of compounds or the number of elements? The number of atoms or the number of isotopes? Explain.
- 84. Analyzing Information** An element has three naturally occurring isotopes. What other information must you know in order to calculate the element's atomic mass?
- 85. Applying Concepts** If atoms are primarily composed of empty space, why can't you pass your hand through a solid object?
- 86. Formulating Models** Sketch a modern atomic model of a typical atom and identify where each type of subatomic particle would be located.
- 87. Applying Concepts** Copper has two naturally occurring isotopes and an atomic mass of 63.546 amu. Cu-63 has a mass of 62.940 amu and an abundance of 69.17%. What is the identity and percent abundance of copper's other isotope?

Writing in Chemistry

- 88.** The Standard Model of particle physics describes all of the known building blocks of matter. Research the particles included in the Standard Model. Write a short report describing the known particles and those thought to exist but not detected experimentally.

- 89.** Individual atoms can be seen using a sophisticated device known as a scanning tunneling microscope. Write a short report on how the scanning tunneling microscope works and create a gallery of scanning tunneling microscope images from sources such as books, magazines, and the Internet.



Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 90.** How is a qualitative observation different from a quantitative observation? Give an example of each. (Chapter 1)
- 91.** A 1.0-cm^3 block of gold can be flattened to a thin sheet that averages 3.0×10^{-8} cm thick. What is the area (in cm^2) of the flattened gold sheet? A letter size piece of paper has an area of 603 cm^2 . How many sheets of paper would the gold cover? (Chapter 2)
- 92.** Classify the following mixtures as heterogeneous or homogeneous. (Chapter 3)
- salt water
 - vegetable soup
 - 14-K gold
 - concrete
- 93.** Are the following changes physical or chemical? (Chapter 3)
- water boils
 - a match burns
 - sugar dissolves in water
 - sodium reacts with water
 - ice cream melts

Use these questions and the test-taking tip to prepare for your standardized test.

- An atom of plutonium _____.
 - can be divided into smaller particles that retain all the properties of plutonium
 - cannot be divided into smaller particles that retain all the properties of plutonium
 - does not possess all the properties of a larger quantity of plutonium
 - cannot be seen using current technology
- Neptunium's only naturally occurring isotope, ${}^{237}_{93}\text{Np}$, decays by emitting one alpha particle, one beta particle, and one gamma ray. What is the new atom formed from this decay?
 - ${}^{233}_{92}\text{U}$
 - ${}^{241}_{93}\text{Np}$
 - ${}^{233}_{90}\text{Th}$
 - ${}^{241}_{92}\text{U}$
- An atom has no net electrical charge because _____.
 - its subatomic particles carry no electrical charges
 - the positively charged protons cancel out the negatively charged neutrons
 - the positively charged neutrons cancel out the negatively charged electrons
 - the positively charged protons cancel out the negatively charged electrons
- ${}^{126}_{52}\text{Te}$ has _____.
 - 126 neutrons, 52 protons, and 52 electrons
 - 74 neutrons, 52 protons, and 52 electrons
 - 52 neutrons, 74 protons, and 74 electrons
 - 52 neutrons, 126 protons, and 126 electrons
- Assume the following three isotopes of element Q exist: ${}^{248}Q$, ${}^{252}Q$, and ${}^{259}Q$. If the atomic mass of Q is 258.63, which of its isotopes is the most abundant?
 - ${}^{248}Q$
 - ${}^{252}Q$
 - ${}^{259}Q$
 - they are all equally abundant
- Based on the table, an atom of neon found in nature would most likely have a mass of _____.
 - 19.992 amu
 - 20.179 amu
 - 20.994 amu
 - 21.991 amu
- In which of the neon isotopes is the number of neutrons the same as the number of protons?
 - ${}^{20}\text{Ne}$
 - ${}^{21}\text{Ne}$
 - ${}^{22}\text{Ne}$
 - none of the above
- The atomic mass of Ne is equal to _____.
 - $\frac{19.922 \text{ amu} + 20.994 \text{ amu} + 21.991 \text{ amu}}{3}$
 - $\frac{1}{3}[(19.992 \text{ amu})(90.48\%) + (20.994 \text{ amu})(0.27\%) + (21.991 \text{ amu})(9.25\%)]$
 - $(19.992 \text{ amu})(90.48\%) + (20.994 \text{ amu})(0.27\%) + (21.991 \text{ amu})(9.25\%)$
 - $19.992 \text{ amu} + 20.994 \text{ amu} + 21.991 \text{ amu}$
- Element X has an unstable nucleus due to an overabundance of neutrons. All of the following are likely to occur EXCEPT _____.
 - element X will undergo radioactive decay
 - element X will eventually become a stable, nonradioactive element
 - element X will gain more protons to balance the neutrons it possesses
 - element X will spontaneously lose energy
- The volume of an atom is made up mostly of _____.
 - protons
 - neutrons
 - electrons
 - empty space

Interpreting Tables Use the table to answer questions 6–8.

Characteristics of Naturally Occurring Neon Isotopes			
Isotope	Atomic number	Mass (amu)	Percent abundance
${}^{20}\text{Ne}$	10	19.992	90.48
${}^{21}\text{Ne}$	10	20.994	0.27
${}^{22}\text{Ne}$	10	21.991	9.25



TEST-TAKING TIP

Skip Around if You Can The questions on some tests start easy and get progressively harder, while other tests mix easy and hard questions. You may want to skip over difficult questions and come back to them later, after you've answered all the easier questions. This will guarantee more points toward your final score. In fact, other questions may help you answer the ones you skipped. Just be sure you fill in the correct ovals on your answer sheet.

Electrons in Atoms

What You'll Learn

- ▶ You will compare the wave and particle models of light.
- ▶ You will describe how the frequency of light emitted by an atom is a unique characteristic of that atom.
- ▶ You will compare and contrast the Bohr and quantum mechanical models of the atom.
- ▶ You will express the arrangements of electrons in atoms through orbital notations, electron configurations, and electron dot structures.

Why It's Important

Why are some fireworks red, some white, and others blue? The key to understanding the chemical behavior of fireworks, and all matter, lies in understanding how electrons are arranged in atoms of each element.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about electrons and atomic structure.

The colorful display from fireworks is due to changes in the electron configurations of atoms.



DISCOVERY LAB



Materials

a wrapped box from your instructor

What's Inside?

It's your birthday, and there are many wrapped presents for you to open. Much of the fun is trying to figure out what's inside the package before you open it. In trying to determine the structure of the atom, chemists had a similar experience. How good are your skills of observation and deduction?

Procedure

1. Obtain a wrapped box from your instructor.
2. Using as many observation methods as you can, and without unwrapping or opening the box, try to figure out what the object inside the box is.
3. Record the observations you make throughout this discovery process.

Analysis

How were you able to determine things such as size, shape, number, and composition of the object in the box? What senses did you use to make your observations? Why is it hard to figure out what type of object is in the box without actually seeing it?

Section

5.1

Light and Quantized Energy

Objectives

- **Compare** the wave and particle models of light.
- **Define** a quantum of energy and explain how it is related to an energy change of matter.
- **Contrast** continuous electromagnetic spectra and atomic emission spectra.

Vocabulary

electromagnetic radiation
wavelength
frequency
amplitude
electromagnetic spectrum
quantum
Planck's constant
photoelectric effect
photon
atomic emission spectrum

Although three subatomic particles had been discovered by the early-1900s, the quest to understand the atom and its structure had really just begun. That quest continues in this chapter, as scientists pursued an understanding of how electrons were arranged within atoms. Perform the **DISCOVERY LAB** on this page to better understand the difficulties scientists faced in researching the unseen atom.

The Nuclear Atom and Unanswered Questions

As you learned in Chapter 4, Rutherford proposed that all of an atom's positive charge and virtually all of its mass are concentrated in a nucleus that is surrounded by fast-moving electrons. Although his nuclear model was a major scientific development, it lacked detail about how electrons occupy the space surrounding the nucleus. In this chapter, you will learn how electrons are arranged in an atom and how that arrangement plays a role in chemical behavior.

Many scientists in the early twentieth century found Rutherford's nuclear atomic model to be fundamentally incomplete. To physicists, the model did not explain how the atom's electrons are arranged in the space around the nucleus. Nor did it address the question of why the negatively charged electrons are not pulled into the atom's positively charged nucleus. Chemists found Rutherford's nuclear model lacking because it did not begin to account for the differences in chemical behavior among the various elements.



Figure 5-1

a Chlorine gas, shown here reacting vigorously with steel wool, reacts with many other atoms as well. **b** Argon gas fills the interior of this incandescent bulb. The nonreactive argon prevents the hot filament from oxidizing, thus extending the life of the bulb. **c** Solid potassium metal is submerged in oil to prevent it from reacting with air or water.

For example, consider the elements chlorine, argon, and potassium, which are found in consecutive order on the periodic table but have very different chemical behaviors. Atoms of chlorine, a yellow-green gas at room temperature, react readily with atoms of many other elements. **Figure 5-1a** shows chlorine atoms reacting with steel wool. The interaction of highly reactive chlorine atoms with the large surface area provided by the steel results in a vigorous reaction. Argon, which is used in the incandescent bulb shown in **Figure 5-1b**, also is a gas. Argon, however, is so unreactive that it is considered a noble gas. Potassium is a reactive metal at room temperature. In fact, as you can see in **Figure 5-1c**, because potassium is so reactive, it must be stored under kerosene or oil to prevent its atoms from reacting with the oxygen and water in the air. Rutherford's nuclear atomic model could not explain why atoms of these elements behave the way they do.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They had observed that certain elements emitted visible light when heated in a flame. Analysis of the emitted light revealed that an element's chemical behavior is related to the arrangement of the electrons in its atoms. In order for you to better understand this relationship and the nature of atomic structure, it will be helpful for you to first understand the nature of light.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Wave Nature of Light

Electromagnetic radiation is a form of energy that exhibits wavelike behavior as it travels through space. Visible light is a type of electromagnetic radiation. Other examples of electromagnetic radiation include visible light from the sun, microwaves that warm and cook your food, X rays that doctors and dentists use to examine bones and teeth, and waves that carry radio and television programs to your home.

All waves can be described by several characteristics, a few of which you may be familiar with from everyday experience. **Figure 5-2a** shows a standing wave created by rhythmically moving the free end of a spring toy. **Figure 5-2b** illustrates several primary characteristics of all waves, wavelength, frequency, amplitude, and speed. **Wavelength** (represented by λ , the Greek letter lambda) is the shortest distance between equivalent points on a continuous wave. For example, in **Figure 5-2b** the wavelength is measured from crest to crest or from trough to trough. Wavelength is usually expressed in meters, centimeters, or nanometers ($1 \text{ nm} = 1 \times 10^{-9} \text{ m}$). **Frequency** (represented by ν , the Greek letter nu) is the number of waves that pass a given

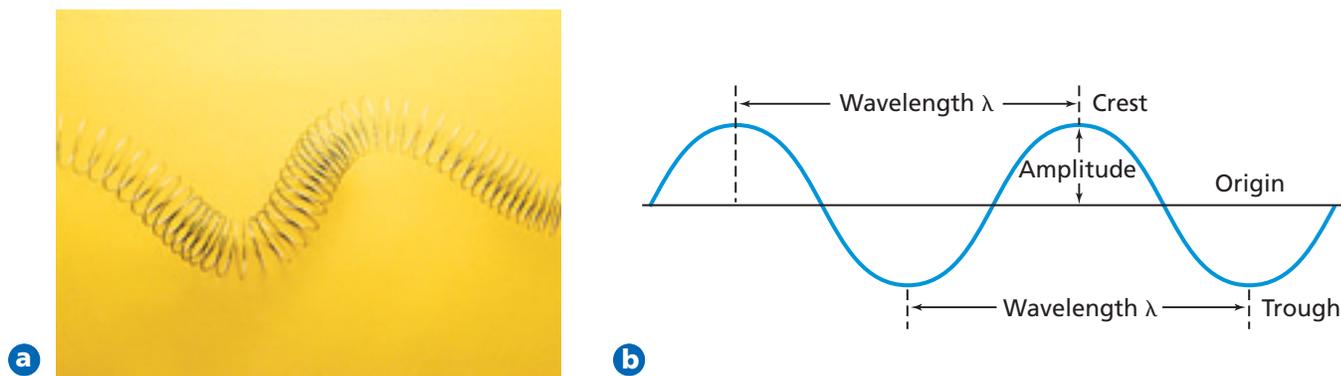


Figure 5-2

a The standing wave produced with this spring toy displays properties that are characteristic of all waves. **b** The primary characteristics of waves are wavelength, frequency, amplitude, and speed. What is the wavelength of the wave in centimeters?

point per second. One hertz (Hz), the SI unit of frequency, equals one wave per second. In calculations, frequency is expressed with units of “waves per second,” $\left(\frac{1}{s}\right)$ or (s^{-1}) , where the term “waves” is understood. For example,

$$652 \text{ Hz} = 652 \text{ waves/second} = \frac{652}{s} = 652 \text{ s}^{-1}$$

The **amplitude** of a wave is the wave’s height from the origin to a crest, or from the origin to a trough. To learn how lightwaves are able to form powerful laser beams, read the **How It Works** at the end of this chapter.

All electromagnetic waves, including visible light, travel at a speed of $3.00 \times 10^8 \text{ m/s}$ in a vacuum. Because the speed of light is such an important and universal value, it is given its own symbol, c . The speed of light is the product of its wavelength (λ) and its frequency (ν).

$$c = \lambda \nu$$

Although the speed of all electromagnetic waves is the same, waves may have different wavelengths and frequencies. As you can see from the equation above, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the red and violet light waves illustrated in **Figure 5-3**. Although both waves travel at the speed of light, you can see that red light has a longer wavelength and lower frequency than violet light.

Sunlight, which is one example of what is called white light, contains a continuous range of wavelengths and frequencies. Sunlight passing through a prism

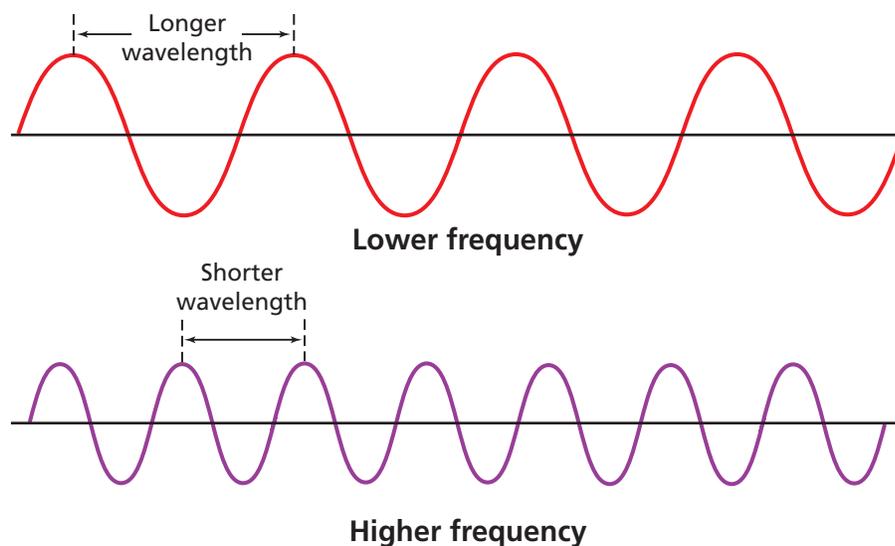
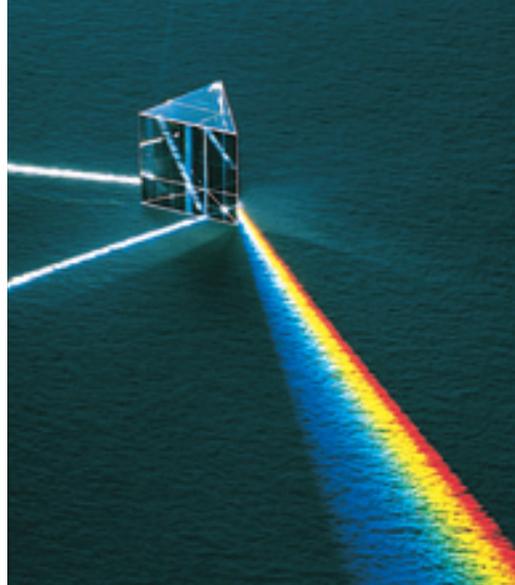


Figure 5-3

The inverse relationship between wavelength and frequency of electromagnetic waves can be seen in these red and violet waves. As wavelength increases, frequency decreases. Wavelength and frequency do not affect the amplitude of a wave. Which wave has the larger amplitude?

Figure 5-4

White light is separated into a continuous spectrum when it passes through a prism.



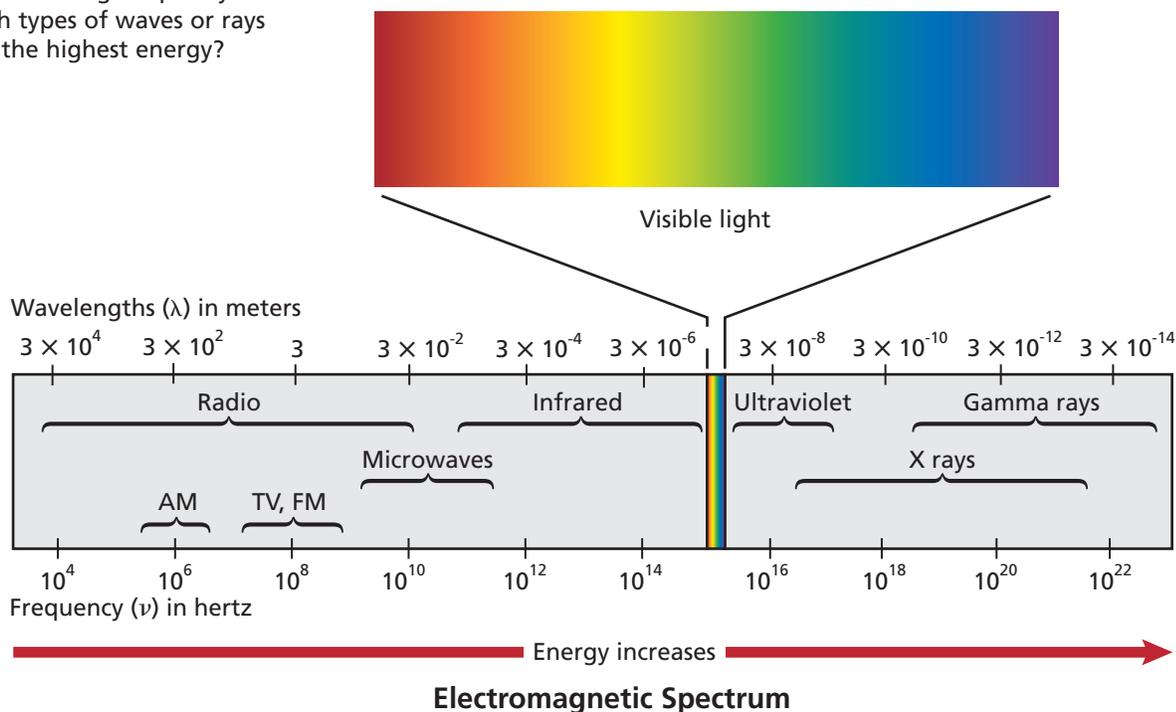
is separated into a continuous spectrum of colors. These are the colors of the visible spectrum. The spectrum is called continuous because there is no portion of it that does not correspond to a unique wavelength and frequency of light. You are already familiar with all of the colors of the visible spectrum from your everyday experiences. And if you have ever seen a rainbow, you have seen all of the visible colors at once. A rainbow is formed when tiny drops of water in the

air disperse the white light from the sun into its component colors, producing a continuous spectrum that arches across the sky.

The visible spectrum of light shown in **Figure 5-4**, however, comprises only a small portion of the complete electromagnetic spectrum, which is illustrated in **Figure 5-5**. The **electromagnetic spectrum**, also called the EM spectrum, encompasses all forms of electromagnetic radiation, with the only differences in the types of radiation being their frequencies and wavelengths. Note in **Figure 5-4** that the short wavelengths bend more than long wavelengths as they pass through the prism, resulting in the sequence of colors red, orange, yellow, green, blue, indigo, and violet. This sequence can be remembered using the fictitious name Roy G. Biv as a memory aid. In examining the energy of the radiation shown in **Figure 5-5**, you should note that energy increases with increasing frequency. Thus, looking back at **Figure 5-3**, the violet light, with its greater frequency, has more energy than the red light. This relationship between frequency and energy will be explained in the next section.

Figure 5-5

The electromagnetic spectrum includes a wide range of wavelengths (and frequencies). Energy of the radiation increases with increasing frequency. Which types of waves or rays have the highest energy?



Because all electromagnetic waves travel at the same speed, you can use the formula $c = \lambda\nu$ to calculate the wavelength or frequency of any wave. Example Problem 5-1 shows how this is done.

EXAMPLE PROBLEM 5-1

Calculating Wavelength of an EM Wave

Microwaves are used to transmit information. What is the wavelength of a microwave having a frequency of 3.44×10^9 Hz?

1. Analyze the Problem

You are given the frequency of a microwave. You also know that because microwaves are part of the electromagnetic spectrum, their speed, frequency, and wavelength are related by the formula $c = \lambda\nu$. The value of c is a known constant. First, solve the equation for wavelength, then substitute the known values and solve.

Known

$$\nu = 3.44 \times 10^9 \text{ Hz}$$

$$c = 3.00 \times 10^8 \text{ m/s}$$

Unknown

$$\lambda = ? \text{ m}$$

2. Solve for the Unknown

Solve the equation relating the speed, frequency, and wavelength of an electromagnetic wave for wavelength (λ).

$$c = \lambda\nu$$

$$\lambda = c/\nu$$

Substitute c and the microwave's frequency, ν , into the equation. Note that hertz is equivalent to $1/\text{s}$ or s^{-1} .

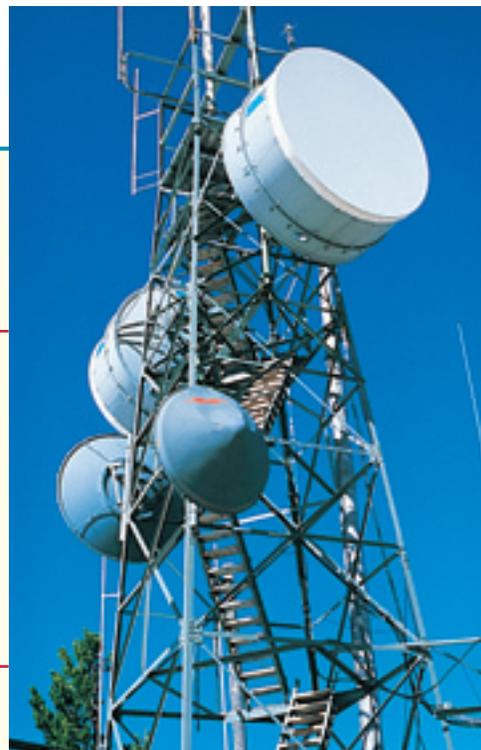
$$\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ s}^{-1}}$$

Divide the values to determine wavelength, λ , and cancel units as required.

$$\lambda = \frac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ s}^{-1}} = 8.72 \times 10^{-2} \text{ m}$$

3. Evaluate the Answer

The answer is correctly expressed in a unit of wavelength (m). Both of the known values in the problem are expressed with three significant figures, so the answer should have three significant figures, which it does. The value for the wavelength is within the wavelength range for microwaves shown in Figure 5-5.



Microwave relay antennas are used to transmit voice and data from one area to another without the use of wires or cables.

PRACTICE PROBLEMS

1. What is the frequency of green light, which has a wavelength of 4.90×10^{-7} m?
2. An X ray has a wavelength of 1.15×10^{-10} m. What is its frequency?
3. What is the speed of an electromagnetic wave that has a frequency of 7.8×10^6 Hz?
4. A popular radio station broadcasts with a frequency of 94.7 MHz. What is the wavelength of the broadcast? (1 MHz = 10^6 Hz)



For more practice with speed, frequency, and wavelength problems, go to **Supplemental Practice Problems** in Appendix A.

Particle Nature of Light

While considering light as a wave does explain much of its everyday behavior, it fails to adequately describe important aspects of light's interactions with matter. The wave model of light cannot explain why heated objects emit only certain frequencies of light at a given temperature, or why some metals emit electrons when colored light of a specific frequency shines on them. Obviously, a totally new model or a revision of the current model of light was needed to address these phenomena.

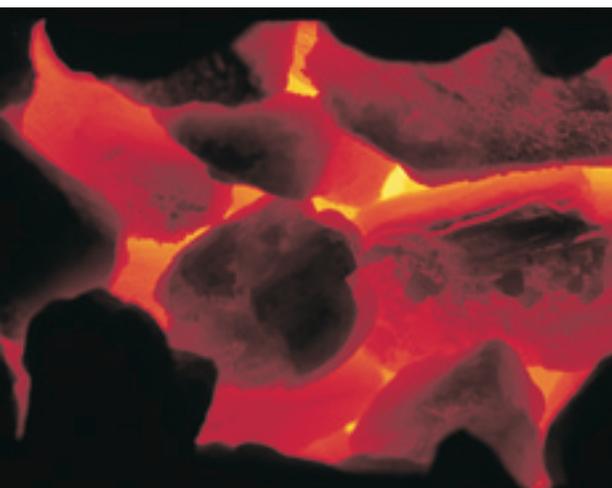
The quantum concept The glowing light emitted by the hot objects shown in **Figure 5-6** are examples of a phenomenon you have certainly seen. Iron provides another example of the phenomenon. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and appears bluish in color at even higher temperatures. As you will learn in greater detail later on in this course, the temperature of an object is a measure of the average kinetic energy of its particles. As the iron gets hotter it possesses a greater amount of energy, and emits different colors of light. These different colors correspond to different frequencies and wavelengths. The wave model could not explain the emission of these different wavelengths of light at different temperatures. In 1900, the German physicist Max Planck (1858–1947) began searching for an explanation as he studied the light emitted from heated objects. His study of the phenomenon led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. That is, a **quantum** is the minimum amount of energy that can be gained or lost by an atom.

Planck and other physicists of the time thought the concept of quantized energy was revolutionary—and some found it disturbing. Prior experience had led scientists to believe that energy could be absorbed and emitted in continually varying quantities, with no minimum limit to the amount. For example, think about heating a cup of water in a microwave oven. It seems that you can add any amount of thermal energy to the water by regulating the power and duration of the microwaves. Actually, the water's temperature increases in infinitesimal steps as its molecules absorb quanta of energy. Because these steps are so small, the temperature seems to rise in a continuous, rather than a stepwise, manner.

The glowing objects shown in **Figure 5-6** are emitting light, which is a form of energy. Planck proposed that this emitted light energy was quantized.

Figure 5-6

These photos illustrate the phenomenon of heated objects emitting different frequencies of light. Matter, regardless of its form, can gain or lose energy only in small “quantized” amounts.



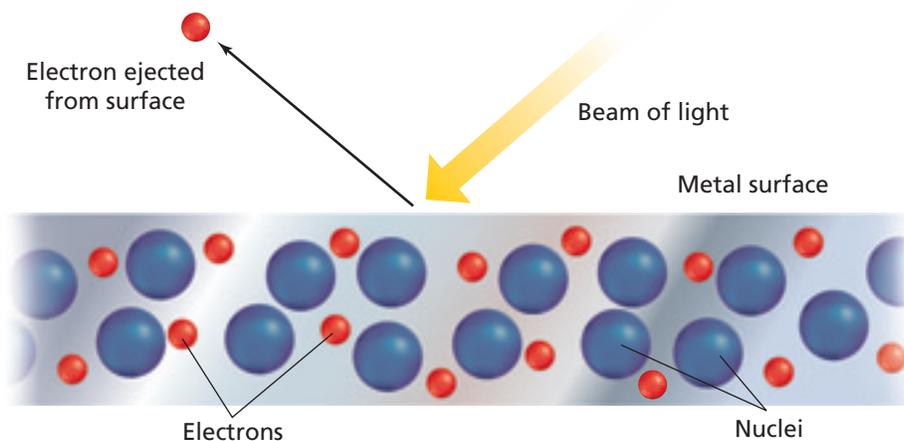


Figure 5-7

In the photoelectric effect, light of a certain minimum frequency (energy) ejects electrons from the surface of a metal. Increasing the intensity of the incident light results in more electrons being ejected. Increasing the frequency (energy) of the incident light causes the ejected electrons to travel faster.

He then went further and demonstrated mathematically that the energy of a quantum is related to the frequency of the emitted radiation by the equation

$$E_{\text{quantum}} = h\nu$$

where E is energy, h is Planck's constant, and ν is frequency. **Planck's constant** has a value of $6.626 \times 10^{-34} \text{ J} \cdot \text{s}$, where J is the symbol for the joule, the SI unit of energy. Looking at the equation, you can see that the energy of radiation increases as the radiation's frequency, ν , increases. This equation explains why the violet light in **Figure 5-3** has greater energy than the red light.

According to Planck's theory, for a given frequency, ν , matter can emit or absorb energy only in whole-number multiples of $h\nu$; that is, $1h\nu$, $2h\nu$, $3h\nu$, and so on. A useful analogy for this concept is that of a child building a wall of wooden blocks. The child can add to or take away height from the wall only in increments of a whole number of blocks. Partial blocks are not possible. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

The photoelectric effect Scientists knew that the wave model (still very popular in spite of Planck's proposal) could not explain a phenomenon called the photoelectric effect. In the **photoelectric effect**, electrons, called photoelectrons, are emitted from a metal's surface when light of a certain frequency shines on the surface, as shown in **Figure 5-7**. Perhaps you've taken advantage of the photoelectric effect by using a calculator, such as the one shown in **Figure 5-8**, that is powered by photoelectric cells. Photoelectric cells in these and many other devices convert the energy of incident light into electrical energy.

The mystery of the photoelectric effect concerns the frequency, and therefore color, of the incident light. The wave model predicts that given enough time, even low-energy, low-frequency light would accumulate and supply enough energy to eject photoelectrons from a metal. However, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intense or how long it shines, light with a frequency less than $1.14 \times 10^{15} \text{ Hz}$ does not eject photoelectrons from silver. But even dim light having a frequency equal to or greater than $1.14 \times 10^{15} \text{ Hz}$ causes the ejection of photoelectrons from silver.

In explaining the photoelectric effect, Albert Einstein proposed in 1905 that electromagnetic radiation has both wavelike and particle-like natures. That is, while a beam of light has many wavelike characteristics, it also can be thought of as a stream of tiny particles, or bundles of energy, called photons. Thus, a **photon** is a particle of electromagnetic radiation with no mass that carries a quantum of energy.

Figure 5-8

The direct conversion of sunlight into electrical energy is a viable power source for low-power consumption devices such as this calculator. The cost of photoelectric cells makes them impractical for large-scale power production.



Extending Planck's idea of quantized energy, Einstein calculated that a photon's energy depends on its frequency.

$$E_{\text{photon}} = h\nu$$

Further, Einstein proposed that the energy of a photon of light must have a certain minimum, or threshold, value to cause the ejection of a photoelectron. That is, for the photoelectric effect to occur a photon must possess, at a minimum, the energy required to free an electron from an atom of the metal. According to this theory, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Although Einstein was able to explain the photoelectric effect by giving electromagnetic radiation particlelike properties, it's important to note that a dual wave-particle model of light was required.

EXAMPLE PROBLEM 5-2



Sunlight bathes Earth in white light—light composed of all of the visible colors of the electromagnetic spectrum.

Calculating the Energy of a Photon

Tiny water drops in the air disperse the white light of the sun into a rainbow. What is the energy of a photon from the violet portion of the rainbow if it has a frequency of $7.23 \times 10^{14} \text{ s}^{-1}$?

1. Analyze the Problem

You are given the frequency of a photon of violet light. You also know that the energy of a photon is related to its frequency by the equation $E_{\text{photon}} = h\nu$. The value of h , Planck's constant, is known. By substituting the known values, the equation can be solved for the energy of a photon of violet light.

Known

$$\begin{aligned}\nu &= 7.23 \times 10^{14} \text{ s}^{-1} \\ h &= 6.626 \times 10^{-34} \text{ J} \cdot \text{s}\end{aligned}$$

Unknown

$$E_{\text{photon}} = ? \text{ J}$$

2. Solve for the Unknown

Substitute the known values for frequency and Planck's constant into the equation relating energy of a photon and frequency.

$$E_{\text{photon}} = h\nu$$

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(7.23 \times 10^{14} \text{ s}^{-1})$$

Multiply the known values and cancel units.

$$E_{\text{photon}} = (6.626 \times 10^{-34} \text{ J} \cdot \cancel{\text{s}})(7.23 \times 10^{14} \cancel{\text{s}^{-1}}) = 4.79 \times 10^{-19} \text{ J}$$

The energy of one photon of violet light is $4.79 \times 10^{-19} \text{ J}$.

3. Evaluate the Answer

The answer is correctly expressed in a unit of energy (J). The known value for frequency has three significant figures, and the answer also is expressed with three significant figures, as it should be. As expected, the energy of a single photon of light is extremely small.

PRACTICE PROBLEMS

- What is the energy of each of the following types of radiation?
 - $6.32 \times 10^{20} \text{ s}^{-1}$
 - $9.50 \times 10^{13} \text{ Hz}$
 - $1.05 \times 10^{16} \text{ s}^{-1}$
- Use Figure 5-5 to determine the types of radiation described in problem 5.

Practice!

For more practice with photon energy problems, go to **Supplemental Practice Problems** in Appendix A.

Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? The process illustrates another phenomenon that cannot be explained by the wave model of light. The light of the neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited and unstable atoms then release energy by emitting light. If the light emitted by the neon is passed through a glass prism, neon's atomic emission spectrum is produced. The **atomic emission spectrum** of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. Neon's atomic emission spectrum consists of several individual lines of color, not a continuous range of colors as seen in the visible spectrum.

Each element's atomic emission spectrum is unique and can be used to determine if that element is part of an unknown compound. For example, when a platinum wire is dipped into a strontium nitrate solution and then inserted into a burner flame, the strontium atoms emit a characteristic red color. You can perform a series of flame tests yourself by doing the **miniLAB** below.

Figure 5-9 on the following page shows an illustration of the characteristic purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen's emission spectrum responsible for producing the glow. Note how the line nature of hydrogen's atomic emission spectrum differs from that of a continuous spectrum. To gain firsthand experience with types of line spectra, you can perform the **CHEMLAB** at the end of this chapter.

miniLAB

Flame Tests

Classifying When certain compounds are heated in a flame, they emit a distinctive color. The color of the emitted light can be used to identify the compound.

Materials Bunsen burner; cotton swabs (6); distilled water; crystals of lithium chloride, sodium chloride, potassium chloride, calcium chloride, strontium chloride, unknown

Procedure

1. Dip a cotton swab into the distilled water. Dip the moistened swab into the lithium chloride so that a few of the crystals stick to the cotton. Put the crystals on the swab into the flame of a Bunsen burner. Observe the color of the flame and record it in your data table.
2. Repeat step 1 for each of the metallic chlorides (sodium chloride, potassium chloride, calcium chloride, and strontium chloride). Be sure to record the color of each flame in your data table.
3. Obtain a sample of unknown crystals from your teacher. Repeat the procedure in step 1 using

Flame Test Results

Compound	Flame color
Lithium chloride	
Sodium chloride	
Potassium chloride	
Calcium chloride	
Strontium chloride	
Unknown	

the unknown crystals. Record the color of the flame produced by the unknown crystals in your data table. Dispose of used cotton swabs as directed by your teacher.

Analysis

1. Each of the known compounds tested contains chlorine, yet each compound produced a flame of a different color. Explain why this occurred.
2. How is the atomic emission spectrum of an element related to these flame tests?
3. What is the identity of the unknown crystals? Explain how you know.

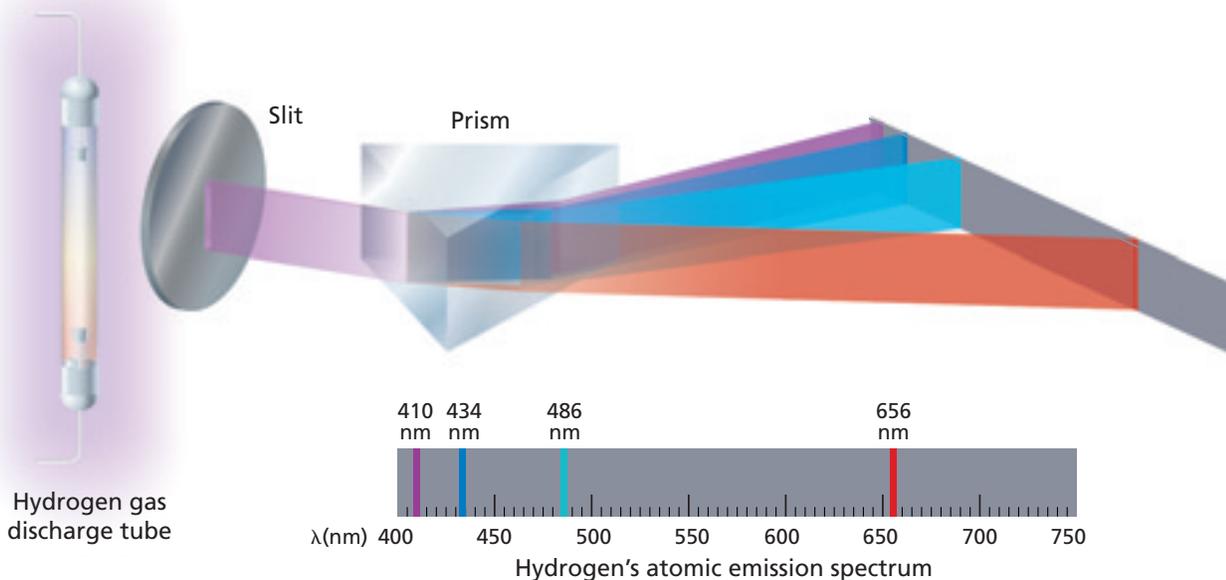


Figure 5-9

The atomic emission spectrum of hydrogen consists of four distinct colored lines of different frequencies. This type of spectrum is also known as a line spectrum. Which line has the highest energy?

An atomic emission spectrum is characteristic of the element being examined and can be used to identify that element. The fact that only certain colors appear in an element's atomic emission spectrum means that only certain specific frequencies of light are emitted. And because those emitted frequencies of light are related to energy by the formula $E_{\text{photon}} = h\nu$, it can be concluded that only photons having certain specific energies are emitted. This conclusion was not predicted by the laws of classical physics known at that time. Scientists found atomic emission spectra puzzling because they had expected to observe the emission of a continuous series of colors and energies as excited electrons lost energy and spiraled toward the nucleus. In the next section, you will learn about the continuing development of atomic models, and how one of those models was able to account for the frequencies of the light emitted by excited atoms.

Section 5.1 Assessment

7. List the characteristic properties of all waves. At what speed do electromagnetic waves travel in a vacuum?
8. Compare the wave and particle models of light. What phenomena can only be explained by the particle model?
9. What is a quantum of energy? Explain how quanta of energy are involved in the amount of energy matter gains and loses.
10. Explain the difference between the continuous spectrum of white light and the atomic emission spectrum of an element.
11. **Thinking Critically** Explain how Einstein utilized Planck's quantum concept in explaining the photoelectric effect.
12. **Interpreting Scientific Illustrations** Use **Figure 5-5** and your knowledge of light to match the numbered items on the right with the lettered items on the left. The numbered items may be used more than once or not at all.

a. longest wavelength	1. gamma rays
b. highest frequency	2. infrared waves
c. greatest energy	3. radio waves

You now know that the behavior of light can be explained only by a dual wave-particle model. Although this model was successful in accounting for several previously unexplainable phenomena, an understanding of the relationships among atomic structure, electrons, and atomic emission spectra still remained to be established.

Bohr Model of the Atom

Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are elements' atomic emission spectra discontinuous rather than continuous? Niels Bohr, a young Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Impressively, Bohr's model also correctly predicted the frequencies of the lines in hydrogen's atomic emission spectrum.

Energy states of hydrogen Building on Planck's and Einstein's concepts of quantized energy (quantized means that only certain values are allowed), Bohr proposed that the hydrogen atom has only certain allowable energy states. The lowest allowable energy state of an atom is called its **ground state**. When an atom gains energy, it is said to be in an excited state. And although a hydrogen atom contains only a single electron, it is capable of having many different excited states.

Bohr went even further with his atomic model by relating the hydrogen atom's energy states to the motion of the electron within the atom. Bohr suggested that the single electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron's orbit, the lower the atom's energy state, or energy level. Conversely, the larger the electron's orbit, the higher the atom's energy state, or energy level. Bohr assigned a quantum number, n , to each orbit and even calculated the orbit's radius. For the first orbit, the one closest to the nucleus, $n = 1$ and the orbit radius is 0.0529 nm; for the second orbit, $n = 2$ and the orbit radius is 0.212 nm; and so on. Additional information about Bohr's description of hydrogen's allowable orbits and energy levels is given in **Table 5-1**.

Table 5-1

Bohr's Description of the Hydrogen Atom				
Bohr atomic orbit	Quantum number	Orbit radius (nm)	Corresponding atomic energy level	Relative energy
First	$n = 1$	0.0529	1	E_1
Second	$n = 2$	0.212	2	$E_2 = 4E_1$
Third	$n = 3$	0.476	3	$E_3 = 9E_1$
Fourth	$n = 4$	0.846	4	$E_4 = 16E_1$
Fifth	$n = 5$	1.32	5	$E_5 = 25E_1$
Sixth	$n = 6$	1.90	6	$E_6 = 36E_1$
Seventh	$n = 7$	2.59	7	$E_7 = 49E_1$

Objectives

- **Compare** the Bohr and quantum mechanical models of the atom.
- **Explain** the impact of de Broglie's wave-particle duality and the Heisenberg uncertainty principle on the modern view of electrons in atoms.
- **Identify** the relationships among a hydrogen atom's energy levels, sublevels, and atomic orbitals.

Vocabulary

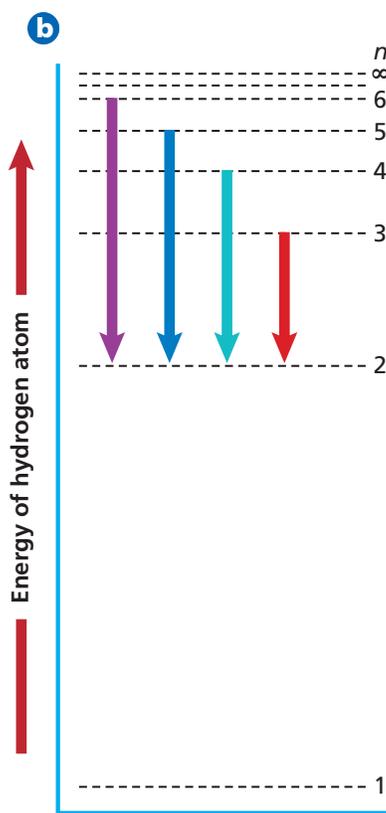
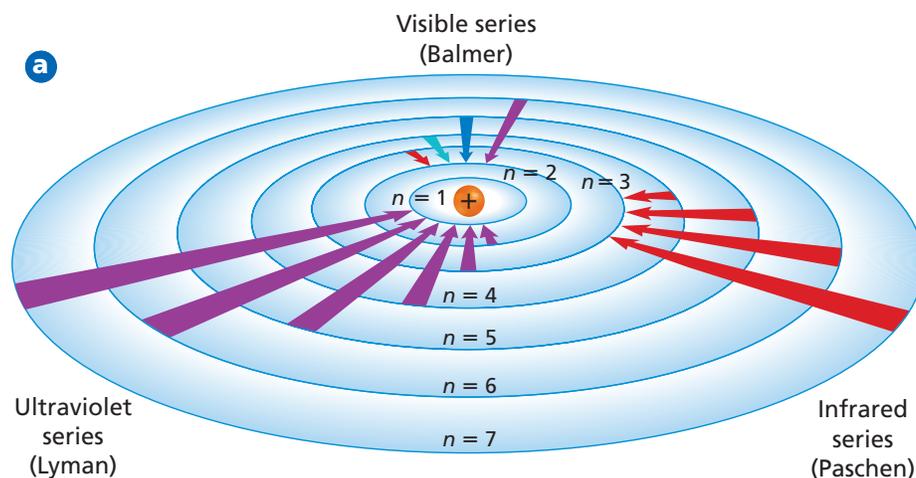
ground state
de Broglie equation
Heisenberg uncertainty principle
quantum mechanical model of the atom
atomic orbital
principal quantum number
principal energy level
energy sublevel

An explanation of hydrogen's line spectrum Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when the electron is in the $n = 1$ orbit. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit such as the $n = 2$ orbit shown in **Figure 5-10a**. Such an electron transition raises the atom to an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit. As a result of this transition, the atom emits a photon corresponding to the difference between the energy levels associated with the two orbits.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

Figure 5-10

a When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon with a specific energy is emitted. Although hydrogen has spectral lines associated with higher energy levels, only the visible, ultraviolet, and infrared series of spectral lines are shown in this diagram. **b** The relative energies of the electron transitions responsible for hydrogen's four visible spectral lines are shown. Note how the energy levels become more closely spaced as n increases.



Note that because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted. You might compare hydrogen's seven atomic orbits to seven rungs on a ladder. A person can climb up or down the ladder only from rung to rung. Similarly, the hydrogen atom's electron can move only from one allowable orbit to another, and therefore, can emit or absorb only certain amounts of energy.

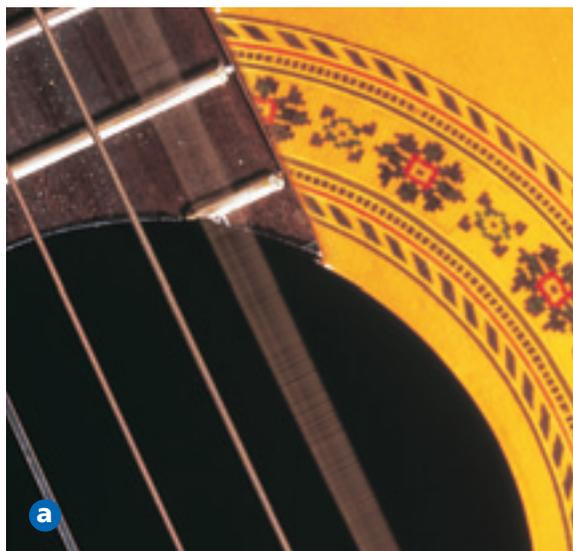
The four electron transitions that account for visible lines in hydrogen's atomic emission spectrum are shown in **Figure 5-10b**. For example, electrons dropping from the third orbit to the second orbit cause the red line. Note that electron transitions from higher-energy orbits to the second orbit account for all of hydrogen's visible lines. This series of visible lines is called the Balmer series. Other electron transitions have been measured that are not visible, such as the Lyman series (ultraviolet) in which electrons drop into the $n = 1$ orbit and the Paschen series (infrared) in which electrons drop into the $n = 3$ orbit. **Figure 5-10b** also shows that unlike rungs on a ladder, the hydrogen atom's energy levels are not evenly spaced. You will be able to see in greater detail how Bohr's atomic model was able to account for hydrogen's line spectrum by doing the **problem-solving LAB** later in this chapter.

Bohr's model explained hydrogen's observed spectral lines remarkably well. Unfortunately, however, the model failed to explain the spectrum of any other element. Moreover, Bohr's model did not fully account for the chemical behavior of atoms. In fact, although Bohr's idea of quantized energy levels laid the groundwork for atomic models to come, later experiments demonstrated that the Bohr model was fundamentally incorrect. The movements of electrons in atoms are not completely understood even now; however, substantial evidence indicates that electrons do not move around the nucleus in circular orbits.

The Quantum Mechanical Model of the Atom

Scientists in the mid-1920s, by then convinced that the Bohr atomic model was incorrect, formulated new and innovative explanations of how electrons are arranged in atoms. In 1924, a young French graduate student in physics named Louis de Broglie (1892–1987) proposed an idea that eventually accounted for the fixed energy levels of Bohr’s model.

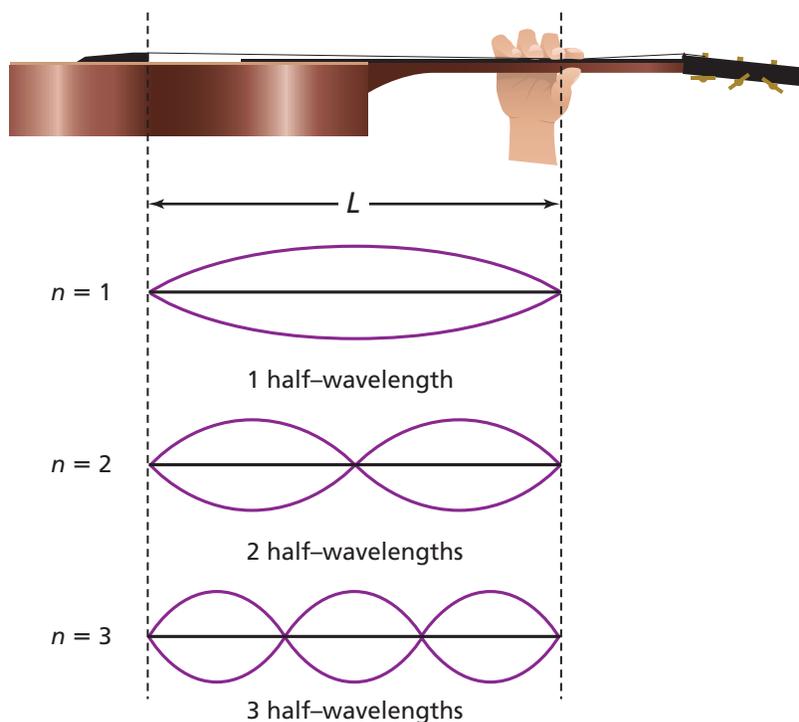
Electrons as waves De Broglie had been thinking that Bohr’s quantized electron orbits had characteristics similar to those of waves. For example, as **Figure 5-11b** shows, only multiples of half-wavelengths are possible on a plucked guitar string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 5-11c**. He also reflected on



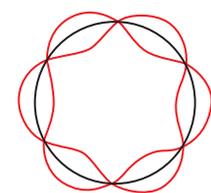
the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question. If waves can have particlelike behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

Figure 5-11

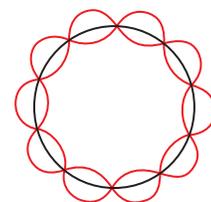
a A vibrating guitar string is constrained to vibrate between two fixed end points. **b** The possible vibrations of the guitar string are limited to multiples of half-wavelengths. Thus, the “quantum” of the guitar string is one-half wavelength. **c** The possible circular orbits of an electron are limited to whole numbers of complete wavelengths.



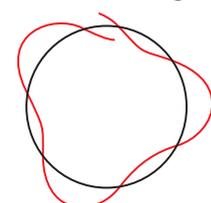
b Vibrating guitar string
Only multiples of half wavelengths allowed



$n = 3$ wavelengths



$n = 5$ wavelengths



$n \neq$ whole number (not allowed)

c Orbiting electron
Only whole numbers of wavelengths allowed

How was Bohr's atomic model able to explain the line spectrum of hydrogen?

Using Models Niels Bohr proposed that electrons must occupy specific, quantized energy levels in an atom. He derived the following equations for hydrogen's electron orbit energies (E_n) and radii (r_n).

$$r_n = (0.529 \times 10^{-10} \text{ m})n^2$$

$$E_n = -(2.18 \times 10^{-18} \text{ J})/n^2$$

Where n = quantum number (1, 2, 3...).

Analysis

Using the orbit radii equation, calculate hydrogen's first seven electron orbit radii and then construct a scale model of those orbits. Use a compass and a metric ruler to draw your scale model on two sheets of paper that have been taped together. (Use caution when handling sharp objects.) Using the orbit energy equation, calculate the energy of each electron orbit and record the values on your model.

Thinking Critically

1. What scale did you use to plot the orbits? How is the energy of each orbit related to its radius?
2. Draw a set of arrows for electron jumps that end at each energy level (quantum number). For example, draw a set of arrows for all transitions that end at $n = 1$, a set of arrows for all transitions that end at $n = 2$, and so on, up to $n = 7$.
3. Calculate the energy released for each of the jumps in step 2, and record the values on your model. The energy released is equal to the difference in the energies of each level.
4. Each set of arrows in step 2 represents a spectral emission series. Label five of the series, from greatest energy change to least energy change, as the Lyman, Balmer, Paschen, Brackett, and Pfund series.
5. Use the energy values in step 3 to calculate the frequency of each photon emitted in each series. Record the frequencies on your model.
6. Using the electromagnetic spectrum as a guide, identify in which range (visible, ultraviolet, infrared, etc.) each series falls.

In considering this question, de Broglie knew that if an electron has wave-like motion and is restricted to circular orbits of fixed radius, the electron is allowed only certain possible wavelengths, frequencies, and energies. Developing his idea, de Broglie derived an equation for the wavelength (λ) of a particle of mass (m) moving at velocity (v).

$$\lambda = \frac{h}{mv}$$

The **de Broglie equation** predicts that all moving particles have wave characteristics. Why, then, you may be wondering, haven't you noticed the wavelength of a fast-moving automobile? Using de Broglie's equation provides an answer. An automobile moving at 25 m/s and having a mass of 910 kg has a wavelength of 2.9×10^{-38} m—a wavelength far too small to be seen or detected, even with the most sensitive scientific instrument. By comparison, an electron moving at the same speed has the easily measured wavelength of 2.9×10^{-5} m. Subsequent experiments have proven that electrons and other moving particles do indeed have wave characteristics.

Step by step, scientists such as Rutherford, Bohr, and de Broglie had been unraveling the mysteries of the atom. However, a conclusion reached by the German theoretical physicist Werner Heisenberg (1901–1976), a contemporary of de Broglie, proved to have profound implications for atomic models.

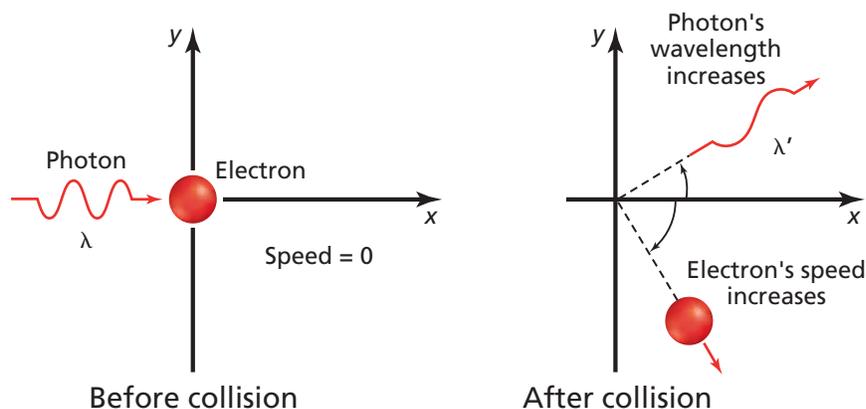


Figure 5-12

A photon that strikes an electron at rest alters the position and velocity of the electron. This collision illustrates the Heisenberg uncertainty principle: It is impossible to simultaneously know both the position and velocity of a particle. Note that after the collision, the photon's wavelength is longer. How has the photon's energy changed?

The Heisenberg Uncertainty Principle

Heisenberg's concluded that it is impossible to make any measurement on an object without disturbing the object—at least a little. Imagine trying to locate a hovering, helium-filled balloon in a completely darkened room. When you wave your hand about, you'll locate the balloon's position when you touch it. However, when you touch the balloon, even gently, you transfer energy to it and change its position. Of course, you could also detect the balloon's position by turning on a flashlight. Using this method, photons of light that reflect from the balloon reach your eyes and reveal the balloon's location. Because the balloon is much more massive than the photons, the rebounding photons have virtually no effect on the balloon's position.

Can photons of light help determine the position of an electron in an atom? As a thought experiment, imagine trying to determine the electron's location by “bumping” it with a high-energy photon of electromagnetic radiation. Unfortunately, because such a photon has about the same energy as an electron, the interaction between the two particles changes both the wavelength of the photon and the position and velocity of the electron, as shown in **Figure 5-12**. In other words, the act of observing the electron produces a significant, unavoidable uncertainty in the position and motion of the electron. Heisenberg's analysis of interactions such as those between photons and electrons led him to his historic conclusion. The **Heisenberg uncertainty principle** states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Although scientists of the time found Heisenberg's principle difficult to accept, it has been proven to describe the fundamental limitations on what can be observed. How important is the Heisenberg uncertainty principle? The interaction of a photon with an object such as a helium-filled balloon has so little effect on the balloon that the uncertainty in its position is too small to measure. But that's not the case with an electron moving at 6×10^6 m/s near an atomic nucleus. The uncertainty in the electron's position is at least 10^{-9} m, about ten times greater than the diameter of the entire atom!

The Schrödinger wave equation In 1926, Austrian physicist Erwin Schrödinger (1887–1961) furthered the wave-particle theory proposed by de Broglie. Schrödinger derived an equation that treated the hydrogen atom's electron as a wave. Remarkably, Schrödinger's new model for the hydrogen atom seemed to apply equally well to atoms of other elements—an area in which Bohr's model failed. The atomic model in which electrons are treated as waves is called the wave mechanical model of the atom or, more commonly, the **quantum mechanical model of the atom**. Like Bohr's model,

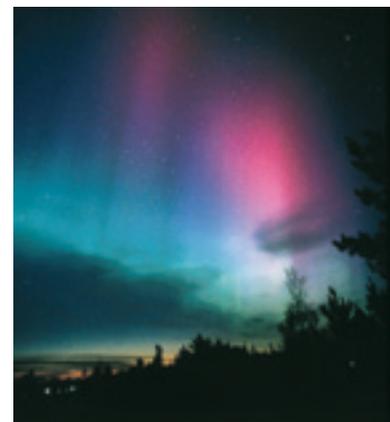
Physics

CONNECTION

People travel thousands of miles to see the aurora borealis (the northern lights) and the aurora australis (the southern lights). Once incorrectly believed to be reflections from the polar ice fields, the auroras occur 100 to 1000 km above Earth.

High-energy electrons and positive ions in the solar wind speed away from the sun at more than one million kilometers per hour. These particles become trapped in Earth's magnetic field and follow along Earth's magnetic field lines.

The electrons interact with and transfer energy to oxygen and nitrogen atoms in the upper atmosphere. The color of the aurora depends on altitude and which atoms become excited. Oxygen emits green light up to about 250 km and red light above 250 km; nitrogen emits blue light up to about 100 km and purple/violet at higher altitudes.



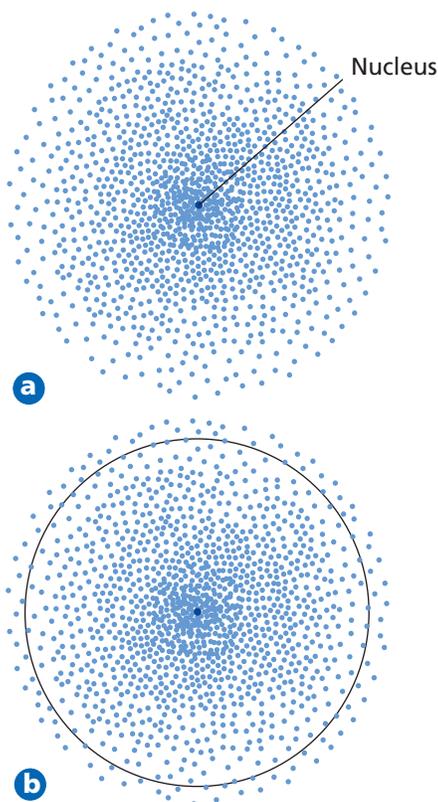


Figure 5-13

a This electron density diagram for a hydrogen atom represents the likelihood of finding an electron at a particular point in the atom. The greater the density of dots, the greater the likelihood of finding hydrogen's electron. **b** The boundary of an atom is defined as the volume that encloses a 90% probability of containing its electrons.

the quantum mechanical model limits an electron's energy to certain values. However, unlike Bohr's model, the quantum mechanical model makes no attempt to describe the electron's path around the nucleus.

The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function. And most importantly, the wave function is related to the probability of finding the electron within a particular volume of space around the nucleus. Recall from your study of math that an event having a high probability is more likely to occur than one having a low probability.

What does the wave function predict about the electron's location in an atom? A three-dimensional region around the nucleus called an **atomic orbital** describes the electron's probable location. You can picture an atomic orbital as a fuzzy cloud in which the density of the cloud at a given point is proportional to the probability of finding the electron at that point. **Figure 5-13a** illustrates the probability map, or orbital, that describes the hydrogen electron in its lowest energy state. It might be helpful to think of the probability map as a time-exposure photograph of the electron moving around the nucleus, in which each dot represents the electron's location at an instant in time. Because the dots are so numerous near the positive nucleus, they seem to form a dense cloud that is indicative of the electron's most probable location. However, because the cloud has no definite boundary, it also is possible that the electron might be found at a considerable distance from the nucleus.

Hydrogen's Atomic Orbitals

Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exactly defined size. To overcome the inherent uncertainty about the electron's location, chemists arbitrarily draw an orbital's surface to contain 90% of the electron's total probability distribution. In other words, the electron spends 90% of the time within the volume defined by the surface, and 10% of the time somewhere outside the surface. The spherical surface shown in **Figure 5-13b** encloses 90% of the lowest-energy orbital of hydrogen.

Recall that the Bohr atomic model assigns quantum numbers to electron orbits. In a similar manner, the quantum mechanical model assigns **principal quantum numbers** (n) that indicate the relative sizes and energies of atomic

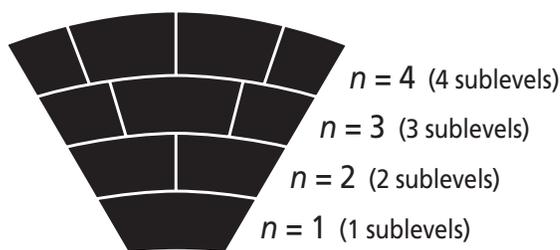


Figure 5-14

Energy sublevels can be thought of as a section of seats in a theater. The rows that are higher up and farther from the stage contain more seats, just as energy levels that are farther from the nucleus contain more sublevels.



orbitals. That is, as n increases, the orbital becomes larger, the electron spends more time farther from the nucleus, and the atom's energy level increases. Therefore, n specifies the atom's major energy levels, called **principal energy levels**. An atom's lowest principal energy level is assigned a principal quantum number of one. When the hydrogen atom's single electron occupies an orbital with $n = 1$, the atom is in its ground state. Up to seven energy levels have been detected for the hydrogen atom, giving n values ranging from 1 to 7.

Principal energy levels contain **energy sublevels**. Principal energy level 1 consists of a single sublevel, principal energy level 2 consists of two sublevels, principal energy level 3 consists of three sublevels, and so on. To better understand the relationship between the atom's energy levels and sublevels, picture the seats in a wedge-shaped section of a theater, as shown in **Figure 5-14**. As you move away from the stage, the rows become higher and contain more seats. Similarly, the number of energy sublevels in a principal energy level increases as n increases.

Sublevels are labeled s, p, d, or f according to the shapes of the atom's orbitals. All s orbitals are spherical and all p orbitals are dumbbell shaped; however, not all d or f orbitals have the same shape. Each orbital may contain at most two electrons. The single sublevel in principal energy level 1 consists of a spherical orbital called the 1s orbital. The two sublevels in principal energy level 2 are designated 2s and 2p. The 2s sublevel consists of the 2s orbital, which is spherical like the 1s orbital but larger in size. See **Figure 5-15a**. The 2p sublevel consists of three dumbbell-shaped p orbitals of equal energy designated $2p_x$, $2p_y$, and $2p_z$. The subscripts x , y , and z merely designate the orientations of p orbitals along the x , y , and z coordinate axes, as shown in **Figure 5-15b**.

Principal energy level 3 consists of three sublevels designated 3s, 3p, and 3d. Each d sublevel consists of five orbitals of equal energy. Four d orbitals have identical shapes but different orientations. However, the fifth, d_{z^2} orbital is shaped and oriented differently from the other four. The shapes and orientations of the five d orbitals are illustrated in **Figure 5-16**. The fourth principal energy level ($n = 4$) contains a fourth sublevel, called the 4f sublevel, which consists of seven f orbitals of equal energy.

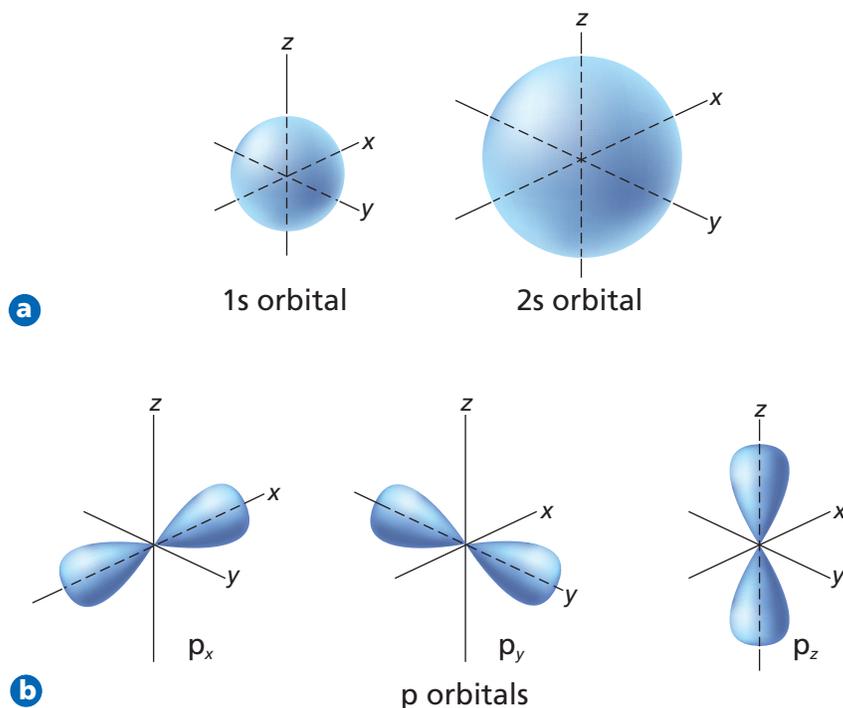


Figure 5-15

Atomic orbitals represent the electron probability clouds of an atom's electrons. **a** The spherical 1s and 2s orbitals are shown here. All s orbitals are spherical in shape and increase in size with increasing principal quantum number. **b** The three dumbbell-shaped p orbitals are oriented along the three perpendicular x , y , and z axes. Each of the p orbitals related to an energy sublevel has equal energy.

Figure 5-16

Four of five equal-energy d orbitals have the same shape. Notice how the d_{xy} orbital lies in the plane formed by the x and y axes, the d_{xz} orbital lies in the plane formed by the x and z axes, and so on. The d_{z^2} orbital has its own unique shape.

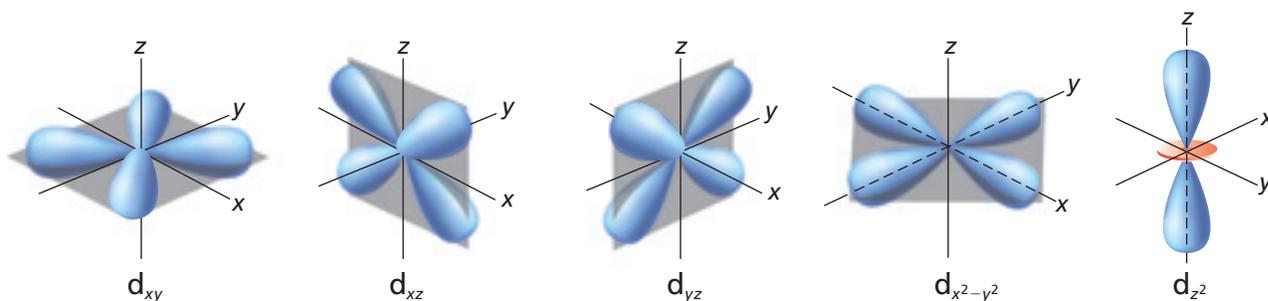


Table 5-2

Hydrogen's First Four Principal Energy Levels			
Principal quantum number (n)	Sublevels (types of orbitals) present	Number of orbitals related to sublevel	Total number of orbitals related to principal energy level (n^2)
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

Hydrogen's first four principal energy levels, sublevels, and related atomic orbitals are summarized in **Table 5-2**. Note that the maximum number of orbitals related to each principal energy level equals n^2 . Because each orbital may contain at most two electrons, the maximum number of electrons related to each principal energy level equals $2n^2$.

Given the fact that a hydrogen atom contains only one electron, you might wonder how the atom can have so many energy levels, sublevels, and related atomic orbitals. At any given time, the atom's electron can occupy just one orbital. So you can think of the other orbitals as unoccupied spaces—spaces available should the atom's energy increase or decrease. For example, when the hydrogen atom is in the ground state, the electron occupies the 1s orbital. However, the atom may gain a quantum of energy that excites the electron to the 2s orbital, to one of the three 2p orbitals, or to another vacant orbital.

You have learned a lot about electrons and quantized energy in this section: how Bohr's orbits explained the hydrogen atom's quantized energy states; how de Broglie's insight led scientists to think of electrons as both particles and waves; and how Schrödinger's wave equation predicted the existence of atomic orbitals containing electrons. In the next section, you'll learn how the electrons are arranged in atomic orbitals of atoms having more than one electron.

Section 5.2 Assessment

- According to the Bohr atomic model, why do atomic emission spectra contain only certain frequencies of light?
- Why is the wavelength of a moving soccer ball not detectable to the naked eye?
- What sublevels are contained in the hydrogen atom's first four energy levels? What orbitals are related to each s sublevel and each p sublevel?
- Thinking Critically** Use de Broglie's wave-particle duality and the Heisenberg uncertainty principle to explain why the location of an electron in an atom is uncertain.
- Comparing and Contrasting** Compare and contrast the Bohr model and quantum mechanical model of the atom.

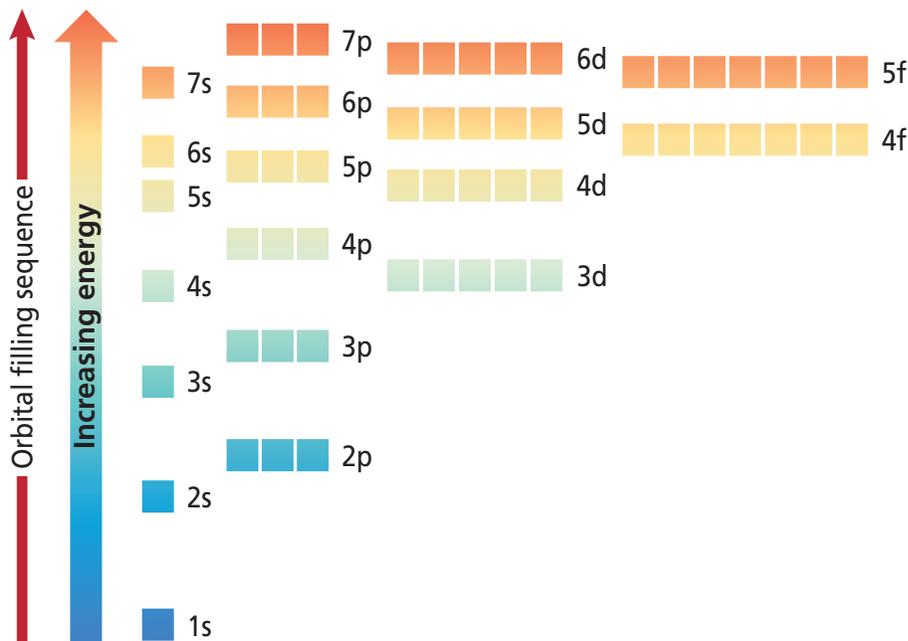
When you consider that atoms of the heaviest elements contain in excess of 100 electrons, that there are numerous principal energy levels and sublevels and their corresponding orbitals, and that each orbital may contain a maximum of two electrons, the idea of determining the arrangement of an atom's electrons seems daunting. Fortunately, the arrangement of electrons in atoms follows a few very specific rules. In this section, you'll learn these rules and their occasional exceptions.

Ground-State Electron Configurations

The arrangement of electrons in an atom is called the atom's **electron configuration**. Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest possible energy. The most stable, lowest-energy arrangement of the electrons in atoms of each element is called the element's ground-state electron configuration. Three rules, or principles—the aufbau principle, the Pauli exclusion principle, and Hund's rule—define how electrons can be arranged in an atom's orbitals.

The aufbau principle The **aufbau principle** states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element's ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in **Figure 5-17**. In the diagram, each box represents an atomic orbital. Several features of the aufbau diagram stand out.

- *All orbitals related to an energy sublevel are of equal energy.* For example, all three 2p orbitals are of equal energy.
- *In a multi-electron atom, the energy sublevels within a principal energy level have different energies.* For example, the three 2p orbitals are of higher energy than the 2s orbital.



Objectives

- **Apply** the Pauli exclusion principle, the aufbau principle, and Hund's rule to write electron configurations using orbital diagrams and electron configuration notation.
- **Define** valence electrons and draw electron-dot structures representing an atom's valence electrons.

Vocabulary

electron configuration
 aufbau principle
 Pauli exclusion principle
 Hund's rule
 valence electron
 electron-dot structure

Figure 5-17

The aufbau diagram shows the energy of each sublevel. Each box on the diagram represents an atomic orbital. Does the 3d or 4s sublevel have greater energy?

Careers Using Chemistry

Spectroscopist

Are you interested in the composition of the materials around you? Do you wonder what stars are made of? Then consider a career as a spectroscopist.

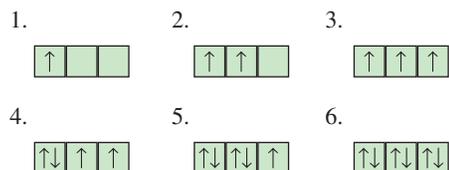
Spectroscopy is the analysis of the characteristic spectra emitted by matter. Spectroscopists perform chemical analyses as part of many research laboratory projects, for quality control in industrial settings, and as part of forensics investigations for law enforcement agencies.

- In order of increasing energy, the sequence of energy sublevels within a principal energy level is *s*, *p*, *d*, and *f*.
- Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level. For example, the orbital related to the atom's 4*s* sublevel has a lower energy than the five orbitals related to the 3*d* sublevel.

Although the aufbau principle describes the sequence in which orbitals are filled with electrons, it's important to know that atoms are not actually built up electron by electron.

The Pauli exclusion principle Each electron in an atom has an associated spin, similar to the way a top spins on its axis. Like the top, the electron is able to spin in only one of two directions. An arrow pointing up (\uparrow) represents the electron spinning in one direction, an arrow pointing down (\downarrow) represents the electron spinning in the opposite direction. The **Pauli exclusion principle** states that a maximum of two electrons may occupy a single atomic orbital, but only if the electrons have opposite spins. Austrian physicist Wolfgang Pauli proposed this principle after observing atoms in excited states. An atomic orbital containing paired electrons with opposite spins is written as $\uparrow\downarrow$.

Hund's rule The fact that negatively charged electrons repel each other has an important impact on the distribution of electrons in equal-energy orbitals. **Hund's rule** states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. For example, let the boxes below represent the 2*p* orbitals. One electron enters each of the three 2*p* orbitals before a second electron enters any of the orbitals. The sequence in which six electrons occupy three *p* orbitals is shown below.



Orbital Diagrams and Electron Configuration Notations

You can represent an atom's electron configuration using two convenient methods. One method is called an orbital diagram. An orbital diagram includes a box for each of the atom's orbitals. An empty box \square represents an unoccupied orbital; a box containing a single up arrow \uparrow represents an orbital with one electron; and a box containing both up and down arrows $\uparrow\downarrow$ represents a filled orbital. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, which contains two electrons in the 1*s* orbital, two electrons in the 2*s* orbital, and 1 electron in two of three separate 2*p* orbitals, is shown below.

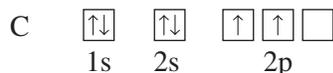
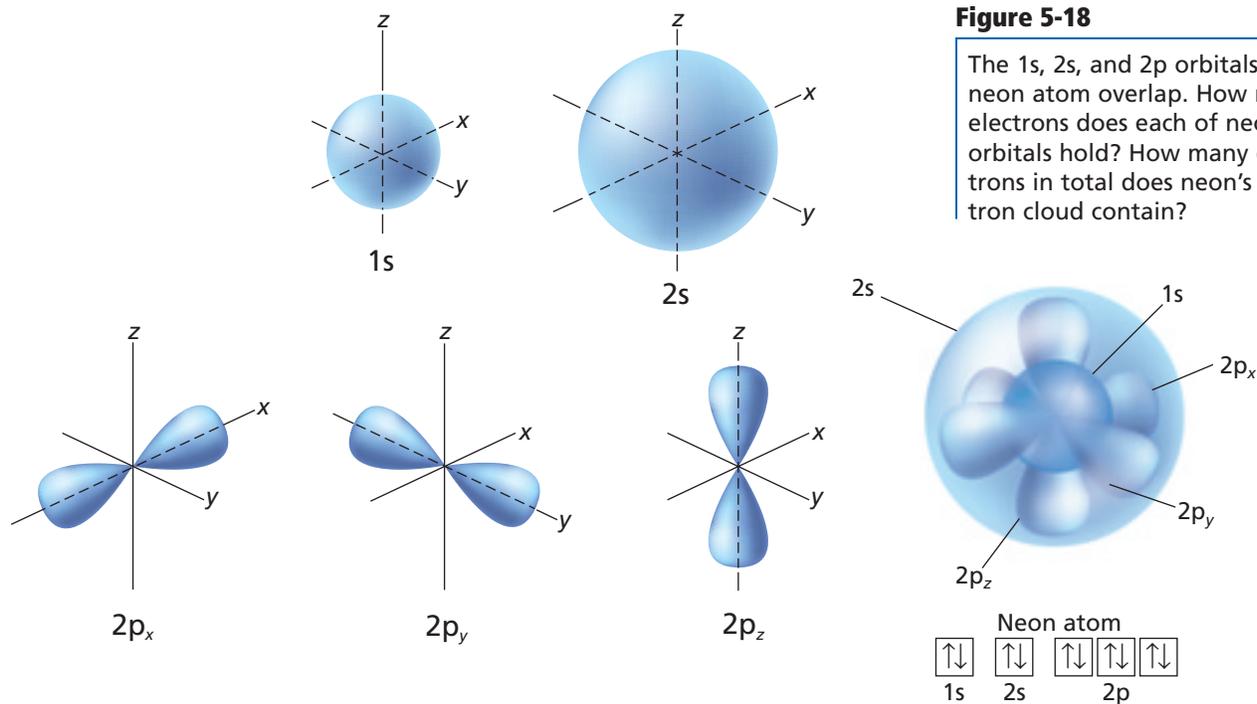


Table 5-3

Electron Configurations and Orbital Diagrams for Elements in the First Two Periods					
Element	Atomic number	Orbital diagram			Electron configuration notation
		1s	2s	2p _x 2p _y 2p _z	
Hydrogen	1	↑			1s ¹
Helium	2	↑↓			1s ²
Lithium	3	↑↓	↑		1s ² 2s ¹
Beryllium	4	↑↓	↑↓		1s ² 2s ²
Boron	5	↑↓	↑↓	↑	1s ² 2s ² 2p ¹
Carbon	6	↑↓	↑↓	↑ ↑	1s ² 2s ² 2p ²
Nitrogen	7	↑↓	↑↓	↑ ↑ ↑	1s ² 2s ² 2p ³
Oxygen	8	↑↓	↑↓	↑↓ ↑ ↑	1s ² 2s ² 2p ⁴
Fluorine	9	↑↓	↑↓	↑↓ ↑↓ ↑	1s ² 2s ² 2p ⁵
Neon	10	↑↓	↑↓	↑↓ ↑↓ ↑↓	1s ² 2s ² 2p ⁶

Recall that the number of electrons in an atom equals the number of protons, which is designated by the element's atomic number. Carbon, which has an atomic number of six, has six electrons in its configuration.

Another shorthand method for describing the arrangement of electrons in an element's atoms is called electron configuration notation. This method designates the principal energy level and energy sublevel associated with each of the atom's orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written 1s²2s²2p². Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 5-3**. To help you visualize the relative sizes and orientations of atomic orbitals, the filled 1s, 2s, 2p_x, 2p_y, and 2p_z orbitals of the neon atom are illustrated in **Figure 5-18**.



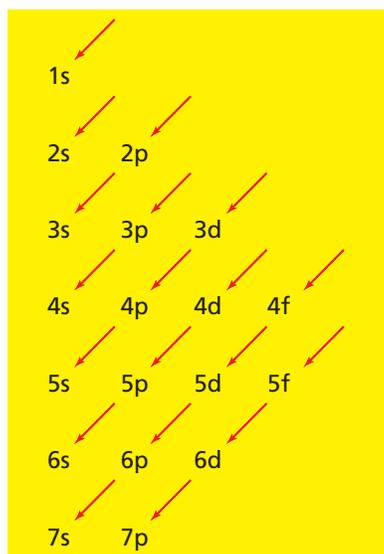
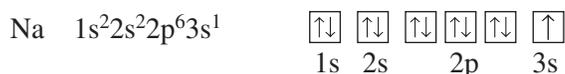


Figure 5-19

This sublevel diagram shows the order in which the orbitals are usually filled. The proper sequence for the first seven orbitals is 1s, 2s, 2p, 3s, 3p, 4s, and 3d. Which is filled first, the 5s or the 4p orbital?

Note that electron configuration notation usually does not show the orbital distributions of electrons related to a sublevel. It's understood that a designation such as nitrogen's $2p^3$ represents the orbital occupancy $2p_x^1 2p_y^1 2p_z^1$.

For sodium, the first ten electrons occupy 1s, 2s, and 2p orbitals. Then, according to the aufbau sequence, the eleventh electron occupies the 3s orbital. The electron configuration notation and orbital diagram for sodium are written



Noble-gas notation is a method of representing electron configurations of noble gases using bracketed symbols. For example, [He] represents the electron configuration for helium, $1s^2$, and [Ne] represents the electron configuration for neon, $1s^2 2s^2 2p^6$. Compare the electron configuration for neon with sodium's configuration above. Note that the inner-level configuration for sodium is identical to the electron configuration for neon. Using noble-gas notation, sodium's electron configuration can be shortened to the form [Ne] $3s^1$. The electron configuration for an element can be represented using the noble-gas notation for the noble gas in the previous period and the electron configuration for the energy level being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in **Table 5-4**.

When writing electron configurations, you may refer to a convenient memory aid called a sublevel diagram, which is shown in **Figure 5-19**. Note that following the direction of the arrows in the sublevel diagram produces the sublevel sequence shown in the aufbau diagram of **Figure 5-17**.

Exceptions to predicted configurations You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, [Ar] $4s^2 3d^4$, and copper, [Ar] $4s^2 3d^9$, would prove to be incorrect. The correct configurations for these two elements are:



The electron configurations for these two elements, as well as those of several elements in other periods, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

Table 5-4

Electron Configurations for Elements in Period Three			
Element	Atomic number	Complete electron configuration	Electron configuration using noble-gas notation
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	[Ne] $3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	[Ne] $3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	[Ne] $3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	[Ne] $3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	[Ne] $3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	[Ne] $3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	[Ne] $3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	[Ne] $3s^2 3p^6$ or [Ar]

EXAMPLE PROBLEM 5-3

Writing Electron Configurations

Germanium (Ge), a semiconducting element, is commonly used in the manufacture of computer chips. What is the ground-state electron configuration for an atom of germanium?

1. Analyze the Problem

You are given the semiconducting element, germanium (Ge). Consult the periodic table to determine germanium's atomic number, which also is equal to its number of electrons. Also note the atomic number of the noble gas element that precedes germanium in the table. Determine the number of additional electrons a germanium atom has compared to the nearest preceding noble gas, and then write out germanium's electron configuration.

2. Solve for the Unknown

From the periodic table, germanium's atomic number is determined to be 32. Thus, a germanium atom contains 32 electrons. The noble gas preceding germanium is argon (Ar), which has an atomic number of 18. Represent germanium's first 18 electrons using the chemical symbol for argon written inside brackets.

[Ar]

The remaining 14 electrons of germanium's configuration need to be written out. Because argon is a noble gas in the third period of the periodic table, it has completely filled 3s and 3p orbitals. Thus, the remaining 14 electrons fill the 4s, 3d, and 4p orbitals in order.

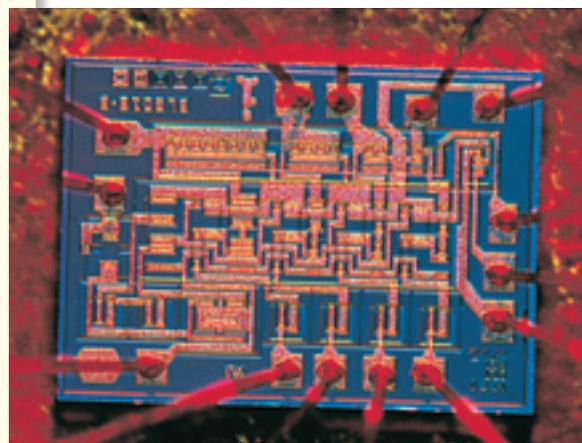
[Ar]4s²3d⁷4p²

Using the maximum number of electrons that can fill each orbital, write out the electron configuration.

[Ar]4s²3d¹⁰4p²

3. Evaluate the Answer

All 32 electrons in a germanium atom have been accounted for. The correct preceding noble gas (Ar) has been used in the notation, and the order of orbital filling for the fourth period is correct (4s, 3d, 4p).



Atoms of boron and arsenic are inserted into germanium's crystal structure in order to produce a semiconducting material that can be used to manufacture computer chips.

PRACTICE PROBLEMS

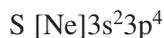
18. Write ground-state electron configurations for the following elements.
 - a. bromine (Br)
 - b. strontium (Sr)
 - c. antimony (Sb)
 - d. rhenium (Re)
 - e. terbium (Tb)
 - f. titanium (Ti)
19. How many electrons are in orbitals related to the third energy level of a sulfur atom?
20. How many electrons occupy p orbitals in a chlorine atom?
21. What element has the following ground-state electron configuration? [Kr]5s²4d¹⁰5p¹
22. What element has the following ground-state electron configuration? [Xe]6s²



For more practice with electron configuration problems, go to **Supplemental Practice Problems** in Appendix A.

Valence Electrons

Only certain electrons, called valence electrons, determine the chemical properties of an element. **Valence electrons** are defined as electrons in the atom's outermost orbitals—generally those orbitals associated with the atom's highest principal energy level. For example, a sulfur atom contains 16 electrons, only six of which occupy the outermost 3s and 3p orbitals, as shown by sulfur's electron configuration. Sulfur has six valence electrons.



Similarly, although a cesium atom contains 55 electrons, it has but one valence electron, the 6s electron shown in cesium's electron configuration.



Francium, which belongs to the same group as cesium, also has a single valence electron.



Electron-dot structures Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method. An atom's **electron-dot structure** consists of the element's symbol, which represents the atomic nucleus and inner-level electrons, surrounded by dots representing the atom's valence electrons. The American chemist G. N. Lewis (1875–1946), devised the method while teaching a college chemistry class in 1902.

In writing an atom's electron-dot structure, dots representing valence electrons are placed one at a time on the four sides of the symbol (they may be placed in any sequence) and then paired up until all are used. The ground-state electron configurations and electron-dot structures for the elements in the second period are shown in **Table 5-5**.

Table 5-5

Electron-Dot Structures for Elements in Period Two			
Element	Atomic number	Electron configuration	Electron-dot structure
Lithium	3	$1s^22s^1$	Li·
Beryllium	4	$1s^22s^2$	·Be·
Boron	5	$1s^22s^22p^1$	·B·
Carbon	6	$1s^22s^22p^2$	·C·
Nitrogen	7	$1s^22s^22p^3$	·N·
Oxygen	8	$1s^22s^22p^4$:Ö·
Fluorine	9	$1s^22s^22p^5$:F·
Neon	10	$1s^22s^22p^6$:Ne:

EXAMPLE PROBLEM 5-4

Writing Electron-Dot Structures

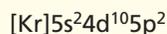
Some sheet glass is manufactured using a process that makes use of molten tin. What is tin's electron-dot structure?

1. Analyze the Problem

You are given the element tin (Sn). Consult the periodic table to determine the total number of electrons an atom of tin has. Write out tin's electron configuration and determine the number of valence electrons it has. Then use the number of valence electrons and the rules for electron-dot structures to draw the electron-dot structure for tin.

2. Solve for the Unknown

From the periodic table, tin is found to have an atomic number of 50. Thus, a tin atom has 50 electrons. Write out the noble-gas form of tin's electron configuration.

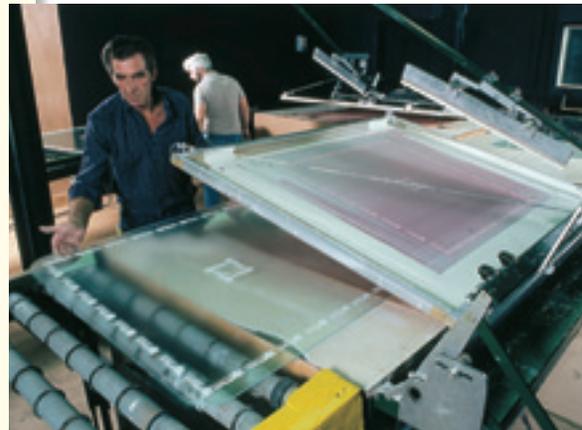


The two 5s and the two 5p electrons (the electrons in the orbitals related to the atom's highest principal energy level) represent tin's four valence electrons. Draw tin's electron-dot structure by representing its four valence electrons with dots, arranged one at a time, around the four sides of tin's chemical symbol (Sn).



3. Evaluate the Answer

The correct symbol for tin (Sn) has been used, and the rules for drawing electron-dot structures have been correctly applied.



Flat-surfaced window glass may be manufactured by floating molten glass on top of molten tin.

PRACTICE PROBLEMS

23. Draw electron-dot structures for atoms of the following elements.

- | | |
|--------------|-------------|
| a. magnesium | d. rubidium |
| b. sulfur | e. thallium |
| c. bromine | f. xenon |



For more practice with electron-dot structure problems, go to **Supplemental Practice Problems** in Appendix A.

Section 5.3 Assessment

24. State the aufbau principle in your own words.
25. Apply the Pauli exclusion principle, the aufbau principle, and Hund's rule to write out the electron configuration and draw the orbital diagram for each of the following elements.
- | | |
|-------------|------------|
| a. silicon | c. calcium |
| b. fluorine | d. krypton |
26. What is a valence electron? Draw the electron-dot structures for the elements in problem 25.
27. **Thinking Critically** Use Hund's rule and orbital diagrams to describe the sequence in which ten electrons occupy the five orbitals related to an atom's d sublevel.
28. **Interpreting Scientific Illustrations** Which of the following is the correct electron-dot structure for an atom of selenium? Explain.
- a. $\cdot \ddot{\text{Se}} :$ b. $\cdot \ddot{\text{Se}} \cdot$ c. $\cdot \ddot{\text{Se}} \cdot$ d. $\cdot \ddot{\text{S}} \cdot$

Line Spectra

You know that sunlight is made up of a continuous spectrum of colors that combine to form white light. You also have learned that atoms of gases can emit visible light of characteristic wavelengths when excited by electricity. The color you see is the sum of all of the emitted wavelengths. In this experiment, you will use a diffraction grating to separate these wavelengths into emission line spectra.

You also will investigate another type of line spectrum—the absorption spectrum. The color of each solution you observe is due to the reflection or transmission of unabsorbed wavelengths of light. When white light passes through a sample and then a diffraction grating, dark lines show up on the continuous spectrum of white light. These lines correspond to the wavelengths of the photons absorbed by the solution.

Problem

What absorption and emission spectra do various substances produce?

Objectives

- **Observe** emission spectra of several gases.
- **Observe** the absorption spectra of various solutions.
- **Analyze** patterns of absorption and emission spectra.

Materials

(For each group)
ring stand with clamp
40-W tubular light bulb
light socket with power cord
275-mL polystyrene culture flask (4)
Flinn C-Spectra® or similar diffraction grating

food coloring (red, green, blue, and yellow)
set of colored pencils
book

(For entire class)
spectrum tubes (hydrogen, neon, and mercury)
spectrum tube power supplies (3)

Safety Precautions



- Always wear safety goggles and a lab apron.
- Use care around the spectrum tube power supplies.
- Spectrum tubes will get hot when used.

Pre-Lab

1. Read the entire CHEMLAB.
2. Explain how electrons in an element's atoms produce an emission spectrum.
3. Distinguish among a continuous spectrum, an emission spectrum, and an absorption spectrum.
4. Prepare your data tables.

Drawings of Emission Spectra

Hydrogen	
Neon	
Mercury	

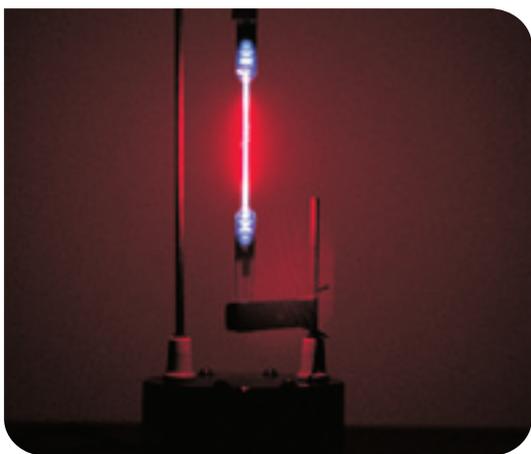
Drawings of Absorption Spectra

Red	
Green	
Blue	
Yellow	

Procedure

1. Use a Flinn C-Spectra® to view an incandescent light bulb. What do you observe? Draw the spectrum using colored pencils.

- Use the Flinn C-Spectra[®] to view the emission spectra from tubes of gaseous hydrogen, neon, and mercury. Use colored pencils to make drawings in the data table of the spectra observed.



- Fill a 275-mL culture flask with about 100-mL water. Add 2 or 3 drops of red food coloring to the water. Shake the solution.
- Repeat step 3 for the green, blue, and yellow food coloring. **CAUTION:** Be sure to thoroughly dry your hands before handling electrical equipment.
- Set up the light 40-W light bulb so that it is near eye level. Place the flask with red food coloring about 8 cm from the light bulb. Use a book or some other object to act as a stage to put the flask on. You should be able to see light from the bulb above the solution and light from the bulb projecting through the solution.



- With the room lights darkened, view the light using the Flinn C-Spectra[®]. The top spectrum viewed will be a continuous spectrum of the white light bulb. The bottom spectrum will be the absorption spectrum of the red solution. The black areas of the absorption spectrum represent the colors absorbed by the red food coloring in the solution. Use colored pencils to make a drawing in the data table of the absorption spectra you observed.
- Repeat steps 5 and 6 using the green, blue, and yellow colored solutions.

Cleanup and Disposal

- Turn off the light socket and spectrum tube power supplies.
- Wait several minutes to allow the incandescent light bulb and the spectrum tubes to cool.
- Follow your teacher's instructions on how to dispose of the liquids and how to store the light bulb and spectrum tubes.

Analyze and Conclude

- Thinking Critically** How can the existence of spectra help to prove that energy levels in atoms exist?
- Thinking Critically** How can the single electron in a hydrogen atom produce all of the lines found in its emission spectrum?
- Predicting** How can you predict the absorption spectrum of a solution by looking at its color?
- Thinking Critically** How can spectra be used to identify the presence of specific elements in a substance?

Real-World Chemistry

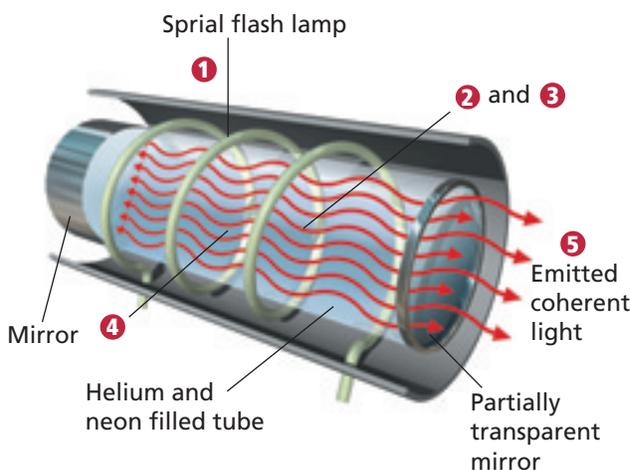
- How can absorption and emission spectra be used by the Hubble space telescope to study the structures of stars or other objects found in deep space?
- The absorption spectrum of chlorophyll *a* indicates strong absorption of red and blue wavelengths. Explain why leaves appear green.

How It Works

Lasers

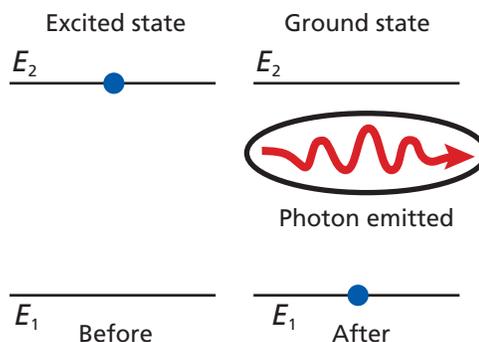
A laser is a device that produces a beam of intense light of a specific wavelength (color). Unlike light from a flashlight, laser light is coherent; that is, it does not spread out as it travels through space. The precise nature of lasers led to their use in pointing and aiming devices, CD players, optical fiber data transmission, and surgery.

- 1 The spiral-wound high-intensity lamp flashes, supplying energy to the helium-neon gas mixture inside the tube. The atoms of the gas absorb the light energy and are raised to an excited energy state.

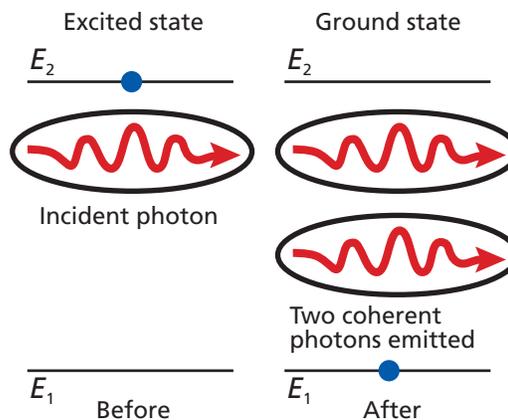


- 4 Photons traveling parallel to the tube are reflected back through the tube by the flat mirrors located at each end. The photons strike additional excited atoms and cause more photons to be released. The intensity of the light in the tube builds.
- 5 Some of the laser's coherent light passes through the partially transparent mirror at one end of the tube and exits the laser. These photons make up the light emitted by the laser.

- 2 The excited atoms begin returning to the ground state, emitting photons in the process. These initial photons travel in all directions.



- 3 The emitted photons hit other excited atoms, causing them to release additional photons. These additional photons are the same wavelength as the photons that struck the excited atoms, and they are coherent (their waves are in sync because they are identical in wavelength and direction).



Thinking Critically

1. Inferring How does the material used in the laser affect the type of light emitted?

2. Relating Cause and Effect Why is one mirror partially transparent?

Summary

5.1 Light and Quantized Energy

- All waves can be described by their wavelength, frequency, amplitude, and speed.
- Light is an electromagnetic wave. In a vacuum, all electromagnetic waves travel at a speed of 3.00×10^8 m/s.
- All electromagnetic waves may be described as both waves and particles. Particles of light are called photons.
- Energy is emitted and absorbed by matter in quanta.
- In contrast to the continuous spectrum produced by white light, an element's atomic emission spectrum consists of a series of fine lines of individual colors.

5.2 Quantum Theory and the Atom

- According to the Bohr model of the atom, hydrogen's atomic emission spectrum results from electrons dropping from higher-energy atomic orbits to lower-energy atomic orbits.
- The de Broglie equation predicts that all moving particles have wave characteristics and relates each particle's wavelength to its mass, its frequency, and Planck's constant.

- The quantum mechanical model of the atom is based on the assumption that electrons are waves.
- The Heisenberg uncertainty principle states that it is not possible to know precisely the velocity and the position of a particle at the same time.
- Electrons occupy three-dimensional regions of space called atomic orbitals. There are four types of orbitals, denoted by the letters s, p, d, and f.

5.3 Electron Configurations

- The arrangement of electrons in an atom is called the atom's electron configuration. Electron configurations are prescribed by three rules: the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- Electrons related to the atom's highest principal energy level are referred to as valence electrons. Valence electrons determine the chemical properties of an element.
- Electron configurations may be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

Key Equations and Relationships

- EM Wave relationship: $c = \lambda\nu$ (p. 119)
- Energy of a quantum: $E_{\text{quantum}} = h\nu$ (p. 123)
- Energy of a photon: $E_{\text{photon}} = h\nu$ (p. 124)
- Energy change of an electron: $\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}}$
 $\Delta E = E_{\text{photon}} = h\nu$ (p. 128)
- de Broglie's equation: $\lambda = \frac{h}{m\nu}$ (p. 130)

Vocabulary

- amplitude (p. 119)
- atomic emission spectrum (p. 125)
- atomic orbital (p. 132)
- aufbau principle (p. 135)
- de Broglie equation (p. 130)
- electromagnetic radiation (p. 118)
- electromagnetic spectrum (p. 120)
- electron configuration (p. 135)
- electron-dot structure (p. 140)
- energy sublevel (p. 133)
- frequency (p. 118)
- ground state (p. 127)
- Heisenberg uncertainty principle (p. 131)
- Hund's rule (p. 136)
- Pauli exclusion principle (p. 136)
- photoelectric effect (p. 123)
- photon (p. 123)
- Planck's constant (p. 123)
- principal energy level (p. 133)
- principal quantum number (p. 132)
- quantum (p. 122)
- quantum mechanical model of the atom (p. 131)
- valence electron (p. 140)
- wavelength (p. 118)

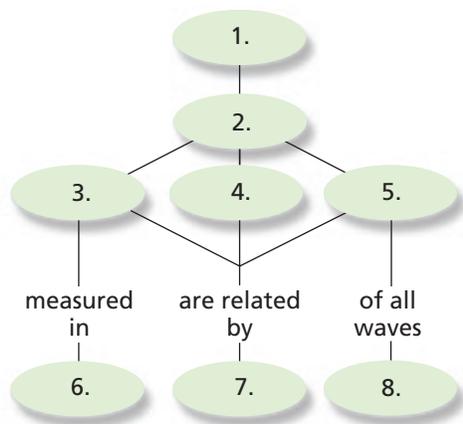


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Concept Mapping

29. Complete the concept map using the following terms: speed, $c = \lambda\nu$, electromagnetic waves, wavelength, characteristic properties, frequency, c , and hertz.



Mastering Concepts

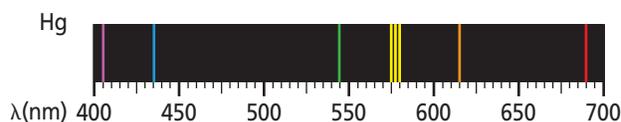
30. Define the following terms.
- frequency (5.1)
 - wavelength (5.1)
 - quantum (5.1)
 - ground state (5.2)
31. Why did scientists consider Rutherford's nuclear model of the atom incomplete? (5.1)
32. Name one type of electromagnetic radiation. (5.1)
33. Explain how the gaseous neon atoms in a neon sign emit light. (5.1)
34. What is a photon? (5.1)
35. What is the photoelectric effect? (5.1)
36. Explain Planck's quantum concept as it relates to energy lost or gained by matter. (5.1)
37. How did Einstein explain the previously unexplainable photoelectric effect? (5.1)
38. Arrange the following types of electromagnetic radiation in order of increasing wavelength. (5.1)
- ultraviolet light
 - microwaves
 - radio waves
 - X rays
39. What is the difference between an atom's ground state and an excited state? (5.2)
40. According to the Bohr model, how do electrons move in atoms? (5.2)
41. What does n designate in Bohr's atomic model? (5.2)
42. Why are you unaware of the wavelengths of moving objects such as automobiles and tennis balls? (5.2)
43. What is the name of the atomic model in which electrons are treated as waves? Who first wrote the electron wave equations that led to this model? (5.2)
44. What is an atomic orbital? (5.2)
45. What is the probability that an electron will be found within an atomic orbital? (5.2)
46. What does n represent in the quantum mechanical model of the atom? (5.2)
47. How many energy sublevels are contained in each of the hydrogen atom's first three energy levels? (5.2)
48. What atomic orbitals are related to a p sublevel? To a d sublevel? (5.2)
49. Which of the following atomic orbital designations are incorrect? (5.2)
- 7f
 - 3f
 - 2d
 - 6p
50. What do the sublevel designations s, p, d, and f specify with respect to the atom's orbitals? (5.2)
51. What do subscripts such as y and xz tell you about atomic orbitals? (5.2)
52. What is the maximum number of electrons an orbital may contain? (5.2)
53. Why is it impossible to know precisely the velocity and position of an electron at the same time? (5.2)
54. What shortcomings caused scientists to finally reject Bohr's model of the atom? (5.2)
55. Describe de Broglie's revolutionary concept involving the characteristics of moving particles. (5.2)
56. How is an orbital's principal quantum number related to the atom's major energy levels? (5.2)
57. Explain the meaning of the aufbau principle as it applies to atoms with many electrons. (5.3)
58. In what sequence do electrons fill the atomic orbitals related to a sublevel? (5.3)
59. Why must the two arrows within a single block of an orbital diagram be written in opposite (up and down) directions? (5.3)
60. How does noble-gas notation shorten the process of writing an element's electron configuration? (5.3)
61. What are valence electrons? How many of a magnesium atom's 12 electrons are valence electrons? (5.3)

- 62.** Light is said to have a dual wave-particle nature. What does this statement mean? (5.3)
- 63.** Describe the difference between a quantum and a photon. (5.3)
- 64.** How many electrons are shown in the electron-dot structures of the following elements? (5.3)
- | | |
|-----------|------------|
| a. carbon | c. calcium |
| b. iodine | d. gallium |
- 74.** A photon has an energy of 1.10×10^{-13} J. What is the photon's wavelength? What type of electromagnetic radiation is it?
- 75.** How long does it take a radio signal from the Voyager spacecraft to reach Earth if the distance between Voyager and Earth is 2.72×10^9 km?
- 76.** If your favorite FM radio station broadcasts at a frequency of 104.5 MHz, what is the wavelength of the station's signal in meters? What is the energy of a photon of the station's electromagnetic signal?

Mastering Problems

Wavelength, Frequency, Speed, and Energy (5.1)

- 65.** What is the wavelength of electromagnetic radiation having a frequency of 5.00×10^{12} Hz? What kind of electromagnetic radiation is this?
- 66.** What is the frequency of electromagnetic radiation having a wavelength of 3.33×10^{-8} m? What type of electromagnetic radiation is this?
- 67.** The laser in a compact disc (CD) player uses light with a wavelength of 780 nm. What is the frequency of this light?
- 68.** What is the speed of an electromagnetic wave having a frequency of 1.33×10^{17} Hz and a wavelength of 2.25 nm?
- 69.** Use **Figure 5-5** to determine each of the following types of radiation.
- | |
|--|
| a. radiation with a frequency of $8.6 \times 10^{11} \text{ s}^{-1}$ |
| b. radiation with a wavelength 4.2 nm |
| c. radiation with a frequency of 5.6 MHz |
| d. radiation that travels at a speed of 3.00×10^8 m/s |
- 70.** What is the energy of a photon of red light having a frequency of 4.48×10^{14} Hz?
- 71.** Mercury's atomic emission spectrum is shown below. Estimate the wavelength of the orange line. What is its frequency? What is the energy of an orange photon emitted by the mercury atom?



- 72.** What is the energy of an ultraviolet photon having a wavelength of 1.18×10^{-8} m?
- 73.** A photon has an energy of 2.93×10^{-25} J. What is its frequency? What type of electromagnetic radiation is the photon?

Electron Configurations (5.3)

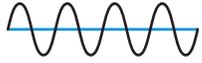
- 77.** List the aufbau sequence of orbitals from 1s to 7p.
- 78.** Write orbital notations and complete electron configurations for atoms of the following elements.
- | |
|--------------|
| a. beryllium |
| b. aluminum |
| c. nitrogen |
| d. sodium |
- 79.** Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols.
- | | |
|-------|-------|
| a. Mn | f. W |
| b. Kr | g. Pb |
| c. P | h. Ra |
| d. Zn | i. Sm |
| e. Zr | j. Bk |
- 80.** What elements are represented by each of the following electron configurations?
- | |
|---|
| a. $1s^2 2s^2 2p^5$ |
| b. $[\text{Ar}] 4s^2$ |
| c. $[\text{Xe}] 6s^2 4f^4$ |
| d. $[\text{Kr}] 5s^2 4d^{10} 5p^4$ |
| e. $[\text{Rn}] 7s^2 5f^{13}$ |
| f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ |
- 81.** Draw electron-dot structures for atoms of each of the following elements.
- | |
|--------------|
| a. carbon |
| b. arsenic |
| c. polonium |
| d. potassium |
| e. barium |
- 82.** An atom of arsenic has how many electron-containing orbitals? How many of the orbitals are completely filled? How many of the orbitals are associated with the atom's $n = 4$ principal energy level?

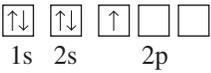
Mixed Review

Sharpen your problem-solving skills by answering the following.

83. What is the frequency of electromagnetic radiation having a wavelength of 1.00 m?
84. What is the maximum number of electrons that can be contained in an atom's orbitals having the following principal quantum numbers?
 a. 3 b. 4 c. 6 d. 7
85. What is the wavelength of light with a frequency of 5.77×10^{14} Hz?
86. Using the waves shown below, identify the wave or waves with the following characteristics.
1. 

2. 

3. 

4. 
- a. longest wavelength c. largest amplitude
 b. greatest frequency d. shortest wavelength
87. How many orientations are possible for the orbitals related to each of the following sublevels?
 a. s b. p c. d d. f
88. Describe the electrons in an atom of nickel in the ground state using the electron configuration notation and the noble-gas notation.
89. Which of the following elements have two electrons in their electron-dot structures: hydrogen, helium, lithium, aluminum, calcium, cobalt, bromine, krypton, and barium?
90. In Bohr's atomic model, what electron orbit transition produces the blue-green line in hydrogen's atomic emission spectrum?
91. A zinc atom contains a total of 18 electrons in its 3s, 3p, and 3d orbitals. Why does its electron-dot structure show only two dots?
92. An X-ray photon has an energy of 3.01×10^{-18} J. What is its frequency and wavelength?
93. Which element has the following orbital diagram?

 1s 2s 2p
94. Which element has the ground-state electron configuration represented by the noble-gas notation $[\text{Rn}]7s^1$?
95. How many photons of infrared radiation having a frequency of 4.88×10^{13} Hz are required to provide an energy of 1.00 J?

Thinking Critically

96. **Comparing and Contrasting** Briefly discuss the difference between an orbit in Bohr's model of the atom and an orbital in the quantum mechanical view of the atom.
97. **Applying Concepts** Scientists use atomic emission spectra to determine the elements in materials of unknown composition. Explain what makes this method possible.
98. **Using Numbers** It takes 8.17×10^{-19} J of energy to remove one electron from a gold surface. What is the maximum wavelength of light capable of causing this effect?
99. **Drawing a Conclusion** The elements aluminum, silicon, gallium, germanium, arsenic, selenium are all used in making various types of semiconductor devices. Write electron configurations and electron-dot structures for atoms of each of these elements. What similarities among the elements' electron configurations do you notice?

Writing in Chemistry

100. In order to make "neon" signs emit a variety of colors, manufacturers often fill the signs with gases other than neon. Research the use of gases in neon signs and specify the colors produced by the gases.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

101. Round 20.561 20 g to three significant figures. (Chapter 2)
102. Identify each of the following as either chemical or physical properties of the substance. (Chapter 3)
 a. mercury is a liquid at room temperature
 b. sucrose is a white, crystalline solid
 c. iron rusts when exposed to moist air
 d. paper burns when ignited
103. Identify each of the following as a pure substance or a mixture. (Chapter 3)
 a. distilled water d. diamond
 b. orange juice with pulp e. milk
 c. smog f. copper metal
104. An atom of gadolinium has an atomic number of 64 and a mass number of 153. How many electrons, protons, and neutrons does it contain? (Chapter 4)

Use these questions and the test-taking tip to prepare for your standardized test.

- Cosmic rays are high-energy radiation from outer space. What is the frequency of a cosmic ray that has a wavelength of 2.67×10^{-13} m when it reaches Earth? (The speed of light is 3.00×10^8 m/s.)
 - $8.90 \times 10^{-22} \text{ s}^{-1}$
 - $3.75 \times 10^{12} \text{ s}^{-1}$
 - $8.01 \times 10^{-5} \text{ s}^{-1}$
 - $1.12 \times 10^{21} \text{ s}^{-1}$
- Wavelengths of light between 5.75×10^{-9} m and 5.85×10^{-9} m appear yellow to the human eye. What is the energy of a photon of yellow light having a frequency of $5.45 \times 10^{16} \text{ s}^{-1}$? (Planck's constant is $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$.)
 - $3.61 \times 10^{-17} \text{ J}$
 - $1.22 \times 10^{-50} \text{ J}$
 - $8.23 \times 10^{49} \text{ J}$
 - $3.81 \times 10^{-24} \text{ J}$

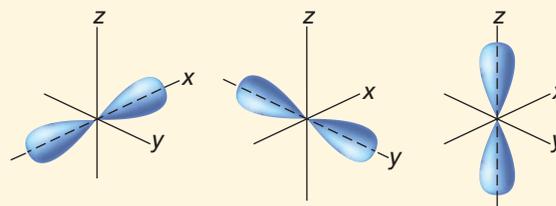
Interpreting Charts Use the periodic table and the chart below to answer questions 3–6.

Electron Configurations for Selected Transition Metals			
Element	Symbol	Atomic number	Electron configuration
Vanadium	V	23	$[\text{Ar}]4s^23d^3$
Yttrium	Y	39	$[\text{Kr}]5s^24d^1$
_____	___	___	$[\text{Xe}]6s^24f^{14}5d^6$
Scandium	Sc	21	$[\text{Ar}]4s^23d^1$
Cadmium	Cd	48	_____

- Using noble-gas notation, the ground-state electron configuration of Cd is _____.
 - $[\text{Kr}]4d^{10}4f^2$
 - $[\text{Ar}]4s^23d^{10}$
 - $[\text{Kr}]5s^24d^{10}$
 - $[\text{Xe}]5s^24d^{10}$
- The element that has the ground-state electron configuration $[\text{Xe}]6s^24f^{14}5d^6$ is _____.
 - La
 - Ti
 - W
 - Os
- The complete electron configuration of a scandium atom is _____.
 - $1s^22s^22p^63s^23p^64s^23d^1$
 - $1s^22s^22p^73s^23p^74s^23d^1$
 - $1s^22s^22p^53s^23p^54s^23d^1$
 - $1s^22s^12p^73s^13p^74s^23d^1$

- Which of the following is the correct orbital diagram for the third and fourth principal energy levels of vanadium?
 -
 -
 -
 -
- Which of the following orbitals has the highest energy?
 - 4f
 - 5p
 - 6s
 - 3d
- What is the electron-dot structure for indium?
 - $\cdot\text{In}$
 - $\cdot\text{In}\cdot$
 - $\cdot\text{In}\cdot$
 - $\cdot\text{In}\cdot$

- The picture below shows all of the orbitals related to one type of sublevel. The type of sublevel to which these orbitals belong is _____.



- s
 - p
 - d
 - f
- What is the maximum number of electrons related to the fifth principal energy level of an atom?
 - 10
 - 20
 - 25
 - 50

Do Some Reconnaissance Find out what the conditions will be for taking the test. Is it timed or untimed? Can you eat a snack at the break? Can you use a calculator or other tools? Will those tools be provided? Will mathematical constants be given? Know these things in advance so that you can practice taking tests under the same conditions.

The Periodic Table and Periodic Law

What You'll Learn

- ▶ You will explain why elements in a group have similar properties.
- ▶ You will relate the group and period trends seen in the periodic table to the electron configuration of atoms.
- ▶ You will identify the s-, p-, d-, and f-blocks of the periodic table.

Why It's Important

The periodic table is the single most powerful chemistry reference tool available to you. Understanding its organization and interpreting its data will greatly aid you in your study of chemistry.

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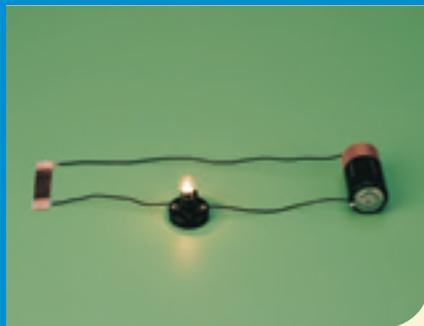


Visit the Chemistry Web site at science.glencoe.com to find links about the periodic table.

The phases of the moon and the cycle of ocean tides are both periodic events, that is, they repeat in a regular manner.



DISCOVERY LAB



Materials

tape
samples of copper
light socket with bulb, wires,
and battery

Versatile Metals

A variety of processes can be used to shape metals into different forms. Because of their physical properties, metals are used in a wide range of applications.

Safety Precautions



Be careful when bending the copper samples, as they may have sharp edges.

Procedure

1. Observe the different types of copper metal that your teacher gives you. Write down as many observations as you can about each of the copper samples.
2. Try gently bending each copper sample (do not break the samples). Record your observations.
3. Connect each copper sample to the circuit as shown in the photo. Record your observations.

Analysis

What properties of copper are similar in all of the samples? How do the samples of copper differ? List several common applications of copper. What properties make metals such as copper so versatile?

Section

6.1

Development of the Modern Periodic Table

Objectives

- **Trace** the development and **identify** key features of the periodic table.

Vocabulary

periodic law
group
period
representative element
transition element
metal
alkali metal
alkaline earth metal
transition metal
inner transition metal
nonmetal
halogen
noble gas
metalloid

You have already learned much in your study of chemistry. Wouldn't it be nice if you could easily organize the chemistry knowledge you are acquiring? You can, with the help of the *periodic* table. It is called a periodic table because, much like the phases of the moon, one of which is shown in the chapter opening photo, the properties of the elements in the table repeat in a periodic way. The periodic table will be an invaluable tool as you continue this course in chemistry. However, before you learn about the modern periodic table, a recounting of the history behind the table's development will help you understand its significance.

History of the Periodic Table's Development

In the late 1790s, French scientist Antoine Lavoisier compiled a list of elements known at the time. The list contained 23 elements. Many of these elements, such as silver, gold, carbon, and oxygen, were known since prehistoric times. The 1800s brought many changes to the world, including an explosion in the number of known elements. The advent of electricity, which was used to break compounds down into their component elements, and the development of the spectrometer, which was used to identify the newly isolated elements, played major roles in the advancement of chemistry. So did the industrial revolution



Figure 6-1

A resident of London, England invented the word smog to describe the city's filthy air, a combination of smoke and natural fog. The quality of London's air became so poor that in 1952 about 4000 Londoners died during a four-day period. This incident led to the passage of England's Clean Air Act in 1956.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Astronomy CONNECTION

The element technetium does not occur naturally on Earth. It has been found in stars. Astronomers analyze the chemical composition of stellar matter by using an instrument called a spectroscope, which separates the light from a star into individual colors, much as a prism does. Although each star has a unique composition of elements, all stars are composed mainly of the gases hydrogen and helium. The Sun, for example, is estimated to be about 70 percent hydrogen and 28 percent helium. A tiny fraction of a star's mass may come from heavier elements such as oxygen, carbon, nitrogen, calcium, or sodium. Two percent of our Sun's mass comes from these heavier elements.

of the mid-1800s, which led to the development of many new chemistry-based industries, such as the manufacture of petrochemicals, soaps, dyes, and fertilizers. By 1870, there were approximately 70 known elements—almost triple the number known in Lavoisier's time. As you can see in **Figure 6-1**, the industrial revolution also created problems, such as increased chemical pollution.

Along with the discovery of new elements came volumes of new scientific data related to the elements and their compounds. Chemists of the time were overwhelmed with learning the properties of so many new elements and compounds. What chemists needed was a tool for organizing the many facts associated with the elements. A significant step toward this goal came in 1860, when chemists agreed upon a method for accurately determining the atomic masses of the elements. Until this time, different chemists used different mass values in their work, making the results of one chemist's work hard to reproduce by another. With newly agreed upon atomic masses for the elements, the search for relationships between atomic mass and elemental properties began in earnest.

John Newlands In 1864, English chemist John Newlands (1837–1898), who is shown in **Figure 6-2**, proposed an organization scheme for the elements. Newlands noticed that when the elements were arranged by increasing atomic mass, their properties repeated every eighth element. In other words, the first and eighth elements had similar properties, the second and ninth elements had similar properties, and so on. A pattern such as this is called periodic because it repeats in a specific manner. Newlands named the periodic relationship that he observed in chemical properties the law of octaves, because an octave is a group of musical notes that repeats every eighth tone. **Figure 6-2** also shows how Newlands organized the first 14 “known” elements (as of the mid-1860s). If you compare Newlands's arrangement of the elements with the modern periodic table on the inside back cover of your textbook, you'll see that some of his rows correspond to columns on the modern periodic table. Acceptance of the law of octaves was hampered because the law did not work for all of the known elements. Also, unfortunately for Newlands, the use of the word octave was harshly criticized by fellow scientists who thought that the musical analogy was unscientific. While Newlands's law was not generally accepted, the passage of a few years would show that he was basically correct; the properties of elements do repeat in a periodic way.

Meyer, Mendeleev, and Moseley In 1869, German chemist Lothar Meyer (1830–1895) and Russian chemist Dmitri Mendeleev (1834–1907) each demonstrated a connection between atomic mass and elemental properties. Mendeleev, however, is generally given more credit than Meyer because he published his organization scheme first and went on to better demonstrate its usefulness. Like Newlands several years earlier, Mendeleev noticed that when the elements were ordered by increasing atomic mass, there was a repetition, or periodic pattern, in their properties. By arranging the elements in order of increasing atomic mass into columns with similar properties, Mendeleev organized the elements into the first periodic table. Mendeleev and part of his periodic table are shown in **Figure 6-3**. Part of the reason Mendeleev's table was widely accepted was that he predicted the existence and properties of undiscovered elements. Mendeleev left blank spaces in the table where he thought the undiscovered elements should go. By noting trends in the properties of known elements, he was able to predict the properties of the yet-to-be discovered elements scandium, gallium, and germanium.



Elements with similar properties
are in the same row

1 octave	A	H 1	A	F 8	— and so on →
	B	Li 2	B	Na 9	→
	C	G 3	C	Mg 10	→
	D	Bo 4	D	Al 11	→
	E	C 5	E	Si 12	→
	F	N 6	F	P 13	→
	G	O 7	G	S 14	→

Figure 6-2

John Newlands noticed that the properties of elements repeated in a manner similar to an octave on a musical scale (A, B, C, D, E, F, G, A, and so on). While there are some similarities between the law of octaves and the modern periodic table, there also are significant differences. You'll notice that some of the chemical symbols do not match. For example, beryllium (Be) was also known as glucinum (G). What similarities and differences can you identify?

Mendeleev's table, however, was not completely correct. After several new elements were discovered and atomic masses of the known elements were more accurately determined, it became apparent that several elements in his table were not in the correct order. Arranging the elements by mass resulted in several elements being placed in groups of elements with differing properties. The reason for this problem was determined in 1913 by English chemist Henry Moseley. As you may recall from Chapter 4, Moseley discovered that atoms of each element contain a unique number of protons in their nuclei—the number of protons being equal to the atom's atomic number. By arranging the elements in order of increasing atomic number instead of increasing atomic mass, as Mendeleev had done, the problems with the order of the elements in the periodic table were solved. Moseley's arrangement of elements by atomic number resulted in a clear periodic pattern of properties. The statement that there is a periodic repetition of chemical and physical properties of the elements when they are arranged by increasing atomic number is called the **periodic law**.

Figure 6-3

Dmitri Mendeleev produced the first useful and widely accepted periodic table. The monument shown on the right is located in St. Petersburg, Russia, and shows an early version of Mendeleev's periodic table. The blank areas on the table show the positions of elements that had not yet been discovered.





Figure 6-4

The modern periodic table arranges the elements by increasing atomic number. The columns are known as groups or families, and the rows are known as periods.

The periodic table became a significant tool for chemists working in the new industries created during the industrial revolution. The table brought order to seemingly unrelated facts. You, too, will find the periodic table a valuable tool. Among other things, it is a useful reference for understanding and predicting the properties of elements and for organizing your knowledge of atomic structure. Do the **problem-solving LAB** on the next page to see how the periodic law can be used to predict unknown elemental properties.

The Modern Periodic Table

The modern periodic table is shown in **Figure 6-4** and on the inside back cover of your textbook. A larger, two-page version of the table appears in **Figure 6-7** on pages 156-157. The table consists of boxes, each containing an element name, symbol, atomic number, and atomic mass. A typical box from the table is shown in **Figure 6-5**. The boxes are arranged in order of increasing atomic number into a series of columns, called **groups** or families, and rows, called **periods**. Beginning with hydrogen in period 1, there are a total of seven periods. Each group is numbered 1 through 8, followed by the letter A or B. For example, scandium (Sc) is in the third column from the left, group 3B. What group is oxygen in? What period contains potassium and calcium? The groups designated with an A (1A through 8A) are often referred to as the main group, or **representative elements** because they possess a wide range of chemical and physical properties. The groups designated with a B (1B through 8B) are referred to as the **transition elements**. A more recent numbering system, which uses the numbers 1 through 18, also appears above each group. The number-and-letter system is used throughout this textbook.

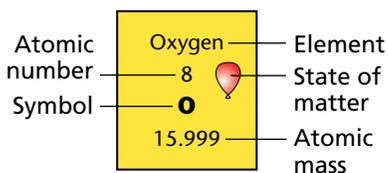


Figure 6-5

A typical box from the periodic table contains important information about an element.

Classifying the elements There are three main classifications for the elements—metals, nonmetals, and metalloids. **Metals** are elements that are generally shiny when smooth and clean, solid at room temperature, and good conductors of heat and electricity. Most metals also are ductile and malleable, meaning that they can be pounded into thin sheets and drawn into wires, respectively. **Figure 6-6** shows several applications that make use of the physical properties of metals.

Most group A elements and all group B elements are metals. If you look at boron (B) in column 3A, you see a heavy stair-step line that zigzags down to astatine (At) at the bottom of group 7A. This stair-step line serves as a visual divider between the metals and the nonmetals on the table. Metals are represented by the light blue boxes in **Figure 6-7**. Except for hydrogen, all of the elements on the left side of the table are metals. The group 1A elements (except for hydrogen) are known as the **alkali metals**; the group 2A elements are known as the **alkaline earth metals**. Both the alkali metals and the alkaline earth metals are chemically reactive, with the alkali metals being the more reactive of the two groups.



Figure 6-6

Metals are used in a wide variety of applications. The excellent electrical conductivity of metals such as copper, makes them a good choice for transmitting electrical power. Ductility and malleability allow metals to be formed into coins, tools, fasteners, and wires.

problem-solving LAB

Francium—solid, liquid or gas?

Predicting Of the first 101 elements, francium is the least stable. Its most stable isotope has a half-life of just 22 minutes! Use your knowledge about the properties of other alkali metals to predict some of francium's properties.

Analysis

In the spirit of Dimitri Mendeleev's prediction of the properties of several, as of then, undiscovered elements, use the given information about the known properties of the alkali metals to devise a method for determining the corresponding property of francium.

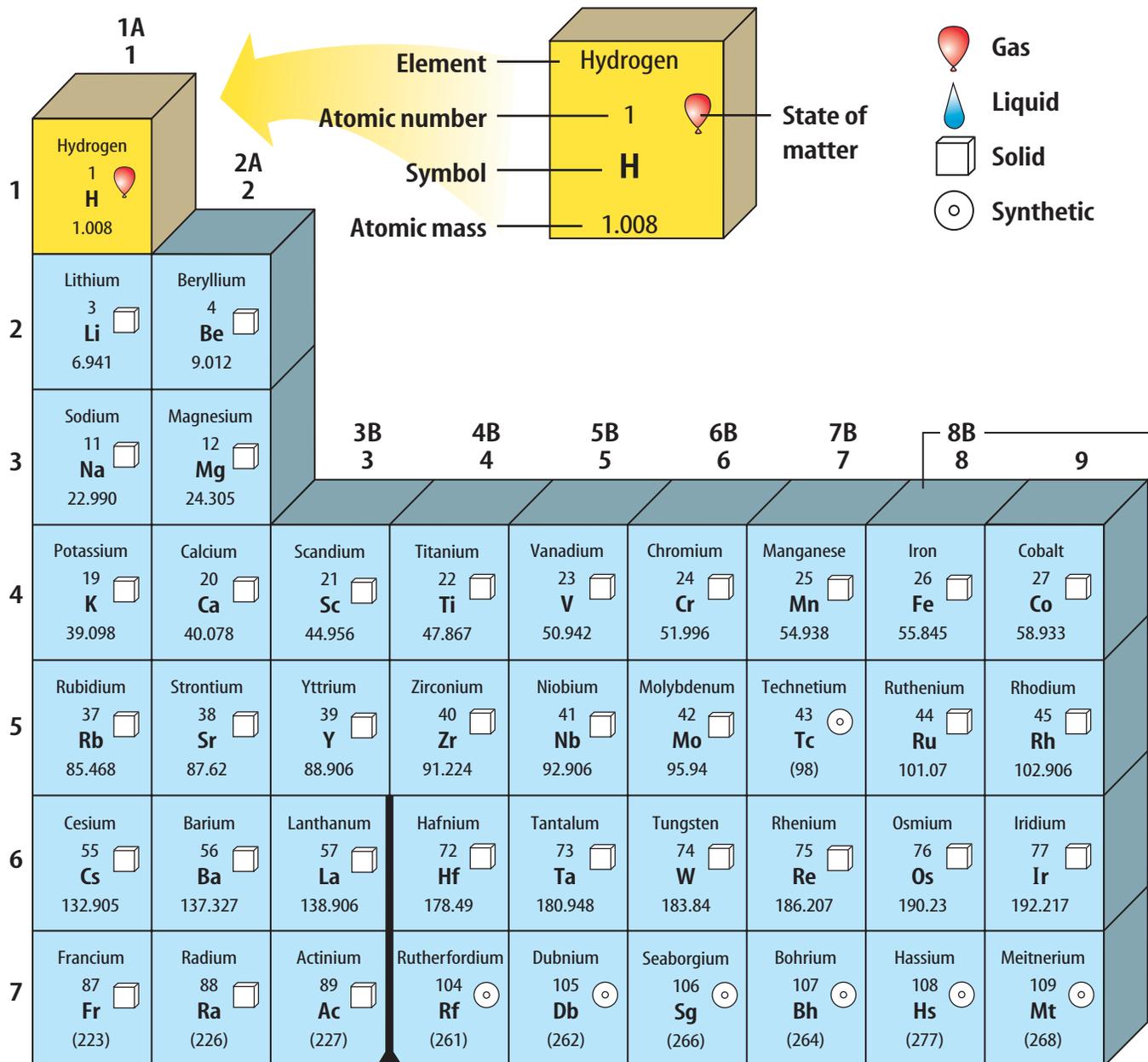
Thinking Critically

- Using the periodic law as a guide, devise an approach that clearly displays the trends for each of the properties given in the table and allows you to extrapolate a value for francium.
- Predict whether francium is a solid, liquid, or gas. How can you support your prediction?
- Which of the given columns of data presents the greatest possible error in making a prediction? Explain.
- Currently, scientists can produce about one million francium atoms per second. Explain why this is still not enough to make basic measurements such as density or melting point.

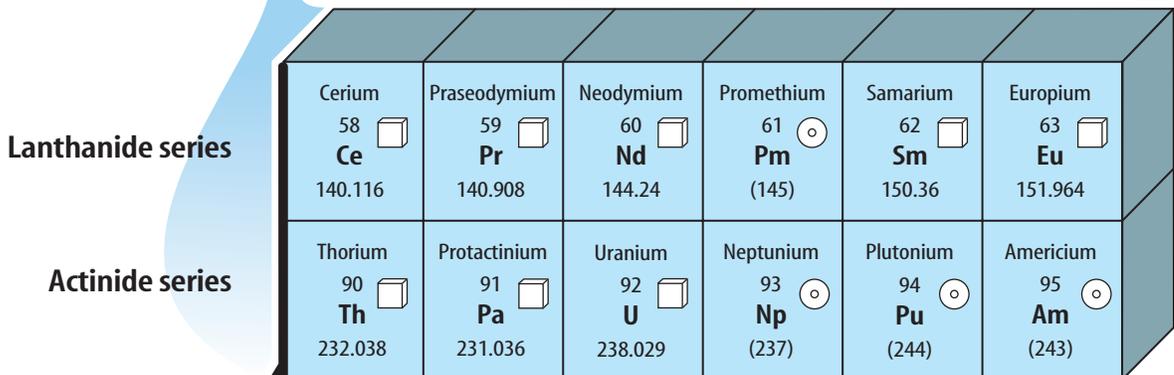
Alkali Metals Data			
Element	Melting point (°C)	Boiling point (°C)	Radius (pm)
lithium	180.5	1347	152
sodium	97.8	897	186
potassium	63.3	766	227
rubidium	39.31	688	248
cesium	28.4	674.8	265
francium	?	?	?

Figure 6-7

PERIODIC TABLE OF THE ELEMENTS



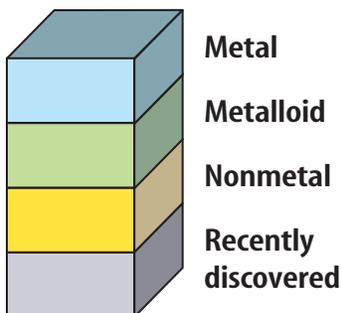
The number in parentheses is the mass number of the longest lived isotope for that element.



CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find updates on the periodic table.



			3A 13	4A 14	5A 15	6A 16	7A 17	8A 18			
			Boron 5 B 10.811	Carbon 6 C 12.011	Nitrogen 7 N 14.007	Oxygen 8 O 15.999	Fluorine 9 F 18.998	Helium 2 He 4.003			
			Aluminum 13 Al 26.982	Silicon 14 Si 28.086	Phosphorus 15 P 30.974	Sulfur 16 S 32.065	Chlorine 17 Cl 35.453	Neon 10 Ne 20.180			
10	1B 11	2B 12	Nickel 28 Ni 58.693	Copper 29 Cu 63.546	Zinc 30 Zn 65.39	Gallium 31 Ga 69.723	Germanium 32 Ge 72.64	Arsenic 33 As 74.922	Selenium 34 Se 78.96	Bromine 35 Br 79.904	Krypton 36 Kr 83.80
			Palladium 46 Pd 106.42	Silver 47 Ag 107.868	Cadmium 48 Cd 112.411	Indium 49 In 114.818	Tin 50 Sn 118.710	Antimony 51 Sb 121.760	Tellurium 52 Te 127.60	Iodine 53 I 126.904	Xenon 54 Xe 131.293
			Platinum 78 Pt 195.078	Gold 79 Au 196.967	Mercury 80 Hg 200.59	Thallium 81 Tl 204.383	Lead 82 Pb 207.2	Bismuth 83 Bi 208.980	Polonium 84 Po (209)	Astatine 85 At (210)	Radon 86 Rn (222)
			Ununnilium * 110 Uun (281)	Unununium * 111 Uuu (272)	Ununbium * 112 Uub (285)		Ununquadium * 114 Uuq (289)		Ununhexium * 116 Uuh (289)		Ununoctium * 118 Uuo (293)

* Names not officially assigned. Discovery of elements 114, 116, and 118 recently reported. Further information not yet available.

Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.925	Dysprosium 66 Dy 162.50	Holmium 67 Ho 164.930	Erbium 68 Er 167.259	Thulium 69 Tm 168.934	Ytterbium 70 Yb 173.04	Lutetium 71 Lu 174.967
Curium 96 Cm (247)	Berkelium 97 Bk (247)	Californium 98 Cf (251)	Einsteinium 99 Es (252)	Fermium 100 Fm (257)	Mendelevium 101 Md (258)	Nobelium 102 No (259)	Lawrencium 103 Lr (262)



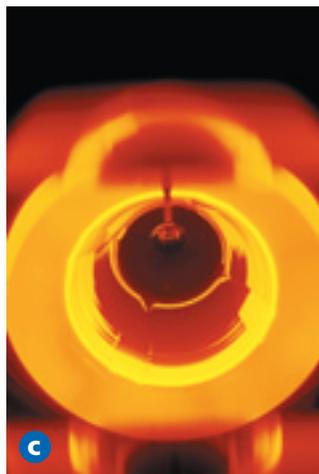


Figure 6-8

a A mountain climber breathes from a container of compressed oxygen gas, a nonmetal. **b** This Persian brass bowl contains inlays of the transition metals silver and gold. **c** Silicon crystals, a metalloid, are grown in an inert atmosphere of argon, a nonmetal. The crystals are used in the manufacture of computer chips.

The group B elements, or transition elements, are divided into **transition metals** and **inner transition metals**. The two sets of inner transition metals, known as the lanthanide and actinide series, are located along the bottom of the periodic table. The rest of the group B elements make up the transition metals. Elements from the lanthanide series are used extensively as phosphors, substances that emit light when struck by electrons. The **How It Works** at the end of the chapter explains more about phosphors and how images are formed on a television screen.

Nonmetals occupy the upper right side of the periodic table. They are represented by the yellow boxes in **Figure 6-7**. **Nonmetals** are elements that are generally gases or brittle, dull-looking solids. They are poor conductors of heat and electricity. The only nonmetal that is a liquid at room temperature is bromine (Br). The highly reactive group 7A elements are known as **halogens**, and the extremely unreactive group 8A elements are commonly called the **noble gases**.

Examine the elements in green boxes bordering the stair-step line in **Figure 6-7**. These elements are called metalloids, or semimetals. **Metalloids** are elements with physical and chemical properties of both metals and nonmetals. Silicon and germanium are two of the most important metalloids, as they are used extensively in computer chips and solar cells. Applications that make use of the properties of nonmetals, transition metals, and metalloids are shown in **Figure 6-8**. Do the **CHEMLAB** at the end of this chapter to observe trends among various elements.

This introduction to the periodic table only touches the surface of its usefulness. In the next section, you will discover how an element's electron configuration, which you learned about in Chapter 5, is related to its position on the periodic table.

Section 6.1 Assessment

- Describe the development of the modern periodic table. Include contributions made by Newlands, Mendeleev, and Moseley.
- Sketch a simplified version of the periodic table and indicate the location of groups, periods, metals, nonmetals, and metalloids.
- Describe the general characteristics of metals, nonmetals, and metalloids.
- Identify each of the following as a representative element or a transition element.

a. lithium (Li)	c. promethium (Pm)
b. platinum (Pt)	d. carbon (C)
- Thinking Critically** For each of the given elements, list two other elements with similar chemical properties.
 - iodine (I)
 - barium (Ba)
 - iron (Fe)
- Interpreting Data** An unknown element has chemical behavior similar to that of silicon (Si) and lead (Pb). The unknown element has a mass greater than that of sulfur (S), but less than that of cadmium (Cd). Use the periodic table to determine the identity of the unknown element.

In Chapter 5, you learned how to write the electron configuration for an atom. This is an important skill because the electron configuration determines the chemical properties of the element. However, the process of writing out electron configurations using the aufbau diagram can be tedious. Fortunately, by noting an atom's position on the periodic table, you can determine its electron configuration and its number of valence electrons.

Organizing the Elements by Electron Configuration

Take a look at the electron configurations for the group 1A elements listed below. These elements comprise the first four periods of group 1A.

Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s^22s^1$	$[\text{He}]2s^1$
Period 3	sodium	$1s^22s^22p^63s^1$	$[\text{Ne}]3s^1$
Period 4	potassium	$1s^22s^22p^63s^23p^64s^1$	$[\text{Ar}]4s^1$

What do the four configurations have in common? The answer is that they all have a single electron in their outermost energy level.

Valence electrons Recall from Chapter 5 that electrons in the highest principal energy level of an atom are called valence electrons. Each of the group 1A elements has one electron in its highest energy level; thus, each element has one valence electron. This is no coincidence. The group 1A elements have similar chemical properties because they all have the same number of valence electrons. This is one of the most important relationships in chemistry; *atoms in the same group have similar chemical properties because they have the same number of valence electrons*. Each group 1A element has a valence electron configuration of s^1 . Likewise, each group 2A element has a valence electron configuration of s^2 . Each column of group A elements on the periodic table has its own unique valence electron configuration.

Valence electrons and period The energy level of an element's valence electrons indicates the period on the periodic table in which it is found. For example, lithium's valence electron is in the second energy level and lithium is found in period 2. Now look at gallium, with its electron configuration of $[\text{Ar}]4s^23d^{10}4p^1$. Gallium's valence electrons are in the fourth energy level, and gallium is found in the fourth period. What is the electron configuration for the group 1A element in the sixth period?

Valence electrons and group number A representative element's group number and the number of valence electrons it contains also are related. Group 1A elements have one valence electron, group 2A elements have two valence electrons, and so on. There are several exceptions to this rule, however. The noble gases in group 8A each have eight valence electrons, with the exception of helium, which has only two valence electrons. Also, the group number rule applies only to the representative elements (the group A elements). See **Figure 6-9** on the next page. The electron-dot structures you learned in Chapter 5 illustrate the connection between group number and number of valence electrons.

Objectives

- **Explain** why elements in the same group have similar properties.
- **Identify** the four blocks of the periodic table based on electron configuration.

Medical Lab Technician

Would you like to analyze blood and tissue samples? How about determining the chemical content of body fluids? If so, you might enjoy being a medical lab technician.

Medical or clinical lab technicians work in large hospitals or independent labs. Under the direction of a technologist, they prepare specimens, conduct tests, and operate computerized analyzers. Technicians need to pay close attention to detail, have good judgement, and be skilled in using computers.

The s-, p-, d-, and f-Block Elements

The periodic table has columns and rows of varying sizes. The reason behind the table's odd shape becomes clear if it is divided into sections, or blocks, representing the atom's energy sublevel being filled with valence electrons. Because there are four different energy sublevels (s, p, d, and f), the periodic table is divided into four distinct blocks, as shown in **Figure 6-10**.

s-block elements The s-block consists of groups 1A and 2A, and the elements hydrogen and helium. In this block, the valence electrons, represented in **Figure 6-9**, occupy only s orbitals. Group 1A elements have partially filled s orbitals containing one valence electron and electron configurations ending in s^1 . Group 2A elements have completely filled s orbitals containing two valence electrons and electron configurations ending in s^2 . Because s orbitals hold a maximum of two electrons, the s-block portion of the periodic table spans two groups.

p-block elements After the s sublevel is filled, the valence electrons, represented in **Figure 6-9**, next occupy the p sublevel and its three p orbitals. The p-block of the periodic table, comprised of groups 3A through 8A, contains elements with filled or partially filled p orbitals. Why are there no p-block elements in period 1? The answer is that the p sublevel does not exist for the first principal energy level ($n = 1$). Thus, the first p-block element is boron (B), in the second period. The p-block spans six groups on the periodic table because the three p orbitals can hold a maximum of six electrons. Together, the s- and p-blocks comprise the representative, or group A, elements.

The group 8A, or noble gas, elements are unique members of the p-block because of their incredible stability. Noble gas atoms are so stable that they undergo virtually no chemical reactions. The reason for their stability lies in their electron configurations. Look at the electron configurations of the first four noble gas elements shown in **Table 6-1**. Notice that *both* the s and p orbitals corresponding to the period's principal energy level are *completely filled*. This arrangement of electrons results in an unusually stable atomic structure. You soon will learn that this stable configuration plays an important role in the formation of ions and chemical bonds.

d-block elements The d-block contains the transition metals and is the largest of the blocks. Although there are a number of exceptions, d-block elements are characterized by a filled outermost s orbital of energy level n , and filled or partially filled d orbitals of energy level $n - 1$. As you move across the period, electrons fill the d orbitals. For example, scandium (Sc), the first d-block element, has an electron configuration of $[\text{Ar}]4s^23d^1$. Titanium, the next element on the table, has an electron configuration of $[\text{Ar}]4s^23d^2$. Note that titanium's filled outermost s orbital has an energy level of $n = 4$, while the partially filled d orbital has an energy level of $n - 1$, or 3. The five d orbitals can hold a total of ten electrons; thus, the d-block spans ten groups on the periodic table.

Figure 6-9

The electron-dot structures of most of the representative elements are shown here. The number of valence electrons is the same for all members of a group. For the group A elements, an atom's number of valence electrons is equal to its group number (in the 1A, 2A, . . . numbering system).

	1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1	H·							He:
2	Li·	Be·	·B·	·C·	·N·	·O·	·F·	·Ne:
3	Na·	Mg·	·Al·	·Si·	·P·	·S·	·Cl·	·Ar:
4	K·	Ca·	·Ga·	·Ge·	·As·	·Se·	·Br·	·Kr:
5	Rb·	Sr·	·In·	·Sn·	·Sb·	·Te·	·I·	·Xe:
6	Cs·	Ba·	·Tl·	·Pb·	·Bi·	·Po·		·Rn:

Table 6-1

Electron Configurations of Helium, Neon, Argon, and Krypton				
Period	Principal energy level	Element	Electron configuration	Electron dot structure
1	$n = 1$	helium	$1s^2$	He:
2	$n = 2$	neon	$[\text{He}]2s^22p^6$: $\ddot{\text{Ne}}$:
3	$n = 3$	argon	$[\text{Ne}]3s^23p^6$: $\ddot{\text{Ar}}$:
4	$n = 4$	krypton	$[\text{Ar}]4s^23d^{10}4p^6$: $\ddot{\text{Kr}}$:

f-block elements The f-block contains the inner transition metals. The f-block elements are characterized by a filled, or partially filled outermost s orbital, and filled or partially filled 4f and 5f orbitals. The electrons of the f sublevel do not fill their orbitals in a predictable manner. Because there are seven f orbitals holding up to a maximum of 14 electrons, the f-block spans 14 columns of the periodic table.

Thus, the s-, p-, d-, and f-blocks determine the shape of the periodic table. As you proceed down through the periods, the principal energy level increases, as does the number of energy sublevels containing electrons. Period 1 contains only s-block elements, periods 2 and 3 contain both s- and p-block elements, periods 4 and 5 contain s-, p-, and d-block elements, and periods 6 and 7 contain s-, p-, d-, and f-block elements.

Figure 6-10

Although electrons fill the orbitals of s- and p-block elements in a predictable manner, there are a number of exceptions in the d- and f-block elements. What is the relationship between the maximum number of electrons an energy sublevel can hold and the size of that block on the diagram?

s block		p block						d block										f block																									
s^1	s^2	p^1	p^2	p^3	p^4	p^5	p^6																																				
1 H		5 B	6 C	7 N	8 O	9 F	10 Ne																																				
3 Li	4 Be	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar																																				
11 Na	12 Mg	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr																								
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe																										
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn																										
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub																																
			58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu																											
			90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr																											

EXAMPLE PROBLEM 6-1



Strontium-containing compounds are used to produce the bright red seen in these road flares.

Electron Configuration and the Periodic Table

Strontium has an electron configuration of $[\text{Kr}]5s^2$. Without using the periodic table, determine the group, period, and block in which strontium is located on the periodic table.

1. Analyze the Problem

You are given the electron configuration of strontium. The energy level of the valence electrons can be used to determine the period in which strontium is located. The electron configuration of the valence electrons can be used to determine the group and the block in which strontium is located.

2. Solve for the Unknown

Group The valence electron configuration of s^2 indicates that strontium is in group 2A. All group 2A elements have the s^2 configuration.

Period The 5 in $5s^2$ indicates that strontium is in period 5.

Block The s^2 indicates that strontium's valence electrons fill the s sublevel. Thus, strontium is in the s-block.

3. Evaluate the Answer

The relationships among electron configuration and position on the periodic table have been correctly applied. The given information identifies a unique position on the table, as it must.



For more practice with electron configuration problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.
a. $[\text{Ne}]3s^2$ b. $[\text{He}]2s^2$ c. $[\text{Kr}]5s^24d^{10}5p^5$
- Write the electron configuration of the element fitting each of the following descriptions.
 - The group 2A element in the fourth period
 - The noble gas in the fifth period
 - The group 2B element in the fourth period
 - The group 6A element in the second period
- What are the symbols for the elements with the following valence electron configurations?
 - s^2d^1 b. s^2p^3 c. s^2p^6

Section 6.2 Assessment

- Explain why elements in the same group on the periodic table have similar chemical properties.
- Given each of the following valence electron configurations, determine which block of the periodic table the element is in.
 - s^2p^4 b. s^1 c. s^2d^1 d. s^2p^1
- Describe how each of the following are related.
 - Group number and number of valence electrons for representative elements
 - Principal energy level of valence electrons and period number
- Without using the periodic table, determine the group, period, and block of an atom with an electron configuration of $[\text{Ne}]3s^23p^4$.
- Thinking Critically** A gaseous element is a poor conductor of heat and electricity, and is extremely nonreactive. Is the element likely to be a metal, nonmetal, or metalloid? Where would the element be located on the periodic table? Explain.
- Formulating Models** Make a simplified sketch of the periodic table and label the s-, p-, d-, and f-blocks.

Many properties of the elements tend to change in a predictable way, known as a trend, as you move across a period or down a group. You will explore several periodic trends in this section. Do the **miniLAB** on the next page to explore several properties that behave periodically.

Atomic Radius

The electron cloud surrounding a nucleus is based on probability and does not have a clearly defined edge. It is true that the outer limit of an electron cloud is defined as the spherical surface within which there is a 90% probability of finding an electron. However, this surface does not exist in a physical way, as the outer surface of a golf ball does. Atomic size is defined by how closely an atom lies to a neighboring atom. Because the nature of the neighboring atom can vary from one substance to another, the size of the atom itself also tends to vary somewhat from substance to substance.

For metals such as sodium, the atomic radius is defined as half the distance between adjacent nuclei in a crystal of the element. See **Figure 6-11a**. For elements that commonly occur as molecules, such as many nonmetals, the

Objectives

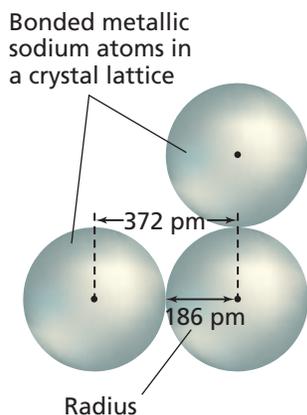
- **Compare** period and group trends of several properties.
- **Relate** period and group trends in atomic radii to electron configuration.

Vocabulary

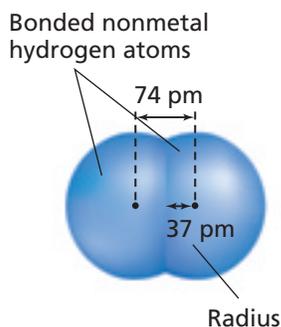
ion
ionization energy
octet rule
electronegativity

Figure 6-11

The table gives atomic radii of the representative elements.



- a** The radius of a metal atom in a metallic crystal is one-half the distance between two adjacent atoms in the crystal.



- b** The radius of a nonmetal atom is often determined from a diatomic molecule of an element.

	1A	2A	3A	4A	5A	6A	7A	8A
1	H 37							He 31
2	Li 152	Be 112	B 85	C 77	N 75	O 73	F 72	Ne 71
3	Na 186	Mg 160	Al 143	Si 118	P 110	S 103	Cl 100	Ar 98
4	K 227	Ca 197	Ga 135	Ge 122	As 120	Se 119	Br 114	Kr 112
5	Rb 248	Sr 215	In 167	Sn 140	Sb 140	Te 142	I 133	Xe 131
6	Cs 265	Ba 222	Tl 170	Pb 146	Bi 150	Po 168	At 140	Rn 140

Chemical symbol
Atomic radius
Relative size

- c** The atomic radii of the representative elements are given in picometers (1×10^{-12} meters) and their relative sizes are shown. The radii for the transition metals have been omitted because they exhibit many exceptions to the general trends shown here. What causes the increase in radii as you move down a group?

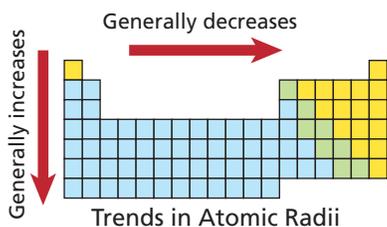


Figure 6-12

This small table provides a summary of the general trends in atomic radii.

atomic radius is defined as half the distance between nuclei of identical atoms that are chemically bonded together. The atomic radius of a nonmetal diatomic hydrogen molecule (H_2) is shown in **Figure 6-11b**.

Trends within periods A pattern in atomic size emerges as you look across a period in **Figure 6-11c**. In general, there is a decrease in atomic radii as you move left-to-right across a period. This trend is caused by the increasing positive charge in the nucleus and the fact that the principal energy level within a period remains the same. Each successive element has one additional proton and electron, and each additional electron is added to the same principal energy level. Moving across a period, no additional electrons come between the valence electrons and the nucleus. Thus, the valence electrons are not shielded from the increased nuclear charge. The result is that the increased nuclear charge pulls the outermost electrons closer to the nucleus.

Trends within groups Atomic radii generally increase as you move down a group. The nuclear charge increases and electrons are added to successively higher principal energy levels. Although you might think the increased nuclear charge would pull the outer electrons toward the nucleus and make the atom smaller, this effect is overpowered by several other factors. Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, making the atom larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the greater pull of the increased nuclear charge. Also, as additional orbitals between the nucleus and the outer electrons are occupied, these electrons shield the outer electrons from the pull of the nucleus. **Figure 6-12** summarizes the group and period trends in atomic radii.

miniLAB

Periodicity of Molar Heats of Fusion and Vaporization

Making and Using Graphs The heats required to melt or to vaporize a mole (a specific amount of matter) of matter are known as the molar heat of fusion (H_f) and the molar heat of vaporization (H_v), respectively. These heats are unique properties of each element. You will investigate if the molar heats of fusion and vaporization for the period 2 and 3 elements behave in a periodic fashion.

Materials either a graphing calculator, a computer graphing program, or graph paper; Appendix **Table C-6** or access to comparable element data references

Procedure

Use **Table C-6** in Appendix C to look up and record the molar heat of fusion and the molar heat of vaporization for the period 3 elements listed in the table. Then, record the same data for the period 2 elements.

Molar Heat Data			
Element	Atomic number	H_f (kJ/mol)	H_v (kJ/mol)
Na	11		
Mg	12		
Al	13		
Si	14		
P	15		
S	16		
Cl	17		
Ar	18		

Analysis

- Graph molar heats of fusion versus atomic number. Connect the points with straight lines and label the curve. Do the same for molar heats of vaporization.
- Do the graphs repeat in a periodic fashion? Describe the graphs to support your answer.

EXAMPLE PROBLEM 6-2

Interpreting Trends in Atomic Radii

Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Do not use **Figure 6-11** to answer the question. Explain your answer in terms of trends in atomic radii.

1. Analyze the Problem

You are given four elements. First, determine the groups and periods the elements occupy. Then apply the general trends in atomic radii to determine which has the largest atomic radius.

2. Solve for the Unknown

From the periodic table, all the elements are found to be in period 2. Ordering the elements from left-to-right across the period yields:

Li, Be, C, F

Applying the trend of decreasing radii across a period means that lithium, the first element in period 2, has the largest radius.

3. Evaluating the Answer

The group trend in atomic radii has been correctly applied. Checking radii values from **Figure 6-11** verifies the answer.

PRACTICE PROBLEMS

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in **Figure 6-11** to answer the questions.

16. Which has the largest radius: magnesium (Mg), silicon (Si), sulfur (S), or sodium (Na)? The smallest?
17. Which has the largest radius: helium (He), xenon (Xe), or argon (Ar)? The smallest?
18. Can you determine which of two unknown elements has the larger radius if the only known information is that the atomic number of one of the elements is 20 greater than the other?



For more practice with periodic trend problems, go to **Supplemental Practice Problems** in Appendix A.

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge. You'll learn about ions in detail in Chapter 8, but for now, let's look at how the formation of an ion affects the size of an atom.

When atoms lose electrons and form positively charged ions, they always become smaller. For example, as shown in **Figure 6-13a** on the next page a sodium atom with a radius of 186 pm shrinks to a radius of 95 pm when it forms a positive sodium ion. The reason for the decrease in size is twofold. The electron lost from the atom will always be a valence electron. The loss of a valence electron may leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now fewer number of remaining electrons decreases, allowing them to be pulled closer to the nucleus.

When atoms gain electrons and form negatively charged ions, they always become larger, as shown in **Figure 6-13b**. The addition of an electron to an

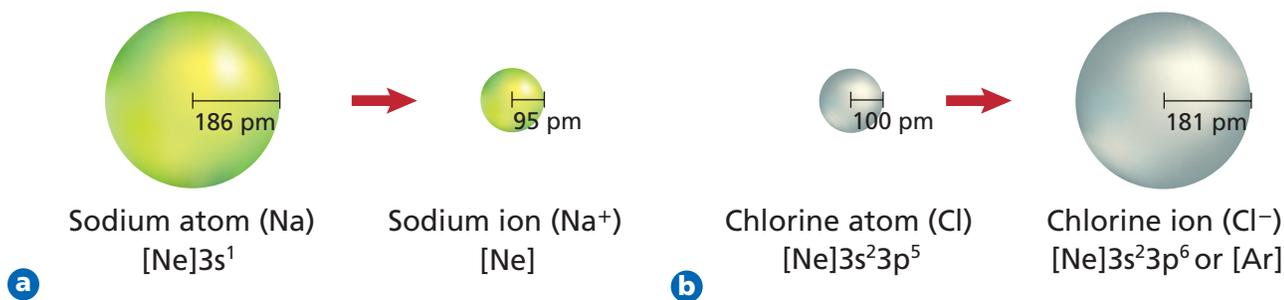


Figure 6-13

Atoms undergo significant changes in size when forming ions. **a** The sodium atom loses an electron and becomes smaller. **b** The chlorine ion gains an electron and becomes larger. How is each ion's electron configuration related to those of the noble gas elements?

atom increases the electrostatic repulsion between the atom's outer electrons, forcing them to move farther apart. The increased distance between the outer electrons results in a larger radius.

Trends within periods The ionic radii of most of the representative elements are shown in **Figure 6-14**. Note that elements on the left side of the table form smaller positive ions, and elements on the right side of the table form larger negative ions. In general, as you move left-to-right across a period, the size of the positive ions gradually decreases. Then, beginning in group 5A or 6A, the size of the much larger negative ions also gradually decreases.

Trends within groups As you move down a group, an ion's outer electrons are in higher principal energy levels, resulting in a gradual increase in ionic size. Thus, the ionic radii of both positive and negative ions increase as you move down a group. **Figure 6-15** on the next page summarizes the group and period trends in ionic radii.

Figure 6-14

The table shows the ionic radii of most of the representative elements. The ion sizes are shown relative to one another, while the actual radii are given in picometers (1×10^{-12} meters). Note that the elements on the left side of the table form positive ions, and those on the right form negative ions.

		1A		2A		3A		4A		5A		6A		7A	
		Li 76		Be 31		B 20		C 15		N 146		O 140		F 133	
		1+		2+		3+		4+		3-		2-		1-	
2		Na 102		Mg 72		Al 54		Si 41		P 212		S 184		Cl 181	
		1+		2+		3+		4+		3-		2-		1-	
3		K 138		Ca 100		Ga 62		Ge 53		As 222		Se 198		Br 195	
		1+		2+		3+		4+		3-		2-		1-	
4		Rb 152		Sr 118		In 81		Sn 71		Sb 62		Te 221		I 220	
		1+		2+		3+		4+		5+		2-		1-	
5		Cs 167		Ba 135		Tl 95		Pb 84		Bi 74					
		1+		2+		3+		4+		5+					
6															

Ionization Energy

To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge in the nucleus and the negative charge of the electron.

Ionization energy is defined as the energy required to remove an electron from a gaseous atom. For example, 8.64×10^{-19} J is required to remove an electron from a gaseous lithium atom. The energy required to remove the first electron from an atom is called the first ionization energy. Therefore, the first ionization energy of lithium equals 8.64×10^{-19} J. The loss of the electron results in the formation of a Li^+ ion. The first ionization energies of the elements in periods 1 through 5 are plotted on the graph in **Figure 6-16**.

Think of ionization energy as an indication of how strongly an atom's nucleus holds onto its valence electrons. A high ionization energy value indicates the atom has a strong hold on its electrons. Atoms with large ionization energy values are less likely to form positive ions. Likewise, a low ionization energy value indicates an atom loses its outer electron easily. Such atoms are likely to form positive ions.

Take a close look at the graph in **Figure 6-16**. Each set of connected points represents the elements in a period. From the graph, it is clear that the group 1A metals have low ionization energies. Thus, group 1A metals (Li, Na, K, Rb) are likely to form positive ions. It also is clear that the group 8A elements (He, Ne, Ar, Kr, Xe) have high ionization energies and are unlikely to form ions. Gases of group 8A are extremely unreactive—their stable electron configuration greatly limits their reactivity.

After removing the first electron from an atom, it is possible to remove additional electrons. The amount of energy required to remove a second electron from a 1+ ion is called the second ionization energy, the amount of energy required to remove a third electron from a 2+ ion is called the third ionization energy, and so on. **Table 6-2** on the next page lists the first through ninth ionization energies for elements in period 2.

Reading across **Table 6-2** from left-to-right, you see that the energy required for each successive ionization always increases. However, the increase in energy does not occur smoothly. Note that for each element there is an ionization for which the required energy jumps dramatically. For example, the second ionization energy of lithium (7300 kJ/mol) is much greater than its first ionization energy (520 kJ/mol). This means a lithium atom is relatively likely to lose its first valence electron, but extremely unlikely to lose its second.

If you examine the table, you'll see that the ionization at which the large jump in energy occurs is related to the atom's number of valence electrons. Lithium has one valence electron and the jump occurs after the first ionization energy. Lithium easily forms the common lithium 1+ ion, but is unlikely to form a lithium 2+ ion. The jump in ionization energy shows that atoms hold

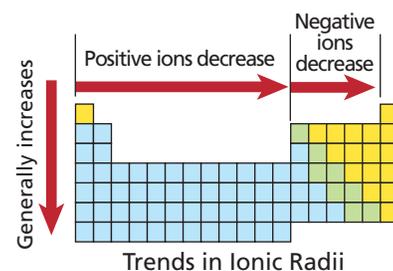


Figure 6-15

This small table provides a summary of the general trends in ionic radii.

Figure 6-16

The graph shows the first ionization energies for elements in periods 1 through 5. Note the high energies required to remove an electron from a noble gas element. What trend in first ionization energies do you observe as you move down a group?

First Ionization Energy of Elements in Periods 1–5

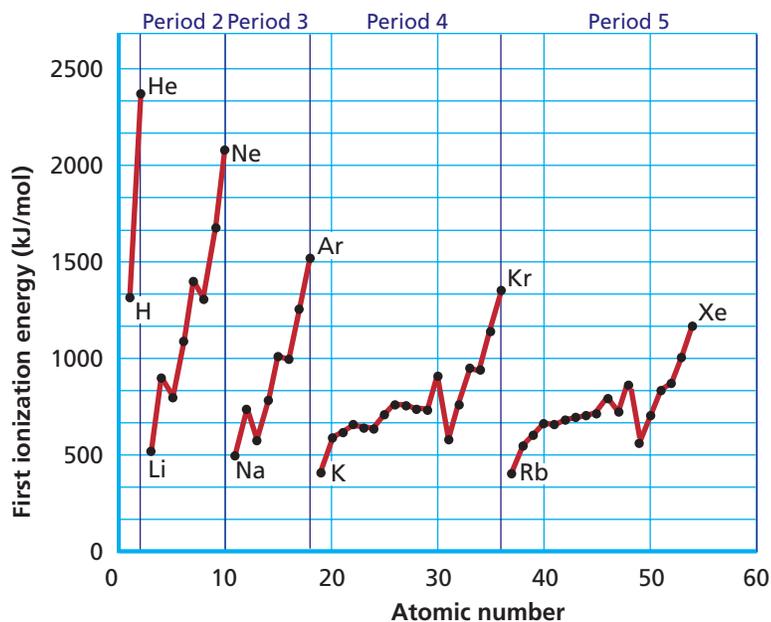


Table 6-2

Successive Ionization Energies for the Period 2 Elements										
Element	Valence electrons	Ionization energy (kJ/mol)*								
		1 st	2 nd	3 rd	4 th	5 th	6 th	7 th	8 th	9 th
Li	1	520	7300							
Be	2	900	1760	14 850						
B	3	800	2430	3660	25 020					
C	4	1090	2350	4620	6220	37 830				
N	5	1400	2860	4580	7480	9440	53 270			
O	6	1310	3390	5300	7470	10 980	13 330	71 330		
F	7	1680	3370	6050	8410	11 020	15 160	17 870	92 040	
Ne	8	2080	3950	6120	9370	12 180	15 240	20 000	23 070	115 380

* mol is an abbreviation for mole, a quantity of matter.

onto their inner core electrons much more strongly than they hold onto their valence electrons. Where does the jump in ionization energy occur for oxygen, an atom with six valence electrons?

Trends within periods As shown in **Figure 6-16** and by the values in **Table 6-2**, first ionization energies generally increase as you move left-to-right across a period. The increased nuclear charge of each successive element produces an increased hold on the valence electrons.

Trends within groups First ionization energies generally decrease as you move down a group. This decrease in energy occurs because atomic size increases as you move down the group. With the valence electrons farther from the nucleus, less energy is required to remove them. **Figure 6-17** summarizes the group and period trends in first ionization energies.

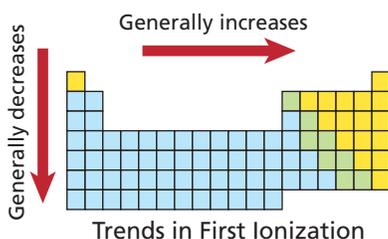


Figure 6-17

This small table provides a summary of the general trends in first ionization energies.

Octet rule When a sodium atom loses its single valence electron to form a 1+ sodium ion, its electron configuration changes as shown below.



Note that the sodium ion has the same electron configuration as neon ($1s^22s^22p^6$), a noble gas. This observation leads to one of the most important principles in chemistry, the octet rule. The **octet rule** states that atoms tend to gain, lose, or share electrons in order to acquire a full set of eight valence electrons. This reinforces what you learned earlier that the electron configuration of filled s and p orbitals of the same energy level (consisting of eight valence electrons) is unusually stable. Note that the first period elements are an exception to the rule, as they are complete with only two valence electrons.

The octet rule is useful for determining the type of ions likely to form. Elements on the right side of the periodic table tend to gain electrons in order to acquire the noble gas configuration; therefore, these elements tend to form negative ions. In a similar manner, elements on the left side of the table tend to lose electrons and form positive ions.

Electronegativity

The **electronegativity** of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. **Figure 6-18** lists the electronegativity values for most of the elements. These values are calculated based upon a number of factors, and are expressed in terms of a numerical

value of 3.98 or less. The units of electronegativity are arbitrary units called Paulings, named after American scientist Linus Pauling (1901–1994).

Note that because the noble gases form very few compounds, they have been left out of **Figure 6-18**. Fluorine is the most electronegative element, with a value of 3.98, and cesium and francium are the least electronegative elements, with values of 0.79 and 0.7, respectively. In a chemical bond, the atom with the greater electronegativity more strongly attracts the bond's electrons. You will use electronegativity values in upcoming chapters to help determine the types of bonds that exist between elements in a compound.

Trends within periods and groups Electronegativity generally decreases as you move down a group, and increases as you move left-to-right across a period; therefore, the lowest electronegativities are found at the lower left side of the periodic table, while the highest electronegativities are found at the upper right.

Figure 6-18

The table shows the electronegativity values for most of the elements. In which areas of the periodic table do the highest electronegativities tend to occur? The lowest?

Increasing electronegativity

													electronegativity < 1.0					
													1.0 ≥ electronegativity < 2.0					
													2.0 ≥ electronegativity < 3.0					
													3.0 ≥ electronegativity < 4.0					

													Decreasing electronegativity					
1 H 2.20														2 He				
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	10 Ne	
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar	
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr	
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe	
55 Cs 0.79	56 Ba 0.89	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn	
87 Fr 0.70	88 Ra 0.90	89 Ac 1.1	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Uuu	112 Uub							

Electronegativity Values in Paulings

Section 6.3 Assessment

- Sketch a simplified periodic table and use arrows and labels to compare period and group trends in atomic and ionic radii, ionization energies, and electronegativities.
- Explain how the period and group trends in atomic radii are related to electron configuration.
- Which has the largest atomic radius: nitrogen (N), antimony (Sb), or arsenic (As)? The smallest?
- For each of the following properties, indicate whether fluorine or bromine has a larger value.

a. electronegativity	c. atomic radius
b. ionic radius	d. ionization energy
- Thinking Critically** Explain why it takes more energy to remove the second electron from a lithium atom than it does to remove the fourth electron from a carbon atom.
- Making and Using Graphs** Graph the atomic radii of the group A elements in periods 2, 3, and 4 versus their atomic numbers. Connect the points of elements in each period, so that there are three separate curves on the graph. Summarize the trends in atomic radii shown on your graph. Explain.

Descriptive Chemistry of the Elements

What do elements look like? How do they behave? Can periodic trends in the properties of elements be observed? You cannot examine all of the elements on the periodic table because of limited availability, cost, and safety concerns. However, you can observe several of the representative elements, classify them, and compare their properties. The observation of the properties of elements is called descriptive chemistry.

Problem

What is the pattern of properties of the representative elements?

Objectives

- **Observe** properties of various elements.
- **Classify** elements as metals, nonmetals, and metalloids.
- **Examine** general trends within the periodic table.

Materials

stoppered test tubes containing small samples of elements
 plastic dishes containing samples of elements
 conductivity apparatus
 1.0M HCl
 test tubes (6)
 test tube rack
 10-mL graduated cylinder
 spatula
 small hammer
 glass marking pencil

Safety Precautions



- Wear safety goggles and a lab apron at all times.
- Do not handle elements with bare hands.
- 1.0M HCl is harmful to eyes and clothing.
- Never test chemicals by tasting.
- Follow any additional safety precautions provided by your teacher.

Pre-Lab

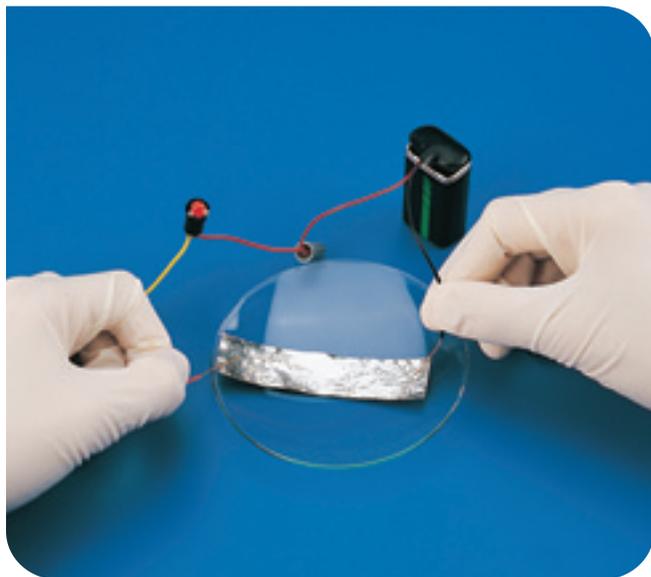
1. Read the entire CHEMLAB.
2. Prepare a data table similar to the one below to record the observations you make during the lab.
3. Examine the periodic table. What is the physical state of most metals? Nonmetals? Metalloids?
4. Look up the definitions of the terms luster, malleability, and electrical conductivity. To what elements do they apply?

Observation of Elements

Element	Appearance and physical state	Malleable or brittle?	Reactivity with HCl	Electrical conductivity	Classification

Procedure

1. Observe and record the appearance of the element sample in each test tube. Observations should include physical state, color, and other characteristics such as luster and texture. **CAUTION:** *Do not remove the stoppers from the test tubes.*
2. Remove a small sample of each of the elements contained in a dish and place it on a hard surface designated by your teacher. Gently tap each element sample with a small hammer. **CAUTION:** *Safety goggles must be worn.* If the element is malleable it will flatten. If it is brittle, it will shatter. Record your observations.
3. Use the conductivity tester to determine which elements conduct electricity. An illuminated light bulb is evidence of electrical conductivity. Record your results in your data table. Clean the electrodes with water and make sure they are dry before testing each element.



4. Label each test tube with the symbol for one of the elements in the plastic dishes. Using a graduated cylinder, add 5 mL of water to each test tube.
5. Use a spatula to put a small amount of each of the six elements (approximately 0.2 g or a 1-cm long ribbon) into the test tube labeled with its chemical symbol. Using a graduated cylinder, add 5 mL of 1.0M HCl to each test tube. Observe each test tube for at least one minute. The formation of bubbles is evidence of a reaction between the acid and the element. Record your observations.

Cleanup and Disposal

Dispose of all materials as instructed by your teacher.

Analyze and Conclude

1. **Interpreting Data** Metals are usually malleable and good conductors of electricity. They are generally lustrous and silver or white in color. Many react with acids. Write the word “metal” beneath the Classification heading in the data table for those element samples that display the general characteristics of metals.
2. **Interpreting Data** Nonmetals can be solids, liquids, or gases. They do not conduct electricity and do not react with acids. If a nonmetal is a solid, it is likely to be brittle and have color (other than white or silver). Write the word “nonmetal” beneath the Classification heading in the data table for those element samples that display the general characteristics of nonmetals.
3. **Interpreting Data** Metalloids combine some of the properties of both metals and nonmetals. Write the word “metalloid” beneath the Classification heading in the data table for those element samples that display the general characteristics of metalloids.
4. **Making a Model** Construct a periodic table and label the representative elements by group (1A through 7A). Using the information in your data table and the periodic table, record the identities of elements observed during the lab in your periodic table.
5. **Interpreting** Describe any trends among the elements you observed in the lab.

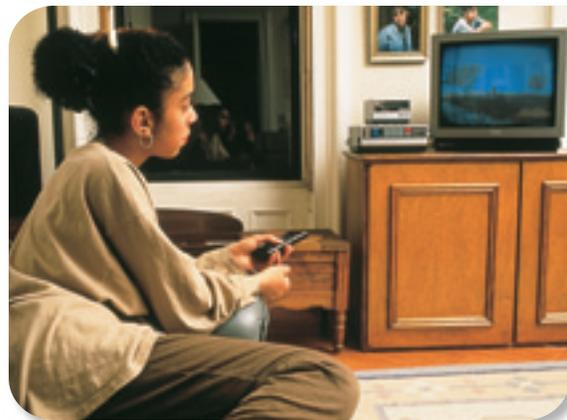
Real-World Chemistry

1. Why did it take so long to discover the first noble gas element?
2. Research one of the most recently discovered elements. New elements are created in particle accelerators and tend to be very unstable. Because of this, many of the properties of a new element can not be determined. Using periodic group trends in melting and boiling point, predict whether the new element you selected is likely to be a solid, liquid, or gas.

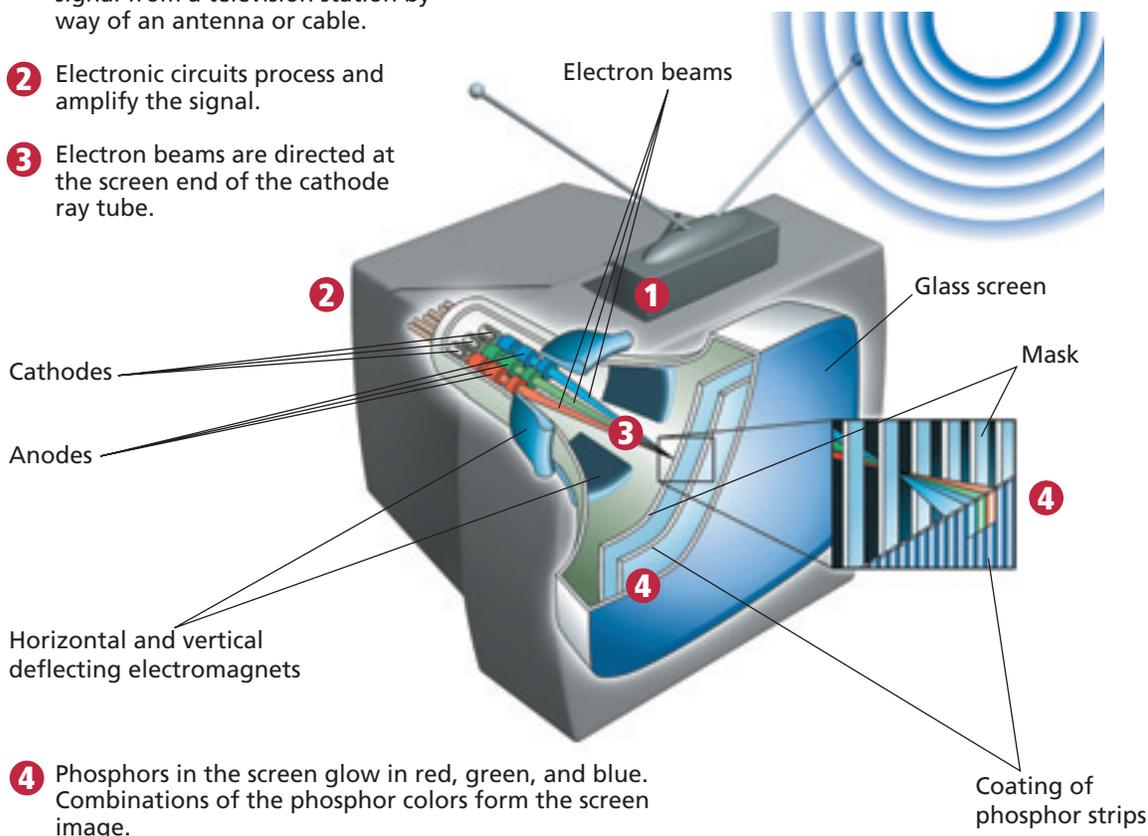
How It Works

Television Screen

Most television screens are part of a cathode ray tube. As you know, a cathode ray tube is an evacuated chamber which produces a beam of electrons, known as a cathode ray. Electronic circuitry inside the television processes an electronic signal received from the television station. The processed signal is used to vary the strength of several electron beams, while magnetic fields are used to direct the beams to different parts of the screen.



- 1 The television receives an electronic signal from a television station by way of an antenna or cable.
- 2 Electronic circuits process and amplify the signal.
- 3 Electron beams are directed at the screen end of the cathode ray tube.



- 4 Phosphors in the screen glow in red, green, and blue. Combinations of the phosphor colors form the screen image.

Thinking Critically

1. **Relating Cause and Effect** Why don't the phosphors in a television screen glow when the television is turned off?
2. **Inferring** Why is the length of time over which a phosphor emits light an important factor to consider when designing a television screen?

Summary

6.1 Development of the Modern Periodic Table

- Periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- Newlands's law of octaves, which was never accepted by fellow scientists, organized the elements by increasing atomic mass. Mendeleev's periodic table, which also organized elements by increasing atomic mass, became the first widely accepted organization scheme for the elements. Moseley fixed the errors inherent in Mendeleev's table by organizing the elements by increasing atomic number.
- The periodic table organizes the elements into periods (rows) and groups (columns) by increasing atomic number. Elements with similar properties are in the same group.
- Elements are classified as either metals, nonmetals, or metalloids. The stair-step line on the table separates metals from nonmetals. Metalloids border the stair-step line.

6.2 Classification of the Elements

- Elements in the same group on the periodic table have similar chemical properties because they have the same valence electron configuration.
- The four blocks of the periodic table can be characterized as follows:
 - s-block: filled or partially filled s orbitals.
 - p-block: filled or partially filled p orbitals.
 - d-block: filled outermost s orbital of energy level n , and filled or partially filled d orbitals of energy level $n - 1$.
 - f-block: filled outermost s orbital, and filled or partially filled 4f and 5f orbitals.

- For the group A elements, an atom's group number equals its number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.
- The s^2p^6 electron configuration of the group 8A elements (noble gases) is exceptionally stable.

6.3 Periodic Trends

- Atomic radii generally decrease as you move left-to-right across a period, and increase as you move down a group.
- Positive ions are smaller than the neutral atoms from which they form. Negative ions are larger than the neutral atoms from which they form.
- Ionic radii of both positive and negative ions decrease as you move left-to-right across a period. Ionic radii of both positive and negative ions increase as you move down a group.
- Ionization energy indicates how strongly an atom holds onto its electrons. After the valence electrons have been removed from an atom, there is a tremendous jump in the ionization energy required to remove the next electron.
- Ionization energies generally increase as you move left-to-right across a period, and decrease as you move down a group.
- The octet rule states that atoms gain, lose, or share electrons in order to acquire the stable electron configuration of a noble gas.
- Electronegativity, which indicates the ability of atoms of an element to attract electrons in a chemical bond, plays a role in determining the type of bond formed between elements in a compound.
- Electronegativity values range from 0.7 to 3.96, and generally increase as you move left-to-right across a period, and decrease as you move down a group.

Vocabulary

- alkali metal (p. 155)
- alkaline earth metal (p. 155)
- electronegativity (p. 168)
- group (p. 154)
- halogen (p. 158)
- inner transition metal (p. 158)
- ion (p. 165)
- ionization energy (p. 167)
- metal (p. 155)
- metalloid (p. 158)
- noble gas (p. 158)
- nonmetal (p. 158)
- octet rule (p. 168)
- period (p. 154)
- periodic law (p. 153)
- representative element (p. 154)
- transition element (p. 154)
- transition metal (p. 158)

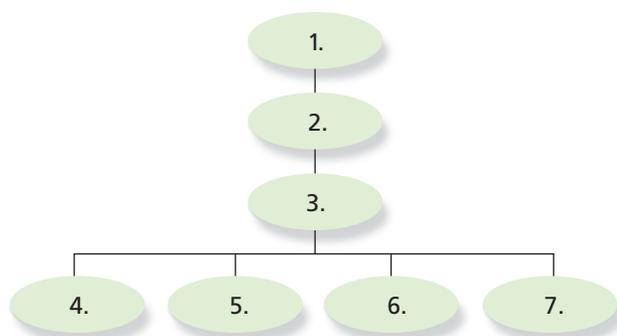


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 6 Assessment.

Concept Mapping

25. Complete the concept map using the following terms: electronegativity, electron configuration, periodic trends, ionic radius, atomic radius, ionization energy, and periodic table.



Mastering Concepts

26. Explain how Mendeleev's periodic table was in error. How was this error fixed? (6.1)
27. Explain the contribution of Newlands's law of octaves to the development of the modern periodic table. (6.1)
28. German chemist Lothar Meyer and Russian chemist Dmitri Mendeleev both proposed similar periodic tables in 1869. Why is Mendeleev generally given credit for the periodic table? (6.1)
29. How was Mendeleev's periodic table organized? (6.1)
30. What is the periodic law? (6.1)
31. Identify each of the following as a metal, nonmetal, or metalloid. (6.1)
- | | |
|--------------|-----------------|
| a. oxygen | d. iron |
| b. barium | e. neon |
| c. germanium | f. praseodymium |
32. Describe the general characteristics of metals. (6.1)
33. Match each numbered item on the right with the lettered item that it is related to on the left. (6.1)
- | | |
|--------------------------|-------------|
| a. alkali metals | 1. group 8A |
| b. halogens | 2. group 1A |
| c. alkaline earth metals | 3. group 2A |
| d. noble gases | 4. group 7A |
34. Identify each of the elements in problem 31 as a representative element or a transition element. (6.1)
35. Sketch a simplified periodic table and use labels to identify the alkali metals, alkaline earth metals, transition metals, inner transition metals, noble gases, and halogens. (6.1)
36. A shiny solid element also is ductile. What side of the periodic table is it likely to be found? (6.1)
37. What are the general properties of a metalloid? List three metalloid elements. (6.1)
38. What is the purpose of the heavy stair-step line on the periodic table? (6.1)
39. Describe the two types of numbering used to identify groups on the periodic table. (6.1)
40. Give the chemical symbol of each of the following elements. (6.1)
- the two elements that are liquids at room temperature
 - the noble gas with the greatest atomic mass
 - any metal from group 4A
 - any inner transition metal
41. Why do the elements chlorine and iodine have similar chemical properties? (6.2)
42. How are the numbers of valence electrons of the group A elements related to the group number? (6.2)
43. How is the energy level of an atom's valence electrons related to the period it is in on the periodic table? (6.2)
44. How many valence electrons do each of the noble gases have? (6.2)
45. What are the four blocks of the periodic table? (6.2)
46. In general, what electron configuration has the greatest stability? (6.2)
47. Determine the group, period, and block in which each of the following elements is located on the periodic table. (6.2)
- | | |
|---------------------------------|--------------------------|
| a. $[\text{Kr}]5s^24d^1$ | c. $[\text{He}]2s^22p^6$ |
| b. $[\text{Ar}]4s^23d^{10}4p^3$ | d. $[\text{Ne}]3s^23p^1$ |
48. Categorize each of the elements in problem 47 as a representative element or a transition metal. (6.2)
49. Explain how an atom's valence electron configuration determines its place on the periodic table. (6.2)
50. Write the electron configuration for the element fitting each of the following descriptions. (6.2)
- the metal in group 5A
 - the halogen in period 3
 - the alkali metal in period 2
 - the transition metal that is a liquid at room temperature

51. Explain why the radius of an atom cannot be measured directly. (6.3)
52. Given any two elements within a group, is the element with the larger atomic number likely to have a larger or smaller atomic radius than the other element? (6.2)
53. Which elements are characterized as having their d orbitals fill with electrons as you move left-to-right across a period? (6.2)
54. Explain why is it harder to remove an inner shell electron than a valence electron from an atom. (6.3)
55. An element forms a negative ion when ionized. On what side of the periodic table is the element located? Explain. (6.3)
56. Of the elements magnesium, calcium, and barium, which forms the ion with the largest radius? The smallest? What periodic trend explains this? (6.3)
57. What is ionization energy? (6.3)
58. Explain why each successive ionization of an electron requires a greater amount of energy. (6.3)
59. Which group has the highest ionization energies? Explain why. (6.3)
60. Define an ion. (6.3)
61. How does the ionic radius of a nonmetal compare with its atomic radius? Explain why the change in radius occurs. (6.3)
62. Explain why atomic radii decrease as you move left-to-right across a period. (6.3)
63. Which element in each pair has the larger ionization energy? (6.3)
- Li, N
 - Kr, Ne
 - Cs, Li
64. Explain the octet rule. (6.3)
65. Use the illustration of spheres A and B to answer each of the following questions. Explain your reasoning for each answer. (6.3)
- If A is an ion and B is an atom of the same element, is the ion a positive or negative ion?



- If A and B represent the atomic radii of two elements in the same period, what is their correct order (left-to-right)?
- If A and B represent the ionic radii of two elements in the same group, what is their correct order (top-to-bottom)?

66. How many valence electrons do elements in each of the following groups have? (6.3)
- group 8A
 - group 3A
 - group 1A
67. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Why? (6.3)

Mixed Review

Sharpen your problem-solving skills by answering the following.

68. Match each numbered item on the right with the lettered item that it is related to on the left.
- | | |
|---------------------|----------------------------|
| a. group A elements | 1. periods |
| b. columns | 2. representative elements |
| c. group B elements | 3. groups |
| d. rows | 4. transition elements |
69. Which element in each pair is more electronegative?
- K, As
 - N, Sb
 - Sr, Be
70. Explain why the s-block of the periodic table is two groups wide, the p-block is six groups wide, and the d-block is ten groups wide.
71. Arrange the elements oxygen, sulfur, tellurium, and selenium in order of increasing atomic radii. Is your order an example of a group trend or a period trend?
72. Identify the elements with the following valence electron configurations.
- | | |
|---------------|---------------|
| a. $5s^1$ | c. $3s^2$ |
| b. $4s^23d^2$ | d. $4s^24p^3$ |
73. Which of the following is not a reason why atomic radii increase as you move down a group?
- shielding of inner electrons
 - valence electrons in larger orbitals
 - increased charge in the nucleus
74. Explain why there are no p-block elements in the first period of the periodic table.
75. Identify each of the following as an alkali metal, alkaline earth metal, transition metal, or inner transition metal.
- | | |
|--------------|--------------|
| a. cesium | d. ytterbium |
| b. zirconium | e. uranium |
| c. gold | f. francium |
76. An element is a brittle solid that does not conduct electricity well. Is the element a metal, nonmetal, or metalloid?

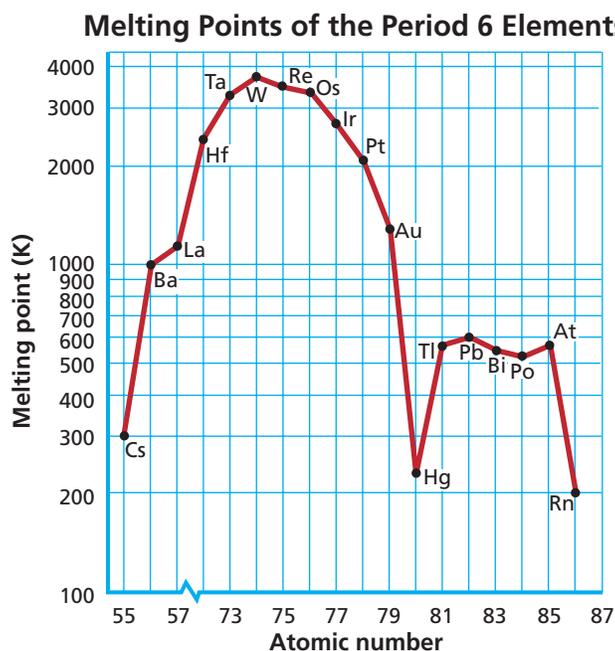
Thinking Critically

77. Interpreting Data Given the following data about an atom's ionization energies, predict its valence electron configuration. Explain your reasoning.

Ionization Data	
Ionization	Ionization Energy (kJ/mol)
First	734
Second	1850
Third	16 432

78. Applying Concepts Sodium forms a 1+ ion, while fluorine forms a 1- ion. Write the electron configuration for each ion. Why don't these two elements form 2+ and 2- ions, respectively?

79. Interpreting Data The melting points of the period 6 elements are plotted versus atomic number in the graph shown below. Determine the trends in melting point by analyzing the graph and the orbital configurations of the elements. Form a hypothesis that explains the trends. (Hint: In Chapter 5, you learned that half-filled sets of orbitals are more stable than other configurations of partially filled orbitals.)



Group 5A Density Data

Element	Atomic Number	Density (g/cm ³)
nitrogen	7	1.25×10^{-3}
phosphorus	15	1.82
arsenic	33	5.73
antimony	51	6.70
bismuth	83	9.78

80. Making and Using Graphs The densities of the group 5A elements are given in the table above. Plot density versus atomic number and state any trends you observe.

Writing in Chemistry

- 81.** In the early 1800s, German chemist J. W. Dobereiner proposed that some elements could be classified into sets of three, called triads. Research and write a report on Dobereiner's triads. What elements comprised the triads? How were the properties of elements within a triad similar?
- 82.** Electron affinity is another periodic property of the elements. Research and write a report on what electron affinity is and describe its group and period trends.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 83.** Define matter. Identify whether or not each of the following is a form of matter. (Chapter 1)
- microwaves
 - helium inside a balloon
 - heat from the Sun
 - velocity
 - a speck of dust
 - the color blue
- 84.** Convert the following mass measurements as indicated. (Chapter 2)
- 1.1 cm to meters
 - 76.2 pm to millimeters
 - 11 Mg to kilograms
 - 7.23 micrograms to kilograms
- 85.** How is the energy of a quantum of emitted radiation related to the frequency of the radiation? (Chapter 5)
- 86.** What element has the ground-state electron configuration of $[\text{Ar}]4s^23d^6$? (Chapter 5).

Use these questions and the test-taking tip to prepare for your standardized test.

- Periodic law states that elements show a _____.
 - repetition of their physical properties when arranged by increasing atomic radius
 - repetition of their chemical properties when arranged by increasing atomic mass
 - periodic repetition of their properties when arranged by increasing atomic number
 - periodic repetition of their properties when arranged by increasing atomic mass
- Elements in the same group of the periodic table have the same _____.
 - number of valence electrons
 - physical properties
 - number of electrons
 - electron configuration
- All of the following are true EXCEPT _____.
 - atomic radius of Na < atomic radius of Mg
 - electronegativity of C > electronegativity of B
 - ionic radius of Br⁻ > atomic radius of Br
 - first ionization energy of K > first ionization energy of Rb
- Which of the following is NOT true of an atom obeying the octet rule?
 - obtains a full set of eight valence electrons
 - acquires the valence configuration of a noble gas
 - possesses eight electrons in total
 - has a s²p⁶ valence configuration
- What is the group, period, and block of an atom with the electron configuration [Ar]4s²3d¹⁰4p⁴?
 - group 4A, period 4, d-block
 - group 6A, period 3, p-block
 - group 4A, period 4, p-block
 - group 6A, period 4, p-block

- Moving down a group on the periodic table, which two atomic properties follow the same trend?
 - atomic radius and ionization energy
 - ionic radius and atomic radius
 - ionization energy and ionic radius
 - ionic radius and electronegativity

Interpreting Tables Use the periodic table and the table at the bottom of the page to answer questions 7 and 8.

- It can be predicted that silicon will experience a large jump in ionization energy after its _____.
 - second ionization
 - third ionization
 - fourth ionization
 - fifth ionization
- Which of the following requires the most energy?
 - second ionization of Li
 - fourth ionization of N
 - first ionization of Ne
 - third ionization of Be
- Niobium (Nb) is a(n) _____.
 - nonmetal
 - transition metal
 - alkali metal
 - halogen
- It can be predicted that element 118 would have properties similar to a(n) _____.
 - alkali earth metal
 - halogen
 - metalloid
 - noble gas



TEST-TAKING TIP

Practice, Practice, Practice Practice to improve your performance on standardized tests. Don't compare yourself to anyone else.

Successive Ionization Energies for the Period 2 Elements

Element	Valence electrons	Ionization energy (kJ/mol)*								
		1 st	2 nd	3 rd	4 th	5 th	6 th	7 th	8 th	9 th
Li	1	520	7300							
Be	2	900	1760	14 850						
B	3	800	2430	3660	25 020					
C	4	1090	2350	4620	6220	37 830				
N	5	1400	2860	4580	7480	9440	53 270			
O	6	1310	3390	5300	7470	10 980	13 330	71 330		
F	7	1680	3370	6050	8410	11 020	15 160	17 870	92 040	
Ne	8	2080	3950	6120	9370	12 180	15 240	20 000	23 070	115 380

* mol is an abbreviation for mole, a quantity of matter.

The Elements

What You'll Learn

- ▶ You will classify elements based on their electron configurations.
- ▶ You will relate electron configurations to the properties of elements.
- ▶ You will identify the sources and uses of selected elements.

Why It's Important

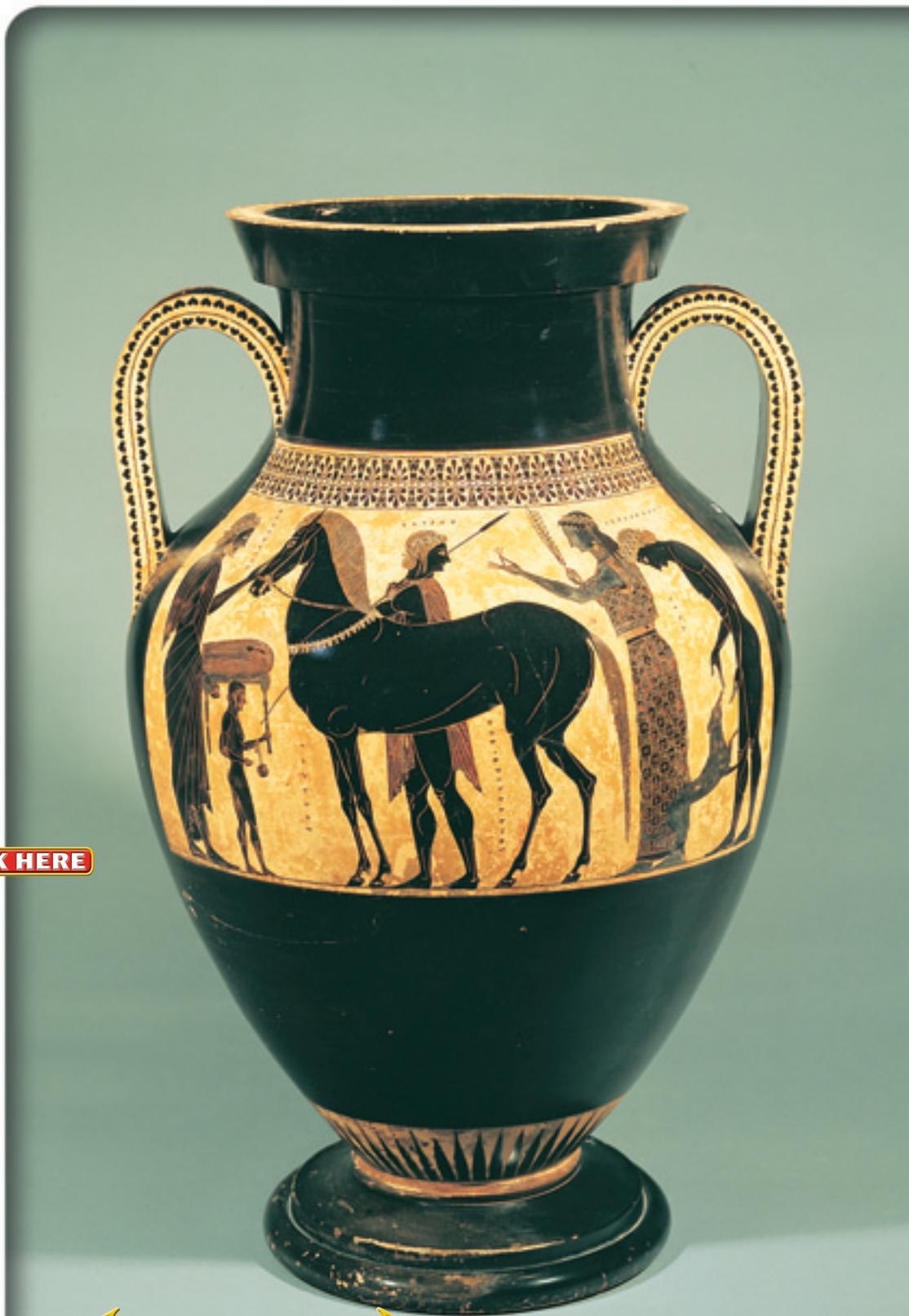
What you know about an element can affect the choices you make. Before all of its properties were known, toxic lead glazes were used to seal clay storage containers. Modern steel cans are lined with tin, which is a non-toxic element similar to lead.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about the properties of elements.

Some historians think that drinking wine from lead-glazed vases contributed to the fall of the Roman Empire.



DISCOVERY LAB



Magnetic Materials

You know that magnets can attract some materials. In this lab, you will classify materials based on their interaction with a magnet and look for a pattern in your data.

Safety Precautions



Always wear your safety goggles during lab.

Procedure

1. Working with a partner, review the properties of magnets. Arrange the bar magnets so that they are attracted to one another. Then arrange them so that they repel one another.
2. Test each item with your magnet. Record your observations.
3. Test as many other items in the classroom as time allows. Predict the results before testing each item.

Analysis

Look at the group of items that were attracted to the magnet. What do these items have in common? In Section 7.3, you will learn more about what determines whether a material is magnetic.

Materials

bar magnet	soda cans
aluminum foil	empty soup cans
paper clips	coins
coins	variety of
hair pin	other items

Section

7.1

Properties of s-Block Elements

Objectives

- **Explain** how elements in a given group are both similar and different.
- **Discuss** the properties of hydrogen.
- **Describe** and **compare** the properties of alkali metals and alkaline earth metals.

Vocabulary

diagonal relationship

If you could travel through the universe to collect samples of matter, you would find 92 naturally occurring elements. The amount of each element would vary from location to location. Helium, which is the second most common element in the universe, is much less abundant here on Earth. Oxygen is the most abundant element on Earth. Per kilogram of Earth's crust, there are 4.64×10^5 mg of oxygen, but only 8×10^{-3} mg of helium. Elements with atomic numbers greater than 92 do not exist in nature. These synthetic elements must be created in laboratories or nuclear reactors.

Representative Elements

Recall from Chapter 6 that no one element can represent the properties of all elements. However, the elements in groups 1A through 8A are called representative elements because, as a group, they display a wide range of physical and chemical properties. For example, groups 1A through 8A include metals and nonmetals; highly reactive elements and some that hardly react at all; and elements that are solids, liquids, and gases at room temperature.

Elements in any given group on the periodic table have the same number of valence electrons. The number and location of valence electrons determine the chemistry of an element. Thus, elements within a group have similar physical and chemical properties. Representative elements display the range of possible valence electrons from one in group 1A to eight in group 8A. The valence electrons of representative elements are in s or p orbitals.

1A 1		2A 2		3B 3	4B 4
Li	Be			B	
	Mg			Al	Si

Figure 7-1

The diagonal arrows connect pairs of elements that have similar properties.

Why are the properties of elements within a group similar but not identical? Although elements within a group have the same number of valence electrons, they have different numbers of nonvalence electrons. Remember what happens as the atomic number increases within a group. As new levels of electrons are added, the atomic radius increases and the shielding effect increases. As a result, the ionization energy decreases. A lower ionization energy makes it easier for an element to lose electrons.

Recall that metals tend to lose electrons. Thus, the lower the ionization energy, the more reactive the metal. For a group of metals, reactivity *increases* as the atomic number increases. The opposite is true for nonmetals because nonmetals tend to gain electrons. The higher the ionization energy of a nonmetal, the more reactive the nonmetal. For a group of nonmetals, reactivity *decreases* as the atomic number increases. Of the representative elements, which is the most reactive metal? Which is the most reactive nonmetal? (Hint: What is the trend for ionization energy across a period?)

Diagonal relationships Some period 2 elements do not behave as predicted by their locations on the table. Often, the lightest element in a group is the least representative. These light elements have more in common with the period-3 element in the next group than with the period-3 element in their own group. These close relationships between elements in neighboring groups are called **diagonal relationships**. Three diagonal relationships that you will study later in this chapter are shown in **Figure 7-1**.

1A
Hydrogen
1
H
1s ¹

Hydrogen

Hydrogen is placed in group 1A because it has one valence electron. This placement does not mean that hydrogen has the same properties as the metals in group 1A. In fact, hydrogen shares many properties with the nonmetals in group 7A. Because hydrogen has both metallic and nonmetallic properties, it is not considered part of any group.

When Henry Cavendish discovered hydrogen in 1766, he called it “flammable air” because it burned when ignited in air. In 1783, Antoine Lavoisier named hydrogen for the water that forms when hydrogen and oxygen combine. Lavoisier used the Greek roots for water (*hydro*) and to form (*genes*). **Figure 7-2** shows the explosive reaction of hydrogen and oxygen that occurred when a passenger airship crashed in 1937. The airship was kept aloft by hydrogen, which is a colorless, odorless, lighter-than-air gas under normal conditions of temperature and pressure. As a part of water and most other compounds found in organisms, hydrogen is essential to life.

The universe contains more than 90% hydrogen by mass. There are three naturally occurring hydrogen isotopes: protium, deuterium, and tritium. The vast majority of hydrogen, 99.985%, is protium (hydrogen-1), which has no neutrons. Deuterium (hydrogen-2), which makes up 0.015% of hydrogen, has one neutron. Tritium (hydrogen-3), which has two neutrons, is a radioactive isotope. It is produced when cosmic rays bombard water in the atmosphere.

The physical properties of isotopes differ slightly because of differences in atomic mass. For example, water that contains deuterium is called heavy water because the neutrons in deuterium add mass to the water molecule. Some nuclear reactors use heavy water to help keep the chain reaction going. The heavy water slows down (or moderates) the neutrons produced during nuclear fission so that they can be absorbed by the uranium fuel. You will learn more about nuclear reactions in Chapter 25.

Figure 7-2

When the *Hindenburg* crashed in New Jersey, the hydrogen that kept the airship aloft came in contact with oxygen and exploded.



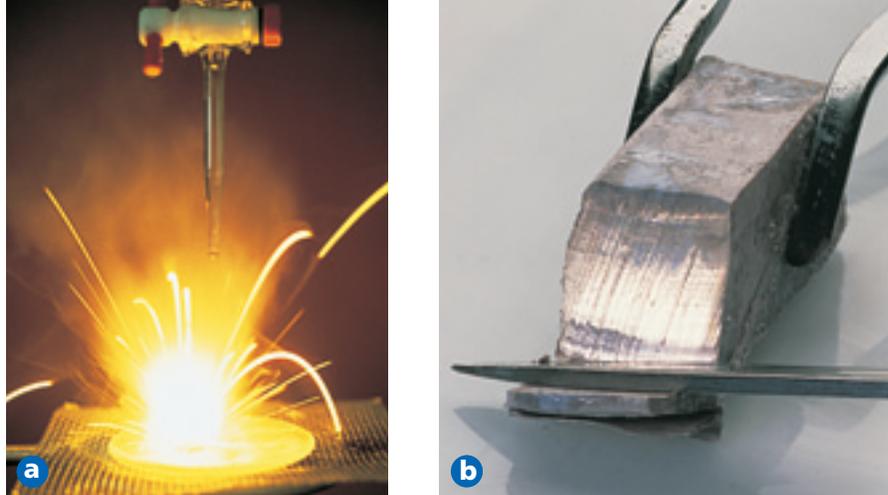


Figure 7-3

- a** The reaction of potassium and water produces hydrogen gas and an alkaline solution of potassium hydroxide. The heat produced by the reaction ignites the hydrogen.
- b** Sodium is as soft as cold butter. The shiny surface dulls as it reacts with oxygen.

Hydrogen's single valence electron explains its unusual set of metallic and nonmetallic properties. When a hydrogen atom acts like a nonmetal, it gains an electron and achieves the stable electron configuration of helium. When hydrogen reacts with a nonmetal such as oxygen, it acts like a metal. Hydrogen loses its single electron and forms a hydrogen ion (H^+). A hydrogen ion is a nucleus with a single proton. Does a hydrogen ion contain any neutrons? In Chapter 19, you will learn the role hydrogen ions play in the chemistry of acids and bases.

Hydrogen can be produced in the laboratory when a metal reacts with an acid or when electricity is used to separate water into hydrogen and oxygen. Large quantities of industrial hydrogen are produced when water reacts with methane, which is the main ingredient in natural gas. The major industrial use of hydrogen is in the production of ammonia from nitrogen and hydrogen gases. Hydrogen also is used to convert liquid vegetable oils into solid fats such as shortening. You will learn more about hydrogenation in Chapter 23, and fats and oils in Chapter 24.

Group 1A: Alkali Metals

People used to pour water over the ashes from a wood-burning fire to produce a compound of sodium called lye. They boiled the lye with animal fat to make soap. Lye, which is the active ingredient in drain cleansers, is an example of an alkaline solution. The term *alkali* comes from the Arabic *al-qili* meaning “ashes of the saltwort plant.” Saltworts grow on beaches or near salt marshes. Because group 1A metals react with water to form alkaline solutions, as shown in **Figure 7-3a**, they are called alkali metals.

Alkali metals easily lose a valence electron and form an ion with a 1+ charge. They are soft enough to cut with a knife. Sodium, shown in **Figure 7-3b**, has the consistency of cold butter. Because alkali metals are highly reactive, they are found combined with other elements in nature. Lab samples are stored in oil to prevent a reaction with oxygen in the air. Alkali metals are good conductors of heat and electricity.

Lithium Trace amounts of lithium, the lightest alkali metal, are found in water, soil, and rocks. Lithium is the least reactive of the alkali metals. Its compounds are less likely to dissolve in water. In these and other properties, lithium is more closely related to magnesium than to the other alkali metals. Lithium has an atomic radius of 152 pm and an ionic radius of 76 pm. Magnesium has an atomic radius of 160 pm and an ionic radius of 72 pm. These similar physical properties lead to similar chemical properties, which is why lithium and magnesium have a diagonal relationship.

1A
Lithium 3 Li [He]2s ¹
Sodium 11 Na [Ne]3s ¹
Potassium 19 K [Ar]4s ¹
Rubidium 37 Rb [Kr]5s ¹
Cesium 55 Cs [Xe]6s ¹
Francium 87 Fr [Rn]7s ¹

Long-lasting lithium batteries may extend the range of electric automobiles. Compounds of lithium are used in dehumidifiers to absorb water. Lithium carbonate is used to strengthen glass and as a drug to treat bipolar disorders. Such disorders involve mood swings from mania to depression. Alloys of lithium, magnesium, and aluminum are used for items such as airplane parts because these parts must be strong, yet lightweight. Remember that an alloy is a solid solution. A chemist can fine-tune the properties of an alloy by varying the amount of each element.

Careers Using Chemistry

Dietician

Would you like to help others improve their eating habits based on sound scientific principles? If so, consider a career as a dietician or nutritionist.

Dieticians work in public health clinics, hospitals, nursing homes, school systems, corporations, doctors' offices, and private practice. They often oversee meal planning and preparation. Dieticians also help people meet specific needs, such as planning meals for diabetic patients. They use diet to promote health and prevent disease.

Sodium and potassium The most abundant alkali metals are sodium and potassium. Sodium is used in sodium vapor lamps and as a heat exchanger in nuclear reactors. Because potassium is more reactive than sodium and more expensive to produce, elemental potassium has fewer industrial uses.

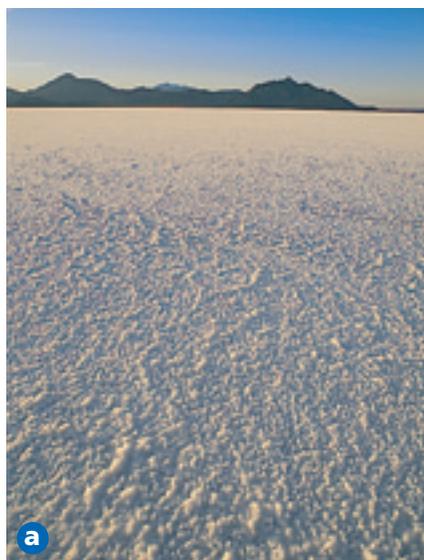
Humans and other vertebrates must have sodium and potassium in their diets because many biological functions are controlled by sodium and potassium ions. Potassium ions are the most common positive ions within cells. Sodium ions are the most common positive ions in the fluid that surrounds cells. When a nerve cell is stimulated, sodium ions flow into the cell and potassium ions flow out. This flow of ions across the cell membrane carries the nerve impulse along the cell. After the impulse passes, a compound attached to the cell membrane uses energy to move the ions back across the membrane so that they are in position for the next impulse.

Figure 7-4a shows a deposit of sodium chloride, or table salt—the most common sodium compound. Table salt occurs naturally in many foods. It is used to keep food from spoiling and to preserve foods for long-term storage, as shown in **Figure 7-4b**. These roles were especially important before the widespread use of refrigeration. Potassium chloride serves as a salt substitute for people whose intake of sodium must be limited. Potassium compounds are included in fertilizers because potassium is an important factor for plant growth and development. Potassium nitrate is used as an explosive for large-scale fireworks displays.

Other alkali metals The most reactive alkali metals—rubidium, cesium, and francium—have little commercial use. Rubidium, with a melting point of only 40°C, melts on a hot day. It will burst into flames if exposed to air. Francium, the most reactive alkali metal, is a rare radioactive element. For which SI base unit is cesium the atomic standard?

Figure 7-4

- a** The Bonneville Salt Flats in Utah formed when a vast land-locked salt lake evaporated.
- b** People started to use salt as a preservative in the days when the only food to which people had access in winter was the food they had dried, cured, or canned.



Group 2A: Alkaline Earth Metals

Medieval alchemists classified solids that did not melt in their fires as “earths.” Group 2A elements form compounds with oxygen, called oxides, that qualify as “earths” by this definition. Except for beryllium oxide, these oxides produce alkaline solutions when they react with water. The label alkaline earth reflects these two properties.

Alkaline earth metals are shiny solids that are harder than alkali metals. Although alkaline earth metals are less reactive than alkali metals, they are usually found combined with oxygen and other nonmetals in Earth’s crust. Alkaline earth metals lose their two valence electrons to form ions with a 2+ charge. Reactions with water reveal the relative reactivity of the alkaline earth metals. Calcium, strontium, and barium react vigorously with room temperature water. Magnesium will react in hot water. Beryllium does not appear to react with water. When exposed to oxygen, alkaline earth metals form a thin oxide coating. Most compounds of alkaline earth metals do not dissolve easily in water. Do the **miniLAB** on the next page to observe some properties of an alkaline earth metal.

Beryllium The lightest member of group 2A, beryllium, is found combined with aluminum, silicon, and oxygen in a material called beryl. **Figure 7-5** shows two highly prized forms of beryl. Finding aluminum and beryllium together is not surprising because these elements have a diagonal relationship and, thus, similar chemical properties. Beryllium is used to moderate neutrons in nuclear reactors. Tools made from an alloy of beryllium and copper are used in situations where a spark from steel tools touching steel equipment could cause a fire or explosion. For example, beryllium–copper tools are used in petroleum refineries.

Calcium Calcium is an essential element for humans, especially in maintaining healthy bones and teeth. Calcium is found widely in nature, mainly combined with carbon and oxygen in calcium carbonate. This compound is the main ingredient in rocks such as limestone, chalk, and marble. Coral reefs build up from calcium carbonate exoskeletons that are created by marine animals called corals. Calcium carbonate is used in antacid tablets and as an abrasive in toothpaste. An abrasive is a hard material used to polish, smooth, or grind a softer material. Emery boards and sandpaper are examples of abrasive materials.

2A	
Beryllium 4 Be [He]2s ²	
Magnesium 12 Mg [Ne]3s ²	
Calcium 20 Ca [Ar]4s ²	
Strontium 38 Sr [Kr]5s ²	
Barium 56 Ba [Xe]6s ²	
Radium 88 Ra [Rn]7s ²	

Figure 7-5

Beryllium-containing **a** emerald and **b** aquamarine get their brilliant colors from impurities—traces of chromium in emerald and traces of scandium in aquamarine.



Figure 7-6

Limestone is a relatively light building material that is easy to work with and durable.



Calcium carbonate as limestone was used to build the Roman aqueduct in **Figure 7-6**. When calcium carbonate decomposes, it forms an oxide of calcium called lime. Lime is one of the most important industrial compounds. For example, lime plays a role in the manufacture of steel, paper, and glass. Gardeners use lime to make soil less acidic. Wastewater treatment plants use lime, as do devices that remove pollutants from smokestacks. Lime is mixed with sand and water to form a paste called mortar.

Magnesium Magnesium is an abundant element that can be formed into almost any shape. Alloys of magnesium with aluminum and zinc are much lighter than steel but equally strong. The backpack frame in **Figure 7-7a** is made from a magnesium alloy, as are many bicycle frames and the “mag” wheels on the sports car in **Figure 7-7b**. The oxide of magnesium has such a high melting point that it is used to line furnaces. Plants cannot function

miniLAB

Properties of Magnesium

Observing and Inferring In this activity, you will mix magnesium with hydrochloric acid and observe the result.

Materials test tube, test tube rack, 10-mL graduated cylinder, hydrochloric acid, magnesium ribbon, sandpaper, cardboard, wood splint, safety matches

Procedure

Record all of your observations.

1. Place your test tube in a test tube rack. For safety, the test tube should remain in the rack throughout the lab.
2. Use a 10-mL graduated cylinder to measure out about 6 mL of hydrochloric acid. Pour the acid slowly into the test tube.
CAUTION: *If acid gets on your skin, flush with cold running water. Use the eyewash station if acid gets in your eye.*
3. Use sandpaper to clean the surface of a 3-cm length of magnesium ribbon.

4. Drop the ribbon into the acid and immediately cover the test tube with a cardboard lid.
5. As the reaction appears to slow down, light a wood splint in preparation for step 6.
6. As soon as the reaction stops, uncover the test tube and drop the burning splint into it.
7. Pour the contents of the test tube into a container specified by your teacher. Then rinse the test tube with water. Do not place your fingers inside the unwashed tube.

Analysis

1. Compare the appearance of the magnesium ribbon before and after you used the sandpaper. What did the sandpaper remove?
2. What happened when you placed the ribbon in the acid? How did you decide when the reaction was over?
3. What did you observe when you placed the burning splint in the test tube?
4. What gas can ignite explosively when exposed to oxygen in the air? (Hint: The gas is lighter than air.)

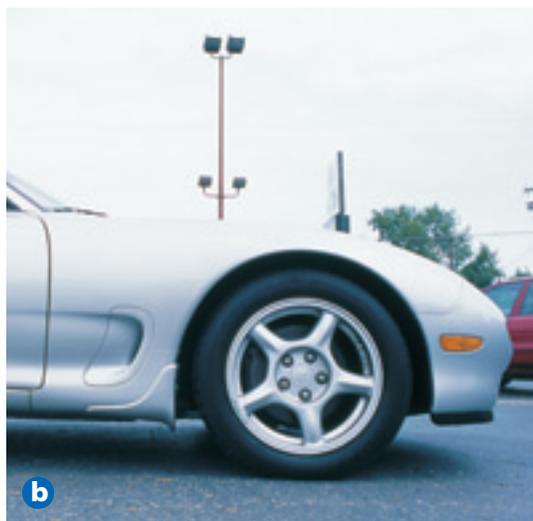


Figure 7-7

- a** The lighter the backpack frame, the more supplies a person can carry.
- b** Reducing a car's weight improves its engine performance, which is why sports cars were the first to use wheels made from a lightweight magnesium alloy.

without a supply of magnesium because each chlorophyll molecule contains a magnesium ion. Your body depends on magnesium ions, too; they play key roles in muscle function and metabolism.

When large quantities of calcium and magnesium ions are found in the water supply, the water is referred to as hard water. Hard water makes it difficult to wash oil from your hair or grease from your dishes because the ions interfere with the action of soaps and detergents. If there are large amounts of hydrogen carbonate ions in the water, they can combine with the calcium and magnesium ions to form deposits that can clog pipes, water heaters, and appliances such as steam irons. Devices called water softeners exchange sodium or hydrogen ions for the calcium and magnesium ions. Do the **CHEMLAB** at the end of the chapter to compare the cleaning ability of hard water and softened water.

Uses of other alkaline earth metals Strontium gives some fireworks their crimson color. Colorful barium compounds are used in paints and some types of glass. Barium also is used as a diagnostic tool for internal medicine. Radium is a highly radioactive element. Radium atoms emit alpha, beta, and gamma rays. Before people understood the danger, they used radium compounds to paint the hands on watches because paint containing radium glows in the dark.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Section 7.1 Assessment

1. Why are elements in groups 1A through 8A called representative elements?
2. What determines the chemical behavior of an element?
3. Why are alkali metals stored in oil?
4. What do group 1A and group 2A elements have in common? Give at least three examples.
5. What types of ions make water hard? What is the main problem with using hard water?
6. Name three factors that make magnesium a good choice for alloys.
7. **Thinking Critically** Lithium behaves more like magnesium than sodium. Use what you learned in Chapter 6 about trends in atomic sizes to explain this behavior.
8. **Applying Concepts** Hydrogen can gain one electron to reach a stable electron configuration. Why isn't hydrogen placed in group 7A with the other elements that share this behavior?

Objectives

- **Describe** and **compare** properties of p-block elements.
- **Define** allotropes and provide examples.
- **Explain** the importance to organisms of selected p-block elements.

Vocabulary

mineral
ore
allotrope

3A
Boron 5 B [He]2s ² 2p ¹
Aluminum 13 Al [Ne]3s ² 3p ¹
Gallium 31 Ga [Ar]4s ² 3d ¹⁰ 4p ¹
Indium 49 In [Kr]5s ² 4d ¹⁰ 5p ¹
Thallium 81 Tl [Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ¹

Differences among the s-block metals are slight in comparison to the range of properties found among the p-block elements. Remember that some p-block elements are metals, some are metalloids, and some are nonmetals. Some are solids and some are gases at room temperature. Even individual p-block elements display a greater range of properties; for example, many form more than one type of ion. As you might expect, the explanation for this property lies with the electron configurations, especially those configurations that are not close to a stable, noble-gas electron configuration.

Group 3A: The Boron Group

Group 3A elements are always found combined with other elements in nature. They are most often found as oxides in Earth's crust. This group contains one metalloid (boron), one familiar and abundant metal (aluminum), and three rare metals (gallium, indium, and thallium). Based on the group number, you would expect group 3A elements to lose three valence electrons to form ions with a 3+ charge. Boron, aluminum, gallium, and indium form such ions. Thallium does not. Thallium is the most metallic member of the group with properties similar to those of alkali metals. Thallium loses only the p valence electron to form ions with a 1+ charge. Gallium and indium also can form ions with a 1+ charge.

Boron Although group 3A is named for the metalloid boron, as with other groups, the lightest member is the least representative. Boron has more in common with silicon in group 4A than with the metallic members of group 3A. Boron and silicon oxides combine to form borosilicate glass, which can withstand extreme differences in temperature without shattering. This property makes borosilicate glass ideal for the laboratory equipment and cookware shown in **Figure 7-8**.

The main source of boron is a complex compound of boron called borax. About half of the world supply of borax comes from a large deposit in California's Mojave Desert. Borax is used as a cleaning agent and as fireproof insulation. Another compound of boron, boric acid, is used as a disinfectant and as an eye wash. A form of boron nitride is the second hardest known material; only diamond is harder. These materials are classified as superabrasives. They are used in grinding wheels, which shape manufactured parts and tools.

Figure 7-8

- How might using borosilicate glassware contribute to lab safety?
- Cookware made from borosilicate glass is designed to go directly from the refrigerator to the oven.



Aluminum Aluminum is the most abundant metal and the third most abundant element in Earth's crust. It usually occurs combined with oxygen or silicon. Because removing aluminum from its ore, bauxite, requires a great deal of energy, it is cost effective to recycle the aluminum in products such as the cans shown in **Figure 7-9**. Aluminum oxide is the major compound in bauxite. It is used as an abrasive, to strengthen ceramics, and in heat-resistant fabrics. Many gems, including ruby and sapphires, are crystals of aluminum oxide with traces of other metals. Chromium gives ruby its red color; iron and titanium give sapphire its bright blue color. The compound aluminum sulfate, known as alum, is used in antiperspirants and to remove suspended particles during water purification.

Gallium Gallium can literally melt in your hand. It is used in some thermometers because it remains a liquid over a wide temperature range—from 30°C to 2403°C. A compound of gallium and arsenic called gallium arsenide produces an electric current when it absorbs light. This property makes gallium arsenide ideal for the semiconductor chips used in light-powered calculators and solar panels. These chips are ten times more efficient than silicon-based chips. Read **How It Works** at the end of the chapter to learn more about semiconductors.

Scientists are using a compound of gallium and nitrogen, gallium nitride, to develop lasers that emit blue rather than red light. Using the shorter wavelengths of blue light would triple the storage capacity of a DVD, making room for three two-hour movies. Blue lasers could increase the speed and resolution of laser printers. Medical devices for detecting cancer cells could be less expensive if they used low-cost, blue-light lasers.

Group 4A: The Carbon Group

Based on the trends discussed in Chapter 6, the metallic properties of the elements in group 4A should increase as the atomic number increases. Carbon is a nonmetal; silicon and germanium are metalloids; tin and lead are metals. With such a wide range of properties, there are few rules that apply to all members of the group. One general trend does apply. The period-2 element, carbon, is not representative of the other elements within the group.

Carbon Group 4A contains one of the most important elements on Earth: carbon. Except for water and ions such as sodium, most substances that control what happens in cells contain carbon. The branch of chemistry that studies these carbon compounds is called organic chemistry. Until 1828 when the first organic compound was synthesized in the laboratory, scientists thought that organic compounds could be created only in organisms. Chapter 9 will help you understand how carbon can form so many different compounds. Chapters 22 through 24 will provide more details about organic chemistry.

The branch of chemistry that deals with all other compounds is called inorganic chemistry, meaning “not organic.” Carbonates, cyanides, carbides, sulfides, and oxides of carbon are classified as inorganic compounds. Geologists call these compounds minerals. A **mineral** is an element or inorganic compound that is found in nature as solid crystals. Minerals usually are found mixed with other materials in ores. An **ore** is a material from which a mineral can be removed at a reasonable cost. In other words, the cost of extraction cannot approach or exceed the economic value of the mineral.

Although both diamond and graphite contain only carbon atoms, they display different properties. Graphite is one of the softest known materials; diamond is one of the hardest. These different forms of the same element are



Figure 7-9

The energy used to recycle aluminum is about 5% of the energy needed to extract aluminum from its ore.

4A
Carbon 6 C $[\text{He}]2s^22p^2$
Silicon 14 Si $[\text{Ne}]3s^23p^2$
Germanium 32 Ge $[\text{Ar}]4s^23d^{10}4p^2$
Tin 50 Sn $[\text{Kr}]5s^24d^{10}5p^2$
Lead 82 Pb $[\text{Xe}]6s^24f^{14}5d^{10}6p^2$

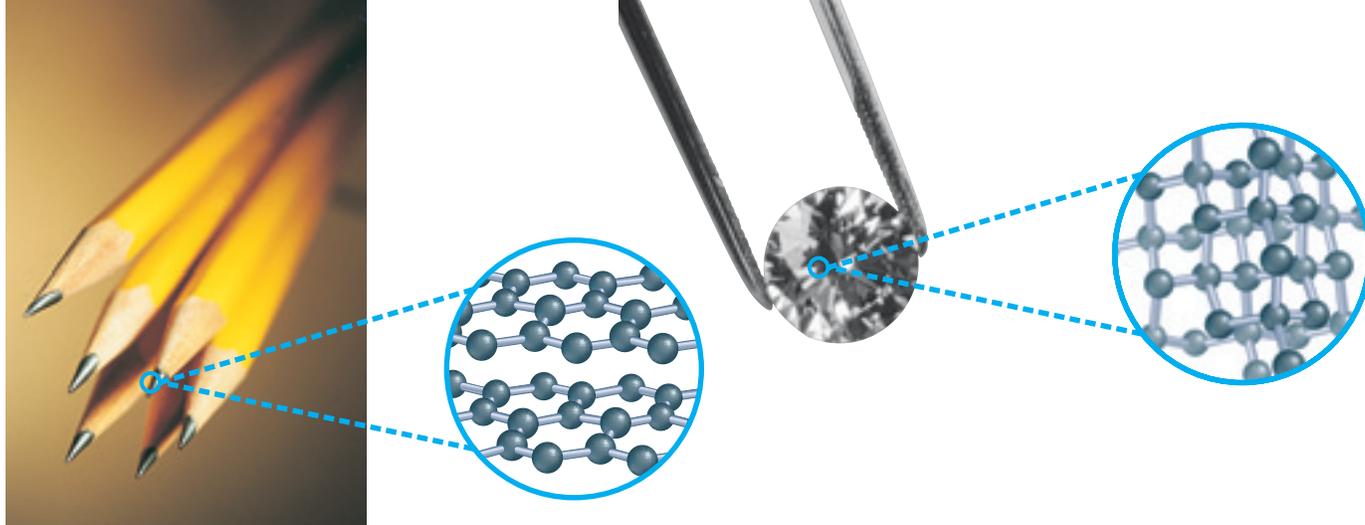


Figure 7-10

Comparing the diagrams to the photos should help explain why graphite and diamond have such different properties.

examples of allotropes. **Allotropes** are forms of an element in the same physical state—solid, liquid, or gas—that have different structures and properties. Because diamond and graphite are both solids made of carbon, they are allotropes. So is the amorphous form of carbon found in coal.

As shown in **Figure 7-10**, each carbon atom in graphite shares electrons with three other carbon atoms to form flat layers that can slide over one another. When you use a pencil, layers of carbon atoms slide onto your paper. This “slippery” property makes graphite a good lubricant. In diamond, each carbon atom shares electrons with four other carbon atoms to form a three-dimensional solid. Because of this arrangement, diamonds on grinding wheels are hard enough to sharpen tools and cut through granite or concrete.

Silicon Silicon, which is used in computer chips and solar cells, is the second most abundant element in Earth’s crust after oxygen. Silicon occurs most often combined with oxygen in the compound silicon dioxide, which also is known as silica. Silica can be found in the quartz crystals, sand, and glass shown in **Figure 7-11**. When rocks containing quartz crystals weather, they produce white sand, which is the type of sand found on most beaches. If white sand is melted and allowed to cool rapidly, glass forms. A glass blower can shape glass into many different forms.

A compound of silicon and carbon, silicon carbide, is a major industrial abrasive. Its common name is carborundum. People who have a home workshop use silicon carbide sticks to sharpen their tools.

Figure 7-11

a Quartz crystals, **b** white sand, and **c** glass all contain silicon dioxide.





Figure 7-12

- a** The bronze used in this statue of a bull located in New York City's financial district is an alloy of tin and copper. **b** This pewter vase is lead free. About 8% antimony is mixed with tin to make the pewter harder.

Lead and tin Recall from the chapter opener that the metals tin and lead are similar with one key difference. Lead is toxic. The uses of pure tin are limited because tin is softer than most metals. For years, tin was used as a coating to keep steel cans from rusting, but aluminum cans are now more common. The decorative items in **Figure 7-12** are made from bronze and pewter. Bronze is an alloy of tin and copper with copper predominating. Some zinc is included to make the alloy harder. Pewter was made from about 40% lead and 60% tin until the toxic effects of lead became known.

Lead may have been the first pure metal obtained from its ores because of its low melting point. Analysis of ancient skeletons shows significant levels of lead. Until people realized the dangers of lead poisoning, lead was used for eating utensils, pipes for plumbing, as an additive in gasoline, and in paint. Paint at the hardware store no longer contains lead, but in older houses with layers of paint, lead may still be present. Lead is especially harmful for preschool children who may inhale it in dust or chew on scraps of peeling paint. In many states, a landlord cannot rent to a family with young children until the lead paint is removed. Currently, the major use of lead is in storage batteries for automobiles, which you will learn about in Chapter 21.

Group 5A: The Nitrogen Group

Although the nonmetals (N and P), metalloids (As and Sb), and metal (Bi) in group 5A each have five valence electrons, they display a wide variety of physical and chemical properties. Nitrogen, which you inhale with every breath, forms some of the most explosive compounds known. Phosphorus has three solid allotropes. Antimony and bismuth expand when they change from a liquid to a solid. Nitrogen can gain three electrons and form ions with a 3− charge; bismuth can lose three electrons and form ions with a 3+ charge.

Nitrogen Nitrogen gas is colorless, odorless, and relatively unreactive. About 78% of Earth's atmosphere is nitrogen. Proteins and other essential organic compounds contain nitrogen. However, most organisms cannot use the nitrogen in air to create these compounds. Bacteria in the soil and roots of plants such as clover convert, or “fix,” molecular nitrogen into nitrogen compounds. These compounds are then used by plants and the animals that consume plants. Some essential organic compounds contain both nitrogen and phosphorus. In Chapter 24, you will learn how these compounds control the movement of energy and transfer of genetic information within cells.

5A
Nitrogen 7 N [He]2s ² 2p ³
Phosphorus 15 P [Ne]3s ² 3p ³
Arsenic 33 As [Ar]4s ² 3d ¹⁰ 4p ³
Antimony 51 Sb [Kr]5s ² 4d ¹⁰ 5p ³
Bismuth 83 Bi [Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ³

History

CONNECTION

The Swedish chemist Alfred Nobel (1833–1896) invented a safe way to use nitroglycerine. He mixed liquid nitroglycerine with sand to form a paste that could be shaped into rods. Nobel patented the mixture, which he called dynamite, and the blasting caps used to detonate the rods. The invention of dynamite greatly reduced the cost of drilling road and railway tunnels through rock. Nobel left most of his money to endow annual prizes in peace, chemistry, physics, medicine, and literature. These prizes are called Nobel Prizes.



As you learned in Section 7.1, the major industrial use of hydrogen is in the production of the nitrogen compound ammonia, which is a colorless gas with an irritating odor. Many cleaning products contain ammonia. Liquid ammonia can be used as a source of nitrogen for plants. It is stored under pressure and pumped directly into the soil. About 25% of ammonia is converted into nitric acid, which is used to produce explosives, dyes, and solid fertilizers. Artists use nitric acid to etch designs into metal plates. When ink is applied to the etched plate and then transferred to paper, a mirror image of the design appears. Some nitrogen compounds are extremely unstable. These include TNT (trinitrotoluene) and nitroglycerine.

Phosphorus Two of the three solid allotropes of phosphorus are shown in **Figure 7-13**. White phosphorus bursts into flames in air and must be stored in water. Red phosphorus is less reactive. It forms when white phosphorus is heated in the absence of air. Red phosphorus is used on the striking surface of matchboxes. The third allotrope, black phosphorus, is produced when either red or white phosphorus is heated under high pressure.

Most phosphorus is used to make phosphoric acid, which in turn is converted into phosphate compounds. These compounds have many uses. Some are found in processed cheese, laxatives, and baking powders. Others are used as a flame-retardant coating for fabrics and as a grease remover in cleaning products. Because phosphorus is essential for plant growth, fertilizers often contain phosphates. Do the **problem-solving LAB** on the next page to explore the role of fertilizers in the economics of farming.

The use of fertilizers containing phosphates can be harmful to the environment. In a lake, bacteria break down waste products and dead organisms into nutrients. Algae feed on these nutrients, which include phosphate ions. Small animals called zooplankton eat the algae; fish eat the zooplankton. If phosphate ions from fertilized fields reach the lake, they cause an increase in algae. Layers of algae form on the surface of the lake and keep light from reaching the algae below, which cannot survive if they do not have light for photosynthesis. Bacterial decay of dead algae uses large amounts of the dissolved oxygen in the water. There is not enough oxygen left to sustain other organisms. Over a long period of time, if algal growth is left unchecked, layers of waste can slowly build up on the lake bottom and the lake can evolve into a swamp. Detergents that are sold in the United States no longer contain phosphate compounds because phosphate ions can be released into water by sewage treatment plants.

Figure 7-13

White phosphorus is toxic and flammable. Why must white phosphorus be stored in water? Red phosphorus is an amorphous solid, which means that there is no definite pattern to the arrangement of its atoms.



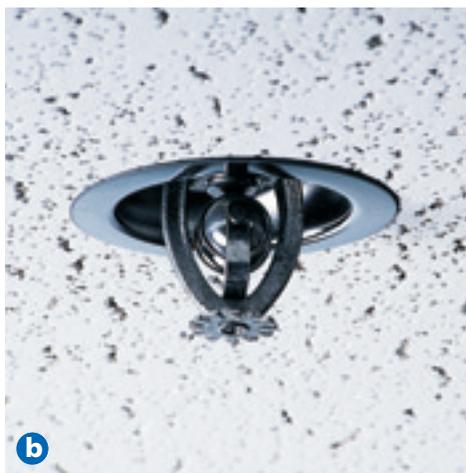


Figure 7-14

a Britannia metal is harder than pewter. This tin-antimony alloy's properties can be varied by the addition of zinc, copper, lead, or bismuth. **b** When heat from a fire melts the Wood's metal plug in a sprinkler head, water that was held back by the plug is freed.

Arsenic, antimony, and bismuth These group 5A elements are less abundant than nitrogen and phosphorus. However, they are among the oldest known elements. Although arsenic is toxic, for centuries physicians used small amounts of an arsenic and sulfur compound, arsenic sulfide, to treat some illnesses. A black compound of antimony and sulfur, antimony sulfide, was used as a cosmetic to darken eyebrows and make the eyes appear larger. Britannia metal, an alloy of tin and antimony, can be shaped by stamping or spinning, and cast in molds, as shown by the tableware in **Figure 7-14a**. Today, lead storage batteries contain 5% antimony. A compound of bismuth is the active ingredient in a popular, pink remedy for diarrhea and nausea. A low-melting alloy of bismuth, lead, tin, and cadmium called Wood's metal is used as a plug in automatic sprinkler systems, as shown in **Figure 7-14b**.

problem-solving LAB

Cost of Fertilizer

Interpreting Data Farmers deduct expenses from income to determine profit. The United States Department of Agriculture (USDA) collects annual data on production expenses from United States farmers, including data on fertilizer.

Analysis

1. What percent of the total expenses was spent on fertilizer and lime for each of the years shown? Round your answers to three significant figures.
2. What trend is revealed by your answers to question 1?

Thinking Critically

3. The table shows that farmers spent more money each year on fertilizer from 1993 to 1996. From this data, can you conclude that farmers increased their use of fertilizer at the same rate? Give a reason for your answer.

Production expenses (in millions)

Year	Total expenses (\$)	Fertilizer* (\$)
1993	160 548	8398
1994	167 547	9180
1995	174 161	10 033
1996	181 303	10 934

* This category includes money spent on lime, which is a soil conditioner.

4. Suppose the USDA data were altered to exclude farmers who raise livestock. Predict how this change would affect expenditures for fertilizer and lime as a percent of total expenses.
5. Use the Internet to find out if data since 1996 has confirmed the trend you identified in question 2.

Group 6A: The Oxygen Group

6A
Oxygen 8 O [He]2s ² 2p ⁴
Sulfur 16 S [Ne]3s ² 3p ⁴
Selenium 34 Se [Ar]4s ² 3d ¹⁰ 4p ⁴
Tellurium 52 Te [Kr]5s ² 4d ¹⁰ 5p ⁴
Polonium 84 Po [Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁴

Polonium is the most metallic member of group 6A. But it is not a typical metal. It is rare, radioactive, and extremely toxic. Polonium is important historically because it was discovered by Marie and Pierre Curie in 1898 and named for Marie's native land, Poland. Selenium and tellurium are metalloids; oxygen and sulfur are nonmetals. There are some trends to note in group 6A. With six valence electrons, the elements act mainly as nonmetals. They tend to gain two electrons to form ions with a 2⁻ charge; they also can share two electrons to achieve a stable electron configuration.

Oxygen Oxygen has two allotropes. You studied ozone (O₃) in Chapter 1. Remember that ozone is an unstable gas with a pungent odor and that it decomposes when exposed to heat or ultraviolet light. The odor of ozone is noticeable during electrical storms and near high-voltage motors such as those used in subway stations. Ozone produced in automobile emissions can irritate eyes, harm lung cells, and affect plant growth.

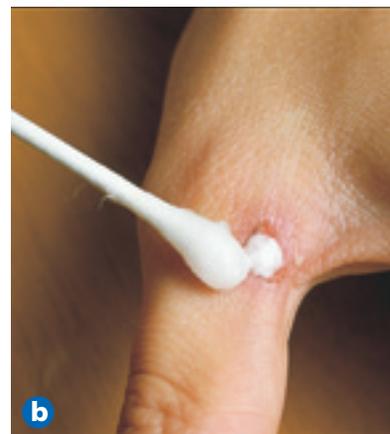
The allotrope that makes up about 21% of Earth's atmosphere is a colorless, odorless gas (O₂). Joseph Priestley (1733–1804) usually is credited with the discovery of oxygen. When he heated an oxide of mercury, the gas produced caused a candle to burn more brightly than it would in air. This experiment showed that air was a mixture and that one gas in air, not air as a whole, was responsible for combustion. When fuels burn, they release energy to heat homes, run automobiles, and operate machinery. The oxygen produced by plants during photosynthesis is used by both plants and animals during cellular respiration—the process that releases energy from carbohydrates. Like all organisms, the fish in **Figure 7-15a** needs a constant supply of oxygen.

Oxygen is separated from the other gases in air through a distillation process that is based on differences in boiling point among the gases. Oxygen is stored as a liquid under pressure in cylinders. These cylinders are too massive for all situations where oxygen may be needed, for example, on an airplane. Instead, a small canister is stored above each seat on the plane. The canister contains chemicals that can react to produce oxygen in an emergency.

Oxygen is the most abundant element in Earth's crust. It is found combined with metals in silicates and carbonates. Based on their names, what other elements are found in these compounds? Oxygen forms at least one oxide with every element except helium, neon, and argon. Oxygen combines with hydrogen to form two oxides—water and hydrogen peroxide, which is used as a bleach. **Figure 7-15b** shows another use of hydrogen peroxide. Oxygen also forms two oxides with carbon—carbon monoxide and carbon dioxide. Which oxide is essential to life and which is life-threatening?

Figure 7-15

- a** Fish use their gills to remove oxygen from water.
- b** A 3% solution of hydrogen peroxide in water can help disinfect an open wound.



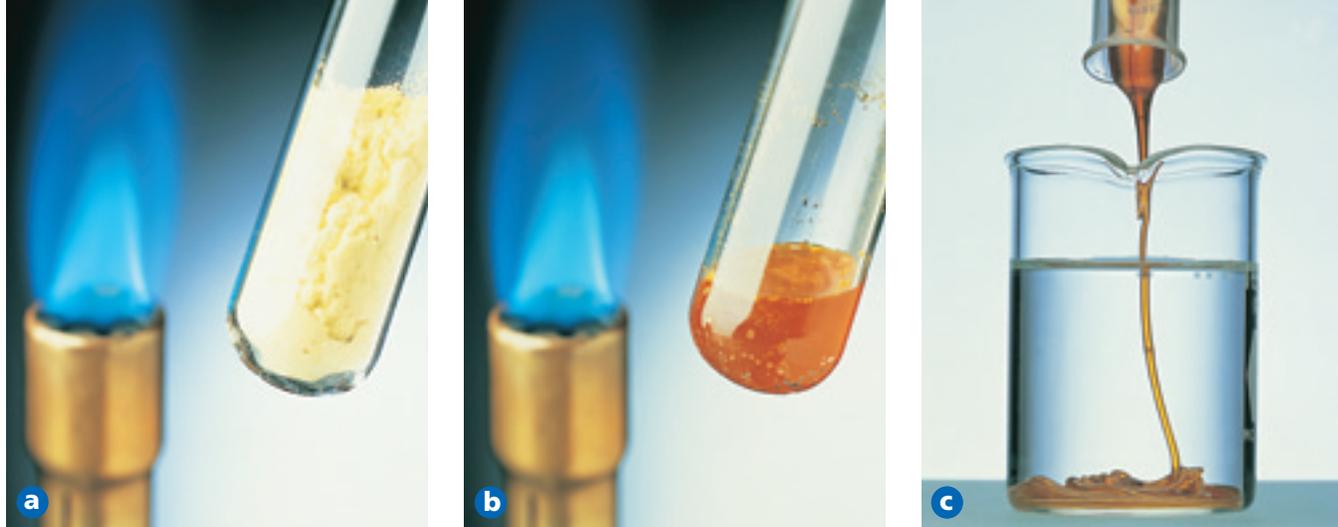


Figure 7-16

- a** When brittle, yellow sulfur is heated, **b** it melts at 112°C .
- c** Further heating produces a thick, red-brown syrup. When the syrup is poured into cold water, an amorphous solid forms. This allotrope is called plastic sulfur because it stretches like a rubber band.

Sulfur Sulfur can be found combined with mercury in cinnabar or with lead in galena. It also is found uncombined in underground deposits. When water heated to 160°C and compressed air are pumped into a deposit, the yellow solid melts and is forced to the surface. Sulfur has ten allotropes, more than carbon and oxygen combined. **Figure 7-16** shows how a brittle allotrope that can be ground into a powder can be changed into one that is elastic.

When sulfur is burned in air, it reacts with oxygen to form sulfur dioxide, which is used to preserve fruit and as an antibacterial agent. Sulfur dioxide released into the atmosphere reacts with water vapor to form one of the acids in acid rain. More than 90% of the sulfur dioxide produced is used to make sulfuric acid. This inexpensive acid is used by so many industries that the amount of it produced can indicate the strength of an economy. About half the sulfuric acid produced is used to make fertilizers. Products ranging from steel to paper to paints also depend on this acid.

Scientists have discovered places in the deep ocean where heat from Earth's interior is released through vents, or openings, in Earth's crust. Because light does not reach the depth shown in **Figure 7-17**, the food chain cannot begin with photosynthesis. Instead, it begins with hydrogen sulfide, which also is released through the vents. (Hydrogen sulfide gives a rotten egg its vile odor.) Some unusual bacteria that live near these vents use hydrogen sulfide as an energy source. Other organisms feed on these bacteria. Volcanoes also release hydrogen sulfide from Earth's interior into the atmosphere. When silver dulls, this is a sign that it has reacted with atmospheric hydrogen sulfide to form silver sulfide, which is called tarnish.



Figure 7-17

Heat and hydrogen sulfide released at deep ocean vents support many organisms.



Figure 7-18

What connection might there be between the name of the locoweed plant and the plant's effect on an animal's behavior?

7A
Fluorine 9 F [He]2s ² 2p ⁵
Chlorine 17 Cl [Ne]3s ² 3p ⁵
Bromine 35 Br [Ar]4s ² 3d ¹⁰ 4p ⁵
Iodine 53 I [Kr]5s ² 4d ¹⁰ 5p ⁵
Astatine 85 At [Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁵

Figure 7-19

Three states of matter are represented by solid iodine (left), chlorine gas (middle), and liquid bromine (right) at room temperature.

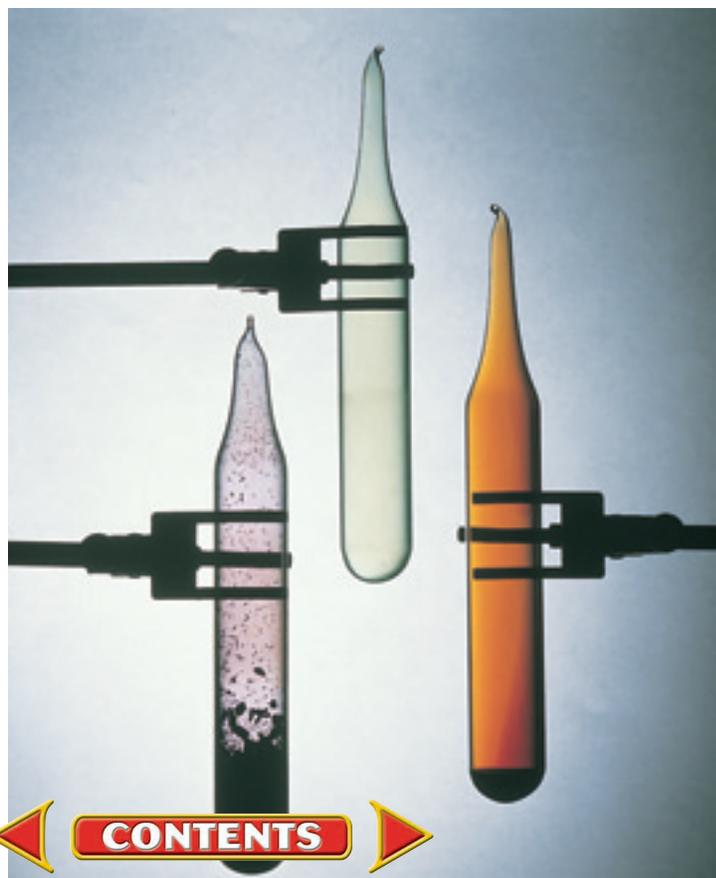
Selenium Some people supplement their diet with tablets that contain essential vitamins and minerals. These supplements may include a small amount of sodium selenate. Selenium also can be found in such foods as fish, eggs, and grains. Selenium works with vitamin E to prevent cell damage. It may help to inhibit the growth of cancer cells. However, in the case of nutrients, more is not always better. The locoweed plant shown in **Figure 7-18** provides an example of this principle related to selenium. When a locoweed plant absorbs selenium from the soil, the concentration of selenium increases to a toxic level. Grazing animals that feed on locoweed can become quite ill.

Because selenium can convert light into electricity, it is used in solar panels. Selenium's ability to conduct electricity increases as its exposure to light increases. Meters that photographers use to measure the level of available light contain selenium. Photocopiers work because charged particles of selenium create an "image" of the item being copied. Selenium also is used in semiconductors, as is tellurium, which is a relatively rare element.

Group 7A: The Halogens

The elements in group 7A are named for their ability to form compounds with almost all metals. Because these compounds are called salts, the group 7A elements are called "salt formers," or halogens. You are familiar with one salt, sodium chloride, which is known as table salt. The halogens differ in their physical properties, as shown in **Figure 7-19**. Chlorine is a gas at room temperature. Bromine is a liquid, but it evaporates easily. Iodine is a solid that can change directly into a vapor.

The chemical behavior of the halogens is quite similar with one exception. Astatine is a radioactive element with no known uses. The other halogens share the following general properties. They are reactive nonmetals that are always found combined with other elements in nature. Because they have seven valence electrons, halogens tend to share one electron or gain one electron to attain a stable, noble-gas electron configuration. They tend to form ions with a 1− charge.





Fluorine Because fluorine is the halogen with the lowest atomic number, it has a small atom that provides little shielding of its valence electrons from the nucleus. Fluorine is the most electronegative element on the periodic table; that is, it has the greatest tendency to attract electrons. Thus, it is logical that fluorine also is the most active of all elements. In fact, it reacts with every element except helium, neon, and argon.

Fluorine comes from the Latin word *fluere*, which means “to flow.” The mineral fluorite, which contains fluorine and calcium, is used to lower the melting points of other minerals to make it easier to separate them from their ores. Fluorine compounds are added to toothpaste and drinking water to protect tooth enamel from decay. A compound of fluorine and carbon provides a non-stick coating for the cookware shown in **Figure 7-20a**. When fluorine reacts with isotopes of uranium, the gases that form are separated by differences in mass. This process is called uranium enrichment; it provides the uranium-235 fuel for nuclear reactors.

Chlorine Chlorine reacts with nearly all of the elements. Although chlorine is a deadly gas, compounds of chlorine have many uses, including some that can save lives. In 1848, a cholera epidemic began in London. About 25 000 people died during the epidemic, which was blamed on raw sewage flowing into the river Thames. In 1855, London became the first city to use chlorine compounds to disinfect sewage.

Chlorine compounds are used as bleaching agents by the textile and paper industries. Homeowners use chlorine bleach to remove stains from clothing. The substances dentists use to block the nerves that carry pain signals to the brain often are chlorine compounds. Hydrochloric acid in your stomach helps you digest food; this acid also is used to remove rust from steel in a process called pickling. Much of the chlorine gas produced is combined with products from oil refineries to make plastics such as polyvinyl chloride (PVC). Floor tiles, pipes for indoor plumbing, and the garden hose in **Figure 7-20b** are a few of the products made from PVC.

Bromine and iodine Compared to chlorine, much less bromine and iodine are produced annually because there are fewer commercial uses for their compounds. Silver bromide and silver iodide are used to coat photographic film. Your body needs iodine to maintain a healthy thyroid gland. This gland produces hormones that control growth and your metabolic rate—the speed at which biochemical reactions occur. A lack of iodine causes the thyroid gland to enlarge, a condition called goiter. Seafood is an excellent source of iodine. So is iodized salt, which contains potassium iodide or sodium iodide in addition to sodium chloride. Because iodine kills bacteria, campers use iodine tablets or crystals to disinfect water.

Figure 7-20

- a** Cookware with a non-stick surface is easy to clean and can reduce the use of cooking oils and fats. **b** A garden hose made from polyvinyl chloride is sturdy and flexible.

8A
Helium 2 He $1s^2$
Neon 10 Ne $[\text{He}]2s^22p^6$
Argon 18 Ar $[\text{Ne}]3s^23p^6$
Krypton 36 Kr $[\text{Ar}]4s^23d^{10}4p^6$
Xenon 54 Xe $[\text{Kr}]5s^24d^{10}5p^6$
Radon 86 Rn $[\text{Xe}]6s^24f^{14}5d^{10}6p^6$

Group 8A: Noble Gases

The noble gases were among the last naturally occurring elements to be discovered because they are colorless and unreactive. Scientists assumed that noble gases could not form compounds. In 1962, the inorganic chemist Neil Bartlett created compounds from xenon and fluorine. Why do you think he chose fluorine for his experiments?

Despite this breakthrough, group 8A elements are still known primarily for their stability. Remember that noble gases have the maximum number of electrons in their outermost energy levels—eight—except for helium, which has two. Noble gases rarely react because of their stable electron configurations. In fact, there are no known compounds of helium, neon, or argon.

Helium The lightest noble gas, helium, was discovered first in the emission spectrum of the Sun. Although helium is light enough to escape Earth's gravity, it can be found on Earth in natural-gas wells. Helium is the lighter-than-air gas used in blimps, airships, and balloons. A mixture of helium and oxygen is used by deep-sea divers. By replacing the nitrogen in air with helium, divers can return to the surface quickly without experiencing a painful condition called the "bends." Liquid helium is used as a coolant for superconducting magnets.

Neon Neon is used in light displays that are commonly referred to as neon lights. When high-voltage electricity passes through the neon gas stored in a gas discharge tube, electrons in the atoms become excited. When the electrons return to a lower energy state, the atoms emit a bright orange light. The color of neon lights is not a constant because gases other than neon can be used in the displays. For example, argon emits blue light and helium emits a pale yellow light.

Argon and krypton Argon is the most abundant of the noble gases on Earth; it makes up about 1% of Earth's atmosphere. Argon provides an inert atmosphere for procedures such as high-temperature welding. This substitution avoids the dangerous mixture of electrical sparks, heat, and oxygen. Argon and krypton are used to prolong the life of filaments in incandescent light bulbs and as a layer of insulation between panes of glass. In Chapter 25, you will learn about another noble gas, the radioactive gas radon, which can be dangerous when inhaled.

Section 7.2 Assessment

- In general, how do p-block elements differ from s-block elements?
- How do animals obtain the nitrogen they need to build compounds such as proteins?
- Explain why noble gases were among the last naturally occurring elements to be discovered.
- What is an allotrope? Describe two allotropes of carbon.
- Compare the physical and chemical properties of fluorine, chlorine, bromine, and iodine.
- How do a mineral and an ore differ?
- Thinking Critically** Although carbon and lead are in the same group, one is a nonmetal and the other a metal. Explain how two elements with the same number of valence electrons can have such different properties.
- Predicting** A place has been left for element 113 on the periodic table. To what group does it belong? How many valence electrons will it have? What element will it most closely resemble?

Properties of d-Block and f-Block Elements

Beginning with period 4, the periodic table is expanded to make room for the elements whose d or f orbitals are being filled. These elements are called transition elements. They are subdivided into d-block and f-block elements—the transition metals and inner transition metals, respectively. Note that the groups of transition elements are labeled “B” to distinguish them from the groups containing representative elements. Recall that a transition metal is any element whose final electron enters a d sublevel. An inner transition metal is any element whose final electron enters an f sublevel.

Figure 7-21 reviews the locations of the d-block and f-block elements on the periodic table. Recall that the f-block elements are placed below the main body of the periodic table with an arrow to indicate their proper locations on the table. The f-block elements are further divided into two series of elements that reflect those locations. The f-block elements from period 6 are named the **lanthanide series** because they follow the element lanthanum. The f-block elements from period 7 are named the **actinide series** because they follow the element actinium.

Transition Metals

Transition metals share properties such as electrical conductivity, luster, and malleability with other metals. There is little variation in atomic size, electronegativity, and ionization energy across a period. However, there are differences in properties among these elements, especially physical properties. For example, silver is the best conductor of electricity. Iron and titanium are used as structural materials because of their relative strength.

The physical properties of transition metals are determined by their electron configurations. Most transition metals are hard solids with relatively high melting and boiling points. Differences in properties among transition metals are based on the ability of unpaired d electrons to move into the valence level. The more unpaired electrons in the d sublevel, the greater the hardness and the higher the melting and boiling points.

d-block transition metals

f-block inner transition metals

Objectives

- **Compare** the electron configurations of transition and inner transition metals.
- **Describe** the properties of transition elements.
- **Explain** why some transition metals form compounds with color and some have magnetic properties.

Vocabulary

lanthanide series
actinide series
ferromagnetism
metallurgy

Figure 7-21

In which periods are the d-block and f-block elements located?

Figure 7-22

From left to right, the compounds contain the following transition metals: scandium (white), titanium (white), vanadium (light blue), chromium (yellow), manganese (light pink), iron (red-orange), cobalt (violet), nickel (green), copper (blue), and zinc (white).



Consider, for example, the period 4 transition metals. Moving from left to right across the period, scandium has one unpaired d electron, titanium has two, vanadium has three, and manganese has five unpaired d electrons. Chromium, with six unpaired electrons—five unpaired d electrons and one unpaired s electron—is the hardest, and has a high melting point. Iron, cobalt, nickel, and copper form pairs of d electrons. Thus, their melting points, and hardness decrease from left to right. Zinc has the lowest melting and boiling points and is a relatively soft metal because its 3d and 4s orbitals are completely filled.

Formation of ions Transition metals can lose two s electrons and form ions with a 2+ charge. Because unpaired d electrons can move to the outer energy level, these elements also can form ions with a charge of 3+ or higher. When a transition metal reacts with a highly electronegative element such as fluorine or oxygen, the positive ions formed can have a charge as high as 6+.

Figure 7-22 shows a compound of each transition metal in period 4 in order from left to right. Note that most of the compounds have color. The metal ions in these compounds have partially filled d sublevels. Electrons in these sublevels can absorb visible light of specific wavelengths. The exceptions are the white compounds that contain scandium, titanium, or zinc. The scandium and titanium ions have an empty d sublevel. The zinc ion has a completely filled and stable d sublevel. Electrons in zinc, scandium, and titanium can be excited to higher levels, but not by wavelengths of visible light. The energy required corresponds to wavelengths in the ultraviolet range of the electromagnetic spectrum.

For those transition metals that can form more than one type of ion, a change from one to another often can be detected by a color change, as shown in **Figure 7-23**. There are vanadium ions in all four solutions. The ions in the yellow solution have a 5+ charge. When zinc is added to the solution, a reaction occurs that changes the charge on vanadium from 5+ to 4+. This change is indicated by a color change in the solution from yellow to blue. If the reaction is allowed to continue, the charge changes from 4+ to 3+ and the color of the solution changes from blue to green. The final change in charge is from 3+ to 2+ with a corresponding color change from green to violet.

Figure 7-23

Each solution contains different vanadium ions.



Magnetism and metals The ability of a substance to be affected by a magnetic field is called *magnetism*. A moving electron creates a magnetic field. Because paired electrons spin in opposite directions, their magnetic fields tend to cancel out. When all of the electrons in atoms or ions are paired, the substance is either unaffected or slightly repelled by a magnetic field. This property is called *diamagnetism*. When there is an unpaired electron in the valence orbital of an atom or ion, the electron is attracted to a magnetic field. This property is called *paramagnetism*. Most substances act as temporary magnets; that is, their magnetic properties disappear after the magnetic field is removed.

Now recall the materials you tested in the **DISCOVERY LAB**. Did any of them contain iron? The transition metals iron, cobalt, and nickel have a property called ferromagnetism. **Ferromagnetism** is the strong attraction of a substance to a magnetic field. **Figure 7-24** shows how the ions in a ferromagnetic metal respond to a magnetic field. The ions align themselves in the direction of the field. When the field is removed, the ions stay aligned and the metal continues to act as a magnet. Thus, iron, cobalt, and nickel can form permanent magnets.

Sources of transition metals Copper, silver, gold, platinum, and palladium are the only transition metals that are unreactive enough to be found in nature uncombined with other elements. All other transition metals are found in nature combined with nonmetals in minerals such as oxides and sulfides. Recall that minerals are mixed with other materials in ores. **Metallurgy** is the branch of applied science that studies and designs methods for extracting metals and their compounds from ores. The methods are divided into those that rely on high temperatures to extract the metal, those that use solutions, and those that rely on electricity. Electricity also is used to purify a metal extracted by high temperatures or solutions.

Figure 7-25 shows a blast furnace in which iron is extracted from its ore. Ore enters the furnace at the top, where the temperature is about 200°C. As the ore travels down through the furnace, carbon monoxide reacts with compounds in iron ore to remove the iron. The temperature increases along the way until molten iron at about 2000°C flows from the bottom. The product, which is called pig iron, contains about 3%–4% carbon. Most pig iron is purified and mixed with other elements in alloys called steel.

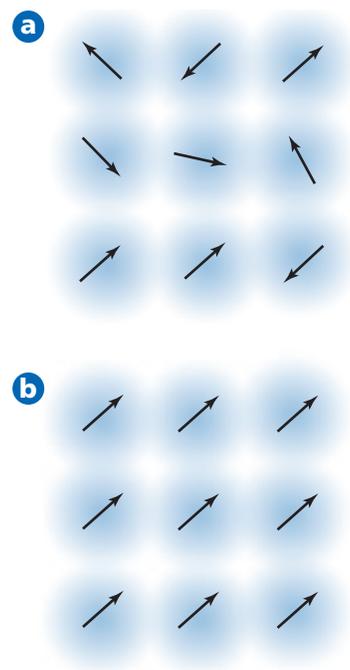


Figure 7-24

a Each ion in a ferromagnetic metal acts as a magnet. **b** If all of the ions are aligned in the same direction, they can form a permanent magnet.



Figure 7-25

This blast furnace is named for the blast of hot air that moves rapidly up the furnace. The hot air carries carbon monoxide, which reacts with compounds in iron ore to remove the iron.



Figure 7-26

Are any strategic materials found in only one location?

Transition metals play a vital role in the economy of many countries because they have a wide variety of uses. As the uses of transition metals increase, so does the demand for these valuable materials. Ores that contain transition metals are located throughout the world, as shown in **Figure 7-26**. The United States now imports more than 60 materials that are classified as “strategic and critical” because the economy and the military are dependent on these materials. The list includes platinum, chromium, cobalt, manganese, and tungsten.

Uses of transition metals Copper is used in electrical wiring, zinc as a protective coating for other metals, and iron in making steel. Many transition metals are found in alloys used to make items such as jet engines, drill bits, surgical instruments, and armor. The plastics, petroleum, and food industries use transition metals such as platinum, palladium, and nickel to control the conditions at which a reaction will occur. In Chapter 17, you will learn how some elements and compounds affect the rate of reactions and increase the amount of products produced in a reaction.

Your body needs large amounts of a few elements to function: carbon, oxygen, hydrogen, nitrogen, sulfur, phosphorus, sodium, potassium, calcium, magnesium, and chlorine. Other elements are essential, but are required only in trace amounts. Many trace elements are transition metals. Except for scandium and titanium, all period 4 transition metals play vital roles in organisms. An iron ion is in the center of each hemoglobin molecule. Hemoglobin picks up oxygen from blood vessels in the lungs and carries it to cells throughout the body. Molecules that help your body digest proteins and eliminate carbon dioxide contain zinc. Manganese and copper are involved in cell respiration. Cobalt is needed for the development of red blood cells.

Table 7-1 shows some of the elements included in vitamin and mineral supplements. The contents and quantities vary among supplements. Daily value (DV) means the amount of each element recommended daily. Which elements would qualify as trace elements?

Table 7-1

Supplement Contents Per Tablet		
Element	Amount	% DV
Calcium	161 mg	16%
Chromium	25 μg	20%
Copper	2 mg	100%
Iodine	150 μg	100%
Manganese	2.5 mg	125%
Magnesium	100 mg	25%
Molybdenum	25 μg	33%
Phosphorus	109 mg	11%
Potassium	40 mg	1%
Selenium	20 μg	28%
Zinc	15 mg	100%



Figure 7-27

- a** What protection do the lanthanide compounds in tinted sunglass lenses provide?
- b** This smoke detector needs to be replaced after about 10 years because the amount of americium steadily decreases.

Inner Transition Metals

The inner transition metals are divided into two groups: the period 6 lanthanide series and the period 7 actinide series.

Lanthanide series Lanthanides are silvery metals with relatively high melting points. Because there is so little variation in properties among inner transition metals, they are found mixed together in nature and are extremely hard to separate. The name of one lanthanide, dysprosium, comes from a Greek word meaning “hard to get at.” Lanthanide ores are mined in Ytterby, Sweden. Which four elements are named for this town?

The glass in welder’s goggles contains neodymium and praseodymium, which absorb high-energy radiation that can damage the eyes. Oxides of yttrium and europium are found in television screens and color computer monitors. The yttrium and europium ions in the oxides emit bright red light when excited by a beam of electrons. An alloy called misch metal, which is 50% cerium, is used by the steel industry to remove carbon from iron and steel. Compounds of lanthanides are used in movie projectors, high-intensity searchlights, lasers, and tinted sunglasses such as those in **Figure 7-27a**.

Actinide series Actinides are radioactive elements. Only three actinides exist in nature. The rest are synthetic elements called transuranium elements. A transuranium element is an element whose atomic number is greater than 92, the atomic number of uranium. Transuranium elements are created in particle accelerators or nuclear reactors. Most transuranium elements decay quickly. One notable exception is plutonium-239. A sample of this isotope can remain radioactive for thousands of years. Plutonium is used as a fuel in nuclear power plants. The home smoke detector in **Figure 7-27b** uses americium.

Section 7.3 Assessment

17. How do the electron configurations of transition and inner transition metals differ?
18. Why do transition metals share properties with other transition metals in their period?
19. How do transuranium elements differ from other inner transition metals?
20. Explain how some transition metals can form ions with more than one charge.
21. What factor determines the magnetic properties of an element and the color of its compounds?
22. What is metallurgy?
23. Compare and contrast the lanthanide series and the actinide series.
24. **Thinking Critically** Why is silver not used for electrical wires if it is such a good conductor of electricity?
25. **Using a Database** Vanadium, manganese, and titanium are used to make different types of steel. Find one use for each type of steel alloy.

Hard Water

The contents of tap water vary among communities. In some areas the water is hard. Hard water is water that contains large amounts of calcium or magnesium ions. Hardness can be measured in milligrams per liter (mg/L) of calcium or magnesium ions. Hard water makes it difficult to get hair, clothes, and dishes clean. In this lab, you will learn how hard water is softened and how softening water affects its ability to clean. You will also collect, test, and classify local sources of water.

Problem

How can hard water be softened? How do hard and soft water differ in their ability to clean?

Objectives

- **Compare** the effect of distilled water, hard water, and soft water on the production of suds.
- **Calculate** the hardness of a water sample.

Materials

3 large test tubes with stoppers	hard water
test tube rack	250-mL beaker
grease pencil	balance
25-mL graduated cylinder	filter paper
distilled water	washing soda
dropper	dish detergent
	metric ruler

Safety Precautions



- Always wear safety goggles and a lab apron.
- Washing soda is a skin and eye irritant.

Pre-Lab

1. Read the entire CHEMLAB.
2. Hypothesize about the effect hard and soft water will have on the ability of a detergent to produce suds. Then, predict the relative sudsiness of the three soap solutions.
3. Prepare a data table similar to the one shown below. Leave space to record your qualitative observations.
4. Are there any other safety precautions you need to consider?
5. Suppose you accidentally add more than one drop of detergent to one of the test tubes. Is there a way to adjust for this error or must you discard the sample and start over?
6. The American Society of Agricultural Engineers, the U.S. Department of the Interior, and the Water Quality Association agree on the following classification of water hardness. GPG stands for grains per gallon. One GPG equals 17.1 mg/L. If a sample of water has 150 mg/L of magnesium ions, what is its hardness in grains per gallon?

Production of Suds	
Sample	Level of suds (cm)
Distilled water	
Hard water	
Soft water	

Classification of Water Hardness		
Classification	mg/L	GPG
Soft	0–60	0–3.5
Moderate	61–120	3.5–7
Hard	121–180	7–10.5
Very hard	> 180	> 10.5

Procedure

1. Use a grease pencil to label three large test tubes D (for distilled water), H (for hard water), and S (for soft water).
2. Use a 25-mL graduated cylinder to measure out 20-mL of distilled water. Pour the water into test tube D. Stopper the tube.
3. Place test tube H next to test tube D and make a mark on test tube H that corresponds to the height of the water in test tube D. Repeat the procedure with test tube S.
4. Obtain about 50-mL of hard water in a beaker from your teacher. Slowly pour hard water into test tube H until you reach the marked height.
5. Place a piece of filter paper on a balance and set the balance to zero. Then measure about 0.2 g of washing soda. Remove the filter paper and washing soda. Reset the balance to zero.
6. Use the filter paper to transfer the washing soda to the beaker containing the remainder of the hard water. Swirl the mixture to soften the water. Record any observations.



7. Slowly pour soft water into test tube S until you reach the marked height.
8. Add one drop of dish detergent to each test tube. Stopper the tubes tightly. Then shake each sample to produce suds. Use a metric ruler to measure the height of the suds.
9. Collect water samples from reservoirs, wells, or rain barrels. Use the sudsiness test to determine the hardness of your samples. If access to a source is restricted, ask a local official to collect the sample.

Cleanup and Disposal

1. Use some of the soapy solutions to remove the grease marks from the test tubes.
2. Rinse all of the liquids down the drain with lots of tap water.

Analyze and Conclude

1. **Comparing and Contrasting** Which sample produced the most suds? Which sample produced the least suds? Set up your own water hardness scale based on your data. What is the relative hardness of the local water samples?
2. **Using Numbers** The hard water you used was prepared by adding 1 gram of magnesium sulfate per liter of distilled water. What is its hardness in grams per gallon?
3. **Drawing a Conclusion** The compound in washing soda is sodium carbonate. How did the sodium carbonate soften the hard water?
4. **Thinking Critically** Remember that most compounds of alkaline earth metals do not dissolve easily in water. What is the white solid that formed when washing soda was added to the solution of magnesium sulfate?
5. **Error Analysis** Could the procedure be changed to make the results more quantitative? Explain.

Real-World Chemistry

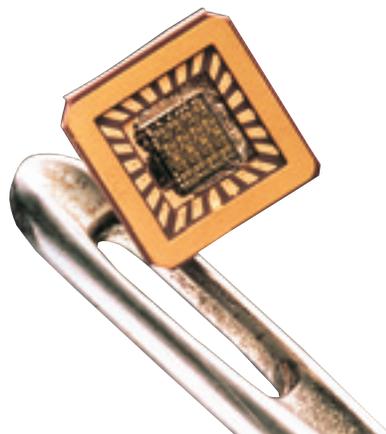
1. Water softeners for washing machines are sold in the detergent section of a store. Look at some of the packages and compare ingredients. Do packages that have different ingredients also have different instructions for how the water softener should be used?
2. Suppose a family notices that the water pressure in their house is not good enough to flush a toilet on the second floor. Other than a leak, what could be interfering with the flow of water?
3. Explain why drinking hard water might be better for your health than drinking soft water. How could a family have the benefit of hard water for drinking and soft water for washing?

How It Works

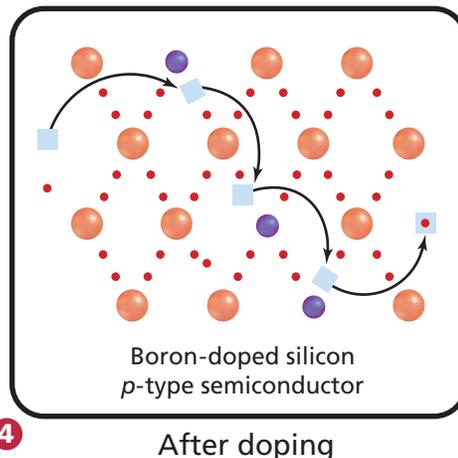
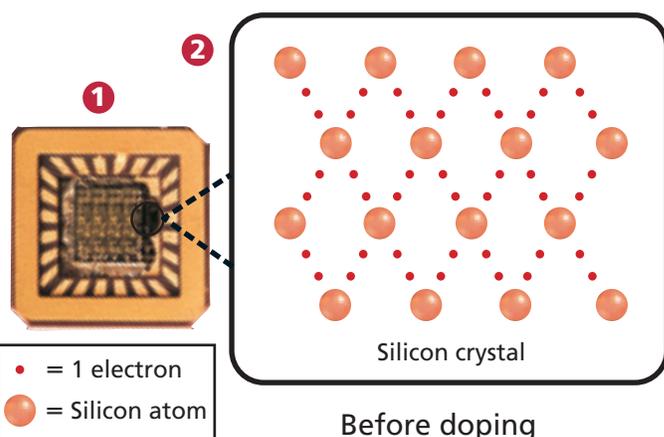
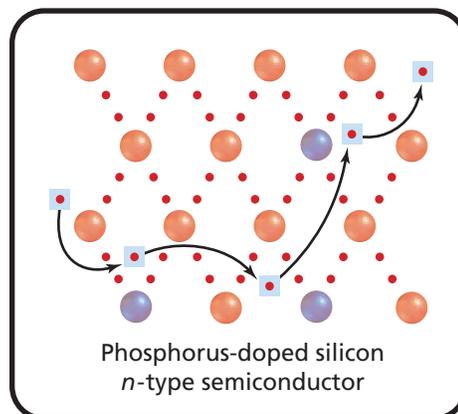
Semiconductor Chips

All of the miniature electrical circuits in your television, computer, calculator, and cell phone depend on elements that are semiconductors. Some metalloids, such as silicon and germanium, are semiconductors.

An electric current is a flow of electrons. This movement of electrons can be used to carry information. Metals are good conductors of electricity; non-metals are poor conductors. Semiconductors fall somewhere in between. Silicon's conductivity can be improved by the addition of phosphorus or boron—a process called doping.



- 1 Hundreds of tiny electronic circuits form an integrated circuit on a thin slice of silicon called a chip.
- 2 Valence electrons hold silicon atoms together. No electrons are available to carry electric current.
- 3 Extra valence electrons from phosphorus are not needed to hold the crystal together. They are free to move throughout the crystal and form an electric current.
- 4 Boron has fewer valence electrons than silicon. When boron is added to silicon, there are spaces lacking electrons. "Holes" are created in the crystal structure. Movement of electrons into and out of these holes produces an electric current.



Thinking Critically

1. Explain why you would expect germanium to have the same type of structure and semi-conducting properties as silicon.
2. What type of semiconductor would you expect arsenic-doped germanium to be?

Summary

7.1 Properties of s-Block Elements

- The number and location of valence electrons determine an element's position on the periodic table and its chemistry.
- Properties within a group are not identical because members have different numbers of inner electrons.
- Similarities between period-2 elements and period-3 elements in neighboring groups are called diagonal relationships.
- The representative elements in groups 1A through 8A have only s and p electrons.
- Because hydrogen has a single electron, it can behave as a metal and lose an electron or behave as a nonmetal and gain an electron.
- The alkali and alkaline earth metals in groups 1A and 2A are the most reactive metals.
- Metals form mixtures called alloys, whose composition can be adjusted to produce different properties.
- Sodium and potassium are the most abundant alkali metals. Many biological functions are controlled by sodium and potassium ions.
- Calcium is essential for healthy teeth and bones. It is most often found as calcium carbonate, which can decompose to form lime—one of the most important industrial compounds.
- Magnesium is used in lightweight, yet strong alloys. Magnesium ions are essential for metabolism, muscle function, and photosynthesis.

7.2 Properties of p-Block Elements

- p-block elements include metals, metalloids, nonmetals, and inert gases.
- Aluminum is the most abundant metal in Earth's crust. Much more energy is needed to extract aluminum from its ore than to recycle aluminum.
- Because a carbon atom can join with up to four other carbon atoms, carbon forms millions of organic compounds.

- Graphite and diamond are allotropes of carbon.
- The most abundant elements in Earth's crust, silicon and oxygen, are usually found in silica, which can be melted and rapidly cooled to form glass.
- Lead, which is still used in storage batteries, was used in pipes, paint, and gasoline until people realized the danger of lead poisoning.
- Nitrogen combines with hydrogen in ammonia, which is used in cleaning products. Nitric acid, which is produced from ammonia, is used to make solid fertilizers, explosives, and dyes.
- Phosphates in fertilizers and cleaning products can harm the environment.
- Sulfur dioxide reacts with water to form one of the acids in acid rain. Most sulfur dioxide is used to make sulfuric acid.
- Halogens are extremely reactive nonmetals. Their compounds are used in toothpaste, disinfectants, and bleaches. Many plastics contain chlorine. Silver bromide or iodide is used to coat photographic film.
- The stable noble gases are used in lighter-than-air blimps, in neon lights, as a substitute for nitrogen in diving tanks, and as an inert atmosphere for welding.

7.3 Properties of d-Block and f-Block Elements

- The d-block transition metals and f-block inner transition metals are more similar across a period than are the s-block and p-block elements.
- The more unpaired electrons in the d sublevel, the harder the transition metal and the higher its melting and boiling points. Ions with partially filled d sublevels often form compounds with color.
- In ferromagnetic metals, ions are permanently aligned in the direction of a magnetic field.
- Many transition metals are strategic materials.
- Lanthanides are silvery metals with high melting points that are found mixed in nature and are hard to separate. Actinides are radioactive elements.

Vocabulary

- actinide series (p. 197)
- ferromagnetism (p. 199)
- mineral (p. 187)
- allotrope (p. 188)
- lanthanide series (p. 197)
- ore (p. 187)
- diagonal relationship (p. 180)
- metallurgy (p. 199)

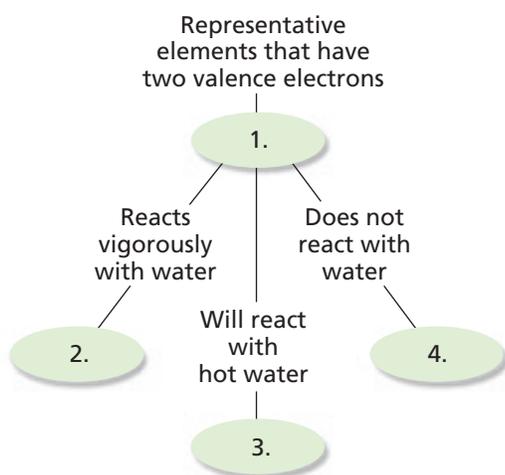


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 7 Assessment.

Concept Mapping

26. Complete the following concept map using the following terms: beryllium; magnesium; calcium, strontium, and barium; alkaline earth metals.



Mastering Concepts

27. Why are elements within a group similar in chemical and physical properties? (7.1)
28. Which groups have representative elements? (7.1)
29. What is a diagonal relationship? (7.1)
30. What happens when hydrogen reacts with a nonmetal element? (7.1)
31. What is heavy water? (7.1)
32. What is the charge on alkali metal ions? On alkaline earth metal ions? (7.1)
33. Identify the element that fits each description. (7.1)
- element in baking soda that turns a flame yellow
 - metallic element found in limestone
 - radioactive alkali metal
34. List some ways group 2A elements differ from group 1A elements. (7.1)
35. Explain why cesium is a more reactive alkali metal than sodium. (7.1)
36. Use their electron configurations to explain why calcium is less reactive than potassium. (7.1)
37. List at least one use for each of the following compounds. (7.1)
- sodium chloride (table salt)
 - calcium oxide (lime)
 - potassium chloride
38. List three types of information that you can obtain from the periodic table. (7.1)
39. Explain why the halogens are extremely reactive non-metals. (7.2)
40. Explain why most carbon compounds are classified as organic compounds. (7.2)
41. What is the charge on halogen ions? (7.2)
42. Argon has only one more proton than chlorine. Explain why these two gases have such different chemical properties. (7.2)
43. Why is red phosphorus classified as an amorphous solid? (7.2)
44. Why is lead no longer used in paints or for plumbing pipes? (7.2)
45. Compare the allotropes of oxygen. (7.2)
46. Identify the element that fits each description. (7.2)
- greenish-yellow gas used to disinfect water
 - main element in emeralds and aquamarines
 - lightweight metal extracted from bauxite ore
 - the most abundant element in Earth's crust
47. What could account for the change in color within the mineral in the photograph below? (7.2)



48. Explain why iodine can be substituted for bromine in some compounds. (7.2)
49. Explain why fluorine reacts with all elements except helium, neon, and argon. (7.2)
50. Name the element that combines with oxygen to form a compound that fits each description. (7.2)
 - a. the compound that can be melted to form glass
 - b. the main compound in ruby and sapphire
 - c. a compound used to preserve fruit and produce an inexpensive acid
51. List at least one use for each compound. (7.2)
 - a. silicon carbide (carborundum)
 - b. aluminum sulfate (alum)
 - c. boric acid
 - d. nitric acid
52. When the metal gallium melts, is the liquid that forms an allotrope? Explain your answer. (7.2)
53. How do transition metals differ from inner transition metals in their electron configurations? (7.3)
54. Explain why compounds of zinc are white but compounds containing copper have a color.
55. Name three general methods for extracting a metal from its ore. (7.3)
56. What does it mean for a metal to be listed as strategic or critical? (7.3)
57. Predict which of the transition metals in period 4 is diamagnetic. Explain your answer. (7.3)
58. Explain how the electron configurations of chromium and copper determine that one is used to strengthen alloys and the other to make jewelry. (7.3)

Mixed Review

Sharpen your problem-solving skills by answering the following.

59. What distinguishes a metal from a nonmetal?
60. Where are the most reactive nonmetals located on the periodic table? The most reactive metals?
61. Which families contain metalloids?
62. What chemical property does zinc share with calcium?
63. Name at least three elements that are commonly found in fertilizers.

64. What physical property is shown by the element in the photograph? Identify the element.



65. Of the period-5 elements, palladium, tin, and silver, which will display noticeably different properties from the other two? Explain your choice.
66. What do rubidium and white phosphorus have in common?
67. Identify the block on the periodic table where you are likely to find:
 - a. a synthetic radioactive element.
 - b. a highly reactive element that forms salts with halogens.
 - c. an element that forms millions of compounds.
68. For each period-3 element, Na through Cl, identify the following:
 - a. the metals, nonmetals, and metalloids.
 - b. the appearance of the element and its state at room temperature.
69. What is the most common ion or ions for each of these elements: sodium, bromine, neon, cadmium, boron, and hydrogen?

Thinking Critically

70. **Applying Concepts** Which element would be easiest to extract from its ore, gold or iron? Explain your choice.
71. **Hypothesizing** Why might countries have different lists of strategic and critical materials?
72. **Applying Concepts** In 1906, Ferdinand Moissan won a Nobel Prize for producing fluorine in its pure elemental form. Why do you think his achievement was considered worthy of such a prize?
73. **Analyze and Conclude** Why were metals such as copper and gold discovered long before gases such as oxygen and nitrogen?

- 74. Drawing a Conclusion** The most important function of blood is to carry oxygen to all cells in the body. What could happen if blood flow to a cell were blocked?
- 75. Using a Database** Find out the following:
- the trend for the melting point of alkali metals
 - the color of copper compounds
 - the colors of the ions of chromium
- 76. Observing and Predicting** If you were given two solutions, one colorless and the other a light blue, which solution would probably contain an ion from a transition metal? Explain your answer.

Writing in Chemistry

- 77.** Research the local city or state regulations for removal of lead paint and write a report on your findings.
- 78.** Use the example of ozone to argue that a single chemical can be both beneficial and harmful.
- 79.** Describe the uses of fluorine, chlorine, and iodine compounds related to drinking water and wastewater.
- 80.** Research and write a report on the use of silver iodide for seeding clouds.
- 81.** Research the history of the term *bromide*, which is used to describe a remedy designed to ease tension.
- 82.** Find out how the daily values for vitamins and minerals are determined. Compare supplements to see whether they all contain the same minerals in the same quantities per tablet. In your report, suggest reasons for any variations you discover.
- 83.** Research the processes of recycling aluminum, plastic, and glass. Design a poster or a multimedia presentation for members of your community about one of the processes. Include the following considerations: Why is it cost effective to recycle? What resources are conserved? What recycling is available in your area? How can the members of your community use this information to make responsible choices about the products they purchase?
- 84.** Find out how the elements are assigned their names. What elements are named after people? Choose one and research why this person was given such an honor.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 85.** Determine the correct number of significant figures in each of the following. (Chapter 2)
- 708.4 mL
 - 1.0050 g
 - 1.000 mg
 - 6.626×10^{-34} s
 - 2000 people
- 86.** A quarter has a mass of 5.627 g. What is its mass in milligrams? (Chapter 2)
- 87.** A solution of sugar has a density of 1.05 g/cm³. If you have 300.0 mL of sugar solution, what is the mass of the solution? (Chapter 2)
- 88.** A substance is said to be volatile if it readily changes from a liquid to a gas at room temperature. Is this a physical change or a chemical change? (Chapter 3)
- 89.** How many electrons, protons, and neutrons are there in the following: (Chapter 4)
- carbon-13
 - chromium-50
 - tin-119
- 90.** An AM radio signal broadcasts at 6.00×10^5 Hz. What is the wavelength of this signal in meters? What is the energy of this signal? (Chapter 5)
- 91.** Determine the energy of a photon with a wavelength of 4.80×10^2 nm. (Chapter 5)
- 92.** Identify which of the following orbitals cannot exist according to quantum theory: 4s, 1p, 2p, 3f, 2d. Briefly explain your answer. (Chapter 5)
- 93.** Mendeleev left a space on his periodic table for the undiscovered element germanium and in 1886 Winkler discovered it. Write the electron configuration for germanium. (Chapter 6)
- 94.** Select the atom or ion in each pair that has the larger radius. (Chapter 6)
- Cs or Fr
 - Br or As
 - O²⁻ or O
- 95.** Arrange the following in order of increasing ionization energy: Li, C, Si, Ne. (Chapter 6)
- 96.** Arrange the following in order of increasing ionization energy: K, Ca, Fr, Mg. (Chapter 6)

Use these questions and the test-taking tip to prepare for your standardized test.

- Which of the following descriptions does NOT apply to gold?
 - transition metal
 - ferromagnetic
 - element
 - nonconductor
- On the periodic table, metalloids are found only in _____.
 - the d-block
 - groups 3A through 6A
 - the f-block
 - groups 1A and 2A
- The majority of _____ are radioactive.
 - actinides
 - lanthanides
 - halogens
 - alkali metals
- Which of the following groups is composed entirely of nonmetals?
 - 1A
 - 3A
 - 5A
 - 7A
- _____ is a major component of organic compounds.
 - Sodium
 - Calcium
 - Carbon
 - Potassium
- Although metals in the same group on the periodic table have the same number of valence electrons, they do not have identical properties. Which of the following is NOT an explanation for differences in properties within a group?
 - The metals have different numbers of nonvalence electrons.
 - The reactivity of metals increase as their ionization energies decrease.
 - The ionization energies of metals decrease as their atomic masses increase.
 - The reactivity of metals increase as their atomic masses decrease.

- Throughout history, dangerous products were used before people understood their harmful effects. Which of the following historic practices was NOT discontinued because of the health hazard it posed?
 - using radium paint to make watch hands glow in the dark
 - lining steel cans with tin to prevent corrosion
 - using arsenic sulfide to treat illnesses
 - adding lead to gasoline to increase engine efficiency

Interpreting Tables Use the table to answer questions 8 through 10.

Characteristics of Elements		
Element	Block	Characteristic
X	s	soft solid; reacts readily with oxygen
Y	p	gas at room temperature; forms salts
Z	—	inert gas

- In which group does Element X most likely belong?
 - 1A
 - 7A
 - 8A
 - 4B
- Element Y is probably _____.
 - an alkaline metal
 - an alkaline earth metal
 - a halogen
 - a transition metal
- Element Z is most likely found in the _____.
 - s-block
 - p-block
 - d-block
 - f-block



TEST-TAKING TIP

Plan Your Work and Work Your Plan

Plan your workload so that you do a little work each day rather than a lot of work all at once. The key to retaining information is repeated review and practice. You will retain more if you study one hour a night for five days in a row instead of cramming for five hours on a Sunday night.

Ionic Compounds

What You'll Learn

- ▶ You will define a chemical bond.
- ▶ You will describe how ions form.
- ▶ You will identify ionic bonding and the characteristics of ionic compounds.
- ▶ You will name and write formulas for ionic compounds.
- ▶ You will relate metallic bonds to the characteristics of metals.

Why It's Important

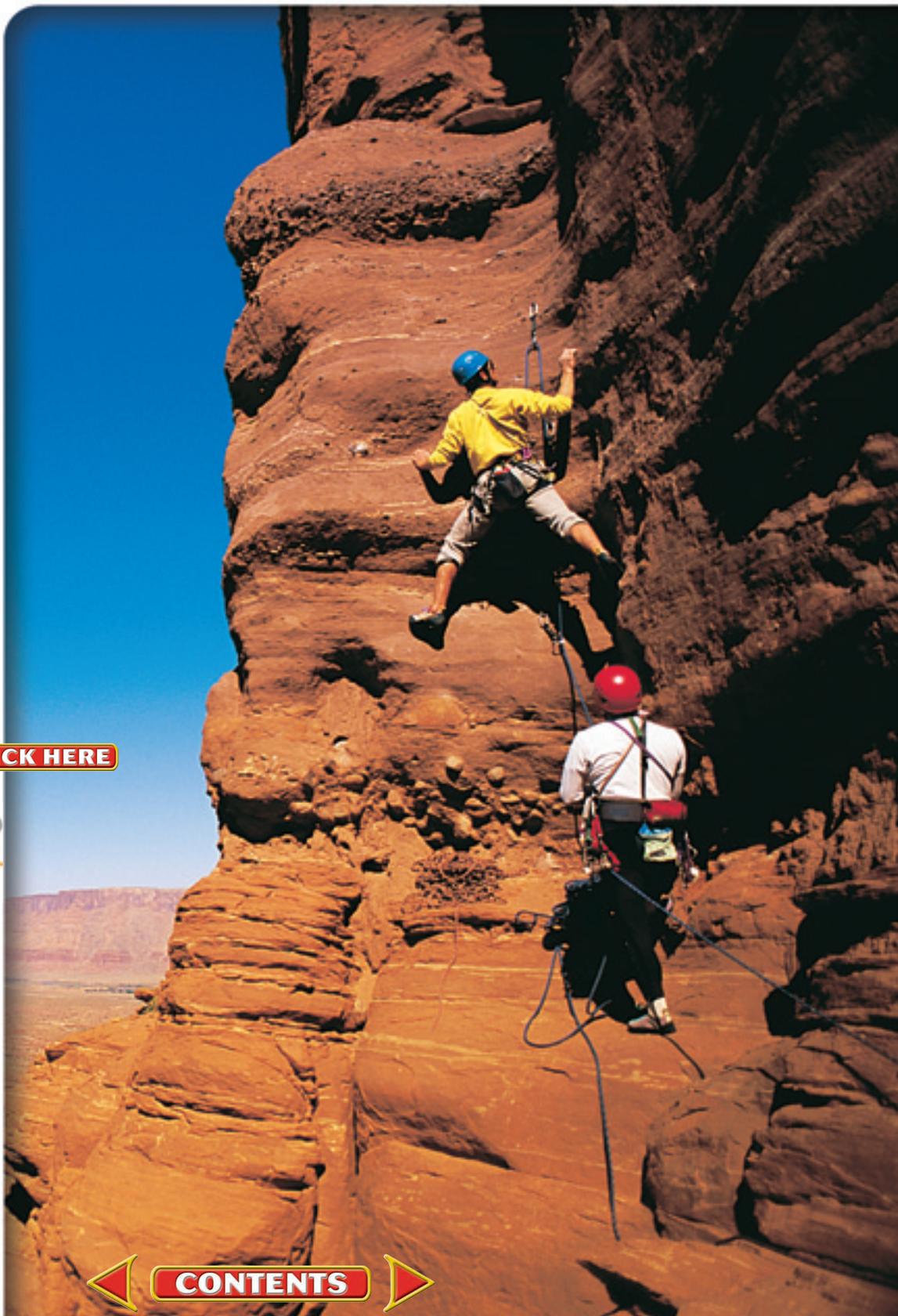
The world around you is composed mainly of compounds. The properties of each compound are based on how the compound is bonded. The salts dissolved in Earth's oceans and the compounds that make up most of Earth's crust are held together by ionic bonds.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about ionic bonding.

The rock surface, the climbers' equipment, the atmosphere, and even the climbers are composed almost entirely of compounds and mixtures of compounds.



DISCOVERY LAB



An Unusual Alloy

Most metals that you encounter are solids. Can a metal melt at a temperature below the boiling point of water? You will use a metal alloy called Onion's Fusible Alloy to answer this question.

Safety Precautions



Use caution around the heat source and the heated beaker and its contents.

Procedure

1. Carefully place a small piece of Onion's Fusible Alloy into a 250-mL beaker. Add about 100 mL of water to the beaker.
2. Heat the beaker and its contents with a laboratory burner or a hot plate. Monitor the temperature with a thermometer. When the temperature rises above 85°C, carefully observe the Onion's Fusible Alloy. Record your observations. Remove the beaker from the heat when the water begins to boil. Allow the contents to cool before handling the Onion's Fusible Alloy.

Analysis

What is unusual about Onion's Fusible Alloy compared to other metals? Onion's Fusible Alloy contains bismuth, lead, and tin. Compare the melting points of these metals to the melting point of Onion's Fusible Alloy.

Materials

Onion's Fusible Alloy
250-mL beaker
laboratory burner or
hot plate
Celsius thermometer

Section

8.1

Forming Chemical Bonds

Objectives

- **Define** chemical bond.
- **Relate** chemical bond formation to electron configuration.
- **Describe** the formation of positive and negative ions.

Vocabulary

chemical bond
cation
anion

Ascending to the summit of a mountain peak, a rock climber can survey the surrounding world. This world is composed of many different kinds of compounds, ranging from simple ones such as the sodium chloride found in the perspiration on the climber's skin to more complex ones such as the calcite or pyrite found in certain rocks. How do these and thousands of other compounds form from the relatively few elements known to exist?

Chemical Bonds

The answer to this question lies in the electron structure of the atoms of the elements involved and the nature of the attractive forces between these atoms. The force that holds two atoms together is called a **chemical bond**. Chemical bonds may form by the attraction between a positive nucleus and negative electrons or the attraction between a positive ion and a negative ion.

In previous chapters, you learned about atomic structure, electron arrangement, and periodic properties of the elements. The elements within a group on the periodic table have similar properties. Many of these properties are due to the number of valence electrons. These same electrons are involved in the formation of chemical bonds between two atoms.

Table 8-1

Electron-Dot Structures								
Group	1A	2A	3A	4A	5A	6A	7A	8A
Diagram	Li·	·Be·	·B·	·C·	·N·	·O·	·F·	·Ne·

Recall that an electron-dot structure is a type of diagram used to keep track of valence electrons and is especially useful when illustrating the formation of chemical bonds. **Table 8-1** shows several examples of electron-dot structures. For example, carbon has an electron configuration of $1s^22s^22p^2$. Its valence electrons are those in the second energy level, as can be seen in the electron-dot structure for carbon in the table.

Recall from Chapter 6 that ionization energy refers to how easily an atom loses an electron. The term electron affinity indicates how much attraction an atom has for electrons. Noble gases, having high ionization energies and low electron affinities, show a general lack of chemical reactivity. Other elements on the periodic table react with each other, forming numerous compounds. The difference in reactivity is directly related to the valence electrons.

All atoms have valence electrons. Why does this difference in reactivity of elements exist? Noble gases have electron configurations that have a full outermost energy level. This level is full with two electrons for helium ($1s^2$). The other noble gases have electron configurations consisting of eight electrons in the outermost energy level, ns^2np^6 . As you will recall, the presence of eight valence electrons in the outer energy level is chemically stable and is called a stable octet. Elements tend to react to acquire the stable electron structure of a noble gas.

Formation of positive ions Recall that a positive ion forms when an atom loses one or more valence electrons in order to attain a noble gas configuration. To understand the formation of a positive ion, compare the electron configurations of the noble gas neon, atomic number 10, and the alkali metal sodium, atomic number 11.

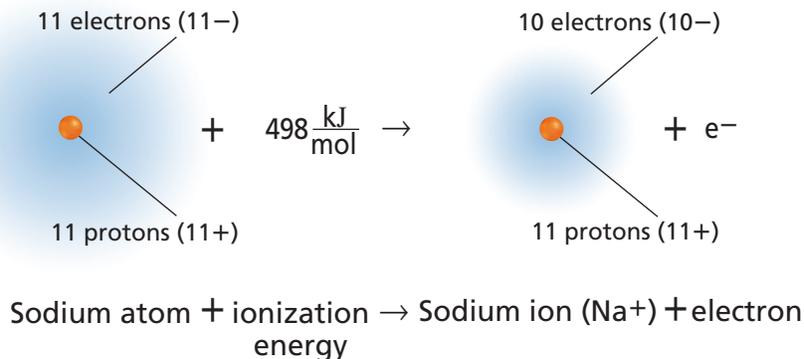
Neon $1s^22s^22p^6$

Sodium $1s^22s^22p^63s^1$

Note that the sodium atom has one 3s valence electron; it differs from the noble gas neon by that single valence electron. If sodium loses this outer valence electron, the resulting electron configuration will be identical to that of neon. **Figure 8-1** shows how a sodium atom loses its valence electron to become a positive sodium ion. A positively charged ion is called a **cation**.

Figure 8-1

In the formation of a positive ion, a neutral atom loses one or more valence electrons. Note that the number of protons is equal to the number of electrons in the uncharged atom, but the ion contains more protons than electrons, making the overall charge on this ion positive.



A sodium ion is more stable than the neutral sodium atom because by losing an electron, the ion acquired the stable outer electron configuration of neon. It is important to understand that although sodium now has the electron configuration of neon, it is not neon. It is a sodium ion with a single positive charge. The 11 protons that establish the character of sodium still remain within its nucleus.

Reactivity of metals is based on the ease with which they lose valence electrons to achieve a stable octet, or noble gas configuration. Group 1A elements, [noble gas] ns^1 , lose their one valence electron, forming an ion with a 1+ charge. Group 2A elements, [noble gas] ns^2 , lose their two valence electrons and form ions with a 2+ charge. For example, potassium, a group 1A element, forms a K^+ ion; magnesium, a group 2A element, forms a Mg^{2+} ion. These two groups contain the most active metals on the periodic table. Some elements in group 3A, [noble gas] ns^2np^1 , also lose electrons and form positive ions. What is the charge on these ions? What is the formula for the aluminum ion?

Recall that, in general, transition metals have an outer energy level of ns^2 . Going from left to right across a period, atoms of each element are filling an inner d sublevel. When forming positive ions, transition metals commonly lose their two valence electrons, forming 2+ ions. However, it is also possible for d electrons to be lost. Thus transition elements also commonly form ions of 3+ or greater, depending on the number of d electrons in the electron structure. It is difficult to predict the number of electrons lost by transition elements. A useful rule of thumb for these metals is that they form ions with a 2+ or 3+ charge.

Although the formation of an octet is the most stable electron configuration, other electron configurations provide some stability. For example, elements in groups 1B through 4A in periods 4 through 6 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as *pseudo-noble gas configurations*. Let's examine the formation of the zinc ion, which is shown in **Figure 8-2**. The zinc atom has the electron configuration of $1s^22s^22p^63s^23p^64s^23d^{10}$. When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^22s^22p^63s^23p^63d^{10}$ results in a pseudo-noble gas configuration.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

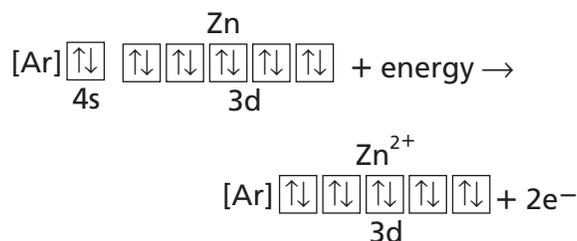
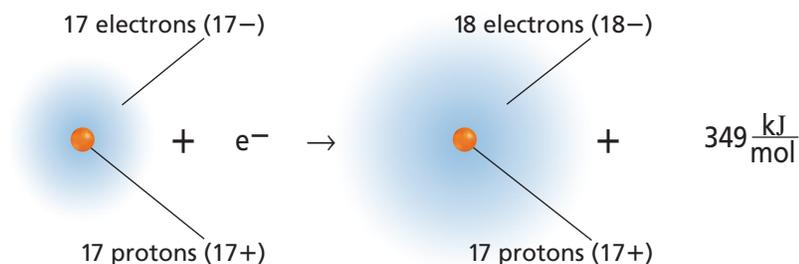


Figure 8-2

Orbital notation provides a convenient way to visualize the loss or gain of valence electrons. When zinc metal reacts with sulfuric acid, the zinc forms a Zn^{2+} ion with a pseudo-noble gas configuration.

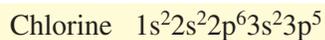
Figure 8-3

In the formation of a negative ion, a neutral atom gains one or more electrons. Again, note that in the neutral atom the number of protons equals the number of electrons. However, the ion contains more electrons than protons, making this overall charge on this ion negative.

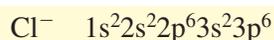
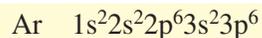


Chlorine atom + electron \rightarrow Chloride ion (Cl^-) + electron affinity

Formation of negative ions Recall that nonmetals, located on the right side of the periodic table, have a great attraction for electrons and form a stable outer electron configuration by gaining electrons. The chlorine atom, a halogen from group 7A, provides a good example.



Examine **Figure 8-3**. To attain a noble gas configuration, chlorine gains one electron, forming a negative ion with a $1-$ charge. By gaining the single electron, the chlorine atom now has the electron configuration of argon.



With the addition of one electron, chlorine becomes an **anion**, which is another name for a negative ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus the anion of chlorine is called the chloride ion. What is the name of the anion formed from nitrogen?

Nonmetals gain the number of electrons that, when added to their valence electrons, equals eight. Phosphorus, a group 5A element with the electron configuration of $[\text{Ne}]3s^23p^3$, has five valence electrons. To form a stable octet, the phosphorus atom may gain three electrons and form the phosphide ion with a $3-$ charge. If an oxygen atom, a group 6A element, gains two electrons, the oxide ion with a charge of $2-$ results.

Some nonmetals can lose or gain other numbers of electrons to form an octet. For example, in addition to gaining three electrons, phosphorus can lose five. However, in general, group 5A elements gain three electrons, group 6A gain two, and group 7A gain one to achieve an octet.

Section 8.1 Assessment

- What is a chemical bond?
- Why do ions form?
- What family of elements is relatively unreactive and why?
- Describe the formation of both positive and negative ions.
- Thinking Critically** Predict the change that must occur in the electron configuration if each of the following atoms is to achieve a noble gas configuration.

a. nitrogen	c. barium
b. sulfur	d. lithium
- Formulating Models** Draw models to represent the formation of the positive calcium ion and the negative bromide ion.

The Formation and Nature of Ionic Bonds

Look at the photos in **Figure 8-4a** and **b**. What do these reactions have in common? As you can see, in both cases, elements react with each other to form a compound. What happens in the formation of a compound?

Formation of an Ionic Bond

Figure 8-4a shows the reaction between the elements sodium and chlorine. During this reaction, a sodium (Na) atom transfers its valence electron to a chlorine (Cl) atom and becomes a positive ion. The chlorine atom accepts the electron into its outer energy level and becomes a negative ion. The compound sodium chloride forms because of the attraction between oppositely charged sodium and chloride ions. The electrostatic force that holds oppositely charged particles together in an ionic compound is referred to as an **ionic bond**. Compounds that contain ionic bonds are ionic compounds. If ionic bonds occur between metals and the nonmetal oxygen, oxides form. Most other ionic compounds are called salts.

Hundreds of compounds contain ionic bonds. Many ionic compounds are binary, which means that they contain only two different elements. Binary ionic compounds contain a metallic cation and a nonmetallic anion. Magnesium oxide, MgO , is a binary compound because it contains the two different elements magnesium and oxygen. However, CaSO_4 is not a binary compound. Can you explain why?

Consider the formation of the ionic compound calcium fluoride from calcium (Ca) and fluorine (F). Calcium, a group 2A metal with the electron configuration $[\text{Ar}]3s^2$, has two valence electrons. Fluorine, a group 7A nonmetal with the electron configuration $[\text{He}]2s^22p^5$, must gain one electron to attain the noble gas configuration of neon.

Because the number of electrons lost must equal the number of electrons gained, it will take two fluorine atoms to gain the two electrons lost from one

Objectives

- **Describe** the formation of ionic bonds.
- **Account** for many of the physical properties of an ionic compound.
- **Discuss** the energy involved in the formation of an ionic bond.

Vocabulary

ionic bond
electrolyte
lattice energy



Figure 8-4

These chemical reactions that produce ionic compounds also release a large amount of energy.

- a** The reaction that occurs between elemental sodium and chlorine gas produces a white crystalline solid.
- b** This sparkler contains iron, which burns in air to produce an ionic compound that contains iron and oxygen.

calcium atom. The compound formed will contain one calcium ion with a charge of 2+ for every two fluoride ions, each with a charge of 1-. Note that the overall charge on one unit of this compound is zero.

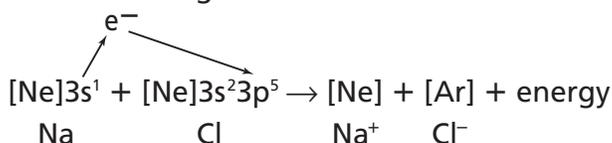
$$1 \text{ Ca ion} \left(\frac{2+}{\text{Ca ion}} \right) + 2 \text{ F ions} \left(\frac{1-}{\text{F ion}} \right) = (2+) + 2(1-) = 0$$

Figure 8-5

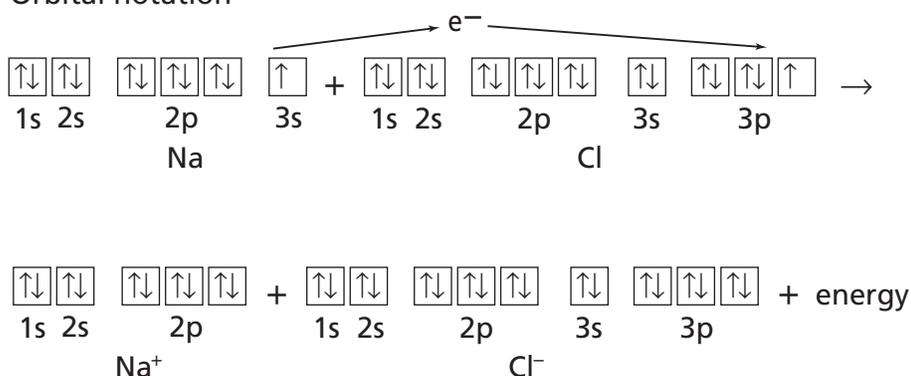
Several methods are used to show how an ionic compound forms.

Figure 8-5 summarizes the formation of an ionic compound from the elements sodium and chlorine using four different methods: electron configuration, orbital notation, electron-dot structures, and atomic models.

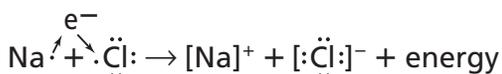
Electron configuration



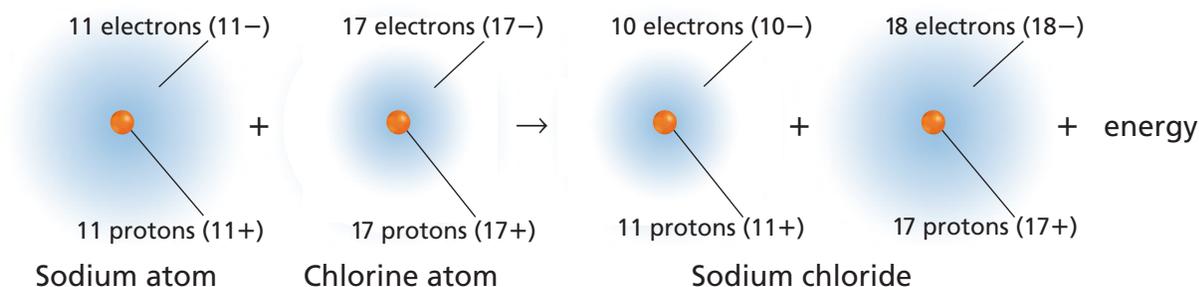
Orbital notation



Electron-dot structures



Atomic models



EXAMPLE PROBLEM 8-1

Formation of an Ionic Compound

Unprotected aluminum metal reacts with oxygen in air, forming the white coating you can observe on aluminum objects such as lawn furniture. Explain the formation of an ionic compound from the elements aluminum and oxygen.

1. Analyze the Problem

You are given that aluminum and oxygen react to form an ionic compound. Aluminum is a group 3A element with three valence electrons, and oxygen is a group 6A element with six valence electrons. To acquire a noble gas configuration, each aluminum atom must lose three electrons and each oxygen atom must gain two electrons.

2. Solve for the Unknown

Remember that the number of electrons lost must equal the number of electrons gained. The smallest number evenly divisible by the three electrons lost by aluminum and the two gained by oxygen is six. Three oxygen atoms are needed to gain the six electrons lost by two aluminum atoms.

3. Evaluate the Answer

The overall charge on one unit of this compound is zero.

$$2 \text{ Al ions } \left(\frac{3+}{\text{Al ion}} \right) + 3 \text{ O ions } \left(\frac{2-}{\text{O ion}} \right) = 2(3+) + 3(2-) = 0$$

PRACTICE PROBLEMS

Explain the formation of the ionic compound composed of each pair of elements.

- sodium and nitrogen
- lithium and oxygen
- strontium and fluorine
- aluminum and sulfur
- cesium and phosphorus



For more practice with forming ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

Properties of Ionic Compounds

The chemical bonds that occur between the atoms in a compound determine many of the physical properties of the compound. During the formation of an ionic compound, the positive and negative ions are packed into a regular repeating pattern that balances the forces of attraction and repulsion between the ions. This particle packing forms an ionic crystal, as shown in **Figure 8-6**. No single unit consisting of only one ion attracting one other ion is formed. Large numbers of positive ions and negative ions exist together in a ratio determined by the number of electrons transferred from the metal to the nonmetal.

Examine the pattern of the ions in the sodium chloride crystal shown in the figure. What shape would you expect a large crystal of this compound to be? This one-to-one ratio of ions produces a cubic crystal. Examine some table salt (NaCl) under a magnifying glass. What shape are these small salt crystals?

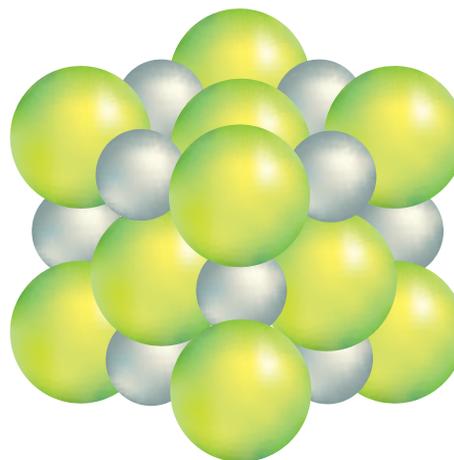


Figure 8-6

In this ionic compound, each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

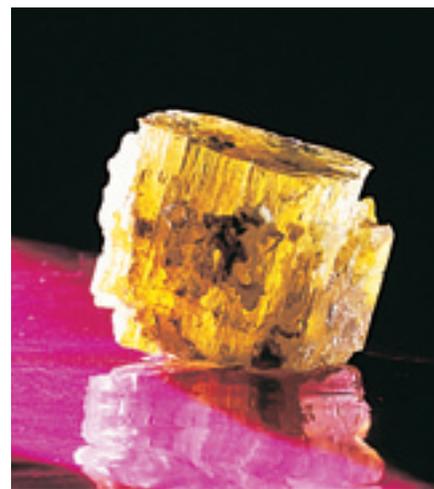


Figure 8-7

Aragonite (CaCO_3), barite (BaSO_4), and beryl ($\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$) are examples of minerals that are ionic compounds. The ions that form them are bonded together in a crystal lattice.

The strong attraction of positive ions and negative ions in an ionic compound results in a crystal lattice. A crystal lattice is a three-dimensional geometric arrangement of particles. In a crystal lattice, each positive ion is surrounded by negative ions and each negative ion is surrounded by positive ions. Ionic crystals vary in shape due to the sizes and relative numbers of the ions bonded, as shown in **Figure 8-7**.

Melting point, boiling point, and hardness are physical properties that depend on how strongly the particles are attracted to each other. Because ionic bonds are relatively strong, the crystals that result require a large amount of energy to be broken apart. Therefore, ionic crystals have high melting points and boiling points, as shown in **Table 8-2**. Their color may be related to their structure. See the **problem-solving LAB** on the next page and **Everyday Chemistry** at the end of this chapter. These crystals are also hard, rigid, and brittle solids due to the strong attractive forces that hold the ions in place. When an external force large enough to overcome the attraction of ions in the crystal is applied, the crystal cracks. The force repositions the like-charged ions next to each other, and the repulsive force cracks the crystal.

Table 8-2

Melting and Boiling Points of Some Ionic Compounds		
Compound	Melting point ($^{\circ}\text{C}$)	Boiling point ($^{\circ}\text{C}$)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
CaCl_2	782	>1600
CaI_2	784	1100
NaCl	801	1413
MgO	2852	3600

Charged particles must be free to move for a material to conduct an electric current. In the solid state, ionic compounds are nonconductors of electricity because of the fixed positions of the ions. However, in a liquid state or when dissolved in water, ionic compounds are electrical conductors because the ions are free to move. An ionic compound whose aqueous solution conducts an electric current is called an **electrolyte**. You will learn more about solutions of electrolytes in Chapter 15.

Energy and the ionic bond During any chemical reaction, energy is either absorbed or released. When energy is absorbed during a chemical reaction, the reaction is endothermic. If energy is released, it is exothermic. As you will see in the **CHEMLAB** at the end of this chapter, energy is released when magnesium reacts with oxygen.

Energy changes also occur during the formation of ionic bonds from the ions formed during a chemical reaction. The formation of ionic compounds from positive and negative ions is almost always exothermic. The attraction of the positive ion for the negative ions close to it forms a more stable system that is lower in energy than the individual ions. If the amount of energy released during bond formation is added to an ionic compound, the bonds that hold the positive and negative ions together break.

You just learned that the ions in an ionic compound are arranged in a pattern in a crystal lattice. The energy required to separate one mole of the ions of an ionic compound is referred to as the **lattice energy**. The strength of the forces holding ions in place is reflected by the lattice energy. The more negative the lattice energy, the stronger the force of attraction.

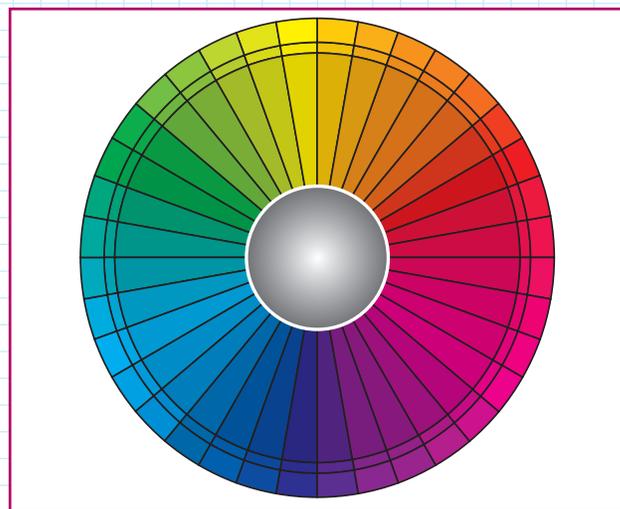
problem-solving LAB

How is color related to a transferred electron?

Factors that Affect Color of an Ionic Compound

Compound	Color	Anion radius (Å)	Visible absorption band
AgF	Yellow	1.36	Blue-violet
AgCl	White	1.81	None
AgBr	Cream	1.95	Violet
AgI	Yellow	2.16	Blue-violet
Ag ₂ S	Black	1.84	All
Al ₂ O ₃	White	1.40	None
Sb ₂ O ₃	White	1.40	None
Bi ₂ O ₃	?	1.40	Violet

Predicting Once an ionic bond is formed, the cation has a tendency to pull the transferred electron toward the nucleus. The appearance of color is directly related to the strength of the pull, which depends upon the size of the ions involved and their oxidation numbers. If visible light of a certain color can send the electron back to the cation momentarily, then the light reflected from a crystal of the compound will be missing this color from its spectrum. The resulting color of the crystal will be the complement of this color of light and can be predicted using the color wheel that is shown here. Complementary colors are across from each other on the color wheel.



Analysis

The information in the table can be used to make some general conclusions about a compound's color and the strength of the cation's pull on the transferred electron. Using this information, the color of a compound can be predicted.

Thinking Critically

1. A larger anion radius results in a more pronounced color. What reason can you give for this fact?
2. Which do you think produces a more pronounced color, a high oxidation state for the anions or a low one? Explain.
3. Use the color wheel to predict the color of Bi₂O₃, the last compound on the list.

Lattice energy is directly related to the size of the ions bonded. Smaller ions generally have a more negative value for lattice energy because the nucleus is closer to and thus has more attraction for the valence electrons. Thus, the lattice energy of a lithium compound is more negative than that of a potassium compound containing the same anion because the lithium ion is smaller than the potassium ion. Which would have a more negative lattice energy, lithium chloride or lithium bromide?

The value of lattice energy is also affected by the charge of the ion. The ionic bond formed from the attraction of ions with larger positive or negative charges generally has a more negative lattice energy. The lattice energy of MgO is almost four times greater than the lattice energy of NaF because the charge of the ions is greater. The lattice energy of SrCl₂ is between the lattice energies of MgO and NaF because SrCl₂ contains ions with both higher and lower charges.

Table 8-3 shows the lattice energies of some ionic compounds. Examine the lattice energies of RbF and KF. How do they confirm that lattice energy is related to ion size? Look at the lattice energies of SrCl₂ and AgCl. How do they show the relationship between lattice energy and the charge of the ions involved?

Table 8-3

Lattice Energies of Some Ionic Compounds			
Compound	Lattice energy (kJ/mol)	Compound	Lattice energy (kJ/mol)
KI	-632	KF	-808
KBr	-671	AgCl	-910
RbF	-774	NaF	-910
NaI	-682	LiF	-1030
NaBr	-732	SrCl ₂	-2142
NaCl	-769	MgO	-3795

Section 8.2 Assessment

- What is an ionic bond?
- How does an ionic bond form?
- List three physical properties associated with an ionic bond.
- Describe the arrangement of ions in a crystal lattice.
- What is lattice energy and how is it involved in an ionic bond?
- Thinking Critically** Using the concepts of ionic radii and lattice energy, account for the trend in melting points shown in the following table.

Trend in Melting Points	
Ionic compound	Melting point in °C
KF	858
KCl	770
KBr	734
KI	681

- Formulating Models** Use electron configurations, orbital notation, and electron-dot structures to represent the formation of an ionic compound from the metal strontium and the nonmetal chlorine.

Names and Formulas for Ionic Compounds

One of the most important requirements of chemistry is communicating information to others. Chemists discuss compounds by using both chemical formulas and names. The chemical formula and the name for the compound must be understood universally. Therefore, a set of rules is used in the naming of compounds. This system of naming allows everyone to write a chemical formula when given a compound name and to name the compound from a given chemical formula.

Formulas for Ionic Compounds

Recall from Section 8.2 that a sample of an ionic compound contains crystals formed from many ions arranged in a pattern. Because no single particle of an ionic compound exists, ionic compounds are represented by a formula that provides the simplest ratio of the ions involved. The simplest ratio of the ions represented in an ionic compound is called a **formula unit**. For example, the formula KBr represents a formula unit for potassium bromide because potassium and bromide ions are in a one-to-one ratio in the compound. A formula unit of magnesium chloride is MgCl_2 because two chloride ions exist for each magnesium ion in the compound. In the compound sodium phosphide, three sodium ions exist for every phosphide ion. What is the formula unit for sodium phosphide?

Because the total number of electrons gained by the nonmetallic atoms must equal the total number of electrons lost by the metallic atoms, the overall charge of a formula unit is zero. The formula unit for MgCl_2 contains one Mg^{2+} ion and two Cl^- ions, for a total charge of zero.

Determining charge Binary ionic compounds are composed of positively charged monatomic ions of a metal and negatively charged monatomic ions of a nonmetal. A **monatomic ion** is a one-atom ion, such as Mg^{2+} or Br^- . **Table 8-4** indicates the charges of common monatomic ions according to the location of their atoms on the periodic table. What is the formula for the beryllium ion? The iodide ion? The nitride ion? Transition metals, which are in groups 3B through 2B, and metals in groups 3A and 4A are not included in this table because of the variance in ionic charges of atoms in the groups. Most transition metals and those in groups 3A and 4A can form several different positive ions.

Table 8-4

Common Ions Based on Groups		
Group	Atoms that commonly form ions	Charge on ions
1A	H, Li, Na, K, Rb, Cs	1+
2A	Be, Mg, Ca, Sr, Ba	2+
5A	N, P, As	3-
6A	O, S, Se, Te	2-
7A	F, Cl, Br, I	1-

Objectives

- **Write** formulas for ionic compounds and oxyanions.
- **Name** ionic compounds and oxyanions.

Vocabulary

formula unit
monatomic ion
oxidation number
polyatomic ion
oxyanion



Careers Using Chemistry

Wastewater Treatment Operator

Would you be interested in a job that assures your community of a safe water supply? Then consider a career in wastewater treatment.

Not only does our water supply have to be safe for human consumption, the water that is returned to rivers and streams must be cleaned of pathogens and suspended solids so it can be used over and over. This career involves testing to identify chemicals, pathogens, and materials in the wastewater, as well as monitoring the multi-step process of their removal.

Table 8-5

Common Ions of Transition Metals and Groups 3A and 4A	
Group	Common ions
3B	Sc^{3+} , Y^{3+} , La^{3+}
4B	Ti^{2+} , Ti^{3+}
5B	V^{2+} , V^{3+}
6B	Cr^{2+} , Cr^{3+}
7B	Mn^{2+} , Mn^{3+} , Tc^{2+}
8B	Fe^{2+} , Fe^{3+}
8B	Co^{2+} , Co^{3+}
8B	Ni^{2+} , Pd^{2+} , Pt^{2+} , Pt^{4+}
1B	Cu^{+} , Cu^{2+} , Ag^{+} , Au^{+} , Au^{3+}
2B	Zn^{2+} , Cd^{2+} , Hg_2^{2+} , Hg^{2+}
3A	Al^{3+} , Ga^{2+} , Ga^{3+} , In^{+} , In^{2+} , In^{3+} , Tl^{+} , Tl^{3+}
4A	Sn^{2+} , Sn^{4+} , Pb^{2+} , Pb^{4+}

The charge of a monatomic ion is its **oxidation number**. Most transition metals and group 3A and 4A metals have more than one oxidation number, as shown in **Table 8-5**. The oxidation numbers given in the table are the most common ones for many of the elements listed but might not be the only ones possible.

The term *oxidation state* is sometimes used and means the same thing as oxidation number. The oxidation number, or oxidation state, of an element in an ionic compound equals the number of electrons transferred from an atom of the element to form the ion. For example, when sodium and chlorine atoms react, the sodium atom transfers one electron to the chlorine atom, forming Na^{+} and Cl^{-} . Thus, in the compound formed, the oxidation state of sodium is 1+ because one electron is transferred from the sodium atom. The oxidation state of chlorine is 1-. One electron is transferred, and the negative sign shows that the electron transferred to, not from, the chlorine atom.

The oxidation numbers of ions are used to determine the formulas for the ionic compounds they form. Recall that in ionic compounds, oppositely charged ions combine chemically in definite ratios to form a compound that has no charge. If you add the oxidation number of each ion multiplied by the number of these ions in a formula unit, the total must be zero.

In the chemical formula for any ionic compound, the symbol of the cation is always written first, followed by the symbol of the anion. Subscripts, which are small numbers to the lower right of a symbol, are used to represent the number of ions of each element in an ionic compound. If no subscript is written, it is assumed to be one.

Suppose you need to determine the formula for one formula unit of the compound that contains sodium and chloride ions. Write the symbol and charge for each ion.



The ratio of ions must be such that the number of electrons lost by the metal is equal to the number of electrons gained by the nonmetal. Because the sum of the oxidation numbers of these ions is zero, these ions must be present in a one-to-one ratio. One sodium ion transfers one electron to one chloride ion, and the formula unit is NaCl.

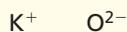
EXAMPLE PROBLEM 8-2

Determining the Formula for an Ionic Compound

The ionic compound formed from potassium and oxygen is used as a dehydrating agent because it reacts readily with water. Determine the correct formula for the ionic compound formed from potassium and oxygen.

1. Analyze the Problem

It is given that potassium and oxygen ions form an ionic compound. The first thing to do is determine the symbol and oxidation number for each ion involved in the ionic compound and write them as shown.



If the charges are not the same, subscripts must be determined to indicate the ratio of positive ions to negative ions.

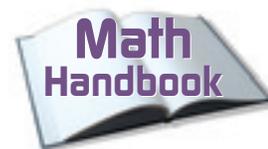
2. Solve for the Unknown

A potassium atom loses one electron while an oxygen atom gains two electrons. If combined in a one-to-one ratio, the number of electrons lost by potassium will not balance the number of electrons gained by oxygen. To have the same number of electrons lost and gained, you must have two potassium ions for every oxide ion. The correct formula is K_2O .

3. Evaluate the Answer

The overall charge on one formula unit of this compound is zero.

$$2 \text{ K-ions } \left(\frac{1+}{\text{K-ion}} \right) + 1 \text{ O-ion } \left(\frac{2-}{\text{O-ion}} \right) = 2(1+) + 1(2-) = 0$$



Review least common multiple in the **Math Handbook** on page 909 of this text.

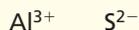
EXAMPLE PROBLEM 8-3

Determining the Formula for an Ionic Compound

Determine the correct formula for the yellowish-gray compound formed from aluminum ions and sulfide ions. This compound decomposes in moist air.

1. Analyze the Problem

You are given that aluminum and sulfur ions form an ionic compound. First, determine the charge of each ion involved.



Each aluminum atom loses three electrons while each sulfur atom gains two. The number of electrons lost must equal the number of electrons gained.

2. Solve for the Unknown

The smallest number that both two and three divide into evenly is six. Therefore, a total of six electrons was transferred. Three sulfur atoms accept the six electrons lost by two aluminum atoms. The correct formula will show two aluminum ions bonded to three sulfur ions, or Al_2S_3 .

3. Evaluate the Answer

The overall charge on one formula unit of this compound is zero.

$$2 \text{ Al-ions } \left(\frac{3+}{\text{Al-ion}} \right) + 3 \text{ S-ions } \left(\frac{2-}{\text{S-ion}} \right) = 2(3+) + 3(2-) = 0$$





For more practice with writing formulas for ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Write the correct formula for the ionic compound composed of the following pairs of ions.

- | | |
|----------------------------|------------------------|
| 19. potassium and iodide | 22. cesium and nitride |
| 20. magnesium and chloride | 23. barium and sulfide |
| 21. aluminum and bromide | |

Compounds that contain polyatomic ions Many ionic compounds contain **polyatomic ions**, which are ions made up of more than one atom. **Table 8-6** lists the formulas and the charges for several polyatomic ions.

The charge given to a polyatomic ion applies to the entire group of atoms. Although an ionic compound containing one or more polyatomic ions contains more than two atoms, the polyatomic ion acts as an individual ion. Therefore, the chemical formula for the compound can be written following the same rules used for a binary compound.

Because a polyatomic ion exists as a unit, never change subscripts of the atoms within the ion. If more than one polyatomic ion is needed, place parentheses around the ion and write the appropriate subscript outside the parentheses. For example, the formula for magnesium chlorate is $\text{Mg}(\text{ClO}_3)_2$. Note that the ammonium ion is the only common polyatomic cation.

How can you determine the formula unit for an ionic compound containing a polyatomic ion? Let's consider the compound formed from the ammonium ion and the chloride ion. First, write each ion along with its charge.



Because the sum of the charges on the ions is zero, the ions are in a one-to-one ratio. The correct formula unit for this compound is NH_4Cl .

Table 8-6

Common Polyatomic Ions			
Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
NO_3^-	nitrate	H_2PO_4^-	dihydrogen phosphate
HSO_4^-	hydrogen sulfate	CO_3^{2-}	carbonate
OH^-	hydroxide	SO_3^{2-}	sulfite
CN^-	cyanide	SO_4^{2-}	sulfate
MnO_4^-	permanganate	$\text{S}_2\text{O}_3^{2-}$	thiosulfate
HCO_3^-	hydrogen carbonate	O_2^{2-}	peroxide
ClO^-	hypochlorite	CrO_4^{2-}	chromate
ClO_2^-	chlorite	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
ClO_3^-	chlorate	HPO_4^{2-}	hydrogen phosphate
ClO_4^-	perchlorate	PO_4^{3-}	phosphate
BrO_3^-	bromate	AsO_4^{3-}	arsenate
IO_3^-	iodate		



EXAMPLE PROBLEM 8-4

Determining the Formula for an Ionic Compound Containing a Polyatomic Ion

The ionic compound formed from the calcium ion and the phosphate ion is a common ingredient in fertilizers. Write the formula for this compound.

1. Analyze the Problem

It is given that calcium and phosphate ions form an ionic compound. You should first write each ion along with its charge.



Because the numerical values of the charges differ, a one-to-one ratio is not possible.

2. Solve for the Unknown

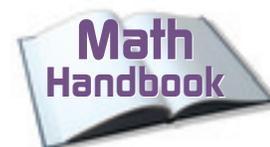
Six is the smallest number evenly divisible by both ionic charges. Therefore, a total of six electrons were transferred. The amount of negative charge of two phosphate ions equals the amount of positive charge of three calcium ions. To use a subscript to indicate more than one unit of a polyatomic ion, you must place the polyatomic ion in parentheses and add the subscript to the outside. The correct formula is $\text{Ca}_3(\text{PO}_4)_2$.

3. Evaluate the Answer

The overall charge on one formula unit of calcium phosphate is zero.



$$3(2+) + 2(-3) = 0$$



Review positive and negative numbers in the **Math Handbook** on page 887 of this text.

PRACTICE PROBLEMS

Determine the correct formula for the ionic compound composed of the following pairs of ions.

24. sodium and nitrate

27. potassium and chromate

25. calcium and chlorate

28. magnesium and carbonate

26. aluminum and carbonate



For more practice with writing formulas for ionic compounds that contain polyatomic ions, go to **Supplemental Practice Problems** in Appendix A.

Naming Ions and Ionic Compounds

You already know how to name monatomic ions. How do you name polyatomic ions? Most polyatomic ions are oxyanions. An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. Many oxyanions contain the same nonmetal and have the same charges but differ in the number of oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the following conventions.

- The ion with more oxygen atoms is named using the root of the nonmetal plus the suffix *-ate*.
- The ion with fewer oxygen atoms is named using the root of the nonmetal plus the suffix *-ite*.

For example:

NO_3^-
nitrate

NO_2^-
nitrite

SO_4^{2-}
sulfate

SO_3^{2-}
sulfite



CONNECTION

Mineralogists, the scientists who study minerals, use various classification schemes to organize the thousands of known minerals. Color; crystal structure; hardness; chemical, magnetic, and electrical properties; and numerous other characteristics are used to classify minerals.

The types of anions minerals contain also can identify them. For example, more than one-third of all known minerals are silicates, which are minerals that contain an anion that is a combination of silicon and oxygen. Halides contain fluoride, chloride, bromide, or iodide ions. Minerals in which the anions contain boron are called borates.



Chlorine in group 7A, the halogens, forms four oxyanions. These oxyanions are named according to the number of oxygen atoms present. The following conventions are used to name these oxyanions.

- *The oxyanion with the greatest number of oxygen atoms is named using the prefix per-, the root of the nonmetal, and the suffix -ate.*
- *The oxyanion with one less oxygen atom is named with the root of the nonmetal and the suffix -ate.*
- *The oxyanion with two fewer oxygen atoms is named using the root of the nonmetal plus the suffix -ite.*
- *The oxyanion with three fewer oxygen atoms is named using the prefix hypo-, the root of the nonmetal, and the suffix -ite.*

ClO_4^-	ClO_3^-	ClO_2^-	ClO^-
perchlorate	chlorate	chlorite	hypochlorite

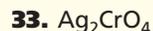
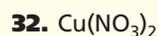
Other halogens form oxyanions that are named similarly to the oxyanions chlorine forms. Bromine forms BrO_3^- , the bromate ion. Iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Naming ionic compounds Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with writing chemical formulas, you will use the following general rules in naming ionic compounds when their formulas are known.

1. *Name the cation first and the anion second.* Remember that the cation is always written first in the formula. For example, CsBr is a compound used in X-ray fluorescent screens. In the formula CsBr , Cs^+ is the cation and is named first. The anion is Br^- and is named second.
2. *Monatomic cations use the element name.* The name of the cation Cs^+ is cesium, the name of the element.
3. *Monatomic anions take their name from the root of the element name plus the suffix -ide.* The compound CsBr contains the bromide anion.
4. *Group 1A and group 2A metals have only one oxidation number. Transition metals and metals on the right side of the periodic table often have more than one oxidation number.* To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation. For example, the compound formed from Fe^{2+} and O^{2-} has the formula FeO and is named iron(II) oxide. The compound formed from Fe^{3+} and O^{2-} has the formula Fe_2O_3 and is named iron(III) oxide.
5. *If the compound contains a polyatomic ion, simply name the ion.* The name of the compound that contains the sodium cation and the polyatomic hydroxide anion, NaOH , is sodium hydroxide. The compound $(\text{NH}_4)_2\text{S}$ is ammonium sulfide.

PRACTICE PROBLEMS

Name the following compounds.



For more practice with naming ionic compounds, go to **Supplemental Practice Problems** in Appendix A.

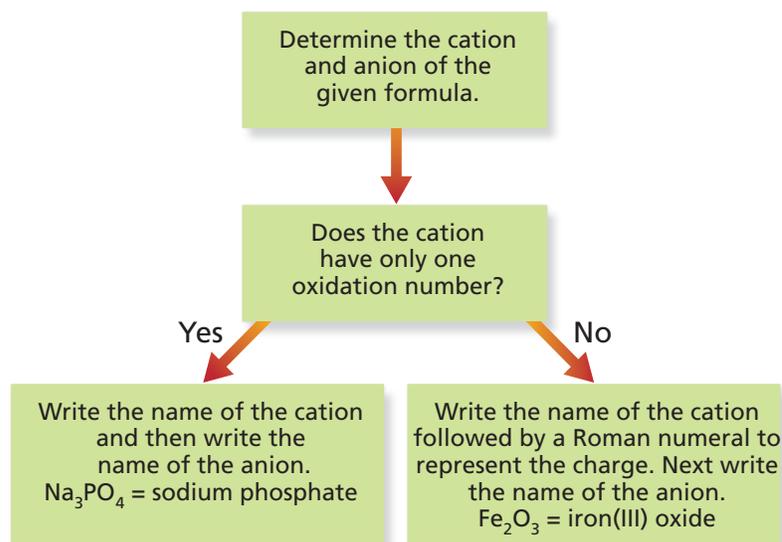


Figure 8-8

This diagram summarizes how to name ionic compounds from their formulas.

Figure 8-8 reviews the steps used in naming ionic compounds if the formula is known. Naming ionic compounds is important in communicating the cation and anion present in a crystalline solid or aqueous solution. How might you change the diagram to help you write the formulas for ionic compounds if you know their names?

All the ion-containing substances you have investigated so far have been ionic compounds. Do any other substances contain ions? Can certain elements contain ions and still be electrically neutral? Do the properties of other ion-containing substances differ from the properties of ionic compounds? In the next section, you will learn the answers to these questions by examining how ions relate to the structure and properties of metals.

Section 8.3 Assessment

34. What is the difference between a monatomic ion and a polyatomic ion? Give an example of each.
35. How do you determine the correct subscripts in a chemical formula?
36. How are metals named in an ionic compound? Nonmetals? Polyatomic ions?
37. What is an oxyanion and how is it named?
38. **Thinking Critically** What subscripts would most likely be used if the following substances formed an ionic compound?
 - a. an alkali metal and a halogen
 - b. an alkali metal and a nonmetal from group 6A
 - c. an alkaline earth metal and a halogen
 - d. an alkaline earth metal and a nonmetal from group 6A
 - e. a metal from group 3A and a halogen

39. **Making and Using Tables** Complete the table below by providing the correct formula for each compound formed from the listed ions.

Formulas for Some Ionic Compounds				
	Oxide	Chloride	Sulfate	Phosphate
Potassium				
Barium				
Aluminum				
Ammonium				

Metallic Bonds and Properties of Metals

Objectives

- **Describe** a metallic bond.
- **Explain** the physical properties of metals in terms of metallic bonds.
- **Define** and **describe** alloys.

Vocabulary

electron sea model
delocalized electrons
metallic bond
alloy

Although metals are not ionic, they share several properties with ionic compounds. Properties of materials are based on bonding, and the bonding in both metals and ionic compounds is based on the attraction of particles with unlike charges.

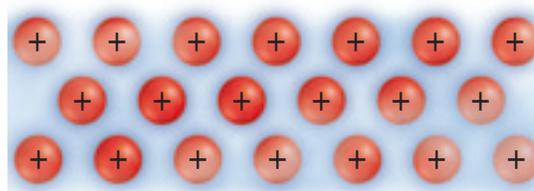
Metallic Bonds

Although metals do not bond ionically, they often form lattices in the solid state. These lattices are similar to the ionic crystal lattices that were discussed in Section 8.2. In such a lattice, eight to 12 other metal atoms surround each metal atom. Although metal atoms have at least one valence electron, they do not share these electrons with neighboring atoms nor do they lose electrons to form ions.

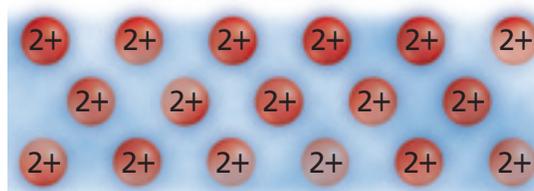
Instead, in this crowded condition, the outer energy levels of the metal atoms overlap. The **electron sea model** proposes that all the metal atoms in a metallic solid contribute their valence electrons to form a “sea” of electrons. The electrons present in the outer energy levels of the bonding metallic atoms are not held by any specific atom and can move easily from one atom to the next. Because they are free to move, they are often referred to as **delocalized electrons**. When the atom’s outer electrons move freely throughout the solid, a metallic cation is formed. Each such ion is bonded to all neighboring metal cations by the “sea” of valence electrons shown in **Figure 8-9**. A **metallic bond** is the attraction of a metallic cation for delocalized electrons.

Properties of metals The typical physical properties of metals can be explained by metallic bonding. These properties provide evidence of the strength of metallic bonds.

The melting points of metals vary greatly. Mercury is a liquid at room temperature, which makes it useful in scientific instruments such as thermometers and barometers. On the other hand, tungsten has a melting point of 3422°C, which makes it useful by itself or in combination with other metals



Group 1A



Group 2A

Figure 8-9

The valence electrons in metals (shown in blue) are evenly distributed among the metallic cations (shown in red). Attractions between the positive cations and negative “sea” hold the metal atoms together in a lattice.

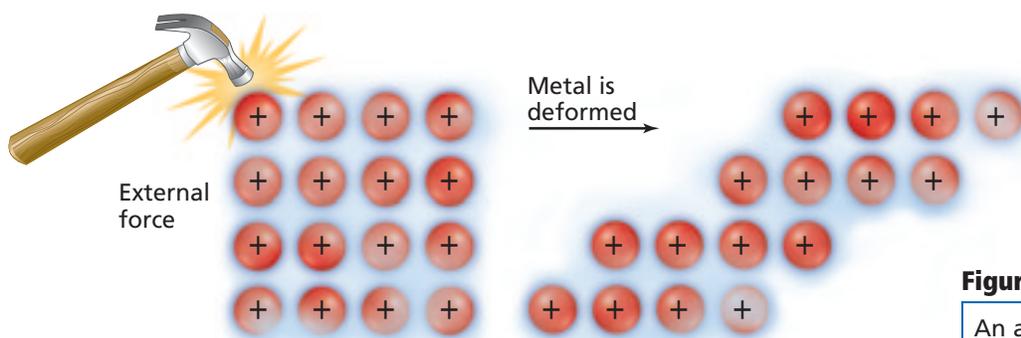


Figure 8-10

An applied force causes metal ions to move through delocalized electrons, making metals malleable and ductile.

for purposes that involve high temperatures or strength. Lightbulb filaments are usually made from tungsten, as are certain spacecraft parts. In general, metals have moderately high melting points and high boiling points, as shown in **Table 8-7**. The melting points are not as extreme as the boiling points because the cations and electrons are mobile in a metal. It does not take an extreme amount of energy for them to be able to move past each other. However, during boiling, atoms must be separated from the group of cations and electrons, which requires much more energy.

Metals are malleable, which means they can be hammered into sheets, and they are ductile, which means they can be drawn into wire. **Figure 8-10** shows how the mobile particles involved in metallic bonding can be pushed or pulled past each other, making metals malleable and ductile.

Metals are generally durable. Although metallic cations are mobile in a metal, they are strongly attracted to the electrons surrounding them and aren't easily removed from the metal.

Delocalized electrons in a metal are free to move, keeping metallic bonds intact. The movement of mobile electrons around positive metallic cations explains why metals are good conductors. The delocalized electrons move heat from one place to another much more quickly than the electrons in a material that does not contain mobile electrons. Mobile electrons easily move as a part of an electric current when electrical potential is applied to a metal. These same delocalized electrons interact with light, absorbing and releasing photons, thereby creating the property of luster in metals.

The mobile electrons in transition metals consist not only of the two outer s electrons but also the inner d electrons. As the number of delocalized electrons increases, so do the properties of hardness and strength. For example, strong metallic bonds are found in transition metals such as chromium, iron, and nickel, whereas alkali metals are considered soft because they have only one delocalized electron, ns^1 .

Table 8-7

Melting Points and Boiling Points of Some Metals		
Element	Melting point (°C)	Boiling point (°C)
Lithium	180	1347
Tin	232	2623
Aluminum	660	2467
Barium	727	1850
Silver	961	2155
Copper	1083	2570

Metal Alloys

Due to the nature of a metallic bond, it is relatively easy to introduce other elements into a metallic crystal, forming an alloy. An **alloy** is a mixture of elements that has metallic properties. **Table 8-8** lists some commercially important alloys and their uses. A company that manufactures trophies probably would use which alloy listed in the table?

The properties of alloys differ somewhat from the properties of the elements they contain. For example, steel is iron mixed with at least one other element. Some properties of iron are present, but steel has additional properties, such as increased strength. Some alloys, such as that used in the **miniLAB**, vary in properties depending on how they are manufactured.

miniLAB

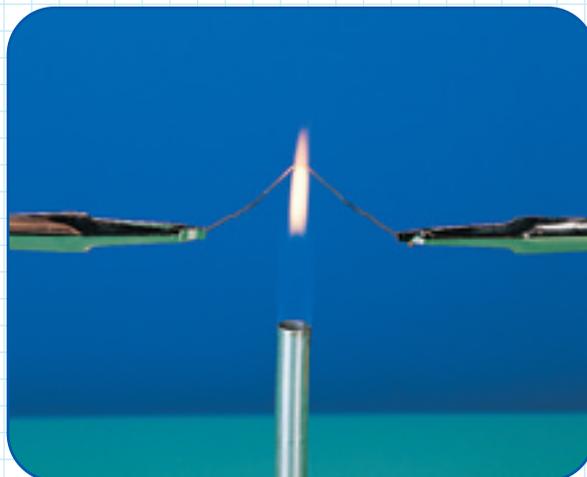
Heat Treatment of Steel

Recognizing Cause and Effect People have treated metals with heat for many centuries. Different properties result when the metal is slowly or rapidly cooled. Can you determine how and why the properties change?

Materials laboratory burner, forceps (2), hairpins (3), 250-mL beaker

Procedure 

1. Examine a property of spring steel by trying to bend open one of the hairpins. Record your observations.
2. Hold each end of a hairpin with forceps. Place the curved central loop in the top of the burner's flame. When it turns red, pull it open into a straight piece of metal. Allow it to cool as you record your observations. Repeat this procedure for the remaining two hairpins.
CAUTION: Do not touch the hot metal.
3. To make softened steel, use forceps to hold all three hairpins vertically in the flame until they glow red all over. Slowly raise the three hairpins straight up and out of the flame so they cool slowly. Slow cooling results in the formation of large crystals.
4. After cooling, bend each of the three hairpins into the shape of the letter J. Record how the metal feels as you bend it.
5. To harden the steel, use tongs to hold two of the bent hairpins in the flame until they are glowing red all over. Quickly plunge the hot metals into a 250-mL beaker containing approximately 200 mL of cold water. Quick-cooling causes the crystal size to be small.
6. Attempt to straighten one of the bends. Record your observations.



7. To temper the steel, use tongs to briefly hold the remaining hardened metal bend above the flame. Slowly move the metal back and forth just above the flame until the gray metal turns to an iridescent blue-gray color. Do not allow the metal to glow red. Slowly cool the metal and then try to unbend it using the end of your finger. Record your observations.

Analysis

1. State a use for spring steel that takes advantage of its unique properties.
2. What are the advantages and disadvantages of using softened steel for body panels on automobiles?
3. What is the major disadvantage of hardened steel? Do you think this form of iron would be wear resistant and retain a sharpened edge?
4. Which two types of steel appear to have their properties combined in tempered steel?
5. State a hypothesis that explains how the different properties you have observed relate to crystal size.

Table 8-8

Some Commercially Important Alloys		
Common name	Composition	Uses
Alnico	Fe 50%, Al 20%, Ni 20%, Co 10%	Magnets
Brass	Cu 67-90%, Zn 10-33%	Plumbing, hardware, lighting
Bronze	Cu 70-95%, Zn 1-25%, Sn 1-18%	Bearings, bells, medals
Cast iron	Fe 96-97%, C 3-4%	Casting
Dental amalgam	Hg 50%, Ag 35%, Sn 15%	Dental fillings
Gold, 10 carat	Au 42%, Ag 12-20%, Cu 38-46%	Jewelry
Lead shot	Pb 99.8%, As (0.2%)	Shotgun shells
Pewter	Sn 70-95%, Sb 5-15%, Pb 0-15%	Tableware
Stainless steel	Fe 73-79%, Cr 14-18%, Ni 7-9%	Instruments, sinks
Sterling silver	Ag 92.5%, Cu 7.5%	Tableware, jewelry

Alloys most commonly form when the elements involved are either similar in size or the atoms of one element are considerably smaller than the atoms of the other. Thus, two basic types of alloys exist, substitutional and interstitial, and many industries depend on their production. A substitutional alloy has atoms of the original metallic solid replaced by other metal atoms of similar size. Sterling silver is an example of a substitutional alloy. When copper atoms replace silver atoms in the original metallic crystal, a solid with properties of both silver and copper is formed. Brass, pewter, and 10-carat gold are all examples of substitutional alloys.

An interstitial alloy is formed when the small holes (interstices) in a metallic crystal are filled with smaller atoms. Forming this type of alloy is similar to pouring sand into a bucket of gravel. Even if the gravel is tightly packed, holes exist between the pieces. The sand does not replace any of the gravel but fills in the spaces. The best-known interstitial alloy is carbon steel. Holes in the iron crystal are filled with carbon atoms, and the physical properties of iron are changed. Iron is relatively soft and malleable. However, the presence of carbon makes the solid harder, stronger, and less ductile than pure iron, increasing its uses.

Section 8.4 Assessment

40. What is a metallic bond?
41. Explain how conductivity of electricity and high melting point of metals are explained by metallic bonding.
42. What is an alloy?
43. How does a substitutional alloy differ from an interstitial alloy?
44. **Thinking Critically** In the laboratory, how could you determine if a solid has an ionic bond or a metallic bond?
45. **Formulating Models** Draw a model to represent the ductility of a metal using the electron sea model shown in **Figure 8-10**.



Making Ionic Compounds

Elements combine to form compounds. If energy is released as the compound is formed, the resulting product is more stable than the reacting elements. In this investigation you will react elements to form two compounds. You will test the compounds to determine several of their properties. Ionic compounds have properties that are different from those of other compounds. You will decide if the products you formed are ionic compounds.

Problem

What are the formulas and names of the products that are formed? Do the properties of these compounds classify them as having ionic bonds?

Objectives

- **Observe** evidence of a chemical reaction.
- **Acquire** and **analyze** information that will enable you to decide if a compound has an ionic bond.
- **Classify** the products as ionic or not ionic.

Materials

magnesium ribbon
crucible
ring stand and ring
clay triangle
laboratory burner
stirring rod

crucible tongs
centigram balance
100-mL beaker
distilled water
conductivity tester

Safety Precautions



- Always wear safety glasses and a lab apron.
- Do not look directly at the burning magnesium. The intensity of the light can damage your eyes.
- Avoid handling heated materials until they have cooled.

Pre-Lab

1. Read the entire procedure. Identify the variables. List any conditions that must be kept constant.
2. Write the electron configuration of the magnesium atom.
 - a. Based on this configuration, will magnesium lose or gain electrons to become a magnesium ion?
 - b. Write the electron configuration of the magnesium ion.
 - c. The magnesium ion has an electron configuration like that of which noble gas?
3. Repeat question 2 for oxygen and nitrogen.
4. Prepare your data table.
5. In your data table, which mass values will be measured directly? Which mass values will be calculated?
6. Explain what must be done to calculate each mass value that is not measured directly.

Mass Data	
Material(s)	Mass (g)
Empty crucible	
Crucible and Mg ribbon before heating	
Magnesium ribbon	
Crucible and magnesium products after heating	
Magnesium products	

Procedure

1. Arrange the ring on the ring stand so that it is about 7 cm above the top of the Bunsen burner. Place the clay triangle on the ring.
2. Measure the mass of the clean, dry crucible, and record the mass in the data table.
3. Roll 25 cm of magnesium ribbon into a loose ball. Place it in the crucible. Measure the mass of the magnesium and crucible and record this mass in the data table.
4. Place the crucible on the clay ring. Heat the crucible with a hot flame, being careful to position the crucible near the top of the flame.



5. When the magnesium metal ignites and begins to burn with a bright white light, immediately turn off the laboratory burner. **CAUTION:** Do not look directly at the burning magnesium. After the magnesium product and crucible have cooled, measure their mass and record it in the data table.
6. Place the dry solid product in a small beaker for further testing.
7. Add 10 mL of distilled water to the dry magnesium product in the beaker and stir. Check the mixture with a conductivity checker, and record your results.

Cleanup and Disposal

1. Wash out the crucible with water.
2. Dispose of the product as directed by your teacher.
3. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Analyzing Data** Use the masses in the table to calculate the mass of the magnesium ribbon and the mass of the magnesium product. Record these masses in the table.
2. **Classifying** What kind of energy was released by the reaction? What can you conclude about the product of this reaction?
3. **Using Numbers** How do you know that the magnesium metal reacts with certain components of the air?
4. **Predicting** Magnesium reacts with both oxygen and nitrogen from the air at the high temperature of the crucible. Predict the binary formulas for both products. Write the names of these two compounds.
5. **Analyzing and Concluding** The product formed from magnesium and oxygen is white, and the product formed from magnesium and nitrogen is yellow. From your observations, which compound makes up most of the product?
6. **Analyzing and Concluding** Did the magnesium compounds and water conduct an electric current? Do the results indicate whether or not the compounds are ionic?
7. **Error Analysis** If the magnesium lost mass instead of gaining mass, what do you think was a possible source of the error?

Real-World Chemistry

1. The magnesium ion plays an important role in a person's biochemistry. Research the role of this electrolyte in your physical and mental health. Is magnesium listed as a component in a multi-vitamin and mineral tablet?
2. Research the use of $\text{Mg}(\text{OH})_2$ in everyday products. What is $\text{Mg}(\text{OH})_2$ commonly called in over-the-counter drugs?

Everyday Chemistry

Colors of Gems

Have you ever wondered what produces the gorgeous colors in a stained-glass window or in the rubies, emeralds, and sapphires mounted on a ring? Compounds of transition elements are responsible for creating the entire spectrum of colors.

Transition elements color gems and glass

Transition elements have many important uses, but one that is often overlooked is their role in giving colors to gemstones and glass. Although not all compounds of transition elements are colored, most inorganic colored compounds contain a transition element such as chromium, iron, cobalt, copper, manganese, nickel, cadmium, titanium, gold, or vanadium. The color of a compound is determined by the identity of the metal, its oxidation number, and the negative ion combined with it.

Impurities give gemstones their color

Crystals have fascinating properties. A clear, colorless quartz crystal is pure silicon dioxide (SiO_2). But a crystal that is colorless in its pure form may exist as a variety of colored gemstones when tiny amounts of transition element compounds, usually oxides, are present. Amethyst (purple), citrine (yellow-brown), and rose quartz (pink) are quartz crystals with transition element impurities scattered throughout. Blue sapphires are composed of aluminum oxide (Al_2O_3) with the impurities iron(II) oxide (FeO) and titanium(IV) oxide (TiO_2). If trace amounts of chromium(III) oxide (Cr_2O_3) are

present in the Al_2O_3 , the resulting gem is a red ruby. A second kind of gemstone is one composed entirely of a colored compound. Most are transition element compounds, such as rose-red rhodochrosite (MnCO_3), black-grey hematite (Fe_2O_3), or green malachite ($\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2$).

How metal ions interact with light to produce color

Why does the presence of Cr_2O_3 in Al_2O_3 make a ruby red? The Cr^{3+} ion absorbs yellow-green colors from white light striking the ruby, and the remaining red-blue light is transmitted, resulting in a deep red color. This same process occurs in all gems. Trace impurities absorb certain colors of light from white light striking or passing through the stone. The remaining colors of light that are reflected or transmitted produce the color of the gem.



Adding transition elements to molten glass for color

Glass is colored by adding transition element compounds to the glass while it is molten. This process is used for stained glass, glass used in glass blowing, and even glass in the form of ceramic glazes. Most of the coloring agents are oxides. When oxides of copper or cobalt are added to molten glass, the glass is blue; oxides of manganese produce purple glass; iron oxides, green; gold oxides, deep ruby red; copper or selenium oxides, red; and antimony oxides, yellow. Some coloring compounds are not oxides. Chromates, for example, produce green glass, and iron sulfide gives a brown color.

Testing Your Knowledge

- 1. Applying** Explain why iron(III) sulfate is yellow, iron(II) thiocyanate is green, and iron(III) thiocyanate is red.
- 2. Acquiring Information** Find out what impurities give amethyst, rose quartz, and citrine their colors.
- 3. Comparing and Contrasting** Conduct research to find the similarities and differences between synthetic and natural gemstones.

Summary

8.1 Forming Chemical Bonds

- A chemical bond is the force that holds two atoms together.
- Atoms that form ions gain or lose valence electrons to achieve the same electron arrangement as that of a noble gas, which is a stable configuration. This noble gas configuration involves a complete outer electron energy level, which usually consists of eight valence electrons.
- A positive ion, or cation, forms when valence electrons are removed and a stable electron configuration is obtained.
- A negative ion, or anion, forms when valence electrons are added to the outer energy level, giving the ion a stable electron configuration.

8.2 The Formation and Nature of Ionic Bonds

- An ionic bond forms when anions and cations close to each other attract, forming a tightly packed geometric crystal lattice.
- Lattice energy is needed to break the force of attraction between oppositely charged ions arranged in a crystal lattice.
- The physical properties of ionic solids, such as melting point, boiling point, hardness, and the ability to conduct electricity in the molten state and as an aqueous solution, are related to the strength of the ionic bonds and the presence of ions.
- An ionic compound is an electrolyte because it conducts an electric current when it is liquid or in aqueous solution.

8.3 Names and Formulas for Ionic Compounds

- Subscripts in an ionic compound indicate the ratio

of cations and anions needed to form electrically neutral compounds. The formula unit represents the ratio of these ions in the crystal lattice.

- If the element that forms the cation has more than one possible oxidation number, Roman numerals are used to indicate the oxidation number present for that element in the compound.
- Ions formed from only one atom are monatomic ions. The charge on a monatomic ion is its oxidation number, or oxidation state.
- Polyatomic ions are two or more atoms bonded together that act as a single unit with a net charge. Many polyatomic ions are oxyanions, containing an atom, usually a nonmetal, and oxygen atoms.
- In a chemical formula, polyatomic ions are placed inside parentheses when using a subscript.
- Ionic compounds are named by the name of the cation followed by the name of the anion.

8.4 Metallic Bonds and Properties of Metals

- Metallic bonds are formed when metal cations attract free valence electrons. A “sea” of electrons moves throughout the entire metallic crystal, producing this attraction.
- The electrons involved in metallic bonding are called delocalized electrons because they are free to move throughout the metal and are not attached to a particular atom.
- The electron sea model can explain the melting point, boiling point, malleability, conductivity, and ductility of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements. The two common types of alloys are substitutional and interstitial.

Vocabulary

- | | | |
|----------------------------------|-------------------------------|-----------------------------|
| • alloy (p. 230) | • electrolyte (p. 218) | • metallic bond (p. 228) |
| • anion (p. 214) | • electron sea model (p. 228) | • monatomic ion (p. 221) |
| • cation (p. 212) | • formula unit (p. 221) | • oxidation number (p. 222) |
| • chemical bond (p. 211) | • ionic bond (p. 215) | • oxyanion (p. 225) |
| • delocalized electrons (p. 228) | • lattice energy (p. 219) | • polyatomic ion (p. 224) |

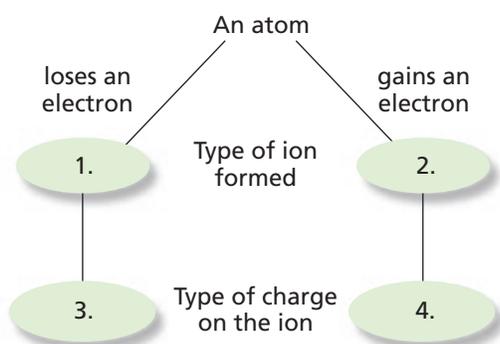


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 8 Assessment.

Concept Mapping

46. Complete the concept map, showing what type of ion is formed in each case and what type of charge the ion has.



Mastering Concepts

47. When do chemical bonds form? (8.1)
48. Why do positive ions and negative ions form? (8.1)
49. Why are halogens and alkali metals likely to form ions? Explain your answer. (8.1)
50. Discuss the importance of electron affinity and ionization energy in the formation of ions. (8.1)
51. Discuss the formation of ionic bonds. (8.2)
52. Briefly discuss three physical properties of ionic solids that are linked to ionic bonds. (8.2)
53. What does the term *electrically neutral* mean when discussing ionic compounds? (8.2)
54. What information is needed to write a correct chemical formula to represent an ionic compound? (8.3)
55. When are subscripts used in formulas for ionic compounds? (8.3)
56. Discuss how an ionic compound is named. (8.3)
57. Describe a metallic bond. (8.4)
58. Briefly explain how malleability and ductility of metals are explained by metallic bonding. (8.4)
59. Compare and contrast the two types of metal alloys. (8.4)

Mastering Problems Ion Formation (8.1)

60. Explain why noble gases are not likely to form chemical bonds.
61. Give the number of valence electrons in an atom of each of the following:

a. cesium	d. zinc
b. rubidium	e. strontium
c. gallium	
62. Discuss the formation of the barium ion.
63. Explain how an anion of nitrogen forms.
64. The more reactive an atom, the higher its potential energy. Which atom has higher potential energy, neon or fluorine? Explain.
65. Predict the reactivity of the following atoms based on their electron configurations.

a. potassium
b. fluorine
c. neon
66. Discuss the formation of the iron ion that has a 3+ oxidation number.

Ionic Bonds and Ionic Compounds (8.2)

67. Determine the ratio of cations to anions for the following ionic compounds.

a. potassium chloride, a salt substitute
b. calcium fluoride, used in the steel industry
c. aluminum oxide, known as corundum in the crystalline form
d. calcium oxide, used to remove sulfur dioxide from power plant exhaust
e. strontium chloride, used in fireworks
68. Using orbital notation, diagram the formation of an ionic bond between aluminum and fluorine.
69. Using electron configurations, diagram the formation of an ionic bond between barium and nitrogen.
70. Discuss the formation of an ionic bond between zinc and oxygen.
71. Under certain conditions, ionic compounds conduct an electric current. Describe these conditions and explain why ionic compounds are not always used as conductors.
72. Which of the following compounds are not likely to occur: CaKr, Na₂S, BaCl₃, MgF? Explain your choices.
73. Using Table 8-2, determine which of the following ionic compounds will have the highest melting point: MgO, KI, or AgCl. Explain your answer.

Formulas and Names for Ionic Compounds (8.3)

- 74.** Give the formula for each of the following ionic compounds.
- a. calcium iodide d. potassium periodate
 b. silver(I) bromide e. silver(I) acetate
 c. copper(II) chloride
- 75.** Name each of the following ionic compounds.
- a. K_2O d. $NaClO$
 b. $CaCl_2$ e. KNO_3
 c. Mg_3N_2
- 76.** Complete **Table 8-9** by placing the symbols, formulas, and names in the blanks.

Table 8-9

Identifying Ionic Compounds			
Cation	Anion	Name	Formula
		ammonium sulfate	
			PbF_2
		lithium bromide	
			Na_2CO_3
Mg^{2+}	PO_4^{3-}		

- 77.** Chromium, a transition metal, forms both the Cr^{2+} and Cr^{3+} ions. Write the formulas for the ionic compounds formed when each of these ions react with
- a. fluorine b. oxygen
- 78.** Which of the following are correct formulas for ionic compounds? For those that are not correct, give the correct formula and justify your answer.
- a. $AlCl$ d. $BaOH_2$
 b. Na_3SO_4 e. Fe_2O
 c. $MgCO_3$
- 79.** Write the formulas for all of the ionic compounds that can be formed by combining each of the cations with each of the anions listed below. Name each compound formed.

Table 8-10

Cations	Anions
K^+	SO_3^{2-}
NH_4^+	I^-
Fe^{3+}	NO_3^-

Metals and Metallic Bonds (8.4)

- 80.** How is a metallic bond different from an ionic bond?
- 81.** Briefly explain why silver is a good conductor of electricity.
- 82.** Briefly explain why iron is used in making the structures of many buildings.
- 83.** The melting point of beryllium is $1287^\circ C$, while that of lithium is $180^\circ C$. Account for the large difference in values.
- 84.** Describe the difference between the metal alloy sterling silver and carbon steel in terms of the types of alloys involved.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 85.** Give the number of valence electrons for atoms of oxygen, sulfur, arsenic, phosphorus, and bromine.
- 86.** Explain why calcium can form a Ca^{2+} ion but not a Ca^{3+} ion.
- 87.** Which of the following ionic compounds would have the most negative lattice energy: $NaCl$, KCl , or $MgCl_2$? Explain your answer.
- 88.** Give the formula for each of the following ionic compounds.
- a. sodium sulfide d. calcium phosphate
 b. iron(III) chloride e. zinc nitrate
 c. sodium sulfate
- 89.** Cobalt, a transition metal, forms both the Co^{2+} and Co^{3+} ions. Write the correct formulas and give the name for the oxides formed by the two different ions.
- 90.** Briefly explain why gold can be used as both a conductor in electronic devices and in jewelry.
- 91.** Discuss the formation of the nickel ion with a 2+ oxidation number.
- 92.** Using electron-dot structure, diagram the formation of an ionic bond between potassium and iodine.
- 93.** Magnesium forms both an oxide and a nitride when burned in air. Discuss the formation of magnesium oxide and magnesium nitride when magnesium atoms react with oxygen and nitrogen atoms.
- 94.** An external force easily deforms sodium metal, while sodium chloride shatters when the same amount of force is applied. Why do these two solids behave so differently?

- 95.** Name each of the following ionic compounds.
- CaO
 - BaS
 - AlPO₄
 - Ba(OH)₂
 - Sr(NO₃)₂
- 96.** Write the formulas for all of the ionic compounds that can be formed by combining each of the cations with each of the anions listed below. Name each compound formed.

Table 8-11

Cations	Anions
Ba ²⁺	S ₂ O ₃ ²⁻
Cu ⁺	Br ⁻
Al ³⁺	NO ₂ ⁻

Thinking Critically

- 97. Concept Mapping** Design a concept map to explain the physical properties of both ionic compounds and metallic solids.
- 98. Predicting** Predict which solid in each of the following will have the higher melting point. Explain your answer.
- NaCl or CsCl
 - Ag or Cu
 - Na₂O or MgO
- 99. Comparing and Contrasting** Compare and contrast cations and anions.
- 100. Observing and Inferring** From the following incorrect formulas and formula names, identify the mistakes and design a flow chart to prevent the mistakes.
- copper acetate
 - Mg₂O₂
 - Pb₂O₅
 - disodium oxide
 - Al₂SO₄₃
- 101. Hypothesizing** Look at the locations of potassium and calcium on the periodic table. Form a hypothesis as to why the melting point of calcium is considerably higher than the melting point of potassium.
- 102. Drawing a Conclusion** Explain why the term *delocalized* is an appropriate term for the electrons involved in metallic bonding.
- 103. Applying Concepts** All uncharged atoms have valence electrons. Explain why elements such as iodine and sulfur don't have metallic bonds.
- 104. Drawing a Conclusion** Explain why lattice energy is a negative quantity.

Writing in Chemistry

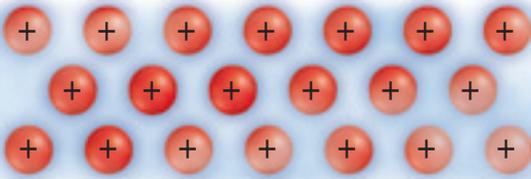
- 105.** Many researchers believe that free radicals are responsible for the effects of aging and cancer. Research free radicals and write about the cause and what can be done to prevent free radicals.
- 106.** Crystals of ionic compounds can be easily grown in the laboratory setting. Research the growth of crystals and try to grow one crystal in the laboratory.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 107.** You are given a liquid of unknown density. The mass of a graduated cylinder containing 2.00 mL of the liquid is 34.68 g. The mass of the empty graduated cylinder is 30.00 g. What is the density of the liquid? (Chapter 2)
- 108.** A mercury atom drops from 1.413×10^{-18} J to 1.069×10^{-18} J. (Chapter 5)
- What is the energy of the photon emitted by the mercury atom?
 - What is the frequency of the photon emitted by the mercury atom?
 - What is the wavelength of the photon emitted by the mercury atom?
- 109.** Which element has the greater ionization energy, chlorine or carbon? (Chapter 6)
- 110.** Compare and contrast the way metals and nonmetals form ions and explain why they are different. (Chapter 6)
- 111.** What are transition elements? (Chapter 6)
- 112.** Write the symbol and name of the element that fits each description. (Chapter 6)
- the second-lightest of the halogens
 - the metalloid with the lowest period number
 - the only group 6A element that is a gas at room temperature
 - the heaviest of the noble gases
 - the group 5A nonmetal that is a solid at room temperature
- 113.** Which group 4A element is (Chapter 7)
- a metalloid that occurs in sand?
 - a nonmetal?
 - used in electrodes in car batteries?
 - a component in many alloys?

Use these questions and the test-taking tip to prepare for your standardized test.

- Which of the following is NOT true of the Sc^{3+} ion?
 - It has the same electron configuration as Ar.
 - It is a scandium ion with three positive charges.
 - It is considered to be a different element than a neutral Sc atom.
 - It was formed by the removal of the valence electrons of Sc.
- Of the salts below, it would require the most energy to break the ionic bonds in _____.
 - BaCl_2
 - LiF
 - NaBr
 - KI
- What is the correct chemical formula for the ionic compound formed by the calcium ion (Ca^{2+}) and the acetate ion ($\text{C}_2\text{H}_3\text{O}_2^-$)?
 - $\text{CaC}_2\text{H}_3\text{O}_2$
 - $\text{CaC}_4\text{H}_6\text{O}_8$
 - $(\text{Ca})_2\text{C}_2\text{H}_3\text{O}_2$
 - $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
- 

The model above has been proposed to explain why _____.

 - metals are shiny, reflective substances
 - metals are excellent conductors of heat and electricity
 - ionic compounds are malleable compounds
 - ionic compounds are good conductors of electricity
- Yttrium, a metallic element with atomic number 39, will form _____.
 - positive ions
 - negative ions
 - both positive and negative ions
 - no ions at all
- The high strength of its ionic bonds results in all of the following properties of NaCl EXCEPT _____.
 - hard crystals
 - high boiling point
 - high melting point
 - low solubility

Interpreting Tables Use the table below to answer questions 7–10.

Properties of Some Ionic Compounds			
Compound	Lattice Energy (kJ/mol)	Melting Point (°C)	Color
Ag_2Se	-2686	?	gray
AlPO_4	?	1460	white
FeI_2	-2439	?	reddish-purple
RbClO_4	?	281	white

- What is the correct name of the compound with the formula RbClO_4 ?
 - rubidium chlorine oxide
 - rubidium chloride tetroxide
 - rubidium perchlorate
 - rubidium chlorate
- Rank the compounds in order of increasing melting point.
 - Ag_2Se , AlPO_4 , FeI_2 , RbClO_4
 - RbClO_4 , FeI_2 , Ag_2Se , AlPO_4
 - AlPO_4 , Ag_2Se , FeI_2 , RbClO_4
 - RbClO_4 , AlPO_4 , Ag_2Se , FeI_2
- Which compound is expected to have the strongest attraction between its ions?
 - Ag_2Se
 - AlPO_4
 - FeI_2
 - RbClO_4
- The anion in AlPO_4 has a charge of _____.
 - 2+
 - 3+
 - 2-
 - 3-



TEST-TAKING TIP

Work Weak Muscles; Maintain Strong Ones

If you're preparing for a standardized test that covers many topics, it's sometimes difficult to focus on all the topics that require your attention. Ask yourself "What's my strongest area?" and "What's my weakest area?" Focus most of your energy on your weaker area and review your stronger topics less frequently.

Covalent Bonding

What You'll Learn

- ▶ You will analyze the nature of a covalent bond.
- ▶ You will name covalently bonded groups of atoms.
- ▶ You will determine the shapes of molecules.
- ▶ You will describe characteristics of covalent molecules.
- ▶ You will compare and contrast polar and nonpolar molecules.

Why It's Important

Most compounds, including those in living organisms, are covalently bonded.

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Visit the Chemistry Web site at science.glencoe.com to find links about electrochemistry.

Herbicides and fertilizers used on crops are covalent compounds.



DISCOVERY LAB



Materials

Beral-type pipette
vinegar
vegetable oil

Oil and Vinegar Dressing

When preparing a meal, you combine different types of food. But when you mix different substances, do they always “mix”? How about making oil and vinegar dressing for tonight’s salad?

Safety Precautions



Procedure

1. Fill the bulb of a Beral-type pipet about 1/3 full of vinegar and 1/3 full of vegetable oil. Shake the pipet and its contents. Record your observations.
2. Allow the contents to sit for about five minutes. Record your observations.

Analysis

Do oil and vinegar mix? What explanation can you give for your observations in this experiment? Why do the instructions on many types of salad dressings read “shake well before using”?

Section

9.1

The Covalent Bond

Objectives

- **Apply** the octet rule to atoms that bond covalently.
- **Describe** the formation of single, double, and triple covalent bonds.
- **Compare** and **contrast** sigma and pi bonds.
- **Relate** the strength of covalent bonds to bond length and bond dissociation energy.

Vocabulary

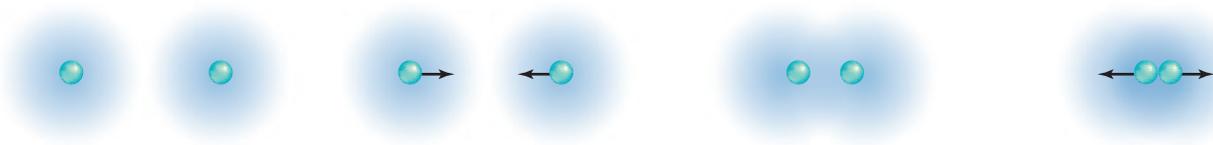
covalent bond
molecule
Lewis structure
sigma bond
pi bond
endothermic
exothermic

Worldwide, scientists are studying ways to increase food supplies, reduce pollution, and prevent disease. Understanding the chemistry of compounds that make up fertilizers, pollutants, and materials that carry genetic information is essential in developing new technologies in these areas. An understanding of the chemistry of compounds requires an understanding of their bonding.

Why do atoms bond?

You learned in Chapter 6 that all noble gases have particularly stable electron arrangements. This stable arrangement consists of a full outer energy level. A full outer energy level consists of two valence electrons for helium and eight valence electrons for all other noble gases. Because of this stability, noble gases, in general, don’t react with other elements to form compounds.

You also learned in Chapter 8 that when metals and nonmetals react to form binary ionic compounds, electrons are transferred, and the resulting ions have noble-gas electron configurations. But sometimes two atoms that both need to gain valence electrons to become stable have a similar attraction for electrons. Sharing of electrons is another way that these atoms can acquire the electron configuration of noble gases. Recall from Chapter 6 that the octet rule states that atoms lose, gain, or share electrons to achieve a stable configuration of eight valence electrons, or an octet. Although exceptions to the octet rule exist, the rule provides a useful framework for understanding chemical bonds.



a The atoms are too far away from each other to have noticeable attraction or repulsion.

b Each nucleus attracts the other atom's electron cloud, but the electron clouds repel each other.

c The distance is right for the attraction of one atom's protons for the other atom's electrons to make the bond stable.

d If the atoms are forced closer together, the nuclei and electrons repel.

Figure 9-1

The overall force between two atoms is the result of electron-electron repulsion, nucleus-nucleus repulsion, and nucleus-electron attraction. The arrows in this diagram show the net force acting on two fluorine atoms as they move toward each other.

What is a covalent bond?

You know that certain atoms, such as magnesium and chlorine, transfer electrons from one atom to another, forming an ionic bond. However, the number of ionic compounds is quite small compared with the total number of known compounds. What type of bonding is found in all these other compounds that are not ionically bonded?

The atoms in these other compounds share electrons. The chemical bond that results from the sharing of valence electrons is a **covalent bond**. In a covalent bond, the shared electrons are considered to be part of the complete outer energy level of both atoms involved. Covalent bonding generally occurs when elements are relatively close to each other on the periodic table. The majority of covalent bonds form between nonmetallic elements.

A **molecule** is formed when two or more atoms bond covalently. The carbohydrates and simple sugars you eat; the proteins, fats, and DNA found in your body; and the wool, cotton, and synthetic fibers in the clothes you wear all consist of molecules formed from covalently bonded atoms.

Formation of a covalent bond Hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2) occur in nature as diatomic molecules, not as single atoms because the molecules formed are more stable than the individual atoms. How do two atoms that do not give up electrons bond with each other?

Consider fluorine (F_2), which has an electron configuration of $1s^22s^22p^5$. Each fluorine atom has seven valence electrons and must have one additional electron to form an octet. As two fluorine atoms approach each other, as shown in **Figure 9-1**, two forces become important. A repulsive force occurs between the like-charged electrons and between the like-charged protons of the two atoms. An attractive force also occurs between the protons of one fluorine atom and the electrons of the other atom. As the fluorine atoms move closer, the attraction of both nuclei for the other atom's electrons increases until the maximum attraction is achieved. At the point of maximum attraction, the attractive forces balance the repulsive forces.

If the two nuclei move even closer, the repulsion between the like-charged nuclei and electron clouds will increase, resulting in repulsive forces that exceed attractive forces. Thus the most stable arrangement of atoms exists at the point of maximum attraction. At that point, the two atoms bond covalently and a molecule forms. Fluorine exists as a diatomic molecule because the sharing of one pair of electrons will give both fluorine atoms stable noble gas configurations. Each fluorine atom in the fluorine molecule has one bonding pair of electrons and three *lone pairs*, which are unshared pairs of electrons.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Single Covalent Bonds

Consider the formation of a hydrogen molecule, which is shown in **Figure 9-2**. Each covalently bonded atom equally attracts one pair of shared electrons. Thus, two electrons shared by two hydrogen nuclei belong to each atom simultaneously. Both hydrogen atoms have the noble gas configuration of helium ($1s^2$). The hydrogen molecule is more stable than individual hydrogen atoms.

When a single pair of electrons is shared, such as in a hydrogen molecule, a single covalent bond forms. The shared electron pair, often referred to as the bonding pair, is represented by either a pair of dots or a line in the Lewis structure for the molecule. **Lewis structures** use electron-dot diagrams to show how electrons are arranged in molecules. For example, a hydrogen molecule is represented as $H:H$ or $H-H$. Hydrogen gas also is represented by the molecular formula H_2 , which reflects the number of atoms in each molecule.

As you have seen, the halogens—group 7A elements—such as fluorine, have seven valence electrons. To attain an octet, one more electron is necessary. Therefore, group 7A elements will form a single covalent bond. You have seen how an atom from group 7A will form a covalent bond with another identical atom. Fluorine exists as F_2 . Similarly, chlorine exists as Cl_2 , bromine as Br_2 , and iodine as I_2 because the molecule formed is more stable than the individual atoms. In addition, such bonds are often formed between the halogen and another element, such as carbon.

Group 6A elements share two electrons to form two covalent bonds. Oxygen is a group 6A element with an electron configuration of $1s^2 2s^2 2p^4$. Water is composed of two hydrogen atoms and one oxygen atom. Each hydrogen atom attains the noble gas configuration of helium as it shares one electron with oxygen. Oxygen, in turn, attains the noble gas configuration of neon by sharing one electron with each hydrogen atom. **Figure 9-3a** shows the Lewis structure for a molecule of water. Notice that two single covalent bonds are formed and two lone pairs remain on the oxygen atom.

Likewise, group 5A elements form three covalent bonds with atoms of nonmetals. Nitrogen is a group 5A element with the electron configuration of $1s^2 2s^2 2p^3$. Ammonia (NH_3) contains three single covalent bonds and one lone pair of electrons on the nitrogen atom. **Figure 9-3b** shows the Lewis structure

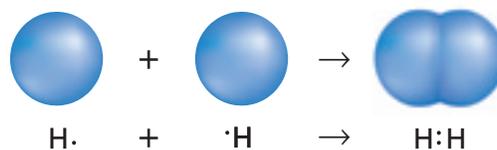


Figure 9-2

By sharing a pair of electrons, these hydrogen atoms have a full outer electron energy level and are stable. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

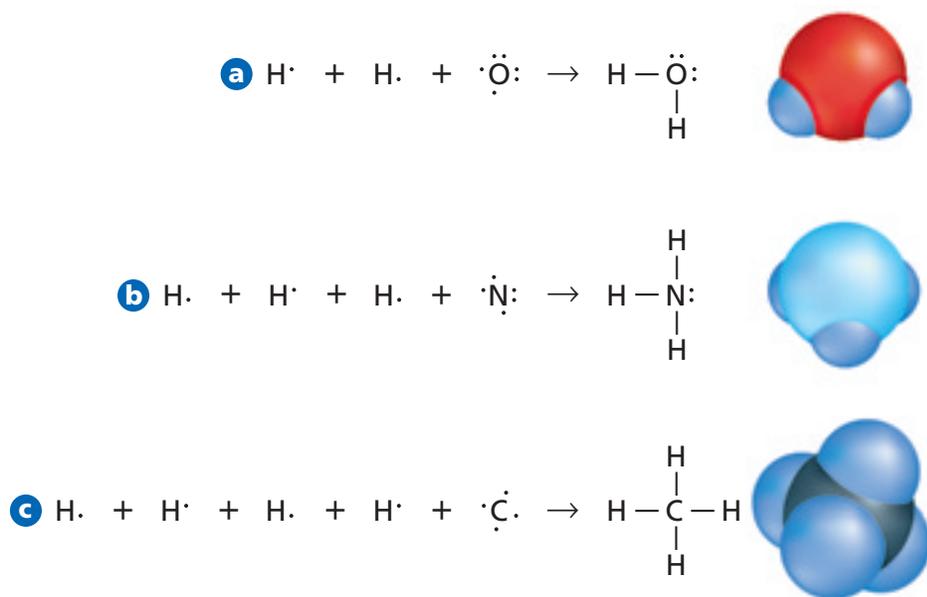


Figure 9-3

These chemical equations show how atoms share electrons to become stable. As can be seen by the Lewis structures (left side) for the molecules, after a reaction, all atoms in the molecules are stable according to the octet rule.

for an ammonia molecule. Nitrogen also forms similar compounds with group 7A elements, such as nitrogen trifluoride (NF₃), nitrogen trichloride (NCl₃), and nitrogen tribromide (NBr₃). Each of these group 7A atoms shares a pair of electrons with the nitrogen atom.

Group 4A elements will form four covalent bonds. A methane molecule (CH₄) is formed when one carbon atom bonds with four hydrogen atoms. Carbon, a group 4A element, has an electron configuration of 1s²2s²2p². With four valence electrons, it must obtain four more electrons for a noble gas configuration. Therefore, when carbon bonds with other atoms, it forms four bonds. Because hydrogen, a group 1A element, has one valence electron, four hydrogen atoms are necessary to provide the four electrons needed by the carbon atom. The Lewis structure for methane is shown in **Figure 9-3c**. Carbon also forms single covalent bonds with other nonmetals, including group 7A elements.

EXAMPLE PROBLEM 9-1

Lewis Structure for a Molecule

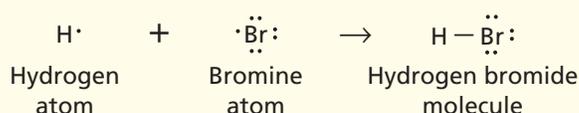
Hydrogen bromide (HBr) is used to manufacture several other bromides and has been medically used as a sedative. Draw the Lewis structure for this molecule.

1. Analyze the Problem

You are given that hydrogen and bromine form the molecule hydrogen bromide. Hydrogen, a group 1A element, has only one valence electron. Therefore, when hydrogen bonds with any nonmetal, it must share one pair of electrons. Bromine, a group 7A element, also needs one electron to complete its octet. Hydrogen and bromine bond with each other by one single covalent bond.

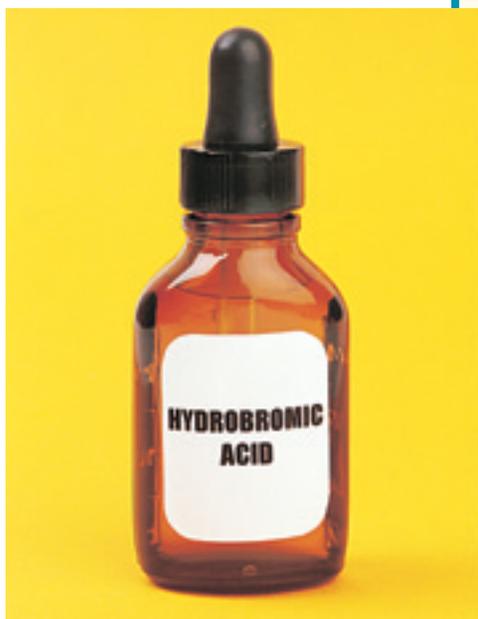
2. Solve for the Unknown

To draw the Lewis structure, first draw the electron-dot structure for each of the two atoms. Then show the sharing of the pairs of electrons by a single line.



3. Evaluate the Answer

Each atom in the molecule has achieved a noble gas configuration and thus is stable.



When hydrogen bromide is dissolved in water, hydrobromic acid forms. This acid must be kept in a dark bottle because it decomposes when exposed to light.

Practice!

For more practice with drawing Lewis structures, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Draw the Lewis structure for each of these molecules.

1. PH₃
2. H₂S
3. HCl
4. CCl₄
5. SiH₄

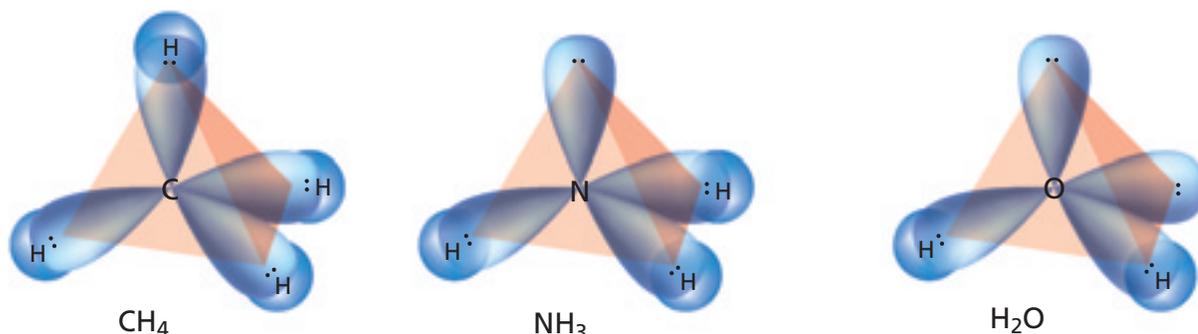


Figure 9-4

The sigma bonds form in these molecules as the atomic orbitals of hydrogen atoms overlap end to end with the orbitals of the central atoms.

The sigma bond Single covalent bonds also are called **sigma bonds**, symbolized by the Greek letter sigma (σ). A sigma bond occurs when the electron pair is shared in an area centered between the two atoms. When two atoms share electrons, the valence atomic orbital of one atom overlaps or merges with the valence atomic orbital of the other atom. A sigma bond results if the atomic orbitals overlap end to end, concentrating the electrons in a bonding orbital between the two atoms. A bonding orbital is a localized region where bonding electrons will most likely be found. **Figure 9-4** indicates the sigma bonds found in methane (CH_4), ammonia (NH_3), and water (H_2O). Sigma bonds can form from the overlap of an s orbital with another s orbital, an s orbital with a p orbital, or a p orbital with another p orbital. Which orbitals overlap in the sigma bonds shown in **Figure 9-4**?

Multiple Covalent Bonds

In many molecules, atoms attain a noble-gas configuration by sharing more than one pair of electrons between two atoms, forming a multiple covalent bond. Atoms of the elements carbon, nitrogen, oxygen, and sulfur most often form multiple bonds. How do you know when two atoms will form a multiple bond? The number of valence electrons of an element is associated with the number of shared electron pairs needed to complete the octet and gives a clue as to how many covalent bonds can form.

Double and triple covalent bonds are examples of multiple bonds. A double covalent bond occurs when two pairs of electrons are shared. The atoms in an oxygen molecule (O_2) share two electron pairs, forming a double bond. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration. If each oxygen atom shares two electrons, a total of two pairs of electrons is shared between the two atoms. A double covalent bond results. See **Figure 9-5a**.

A triple covalent bond is formed when three pairs of electrons are shared between two atoms. Nitrogen (N_2) shares three electron pairs, producing a triple bond. One nitrogen atom needs three additional electrons to attain a noble-gas configuration. **Figure 9-5b** shows the triple bond formed between two nitrogen atoms.

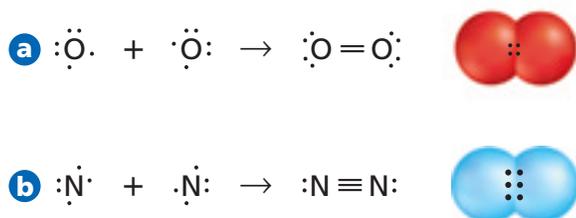
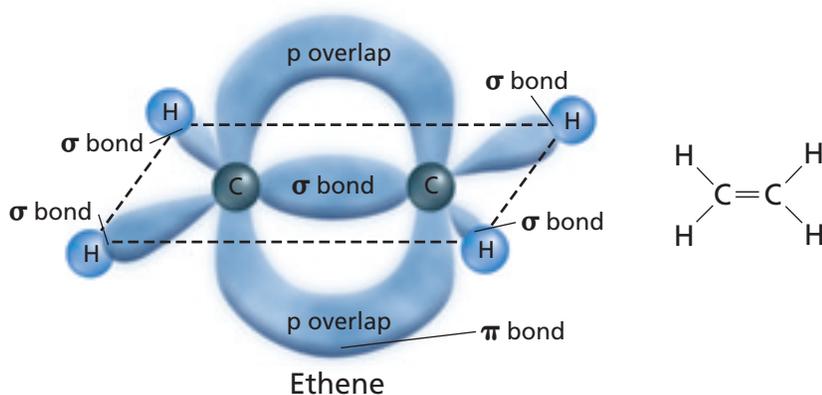


Figure 9-5

Multiple bonds form when more than one pair of electrons are shared. **a** The two pairs of electrons shared by oxygen atoms form a double bond. **b** Nitrogen shares three pairs of electrons, forming a triple bond.

Figure 9-6

Notice how the multiple bond between carbon atoms in ethene (C_2H_4) consists of one sigma bond and one pi bond. The carbon atoms are close enough that the side-by-side p orbitals overlap. The p-overlap that forms the pi bond results in a donut-shaped cloud around the sigma bond.



The pi bond A **pi bond**, denoted by the Greek symbol pi (π), is formed when parallel orbitals overlap to share electrons, as shown in **Figure 9-6**. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together. A multiple bond consists of one sigma bond and at least one pi bond. A double covalent bond has one sigma bond and one pi bond. A triple covalent bond consists of one sigma bond and two pi bonds. See **Figure 9-6**. A pi bond always accompanies a sigma bond when forming double and triple bonds.

Strength of Covalent Bonds

Remember that a covalent bond involves attractive and repulsive forces. In a molecule, nuclei and electrons attract each other, but nuclei repel other nuclei, and electrons repel other electrons. When this balance of forces is upset, the covalent bond can be broken. Some covalent bonds are broken more easily than others because they differ in strength. Several factors control the strength of covalent bonds.

The strength of a covalent bond depends on how much distance separates bonded nuclei. The distance between the two bonding nuclei at the position of maximum attraction is called *bond length*, which is determined by the size of the atoms and how many electron pairs are shared. See **Figure 9-7**. The bond length of the single bond in an F_2 molecule is 1.43×10^{-10} m. The bond length of the double bond in O_2 is 1.21×10^{-10} m. The bond length of the triple bond in N_2 is 1.10×10^{-10} m. Although the sizes of the atoms are not the same, not much difference exists in the size of these molecules. How does the number of pairs of electrons shared in F_2 , O_2 , and N_2 relate to the bond lengths in each of these molecules? As the number of shared electron pairs increases, bond length decreases. A triple bond, sharing three electron pairs, has a shorter bond length than a single bond where only two electrons are shared. The shorter the bond length, the stronger the bond. Thus, single bonds, such as those in F_2 , are weaker than double bonds, such as those in O_2 . Double bonds are weaker than triple bonds, such as those in N_2 .

An energy change accompanies the forming or breaking of a bond between atoms in a molecule. Energy is released when a bond forms. Energy must be added to break the bonds in a molecule. The amount of energy required to break a specific covalent bond is called *bond dissociation energy*. Breaking bonds always requires the addition of energy. Thus, bond dissociation energy is always a positive value. The bond dissociation energy of F_2 is 159 kJ/mol, of O_2 is 498 kJ/mol, and of N_2 is 945 kJ/mol. The sum of the bond dissociation energy values for all bonds in a compound is used to determine the

Figure 9-7

Bond length is the distance from the center of one nucleus to the center of the other nucleus of two bonded atoms.



amount of chemical potential energy available in a molecule of that compound.

Bond dissociation energy indicates the strength of a chemical bond because a direct relationship exists between bond energy and bond length. As two atoms are bonded closer together, greater amounts of bond energy are needed to separate them. Think back to the relative bond lengths of F_2 , O_2 , and N_2 . Based on bond length, which of these three molecules would you predict to have the greatest bond energy? The least? Do the bond dissociation energies supplied in the previous paragraph confirm your predictions?

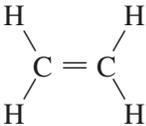
In chemical reactions, bonds in reactant molecules are broken and new bonds are formed as product molecules form. See **Figure 9-8**. The total energy change of the chemical reaction is determined from the energy of the bonds broken and formed. **Endothermic** reactions occur when a greater amount of energy is required to break the existing bonds in the reactants than is released when the new bonds form in the product molecules. **Exothermic** reactions occur when more energy is released forming new bonds than is required to break bonds in the initial reactants. You will learn more about energy and chemical processes in Chapter 16.



Figure 9-8

Energy is used to break C–C bonds in coal and O–O bonds of oxygen in air. Energy is released as heat and light due to the formation of CO_2 . Coal burning is an exothermic reaction.

Section 9.1 Assessment

- What is a covalent bond? How does it differ from an ionic bond?
- What is a single covalent bond? Why does it form?
- Why do multiple bonds form?
- What is the difference between a sigma bond and a pi bond?
- How is bond length related to bond dissociation energy?
- Thinking Critically** From the following structures, predict the relative bond energies needed to break all of the bonds present.
 - $H-C \equiv C-H$
 - 
- Making Predictions** Draw the electron-dot diagrams for the elements sulfur, carbon, bromine, oxygen, and hydrogen. Using Lewis structures, predict the number of covalent bonds formed when
 - one atom of sulfur bonds with two atoms of hydrogen.
 - one atom of carbon bonds with two atoms of sulfur.
 - two atoms of bromine bond with one atom of sulfur.
 - one atom of carbon bonds with four atoms of bromine
 - one atom of sulfur bonds with two atoms of oxygen.

Objectives

- **Identify** the names of binary molecular compounds from their formulas.
- **Name** acidic solutions.

Vocabulary

oxyacid

You know that many atoms covalently bond to form molecules that behave as a single unit. These units can be represented by chemical formulas and names that are used to identify them. When naming molecules, the system of rules is similar to the one you used to name ionic compounds.

Naming Binary Molecular Compounds

The anesthetic dinitrogen oxide (N_2O), commonly known as nitrous oxide, is a covalently bonded compound. Because it contains only two different elements, it is a binary molecular compound. Binary molecular compounds are composed of two different nonmetals and do not contain metals or ions. Although many of these compounds have common names, they also have scientific names that reveal their composition. Use the following simple rules to name binary molecular compounds.

1. The first element in the formula is always named first, using the entire element name.
2. The second element in the formula is named using the root of the element and adding the suffix *-ide*.
3. Prefixes are used to indicate the number of atoms of each type that are present in the compound. The most common prefixes are shown in Table 9-1.

Table 9-1

Prefixes in Covalent Compounds			
Number of atoms	Prefix	Number of atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

One exception to using these prefixes is that the first element in the formula never uses the prefix *mono-*. Also, to avoid awkward pronunciation, drop the final letter in the prefix when the element name begins with a vowel. For example, CO is carbon monoxide, not monocarbon monooxide.

EXAMPLE PROBLEM 9-2

Naming Binary Molecular Compounds

Name the compound P_2O_5 , which is used as a drying and dehydrating agent.

1. Analyze the Problem

You are given the formula for a compound. This formula reveals what elements are present and how many atoms of each element exist in a molecule. Because only two different elements are present and both

are nonmetals, the compound can be named using the rules for naming binary molecular compounds.

2. Solve for the Unknown

Name the first element present in the compound.

phosphorus

The second element is oxygen. The root of this name is *ox-*, so the second part of the name is *oxide*.

phosphorus oxide

From the formula P_2O_5 , two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From **Table 9-1**, *di-* is the prefix for two, and *penta-* is the prefix for five.

diphosphorus pentoxide.

The *a-* of *penta-* is not used because *oxide* begins with a vowel.

3. Evaluate the Answer

The name diphosphorus pentoxide shows that a molecule of the compound contains two phosphorus atoms and five oxygen atoms, which agrees with the chemical formula for the compound, P_2O_5 .



Diphosphorus pentoxide releases energy and produces fumes when it dissolves in water to form phosphoric acid.

PRACTICE PROBLEMS

Name the following binary covalent compounds.

13. CCl_4

16. SO_2

14. As_2O_3

17. NF_3

15. CO



For more practice with naming binary covalent compounds, go to **Supplemental Practice Problems** in Appendix A.

Common names of some molecular compounds How frequently have you drunk an icy, cold glass of dihydrogen monoxide? Quite frequently, but you probably didn't call it that. You called it by its more common name, which is water. Remember from Chapter 8 that many ionic compounds have common names in addition to their scientific ones. Baking soda is sodium hydrogen carbonate and common table salt is sodium chloride. Many covalent compounds also have both common and scientific names.

Many binary molecular compounds were discovered and given common names long before the modern naming system was developed. **Table 9-2** lists some of these molecules, their common names, and the binary molecular compound names.

Table 9-2

Formulas and Names of Some Covalent Compounds		
Formula	Common name	Molecular compound name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
N_2O	nitrous oxide (laughing gas)	dinitrogen monoxide
NO	nitric oxide	nitrogen monoxide



Organic Chemist

Would you like to create molecular models, using state-of-the-art computers? That is one role of organic chemists, who study molecules based on carbon.

Organic chemists and chemists in related fields use computers to analyze the structure and motion of molecules. Sometimes they build a visual model of one molecule. Other times they examine entire molecular systems. This information is essential in finding new drugs and in mapping human genes. Molecular modeling is one of the fastest-growing fields in science.

Naming Acids

Water solutions of some molecules are acidic and are named as acids. Acids are important compounds with specific properties that will be discussed at length in Chapter 19. If the compound produces hydrogen ions (H^+) in solution, it is an acid. For example, HCl produces H^+ in solution and is an acid. Two common types of acids exist—binary acids and oxyacids.

Naming binary acids A binary acid contains hydrogen and one other element. When naming a binary acid, use the prefix *hydro-* to name the hydrogen part of the compound. The rest of the name consists of a form of the root of the second element plus the suffix *-ic*, followed by the word *acid*. For example, HBr in a water solution is called hydrobromic acid.

Although the term *binary* indicates exactly two elements, a few acids that contain more than two elements are named according to the rules for naming binary acids. If no oxygen is present in the formula for the acidic compound, the acid is named in the same way as a binary acid, except that the root of the second part of the name is the root of the polyatomic ion that the acid contains. For example, HCN , which is composed of hydrogen and the cyanide ion, is called hydrocyanic acid.

Naming oxyacids Another set of rules is used to name an acid that contains an oxyanion. An oxyanion is a polyatomic ion that contains oxygen. Any acid that contains hydrogen and an oxyanion is referred to as an **oxyacid**.

Because the name of an oxyacid depends on the oxyanion present in the acid, you must first identify the anion present. The name of an oxyacid consists of a form of the root of the anion, a suffix, and the word *acid*. If the anion suffix is *-ate*, it is replaced with the suffix *-ic*. When the anion ends in *-ite*, the suffix is replaced with *-ous*. Consider the oxyacid HNO_3 . Its oxyanion is nitrate (NO_3^-). Following this rule, HNO_3 is named nitric acid. The anion of HNO_2 is the nitrite ion (NO_2^-). HNO_2 is nitrous acid. Notice that the hydrogen in an oxyacid is not part of the name.

It's important to remember that these hydrogen-containing compounds are named as acids only when they are in water solution. For example, at room temperature and pressure HCl is hydrogen chloride, a gas. When HCl is dissolved in water, it is hydrochloric acid.

PRACTICE PROBLEMS

Name the following acids. Assume each compound is dissolved in water.

- | | |
|---------------------|-----------------------------|
| 18. HI | 21. H_2SO_4 |
| 19. HClO_3 | 22. H_2S |
| 20. HClO_2 | |

Writing Formulas from Names

The name of a molecular compound reveals its composition and is important in communicating the nature of the compound. **Figure 9-9** can help you determine the name of a molecular covalent compound.

The name of any binary molecule allows you to write the correct formula with ease. Subscripts are determined from the prefixes used in the name because the name indicates the exact number of each atom present in the molecule. The formula for an acid can be derived from the name as well.



For more practice with naming acids, go to **Supplemental Practice Problems** in Appendix A.

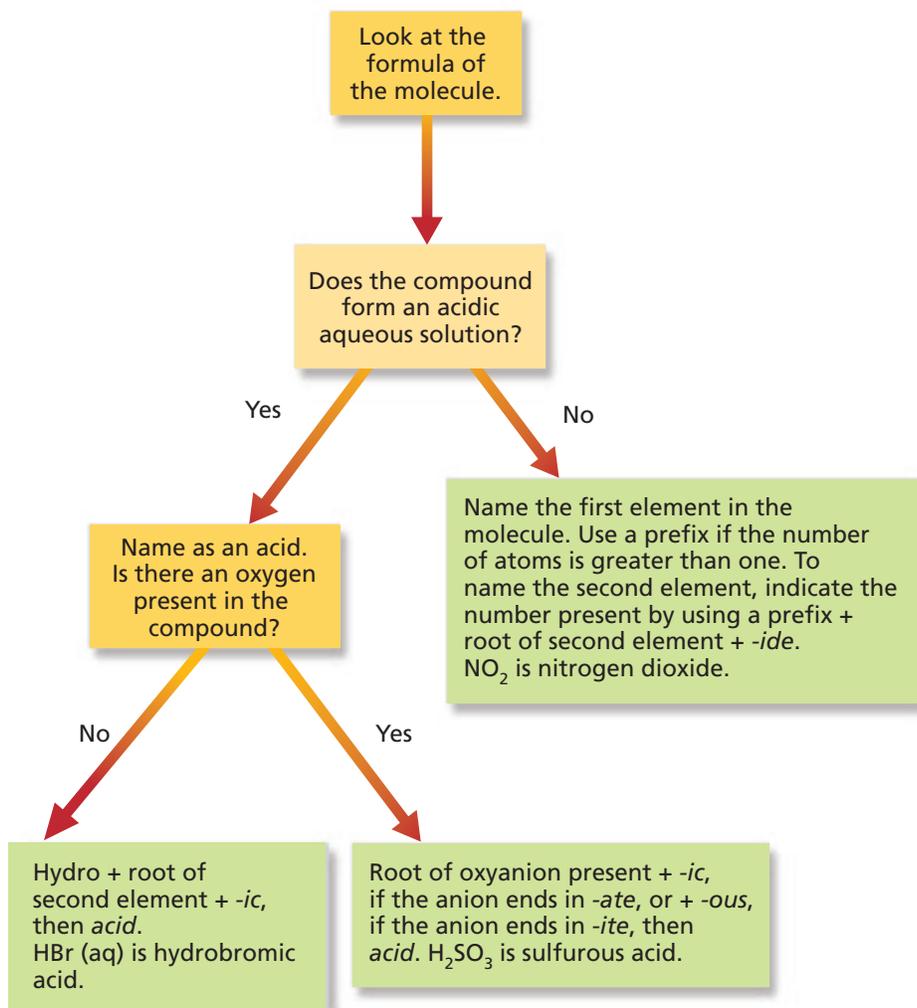


Figure 9-9

This flow chart summarizes how to name molecular compounds when their formulas are known.

Section 9.2 Assessment

23. What is a binary molecular compound?
24. Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N₂O₄.
25. Compare and contrast naming binary acids and naming other binary covalent molecules.
26. What is the difference between a binary acid and an oxyacid?
27. Write the molecular formula for each of the following compounds.
 - a. disulfur trioxide
 - b. iodic acid
 - c. dinitrogen monoxide
 - d. hydrofluoric acid
 - e. phosphorus pentachloride

28. **Thinking Critically** Write the molecular formula for each listed compound.
 - a. dinitrogen trioxide and nitrogen monoxide
 - b. hydrochloric acid and chloric acid
 - c. sulfuric acid and sulfurous acid
29. **Making and Using Tables** Complete the following table.

Formulas and Names of Covalent Compounds	
Formula	Name
PCl ₅	
	hydrobromic acid
H ₃ PO ₄	
	oxygen difluoride
SO ₂	

Objectives

- **List** five basic steps used in drawing Lewis structures.
- **Explain** why resonance occurs, and **identify** resonance structures.
- **Explain** three exceptions to the octet rule, and **identify** molecules in which these exceptions occur.

Vocabulary

structural formula
resonance
coordinate covalent bond

You can now identify atoms that bond covalently and name the molecular compounds formed through covalent bonding. In order to predict the arrangement of atoms in each molecule, a model, or representation is used. Several different models can be used, as shown in **Figure 9-10**. Note that in the ball-and-stick and space-filling molecular models, atoms of each specific element are represented by spheres of a representative color, as shown in **Table C-1** in Appendix C. These colors are used for identification of the atoms if the chemical symbol of the element is not present.

Structural Formulas

One of the most useful molecular models is the **structural formula**, which uses letter symbols and bonds to show relative positions of atoms. The structural formula can be predicted for many molecules by drawing the Lewis structure. You have already seen some simple examples of Lewis structures, but more involved structures are needed to help you determine the shapes of molecules.

Although it is fairly easy to draw Lewis structures for most compounds formed by nonmetals, it is a good idea to follow a regular procedure. The following steps should be used to determine Lewis structures.

1. Predict the location of certain atoms.
 - a. Hydrogen is always a terminal, or end, atom. Because it can share only one pair of electrons, hydrogen can be connected to only one other atom.
 - b. The atom with the least attraction for shared electrons in the molecule is the central atom. This element usually is the one closer to the left on the periodic table. The central atom is located in the center of the molecule, and all other atoms become terminal atoms.
2. Find the total number of electrons available for bonding. This total is the number of valence electrons in the atoms in the molecule.
3. Determine the number of bonding pairs by dividing the number of electrons available for bonding by two.
4. Place one bonding pair (single bond) between the central atom and each of the terminal atoms.

PH_3
Molecular formula

$\begin{array}{c} \text{H} - \ddot{\text{P}} - \text{H} \\ | \\ \text{H} \end{array}$
Lewis structure


Space-filling
molecular model

$\begin{array}{c} \text{H} - \text{P} - \text{H} \\ | \\ \text{H} \end{array}$
Structural formula

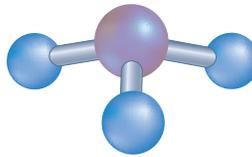

Ball-and-stick
molecular model

Figure 9-10

All of these models can be used to show the relative locations of atoms and electrons in the phosphorus trihydride (phosphine) molecule.

- Subtract the number of pairs you used in step 4 from the number of bonding pairs you determined in step 3. The remaining electron pairs include lone pairs as well as pairs used in double and triple bonds. Place lone pairs around each terminal atom bonded to the central atom to satisfy the octet rule. Any remaining pairs are assigned to the central atom.
- If the central atom is not surrounded by four electron pairs, it does not have an octet. You must convert one or two of the lone pairs on the terminal atoms to a double bond or a triple bond between the terminal atom and the central atom. These pairs are still associated with the terminal atom as well as with the central atom. Remember that, in general, carbon, nitrogen, oxygen, and sulfur can form double or triple bonds with the same element or with another element.

EXAMPLE PROBLEM 9-3

Lewis Structure: Covalent Compound with Single Bonds

Ammonia is a raw material for the manufacture of many materials, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

1. Analyze the Problem

You are given that the ammonia molecule consists of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2. Solve for the Unknown

Find the total number of valence electrons.

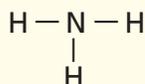
$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}}$$

$$= 8 \text{ valence electrons}$$

Determine the total number of bonding pairs.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

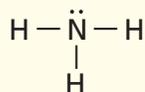
Draw single bonds from each H to the N.



Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

$$4 \text{ pairs total} - 3 \text{ pairs used} = 1 \text{ pair available}$$

One lone pair remains to be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs. Place the remaining lone pair on the central atom, N.



3. Evaluate the Answer

Each hydrogen atom shares one pair of electrons, as required, and the central nitrogen atom shares three pairs of electrons and has one lone pair, providing a stable octet.



Ammonia is a common ingredient in many cleaning products.

EXAMPLE PROBLEM 9-4

Lewis Structure: Covalent Compound with Multiple Bonds

Carbon dioxide is a product of all cellular respiration. Draw the Lewis structure for carbon dioxide (CO₂).

1. Analyze the Problem

You are given that the carbon dioxide molecule consists of one carbon atom and two oxygen atoms. Because carbon has less attraction for shared electrons, carbon is the central atom, and the two oxygen atoms are terminal.

2. Solve for the Unknown

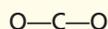
Find the total number of valence electrons.

$$1 \text{ C atom} \times \frac{4 \text{ valence electrons}}{\text{C atom}} + 2 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{\text{O atom}}$$
$$= 16 \text{ valence electrons}$$

Determine the total number of bonding pairs.

$$\frac{16 \text{ electrons}}{2 \text{ electrons/pair}} = 8 \text{ pairs}$$

Draw single bonds from each O to the central C atom.



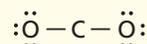
Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

$$8 \text{ pairs total} - 2 \text{ pairs used} = 6 \text{ pairs available}$$

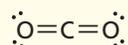
Add three lone pairs to each terminal oxygen. Subtract the lone pairs used from the pairs available.

$$6 \text{ pairs available} - 6 \text{ lone pairs used} = 0$$

No electron pairs remain available for the carbon atom.



Carbon does not have an octet, so use a lone pair from each oxygen atom to form a double bond with the carbon atom.



3. Evaluate the Answer

Both carbon and oxygen now have an octet, which satisfies the octet rule.



As a waste product, carbon dioxide is one of the gases exhaled.

The polyatomic ions you learned about in Chapter 8 are related to covalent compounds. Although the unit acts as an ion, the atoms within the ion itself are covalently bonded. The structures of these ions can also be represented by Lewis structures.

The procedure for drawing Lewis structures for polyatomic ions is similar to drawing them for covalent compounds. The main difference is in finding the total number of electrons available for bonding. Compared to the number of valence electrons present in the atoms that make up the ion, more electrons are present if the ion is negatively charged and fewer are present if the ion is positive. To find the total number of electrons available for bonding, first find the number available in the atoms present in the ion. Then, subtract the ion charge if the ion is positive, and add the ion charge if the ion is negative.

EXAMPLE PROBLEM 9-5

Lewis Structure: Polyatomic Ion

Draw the correct Lewis structure for the polyatomic ion phosphate (PO_4^{3-}).

1. Analyze the Problem

You are given that the phosphate ion consists of one phosphorus atom and four oxygen atoms and has a charge of -3 . Phosphorus has less attraction for shared electrons, so it is the central atom, and the four oxygen atoms are terminal.

2. Solve for the Unknown

Find the total number of valence electrons.

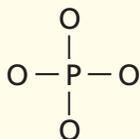
$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{\text{P atom}} + 4 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{\text{O atom}}$$

$$+ 3 \text{ electrons from the negative charge} = 32 \text{ valence electrons}$$

Determine the total number of bonding pairs.

$$\frac{32 \text{ electrons}}{2 \text{ electrons/pair}} = 16 \text{ pairs}$$

Draw single bonds from each O to the central P.



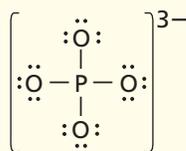
Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

$$16 \text{ pairs total} - 4 \text{ pairs used} = 12 \text{ pairs available}$$

Add three lone pairs to each terminal oxygen. Subtract the lone pairs used from the pairs available.

$$12 \text{ pairs available} - 12 \text{ lone pairs used} = 0$$

No electron pairs remain available for the phosphorus atom, resulting in the following Lewis structure for the phosphate ion.



3. Evaluate the Answer

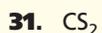
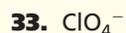
Phosphorus and all four oxygen atoms have an octet, and the group of atoms has a net charge of -3 .



One phosphate compound, $\text{MgH}_4(\text{PO}_4)_2$, is used to fireproof wood. This compound contains the polyatomic ion phosphate.

PRACTICE PROBLEMS

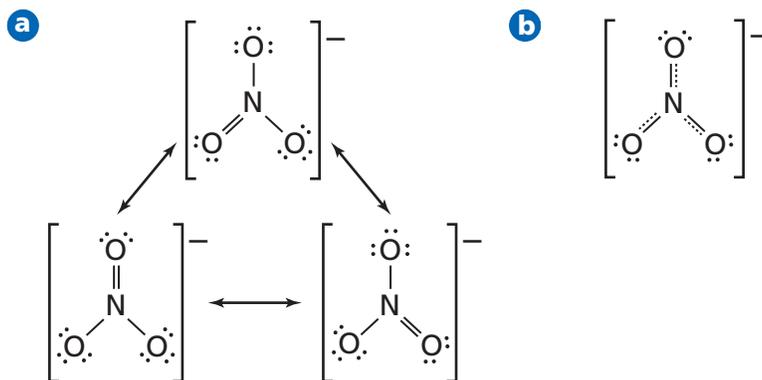
Draw a Lewis structure for each of the following:



For more practice with drawing Lewis structures, go to **Supplemental Practice Problems** in Appendix A.

Figure 9-11

The nitrate ion (NO_3^-) exhibits resonance. **a** These resonance structures differ only in the location of the double bond. The locations of the nitrogen and oxygen atoms stay the same. **b** The actual nitrate ion is like an average of the three resonance structures in **a**. The dotted line indicates possible locations of the double bond.



Resonance Structures

Using the same sequence of atoms, it is possible to have more than one correct Lewis structure when a molecule or polyatomic ion has both a double bond and a single bond. Consider the polyatomic ion nitrate (NO_3^-) shown in **Figure 9-11a**. Three equivalent structures can be used to represent the nitrate ion. **Resonance** is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion. The two or more correct Lewis structures that represent a single molecule or ion are often referred to as *resonance structures*. Resonance structures differ only in the position of the electron pairs, never the atom positions. The location of the lone pairs and bonding pairs differs in resonance structures. The molecule O_3 and the polyatomic ions NO_3^- , NO_2^- , SO_3^{2-} and CO_3^{2-} commonly form resonance structures.

It is important to note that each actual molecule or ion that undergoes resonance behaves as if it has only one structure. See **Figure 9-11b**. Experimentally measured bond lengths show that the bonds are identical to each other. They are shorter than single bonds but longer than double bonds. The actual bond length is an average of the bonds in the resonance structures.



For more practice with drawing Lewis resonance structures, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

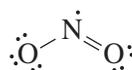
Draw the Lewis resonance structures for the following.

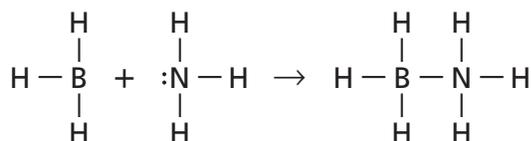
- | | |
|-------------------|---------------------|
| 35. SO_3 | 37. O_3 |
| 36. SO_2 | 38. NO_2^- |

Exceptions to the Octet Rule

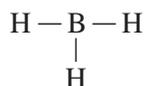
The Lewis structure is focused on the attainment of an octet by all atoms when they bond with other elements. Some molecules and ions, however, do not obey the octet rule. Three reasons exist for these exceptions.

First, a small group of molecules has an odd number of valence electrons and cannot form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs. ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.





Second, some compounds form with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 3A nonmetal, forms three covalent bonds with other nonmetallic atoms.



A total of six electrons is shared by the boron atom, which is two less than the number needed for an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom. When one atom donates a pair of electrons to be shared with an atom or ion that needs two electrons to become stable, a **coordinate covalent bond** forms, as in **Figure 9-12**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an *expanded octet*. An expanded octet can be explained by considering the d orbital that occurs in the energy levels of elements in period three or higher. An example of an expanded octet is the bond formation in the molecule PCl_5 . Five bonds are formed with ten electrons shared in one s orbital, three p orbitals, and one d orbital. Another example is the molecule SF_6 , which has six bonds sharing 12 electrons in an s orbital, three p orbitals, and two d orbitals. When you draw the Lewis structure for these compounds, extra lone pairs are added to the central atom or more than four bonding atoms are present in the molecule.

Figure 9-12

In this reaction between boron trihydride (BH_3) and ammonia (NH_3), the nitrogen atom donates both electrons that are shared by boron and ammonia, forming a coordinate covalent bond.

EXAMPLE PROBLEM 9-6

Lewis Structure: Exception to the Octet Rule

Xenon is a noble gas that will form a few compounds with nonmetals that strongly attract electrons. Draw the correct Lewis structure for xenon tetrafluoride (XeF_4).

1. Analyze the Problem

You are given that a molecule of xenon tetrafluoride consists of one xenon atom and four fluorine atoms. Xenon has less attraction for electrons, so it is the central atom.

2. Solve for the Unknown

Find the total number of valence electrons.

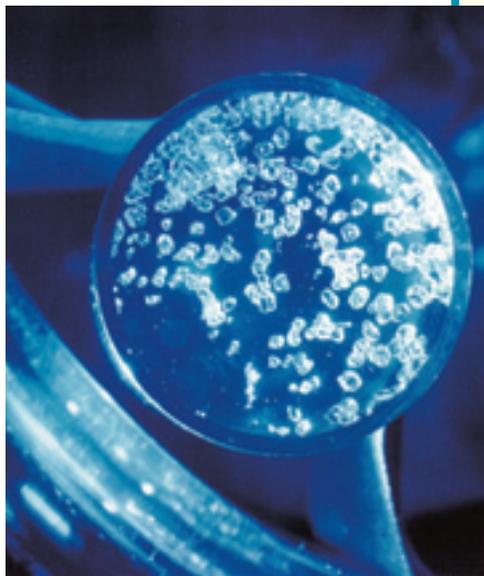
$$1 \text{ Xe atom} \times \frac{8 \text{ valence electrons}}{\text{Xe atom}} + 4 \text{ F atoms} \times \frac{7 \text{ valence electrons}}{\text{F atom}}$$

$$= 36 \text{ valence electrons}$$

Determine the number of bonding pairs.

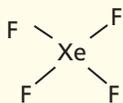
$$\frac{36 \text{ electrons}}{2 \text{ electrons/pair}} = 18 \text{ pairs}$$

Continued on next page



Xenon compounds, such as the XeF_4 shown here, are highly toxic because they are so reactive.

Use four bonding pairs to bond the four F atoms to the central Xe atom.



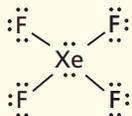
Determine the number of remaining pairs.

$$18 \text{ pairs available} - 4 \text{ pairs used} = 14 \text{ pairs available}$$

Add three pairs to each F atom to obtain an octet. Determine how many pairs remain.

$$14 \text{ pairs} - 4 \text{ F atoms} \times \frac{3 \text{ pairs}}{\text{F atom}} = 2 \text{ pairs unused}$$

Place the two remaining pairs on the central Xe atom.



3. Evaluate the Answer

This structure gives xenon 12 total electrons, an expanded octet, for a total of six bond positions.

Practice!

For more practice with drawing Lewis structures, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Draw the correct Lewis structures for the following molecules, which contain expanded octets.

39. SF_6
40. PCl_5
41. ClF_3

You now know how to draw Lewis structures for molecules and polyatomic ions. You can use them to determine the number of bonding pairs between atoms and the number of lone pairs present. Next, you will learn to describe molecular structure and predict the angles in a molecule, both of which determine the three-dimensional molecular shape.

Section 9.3 Assessment

42. What is the role of the central atom when drawing the Lewis structure for a molecule?
43. What is resonance?
44. List three exceptions to the octet rule.
45. What is a coordinate covalent bond?
46. What is an expanded octet?
47. **Thinking Critically** Draw the resonance structures for the N_2O molecule.
48. **Formulating Models** Draw the Lewis structures for the following molecules and ions.
 - a. CN^-
 - b. SiF_4
 - c. HCO_3^-
 - d. AsF_6^-

The shape of a molecule determines many of its physical and chemical properties. Molecular shape, in turn, is determined by the overlap of orbitals that share electrons. Theories have been developed to explain the overlap of bonding orbitals and are used to predict the shape of the molecule.

VSEPR Model

Many chemical reactions, especially those in living things, depend on the ability of two compounds to contact each other. The shape of the molecule determines whether or not molecules can get close enough to react.

Once a Lewis structure is drawn, you can determine the molecular geometry, or shape, of the molecule. The model used to determine the molecular shape is referred to as the **Valence Shell Electron Pair Repulsion model**, or **VSEPR model**. This model is based on an arrangement that minimizes the repulsion of shared and unshared pairs of electrons around the central atom.

The repulsions among electron pairs in a molecule result in atoms existing at fixed angles to each other. The angle formed by any two terminal atoms and the central atom is a *bond angle*. Bond angles predicted by VSEPR are supported by experimental evidence. Shared electron pairs repel one another. Lone pairs of electrons also are important in determining the shape of the molecule. Because lone pairs are not shared between two nuclei, they occupy a slightly larger orbital than shared electrons. Shared bonding orbitals are pushed together slightly by lone pairs.

To make sense of the VSEPR model, consider balloons of similar size tied together, as shown in **Figure 9-13**. Each balloon represents an orbital and, therefore, the repulsive force that keeps other electrons from entering this space. When each set of balloons is connected at a central point that represents the central atom, the balloons naturally form a shape that minimizes interactions between the balloons. You will build additional examples of VSEPR models in the **miniLAB** on page 261.

Examine **Table 9-3**, which indicates some common shapes of molecules. First, consider molecules that contain no lone pairs of electrons. When only two pairs of electrons are shared off the central atom (BeCl_2), these bonding electrons will seek the maximum separation, which is a bond angle of 180° , and the molecular shape is linear. Three bonding electron pairs also will separate as much as possible (AlCl_3), forming a trigonal planar shape with 120° bond angles. When the central atom has four pairs of bonding electrons (CH_4), the shape is tetrahedral with bond angles of 109.5° .

It is important to remember that a lone pair also occupies a position in space. Recall phosphine (PH_3) shown in **Figure 9-10**, which has three single covalent bonds and one lone pair. You might expect the structure to be tetrahedral because of the four bonding positions around the central atom. However, the lone pair takes up a greater amount of space than the shared pairs. The geometry of PH_3 is trigonal pyramidal, and the bond angles are 107.3° .

Now consider a water molecule (H_2O), which has two single covalent bonds and two lone pairs according to its Lewis structure. Although a water molecule has four electron pairs off the central atom, it is not tetrahedral because the two lone pairs occupy more space than do the paired electrons. The water molecule has a bent shape with a bond angle of 104.5° .

Objectives

- **Discuss** the VSEPR bonding theory.
- **Predict** the shape of and the bond angles in a molecule.
- **Define** hybridization.

Vocabulary

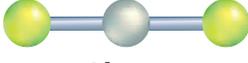
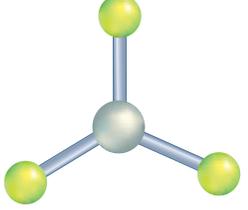
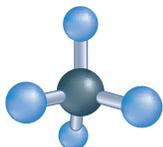
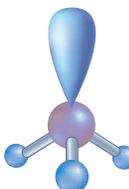
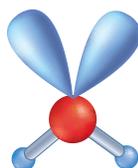
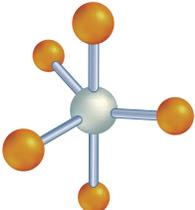
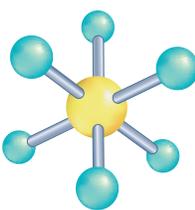
VSEPR model
hybridization



Figure 9-13

Electrons in a molecule are located as far apart as they can be, just as these balloons are arranged.

Table 9-3

Molecular Shapes						
Example	Total pairs	Shared pairs	Lone pairs	*Molecular shape	Bond Angle	Hybrid Orbitals
BeCl ₂	2	2	0	 Linear	180°	sp
AlCl ₃	3	3	0	 Trigonal planar	120°	sp ²
CH ₄	4	4	0	 Tetrahedral	109.5°	sp ³
PH ₃	4	3	1	 Trigonal pyramidal	107.3°	sp ³
H ₂ O	4	2	2	 Bent	104.5°	sp ³
NbBr ₅	5	5	0	 Trigonal bipyramidal	(vertical to horizontal) 90° / 120° (horizontal to horizontal)	sp ³ d
SF ₆	6	6	0	 Octahedral	90°	sp ³ d ²

*Balls represent atoms; sticks represent bonds; and lobes represent lone pairs of electrons.

Hybridization

A hybrid results from combining two of the same type of object, and it has characteristics of both. Atomic orbitals undergo hybridization during bonding. Let's consider the bonding involved in the methane molecule (CH_4). The carbon atom has four valence electrons with the electron configuration of $[\text{He}]2s^22p^2$. You might expect the two unpaired p electrons to bond with other atoms and the 2s electrons to remain a lone pair. However, carbon atoms undergo **hybridization**, a process in which atomic orbitals are mixed to form new, identical *hybrid orbitals*. Each hybrid orbital contains one electron that it can share with another atom, as shown in **Figure 9-14**. Carbon is the most common element that undergoes hybridization. Because the four hybrid orbitals form from one s and three p orbitals, this hybrid orbital is called an sp^3 orbital.

Look again at **Table 9-3** on the previous page. Notice that the number of atomic orbitals mixed to form the hybrid orbital equals the total number of pairs of electrons. In addition, the number of hybrid orbitals formed equals the number of atomic orbitals mixed. For example, AlCl_3 has a total of three pairs of electrons and VSEPR predicts a trigonal planar molecular shape. To have this shape, one s and two p orbitals on the central atom Al must mix to form three identical sp^2 hybrid orbitals.

Lone pairs also occupy hybrid orbitals. Compare the hybrid orbitals of BeCl_2 and H_2O in **Table 9-3**. Both compounds contain three atoms. Why is H_2O sp^3 ? There are two lone pairs on the central atom (oxygen) in H_2O . Therefore, there must be four hybrid orbitals—two for bonding and two for the lone pairs.

Recall from Section 9.1 that multiple covalent bonds consist of one sigma bond and one or more pi bonds. Only the two electrons in the sigma bond occupy hybrid orbitals such as sp and sp^2 . The remaining unhybridized p orbitals overlap to form pi bonds.

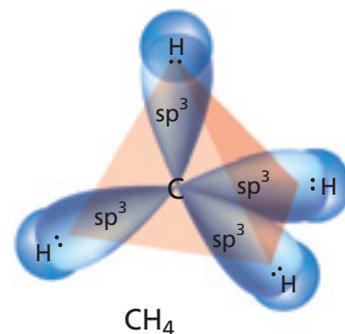


Figure 9-14

One s and three p orbitals hybridize to form four sp^3 orbitals. According to VSEPR, a tetrahedral shape minimizes repulsion between the orbitals.

miniLAB

Building VSEPR Models

Formulating Models The VSEPR model states that pairs of valence electrons on a central atom repel each other and are arranged so that the repulsions are as small as possible. In this **miniLAB**, you will use marshmallows and gumdrops to build models of substances, showing examples of the VSEPR model.

Materials regular-sized marshmallows (3); mini-sized marshmallows (9); small gumdrops (3); toothpicks, cut in half

Procedure

1. Draw Lewis structures for methane (CH_4), ammonia (NH_3), and water (H_2O). Notice the location of each shared and unshared pair of electrons.
2. Using your Lewis structures, build a VSEPR

model for each molecule. Use a mini-marshmallow to represent both the hydrogen atom and the region of space containing the pair of electrons shared by hydrogen and the central atom. Use a regular-sized marshmallow to represent the space occupied by an unshared pair of electrons and a small gumdrop to represent a central atom. Use small pieces of toothpicks to attach the marshmallows and gumdrops to each other. Sketch each of your models.

Analysis

1. How did drawing a Lewis structure help you to determine the geometry of each of your substances?
2. Why was a mini-marshmallow used to show a shared pair of electrons and a regular marshmallow an unshared pair?
3. How can the VSEPR model help to predict the bond angles for these substances?

EXAMPLE PROBLEM 9-7

Finding the Shape of a Molecule

Phosphorus trihydride is produced when organic materials rot, and it smells like rotten fish. What is the shape of a molecule of phosphorus trihydride? Determine the bond angle, and identify the type of hybrid.

1. Analyze the Problem

You are given a molecule of phosphorus trihydride that contains one phosphorus atom bonded to three hydrogen atoms. The hydrogen atoms are terminal atoms, and phosphorus is the central atom.

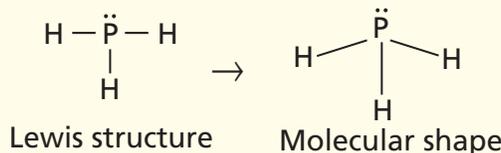
2. Solve for the Unknown

Find the total number of valence electrons and the number of electron pairs.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{\text{P atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{\text{H atom}} \\ = 8 \text{ valence electrons}$$

$$\frac{8 \text{ available electrons}}{2 \text{ electrons/pair}} = 4 \text{ available pairs}$$

Draw the Lewis structure, using one pair of electrons to bond each H to the central P and assigning the lone pair to the phosphorus atom.



The molecular geometry is trigonal pyramidal with a bond angle of 107° . With four bonding positions, the molecule is an sp^3 hybrid.

4. Evaluate the Answer

Each atom has a stable electron configuration, and all electron pairs are accounted for.



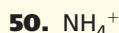
Phosphine, or phosphorus trihydride, is formed when phosphorus-containing organic material, such as this potato, rots.



For more practice with molecular geometry and hybridization, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Determine the molecular geometry, bond angle, and type of hybridization for the following.



Section 9.4 Assessment

- What is the VSEPR model?
- What are the bond angles in a molecule with a tetrahedral shape?
- What is hybridization?
- What are the hybrid orbitals in a molecule with a tetrahedral shape?
- Thinking Critically** Compare the molecules PF_3 and PF_5 . What is the molecular shape of each of

the two molecules? What type of hybrid orbital is in each molecule? Why is the shape different?

- Making and Using Tables** Make a table that contains the Lewis structure, molecular shape, bond angle, and type of hybrid for the following molecules: CS_2 , CH_2O , H_2Se , CCl_2F_2 , and NCl_3 .

You now know that the type of bond that forms when two elements react depends on which elements are involved. What makes one type of bond form when carbon burns and another type form when iron corrodes? The answer lies in how much attraction each type of atom has for electrons.

Electronegativity Difference and Bond Character

Electron affinity is a measure of the tendency of an atom to accept an electron. Excluding noble gases, electron affinity increases as the atomic number increases within a given period and decreases with an increase in atomic number within a group. The scale of electronegativities allows a chemist to evaluate the electron affinity of specific atoms when they are incorporated into a compound. Recall from Chapter 6 that electronegativity indicates the relative ability of an atom to attract electrons in a chemical bond.

Look at the electronegativity values shown in **Figure 9-15**. Can you observe any trends? Note that fluorine has the highest electronegativity value (3.98), while francium has the lowest (0.7). The same trends appear with electronegativities that can be observed with electron affinities. Because noble gases do not generally form compounds, individual electronegativity values for helium, neon, and argon are not given. However, larger noble gases like xenon do occasionally bond with highly electronegative atoms such as fluorine.

Objectives

- **Describe** how electronegativity is used to determine bond type.
- **Compare** and **contrast** polar and nonpolar covalent bonds and polar and nonpolar molecules.
- **Describe** the characteristics of compounds that are covalently bonded.

Vocabulary

polar covalent

Figure 9-15

Electronegativity values are not measured quantities. They are values assigned by Linus Pauling comparing the abilities of atoms to attract shared electrons with the ability of fluorine to do so.

1 H 2.20																	5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	
3 Li 0.98	4 Be 1.57																	13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
11 Na 0.93	12 Mg 1.31	19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96				
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66						
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2						
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1	Elements not included on this table have no measured electronegativity or form relatively few bonds.																			
Lanthanide series			58 Ce 1.12	59 Pr 1.13	60 Nd 1.14	61 Pm —	62 Sm 1.17	63 Eu —	64 Gd 1.20	65 Tb —	66 Dy 1.22	67 Ho 1.23	68 Er 1.24	69 Tm 1.25	70 Yb —	71 Lu 1.0						
Actinide series			90 Th 1.3	91 Pa 1.5	92 U 1.7	93 Np 1.3	94 Pu 1.3															

History

CONNECTION

In the 1940s, while experimenting with a device called a magnetron that generates microwaves, a scientist noticed that a candy bar in his pocket was melting. This accidental discovery led to the creation of the microwave oven, which cooks foods based on the polarity of the molecules involved. Early microwave ovens did not sell well because they were about the size and weight of a refrigerator and cost several thousand dollars. Eventually, smaller, lower-cost versions appeared for the home kitchen.



The character and type of a chemical bond can be predicted using the electronegativity difference of the elements that are bonded. For identical atoms, which have an electronegativity difference of zero, the electrons in the bond are equally shared between the two atoms and the bond is considered nonpolar covalent, which is a pure covalent bond. Chemical bonds between atoms of different elements are never completely ionic or covalent, and the character of the bond depends on how strongly the bonded atoms attract electrons. A covalent bond formed between atoms of different elements does not have equal sharing of the electron pair because there is a difference in electronegativity. Unequal sharing results in a **polar covalent** bond. Large differences in electronegativity indicate that an electron was transferred from one atom to another, resulting in bonding that is primarily ionic.

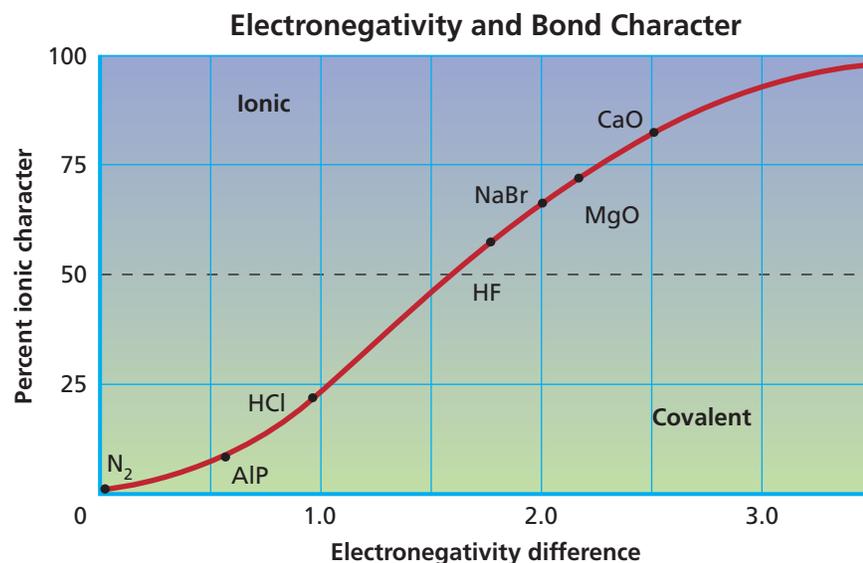
Bonding often is not clearly ionic or covalent. As the difference in electronegativity increases, the bond becomes more ionic in character. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. Generally, ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together. **Figure 9-16** summarizes the range of chemical bonding between two atoms. What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? Where would LiBr be plotted on the graph?

Polar Covalent Bonds

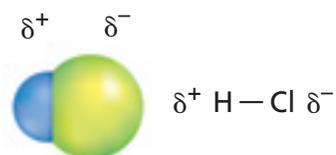
Why are some bonds polar covalent? Sharing is not always equal. Consider a tug-of-war when the two sides are not of equal strength. Although both sides share the rope, one side pulls more of the rope toward its side. Polar covalent bonds form because not all atoms that share electrons attract them equally. When a polar bond forms, the shared pair of electrons is pulled toward one of the atoms. The electrons spend more time around that atom than they do around the other atom. Partial charges occur at the ends of the bond. Using the symbols δ^- , partially negative, and δ^+ , partially positive, next to a model of the molecule indicates the polarity of a polar covalent bond. See **Figure 9-17**. The more electronegative atom is located at the partially negative end, while the less electronegative atom is found at the partially positive end. The resulting polar bond often is referred to as a dipole (two poles).

Figure 9-16

This graph shows how the percent ionic character of a bond depends on the difference in electronegativity of the atoms that form it. Above 50% ionic character, bonds are mostly ionic. What is the percent ionic character of a pure covalent bond?



Electronegativity Cl = 3.16
 Electronegativity H = 2.20
 Difference = 0.96



Molecular polarity Molecules are either nonpolar or polar, depending on the location and nature of the covalent bonds they contain. One way to distinguish polar from nonpolar molecules is that nonpolar molecules are not attracted by an electric field. Polar molecules tend to align with an electric field because polar molecules have a greater electron density on one side of the molecule. A polar molecule has a partial negative charge on one side, while the other side of the molecule has a partial positive charge. The molecule is a dipole because of the two partial charges.

Polar molecule or not? Let's compare water (H₂O) and carbon tetrachloride (CCl₄) molecules to see why some molecules are polar and some are not. Both molecules contain polar covalent bonds. Using **Figure 9-15** on page 263, you can see that the electronegativity difference between one hydrogen atom and one oxygen atom is 1.24. The electronegativity difference between one chlorine atom and one carbon atom is 0.61. Although these electronegativity differences vary quite a bit, both the H—O and the C—Cl bonds are considered to be polar covalent.



Both molecules contain more than one polar covalent bond. But water molecules are polar and carbon tetrachloride molecules are not. Examine the geometry of the molecules to see the reason for this difference.

The shape of H₂O determined by VSEPR is bent because there are two lone pairs of electrons on the central oxygen atom. Because the polar H—O bonds are not symmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.



The shape of CCl₄ is tetrahedral. This molecule is symmetric. The electrical charge measured at any distance from its center is identical to the charge measured at the same distance on the opposite side. The average center of the negative charge is located on the carbon atom. The positive center also is located on the carbon atom. Because the partial charges are balanced, CCl₄ is a nonpolar molecule. Note that symmetric molecules are usually nonpolar, and molecules that are asymmetric are polar as long as the bond type is polar.

Is the molecule of ammonia (NH₃) shown in **Figure 9-18** polar or not? It contains a central nitrogen atom and three terminal hydrogen atoms. It has a trigonal pyramidal shape because of the lone pair of electrons present on the nitrogen atom. Using **Figure 9-15**, you can find that the electronegativity difference of hydrogen and nitrogen is 0.84. So, each N—H bond is polar covalent.

Figure 9-17

In a molecule containing hydrogen and chlorine, chlorine has the higher electronegativity. Therefore, the shared pair of electrons is with the chlorine atom more often than it is with the hydrogen atom. The symbols indicating partial charge of each end of the molecule reflect this unequal sharing of electrons.

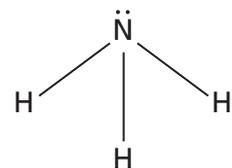


Figure 9-18

Ammonia, represented by the Lewis structure, is used in the manufacture of certain synthetic fibers.

The charge distribution is unequal because the molecule is not symmetric. Thus, the molecule is polar. The **CHEMLAB** and the **How it Works** feature at the end of this chapter are based on the polarity of molecules.

Solubility of polar molecules The ability of a substance to dissolve in another substance is known as the physical property solubility. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds usually are soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances. See **Figure 9-19**. Solubility will be covered in detail in Chapter 15.

Properties of Covalent Compounds

You've probably noticed how similar table salt, an ionic solid, and table sugar, a covalent solid, are in appearance. But if you heat salt on the stove, it won't melt, even if the temperature is high. Sugar, on the other hand, melts at a relatively low temperature. Does type of bonding affect properties?

Differences in properties are a result of differences in attractive forces. In a covalent compound, the covalent bond between atoms in molecules is quite strong, but the attraction between individual molecules is relatively weak. The weak forces of attraction between individual molecules are known as *intermolecular forces*, or *van der Waals forces*. Intermolecular forces, which are discussed at length in Chapter 13, vary in strength but are weaker than the bonds that join atoms in a molecule or ions in an ionic compound.

There are different types of intermolecular forces. For nonpolar substances, the attraction between the molecules is weak and is called a dispersion force, or induced dipole. The effect of dispersion forces is investigated in the **problem-solving LAB** on the next page. The force between polar molecules is stronger and is called a dipole-dipole force. This force is the attraction of one end of the dipole to the oppositely charged end of another dipole. The more polar the molecule, the stronger the dipole-dipole force. The third force, a hydrogen bond, is an especially strong intermolecular force that is formed between the hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.

Many physical properties of covalent molecular solids are due to intermolecular forces. The melting and boiling points of molecular substances are relatively low compared with those of ionic substances. That's why salt doesn't melt when you heat it but sugar does. Many molecular substances exist as gases or vaporize readily at room temperature. Oxygen (O_2), carbon dioxide (CO_2), and hydrogen sulfide (H_2S) are examples of covalent gases. Hardness is also due to the intermolecular forces between individual molecules, so covalent molecules form relatively soft solids. Paraffin is a common example of a covalent solid.

In the solid state, molecules line up in a pattern forming a crystal lattice similar to that of an ionic solid, but with less attraction between particles. The structure of the crystal lattice depends on the shape of the molecule and the type of intermolecular force. Most information about molecules, including properties, molecular shape, bond length, and bond angle, has been determined by studying molecular solids.



Figure 9-19

Oil, most petroleum products, and other symmetric covalent molecules are nonpolar, whereas water and other asymmetric molecules are usually polar. When polar and nonpolar substances are mixed, they separate into two layers, as seen when oil floats on water.



For more practice with determining polarity, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Decide whether each of the following molecules is polar or nonpolar.



How do dispersion forces determine the boiling point of a substance?

Making and Using Graphs The strength of the dispersion force between nonpolar molecules determines the temperature at which these substances boil. Measuring the boiling point allows one to get a fairly good estimate of the relative strength of this force among different molecules.

Analysis

The table lists the molecular masses and boiling points for eight compounds composed of carbon and hydrogen atoms. Construct a graph showing the relationship between these quantities. Infer from the graph how dispersion forces change with increasing molecular mass. Then, draw the molecular structure for each molecule.

Thinking Critically

- How do the relationship demonstrated by the graph and the geometry of the different hydrocarbons help to explain the effect the dispersion force has on molecules of different size?
- How would you use your graph to predict the properties of similar but larger compounds, such as decane ($C_{10}H_{22}$)?

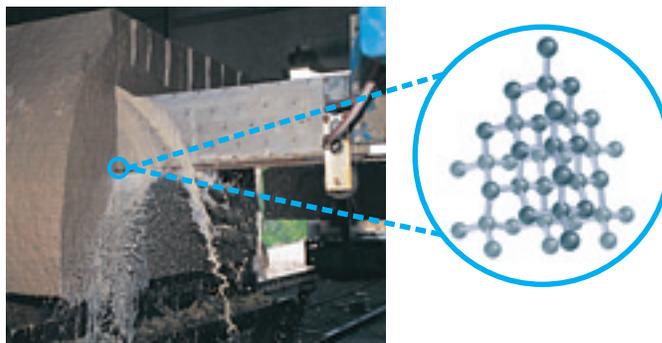
Name	Formula	M (amu)	T_{bp} ($^{\circ}C$)
Methane	CH_4	16	-161.48
Ethane	C_2H_6	30	-88.6
Propane	C_3H_8	44	-42.1
Butane	C_4H_{10}	58	-0.5
Pentane	C_5H_{12}	72	36.06
Hexane	C_6H_{14}	86	68.73
Heptane	C_7H_{16}	100	98.5
Octane	C_8H_{18}	114	125.7

Covalent Network Solids

A number of solids are composed only of atoms interconnected by a network of covalent bonds. These solids are often called covalent network solids. Quartz is a network solid, as is diamond. See **Figure 9-20**. In contrast to molecular solids, network solids are typically brittle, nonconductors of heat or electricity, and extremely hard. In a diamond, four other carbon atoms surround each carbon atom. This tetrahedral arrangement forms a strongly bonded crystal system that is extremely hard and has a very high melting point.

Figure 9-20

In diamond, each carbon atom is bonded to four other carbon atoms. Network solids, such as diamond, are often used in cutting tools because of their hardness.



Section 9.5 Assessment

- Define electronegativity.
- How is electronegativity difference used in determining the type of bond that occurs between two atoms?
- Describe a polar covalent bond.
- What is a polar molecule?
- List three properties of a covalent compound.
- Thinking Critically** Predict the type of bond that will form between the following atoms.
 - H and S
 - C and H
 - Na and S
- Drawing Conclusions** Draw the Lewis structure for the SF_4 and SF_6 molecules and determine if each molecule is polar or nonpolar.

Chromatography

Paper chromatography is a common way to separate various components of a mixture. The components of the mixture separate because different substances are selectively absorbed by paper due to differences in polarity. In this field or laboratory investigation, you will separate the various pigments found in leaves. You also will calculate the ratio called R_f for each of them. The ratio R_f compares the distance traveled by a substance, D_s , to the distance traveled by the solvent, D_f . The ratio is written as $R_f = D_s / D_f$.

Problem

How can a mixture be separated based on the polarity of substances in the mixture?

Objectives

- **Separate** pigments found in leaves.
- **Determine** the R_f value for each of the pigments in the leaves.

Materials

chromatography paper (3 pieces)	aluminum foil
2-L plastic soft drink bottle	acetone
pencils (2)	fresh leaf samples from three different species of deciduous trees or outdoor plants.
metric ruler	
tape	
scissors or metal snips	

Safety Precautions



- Acetone is a flammable liquid. Do not use near flames or sparks.
- Do not allow acetone to contact skin.
- Perform procedure in an area with proper ventilation.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your observations.
3. What is polarity? How is polarity related to how chromatography works?
4. Predict what will happen when a mixture of leaf pigments is placed on a piece of paper and a solvent is allowed to move through the paper, moving the pigment with it.
5. Suppose that the pigments in two samples contain red pigment and that the red pigment in sample A is more soluble in acetone than the red pigment in sample B. Form a hypothesis regarding which red pigment has the higher R_f value. Explain your answer.

Paper Chromatography				
Leaf sample	D_f (cm)	Colors	D_s (cm)	R_f
1				
2				
3				

Procedure

1. For each leaf sample, crush the leaves and soak them in a small amount of acetone to make a concentrated solution of the pigments in the leaves.
2. Cut the top off a 2-liter bottle. Cut small notches, as shown in the figure, so that a pencil can rest across the top of the bottle.
3. Cut three pieces of 3-cm wide chromatography paper to a length of about 18 cm. Label the top of each paper with a number. Assign a number to each pigment sample used. Draw pencil lines about 5 cm from the bottom of the end of each paper.
4. On the pencil line of paper 1, put a dot from the first sample. Make sure the dot is concentrated but not wide. Do the same for the other samples on their respective papers. Tape the papers to the pencil, as shown in the figure.
5. Put enough acetone in the 2-liter bottle so that when the papers are put in the bottle, the solvent touches only the bottom 1 cm of each paper, as shown in the figure. **CAUTION:** Do not allow acetone to come in contact with skin. Use in area with proper ventilation.
6. Carefully lower the chromatography papers into the acetone and put the pencil into the notches at the top of the bottle. Cover the top with aluminum foil. Allow the chromatograms to develop for about 35-40 minutes.
7. When the chromatograms are finished, remove them from the bottle. Mark the highest point reached by the solvent. Then, allow the papers to air dry.



Cleanup and Disposal

1. Dispose of the acetone as directed by your teacher.
2. Throw the chromatography paper in the trash can.

Analyze and Conclude

1. **Observing and Inferring** Record in the data table the colors that are found in each of the chromatograms. Space is allowed for three colors, but some samples may contain fewer or more than three colors.
2. **Measuring** For each strip, measure the distance the solvent traveled from the pencil line (D_f). For each color, measure from the top of the original marker dot to the farthest point the color traveled (D_s). Record these values in your data table.
3. **Interpreting Data** Calculate the R_f values for each of the pigments in each chromatogram and record them in the data table.
4. **Comparing and Contrasting** Describe the differences between the pigments in each of the samples.
5. **Applying Concepts** Will a polar solvent, such as water, cause a difference in how the pigments are separated? Explain your answer.
6. **Error Analysis** What could be done to improve the measurements you used to calculate R_f ?

Real-World Chemistry

1. Use your results to explain what happens to leaves in autumn.
2. How might chromatography be used to analyze the composition of the dye in a marker?

How It Works

Microwave Oven

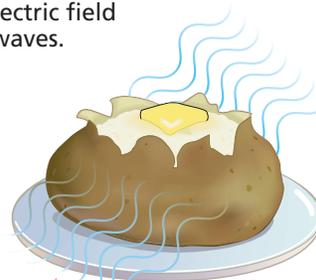
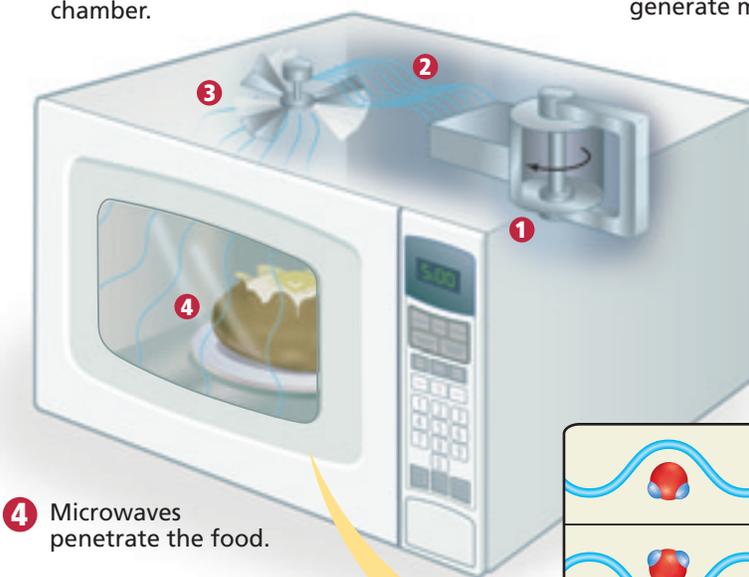
Today, about 90 percent of home kitchens in the United States have a microwave oven. The polar nature and small size of water molecules allow a microwave oven to cook food without a conventional source of heat. Many large molecules, such as those in sugars and fats, are nonpolar and are not easily heated by microwaves because the molecules cannot easily and rapidly realign themselves. Also, substances that are in the form of crystals usually do not heat well because the bonds that form the crystal structure prevent the molecules from easily realigning themselves.



3 Stirrer distributes microwaves throughout the cooking chamber.

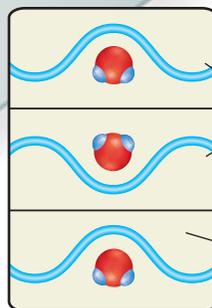
2 Microwaves travel through a hollow tube called a waveguide.

1 Electrons moving in both a magnetic and electric field generate microwaves.



7 Increased molecular motion generates heat — raising the temperature of the food.

4 Microwaves penetrate the food.



6 Water molecules flip back and forth as the microwave field oscillates.

5 Polar water molecules align themselves with the microwave field.

Thinking Critically

1. Predicting How well would you expect a microwave oven to heat liquid carbon dioxide (CO_2)? CO_2 is a small molecule and its structure is $\text{O}=\text{C}=\text{O}$. Explain your reasoning.

2. Predicting Would microwaves have the same heating effect on ice as they have on liquid water? What about table salt (NaCl)? Justify your predictions.

Summary

9.1 The Covalent Bond

- A covalent bond is formed when atoms share one or more pairs of electrons.
- Molecules, formed when atoms share electrons, are more stable than their constituent atoms.
- Sharing a single pair of electrons results in a single covalent bond. Two atoms sharing more than one pair of electrons results in a multiple bond.
- A double covalent bond results when two pairs of electrons are shared between atoms. Sharing three pairs of electrons results in a triple covalent bond.
- When an electron pair is shared by the direct overlap of bonding orbitals, a sigma bond results. The overlap of parallel orbitals forms a pi bond. Single bonds are sigma bonds. Multiple bonds involve both sigma and pi bonds.
- Bond length depends on the sizes of the bonded atoms and the number of electron pairs they share. Bond dissociation energy is the energy needed to break a covalent bond. Bond length and bond dissociation energy are directly related.

9.2 Naming Molecules

- Names of covalent molecular compounds include prefixes that tell the number of each atom present.
- Molecules that produce hydrogen ions in solution are acids and are named accordingly.

9.3 Molecular Structures

- The Lewis structure is used to show the distribution of shared and lone pairs of electrons in a molecule.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur when an odd number of valence electrons exists between the bonding atoms, not enough electrons are available for an octet, or more than eight electrons are shared.

- Coordinate covalent bonding occurs when one atom of the bonding pair supplies both shared electrons.

9.4 Molecular Shape

- The valence shell electron pair repulsion, or VSEPR, model can be used to predict the three-dimensional shape of a molecule. Electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.
- Two orbitals form two sp hybrid orbitals, and the molecule is linear. Three orbitals, forming three sp² hybrid orbitals, form a molecule that is trigonal planar. Four orbitals, forming four sp³ hybrid orbitals, form a molecule that is tetrahedral.

9.5 Electronegativity and Polarity

- Electronegativity is the tendency of an atom to attract electrons and is related to electron affinity. The electronegativity difference between two bonded atoms is used to determine the type of bond that most likely occurs.
- Polar bonds occur when electrons are not shared equally, resulting in an unequal distribution of charge and the formation of a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Weak intermolecular forces, also called van der Waals forces, hold molecules together in the liquid and solid phases. These weak attractive forces determine properties. Molecular solids tend to be soft and have low melting and boiling points.
- Covalent network solids result when each atom is covalently bonded to many other atoms in the solid. These solids are hard and have high melting points.

Vocabulary

- coordinate covalent bond (p. 257)
- covalent bond (p. 242)
- endothermic (p. 247)
- exothermic (p. 247)
- hybridization (p. 261)
- Lewis structure (p. 243)
- molecule (p. 242)
- oxyacid (p. 250)
- pi bond (p. 246)
- polar covalent (p. 264)
- resonance (p. 256)
- sigma bond (p. 245)
- structural formula (p. 252)
- VSEPR model (p. 259)

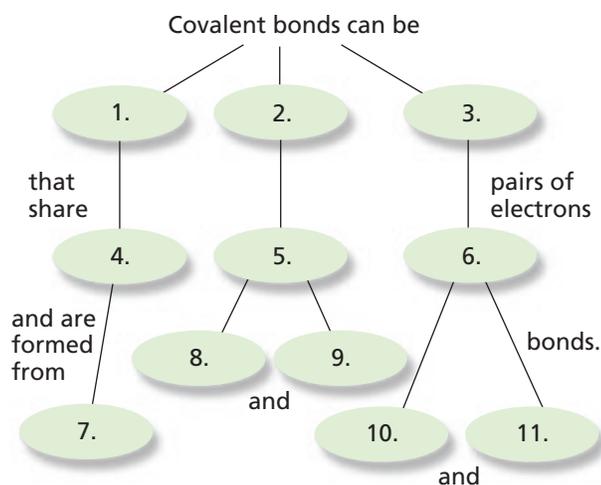


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 9 Assessment.

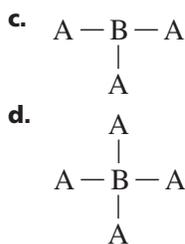
Concept Mapping

- 71.** Complete the concept map using the following terms: double bonds, one, pi, sigma, single bonds, three, triple bonds, two. Each term can be used more than once.



Mastering Concepts

- 72.** What is the octet rule, and how is it used in covalent bonding? (9.1)
- 73.** Describe the formation of a covalent bond. (9.1)
- 74.** Describe the bonding in molecules. (9.1)
- 75.** Describe the forces, both attractive and repulsive, that occur as two atoms come closer together. (9.1)
- 76.** How could you predict the presence of a sigma or pi bond in a molecule? (9.1)
- 77.** Explain how molecular compounds are named. (9.2)
- 78.** When is a molecular compound named as an acid? (9.2)
- 79.** What must be known in order to draw the Lewis structure for a molecule? (9.3)
- 80.** On what is the VSEPR model based? (9.4)
- 81.** What is the molecular shape of each of the following molecules? Estimate the bond angle for each assuming no lone pair. (9.4)
- A—B
 - A—B—A

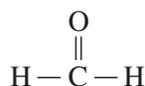


- 82.** What is the maximum number of hybrid orbitals a carbon atom can form? (9.4)
- 83.** Explain the theory of hybridization and determine the number of hybrid orbitals present in the molecule PCl_5 . (9.4)
- 84.** Describe the trends in electronegativity in the periodic table. (9.5)
- 85.** Explain the difference between nonpolar molecules and polar molecules. (9.5)
- 86.** Compare the location of bonding electrons in a polar covalent bond with those in a nonpolar covalent bond. Explain your answer. (9.5)
- 87.** What is the difference between a covalent molecular solid and a covalent network solid? Do their physical properties differ? Explain your answer. (9.5)

Mastering Problems

Covalent Bonds (9.1)

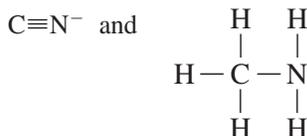
- 88.** Give the number of valence electrons in N, As, Br, and Se. Predict the number of covalent bonds needed for each of these elements to satisfy the octet rule.
- 89.** Locate the sigma and pi bonds in the following molecule.



- 90.** Locate the sigma and pi bonds in the following molecule.
- $$\text{H} - \text{C} \equiv \text{C} - \text{H}$$

Bond Length (9.1)

- 91.** Consider the molecules CO , CO_2 , and CH_2O . Which C—O bond is shorter? In which molecule is the C—O bond stronger?
- 92.** Consider the carbon-nitrogen bonds in the following:



Which bond is shorter? Which is stronger?

93. Rank each of the molecules below in order of the shortest to the longest sulfur-oxygen bond length.

- a. SO_2 b. SO_3^{2-} c. SO_4^{2-}

Naming Covalent Compounds (9.2)

94. Name each of the following solutions as an acid.

- a. HClO_2 c. H_2Se
b. H_3PO_4 d. HClO_3

95. Name each of the following molecules.

- a. NF_3 c. SO_3
b. NO d. SiF_4

96. Name each of the following molecules.

- a. SeO_2 c. N_2F_4
b. SeO_3 d. S_4N_4

Writing Formulas (9.2)

97. Write the formula for each of the following.

- a. sulfur difluoride
b. silicon tetrachloride
c. carbon tetrafluoride
d. sulfurous acid

98. Write the formula for each of the following.

- a. silicon dioxide
b. bromous acid
c. chlorine trifluoride
d. hydrobromic acid

Lewis Structures (9.3)

99. Draw the Lewis structure for each of these molecules or ions.

- a. H_2S c. SO_2
b. BF_4^- d. SeCl_2

100. Draw the Lewis structure for each of these molecules or ions.

- a. SeF_2 d. POCl_3
b. ClO_2^- e. GeF_4
c. PO_3^{3-}

101. Which of the following elements are capable of forming molecules in which an atom has an expanded octet? Explain your answer.

- a. B d. O
b. C e. Se
c. P

102. Draw three resonance structures for the polyatomic ion CO_3^{2-} .

103. Draw two resonance structures for the polyatomic ion CHO_2^- .

104. Draw the Lewis structure for each of the following molecules that have central atoms that do not obey the octet rule.

- a. PCl_5 c. ClF_5
b. BF_3 d. BeH_2

Molecular Shape (9.4)

105. Predict the molecular shape and bond angle, and identify the hybrid orbitals for each of the following. Drawing the Lewis structure may help you.

- a. SCl_2
b. NH_2Cl
c. HOF
d. BF_3

106. For each of the following, predict the molecular shape.

- a. COS b. CF_2Cl_2

107. Identify the expected hybrid on the central atom for each of the following. Drawing the Lewis structure may help you.

- a. XeF_4 c. KrF_2
b. TeF_4 d. OF_2

Electronegativity and Polarity (9.5)

108. For each pair, indicate the more polar bond by circling the negative end of its dipole.

- a. $\text{C}-\text{S}$, $\text{C}-\text{O}$
b. $\text{C}-\text{F}$, $\text{C}-\text{N}$
c. $\text{P}-\text{H}$, $\text{P}-\text{Cl}$

109. For each of the bonds listed, tell which atom is more negatively charged.

- a. $\text{C}-\text{H}$ c. $\text{C}-\text{S}$
b. $\text{C}-\text{N}$ d. $\text{C}-\text{O}$

110. Predict which of the following bonds is the most polar.

- a. $\text{C}-\text{O}$ c. $\text{C}-\text{Cl}$
b. $\text{Si}-\text{O}$ d. $\text{C}-\text{Br}$

111. Rank the following bonds according to increasing polarity.

- a. $\text{C}-\text{H}$ d. $\text{O}-\text{H}$
b. $\text{N}-\text{H}$ e. $\text{Cl}-\text{H}$
c. $\text{Si}-\text{H}$

112. Consider the following and determine if they are polar. Explain your answers.

- a. H_3O^+ c. H_2S
b. PCl_5 d. CF_4

113. Why is the CF_4 molecule nonpolar even though it contains polar bonds?

- 114.** Use Lewis structures to predict the molecular polarities for sulfur difluoride, sulfur tetrafluoride, and sulfur hexafluoride.

Mixed Review

Sharpen your problem solving skills by answering the following.

- 115.** Consider the following molecules and determine which of the molecules are polar. Explain your answer.
- CH_3Cl
 - ClF
 - NCl_3
 - BF_3
 - CS_2
- 116.** Arrange the following bonds in order of least to greatest polar character.
- $\text{C}-\text{O}$
 - $\text{Si}-\text{O}$
 - $\text{Ge}-\text{O}$
 - $\text{C}-\text{Cl}$
 - $\text{C}-\text{Br}$
- 117.** Draw the Lewis structure for ClF_3 and identify the hybrid orbitals.
- 118.** Use the Lewis structure for SF_4 to predict the molecular shape and identify the hybrid orbitals.
- 119.** Write the formula for each of these molecules.
- chlorine monoxide
 - arsenic acid
 - phosphorus pentachloride
 - hydrosulfuric acid
- 120.** Name each of the following molecules.
- PCl_3
 - Cl_2O_7
 - P_4O_6
 - NO

Thinking Critically

- 121. Concept Mapping** Design a concept map that will link both the VSEPR model and the hybridization theory to molecular shape.
- 122. Making and Using Tables** Complete the table using Chapters 8 and 9.

Table 9-4

Properties and Bonding			
Solid	Bond description	Characteristic of solid	Example
Ionic			
Covalent molecular			
Metallic			
Covalent network			

- 123. Drawing Conclusions** Consider each of the following characteristics and determine whether the molecule is polar or nonpolar.
- a solid at room temperature
 - a gas at room temperature
 - attracted to an electric current

Writing in Chemistry

- 124.** Research chromatography and write a paper discussing how it is used to separate mixtures.
- 125.** Research laundry detergents. Write a paper to explain why they are used to clean oil and grease out of fabrics.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 126.** The mass of the same liquid is given in the table below for the following volumes. Graph the volume on the x -axis and the mass on the y -axis. Calculate the slope of the graph. What information will the slope give you? (Chapter 2)

Table 9-5

Volume vs. Mass	
Volume	Mass
4.1 mL	9.36 g
6.0 mL	14.04 g
8.0 mL	18.72 g
10.0 mL	23.40 g

- 127.** Which group 3A element is expected to exhibit properties that are significantly different from the remaining family members? (Chapter 7)
- 128.** Write the correct chemical formula or name the following compounds. (Chapter 8)
- NaI
 - calcium carbonate
 - $\text{Fe}(\text{NO}_3)_3$
 - $\text{Sr}(\text{OH})_2$
 - potassium chlorate
 - copper(II) sulfate
 - CoCl_2
 - ammonium phosphate
 - silver acetate
 - $\text{Mg}(\text{BrO}_3)_2$

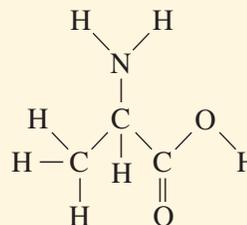
Use the questions and the test-taking tip to prepare for your standardized test.

- The common name of SiI_4 is tetraiodosilane. What is its molecular compound name?
 - silane tetraiodide
 - silane tetraiodine
 - silicon iodide
 - silicon tetraiodide
- Which of the following compounds contains at least one pi bond?
 - CO_2
 - CHCl_3
 - AsI_3
 - BeF_2
- The Lewis structure for silicon disulfide is _____.
 - $\text{Si} : \ddot{\text{S}} :$
 - $\ddot{\text{S}} :: \text{Si} :: \ddot{\text{S}} :$
 - $\ddot{\text{S}} : \text{Si} : \ddot{\text{S}} :$
 - $:\text{S} :: \text{Si} :: \text{S} :$
- The central selenium atom in selenium hexafluoride forms an expanded octet. How many electron pairs surround the central Se atom?
 - 4
 - 5
 - 6
 - 7
- Chloroform (CHCl_3) was one of the first anesthetics used in medicine. The chloroform molecule contains 26 valence electrons in total. How many of these valence electrons take part in covalent bonds?
 - 26
 - 13
 - 8
 - 4
- Which is the strongest type of intermolecular bond?
 - ionic bond
 - dipole-dipole force
 - dispersion force
 - hydrogen bond
- All of the following compounds have bent molecular shapes EXCEPT _____.
 - BeH_2
 - H_2S
 - H_2O
 - SeH_2
- Which of the following compounds is NOT polar?
 - H_2S
 - CCl_4
 - SiH_3Cl
 - AsH_3

Interpreting Tables Use the table to answer the following questions.

Bond Dissociation Energies at 298 K			
Bond	kJ/mol	Bond	kJ/mol
Cl—Cl	242	$\text{N}\equiv\text{N}$	945
C—C	345	O—H	467
C—H	416	C—O	358
C—N	305	C=O	745
H—I	299	O=O	498
H—N	391		

- Which of the following diatomic gases has the shortest bond between its two atoms?
 - HI
 - O_2
 - Cl_2
 - N_2
- Approximately how much energy will it take to break all of the bonds present in the molecule below?



- 5011 kJ/mol
- 3024 kJ/mol
- 4318 kJ/mol
- 4621 kJ/mol



TEST-TAKING TIP

Wear A Watch If you are taking a timed test, you should make sure that you pace yourself and do not spend too much time on any one question, but you shouldn't spend time staring at the clock. When each section of the test begins, set your watch for noon. This will make it very easy for you to figure out how many minutes have passed. After all, it is much easier to know that you started at 12:00 (according to your watch) and you'll be done at 12:30 than it is to figure out that you started at 10:42 and that time will run out at 11:12.

Chemical Reactions

What You'll Learn

- ▶ You will write chemical equations to describe chemical reactions.
- ▶ You will classify and identify chemical reactions.
- ▶ You will write ionic equations for reactions that occur in aqueous solutions.

Why It's Important

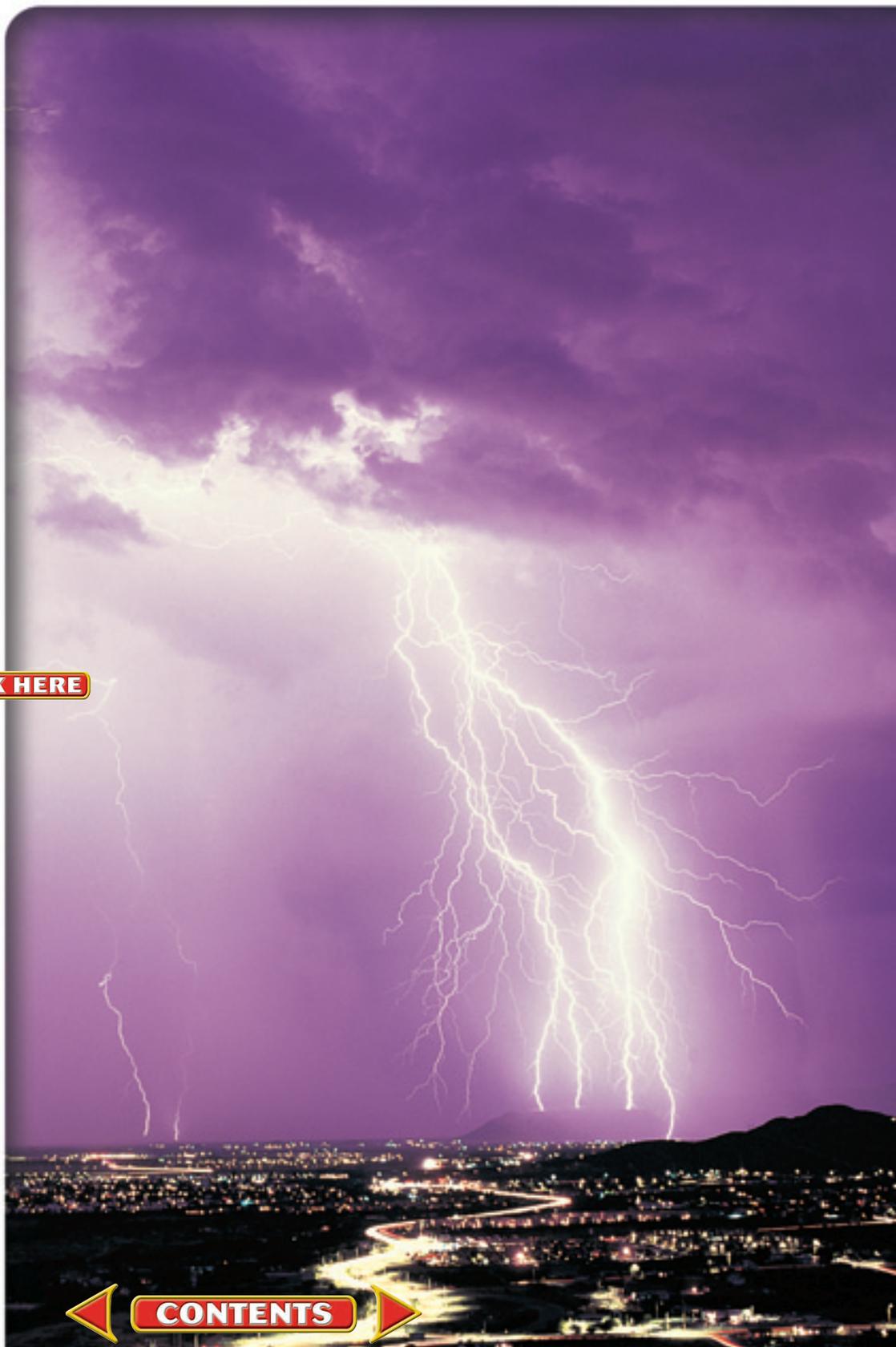
Chemical reactions affect you every second of every day. For example, life-sustaining chemical reactions occur continuously in your body. Other chemical reactions occur in less likely situations, such as in a thunderstorm.



CLICK HERE

Visit the Chemistry Web site at science.glencoe.com to find links about chemical reactions.

The electricity of a lightning bolt provides the energy that sparks chemical reactions among substances in the atmosphere.



DISCOVERY LAB



Observing a Change

An indicator is a chemical that shows when change occurs during a chemical reaction.

Safety Precautions



Always wear goggles and an apron in the laboratory.

Procedure

1. Measure 10.0 mL distilled water in a graduated cylinder and pour it into the beaker. Add one drop of 0.1M ammonia to the water.
2. Stir 15 drops of indicator into the solution with the stirring rod. Observe the solution's color. Measure its temperature with the thermometer.
3. Drop the effervescent tablet into the solution. Observe what happens. Record your observations, including any temperature change.

Analysis

Did a color change and a temperature change occur? Was a gas produced? Did a physical change or a chemical change occur? Explain.

Materials

distilled water	universal indicator
25-mL graduated cylinder	stirring rod
100-mL beaker	thermometer
pipettes (2)	effervescent tablet
0.1M ammonia	

Section

10.1

Reactions and Equations

Objectives

- **Recognize** evidence of chemical change.
- **Represent** chemical reactions with equations.

Vocabulary

chemical reaction
reactant
product
chemical equation
coefficient

Do you know that the foods you eat, the fibers in your clothes, and the plastic in your CDs have something in common? Foods, fibers, and plastic are produced when the atoms in substances are rearranged to form different substances. Atoms are rearranged during the flash of lightning shown in the photo on the opposite page. They were also rearranged when you dropped the effervescent tablet into the beaker of water and indicator in the **DISCOVERY LAB**.

Evidence of Chemical Reactions

The process by which the atoms of one or more substances are rearranged to form different substances is called a **chemical reaction**. A chemical reaction is another name for a chemical change, which you read about in Chapter 3. Chemical reactions affect every part of your life. They break down your food, producing the energy you need to live. They produce natural fibers such as cotton and wool in the bodies of plants and animals. In factories, they produce synthetic fibers such as nylon and polyesters. Chemical reactions in the engines of cars and buses provide the energy to power the vehicles.

How can you tell when a chemical reaction has taken place? Although some chemical reactions are hard to detect, many reactions provide evidence that they have occurred. A temperature change can indicate a chemical reaction. Many reactions, such as those that occur during a forest fire, release energy in the form of heat and light. Other reactions absorb heat.

In addition to a temperature change, other types of evidence may indicate that a chemical reaction has occurred. One indication of a chemical reaction is

Figure 10-1

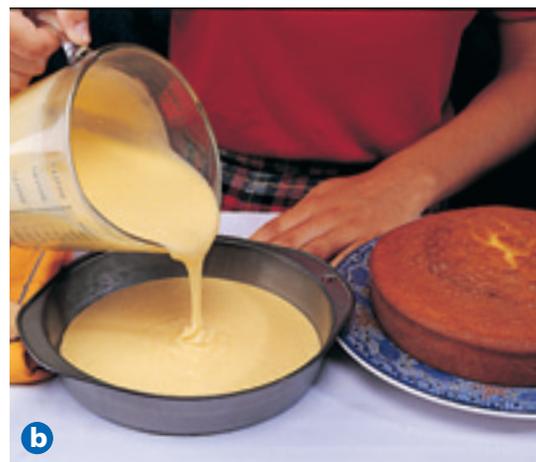
Each of these photos illustrates evidence of a chemical reaction.

a Reactions that happened when the marshmallow was burned are obvious by the color change.

b Chemical reactions occur in the oven when a cake mix is baked, namely, the formation of gas bubbles that cause the cake to rise.

c The tarnish that appears on silver and other metals is actually a solid that forms as a result of chemical reactions that take place when the metal is exposed to traces of sulfur compounds in the air.

d Numerous chemical reactions happen in an explosion. The appearance of smoke, the release of energy in the form of heat, and the permanent color change of materials involved are all evidence of chemical reactions.



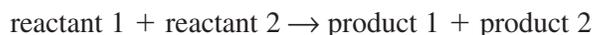
a color change. For example, you may have noticed that the color of some nails that are left outside changes from silver to orange-brown in a short time. The color change is evidence that a chemical reaction occurred between the iron in the nail and the oxygen in air. Odor, gas bubbles, and/or the appearance of a solid are other indications of chemical change. Each of the photographs in **Figure 10-1** shows evidence of a chemical reaction. Do you recognize the evidence in each?

Table 10-1

Symbols Used in Equations	
Symbol	Meaning
+	Separates two or more reactants or products
→	Separates reactants from products
(s)	Identifies solid state
(l)	Identifies liquid state
(g)	Identifies gaseous state
(aq)	Identifies water solution

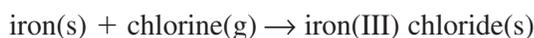
Representing Chemical Reactions

Chemists use statements called equations to represent chemical reactions. Their equations show a reaction's **reactants**, which are the starting substances, and **products**, which are the substances formed during the reaction. Chemical equations do not express numerical equalities as do mathematical equations because during chemical reactions the reactants are used up as the products form. Instead, the equations used by chemists show the direction in which the reaction progresses. Therefore, an arrow rather than an equal sign is used to separate the reactants from the products. You read the arrow as “react to produce” or “yield”. The reactants are written to the arrow's left, and the products are written to its right. When there are two or more reactants, or two or more products, a plus sign separates each reactant or each product. These elements of equation notation are shown below.



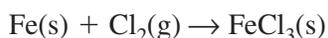
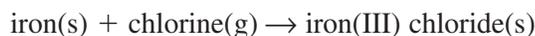
In equations, symbols are used to show the physical states of the reactants and products. Reactants and products can exist as solids, liquids, and gases. When they are dissolved in water, they are said to be aqueous. It is important to show the physical states of a reaction's reactants and products in an equation because the physical states provide clues about how the reaction occurs. Some basic symbols used in equations are shown in **Table 10-1**.

Word equations You can use statements called word equations to indicate the reactants and products of chemical reactions. The word equation below describes the reaction between iron and chlorine, which is shown in **Figure 10-2**. Iron is a solid and chlorine is a gas. The brown cloud in the photograph is composed of the reaction's product, which is solid particles of iron(III) chloride.



This word equation is read, "Iron and chlorine react to produce iron(III) chloride."

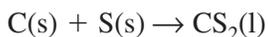
Skeleton equations Although word equations help to describe chemical reactions, they are cumbersome and lack important information. A skeleton equation uses chemical formulas rather than words to identify the reactants and the products. For example, the skeleton equation for the reaction between iron and chlorine uses the formulas for iron, chlorine, and iron(III) chloride in place of the words.



How would you write the skeleton equation that describes the reaction between carbon and sulfur to form carbon disulfide? Carbon and sulfur are solids. First, write the chemical formulas for the reactants to the left of an arrow. Then, separate the reactants with a plus sign and indicate their physical states.



Finally, write the chemical formula for the product, liquid carbon disulfide, to the right of the arrow and indicate its physical state. The result is the skeleton equation for the reaction.



This skeleton equation tells us that the reaction of carbon in the solid state reacts with sulfur in the solid state to produce carbon disulfide, which is in the liquid state.



Figure 10-2

Science, like all other disciplines, has a specialized language that allows specific information to be communicated in a uniform manner. This reaction between iron and chlorine can be described by a word equation, skeleton equation, or balanced chemical equation.

PRACTICE PROBLEMS

Write skeleton equations for the following word equations.

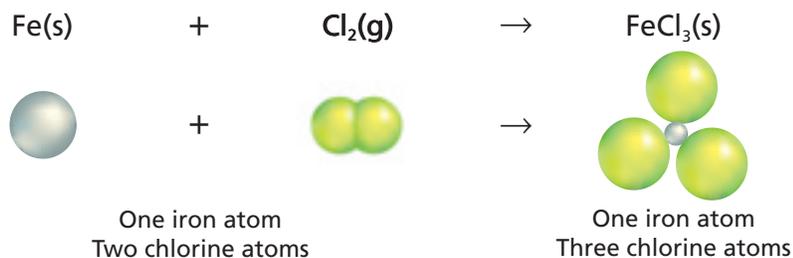
1. hydrogen(g) + bromine(g) \rightarrow hydrogen bromide(g)
2. carbon monoxide(g) + oxygen(g) \rightarrow carbon dioxide(g)
3. potassium chlorate(s) \rightarrow potassium chloride(s) + oxygen(g)



For more practice with writing skeleton equations, go to **Supplemental Practice Problems** in Appendix A.

Figure 10-3

The information conveyed by skeleton equations is limited. In this case, the skeleton equation (top) is correct, but it does not show the exact number of atoms that actually interact. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



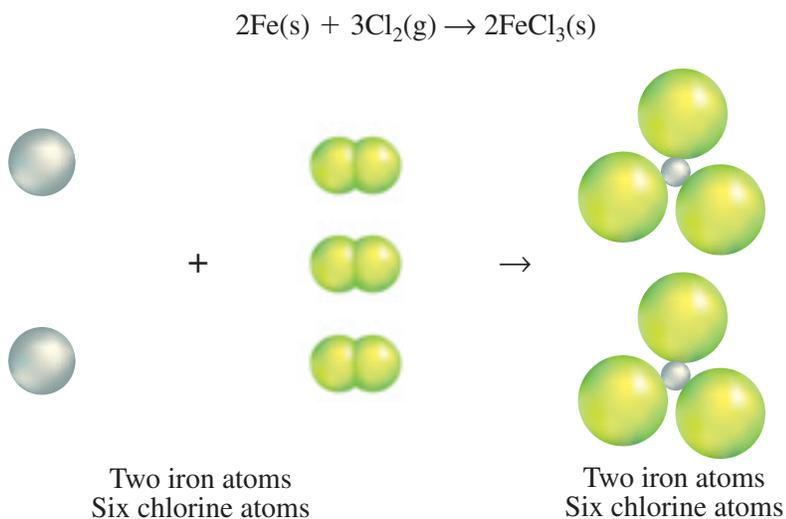
Chemical equations Writing a skeleton equation is an important step toward using an equation to completely describe a chemical reaction. But, like word equations, skeleton equations also lack important information about reactions. Recall from Chapter 3 that the law of conservation of mass states that in a chemical change, matter is neither created nor destroyed. Chemical equations must show that matter is conserved during a reaction, and skeleton equations lack that information.

Look at **Figure 10-3**. The skeleton equation for the reaction between iron and chlorine shows that one iron atom and two chlorine atoms react to produce a substance containing one iron atom and three chlorine atoms. Was a chlorine atom created in the reaction? Atoms are not created in chemical reactions, and to accurately show what happened, more information is needed.

To accurately represent a chemical reaction by an equation, the equation must show how the law of conservation of mass is obeyed. In other words, the equation must show that the number of atoms of each reactant and each product is equal on both sides of the arrow. Such an equation is called a balanced **chemical equation**. A chemical equation is a statement that uses chemical formulas to show the identities and relative amounts of the substances involved in a chemical reaction. It is chemical equations that chemists use most often to represent chemical reactions.

Balancing Chemical Equations

The balanced equation for the reaction between iron and chlorine, shown below, reflects the law of conservation of mass.



To balance an equation, you must find the correct coefficients for the chemical formulas in the skeleton equation. A **coefficient** in a chemical equation is the number written in front of a reactant or product. Coefficients are usually whole numbers, and are usually not written if the value is 1. A coefficient

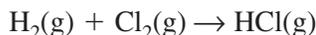


Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

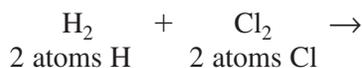
tells you the smallest number of particles of the substance involved in the reaction. That is, the coefficients in a balanced equation describe the lowest whole-number ratio of the amounts of all of the reactants and products.

Steps for balancing equations Most chemical equations can be balanced by following the steps given below. For example, you can use these steps to write the chemical equation for the reaction between hydrogen and chlorine that produces hydrogen chloride.

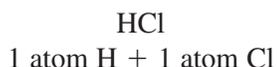
Step 1 Write the skeleton equation for the reaction. Make sure that the chemical formulas correctly represent the substances. An arrow separates the reactants from the products, and a plus sign separates multiple reactants and products. Show the physical states of all reactants and products.



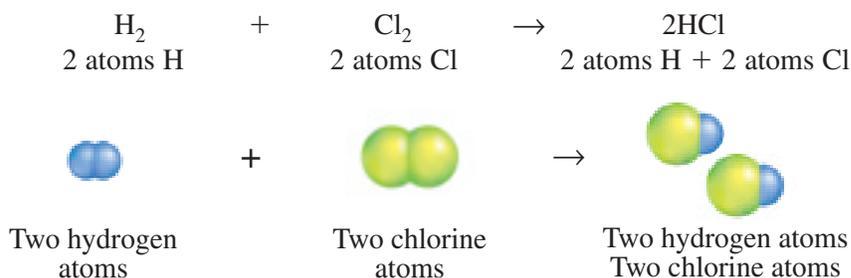
Step 2 Count the atoms of the elements in the reactants. If a reaction involves identical polyatomic ions in the reactants and products, count the ions as if they are elements. This reaction does not involve any polyatomic ions. Two atoms of hydrogen and two atoms of chlorine are reacting.



Step 3 Count the atoms of the elements in the products. One atom of hydrogen and one atom of chlorine are produced.



Step 4 Change the coefficients to make the number of atoms of each element equal on both sides of the equation. Never change a subscript in a chemical formula to balance an equation because doing so changes the identity of the substance.



Step 5 Write the coefficients in their lowest possible ratio. The coefficients should be the smallest possible whole numbers. The ratio 1 hydrogen to 1 chlorine to 2 hydrogen chloride (1:1:2) is the lowest possible ratio because the coefficients cannot be reduced and still remain whole numbers.

Step 6 Check your work. Make sure that the chemical formulas are written correctly. Then, check that the number of atoms of each element is equal on both sides of the equation.

Earth Science

CONNECTION

Weathering is the general term used to describe the ways in which rock is broken down at or near Earth's surface. Soils are the result of weathering and the activities of plants and animals.

Physical weathering, also called mechanical weathering, involves expansion and contraction with changes in temperature, pressure, and the growth of plants and organisms in the rock. Water in rock fissures and crevices cause rock to fracture when water expands during freezing. Freeze-thaw physical weathering is more likely to occur in sub-Arctic climates.

Chemical weathering involves the break down of rock by chemical reactions. The mineral composition of the rock is changed, reorganized, or redistributed. For example, minerals that contain iron may react with oxygen in the air. Water in which carbon dioxide is dissolved will dissolve limestone. Chemical weathering is more likely to take place in humid tropical climates.



EXAMPLE PROBLEM 10-1

Writing a Balanced Chemical Equation

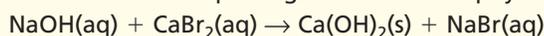
Write the balanced chemical equation for the reaction in which sodium hydroxide and calcium bromide react to produce solid calcium hydroxide and sodium bromide. The reaction occurs in water.

1. Analyze the Problem

You are given the reactants and products in a chemical reaction. Start with a skeleton equation and use the steps given in the text for balancing chemical equations.

2. Solve for the Unknown

Step 1 Write the skeleton equation. Be sure to put the reactants on the left side of an arrow and the products on the right. Separate the substances with plus signs and indicate physical states.



Step 2 Count the atoms of each element in the reactants.

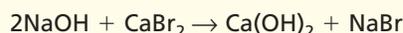
1 Na, 1 O, 1 H, 1 Ca, 2 Br

Step 3 Count the atoms of each element in the products.

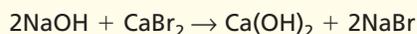
1 Na, 2 O, 2 H, 1 Ca, 1 Br

Step 4 Adjust the coefficients.

Insert the coefficient 2 in front of NaOH to balance the hydroxide ions.



Insert the coefficient 2 in front of NaBr to balance the Na and Br atoms.



Step 5 Write the coefficients in their lowest possible ratio.

The ratio of the coefficients is 2:1:1:2.

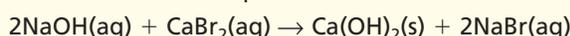
Step 6 Check to make sure that the number of atoms of each element is equal on both sides of the equation.

Reactants: 2 Na, 2 OH, 1 Ca, 2 Br

Products: 2 Na, 2 OH, 1 Ca, 2 Br

3. Evaluate the Answer

The chemical formulas for all substances are written correctly. The number of atoms of each element is equal on both sides of the equation. The coefficients are written in the lowest possible ratio. The balanced chemical equation for the reaction is



The brick and mortar used in many construction applications involves chemical reactions with calcium and other substances. Although these materials have been used for centuries, chemistry continues to improve their durability and performance.

PRACTICE PROBLEMS

Write chemical equations for each of the following reactions.

- In water, iron(III) chloride reacts with sodium hydroxide, producing solid iron(III) hydroxide and sodium chloride.
- Liquid carbon disulfide reacts with oxygen gas, producing carbon dioxide gas and sulfur dioxide gas.
- Solid zinc and aqueous hydrogen sulfate react to produce hydrogen gas and aqueous zinc sulfate.

Practice!

For more practice writing chemical equations, go to **Supplemental Practice Problems** in Appendix A.

Balancing Chemical Equations

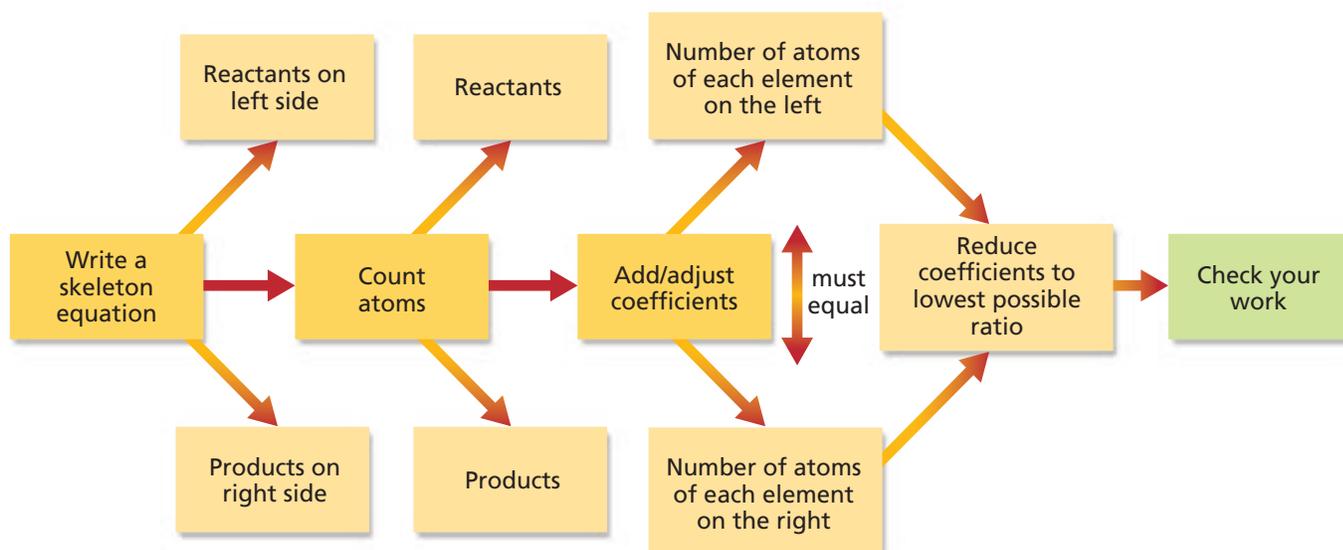


Figure 10-4

It is imperative to your study of chemistry to be able to balance chemical equations. Use this flowchart to help you master the skill.

Probably the most fundamental concept of chemistry is the law of conservation of mass that you first encountered in Chapter 3. All chemical reactions obey the law that matter is neither created nor destroyed. Therefore, it is also fundamental that the equations that represent chemical reactions include sufficient information to show that the reaction obeys the law of conservation of mass. You have learned how to show this relationship with balanced chemical equations. The flowchart shown in **Figure 10-4** summarizes the steps for balancing equations. You will undoubtedly find that some chemical equations can be balanced easily, whereas others are more difficult to balance. All chemical equations, however, can be balanced by the process you learned in this section.

Section 10.1 Assessment

- List three types of evidence that indicate a chemical reaction has occurred.
- Compare and contrast a skeleton equation and a chemical equation.
- Why is it important that a chemical equation be balanced?
- When balancing a chemical equation, can you adjust the number that is subscripted to a substance formula? Explain your answer.
- Why is it important to reduce coefficients in a balanced equation to the lowest possible whole-number ratio?
- Thinking Critically** Explain how an equation can be balanced even if the number of reactant particles differs from the number of product particles.
- Using Numbers** Is the following equation balanced? If not, correct the coefficients.
$$2\text{K}_2\text{CrO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbCrO}_4(\text{s})$$

Objectives

- **Classify** chemical reactions.
- **Identify** the characteristics of different classes of chemical reactions.

Vocabulary

synthesis reaction
 combustion reaction
 decomposition reaction
 single-replacement reaction
 double-replacement reaction
 precipitate

Figure 10-5

Without some level of organization, it would be difficult to shelve and maintain the huge number of books that have been published. Chemistry, too, hinges on strict rules of organization. Chemical reactions are classified as synthesis, combustion, decomposition, single-replacement, and double-replacement reactions.

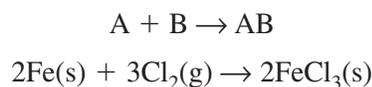


How long do you think it would take you to find your favorite author's new novel in an unorganized book store? Because there are so many books in book stores, it could take you a very long time. Book stores, such as the store shown in **Figure 10-5**, supermarkets, and music stores are among the many places where things are classified and organized. Chemists classify chemical reactions in order to organize the many reactions that occur daily in living things, laboratories, and industry. Knowing the categories of chemical reactions can help you remember and understand them. It also can help you recognize patterns and predict the products of many chemical reactions.

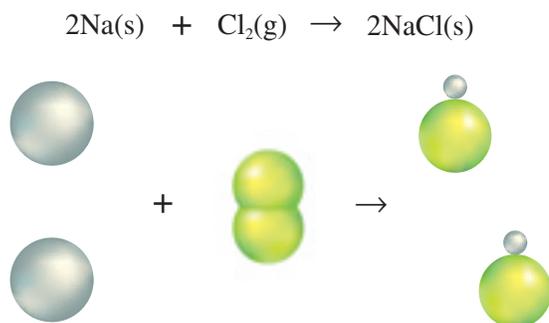
Chemists classify reactions in different ways. One way is to distinguish among five types of chemical reactions: synthesis, combustion, decomposition, single-replacement, and double-replacement reactions. Some reactions fit equally well into more than one of these classes.

Synthesis Reactions

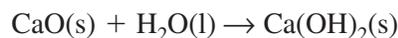
In the previous section, you read about the reaction that occurs between iron and chlorine gas to produce iron(III) chloride. In this reaction, two elements (A and B) combine to produce one new compound (AB).



The reaction between iron and chlorine gas is an example of a **synthesis reaction**—a chemical reaction in which two or more substances react to produce a single product. When two elements react, the reaction is always a synthesis reaction. Another example of a synthesis reaction is shown below. In this reaction, sodium and chlorine react to produce sodium chloride.



Just as two elements can combine, two compounds can also combine to form one compound. For example, the reaction between calcium oxide and water to form calcium hydroxide is a synthesis reaction.



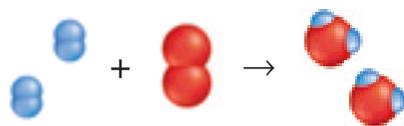
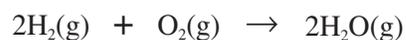
Another type of synthesis reaction may involve a reaction between a compound and an element, as happens when sulfur dioxide gas reacts with oxygen gas to form sulfur trioxide.



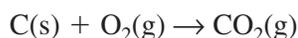
Combustion Reactions

The synthesis reaction between sulfur dioxide and oxygen can be classified also as a combustion reaction. In a **combustion reaction**, oxygen combines with a substance and releases energy in the form of heat and light. Oxygen can combine in this way with many different substances, making combustion reactions common.

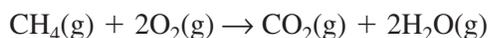
A combustion reaction, such as the one shown in **Figure 10-6**, occurs between hydrogen and oxygen when hydrogen is heated. Water is formed during the reaction and a large amount of energy is released.



Another important combustion reaction occurs when coal is burned to produce energy. Coal is called a fossil fuel because it contains the remains of plants that lived long ago. It is composed primarily of the element carbon. Coal-burning power plants generate electric power in many parts of the United States. The primary reaction that occurs in these plants is between carbon and oxygen.



Note that the combustion reactions just mentioned are also synthesis reactions. However, not all combustion reactions are synthesis reactions. For example, the reaction involving methane gas, CH_4 , and oxygen illustrates a combustion reaction in which one substance replaces another in the formation of products.



Methane, which belongs to a group of substances called hydrocarbons, is the major component of natural gas. All hydrocarbons contain carbon and hydrogen and burn in oxygen to yield the same products as methane does—carbon dioxide and water. For example, most cars and trucks are powered by gasoline, which contains hydrocarbons. In engines, gasoline is combined with oxygen, producing carbon dioxide, water, and energy that powers the vehicles. You will learn more about hydrocarbons in Chapter 22.



Figure 10-6

The tragedy of the *Challenger* space mission was the result of a combustion reaction between oxygen and hydrogen.

PRACTICE PROBLEMS

Write chemical equations for the following reactions. Classify each reaction into as many categories as possible.

14. The solids aluminum and sulfur react to produce aluminum sulfide.
15. Water and dinitrogen pentoxide gas react to produce aqueous hydrogen nitrate.
16. The gases nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.
17. Ethane gas (C_2H_6) burns in air, producing carbon dioxide gas and water vapor.



For more practice with writing synthesis and combustion equations, go to **Supplemental Practice Problems** in Appendix A.

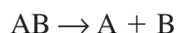
Figure 10-7

The decomposition of sodium azide, which produces a gas, is the key to inflatable air bags.



Decomposition Reactions

Some chemical reactions are essentially the opposite of synthesis reactions. These reactions are classified as decomposition reactions. A **decomposition reaction** is one in which a single compound breaks down into two or more elements or new compounds. In generic terms, decomposition reactions look like the following.



Decomposition reactions often require an energy source, such as heat, light, or electricity, to occur. For example, ammonium nitrate breaks down into dinitrogen monoxide and water when the reactant is heated to high temperature.



You can see that this decomposition reaction involves one reactant breaking down into more than one product.

The outcome of another decomposition reaction is shown in **Figure 10-7**. Automobile safety air bags inflate rapidly as sodium azide pellets decompose. A device that can provide an electric signal to start the reaction is packaged inside air bags along with the sodium azide pellets. When the device is activated, sodium azide decomposes, producing nitrogen gas that quickly inflates the safety bag.



PRACTICE PROBLEMS

Write chemical equations for the following decomposition reactions.

18. Aluminum oxide(s) decomposes when electricity is passed through it.
19. Nickel(II) hydroxide(s) decomposes to produce nickel(II) oxide(s) and water.
20. Heating sodium hydrogen carbonate(s) produces sodium carbonate(aq), carbon dioxide(g), and water.

Practice!

For more practice with writing decomposition equations, go to **Supplemental Practice Problems** in Appendix A.

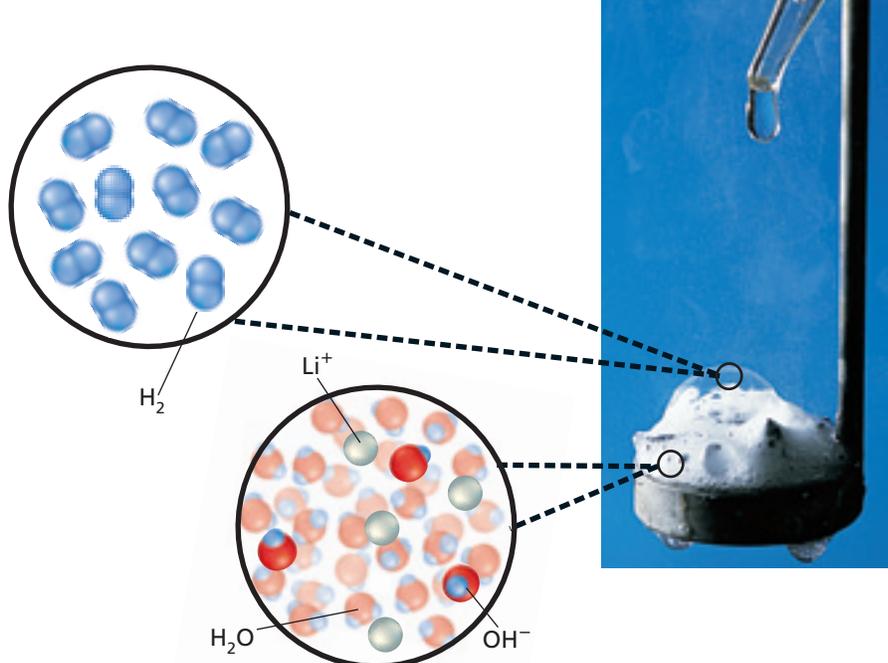


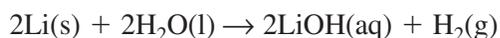
Figure 10-8

The reaction of lithium and water is a single-replacement reaction. Lithium replaces a hydrogen in water, and the products of the reaction are aqueous lithium hydroxide and hydrogen gas. Lithium hydroxide exists as lithium and hydroxide ions in solution.

Replacement Reactions

In contrast to synthesis, combustion, and decomposition reactions, many chemical reactions involve the replacement of an element in a compound. There are two types of replacement reactions: single-replacement reactions and double-replacement reactions.

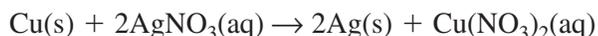
Single-replacement reactions Now that you've seen how atoms and molecules rearrange in synthesis and combustion reactions, look closely at the reaction between lithium and water that is shown in **Figure 10-8**. The expanded view of the reaction at the molecular level shows that a lithium atom replaces one of the hydrogen atoms in a water molecule. The following chemical equation describes this activity.



A reaction in which the atoms of one element replace the atoms of another element in a compound is called a **single-replacement reaction**.



The reaction between lithium and water is one type of single-replacement reaction in which a metal replaces a hydrogen in a water molecule. Another type of single-replacement reaction occurs when one metal replaces another metal in a compound dissolved in water. For example, **Figure 10-9** shows a single-replacement reaction occurring when a spiral of pure copper wire is placed in aqueous silver nitrate. The shiny crystals that are accumulating on the copper wire are the silver atoms that the copper atoms replaced.



A metal will not always replace another metal in a compound dissolved in water. This is because metals differ in their reactivities. A metal's reactivity is its ability to react with another substance. In **Figure 10-10** you see an activity series of some metals. This series orders metals by their reactivity with other metals. Single-replacement reactions like the one between copper and aqueous silver nitrate determine a metal's position on the list. The most active



Figure 10-9

The chemical equation for the single-replacement reaction involving copper and silver nitrate clearly describes the replacement of silver by copper, but the visual evidence of this chemical reaction is a solid precipitate.

Most active	↓	METALS
		Lithium
		Rubidium
		Potassium
		Calcium
		Sodium
		Magnesium
		Aluminum
		Manganese
		Zinc
		Iron
		Nickel
		Tin
		Lead
		Copper
		Silver
		Platinum
Gold		
Least active	↓	HALOGENS
		Fluorine
		Chlorine
		Bromine
Most active	↓	Iodine

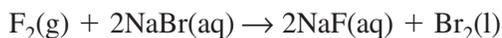
metals, which are those that do replace the metal in a compound, are at the top of the list. The least active metals are at the bottom.

You can use **Figure 10-10** to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, you saw in **Figure 10-9** that copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper(II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series and no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.



The **CHEMLAB** at the end of this chapter gives you an opportunity to explore the activities of metals in the laboratory.

A third type of single-replacement reaction involves the replacement of a nonmetal in a compound by another nonmetal. Halogens are frequently involved in these types of reactions. Like metals, halogens exhibit different activity levels in single-replacement reactions. The reactivities of halogens, determined by single-replacement reactions, are also shown in **Figure 10-10**. The most active halogen is fluorine, and the least active is iodine. A more reactive halogen replaces a less reactive halogen that is part of a compound dissolved in water. For example, fluorine replaces bromine in water containing dissolved sodium bromide. However, bromine does not replace fluorine in water containing dissolved sodium fluoride.



The **problem-solving LAB** below will help you to relate periodic trends of the halogens to their reactivities.

Figure 10-10

An activity series, similar to the series shown here for various metals and halogens, is a useful tool for determining whether a chemical reaction will take place and for determining the result of a replacement reaction.

problem-solving LAB

Can you predict the reactivities of halogens?

Analyzing and Concluding The location of all the halogens in group 7A in the periodic table tells you that halogens have common characteristics. Indeed, halogens are all nonmetals and have seven electrons in their outermost orbitals. However, each halogen has its own characteristics, too, such as its ability to react with other substances.

Analysis

Examine the accompanying table. It includes data about the atomic radii, ionization energies, and electronegativities of the halogens.

Thinking Critically

1. Describe any periodic trends that you identify in the table data.

Properties of Halogens			
Halogen	Atomic radius (pm)	Ionization energy (kJ/mol)	Electronegativity
Fluorine	72	1681	3.98
Chlorine	100	1251	3.16
Bromine	114	1140	2.96
Iodine	133	1008	2.66
Astatine	140	—	2.2

2. Relate any periodic trends that you identify among the halogens to the activity series of halogens shown in **Figure 10-10**.

3. Predict the location of the element astatine in the activity series of halogens. Explain your answer.

EXAMPLE PROBLEM 10-2

Single-Replacement Reactions

Predict the products that will result when these reactants combine and write a balanced chemical equation for each reaction.



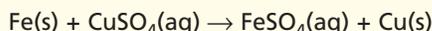
1. Analyze the Problem

You are given three sets of reactants. Using **Figure 10-10**, you must first determine if each reaction takes place. Then, if a reaction is predicted, you can determine the product(s) of the reaction. With this information you can write a skeleton equation for the reaction.

Finally, you can use the steps for balancing chemical equations to write the complete balanced chemical equation.

2. Solve for the Unknown

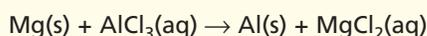
Iron is listed above copper in the metals activity series. Therefore, the first reaction will take place because iron is more reactive than copper. In this case, iron will replace copper. The skeleton equation for this reaction is



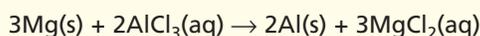
This equation is balanced.

In the second reaction, chlorine is more reactive than bromine because bromine is listed below chlorine in the halogen activity series. Therefore, the reaction will not take place. The skeleton equation for this situation is $\text{Br(l)} + \text{MgCl}_2\text{(aq)} \rightarrow \text{NR}$. No balancing is required.

Magnesium is listed above aluminum in the metals activity series. Therefore, the third reaction will take place because magnesium is more reactive than aluminum. In this case, magnesium will replace aluminum. The skeleton equation for this reaction is



This equation is not balanced. The balanced equation is



3. Evaluate the Answer

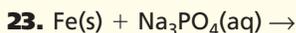
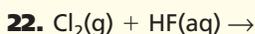
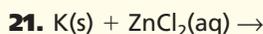
The activity series shown in **Figure 10-10** supports the reaction predictions. The chemical equations are balanced because the number of atoms of each substance is equal on both sides of the equation.



Magnesium is an essential element for the human body. You can ensure an adequate magnesium intake by eating magnesium-rich foods.

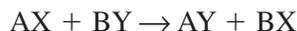
PRACTICE PROBLEMS

Predict if the following single-replacement reactions will occur. If a reaction occurs, write a balanced equation for the reaction.

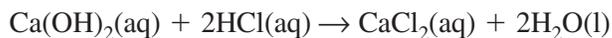


For more practice with predicting if single-replacement reactions will occur, go to **Supplemental Practice Problems** in Appendix A.

Double-replacement reactions The final type of replacement reaction which involves an exchange of ions between two compounds is called a **double-replacement reaction**.

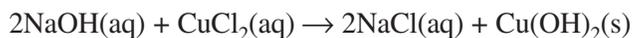


In this generic equation, A and B represent positively charged ions (cations), and X and Y represent negatively charged ions (anions). You can see that the anions have switched places and are now bonded to the other cations in the reaction. In other words, X replaces Y and Y replaces X—a double replacement. More simply, you might say that the positive and negative ions of two compounds switch places. The reaction between calcium hydroxide and hydrochloric acid is a double-replacement reaction.



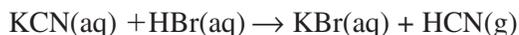
The ionic components of the reaction are Ca^{2+} , OH^- , H^+ , and Cl^- . Knowing this, you can now see the two replacements of the reaction. The anions (OH^- and Cl^-) have changed places and are now bonded to the other cations (Ca^{2+} and H^+) in the reaction.

The reaction between sodium hydroxide and copper(II) chloride in solution is also a double-replacement reaction.



In this case, the anions (OH^- and Cl^-) changed places and are now associated with the other cations (Na^+ and Cu^{2+}). The result of this reaction is a solid product, copper(II) hydroxide. A solid produced during a chemical reaction in a solution is called a **precipitate**.

One of the key characteristics of double-replacement reactions is the type of product that is formed when the reaction takes place. All double-replacement reactions produce either a precipitate, a gas, or water. An example of a double-replacement reaction that forms a gas is that of potassium cyanide and hydrobromic acid.



It is important to be able to evaluate the chemistry of double-replacement reactions and predict the products of these reactions. The basic steps to do this are given in **Table 10-2**.

Table 10-2

Guidelines for Double-Replacement Reactions	
Step	Example
1. Write the components of the reactants in a skeleton equation.	$\text{Al(NO}_3)_3 + \text{H}_2\text{SO}_4$
2. Identify the cations and anions in each compound.	$\text{Al(NO}_3)_3$ has Al^{3+} and NO_3^- H_2SO_4 has H^+ and SO_4^{2-}
3. Pair up each cation with the anion from the other compound.	Al^{3+} pairs with SO_4^{2-} H^+ pairs with NO_3^-
4. Write the formulas for the products using the pairs from step 3.	$\text{Al}_2(\text{SO}_4)_3$ HNO_3
5. Write the complete equation for the double-replacement reaction.	$\text{Al(NO}_3)_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + \text{HNO}_3$
6. Balance the equation.	$2\text{Al(NO}_3)_3 + 3\text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 6\text{HNO}_3$



Now use this information to work the following practice problems.

PRACTICE PROBLEMS

Write the balanced chemical equations for the following double-replacement reactions.

24. Aqueous lithium iodide and aqueous silver nitrate react to produce solid silver iodide and aqueous lithium nitrate.
25. Aqueous barium chloride and aqueous potassium carbonate react to produce solid barium carbonate and aqueous potassium chloride.
26. Aqueous sodium oxalate and aqueous lead(II) nitrate react to produce solid lead(II) oxalate and aqueous sodium nitrate.



For more practice with writing double-replacement equations, go to **Supplemental Practice Problems** in Appendix A.

Now that you have learned about the various types of chemical reactions, you can use **Table 10-3** to help you organize them in a way such that you can identify each and predict its products.

As the table indicates, the components of double-replacement reactions are dissolved in water. As you continue with Section 10.3, you will learn more about double-replacement reactions in aqueous solutions.

Table 10-3

Predicting Products of Chemical Reactions		
Class of reaction	Reactants	Probable products
Synthesis	Two or more substances	One compound
Combustion	A metal and oxygen A nonmetal and oxygen A compound and oxygen	The oxide of the metal The oxide of the nonmetal Two or more oxides
Decomposition	One compound	Two or more elements and/or compounds
Single-replacement	A metal and a compound A nonmetal and a compound	A new compound and the replaced metal A new compound and the replaced nonmetal
Double-replacement	Two compounds	Two different compounds, one of which is often a solid, water, or a gas

Section 10.2 Assessment

27. What are the five classes of chemical reactions?
28. Identify two characteristics of combustion reactions.
29. Compare and contrast single-replacement reactions and double-replacement reactions.
30. Describe the result of a double-replacement reaction.
31. **Thinking Critically** Does the following reaction occur? Explain your answer.
$$3\text{Ni} + 2\text{AuBr}_3 \rightarrow 3\text{NiBr}_2 + 2\text{Au}$$
32. **Classifying** What type of reaction is most likely to occur when barium reacts with fluorine? Write the chemical equation for the reaction.

Objectives

- **Describe** aqueous solutions.
- **Write** complete ionic and net ionic equations for chemical reactions in aqueous solutions.
- **Predict** whether reactions in aqueous solutions will produce a precipitate, water, or a gas.

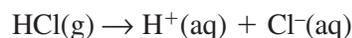
Vocabulary

solute
solvent
aqueous solution
complete ionic equation
spectator ion
net ionic equation

Many of the reactions discussed in the previous section involve substances dissolved in water. When a substance dissolves in water, a solution forms. You learned in Chapter 3 that a solution is a homogeneous mixture. A solution contains one or more substances called **solutes** dissolved in the water. In this case, water is the **solvent**, the most plentiful substance in the solution. An **aqueous solution** is a solution in which the solvent is water. Read the **How It Works** feature at the end of this chapter to see how aqueous solutions are used in hot and cold packs.

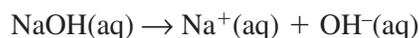
Aqueous Solutions

Although water is always the solvent in aqueous solutions, there are many possible solutes. Some solutes, such as sucrose (table sugar) and ethanol (grain alcohol), are molecular compounds that exist as molecules in aqueous solutions. Other solutes are molecular compounds that form ions when they dissolve in water. For example, the molecular compound hydrogen chloride forms hydrogen ions and chloride ions when it dissolves in water, as shown in **Figure 10-11**. An equation can be used to show this process.



Compounds such as hydrogen chloride that produce hydrogen ions in aqueous solution are acids. In fact, an aqueous solution of hydrogen chloride is often referred to as hydrochloric acid. You'll learn more about acids in Chapter 19.

In addition to molecular compounds, ionic compounds may be solutes in aqueous solutions. Recall from Chapter 8 that ionic compounds consist of positive ions and negative ions held together by ionic bonds. When ionic compounds dissolve in water, their ions can separate. The equation below shows an aqueous solution of the ionic compound sodium hydroxide.



When two aqueous solutions that contain ions as solutes are combined, the ions may react with one another. These reactions are always double-replacement reactions. The solvent molecules, which are all water molecules, do not usually react. Three types of products can form from the double-replacement reaction: precipitate, water, or gas. You can observe a precipitate forming when you do the **miniLAB** for this chapter.

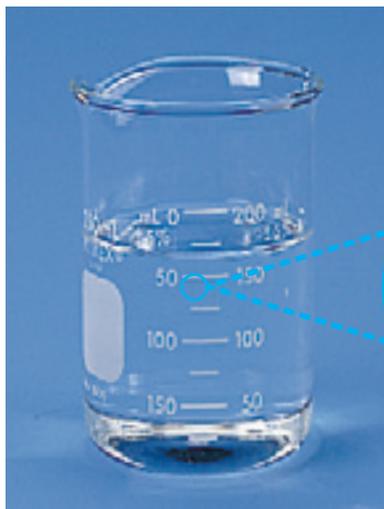
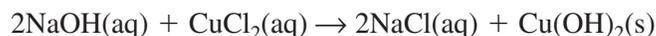


Figure 10-11

In an aqueous solution, hydrogen chloride (HCl) breaks apart into hydrogen ions (H^+) and chloride ions (Cl^-).

Reactions That Form Precipitates

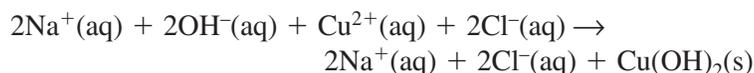
Some reactions that occur in aqueous solutions produce precipitates. For example, when aqueous solutions of sodium hydroxide and copper(II) chloride are mixed, a double-replacement reaction occurs in which the precipitate copper(II) hydroxide forms.



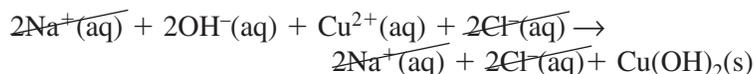
This is shown in **Figure 10-12**.

Note that the chemical equation does not show some details of this reaction. Sodium hydroxide and copper(II) chloride are ionic compounds. Therefore, in aqueous solutions they exist as Na^+ , OH^- , Cu^{2+} , and Cl^- ions. When their solutions are combined, Cu^{2+} ions in one solution and OH^- ions in the other solution react to form the precipitate copper(II) hydroxide, $\text{Cu}(\text{OH})_2(\text{s})$. The Na^+ and Cl^- ions remain dissolved in the new solution.

To show the details of reactions that involve ions in aqueous solutions, chemists use ionic equations. Ionic equations differ from chemical equations in that substances that are ions in solution are written as ions in the equation. Look again at the reaction between aqueous solutions of sodium hydroxide and copper(II) chloride. To write the ionic equation for this reaction, you must show the reactants $\text{NaOH}(\text{aq})$ and $\text{CuCl}_2(\text{aq})$ and the product $\text{NaCl}(\text{aq})$ as ions.



An ionic equation that shows all of the particles in a solution as they realistically exist is called a **complete ionic equation**. Note that the sodium ions and the chloride ions are both reactants and products. Because they are both reactants and products, they do not participate in the reaction. Ions that do not participate in a reaction are called **spectator ions** and usually are not shown in ionic equations. Ionic equations that include only the particles that participate in the reaction are called **net ionic equations**. Net ionic equations are written from complete ionic equations by crossing out all spectator ions. For example, a net ionic equation is what remains after the sodium and chloride ions are crossed out of this complete ionic equation.



Only the hydroxide and copper ions are left in the net ionic equation shown below.

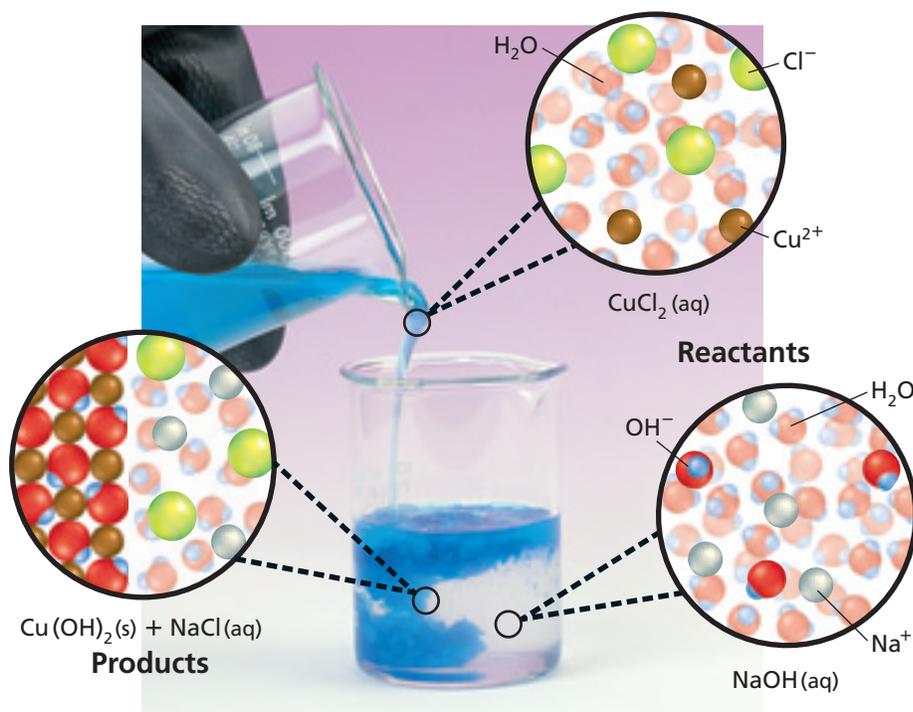
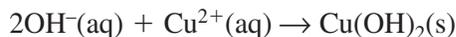


Figure 10-12

Like the aqueous solution of HCl you saw in **Figure 10-11**, sodium hydroxide (NaOH) in an aqueous solution dissociates into its ions sodium (Na^+) and hydroxide (OH^-). Copper chloride (CuCl_2) also dissociates into its ions, Cu^{2+} and Cl^- .

EXAMPLE PROBLEM 10-3



Barium carbonate (BaCO_3) is used in the rubber that is eventually processed into tires.

Reactions That Form a Precipitate

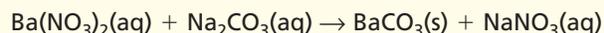
Write the chemical, complete ionic, and net ionic equations for the reaction between aqueous solutions of barium nitrate and sodium carbonate that forms the precipitate barium carbonate.

1. Analyze the Problem

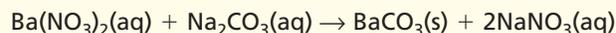
You are given the word equation for the reaction between barium nitrate and sodium carbonate. You must determine the chemical formulas and relative amounts of all reactants and products to write the chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation you can write the net ionic equation. The net ionic equation will include fewer substances than the other equations.

2. Solve for the Unknown

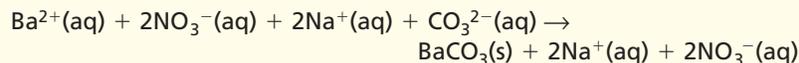
Write the correct chemical formulas and physical states for all substances involved in the reaction.



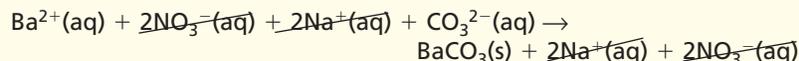
Balance the skeleton equation.



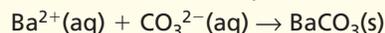
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.



3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only the reacting particles. The particles that compose the solid precipitate that is the result of the reaction are no longer ions.

PRACTICE PROBLEMS

Write chemical, complete ionic, and net ionic equations for the following reactions that may produce precipitates. Use NR to indicate that no reaction occurs.

33. Aqueous solutions of potassium iodide and silver nitrate are mixed, forming the precipitate silver iodide.
34. Aqueous solutions of ammonium phosphate and sodium sulfate are mixed. No precipitate forms and no gas is produced.
35. Aqueous solutions of aluminum chloride and sodium hydroxide are mixed, forming the precipitate aluminum hydroxide.
36. Aqueous solutions of lithium sulfate and calcium nitrate are mixed, forming the precipitate calcium sulfate.
37. Aqueous solutions of sodium carbonate and manganese(V) chloride are mixed, forming the precipitate manganese(V) carbonate.

Practice!

For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

Observing a Precipitate-Forming Reaction

Applying Concepts When two clear, colorless solutions are mixed, a chemical reaction may occur, resulting in the formation of a precipitate.

Materials 150-mL beakers (2); 100-mL graduated cylinder; stirring rod (2); spatula (2); weighing paper (2); NaOH; Epsom salts ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$); distilled water, balance

Procedure

- CAUTION:** Use gloves when working with NaOH. Measure about 4 g NaOH and place it in a 150-mL beaker. Add 50 mL distilled water to the NaOH. Mix with a stirring rod until the NaOH dissolves.
- Measure about 6 g Epsom salts and place it in another 150-mL beaker. Add 50 mL distilled water to the Epsom salts. Mix with another stirring rod until the Epsom salts dissolve.
- Slowly pour the Epsom salts solution into the NaOH solution. Record your observations.
- Stir the new solution. Record your observations.
- Allow the precipitate to settle, then decant the



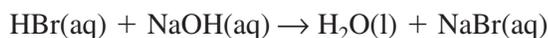
liquid from the solid. Dispose of the solid as your teacher instructs.

Analysis

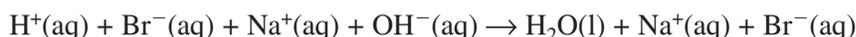
- Write a chemical equation for the reaction between the NaOH and MgSO_4 . Most sulfate compounds exist as ions in aqueous solutions.
- Write the complete ionic equation for this reaction.
- Write the net ionic equation for this reaction.

Reactions That Form Water

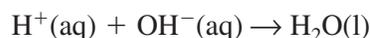
Another type of double-replacement reaction that occurs in an aqueous solution produces water molecules. The water molecules produced in the reaction increase the number of solvent particles. Unlike reactions in which a precipitate forms, no evidence of a chemical reaction is observable because water is colorless and odorless and already makes up most of the solution. For example, when you mix hydrobromic acid with a sodium hydroxide solution, a double-replacement reaction occurs and water is formed.



In this case, the reactants and the product sodium bromide exist as ions in an aqueous solution. The complete ionic equation for this reaction shows these ions.



Look carefully at the complete ionic equation. The reacting solute ions are the hydrogen and hydroxide ions because the sodium and bromine ions are both spectator ions. If you cross out the spectator ions, you are left with the ions that take part in the reaction.



This equation is the net ionic equation for the reaction.

EXAMPLE PROBLEM 10-4



Hearing aids often use lithium batteries because batteries made with lithium are lightweight and have a long lifetime.

Reactions That Form Water

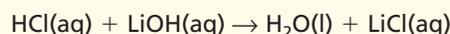
Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous lithium hydroxide, which produces water.

1. Analyze the Problem

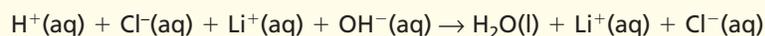
You are given the word equation for the reaction that occurs between hydrochloric acid and lithium hydroxide. You must determine the chemical formulas for and relative amounts of all reactants and products to write the chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation you can write the net ionic equation.

2. Solve for the Unknown

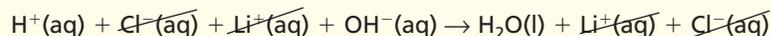
Write the skeleton equation for the reaction and balance it.



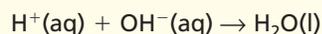
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation.



3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces water. The particles that compose the product water are no longer ions.

PRACTICE PROBLEMS

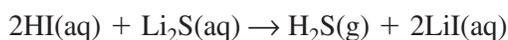
Write chemical, complete ionic, and net ionic equations for the reactions between the following substances, which produce water.

38. Sulfuric acid (H_2SO_4) and aqueous potassium hydroxide
39. Hydrochloric acid (HCl) and aqueous calcium hydroxide
40. Nitric acid (HNO_3) and aqueous ammonium hydroxide
41. Hydrosulfuric acid (H_2S) and aqueous calcium hydroxide
42. Phosphoric acid (H_3PO_4) and aqueous magnesium hydroxide

Reactions That Form Gases

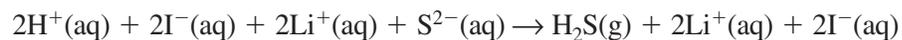
A third type of double-replacement reaction that occurs in aqueous solutions results in the formation of a gas. Some gases commonly produced in these reactions are carbon dioxide, hydrogen cyanide, and hydrogen sulfide.

A gas-producing reaction occurs when you mix hydroiodic acid (HI) with an aqueous solution of lithium sulfide. Bubbles of hydrogen sulfide gas form in the container during the reaction. Lithium iodide is also produced in this reaction and remains dissolved in the solution.

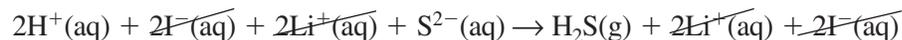


For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

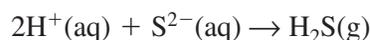
The reactants hydroiodic acid and lithium sulfide exist as ions in aqueous solution. Therefore, you can write an ionic equation for this reaction. The complete ionic equation includes all of the substances in the solution.



Note that there are many spectator ions in the equation. When the spectator ions are crossed out, only the substances involved in the reaction remain in the equation.



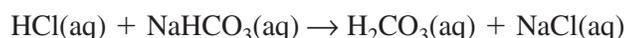
This is the net ionic equation.



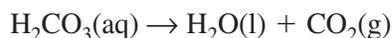
You observed another gas-producing reaction in the **DISCOVERY LAB** at the beginning of this chapter. In that reaction carbon dioxide gas was produced and bubbled out of the solution. Another reaction that produces carbon dioxide gas occurs in your kitchen when you mix vinegar and baking soda. Vinegar is an aqueous solution of acetic acid and water. Baking soda essentially consists of sodium hydrogen carbonate. Rapid bubbling occurs when vinegar and baking soda are combined. The bubbles are carbon dioxide gas escaping from the solution. You can see this reaction occurring in **Figure 10-13**.

A reaction similar to the one between vinegar and baking soda occurs when you combine any acidic solution and sodium hydrogen carbonate. In all cases, two reactions must occur almost simultaneously in the solution to produce the carbon dioxide gas. One reaction is double-replacement and the other is decomposition.

For example, when you dissolve sodium hydrogen carbonate in hydrochloric acid, a gas-producing double-replacement reaction occurs. The hydrogen in the hydrochloric acid and the sodium in the sodium hydrogen carbonate replace each other.



Sodium chloride is an ionic compound and its ions remain separate in the aqueous solution. However, as the carbonic acid (H_2CO_3) forms, it decomposes immediately into water and carbon dioxide.



Careers Using Chemistry

Pastry Chef

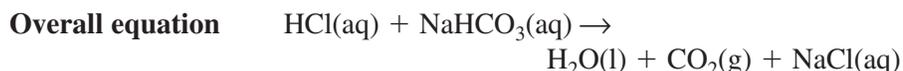
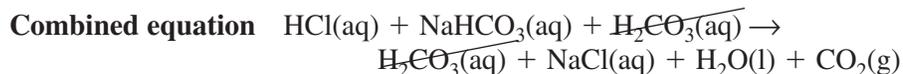
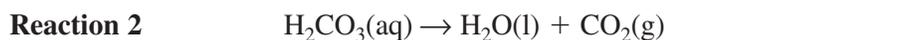
Do you like to make desserts that are as wonderful to look at as they are to eat? Then consider a career as a pastry chef.

Pastry chefs work in hotels, bakeries, catering, and manufacturing. Like the owner of a business, a pastry chef also must take on financial and managerial tasks. Creativity as well as a solid understanding of the chemistry of cooking is a must.

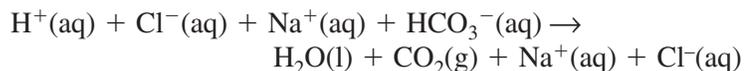
Figure 10-13

When vinegar and baking soda (sodium hydrogen carbonate, NaHCO_3) combine, the result is a vigorous bubbling that releases carbon dioxide (CO_2).

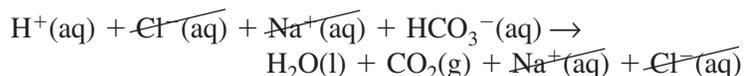
The two reactions can be combined and represented by one chemical equation in a process similar to adding mathematical equations. An equation that combines two reactions is called an overall equation. To write an overall equation, the reactants in the two reactions are written on the reactant side of the combined equation, and the products of the two reactions are written on the product side. Then any substances that are on both sides of the equation are crossed out.



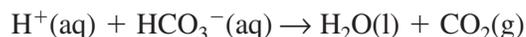
In this case, the reactants in the overall equation exist as ions in aqueous solutions. Therefore, a complete ionic equation can be written for the reaction.



Note that the sodium and chloride ions are the spectator ions. When you cross them out only the substances that take part in the reaction remain.



The net ionic equation shows that both water and carbon dioxide gas are produced in this reaction.



This is an important reaction in your life. This reaction is occurring in the blood vessels of your lungs as you read these words. The carbon dioxide gas produced in your cells is transported in your blood in the form of the bicarbonate ion (HCO_3^-). In the blood vessels of your lungs, the HCO_3^- ions combine with H^+ ions to produce CO_2 , which you exhale. This reaction also occurs in sodium bicarbonate products, such as those shown in **Figure 10-14**, that are made with baking soda.



Figure 10-14

The reactions of sodium bicarbonate in aqueous solution have many household applications. This sampling of products shows a variety of products that involve a chemical reaction of sodium bicarbonate.

EXAMPLE PROBLEM 10-5

Reactions That Form Gases

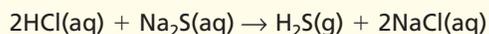
Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous sodium sulfide, which produces hydrogen sulfide gas.

1. Analyze the Problem

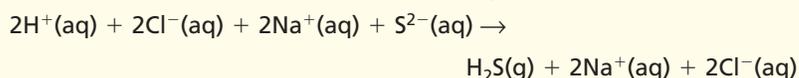
You are given the word equation for the reaction between hydrochloric acid and sodium sulfide. You must write the skeleton equation and balance it. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions in the complete ionic equation, you can write the net ionic equation.

2. Solve for the Unknown

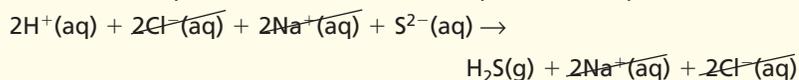
Write the correct skeleton equation for the reaction and balance it.



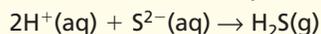
Show the ionic states of the reactants and products.



Cross out the spectator ions from the complete ionic equation.



Write the net ionic equation in its smallest whole number ratio.



3. Evaluate the Answer

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces hydrogen sulfide. The particles that compose the product are no longer ions.



Many chemical reactions are obvious by the results that can be detected by the senses. The chemical reaction that occurs when eggs rot gives off hydrogen sulfide (H_2S) gas, which has a pungent, distinctive odor.

PRACTICE PROBLEMS

Write chemical, complete ionic, and net ionic equations for these reactions.

- Perchloric acid (HClO_4) reacts with aqueous potassium carbonate.
- Sulfuric acid (H_2SO_4) reacts with aqueous sodium cyanide.
- Hydrobromic acid (HBr) reacts with aqueous ammonium carbonate.
- Nitric acid (HNO_3) reacts with aqueous potassium rubidium sulfide.



For more practice with writing ionic equations, go to **Supplemental Practice Problems** in Appendix A.

Section 10.3 Assessment

- Describe an aqueous solution.
- Distinguish between a complete ionic equation and a net ionic equation.
- What are three common types of products produced by reactions that occur in aqueous solutions?
- Thinking Critically** Explain why net ionic equations communicate more than chemical equations about reactions in aqueous solutions.
- Communicating** Describe the reaction of aqueous solutions of sodium sulfide and copper(II) sulfate, producing the precipitate copper(II) sulfide.

Activities of Metals

Some metals are more reactive than others. By comparing how different metals react with the same ions in aqueous solutions, an activity series for the tested metals can be developed. The activity series will reflect the relative reactivity of the tested metals. It can be used to predict whether reactions will occur.

Problem

Which is the most reactive metal tested? Which is the least reactive metal tested? Can this information be used to predict whether reactions will occur?

Objectives

- **Observe** chemical reactions.
- **Sequence** the activities of some metals.
- **Predict** if reactions will occur between certain substances.

Materials

1.0M $\text{Zn}(\text{NO}_3)_2$	Al wire
1.0M $\text{Al}(\text{NO}_3)_3$	Mg ribbon
1.0M $\text{Cu}(\text{NO}_3)_2$	Zn metal strips (4)
1.0M $\text{Mg}(\text{NO}_3)_2$	emery cloth or fine sandpaper
pipettes (4)	24-well microscale reaction plate
wire cutters	
Cu wire	

Safety Precautions



- Always wear safety goggles and a lab apron.
- Use caution when using sharp and coarse equipment.

Pre-Lab

1. Read the entire CHEMLAB.
2. Make notes about procedures and safety precautions to use in the laboratory.
3. Prepare your data table.

Reactions Between Solutions and Metals

	$\text{Al}(\text{NO}_3)_3$	$\text{Mg}(\text{NO}_3)_2$	$\text{Zn}(\text{NO}_3)_2$	$\text{Cu}(\text{NO}_3)_2$
Al				
Mg				
Zn				
Cu				

4. Form a hypothesis about what reactions will occur.
5. What are the independent and dependent variables?
6. What gas is produced when magnesium and hydrochloric acid react? Write the chemical equation for the reaction.
7. Why is it important to clean the magnesium ribbon? How might not polishing a piece of metal affect the reaction involving that metal?

Procedure

1. Use a pipette to fill each of the four wells in column 1 of the reaction plate with 2 mL of 1.0M $\text{Al}(\text{NO}_3)_3$ solution.
2. Repeat the procedure in step 1 to fill the four wells in column 2 with 2 mL of 1.0M $\text{Mg}(\text{NO}_3)_2$ solution.
3. Repeat the procedure in step 1 to fill the four wells in column 3 with 2 mL of 1.0M $\text{Zn}(\text{NO}_3)_2$ solution.
4. Repeat the procedure in step 1 to fill the four wells in column 4 with 2 mL of 1.0M $\text{Cu}(\text{NO}_3)_2$ solution.
5. With the emery paper or sandpaper, polish 10 cm of aluminum wire until it is shiny. Use wire cutters to cut the lead wire into four 2.5-cm pieces. Place a piece of the aluminum wire in each row A well that contains solution.
6. Repeat the procedure in step 5 using 10 cm of magnesium ribbon. Place a piece of the Mg ribbon in each row B well that contains solution.
7. Use the emery paper or sandpaper to polish small strips of zinc metal. Place a piece of Zn metal in each row C well that contains solution.



- Repeat the procedure in step 5 using 10 cm of copper wire. Place a piece of Cu wire in each row D well that contains solution.
- Observe what happens in each cell. After five minutes, record your observations on the data table you made.

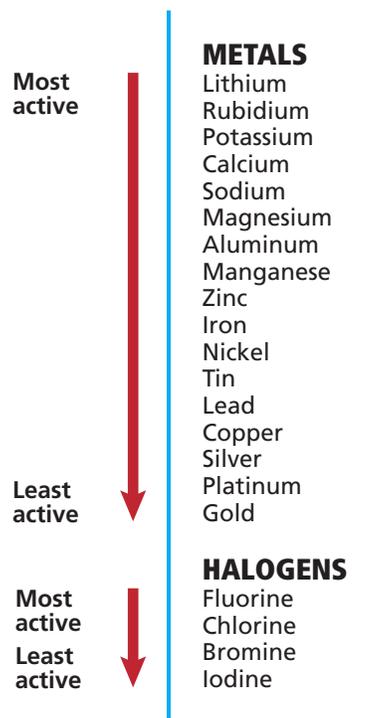
Cleanup and Disposal

- Dispose of all chemicals and solutions as directed by your teacher.
- Clean your equipment and return it to its proper place.
- Wash your hands thoroughly before you leave the lab.

Analyze and Conclude

- Observing and Inferring** In which wells of the reaction plate did chemical reactions occur? Which metal reacted with the most solutions? Which metal reacted with the fewest solutions? Which metal is the most reactive?
- Sequencing** The most active metal reacted with the most solutions. The least active metal reacted with the fewest solutions. Order the four metals from the most active to the least active.
- Comparing and Contrasting** Compare your activity series with the activity series shown here. How does the order you determined for the four metals you tested compare with the order of these metals?

- Applying Concepts** Write a chemical equation for each single-replacement reaction that occurred on your reaction plate.
- Predicting** Use the diagram below to predict if a single-replacement reaction will occur between the following reactants. Write a chemical equation for each reaction that will occur.
 - Ca and $\text{Sn}(\text{NO}_3)_2$
 - Ag and $\text{Ni}(\text{NO}_3)_2$
 - Cu and $\text{Pb}(\text{NO}_3)_3$
- Error Analysis** If the activity series you sequenced does not agree with the order in the diagram below, propose a reason for the disagreement.



Real-World Chemistry

- Under what circumstances might it be important to know the activity tendencies of a series of elements?
- Describe some of the environmental impacts of nitrates.

How It Works

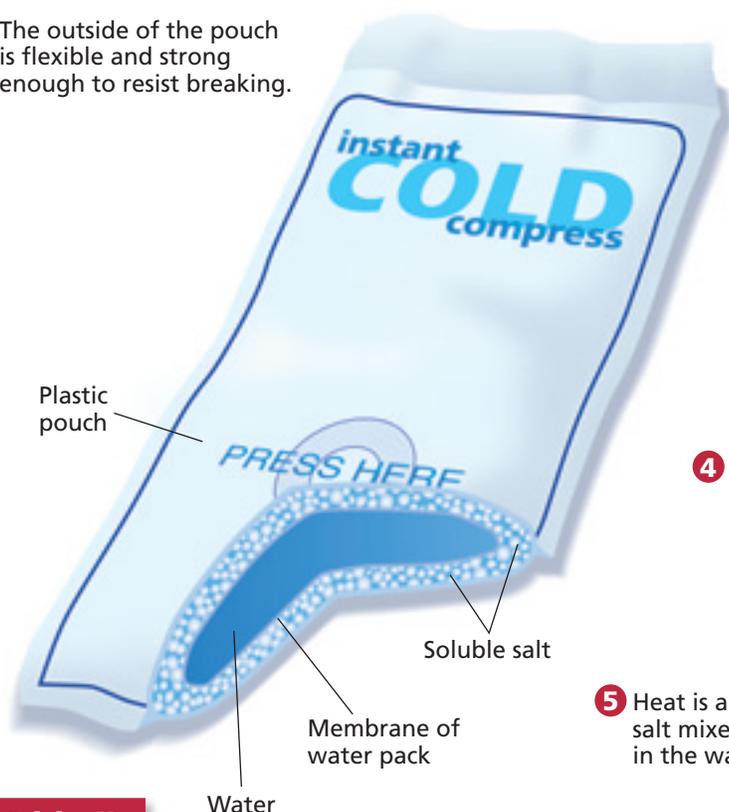
Hot and Cold Packs

Athletes know that the application of heat or cold to a strain or sprain usually relieves the pain and may lessen the severity of the injury. Instant hot and cold packs allow you to quickly and easily apply the appropriate remedy to the injury.

Hot and cold packs create aqueous solutions of a soluble salt. A salt such as ammonium nitrate is used in the cold pack and heat is absorbed as the salt dissolves in the water. Hot packs release heat when a salt such as calcium chloride dissolves in the water.



- 1 The outside of the pouch is flexible and strong enough to resist breaking.



- 2 The water is stored in an inner chamber made from thin, easily broken plastic membrane.

- 3 Salt is stored between the outer covering and the water container.

- 4 The water is released from the inner chamber when you squeeze or strike the pack.

- 5 Heat is absorbed as the salt mixes with and dissolves in the water.

Thinking Critically

1. Predicting An aqueous solution of sodium thiosulfate releases heat when it crystallizes. What would happen when sodium thiosulfate crystals dissolve in water?

2. Hypothesizing One type of heat pack contains fine iron particles. These packs are kept in a sealed container and release heat when they are exposed to air. How does this type of pack work?

Summary

10.1 Reactions and Equations

- Some chemical reactions release energy in the form of heat and light, and some absorb energy.
- Changes in temperature, color, odor, and physical state are all types of evidence that indicate a chemical reaction has occurred.
- Word and skeleton equations provide important information about a chemical reaction, such as the reactants and products involved in the reaction and their physical states.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction. Chemical equations are balanced.
- Balancing an equation involves adjusting the coefficients of the chemical formulas in the skeleton equation until the number of atoms of each element is equal on both sides of the equation.

10.2 Classifying Chemical Reactions

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Synthesis, combustion, decomposition, single-replacement, and double-replacement reactions are five classes of chemical reactions.
- A synthesis reaction occurs when two substances react to yield a single product. The substances that react can be two elements, a compound and an element, or two compounds.
- A combustion reaction occurs when a substance reacts with oxygen, producing heat and light.
- A decomposition reaction occurs when a single compound breaks down into two or more elements or new compounds.

- A single-replacement reaction occurs when the atoms of one element replace the atoms of another element in a compound.
- In single-replacement reactions, a metal may replace hydrogen in water, a metal may replace another metal in a compound dissolved in water, and a non-metal may replace another nonmetal in a compound.
- Metals and halogens can be ordered according to their reactivities. These listings, which are called activity series, can be used to predict if single-replacement reactions will occur.
- A double-replacement reaction involves the exchange of positive ions between two compounds.

10.3 Reactions in Aqueous Solutions

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When most ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions may react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.
- Three types of products produced during reactions in aqueous solutions are precipitates, water, and gases.
- An ionic equation shows the details of reactions in aqueous solutions. A complete ionic equation shows all the particles in a solution as they exist. A net ionic equation includes only the particles that participate in a reaction in a solution.

Vocabulary

- aqueous solution (p. 292)
- chemical equation (p. 280)
- chemical reaction (p. 277)
- coefficient (p. 280)
- combustion reaction (p. 285)
- complete ionic equation (p.293)
- decomposition reaction (p. 286)
- double-replacement reaction (p. 290)
- net ionic equation (p. 293)
- precipitate (p. 290)
- product (p. 278)
- reactant (p. 278)
- single-replacement reaction (p. 287)
- solute (p. 292)
- solvent (p. 292)
- spectator ion (p. 293)
- synthesis reaction (p. 284)

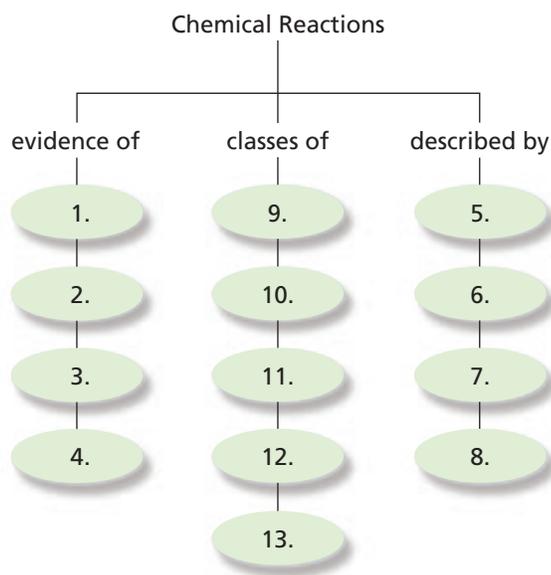


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 10 Assessment.

Concept Mapping

52. Use the following terms and phrases to complete the concept map: synthesis, net ionic equation, change in energy, change in physical state, single-replacement, word equation, decomposition, complete ionic equation, double-replacement, combustion, change in odor, chemical equation, change in color.



Mastering Concepts

53. Explain the difference between reactants and products. (10.1)
54. What do the arrows and coefficients used by chemists in equations communicate? (10.1)
55. Write formulas for the following substances and designate their physical states. (10.1)
- nitrogen dioxide gas
 - liquid gallium
 - barium chloride dissolved in water
 - solid ammonium carbonate
56. Identify the reactants in the following reaction: When potassium is dropped into aqueous zinc nitrate, zinc and aqueous potassium nitrate form. (10.1)

57. When gasoline is burned in an automobile engine, what evidence indicates that a chemical change has occurred? (10.1)
58. Write the word equation for this skeleton equation. (10.1)
- $$\text{Mg(s)} + \text{FeCl}_3\text{(aq)} \rightarrow \text{Fe(s)} + \text{MgCl}_2\text{(aq)}$$
59. Balance the equation in question 58. (10.1)
60. What are five classes of chemical reactions? (10.2)
61. How would you classify a chemical reaction between two reactants that produces one product? (10.2)
62. Explain the difference between a single-replacement reaction and a double-replacement reaction. (10.2)
63. Under what conditions does a precipitate form in a chemical reaction? (10.2)
64. Classify the chemical reaction in question 58. (10.2)
65. In each of the following pairs, which element will replace the other in a reaction? (10.2)
- tin and sodium
 - fluorine and iodine
 - lead and silver
 - copper and nickel
66. When reactions occur in aqueous solutions what common types of products are produced? (10.3)
67. Compare and contrast chemical equations and ionic equations. (10.3)
68. What is a net ionic equation? How does it differ from a complete ionic equation? (10.3)
69. Define spectator ion. (10.3)
70. Write the net ionic equation for a chemical reaction that occurs in an aqueous solution and produces water. (10.3)

Mastering Problems

Balancing Chemical Equations (10.1)

71. Write skeleton equations for these reactions.
- hydrogen iodide(g) \rightarrow hydrogen(g) + iodine(g)
 - aluminum(s) + iodine(s) \rightarrow aluminum iodide(s)
 - iron(II) oxide(s) + oxygen(g) \rightarrow iron(III) oxide(s)
72. Write skeleton equations for these reactions.
- butane (C₄H₁₀)(l) + oxygen(g) \rightarrow carbon dioxide(g) + water(l)
 - aluminum carbonate(s) \rightarrow aluminum oxide(s) + carbon dioxide(g)
 - silver nitrate(aq) + sodium sulfide(aq) \rightarrow silver sulfide(s) + sodium nitrate(aq)

- 73.** Write skeleton equations for these reactions.
- iron(s) + fluorine(g) \rightarrow iron(III) fluoride(s)
 - sulfur trioxide(g) + water(l) \rightarrow sulfuric acid(aq)
 - sodium(s) + magnesium iodide(aq) \rightarrow sodium iodide(aq) + magnesium(s)
 - vanadium(s) + oxygen(g) \rightarrow vanadium(V) oxide(s)
- 74.** Write skeleton equations for these reactions.
- lithium(s) + gold(III) chloride(aq) \rightarrow lithium chloride(aq) + gold(s)
 - iron(s) + tin(IV) nitrate(aq) \rightarrow iron(III) nitrate(aq) + tin(s)
 - nickel(II) chloride(s) + oxygen(g) \rightarrow nickel(II) oxide(s) + dichlorine pentoxide(g)
 - lithium chromate(aq) + barium chloride(aq) \rightarrow lithium chloride(aq) + barium chromate(s)
- 75.** Balance the skeleton equations for the reactions described in question 71.
- 76.** Balance the skeleton equations for the reactions described in question 72.
- 77.** Balance the skeleton equations for the reactions described in question 73.
- 78.** Balance the skeleton equations for the reactions described in question 74.
- 79.** Write chemical equations for these reactions.
- When solid naphthalene ($C_{10}H_8$) burns in air, the products are gaseous carbon dioxide and liquid water.
 - Bubbling hydrogen sulfide gas through manganese(II) chloride dissolved in water results in the formation of the precipitate manganese(II) sulfide and hydrochloric acid.
 - Solid magnesium reacts with nitrogen gas to produce solid magnesium nitride.
 - Heating oxygen difluoride gas yields oxygen gas and fluorine gas.
- boron + fluorine \rightarrow
 - germanium + sulfur \rightarrow
 - zirconium + nitrogen \rightarrow
 - tetraphosphorus decoxide + water \rightarrow phosphoric acid
- 86.** Write a chemical equation for the combustion of each of the following substances. If a compound contains the elements carbon and hydrogen, assume that carbon dioxide gas and liquid water are produced.
- solid barium
 - solid boron
 - liquid acetone (C_3H_6O)
 - liquid octane (C_8H_{18})
- 87.** Write chemical equations for each of the following decomposition reactions. One or more products may be identified.
- magnesium bromide \rightarrow
 - cobalt(II) oxide \rightarrow
 - titanium(IV) hydroxide \rightarrow titanium(IV) oxide + water
 - barium carbonate \rightarrow barium oxide + carbon dioxide
- 88.** Write chemical equations for the following single-replacement reactions that may occur in water. If no reaction occurs, write NR in place of the products.
- nickel + magnesium chloride \rightarrow
 - calcium + copper(II) bromide \rightarrow
 - potassium + aluminum nitrate \rightarrow
 - magnesium + silver nitrate \rightarrow
- 89.** Write chemical equations for each of the following double-replacement reactions that occur in water.
- rubidium iodide + silver nitrate \rightarrow
 - sodium phosphate + manganese(II) chloride \rightarrow
 - lithium carbonate + molybdenum(VI) bromide \rightarrow
 - calcium nitrate + aluminum hydroxide \rightarrow

Classifying Chemical Reactions (10.2)

- 80.** Classify each of the reactions represented by the chemical equations in question 75.
- 81.** Classify each of the reactions represented by the chemical equations in question 76.
- 82.** Classify each of the reactions represented by the chemical equations in question 77.
- 83.** Classify each of the reactions represented by the chemical equations in question 78.
- 84.** Classify each of the reactions represented by the chemical equations in question 79.
- 85.** Write chemical equations for each of the following synthesis reactions.

Reactions in Aqueous Solutions (10.3)

- 90.** Write complete ionic and net ionic equations for each of the following reactions.
- $K_2S(aq) + CoCl_2(aq) \rightarrow 2KCl(aq) + CoS(s)$
 - $H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaSO_4(s)$
 - $2HClO(aq) + Ca(OH)_2(aq) \rightarrow 2H_2O(l) + Ca(ClO)_2(aq)$
- 91.** A reaction occurs when hydrosulfuric acid (H_2S) is mixed with an aqueous solution of iron(III) bromide. Solid iron(III) sulfide is produced. Write the chemical and net ionic equations for the reaction.

92. Write complete ionic and net ionic equations for each of the following reactions.

- $\text{H}_3\text{PO}_4(\text{aq}) + 3\text{RbOH}(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \text{Rb}_3\text{PO}_4(\text{aq})$
- $\text{HCl}(\text{aq}) + \text{NH}_4\text{OH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{NH}_4\text{Cl}(\text{aq})$
- $2\text{HI} + (\text{NH}_4)_2\text{S}(\text{aq}) \rightarrow \text{H}_2\text{S}(\text{g}) + 2\text{NH}_4\text{I}(\text{aq})$
- $\text{HNO}_3(\text{aq}) + \text{KCN}(\text{aq}) \rightarrow \text{HCN}(\text{g}) + \text{KNO}_3(\text{aq})$

93. A reaction occurs when sulfurous acid (H_2SO_3) is mixed with an aqueous solution of sodium hydroxide. Aqueous sodium sulfite is produced. Write the chemical and net ionic equations for the reaction.

94. A reaction occurs when nitric acid (HNO_3) is mixed with an aqueous solution of potassium hydrogen carbonate. Aqueous potassium nitrate is produced. Write the chemical and net ionic equations for the reaction.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- Identify the products in the following reaction that occurs in plants: Carbon dioxide and water react to produce glucose and oxygen.
- How will aqueous solutions of sucrose and hydrogen chloride differ?
- Write the word equation for each of these skeleton equations. C_6H_6 is the formula for benzene.
 - $\text{C}_6\text{H}_6(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
- Write skeleton equations for the following reactions.
 - ammonium phosphate(aq) + chromium(III) bromide(aq) \rightarrow ammonium bromide(aq) + chromium(III) phosphate(s)
 - chromium(VI) hydroxide(s) \rightarrow chromium(VI) oxide(s) + water(l)
 - aluminum(s) + copper(I) chloride(aq) \rightarrow aluminum chloride(aq) + copper(s)
 - potassium iodide(aq) + mercury(I) nitrate(aq) \rightarrow potassium nitrate(aq) + mercury(I) iodide(s)
- Balance the skeleton equations for the reactions described in question 98.
- Classify each of the reactions represented by the chemical equations in question 99.
- Predict whether each of the following reactions will occur in aqueous solutions. If you predict that a reaction will not occur, explain your reasoning. Note: Barium sulfate and silver bromide precipitate in aqueous solutions.

- sodium hydroxide + ammonium sulfate \rightarrow
- niobium(V) sulfate + barium nitrate \rightarrow
- strontium bromide + silver nitrate \rightarrow

Thinking Critically

102. Predicting A piece of aluminum metal is placed in aqueous KCl. Another piece of aluminum is placed in an aqueous AgNO_3 solution. Explain why a chemical reaction does or does not occur in each instance.

103. Designing an Experiment You suspect that the water in a lake close to your school may contain lead in the form of $\text{Pb}^{2+}(\text{aq})$ ions. Formulate your suspicion as a hypothesis and design an experiment to test your theory. Write the net ionic equations for the reactions of your experiment. (Hint: In aqueous solution, Pb^{2+} forms compounds that are solids with Cl^- , Br^- , I^- , and SO_4^{2-} ions.)

104. Applying Concepts Write the chemical equations and net ionic equations for each of the following reactions that may occur in aqueous solutions. If a reaction does not occur, write NR in place of the products. Magnesium phosphate precipitates in an aqueous solution.

- $\text{KNO}_3 + \text{CsCl} \rightarrow$
- $\text{Ca}(\text{OH})_2 + \text{KCN} \rightarrow$
- $\text{Li}_3\text{PO}_4 + \text{MgSO}_4 \rightarrow$
- $\text{HBrO} + \text{NaOH} \rightarrow$

Writing in Chemistry

- Prepare a poster describing types of chemical reactions that occur in the kitchen.
- Write a report in which you compare and contrast chemical and mathematical equations.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- Distinguish among a mixture, a solution, and a compound. (Chapter 3)
- Write the formula for the compounds made from each of the following pairs of ions. (Chapter 9)
 - copper(I) and sulfite
 - tin(IV) and fluoride
 - gold(III) and cyanide
 - lead(II) and sulfide

Use these questions and the test-taking tip to prepare for your standardized test.

- Potassium chromate and lead(II) acetate are both dissolved in a beaker of water, where they react to form solid lead(II) chromate. What is the balanced net ionic equation describing this reaction?
 - $\text{Pb}^{2+}(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2(\text{s})$
 - $\text{Pb}^{2+}(\text{aq}) + 2\text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{Pb}(\text{CrO}_4)_2(\text{s})$
 - $\text{Pb}^{2+}(\text{aq}) + \text{CrO}_4^{2-}(\text{aq}) \rightarrow \text{PbCrO}_4(\text{s})$
 - $\text{Pb}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{PbC}_2\text{H}_3\text{O}_2(\text{s})$
- What type of reaction is described by the following equation?

$$\text{Cs}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CsOH}(\text{aq}) + \text{H}_2(\text{g})$$
 - synthesis
 - combustion
 - decomposition
 - replacement
- Which of the following reactions between halogens and halide salts will occur?
 - $\text{F}_2(\text{g}) + \text{FeI}_2(\text{aq}) \rightarrow \text{FeF}_2(\text{aq}) + \text{I}_2(\text{l})$
 - $\text{I}_2(\text{s}) + \text{MnBr}_2(\text{aq}) \rightarrow \text{MnI}_2(\text{aq}) + \text{Br}_2(\text{g})$
 - $\text{Cl}_2(\text{s}) + \text{SrF}_2(\text{aq}) \rightarrow \text{SrCl}_2(\text{aq}) + \text{F}_2(\text{g})$
 - $\text{Br}_2(\text{l}) + \text{CoCl}_2(\text{aq}) \rightarrow \text{CoBr}_2(\text{aq}) + \text{Cl}_2(\text{g})$

Interpreting Tables Use the table to answer questions 4–6.

Physical Properties of Select Ionic Compounds

Compound	Name	Physical state at room temp.	Soluble in water?	Melting point (°C)
NaClO_3	sodium chlorate	solid	yes	248
Na_2SO_4	sodium sulfate	solid	yes	884
NiCl_2	nickel(II) chloride	solid	yes	1009
$\text{Ni}(\text{OH})_2$	nickel(II) hydroxide	solid	no	230
AgNO_3	silver nitrate	solid	yes	212

- An aqueous solution of nickel(II) sulfate is mixed with aqueous sodium hydroxide. Will a visible reaction occur?
 - No, solid nickel(II) hydroxide is soluble in water.
 - No, solid sodium sulfate is soluble in water.

- Yes, solid sodium sulfate will precipitate out of solution.
 - Yes, solid nickel(II) hydroxide will precipitate out of solution.
- When $\text{AgClO}_3(\text{aq})$ and $\text{NaNO}_3(\text{aq})$ are mixed, _____.
 - no visible reaction occurs
 - solid NaClO_3 precipitates out of solution
 - NO_2 gas is released from the reaction
 - solid Ag metal is produced
 - Finely ground nickel(II) hydroxide is placed in a beaker of water. It sinks to the bottom of the beaker and remains unchanged. An aqueous solution of hydrochloric acid (HCl) is then added the beaker, and the $\text{Ni}(\text{OH})_2$ disappears. Which of the following equations best describes what occurred in the beaker?
 - $\text{Ni}(\text{OH})_2(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{NiO}(\text{aq}) + \text{H}_2(\text{g}) + \text{HCl}(\text{aq})$
 - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{NiCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
 - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NiCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
 - $\text{Ni}(\text{OH})_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NiCl}_2(\text{aq}) + 3\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
 - The combustion of ethanol, $\text{C}_2\text{H}_6\text{O}$, produces carbon dioxide and water vapor. The equation that best describes this process is _____.
 - $\text{C}_2\text{H}_6\text{O}(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
 - $\text{C}_2\text{H}_6\text{O}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{g})$
 - $\text{C}_2\text{H}_6\text{O}(\text{l}) \rightarrow 3\text{O}_2(\text{l}) + 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
 - What is the product of this synthesis reaction?

$$\text{Cl}_2(\text{g}) + 2\text{NO}(\text{g}) \rightarrow ?$$
 - NCl_2
 - 2NOCl
 - N_2O_2
 - 2ClO



TEST-TAKING TIP

Tables If a test question involves a table, skim the table before reading the question. Read the title, column heads, and row heads. Then read the question and interpret the information in the table.

The Mole

What You'll Learn

- ▶ You will use the mole and molar mass to make conversions among moles, mass, and number of representative particles.
- ▶ You will determine the percent composition of the components of compounds.
- ▶ You will calculate the empirical and molecular formulas for compounds and determine the formulas for hydrates.

Why It's Important

New materials, new products, new consumer goods of all kinds come on the market regularly. But before manufacturing begins on most new products, calculations involving the mole must be done.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links to the mole and molar calculations.

Florists often sell flowers, such as roses, carnations, and tulips, by the dozen. A dozen is a counting unit for 12 items.



DISCOVERY LAB



Materials

centimeter ruler
paper clip

How much is a mole?

Counting large numbers of items is easier when you use counting units like the dozen. Chemists use a counting unit called the mole.

Procedure

1. Measure the length of a paper clip to the nearest 0.1 cm.
2. If a mole is 6.02×10^{23} items, how far will a mole of paper clips, placed end to end lengthwise, reach into space?

Analysis

How many light-years (ly) would the paper clips extend into space? (1 light-year = 9.46×10^{15} m). How does the distance you calculated compare with the following astronomical distances: nearest star (other than the sun) = 4.3 ly, center of our galaxy = 30 000 ly, nearest galaxy = 2×10^6 ly?

Section

11.1

Measuring Matter

Objectives

- **Describe** how a mole is used in chemistry.
- **Relate** a mole to common counting units.
- **Convert** moles to number of representative particles and number of representative particles to moles.

Vocabulary

mole
Avogadro's number

If you were buying a bouquet of roses for a special occasion, you probably wouldn't ask for 12 or 24; you'd ask for one or two dozen. Similarly, you might buy a pair of gloves, a ream of paper for your printer, or a gross of pencils. Each of the units shown in **Figure 11-1**—a pair, a dozen, a gross, and a ream—represents a specific number of items. These units make counting objects easier. It's easier to buy and sell paper by the ream—500 sheets—than by the individual sheet.

Counting Particles

Each of the counting units shown in **Figure 11-1** is appropriate for certain kinds of objects depending primarily on their size and the use they serve. But regardless of the object—boots, eggs, pencils, paper—the number that the unit represents is always constant.



Figure 11-1

A pair is always two objects, a dozen is 12, a gross is 144, and a ream is 500. Can you think of any other counting units?



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Chemists also need a convenient method for counting accurately the number of atoms, molecules, or formula units in a sample of a substance. As you know, atoms and molecules are extremely small. There are so many of them in even the smallest sample that it's impossible to actually count them. That's why chemists created their own counting unit called the mole. In the **DISCOVERY LAB**, you found that a mole of paper clips is an enormous number of items.

What is a mole? The **mole**, commonly abbreviated mol, is the SI base unit used to measure the amount of a substance. It is the number of representative particles, carbon atoms, in exactly 12 g of pure carbon-12. Through years of experimentation, it has been established that a mole of anything contains $6.022\,136\,7 \times 10^{23}$ representative particles. A representative particle is any kind of particle such as atoms, molecules, formula units, electrons, or ions. The number $6.022\,136\,7 \times 10^{23}$ is called **Avogadro's number** in honor of the Italian physicist and lawyer Amedeo Avogadro who, in 1811, determined the volume of one mole of a gas. In this book, Avogadro's number will be rounded to three significant figures— 6.02×10^{23} .

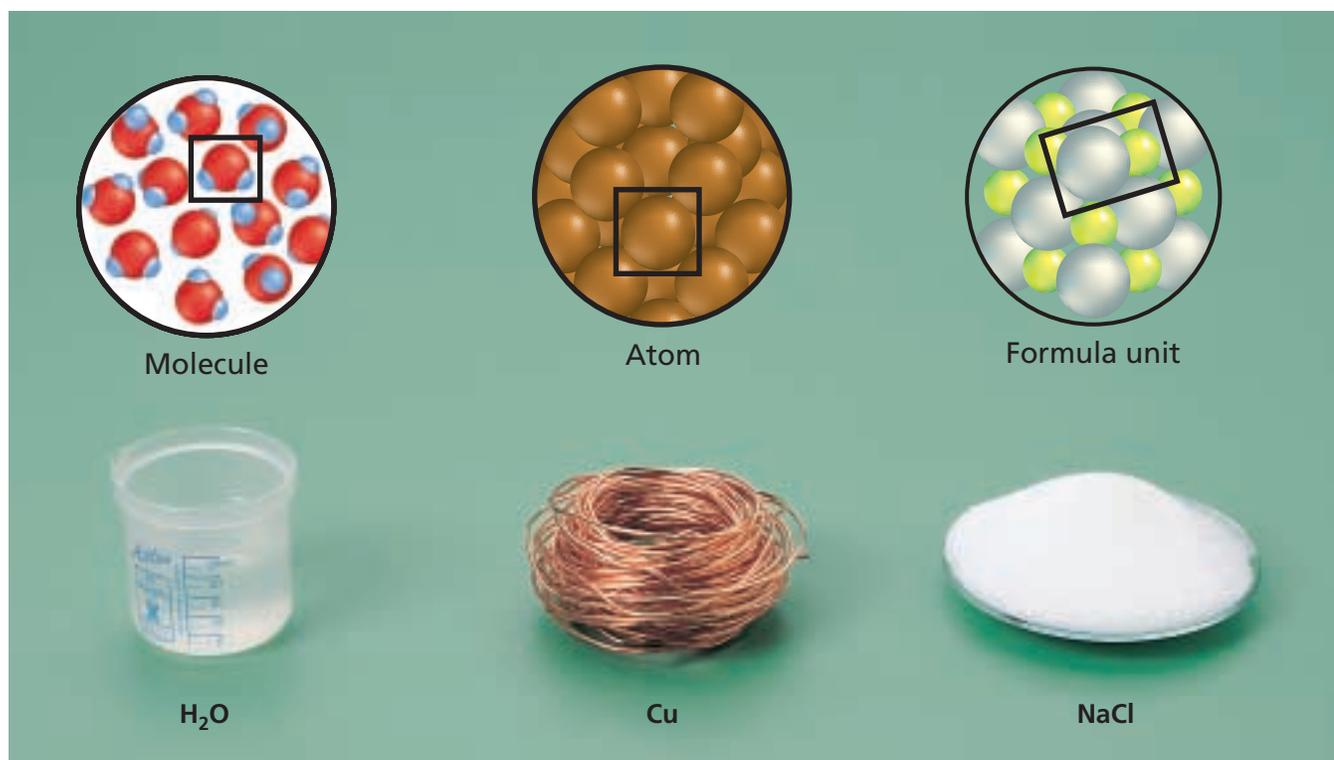
If you write out Avogadro's number, it looks like this.

602 000 000 000 000 000 000 000

Avogadro's number is an enormous number, as it must be in order to count extremely small particles. As you can imagine, Avogadro's number would not be convenient for measuring a quantity of marbles. Avogadro's number of marbles would cover the surface of Earth to a depth of more than six kilometers! But you can see in **Figure 11-2** that it is convenient to use the mole to measure substances. One-mole quantities of three substances are shown, each with a different representative particle. The representative particle in a mole of water is the water molecule, the representative particle in a mole of copper is the copper atom, and the representative particle in a mole of sodium chloride is the formula unit.

Figure 11-2

The amount of each substance shown is 6.02×10^{23} or one mole of representative particles. The representative particle for each substance is shown in a box. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



Converting Moles to Particles and Particles to Moles

Suppose you buy three and a half dozen roses and want to know how many roses you have. Recall what you have learned about conversion factors. You can multiply the known quantity (3.5 dozen roses) by a conversion factor to express the quantity in the units you want (number of roses). You must set up your calculation as shown here so that all units cancel except those required for the answer.

$$\text{Conversion factor: } \frac{12 \text{ roses}}{1 \text{ dozen}}$$
$$3.5 \cancel{\text{ dozen}} \times \frac{12 \text{ roses}}{1 \cancel{\text{ dozen}}} = 42 \text{ roses}$$

Note that the units cancel and the answer tells you that 42 roses are in 3.5 dozen.

Now, suppose you want to determine how many particles of sucrose are in 3.50 moles of sucrose. You know that one mole contains 6.02×10^{23} representative particles. Therefore, you can write a conversion factor, Avogadro's number, that relates representative particles to moles of a substance.

$$\text{Conversion factor: } \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mole}}$$

You can find the number of representative particles in a number of moles just as you found the number of roses in 3.5 dozen.

$$\text{number of moles} \times \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mole}}$$
$$= \text{number of representative particles}$$

For sucrose, the representative particle is a molecule, so the number of molecules of sucrose is obtained by multiplying 3.50 moles of sucrose by the conversion factor, Avogadro's number.

$$3.50 \cancel{\text{ mol sucrose}} \times \frac{6.02 \times 10^{23} \text{ molecules sucrose}}{1 \cancel{\text{ mol sucrose}}}$$
$$= 2.11 \times 10^{24} \text{ molecules sucrose}$$

There are 2.11×10^{24} molecules of sucrose in 3.50 moles.

PRACTICE PROBLEMS

1. Determine the number of atoms in 2.50 mol Zn.
2. Given 3.25 mol AgNO_3 , determine the number of formula units.
3. Calculate the number of molecules in 11.5 mol H_2O .

Now, suppose you want to find out how many moles are represented by a certain number of representative particles. You can use the inverse of Avogadro's number as a conversion factor.

$$\text{number of representative particles} \times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ representative particles}}$$
$$= \text{number of moles}$$

History

CONNECTION

Lorenzo Romano Amedeo Carlo Avogadro, Conte di Quaregna e Ceretto was born in Turin, Italy in 1776 and was educated as a church lawyer. During the early 1800s, he studied mathematics and physics and was appointed to a professorship at the Royal College of Vercelli where he produced his hypothesis on gases. From 1820 until his death, Avogadro was professor of physics at the University of Turin where he conducted research on electricity and the physical properties of liquids.

Avogadro's hypothesis did not receive recognition for more than fifty years. Although Avogadro did nothing to measure the number of particles in equal volumes of gases, his hypothesis did lead to the eventual calculation of the number, 6.02×10^{23} .

National Mole Day is celebrated on October 23 (10/23) from 6:02 A.M. to 6:02 P.M. to commemorate Avogadro's contribution to modern chemistry.



For more practice converting from moles to representative particles, go to

Supplemental Practice Problems in Appendix A.

The number of moles of substance is obtained by multiplying the number of particles by this factor, as you will see in Example Problem 11-1.

EXAMPLE PROBLEM 11-1



Ointments containing zinc oxide provide protection from sunburn and are used to treat some skin diseases.

Converting Number of Representative Particles to Moles

Zinc is used as a corrosion-resistant coating on iron and steel. It is also an essential trace element in your diet. Calculate the number of moles that contain 4.50×10^{24} atoms of zinc (Zn).

1. Analyze the Problem

You are given the number of atoms of zinc and must find the equivalent number of moles. If you compare 4.50×10^{24} atoms Zn with 6.02×10^{23} , the number of atoms in one mole, you can predict that the answer should be less than 10 moles.

Known

number of atoms = 4.50×10^{24} atoms Zn

1 mol Zn = 6.02×10^{23} atoms Zn

Unknown

mol Zn = ? mol

2. Solve for the Unknown

Multiply the number of zinc atoms by the conversion factor that is the inverse of Avogadro's number.

$$\text{number of atoms} \times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ atoms}} = \text{number of moles}$$

$$4.50 \times 10^{24} \text{ atoms Zn} \times \frac{1 \text{ mol Zn}}{6.02 \times 10^{23} \text{ atoms Zn}} = 7.48 \text{ mol Zn}$$

3. Evaluate the Answer

The number of atoms of zinc and Avogadro's number have three significant figures. Therefore, the answer is expressed correctly with three digits. The answer is less than 10 moles, as predicted, and has the correct unit.

Practice!

For more practice converting from representative particles to moles, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

4. How many moles contain each of the following?

a. 5.75×10^{24} atoms Al

c. 3.58×10^{23} formula units ZnCl_2

b. 3.75×10^{24} molecules CO_2

d. 2.50×10^{20} atoms Fe

Section 11.1 Assessment

- How is a mole similar to a dozen?
- What is the relationship between Avogadro's number and one mole?
- Explain how you can convert from the number of representative particles of a substance to moles of that substance.
- Explain why chemists use the mole.
- Thinking Critically** Arrange the following from

the smallest number of representative particles to the largest number of representative particles: 1.25×10^{25} atoms Zn; 3.56 mol Fe; 6.78×10^{22} molecules glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).

- Using Numbers** Determine the number of representative particles in each of the following and identify the representative particle: 11.5 mol Ag; 18.0 mol H_2O ; 0.150 mol NaCl.

You wouldn't expect a dozen limes to have the same mass as a dozen eggs. Eggs and limes differ in size and composition, so it's not surprising that they have different masses, as **Figure 11-3** shows. Moles of substances also have different masses for the same reason—the substances have different compositions. If you put a mole of carbon on a balance beside a mole of metallic copper, you would see a difference in mass just as you do for a dozen eggs and a dozen limes. Carbon atoms differ from copper atoms. Thus, the mass of 6.02×10^{23} atoms of carbon does not equal the mass of 6.02×10^{23} atoms of copper. How do you determine the mass of a mole?



The Mass of a Mole

In Chapter 4, you learned that the relative scale of atomic masses uses the isotope carbon-12 as the standard. Each atom of carbon-12 has a mass of 12 atomic mass units (amu). The atomic masses of all other elements are established relative to carbon-12. For example, an atom of hydrogen-1 has a mass of 1 amu. The mass of an atom of helium-4 is 4 amu. Therefore, the mass of one atom of hydrogen-1 is one-twelfth the mass of one atom of carbon-12. The mass of one atom of helium-4 is one-third the mass of one atom of carbon-12.

You can find atomic masses on the periodic table, but notice that the values shown are not exact integers. For example, you'll find 12.011 amu for carbon, 1.008 amu for hydrogen, and 4.003 amu for helium. These differences occur because the recorded values are weighted averages of the masses of all the naturally occurring isotopes of each element.

How does the mass of one atom relate to the mass of a mole of that atom? You know that the mole is defined as the number of representative particles, or carbon-12 atoms, in exactly 12 g of pure carbon-12. Thus, the mass of one mole of carbon-12 atoms is 12 g. What about other elements? Whether you are considering a single atom or Avogadro's number of atoms (a mole), the masses of all atoms are established relative to the mass of carbon-12. The mass of a mole of hydrogen-1 is one-twelfth the mass of a mole of carbon-12 atoms, or 1.0 g. The mass of a mole of helium-4 atoms is one-third the mass of a mole of carbon-12 atoms, or 4.0 g. The mass in grams of one mole of any pure substance is called its **molar mass**. The molar mass of any element is numerically equal to its atomic mass and has the units g/mol. An atom of manganese has an atomic mass of 54.94 amu. Therefore, the molar mass of manganese is 54.94 g/mol. When you measure 54.94 g of manganese on a balance,

Objectives

- **Relate** the mass of an atom to the mass of a mole of atoms.
- **Calculate** the number of moles in a given mass of an element and the mass of a given number of moles of an element.
- **Calculate** the number of moles of an element when given the number of atoms of the element.
- **Calculate** the number of atoms of an element when given the number of moles of the element.

Vocabulary

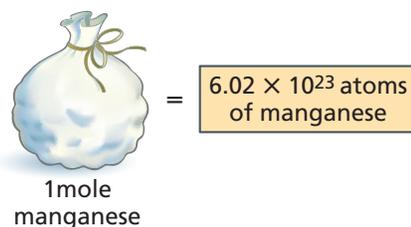
molar mass

Figure 11-3

A dozen limes has approximately twice the mass of a dozen eggs. The difference in mass is reasonable because limes are different from eggs in composition and size.

Figure 11-4

One mole of manganese, represented by a bag of particles, contains Avogadro's number of atoms and has a mass equal to its atomic mass in grams. The same is true for all the elements.



you indirectly count 6.02×10^{23} atoms of manganese. **Figure 11-4** shows the relationship between molar mass and one mole of an element. The **problem-solving LAB** will further clarify these relationships.

Using Molar Mass

Imagine that your class bought jellybeans in bulk to sell by the dozen at a candy sale. You soon realize that it's too much work counting out each dozen, so instead you decide to measure the jellybeans by mass. You find that 1 dozen jellybeans has a mass of 35 g. What mass of jellybeans should you measure if a customer wants 5 dozen? The conversion factor that relates mass and dozens of jellybeans is

$$\frac{35 \text{ g jellybeans}}{1 \text{ dozen}}$$

problem-solving LAB

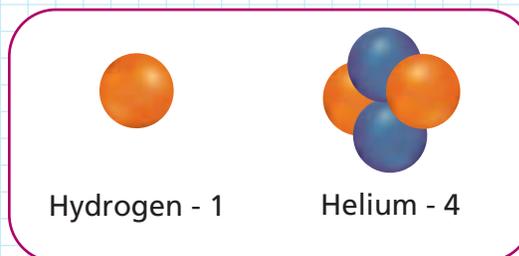
Molar Mass, Avogadro's Number and the Atomic Nucleus

Formulating models A nuclear model of mass can provide a simple picture of the connections between the mole, molar mass, and the number of representative particles in a mole.

Analysis

The diagram shows models of the nuclei of hydrogen-1 and helium-4. The hydrogen-1 nucleus contains one proton with a mass of 1.007 amu. The mass of the proton in grams has been found experimentally to be 1.672×10^{-24} g. Helium-4 contains two protons and two neutrons and has a mass of approximately 4 amu.

1. What is the mass in grams of one helium atom? (The mass of a neutron is approximately the same as the mass of a proton.)
2. Carbon-12 contains six protons and six neutrons. Draw a model of the nucleus of



carbon-12 and calculate the mass of one atom in amu and in grams.

Thinking Critically

1. How many atoms of hydrogen-1 are in a 1.007-g sample? Recall that 1.007 amu is the mass of one atom of hydrogen-1. Round your answer to two significant digits.
2. If you had samples of helium and carbon that contained the same number of atoms as you calculated in question 1, what would be the mass in grams of each sample?
3. What can you conclude about the relationship between the number of atoms and the mass of each sample?

You would multiply the number of dozens to be sold by this conversion factor.

$$5 \cancel{\text{dozen}} \times \frac{35 \text{ g jellybeans}}{1 \cancel{\text{dozen}}} = 175 \text{ g jellybeans}$$

Note how the units cancel to give you the mass of 5 dozen jellybeans.

Now, suppose that while working in chemistry lab, you need 3.00 moles of manganese (Mn) for a chemical reaction. How can you measure that amount? Like the 5 dozen jellybeans, the number of moles of manganese can be converted to an equivalent mass and measured on a balance. To calculate mass from the number of moles, you need to multiply the number of moles of manganese required in the reaction (3.00 moles of Mn) by a conversion factor that relates mass and moles of manganese. That conversion factor is the molar mass of manganese (54.9 g/mol).

$$\text{number of moles} \times \frac{\text{number of grams}}{1 \text{ mole}} = \text{mass}$$

$$3.00 \cancel{\text{mol Mn}} \times \frac{54.9 \text{ g Mn}}{1 \cancel{\text{mol Mn}}} = 165 \text{ g Mn}$$

If you measure 165 g of manganese on a balance, you will have the 3.00 moles of manganese you need for the reaction. The reverse conversion—from mass to moles—also involves the molar mass as a conversion factor, but it is the inverse of the molar mass that is used. Can you explain why?

EXAMPLE PROBLEM 11-2

Mole to Mass Conversion

Chromium (Cr) is a transition element used as a coating on metals and in steel alloys to control corrosion. Calculate the mass in grams of 0.0450 moles of chromium.

1. Analyze the Problem

You are given the number of moles of chromium and must convert it to an equivalent mass using the molar mass of chromium from the periodic table. Because the sample is less than one-tenth mole, the answer should be less than one-tenth the molar mass.

Known

number of moles = 0.0450 mol Cr
molar mass Cr = 52.00 g/mol Cr

Unknown

mass = ? g Cr

2. Solve for the Unknown

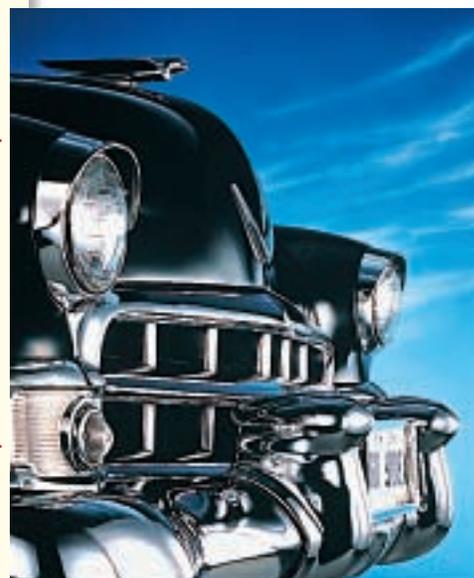
Multiply the known number of moles of chromium by the conversion factor that relates grams of chromium to moles of chromium, the molar mass.

$$\text{moles Cr} \times \frac{\text{grams Cr}}{1 \text{ mol Cr}} = \text{grams Cr}$$

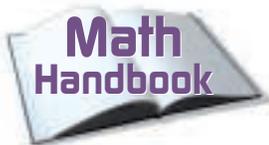
$$0.0450 \cancel{\text{mol Cr}} \times \frac{52.00 \text{ g Cr}}{1 \cancel{\text{mol Cr}}} = 2.34 \text{ g Cr}$$

3. Evaluate the Answer

The known number of moles of chromium has the smallest number of significant figures (3), so the answer is correctly stated with three digits. The answer is less than one-tenth the mass of one mole as predicted and has the correct unit.



Chromium resists corrosion, which means it doesn't react readily with oxygen in the air. It was used in this 1948 Cadillac to protect the steel and add glitter.



Review the meaning of inverse in the **Math Handbook** on page 905 of this text.

PRACTICE PROBLEMS

11. Determine the mass in grams of each of the following.
- | | |
|----------------|----------------|
| a. 3.57 mol Al | c. 3.45 mol Co |
| b. 42.6 mol Si | d. 2.45 mol Zn |

EXAMPLE PROBLEM 11-3

Mass to Mole Conversion

Calcium, the fifth most abundant element on Earth, is always found combined with other elements because of its high reactivity. How many moles of calcium are in 525 g calcium (Ca)?

1. Analyze the Problem

You are given the mass of calcium and must convert the mass to moles of calcium. The mass of calcium is more than ten times larger than the molar mass. Therefore, the answer should be greater than ten moles.

Known

mass = 525 g Ca
molar mass Ca = 40.08 g/mol Ca

Unknown

number of moles = ? mol Ca

2. Solve for the Unknown

Multiply the known amount of calcium by the conversion factor that relates moles of calcium to grams of calcium, the inverse of molar mass.

$$\text{mass} \times \frac{1 \text{ mole}}{\text{number of grams}} = \text{number of moles}$$

$$525 \text{ g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} = 13.1 \text{ mol Ca}$$

3. Evaluate the Answer

The mass of calcium has the smaller number of significant figures (3), so the answer is expressed correctly with three digits. As predicted, the answer is greater than 10 moles and has the expected unit.

PRACTICE PROBLEMS

12. Determine the number of moles in each of the following.
- | | |
|--------------|---------------|
| a. 25.5 g Ag | c. 125 g Zn |
| b. 300.0 g S | d. 1.00 kg Fe |

Conversions from mass to atoms and atoms to mass So far, you have learned how to convert mass to the number of moles and the number of moles to mass. You can go one step further and convert mass to the number of atoms. Recall the jellybeans you were selling at the candy sale. At the end of the day, you find that 550 g of jellybeans are left unsold. Without counting, can you determine how many jellybeans this is? You know that one dozen jellybeans has a mass of 35 g and that 1 dozen contains 12 jellybeans. Thus, you can first convert the 550 g to dozens of jellybeans by using the conversion factor that relates dozens and mass.

$$550 \text{ g jellybeans} \times \frac{1 \text{ dozen jellybeans}}{35 \text{ g jellybeans}} = 16 \text{ dozen jellybeans}$$



For more practice with mass and mole conversions, go to **Supplemental Practice Problems** in Appendix A.



Next, you can determine how many jellybeans are in 16 dozen by multiplying by the conversion factor that relates number of particles (jellybeans) and dozens.

$$16 \cancel{\text{dozen}} \times \frac{12 \text{ jellybeans}}{1 \cancel{\text{dozen}}} = 192 \text{ jellybeans}$$

The 550 g of leftover jellybeans is equal to 192 jellybeans.

Just as you cannot make a direct conversion from the mass of jellybeans to the number of jellybeans, you cannot make a direct conversion from the mass of a substance to the number of representative particles in that substance. You must first convert the mass to moles by multiplying by a conversion factor that relates moles and mass. Can you identify the conversion factor? The number of moles must then be multiplied by a conversion factor that relates the number of representative particles to moles. That conversion factor is Avogadro's number.

EXAMPLE PROBLEM 11-4

Mass to Atoms Conversion

Gold is one of a group of metals called the coinage metals (copper, silver, and gold). How many atoms of gold (Au) are in a pure gold nugget having a mass of 25.0 g.

1. Analyze the Problem

You are given a mass of gold and must determine how many atoms it contains. Because you cannot go directly from mass to the number of atoms, you must first convert mass to moles using molar mass. Then, you can convert moles to the number of atoms using Avogadro's number. The given mass of the gold nugget is about one-eighth the molar mass of gold (196.97 g/mol), so the number of gold atoms should be approximately one-eighth Avogadro's number.

Known

mass = 25.0 g Au

molar mass Au = 196.97 g/mol Au

Unknown

number of atoms = ? atoms Au

2. Solve for the Unknown

Multiply the known amount of gold by the inverse of the molar mass as the conversion factor.

$$\text{mass Au} \times \frac{1 \text{ mole Au}}{\text{number of grams Au}} = \text{moles Au}$$

$$25.0 \cancel{\text{g Au}} \times \frac{1 \text{ mol Au}}{196.97 \cancel{\text{g Au}}} = 0.127 \text{ mol Au}$$

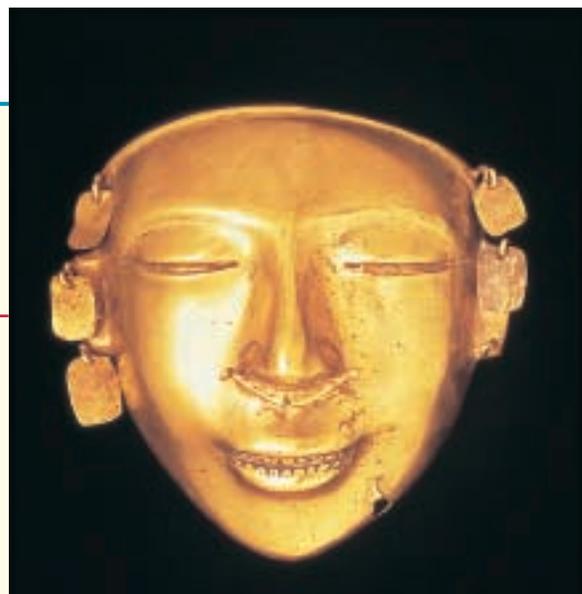
Multiply the calculated number of moles of gold by Avogadro's number as a conversion factor.

$$\text{moles Au} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \text{ mole Au}} = \text{atoms Au}$$

$$0.127 \cancel{\text{mol Au}} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \cancel{\text{mol Au}}} = 7.65 \times 10^{22} \text{ atoms Au}$$

3. Evaluate the Answer

The mass of gold has the smallest number of significant figures (3), so the answer is expressed correctly with three digits. The answer is approximately one-eighth Avogadro's number as predicted, and the unit is correct.



Gold is called a noble metal because it doesn't react readily with other elements. Early civilizations used nearly pure gold for coins and ornaments such as this gold mask from Quimbaya, Columbia, A.D. 1000-1500.

PRACTICE PROBLEMS

13. How many atoms are in each of the following samples?
- 55.2 g Li
 - 0.230 g Pb
 - 11.5 g Hg
 - 45.6 g Si
 - 0.120 kg Ti

EXAMPLE PROBLEM 11-5

Atoms to Mass Conversion

Helium is an unreactive noble gas often found in underground deposits mixed with methane. The mixture is separated by cooling the gaseous mixture until all but the helium has liquified.

A party balloon contains 5.50×10^{22} atoms of helium (He) gas. What is the mass in grams of the helium?

1. Analyze the Problem

You are given the number of atoms of helium and must find the mass of the gas.

Known

number of atoms = 5.50×10^{22} atoms He
molar mass He = 4.00 g/mol He

Unknown

mass = ? g He

2. Solve for the Unknown

Multiply the number of atoms of helium by the inverse of Avogadro's number as a conversion factor.

$$\text{atoms He} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \text{ atoms He}} = \text{moles He}$$

$$5.50 \times 10^{22} \text{ atoms He} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \text{ atoms He}} = 0.0914 \text{ mol He}$$

Multiply the calculated number of moles of helium by the conversion factor that relates mass of helium to moles of helium, molar mass.

$$\text{moles He} \times \frac{\text{number of grams He}}{1 \text{ mole He}} = \text{mass He}$$

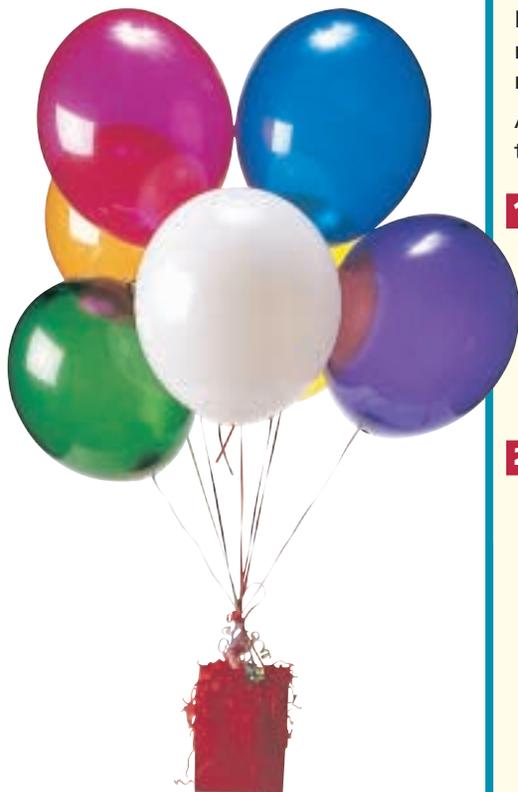
$$0.0914 \text{ mol He} \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} = 0.366 \text{ g He}$$

3. Evaluate the Answer

The answer is expressed correctly with three significant figures and has the expected unit.

PRACTICE PROBLEMS

14. What is the mass in grams of each of the following?
- 6.02×10^{24} atoms Bi
 - 1.00×10^{24} atoms Mn
 - 3.40×10^{22} atoms He
 - 1.50×10^{15} atoms N
 - 1.50×10^{15} atoms U

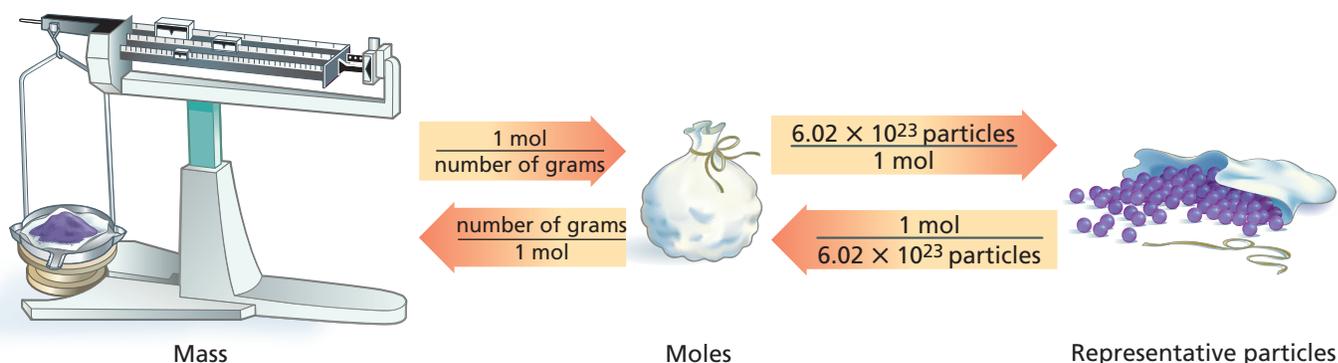


Helium gas, used in party balloons, is heavier than hydrogen gas but safer because it is unreactive and will not burn as hydrogen does.

Practice!

For more practice with mass and number of atoms conversions, go to **Supplemental Practice Problems** in Appendix A.

Now that you have learned about and practiced conversions between mass, moles, and representative particles, you can see that the mole is at the center of these calculations. Mass must always be converted to moles before being converted to atoms, and atoms must similarly be converted to moles before calculating their mass. **Figure 11-5** shows the steps to follow as you work with these conversions.



In **Figure 11-5**, mass is represented by a laboratory balance, moles are represented by a bag or bundle of particles, and representative particles are represented by the contents that are spilling out of the bag. You can see that two steps are needed to convert from mass on the left to representative particles on the right or to convert from representative particles on the right to mass on the left. The conversion factors for these conversions are given on the arrows pointing left and right. In the Example Problems, you have been making each of these conversions in separate steps, but you could make the same conversions in one calculation. For example, suppose you want to find out how many molecules of water are in 1.00 g of water. This calculation involves the conversion factors on the arrows pointing to the right. You could set up your calculation like this.

$$1.00 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 3.34 \times 10^{22} \text{ molecules H}_2\text{O}$$

Note that the units cancel to give the answer in molecules of water. Do the reverse calculation yourself using the conversion factors on the arrows pointing from right to left. What is the mass of 3.34×10^{22} molecules of water? What answer should you expect? What unit?

Figure 11-5

The mole is at the center of conversions between mass and particles. Two steps are needed to go from mass to representative particles or the reverse.

Section 11.2 Assessment

- Explain what is meant by molar mass.
- What conversion factor should be used to convert from mass to moles? Moles to mass?
- Explain the steps needed to convert the mass of an element to the number of atoms of the element.
- Thinking Critically** The mass of a single atom is usually given in the unit amu. Would it be possible to express the mass of a single atom in grams? Explain.
- Sequencing** Arrange the following in order of mass from the smallest mass to the largest: 1.0 mol Ar, 3.0×10^{24} atoms Ne, 20 g Kr.

Objectives

- **Recognize** the mole relationships shown by a chemical formula.
- **Calculate** the molar mass of a compound.
- **Calculate** the number of moles of a compound from a given mass of the compound, and the mass of a compound from a given number of moles of the compound.
- **Determine** the number of atoms or ions in a mass of a compound.

You have learned that different kinds of representative particles are counted using the mole, but so far you have applied this counting unit only to atoms of elements. Can you make similar conversions for compounds and ions? If so, you will need to know the molar mass of the compounds and ions.

Chemical Formulas and the Mole

Recall that the chemical formula for a compound indicates the types of atoms and the number of each contained in one unit of the compound. For example, freon has the formula CCl_2F_2 . The subscripts in the formula tell you that one molecule of CCl_2F_2 consists of one atom of carbon, two atoms of chlorine, and two atoms of fluorine that have chemically combined. The ratio of carbon to chlorine to fluorine is 1 : 2 : 2.

But suppose you had a mole of freon. The representative particles would be molecules of freon. A mole of freon would contain Avogadro's number of freon molecules, which means that instead of one carbon atom, you would have a mole of carbon atoms. And instead of two chlorine atoms and two fluorine atoms, you would have two moles of chlorine atoms and two moles of fluorine atoms. The ratio of carbon to chlorine to fluorine in one mole of freon would still be 1 : 2 : 2, as it is in one molecule of freon.

Figure 11-6 illustrates this principle for a dozen freon molecules. Check for yourself that a dozen freon molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. The chemical formula CCl_2F_2 not only represents an individual molecule of freon, it also represents a mole of the compound.

In some chemical calculations, you may need to convert from moles of a compound to moles of individual atoms in the compound or from moles of individual atoms in a compound to moles of the compound. The following conversion factors can be written for use in these calculations for the molecule freon.

$$\frac{1 \text{ mol C atoms}}{1 \text{ mol CCl}_2\text{F}_2} \quad \frac{2 \text{ mol Cl atoms}}{1 \text{ mol CCl}_2\text{F}_2} \quad \frac{2 \text{ mol F atoms}}{1 \text{ mol CCl}_2\text{F}_2}$$

To find out how many moles of fluorine atoms are in 5.50 moles of freon, you would multiply the moles of freon by the conversion factor that relates moles of fluorine atoms to moles of CCl_2F_2 .

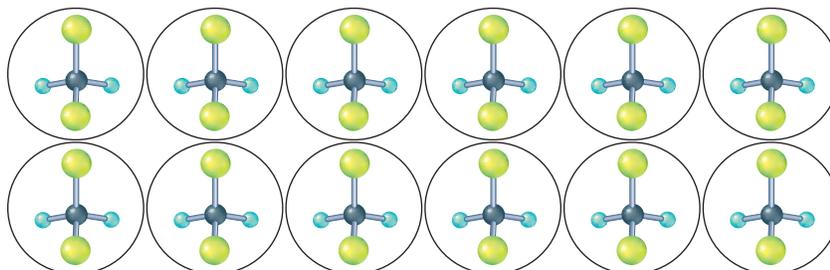
$$\text{moles CCl}_2\text{F}_2 \times \frac{\text{moles F atoms}}{1 \text{ mole CCl}_2\text{F}_2} = \text{moles F atoms}$$

$$5.50 \text{ mol CCl}_2\text{F}_2 \times \frac{2 \text{ mol F atoms}}{1 \text{ mol CCl}_2\text{F}_2} = 11.0 \text{ mol F atoms}$$

Therefore, 11.0 mol F atoms are in 5.50 mol CCl_2F_2 .

Figure 11-6

A dozen freon molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. How many of each kind of atom are contained in one mole of freon?



Conversion factors such as the one just used for fluorine can be written for any element in a compound. The number of moles of the element that goes in the numerator of the conversion factor is the subscript for that element in the chemical formula.

EXAMPLE PROBLEM 11-6

Mole Relationships from a Chemical Formula

Aluminum oxide (Al_2O_3), often called alumina, is the principal raw material for the production of aluminum. Alumina occurs in the minerals corundum and bauxite. Determine the moles of aluminum ions (Al^{3+}) in 1.25 moles of aluminum oxide.

1. Analyze the Problem

You are given the number of moles of Al_2O_3 and must determine the number of moles of Al^{3+} ions. Use a conversion factor based on the chemical formula that relates moles of Al^{3+} ions to moles of Al_2O_3 . Every mole of Al_2O_3 contains two moles of Al^{3+} ions. Thus, the answer should be two times the number of moles of Al_2O_3 .

Known

number of moles = 1.25 mol Al_2O_3

Unknown

number of moles = ? mol Al^{3+} ions

2. Solve for the Unknown

1 mol Al_2O_3 contains 2 mol Al^{3+} ions. Determine the conversion factor relating moles of Al^{3+} ions to moles of Al_2O_3 .

$$\frac{2 \text{ mol Al}^{3+} \text{ ions}}{1 \text{ mol Al}_2\text{O}_3}$$

Multiply the known number of moles of Al_2O_3 by the conversion factor.

$$\text{moles Al}_2\text{O}_3 \times \frac{\text{moles Al}^{3+} \text{ ions}}{\text{mole Al}_2\text{O}_3} = \text{moles Al}^{3+} \text{ ions}$$

$$1.25 \text{ mol Al}_2\text{O}_3 \times \frac{2 \text{ mol Al}^{3+} \text{ ions}}{1 \text{ mol Al}_2\text{O}_3} = 2.50 \text{ mol Al}^{3+} \text{ ions}$$

3. Evaluate the Answer

Because the conversion factor is a ratio of whole numbers, the number of significant digits is based on the moles of Al_2O_3 . Therefore, the answer is expressed correctly with three significant figures. As predicted, the answer is twice the number of moles of Al_2O_3 .



The capstone placed at the top of the Washington Monument in 1884 is a 22.86-cm pyramid of pure aluminum. Until an inexpensive purification process was developed, aluminum was considered a rare and precious metal.

PRACTICE PROBLEMS

- Determine the number of moles of chloride ions in 2.50 mol ZnCl_2 .
- Calculate the number of moles of each element in 1.25 mol glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).
- Determine the number of moles of sulfate ions present in 3.00 mol iron(III) sulfate ($\text{Fe}_2(\text{SO}_4)_3$).
- How many moles of oxygen atoms are present in 5.00 mol diphosphorus pentoxide (P_2O_5)?
- Calculate the number of moles of hydrogen atoms in 11.5 mol water.



For more practice calculating the number of moles of atoms or ions in a given number of moles of a compound, go to **Supplemental Practice Problems** in Appendix A.

The Molar Mass of Compounds

The mass of your backpack is the sum of the mass of the pack plus the masses of the books, notebooks, pencils, lunch, and miscellaneous items you put into it. You could find its mass by determining the mass of each item separately and adding them together. Similarly, the mass of a mole of a compound equals the sum of the masses of every particle that makes up the compound. You know how to use the molar mass of an element as a conversion factor in calculations. You also know that a chemical formula indicates the number of moles of each element in a compound. With this information, you can now determine the molar mass of a compound.

Suppose you want to determine the molar mass of potassium chromate (K_2CrO_4). Using the periodic table, the mass of one mole of each element present in potassium chromate can be determined. That mass is then multiplied by the number of moles of that element in the chemical formula. Adding the masses of all elements present will yield the molar mass of K_2CrO_4 .

number of moles \times molar mass = number of grams

$$2.000 \text{ mol K} \times \frac{39.10 \text{ g K}}{1 \text{ mol K}} = 78.20 \text{ g}$$

$$1.000 \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 52.00 \text{ g}$$

$$4.000 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 64.00 \text{ g}$$

$$\text{molar mass } \text{K}_2\text{CrO}_4 = 194.20 \text{ g}$$

Practice!

For more practice calculating the molar mass of a compound, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- Determine the molar mass of each of the following ionic compounds: NaOH , CaCl_2 , $\text{KC}_2\text{H}_3\text{O}_2$, $\text{Sr}(\text{NO}_3)_2$, and $(\text{NH}_4)_3\text{PO}_4$.
- Calculate the molar mass of each of the following molecular compounds: $\text{C}_2\text{H}_5\text{OH}$, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, HCN , CCl_4 , and H_2O .

The molar mass of a compound demonstrates the law of conservation of mass. The sum of the masses of the elements that reacted to form the compound equals the mass of the compound. **Figure 11-7** shows 194 g, or one mole, of K_2CrO_4 and masses equal to one mole of two other substances.

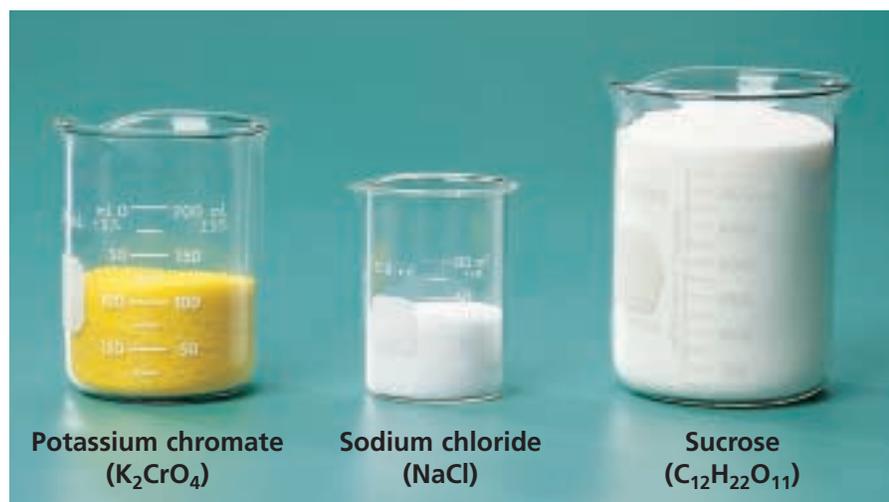


Figure 11-7

Each substance contains different numbers and kinds of atoms so their molar masses are different. The molar mass of each compound is the sum of the masses of all the elements contained in the compound.

Converting Moles of a Compound to Mass

Suppose you need to measure a certain number of moles of a compound for an experiment. First, you must calculate the mass in grams that corresponds to the necessary number of moles. Then, that mass can be measured on a balance. In Example Problem 11-2, you learned how to convert the number of moles of elements to mass using molar mass as the conversion factor. The procedure is the same for compounds except that you must first calculate the molar mass of the compound.

EXAMPLE PROBLEM 11-7

Mole-to-Mass Conversion for Compounds

The characteristic odor of garlic is due to the compound allyl sulfide ((C₃H₅)₂S). What is the mass of 2.50 moles of allyl sulfide?

1. Analyze the Problem

You are given 2.50 mol (C₃H₅)₂S and must convert the moles to mass using the molar mass as a conversion factor. The molar mass is the sum of the molar masses of all the elements in (C₃H₅)₂S.

Known

number of moles = 2.50 mol (C₃H₅)₂S

Unknown

molar mass (C₃H₅)₂S = ? g/mol (C₃H₅)₂S

mass = ? g (C₃H₅)₂S

2. Solve for the Unknown

Calculate the molar mass of (C₃H₅)₂S.

$$1 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 32.07 \text{ g S}$$

$$6 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 72.06 \text{ g C}$$

$$10 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 10.08 \text{ g H}$$

$$\text{molar mass (C}_3\text{H}_5\text{)}_2\text{S} = 114.21 \text{ g/mol (C}_3\text{H}_5\text{)}_2\text{S}$$

Convert mol (C₃H₅)₂S to g (C₃H₅)₂S by using the molar mass as a conversion factor.

$$\text{moles (C}_3\text{H}_5\text{)}_2\text{S} \times \frac{\text{number of grams (C}_3\text{H}_5\text{)}_2\text{S}}{1 \text{ mole (C}_3\text{H}_5\text{)}_2\text{S}} = \text{mass (C}_3\text{H}_5\text{)}_2\text{S}$$

$$2.50 \text{ mol (C}_3\text{H}_5\text{)}_2\text{S} \times \frac{114.21 \text{ g (C}_3\text{H}_5\text{)}_2\text{S}}{1 \text{ mol (C}_3\text{H}_5\text{)}_2\text{S}} = 286 \text{ g (C}_3\text{H}_5\text{)}_2\text{S}$$

3. Evaluate the Answer

Mol (C₃H₅)₂S has the smaller number of significant figures (3), so the answer is expressed correctly with three digits. The unit, g, is correct.



The pungent odor of garlic is characteristic of sulfides. Sulfides, including hydrogen sulfide, are noted for their strong, often unpleasant odors. The sulfur atom in allyl sulfide forms a chemical bond to each of the two C₃H₅ groups in the molecule.

PRACTICE PROBLEMS

27. What is the mass of 3.25 moles of sulfuric acid (H₂SO₄)?
28. What is the mass of 4.35×10^{-2} moles of zinc chloride (ZnCl₂)?
29. How many grams of potassium permanganate are in 2.55 moles?



For more practice converting moles of a compound to mass, go to **Supplemental Practice Problems** in Appendix A.

Converting the Mass of a Compound to Moles

Imagine that the experiment you are doing in the laboratory produces 5.55 g of a compound. How many moles is this? To find out, you calculate the molar mass of the compound and determine it to be 185.0 g/mol. The molar mass relates grams and moles, but this time you need the inverse of the molar mass as the conversion factor.

$$5.50 \text{ g compound} \times \frac{1 \text{ mol compound}}{185.0 \text{ g compound}} = 0.0297 \text{ mol compound}$$

EXAMPLE PROBLEM 11-8

Mass-to-Mole Conversion for Compounds

Calcium hydroxide ($\text{Ca}(\text{OH})_2$) is used to remove sulfur dioxide from the exhaust gases emitted by power plants and for softening water by the elimination of Ca^{2+} and Mg^{2+} ions. Calculate the number of moles of calcium hydroxide in 325 g.

1. Analyze the Problem

You are given 325 g $\text{Ca}(\text{OH})_2$ and are solving for the number of moles of $\text{Ca}(\text{OH})_2$. You must first calculate the molar mass of $\text{Ca}(\text{OH})_2$.

Known

$$\text{mass} = 325 \text{ g Ca}(\text{OH})_2$$

Unknown

$$\text{molar mass} = ? \text{ g/mol Ca}(\text{OH})_2$$

$$\text{number of moles} = ? \text{ mol Ca}(\text{OH})_2$$

2. Solve for the Unknown

Determine the molar mass of $\text{Ca}(\text{OH})_2$.

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g}$$

$$2 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 2.016 \text{ g}$$

$$\text{molar mass of Ca}(\text{OH})_2 = 74.096 \text{ g/mol} = 74.10 \text{ g/mol}$$

Use the inverse of molar mass as the conversion factor to calculate moles.

$$325 \text{ g Ca}(\text{OH})_2 \times \frac{1 \text{ mol Ca}(\text{OH})_2}{74.10 \text{ g Ca}(\text{OH})_2} = 4.39 \text{ mol Ca}(\text{OH})_2$$

3. Evaluate the Answer

The given mass of $\text{Ca}(\text{OH})_2$ has fewer digits than any other value in the calculations so it determines the number of significant figures in the answer (3). To check the reasonableness of the answer, round off the molar mass of $\text{Ca}(\text{OH})_2$ to 75 g/mol and the given mass of $\text{Ca}(\text{OH})_2$ to 300 g. Seventy-five is contained in 300 four times. Thus, the answer is reasonable.



This compound, commonly called lime, is calcium oxide (CaO). Calcium oxide reacts with water to produce calcium hydroxide. Lime is a component of cement and is used to counteract excess acidity in soil.

Practice!

For more practice converting the mass of a compound to moles, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

30. Determine the number of moles present in each of the following.

a. 22.6 g AgNO_3

d. 25.0 g Fe_2O_3

b. 6.50 g ZnSO_4

e. 254 g PbCl_4

c. 35.0 g HCl

Converting the Mass of a Compound to Number of Particles

Example Problem 11-8 illustrated how to find the number of moles of a compound contained in a given mass. Now, you will learn how to calculate the number of representative particles—molecules or formula units—contained in a given mass and, in addition, the number of atoms or ions. Recall that no direct conversion is possible between mass and number of particles. You must first convert the given mass to moles by multiplying by the inverse of the molar mass. Then, you can convert moles to the number of representative particles by multiplying by Avogadro's number. To determine numbers of atoms or ions in a compound, you will need conversion factors that are ratios of the number of atoms or ions in the compound to one mole of compound. These are based on the chemical formula. Example Problem 11-9 provides practice in solving this type of problem.

EXAMPLE PROBLEM 11-9

Conversion from Mass to Moles to Particles

Aluminum chloride is used in refining petroleum and manufacturing rubber and lubricants. A sample of aluminum chloride (AlCl_3) has a mass of 35.6 g.

- How many aluminum ions are present?
- How many chloride ions are present?
- What is the mass in grams of one formula unit of aluminum chloride?

1. Analyze the Problem

You are given 35.6 g AlCl_3 and must calculate the number of Al^{3+} ions, the number of Cl^- ions, and the mass in grams of one formula unit of AlCl_3 . Molar mass, Avogadro's number, and ratios from the chemical formula are the necessary conversion factors. The ratio of Al^{3+} ions to Cl^- ions in the chemical formula is 1:3. Therefore, the calculated numbers of ions should be in that ratio. The mass of one formula unit in grams should be an extremely small number.

Known

mass = 35.6 g AlCl_3

Unknown

number of ions = ? Al^{3+} ions

number of ions = ? Cl^- ions

mass = ? g/formula unit AlCl_3

2. Solve for the Unknown

Determine the molar mass of AlCl_3 .

$$1 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 26.98 \text{ g Al}$$

$$3 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 106.35 \text{ g Cl}$$

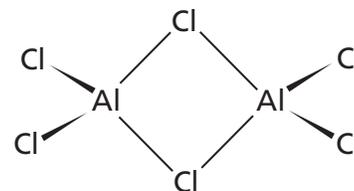
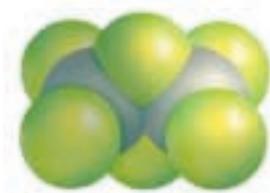
$$\text{Molar mass of } \text{AlCl}_3 = 133.33 \text{ g/mol } \text{AlCl}_3$$

Multiply by the inverse of the molar mass as a conversion factor to convert the mass of AlCl_3 to moles.

$$\text{grams } \text{AlCl}_3 \times \frac{1 \text{ mol } \text{AlCl}_3}{\text{grams } \text{AlCl}_3} = \text{moles } \text{AlCl}_3$$

$$35.6 \text{ g } \text{AlCl}_3 \times \frac{1 \text{ mol } \text{AlCl}_3}{133.33 \text{ g } \text{AlCl}_3} = 0.267 \text{ mol } \text{AlCl}_3$$

Continued on next page



At ordinary temperatures, aluminum chloride is a solid with the formula AlCl_3 . In the vapor phase, however, aluminum chloride exists as a doubled molecule, or dimer, with the formula Al_2Cl_6 .

Multiply by Avogadro's number to calculate the number of formula units of AlCl_3 .

$$0.267 \text{ mol AlCl}_3 \times \frac{6.02 \times 10^{23} \text{ formula units}}{1 \text{ mol AlCl}_3} = 1.61 \times 10^{23} \text{ formula units AlCl}_3$$

To calculate the number of Al^{3+} and Cl^- ions, use the ratios from the chemical formula as conversion factors.

$$1.61 \times 10^{23} \text{ AlCl}_3 \text{ formula unit} \times \frac{1 \text{ Al}^{3+} \text{ ion}}{1 \text{ AlCl}_3 \text{ formula unit}} = 1.61 \times 10^{23} \text{ Al}^{3+} \text{ ions}$$

$$1.61 \times 10^{23} \text{ AlCl}_3 \text{ formula unit} \times \frac{3 \text{ Cl}^- \text{ ions}}{1 \text{ AlCl}_3 \text{ formula unit}} = 4.83 \times 10^{23} \text{ Cl}^- \text{ ions}$$

Calculate the mass in grams of one formula unit of AlCl_3 . Start with molar mass and use the inverse of Avogadro's number as a conversion factor.

$$\frac{133.33 \text{ g AlCl}_3}{1 \text{ mol}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ formula unit}} = 2.21 \times 10^{-22} \text{ g AlCl}_3/\text{formula unit}$$

3. Evaluate the Answer

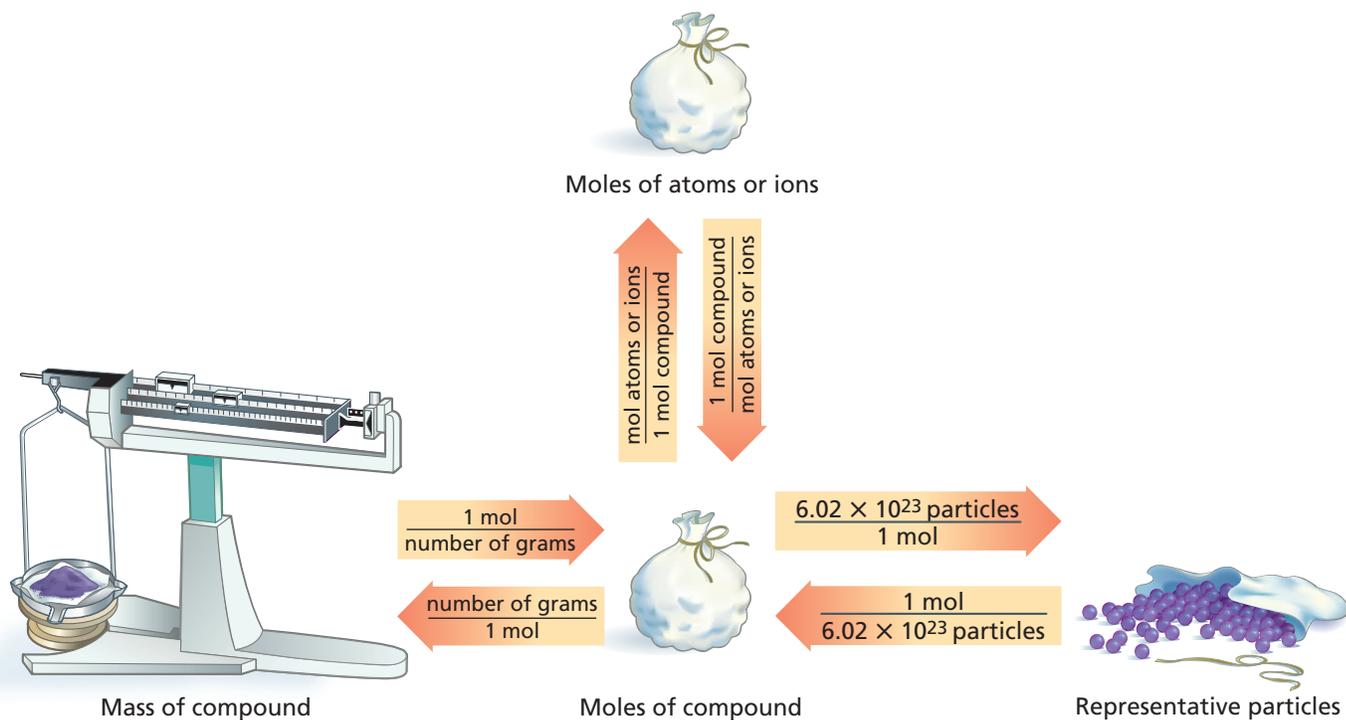
A minimum of three significant figures is used in each value in the calculations. Therefore, the answers have the correct number of digits. The number of Cl^- ions is three times the number of Al^{3+} ions, as predicted. The mass of a formula unit of AlCl_3 can be checked by calculating it in a different way: Divide the mass of AlCl_3 (35.6 g) by the number of formula units contained in the mass (1.61×10^{23} formula units) to obtain the mass of one formula unit. The two answers are the same.

Practice!

For more practice calculating the number of ions or atoms in a mass of a compound and the mass in grams of one formula unit, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- 31.** A sample of silver chromate (Ag_2CrO_4) has a mass of 25.8 g.
 - a. How many Ag^+ ions are present?
 - b. How many CrO_4^{2-} ions are present?
 - c. What is the mass in grams of one formula unit of silver chromate?
- 32.** What mass of sodium chloride contains 4.59×10^{24} formula units?
- 33.** A sample of ethanol ($\text{C}_2\text{H}_5\text{OH}$) has a mass of 45.6 g.
 - a. How many carbon atoms does the sample contain?
 - b. How many hydrogen atoms are present?
 - c. How many oxygen atoms are present?
- 34.** A sample of sodium sulfite (Na_2SO_3) has a mass of 2.25 g.
 - a. How many Na^+ ions are present?
 - b. How many SO_3^{2-} ions are present?
 - c. What is the mass in grams of one formula unit of Na_2SO_3 ?
- 35.** A sample of carbon dioxide has a mass of 52.0 g.
 - a. How many carbon atoms are present?
 - b. How many oxygen atoms are present?
 - c. What is the mass in grams of one molecule of CO_2 ?



Conversions among mass, moles, and the number of particles are summarized in **Figure 11-8**. Refer to this diagram often until you become familiar with the calculations. Note that the molar mass (number of grams/1 mol) and the inverse of molar mass (1 mol/number of grams) are the conversion factors between the mass of a substance and the number of moles of the substance. Avogadro's number and its inverse are the conversion factors between the moles of a substance and the number of representative particles. To convert between the number of moles of a compound and the number of moles of atoms or ions contained in the compound, you need the ratio of moles of atoms or ions to 1 mole of compound or its inverse, which are shown on the upward and downward arrows in **Figure 11-8**. These ratios are derived from the subscripts in the chemical formula. What ratio would you use to find the moles of hydrogen atoms in four moles of water?

Figure 11-8

Note the central position of the mole. To go from the left, right, or top of the diagram to any other place, you must go through the mole. The conversion factors on the arrows provide the means for making the conversions.

Section 11.3 Assessment

36. Describe how you can determine the molar mass of a compound.
37. What three conversion factors are often used in mole conversions?
38. Explain how you can determine the number of atoms or ions in a given mass of a compound.
39. If you know the mass in grams of a molecule of a substance, could you obtain the mass of a mole of that substance? Explain.
40. **Thinking Critically** Design a bar graph that will show the number of moles of each element present in 500 g dioxin ($\text{C}_{12}\text{H}_4\text{Cl}_4\text{O}_2$), a powerful poison.
41. **Applying Concepts** The recommended daily allowance of calcium is 1000 mg of Ca^{2+} ions. Calcium carbonate is used to supply the calcium in vitamin tablets. How many moles of calcium ions does 1000 mg represent? How many moles of calcium carbonate are needed to supply the required amount of calcium ions? What mass of calcium carbonate must each tablet contain?

Empirical and Molecular Formulas

Objectives

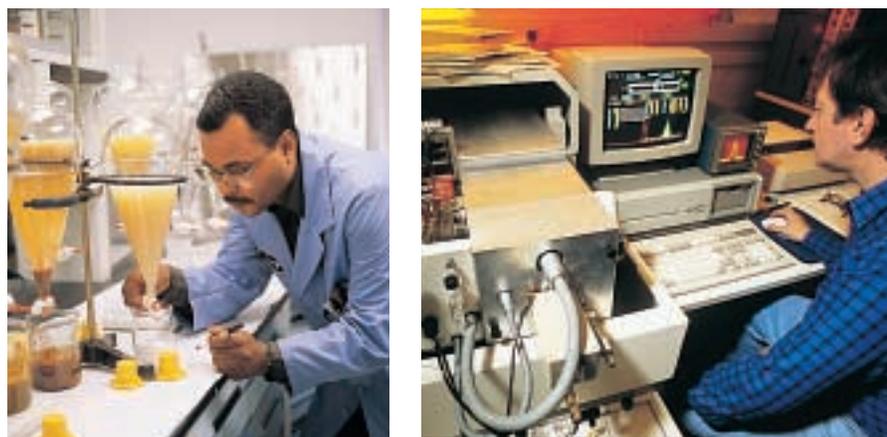
- **Explain** what is meant by the percent composition of a compound.
- **Determine** the empirical and molecular formulas for a compound from mass percent and actual mass data.

Vocabulary

percent composition
empirical formula
molecular formula

Figure 11-9

New compounds are first made on a small scale by a synthetic chemist like the one on the left. Then, an analytical chemist, like the one on the right, analyzes the compound to verify its structure and percent composition.



Percent Composition

It's the analytical chemist's job to identify the elements a compound contains and determine their percent by mass. For example, a 100-g sample of a new compound contains 55 g of element X and 45 g of element Y. The percent by mass of any element in a compound can be found by dividing the mass of the element by the mass of the compound and multiplying by 100.

$$\frac{\text{mass of element}}{\text{mass of compound}} \times 100 = \text{percent by mass}$$

$$\frac{55 \text{ g element X}}{100 \text{ g compound}} \times 100 = 55\% \text{ element X}$$

$$\frac{45 \text{ g element Y}}{100 \text{ g compound}} \times 100 = 45\% \text{ element Y}$$

Because percent means parts per 100, the percents by mass of all the elements of a compound must always add up to 100. The percent composition of the compound is 55% X and 45% Y. The percent by mass of each element in a compound is called the **percent composition** of a compound.

Percent composition from the chemical formula If you already know the chemical formula for a compound such as water (H_2O), can you calculate its percent composition? The answer is yes. You can use the chemical formula to calculate the molar mass of water (18.02 g/mol) and assume you have an 18.02-g sample. Because the percent composition of a compound is always the same, no matter the size of the sample, you can assume that the sample

size is one mole. To find the mass of each element in a mole of water, multiply the molar mass of the element by its subscript in the chemical formula. Because one mole of water contains two moles of hydrogen atoms, the mass of hydrogen in a mole of water is $(2 \text{ mol})(1.01 \text{ g/mol}) = 2.02 \text{ g}$. To find the percent by mass of hydrogen in water, divide the mass of hydrogen by the molar mass of water (18.02 g/mol) and multiply by 100.

$$\frac{2.02 \text{ g H}}{18.02 \text{ g H}_2\text{O}} \times 100 = 11.2\% \text{ H}$$

One mole of water contains one mole of oxygen. Thus, the mass of oxygen in one mole of water is 16.00 g . The percent by mass of oxygen is

$$\frac{16.00 \text{ g O}}{18.02 \text{ g H}_2\text{O}} \times 100 = 88.80\% \text{ O}$$

The percent composition of water is 11.2% hydrogen and 88.80% oxygen.

The general equation for calculating the percent by mass of any element in a compound is

$$\frac{\text{mass of element in 1 mol compound}}{\text{molar mass of compound}} \times 100 = \% \text{ by mass element}$$

The **miniLAB** provides an opportunity to practice calculating percents.

Careers Using Chemistry

Analytical Chemist

Would you like to work in a laboratory, solving problems and making precise measurements with state-of-the-art equipment? If so, consider a career as an analytical chemist.

Analytical chemists identify and measure the elements and compounds found in substances. They might determine the composition of the raw materials used in manufacturing or help physicians diagnose diseases. They might identify air or water pollutants or determine which nutrients are in certain foods. Analytical chemists work in fields as varied as archeology, crime, and space science.

miniLAB

Percent Composition and Gum

Interpreting Data Water soluble sweeteners and flavorings are added to chewing gum. Are these chemicals added as an outside coating or are they mixed throughout the gum?

Materials balance, weighing paper, 250-mL beakers (2), pieces of chewing gum (2), stirring rod, paper towels, window screen ($10 \text{ cm} \times 10 \text{ cm}$), scissors, clock or timer

Procedure



CAUTION: Do not taste or eat any items used in the lab.

1. Unwrap two pieces of chewing gum. Measure the mass of each separately on a piece of weighing paper. Label the weighing papers with the masses to avoid mixing up your data. Record the masses.
2. Add 150 mL of cold tap water to a 250-mL beaker. Place one piece of chewing gum in the water and stir for two minutes.
3. Remove the gum from the water and pat dry using paper towels. Measure and record the mass of the dried gum.

4. Use scissors to cut the second piece of gum into small pieces, each about the width of a pea. Repeat step 2 using fresh water. Use the stirring rod to keep the pieces of gum from clumping together.
5. Use the window screen to strain the water from the gum. Pat the gum dry using paper towels. Measure and record the mass of the dried gum.
6. Discard the gum in a waste container.

Analysis

1. For the uncut piece of gum, calculate the mass of sweeteners and flavorings that dissolved in the water. The mass of sweeteners and flavorings is the difference between the original mass of the gum and the mass of the dried gum.
2. For the gum that was in small pieces, calculate the mass of dissolved sweeteners and flavorings.
3. For both pieces of gum, calculate the percent of the original mass that was soluble sweeteners and flavorings. For help, refer to *Percents* in the **Math Handbook** on page 909 of this text.
4. What can you infer from the two percentages? Is the gum sugar-coated or are the sweeteners and flavorings mixed throughout?

PRACTICE PROBLEMS

- Determine the percent by mass of each element in calcium chloride.
- Calculate the percent composition of sodium sulfate.
- Which has the larger percent by mass of sulfur, H_2SO_3 or $\text{H}_2\text{S}_2\text{O}_8$?
- What is the percent composition of phosphoric acid (H_3PO_4)?



For more practice calculating percent composition, go to **Supplemental Practice Problems** in Appendix A.

Empirical Formula

Suppose that the identities of the elements in a sample of a new compound have been determined and the compound's percent composition is known. These data can be used to find the formula for the compound. First, you must determine the smallest whole number ratio of the moles of the elements in the compound. This ratio provides the subscripts in the empirical formula. The **empirical formula** for a compound is the formula with the smallest whole-number mole ratio of the elements. The empirical formula may or may not be the same as the actual molecular formula. If the two formulas are different, the molecular formula will always be a simple multiple of the empirical formula. The empirical formula for hydrogen peroxide is HO ; the molecular formula is H_2O_2 . In both formulas, the ratio of oxygen to hydrogen is 1:1.

The data used to determine the chemical formula for a compound may be in the form of percent composition or it may be the actual masses of the elements in a given mass of the compound. If percent composition is given, you can assume that the total mass of the compound is 100.00 g and that the percent by mass of each element is equal to the mass of that element in grams. For example, the percent composition of an oxide of sulfur is 40.05% S and 59.95% O. Thus, as you can see in **Figure 11-10**, 100.00 g of the oxide contains 40.05 g S and 59.95 g O. The mass of each element can be converted to a number of moles by multiplying by the inverse of the molar mass. Recall that the number of moles of S and O are calculated in this way.

$$40.05 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 1.249 \text{ mol S}$$

$$59.95 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.747 \text{ mol O}$$

The mole ratio of S atoms to O atoms in the oxide is 1.249 : 3.747. As you can see, these values are not whole numbers and cannot be used as subscripts in a chemical formula.

How, then, can the mole ratio be converted to whole numbers? As a starting point, recognize that the element with the smaller number of moles, in this case sulfur, might have the smallest subscript possible, 1. You can make the mole value of sulfur equal to 1 if you divide both mole values by the value of sulfur (1.249). In doing so, you do not change the ratio between the two elements because both are divided by the same number.

$$\frac{1.249 \text{ mol S}}{1.249} = 1 \text{ mol S}$$

$$\frac{3.747 \text{ mol O}}{1.249} = 3 \text{ mol O}$$

The simplest whole number mole ratio of S atoms to O atoms is 1:3. Thus, the empirical formula for the oxide of sulfur is SO_3 .

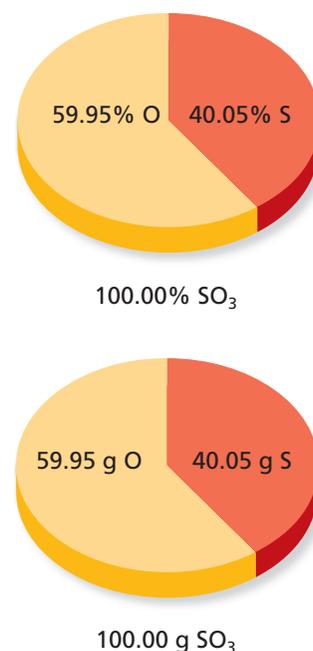
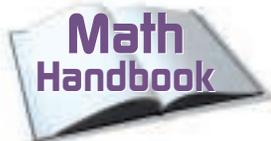


Figure 11-10

Keep this figure in mind when doing problems using percent composition. You can always assume that you have a 100-g sample of the compound and use the percents of the elements as masses of the elements.

Often in determining empirical formulas, the calculated mole values are still not whole numbers, as they are in the preceding example. In such cases, all the mole values must be multiplied by the smallest factor that will make them whole numbers. This is shown in Example Problem 11-11.



Review ratios in the **Math Handbook** on page 908 of this text.

EXAMPLE PROBLEM 11-11

Calculating an Empirical Formula from Percent Composition

Methyl acetate is a solvent commonly used in some paints, inks, and adhesives. Determine the empirical formula for methyl acetate, which has the following chemical analysis: 48.64% carbon, 8.16% hydrogen, and 43.20% oxygen.

1. Analyze the Problem

You are given the percent composition of methyl acetate and must find the empirical formula. Because you can assume that each percent by mass represents the mass of the element in a 100.00-g sample, the percent sign can be replaced with the unit grams. Then, you can convert from grams to moles using the molar mass and find the smallest whole-number ratio of moles of the elements.

Known

percent by mass = 48.64% C
percent by mass = 8.16% H
percent by mass = 43.20% O

Unknown

empirical formula = ?

2. Solve for the Unknown

The mass of C is 48.64 g, the mass of H is 8.16 g, and the mass of O is 43.20 g. Multiply the mass of each element by the conversion factor that relates moles to grams based on molar mass.

$$48.64 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.050 \text{ mol C}$$

$$8.16 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 8.10 \text{ mol H}$$

$$43.20 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.700 \text{ mol O}$$

Methyl acetate has a mole ratio of 4.050 mol C : 8.10 mol H : 2.700 mol O.

Calculate the simplest ratio of moles of the elements by dividing each number of moles by the smallest value in the mole ratio.

$$\frac{4.050 \text{ mol C}}{2.700} = 1.500 \text{ mol C} = 1.5 \text{ mol C}$$

$$\frac{8.10 \text{ mol H}}{2.700} = 3.00 \text{ mol H} = 3 \text{ mol H}$$

$$\frac{2.700 \text{ mol O}}{2.700} = 1.000 \text{ mol O} = 1 \text{ mol O}$$

The simplest ratio is 1.5 mol C : 3 mol H : 1 mol O.

Multiply the numbers of moles in the ratio by the smallest number that will produce a ratio of whole numbers.

$$2 \times 1.5 \text{ mol C} = 3 \text{ mol C}$$

$$2 \times 3 \text{ mol H} = 6 \text{ mol H}$$

$$2 \times 1 \text{ mol O} = 2 \text{ mol O}$$

The simplest whole-number ratio of C atoms to H atoms to O atoms is 3 : 6 : 2. The empirical formula is $\text{C}_3\text{H}_6\text{O}_2$.



3. Evaluate the Answer

The calculations are correct and significant figures have been observed. To check that the formula is correct, the percent composition represented by the formula can be calculated. The percent composition checks exactly with the data given in the problem.

PRACTICE PROBLEMS

46. A blue solid is found to contain 36.84% nitrogen and 63.16% oxygen. What is the empirical formula for this solid?
47. Determine the empirical formula for a compound that contains 35.98% aluminum and 64.02% sulfur.
48. Propane is a hydrocarbon, a compound composed only of carbon and hydrogen. It is 81.82% carbon and 18.18% hydrogen. What is the empirical formula?
49. The chemical analysis of aspirin indicates that the molecule is 60.00% carbon, 4.44% hydrogen, and 35.56% oxygen. Determine the empirical formula for aspirin.
50. What is the empirical formula for a compound that contains 10.89% magnesium, 31.77% chlorine, and 57.34% oxygen?



For more practice calculating an empirical formula from percent composition, go to **Supplemental Practice Problems** in Appendix A.

Molecular Formula

Would it surprise you to learn that two or more substances with distinctly different properties can have the same percent composition and the same empirical formula? How is this possible? Remember that the subscripts in an empirical formula indicate the simplest whole-number ratio of moles of the elements in the compound. But the simplest ratio does not always indicate the actual number of moles in the compound. To identify a new compound, a chemist must go one step further and determine the **molecular formula**, which specifies the actual number of atoms of each element in one molecule or formula unit of the substance. **Figure 11-11** shows an important use of the gas, acetylene. It has the same percent composition and empirical formula, CH, as benzene which is a liquid. Yet chemically and structurally acetylene and benzene are very different.

To determine the molecular formula for a compound, the molar mass of the compound must be determined through experimentation and compared with the mass represented by the empirical formula. For example, the molar mass of acetylene is 26.04 g/mol and the mass of the empirical formula, CH, is 13.02 g/mol. Dividing the actual molar mass by the mass of the empirical formula indicates that the molar mass of acetylene is two times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of acetylene}}{\text{mass of empirical formula CH}} = \frac{26.04 \text{ g/mol}}{13.02 \text{ g/mol}} = 2.000$$

The result shows that the molar mass of acetylene is two times the mass represented by the empirical formula. Thus, the molecular formula of acetylene must contain twice the number of carbon atoms and twice the number of hydrogen atoms represented by the empirical formula.

Figure 11-11

Acetylene is a gas used for welding because of the high-temperature flame produced when it is burned with oxygen.



Similarly, when the experimentally determined molar mass of benzene, 78.12 g/mol, is compared with the mass of the empirical formula, the molar mass of benzene is found to be six times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of benzene}}{\text{mass of empirical formula CH}} = \frac{78.12 \text{ g/mol}}{13.02 \text{ g/mol}} = 6.000$$

The molar mass of benzene is six times the mass represented by the empirical formula, so the molecular formula for benzene must represent six times the number of carbon and hydrogen atoms shown in the empirical formula. You can conclude that the molecular formula for acetylene is (CH)₂ or C₂H₂ and the molecular formula for benzene is (CH)₆ or C₆H₆.

A molecular formula can be represented as the empirical formula multiplied by an integer *n*.

$$\text{molecular formula} = (\text{empirical formula})_n$$

The integer is the factor (6 in the example above) by which the subscripts in the empirical formula must be multiplied to obtain the molecular formula.

EXAMPLE PROBLEM 11-12

Determining a Molecular Formula

Succinic acid is a substance produced by lichens. Chemical analysis indicates it is composed of 40.68% carbon, 5.08% hydrogen, and 54.24% oxygen and has a molar mass of 118.1 g/mol. Determine the empirical and molecular formulas for succinic acid.

1. Analyze the Problem

You are given the percent composition that allows you to calculate the empirical formula. Assume that each percent by mass represents the mass of the element in a 100.00-g sample. You can compare the given molar mass with the mass represented by the empirical formula to find *n*.

Known

percent by mass = 40.68% C

percent by mass = 5.08% H

percent by mass = 54.24% O

molar mass = 118.1 g/mol succinic acid

Unknown

empirical formula = ?

molecular formula = ?

2. Solve for the Unknown

Use the percents by mass as grams of elements and convert to the number of moles by multiplying by the conversion factor that relates moles to mass based on molar mass.

$$40.68 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 3.387 \text{ mol C}$$

$$5.08 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.04 \text{ mol H}$$

$$54.24 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.390 \text{ mol O}$$

Succinic acid has a mole ratio of C : H : O of 3.387 : 5.04 : 3.390.



Succinic acid occurs naturally in fossils and fungi, and in lichens such as those shown. Succinic acid produced commercially is used to make compounds used in perfumes (esters) and in lacquers and dyes.

Calculate the simplest ratio among the elements by dividing each mole value by the smallest value in the mole ratio.

$$\frac{3.387 \text{ mol C}}{3.387} = 1 \text{ mol C}$$

$$\frac{5.040 \text{ mol H}}{3.387} = 1.49 \text{ mol H} = 1.5 \text{ mol H}$$

$$\frac{3.390 \text{ mol O}}{3.387} = 1.001 \text{ mol O} = 1 \text{ mol O}$$

The simplest ratio is 1 mol C : 1.5 mol H : 1 mol O.

The simplest mol ratio includes a fractional value that cannot be used as a subscript in a formula. Multiply all mole values by 2.

$$2 \times 1 \text{ mol C} = 2 \text{ mol C}$$

$$2 \times 1.5 \text{ mol H} = 3 \text{ mol H}$$

$$2 \times 1 \text{ mol O} = 2 \text{ mol O}$$

The simplest whole-number ratio of C atoms to H atoms to O atoms is 2 : 3 : 2. The empirical formula is $\text{C}_2\text{H}_3\text{O}_2$.

Calculate the empirical formula mass using the molar mass of each element.

$$2 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 24.02 \text{ g C}$$

$$3 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 3.024 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 32.00 \text{ g O}$$

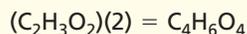
$$\text{molar mass } \text{C}_2\text{H}_3\text{O}_2 = 59.04 \text{ g/mol } \text{C}_2\text{H}_3\text{O}_2$$

Divide the experimentally determined molar mass of succinic acid by the mass of the empirical formula to determine n .

$$n = \frac{\text{molar mass of succinic acid}}{\text{molar mass of } \text{C}_2\text{H}_3\text{O}_2}$$

$$n = \frac{118.1 \text{ g/mol}}{59.04 \text{ g/mol}} = 2.000$$

Multiply the subscripts in the empirical formula by 2 to determine the actual subscripts in the molecular formula.



The molecular formula for succinic acid is $\text{C}_4\text{H}_6\text{O}_4$.

3. Evaluate the Answer

Calculation of the molar mass from the molecular formula gives the same result as the experimental molar mass.

PRACTICE PROBLEMS

51. Analysis of a chemical used in photographic developing fluid indicates a chemical composition of 65.45% C, 5.45% H, and 29.09% O. The molar mass is found to be 110.0 g/mol. Determine the molecular formula.
52. A compound was found to contain 49.98 g carbon and 10.47 g hydrogen. The molar mass of the compound is 58.12 g/mol. Determine the molecular formula.
53. A colorless liquid composed of 46.68% nitrogen and 53.32% oxygen has a molar mass of 60.01 g/mol. What is the molecular formula?



For more practice calculating a molecular formula from percent composition, go to **Supplemental Practice Problems** in Appendix A.

EXAMPLE PROBLEM 11-13



Because titanium is stable under severe conditions of heat and cold, it is used in aircraft engines, missiles, and space vehicles.

Calculating an Empirical Formula from Mass Data

Although the mineral ilmenite contains more iron than titanium, the ore is usually mined and processed for titanium, a strong, light, and flexible metal. A sample of ilmenite is found to contain 5.41 g iron, 4.64 g titanium, and 4.65 g oxygen. Determine the empirical formula for ilmenite.

1. Analyze the Problem

You are given the masses of the elements found in a known mass of ilmenite and must determine the empirical formula of the mineral. Convert the known masses of each element to moles using the conversion factor that relates moles to grams, the inverse of molar mass. Then, find the smallest whole-number ratio of the moles of the elements.

Known

mass of iron = 5.41 g Fe
mass of titanium = 4.64 g Ti
mass of oxygen = 4.65 g O

Unknown

empirical formula = ?

2. Solve for the Unknown

Multiply the known mass of each element by the conversion factor that relates moles to grams based on the molar mass.

$$5.41 \cancel{\text{g Fe}} \times \frac{1 \text{ mol Fe}}{55.85 \cancel{\text{g Fe}}} = 0.0969 \text{ mol Fe}$$

$$4.64 \cancel{\text{g Ti}} \times \frac{1 \text{ mol Ti}}{47.88 \cancel{\text{g Ti}}} = 0.0969 \text{ mol Ti}$$

$$4.65 \cancel{\text{g O}} \times \frac{1 \text{ mol O}}{16.00 \cancel{\text{g O}}} = 0.291 \text{ mol O}$$

Ilmenite has a mole ratio of Fe : Ti : O of 0.0969 : 0.0969 : 0.291. Calculate the simplest ratio by dividing each mol value by the smallest value in the ratio. Fe and Ti have the same lower value (0.0969).

$$\frac{0.0969 \text{ mol Fe}}{0.0969} = 1 \text{ mol Fe}$$

$$\frac{0.0969 \text{ mol Ti}}{0.0969} = 1 \text{ mol Ti}$$

$$\frac{0.291 \text{ mol O}}{0.0969} = 3 \text{ mol O}$$

All the mol values are whole numbers. Thus, the simplest whole-number ratio of Fe : Ti : O is 1 : 1 : 3. The empirical formula for ilmenite is FeTiO_3 .

3. Evaluate the Answer

The mass of iron is slightly greater than the mass of titanium, but the molar mass of iron is also slightly greater than that of titanium. Thus, it is reasonable that the number of moles of iron and titanium are equal. The mass of titanium is approximately the same as the mass of oxygen, but the molar mass of oxygen is about 1/3 that of titanium. Thus, a 3:1 ratio of oxygen to titanium is reasonable.

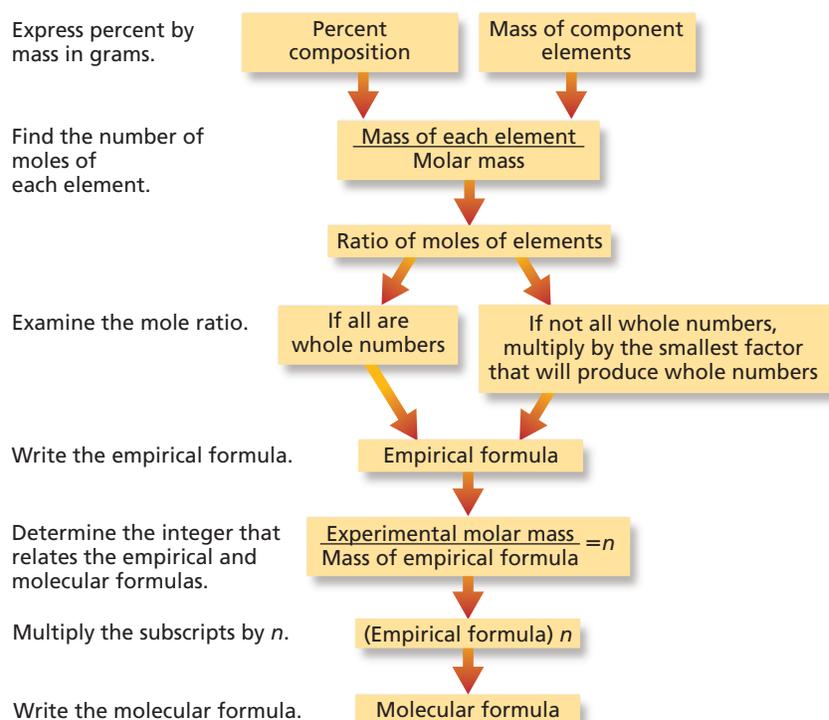
PRACTICE PROBLEMS

54. When an oxide of potassium is decomposed, 19.55 g K and 4.00 g O are obtained. What is the empirical formula for the compound?
55. Analysis of a compound composed of iron and oxygen yields 174.86 g Fe and 75.14 g O. What is the empirical formula for this compound?
56. The pain reliever morphine contains 17.900 g C, 1.680 g H, 4.225 g O, and 1.228 g N. Determine the empirical formula.
57. An oxide of aluminum contains 0.545 g Al and 0.485 g O. Find the empirical formula for the oxide.



For more practice calculating an empirical formula from mass data, go to **Supplemental Practice Problems** in Appendix A.

The steps in determining empirical and molecular formulas from percent composition or mass data are outlined below. As in other calculations, the route leads from mass through moles because formulas are based on the relative numbers of moles of elements in each mole of compound.



Section 11.4 Assessment

58. Explain how percent composition data for a compound are related to the masses of the elements in the compound.
59. What is the difference between an empirical formula and a molecular formula?
60. Explain how you can find the mole ratio in a chemical compound.
61. **Thinking Critically** An analysis for copper was performed on two pure solids. One solid was found to contain 43.0% copper; the other contained 32.0% copper. Could these solids be samples of the same copper-containing compound? Explain your answer.
62. **Inferring** Hematite (Fe_2O_3) and magnetite (Fe_3O_4) are two ores used as sources of iron. Which ore provides the greater percent of iron per kilogram?

Objectives

- **Explain** what a hydrate is and how its name reflects its composition.
- **Determine** the formula for a hydrate from laboratory data.

Vocabulary

hydrate

Figure 11-12

The presence of water and various mineral impurities account for the variety of different colored opals. Further changes in color occur when opals are allowed to dry out.

Have you ever watched crystals slowly form from a water solution? Sometimes water molecules adhere to the ions as the solid forms. These water molecules become part of the crystal and are called water of hydration. Solids in which water molecules are trapped are called hydrates. A **hydrate** is a compound that has a specific number of water molecules bound to its atoms. **Figure 11-12** shows examples of a common gemstone called opal, which is composed of silicon dioxide (SiO_2). The unusual coloring is the result of the presence of water in the mineral.



Naming Hydrates

In the formula for a hydrate, the number of water molecules associated with each formula unit of the compound is written following a dot: for example, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$. This compound is called sodium carbonate decahydrate. In the word *decahydrate*, the prefix *deca-* means ten and the root word *hydrate* refers to water. Decahydrate means that ten molecules of water are associated with one formula unit of compound. The mass of water associated with a formula unit must be included in molar mass calculations. Hydrates are found with a variety of numbers of water molecules. **Table 11-1** lists some common hydrates.

Table 11-1

Formulas for Hydrates			
Prefix	Molecules H_2O	Formula	Name
Mono-	1	$(\text{NH}_4)_2\text{C}_2\text{O}_4 \cdot \text{H}_2\text{O}$	Ammonium oxalate monohydrate
Di-	2	$\text{CaCl}_2 \cdot 2\text{H}_2\text{O}$	Calcium chloride dihydrate
Tri-	3	$\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3\text{H}_2\text{O}$	Sodium acetate trihydrate
Tetra-	4	$\text{FePO}_4 \cdot 4\text{H}_2\text{O}$	Iron(III) phosphate tetrahydrate
Penta-	5	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	Copper(II) sulfate pentahydrate
Hexa-	6	$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	Cobalt(II) chloride hexahydrate
Hepta-	7	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	Magnesium sulfate heptahydrate
Octa-	8	$\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$	Barium hydroxide octahydrate
Deca-	10	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	Sodium carbonate decahydrate

Analyzing a Hydrate

To analyze hydrates, you must drive off the water of hydration. Often this is done by heating the compound. The substance remaining after heating is anhydrous, or “without water.” For example, hydrated cobalt(II) chloride is a pink solid that turns a deep blue when the water of hydration is driven off and anhydrous cobalt(II) chloride is produced. See **Figure 11-13**.



Figure 11-13

Hydrated CoCl_2 , shown on the left, is pink. Its anhydrous form, on the right, is blue. The transition from pink to blue was accomplished by heating the hydrate until all water of hydration was removed.

Formula for a hydrate How can you determine the formula for a hydrate? You must find the number of moles of water associated with one mole of the hydrate. Suppose you have a 5.00-g sample of a hydrate of barium chloride. You know that the formula is $\text{BaCl}_2 \cdot x\text{H}_2\text{O}$. You must determine x , the coefficient of H_2O in the hydrate formula that indicates the number of moles of water associated with one mole of BaCl_2 . To find x , you would heat the sample of the hydrate to drive off the water of hydration. After heating, the dried substance, which is anhydrous BaCl_2 , has a mass of 4.26 g. The mass of the water of hydration is the difference between the mass of the hydrate (5.00 g) and the mass of the anhydrous compound (4.26 g).

$$5.00 \text{ g BaCl}_2 \text{ hydrate} - 4.26 \text{ g anhydrous BaCl}_2 = 0.74 \text{ g H}_2\text{O}$$

You now know the masses of BaCl_2 and H_2O in the sample. You can convert these masses to moles using the molar masses. The molar mass of BaCl_2 is 208.23 g/mol and the molar mass of H_2O is 18.02 g/mol.

$$4.26 \text{ g BaCl}_2 \times \frac{1 \text{ mol BaCl}_2}{208.23 \text{ g BaCl}_2} = 0.0205 \text{ mol BaCl}_2$$

$$0.74 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 0.041 \text{ mol H}_2\text{O}$$

Now that the moles of BaCl_2 and H_2O have been determined, you can calculate the ratio of moles of H_2O to moles of BaCl_2 which is x , the coefficient that precedes H_2O in the formula for the hydrate.

$$x = \frac{\text{moles H}_2\text{O}}{\text{moles BaCl}_2} = \frac{0.041 \text{ mol H}_2\text{O}}{0.0205 \text{ mol BaCl}_2} = \frac{2.0 \text{ mol H}_2\text{O}}{1.00 \text{ mol BaCl}_2} = \frac{2}{1}$$

The ratio of moles of H_2O to moles of BaCl_2 is 2:1, so two moles of water are associated with one mole of barium chloride. The value of the coefficient x is 2 and the formula for the hydrate is $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$. What is the name of the hydrate? The **CHEMLAB** at the end of this chapter will give you experience determining the formula of a hydrate.

EXAMPLE PROBLEM 11-14

Determining the Formula for a Hydrate

A mass of 2.50 g of blue, hydrated copper sulfate ($\text{CuSO}_4 \cdot x\text{H}_2\text{O}$) is placed in a crucible and heated. After heating, 1.59 g white anhydrous copper sulfate (CuSO_4) remains. What is the formula for the hydrate? Name the hydrate.

1. Analyze the Problem

You are given a mass of hydrated copper sulfate. The mass after heating is the mass of the anhydrous compound. You know the formula for the compound except for x , the number of moles of water of hydration.

Known

mass of hydrated compound = 2.50 g $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$

mass of anhydrous compound = 1.59 g CuSO_4

molar mass = 18.02 g/mol H_2O

molar mass = 159.6 g/mol CuSO_4

Unknown

formula for hydrate = ?

name of hydrate = ?

2. Solve for the Unknown

Subtract the mass of the anhydrous copper sulfate from the mass of the hydrated copper sulfate to determine the mass of water lost.

mass of hydrated copper sulfate	2.50 g
mass of anhydrous copper sulfate	<u>-1.59 g</u>
mass of water lost	0.91 g

Calculate the number of moles of H_2O and anhydrous CuSO_4 using the conversion factor that relates moles and mass based on the molar mass.

$$1.59 \text{ g } \text{CuSO}_4 \times \frac{1 \text{ mol } \text{CuSO}_4}{159.6 \text{ g } \text{CuSO}_4} = 0.00996 \text{ mol } \text{CuSO}_4$$

$$0.91 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} = 0.050 \text{ mol } \text{H}_2\text{O}$$

Determine the value of x .

$$x = \frac{\text{moles } \text{H}_2\text{O}}{\text{moles } \text{CuSO}_4}$$

$$x = \frac{0.050 \text{ mol } \text{H}_2\text{O}}{0.00996 \text{ mol } \text{CuSO}_4} = \frac{5.0 \text{ mol } \text{H}_2\text{O}}{1.00 \text{ mol } \text{CuSO}_4} = 5$$

The ratio of H_2O to CuSO_4 is 5 : 1, so the formula for the hydrate is $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, copper(II) sulfate pentahydrate.

3. Evaluate the Answer

Copper(II) sulfate pentahydrate is listed as a hydrate in Table 11-1.



Heating blue hydrated copper sulfate drives off the water of hydration and converts it to white anhydrous copper sulfate. How could you convert anhydrous copper sulfate to its blue hydrated form?

PRACTICE PROBLEMS

63. A hydrate is found to have the following percent composition: 48.8% MgSO_4 and 51.2% H_2O . What is the formula and name for this hydrate?
64. Figure 11-13 shows a common hydrate of cobalt(II) chloride. If 11.75 g of this hydrate is heated, 9.25 g of anhydrous cobalt chloride remains. What is the formula and name for this hydrate?

Practice!

For more practice calculating the formula for a hydrate, go to **Supplemental Practice Problems** in Appendix A.

Uses of hydrates The ability of the anhydrous form of a hydrate to absorb water into its crystal structure has some important applications. Anhydrous calcium chloride and calcium sulfate are used as desiccants or drying agents in the laboratory because they can absorb water from the air or from their liquid surroundings. For example, calcium sulfate is often added to solvents such as ethanol and ethyl ether to keep them free of water. Anhydrous calcium chloride is placed in the bottom of tightly sealed containers called desiccators. The calcium chloride absorbs moisture from the air inside the desiccator, thus creating a dry atmosphere in which other substances can be placed to be kept dry. Calcium chloride forms a monohydrate, a dihydrate, and a hexahydrate. Electronic and optical equipment, particularly that transported overseas by ship, is packaged with packets of desiccants that absorb water from the air and prevent moisture from interfering with sensitive circuitry. Some of these uses are illustrated in **Figure 11-14**.

Some hydrates, sodium sulfate decahydrate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$) for example, are used to store solar energy. When the Sun's energy heats the hydrate to a temperature greater than 32°C , the single formula unit of Na_2SO_4 in the hydrate dissolves in the 10 moles of water of hydration. In the process, energy is absorbed by the hydrate. This solar energy, stored in the solution of the hydrate, is released when the temperature decreases and the hydrate crystallizes again.



Figure 11-14

Calcium sulfate is not soluble in ethanol so it remains on the bottom of the ethanol bottle and absorbs any water dissolved in the ethanol. Calcium chloride, in the bottom of the desiccator, keeps the air inside the desiccator dry. Porous packets of desiccants can be packaged with materials that need to be kept moisture free.

Section 11.5 Assessment

65. What is a hydrate? What does its name indicate about its composition?
66. Describe the experimental procedure for determining the formula for a hydrate. Explain the reason for each step.
67. Name the compound having the formula $\text{SrCl}_2 \cdot 6\text{H}_2\text{O}$.
68. **Thinking Critically** Explain how the hydrate illustrated in **Figure 11-13** might be used as a means of roughly determining the probability of rain.
69. **Sequencing** Arrange these hydrates in order of increasing percent water content: $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$, $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$.

Hydrated Crystals

Hydrates are compounds that incorporate water molecules in their crystalline structures. The ratio of moles of water to one mole of the compound is a small whole number. For example, in the hydrated compound copper(II) sulfate pentahydrate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$), the ratio is 5:1. The ratio of moles of water to one mole of a hydrate can be determined experimentally by heating the hydrate to remove water.

Problem

How can you determine the moles of water in a mole of a hydrated compound?

Objectives

- **Heat** a known mass of hydrated compound until the water is removed.
- **Calculate** the formula for a hydrate using the mass of the hydrated compound and the mass of the anhydrous compound.

Materials

Bunsen burner
ring stand and ring
crucible and lid
clay triangle
crucible tongs
balance
Epsom salts (hydrated MgSO_4)
spatula
spark lighter or matches

Safety Precautions



- Always wear safety goggles and a lab apron.
- Hot objects will not appear to be hot.
- Use the Bunsen burner carefully.
- Turn off the Bunsen burner when not in use.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Explain how you will obtain the mass of water and the mass of anhydrous MgSO_4 contained in the hydrate.
4. How will you convert the masses of anhydrous MgSO_4 and water to moles?
5. How can you obtain the formula for the hydrate from the moles of anhydrous MgSO_4 and the moles of water?



Mass Data and Observations of Epsom Salts

Observations of hydrated MgSO_4	
Mass of crucible and lid	
Mass of crucible, lid, and hydrated MgSO_4	
Mass of hydrated MgSO_4	
Mass of crucible, lid, and anhydrous MgSO_4	
Mass of anhydrous MgSO_4	
Mass of water in hydrated MgSO_4	
Moles of anhydrous MgSO_4	
Moles of water in hydrated MgSO_4	
Observation of anhydrous MgSO_4	

Procedure

1. Measure to the nearest 0.01 g the mass of a clean, dry crucible with a lid. Record the mass.
2. Add about 3 g hydrated MgSO_4 to the crucible. Measure the mass of the crucible, lid, and hydrate to the nearest 0.01 g and record the mass.
3. Record your observations of the hydrate.
4. Place the triangle on the ring of the ring stand. Carefully place the crucible in the triangle as shown in the photo.
5. Place the crucible lid on the crucible slightly cocked to help prevent spattering and allow vapor to escape. Begin heating with a low flame, then gradually progress to a stronger flame. Heat for about 10 minutes.
6. When heating is complete, remove the crucible using tongs. Place the lid on the crucible and allow the crucible and contents to cool.
7. Measure the mass of the crucible, lid, and MgSO_4 and record the mass in the data table.
8. Observe the anhydrous MgSO_4 and record your observations.

Cleanup and Disposal

1. Discard the anhydrous MgSO_4 in a trash container or as directed by your teacher.
2. Return all lab equipment to its proper place and clean your lab station.
3. Wash your hands thoroughly when all lab work and cleanup are complete.

Analyze and Conclude

1. **Using Numbers** Use your experimental data to calculate the formula for hydrated MgSO_4 .
2. **Observing and Inferring** How did your observations of the hydrated MgSO_4 crystals compare with those of the anhydrous MgSO_4 crystals?
3. **Drawing Conclusions** Why might the method used in this experiment not be suitable for determining the water of hydration for all hydrates?
4. **Error Analysis** What is the percent error of your calculation of the water of hydration for MgSO_4 if the formula for the hydrate is $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$? What changes would you make in the procedure to reduce error?
5. **Predicting** What might you observe if the anhydrous crystals were left uncovered overnight?

Real-World Chemistry

1. Packets of the anhydrous form of a hydrate are sometimes used to keep cellars from being damp. Is there a limit to how long a packet could be used?
2. Gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) is a mineral used for making wallboard for construction. The mineral is stripped of three-quarters of its water of hydration in a process called calcining. Then, after mixing with water, it hardens to a white substance called plaster of paris. Infer what happens as calcined gypsum becomes plaster of paris.

CHEMISTRY and Technology

Making Medicines

In 1859, a chemist named Herman Kolbe identified the chemical structure of the compound that made willow bark an effective herbal medicine. Unfortunately, many people who took the medicine suffered upset stomachs from it. One such person was the father of a chemist named Felix Hoffman. To help his father, Hoffman changed the compound in such a way that it eased his father's arthritis without irritating his stomach. The new compound was acetylsalicylic acid, now known as aspirin.

Chemical Synthesis

Like aspirin, many medicines are compounds synthesized in the laboratory. During chemical synthesis, chemists combine simple compounds to form more complex compounds. Often, several chemical reactions must occur in a particular sequence to produce the desired product. The process is like putting building blocks together one at a time as the chemist carries out several chemical reactions in a particular sequence. The final product is then tested, or screened, to see if it has the desired effect. In the case of a medicine, the compound must be designed to interact with a molecule in the body known as the target molecule. The compound must have a particular shape and characteristics to be effective against the target molecule. If chemists know the shape and chemical properties of the target molecule, they try to design a compound to fit it exactly. More often, chemists must try as many compounds as possible until they find one that works.

Combinatorial Chemistry

In the past, finding a potentially successful medicine would take many years. Now, chemists may be able to speed up the process by turning to the techniques of combinatorial chemistry. In combinatorial chemistry, chemists use robots to put chemical building blocks together in every possible combination at the same time. Instead of chemists making one compound per week, combinatorial chemistry enables chemists to produce as many as 100 compounds a day and up to tens of thousands of compounds a year!



Looking to the Future

The goal of combinatorial chemistry is not to produce as many compounds as possible but to try to identify a basic structure that is effective and then develop variations of that structure. It's too soon to tell if combinatorial chemistry will revolutionize the way medicines are developed because the process has been used only within the last decade. However, some promising leads have been generated in months rather than years, and compounds produced from combinatorial chemistry are currently being tested.

Investigating the Technology

- 1. Controlling Variables** Why would it be important for chemists to control the variables of the reactions involved in combinatorial chemistry?
- 2. Using Resources** Research the two main approaches to combinatorial chemistry. Find out how they are alike and how they are different.



Visit the Chemistry Web site at science.glencoe.com to find links to more information about combinatorial chemistry.

Summary

11.1 Measuring Matter

- The mole is a unit used to count particles indirectly. One mole is the amount of a pure substance that contains 6.02×10^{23} representative particles.
- One mole of carbon-12 atoms contains 12 grams of the isotope carbon-12.

11.2 Mass and the Mole

- The molar mass of an element is the numerical equivalent of the atomic mass (amu) in grams.
- The molar mass of any substance is the mass in grams of Avogadro's number of representative particles of the substance.
- Molar mass is used to convert from moles of an element to mass, and the inverse of molar mass is used to convert from mass of an element to moles.

11.3 Moles of Compounds

- Subscripts in a chemical formula indicate how many moles of each element are in one mole of the compound.
- The molar mass of a compound is the sum of the masses of all the moles of elements present in the compound.

11.4 Empirical and Molecular Formulas

- The percent composition of a known compound can be calculated by dividing the mass of each element in one mole by the mass of a mole of the compound and multiplying by 100.
- The subscripts in an empirical formula are in a ratio of the smallest whole numbers of moles of the elements in the compound.
- The molecular formula for a compound can be determined by finding the integer by which the mass of the empirical formula differs from the molar mass of the compound.

11.5 The Formula for a Hydrate

- The formula for a hydrate consists of the formula for the ionic compound and the number of water molecules associated with one formula unit.
- The name of a hydrate consists of the compound name followed by the word *hydrate* with a prefix indicating the number of water molecules associated with one mole of compound.
- Anhydrous compounds are formed when hydrates are heated and the water of hydration is driven off.

Key Equations and Relationships

- number of representative particles = number of moles $\times \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mole}}$
(p. 311)
- number of moles = number of representative particles $\times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ representative particles}}$
(p. 311)
- mass = number of moles $\times \frac{\text{number of grams}}{1 \text{ mole}}$
(p. 315)
- number of moles = mass $\times \frac{1 \text{ mole}}{\text{number of grams}}$
(p. 316 Example Problem)
- percent by mass = $\frac{\text{mass of element}}{\text{mass of compound}} \times 100$
(p. 328)
- molecular formula = (empirical formula) n
(p. 334)

Vocabulary

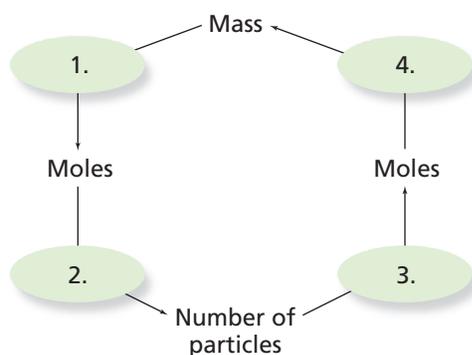
- Avogadro's number (p. 310)
- empirical formula (p. 331)
- hydrate (p. 338)
- molar mass (p. 313)
- mole (p. 310)
- molecular formula (p. 333)
- percent composition (p. 328)



Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 11 Assessment.

Concept Mapping

70. Complete this concept map by placing in each box the conversion factor needed to convert from each measure of matter to the next.



Mastering Concepts

71. Why is the mole an important unit to chemists? (11.1)
72. How is a mole similar to a dozen? (11.1)
73. What is the numerical value of Avogadro's number? (11.1)
74. What is molar mass? (11.2)
75. Which contains more atoms, a mole of silver atoms or a mole of gold atoms? Explain your answer. (11.2)
76. Discuss the relationships that exist between the mole, molar mass, and Avogadro's number. (11.2)
77. Which has a greater mass, a mole of silver atoms or a mole of gold atoms? Explain your answer. (11.2)
78. Explain the difference between atomic mass (amu) and molar mass (gram). (11.2)
79. If you divide the molar mass of an element by Avogadro's number, what is the meaning of the quotient? (11.2)
80. List three conversion factors used in molar conversions. (11.3)
81. What information is provided by the formula for potassium chromate (K_2CrO_4)? (11.3)
82. Which of the following molecules contains the most moles of carbon atoms per mole of the compound: ascorbic acid ($C_6H_8O_6$), glycerin ($C_3H_8O_3$), or vanillin ($C_8H_8O_3$)? Explain. (11.3)
83. Explain what is meant by percent composition. (11.4)
84. What is the difference between an empirical formula and a molecular formula? Use an example to illustrate your answer. (11.4)
85. Do all pure samples of a given compound have the same percent composition? Explain. (11.4)
86. What information must a chemist obtain in order to determine the empirical formula of an unknown compound? (11.4)
87. What is a hydrated compound? Use an example to illustrate your answer. (11.5)
88. Explain how hydrates are named. (11.5)

Mastering Problems

Mole-Particle Conversions (11.1)

89. Determine the number of representative particles in each of the following.
- 0.250 mol silver
 - 8.56×10^{-3} mol sodium chloride
 - 35.3 mol carbon dioxide
 - 0.425 mol nitrogen (N_2)
90. Determine the number of moles in each of the following.
- 3.25×10^{20} atoms lead
 - 4.96×10^{24} molecules glucose
 - 1.56×10^{23} formula units sodium hydroxide
 - 1.25×10^{25} copper(II) ions
91. Make the following conversions.
- 1.51×10^{15} atoms Si to mol Si
 - 4.25×10^{-2} mol H_2SO_4 to molecules H_2SO_4
 - 8.95×10^{25} molecules CCl_4 to mol CCl_4
 - 5.90 mol Ca to atoms Ca
92. How many molecules are contained in each of the following?
- 1.35 mol carbon disulfide (CS_2)
 - 0.254 mol diarsenic trioxide (As_2O_3)
 - 1.25 mol water
 - 150.0 mol HCl
93. How many moles contain each of the following?
- 1.25×10^{15} molecules carbon dioxide
 - 3.59×10^{21} formula units sodium nitrate
 - 2.89×10^{27} formula units calcium carbonate

- 94.** A bracelet containing 0.200 mol of metal atoms is 75% gold. How many particles of gold atoms are in the bracelet?
- 95.** A solution containing 0.250 mol Cu^{2+} ions is added to another solution containing 0.130 mol Ca^{2+} ions. What is the total number of metal ions in the combined solution?
- 96.** If a snowflake contains 1.9×10^{18} molecules of water, how many moles of water does it contain?
- 97.** If you could count two atoms every second, how long would it take you to count a mole of atoms? Assume that you counted continually 24 hours every day. How does the time you calculated compare with the age of Earth, which is estimated to be 4.5×10^9 years old?

Mole-Mass Conversions (11.2)

- 98.** Calculate the mass of the following.
- a. 5.22 mol He c. 2.22 mol Ti
b. 0.0455 mol Ni d. 0.00566 mol Ge
- 99.** Make the following conversions.
- a. 3.50 mol Li to g Li
b. 7.65 g Co to mol Co
c. 5.62 g Kr to mol Kr
d. 0.0550 mol As to g As
- 100.** Determine the mass in grams of the following.
- a. 1.33×10^{22} mol Sb
b. 4.75×10^{14} mol Pt
c. 1.22×10^{23} mol Ag
d. 9.85×10^{24} mol Cr

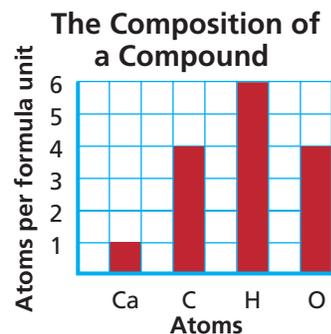
Particle-Mass Conversions (11.2)

- 101.** Convert the following to mass in grams.
- a. 4.22×10^{15} atoms U
b. 8.65×10^{25} atoms H
c. 1.25×10^{22} atoms O
d. 4.44×10^{23} atoms Pb
- 102.** A sensitive balance can detect masses of 1×10^{-8} g. How many atoms of silver would be in a sample having this mass?
- 103.** Calculate the number of atoms in each of the following.
- a. 25.8 g Hg c. 150 g Ar
b. 0.0340 g Zn d. 0.124 g Mg
- 104.** Which has more atoms, 10.0 g of carbon or 10.0 g of calcium? How many atoms does each have?
- 105.** Which has more atoms, 10.0 moles of carbon or 10.0 moles of calcium? How many does each have?

- 106.** A mixture contains 0.250 mol Fe and 1.20 g C. What is the total number of atoms in the mixture?

Chemical Formulas (11.3)

- 107.** In the formula for sodium phosphate (Na_3PO_4) how many moles of sodium are represented? How many moles of phosphorus? How many moles of oxygen?
- 108.** How many moles of oxygen atoms are contained in the following?
- a. 2.50 mol KMnO_4
b. 45.9 mol CO_2
c. 1.25×10^{-2} mol $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
- 109.** The graph shows the numbers of atoms of each element in a compound. What is its formula? What is its molar mass?



- 110.** How many carbon tetrachloride molecules are in 3.00 mol carbon tetrachloride (CCl_4)? How many carbon atoms? How many chlorine atoms? How many total atoms?

Molar Mass (11.3)

- 111.** Determine the molar mass of each of the following.
- a. nitric acid (HNO_3)
b. ammonium nitrate (NH_4NO_3)
c. zinc oxide (ZnO)
d. cobalt chloride (CoCl_2)
- 112.** Calculate the molar mass of each of the following.
- a. ascorbic acid ($\text{C}_6\text{H}_8\text{O}_6$)
b. sulfuric acid (H_2SO_4)
c. silver nitrate (AgNO_3)
d. saccharin ($\text{C}_7\text{H}_5\text{NO}_3\text{S}$)
- 113.** Determine the molar mass of allyl sulfide, the compound responsible for the smell of garlic. The chemical formula of allyl sulfide is $(\text{C}_3\text{H}_5)_2\text{S}$.

Mass-Mole Conversions (11.3)

- 114.** How many moles are in 100.0 g of each of the following compounds?
- dinitrogen oxide (N_2O)
 - methanol (CH_3OH)
- 115.** What is the mass of each of the following?
- 4.50×10^{-2} mol CuCl_2
 - 1.25×10^2 mol $\text{Ca}(\text{OH})_2$
- 116.** Determine the number of moles in each of the following.
- 1.25×10^2 g Na_2S
 - 0.145 g H_2S
- 117.** Benzoyl peroxide is a substance used as an acne medicine. What is the mass in grams of 3.50×10^{-2} moles of benzoyl peroxide ($\text{C}_{14}\text{H}_{10}\text{O}_4$)?
- 118.** Hydrofluoric acid is a substance used to etch glass. Determine the mass of 4.95×10^{25} HF molecules.
- 119.** How many moles of aluminum ions are in 45.0 g of aluminum oxide?
- 120.** How many moles of ions are in the following?
- 0.0200 g AgNO_3
 - 0.100 mol K_2CrO_4
 - 0.500 g $\text{Ba}(\text{OH})_2$
 - 1.00×10^{-9} mol Na_2CO_3

Mass-Particle Conversions (11.3)

- 121.** Calculate the values that will complete the table.

Table 11-2

Moles, Mass, and Representative Particles			
Compound	Number of moles	Mass (g)	Representative particles
Silver acetate $\text{Ag}(\text{C}_2\text{H}_3\text{O}_2)$	2.50		
Glucose $\text{C}_6\text{H}_{12}\text{O}_6$		324.0	
Benzene C_6H_6			5.65×10^{21}
Lead(II) sulfide PbS		100.0	

- 122.** How many formula units are present in 500.0 g lead(II) chloride?
- 123.** Determine the number of atoms in 3.50 g gold.
- 124.** Calculate the mass of 3.62×10^{24} molecules of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).
- 125.** Determine the number of molecules of ethanol ($\text{C}_2\text{H}_5\text{OH}$) in 47.0 g.
- 126.** What mass of iron(III) chloride contains 2.35×10^{23} chloride ions?
- 127.** How many moles of iron can be recovered from 100.0 kg Fe_3O_4 ?
- 128.** The mass of an electron is 9.11×10^{-28} g. What is the mass of a mole of electrons?
- 129.** Vinegar is 5.0% acetic acid (CH_3COOH). How many molecules of acetic acid are present in 25.0 g vinegar?
- 130.** The density of lead is 11.3 g/cm^3 . Calculate the volume of one mole lead.
- 131.** Calculate the moles of aluminum ions present in 250.0 g aluminum oxide (Al_2O_3).
- 132.** Determine the number of chloride ions in 10.75 g of magnesium chloride.
- 133.** Acetaminophen, a common aspirin substitute, has the formula $\text{C}_8\text{H}_9\text{NO}_2$. Determine the number of molecules of acetaminophen in a 500 mg tablet.
- 134.** Calculate the number of sodium ions present in 25.0 g sodium chloride.
- 135.** Determine the number of oxygen atoms present in 25.0 g carbon dioxide.

Percent Composition (11.4)

- 136.** Express the composition of each of the following as the mass percent of its elements (percent composition).
- sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$)
 - magnetite (Fe_3O_4)
 - aluminum sulfate ($\text{Al}_2(\text{SO}_4)_3$)
- 137.** Which of the following iron compounds contain the greatest percentage of iron: pyrite (FeS_2), hematite (Fe_2O_3), or siderite (FeCO_3)?
- 138.** Determine the empirical formula for each of the following compounds.
- ethylene (C_2H_4)
 - ascorbic acid ($\text{C}_6\text{H}_8\text{O}_6$)
 - naphthalene (C_{10}H_8)
- 139.** Caffeine, a stimulant found in coffee, has the chemical formula $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$.
- Calculate the molar mass of caffeine.
 - Determine the percent composition of caffeine.
- 140.** Which of the titanium-containing minerals rutile (TiO_2) or ilmenite (FeTiO_3) has the larger percent of titanium?
- 141.** Vitamin E, found in many plants, is thought to retard the aging process in humans. The formula for vitamin E is $\text{C}_{29}\text{H}_{50}\text{O}_2$. What is the percent composition of vitamin E?

142. Aspartame, an artificial sweetener, has the formula $C_{14}H_{18}N_2O_5$. Determine the percent composition of aspartame.

Empirical and Molecular Formulas (11.4)

- 143.** The hydrocarbon used in the manufacture of foam plastics is called styrene. Analysis of styrene indicates the compound is 92.25% C and 7.75% H and has a molar mass of 104 g/mol. Determine the molecular formula for styrene.
- 144.** Monosodium glutamate (MSG) is sometimes added to food to enhance flavor. Analysis determined this compound to be 35.5% C, 4.77% H, 8.29% N, 13.6% Na, and 37.9% O. What is the empirical formula for MSG?
- 145.** Determine the molecular formula for ibuprofen, a common headache remedy. Analysis of ibuprofen yields a molar mass of 206 g/mol and a percent composition of 75.7% C, 8.80% H and 15.5% O.
- 146.** Vanadium oxide is used as an industrial catalyst. The percent composition of this oxide is 56.0% vanadium and 44.0% oxygen. Determine the empirical formula for vanadium oxide.
- 147.** What is the empirical formula of a compound that contains 10.52 g Ni, 4.38 g C, and 5.10 g N?
- 148.** The Statue of Liberty turns green in air because of the formation of two copper compounds, $Cu_3(OH)_4SO_4$ and $Cu_4(OH)_6SO_4$. Determine the mass percent of copper in these compounds.
- 149.** Analysis of a compound containing chlorine and lead reveals that the compound is 59.37% lead. The molar mass of the compound is 349.0 g/mol. What is the empirical formula for the chloride? What is the molecular formula?
- 150.** Glycerol is a thick, sweet liquid obtained as a byproduct of the manufacture of soap. Its percent composition is 39.12% carbon, 8.75% hydrogen, and 52.12% oxygen. The molar mass is 92.11 g/mol. What is the molecular formula for glycerol?

The Formula for a Hydrate (11.5)

- 151.** Determine the mass percent of anhydrous sodium carbonate (Na_2CO_3) and water in sodium carbonate decahydrate ($Na_2CO_3 \cdot 10H_2O$).
- 152.** What is the formula and name of a hydrate that is 85.3% barium chloride and 14.7% water?
- 153.** Gypsum is hydrated calcium sulfate. A 4.89-g sample of this hydrate was heated, and after the water was driven off, 3.87 g anhydrous calcium sulfate remained. Determine the formula of this hydrate and name the compound.

154. The table shows data from an experiment to determine the formulas of hydrated barium chloride. Determine the formula for the hydrate and its name.

Table 11-3

Data for $BaCl_2 \cdot xH_2O$	
Mass of empty crucible	21.30 g
Mass of hydrate + crucible	31.35 g
Initial mass of hydrate	
Mass after heating 5 min	29.87 g
Mass of anhydrous solid	

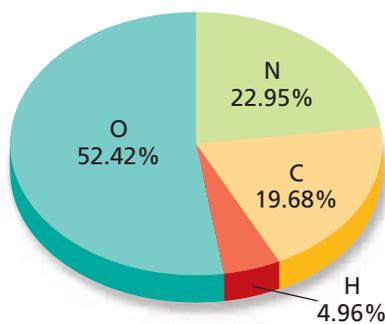
- 155.** A 1.628-g sample of a hydrate of magnesium iodide is heated until its mass is reduced to 1.072 g and all water has been removed. What is the formula of the hydrate?
- 156.** Hydrated sodium tetraborate ($Na_2B_4O_7 \cdot xH_2O$) is commonly called borax. Chemical analysis indicates that this hydrate is 52.8% sodium tetraborate and 47.2% water. Determine the formula and name the hydrate.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 157.** Determine the following:
- the number of representative particles in 3.75 g Zn
 - the mass of 4.32×10^{22} atoms Ag
 - the number of sodium ions in 25.0 g of Na_2O
- 158.** Which of the following has the greatest number of oxygen atoms?
- 17.63 g CO_2
 - 3.21×10^{22} molecules CH_3OH
 - 0.250 mol $C_6H_{12}O_6$
- 159.** Which of the following compounds has the greatest percent of oxygen by mass: TiO_2 , Fe_2O_3 , Al_2O_3 ?
- 160.** Naphthalene, commonly known as moth balls, is composed of 93.7% carbon and 6.3% hydrogen. The molar mass of naphthalene is 128 g/mol. Determine the empirical and molecular formulas for naphthalene.
- 161.** Which of the following molecular formulas are also empirical formulas: ethyl ether ($C_4H_{10}O$), aspirin ($C_9H_8O_4$), butyl dichloride ($C_4H_8Cl_2$), glucose ($C_6H_{12}O_6$).

- 162.** Calculate each of the following:
- the number of moles in 15.5 g Na_2SO_4
 - the number of formula units in 0.255 mol NaCl
 - the mass in grams of 0.775 mol SF_6
 - the number of Cl^- ions in 14.5 g MgCl_2 .
- 163.** The graph shows the percent composition of a compound containing carbon, hydrogen, oxygen, and nitrogen. How many grams of each element are present in 100 g of the compound?



- 164.** A party balloon was filled with 9.80×10^{22} atoms of helium. After 24 hours, 45% of the helium had escaped. How many atoms of helium remained?
- 165.** Tetrafluoroethylene, which is used in the production of Teflon, is composed of 24.0% carbon and 76.0% fluorine and has a molar mass of 100.0 g/mol. Determine the empirical and molecular formulas of this compound.
- 166.** Calculate the mass in grams of one atom of lead.
- 167.** Diamond is a naturally occurring form of carbon. If you have a 0.25-carat diamond, how many carbon atoms are present? (1 carat = 0.200 g)
- 168.** How many molecules of isooctane (C_8H_{18}) are present in 1.00 L? (density of isooctane = 0.680 g/mL)
- 169.** Calculate the number of molecules of water in a swimming pool which is 40.0 m in length, 20.0 m in width, and 5.00 m in depth. Assume that the density of water is 1.00 g/cm³.

Thinking Critically

- 170. Analyze and Conclude** A mining company has two possible sources of copper: chalcopyrite (CuFeS_2) and chalcocite (Cu_2S). If the mining conditions and the extraction of copper from the ore were identical for each of the ores, which ore would yield the greater quantity of copper? Explain your answer.

- 171. Designing an Experiment** Design an experiment that can be used to determine the amount of water in alum ($\text{KAl}(\text{SO}_4)_2 \cdot x\text{H}_2\text{O}$).
- 172. Concept Mapping** Design a concept map that illustrates the mole concept. Include moles, Avogadro's number, molar mass, number of particles, percent composition, empirical formula, and molecular formula.
- 173. Communicating** If you use the expression "a mole of nitrogen," is it perfectly clear what you mean, or is there more than one way to interpret the expression? Explain. How could you change the expression to make it more precise?

Writing in Chemistry

- 174.** Octane ratings are used to identify certain grades of gasoline. Research *octane rating* and prepare a pamphlet for consumers identifying the different types of gasoline, the advantages of each, and when each grade is used.
- 175.** Research the life of the Italian chemist Amedeo Avogadro (1776–1856) and how his work led scientists to the number of particles in a mole.

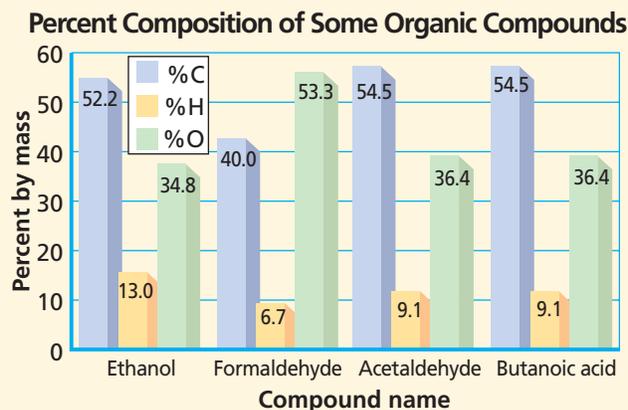
Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 176.** Express the following answers with the correct number of significant figures. (Chapter 2)
- $18.23 - 456.7$
 - $4.233 \div 0.0131$
 - $(82.44 \times 4.92) + 0.125$
- 177.** Distinguish between atomic number and mass number. How do these two numbers compare for isotopes of an element? (Chapter 4)
- 178.** Write balanced equations for the following reactions. (Chapter 10)
- Magnesium metal and water combine to form solid magnesium hydroxide and hydrogen gas.
 - Dinitrogen tetroxide gas decomposes into nitrogen dioxide gas.
 - Aqueous solutions of sulfuric acid and potassium hydroxide undergo a double replacement reaction.
- 179.** How can you tell if a chemical equation is balanced? (Chapter 10)

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Graphs Use the graph to answer questions 1–4.



- Acetaldehyde and butanoic acid must have the same _____.
 - molecular formula
 - empirical formula
 - molar mass
 - chemical properties
- If the molar mass of butanoic acid is 88.1 g/mol, then what is its molecular formula?
 - CH_2O
 - $\text{C}_2\text{H}_4\text{O}$
 - C_6HO_2
 - $\text{C}_4\text{H}_8\text{O}_2$
- What is the empirical formula of ethanol?
 - C_4HO_3
 - $\text{C}_{52}\text{H}_{13}\text{O}_{35}$
 - $\text{C}_2\text{H}_6\text{O}$
 - $\text{C}_4\text{H}_{13}\text{O}_2$
- The empirical formula of formaldehyde is the same as its molecular formula. How many grams are in 2.000 moles of formaldehyde?
 - 30.00 g
 - 60.06 g
 - 182.0 g
 - 200.0 g
- A mole is all of the following EXCEPT _____.
 - the atomic or molar mass of an element or compound
 - Avogadro's number of molecules of a compound
 - the number of atoms in exactly 12 g of pure ^{12}C
 - the SI measurement unit for the amount of a substance
- How many atoms are in 0.625 moles of Ge (atomic mass = 72.59 amu)?
 - 2.73×10^{25}
 - 6.99×10^{25}
 - 3.76×10^{23}
 - 9.63×10^{23}
- The molar mass of fluorapatite ($\text{Ca}_5(\text{PO}_4)_3\text{F}$) is _____.
 - 314 g/mol
 - 344 g/mol
 - 442 g/mol
 - 504 g/mol
- How many moles of cobalt(III) titanate (Co_2TiO_4) are in 7.13 g of the compound?
 - 2.39×10^1 mol
 - 3.14×10^{-2} mol
 - 3.22×10^1 mol
 - 4.17×10^{-2} mol
- Magnesium sulfate (MgSO_4) is often added to water-insoluble liquid products of chemical reactions to remove any unwanted water. MgSO_4 readily absorbs water to form two different hydrates. One of these hydrates is found to contain 13.0% H_2O and 87.0% MgSO_4 . What is the name of this hydrate?
 - magnesium sulfate monohydrate
 - magnesium sulfate dihydrate
 - magnesium sulfate hexahydrate
 - magnesium sulfate heptahydrate
- The mass of one molecule of barium hexafluorosilicate (BaSiF_6) is _____.
 - 1.68×10^{26} g
 - 2.16×10^{21} g
 - 4.64×10^{-22} g
 - 6.02×10^{-23} g



TEST-TAKING TIP

Your Answers Are Better Than The Test's When you know how to answer a question, answer it in your own words before looking at the answer choices. Often more than one answer choice will look good, so do the calculations first, and arm yourself with your answer before looking. If you need to, cover the choices with one hand until you're ready.

Stoichiometry

What You'll Learn

- ▶ You will write mole ratios from balanced chemical equations.
- ▶ You will calculate the number of moles and the mass of a reactant or product when given the number of moles or the mass of another reactant or product.
- ▶ You will identify the limiting reactant in a chemical reaction.
- ▶ You will determine the percent yield of a chemical reaction.

Why It's Important

The cost of the things you buy is lower because chemists use stoichiometric calculations to increase efficiency in laboratories, decrease waste in manufacturing, and produce products more quickly.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links to stoichiometry.

The candle will continue to burn as long as oxygen and candle wax are present.



DISCOVERY LAB



Materials

10-mL graduated cylinder
100-mL beaker
stirring rod
0.01M potassium permanganate (KMnO_4)
0.01M sodium hydrogen sulfite (NaHSO_3)

Observing a Chemical Reaction

Reactants are consumed in a chemical reaction as products are produced. What evidence can you observe that a reaction takes place?

Safety Precautions



Procedure

1. Measure 5.0 mL of 0.01M potassium permanganate solution (KMnO_4) and pour it into a 100-mL beaker.
2. Add 5.0 mL of 0.01M sodium hydrogen sulfite solution (NaHSO_3) to the potassium permanganate solution while stirring. Record your observations.
3. Slowly add additional 5.0-mL portions of the NaHSO_3 solution until the KMnO_4 solution turns colorless. Record your observations.
4. Record the total volume of the NaHSO_3 solution you used to cause the beaker's contents to become colorless.

Analysis

What evidence do you have that a reaction occurred? Would anything more have happened if you continued to add NaHSO_3 solution to the beaker? Explain.

Section

12.1

What is stoichiometry?

Objectives

- **Identify** the quantitative relationships in a balanced chemical equation.
- **Determine** the mole ratios from a balanced chemical equation.

Vocabulary

stoichiometry
mole ratio

Were you surprised when, in doing the **DISCOVERY LAB**, you saw the purple color of potassium permanganate disappear as you added sodium hydrogen sulfite? If you concluded that the potassium permanganate had been used up and the reaction had stopped, you are right. The photo on the opposite page shows the combustion of a candle using the oxygen in the surrounding air. What would happen if a bell jar was lowered over the burning candle blocking off the supply of oxygen? You know that oxygen is needed for the combustion of candle wax, so when the oxygen inside the bell jar is used up, the candle will go out.

Chemical reactions, such as the reaction of potassium permanganate with sodium hydrogen sulfite and the combustion of a candle, stop when one of the reactants is used up. Thus, in planning the reaction of potassium permanganate and sodium hydrogen sulfite, a chemist needs to know how many grams of potassium permanganate are needed to react completely with a known mass of sodium hydrogen sulfite. You might ask, "How much oxygen is required to completely burn a candle of known mass, or how much product will be produced if a given amount of a reactant is used?" Stoichiometry is the tool for answering these questions.

Careers Using Chemistry

Pharmacist

Are you interested in taking an active role in the health care of others? Would you like to advise physicians as well as patients? Then consider a career as a pharmacist.

Pharmacists must understand the composition and use of prescribed drugs and medicines, and over-the-counter medications. They advise doctors and patients about proper use, harmful combinations, and possible side-effects. Although pharmaceutical companies supply most medicines, pharmacists may do the actual mixing of ingredients to form powders, tablets, capsules, ointments, and solutions.



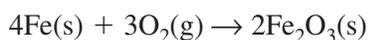
Figure 12-1

If you know the equation for this reaction between iron and oxygen, you can calculate the number of moles and the mass of each reactant and product.

Mole-Mass Relationships in Chemical Reactions

The study of quantitative relationships between amounts of reactants used and products formed by a chemical reaction is called **stoichiometry**. Stoichiometry is based on the law of conservation of mass, which was introduced by Antoine Lavoisier in the eighteenth century. The law states that matter is neither created nor destroyed in a chemical reaction. Chemical bonds in reactants break and new chemical bonds form to produce products, but the amount of matter present at the end of the reaction is the same as was present at the beginning. Therefore, the mass of the reactants equals the mass of the products.

Stoichiometry and the balanced chemical equation Look at the reaction of powdered iron with oxygen shown in **Figure 12-1**. As tiny particles of iron react with oxygen in the air, iron(III) oxide (Fe_2O_3) is produced.



You can interpret this equation in terms of representative particles by saying that four atoms of iron react with three molecules of oxygen to produce two formula units of iron(III) oxide. But, remember that coefficients in an equation represent not only numbers of individual particles but also numbers of moles of particles. Therefore, you can also say that four moles of iron react with three moles of oxygen to produce two moles of iron(III) oxide.

Does the chemical equation tell you anything about the masses of the reactants and products? Not directly. But as you learned in Chapter 11, the mass of any substance can be determined by multiplying the number of moles of the substance by the conversion factor that relates mass and number of moles, which is the molar mass. Thus, the mass of the reactants can be calculated in this way.

$$4 \cancel{\text{mol Fe}} \times \frac{55.85 \text{ g Fe}}{1 \cancel{\text{mol Fe}}} = 223.4 \text{ g Fe}$$

$$3 \cancel{\text{mol O}_2} \times \frac{32.00 \text{ g O}_2}{1 \cancel{\text{mol O}_2}} = 96.00 \text{ g O}_2$$

The total mass of the reactants = 319.4 g

Similarly, the mass of the product is

$$2 \cancel{\text{mol Fe}_2\text{O}_3} \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \cancel{\text{mol Fe}_2\text{O}_3}} = 319.4 \text{ g}$$

The total mass of the reactants equals the mass of the product, as predicted by the law of conservation of mass. **Table 12-1** summarizes the relationships that can be determined from a balanced chemical equation.

Table 12-1

Relationships Derived from a Balanced Chemical Equation			
Iron	+	Oxygen	→ Iron(III) oxide
$4\text{Fe}(\text{s})$	+	$3\text{O}_2(\text{g})$	→ $2\text{Fe}_2\text{O}_3(\text{s})$
4 atoms Fe	+	3 molecules O_2	→ 2 formula units Fe_2O_3
4 moles Fe	+	3 moles O_2	→ 2 moles Fe_2O_3
223.4 g Fe	+	96.0 g O_2	→ 319.4 g Fe_2O_3
319.4 g reactants			→ 319.4 g product



EXAMPLE PROBLEM 12-1

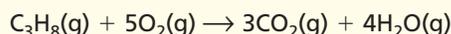
Interpreting Chemical Equations

The combustion of propane (C_3H_8) provides energy for heating homes, cooking food, and soldering metal parts. Interpret the equation for the combustion of propane in terms of representative particles, moles, and mass. Show that the law of conservation of mass is observed.

1. Analyze the Problem

The formulas in the equation represent both representative particles (molecules) and moles. Therefore, the equation can be interpreted in terms of molecules and moles. The law of conservation of mass will be verified if the masses of the reactants and products are equal.

Known



Unknown

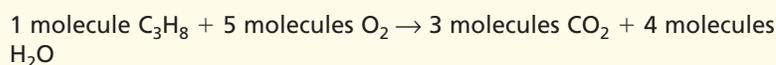
The equation in terms of molecules = ?

The equation in terms of moles = ?

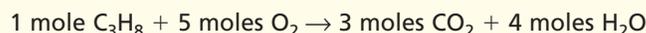
The equation in terms of mass = ?

2. Solve for the Unknown

The coefficients indicate the number of molecules.



The coefficients indicate the number of moles.



Calculate the mass of each reactant and product by multiplying the number of moles by the conversion factor molar mass.

$$\text{moles reactant or product} \times \frac{\text{grams reactant or product}}{1 \text{ mole reactant or product}} = \text{grams reactant or product}$$

$$1 \text{ mol } \text{C}_3\text{H}_8 \times \frac{44.09 \text{ g } \text{C}_3\text{H}_8}{1 \text{ mol } \text{C}_3\text{H}_8} = 44.09 \text{ g } \text{C}_3\text{H}_8$$

$$5 \text{ mol } \text{O}_2 \times \frac{32.00 \text{ g } \text{O}_2}{1 \text{ mol } \text{O}_2} = 160.0 \text{ g } \text{O}_2$$

$$3 \text{ mol } \text{CO}_2 \times \frac{44.01 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} = 132.0 \text{ g } \text{CO}_2$$

$$4 \text{ mol } \text{H}_2\text{O} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} = 72.08 \text{ g } \text{H}_2\text{O}$$

Add the masses of the reactants.

$$44.09 \text{ g } \text{C}_3\text{H}_8 + 160.0 \text{ g } \text{O}_2 = 204.1 \text{ g reactants}$$

Add the masses of the products.

$$132.0 \text{ g } \text{CO}_2 + 72.08 \text{ g } \text{H}_2\text{O} = 204.1 \text{ g products}$$

$$204.1 \text{ g reactants} = 204.1 \text{ g products}$$

The law of conservation of mass is observed.

3. Evaluate the Answer

The sums of the reactants and the products are correctly stated to the first decimal place because each mass is accurate to the first decimal place. The mass of reactants equals the mass of products as predicted by the law of conservation of mass.



Because propane gas is readily liquified, it can be stored in tanks and transported to wherever it is needed.

Practice!

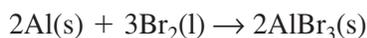
For more practice interpreting chemical equations, go to

Supplemental Practice Problems in Appendix A.

PRACTICE PROBLEMS

- Interpret the following balanced chemical equations in terms of particles, moles, and mass. Show that the law of conservation of mass is observed.
 - $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
 - $\text{HCl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $4\text{Zn}(\text{s}) + 10\text{HNO}_3(\text{aq}) \rightarrow 4\text{Zn}(\text{NO}_3)_2(\text{aq}) + \text{N}_2\text{O}(\text{g}) + 5\text{H}_2\text{O}(\text{l})$
 - $2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{MgO}(\text{s})$
 - $2\text{Na}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$

Mole ratios You have seen that the coefficients in a chemical equation indicate the relationships among moles of reactants and products. For example, return to the reaction between iron and oxygen described in **Table 12-1**. The equation indicates that four moles of iron react with three moles of oxygen. It also indicates that four moles of iron react to produce two moles of iron(III) oxide. How many moles of oxygen react to produce two moles of iron(III) oxide? You can use the relationships between coefficients to write conversion factors called mole ratios. A **mole ratio** is a ratio between the numbers of moles of any two substances in a balanced chemical equation. As another example, consider the reaction shown in **Figure 12-2**. Aluminum reacts with bromine to form aluminum bromide. Aluminum bromide is used as a catalyst to speed up a variety of chemical reactions.



What mole ratios can be written for this reaction? Starting with the reactant aluminum, you can write a mole ratio that relates the moles of aluminum to the moles of bromine. Another mole ratio shows how the moles of aluminum relate to the moles of aluminum bromide.

$$\frac{2 \text{ mol Al}}{3 \text{ mol Br}_2} \quad \text{and} \quad \frac{2 \text{ mol Al}}{2 \text{ mol AlBr}_3}$$

Two other mole ratios show how the moles of bromine relate to the moles of the other two substances in the equation, aluminum and aluminum bromide.

$$\frac{3 \text{ mol Br}_2}{2 \text{ mol Al}} \quad \text{and} \quad \frac{3 \text{ mol Br}_2}{2 \text{ mol AlBr}_3}$$

Similarly, two ratios relate the moles of aluminum bromide to the moles of aluminum and bromine.

$$\frac{2 \text{ mol AlBr}_3}{2 \text{ mol Al}} \quad \text{and} \quad \frac{2 \text{ mol AlBr}_3}{3 \text{ mol Br}_2}$$

Six ratios define all the mole relationships in this equation. Each of the three substances in the equation forms a ratio with the two other substances.

What mole ratios can be written for the decomposition of potassium chlorate (KClO_3)? This reaction is sometimes used to obtain small amounts of oxygen in the laboratory.

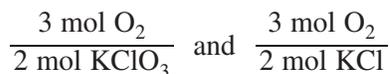
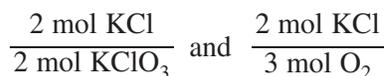
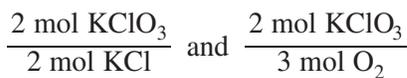


Each substance forms a mole ratio with the two other substances in the reaction. Thus, each substance should be the numerator of two mole ratios.



Figure 12-2

Bromine is one of the two elements that are liquids at room temperature. Mercury is the other. Aluminum is a lightweight metal that resists corrosion. Aluminum and bromine react vigorously to form the ionic compound aluminum bromide.



Can you explain why each reactant and product is in the numerator two times?

Six mole ratios define all the relationships in this reaction, which has three participating species. How many could you write for a reaction involving a total of four reactants and products? A simple way to find out is to multiply the number of species in the equation by the next lower number. Thus, you could write 12 mole ratios for a reaction involving four species.

PRACTICE PROBLEMS

2. Determine all possible mole ratios for the following balanced chemical equations.

- $4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)}$
- $3\text{Fe(s)} + 4\text{H}_2\text{O(l)} \rightarrow \text{Fe}_3\text{O}_4\text{(s)} + 4\text{H}_2\text{(g)}$
- $2\text{HgO(s)} \rightarrow 2\text{Hg(l)} + \text{O}_2\text{(g)}$

3. Balance the following equations and determine the possible mole ratios.

- $\text{ZnO(s)} + \text{HCl(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{O(l)}$
- butane (C_4H_{10}) + oxygen \rightarrow carbon dioxide + water

You may be wondering why you need to learn to write mole ratios. As you will see in the next section, mole ratios are the key to calculations based upon a chemical equation. Suppose you know the amount of one reactant you will use in a chemical reaction. With the chemical equation and the mole ratios, you can calculate the amount of any other reactant in the equation and the maximum amount of product you can obtain.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.



For more practice writing mole ratios, go to **Supplemental Practice Problems** in Appendix A.

Section 12.1 Assessment

- What is stoichiometry?
- List three ways in which a balanced chemical equation can be interpreted.
- What is a mole ratio?
- Thinking Critically** Write a balanced chemical equation for each reaction and determine the possible mole ratios.
 - Nitrogen reacts with hydrogen to produce ammonia.
 - Hydrogen peroxide (H_2O_2) decomposes to produce water and oxygen.
 - Pieces of zinc react with a phosphoric acid solution to produce solid zinc phosphate and hydrogen gas.
- Formulating Models** Use the balanced chemical equation to determine the mole ratios for the reaction of hydrogen and oxygen, $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O}$. Make a drawing showing six molecules of hydrogen reacting with the correct number of oxygen molecules. Show the number of molecules of water produced.



Objectives

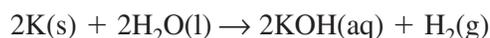
- **Explain** the sequence of steps used in solving stoichiometric problems.
- **Use** the steps to solve stoichiometric problems.

Suppose a chemist needs to obtain a certain amount of product from a reaction. How much reactant must be used? Or, suppose the chemist wants to know how much product will form if a certain amount of reactant is used. Chemists use stoichiometric calculations to answer these questions.

Using Stoichiometry

Recall that stoichiometry is the study of quantitative relationships between the amounts of reactants used and the amounts of products formed by a chemical reaction. What are the tools needed for stoichiometric calculations? All stoichiometric calculations begin with a balanced chemical equation, which indicates relative amounts of the substances that react and the products that form. Mole ratios based on the balanced chemical equation are also needed. You learned to write mole ratios in Section 12.1. Finally, mass-to-mole conversions similar to those you learned about in Chapter 11 are required.

Stoichiometric mole-to-mole conversion The vigorous reaction between potassium and water is shown in **Figure 12-3**. How can you determine the number of moles of hydrogen produced when 0.0400 mole of potassium is used? Start by writing the balanced chemical equation.



Then, identify the substance that you know and the substance that you need to determine. The given substance is 0.0400 mole of potassium. The unknown is the number of moles of hydrogen. Because the quantity of the given substance is in moles and the unknown substance is to be determined in moles, this problem is a mole-to-mole conversion.

To solve the problem, you need to know how the unknown moles of hydrogen are related to the known moles of potassium. In Section 12.1 you learned to use the balanced chemical equation to write mole ratios that describe mole relationships. Mole ratios are used as conversion factors to convert a known number of moles of one substance to moles of another substance in the same chemical reaction. What mole ratio could be used to convert moles of potassium to moles of hydrogen? In the correct mole ratio, the moles of unknown (H_2) should be the numerator and the moles of known (K) should be the denominator. The correct mole ratio is

$$\frac{1 \text{ mol H}_2}{2 \text{ mol K}}$$

This mole ratio can be used to convert the known number of moles of potassium to a number of moles of hydrogen. Remember that when you use a conversion factor, the units must cancel.

$$\text{moles of known} \times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$$

$$0.0400 \text{ mol K} \times \frac{1 \text{ mol H}_2}{2 \text{ mol K}} = 0.0200 \text{ mol H}_2$$

If you put 0.0400 mol K into water, 0.0200 mol H_2 will be produced. The **How It Works** feature at the end of this chapter shows the importance of mole ratios.



Figure 12-3

Potassium metal reacts vigorously with water, releasing so much heat that the hydrogen gas formed in the reaction catches fire.

EXAMPLE PROBLEM 12-2

Stoichiometric Mole-to-Mole Conversion

One disadvantage of burning propane (C_3H_8) is that carbon dioxide (CO_2) is one of the products. The released carbon dioxide increases the growing concentration of CO_2 in the atmosphere. How many moles of carbon dioxide are produced when 10.0 moles of propane are burned in excess oxygen in a gas grill?

1. Analyze the Problem

You are given moles of the reactant propane, and moles of the product carbon dioxide must be found. The balanced chemical equation must be written. Conversion from moles of C_3H_8 to moles of CO_2 is required. The correct mole ratio has moles of unknown substance in the numerator and moles of known substance in the denominator.

Known

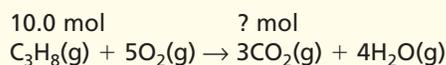
moles of propane = 10.0 mol C_3H_8

Unknown

moles of carbon dioxide = ? mol CO_2

2. Solve for the Unknown

Write the balanced chemical equation. Label the known substance and the unknown substance.



Determine the mole ratio that relates mol CO_2 to mol C_3H_8 .

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$$

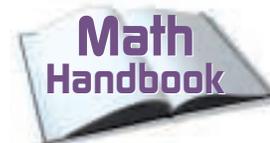
Multiply the known number of moles of C_3H_8 by the mole ratio.

$$10.0 \text{ mol } C_3H_8 \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} = 30.0 \text{ mol } CO_2$$

Burning 10.0 mol C_3H_8 produces 30.0 mol CO_2 .

3. Evaluate the Answer

The given number of moles has three significant figures. Therefore, the answer must have three digits. The balanced chemical equation indicates that 1 mol C_3H_8 produces 3 mol CO_2 . Thus, 10.0 mol C_3H_8 would produce three times as many moles of CO_2 , or 30.0 mol.



Review dimensional analysis in the **Math Handbook** on page 900 of this text.

PRACTICE PROBLEMS

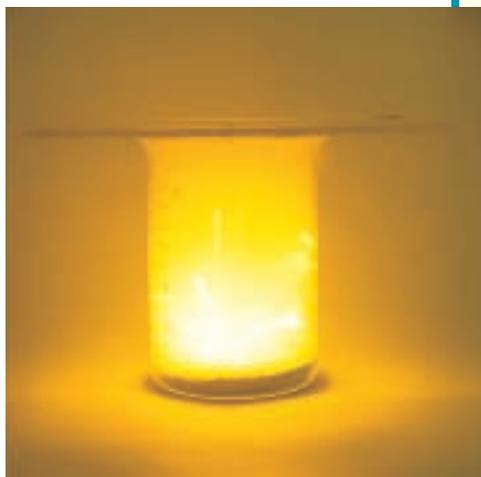
9. Sulfuric acid is formed when sulfur dioxide reacts with oxygen and water. Write the balanced chemical equation for the reaction. If 12.5 mol SO_2 reacts, how many mol H_2SO_4 can be produced? How many mol O_2 is needed?
10. A reaction between methane and sulfur produces carbon disulfide (CS_2), a liquid often used in the production of cellophane.
- $$\text{___ } CH_4(g) + \text{___ } S_8(s) \rightarrow \text{___ } CS_2(l) + \text{___ } H_2S(g)$$
- Balance the equation.
 - Calculate the mol CS_2 produced when 1.50 mol S_8 is used.
 - How many mol H_2S is produced?



For more practice converting from moles of one substance to moles of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

Stoichiometric mole-to-mass conversion Now, suppose you know the number of moles of a reactant or product in a reaction and you want to calculate the mass of another product or reactant. This situation is an example of a mole-to-mass conversion.

EXAMPLE PROBLEM 12-3



The reaction of sodium and chlorine to form sodium chloride releases a large amount of energy in the form of light and heat. It should not surprise you, then, that a large amount of energy is required to decompose sodium chloride.

Stoichiometric Mole-to-Mass Conversion

Determine the mass of sodium chloride or table salt (NaCl) produced when 1.25 moles of chlorine gas reacts vigorously with sodium.

1. Analyze the Problem

You are given the moles of the reactant Cl_2 and must determine the mass of the product NaCl. You must convert from moles of Cl_2 to moles of NaCl using the mole ratio from the equation. Then, you need to convert moles of NaCl to grams of NaCl using the molar mass as the conversion factor.

Known

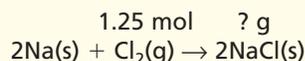
moles of chlorine = 1.25 mol Cl_2

Unknown

mass of sodium chloride = ? g NaCl

2. Solve for the Unknown

Write the balanced chemical equation and identify the known and unknown substances.



Write the mole ratio that relates mol NaCl to mol Cl_2 .

$$\frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2}$$

Multiply the number of moles of Cl_2 by the mole ratio.

$$1.25 \text{ mol Cl}_2 \times \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2} = 2.50 \text{ mol NaCl}$$

Multiply mol NaCl by the molar mass of NaCl.

$$2.50 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 146 \text{ g NaCl}$$

3. Evaluate the Answer

The given number of moles has three significant figures, so the mass of NaCl is correctly stated with three digits. The computations are correct and the unit is as expected.

PRACTICE PROBLEMS

- Titanium is a transition metal used in many alloys because it is extremely strong and lightweight. Titanium tetrachloride (TiCl_4) is extracted from titanium oxide using chlorine and coke (carbon).
 $\text{TiO}_2\text{(s)} + \text{C(s)} + 2\text{Cl}_2\text{(g)} \rightarrow \text{TiCl}_4\text{(s)} + \text{CO}_2\text{(g)}$
 If you begin with 1.25 mol TiO_2 , what mass of Cl_2 gas is needed?
- Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How many grams of chlorine gas can be obtained from 2.50 mol NaCl?

Practice!

For more practice converting from moles of one substance to mass of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

Stoichiometric mass-to-mass conversion If you were preparing to carry out a chemical reaction in the laboratory, you would need to know how much of each reactant to use in order to produce the mass of product you required. Example Problem 12-4 will demonstrate how you can use a measured mass of the known substance, the balanced chemical equation, and mole ratios from the equation to find the mass of the unknown substance. The **CHEMLAB** at the end of this chapter will provide you with laboratory experience determining a mole ratio.

EXAMPLE PROBLEM 12-4

Stoichiometric Mass-to-Mass Conversion

Ammonium nitrate (NH_4NO_3), an important fertilizer, produces N_2O gas and H_2O when it decomposes. Determine the mass of water produced from the decomposition of 25.0 g of solid ammonium nitrate.

1. Analyze the Problem

You are given the mass of the reactant and will need to write the balanced chemical equation. You then must convert from the mass of the reactant to moles of the reactant. You will next use a mole ratio to relate moles of the reactant to moles of the product. Finally, you will use the molar mass to convert from moles of the product to the mass of the product.

Known

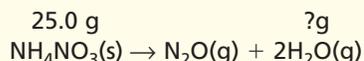
mass of ammonium nitrate = 25.0 g NH_4NO_3

Unknown

mass of water = ? g H_2O

2. Solve for the Unknown

Write the balanced chemical equation for the reaction and identify the known and unknown substances.



Convert grams of NH_4NO_3 to moles of NH_4NO_3 using the inverse of molar mass as the conversion factor.

$$25.0 \text{ g } \cancel{\text{NH}_4\text{NO}_3} \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{80.04 \text{ g } \cancel{\text{NH}_4\text{NO}_3}} = 0.312 \text{ mol } \text{NH}_4\text{NO}_3$$

Determine from the equation the mole ratio of mol H_2O to mol NH_4NO_3 . The unknown quantity is the numerator.

$$\frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{NH}_4\text{NO}_3}$$

Multiply mol NH_4NO_3 by the mole ratio.

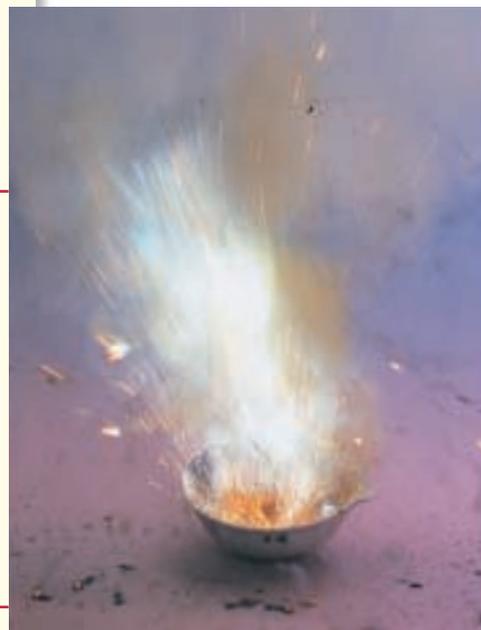
$$0.312 \text{ mol } \cancel{\text{NH}_4\text{NO}_3} \times \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \cancel{\text{NH}_4\text{NO}_3}} = 0.624 \text{ mol } \text{H}_2\text{O}$$

Calculate the mass of H_2O using molar mass as the conversion factor.

$$0.624 \text{ mol } \cancel{\text{H}_2\text{O}} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \cancel{\text{H}_2\text{O}}} = 11.2 \text{ g } \text{H}_2\text{O}$$

3. Evaluate the Answer

The number of significant figures in the answer, three, is determined by the given moles of ammonium nitrate. The calculations are correct and the unit is appropriate.



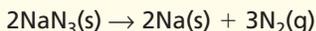
The vigorous decomposition of ammonium nitrate attests to its use in explosives. However, ammonium nitrate is also widely used as a fertilizer because it is 100 percent available for plant use.

Practice!

For more practice converting from mass of one substance to mass of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- 13.** One in a series of reactions that inflate air bags in automobiles is the decomposition of sodium azide (NaN_3).



Determine the mass of N_2 produced if 100.0 g NaN_3 is decomposed.

- 14.** In the formation of acid rain, sulfur dioxide reacts with oxygen and water in the air to form sulfuric acid. Write the balanced chemical equation for the reaction. If 2.50 g SO_2 react with excess oxygen and water, how many grams of H_2SO_4 are produced?

The steps you followed in Example Problem 12-4 are illustrated in **Figure 12-4** and described below it. Use the steps as a guide when you do stoichiometric calculations until you become thoroughly familiar with the procedure. Study **Figure 12-4** as you read.

The specified unit of the given substance determines at what point you will start your calculations. If the amount of the given substance is in moles, step 2 is omitted and step 3, mole-to-mole conversion, becomes the starting point for the calculations. However, if mass is the starting unit, calculations begin with step 2. The end point of the calculation depends upon the specified unit of the unknown substance. If the answer is to be obtained in moles, the calculation is finished with step 3. If the mass of the unknown is to be determined, you must go on to step 4.

Like any other type of problem, stoichiometric calculations require practice. You can begin to practice your skills in the **miniLAB** that follows.

miniLAB

Baking Soda Stoichiometry

Predicting When baking soda is an ingredient in your recipe, its purpose is to make the batter rise and produce a product with a light and fluffy texture. That's because baking soda, or sodium hydrogen carbonate (NaHCO_3), decomposes upon heating to form carbon dioxide gas.



Predict how much sodium carbonate (Na_2CO_3) is produced when baking soda decomposes.

Materials ring stand, ring, clay triangle, crucible, crucible tongs, Bunsen burner, balance, 3.0 g baking soda (NaHCO_3)

Procedure



1. Measure the mass of a clean, dry crucible. Add about 3.0 g of NaHCO_3 and measure the combined mass of the crucible and NaHCO_3 . Record both masses in your data table and calculate the mass of the NaHCO_3 .

2. Use this starting mass of baking soda and the balanced chemical equation to calculate the mass of Na_2CO_3 that will be produced.
3. Set up a ring stand with a ring and clay triangle for heating the crucible.
4. Heat the crucible slowly at first and then with a stronger flame for 7–8 min. Use tongs to remove the hot crucible. Record your observations during the heating.
5. Allow the crucible to cool and then obtain the mass of the crucible and sodium carbonate.

Analysis

1. What were your observations during the heating of the baking soda?
2. How did your calculated mass of sodium carbonate compare with the actual mass you obtained from the experiment? If the two masses are different, suggest reasons for the difference.

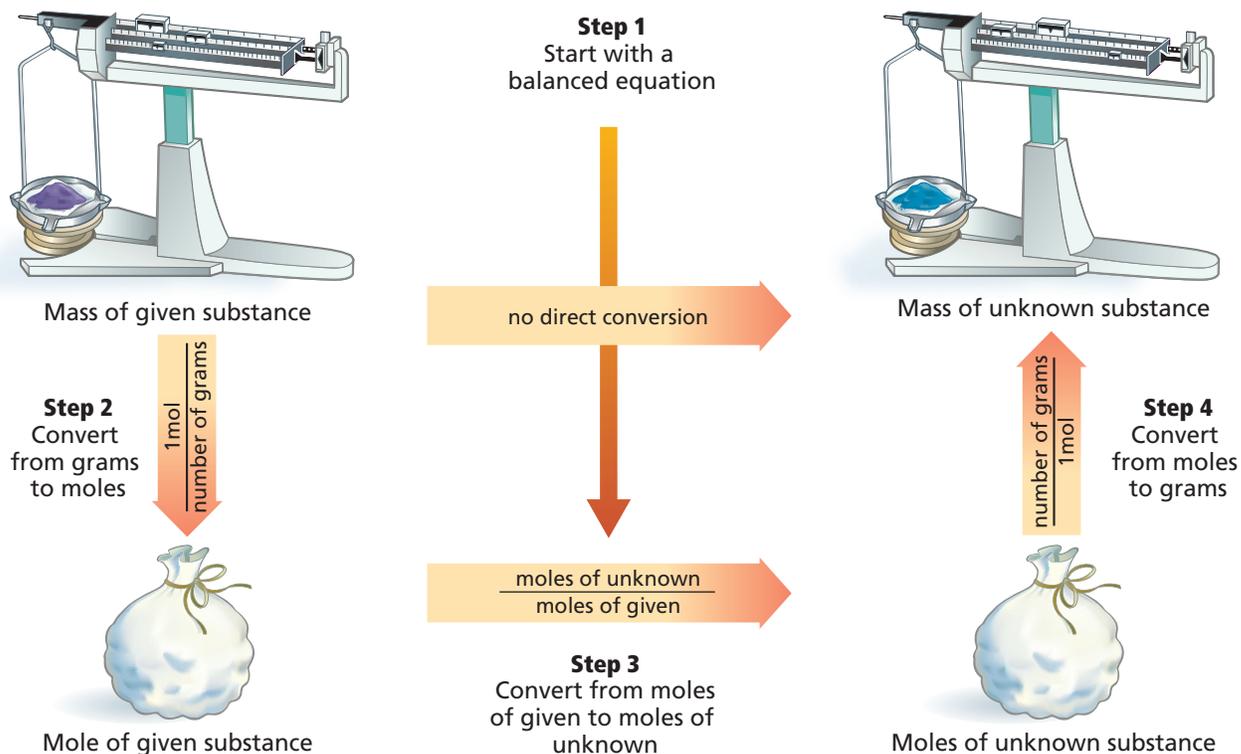


Figure 12-4

Follow the steps from the balanced equation to the mass of the unknown. Note that there is no shortcut from the mass of the given substance to the mass of the unknown substance. The route goes through the mole. However, you can follow the arrow from step 1 to step 3 if the amount of the given substance is in moles.

Steps in Stoichiometric Calculations

1. Write a balanced chemical equation. Interpret the equation in terms of moles.
2. Determine the moles of the given substance using a mass-to-mole conversion. Use the inverse of the molar mass as the conversion factor.
3. Determine the moles of the unknown substance from the moles of the given substance. Use the appropriate mole ratio from the balanced chemical equation as the conversion factor.
4. From the moles of the unknown substance, determine the mass of the unknown substance using a mole-to-mass conversion. Use the molar mass as the conversion factor.

Section 12.2 Assessment

15. Why is a balanced chemical equation needed in solving stoichiometric calculations?
16. When solving stoichiometric problems, how is the correct mole ratio expressed?
17. List the four steps used in solving stoichiometric problems.
18. **Thinking Critically** In a certain industrial process, magnesium reacts with liquid bromine. How would a chemical engineer determine the mass of bromine needed to react completely with a given mass of magnesium?
19. **Concept Mapping** Many cities use calcium chloride to prevent ice from forming on roadways. To produce calcium chloride, calcium carbonate (limestone) is reacted with hydrochloric acid according to this equation.

$$\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$
 Create a concept map that describes how you can determine the mass of calcium chloride produced if the mass of hydrochloric acid is given.

Objectives

- **Identify** the limiting reactant in a chemical equation.
- **Identify** the excess reactant and **calculate** the amount remaining after the reaction is complete.
- **Calculate** the mass of a product when the amounts of more than one reactant are given.

Vocabulary

limiting reactant
excess reactant

At a school dance, the music begins and boys and girls pair up to dance. If there are more boys than girls, some boys will be left without partners. The same is true of reactants in a chemical reaction. Rarely in nature are reactants in a chemical reaction present in the exact ratios specified by the balanced equation. Generally, one or more reactants are in excess and the reaction proceeds until all of one reactant is used up.

Why do reactions stop?

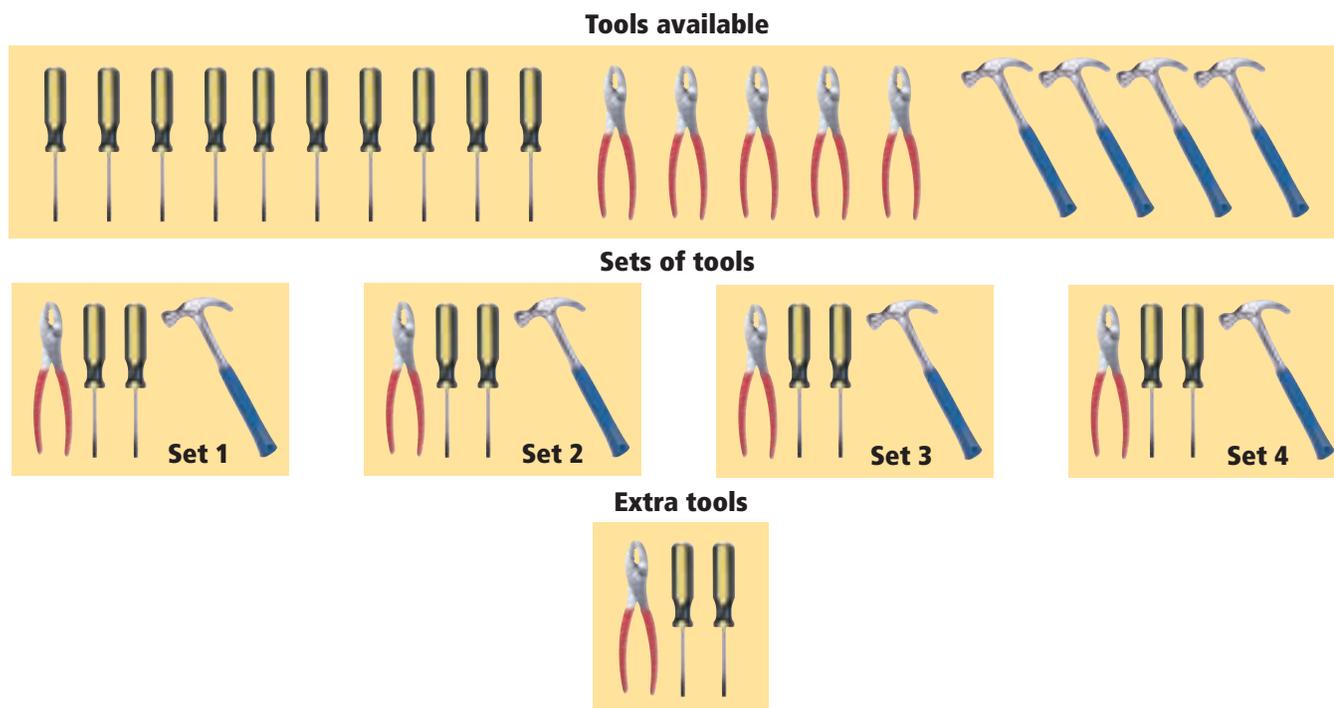
When a chemical reaction is carried out in the laboratory, the same principle applies. Usually, one or more reactants are in excess, while one is limited. The amount of product depends upon the reactant that is limited.

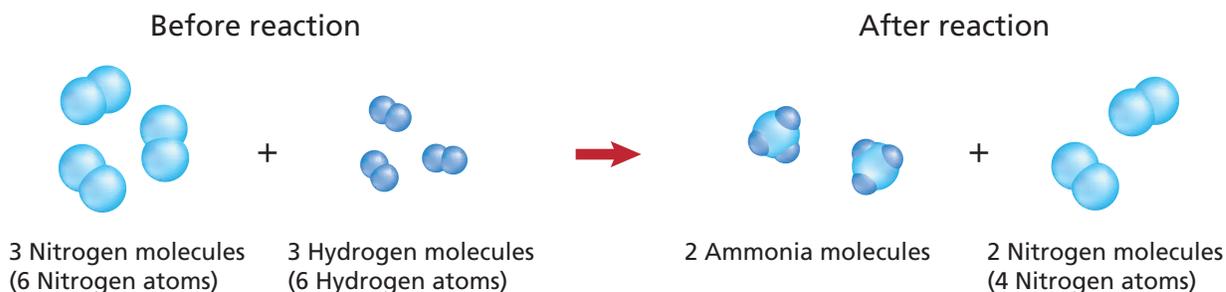
Remember the reaction between potassium permanganate and sodium hydrogen sulfite in the **DISCOVERY LAB**. As you added colorless sodium hydrogen sulfite to purple potassium permanganate, the color faded as a reaction took place. Finally, the solution was colorless. You could have continued adding sodium hydrogen sulfite, but would any further reaction have taken place? You are correct if you said that no further reaction could take place because no potassium permanganate was available to react. Potassium permanganate was a limiting reactant. As the name implies, the **limiting reactant** limits the extent of the reaction and, thereby, determines the amount of product. A portion of all of the other reactants remains after the reaction stops. These left-over reactants are called **excess reactants**. What was the excess reactant in the reaction of potassium permanganate and sodium hydrogen sulfite?

To help you understand limiting reactants, consider the analogy in **Figure 12-5**. How many tool sets can be assembled from the items shown if each complete tool set consists of one pair of pliers, one hammer, and two screwdrivers? You can see that four complete tool sets can be assembled. The number of tool sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess. Chemical reactions work in a similar way.

Figure 12-5

Each tool set must have one hammer so only four sets can be assembled. Which tool is limiting? Which tools are in excess?





The calculations you did in Section 12.2 were based on having the reactants present in the ratio described by the balanced chemical equation. How can you calculate the amount of product formed when one reactant limits the amount of product and the other is in excess? The first thing you must do is determine which reactant is the limiting reactant.

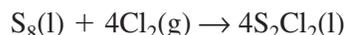
Consider the reaction shown in **Figure 12-6** in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). You can visualize that in the first step of the reaction, all the nitrogen molecules and hydrogen molecules are separated into individual atoms. These are the atoms available for reassembling into ammonia molecules just like the tools in **Figure 12-5** before they were assembled into tool kits. How many molecules of ammonia will be produced from the available atoms? Four tool kits could be assembled from the tools because only four hammers were available. Two ammonia molecules can be assembled from the hydrogen and nitrogen atoms because only six hydrogen atoms are available, three for each ammonia molecule. When the hydrogen is gone, two molecules of nitrogen remain unreacted. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It's important to know which reactant is the limiting reactant because, as you have just learned, the amount of product formed depends upon this reactant.

Figure 12-6

Refer to **Table C-1** in Appendix C for a key to atom color conventions. Check to see whether all the atoms present before the reaction are present after the reaction. Some nitrogen molecules were unchanged in the reaction. Which reactant is in excess?

Calculating the Product When a Reactant Is Limited

How can you determine which reactant is limited? As an example, consider the formation of disulfur dichloride (S_2Cl_2). Disulfur dichloride is used to vulcanize rubber, a process that makes rubber harder, stronger, and less likely to become soft when hot or brittle when cold. In the production of disulfur dichloride, molten sulfur reacts with chlorine gas according to this equation.



If 200.0 g of sulfur reacts with 100.0 g of chlorine, what mass of disulfur dichloride is produced?

Masses of both reactants are given. You must first determine which one is the limiting reactant because the reaction will stop producing product when the limiting reactant is used up. Identifying the limiting reactant involves finding the number of moles of each reactant. This is done by converting the masses of chlorine and sulfur to moles. Multiply each mass by the conversion factor that relates moles and mass, the inverse of the molar mass.

$$100.0 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{70.91 \text{ g } Cl_2} = 1.410 \text{ mol } Cl_2$$

$$200.0 \text{ g } S_8 \times \frac{1 \text{ mol } S_8}{256.5 \text{ g } S_8} = 0.7797 \text{ mol } S_8$$

CONNECTION

Vitamins and minerals are needed in small amounts in your daily diet. These substances are essential to the reactions in normal metabolism, growth and development, and regulation of the functioning of your body's cells.

Vitamin B-12 is necessary for red blood cells to mature properly in the bone marrow. People with pernicious anemia lose their ability to make intrinsic factor, a substance secreted by the stomach's membrane lining that enables vitamin B-12 to be absorbed from the intestine. Without an adequate amount of vitamin B-12, the body is unable to synthesize DNA properly, which affects red blood cell production.



The next step involves determining whether the two reactants are in the correct mole ratio as given in the balanced chemical equation. The coefficients in the balanced chemical equation indicate that four moles of chlorine are needed to react with one mole of sulfur. This 4:1 ratio from the equation must be compared with the actual ratio of the moles of available reactants just calculated above. To determine the actual ratio of moles, divide the available moles of chlorine by the available moles of sulfur.

$$\frac{1.410 \text{ mol Cl}_2 \text{ available}}{0.7797 \text{ mol S}_8 \text{ available}} = \frac{1.808 \text{ mol Cl}_2 \text{ available}}{1 \text{ mol S}_8 \text{ available}}$$

Only 1.808 mol of chlorine is actually available for every 1 mol of sulfur instead of the 4 mol of chlorine required by the balanced chemical equation. Therefore, chlorine is the limiting reactant.

After the limiting reactant has been determined, the amount of product in moles can be calculated by multiplying the given number of moles of the limiting reactant (1.410 mol Cl₂) by the mole ratio that relates disulfur dichloride and chlorine.

$$1.410 \text{ mol Cl}_2 \times \frac{4 \text{ mol S}_2\text{Cl}_2}{4 \text{ mol Cl}_2} = 1.410 \text{ mol S}_2\text{Cl}_2$$

Then, moles of S₂Cl₂ is converted to grams of S₂Cl₂ by multiplying by the conversion factor that relates mass and moles, molar mass.

$$1.410 \text{ mol S}_2\text{Cl}_2 \times \frac{135.0 \text{ g S}_2\text{Cl}_2}{1 \text{ mol S}_2\text{Cl}_2} = 190.4 \text{ g S}_2\text{Cl}_2$$

These two calculations can be combined into one like this.

$$1.410 \text{ mol Cl}_2 \times \frac{4 \text{ mol S}_2\text{Cl}_2}{4 \text{ mol Cl}_2} \times \frac{135.0 \text{ g S}_2\text{Cl}_2}{1 \text{ mol S}_2\text{Cl}_2} = 190.4 \text{ g S}_2\text{Cl}_2$$

Now you know that 190.4 g S₂Cl₂ is produced when 1.410 mol Cl₂ reacts with an excess of S₈.

What about the reactant sulfur, which you know is in excess? How much of it actually reacted? You can calculate the mass of sulfur needed to react completely with 1.410 mol of chlorine using a mole-to-mass calculation. The first step is to multiply the moles of chlorine by the mole ratio of sulfur to chlorine to obtain the number of moles of sulfur. Remember, the unknown is the numerator and the known is the denominator.

$$1.410 \text{ mol Cl}_2 \times \frac{1 \text{ mol S}_8}{4 \text{ mol Cl}_2} = 0.3525 \text{ mol S}_8$$

Now, to obtain the mass of sulfur needed, 0.3525 mol S₈ is multiplied by the conversion factor that relates mass and moles, molar mass.

$$0.3525 \text{ mol S}_8 \times \frac{256.5 \text{ g S}_8}{1 \text{ mol S}_8} = 90.42 \text{ g S}_8 \text{ needed}$$

Knowing that 90.42 g S₈ is needed, you can calculate the amount of sulfur left unreacted when the reaction ends. Because 200.0 g of sulfur is available and only 90.42 g is needed, the mass in excess is

$$200.0 \text{ g S}_8 \text{ available} - 90.42 \text{ g S}_8 \text{ needed} = 109.6 \text{ g S}_8 \text{ in excess.}$$

EXAMPLE PROBLEM 12-5

Determining the Limiting Reactant

The reaction between solid white phosphorus and oxygen produces solid tetraphosphorus decoxide (P_4O_{10}). This compound is often called diphosphorus pentoxide because its empirical formula is P_2O_5 .

- Determine the mass of tetraphosphorus decoxide formed if 25.0 g of phosphorus (P_4) and 50.0 g of oxygen are combined.
- How much of the excess reactant remains after the reaction stops?

1. Analyze the Problem

You are given the masses of both reactants so the limiting reactant must be identified and used for finding the mass of product. From moles of limiting reactant, the moles of the excess reactant used in the reaction can be determined. The number of moles of excess reactant that actually reacted can be converted to mass and subtracted from the given mass to find the amount in excess.

Known

mass of phosphorus = 25.0 g P_4

mass of oxygen = 50.0 g O_2

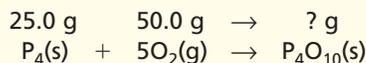
Unknown

mass of tetraphosphorus decoxide = ? g P_4O_{10}

mass of excess reactant = ? g excess reactant

2. Solve for the Unknown

- Write the balanced chemical equation and identify the knowns and the unknown.



Determine the number of moles of the reactants by multiplying each mass by the conversion factor that relates moles and mass, the inverse of molar mass.

$$25.0 \text{ g } P_4 \times \frac{1 \text{ mol } P_4}{123.9 \text{ g } P_4} = 0.202 \text{ mol } P_4$$

$$50.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 1.56 \text{ mol } O_2$$

Calculate the actual ratio of available moles of O_2 and available moles of P_4 .

$$\frac{1.56 \text{ mol } O_2}{0.202 \text{ mol } P_4} = \frac{7.72 \text{ mol } O_2}{1 \text{ mol } P_4}$$

Determine the mole ratio of the two reactants from the balanced chemical equation.

$$\frac{5 \text{ mol } O_2}{1 \text{ mol } P_4}$$

Because 7.72 mol O_2 is available but only 5 mol is needed to react with 1 mol P_4 , O_2 is in excess and P_4 is the limiting reactant. Use the moles of P_4 to determine the moles of P_4O_{10} that will be produced.

Multiply the number of moles of P_4 by the mole ratio of P_4O_{10} (the unknown) to P_4 (the known).

$$0.202 \text{ mol } P_4 \times \frac{1 \text{ mol } P_4O_{10}}{1 \text{ mol } P_4} = 0.202 \text{ mol } P_4O_{10}$$

Continued on next page



Tiny pieces of white phosphorus deposited on filter paper burst into flame on contact with air. That's why white phosphorus is never found free in nature. Phosphorus is an essential element in living systems; for example, phosphate groups occur regularly along strands of DNA.

To calculate the mass of P_4O_{10} , multiply moles of P_4O_{10} by the conversion factor that relates mass and moles, molar mass.

$$0.202 \text{ mol } \cancel{\text{P}_4\text{O}_{10}} \times \frac{283.9 \text{ g } \text{P}_4\text{O}_{10}}{1 \text{ mol } \cancel{\text{P}_4\text{O}_{10}}} = 57.3 \text{ g } \text{P}_4\text{O}_{10}$$

- b. Because O_2 is in excess, only part of the available O_2 is consumed. Use the limiting reactant, P_4 , to determine the moles and mass of O_2 used.

$$0.202 \text{ mol } \cancel{\text{P}_4} \times \frac{5 \text{ mol } \text{O}_2}{1 \text{ mol } \cancel{\text{P}_4}} = 1.01 \text{ mol } \text{O}_2 \text{ (moles needed)}$$

Multiply moles of O_2 by the conversion factor that relates mass and moles, molar mass.

$$1.01 \text{ mol } \cancel{\text{O}_2} \times \frac{32.00 \text{ g } \text{O}_2}{1 \text{ mol } \cancel{\text{O}_2}} = 32.3 \text{ g } \text{O}_2 \text{ (mass needed)}$$

Subtract the mass of O_2 needed from the mass available to calculate excess O_2 .

$$50.0 \text{ g } \text{O}_2 \text{ available} - 32.3 \text{ g } \text{O}_2 \text{ needed} = 17.7 \text{ g } \text{O}_2 \text{ in excess}$$

3. Evaluate the Answer

All values have a minimum of three significant figures, so the mass of P_4O_{10} is correctly stated with three digits. The mass of excess O_2 (17.7 g) is found by subtracting two numbers that are accurate to the first decimal place. Therefore, the mass of excess O_2 correctly shows one decimal place. The sum of the oxygen that was consumed (32.3 g) and the given mass of phosphorus (25.0 g) is 57.3 g, the calculated mass of the product phosphorus decoxide.

Practice!

For more practice determining the limiting reactant in a chemical reaction, go to

Supplemental Practice Problems in Appendix A.

PRACTICE PROBLEMS

20. The reaction between solid sodium and iron(III) oxide is one in a series of reactions that inflates an automobile airbag.
- $$6\text{Na(s)} + \text{Fe}_2\text{O}_3\text{(s)} \rightarrow 3\text{Na}_2\text{O(s)} + 2\text{Fe(s)}$$
- If 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine
- the limiting reactant.
 - the reactant in excess.
 - the mass of solid iron produced.
 - the mass of excess reactant that remains after the reaction is complete.
21. Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) and oxygen. Write the balanced chemical equation for the reaction. If a plant has 88.0 g carbon dioxide and 64.0 g water available for photosynthesis, determine
- the limiting reactant.
 - the excess reactant and the mass in excess.
 - the mass of glucose produced.

Why use an excess of a reactant? Why are reactions usually not carried out using amounts of reactants in the exact mole ratios given in the balanced equation? Some reactions do not continue until all the reactants are used up. Instead, they appear to stop while portions of the reactants are still present in the reaction mixture. Because this is inefficient and wasteful, chemists have found that by using an excess of one reactant—often the least expensive

one—reactions can be driven to continue until all of the limiting reactant is used up. Using an excess of one reactant can also speed up a reaction.

In **Figure 12-7**, you can see an example of how controlling the amount of a reactant can increase efficiency. Your school laboratory may have the kind of Bunsen burner shown in the figure. If so, you probably know that this type of burner has a control that can vary the amount of air (oxygen) that mixes with the gas. How efficiently the burner operates depends upon the ratio of oxygen to methane gas in the fuel mixture. When the amount of air is limited, the resulting flame is yellow because of glowing bits of unburned fuel, which deposit on glassware as soot (carbon). Fuel is wasted because the amount of energy released is less than the amount that could have been produced if enough oxygen were available. When sufficient oxygen is present in the combustion mixture, the burner produces a hot, intense blue flame. No soot is deposited because the fuel is completely converted to carbon dioxide and water vapor.

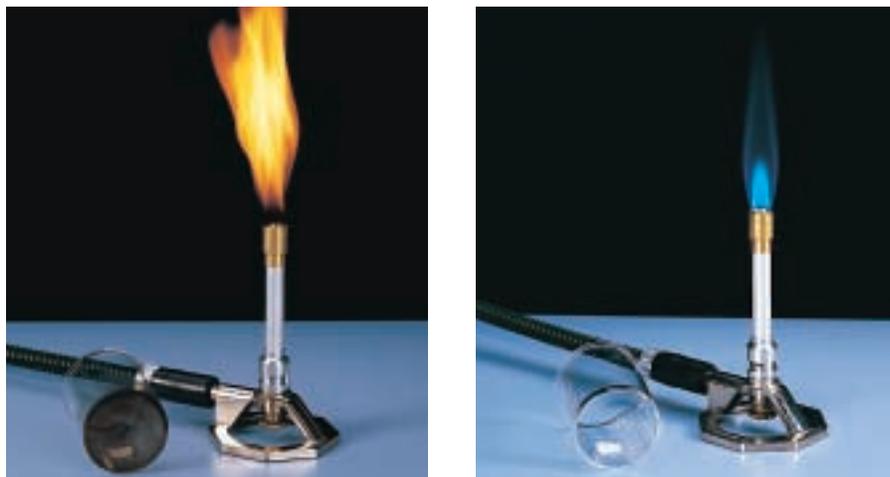


Figure 12-7

With insufficient oxygen, the burner on the left burns with a yellow, sooty flame. The burner on the right burns hot and clean because an excess of oxygen is available to react completely with the methane gas.

Section 12.3 Assessment

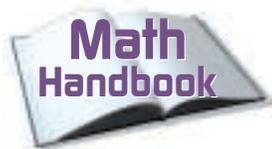
- 22.** What is meant by the limiting reactant? Why is it necessary to identify the limiting reactant when you want to know how much product will form in a chemical reaction?
- 23.** Describe how the mass of the product can be calculated when one reactant is in excess.
- 24.** Are limiting reactants present in all reactions? Explain.
- 25. Thinking Critically** For the following reactions, identify the limiting reactant and the excess reactant. Give reasons for your choices.
- wood burning in a campfire
 - sulfur in the air reacting with silver flatware to produce tarnish, or silver sulfide
 - baking powder in cake batter decomposing to produce carbon dioxide, which makes the cake rise
- 26. Analyze and Conclude** The equation representing the production of tetraphosphorus trisulfide (P_4S_3), a substance used in some match heads, is
- $$8P_4 + 3S_8 \rightarrow 8P_4S_3$$
- Determine if each of the following statements is correct. If the statement is incorrect, rewrite it to make it correct.
- To produce 4 mol P_4S_3 , 4 mol P_4 must react with 1.5 mol S_8 .
 - When 4 mol P_4 reacts with 4 mol S_8 , sulfur is the limiting reactant.
 - When 6 mol P_4 and 6 mol S_8 react, 1320 g P_4S_3 is produced.

Objectives

- **Calculate** the theoretical yield of a chemical reaction from data.
- **Determine** the percent yield for a chemical reaction.

Vocabulary

theoretical yield
actual yield
percent yield



Review the meaning and calculation of percents in the **Math Handbook** on page 907 in this text.

Suppose you were determined to improve your jump shot and took time each afternoon to practice. One afternoon, you succeeded in getting the ball through the hoop 49 times out of a total of 75 tries. Theoretically, you could have been successful 75 times, but in actuality that usually doesn't happen. (That's why you practice.) But how successful were you? You could calculate your efficiency as a percent by dividing the number of successful tries by the total number of tries and multiplying by 100.

$$\frac{49 \text{ actual successes}}{75 \text{ theoretical successes}} \times 100 = 65\% \text{ successful shots}$$

Sixty-five percent successful jump shots means that you could expect to get the ball into the basket 65 times if you made 100 attempts.

Similar calculations are made to determine the success of chemical reactions because most reactions never succeed in producing the predicted amount of product. Although your work with stoichiometric problems so far may have led you to think that chemical reactions proceed according to the balanced equation without any difficulties and always produce the calculated amount of product, this is not the case! Not every reaction goes cleanly or completely. Many reactions stop before all of the reactants are used up, so the actual amount of product is less than expected. Liquid reactants or products may adhere to the surfaces of containers or evaporate, and solid product is always left behind on filter paper or lost in the purification process. In some instances, products other than the intended ones may be formed by competing reactions, thus reducing the yield of the desired product.

How much product?

In many of the calculations you have been practicing, you have been asked to calculate the amount of product that can be produced from a given amount of reactant. The answer you obtained is called the theoretical yield of the reaction. The **theoretical yield** is the maximum amount of product that can be produced from a given amount of reactant. A chemical reaction rarely produces the theoretical yield of product. A chemist determines the actual yield of a reaction through a careful experiment in which the mass of the product is measured. The **actual yield** is the amount of product actually produced when the chemical reaction is carried out in an experiment.

Chemists need to know how efficient a reaction is in producing the desired product. One way of measuring efficiency is by means of percent yield. Just as you calculated your percent of successful jump shots, a chemist can calculate what percent of the amount of product that could theoretically be produced was actually produced. **Percent yield** of product is the ratio of the actual yield to the theoretical yield expressed as a percent.

$$\text{Percent yield} = \frac{\text{actual yield (from an experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100$$

The **problem-solving LAB** on page 372 will help you understand the importance of percent yield in chemical reactions and the kind of factors that may determine the size of the percent yield.

EXAMPLE PROBLEM 12-6

Calculating Percent Yield

When potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g silver nitrate ($AgNO_3$), solid silver chromate (Ag_2CrO_4) is formed.

- Determine the theoretical yield of the silver chromate precipitate.
- If 0.455 g of silver chromate is obtained, calculate the percent yield.

1. Analyze the Problem

You are given the mass of the reactant $AgNO_3$ and the actual yield of the product Ag_2CrO_4 . You need to write the balanced chemical equation and calculate the theoretical yield by making these conversions: grams of silver nitrate to moles of silver nitrate, moles of silver nitrate to moles of silver chromate, moles of silver chromate to grams of silver chromate. The percent yield can be calculated from the actual yield of product and the calculated theoretical yield.

Known

mass of silver nitrate = 0.500 g $AgNO_3$

actual yield = 0.455 g Ag_2CrO_4

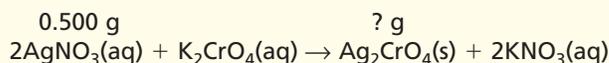
Unknown

theoretical yield = ? g Ag_2CrO_4

percent yield = ? % Ag_2CrO_4

2. Solve for the Unknown

Write the balanced chemical equation and indicate the known and unknown quantities.



Convert grams of $AgNO_3$ to moles of $AgNO_3$ using the inverse of molar mass.

$$0.500 \text{ g } AgNO_3 \times \frac{1 \text{ mol } AgNO_3}{169.9 \text{ g } AgNO_3} = 2.94 \times 10^{-3} \text{ mol } AgNO_3$$

Use the appropriate mole ratio to convert mol $AgNO_3$ to mol Ag_2CrO_4 .

$$2.94 \times 10^{-3} \text{ mol } AgNO_3 \times \frac{1 \text{ mol } Ag_2CrO_4}{2 \text{ mol } AgNO_3} = 1.47 \times 10^{-3} \text{ mol } Ag_2CrO_4$$

Calculate the mass of Ag_2CrO_4 (the theoretical yield) by multiplying mol Ag_2CrO_4 by the molar mass.

$$1.47 \times 10^{-3} \text{ mol } Ag_2CrO_4 \times \frac{331.7 \text{ g } Ag_2CrO_4}{1 \text{ mol } Ag_2CrO_4} = 0.488 \text{ g } Ag_2CrO_4$$

Divide the actual yield by the theoretical yield and multiply by 100.

$$\frac{0.455 \text{ g } Ag_2CrO_4}{0.488 \text{ g } Ag_2CrO_4} \times 100 = 93.2\% \text{ } Ag_2CrO_4$$

3. Evaluate the Answer

All quantities have three significant figures so the percent is correctly stated with three digits. The molar mass of Ag_2CrO_4 is about twice the molar mass of $AgNO_3$, and the ratio of mol $AgNO_3$ to mol Ag_2CrO_4 in the equation is 2:1. Therefore, 0.500 g $AgNO_3$ should produce about the same mass of Ag_2CrO_4 . The actual yield of Ag_2CrO_4 is close to 0.500 g, so a percent yield of 93.2% is reasonable.



Like silver chromate, all chromium compounds are colored. Although silver chromate is insoluble in water, potassium chromate (K_2CrO_4 , yellow) and potassium dichromate ($K_2Cr_2O_7$, orange) are soluble.

Practice!

For more practice calculating percent yield, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- 27.** Aluminum hydroxide is often present in antacids to neutralize stomach acid (HCl). If 14.0 g aluminum hydroxide is present in an antacid tablet, determine the theoretical yield of aluminum chloride produced when the tablet reacts with stomach acid. If the actual yield of aluminum chloride from this tablet is 22.0 g, what is the percent yield?
- $$\text{Al}(\text{OH})_3(\text{s}) + 3\text{HCl}(\text{aq}) \rightarrow \text{AlCl}_3(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$$
- 28.** When copper wire is placed into a silver nitrate solution, silver crystals and copper(II) nitrate solution form. Write the balanced chemical equation for the reaction. If a 20.0-g sample of copper is used, determine the theoretical yield of silver. If 60.0 g silver is actually recovered from the reaction, determine the percent yield of the reaction.
- 29.** Zinc reacts with iodine in a synthesis reaction. Write the balanced chemical equation for the reaction. Determine the theoretical yield if a 125.0-g sample of zinc was used. Determine the percent yield if 515.6 g product is recovered.

problem-solving LAB

How does the surface area of a solid reactant affect percent yield?

Designing an Experiment The cost of every manufactured item you buy is based largely on the cost of producing the item. Manufacturers compete to reduce costs and increase profits. This means increasing the percent yield of the manufacturing process by producing the most product for the amount of reactant used. If you were going to produce iron oxide (Fe_2O_3) from steel wool, how could you design an experiment to determine what gauge (diameter) steel wool will produce the highest yield? Write the equation for the reaction upon which you will base your experiment.

Analysis

In the first photo, different gauges of steel wool are shown. The second photo shows the combustion of a sample of steel wool using a Bunsen burner.

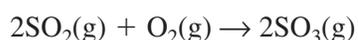
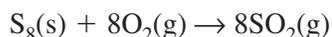
Thinking Critically

1. What quantities must be used to calculate the percent yield of Fe_2O_3 when iron is burned? How will these quantities be measured and how many measurements should be made?
2. What quantities should be kept constant in the experiment?
3. How will the resulting data be analyzed?
4. Are there any obvious errors in the design that could significantly affect the results? If so, how could they be avoided?

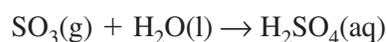


Percent yield in the marketplace You learned in the **problem-solving LAB** that in order to compete, manufacturers must reduce the cost of making their products to the lowest level possible. Percent yield is important in the calculation of overall cost effectiveness in industrial processes. For example, sulfuric acid (H_2SO_4) is made using mined sulfur, **Figure 12-8**. Sulfuric acid is an important chemical because it is a raw material for products such as fertilizers, detergents, pigments, and textiles. The cost of sulfuric acid affects the cost of many of the consumer items you use every day.

A two-step process called the contact process is often used for the manufacture of sulfuric acid. Over time, the process has been improved by chemical engineers to produce the maximum yield of product and, at the same time, comply with environmental standards for clean air. The two steps in the contact process are



A final step, the combination of SO_3 with water, produces the product, H_2SO_4 .



The first step, the combustion of sulfur, produces almost 100% yield. The second step also produces a high yield if a catalyst is used at the relatively low temperature of 400°C . A catalyst is a substance that speeds a reaction but does not appear in the chemical equation. Under these conditions, the reaction is slow. Raising the temperature speeds up the reaction but the yield decreases.

To maximize yield and minimize time in the second step, engineers have devised a system in which the reactants, O_2 and SO_2 , are passed over a catalyst at 400°C . Because the reaction releases a great deal of heat, the temperature gradually increases with an accompanying decrease in yield. Thus, when the temperature reaches approximately 600°C , the mixture is cooled and then passed over the catalyst again. A total of four passes over the catalyst with cooling between passes results in a yield greater than 98%. This four-pass procedure maximizes the yield at temperatures near 400°C , and uses the modest increase in temperature to increase the rate and minimize the time.



Figure 12-8

The large tonnage of sulfur needed to make sulfuric acid and other products is often obtained by forcing hot water into underground deposits to melt the sulfur and then pumping the liquid sulfur to the surface.

Section 12.4 Assessment

30. Distinguish between the theoretical yield and the actual yield of a chemical reaction.

31. Give several reasons why the actual yield is not usually equal to the theoretical yield.

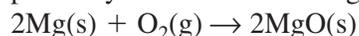
32. Explain how percent yield is calculated.

33. Thinking Critically In an experiment, you are to combine iron with an excess of sulfur and heat the mixture to obtain iron(III) sulfide.



What experimental information must you collect in order to calculate the percent yield of this reaction?

34. Interpreting Data Use the data to determine the percent yield of the following reaction.



Oxygen is in excess.

Reaction Data	
Mass of crucible	35.67 g
Mass of crucible + Mg	38.06 g
Mass of Mg	
Mass of crucible + MgO	39.15 g
Mass of MgO	

A Mole Ratio

Iron reacts with copper(II) sulfate in a single replacement reaction. By measuring the mass of iron that reacts and the mass of copper metal produced, you can calculate the ratio of moles of reactant to moles of product. This mole ratio can be compared to the ratio found in the balanced chemical equation.

Problem

Which reactant is the limiting reactant? How does the experimental mole ratio of Fe to Cu compare with the mole ratio in the balanced chemical equation? What is the percent yield?

Objectives

- **Observe** a single replacement reaction.
- **Measure** the masses of iron and copper.
- **Calculate** the moles of each metal and the mole ratio.

Materials

iron metal	400-mL beaker
filings, 20 mesh	100-mL graduated cylinder
copper(II) sulfate pentahydrate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)	weighing paper
distilled water	balance
stirring rod	hot plate
150-mL beaker	beaker tongs

Safety Precautions



- Always wear safety glasses and a lab apron.
- Hot objects will not appear to be hot.
- Do not heat broken, chipped, or cracked glassware.
- Turn off the hot plate when not in use.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Is it important that you know you are using the hydrated form of copper(II) sulfate? Would it be possible to use the anhydrous form? Why or why not?

Data for the Reaction of Copper(II) Sulfate and Iron

Mass of empty 150-mL beaker	
Mass of 150-mL beaker + $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	
Mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	
Mass of iron filings	
Mass of 150-mL beaker and dried copper	
Mass of dried copper	
Observations	

Procedure

1. Measure and record the mass of a clean, dry 150-mL beaker.
2. Place approximately 12 g of copper(II) sulfate pentahydrate into the 150-mL beaker and measure and record the combined mass.
3. Add 50 mL of distilled water to the copper(II) sulfate pentahydrate and heat the mixture on the hot plate at a medium setting. Stir until all of the solid is dissolved, but do not boil. Using tongs, remove the beaker from the hot plate.
4. Measure approximately 2 g of iron metal filings onto a piece of weighing paper. Measure and record the exact mass of the filings.
5. While stirring, slowly add the iron filings to the hot copper(II) sulfate solution.
6. Allow the reaction mixture to stand, without stirring, for five minutes to ensure complete reaction. The solid copper metal will settle to the bottom of the beaker.
7. Use the stirring rod to decant (pour off) the liquid into a 400-mL beaker. Be careful to decant only the liquid.

8. Add 15 mL of distilled water to the copper solid and carefully swirl the beaker to wash the copper. Decant the liquid into the 400-mL beaker.
9. Repeat step 8 two more times.
10. Place the 150-mL beaker containing the wet copper on the hot plate. Use low heat to dry the copper.
11. Remove the beaker from the hot plate and allow it to cool.
12. Measure and record the mass of the cooled 150-mL beaker and the copper.

Cleanup and Disposal

1. Make sure the hot plate is off.
2. The dry copper can be placed in a waste container. Wet any residue that sticks to the beaker and wipe it out using a paper towel. Pour the unreacted copper(II) sulfate and iron(II) sulfate solutions into a large beaker in the fume hood.
3. Return all lab equipment to its proper place.
4. Wash your hands thoroughly after all lab work and cleanup is complete.

Analyze and Conclude

1. **Observing and Inferring** What evidence did you observe that confirms that a chemical reaction occurred?
2. **Applying Concepts** Write a balanced chemical equation for the single-replacement reaction that occurred.
3. **Interpreting Data** From your data, determine the mass of copper produced.
4. **Using Numbers** Use the mass of copper to calculate the moles of copper produced.
5. **Using Numbers** Calculate the moles of iron used in the reaction.
6. **Using Numbers** Determine the whole number ratio of moles of iron to moles of copper.
7. **Comparing and Contrasting** Compare the ratio of moles of iron to moles of copper from the balanced chemical equation to the mole ratio calculated using your data.



8. **Evaluating Results** Use the balanced chemical equation to calculate the mass of copper that should have been produced from the sample of iron you used. Use this number and the mass of copper you actually obtained to calculate the percent yield.
9. **Error Analysis** What was the source of any deviation from the mole ratio calculated from the chemical equation? How could you improve your results?
10. **Drawing a Conclusion** Which reactant is the limiting reactant? Explain.

Real-World Chemistry

1. A furnace that provides heat by burning methane gas (CH_4) must have the correct mixture of air and fuel to operate efficiently. What is the mole ratio of air to methane gas in the combustion of methane? Hint: Air is 20% oxygen.
2. Automobile air bags inflate on impact because a series of gas-producing chemical reactions are triggered. To be effective in saving lives, the bags must not overinflate or underinflate. What factors must automotive engineers take into account in the design of air bags?

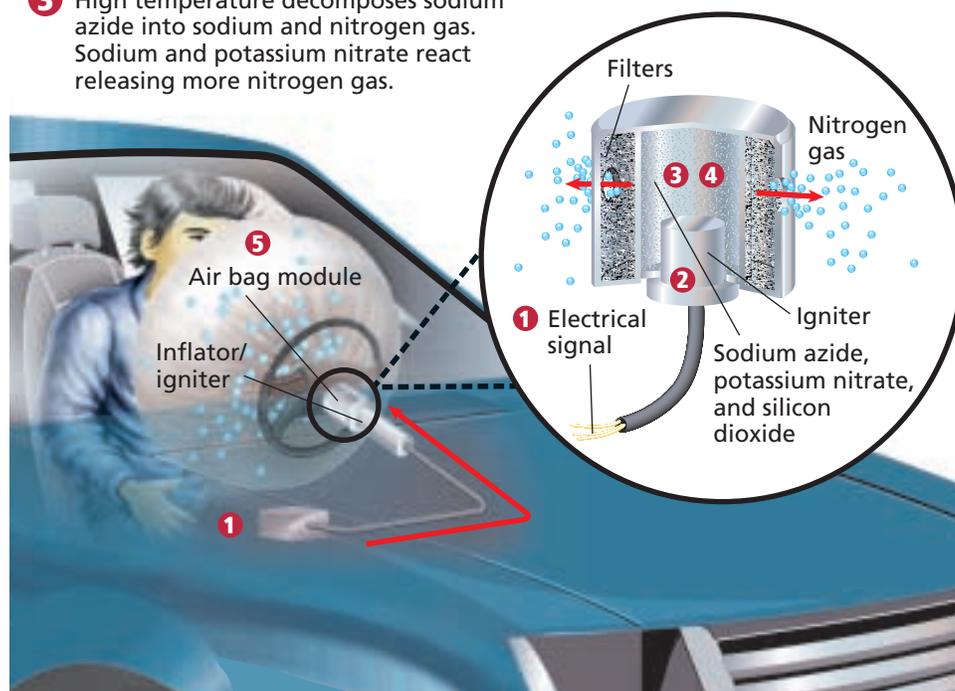
How It Works

Air Bags

Air bags fill with nitrogen gas as they deploy in automobile crashes. The source of the nitrogen gas is the chemical compound sodium azide (NaN_3). Hazardous sodium metal is produced along with the nitrogen, so potassium nitrate is added to convert the sodium into less hazardous sodium oxide

(Na_2O). Stoichiometric calculations are needed to determine the precise quantity of sodium azide that will produce the volume of nitrogen gas required to inflate the air bag. If too much gas is produced, the air bag may be so rigid that hitting it would be the same as hitting a solid wall.

- 1 Crash sensor detects rapid deceleration and sends signal to air bag module.
- 2 Igniter explodes, heating sodium azide in the inflator.
- 3 High temperature decomposes sodium azide into sodium and nitrogen gas. Sodium and potassium nitrate react releasing more nitrogen gas.
- 4 Silicon dioxide (sand), sodium oxide, and potassium oxide formed in the inflator fuse into glass.
- 5 Nitrogen gas inflates air bag.



Thinking Critically

1. **Predicting** Which starting material used in the air bag inflator is the least important for the proper inflation of the air bag? Would it be necessary to have it present in a precise stoichiometric ratio? Why or why not?
2. **Analyze and Conclude** What is the correct stoichiometric ratio between NaN_3 and KNO_3 to ensure no sodium is unreacted? What would be the consequences of an excess of KNO_3 to the operation of an air bag?

Summary

12.1 What is stoichiometry?

- Balanced chemical equations can be interpreted in terms of representative particles (atoms, molecules, formula units), moles, and mass.
- The law of conservation of mass, as applied to chemical reactions, means that the total mass of the reactants is equal to the total mass of the products.
- Mole ratios are central to stoichiometric calculations. They are derived from the coefficients in a balanced chemical equation. To write mole ratios, the number of moles of each reactant and product is placed, in turn, in the numerator of the ratio with the moles of each other reactant and product placed in the denominator.

12.2 Stoichiometric Calculations

- Stoichiometric calculations allow a chemist to predict the amount of product that can be obtained from a given amount of reactant or to determine how much of two or more reactants must be used to produce a specified amount of product.
- The four steps in stoichiometric calculations begin with the balanced chemical equation.
- Mole ratios used in the calculations are determined from the balanced chemical equation.
- The mass of the given substance is converted to moles of the given substance. Then, moles of the given substance are converted by means of a mole

ratio to moles of the unknown substance. Finally, moles of the unknown substance are converted to the mass of the unknown substance.

12.3 Limiting Reactants

- The limiting reactant is the reactant that is completely consumed during a chemical reaction. Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the given amount of the limiting reactant.

12.4 Percent Yield

- The theoretical yield of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The actual yield is the amount of product actually produced. Actual yield must be obtained through experimentation.
- Percent yield is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.

Key Equations and Relationships

- moles of known $\times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$
(p. 358)
- $\frac{\text{actual yield (from experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100 = \text{percent yield}$
(p. 370)

Vocabulary

- actual yield (p. 370)
- excess reactant (p. 364)
- limiting reactant (p. 364)
- mole ratio (p. 356)
- percent yield (p. 370)
- stoichiometry (p. 354)
- theoretical yield (p. 370)

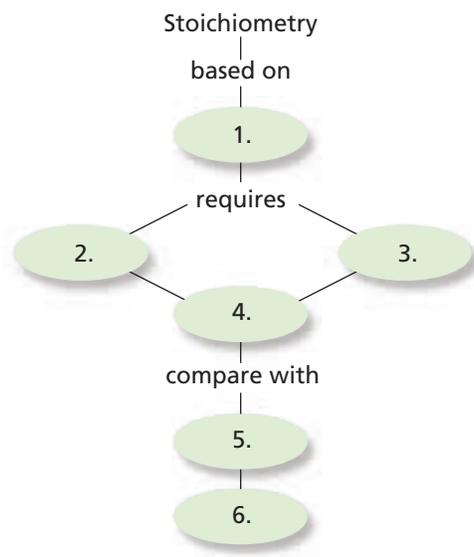




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Concept Mapping



35. Fill in the ovals with the following terms to create a concept map: actual yield, balanced chemical equation, molar mass, mole ratio, percent yield, and theoretical yield.

Mastering Concepts

36. What relationships can be determined from a balanced chemical equation? (12.1)
37. Explain how the law of conservation of mass allows you to interpret a balanced chemical equation in terms of mass. (12.1)
38. Explain why mole ratios are central to stoichiometric calculations. (12.1)
39. What is the mole ratio that can convert from moles of A to moles of B? (12.1)
40. What is the first step in all stoichiometric calculations? (12.2)
41. How is molar mass used in some stoichiometric calculations? (12.2)
42. What information must you have in order to calculate the mass of product formed in a chemical reaction? (12.2)
43. What is meant by limiting reactant? Excess reactant? (12.3)
44. How are mole ratios used in finding the limiting reactant in a reaction? (12.3)
45. What is the difference between actual yield and theoretical yield? (12.4)
46. How are actual yield and theoretical yield determined? (12.4)
47. Can the percent yield of a chemical reaction be more than 100%? Explain your answer. (12.4)
48. What relationship is used to determine the percent yield of a chemical reaction? (12.4)
49. What experimental information do you need in order to calculate both the theoretical and percent yield of any chemical reaction? (12.4)
50. A metal oxide reacts with water to produce a metal hydroxide. What additional information would you need to determine the percent yield of metal hydroxide from this reaction? (12.4)

Mastering Problems

Interpreting Equations (12.1)

51. Interpret the following equation in terms of particles, moles, and mass.
- $$4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)$$
52. When tin(IV) oxide is heated with carbon in a process called smelting, the element tin can be extracted.
- $$\text{SnO}_2(s) + 2\text{C}(s) \rightarrow \text{Sn}(l) + 2\text{CO}(g)$$
- Interpret the equation in terms of particles, moles, and mass.
53. When hydrochloric acid solution reacts with lead(II) nitrate solution, lead(II) chloride precipitates and a solution of nitric acid is produced.
- Write the balanced chemical equation for the reaction.
 - Interpret the equation in terms of molecules and formula units, moles, and mass.

Mole Ratios (12.1)

54. When solid copper is added to nitric acid, copper(II) nitrate, nitrogen dioxide, and water are produced. Write the balanced chemical equation for the reaction. List six mole ratios for the reaction.
55. When aluminum is mixed with iron(III) oxide, iron metal and aluminum oxide are produced along with a large quantity of heat. What mole ratio would you use to determine mol Fe if mol Fe_2O_3 is known?
- $$\text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow 2\text{Fe}(s) + \text{Al}_2\text{O}_3(s) + \text{heat}$$

- 56.** Solid silicon dioxide, often called silica, reacts with hydrofluoric acid (HF) solution to produce the gas silicon tetrafluoride and water.
- Write the balanced chemical equation for the reaction.
 - List three mole ratios and explain how you would use them in stoichiometric calculations.
- 57.** Determine the mole ratio necessary to convert moles of aluminum to moles of aluminum chloride when aluminum reacts with chlorine.
- 58.** Chromite (FeCr_2O_4) is the most important commercial ore of chromium. One of the steps in the process used to extract chromium from the ore is the reaction of chromite with coke (carbon) to produce ferrochrome (FeCr_2).
- $$2\text{C}(\text{s}) + \text{FeCr}_2\text{O}_4(\text{s}) \rightarrow \text{FeCr}_2(\text{s}) + 2\text{CO}_2(\text{g})$$
- What mole ratio would you use to convert from moles of chromite to moles of ferrochrome?
- 59.** The air pollutant SO_2 is removed from the air by means of a reaction among sulfur dioxide, calcium carbonate, and oxygen. The products of this reaction are calcium sulfate and carbon dioxide. Determine the mole ratio you would use to convert mol SO_2 to mol CaSO_4 .

Reaction Data			
Moles of reactants		Moles of products	
W	X	Y	Z
0.90	0.30	0.60	1.20

- 60.** Two substances, W and X, react to form the products Y and Z. The table shows the numbers of moles of the reactants and products involved when the reaction was carried out in one experiment. Use the data to determine the coefficients that will balance the equation $\text{W} + \text{X} \rightarrow \text{Y} + \text{Z}$.

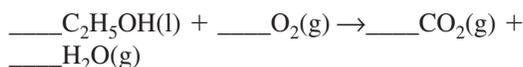
Stoichiometric Mole-to-Mole Conversions (12.2)

- 61.** If 5.50 mol calcium carbide (CaC_2) reacts with an excess of water, how many moles of acetylene (C_2H_2) will be produced?
- $$\text{CaC}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{OH})_2(\text{aq}) + \text{C}_2\text{H}_2(\text{g})$$
- 62.** When an antacid tablet dissolves in water, the fizz is due to a reaction between sodium hydrogen carbonate (sodium bicarbonate, NaHCO_3) and citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$).
- $$3\text{NaHCO}_3(\text{aq}) + \text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) \rightarrow 3\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l}) + \text{Na}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq})$$
- How many moles of carbon dioxide can be produced if one tablet containing 0.0119 mol NaHCO_3 is dissolved?
- 63.** One of the main components of pearls is calcium carbonate. If pearls are put in acidic solution, they dissolve.
- $$\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$
- How many mol CaCO_3 can be dissolved in 0.0250 mol HCl?
- ### Stoichiometric Mole-to-Mass Conversions (12.2)
- 64.** Citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) is a product of the fermentation of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) in air.
- $$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$$
- Determine the mass of citric acid produced when 2.50 mol $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ is used.
- 65.** Esterification is a reaction between an organic acid and an alcohol that forms an ester and water. The ester ethyl butanoate ($\text{C}_3\text{H}_7\text{COOC}_2\text{H}_5$), which is responsible for the fragrance of pineapples, is formed when the alcohol ethanol ($\text{C}_2\text{H}_5\text{OH}$) and butanoic acid ($\text{C}_3\text{H}_7\text{COOH}$) are heated in the presence of sulfuric acid.
- $$\text{C}_2\text{H}_5\text{OH}(\text{l}) + \text{C}_3\text{H}_7\text{COOH}(\text{l}) \rightarrow \text{C}_3\text{H}_7\text{COOC}_2\text{H}_5(\text{l}) + \text{H}_2\text{O}(\text{l})$$
- Determine the mass of ethyl butanoate produced if 4.50 mol ethanol is used.
- 66.** Carbon dioxide is released into the atmosphere through the combustion of octane (C_8H_{18}) in gasoline. Write the balanced chemical equation for the combustion of octane and calculate the mass of octane needed to release 5.00 mol CO_2 .
- 67.** A solution of potassium chromate reacts with a solution of lead(II) nitrate to produce a yellow precipitate of lead(II) chromate and a solution of potassium nitrate.
- Write the balanced chemical equation.
 - Starting with 0.250 mol potassium chromate, determine the mass of lead chromate that can be obtained.
- 68.** The exothermic reaction between liquid hydrazine (N_2H_2) and liquid hydrogen peroxide (H_2O_2) is used to fuel rockets. The products of this reaction are nitrogen gas and water.
- Write the balanced chemical equation.
 - How many grams of hydrazine are needed to produce 10.0 mol nitrogen gas?

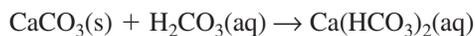
Stoichiometric Mass-to-Mass Conversions (12.2)

- 69.** Chloroform (CHCl_3), an important solvent, is produced by a reaction between methane and chlorine.
- $$\text{CH}_4(\text{g}) + 3\text{Cl}_2(\text{g}) \rightarrow \text{CHCl}_3(\text{g}) + 3\text{HCl}(\text{g})$$
- How many g CH_4 is needed to produce 50.0 g CHCl_3 ?

- 70.** Gasohol is a mixture of ethanol and gasoline. Balance the equation and determine the mass of CO_2 produced from the combustion of 100.0 g ethanol.

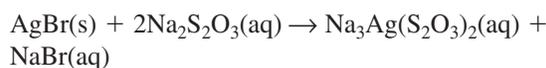


- 71.** When surface water dissolves carbon dioxide, carbonic acid (H_2CO_3) is formed. When the water moves underground through limestone formations, the limestone dissolves and caves are sometimes produced.

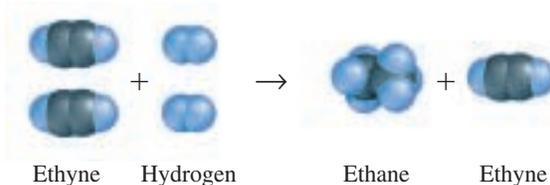


What mass of limestone must have dissolved if 3.05×10^{10} kg of calcium hydrogen carbonate was produced?

- 72.** Car batteries use solid lead and lead(IV) oxide with sulfuric acid solution to produce an electric current. The products of this reaction are lead(II) sulfate in solution and water.
- Write the balanced chemical equation for this reaction.
 - Determine the mass of lead(II) sulfate produced when 25.0 g lead reacts with an excess of lead(IV) oxide and sulfuric acid.
- 73.** The fuel methanol (CH_3OH) is made by the reaction of carbon monoxide and hydrogen.
- Write the balanced chemical equation.
 - How many grams of hydrogen are needed to produce 45.0 grams of methanol?
- 74.** To extract gold from its ore, the ore is treated with sodium cyanide solution in the presence of oxygen and water.
- $$4\text{Au}(\text{s}) + 8\text{NaCN}(\text{aq}) + \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{NaAu}(\text{CN})_2(\text{aq}) + 4\text{NaOH}(\text{aq})$$
- Determine the mass of gold that can be extracted if 25.0 g sodium cyanide is used.
 - If the mass of the ore from which the gold was extracted is 150.0 g, what percentage of the ore is gold?
- 75.** Photographic film contains silver bromide in gelatin. Once exposed, some of the silver bromide decomposes producing fine grains of silver. The unexposed silver bromide is removed by treating the film with sodium thiosulfate. Soluble sodium silver thiosulfate ($\text{Na}_3\text{Ag}(\text{S}_2\text{O}_3)_2$) is produced.

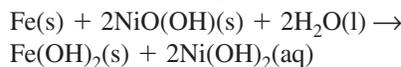


Determine the mass of $\text{Na}_3\text{Ag}(\text{S}_2\text{O}_3)_2$ produced if 0.275 g AgBr is removed.

Limiting Reactants (12.3)


- 76.** The illustration shows the reaction between ethyne (acetylene, C_2H_2) and hydrogen. The product is ethane (C_2H_6). Which is the limiting reactant? Which is the excess reactant? Explain.

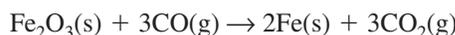
- 77.** This reaction takes place in a nickel-iron battery.



Determine the number of moles of iron(II) hydroxide ($\text{Fe}(\text{OH})_2$) produced if 5.00 mol Fe and 8.00 mol $\text{NiO}(\text{OH})$ react.

- 78.** How many moles of cesium xenon heptafluoride (CsXeF_7) can be produced from the reaction of 12.5 mol cesium fluoride with 10.0 mol xenon hexafluoride?
- $$\text{CsF}(\text{s}) + \text{XeF}_6(\text{s}) \rightarrow \text{CsXeF}_7(\text{s})$$

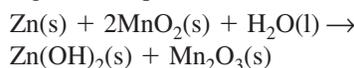
- 79.** Iron is obtained commercially by the reaction of hematite (Fe_2O_3) with carbon monoxide. How many grams of iron are produced if 25.0 moles of hematite react with 30.0 moles of carbon monoxide?



- 80.** Under certain conditions of temperature and pressure, hydrogen and nitrogen react to produce ammonia (NH_3). Write the balanced chemical equation and determine the mass of ammonia produced if 3.50 mol H_2 reacts with 5.00 mol N_2 .

- 81.** The reaction of chlorine gas with solid phosphorus (P_4) produces solid phosphorus pentachloride. When 16.0 g chlorine reacts with 23.0 g P_4 , which reactant limits the amount of phosphorus pentachloride produced? Which reactant is in excess?

- 82.** An alkaline battery produces electrical energy according to this equation.



- Determine the limiting reactant if 25.0 g Zn and 30.0 g MnO_2 are used.
- Determine the mass of $\text{Zn}(\text{OH})_2$ produced.

- 83.** Lithium reacts spontaneously with bromine to produce lithium bromide. Write the balanced chemical equation for the reaction. If 25.0 g of lithium and 25.0 g of bromine are present at the beginning of the reaction, determine
- the limiting reactant
 - the mass of lithium bromide produced
 - the excess reactant and the mass in excess.

Percent Yield (12.4)

- 84.** Ethanol (C_2H_5OH) is produced from the fermentation of sucrose in the presence of enzymes.
- $$C_{12}H_{22}O_{11}(aq) + H_2O(g) \rightarrow 4C_2H_5OH(l) + 4CO_2(g)$$
- Determine the theoretical and percent yields of ethanol if 684 g sucrose undergoes fermentation and 349 g ethanol is obtained.
- 85.** Lead(II) oxide is obtained by roasting galena, lead(II) sulfide, in air.
- $$\begin{aligned} & ___ PbS(s) + ___ O_2(g) \rightarrow \\ & ___ PbO(s) + ___ SO_2(g) \end{aligned}$$
- Balance the equation and determine the theoretical yield of PbO if 200.0 g PbS is heated.
 - What is the percent yield if 170.0 g PbO is obtained?
- 86.** Upon heating, calcium carbonate decomposes to produce calcium oxide and carbon dioxide.
- Determine the theoretical yield of CO_2 if 235.0 g $CaCO_3$ is heated.
 - What is the percent yield of CO_2 if 97.5 g CO_2 is collected?
- 87.** Hydrofluoric acid solutions cannot be stored in glass containers because HF reacts readily with silica in glass to produce hexafluorosilicic acid (H_2SiF_6).
- $$SiO_2(s) + 6HF(aq) \rightarrow H_2SiF_6(aq) + 2H_2O(l)$$
- If 40.0 g SiO_2 and 40.0 g of HF react
- determine the limiting reactant.
 - determine the mass of the excess reactant.
 - determine the theoretical yield of H_2SiF_6 .
 - determine the percent yield if the actual yield is 45.8 g H_2SiF_6 .
- 88.** Pure zirconium is obtained using the two-step Van Arkel process. In the first step, impure zirconium and iodine are heated to produce zirconium iodide (ZrI_4). In the second step, ZrI_4 is decomposed to produce pure zirconium.
- $$ZrI_4(s) \rightarrow Zr(s) + 2I_2(g)$$
- Determine the percent yield of zirconium if 45.0 g ZrI_4 is decomposed and 5.00 g pure Zr is obtained.

- 89.** Phosphorus is commercially prepared by heating a mixture of calcium phosphate, sand, and coke in an electric furnace. The process involves two reactions.
- $$2Ca_3(PO_4)_2(s) + 6SiO_2(s) \rightarrow 6CaSiO_3(l) + P_4O_{10}(g)$$
- $$P_4O_{10}(g) + 10C(s) \rightarrow P_4(g) + 10CO(g)$$

The P_4O_{10} produced in the first reaction reacts with an excess of coke (C) in the second reaction. Determine the theoretical yield of P_4 if 250.0 g $Ca_3(PO_4)_2$ and 400.0 g SiO_2 are heated. If the actual yield of P_4 is 45.0 g, determine the percent yield of P_4 .

- 90.** Chlorine can be prepared by the reaction of manganese(IV) oxide with hydrochloric acid.
- $$\begin{aligned} & ___ MnO_2(s) + ___ HCl(aq) \rightarrow \\ & ___ MnCl_2(aq) + ___ Cl_2(g) + ___ H_2O(l) \end{aligned}$$

Balance the equation and determine the theoretical and percent yields of chlorine if 86.0 g MnO_2 and 50.0 g HCl react. The actual yield of chlorine is 20.0 g.

Mixed Review

Sharpen your problem-solving skills by answering the following.

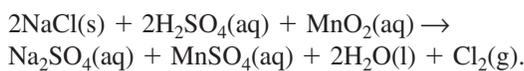
- 91.** Ammonium sulfide reacts with copper(II) nitrate in a double replacement reaction. What mole ratio would you use to determine the moles of NH_4NO_3 produced if the moles of CuS are known.
- 92.** One method for producing nitrogen in the laboratory is to react ammonia with copper(II) oxide.
- $$\begin{aligned} & ___ NH_3(g) + ___ CuO(s) \rightarrow \\ & ___ Cu(s) + ___ H_2O(l) + ___ N_2(g) \end{aligned}$$
- Balance the equation.
 - If 40.0 g NH_3 is reacted with 80.0 g CuO, determine the limiting reactant.
 - Determine the mass of N_2 produced by this reaction.
 - Which reactant is in excess? How much remains after the reaction?
- 93.** The compound calcium cyanamide ($CaNCN$) can be used as a fertilizer. To obtain this compound, calcium carbide is reacted with nitrogen at high temperatures.
- $$CaC_2(s) + N_2(g) \rightarrow CaNCN(s) + C(s)$$
- What mass of CaNCN can be produced if 7.50 mol CaC_2 reacts with 5.00 mol N_2 ?
- 94.** When copper(II) oxide is heated in the presence of hydrogen gas, elemental copper and water are produced. What mass of copper can be obtained if 32.0 g copper(II) oxide is used?

95. Nitrogen oxide is present in urban pollution but it is immediately converted to nitrogen dioxide as it reacts with oxygen.

- Write the balanced chemical equation for the formation of nitrogen dioxide from nitrogen oxide.
- What mole ratio would you use to convert from moles of nitrogen oxide to moles of nitrogen dioxide?

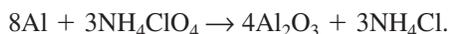
96. Determine the theoretical and percent yield of hydrogen gas if 36.0 g water undergoes electrolysis to produce hydrogen and oxygen and 3.80 g hydrogen is collected.

97. The Swedish chemist Karl Wilhlm was first to produce chlorine in the laboratory.



What mole ratio could be used to find the moles of chlorine produced from 4.85 moles of sodium chloride? Determine the moles of chlorine produced. Determine the mass of chlorine produced.

98. The solid booster rockets of the space shuttle contain ammonium perchlorate (NH_4ClO_4) and powdered aluminum as the propellant.



Determine the percent yield if 6.00×10^5 kg NH_4ClO_4 produces 6.56×10^5 kg aluminum oxide.

102. Formulating Models Copper reacts with chlorine to produce copper(II) chloride. Draw a diagram that represents eight atoms of copper reacting with six molecules of chlorine. Make sure you include the particles before the reaction and after the reaction. Include any excess reactants.

103. Applying Concepts When your campfire begins to die down and smolder, it helps to fan it. Explain in terms of stoichiometry why the fire begins to flare up again.

Writing in Chemistry

104. Research the air pollutants produced by using gasoline in internal combustion engines. Discuss the common pollutants and the reaction that produces them. Show, through the use of stoichiometry, how each pollutant could be reduced if more people used mass transit.

105. The percent yield of ammonia produced when hydrogen and nitrogen are combined under ordinary conditions is extremely small. However, the Haber Process combines the two gases under a set of conditions designed to maximize yield. Research the conditions used in the Haber Process and find out why the development of the process was of great importance.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

106. You observe that sugar dissolves more quickly in hot tea than in iced tea. You state that higher temperatures increase the rate at which sugar dissolves in water. Is this statement a hypothesis or theory and why? (Chapter 1)

107. Write the electron configuration for each of the following atoms. (Chapter 5)

- fluorine
- aluminum
- titanium
- radon

108. Explain why the gaseous nonmetals exist as diatomic molecules, but other gaseous elements exist as single atoms. (Chapter 9)

109. Write a balanced equation for the reaction of potassium with oxygen. (Chapter 10)

110. What is the molecular mass of UF_6 ? What is the molar mass of UF_6 ? (Chapter 11)

Thinking Critically

99. Analyze and Conclude In an experiment, you obtain a percent yield of product of 108%. Is such a percent yield possible? Explain. Assuming that your calculation is correct, what reasons might explain such a result?

100. Observing and Inferring Determine whether the following reactions depend upon a limiting reactant. Explain why or why not and identify the limiting reactant.

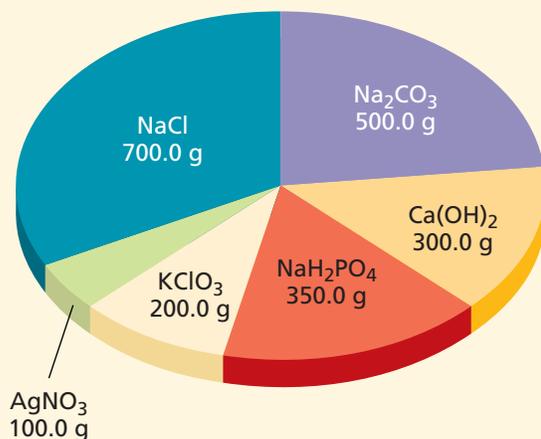
- Potassium chlorate decomposes to form potassium chloride and oxygen.
- Silver nitrate and hydrochloric acid react to produce silver chloride and nitric acid.
- Propane (C_3H_8) burns in excess oxygen to produce carbon dioxide and water.

101. Designing an Experiment Design an experiment that can be used to determine the percent yield of anhydrous copper(II) sulfate when copper(II) sulfate pentahydrate is heated to remove water.

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Graphs Use the graph below to answer questions 1–4.

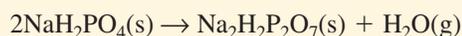
Supply of Various Chemicals in Dr. Raitano's Laboratory



- Pure silver metal can be made using the reaction shown below:

$$\text{Cu(s)} + 2\text{AgNO}_3\text{(aq)} \rightarrow 2\text{Ag(s)} + \text{Cu(NO}_3)_2\text{(aq)}$$
 How many grams of copper metal will be needed to use up all of the AgNO₃ in Dr. Raitano's laboratory?
 a. 18.70 g c. 74.7 g
 b. 37.3 g d. 100 g
- $$\text{Na}_2\text{CO}_3\text{(aq)} + \text{Ca(OH)}_2\text{(aq)} \rightarrow 2\text{NaOH(aq)} + \text{CaCO}_3\text{(s)}$$
 The LeBlanc process, shown above, is the traditional method of manufacturing sodium hydroxide. Using the amounts of chemicals available in Dr. Raitano's lab, the maximum number of moles of NaOH that can be produced is _____.
 a. 4.05 mol c. 8.097 mol
 b. 4.72 mol d. 9.43 mol
- Pure O₂ gas can be generated from the decomposition of potassium chlorate (KClO₃):

$$2\text{KClO}_3\text{(s)} \rightarrow 2\text{KCl(s)} + 3\text{O}_2\text{(g)}$$
 If half of the KClO₃ in the lab is used and 12.8 g of oxygen gas are produced, the percent yield of this reaction is _____.
 a. 12.8% c. 65.6%
 b. 32.7% d. 98.0%
- Sodium dihydrogen pyrophosphate (Na₂H₂P₂O₇), more commonly known as baking powder, is manufactured by heating NaH₂PO₄ at high temperatures:



If 444.0 g of Na₂H₂P₂O₇ are needed, how much more NaH₂PO₄ will Dr. Raitano have to buy to make enough Na₂H₂P₂O₇?

- 94.0 g
 - 130.0 g
 - 480 g
 - none—the lab already has enough
- Stoichiometry is based on the law of _____.
 a. constant mole ratios
 b. Avogadro's constant
 c. conservation of energy
 d. conservation of mass
 - Red mercury(II) oxide decomposes at high temperatures to form mercury metal and oxygen gas:

$$2\text{HgO(s)} \rightarrow 2\text{Hg(l)} + \text{O}_2\text{(g)}$$
 If 3.55 moles of HgO decompose to form 1.54 moles of O₂ and 618 g of Hg, what is the percent yield of this reaction?
 a. 13.2% c. 56.6%
 b. 42.5% d. 86.8%
 - Dimethyl hydrazine (CH₃)₂N₂H₂ ignites spontaneously upon contact with dinitrogen tetroxide (N₂O₄):

$$(\text{CH}_3)_2\text{N}_2\text{H}_2\text{(l)} + 2\text{N}_2\text{O}_4\text{(l)} \rightarrow 3\text{N}_2\text{(g)} + 4\text{H}_2\text{O(g)} + 2\text{CO}_2\text{(g)}$$
 Because this reaction produces an enormous amount of energy from a small amount of reactants, it was used to drive the rockets on the Lunar Excursion Modules (LEM's) of the Apollo space program. If 2.0 moles of dimethyl hydrazine are mixed with 4.0 moles of dinitrogen tetroxide, and the reaction achieves an 85% yield, how many moles of N₂, H₂O, and CO₂ will be formed?
 a. 0.57 mol N₂, 0.43 mol H₂O, 0.85 mol CO₂
 b. 2.6 mol N₂, 3.4 mol H₂O, 1.7 mol CO₂
 c. 5.1 mol N₂, 6.8 mol H₂O, 3.4 mol CO₂
 d. 6.0 mol N₂, 8.0 mol H₂O, 4.0 mol CO₂



TEST-TAKING TIP

Calculators Are Only Machines If your test allows you to use a calculator, use it wisely. The calculator can't figure out what the question is asking. That's still *your* job. Figure out which numbers are relevant, and determine the best way to solve the problem before you start punching keys.

States of Matter

What You'll Learn

- ▶ You will use the kinetic-molecular theory to explain the physical properties of gases, liquids, and solids.
- ▶ You will compare types of intermolecular forces.
- ▶ You will explain how kinetic energy and intermolecular forces combine to determine the state of a substance.
- ▶ You will describe the role of energy in phase changes.

Why It's Important

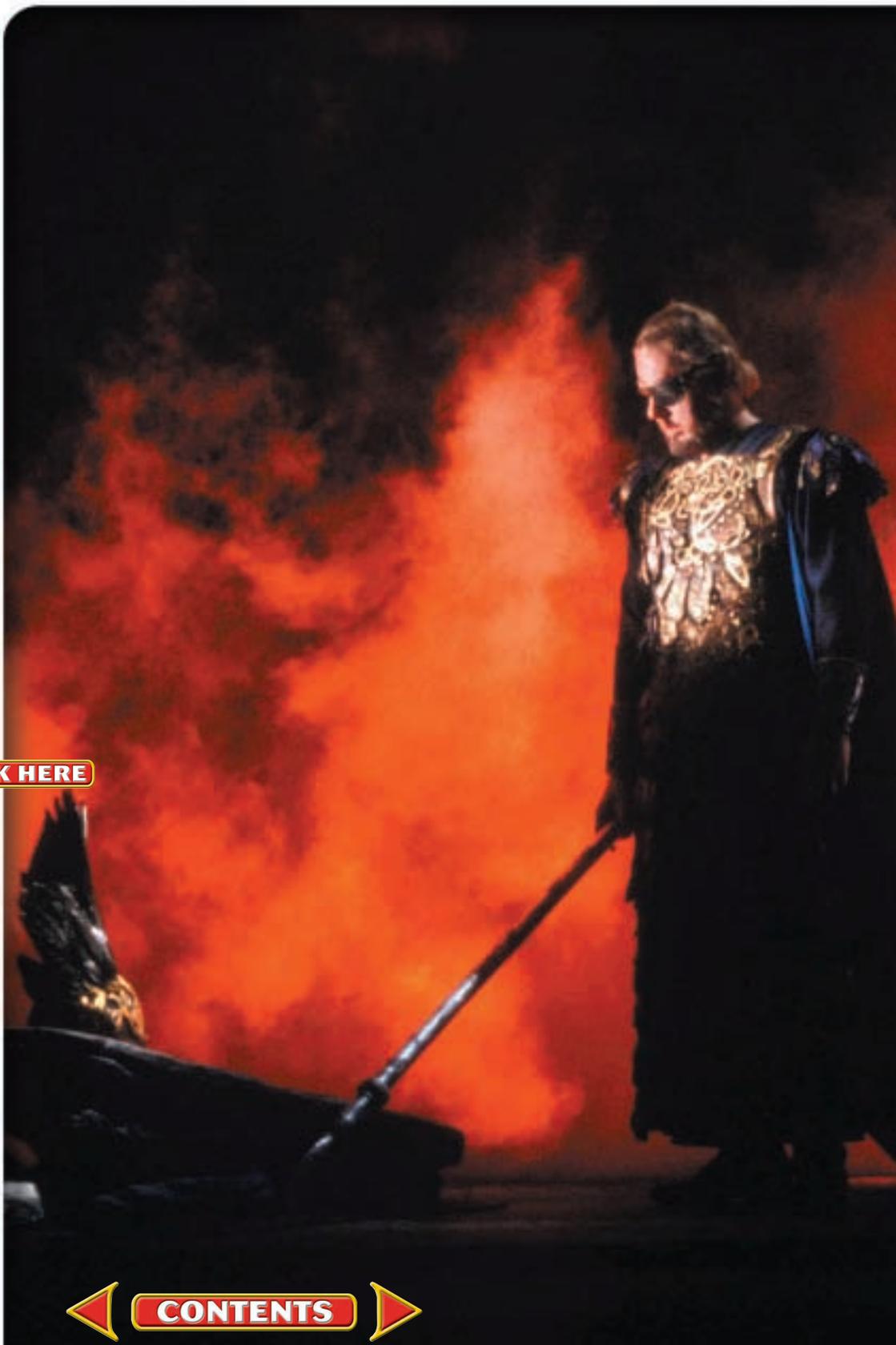
Water collects on a bathroom mirror as you shower, a full bottle of water shatters in a freezer, and a glass object breaks when it is dropped. You will be able to explain such familiar events after you learn more about the different states of matter.

CLICK HERE

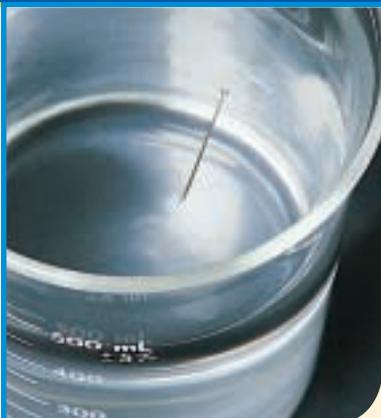


Visit the Chemistry Web site at science.glencoe.com to find links about the states of matter and phase changes.

Solid carbon dioxide is called dry ice. At room temperature, dry ice is used to create the illusion of fog on stage.



DISCOVERY LAB



Materials

pin
600-mL beaker
400 mL water
detergent
dropper

Defying Density

You know that an object sinks or floats in water based on its density. In this activity, you will explore an exception to this rule.

Safety Precautions



Be careful handling the pin, which has a sharp point.

Procedure

1. Pour about 400 mL of water into a 600-mL beaker. Float the pin on the surface of the water.
2. Use a dropper to add one drop of water containing detergent to the beaker. Place the drop on the water surface near the wall of the beaker. Observe what happens.

Analysis

Is a metal pin likely to be more or less dense than water? How does the shape of the pin help it to float? Hypothesize about the reason for the pin's behavior before and after you added the detergent.

Section

13.1

Gases

Objectives

- **Use** the kinetic-molecular theory to explain the behavior of gases.
- **Describe** how mass affects the rates of diffusion and effusion.
- **Explain** how gas pressure is measured and **calculate** the partial pressure of a gas.

Vocabulary

kinetic-molecular theory
elastic collision
temperature
diffusion
Graham's law of effusion
pressure
barometer
pascal
atmosphere
Dalton's law of partial pressures

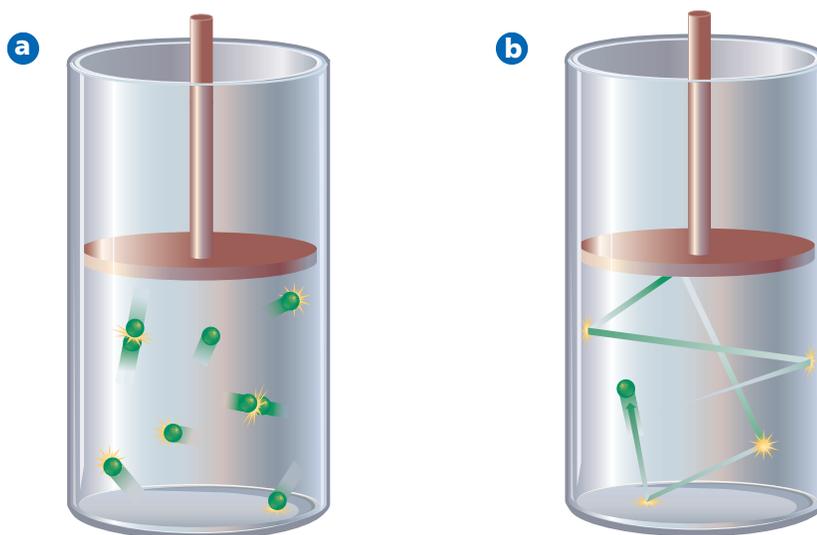
You have learned that the types of atoms present (composition) and their arrangement (structure) determine the chemical properties of matter. Composition and structure also affect the physical properties of liquids and solids. Based solely on physical appearance, you can distinguish water from mercury or gold from graphite. By contrast, substances that are gases at room temperature usually display similar physical properties despite their different compositions. Why is there so little variation in behavior among gases? Why are the physical properties of gases different from those of liquids and solids?

The Kinetic-Molecular Theory

Flemish physician Jan Baptista Van Helmont (1577–1644) used the Greek word *chaos*, which means without order, to describe those products of reactions that had no fixed shape or volume. From the word *chaos* came the term *gas*. By the eighteenth century, scientists knew how to collect gaseous products by displacing water. Now they could observe and measure properties of individual gases. About 1860, Ludwig Boltzmann and James Maxwell, who were working in different countries, each proposed a model to explain the properties of gases. That model is the kinetic-molecular theory. Because all of the gases known to Boltzmann and Maxwell contained molecules, the name of the model refers to molecules. The word *kinetic* comes from a Greek word meaning “to move.” Objects in motion have energy called kinetic energy. The **kinetic-molecular theory** describes the behavior of gases in terms of particles in motion. The model makes several assumptions about the size, motion, and energy of gas particles.

Figure 13-1

- a** Based on its size, is a gas particle more likely to collide with another particle or with the walls of its container?
- b** What happens to the path of a gas particle after a collision?



Particle size Gases consist of small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space. Because gas particles are far apart, there are no significant attractive or repulsive forces among them.

Particle motion Gas particles are in constant, random motion. Particles move in a straight line until they collide with other particles or with the walls of their container, as shown in **Figure 13-1**. Collisions between gas particles are elastic. An **elastic collision** is one in which no kinetic energy is lost. Kinetic energy may be transferred between colliding particles, but the total kinetic energy of the two particles does not change.

Figure 13-2

The air in a life jacket allows the person wearing the jacket to float on the water.

Particle energy Two factors determine the kinetic energy of a particle: mass and velocity. The kinetic energy of a particle can be represented by the equation

$$KE = \frac{1}{2}mv^2$$

in which KE is kinetic energy, m is the mass of the particle, and v is its velocity. Velocity reflects both the speed and the direction of motion. In a sample of a single gas, all particles have the same mass but all particles do not have the same velocity. Therefore, all particles do not have the same kinetic energy. Kinetic energy and temperature are related. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter. At a given temperature, all gases have the same average kinetic energy.

Explaining the Behavior of Gases

Kinetic-molecular theory can help explain the behavior of gases. For example, the constant motion of gas particles allows a gas to expand until it fills its container, such as the flotation device in **Figure 13-2**. What property of gases makes it possible for an air-filled flotation device to work?

Low density Remember that density is mass per unit volume. The density of chlorine gas is 2.95×10^{-3} g/mL at 20°C ; the density of solid gold is 19.3 g/mL. Gold is more than 6500 times as dense as chlorine. This large difference cannot be due only to the difference in mass between gold atoms and chlorine molecules



(about 3:1). As the kinetic-molecular theory states, a great deal of space exists between gas particles. Thus, there are fewer chlorine molecules than gold atoms in the same volume.

Compression and expansion If you squeeze a pillow made of foam, you can compress it; that is, you can reduce its volume. The foam contains air pockets. The large amount of empty space between the air particles in those pockets allows the air to be easily pushed into a smaller volume. When you stop squeezing, the random motion of air particles fills the available space and the pillow expands to its original shape. **Figure 13-3** illustrates what happens to the density of a gas in a container as it is compressed and as it is allowed to expand.

Diffusion and effusion According to the kinetic-molecular theory, there are no significant forces of attraction between gas particles. Thus, gas particles can flow easily past each other. Often, the space into which a gas flows is already occupied by another gas. The random motion of the gas particles causes the gases to mix until they are evenly distributed. **Diffusion** is the term used to describe the movement of one material through another. The term may be new, but you are probably familiar with the process. If you are in the den, can you tell when someone sprays perfume in the bedroom? Perfume particles released in the bedroom diffuse through the air until they reach the den. Particles diffuse from an area of high concentration (the bedroom) to one of low concentration (the den).

The rate of diffusion depends mainly on the mass of the particles involved. Lighter particles diffuse more rapidly than heavier particles. Recall that different gases at the same temperature have the same average kinetic energy as described by the equation $KE = 1/2mv^2$. However, the mass of gas particles varies from gas to gas. For lighter particles to have the same average kinetic energy as heavier particles, they must, on average, have a greater velocity.

Effusion is a process related to diffusion. During effusion, a gas escapes through a tiny opening. What happens when you puncture a container such as a balloon or a tire? In 1846, Thomas Graham did experiments to measure the rates of effusion for different gases at the same temperature. Graham designed his experiment so that the gases effused into a vacuum—a space containing no matter. He discovered an inverse relationship between effusion rates and molar mass. **Graham's law of effusion** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

Graham's law also applies to rates of diffusion, which is logical because heavier particles diffuse more slowly than lighter particles at the same temperature. Using Graham's law, you can set up a proportion to compare the diffusion rates for two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

For example, consider the gases ammonia (NH_3) and hydrogen chloride (HCl). Which gas diffuses faster? Example Problem 13-1 and the photograph on the next page provide quantitative and visual comparisons of the diffusion rates for ammonia and hydrogen chloride.

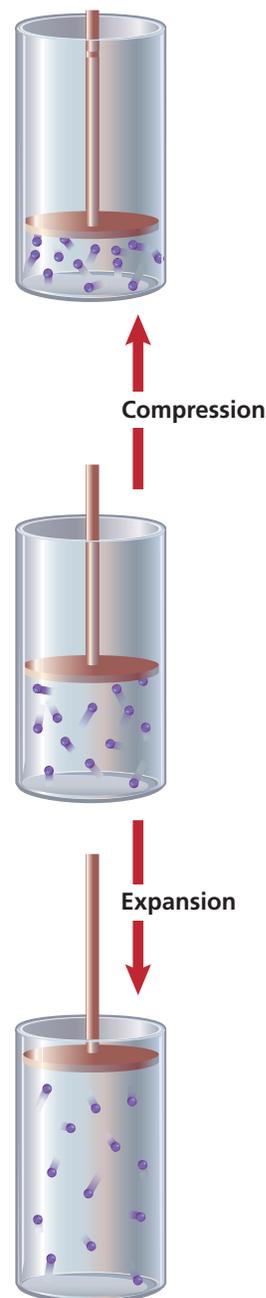


Figure 13-3

The volume of the top cylinder is half that of the middle cylinder. The volume of the bottom cylinder is twice that of the middle cylinder. Compare the density of the gas in the top cylinder with its density in the bottom cylinder.

EXAMPLE PROBLEM 13-1

Finding a Ratio of Diffusion Rates

Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?



When molecules of gaseous HCl (from a bottle of hydrochloric acid) and NH_3 (from a bottle of aqueous ammonia) meet, they react to form the white solid ammonium chloride, NH_4Cl .

1. Analyze the Problem

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known

molar mass_{HCl} = 36.5 g/mol

molar mass_{NH₃} = 17.0 g/mol

Unknown

ratio of diffusion rates = ?

2. Solve for the Unknown

Substitute the known values into Graham's equation and solve.

$$\begin{aligned}\frac{\text{Rate}_{\text{NH}_3}}{\text{Rate}_{\text{HCl}}} &= \sqrt{\frac{\text{molar mass}_{\text{HCl}}}{\text{molar mass}_{\text{NH}_3}}} \\ &= \sqrt{\frac{36.5 \text{ g/mol}}{17.0 \text{ g/mol}}} = \sqrt{2.15} = 1.47\end{aligned}$$

The ratio of diffusion rates is 1.47.

3. Evaluate the Answer

Ammonia molecules diffuse about 1.5 times as fast as hydrogen chloride molecules. This ratio is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer does, too.

PRACTICE PROBLEMS

1. Calculate the ratio of effusion rates for nitrogen (N_2) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide (CO) and carbon dioxide (CO_2).
3. What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

Gas Pressure

If you were to walk across deep snow in boots, you would probably sink into the snow with each step. With a pair of snowshoes like those in **Figure 13-4**, you would be less likely to sink. In each case, the force with which you press down on the snow is related to your mass. With snowshoes, the force would be spread out over a larger area. Therefore, the pressure on any given area of snow would be reduced. **Pressure** is defined as force per unit area. How do the size and style of shoes you wear affect the force you exert on a surface?

Gas particles exert pressure when they collide with the walls of their container. Because an individual gas particle has little mass, it can exert little pressure. However, there are about 10^{22} gas particles in a liter container. With this many particles colliding, the pressure can be substantial. In Chapter 14, you will learn how temperature, volume, and number of moles affect the pressure that a gas exerts.

Practice!

For more practice with diffusion problems, go to **Supplemental Practice Problems** in Appendix A.

Earth is surrounded by an atmosphere that extends into space for hundreds of kilometers. Because the particles in air move in every direction, they exert pressure in all directions. This pressure is called atmospheric pressure, or air pressure. At the surface of Earth, air pressure is approximately equal to the pressure exerted by a 1-kilogram mass on a square centimeter. Air pressure varies at different points on Earth. At the top of a mountain, the mass of the air column pressing down on a square centimeter of Earth is less than the mass of the air column at sea level. Thus, the air pressure at higher altitudes is slightly lower than at sea level.

Measuring air pressure Italian physicist Evangelista Torricelli (1608–1647) was the first to demonstrate that air exerted pressure. He had noticed that water pumps were unable to pump water higher than about ten meters. He hypothesized that the height of a column of liquid would vary with the density of the liquid. To test this idea, Torricelli designed the equipment shown in **Figure 13-5a**. He filled a thin glass tube that was closed at one end with mercury. While covering the open end so that air could not enter, he inverted the tube and placed it (open end down) in a dish of mercury. The open end was below the surface of the mercury in the dish. The height of the mercury in the tube fell to about 75 cm. Mercury is about 13.6 times denser than water. Did the results of his experiment support Torricelli’s hypothesis?

The device that Torricelli invented is called a barometer. A **barometer** is an instrument used to measure atmospheric pressure. As Torricelli showed, the height of the mercury in a barometer is always about 760 mm. The exact height of the mercury is determined by two forces. Gravity exerts a constant downward force on the mercury. This force is opposed by an upward force exerted by air pressing down on the surface of the mercury. Changes in air temperature or humidity cause air pressure to vary. An increase in air pressure causes the mercury to rise; a decrease causes the mercury to fall.

A manometer is an instrument used to measure gas pressure in a closed container. In a manometer, a flask is connected to a U-tube that contains mercury. In **Figure 13-5b**, there is no gas in the flask. The mercury is at the same height in each arm of the U-tube. In **Figure 13-5c**, there is gas in the flask. When the valve between the flask and the U-tube is opened, gas particles diffuse out of the flask into the U-tube. The released gas particles push down on the mercury in the tube. What happens to the height of the mercury in each arm of the U-tube? The difference in the height of the mercury in the two arms is used to calculate the pressure of the gas in the flask.

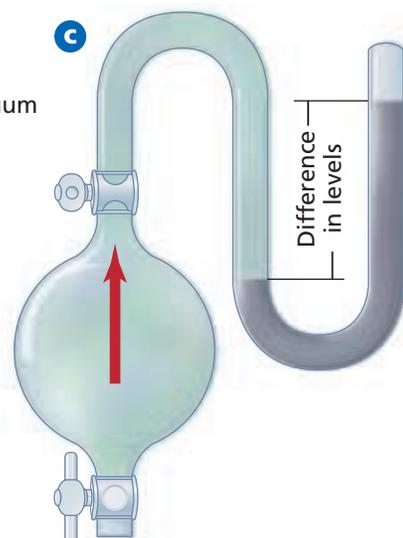
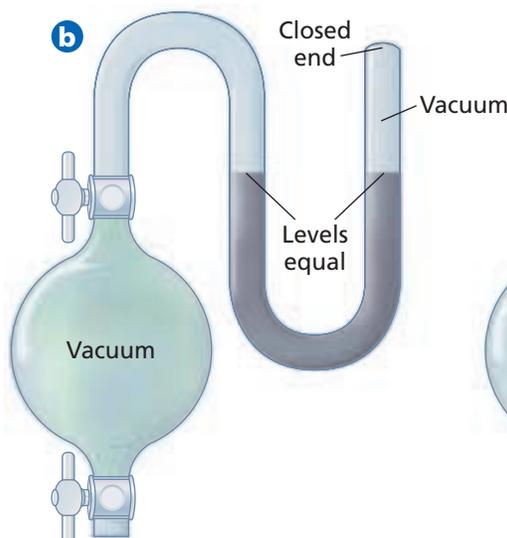
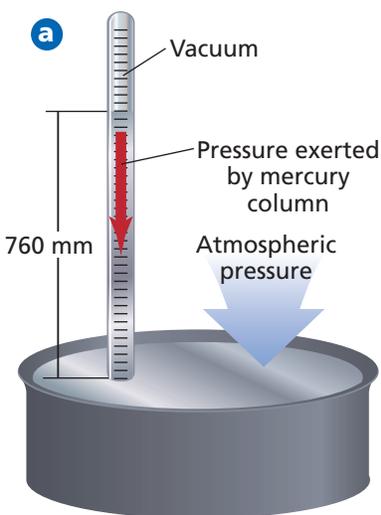


Figure 13-4

Snowshoes reduce the force on a given area. What devices other than snowshoes can people use to travel across deep snow?

Figure 13-5

a A barometer measures air pressure. **b** A manometer measures the pressure of an enclosed gas. Before gas is released into the U-tube, the mercury is at the same height in each arm. **c** After gas is released into the U-tube, the heights in the two arms are no longer equal.

Table 13-1

Comparison of Pressure Units		
Unit	Compared with 1 atm	Compared with 1 kPa
kilopascal (kPa)	1 atm = 101.3 kPa	
millimeters of mercury (mm Hg)	1 atm = 760 mm Hg	1 kPa = 7.501 mm Hg
torr	1 atm = 760 torr	1 kPa = 7.501 torr
pounds per square inch (psi or lb/in ²)	1 atm = 14.7 psi	1 kPa = 0.145 psi
atmosphere (atm)		1 kPa = 0.009 869 atm

Units of pressure The SI unit of pressure is the pascal (Pa). It is named for Blaise Pascal, a French mathematician and philosopher. The pascal is derived from the SI unit of force, the newton (N), which is derived from three SI base units: the kilogram, the meter, and the second. One **pascal** is equal to a force of one newton per square meter: $1 \text{ Pa} = 1 \text{ N/m}^2$. Many fields of science still use more traditional units of pressure. For example, engineers often report pressure as pounds per square inch (psi). The pressures measured by barometers and manometers can be reported in millimeters of mercury (mm Hg). There also is a unit called the torr, which is named to honor Torricelli. One torr is equal to one mm Hg.

At sea level, the average air pressure is 760 mm Hg when the temperature is 0°C. Air pressure often is reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mm Hg or 760 torr or 101.3 kilopascals (kPa). **Table 13-1** compares different units of pressure. Because the units 1 atm, 760 mm Hg, and 760 torr are defined units, they have as many significant figures as needed when used in calculations. Do the **problem-solving LAB** to see how the combined pressure of air and water affects divers.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

problem-solving LAB

How are the depth of a dive and pressure related?

Making and Using Graphs For centuries, people have dived deep into the sea to collect items such as pearls. Nowadays, single-breath diving has become competitive. On January 18, 2000, Francisco "Pipin" Ferreras from Cuba set a new record with a dive of 162 meters. He was underwater for 3.2 minutes.

Analysis

Use the data in the table to make a graph of pressure versus depth.

- How are pressure and depth related?
- What would the pressure be at the surface of the water? What does this value represent?
- What was the pressure on Francisco at 162 m?

Pressure Versus Depth

Depth of dive (m)	Pressure (atm)
10	2.0
20	3.0
30	4.0
40	5.0
50	6.0
60	7.0

Thinking Critically

- Using the equation for slope, calculate the slope of your graph. Then write an equation to express the relationship between pressure and depth. $\left(\text{slope} = \frac{\Delta y}{\Delta x} = \frac{y_2 - y_1}{x_2 - x_1}\right)$

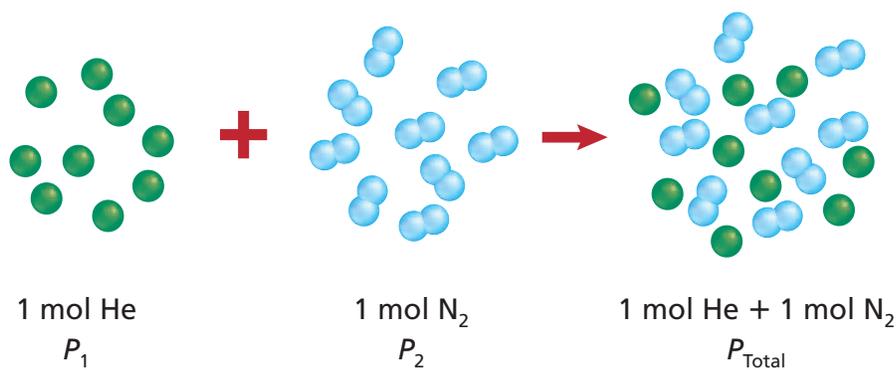


Figure 13-6

How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container? Refer to **Table C-1** in Appendix C for a key to atom color conventions.

Dalton's law of partial pressures When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of one mole of any gas is the same. Dalton's law of partial pressures can be summarized as

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots P_n$$

P_{total} represents the total pressure of a mixture of gases. P_1 , P_2 , and so on represent the partial pressures of each gas in the mixture. **Figure 13-6** shows what happens when one mole of helium and one mole of nitrogen are combined in a single closed container.

EXAMPLE PROBLEM 13-2

Finding the Partial Pressure of a Gas

A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2 , if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1. Analyze the Problem

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$$\begin{aligned} P_{\text{N}_2} &= 0.12 \text{ atm} \\ P_{\text{CO}_2} &= 0.70 \text{ atm} \\ P_{\text{total}} &= 0.97 \text{ atm} \end{aligned}$$

Unknown

$$P_{\text{O}_2} = ? \text{ atm}$$

2. Solve for the Unknown

Rearrange the equation to solve for the unknown value, P_{O_2} .

$$\begin{aligned} P_{\text{O}_2} &= P_{\text{total}} - P_{\text{CO}_2} - P_{\text{N}_2} \\ P_{\text{O}_2} &= 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm} \\ P_{\text{O}_2} &= 0.15 \text{ atm} \end{aligned}$$

3. Evaluate the Answer

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure 0.97 atm. The answer has two significant figures to match the data.

Practice!

For more practice with partial pressure problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mm Hg and the partial pressure of helium is 439 mm Hg?
- Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.
- Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.

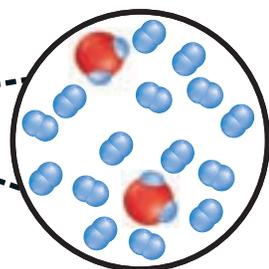
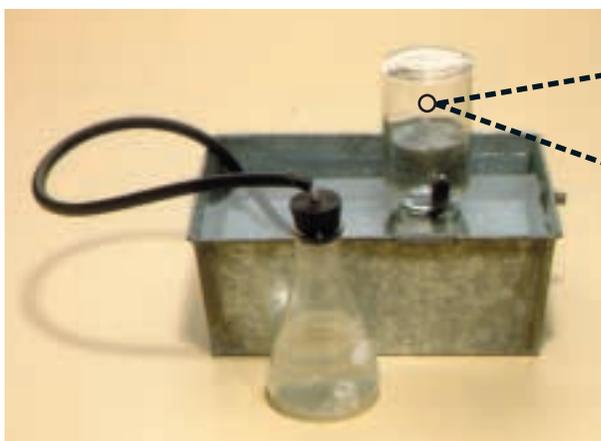
Dalton's law of partial pressures can be used to determine the amount of gas produced by a reaction. The gas produced is bubbled into an inverted container of water, as shown in **Figure 13-7**. As the gas collects, it displaces the water. The gas collected in the container will be a mixture of hydrogen and water vapor. Therefore, the total pressure inside the container will be the sum of the partial pressures of hydrogen and water vapor.

The partial pressures of gases at the same temperature are related to their concentration. The partial pressure of water vapor has a fixed value at a given temperature. You can look up the value in a reference table. At 20°C, the partial pressure of water vapor is 2.3 kPa. You can calculate the partial pressure of hydrogen by subtracting the partial pressure of water vapor from the total pressure. If the total pressure of the hydrogen and water mixture is 95.0 kPa, what is the partial pressure of hydrogen at 20°C?

As you will learn in Chapter 14, knowing the pressure, volume, and temperature of a gas allows you to calculate the number of moles of the gas. Temperature and volume can be measured during an experiment. Once the temperature is known, the partial pressure of water vapor is used to calculate the pressure of the gas. The known values for volume, temperature, and pressure are then used to find the number of moles.

Figure 13-7

In the flask, sulfuric acid (H_2SO_4) reacts with zinc to produce hydrogen gas. The hydrogen is collected at 20°C.



Section 13.1 Assessment

- What assumption of the kinetic-molecular theory explains why a gas can expand to fill a container?
- How does the mass of a gas particle affect its rate of effusion?
- Suppose two gases in a container have a total pressure of 1.20 atm. What is the pressure of gas B if the partial pressure of gas A is 0.75 atm?
- Explain how changes in atmospheric pressure affect the height of the column of mercury in a barometer.
- Recognizing Cause and Effect** Explain why a tire or balloon expands when air is added.
- Thinking Critically** Explain why the container of water must be inverted when a gas is collected by displacement of water.

If all particles of matter at room temperature have the same average kinetic energy, why are some materials gases while others are liquids or solids? The answer lies with the attractive forces within and between particles. The attractive forces that hold particles together in ionic, covalent, and metallic bonds are called intramolecular forces. The prefix *intra-* means “within.” For example, intramural sports are competitions among teams from within a single school. The term *molecular* can refer to atoms, ions, or molecules. **Table 13-2** summarizes what you learned about intramolecular forces in Chapters 8 and 9.

Intermolecular Forces

Intramolecular forces do not account for all attractions between particles. There are forces of attraction called intermolecular forces. The prefix *inter-* means “between” or “among.” For example, an interview is a conversation between two people. Intermolecular forces can hold together identical particles, such as water molecules in a drop of water, or two different types of particles, such as carbon atoms in graphite and the cellulose particles in paper. The three intermolecular forces that will be discussed in this section are dispersion forces, dipole–dipole forces, and hydrogen bonds. Although some intermolecular forces are stronger than others, all intermolecular forces are weaker than intramolecular, or bonding, forces.

Dispersion forces Recall that oxygen molecules are nonpolar because electrons are evenly distributed between the equally electronegative oxygen atoms. Under the right conditions, however, oxygen molecules can be compressed into a liquid. For oxygen to be compressed, there must be some force of attraction between its molecules. The force of attraction between oxygen molecules is called a dispersion force. **Dispersion forces** are weak forces that result from temporary shifts in the density of electrons in electron clouds. Dispersion forces are sometimes called London forces after the German-American physicist who first described them, Fritz London.

Remember that the electrons in an electron cloud are in constant motion. When two nonpolar molecules are in close contact, especially when they collide, the

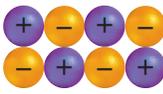
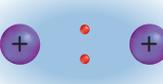
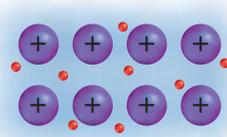
Objectives

- **Describe** and **compare** intramolecular and intermolecular forces.
- **Distinguish** among intermolecular forces.

Vocabulary

dispersion force
dipole–dipole force
hydrogen bond

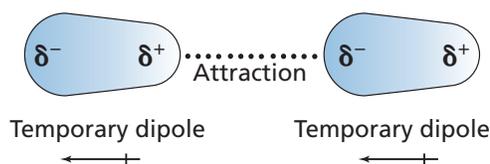
Table 13-2

Comparison of Intramolecular Forces			
Force	Model	Basis of attraction	Example
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H ₂
Metallic		metal cations and mobile electrons	Fe

electron cloud of one molecule repels the electron cloud of the other molecule. The electron density around each nucleus is, for a moment, greater in one region of each cloud. Each molecule forms a temporary dipole.

Figure 13-8

What do the δ^- and δ^+ signs on a temporary dipole represent?



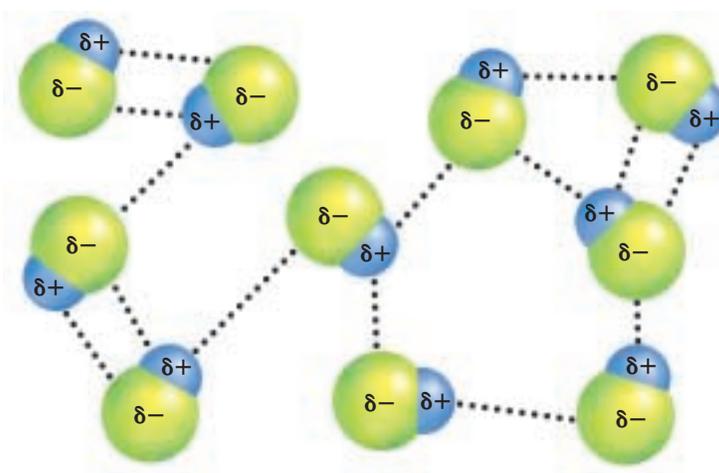
When temporary dipoles are close together, a weak dispersion force exists between oppositely charged regions of the dipoles, as shown in **Figure 13-8**.

Due to the temporary nature of the dipoles, dispersion forces are the weakest intermolecular force. Dispersion forces exist between all particles, but they play a significant role only when there are no stronger forces of attraction acting on particles. Dispersion forces are the dominant force of attraction between identical nonpolar molecules. These forces can have a noticeable effect as the number of electrons involved increases. For example, fluorine, chlorine, bromine, and iodine exist as diatomic molecules. Recall that the number of nonvalence electrons increases from fluorine to chlorine to bromine to iodine. Because the larger halogen molecules have more electrons, there can be a greater difference between the positive and negative regions of their temporary dipoles and, thus, stronger dispersion forces. This difference in dispersion forces explains why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature.

Dipole–dipole forces Polar molecules contain permanent dipoles; that is, some regions of a polar molecule are always partially negative and some regions of the molecule are always partially positive. Attractions between oppositely charged regions of polar molecules are called **dipole–dipole forces**. Neighboring polar molecules orient themselves so that oppositely charged regions line up. When hydrogen chloride gas molecules approach, the partially positive hydrogen atom in one molecule is attracted to the partially negative chlorine atom in another molecule. **Figure 13-9** shows multiple attractions among hydrogen chloride molecules.

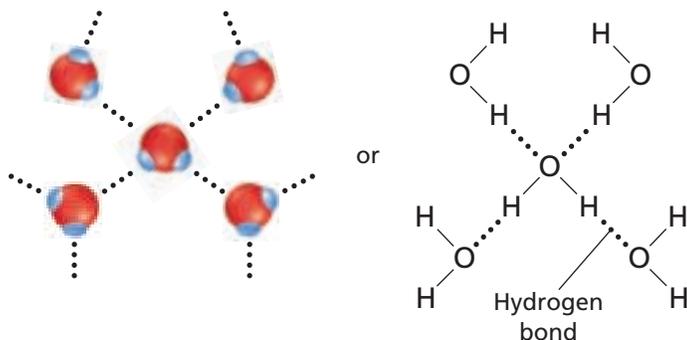
Figure 13-9

The degree of polarity in a polar molecule depends on the relative electronegativity values of the elements in the molecule. What are the electronegativity values for hydrogen and chlorine?



Because the dipoles are permanent, you might expect dipole–dipole forces to be stronger than dispersion forces. This prediction will hold true as long as the molecules being compared have approximately the same mass.

Hydrogen bonds One special type of dipole–dipole attraction is called a hydrogen bond. A **hydrogen bond** is a dipole–dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair. For a hydrogen bond to form, hydrogen must be bonded to either a fluorine, oxygen, or nitrogen atom. These atoms are electronegative enough to cause a large partial positive charge on the hydrogen atom, yet small enough that their lone pairs of electrons can come close to hydrogen atoms. Consider, for example, a water molecule. In a water molecule, the hydrogen atoms have a large partial positive charge and the oxygen atom has a large partial negative charge. When water molecules approach, a hydrogen atom on one molecule is attracted to the oxygen atom on the other molecule, as shown in **Figure 13-10**.



Hydrogen bonds explain why water is a liquid at room temperature while compounds of comparable mass are gases. Look at the data in **Table 13-3**. The difference between methane and water is easy to explain. Because methane molecules are nonpolar, the only forces holding the molecules together are relatively weak dispersion forces. The difference between ammonia and water is not as obvious. Molecules of both compounds can form hydrogen bonds. Yet, ammonia is a gas at room temperature, which indicates that the attractive forces between ammonia molecules are not as strong. Because oxygen atoms are more electronegative than nitrogen atoms, the O–H bonds in water are more polar than the N–H bonds in ammonia. As a result, the hydrogen bonds between water molecules are stronger than those between ammonia molecules.

Table 13-3

Properties of Three Molecular Compounds		
Compound	Molar mass (g)	Boiling point (°C)
Water (H ₂ O)	18.0	100
Methane (CH ₄)	16.0	–164
Ammonia (NH ₃)	17.0	–33.4

Figure 13-10

The hydrogen bonds between water molecules are stronger than typical dipole–dipole attractions because the bond between hydrogen and oxygen is highly polar.

Section 13.2 Assessment

- Why are dipole–dipole forces stronger than dispersion forces for molecules of comparable mass?
- Which of the molecules listed below can form hydrogen bonds? For which of the molecules would dispersion forces be the only intermolecular force? Give reasons for your answers.
 - H₂
 - NH₃
 - HCl
 - HF
- Predicting** Make a prediction about the relative boiling points of the noble gases. Give a reason for your answer.
- Thinking Critically** In a methane molecule (CH₄), there are 4 single covalent bonds. In an octane molecule (C₈H₁₈), there are 25 single covalent bonds. How does the number of bonds affect the dispersion forces in samples of methane and octane? Which compound is a gas at room temperature? Which is a liquid?

Objectives

- **Apply** kinetic-molecular theory to the behavior of liquids and solids.
- **Relate** properties such as viscosity, surface tension, and capillary action to intermolecular forces.
- **Compare** the structures and properties of different types of solids.

Vocabulary

viscosity
surface tension
surfactant
crystalline solid
unit cell
amorphous solid

Although the kinetic-molecular theory was developed to explain the behavior of gases, the model can be applied to liquids and solids. When applying the kinetic-molecular theory to these states of matter, you must consider the forces of attraction between particles as well as their energy of motion.

Liquids

In Chapter 3, you learned that a liquid can take the shape of its container but that its volume is fixed. In other words, the particles in a liquid can flow to adjust to the shape of a container, but the liquid cannot expand to fill its container. Kinetic-molecular theory predicts the constant motion of the liquid particles. Individual liquid molecules do not have fixed positions in the liquid. However, forces of attraction between liquid particles limit their range of motion so that the particles remain closely packed in a fixed volume.

Density and compression At 25°C and one atmosphere of air pressure, liquids are much denser than gases. The density of a liquid is much greater than that of its vapor at the same conditions. For example, liquid water is about 1250 times denser than water vapor at 25°C and one atmosphere of pressure. Because they are at the same temperature, both gas and liquid particles have the same average kinetic energy. Thus, the higher density of liquids must be traced to the intermolecular forces that hold particles together.

Like gases, liquids can be compressed. But the change in volume for liquids is much smaller because liquid particles are already tightly packed together. An enormous amount of pressure must be applied to reduce the volume of a liquid by even a few percent.

Fluidity Fluidity is the ability to flow. Gases and liquids are classified as fluids because they can flow. A liquid can diffuse through another liquid as shown in **Figure 13-11**. A liquid diffuses more slowly than a gas at the same

Figure 13-11

How much time do you think it took for the food coloring to completely mix with the water?



temperature, however, because intermolecular attractions interfere with the flow. Thus, liquids are less fluid than gases. A comparison between water and natural gas can illustrate this difference. When there is a leak in a basement water pipe, the water remains in the basement unless the amount of water released exceeds the volume of the basement. Damage can be limited if the water supply is shut off quickly.

Natural gas, or methane, is the fuel burned in gas furnaces, gas hot-water heaters, and gas stoves. Gas that leaks from a gas pipe can diffuse throughout a house. Because natural gas is odorless, companies that supply the fuel include a compound with a distinct odor to warn customers of a leak. If the gas is effusing through a small hole in a pipe, the customer has time to shut off the gas supply, open windows to allow the gas to diffuse, and call the gas company to report the leak. If there is a break in a gas line, the customer must leave the house immediately because natural gas can explode when it comes in contact with an open flame or spark.

Viscosity Do you know the meaning of the phrase “slow as molasses”? Have you ever tried to get ketchup to flow out of a bottle? If so, you are already familiar with the concept of viscosity. **Viscosity** is a measure of the resistance of a liquid to flow. The particles in a liquid are close enough for attractive forces to slow their movement as they flow past one another. The viscosity of a liquid is determined by the type of intermolecular forces involved, the shape of the particles, and the temperature.

The stronger the attractive forces, the higher the viscosity. If you have used glycerol in the laboratory to help insert a glass tube into a rubber stopper, you know that glycerol is a viscous liquid. **Figure 13-12** uses structural formulas to show the hydrogen bonding that makes glycerol so viscous. Because it has three hydrogen atoms attached to oxygen atoms, a glycerol molecule can form three hydrogen bonds with other glycerol molecules.

The size and shape of particles affect viscosity. Recall that the overall kinetic energy of a particle is determined by its mass and velocity. Suppose the attractive forces between molecules in liquid A and liquid B are similar. If the molecules in liquid A are more massive than the molecules in liquid B, liquid A will have a greater viscosity. Liquid A’s molecules will, on average, move more slowly than the molecules in liquid B. Molecules with long chains have a higher viscosity than shorter, more compact molecules, assuming the molecules exert the same type of attractive forces. Within the long chains, there is less distance between atoms on neighboring molecules and, thus, a greater chance for attractions between atoms. The long molecules in cooking oils and motor oils make these liquids thick and slow to pour.

Viscosity and temperature Viscosity decreases with temperature. When you pour a tablespoon of cooking oil into a frying pan, the oil tends not to spread across the bottom of the pan until you heat the oil. With the increase in temperature, there is an increase in the average kinetic energy of the oil molecules. The added energy makes it easier for the molecules to overcome the intermolecular forces that keep the molecules from flowing.

In **Figure 13-13**, the motor oil that keeps the moving parts of an internal combustion engine lubricated is being replaced.

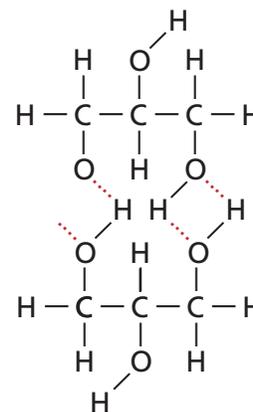


Figure 13-12

How many hydrogen bonds can a glycerol molecule form?

Figure 13-13

The viscosity of motor oil increases in summer because additives in the oil change their shape as the temperature rises.



What could happen to an engine if there were no motor oil to lubricate its parts? Because temperature changes affect the viscosity of motor oil, people often used different motor oil blends in winter and summer. The motor oil used in winter was designed to keep flowing at low temperatures. The motor oil used in summer was more viscous so that it would maintain sufficient viscosity on extremely hot days or during long trips. Today, additives in motor oil can help adjust the viscosity so that the same oil blend can be used all year. Molecules in the additives are compact spheres with relatively low viscosity at cool temperatures. At high temperatures, the shape of the additive molecules changes to long strands. These strands get tangled with the oil molecules, which increases the viscosity of the oil.

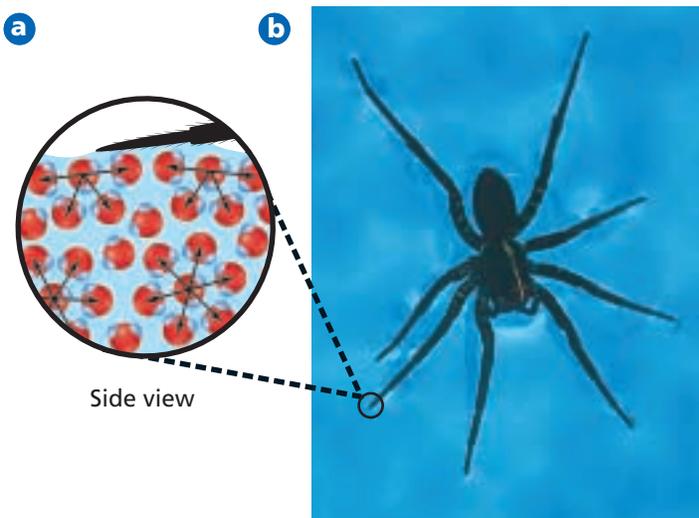
Surface tension Intermolecular forces do not have an equal effect on all particles in a liquid, as shown in **Figure 13-14a**. Particles in the middle of the liquid can be attracted to particles above them, below them, and to either side. For particles at the surface of the liquid, there are no attractions from above to balance the attractions from below. Thus, there is a net attractive force pulling down on particles at the surface. The surface tends to have the smallest possible area and to act as though it is stretched tight like the head of a drum. For the surface area to increase, particles from the interior must move to the surface. It takes energy to overcome the attractions holding these particles in the interior. The energy required to increase the surface area of a liquid by a given amount is called **surface tension**. Surface tension is a measure of the inward pull by particles in the interior.

In general, the stronger the attractions between particles, the greater the surface tension. Water has a high surface tension because its molecules can form multiple hydrogen bonds. Drops of water are shaped like spheres because the surface area of a sphere is smaller than the surface area of any other shape of similar volume. Water's high surface tension is what allows the spider in **Figure 13-14b** to walk on the surface of a pond.

It is difficult to remove dirt from skin or clothing using only water. Because dirt particles cannot penetrate the surface of the water drops, the water cannot remove the dirt. What happened when you added a drop of detergent to the beaker in the **DISCOVERY LAB**? Soaps and detergents decrease the surface tension of water by disrupting the hydrogen bonds between water molecules. When the bonds are broken, the water spreads out. Compounds that lower the surface tension of water are called surface active agents or **surfactants**.

Figure 13-14

- a** Particles at the surface are drawn toward the interior of a liquid until attractive and repulsive forces are balanced.
- b** How does the spider's structure help it stay afloat on water?



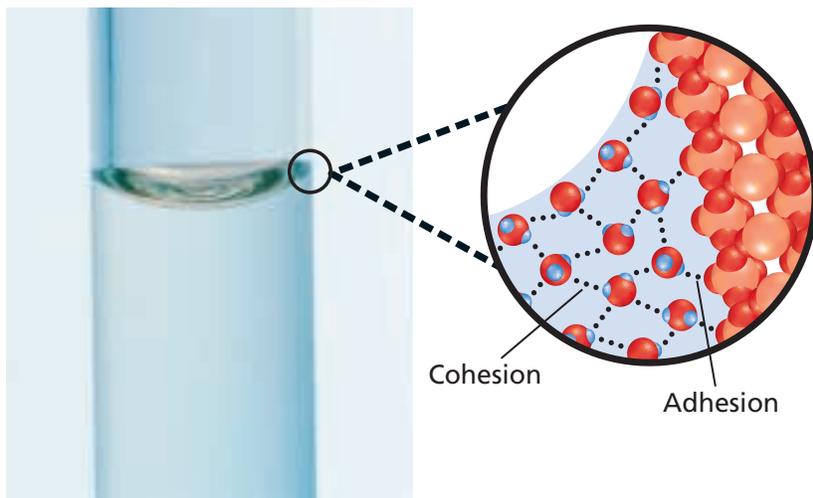


Figure 13-15

The surface of the water in a graduated cylinder is concave because water molecules are more strongly attracted to the silicon dioxide in glass than to other water molecules.

Capillary action When water is placed in a narrow container such as a graduated cylinder, you can see that the surface of the water is not straight. The surface forms a concave meniscus; that is, the surface dips in the center. **Figure 13-15** models what is happening to the water at the molecular level. There are two types of forces at work: cohesion and adhesion. Cohesion describes the force of attraction between identical molecules. Adhesion describes the force of attraction between molecules that are different. Because the adhesive forces between water molecules and the silicon dioxide in glass are greater than the cohesive forces between water molecules, the water rises along the inner walls of the cylinder.

If the cylinder is extremely narrow, a thin film of water will be drawn upward. Narrow tubes are called capillary tubes. This movement of a liquid such as water is called capillary action, or capillarity. Capillary action helps explain how paper towels can absorb large amounts of water. The water is drawn into the narrow spaces between the cellulose fibers in paper towels by capillary action. In addition, the water molecules form hydrogen bonds with cellulose molecules. These same factors account for the absorbent properties of disposable diapers. Water is drawn from the surface of the diaper to the interior by capillary action. The diaper can absorb about 200 times its mass in fluid.

Solids

According to the kinetic-molecular theory, a mole of solid particles has as much kinetic energy as a mole of liquid particles at the same temperature. By definition, the particles in a solid must be in constant motion. So why do solids have a definite shape and volume? For a substance to be a solid rather than a liquid at a given temperature, there must be strong attractive forces acting between particles in the solid. These forces limit the motion of the particles to vibrations around fixed locations in the solid. Thus, there is more order in a solid than in a liquid. Because of this order, solids are much less fluid than liquids and gases. In fact, solids are not classified as fluids.

Density of solids In general, the particles in a solid are more closely packed than those in a liquid. Thus, most solids are more dense than most liquids. When the liquid and solid states of a substance coexist, the solid almost always sinks in the liquid. In **Figure 13-16**, solid cubes of benzene sink in liquid benzene because solid benzene is more dense than liquid benzene. There is about a 10% difference in density between the solid and

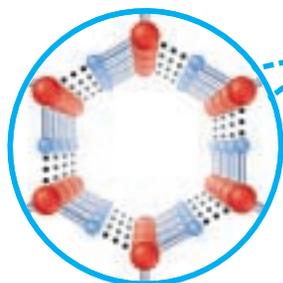
Figure 13-16

Ice cubes float in water. Benzene cubes sink in liquid benzene because solid benzene is more dense than liquid benzene.



Figure 13-17

An iceberg can float because the rigid, three-dimensional structure of ice keeps water molecules farther apart than they are in liquid water.



liquid states of most substances. Because the particles in a solid are closely packed, ordinary amounts of pressure will not change the volume of a solid.

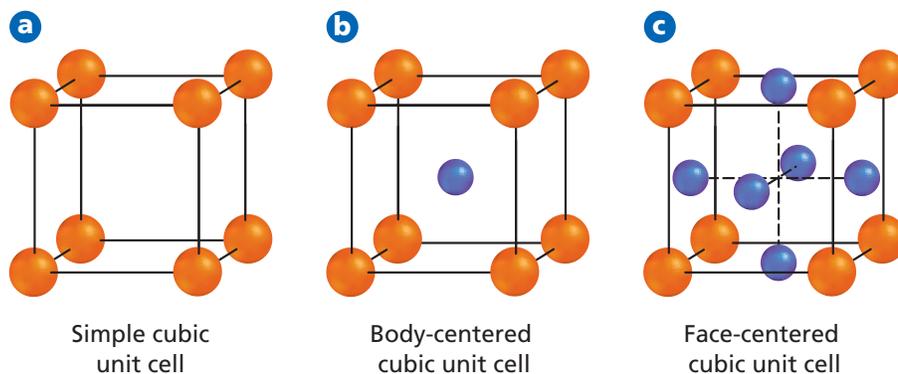
You cannot predict the relative densities of ice and liquid water based on benzene. Ice cubes and icebergs float because water is less dense as a solid than it is as a liquid. **Figure 13-17** shows the reason for the exception. The water molecules in ice are less closely packed together than in liquid water; that is, there is more space between the molecules in ice. As a result, there are more particles per unit volume in liquid water than in solid water.

Crystalline solids Although ice is unusual in its density, ice is typical of most solids in that its molecules are packed together in a predictable way. A **crystalline solid** is a solid whose atoms, ions, or molecules are arranged in an orderly, geometric, three-dimensional structure. The individual pieces of a crystalline solid are called crystals. In Chapter 8, you learned that the locations of ions in an ionic solid can be represented as points on a framework called a crystal lattice. A **unit cell** is the smallest arrangement of connected points that can be repeated in three directions to form the lattice. The relationship of a unit cell to a crystal lattice is similar to that of a formula unit to an ionic compound. Both the unit cell and the formula unit are small, representative parts of a much larger whole. **Figure 13-18** shows three ways that atoms or ions can be arranged in a cubic unit cell.

The shape of a crystalline solid is determined by the type of unit cell from which its lattice is built. **Figure 13-19** shows seven categories of crystals based on shape. Crystal shapes differ because the surfaces, or faces, of unit cells do not always meet at right angles and the edges of the faces vary in length. In **Figure 13-19**, the edges are labeled a , b , and c ; the angles at which the faces meet are labeled α , β , and γ . Use the drawing as you model crystal shapes in the **miniLAB**.

Figure 13-18

These drawings show three of the ways particles are arranged in crystal lattices. Each sphere represents a particle. **a** Particles are arranged only at the corners of the cube. **b** There is a particle in the center of the cube. **c** There are particles in the center of each of the six cubic faces, but no particle in the center of the cube itself.



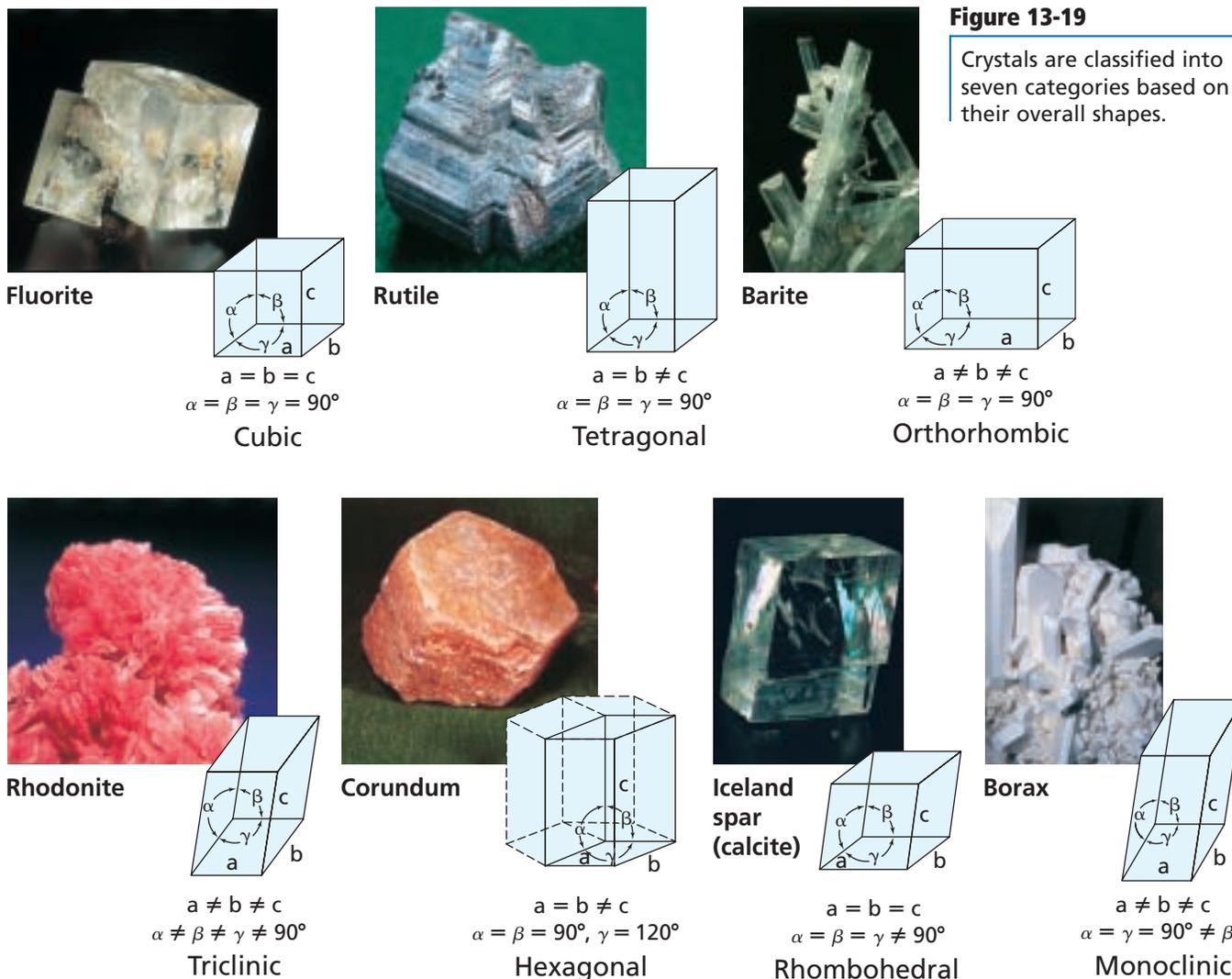


Figure 13-19
Crystals are classified into seven categories based on their overall shapes.

miniLAB

Crystal Unit Cell Models

Formulating Models You can make physical models that illustrate the structures of crystals.

Materials 12 plastic or paper soda straws, 22- or 26-gauge wire, scissors

Procedure 

1. Cut four soda straws into thirds. Wire the pieces to make a cube. All angles are 90°.
2. To model a rhombohedral crystal, deform the cube from step 1 until no angles are 90°.
3. To model a hexagonal crystal, flatten the model from step 2 until it looks like a pie with six slices.
4. To model a tetragonal crystal, cut four straws in half. Cut four of the pieces in half again. Wire the eight shorter pieces to make four

square ends. Use the longer pieces to connect the square ends.

5. To model the orthorhombic crystal, cut four straws in half. Cut 1/3 off four of the halves. Connect the four long, four medium, and four short pieces so that each side is a rectangle.
6. To model the monoclinic crystal, deform the model from step 5 along one axis. To model the triclinic crystal, deform the model from step 5 until it has no 90° angles.

Analysis

1. Which two models have three axes of equal length? How do these models differ?
2. Which model includes a square and rectangle?
3. Which models have three unequal axes?
4. Do you think crystals are perfect or do they have defects? Explain your answer.

Table 13-4

Types of Crystalline Solids			
Type	Unit particles	Characteristics of solid phase	Examples
atomic	atoms	soft to very soft; very low melting points; poor conductivity	group 8A elements
molecular	molecules	fairly soft; low to moderately high melting points; poor conductivity	I ₂ , H ₂ O, NH ₃ , CO ₂ , C ₁₂ H ₂₂ O ₁₁ (table sugar)
covalent network	atoms connected by covalent bonds	very hard; very high melting points; often poor conductivity	diamond (C) and quartz (SiO ₂)
ionic	ions	hard; brittle; high melting points; poor conductivity	NaCl, KBr, CaCO ₃
metallic	atoms surrounded by mobile valence electrons	soft to hard; low to very high melting points; malleable and ductile; excellent conductivity	all metallic elements

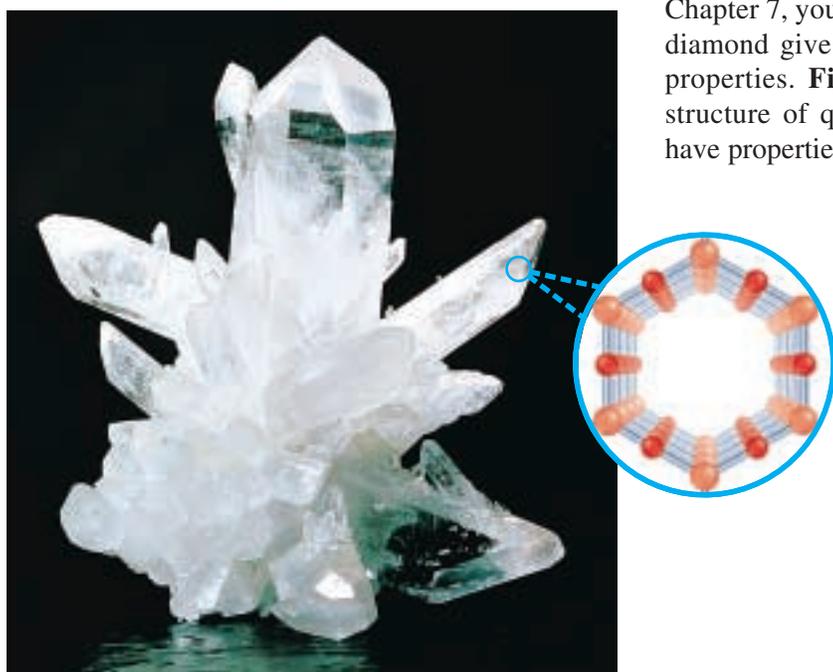
Crystalline solids can be classified into five categories based on the types of particles they contain: atomic solids, molecular solids, covalent network solids, ionic solids, and metallic solids. **Table 13-4** summarizes the general characteristics of each category and provides examples. The only atomic solids are noble gases. Their properties reflect the weak dispersion forces between the atoms.

Molecular solids In molecular solids, the molecules are held together by dispersion forces, dipole–dipole forces, or hydrogen bonds. Most molecular compounds are not solids at room temperature. Even water, which can form strong hydrogen bonds, is a liquid at room temperature. Molecular compounds such as sugar are solids at room temperature because of their large molar masses. With larger molecules, many weak attractions can combine to hold the molecules together. Because they contain no ions, molecular solids are poor conductors of heat and electricity.

Covalent network solids Atoms such as carbon and silicon, which can form multiple covalent bonds, are able to form covalent network solids. In Chapter 7, you learned how the structures of graphite and diamond give those solid allotropes of carbon different properties. **Figure 13-20** shows the covalent network structure of quartz. Based on its structure, will quartz have properties similar to diamond or graphite?

Figure 13-20

The most common kind of quartz has a hexagonal crystal structure.



Ionic solids Remember that each ion in an ionic solid is surrounded by ions of opposite charge. The type of ions and the ratio of ions determine the structure of the lattice and the shape of the crystal. The network of attractions that extends throughout an ionic crystal gives these compounds their high melting points and hardness. Ionic crystals are strong, but brittle. When ionic crystals are struck, the cations and anions are shifted from their fixed positions. Repulsions between ions of like charge cause the crystal to shatter.

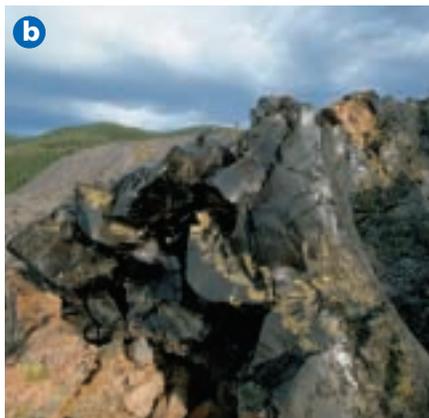


Figure 13-21

a The material a wire is made from, its thickness, and its length are variables that affect the flow of electrons through the wire. **b** Native Americans used the obsidian that forms when lava cools to make arrowheads and knives.

Metallic solids Recall from Chapter 8 that metallic solids consist of positive metal ions surrounded by a sea of mobile electrons. The strength of the metallic bonds between cations and electrons varies among metals and accounts for their wide range of physical properties. For example, tin melts at 232°C , but nickel melts at 1455°C . The mobile electrons make metals malleable—easily hammered into shapes—and ductile—easily drawn into wires. When force is applied to a metal, the electrons shift and thereby keep the metal ions bonded in their new positions. Read **Everyday Chemistry** at the end of the chapter to learn about shape-memory metals. Mobile electrons make metals good conductors of heat and electricity. Power lines carry electricity from power plants to homes and businesses and to the electric train shown in **Figure 13-21a**.

Amorphous solids Not all solids contain crystals. An **amorphous solid** is one in which the particles are not arranged in a regular, repeating pattern. The term amorphous is derived from a Greek word that means “without shape.” An amorphous solid often forms when a molten material cools too quickly to allow enough time for crystals to form. **Figure 13-21b** shows an example of an amorphous solid.

Glass, rubber, and many plastics are amorphous solids. Recent studies have shown that glass may have some structure. When X-ray diffraction is used to study glass, there appears to be no pattern to the distribution of atoms. When neutrons are used instead, an orderly pattern of silicate units can be detected in some regions. Researchers hope to use this new information to control the structure of glass for optical applications and to produce glass that can conduct electricity.

Careers Using Chemistry

Materials Engineer

Would you like to change the atomic and molecular structure of substances to create products ranging from computer chips to skis? If so, consider a career as a materials engineer.

Materials engineers specialize in metals or ceramics. They create new materials to meet certain mechanical, electrical, or chemical standards. They also evaluate existing materials for new applications. Materials engineers work for the government and in the aviation, metal processing, electronics, automotive, and other industries in which materials play key roles.

Section 13.3 Assessment

17. Explain how hydrogen bonds affect the viscosity of a liquid. How does a change in temperature affect viscosity?
18. What effect does soap have on the surface tension of water?
19. How are a unit cell and a crystal lattice related?
20. Explain why solids are not classified as fluids.
21. What is the difference between a molecular solid and a covalent network solid?
22. Explain why most solids are denser than most liquids at the same temperature.
23. **Thinking Critically** Hypothesize why the surface of mercury in a thermometer is convex; that is, the surface is higher at the center.

Objectives

- **Explain** how the addition and removal of energy can cause a phase change.
- **Interpret** a phase diagram.

Vocabulary

melting point
 vaporization
 evaporation
 vapor pressure
 boiling point
 sublimation
 condensation
 deposition
 freezing point
 phase diagram
 triple point

Suppose you take a glass of ice water outside on a hot day. You set it down and rush inside to answer the phone. When you return much later, you find that the ice cubes have melted and there is less water in your glass. Can you use the kinetic-molecular theory to explain what happened to your drink?

Most substances can exist in three states depending on the temperature and pressure. A few substances, such as water, exist in all three states under ordinary conditions. States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture. Ice water is a heterogeneous mixture with two phases, solid ice and liquid water. When energy is added or removed from a system, one phase can change into another. As you read this section, use what you know about the kinetic-molecular theory to help explain the phase changes summarized in **Figure 13-22**.

Phase Changes That Require Energy

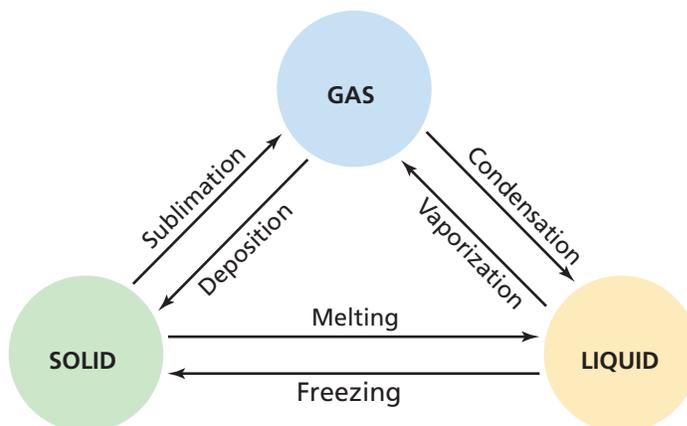
Because you are familiar with the phases of water—ice, liquid water, and water vapor—and have observed changes between those phases, we can use water as the primary example in the discussion of phase changes.

Melting What does happen to ice cubes in a glass of ice water? When ice cubes are placed in water, the water is at a higher temperature than the ice. Heat flows from the water to the ice. Heat is the transfer of energy from an object at a higher temperature to an object at a lower temperature. The energy absorbed by the ice is not used to raise the temperature of the ice. Instead, it disrupts the hydrogen bonds holding the water molecules together in the ice crystal. When molecules on the surface of the ice absorb enough energy to break the hydrogen bonds, they move apart and enter the liquid phase. As molecules are removed, the ice cube shrinks. The process continues until all of the ice melts. If a tray of ice cubes is left on a counter, where does the energy to melt the cubes come from?

The amount of energy required to melt one mole of a solid depends on the strength of the forces keeping the particles together in the solid. Because hydrogen bonds between water molecules are strong, a relatively large amount of energy is required. However, the energy required to melt ice is much less than the energy required to melt table salt because the ionic bonds in sodium chloride are much stronger than the hydrogen bonds in ice.

Figure 13-22

The six possible transitions between phases. What phase changes can occur between solids and liquids?



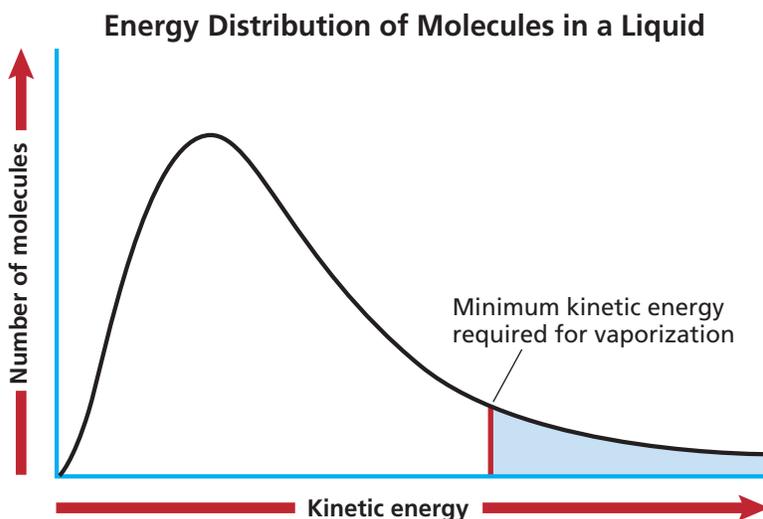


Figure 13-23

This graph shows a typical distribution of kinetic energies for a liquid at 25°C. The most probable kinetic energy lies at the peak of the curve. How would the curve look for the same liquid at 30°C?

The temperature at which the liquid phase and the solid phase of a given substance can coexist is a characteristic physical property of many solids. The **melting point** of a crystalline solid is the temperature at which the forces holding its crystal lattice together are broken and it becomes a liquid. It is difficult to specify an exact melting point for an amorphous solid because these solids tend to act like liquids when they are still in the solid state.

Vaporization While ice melts, the temperature of the ice–water mixture remains constant. Once all of the ice has melted, additional energy added to the system increases the kinetic energy of the liquid molecules. The temperature of the system begins to rise. In liquid water, some molecules will have more kinetic energy than other molecules. **Figure 13-23** shows how energy is distributed among the molecules in a liquid at 25°C. The shaded portion indicates those molecules that have the energy required to overcome the forces of attraction holding the molecules together in the liquid.

Particles that escape from the liquid enter the gas phase. For a substance that is ordinarily a liquid at room temperature, the gas phase is called a vapor. **Vaporization** is the process by which a liquid changes to a gas or vapor. If the input of energy is gradual, the molecules tend to escape from the surface of the liquid. Remember that molecules at the surface are attracted to fewer other molecules than are molecules in the interior. When vaporization occurs only at the surface of a liquid, the process is called **evaporation**. Even at cold temperatures, some water molecules have enough energy to evaporate. As the temperature rises, more and more molecules achieve the minimum energy required to escape from the liquid.

Evaporation, which requires energy, is the method by which your body controls its temperature. When you sit outside on a hot day or when you exercise, your body releases an aqueous solution called sweat from glands in your skin. Water molecules in sweat can absorb heat energy from your skin and evaporate. Excess heat is carried from all parts of your body to your skin by your blood. Evaporation of water also explains why a swim in cool ocean water is so refreshing. Not only is heat transferred from you to the cooler water while you are in the water, but water molecules left on your skin continue to cool you by evaporation once you come out of the water. The salts that remain when sea water evaporates often leave a white residue on your skin. To compare the rates of evaporation for different liquids, do the **CHEMLAB** at the end of the chapter.

Figure 13-24

Evaporation occurs in both open and closed containers.

a In an open container, water molecules that evaporate can escape from the container.

b Water vapor collects above the liquid in a closed container.

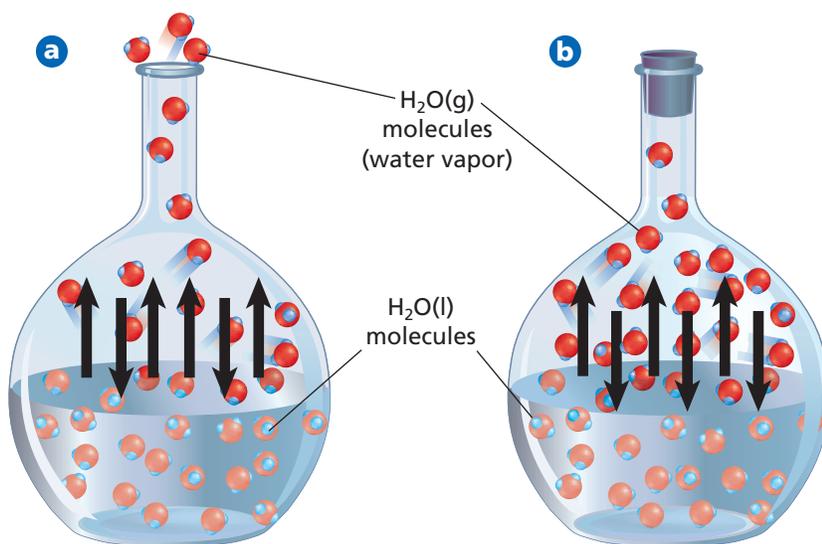


Figure 13-24 compares evaporation in an open container with evaporation in a closed container. If water is in an open container, all the molecules will eventually evaporate. The time it takes for them to evaporate depends on the amount of water and the available energy. How does temperature affect the rate of evaporation? In a partially filled, closed container, the situation is different. Water vapor collects above the liquid and exerts pressure on the surface of the liquid. The pressure exerted by a vapor over a liquid is called **vapor pressure**. How would a rise in temperature affect vapor pressure?

The temperature at which the vapor pressure of a liquid equals the external or atmospheric pressure is called the **boiling point**. Use **Figure 13-25** to compare what happens to a liquid at temperatures below its boiling point with what happens to a liquid at its boiling point. At the boiling point, molecules throughout the liquid have enough energy to vaporize. Bubbles of vapor collect below the surface of the liquid and rise to the surface. If a container has smooth walls and there are no dust particles to provide a site for bubble formation, a liquid can be heated to a temperature above its boiling point. When the liquid finally boils, the eruption of bubbles can cause the liquid to spatter. This problem can be avoided if a stone or ceramic chip is added to the liquid.

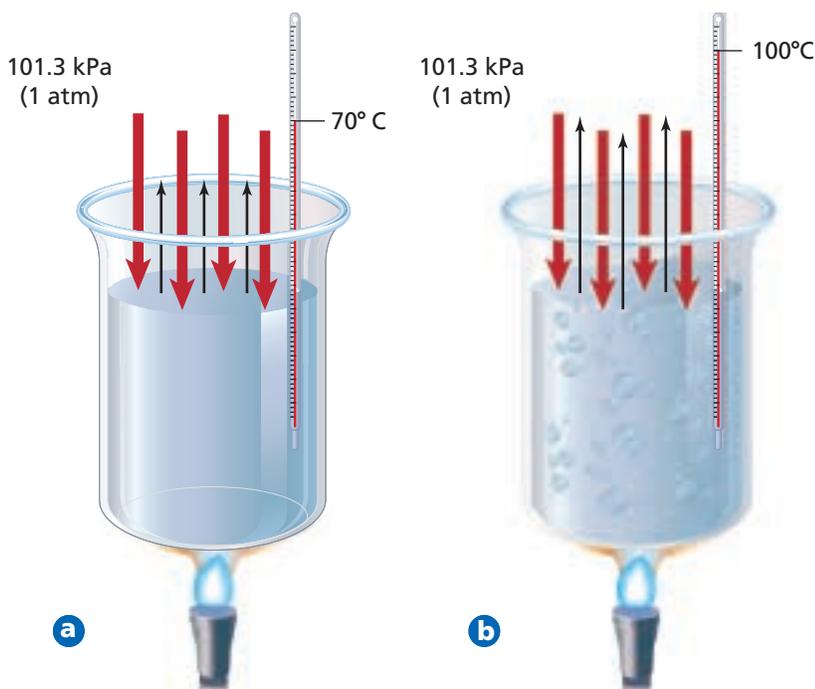


Figure 13-25

a As temperature increases, water molecules gain kinetic energy. Vapor pressure increases (black arrows) but is less than atmospheric pressure (red arrows).

b A liquid has reached its boiling point when its vapor pressure is equal to atmospheric pressure. At sea level, the boiling point of water is 100°C.



Sublimation Many substances have the ability to change directly from the solid phase to the gas phase. **Sublimation** is the process by which a solid changes directly to a gas without first becoming a liquid. Solid iodine and solid carbon dioxide (dry ice) sublime at room temperature. The special effect shown in the chapter opener is caused when dry ice sublimates and cools water vapor in the air. Dry ice keeps objects that could be damaged by melting water cold during shipping. Moth balls, which contain the compounds naphthalene or *p*-dichlorobenzene, sublime. So do solid air fresheners.

If ice cubes are left in a freezer for a long time, they shrink because the ice sublimates. At extremely low pressures, ice sublimates in a much shorter time period. This property of ice is used in a process called freeze drying. **Figure 13-26a** shows equipment used to produce freeze-dried food. Fresh food is frozen and placed in a container that is attached to a vacuum pump. As the pressure in the container is reduced, the ice sublimates and is removed from the container. The mass of the food is greatly reduced when water is removed. Backpackers often use freeze-dried foods because they want to carry as light a load as possible on a long hike. The astronauts shown in **Figure 13-26b** also have concerns about the mass of their cargo. More importantly, freeze-dried foods contain no water to support the growth of bacteria. They can be stored for a long time without refrigeration.

Phase Changes That Release Energy

Have you ever awakened on a chilly morning to see dew on the grass or frost on your windows? When you set a glass of ice water on a picnic table, do you notice beads of water on the outside of the glass? These events are examples of phase changes that release energy into the surroundings.

Condensation When a water vapor molecule loses energy, its velocity is reduced. The vapor molecule is more likely to interact and form a hydrogen bond when it collides with another water molecule. The formation of hydrogen bonds signals the change from the vapor phase to the liquid phase. Because liquid molecules are more dense than vapor molecules, the process by which a gas or a vapor becomes a liquid is called **condensation**. Condensation is the reverse of vaporization. When hydrogen bonds form in liquid water, energy is released.

There are different causes for the condensation of water vapor. All involve a transfer of energy. The vapor molecules can come in contact with a cold surface such as the outside of a glass containing ice water. Heat from the vapor

Figure 13-26

- a** Less flavor is lost when water is removed by freeze drying than when water is removed by evaporation. **b** These astronauts mix their freeze-dried food with water before eating.

Figure 13-27

As water vapor cools, it can condense on a car or it can be deposited as six-sided snow crystals.



molecules is transferred to the glass as the water vapor condenses. The water vapor that condenses on blades of grass or the car shown in **Figure 13-27** forms liquid droplets called dew. When a layer of air near the ground cools, water vapor in the air condenses and produces fog. Dew and fog evaporate when exposed to sunlight. Clouds form when layers of air high above the surface of Earth cool. Clouds are made entirely of water droplets. When the drops grow large enough, they fall to the ground as rain.

Earth Science

CONNECTION

Sometimes, too much rain falls in one location and too little rain falls in another. People have been trying to produce rain on demand for centuries. Because most clouds exist at temperatures below the freezing point of water, rain often begins when water vapor deposits on ice crystals. (When ice crystals approach the surface of Earth, they melt and fall as rain if the temperature of the air near the surface is above freezing.) Rainmakers focus on the crucial role played by ice crystals. For example, dry ice pellets can be dropped into a cloud. The cold dry ice cools the water vapor in the cloud and ice crystals form. Sometimes tiny silver iodide crystals are sprayed into a cloud to serve as artificial “ice pellets.”

Deposition Some substances can change directly into a solid without first forming a liquid. When water vapor comes in contact with a cold window in winter, it forms a solid deposit on the window called frost. **Deposition** is the process by which a substance changes from a gas or vapor to a solid without first becoming a liquid. Deposition is the reverse of sublimation. The snowflakes shown in **Figure 13-27** form when water vapor high up in the atmosphere changes directly into solid ice crystals. Energy is released as the crystals form.

Freezing Suppose you place liquid water in an ice tray in a freezer. As heat is removed from the water, the molecules lose kinetic energy and their velocity decreases. The molecules are less likely to flow past one another. When enough energy has been removed, the hydrogen bonds between water molecules keep the molecules fixed, or frozen, into set positions. Freezing is the reverse of melting. The **freezing point** is the temperature at which a liquid is converted into a crystalline solid. How do the melting point and freezing point of a given substance compare?

Phase Diagrams

There are two variables that combine to control the phase of a substance: temperature and pressure. These variables can have opposite effects on a substance. For example, a temperature increase causes more liquid to vaporize, but an increase in pressure causes more vapor to condense. A **phase diagram** is a graph of pressure versus temperature that shows in which phase a substance exists under different conditions of temperature and pressure.

Figure 13-28a shows the phase diagram for water. You can use this graph to predict what phase water will be in for any combination of temperature and pressure. Note that there are three regions representing the solid, liquid, and vapor phases of water and three curves that separate the regions from one another. At points that fall along the curves, two phases of water can coexist. At what point on the graph do liquid water and water vapor exist at

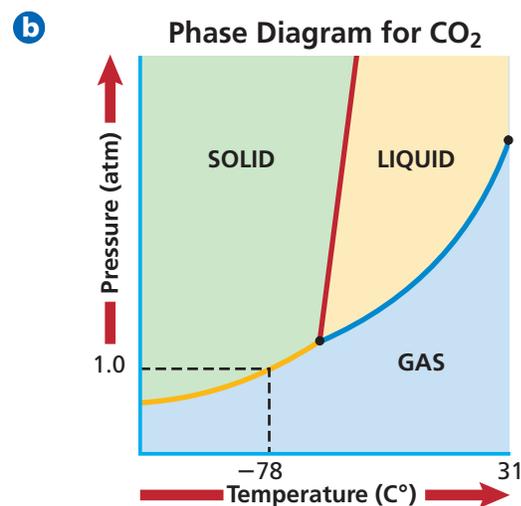
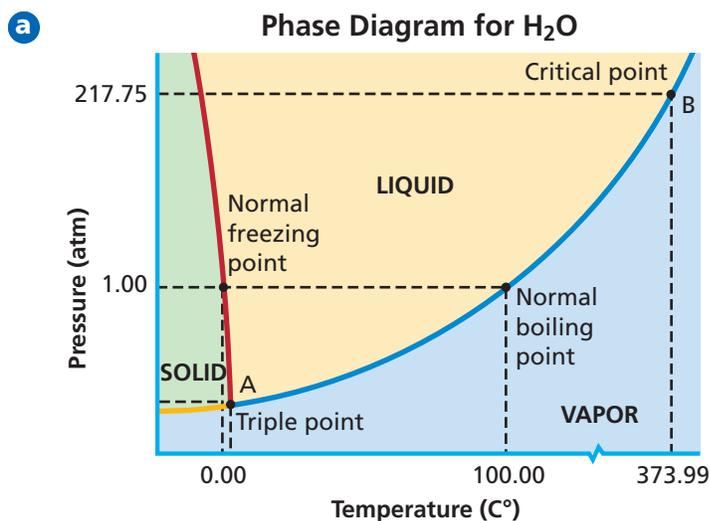


Figure 13-28

Compare the temperature and pressure values at the triple point for water and carbon dioxide.

1.00 atm and 100.00°C? The short, yellow curve shows the temperature and pressure conditions under which solid water and water vapor can coexist. The long, blue curve shows the temperature and pressure conditions under which liquid water and water vapor can coexist. The red curve shows the temperature and pressure conditions under which solid water and liquid water can coexist. What is the normal freezing point of water?

Point A on the phase diagram of water is the triple point for water. The **triple point** is the point on a phase diagram that represents the temperature and pressure at which three phases of a substance can coexist. All six phase changes can occur at the triple point: freezing and melting; evaporation and condensation; sublimation and deposition. Point B is called the critical point. This point indicates the critical pressure and critical temperature above which water cannot exist as a liquid. If water vapor is at the critical temperature, an increase in pressure will not change the vapor into a liquid.

The phase diagram for each substance is different because the normal boiling and freezing points of substances are different. However, each diagram will supply the same type of data for the phases, including a triple point. Of course, the range of temperatures chosen will vary to reflect the physical properties of the substance. Consider the phase diagram for carbon dioxide, which is shown in **Figure 13-28b**. If this diagram used the same temperature range chosen for water, there would be no solid region and an extremely small liquid region. Find the triple point for carbon dioxide. Estimate the pressure and temperature conditions at the triple point.

Section 13.4 Assessment

24. What information does a phase diagram supply?
25. What is the major difference between the processes of melting and freezing?
26. Explain what the triple point and the critical point on a phase diagram represent.
27. **Comparing and Contrasting** Compare what happens to the energy, order, and spacing of particles when a solid other than ice changes to a liquid with what happens to the energy, order, and spacing of particles when a gas changes to a liquid.
28. **Thinking Critically** Aerosol cans contain compressed gases that, when released, help propel the contents out of the can. Why is it important to keep aerosol cans from overheating?

Comparing Rates of Evaporation

Several factors determine how fast a sample of liquid will evaporate. The volume of the sample is a key factor. A drop of water takes less time to evaporate than a liter of water. The amount of energy supplied to the sample is another factor. In this lab, you will investigate how the type of liquid and temperature affect the rate of evaporation.

Problem

How do intermolecular forces affect the evaporation rates of liquids?

Objectives

- **Measure** and **compare** the rates of evaporation for different liquids.
- **Classify** liquids based on their rates of evaporation.
- **Predict** which intermolecular forces exist between the particles of each liquid.

Materials

distilled water	grease pencil or masking tape and
ethanol	a marking pen
isopropyl alcohol	paper towel
acetone	square of waxed paper
household ammonia	stopwatch
5 droppers	
5 small plastic cups	

Safety Precautions



- Always wear safety goggles and a lab apron.
- Wear gloves because some of the liquids can dry out your skin.
- Avoid inhaling any of the vapors, especially ammonia.
- There should be no open flames in the lab; some of the liquids are flammable.

Pre-Lab

1. Read the entire CHEMLAB. Prepare a data table similar to the one shown.
2. What is evaporation? Describe what happens at the molecular level during evaporation.
3. List the three possible intermolecular forces. Which force is the weakest? Which force is the strongest?
4. Look at the materials list for this lab. Consider the five liquids you will test. Predict which liquids will evaporate quickly and which will take longer to evaporate. Give reasons for your predictions.
5. To calculate an evaporation rate, you would divide the evaporation time by the quantity of liquid used. Explain why it is possible to use the evaporation times from this lab as evaporation rates.
6. Make sure you know how to use the stopwatch provided. Will you need to convert the reading on the stopwatch to seconds?

Evaporation Data

Liquid	Evaporation time (s)	Shape of liquid drop
distilled water		
ethanol		
warm ethanol		
isopropyl alcohol		
acetone		
household ammonia		

Procedure

1. Use a grease pencil or masking tape to label each of 5 small plastic cups. Use A for distilled water, B for ethanol, C for isopropyl alcohol, D for acetone, and E for household ammonia.
2. Place the plastic cups on a paper towel.

- Use a dropper to collect about 1 mL of distilled water and place the water in the cup labeled A. Place the dropper on the paper towel directly in front of the cup. Repeat with the other liquids.



- Place a square of waxed paper on your lab surface. Plan where on the waxed paper you will place each of the 5 drops that you will test. The drops must be as far apart as possible to avoid mixing.
- Have your stopwatch ready. Collect some water in your water dropper and place a single drop on the waxed paper. Begin timing. Time how long it takes for the drop to completely evaporate. While you wait, make two drawings of the drop. One drawing should show the shape of the drop as viewed from above. The other drawing should be a side view at eye level. If the drop takes longer than 5 minutes to evaporate, record > 300 in your data table.
- Repeat step 5 with the four other liquids.
- Use the above procedure to design an experiment in which you can observe the effect of temperature on the rate of evaporation of ethanol. Your teacher will provide a sample of warm ethanol. Record your observations.

Cleanup and Disposal

- Crumple up the waxed paper and place it in the container assigned by your teacher.
- Place unused liquids in the containers specified by your teacher.
- Wash out all droppers and test tubes except those used for distilled water.

Analyze and Conclude

- Classifying** Which liquids evaporated quickly? Which liquids were slow to evaporate?
- Drawing a Conclusion** Based on your data, in which liquid(s) are the attractive forces between molecules most likely to be dispersion forces?
- Interpreting Data** Make a generalization about the shape of a liquid drop and the evaporation rate of the liquid.
- Recognizing Cause and Effect** What is the relationship between surface tension and the shape of a liquid drop? What are the attractive forces that increase surface tension?
- Applying Concepts** The isopropyl alcohol you used is a mixture of isopropyl alcohol and water. Would pure isopropyl alcohol evaporate more quickly or more slowly compared to the alcohol and water mixture? Give a reason for your answer.
- Thinking Critically** Household ammonia is a mixture of ammonia and water. Based on the data you collected, is there more ammonia or more water in the mixture? Use what you learned about the relative strengths of the attractive forces in ammonia and water to support your conclusion.
- Drawing a Conclusion** How does the rate of evaporation of warm ethanol compare to ethanol at room temperature? Use kinetic-molecular theory to explain your observations.
- Error Analysis** How could you change the procedure to make it more precise?

Real-World Chemistry

- The vapor phases of liquids such as acetone and alcohol are more flammable than their liquid phases. For flammable liquids, what is the relationship between evaporation rate and the likelihood that the liquid will burn?
- Suggest why a person who has a higher than normal temperature might be given a rubdown with rubbing alcohol (70% isopropyl alcohol).
- Table salt can be collected from salt water by evaporation. The water is placed in large shallow containers. What advantage do these shallow containers have over deep containers with the same overall volume?

Everyday Chemistry

Metals with a Memory

How can the eyeglass frames shown in the photo "remember" their original shape? How do braces move your teeth? These items may be made of a shape-memory alloy that has a remarkable property. It reverts to its previous shape when heated or when the stress that caused its shape is removed.

Two solid phases?

Melting is the transition of a material from a solid to a liquid. Transitions from one phase to another also can take place within a solid. A solid can have two phases if it has two possible crystal structures. It is the ability to undergo these changes in crystalline structure that gives shape-memory alloys their properties.

For example, an alloy that contains equal amounts of two metals may have phases called austenite and martensite. In the austenite phase, the metal ions are arranged in a rigid crystalline cubic structure. In the martensite phase, the metal ions are arranged in a more flexible structure. When the internal organization changes, the properties of the alloy change.

The austenite phase can be changed into the martensite phase if the alloy is cooled below the transition temperature under controlled conditions. Copper-aluminum-nickel, nickel-titanium, and copper-zinc-aluminum are the most common shape-memory alloys.

Nitinol

Nitinol is an alloy of nickel and titanium that has the austenite phase structure. Each nickel atom is at the center of a cube of titanium, and a titanium atom is at the center of each cube of nickel atoms. If nitinol is shaped into a straight wire, heated, and cooled past the transition temperature, it converts into the martensite phase. Now its structure allows the nitinol to be bent by an external stress. If the wire is heated, the stress is released and it reverts back to its initial shape.

Applications of nitinol

Nitinol is a corrosion resistant, light-weight material that generates a large force when it returns to its original shape. It also is safe to use inside the body, that is, nitinol is biocompatible. You may need fewer trips to the orthodontist because the nitinol wires in your braces need to be tightened and adjusted less frequently. Staples to repair human bones, devices that trap blood clots, and anchors used to reattach tendons to bone in an injured shoulder may be made of nitinol. Nitinol is used in a variety of military, safety, and robotics applications. Life-saving devices such as antiscalding valves that automatically shut off water flow and fire sprinklers that respond more quickly to heat rely on shape-memory alloys. So do vibration-control devices in buildings and bridges.



Testing Your Knowledge

- 1. Inferring** Twisted nitinol-wire eyeglass frames unbend at room temperature. Is the transition temperature of the frames above or below room temperature?
- 2. Applying** Design a simple lever that could be raised and lowered smoothly using nitinol wires.

Summary

13.1 Gases

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Because of the space between gas particles, the density of gases is low and gases can be compressed.
- Because there are no significant forces of attraction between gas particles, gases can diffuse and effuse at rates determined by the mass of the particles.
- A barometer measures the pressure gas particles in Earth's atmosphere exert against Earth's surface.
- The total pressure of a gas mixture is the sum of the partial pressures of each gas in the mixture.

13.2 Forces of Attraction

- The intramolecular forces that hold together ionic, covalent, and metallic bonds are stronger than intermolecular forces.
- Dispersion forces are weak intermolecular forces between temporary dipoles of nonpolar molecules.
- Dipole–dipole forces occur between polar molecules. A hydrogen bond is a strong dipole–dipole force between molecules in which hydrogen atoms are bonded to highly electronegative atoms.

13.3 Liquids and Solids

- Particles are in motion in liquids and solids, but the range of motion is limited by intermolecular forces. Because liquids and solids are denser than gases, they are not easily compressed.
- In general, viscosity increases as the temperature decreases and as intermolecular forces increase.
- Surface tension results from an uneven distribution of attractive forces. A liquid displays capillarity when adhesive forces are stronger than cohesive forces.
- Except for amorphous solids, solids are more ordered than liquids. Crystalline solids can be classified by shape and composition.

13.4 Phase Changes

- Melting, vaporization, and sublimation are changes that require energy. Freezing, condensation, and deposition are changes that release energy. The temperature of a system remains constant during a phase change.
- Evaporation happens at the surface of a liquid. Boiling occurs when a liquid's vapor pressure is equal to atmospheric pressure.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

Key Equations and Relationships

Kinetic energy: $KE = 1/2mv^2$
(p. 386)

Graham's law of effusion:
(p. 387)

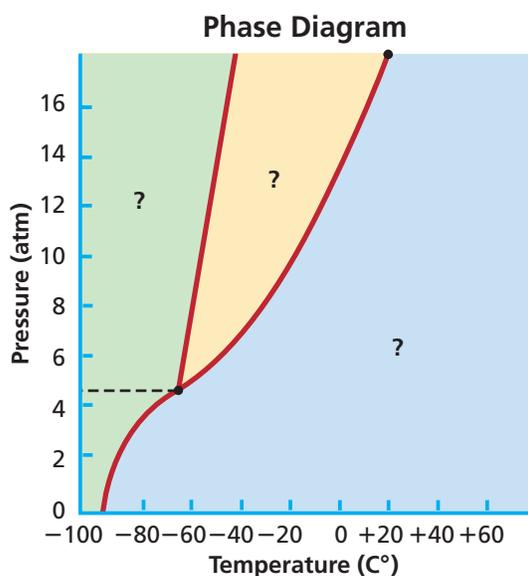
$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

Dalton's law of partial pressures: $P_{\text{total}} = P_1 + P_2 + P_3 + \dots P_n$
(p. 391)

Vocabulary

- amorphous solid (p. 403)
- atmosphere (p. 390)
- barometer (p. 389)
- boiling point (p. 406)
- condensation (p. 407)
- crystalline solid (p. 400)
- Dalton's law of partial pressures (p. 391)
- deposition (p. 408)
- diffusion (p. 387)
- dipole–dipole forces (p. 394)
- dispersion forces (p. 393)
- elastic collision (p. 386)
- evaporation (p. 405)
- freezing point (p. 408)
- Graham's law of effusion (p. 387)
- hydrogen bond (p. 395)
- kinetic-molecular theory (p. 385)
- melting point (p. 405)
- pascal (p. 390)
- phase diagram (p. 408)
- pressure (p. 388)
- sublimation (p. 407)
- surface tension (p. 398)
- surfactant (p. 398)
- temperature (p. 386)
- triple point (p. 409)
- unit cell (p. 400)
- vaporization (p. 405)
- vapor pressure (p. 406)
- viscosity (p. 397)

52. Explain why water has a higher surface tension than benzene, whose molecules are nonpolar. (13.3)
53. How does sublimation differ from deposition? (13.4)
54. Compare boiling and evaporation. (13.4)
55. Define melting point. (13.4)
56. Explain the relationships among vapor pressure, atmospheric pressure, and boiling point. (13.4)
57. Explain why dew forms on cool mornings. (13.4)
58. Label the solid, liquid, and gas phases, triple point, and critical point on the phase diagram shown. (13.4)



59. Why does it take more energy to boil 10 g of liquid water than to melt an equivalent mass of ice? (13.4)
60. Why does a pile of snow slowly shrink even on days when the temperature never rises above the freezing point of water? (13.4)
61. Examine the phase diagram for carbon dioxide in **Figure 13-28**. Notice that at 1 atm pressure, the solid sublimates to a gas. What happens to the solid at much higher pressures? (13.4)

Mastering Problems

Graham's Law of Effusion (13.1)

62. What is the molar mass of a gas that takes three times longer to effuse than helium?
63. What is the ratio of effusion rates of krypton and neon at the same temperature and pressure?
64. Calculate the molar mass of a gas that diffuses three times faster than oxygen under similar conditions.

Dalton's Law of Partial Pressures (13.1)

65. What is the partial pressure of water vapor in an air sample when the total pressure is 1.00 atm, the partial pressure of nitrogen is 0.79 atm, the partial pressure of oxygen is 0.20 atm, and the partial pressure of all other gases in air is 0.0044 atm?
66. What is the total gas pressure in a sealed flask that contains oxygen at a partial pressure of 0.41 atm and water vapor at a partial pressure of 0.58 atm?
67. Find the partial pressure of oxygen in a sealed vessel that has a total pressure of 2.6 atm and also contains carbon dioxide at 1.3 atm and helium at 0.22 atm.

Converting Pressure Units (13.1)

68. What is the total pressure in atmospheres of a mixture of three gases with partial pressures of 12.0 kPa, 35.6 kPa, and 22.2 kPa?
69. The pressure atop the world's highest mountain, Mount Everest, is usually about 33.6 kPa. Convert the pressure to atmospheres. How does the pressure compare with the pressure at sea level?
70. The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?
71. At an ocean depth of 250 feet, the pressure is about 8.4 atm. Convert the pressure to mm Hg and kPa units.

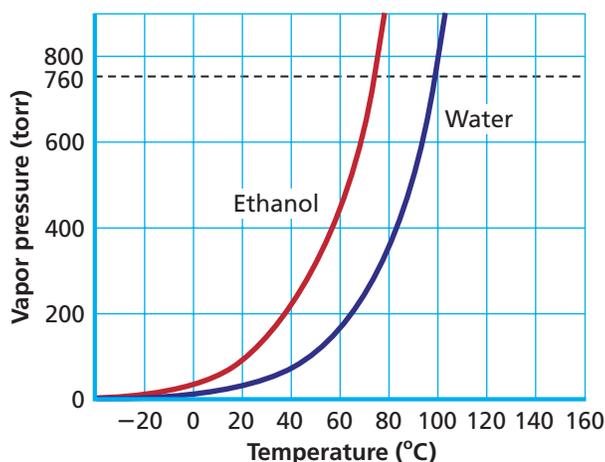
Mixed Review

Sharpen your problem solving skills by answering the following.

72. Use the kinetic-molecular theory to explain why both gases and liquids are fluids.
73. Use intermolecular forces to explain why oxygen is a gas at room temperature and water is a liquid.
74. Use the kinetic-molecular theory to explain why gases are easier to compress than liquids or solids.
75. The density of mercury at 25°C and a pressure of 760 mm Hg is 13.5 g/mL; water at the same temperature and pressure has a density of 1.00 g/mL. Explain this difference in terms of intermolecular forces and the kinetic-molecular theory.
76. Two flasks of equal size are connected by a narrow tube that is closed in the middle with a stopcock. One flask has no gas particles; the other flask contains 0.1 mol of hydrogen gas at a pressure of 2.0 atm.
 - a. Describe what happens to the gas molecules after the stopcock is opened.
 - b. What will happen to the gas pressure after the stopcock is opened?

Thinking Critically

- 77. Interpreting Graphs** Examine the graph below, which plots vapor pressure versus temperature for water and ethyl alcohol.
- What is the boiling point of water at 1 atm?
 - What is the boiling point of ethyl alcohol at 1 atm?
 - Describe the relationship between temperature and vapor pressure for water and alcohol.
 - Estimate the temperature at which water will boil when the atmospheric pressure is 0.80 atm.



- 78. Applying Concepts** A solid being heated stays at a constant temperature until it is completely melted. What happens to the heat energy put into the system during that time?
- 79. Comparing and Contrasting** An air compressor uses energy to squeeze air particles together. When the air is released, it expands, allowing the energy to be used for purposes such as gently cleaning surfaces without using a more abrasive liquid or solid. Hydraulic systems essentially work the same way, but involve compression of liquid water rather than air. What do you think are some advantages and disadvantages of these two types of technology?
- 80. Hypothesizing** What type of crystalline solid do you predict would best suit the following needs?
- a material that can be melted and reformed at a low temperature
 - a material that can be drawn into long, thin wires
 - a material that conducts electricity when molten
 - an extremely hard material that is non-conductive
- 81. Communicating** Which process is responsible for your being able to smell perfume from an open bottle that is located across the room from you, effusion or diffusion? Explain.

- 82. Inferring** A laboratory demonstration involves pouring bromine vapors, which are a deep red color, into a flask of air and then tightly sealing the top of the flask. The bromine is observed to first sink to the bottom of the beaker. After several hours have passed, the red color is distributed equally throughout the flask.
- Is bromine gas more or less dense than air?
 - Would liquid bromine diffuse more or less quickly than gaseous bromine after you pour it into another liquid?

Writing in Chemistry

- 83.** Propane gas is a commonly used heating fuel for gas grills and homes. It is not packaged as a gas, however. It is liquefied and referred to as liquid propane or “LP gas.” What advantages are there to storing and transporting propane as a liquid rather than a gas? Are there any disadvantages?
- 84.** Find out what your birthstone is if you don’t already know, and write a brief report about the chemistry of that gem. Find out its chemical composition, which category its unit cell is in, how hard and durable it is, and its approximate cost at the present time.
- 85.** What would happen to life on Earth if ice were denser than liquid water? Would life be possible? Write an essay on this topic.

Cumulative Review

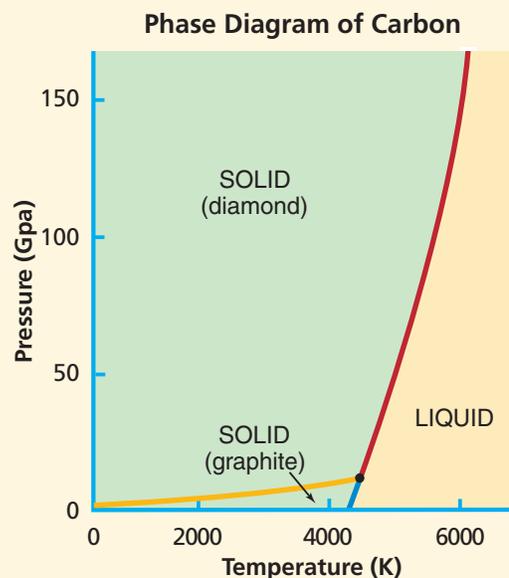
Refresh your understanding of the previous chapters by answering the following.

- 86.** Identify the following as an element, compound, homogeneous mixture, or heterogeneous mixture. (Chapter 3)
- | | |
|-------------|------------|
| a. air | e. ammonia |
| b. blood | f. mustard |
| c. antimony | g. water |
| d. brass | h. tin |
- 87.** Use the periodic table to separate these ten elements into five pairs of elements having similar properties. (Chapter 7)
- S, Ne, Li, O, Mg, Ag, Sr, Kr, Cu, Na
- 88.** You are given two clear, colorless aqueous solutions. You are told that one solution contains an ionic compound and one contains a covalent compound. How could you determine which is an ionic solution and which is a covalent solution? (Chapter 9)
- 89.** Determine the number of atoms in 56.1 g Al. (Chapter 11)

Use these questions and the test-taking tip to prepare for your standardized test.

- Water has an extremely high boiling point compared to other compounds of similar molar mass because of _____.
 - hydrogen bonding
 - adhesive forces
 - covalent bonding
 - dispersion forces
- The ratio of effusion rates for nitric oxide (NO) and nitrogen tetroxide (N₂O₄) is _____.
 - 0.326
 - 0.571
 - 1.751
 - 3.066
- Which of the following is NOT an assumption of the kinetic-molecular theory?
 - Collisions between gas particles are elastic.
 - All the gas particles in a sample have the same velocity.
 - A gas particle is not significantly attracted or repelled by other gas particles.
 - All gases at a given temperature have the same average kinetic energy.
- Which of the following statements does NOT describe what happens as a liquid boils?
 - The temperature of the system rises.
 - Energy is absorbed by the system.
 - The vapor pressure of the liquid is equal to atmospheric pressure.
 - The liquid is entering the gas phase.
- The solid phase of a compound has a definite shape and volume because its particles _____.
 - are not in constant motion.
 - are always packed more tightly than particles in the compound's liquid phase.
 - can only vibrate around fixed points.
 - are held together by strong intramolecular forces.
- A sealed flask contains neon, argon, and krypton gas. If the total pressure in the flask is 3.782 atm, the partial pressure of Ne is 0.435 atm, and the partial pressure of Kr is 1.613 atm, what is the partial pressure of Ar?
 - 2.048 torr
 - 1.734 torr
 - 1556 torr
 - 1318 torr
- Which of the following does not affect the viscosity of a liquid?
 - intermolecular attractive forces
 - size and shape of molecules
 - temperature of the liquid
 - capillary action

Interpreting Graphs Use the graph to answer the following questions.



- Diamond is most likely to form at _____.
 - temperatures > 5000 K and pressures < 100 GPa.
 - temperatures > 6000 K and pressures > 25 GPa.
 - temperatures < 4000 K and pressures > 25 GPa.
 - temperatures < 4500 K and pressures < 10 GPa.
- Find the point on the graph at which carbon exists in three phases: solid graphite, solid diamond, and liquid carbon. The temperature and pressure at that point are _____.
 - 4700 K and 15 GPa.
 - 3000 K and 10 GPa.
 - 5100 K and 50 GPa.
 - 14500 K and 5 GPa.
- In what form or forms does carbon exist at 6000 K and 75 GPa?
 - diamond only
 - liquid carbon only
 - diamond and liquid carbon
 - liquid carbon and graphite



TEST-TAKING TIP

Focus When you take a test, pay no attention to anyone other than the proctor. If students near you are talking, move to a different seat. If someone other than the proctor talks to you during a test, don't respond. Not only is talking a distraction, but the proctor may think that you are cheating. Don't take the chance. Focus on the test, and nothing else.

Gases

What You'll Learn

- ▶ You will use gas laws to calculate how pressure, temperature, volume, and number of moles of a gas will change when one or more of these variables is altered.
- ▶ You will compare properties of real and ideal gases.
- ▶ You will apply the gas laws and Avogadro's principle to chemical equations.

Why It's Important

From barbecuing on a gas grill to taking a ride in a hot-air balloon, many activities involve gases. It is important to be able to predict what effect changes in pressure, temperature, volume, or amount, will have on the properties and behavior of a gas.

CLICK HERE



Visit the Glencoe Chemistry Web site at science.glencoe.com to find links about the behavior of gases.

Firefighters breathe air that has been compressed into tanks that they can wear on their backs.



DISCOVERY LAB



Materials

5-gal bucket
round balloon
ice
string

More Than Just Hot Air

How does a temperature change affect the air in a balloon?

Safety Precautions



Always wear goggles to protect eyes from broken balloons.

Procedure

1. Inflate a round balloon and tie it closed.
2. Fill the bucket about half full of cold water and add ice.
3. Use a string to measure the circumference of the balloon.
4. Stir the water in the bucket to equalize the temperature. Submerge the balloon in the ice water for 15 minutes.
5. Remove the balloon from the water. Measure the circumference.

Analysis

What happens to the size of the balloon when its temperature is lowered? What might you expect to happen to its size if the temperature is raised?

Section

14.1

The Gas Laws

Objectives

- **State** Boyle's law, Charles's law, and Gay-Lussac's law.
- **Apply** the three gas laws to problems involving the pressure, temperature, and volume of a gas.

Vocabulary

Boyle's law
Charles's law
Gay-Lussac's law

The manufacturer of the air tank in the photo on the opposite page had to understand the nature of the gases the tank contains. Understanding gases did not happen accidentally. The work of many scientists over many years has contributed to our present knowledge of the nature of gases. The work of three scientists in particular was valuable enough that laws describing gas behavior were named in their honor. In this section, you'll study three important gas laws: Boyle's law, Charles's law, and Gay-Lussac's law. Each of these laws relates two of the variables that determine the behavior of gases—pressure, temperature, volume, and amount of gas present.

Kinetic Theory

You can't understand gases without understanding the movement of gas particles. Remember from your study of the kinetic-molecular theory in Chapter 13 that gas particles behave differently than those of liquids and solids. The kinetic theory provides a model that is used to explain the properties of solids, liquids, and gases in terms of particles that are always in motion and the forces that exist between them. The kinetic theory assumes the following concepts about gases are true.

- *Gas particles do not attract or repel each other.* Gases are free to move within their containers without interference from other particles.
- *Gas particles are much smaller than the distances between them.* You saw in the **DISCOVERY LAB** that gas has volume. However, the kinetic theory assumes that gas particles themselves have virtually no volume.

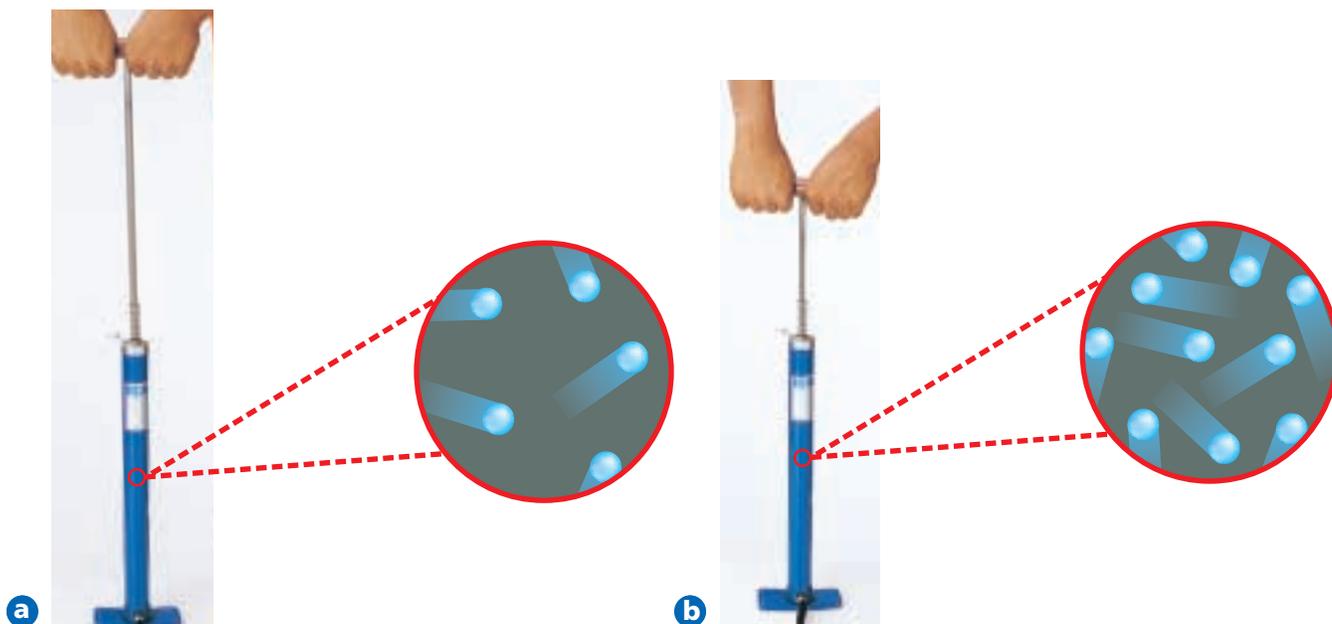


Figure 14-1

The kinetic theory relates pressure and the number of collisions per unit time for a gas.

a When the bicycle pump is pulled out as far as it will go, the pressure of the air inside the pump equals that of the atmosphere.

b If the piston is pushed down half the length of the pump, the air particles are squeezed into a space half the original size. Pressure doubles because the frequency of collisions between the gas particles and the inner wall of the pump has doubled.

Almost all the volume of a gas is empty space. Gases can be compressed by moving gas particles closer together because of this low density of particles.

- *Gas particles are in constant, random motion.* Gas particles spread out and mix with each other because of this motion. The particles move in straight lines until they collide with each other or with the walls of their container.
- *No kinetic energy is lost when gas particles collide with each other or with the walls of their container.* Such collisions are completely elastic. As long as the temperature stays the same, the total kinetic energy of the system remains constant.
- *All gases have the same average kinetic energy at a given temperature.* As temperature increases, the total energy of the gas system increases. As temperature decreases, the total energy of the gas system decreases.

The nature of gases Actual gases don't obey all the assumptions made by the kinetic theory. But for many gases, their behavior approximates the behavior assumed by the kinetic theory. You will learn more about real gases and how they vary from these assumptions in Section 14.3.

Notice how all the assumptions of the kinetic theory are based on the four factors previously mentioned—the number of gas particles present and the temperature, the pressure, and the volume of the gas sample. These four variables all work together to determine the behavior of gases. When one variable changes, it affects the other three. Look at the following example of how a change in one variable affects at least one other variable.

What happens to the gas in a plastic balloon if you squeeze it, decreasing its volume? Because the balloon is closed, the amount of gas is constant. Assume the temperature is held constant. Decreasing the volume pushes the gas particles closer together. Recall from the kinetic-molecular theory that as gas particles are pushed closer together, the number of collisions between particles themselves and between the particles and the walls of their container increases. As the number of collisions per unit time increases, so does the observed pressure. Therefore, as the volume of a gas decreases, its pressure increases. Similarly, if the balloon is no longer squeezed, the volume increases and the pressure decreases. You can see another example of this principle in **Figure 14-1**. The interdependence of the variables of volume, pressure, temperature, and amount of gas is the basis for the following gas laws.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Boyle's Law

Robert Boyle (1627–1691), an Irish chemist, did experiments like the one shown in **Figure 14-2** to study the relationship between the pressure and the volume of a gas. By taking careful quantitative measurements, he showed that if the temperature is constant, doubling the pressure of a fixed amount of gas decreases its volume by one-half. On the other hand, reducing the pressure by half results in a doubling of the volume. A relationship in which one variable increases as the other variable decreases is referred to as an inversely proportional relationship. For help with understanding inverse relationships, see the **Math Handbook** page 905.

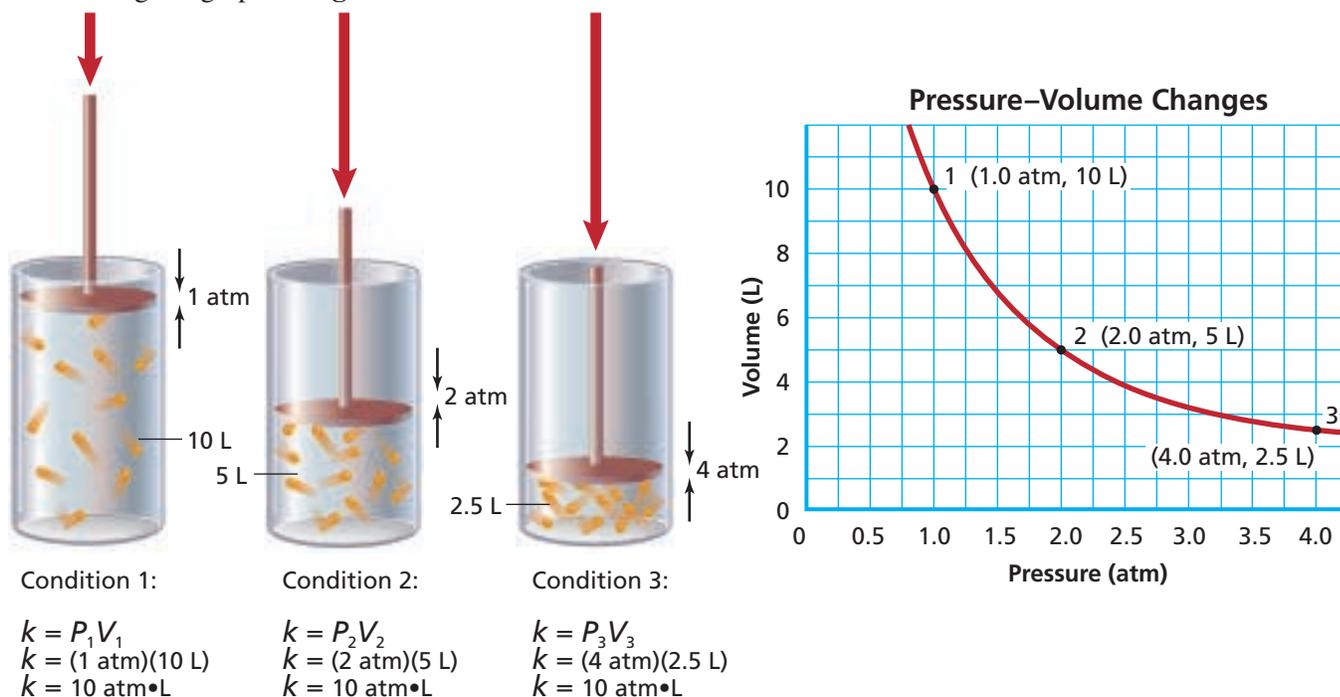
Boyle's law states that the volume of a given amount of gas held at a constant temperature varies inversely with the pressure. Look at the graph in **Figure 14-2** in which pressure versus volume is plotted for a gas. The plot of an inversely proportional relationship results in a downward curve. If you choose any two points along the curve and multiply the pressure times the volume at each point, how do your two answers compare? Note that the product of the pressure and the volume for each of points 1, 2, and 3 is 10 atm·L. From the graph, what would the volume be if the pressure is 2.5 atm? What would the pressure be if the volume is 2 L?

The products of pressure times volume for any two sets of conditions are equal, so Boyle's law can be expressed mathematically as follows.

$$P_1V_1 = P_2V_2$$

P_1 and V_1 represent a set of initial conditions for a gas and P_2 and V_2 represent a set of new conditions. If you know any three of these four values for a gas at constant temperature, you can solve for the fourth by rearranging the equation. For example, if P_1 , V_1 , and P_2 are known, dividing both sides of the equation by P_2 will isolate the unknown variable V_2 .

Use the equation for Boyle's law to calculate the volume that corresponds to a pressure of 2.5 atm, assuming that the amount of gas and temperature are constant. Then find what pressure corresponds to a volume of 2.0 L. Use 2.0 atm for P_1 and 5 L for V_1 . How do these answers compare to those you found using the graph in **Figure 14-2**?



Careers Using Chemistry

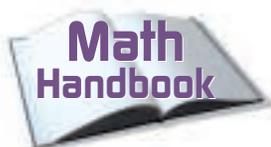
Meteorologist

Would you like to be able to plan your days better because you know what the weather will be? Then consider a career as a meteorologist.

Most weather is caused by the interaction of energy and air. Meteorologists study how these interactions affect and are affected by changes in temperature and pressure of the air. For example, winds and fronts are direct results of pressure changes caused by uneven heating of Earth's atmosphere by the sun.

Figure 14-2

The gas particles in this cylinder take up a given volume at a given pressure. As pressure increases, volume decreases. The graph shows that pressure and volume have an inverse relationship, which means that as pressure increases, volume decreases. This relationship is illustrated by a downward curve in the line from condition 1 to condition 2 to condition 3.



Review using inverse relationships in the **Math Handbook** on page 905 of your textbook.

EXAMPLE PROBLEM 14-1

Boyle's Law

A sample of helium gas in a balloon is compressed from 4.0 L to 2.5 L at a constant temperature. If the pressure of the gas in the 4.0-L volume is 210 kPa, what will the pressure be at 2.5 L?

1. Analyze the Problem

You are given the initial and final volumes and the initial pressure of a sample of helium. Boyle's law states that as volume decreases, pressure increases if temperature remains constant. Because the volume in this problem is decreasing, the pressure will increase. So the initial pressure should be multiplied by a volume ratio greater than one.

Known

$$V_1 = 4.0 \text{ L}$$

$$V_2 = 2.5 \text{ L}$$

$$P_1 = 210 \text{ kPa}$$

Unknown

$$P_2 = ? \text{ kPa}$$

2. Solve for the Unknown

Divide both sides of the equation for Boyle's law by V_2 to solve for P_2 .

$$P_1 V_1 = P_2 V_2$$

$$P_2 = P_1 \left(\frac{V_1}{V_2} \right)$$

Substitute the known values into the rearranged equation.

$$P_2 = 210 \text{ kPa} \left(\frac{4.0 \text{ L}}{2.5 \text{ L}} \right)$$

Multiply and divide numbers and units to solve for P_2 .

$$P_2 = 210 \text{ kPa} \left(\frac{4.0 \cancel{\text{L}}}{2.5 \cancel{\text{L}}} \right) = 340 \text{ kPa}$$

3. Evaluate the Answer

When the volume is decreased by almost half, the pressure is expected to almost double. The calculated value of 340 kPa is reasonable. The unit in the answer is kPa, a pressure unit.

PRACTICE PROBLEMS

Assume that the temperature and the amount of gas present are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?
3. Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume of air when the pressure is increased to 1.43 atm by applying force to the piston?
4. If it takes 0.0500 L of oxygen gas kept in a cylinder under pressure to fill an evacuated 4.00-L reaction vessel in which the pressure is 0.980 atm, what was the initial pressure of the gas in the cylinder?
5. A sample of neon gas occupies 0.220 L at 0.860 atm. What will be its volume at 29.2 kPa pressure?



For more practice with Boyle's law problems, go to **Supplemental Practice Problems** in Appendix A.



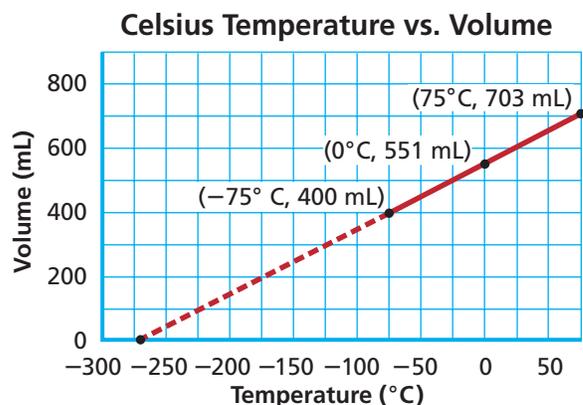
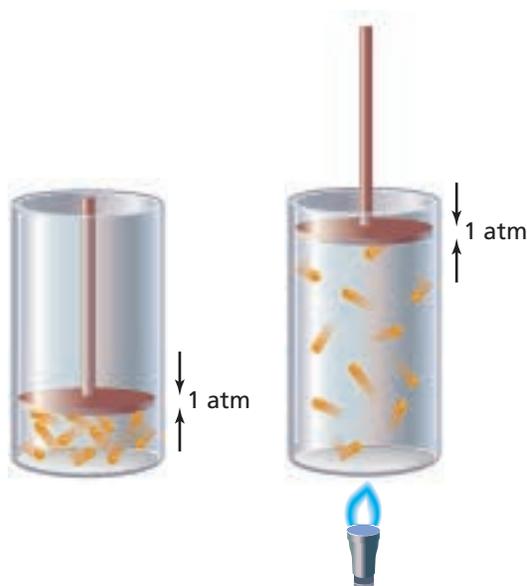
Charles's Law

Have you ever noticed that on a cold day, a tire on a car might look as if it's low on air? However, after driving the car for awhile, the tire warms up and looks less flat. What made the difference in the tire? When canning vegetables at home, why are they often packed hot in the jars and then sealed? These questions can be answered by applying another of the gas laws, Charles's law. Another example of how gases are affected by temperature is shown in the **problem-solving LAB**. If kelvin temperature is doubled, so is the volume.

How are gas temperature and volume related? The French physicist Jacques Charles (1746–1823) studied the relationship between volume and temperature. In his experiments, he observed that as temperature increases, so does the volume of a gas sample when the pressure is held constant. This property can be explained by the kinetic-molecular theory; at a higher temperature, gas particles move faster, striking each other and the walls of their container more frequently and with greater force. For the pressure to stay constant, volume must increase so that the particles have farther to travel before striking the walls. Having to travel farther decreases the frequency with which the particles strike the walls of the container.

Look at **Figure 14-3**, which includes a graph of volume versus temperature for a gas sample kept at a constant pressure. Note that the resulting plot is a straight line. Note also that you can predict the temperature at which the volume will reach a value of zero liters by extrapolating the line at temperatures below which values were actually measured. The temperature that corresponds to zero volume is -273.15°C , or 0 on the kelvin (K) temperature scale. This temperature is referred to as absolute zero, and it is the lowest possible theoretical temperature. Theoretically, at absolute zero, the kinetic energy of particles is zero, so all motion of gas particles at that point ceases.

Examine the relationship between temperature and volume shown by the cylinders in **Figure 14-3**. In the graph, note that 0°C does not correspond to zero volume. Although the relationship is linear, it is not direct. For example, you can see from the graph that increasing the temperature from 25°C to 50°C does not double the volume of the gas. If the kelvin temperature is plotted instead, a direct proportion is the result. See **Figure 14-4** on the next page.



History

CONNECTION

Jacques Charles's interest in the behavior of gases was sparked by his involvement in ballooning. He built the first balloon that was filled with hydrogen instead of hot air. Charles's investigations attracted the attention of King Louis XVI, who allowed Charles to establish a laboratory in the Louvre, which is a museum in Paris.

Figure 14-3

The gas particles in this cylinder take up a given volume at a given temperature. When the cylinder is heated, the kinetic energy of the particles increases. The volume of the gas increases, pushing the piston outward. Thus the distance that the piston moves is a measure of the increase in volume of the gas as it is heated. Note that the graph of volume versus temperature extrapolates to -273.15°C , or 0 K.

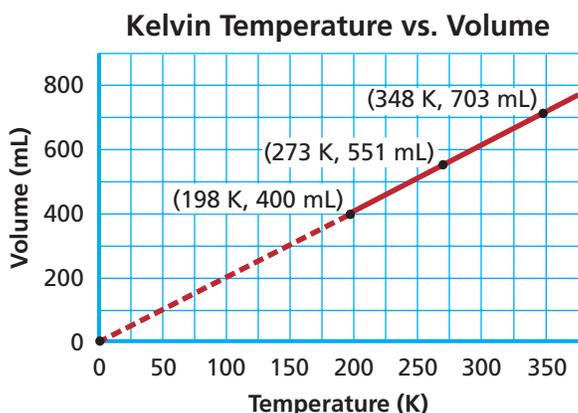


Figure 14-4

This graph illustrates the directly proportional relationship between the volume and the kelvin temperature of a gas held at constant pressure. When the kelvin temperature doubles, the volume doubles.

Charles's law states that the volume of a given mass of gas is directly proportional to its kelvin temperature at constant pressure. For help with understanding direct relationships, see the **Math Handbook**, page 905. So for any two sets of conditions, Charles's law can be expressed as

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Here V_1 and T_1 represent any initial pair of conditions, while V_2 and T_2 are any new set of conditions. As with Boyle's law, if you know any three of the four values, you can calculate the fourth using the equation.

The temperature must be expressed in kelvin units when using the equation for Charles's law. The kelvin scale starts at absolute zero, which corresponds to -273.15°C and is 0 K. Because a Celsius degree and a kelvin unit are the same size, it is easy to convert a temperature in Celsius to kelvin units. Round 273.15 to 273, and add it to the Celsius temperature.

$$T_K = 273 + T_C$$

problem-solving LAB

How is turbocharging in a car engine maximized?

Interpreting Scientific Illustrations After gasoline and air are burned in the combustion chamber of an automobile, the resulting hot gases are exhausted out the tailpipe. The horsepower of an automobile engine can be significantly improved if the energy of these exhaust gases is used to operate a compressor that forces additional air into the combustion chamber. Outside air is then blown over this compressed air to cool it before it enters the engine. Increasing the power of an engine in this manner is known as turbocharging.

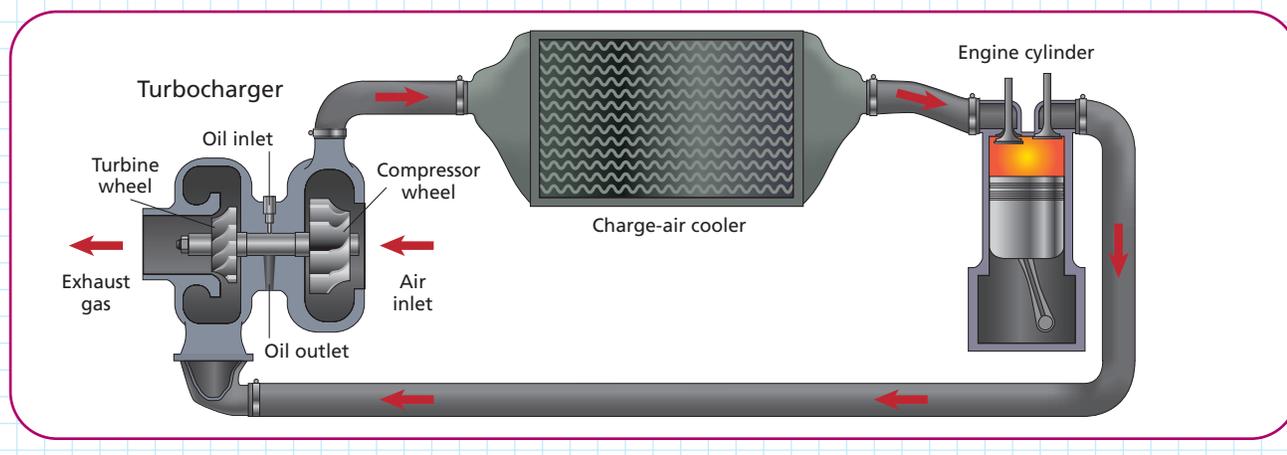
Analysis

Examine the illustration of an engine fitted with

a turbocharging system. The paths of the exhaust gas, entering combustion air, and the cooling air are shown.

Thinking Critically

1. What property of the exhaust gas is being used to turn the turbine that runs the compressor? Explain.
2. If more power is to be gained from this design, what must also accompany the extra supply of oxygen to the combustion chamber?
3. What property does the compressor alter so that more air can be injected into the combustion chamber? Explain.
4. Why does the air in the compressor get hot, and why does cooling help to improve the power of the engine?



EXAMPLE PROBLEM 14-2

Charles's Law

A gas sample at 40.0°C occupies a volume of 2.32 L. If the temperature is raised to 75.0°C, what will the volume be, assuming the pressure remains constant?

1. Analyze the Problem

You are given the initial temperature and volume of a sample of gas. Charles's law states that as the temperature increases, so does the volume, assuming the pressure is constant. Because the temperature in this problem is increasing, the volume will increase. So the initial volume should be multiplied by a volume ratio greater than one.

Known

$$T_1 = 40.0^\circ\text{C}$$

$$V_1 = 2.32 \text{ L}$$

$$T_2 = 75.0^\circ\text{C}$$

Unknown

$$V_2 = ? \text{ L}$$

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$$T_K = 273 + T_C$$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$$T_1 = 273 + 40.0^\circ\text{C} = 313 \text{ K}$$

$$T_2 = 273 + 75.0^\circ\text{C} = 348 \text{ K}$$

Multiply both sides of the equation for Charles's law by T_2 to solve for V_2 .

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = V_1 \left(\frac{T_2}{T_1} \right)$$

Substitute the known values into the rearranged equation.

$$V_2 = 2.32 \text{ L} \left(\frac{348 \text{ K}}{313 \text{ K}} \right)$$

Multiply and divide numbers and units to solve for V_2 .

$$V_2 = 2.32 \text{ L} \left(\frac{348 \cancel{\text{K}}}{313 \cancel{\text{K}}} \right) = 2.58 \text{ L}$$

3. Evaluate the Answer

The increase in kelvin units is relatively small. Therefore, you expect the volume to show a small increase, which agrees with the answer. The unit of the answer is liters, a volume unit.



Review using direct relationships in the **Math Handbook** on page 905 of this textbook.

PRACTICE PROBLEMS

Assume that the pressure and the amount of gas present remain constant in the following problems.

6. A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?
7. The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?
8. What is the volume of the air in a balloon that occupies 0.620 L at 25°C if the temperature is lowered to 0.00°C?



For more practice with Charles's law problems, go to **Supplemental Practice Problems** in Appendix A.

Kelvin Temperature vs. Pressure

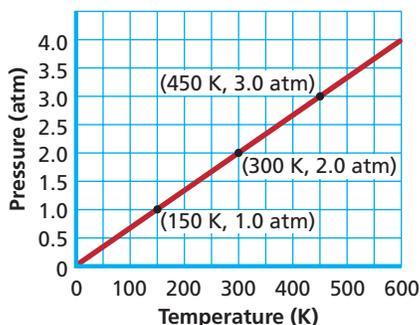


Figure 14-5

Compare the relationship between pressure and kelvin temperature as shown in this graph and the relationship between volume and kelvin temperature as shown in **Figure 14-4**. Notice that both show direct relationships.

Gay-Lussac's Law

Boyle's law relates pressure and volume of a gas, and Charles's law states the relationship between a gas's temperature and volume. What is the relationship between pressure and temperature? Pressure is a result of collisions between gas particles and the walls of their container. An increase in temperature increases collision frequency and energy, so raising the temperature should also raise the pressure if the volume is not changed.

How are temperature and pressure of a gas related? Joseph Gay-Lussac explored the relationship between temperature and pressure of a contained gas at a fixed volume. He found that a direct proportion exists between the kelvin temperature and the pressure, such as that illustrated in **Figure 14-5**. **Gay-Lussac's law** states that the pressure of a given mass of gas varies directly with the kelvin temperature when the volume remains constant. It can be expressed mathematically.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

As with Boyle's and Charles's laws, if you know any three of the four variables, you can calculate the fourth using this equation. Remember that temperature must be in kelvin units whenever it is used in a gas law equation.

EXAMPLE PROBLEM 14-3

Gay-Lussac's Law

The pressure of a gas in a tank is 3.20 atm at 22.0°C. If the temperature rises to 60.0°C, what will be the gas pressure in the tank?

1. Analyze the Problem

You are given the initial pressure and the initial and final temperatures of a gas sample. Gay-Lussac's law states that if the temperature of a gas increases, so does its pressure. Because the temperature in this problem is increasing, the pressure will increase. So the initial pressure should be multiplied by a volume ratio greater than one.

Known

$$\begin{aligned} P_1 &= 3.20 \text{ atm} \\ T_1 &= 22.0^\circ\text{C} \\ T_2 &= 60.0^\circ\text{C} \end{aligned}$$

Unknown

$$P_2 = ? \text{ atm}$$

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$$T_K = 273 + T_C$$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$$T_1 = 273 + 22.0^\circ\text{C} = 295 \text{ K}$$

$$T_2 = 273 + 60.0^\circ\text{C} = 333 \text{ K}$$

Multiply both sides of the equation for Gay-Lussac's law by T_2 to solve for P_2 .

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$P_2 = P_1 \left(\frac{T_2}{T_1} \right)$$



This cooker is sealed so that the volume is constant. Pressure increases in the cooker as temperature increases.

Substitute the known values into the rearranged equation.

$$P_2 = 3.20 \text{ atm} \left(\frac{333 \text{ K}}{295 \text{ K}} \right)$$

Multiply and divide numbers and units to solve for P_2 .

$$P_2 = 3.20 \text{ atm} \left(\frac{333 \cancel{\text{K}}}{295 \cancel{\text{K}}} \right) = 3.61 \text{ atm}$$

3. Evaluate the Answer

Gay-Lussac's law states that pressure and temperature are directly proportional. Kelvin temperature shows a small increase, so you expect the pressure to show a small increase, which agrees with the answer calculated. The unit is atm, a pressure unit.

PRACTICE PROBLEMS

Assume that the volume and the amount of gas are constant in the following problems.

9. A gas in a sealed container has a pressure of 125 kPa at a temperature of 30.0°C. If the pressure in the container is increased to 201 kPa, what is the new temperature?
10. The pressure in an automobile tire is 1.88 atm at 25.0°C. What will be the pressure if the temperature warms up to 37.0°C?
11. Helium gas in a 2.00-L cylinder is under 1.12 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature of the gas in the cylinder?
12. If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how much does the temperature have to decrease to lower the pressure to 28.4 kPa?
13. A rigid plastic container holds 1.00 L methane gas at 660 torr pressure when the temperature is 22.0°C. How much more pressure will the gas exert if the temperature is raised to 44.6°C?



For more practice with Gay-Lussac's law problems, go to **Supplemental Practice Problems** in Appendix A.

You have seen how the variables of temperature, pressure, and volume affect a gas sample. The gas laws covered in this section each relate two of these three variables if the other variable remains constant. What happens when all three of these variables change? You'll investigate this situation in the next section.

Section 14.1 Assessment

14. State Boyle's, Charles's, and Gay-Lussac's laws using sentences, then equations.
15. A weather balloon of known initial volume is released. The air pressures at its initial and final altitudes are known. Why can't you find its new volume by using these known values and Boyle's law?
16. Which of the three variables that apply to equal amounts of gases are directly proportional? Which are inversely proportional?
17. **Thinking Critically** Explain why gases such as the oxygen found in tanks used at hospitals are compressed. Why must care be taken to prevent compressed gases from reaching a high temperature?
18. **Concept Mapping** Draw a concept map that shows the relationship among pressure, volume, and temperature variables for gases and Boyle's, Charles's, and Gay-Lussac's laws.

The Combined Gas Law and Avogadro's Principle

Objectives

- **State** the relationship among temperature, volume, and pressure as the combined gas law.
- **Apply** the combined gas law to problems involving the pressure, temperature, and volume of a gas.
- **Relate** numbers of particles and volumes by using Avogadro's principle.

Vocabulary

combined gas law
Avogadro's principle
molar volume

In the previous section, you applied the three gas laws covered so far to problems in which either pressure, volume, or temperature of a gas sample was held constant as the other two changed. As illustrated in **Figure 14-6**, in a number of applications involving gases, all three variables change. If all three variables change, can you calculate what their new values will be? In this section you will see that Boyle's, Charles's, and Gay-Lussac's laws can be combined into a single equation that can be used for just that purpose.

The Combined Gas Law

Boyle's, Charles's, and Gay-Lussac's laws can be combined into a single law. This **combined gas law** states the relationship among pressure, volume, and temperature of a fixed amount of gas. All three variables have the same relationship to each other as they have in the other gas laws: Pressure is inversely proportional to volume and directly proportional to temperature, and volume is directly proportional to temperature. The equation for the combined gas law can be expressed as

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

As with the other gas laws, this equation allows you to use known values for the variables under one set of conditions to find a value for a missing variable under another set of conditions. Whenever five of the six values from the two sets of conditions are known, the sixth can be calculated using this expression for the combined gas law.

This combined law lets you work out problems involving more variables that change, and it also provides a way for you to remember the other three laws without memorizing each equation. If you can write out the combined gas law equation, equations for the other laws can be derived from it by remembering which variable is held constant in each case.

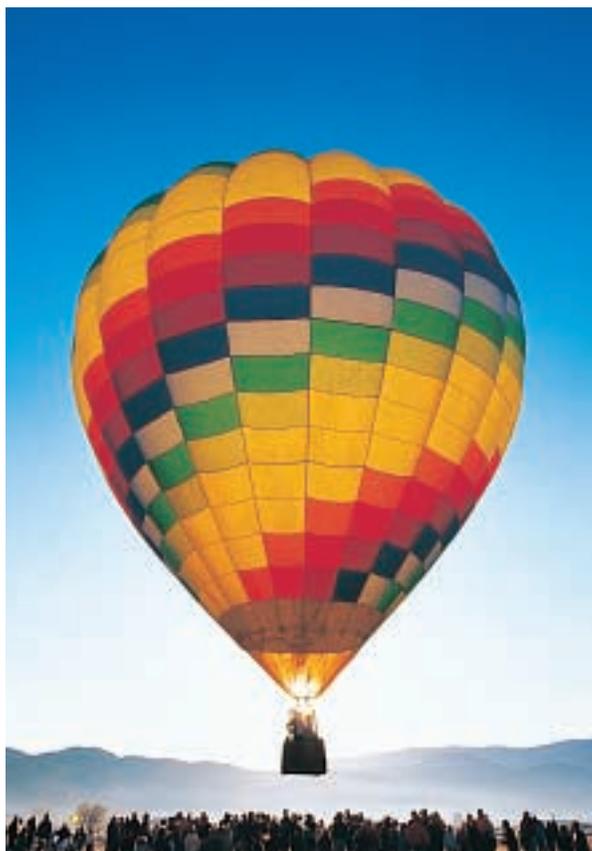
For example, if temperature remains constant as pressure and volume vary, then $T_1 = T_2$. After simplifying the combined gas law under these conditions, you are left with

$$P_1 V_1 = P_2 V_2$$

You should recognize this equation as the equation for Boyle's law. See whether you can derive Charles's and Gay-Lussac's laws from the combined gas law.

Figure 14-6

Constructing an apparatus that uses gases must take into account the changes in gas variables such as pressure, volume, and temperature that can take place. As a hot-air balloonist ascends in the sky, pressure and temperature both decrease, and the volume of the gas in the balloon is affected by those changes.



EXAMPLE PROBLEM 14-4

The Combined Gas Law

A gas at 110 kPa and 30.0°C fills a flexible container with an initial volume of 2.00 L. If the temperature is raised to 80.0°C and the pressure increased to 440 kPa, what is the new volume?

1. Analyze the Problem

You are given the initial pressure, temperature, and volume of a gas sample as well as the final pressure and temperature. The volume of a gas is directly proportional to kelvin temperature, so volume increases as temperature increases. Therefore the volume should be multiplied by a temperature factor greater than one. Volume is inversely proportional to pressure, so as pressure increases, volume decreases. Therefore the volume should be multiplied by a pressure factor that is less than one.

Known

$$\begin{aligned}P_1 &= 110 \text{ kPa} \\T_1 &= 30.0^\circ\text{C} \\V_1 &= 2.00 \text{ L} \\T_2 &= 80.0^\circ\text{C} \\P_2 &= 440 \text{ kPa}\end{aligned}$$

Unknown

$$V_2 = ? \text{ L}$$

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$$T_K = 273 + T_C$$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$$T_1 = 273 + 30.0^\circ\text{C} = 303 \text{ K}$$

$$T_2 = 273 + 80.0^\circ\text{C} = 353 \text{ K}$$

Multiply both sides of the equation for the combined gas law by T_2 and divide it by P_2 to solve for V_2 .

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$V_2 = V_1 \left(\frac{P_1}{P_2} \right) \left(\frac{T_2}{T_1} \right)$$

Substitute the known values into the rearranged equation.

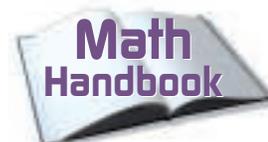
$$V_2 = 2.00 \text{ L} \left(\frac{110 \text{ kPa}}{440 \text{ kPa}} \right) \left(\frac{353 \text{ K}}{303 \text{ K}} \right)$$

Multiply and divide numbers and units to solve for V_2 .

$$V_2 = 2.00 \text{ L} \left(\frac{110 \text{ kPa}}{440 \text{ kPa}} \right) \left(\frac{353 \text{ K}}{303 \text{ K}} \right) = 0.58 \text{ L}$$

3. Evaluate the Answer

Increasing the temperature causes the volume to increase, but increasing the pressure causes the volume to decrease. Because the pressure change is much greater than the temperature change, the volume undergoes a net decrease. The calculated answer agrees with this. The unit is L, a volume unit.



Review rearranging algebraic equations in the **Math Handbook** on page 897 of this textbook.



For more practice with combined gas law problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Assume that the amount of gas is constant in the following problems.

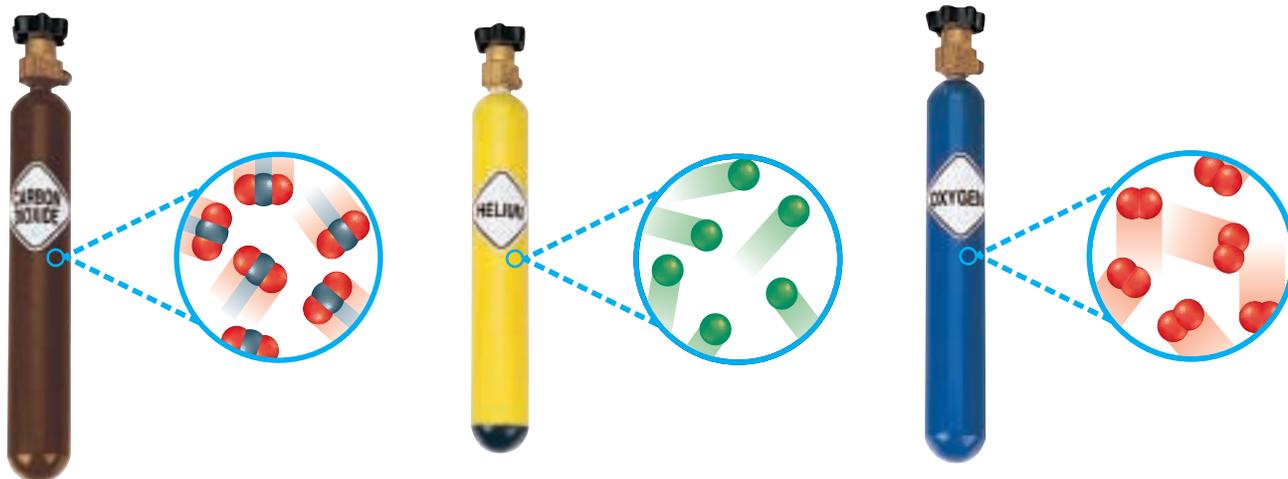
19. A helium-filled balloon at sea level has a volume of 2.1 L at 0.998 atm and 36°C. If it is released and rises to an elevation at which the pressure is 0.900 atm and the temperature is 28°C, what will be the new volume of the balloon?
20. At 0.00°C and 1.00 atm pressure, a sample of gas occupies 30.0 mL. If the temperature is increased to 30.0°C and the entire gas sample is transferred to a 20.0-mL container, what will be the gas pressure inside the container?
21. A sample of air in a syringe exerts a pressure of 1.02 atm at a temperature of 22.0°C. The syringe is placed in a boiling water bath at 100.0°C. The pressure of the air is increased to 1.23 atm by pushing the plunger in, which reduces the volume to 0.224 mL. What was the original volume of the air?
22. An unopened, cold 2.00-L bottle of soda contains 46.0 mL of gas confined at a pressure of 1.30 atm at a temperature of 5.0°C. If the bottle is dropped into a lake and sinks to a depth at which the pressure is 1.52 atm and the temperature is 2.09°C, what will be the volume of gas in the bottle?
23. A sample of gas of unknown pressure occupies 0.766 L at a temperature of 298 K. The same sample of gas is then tested under known conditions and has a pressure of 32.6 kPa and occupies 0.644 L at 303 K. What was the original pressure of the gas?

Avogadro's Principle

The particles making up different gases can vary greatly in size. However, according to the kinetic-molecular theory, the particles in a gas sample are usually far enough apart that size has a negligible influence on the volume occupied by a fixed number of particles, as shown in **Figure 14-7**. For example, 1000 relatively large krypton gas particles occupy the same volume as 1000 much smaller helium gas particles at the same temperature and pressure. It was Avogadro who first proposed this idea in 1811. Today, it is known as **Avogadro's principle**, which states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

Figure 14-7

Compressed gas tanks of equal volume that are at the same pressure and temperature contain equal numbers of gas particles, regardless of which gas they contain. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



Remember from Chapter 11 that the most convenient unit for counting numbers of atoms or molecules is the mole. One mole contains 6.02×10^{23} particles. The **molar volume** for a gas is the volume that one mole occupies at 0.00°C and 1.00 atm pressure. These conditions of temperature and pressure are known as standard temperature and pressure (STP). Avogadro showed experimentally that one mole of any gas will occupy a volume of 22.4 L at STP. The fact that this value is the same for all gases greatly simplifies many gas law calculations. Because the volume of one mole of a gas at STP is 22.4 L , you can use the following conversion factor to find the number of moles, the mass, and even the number of particles in a gas sample.

$$\text{Conversion factor: } \frac{22.4\text{ L}}{1\text{ mol}}$$

Suppose you want to find the number of particles in a sample of gas that has a volume of 3.72 L at STP. First, find the number of moles of gas in the sample.

$$3.72\cancel{\text{ L}} \times \frac{1\text{ mol}}{22.4\cancel{\text{ L}}} = 0.166\text{ mol}$$

A mole of gas contains 6.02×10^{23} gas particles. Use this definition to convert the number of moles to the number of particles.

$$0.166\text{ mol} \times \frac{6.02 \times 10^{23}\text{ particles}}{1\text{ mol}} = 9.99 \times 10^{22}\text{ particles}$$

The following example problems show you how to use molar volume in other ways.

EXAMPLE PROBLEM 14-5

Avogadro's Principle—Using Moles

Calculate the volume that 0.881 mol of gas at standard temperature and pressure (STP) will occupy.

1. Analyze the Problem

You are given the temperature and pressure of a gas sample and the amount of gas the sample contains. According to Avogadro's principle, 1 mol of gas occupies 22.4 L at STP. The number of moles of gas should be multiplied by the conversion factor $\frac{22.4\text{ L}}{1\text{ mol}}$ to find the volume.

Known

$n = 0.881\text{ moles}$
 $T = 0.00^\circ\text{C}$
 $P = 1.00\text{ atm}$

Unknown

$V = ?\text{ L}$

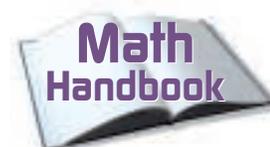
2. Solve for the Unknown

Because conditions are already at STP, multiply the known number of moles by the conversion factor that relates liters to moles to solve for the unknown volume.

$$V = 0.881\text{ mol} \left(\frac{22.4\text{ L}}{1\text{ mol}} \right) = 19.7\text{ L}$$

3. Evaluate the Answer

Because the amount of gas present is slightly less than one mole, you expect the answer to be slightly less than 22.4 L . The answer agrees with that prediction. The unit in the answer is L , a volume unit.



Review unit conversion in the **Math Handbook** on page 901 of this textbook.





For more practice with Avogadro's principle problems that use moles, go to

Supplemental Practice Problems in Appendix A.

PRACTICE PROBLEMS

24. Determine the volume of a container that holds 2.4 mol of gas at STP.
25. What size container do you need to hold 0.0459 mol N_2 gas at STP?
26. What volume will 1.02 mol of carbon monoxide gas occupy at STP?
27. How many moles of nitrogen gas will be contained in a 2.00-L flask at STP?
28. If a balloon will rise off the ground when it contains 0.0226 mol of helium in a volume of 0.460 L, how many moles of helium are needed to make the balloon rise when its volume is 0.865 L? Assume that temperature and pressure stay constant.

EXAMPLE PROBLEM 14-6

Avogadro's Principle—Using Mass

Calculate the volume that 2.0 kg of methane gas (CH_4) will occupy at STP.

1. Analyze the Problem

You are given the temperature and pressure of a gas sample and the mass of gas the sample contains. One mole of a gas occupies 22.4 L at STP. The number of moles can be calculated by dividing the mass of the sample, m , by its molecular mass, M .

Known

$$\begin{aligned} m &= 2.00 \text{ kg} \\ T &= 0.00^\circ\text{C} \\ P &= 1.00 \text{ atm} \end{aligned}$$

Unknown

$$V = ? \text{ L}$$

2. Solve for the Unknown

Use atomic masses and numbers of each type of atom to determine molecular mass for methane. Express that molecular mass in grams per mole to determine molar mass.

$$\begin{aligned} M &= \frac{1 \text{ C-atom} \times 12.01 \text{ amu}}{\text{C-atom}} + \frac{4 \text{ H-atoms} \times 1.01 \text{ amu}}{\text{H-atom}} \\ &= 12.01 \text{ amu} + 4.04 \text{ amu} = 16.05 \text{ amu}; 16.05 \text{ g/mol} \end{aligned}$$

Multiply the mass of methane by a conversion factor to change it from kg to g.

$$2.00 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) = 2.00 \times 10^3 \text{ g}$$

Divide the mass in grams of methane by its molecular mass to find the number of moles.

$$\frac{m}{M} = \frac{2.00 \times 10^3 \text{ g}}{16.05 \text{ g/mol}} = 125 \text{ mol}$$

Because conditions are already at STP, multiply the known number of moles by the conversion factor of 22.4 L/1 mol to solve for the unknown volume.

$$V = 125 \text{ mol} \left(\frac{22.4 \text{ L}}{1 \text{ mol}} \right) = 2.80 \times 10^3 \text{ L}$$

3. Evaluate the Answer

The mass of methane present is much more than 1 mol, so you expect a large volume, which is in agreement with the answer. The units are L, a volume unit.



Avogadro's principle is essential to manufacturers that use gases. This factory produces ammonia.

PRACTICE PROBLEMS

29. How many grams of carbon dioxide gas are in a 1.0-L balloon at STP?
30. What volume in milliliters will 0.00922 g H_2 gas occupy at STP?
31. What volume will 0.416 g of krypton gas occupy at STP?
32. A flexible plastic container contains 0.860 g of helium gas in a volume of 19.2 L. If 0.205 g of helium is removed without changing the pressure or temperature, what will be the new volume?
33. Calculate the volume that 4.5 kg of ethylene gas (C_2H_4) will occupy at STP.



For more practice with Avogadro's principle problems that use mass, go to **Supplemental Practice Problems** in Appendix A.

Why aren't weather balloons completely inflated when they are released? How strong must the walls of a scuba tank be? Look at the example of the combined gas law and Avogadro's principle shown in **Figure 14-8**. Using the combined gas law and Avogadro's principle together will help you understand how gases are affected by pressure, temperature, and volume.



Figure 14-8

The combined gas law and Avogadro's principle have many practical applications that scientists and manufacturers must consider.

Section 14.2 Assessment

34. State the combined gas law using a sentence and then an equation.
35. What variable is assumed to be constant when using the combined gas law?
36. What three laws are used to make the combined gas law?
37. Explain why Avogadro's principle holds true for gases that have large particles and also for gases that have small particles.
38. Why must conditions of temperature and pressure be stated to do calculations involving molar volume?
39. **Thinking Critically** Think about what happens when a bottle of carbonated soft drink is shaken before being opened. Use the gas laws to explain whether the effect will be greater when the liquid is warm or cold.
40. **Applying Concepts** Imagine that you are going on an airplane trip in an unpressurized plane. You are bringing aboard an air-filled pillow that you have inflated fully. Predict what will happen when you try to use the pillow while the plane is at its cruising altitude.

Objectives

- **Relate** the amount of gas present to its pressure, temperature, and volume by using the ideal gas law.
- **Compare** the properties of real and ideal gases.

Vocabulary

ideal gas constant (R)
ideal gas law

In the last section, you learned that Avogadro noted the importance of being able to calculate the number of moles of a gas present under a given set of conditions. The laws of Avogadro, Boyle, Charles, and Gay-Lussac can be combined into a single mathematical statement that describes the relationship among pressure, volume, temperature, and number of moles of a gas. This formula is called the ideal gas law because it works best when applied to problems involving gases contained under certain conditions. The particles in an ideal gas are far enough apart that they exert minimal attractive or repulsive forces on one another and occupy a negligible volume.

The Ideal Gas Law

The number of moles is a fourth variable that can be added to pressure, volume, and temperature as a way to describe a gas sample. Recall that as the other gas laws were presented, care was taken to state that the relationships hold true for a “fixed mass” or a “given amount” of a gas sample. Changing the number of gas particles present will affect at least one of the other three variables.

As **Figure 14-9** illustrates, increasing the number of particles present in a sample will raise the pressure if the volume and temperature are kept constant. If the pressure and temperature are constant and more gas particles are added, the volume will increase.

Because pressure, volume, temperature, and the number of moles present are all interrelated, it would be helpful if one equation could describe their relationship. Remember that the combined gas law relates volume, temperature, and pressure of a sample of gas.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

For a specific sample of gas, you can see that this relationship of pressure, volume, and temperature is always the same. You could say that

$$\frac{PV}{T} = k$$

where k is a constant based on the amount of gas present, n . Experiments using known values of P , T , V , and n show that

$$k = nR$$

where R represents an experimentally determined constant that is referred to as the **ideal gas constant**. Therefore, the **ideal gas law**,

$$PV = nRT$$

describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present. The ideal gas law is used in the **CHEMLAB** in this chapter.

The ideal gas constant In the ideal gas equation, the value of R depends on the units used for pressure. **Table 14-1** shows the numerical value of R for different units of pressure.

Figure 14-9

The volume and temperature of this tire stay the same as air is added. However, the pressure in the tire increases as the amount of air present increases.



Table 14-1

Numerical Values of the Gas Constant, R					
Units of R	Numerical value of R	Units of P	Units of V	Units of T	Units of n
$\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$	0.0821	atm	L	K	mol
$\frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}$	8.314	kPa	L	K	mol
$\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$	62.4	mm Hg	L	K	mol

The R value you will probably find most useful is the first one listed in the table, $0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$. Use this R in problems in which the unit of volume is liters, the pressure is in atmospheres, and the temperature is in kelvins.

Real versus ideal gases What does the term *ideal gas* mean? An ideal gas is one whose particles take up no space and have no intermolecular attractive forces. An ideal gas follows the gas laws under all conditions of temperature and pressure.

In the real world, no gas is truly ideal. All gas particles have some volume, however small it may be, because of the sizes of their atoms and the lengths of their bonds. All gas particles also are subject to intermolecular interactions. Despite that, most gases will behave like ideal gases at many temperature and pressure levels. Under the right conditions of temperature and pressure, calculations made using the ideal gas law closely approximate actual experimental measurements.

When is the ideal gas law not likely to work for a real gas? Real gases deviate most from ideal gas behavior at extremely high pressures and low temperatures. As the amount of space between particles and the speed at which the particles move decrease, the effects of the volume of gas particles and intermolecular attractive forces become increasingly important. The gas behaves as a real gas in **Figure 14-10a**. Lowering the temperature of nitrogen gas results in less kinetic energy of the gas particles, which means their intermolecular attractive forces are strong enough to bond them more closely together. When the temperature is low enough, this real gas condenses to form a liquid. The gas in **Figure 14-10b** also behaves as a real gas. Increasing the pressure on a gas such as propane lowers the volume and forces the gas particles closer together until their volume is no longer negligible compared to the volume of the tank. Real gases such as propane will liquefy if enough pressure is applied.

Figure 14-10

Real gases deviate most from ideal behavior at low temperatures and high pressures.

- a** Liquid nitrogen is used to store biological tissue samples at low temperatures.
- b** Increased pressure allows a larger mass of propane to fit into a smaller volume for easier transport. Propane is sold as LP (liquid propane) for this reason, although it is actually burned for fuel as a gas.

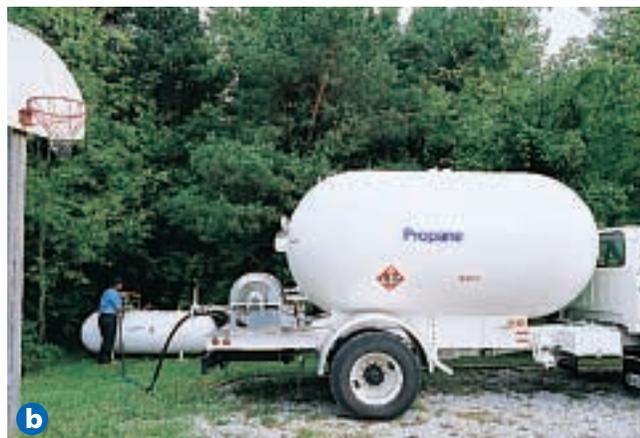
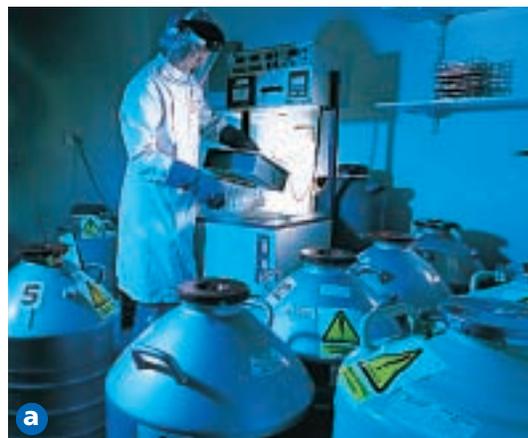
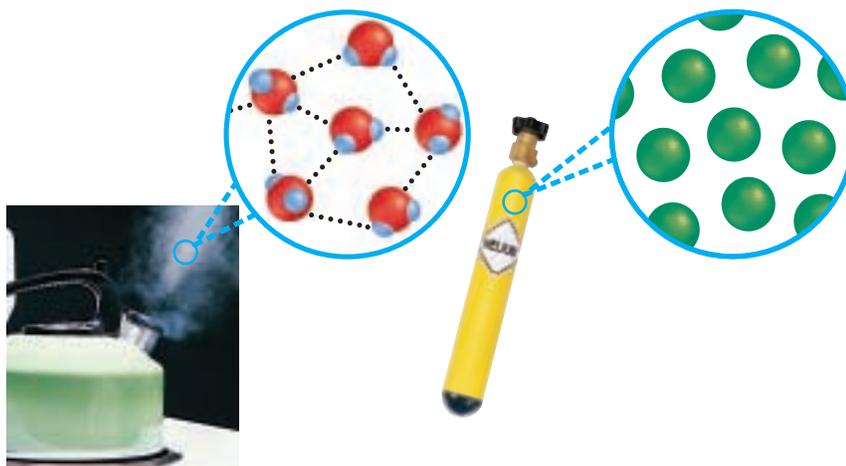


Figure 14-11

In polar gas molecules, such as water vapor, oppositely charged poles attract each other through electrostatic forces.

A nonpolar gas, such as helium, is uncharged. Thus nonpolar gases are more likely to behave like ideal gases than are polar gases.



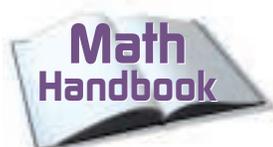
The nature of the particles making up a gas also affects how ideally the gas behaves. For example, polar gas molecules such as water vapor generally have larger attractive forces between their particles than nonpolar molecules such as chlorine gas. The oppositely charged ends of polar molecules are pulled together through electrostatic forces, as shown in **Figure 14-11**. Therefore polar gases do not behave as ideal gases. Also, the particles of gases composed of molecules such as butane (C_4H_{10}) occupy more actual volume than an equal number of gas particles of smaller molecules such as helium (He). Therefore, larger gas molecules tend to cause a greater departure from ideal behavior than do smaller gas molecules.

Applying the Ideal Gas Law

Look again at the combined gas law on page 428. Notice that it cannot be used to find n , the number of moles of a gas. However, the ideal gas law can be used to solve for the value of any one of the four variables P , V , T , or n if the values of the other three are known. Rearranging the $PV = nRT$ equation allows you to also calculate the molar mass and density of a gas sample if the mass of the sample is known.

To find the molar mass of a gas sample, the mass, temperature, pressure, and volume of the gas must be known. Remember from Chapter 12 that the number of moles of a gas (n) is equal to the mass (m) divided by the molecular mass (M). Therefore, the n in the equation can be replaced by m/M .

$$PV = \frac{mRT}{M} \text{ or } M = \frac{mRT}{PV}$$



Review fractions in the **Math Handbook** on page 907 of this textbook.

EXAMPLE PROBLEM 14-7

The Ideal Gas Law—Using Moles

Calculate the number of moles of gas contained in a 3.0-L vessel at 3.00×10^2 K with a pressure of 1.50 atm.

1 Analyze the Problem

You are given the volume, temperature, and pressure of a gas sample. When using the ideal gas law to solve for n , choose the value of R that contains the pressure and temperature units given in the problem.

Known

$$V = 3.0 \text{ L}$$

$$T = 3.00 \times 10^2 \text{ K}$$

$$P = 1.50 \text{ atm}$$

$$R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

Unknown

$$n = ? \text{ mol}$$

2. Solve for the Unknown

Divide both sides of the ideal gas law equation by RT to solve for n .

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

Substitute the known values into the rearranged equation.

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(3.00 \times 10^2 \text{ K})}$$

Multiply and divide numbers and units to solve for n .

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(3.00 \times 10^2 \text{ K})} = 0.18 \text{ mol}$$

3. Evaluate the Answer

One mole of a gas at STP occupies 22.4 L. In this problem, the volume is much smaller than 22.4 L while the temperature and pressure values are not too different from those at STP. The answer agrees with the prediction that the number of moles present will be significantly less than one mole. The unit of the answer is the mole.

PRACTICE PROBLEMS

41. If the pressure exerted by a gas at 25°C in a volume of 0.044 L is 3.81 atm, how many moles of gas are present?
42. Determine the Celsius temperature of 2.49 moles of gas contained in a 1.00-L vessel at a pressure of 143 kPa.
43. Calculate the volume that a 0.323-mol sample of a gas will occupy at 265 K and a pressure of 0.900 atm.
44. What is the pressure in atmospheres of a 0.108-mol sample of helium gas at a temperature of 20.0°C if its volume is 0.505 L?
45. Determine the kelvin temperature required for 0.0470 mol of gas to fill a balloon to 1.20 L under 0.988 atm pressure.



For more practice with ideal gas law problems that use moles, go to **Supplemental Practice Problems** in Appendix A.

Recall from Chapter 2 that density (D) is defined as mass (m) per unit volume (V). After rearranging the ideal gas equation to solve for molar mass, D can be substituted for m/V .

$$M = \frac{mRT}{PV} = \frac{DRT}{P}$$

This equation can be rearranged to solve for the density of a gas.

$$D = \frac{MP}{RT}$$

Why might you need to know the density of a gas? Consider what requirements are necessary to fight a fire. One way to put out a fire is to remove its oxygen source by covering it with another gas that will neither burn nor support combustion. This gas must have a greater density than oxygen so that it will fall to the level of the fire. You can observe applications of density when you do the **miniLAB** later in this section and read the **Chemistry and Technology** feature.



EXAMPLE PROBLEM 14-8

The Ideal Gas Law—Using Molar Mass

What is the molar mass of a pure gas that has a density of 1.40 g/L at STP?

1. Analyze the Problem

You are given the density, temperature, and pressure of a sample of gas. Because density is known and mass and volume are not, use the form of the ideal gas equation that involves density.

Known

$$D = 1.40 \frac{\text{g}}{\text{L}}$$

$$T = 0.00^\circ\text{C}$$

$$P = 1.00 \text{ atm}$$

$$R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

Unknown

$$M = ? \frac{\text{g}}{\text{mol}}$$

2. Solve for the Unknown

Convert the standard T to kelvin units.

$$T_{\text{K}} = 273 + T_{\text{C}}$$

$$T_{\text{K}} = 273 + 0.00^\circ\text{C} = 273 \text{ K}$$

Use the form of the ideal gas law that includes density (D) and solves for M .

$$M = \frac{DRT}{P}$$

Substitute the known values into the equation.

$$M = \frac{\left(1.40 \frac{\text{g}}{\text{L}}\right)\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})}{1 \text{ atm}}$$

Multiply and divide numbers and units to solve for M .

$$M = \frac{\left(1.40 \frac{\text{g}}{\text{L}}\right)\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})}{1 \text{ atm}} = 31.4 \text{ g/mol}$$

3. Evaluate the Answer

You would expect the molar mass of a gas to fall somewhere between that of one of the lightest gases under normal conditions, such as 2 g/mol for H_2 , and that of a relatively heavy gas, such as 222 g/mol for Rn . The answer seems reasonable. The unit is g/mol, which is the molar mass unit.



The density of a gas can be used to identify it. This gas is denser than air and can be poured from one container to another.

Practice!

For more practice with ideal gas law problems that use molar mass, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- How many grams of gas are present in a sample that has a molar mass of 70.0 g/mol and occupies a 2.00-L container at 117 kPa and 35.1°C?
- Calculate the grams of N_2 gas present in a 0.600-L sample kept at 1.00 atm pressure and a temperature of 22.0°C.
- What is the density of a gas at STP that has a molar mass of 44.0 g/mol?
- What is the molar mass of a sample of gas that has a density of 1.09 g/mL at 1.02 atm pressure and 25.0°C?
- Calculate the density a gas will have at STP if its molar mass is 39.9 g/mol.

The Density of Carbon Dioxide

Hypothesizing Air is a mixture of mostly nitrogen and oxygen. Use observations to form a hypothesis about which has greater density, air or carbon dioxide.

Materials masking tape, aluminum foil, metric ruler, 1-L beaker, candle, matches, thermometer, barometer or weather radio, baking soda (NaHCO_3), vinegar (5% CH_3COOH)

Procedure

1. Record the temperature and the barometric pressure of the air in the classroom.
2. Roll a 23-cm \times 30-cm piece of aluminum foil into a cylinder that is 6 cm \times 30 cm. Tape the edges with masking tape.
3. Use matches to light a candle. **CAUTION:** Run water over the extinguished match before throwing it away. Keep all hair and loose clothing away from the flame.
4. Place 30 g of baking soda in the bottom of a large beaker. Add 40 mL of vinegar.
5. Quickly position the foil cylinder at approximately 45° up and away from the top of the candle flame.
6. While the reaction in the beaker is actively producing CO_2 gas, carefully pour the gas, but not the liquid, out of the beaker and into the top of the foil tube. Record your observations.



Analysis

1. Based on your observations, state a hypothesis about whether CO_2 is heavier or lighter than air.
2. Use the combined gas law to calculate molar volume at room temperature and atmospheric pressure.
3. Carbon dioxide gas (CO_2) has a molar mass of 44 g/mol. The two major components of air, which are oxygen and nitrogen, have molar masses of 32 g/mol and 28 g/mol, respectively. Calculate the room-temperature densities in g/L of nitrogen (N_2), oxygen (O_2), and carbon dioxide (CO_2) gases.
4. Do these calculations confirm your hypothesis? Explain.

Section 14.3 Assessment

51. Write the equation for the ideal gas law.
52. Use the kinetic-molecular theory to analyze and evaluate the ideal gas law's applicability to real gases.
53. List common units for each variable in the ideal gas law.
54. **Thinking Critically** Which of the following gases would you expect to behave most like an ideal gas at room temperature and atmospheric pressure: water vapor, carbon dioxide, helium, or hydrogen? Explain.

55. **Making and Using Graphs** The accompanying data show the volume of hydrogen gas collected at a number of different temperatures. Illustrate these data with a graph and use them to determine the temperature at which the volume will reach a value of 0 mL. What is this temperature called? For more help, refer to **Drawing Line Graphs** in the **Math Handbook** on page 903 of this text.

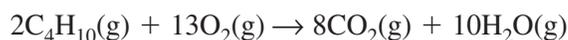
Volume of H_2 Collected at Different Temperatures

Trial	1	2	3	4	5	6
T ($^{\circ}\text{C}$)	300	175	110	0	-100	-150
V (mL)	48	37	32	22	15	11

Objectives

- **Determine** volume ratios for gaseous reactants and products by using coefficients from a chemical equation.
- **Calculate** amounts of gaseous reactants and products in a chemical reaction using the gas laws.

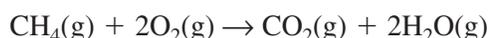
All the laws you have learned so far involving gases can be applied to calculate the stoichiometry of reactions in which gases are reactants or products. Recall that the coefficients in chemical equations represent molar amounts of substances taking part in the reaction. For example, when butane gas burns, the reaction is represented by the following chemical equation.



From the balanced chemical equation, you know that 2 mol of butane reacts with 13 mol of oxygen, producing 8 mol of carbon dioxide and 10 mol of water vapor. By examining this balanced equation, you are able to find mole ratios of substances in this reaction. Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. Thus, when gases are involved, the coefficients in a balanced chemical equation represent not only molar amounts but also relative volumes. For example, if 2 L of butane reacts, the reaction involves 13 L of oxygen and produces 8 L of carbon dioxide and 10 L of water vapor.

Calculations Involving Only Volume

To find the volume of a gaseous reactant or product in a reaction, you must know the balanced chemical equation for the reaction and the volume of at least one other gas involved in the reaction. Examine the reaction showing the combustion of methane, which takes place every time you light a Bunsen burner.



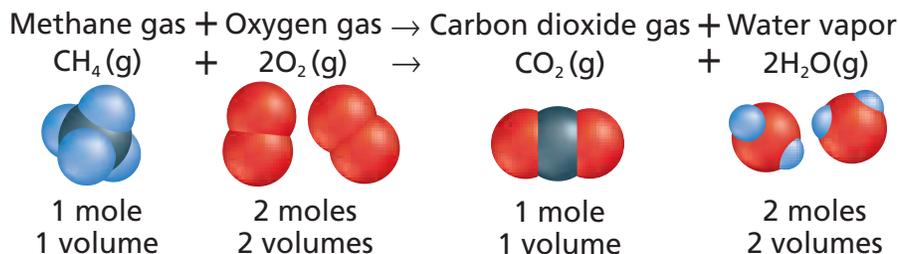
Because the coefficients represent volume ratios for gases taking part in the reaction, you can determine that it takes 2 L of oxygen to react completely with 1 L of methane. The complete combustion of 1 L of methane will produce 1 L of carbon dioxide and 2 L of water vapor, as shown in **Figure 14-12**.

What volume of methane is needed to produce 26 L of water vapor? Because the volume ratio of methane and water vapor is 1:2, the volume of methane needed is half that of the water vapor. Thus, 13 L of methane is needed to produce 26 L of water vapor. What volume of oxygen is needed to produce 6.0 L of carbon dioxide?

Note that no conditions of temperature and pressure are listed. They are not needed as part of the calculation because after mixing, each gas is at the same temperature and pressure. The same temperature and pressure affect all gases in the same way, so these conditions don't need to be considered.

Figure 14-12

The coefficients in a balanced equation show the relationships among numbers of moles of all reactants and products. The coefficients also show the relationships among volumes of any gaseous reactants or products. From these coefficients, volume ratios can be set up for any pair of gases in the reaction.



EXAMPLE PROBLEM 14-9

Volume-Volume Problems

What volume of oxygen gas is needed for the complete combustion of 4.00 L of propane gas (C_3H_8)? Assume constant pressure and temperature.

1 Analyze the Problem

You are given the volume of a gaseous reactant in a chemical reaction. Remember that the coefficients in a balanced chemical equation provide the volume relationships of gaseous reactants and products.

Known

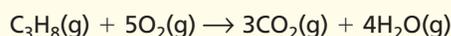
$$V_{C_3H_8} = 4.00 \text{ L}$$

Unknown

$$V_{O_2} = ? \text{ L}$$

2. Solve for the Unknown

Write the balanced equation for the combustion of C_3H_8 .



Use the balanced equation to find the volume ratio for O_2 and C_3H_8 .

$$\frac{5 \text{ volumes } O_2}{1 \text{ volume } C_3H_8}$$

Multiply the known volume of C_3H_8 by the volume ratio to find the volume of O_2 .

$$(4.00 \text{ L } C_3H_8) \times \frac{5 \text{ volumes } O_2}{1 \text{ volume } C_3H_8} = 20.0 \text{ L } O_2$$

3. Evaluate the Answer

The coefficients in the combustion equation show that a much larger volume of O_2 than C_3H_8 is used up in the reaction, which is in agreement with the calculated answer. The unit of the answer is L, a unit of volume.



Correct proportions of gases are needed for many chemical reactions. The combustion of propane heats the air that inflates the balloon.

PRACTICE PROBLEMS

56. What volume of oxygen is needed to react with solid sulfur to form 3.5 L SO_2 ?
57. Determine the volume of hydrogen gas needed to react completely with 5.00 L of oxygen gas to form water.
58. How many liters of propane gas (C_3H_8) will undergo complete combustion with 34.0 L of oxygen gas?
59. What volume of oxygen is needed to completely combust 2.36 L of methane gas (CH_4)?



For more practice with volume-volume problems, go to **Supplemental Practice Problems** in Appendix A.

Calculations Involving Volume and Mass

To do stoichiometric calculations that involve both gas volumes and masses, you must know the balanced equation for the reaction involved, at least one mass or volume value for a reactant or product, and the conditions under which the gas volumes have been measured. Then the ideal gas law can be used along with volume or mole ratios to complete the calculation.

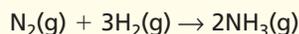
In doing this type of problem, remember that the balanced chemical equation allows you to find ratios for moles and gas volumes only—not for masses. All masses given must be converted to moles or volumes before being used as part of a ratio. Also remember that the temperature units used must be kelvin.



EXAMPLE PROBLEM 14-10

Volume-Mass Problems

Ammonia is synthesized from hydrogen and nitrogen gases.



If 5.00 L of nitrogen reacts completely by this reaction at a constant pressure and temperature of 3.00 atm and 298 K, how many grams of ammonia are produced?

1. Analyze the Problem

You are given the volume, pressure, and temperature of a gas sample. The mole and volume ratios of gaseous reactants and products are given by the coefficients in the balanced chemical equation. Volume can be converted to moles and thus related to mass by using molar mass and the ideal gas law.

Known

$$V_{\text{N}_2} = 5.00 \text{ L}$$

$$P = 3.00 \text{ atm}$$

$$T = 298 \text{ K}$$

Unknown

$$m_{\text{NH}_3} = ? \text{ g}$$

2. Solve for the Unknown

Determine volume ratios from the balanced chemical equation.

$$\frac{1 \text{ volume N}_2}{2 \text{ volumes NH}_3}$$

Use this ratio to determine how many liters of gaseous ammonia will be made from 5.00 L of nitrogen gas.

$$5.00 \text{ L N}_2 \left(\frac{2 \text{ volumes NH}_3}{1 \text{ volume N}_2} \right) = 10.0 \text{ L NH}_3$$

Rearrange the equation for the ideal gas law to solve for n .

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

Substitute the known values into the rearranged equation using the volume of NH_3 for V . Multiply and divide numbers and units to solve for the number of moles of NH_3 .

$$n = \frac{(3.00 \text{ atm})(10.0 \text{ L})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298 \text{ K})} = 1.23 \text{ mol NH}_3$$

Find the molar mass, M , of NH_3 by finding the molecular mass and expressing it in units of g/mol.

$$M = \left(\frac{1 \text{ N atom} \times 14.01 \text{ amu}}{\text{N atom}} \right) + \left(\frac{3 \text{ H atoms} \times 1.01 \text{ amu}}{\text{H atom}} \right)$$
$$= 17.04 \text{ amu}$$

$$M = 17.04 \frac{\text{g}}{\text{mol}}$$

Convert moles of ammonia to grams of ammonia using molar mass of ammonia as a conversion factor.

$$1.23 \text{ mol NH}_3 \times 17.04 \frac{\text{g}}{\text{mol}} = 21.0 \text{ g NH}_3$$

3. Evaluate the Answer

To check your answer, calculate the volume of reactant nitrogen at STP. Then use molar volume and the mole ratio between N_2 and NH_3 to determine how many moles of NH_3 were produced. You can convert the answer to grams using the molar mass of NH_3 . All data was given in three significant digits as is the answer.



Ammonia is essential in the production of chemical fertilizers.

PRACTICE PROBLEMS

- 60.** Ammonium nitrate is a common ingredient in chemical fertilizers. Use the reaction shown to calculate the mass of solid ammonium nitrate that must be used to obtain 0.100 L of dinitrogen oxide gas at STP.

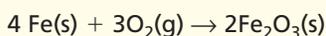


- 61.** Calcium carbonate forms limestone, one of the most common rocks on Earth. It also forms stalactites, stalagmites, and many other types of formations found in caves. When calcium carbonate is heated, it decomposes to form solid calcium oxide and carbon dioxide gas.



How many liters of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

- 62.** Determine how many moles of water vapor will be produced at 1.00 atm and 200°C by the complete combustion of 10.5 L of methane gas (CH_4).
- 63.** When iron rusts, it undergoes a reaction with oxygen to form iron(III) oxide.



Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

- 64.** Solid potassium metal will react with Cl_2 gas to form ionic potassium chloride. How many liters of Cl_2 gas are needed to completely react with 0.204 g of potassium at STP?



For more practice with volume-mass problems, go to **Supplemental Practice Problems** in Appendix A.

Stoichiometric problems such as these are considered in industrial processes that involve gases. How much of a reactant should be purchased? How much of a product will be produced? What conditions of temperature and pressure are necessary? Answers to these questions are essential to effective production of a product.

Section 14.4 Assessment

- 65.** How do mole ratios compare to volume ratios for gaseous reactants and products in a balanced chemical equation?
- 66.** Is the volume of a gas directly or indirectly proportional to the number of moles of a gas at constant temperature and pressure? Explain.
- 67.** Determine the volume ratio of ammonia to nitrogen in the reaction shown.
- $$3\text{H}_2 + \text{N}_2 \rightarrow 2\text{NH}_3$$
- Which will occupy a larger volume at a given temperature and pressure: one mole of H_2 or one mole of NH_3 ?
- 68. Thinking Critically** One mole of a gas occupies a volume of 22.4 L at STP. Calculate the temperature and pressure conditions needed to fit two moles of a gas into a volume of 22.4 L.
- 69. Predicting** Using what you have learned about gases, predict what will happen to the size of the reaction vessel you need to carry out a reaction involving gases if the temperature is doubled and the pressure is held constant.

Using the Ideal Gas Law

The ideal gas law is a powerful tool that the chemist—and now you—can use to determine the molar mass of an unknown gas. By measuring the temperature, pressure, volume, and mass of a gas sample, you can calculate the molar mass of the gas.

Problem

How can the equation for the ideal gas law be used to calculate the molar mass of a gas?

Objectives

- **Measure** the mass, volume, temperature, and pressure of an insoluble gas collected over water.
- **Calculate** the molar mass of an unknown gas using the ideal gas equation.

Materials

aerosol can of duster	plastic microtip pipette
600-mL graduated beaker	latex tubing
bucket or bowl	glass tubing
thermometer (°C)	scissors
barometer or weather radio	electrical or duct tape
	balance

Safety Precautions



- Read and observe all cautions listed on the aerosol can of office equipment duster.
- Do not have any open flames in the room.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Because you will collect the aerosol gas over water, the beaker contains both the aerosol gas and water vapor. Form a hypothesis about how the presence of water vapor will affect the calculated value of the molar mass of the gas. Explain.
4. The following gases are or have been used in aerosol cans, some as propellants. Use the gases' molecular formulas to calculate their molar masses.
 - a. propane, C_3H_8
 - b. butane, C_4H_{10}
 - c. dichlorodifluoromethane, CCl_2F_2
 - d. tetrafluoroethane, $C_2H_2F_4$
5. Given the following data for a gas, use the equation for the ideal gas law to calculate the molar mass.
 - a. mass = 0.810 g
 - b. pressure = 0.954 atm
 - c. volume = 0.461 L
 - d. temperature = 291 K

Data and Calculations

Mass of can before release of gas (g)	
Mass of can after release of gas (g)	
Mass of gas released (g)	
Air temperature (°C)	
Air temperature (K)	
Air pressure (list what unit was used)	
Air pressure (atm)	
Volume of gas collected (L)	

Procedure

1. Place the bucket in the sink and fill it with water.
2. Submerge the beaker in the water. Then, invert it in the bucket, being careful to keep it completely filled with water.
3. Measure the mass of an aerosol can of office equipment duster. Record the mass in the data table.

- Use scissors to cut the stem from a plastic microtip pipette.
- Fit the pipette stem over the long plastic spray tip that comes with the aerosol can to extend the length of the tip and enlarge the diameter.
- Connect one end of 30 cm of latex tubing to glass tubing that is 8 cm long.
- Connect the other end to the pipette stem that is attached to the aerosol can. If necessary, tape any connections so that they don't leak.
- Place the end of the glass tubing under the pour spout of the inverted beaker as shown in the photo.



- Hold the beaker down while you slowly release the gas from the aerosol can. Collect between 400 and 500 mL of the gas by water displacement.
- To equalize the air pressure, lift the beaker so that the water level inside and outside the beaker is the same.
- Carefully read the volume of the gas collected using the graduations on the beaker.
- Record this volume of the gas collected in the data table.
- Remove the tubing from the aerosol can.
- Measure the mass of the can and record it in the data table.
- Using a barometer or weather radio, record the atmospheric pressure in the data table.

- Using a thermometer, determine air temperature. Record it in the data table.

Cleanup and Disposal

- Dispose of the empty can according to the instructions on its label.
- Pour the water down the drain.
- Discard any tape and the pipettes in the trash can.
- Return all lab equipment to its proper place.

Analyze and Conclude

- Using Numbers** Fill in the remainder of the data table by calculating the mass of the gas that was released from the aerosol can, converting the atmospheric pressure from the units measured into atmospheres, and converting the air temperature into kelvins. Substitute your data from the table into the form of the ideal gas equation that solves for M . Calculate the molar mass of the gas in the can using the appropriate value for R .
- Using Numbers** Read the contents of the can and determine which of the gases from step 4 in the Pre-Lab is the most likely propellant.
- Error Analysis** Remember that you are collecting the gas after it has bubbled through water. What might happen to some of the gas as it goes through the water? What might be present in the gas in the beaker in addition to the gas from the can? Calculate the percent error using your calculated molar mass compared to the molar mass of the gas in the aerosol can.
- Interpreting Data** Were your data consistent with the ideal gas law? Evaluate the pressure and temperature at which your experiment was done, and the polarity of the gas. Would you expect the gas in your experiment to behave as an ideal gas or a real gas?

Real-World Chemistry

- Explain why the label on an aerosol can warns against exposing the can to high heat.
- Use the ideal gas law to explain why the wind blows.

CHEMISTRY and Technology

Giving a Lift to Cargo

Have you ever been caught in a traffic jam caused by a truck carrying an oversized piece of industrial equipment? The need to transport heavy loads more efficiently is one factor leading to the revival of a technology that most people considered to be dead—the airship, sometimes called a dirigible or zeppelin. The burning of the zeppelin *Hindenburg* in May 1937, followed closely by World War II, ended nearly all commercial use of these “lighter-than-air” ships. But by the 1990s, chemists and engineers had developed strong, lightweight alloys and tough, fiber-reinforced composite plastics. These materials, along with computerized control and satellite navigation systems are making commercial airships practical again.

Modern airships

Modern airships use helium to provide lift. The *Hindenburg* used hydrogen gas, which provides about twice the lift of helium but is no longer used because it is extremely flammable. The helium is contained in bags made of a space-age, leakproof fabric called Tedlar, a type of plastic. The bags are loosely inflated so that the helium pressure is about the same as atmospheric pressure. The quantity of helium determines the lifting ability of the ship.

Airships and Boyle’s law

As the airship rises, atmospheric pressure decreases and the helium expands, as Boyle’s law predicts. As the ship reaches the desired altitude, air is pumped into another bag called a *ballonet*. The pressure of the ballonet prevents the helium bags from expanding further, thus keeping the ship at that altitude. To descend, more air is pumped into the ballonet. This added pressure squeezes the helium bags, causing the volume to decrease and the density of the helium to increase. Also, the compressed air adds weight to the ship. The lifting power of the helium is reduced and the airship moves downward.

Uses for modern airships

The old airships of 1900 to 1940 were used mostly for luxury passenger service. Among modern airships, the German Zeppelin-NT and a ship being



developed by the Hamilton Airship Company in South Africa are designed to carry passengers.

Probably the most interesting new airship is the huge CargoLifter from a German-American company. Its length is slightly less than the length of three football fields. CargoLifter’s skeleton is constructed of a strong carbon fiber composite material that is much less dense than metals. At the mooring mast, it is as tall as a 27-story building and contains 450 000 m³ of helium. It is designed to carry loads up to 160 metric tons (352 000 pounds). It can pick up and deliver objects such as large turbines and entire locomotives. Because modern roads are not needed, equipment can be delivered to locations in developing countries that would otherwise be impossible to reach.

Investigating the Technology

- Thinking Critically** No plastics, fabrics, or metals that are “lighter-than-air” exist. Yet these materials are used to make airships. Why is it possible to describe an airship as a lighter-than-air craft?



Visit the Chemistry Web site at science.glencoe.com to find links to more information about airships.

Summary

14.1 The Gas Laws

- Boyle's law states that the pressure and volume of a contained gas are inversely proportional if temperature is constant.
- Charles's law states that the volume and kelvin temperature of a contained gas are directly proportional if pressure is constant.
- Gay-Lussac's law states that the pressure and kelvin temperature of a contained gas are directly proportional if volume is constant.

14.2 The Combined Gas Law and Avogadro's Principle

- Boyle's, Charles's, and Gay-Lussac's laws are brought together in the combined gas law, which permits calculations involving changes in the three gas variables of pressure, volume, and temperature.
- Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. The volume of one mole of a gas at STP is 22.4 L.

14.3 The Ideal Gas Law

- The combined gas law and Avogadro's principle are used together to form the ideal gas law. In the equation for the ideal gas law, R is the ideal gas constant.
- The ideal gas law allows you to determine the number of moles of a gas when its pressure, temperature, and volume are known.
- Real gases deviate from behavior predicted for ideal gases because the particles of a real gas occupy volume and are subject to intermolecular forces.
- The ideal gas law can be used to find molar mass if the mass of the gas is known, or the density of the gas if its molar mass is known.

14.4 Gas Stoichiometry

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.

Key Equations and Relationships

- Boyle's law: $P_1V_1 = P_2V_2$, constant temperature (p. 421)
- Charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, constant pressure (p. 424)
- Gay-Lussac's law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, constant volume (p. 426)
- Combined gas law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ (p. 428)
- Ideal gas law: $PV = nRT$ (p. 434)
- Finding molar mass: $M = \frac{mRT}{PV}$ (p. 437)
- Finding density: $D = \frac{MP}{RT}$ (p. 437)

Vocabulary

- Avogadro's principle (p. 430)
- Boyle's law (p. 421)
- Charles's law (p. 424)
- combined gas law (p. 428)
- Gay-Lussac's law (p. 426)
- ideal gas constant (R) (p. 434)
- ideal gas law (p. 434)
- molar volume (p. 431)

CHEMISTRY
Online

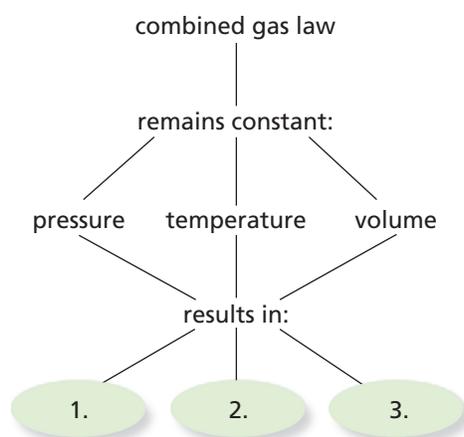


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Go to the Glencoe Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 14 Assessment.

Concept Mapping

- 70.** Complete the following concept map that shows how Boyle's law, Charles's law, and Gay-Lussac's law are derived from the combined gas law.



Mastering Concepts

- 71.** State the laws of Boyle, Charles, and Gay-Lussac as equations. (14.1)
- 72.** What will happen to the pressure of a contained gas if its temperature is lowered? (14.1)
- 73.** Why is it important not to puncture an aerosol can? (14.1)
- 74.** Explain why an unopened bag of potato chips left in a hot car appears to become larger. (14.1)
- 75.** If two variables have an inverse relationship, what happens to the value of one as the value of the other increases? (14.1)
- 76.** If two variables have a direct relationship, what happens to the value of one as the value of the other is increased? (14.1)
- 77.** Label the following examples as being generally representative of direct or inverse relationship. (14.1)
- popularity of a musical group versus how hard it is to get tickets for its concert
 - number of carats a diamond weighs versus its cost
 - number of people helping versus how long it takes to clean up after a party

- 78.** Label the following examples as being representative of a direct or inverse relationship. (14.1)
- pressure versus volume of a gas
 - volume versus temperature of a gas
 - pressure versus temperature of a gas
- 79.** What four variables are used to describe gases? (14.1)
- 80.** List the standard conditions for gas measurements. (14.2)
- 81.** Write the equation for the combined gas law. Identify the units most commonly used with each variable. (14.2)
- 82.** State Avogadro's principle. (14.2)
- 83.** What volume is occupied by one mole of a gas at STP? What volume do two moles occupy at STP? (14.2)
- 84.** What units must be used to express the temperature in the equation for the ideal gas law? Explain. (14.3)
- 85.** List two conditions under which a gas is least likely to behave ideally. (14.3)
- 86.** Write the value and units for the gas constant R in two common forms. (14.3)
- 87.** What information is needed to solve a volume-mass problem that involves gases? (14.4)

Mastering Problems

The Gas Laws (14.1)

- 88.** Use Boyle's, Charles's, or Gay-Lussac's law to calculate the missing value in each of the following.
- $V_1 = 2.0 \text{ L}$, $P_1 = 0.82 \text{ atm}$, $V_2 = 1.0 \text{ L}$, $P_2 = ?$
 - $V_1 = 250 \text{ mL}$, $T_1 = ?$, $V_2 = 400 \text{ mL}$, $T_2 = 298 \text{ K}$
 - $V_1 = 0.55 \text{ L}$, $P_1 = 740 \text{ mm Hg}$, $V_2 = 0.80 \text{ L}$, $P_2 = ?$
 - $T_1 = 25^\circ\text{C}$, $P_1 = ?$, $T_2 = 37^\circ\text{C}$, $P_2 = 1.0 \text{ atm}$
- 89.** What is the pressure of a fixed volume of a gas at 30.0°C if it has a pressure of 1.11 atm at 15.0°C ?
- 90.** A fixed amount of oxygen gas is held in a 1.00-L tank at a pressure of 3.50 atm . The tank is connected to an empty 2.00-L tank by a tube with a valve. After this valve has been opened and the oxygen is allowed to flow freely between the two tanks at a constant temperature, what is the final pressure in the system?
- 91.** Hot-air balloons rise because the hot air inside the balloon is less dense than the cooler air outside. Calculate the volume an air sample will occupy inside a balloon at 43.0°C if it occupies 2.50 L at the outside air temperature of 22.0°C , assuming the pressure is the same at both locations.

The Combined Gas Law (14.2)

- 92.** A sample of nitrogen gas is stored in a 500.0-mL flask at 108 kPa and 10.0°C. The gas is transferred to a 750.0-mL flask at 21.0°C. What is the pressure of nitrogen in the second flask?
- 93.** The air in a dry, sealed 2-L soda bottle has a pressure of 0.998 atm at sea level at a temperature of 34.0°C. What will be its pressure if it is brought to a higher altitude where the temperature is only 23.0°C?
- 94.** A weather balloon is filled with helium that occupies a volume of 5.00×10^4 L at 0.995 atm and 32.0°C. After it is released, it rises to a location where the pressure is 0.720 atm and the temperature is -12.0°C . What is the volume of the balloon at that new location?

Avogadro's Principle (14.2)

- 95.** Propane, C_3H_8 , is a gas commonly used as a home fuel for cooking and heating.
- Calculate the volume that 0.540 mol of propane occupies at STP.
 - Think about the size of this volume compared to the amount of propane that it contains. Why do you think propane is usually liquefied before it is transported?
- 96.** Carbon monoxide, CO, is a product of incomplete combustion of fuels. Find the volume that 42 g of carbon monoxide gas occupies at STP.

The Ideal Gas Law (14.3)

- 97.** The lowest pressure achieved in a laboratory is about 1.0×10^{-15} mm Hg. How many molecules of gas are present in a 1.00-L sample at that pressure and a temperature of 22.0°C?
- 98.** Determine the density of chlorine gas at 22.0°C and 1.00 atm pressure.
- 99.** Geraniol is a compound found in rose oil that is used in perfumes. What is the molar mass of geraniol if its vapor has a density of $0.480 \frac{\text{g}}{\text{L}}$ at a temperature of 260.0°C and a pressure of 0.140 atm?
- 100.** A 2.00-L flask is filled with propane gas (C_3H_8) at 1.00 atm and -15.0°C . What is the mass of the propane in the flask?

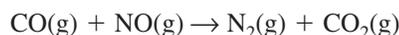
Gas Stoichiometry (14.4)

- 101.** Ammonia is formed industrially by reacting nitrogen and hydrogen gases. How many liters of ammonia gas can be formed from 13.7 L of hydrogen gas at 93.0°C and a pressure of 40.0 kPa?

- 102.** When 3.00 L of propane gas is completely combusted to form water vapor and carbon dioxide at a temperature of 350°C and a pressure of 0.990 atm, what mass of water vapor will result?

- 103.** When heated, solid potassium chlorate (KClO_3) decomposes to form solid potassium chloride and oxygen gas. If 20.8 g of potassium chlorate decomposes, how many liters of oxygen gas will form at STP?

- 104.** Use the reaction shown below to answer these questions.



- Balance the equation.
- What is the volume ratio of carbon monoxide to carbon dioxide in the balanced equation?
- If 42.7 g CO is reacted completely at STP, what volume of N_2 gas will be produced?

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 105.** Gaseous methane (CH_4) undergoes complete combustion by reacting with oxygen gas to form carbon dioxide and water vapor.
- Write a balanced equation for this reaction.
 - What is the volume ratio of methane to water in this reaction?
- 106.** If 2.33 L of propane at 24°C and 67.2 kPa is completely burned in excess oxygen, how many moles of carbon dioxide will be produced?
- 107.** Use Boyle's, Charles's, or Gay-Lussac's law to calculate the missing value in each of the following.
- $V_1 = 1.4$ L, $P_1 = ?$, $V_2 = 3.0$ L, $P_2 = 1.2$ atm
 - $V_1 = 705$ mL, $T_1 = 273$ K, $V_2 = ?$, $T_2 = 323$ K
 - $V_1 = 0.540$ L, $P_1 = ?$, $V_2 = 0.990$ L, $P_2 = 775$ mm Hg
 - $T_1 = 37^\circ\text{C}$, $P_1 = 5.0$ atm, $P_2 = 2.5$ atm, $T_2 = ?$
- 108.** Determine the pressure inside a television picture tube with a volume of 3.50 L that contains 2.00×10^{-5} g of nitrogen gas at 22.0°C.
- 109.** Determine how many liters 8.80 g of carbon dioxide gas would occupy at:
- STP
 - 160°C and 3.00 atm
 - 288 K and 118 kPa
- 110.** If 5.00 L of hydrogen gas, measured at 20.0°C and 80.1 kPa is burned in excess oxygen to form water, what mass of oxygen (measured at the same temperature and pressure) will be consumed?

Thinking Critically

- 111. Making and Using Graphs** Automobile tires become underinflated as temperatures drop during the winter months if no additional air is added to the tires at the start of the cold season. For every 10°F drop in temperature, the air pressure in a car's tires goes down by about 1 psi (14.7 psi equals 1.00 atm). Complete the following table. Then make a graph illustrating how the air pressure in a tire changes over the temperature range from 40°F to -10°F , assuming you start with a pressure of 30.0 psi at 40°F .

Tire Inflation Based on Temperature	
Temperature ($^{\circ}\text{F}$)	Pressure (psi)
40	
30	
20	
10	
0	
-10	

- 112. Applying Concepts** When nitroglycerin ($\text{C}_3\text{H}_5\text{N}_3\text{O}_9$) explodes, it decomposes into the following gases: CO_2 , N_2 , NO , and H_2O . If 239 g of nitroglycerin explodes, what volume will the mixture of gaseous products occupy at 1.00 atm pressure and 2678°C ?
- 113. Analyze and Conclude** What is the numerical value of the ideal gas constant (R) in $\frac{\text{cm}^3 \cdot \text{Pa}}{\text{K} \cdot \text{mol}}$?
- 114. Applying Concepts** Calculate the pressure of a mixture of two gases that contains 4.67×10^{22} molecules CO and 2.87×10^{24} molecules of N_2 in a 6.00-L container at 34.8°C .

Writing in Chemistry

- 115.** It was the dream of many early balloonists to complete a trip around the world in a hot-air balloon, a goal not achieved until 1999. Write about what you imagine a trip in a balloon would be like, including a description of how manipulating air temperature would allow you to control altitude.
- 116.** Investigate and explain the function of the regulators on the air tanks used by scuba divers.



Cumulative Review

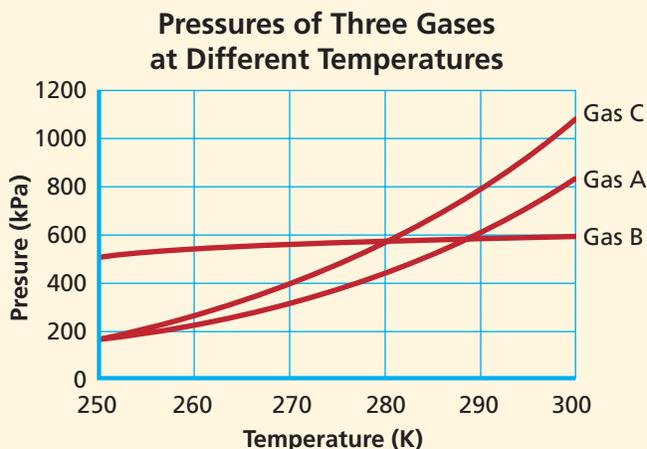
Refresh your understanding of previous chapters by answering the following.

- 117.** Convert each of the following mass measurements to its equivalent in kilograms. (Chapter 2)
- 247 g
 - 53 Mg
 - $7.23 \mu\text{g}$
 - 975 mg
- 118.** How many atoms of each element are present in five formula units of calcium permanganate? (Chapter 8)
- 119.** Terephthalic acid is an organic compound used in the formation of polyesters. It contains 57.8 percent C, 3.64 percent H, and 38.5 percent O. The molar mass is known to be approximately 166 g/mol. What is the molecular formula of terephthalic acid? (Chapter 11)
- 120.** The particles of which of the following gases have the highest average speed? The lowest average speed? (Chapter 13)
- carbon monoxide at 90°C
 - nitrogen trifluoride at 30°C
 - methane at 90°C
 - carbon monoxide at 30°C

Use these questions and the test-taking tip to prepare for your standardized test.

- The kinetic-molecular theory describes the microscopic behavior of gases. One main point of the theory is that within a sample of gas, the frequency of collisions between individual gas particles and between the particles and the walls of their container increases if the sample is compressed. The gas law that states this relationship in mathematical terms is _____.
 - Gay-Lussac's Law
 - Charles's Law
 - Boyle's Law
 - Avogadro's Law
- Three 2.0-L containers are placed in a 50°C room. Samples of 0.5 mol N₂, 0.5 mol Xe, and 0.5 mol ethene (C₂H₄) are pumped into Containers 1, 2, and 3, respectively. Inside which container will the pressure, be greatest?
 - Container 2
 - Container 3
 - Containers 2 and 3 have the same, higher pressure
 - Containers 1, 2, and 3 have equal pressures

Interpreting Graphs Use the graph to answer questions 3–5.



- It can be seen from the graph that _____.
 - as temperature increases, pressure decreases
 - as pressure increases, volume decreases
 - as temperature decreases, moles decrease
 - as pressure decreases, temperature decreases
- Which of these gases is an ideal gas?
 - Gas A
 - Gas B
 - Gas C
 - none of the above

- The predicted pressure of Gas B at 310 K is _____.
 - 260 kPa
 - 620 kPa
 - 1000 kPa
 - 1200 kPa
- What volume will 0.875 moles of SF₄ occupy at STP?
 - 19.6 L
 - 21.4 L
 - 22.4 L
 - 32.7 L
- While it is on the ground, a blimp is filled with 5.66×10^6 L of He gas. The pressure inside the grounded blimp, where the temperature is 25°C, is 1.10 atm. Modern blimps are non-rigid, which means that their volume is changeable. If the pressure inside the blimp remains the same, what will be the volume of the blimp at a height of 2300 m, where the temperature is 12°C?
 - 5.66×10^6 L
 - 2.72×10^6 L
 - 5.4×10^6 L
 - 5.92×10^6 L
- The reaction that provides blowtorches with their intense flame is the combustion of acetylene (C₂H₂) to form carbon dioxide and water vapor. Assuming that the pressure and temperature of the reactants are the same, what volume of oxygen gas is required to completely burn 5.60 L of acetylene?
 - 2.24 L
 - 5.60 L
 - 11.2 L
 - 14.0 L
- A sample of argon gas is compressed into a volume of 0.712 L by a piston exerting 3.92 atm of pressure. The piston is slowly released until the pressure of the gas is 1.50 atm. The new volume of the gas is _____.
 - 0.272 L
 - 3.67 L
 - 1.86 L
 - 4.19 L
- Assuming ideal behavior, how much pressure will 0.0468 g of ammonia (NH₃) gas exert on the walls of a 4.00-L container at 35.0°C?
 - 0.0174 atm
 - 0.296 atm
 - 0.00198 atm
 - 0.278 atm



TEST-TAKING TIP

Ask Questions If you've got a question about what will be on the test, the way the test is scored, the time limits placed on each section, or anything else. . . by all means ask! Will you be required to know the specific names of the gas laws, such as Boyle's law and Charles's law?

Solutions

What You'll Learn

- ▶ You will describe and categorize solutions.
- ▶ You will calculate concentrations of solutions.
- ▶ You will analyze the colligative properties of solutions.
- ▶ You will compare and contrast heterogeneous mixtures.

Why It's Important

The air you breathe, the fluids in your body, and some of the foods you ingest are solutions. Because solutions are so common, learning about their behavior is fundamental to understanding chemistry.

CHEMISTRY
Online



CLICK HERE

Visit the Chemistry Web site at science.glencoe.com to find links about solutions.

Though it isn't apparent, there are at least three different solutions in this photo; the air, the water in the harbor, and the steel used in the construction of the bridge are all solutions.



DISCOVERY LAB



Materials

balance
50-mL graduated cylinder
100-mL beaker (2)
stirring rod
ammonium chloride (NH_4Cl)
calcium chloride (CaCl_2)
water

Solution Formation

The intermolecular forces among dissolving particles and the attractive forces between solute and solvent particles result in an overall energy change. Can this change be observed?

Safety Precautions



Dispose of solutions by flushing them down a drain with excess water.

Procedure

1. Measure 10 g of ammonium chloride (NH_4Cl) and place it in a 100-mL beaker.
2. Add 30 mL of water to the NH_4Cl , stirring with your stirring rod.
3. Feel the bottom of the beaker and record your observations.
4. Repeat the procedure with calcium chloride (CaCl_2).

Analysis

Which dissolving process is exothermic? Endothermic? Suggest some practical applications for dissolving processes that are exothermic and for those that are endothermic.

Section

15.1

What are solutions?

Objectives

- Describe the characteristics of solutions and identify the various types.
- Relate intermolecular forces and the process of solvation.
- Define solubility and identify factors affecting it.

Vocabulary

soluble
insoluble
immiscible
miscible
solvation
heat of solution
solubility
saturated solution
unsaturated solution
supersaturated solution
Henry's law

Have you ever thought of the importance of solutions? Even if you haven't, the fact is that solutions are all around you. They are even inside you; you and all other organisms are composed of cells containing solutions that support life. You inhale a solution when you breathe. You are immersed in a solution whether standing in a room or swimming in a pool. Structures such as the ones shown in the photo on the previous page would not be possible without steel, yet another solution.

Characteristics of Solutions

Cell solutions, ocean water, and steel may appear quite dissimilar, but they share certain characteristics. In Chapter 3, you learned that solutions are homogeneous mixtures containing two or more substances called the solute and the solvent. The solute is the substance that dissolves. The solvent is the dissolving medium. When you look at a solution, it is not possible to distinguish the solute from the solvent.

A solution may exist as a gas, liquid, or solid depending on the state of its solvent, as shown in **Figure 15-1** and **Table 15-1** on the next page. Air is a gaseous solution, and its solvent is nitrogen gas; braces may be made of nitinol, a solid solution of titanium in nickel. Most solutions, however, are liquids. You learned in Chapter 10 that reactions can take place in aqueous solutions, that is, solutions in which reactants and products are mixed in water. In fact, water is the most common solvent among liquid solutions.

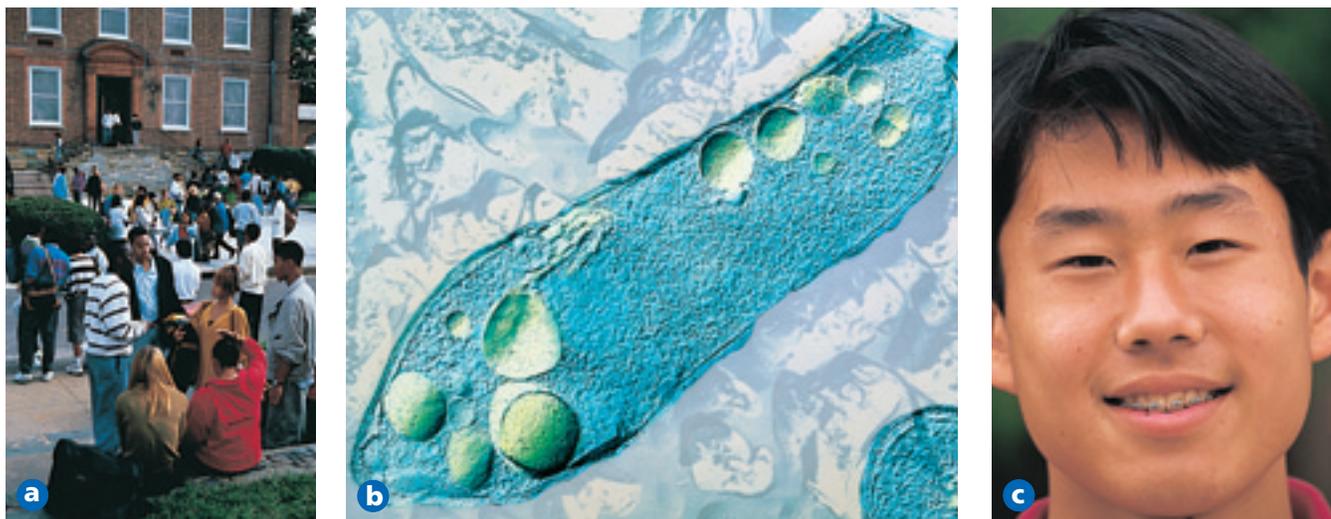


Figure 15-1

- a** The air you breathe is a gas solution primarily containing oxygen, nitrogen, and argon.
- b** The biological reactions necessary for life occur in aqueous solutions within cells. **c** A solid solution of titanium and nickel is commonly used for braces in orthodontia.

Table 15-1 lists examples of different types of solutions. Note that the solutes in the solutions may be gases, liquids, or solids. Which solutions contain gaseous solutes? Which solutions are aqueous? Note also that solutions such as ocean water can contain more than one solute.

Some combinations of substances readily form solutions and others do not. A substance that dissolves in a solvent is said to be **soluble** in that solvent. For example, sugar is soluble in water, a fact you probably learned by dissolving sugar in flavored water to make a sweetened beverage such as tea, lemonade, or fruit punch. A substance that does not dissolve in a solvent is said to be **insoluble** in that solvent. Sand is insoluble in water. Have you ever shaken a bottle of oil and vinegar when making salad dressing? If so, what happens to the liquids shortly after you stop mixing them? You are correct if you answered that they separate, or cease to mix. Oil is insoluble in vinegar; and thus, oil and vinegar are said to be **immiscible**. Two liquids that are soluble in each other, such as those that form the antifreeze listed in **Table 15-1**, are said to be **miscible**. Are water and acetic acid miscible?

Table 15-1

Types and Examples of Solutions			
Type of solution	Example	Solvent	Solute
Gas			
Gas in gas	Air	Nitrogen (gas)	Oxygen (gas)
Liquid			
Gas in liquid	Carbonated water	Water (liquid)	Carbon dioxide (gas)
Gas in liquid	Ocean water	Water (liquid)	Oxygen gas (gas)
Liquid in liquid	Antifreeze	Water (liquid)	Ethylene glycol (liquid)
Liquid in liquid	Vinegar	Water (liquid)	Acetic acid (liquid)
Solid in liquid	Ocean water	Water (liquid)	Sodium chloride (solid)
Solid			
Liquid in solid	Dental amalgam	Silver (solid)	Mercury (liquid)
Solid in solid	Steel	Iron (solid)	Carbon (solid)



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.



Solvation in Aqueous Solutions

Why are some substances soluble in one another whereas others are not? To form a solution, solute particles must separate from one another and the solute and solvent particles must mix. Recall from Chapter 13 that attractive forces exist among the particles of all substances. Attractive forces exist between the pure solute particles, between the pure solvent particles, and between the solute and solvent particles. When a solid solute is placed in a solvent, the solvent particles completely surround the surface of the solid solute. If the attractive forces between the solvent and solute particles are greater than the attractive forces holding the solute particles together, the solvent particles pull the solute particles apart and surround them. These surrounded solute particles then move away from the solid solute, out into the solution. The process of surrounding solute particles with solvent particles to form a solution is called **solvation**. Solvation in water is called hydration.

“Like dissolves like” is the general rule used to determine whether solvation will occur in a specific solvent. To determine whether a solvent and solute are alike, you must examine the bonding and the polarity of the particles and the intermolecular forces between particles.

Aqueous solutions of ionic compounds Examine **Figure 15-2** to review the polar nature of water. Recall that water molecules are dipoles with partially positive and partially negative ends. Water molecules are in constant motion as described by the kinetic-molecular theory. When a crystal of an ionic compound, such as sodium chloride (NaCl), is placed in a beaker of water, the water molecules collide with the surface of the crystal. The charged ends of the water molecules attract the positive sodium ions and negative chloride ions. This attraction between the dipoles and the ions is greater than the attraction among the ions in the crystal, and so the ions break away from the surface. The water molecules surround the ions and the solvated ions move into solution, as shown in **Figure 15-3**. This exposes more ions on the surface of the crystal. Solvation continues until the entire crystal has dissolved and all ions are distributed throughout the solvent.

Gypsum is a compound composed of calcium ions and sulfate ions. It is mixed with water to make plaster. Plaster is a mixture, not a solution. Gypsum is insoluble in water because the attractive forces among the ions in calcium sulfate are so strong that they cannot be overcome by the attractive forces exerted by the water molecules. Solvation does not occur.

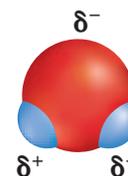


Figure 15-2

The bent shape of a water molecule results in dipoles that do not cancel each other out. The molecule has a net polarity, with the oxygen end being partially negative and the hydrogen ends being partially positive. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

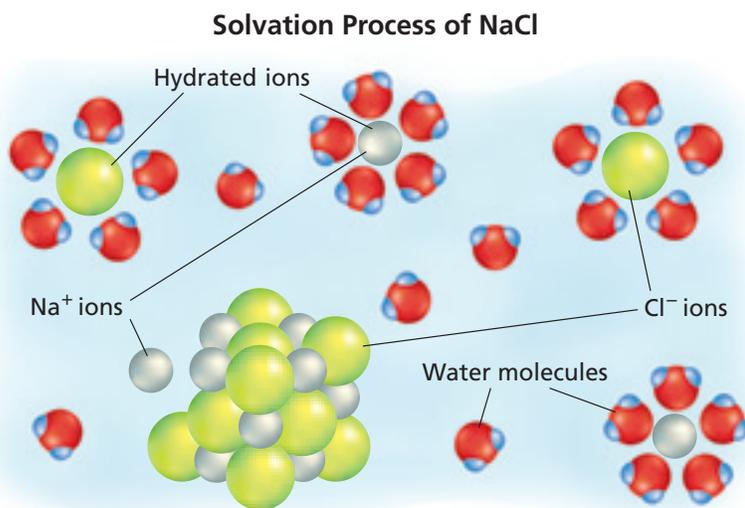


Figure 15-3

Solid sodium chloride dissolves as its ions are surrounded by solvent water molecules. Note how the polar water molecules orient themselves differently around the positive and negative ions.

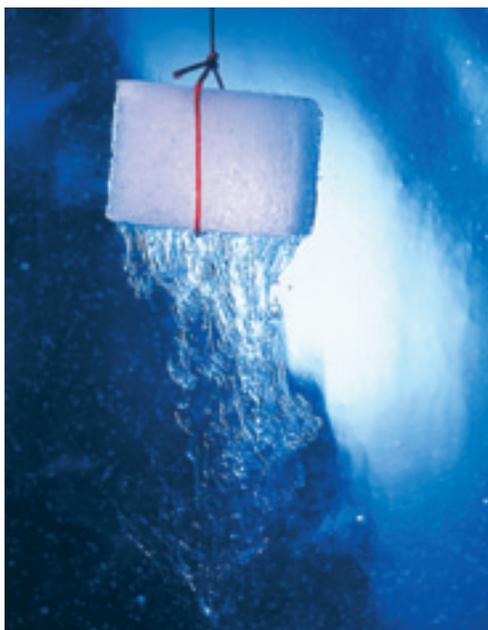
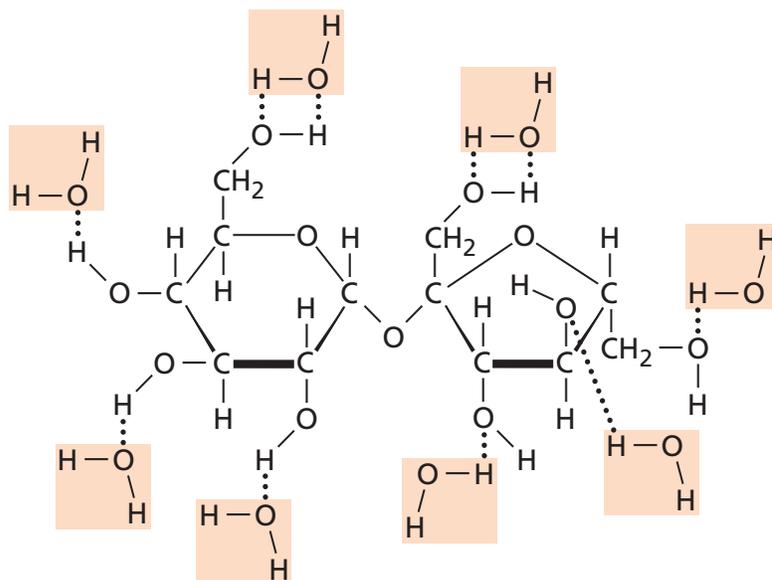


Figure 15-4

The solvation of a sugar cube (sucrose) can be seen in the above photo. Polar sucrose ($C_{12}H_{22}O_{11}$) molecules contain eight O–H bonds. Forces between polar water molecules and polar sucrose molecules break these O–H bonds and the sucrose dissolves.



Aqueous solutions of molecular compounds Water also is a good solvent for many molecular compounds. You know that table sugar dissolves in water. Table sugar is the molecular compound sucrose. See **Figure 15-4**. Note that its structure has a number of O–H bonds. Sucrose molecules are polar. When water is added to sucrose, each O–H bond becomes a site for hydrogen bonding with water. As soon as the sugar crystals contact the water, water molecules collide with the outer surface of the crystal. The attractive forces among sucrose molecules are overcome by the attractive forces between polar water molecules and polar sucrose molecules. Sugar molecules leave the crystal and become solvated by water molecules.

The oil in **Figure 15-5** is a substance made up of primarily carbon and hydrogen. It does not form a solution with water. Why are oil and water immiscible? There is little attraction between the polar water molecules and the nonpolar oil molecules. However, oil spills can be cleaned up with a nonpolar solvent. Nonpolar solutes are more readily dissolved in nonpolar solvents.



Figure 15-5

Nonpolar oil molecules do not mix with polar water molecules. Because oil is less dense than water, it floats on the water's surface. For this reason, oil spills at sea often wash ashore.

Factors that affect rate of solvation As you have just learned, solvation occurs only when and where the solute and solvent particles come in contact with each other. There are three common ways to increase the collisions between solute and solvent particles, and thus increase the rate at which the solute dissolves: agitating the mixture, increasing the surface area of the solute, and increasing the temperature of the solvent.

Agitating the mixture by stirring and shaking moves dissolved solute particles away from the contact surfaces more quickly and thereby allows new collisions between solute and solvent particles to occur. Without stirring or shaking, solvated particles move away from the contact areas slowly. Breaking the solute into small pieces increases its surface area. A greater surface area allows more collisions to occur. This is why a teaspoon of granulated sugar dissolves more quickly than an equal amount of sugar in a cube. Raising the temperature of the solvent increases the kinetic energy of its particles, resulting in more frequent collisions and collisions with greater energy than those that occur at lower temperatures. Can you think of an example of how an increase in temperature affects the rate of dissolving?

CONNECTION

Water is the most abundant solvent on Earth. The amount of oxygen contained in water is called dissolved oxygen (DO). Dissolved oxygen is an indicator of water quality because oxygen is necessary for fish and other aquatic life. Oxygen enters a body of water by photosynthesis of aquatic plants and by transfer of oxygen across the air-water boundary.

Thermal pollution is a reduction in water quality due to an increase in water temperature. The discharge of heated water into a lake, river, or other body of water by factories or power plants causes the solubility of oxygen in water to decrease. In addition to not having enough oxygen to support life, as shown below, the lower oxygen level magnifies the effects of toxic and organic pollutants.

To reduce thermal pollution, measures must be taken before the heated water is released into the environment. The water can be discharged into holding lakes or canals and allowed to cool. In addition, many facilities use cooling towers to dissipate heat into the air.



Heat of solution During the process of solvation, the solute must separate into particles. Solvent particles also must move apart in order to allow solute particles to come between them. Energy is required to overcome the attractive forces within the solute and within the solvent, so both steps are endothermic. When solute and solvent particles mix, the particles attract each other and energy is released. This step in the solvation process is exothermic. The overall energy change that occurs during the solution formation process is called the **heat of solution**.

As you observed in the **DISCOVERY LAB** at the beginning of this chapter, some solutions release energy as they form, whereas others absorb energy during formation. For example, after ammonium nitrate dissolves in water, its container feels cool. In contrast, after calcium chloride dissolves in water, its container feels warm. You will learn more about the heat of solution in the next chapter.

Solubility

If you have ever added so much sugar to a sweetened beverage that sugar crystals accumulated on the container's bottom, then you know that only a limited amount of solute can dissolve in a solvent at a given set of conditions. In fact, every solute has a characteristic solubility. **Solubility** refers to the maximum amount of solute that will dissolve in a given amount of solvent at a specified temperature and pressure. As you can see from **Table 15-2**, solubility is usually expressed in grams of solute per 100 g of solvent.

Table 15-2

Solubilities of Some Solutes in Water at Various Temperatures					
Substance	Formula	Solubility (g/100 g H ₂ O)*			
		0°C	20°C	60°C	100°C
Aluminum sulfate	Al ₂ (SO ₄) ₃	31.2	36.4	59.2	89.0
Ammonium chloride	NH ₄ Cl	29.4	37.2	55.3	77.3
Barium hydroxide	Ba(OH) ₂	1.67	3.89	20.94	—
Barium nitrate	Ba(NO ₃) ₂	4.95	9.02	20.4	34.4
Calcium hydroxide	Ca(OH) ₂	0.189	0.173	0.121	0.076
Lead(II) chloride	PbCl ₂	0.67	1.00	1.94	3.20
Lithium sulfate	Li ₂ SO ₄	36.1	34.8	32.6	—
Potassium chloride	KCl	28.0	34.2	45.8	56.3
Potassium sulfate	K ₂ SO ₄	7.4	11.1	18.2	24.1
Sodium chloride	NaCl	35.7	35.9	37.1	39.2
Silver nitrate	AgNO ₃	122	216	440	733
Sucrose	C ₁₂ H ₂₂ O ₁₁	179.2	203.9	287.3	487.2
Ammonia*	NH ₃	1130	680	200	—
Carbon dioxide*	CO ₂	1.713	0.878	0.359	—
Oxygen*	O ₂	0.048	0.031	0.019	—

* L/1 L H₂O of gas at standard pressure (101 kPa)

— No value available

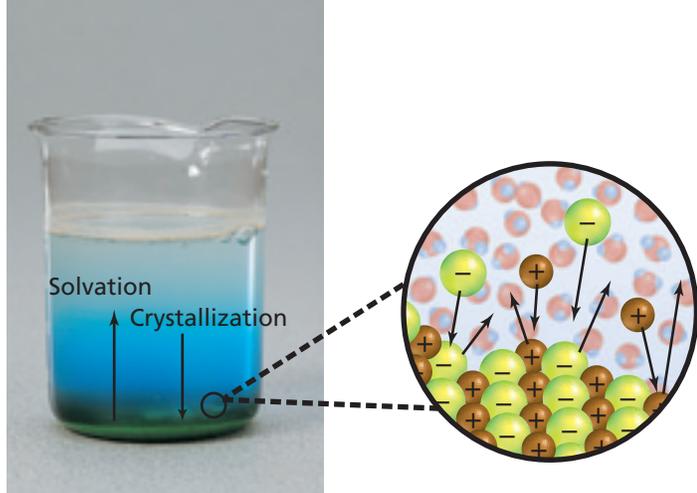


Figure 15-6

A dynamic equilibrium exists in a saturated solution. That is, the rate at which solute particles in the crystal are solvated is equal to the rate at which solvated solute particles rejoin the crystal.

Just as solvation can be understood at the particle level, so can solubility. When a solute is added to a solvent, solvent particles collide with the solute's surface particles; solute particles begin to mix randomly among the solvent particles. At first, the solute particles are carried away from the crystal. However, as the number of solvated particles increases, the same random mixing results in increasingly frequent collisions between solvated solute particles and the remaining crystal. Some colliding solute particles rejoin the crystal, or crystallize as you can see in **Figure 15-6**. As solvation continues, the crystallization rate increases while the solvation rate remains constant. As long as the solvation rate is greater than the crystallization rate, the net effect is continuing solvation.

Depending on the amount of solute present, the rates of solvation and crystallization may eventually equalize: no more solute appears to dissolve and a state of dynamic equilibrium exists between crystallization and solvation (as long as the temperature remains constant). Although solute particles continue to dissolve and crystallize in solutions that reach equilibrium, the overall amount of dissolved solute in the solution remains constant. Such a solution is said to be a **saturated solution**; it contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure. An **unsaturated solution** is one that contains less dissolved solute for a given temperature and pressure than a saturated solution. In other words, more solute can be dissolved in an unsaturated solution.

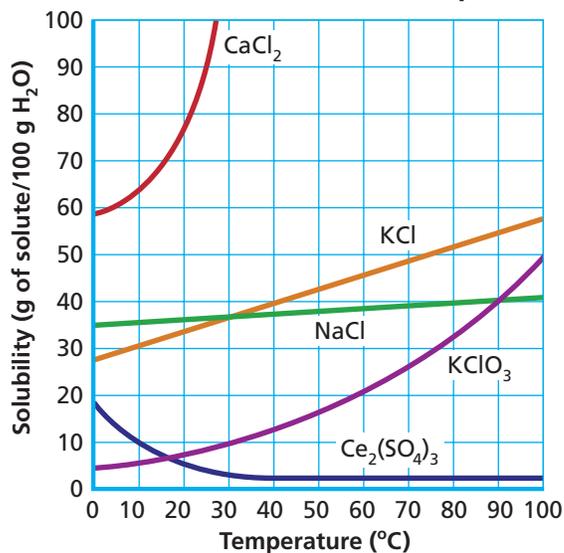
Factors That Affect Solubility

Pressure affects the solubility of gaseous solutes and gaseous solutions. The solubility of a solute also depends on the nature of the solute and solvent. Temperature affects the solubility of all substances. To learn how a blood disorder called sickle-cell disease can affect oxygen's solubility in blood, read the **Chemistry and Society** feature at the end of this chapter.

Figure 15-7

This graph shows the solubility of several substances as a function of temperature. What is the solubility of KClO_3 at 60°C ?

Solubilities as a Function of Temperature



Temperature and solubility Many substances are more soluble at high temperatures than at low temperatures, as you can see by the data graphed in **Figure 15-7**. For example, calcium chloride (CaCl_2) has a solubility of about 64 g CaCl_2 per 100 g H_2O at 10°C . Increasing the temperature to approximately 27°C increases the solubility by 50%, to 100 g CaCl_2 per 100 g H_2O . This fact also is illustrated by the data in **Table 15-2**, on the previous page. From the table you can see that at 20°C , 203.9 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolves in 100 g of water. At 100°C , 487.2 g of sucrose dissolves in 100 g of water, nearly a 140% increase in solubility over an 80°C temperature range. According to **Figure 15-7** and **Table 15-2**, what substances decrease in solubility as temperature increases?



Figure 15-8

Sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) is commonly used in the preparation of a supersaturated solution. When a seed crystal is added **(a)**, the excess sodium acetate quickly crystallizes out of the solution, **(b)** and **(c)**.

As you can see from **Table 15-2**, the gases oxygen and carbon dioxide are less soluble at higher temperatures than at lower temperatures. This is a predictable trend for all gaseous solutes in liquid solvents. Can you explain why? Recall from Chapter 13 that the kinetic energy of gas particles allows them to escape from a solution more readily at higher temperatures. Thus, as a solution's temperature increases, the solubility of a gaseous solute decreases.

The fact that solubility changes with temperature and that some substances become more soluble with increasing temperature, is the key to forming supersaturated solutions. A **supersaturated solution** contains more dissolved solute than a saturated solution at the same temperature. To make a supersaturated solution, a saturated solution is formed at a high temperature and then cooled slowly. The slow cooling allows the excess solute to remain dissolved in solution at the lower temperature, as shown in **Figure 15-8a**. Supersaturated solutions are unstable. If a tiny amount of solute, called a seed crystal, is added to a supersaturated solution, the excess solute precipitates quickly, as shown in **Figures 15-8b** and **15-8c**. Crystallization can also occur if the inside of the container is scratched or the supersaturated solution undergoes a physical shock such as stirring or tapping the container. Using crystals of silver iodide (AgI) to seed air supersaturated with water vapor causes the water particles to come together and form droplets that may fall to Earth as rain. This often-performed technique is called cloud seeding. The rock candy and mineral deposits at the edges of mineral springs, shown in **Figure 15-9**, are formed from supersaturated solutions.

Figure 15-9

Examples of crystals that formed from supersaturated solutions include rock candy and hot spring mineral deposits.





Figure 15-10

a When the cap on the soda bottle is closed, pressure above the solution keeps excess carbon dioxide (CO_2) from escaping the solution. **b** When the cap is removed, the decreased pressure above the solution results in the decreased solubility of the carbon dioxide—the carbon dioxide escapes the solution.

Pressure and solubility Pressure affects the solubility of gaseous solutes. The solubility of a gas in any solvent increases as its external pressure (the pressure above the solution) increases. Carbonated beverages depend on this fact. Carbonated beverages contain carbon dioxide gas dissolved in an aqueous solution. The dissolved gas gives the beverage its fizz. In bottling the beverage, carbon dioxide is dissolved in the solution at a pressure higher than atmospheric pressure. When the beverage container is opened, the pressure of the carbon dioxide gas in the space above the liquid (in the neck of the bottle) decreases. As a result, bubbles of carbon dioxide gas form in the solution, rise to the top, and escape. See **Figure 15-10**. Unless the cap is placed back on the bottle, the process will continue until the solution loses almost all of its carbon dioxide gas and goes flat.

Henry's law The decreased solubility of the carbon dioxide contained in the beverage after its cap is removed can be described by Henry's law. **Henry's law** states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid. You can express this relationship in the following way

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

where S_1 is the solubility of a gas at a pressure P_1 and S_2 is the solubility of the gas at the new pressure P_2 .

You often will solve Henry's law for the solubility S_2 at a new pressure P_2 , where P_2 is known. The basic rules of algebra can be used to solve Henry's law for any one specific variable. To solve for S_2 , begin with the standard form of Henry's law.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Cross-multiplying yields,

$$S_1P_2 = P_1S_2$$

Dividing both sides of the equation by P_1 yields the desired result, the equation solved for S_2 .

$$\frac{S_1P_2}{P_1} = \frac{\cancel{P_1}S_2}{\cancel{P_1}} \quad S_2 = \frac{S_1P_2}{P_1}$$

EXAMPLE PROBLEM 15-1

Using Henry's Law

If 0.85 g of a gas at 4.0 atm of pressure dissolves in 1.0 L of water at 25°C, how much will dissolve in 1.0 L of water at 1.0 atm of pressure and the same temperature?

1. Analyze the Problem

You are given the solubility of a gas at an initial pressure. The temperature of the gas remains constant as the pressure changes. Because decreasing pressure reduces a gas's solubility, less gas should dissolve at the lower pressure.

Known

$$S_1 = 0.85 \text{ g/L}$$

$$P_1 = 4.0 \text{ atm}$$

$$P_2 = 1.0 \text{ atm}$$

Unknown

$$S_2 = ? \text{ g/L}$$

2. Solve for the Unknown

Rearrange Henry's law to solve for S_2 .

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}, \quad S_2 = S_1 \left(\frac{P_2}{P_1} \right)$$

Substitute the known values into the equation and solve.

$$S_2 = (0.85 \text{ g/L}) \left(\frac{1.0 \text{ atm}}{4.0 \text{ atm}} \right) = 0.21 \text{ g/L}$$

3. Evaluate the Answer

The answer is correctly expressed to two significant figures. The solubility decreased as expected. The pressure on the solution was reduced from 4.0 atm to 1.0 atm, so the solubility should be reduced to one-fourth its original value, which it is.

PRACTICE PROBLEMS

1. If 0.55 g of a gas dissolves in 1.0 L of water at 20.0 kPa of pressure, how much will dissolve at 110.0 kPa of pressure?
2. A gas has a solubility of 0.66 g/L at 10.0 atm of pressure. What is the pressure on a 1.0-L sample that contains 1.5 g of gas?



For more practice with Henry's law problems, go to **Supplemental Practice Problems** in Appendix A.

Section 15.1 Assessment

3. Describe the characteristics of a solution and identify the various types.
4. How do intermolecular forces affect solvation?
5. What is solubility? Describe two factors that affect solubility.
6. **Thinking Critically** If a seed crystal was added to a supersaturated solution, how would you characterize the resulting solution?
7. **Making and Using Graphs** Use the information in  **Table 15-2** to graph the solubilities of aluminum sulfate, lithium sulfate, and potassium sulfate at 0°C, 20°C, 60°C, and 100°C. Which substance's solubility is most affected by increasing temperature?

Objectives

- **State** the concentrations of solutions in different ways.
- **Calculate** the concentrations of solutions.

Vocabulary

concentration
molarity
molality
mole fraction

You have learned about the process of solvation and the factors that affect solubility. The **concentration** of a solution is a measure of how much solute is dissolved in a specific amount of solvent or solution. How would you describe the concentration of the solutions in **Figure 15-11**? Concentration may be described qualitatively using the words *concentrated* or *dilute*. In general a concentrated solution, as shown on the left in **Figure 15-11**, contains a large amount of solute. Conversely, a dilute solution contains a small amount of solute. How do you know that the tea on the right in **Figure 15-11** is a more dilute solution than the tea on the left?

Expressing Concentration

Although qualitative descriptions of concentration can be useful, solutions are more often described quantitatively. Some commonly used quantitative descriptions are percent by either mass or volume, molarity, and molality. These descriptions express concentration as a ratio of measured amounts of solute and solvent or solution. **Table 15-3** lists each ratio's description.

Table 15-3

Concentration Ratios	
Concentration description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$

You may be wondering if one description is preferable to another. The description used depends on the type of solution analyzed and the reason for describing it. For example, a chemist working with a reaction in an aqueous solution most likely refers to the molarity of the solution, because he or she needs to know the number of particles involved in the reaction.

Figure 15-11

The strength of the tea corresponds to its concentration. The darker cup of tea is more concentrated than the lighter cup.



Using Percent to Describe Concentration

Concentration expressed as a percent is a ratio of a measured amount of solute to a measured amount of solution. Percent by mass usually describes solutions in which a solid is dissolved in a liquid, such as sodium chloride in water. The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution equals the sum of the masses of the solute and the solvent.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

EXAMPLE PROBLEM 15-2

Calculating Percent by Mass

In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

1. Analyze the Problem

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

Known

mass of solute = 3.6 g NaCl
mass of solvent = 100.0 g H₂O

Unknown

percent by mass = ?

2. Solve for the Unknown

Find the mass of the solution.

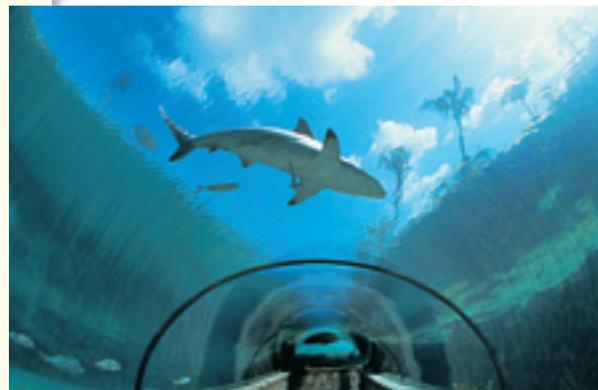
$$\begin{aligned}\text{Mass of solution} &= \text{grams of solute} + \text{grams of solvent} \\ &= 3.6 \text{ g} + 100.0 \text{ g} \\ &= 103.6 \text{ g}\end{aligned}$$

Substitute the known values into the percent by mass equation.

$$\begin{aligned}\text{Percent by mass} &= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100 \\ &= \frac{3.6 \text{ g}}{103.6 \text{ g}} \times 100 \\ &= 3.5\%\end{aligned}$$

3. Evaluate the Answer

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures, therefore, the answer also is expressed with two significant figures.



Maintaining the proper saline (salt) concentration is important to the health of saltwater fish.

PRACTICE PROBLEMS

- What is the percent by mass of NaHCO₃ in a solution containing 20 g NaHCO₃ dissolved in 600 mL H₂O?
- You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite, NaOCl, is 3.62%. How many grams of NaOCl are in the solution?
- In question 9, how many grams of solvent are in the solution?



For more practice with percent by mass problems, go to **Supplemental Practice Problems** in Appendix A.

Percent by volume usually describes solutions in which both solute and solvent are liquids. The percent by volume is the ratio of the volume of the solute to the volume of the solution expressed as a percent. The volume of the solution is the sum of the volumes of the solute and the solvent. Calculations are similar to those involving percent by mass.

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Rubbing alcohol is an aqueous solution of liquid isopropyl alcohol. The label on a typical container, such as the one shown in **Figure 15-12**, usually states that the rubbing alcohol is 70% isopropyl alcohol. This value is a percent by volume. It tells you that 70 volumes of isopropyl alcohol are dissolved in every 100 volumes of solution. Because a solution's volume is the sum of the volumes of solute and solvent, there must be 30 volumes of water (solvent) in every 100 volumes of the rubbing alcohol.



For more practice with percent by volume problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

11. What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 115 mL of water?
12. If you have 100.0 mL of a 30.0% aqueous solution of ethanol, what volumes of ethanol and water are in the solution?
13. What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?

Molarity

As you have learned, percent by volume and percent by mass are only two of the commonly used ways to quantitatively describe the concentrations of liquid solutions. One of the most common units of solution concentration is molarity. **Molarity** (M) is the number of moles of solute dissolved per liter of solution. Molarity also is known as molar concentration. The unit M is read as molar. A liter of solution containing one mole of solute is a $1M$ solution, which is read as a one molar solution. A liter of solution containing 0.1 mole of solute is a $0.1M$ solution.

To calculate a solution's molarity, you must know the volume of the solution and the amount of dissolved solute.

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

For example, suppose you need to calculate the molarity of 100.0 mL of an aqueous solution containing 0.085 mole of dissolved potassium chloride (KCl). You would first convert the volume of the solution from milliliters to liters using the conversion factor $1 \text{ L} = 1000 \text{ mL}$.

$$(100 \text{ mL}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1000 \text{ L}$$

Then, to determine the molarity, you would divide the number of moles of solute by the solution volume in liters.

$$\frac{0.085 \text{ mol KCl}}{0.1000 \text{ L solution}} = \frac{0.85 \text{ mol}}{\text{L}} = 0.85M$$



Figure 15-12

The composition of this isopropyl alcohol is given in percent by volume, which is often expressed as % (v/v). What does each v in the expression 70% (v/v) refer to?



Do the **CHEMLAB** at the end of this chapter to learn about an experimental technique for determining solution concentration.

EXAMPLE PROBLEM 15-3

Calculating Molarity

A 100.5-mL intravenous (IV) solution contains 5.10 g of glucose ($C_6H_{12}O_6$). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

1. Analyze the Problem

You are given the mass of glucose dissolved in a volume of solution. The molarity of the solution is the ratio of moles of solute per liter of solution. Glucose is the solute and water is the solvent.

Known

mass of solute = 5.10 g $C_6H_{12}O_6$
molar mass of $C_6H_{12}O_6$ = 180.16 g/mol
volume of solution = 100.5 mL

Unknown

solution concentration = ?M

2. Solve for the Unknown

Use the molar mass to calculate the number of moles of $C_6H_{12}O_6$.

$$(5.10 \text{ g } C_6H_{12}O_6) \left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6} \right) = 0.0283 \text{ mol } C_6H_{12}O_6$$

Use the conversion factor $\frac{1 \text{ L}}{1000 \text{ mL}}$ to convert the volume of H_2O in milliliters to liters.

$$(100.5 \text{ mL solution}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1005 \text{ L solution}$$

Substitute the known values into the equation for molarity and solve.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{molarity} = \frac{0.0283 \text{ mol } C_6H_{12}O_6}{0.1005 \text{ L solution}} = \frac{0.282 \text{ mol } C_6H_{12}O_6}{\text{L solution}} = 0.282M$$

3. Evaluate the Answer

The molarity is a small value, which is expected because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem contained three significant figures, and therefore, the value of the molarity also has three significant figures.



To prevent dehydration, intravenous (IV) drips are administered to many hospital patients. A solution containing sodium chloride and glucose is commonly used.

PRACTICE PROBLEMS

14. What is the molarity of an aqueous solution containing 40.0 g of glucose ($C_6H_{12}O_6$) in 1.5 L of solution?
15. What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?
16. Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.



For more practice with molarity problems, go to **Supplemental Practice Problems** in Appendix A.



Figure 15-13

Accurately preparing a solution takes care. **a** In step 1, the mass of solute to be used is measured out. **b** In step 2, the solute is placed in a volumetric flask of the correct volume. **c** In step 3, distilled water is added to the flask to bring the solution level up to the calibration mark on the flask.

Preparing Molar Solutions

Now that you know how to calculate the molarity of a solution, how do you think you would prepare 1 L of a 1.50M aqueous solution of sucrose ($C_{12}H_{22}O_{11}$) for an experiment? A 1.50M aqueous solution of sucrose contains 1.50 moles of sucrose dissolved in a liter of solution. The molar mass of sucrose is 342 g. Thus, 1.50 moles of sucrose has a mass of 513 g, an amount that you can measure on a balance.

$$\frac{1.50 \text{ mol } C_{12}H_{22}O_{11}}{1 \text{ L solution}} \times \frac{342 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ mol } C_{12}H_{22}O_{11}} = \frac{513 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ L solution}}$$

Unfortunately, you cannot simply add 513 g of sugar to one liter of water to make the 1.50M solution. Do you know why? Like all substances, sugar takes up space and will add volume to the solution. Therefore, you must use slightly less than one liter of water to make one liter of solution. Follow the steps shown in **Figure 15-13** to learn how to prepare the correct volume of the solution.

You often will do experiments that call for only small quantities of solution. For example, you may need only 100 mL of a 1.50M sucrose solution for an experiment. How do you determine the amount of sucrose to use? Look again at the definition of molarity. As calculated above, 1.50M solution of sucrose contains 1.50 mol of sucrose per one liter of solution. Therefore, one liter of solution contains 513 g of sucrose.

This relationship can be used as a conversion factor to calculate how much solute you need for your experiment.

$$100 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{513 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ L solution}} = 51.3 \text{ g } C_{12}H_{22}O_{11}$$

Thus, you would need to measure out 51.3 g of sucrose to make 100 mL of a 1.50M solution.

PRACTICE PROBLEMS

17. How many grams of $CaCl_2$ would be dissolved in 1.0 L of a 0.10M solution of $CaCl_2$?
18. A liter of 2M NaOH solution contains how many grams of NaOH?
19. How many grams of $CaCl_2$ should be dissolved in 500.0 mL of water to make a 0.20M solution of $CaCl_2$?
20. How many grams of NaOH are in 250 mL of a 3.0M NaOH solution?

Practice!

For more practice with molarity problems, go to **Supplemental Practice Problems** in Appendix A.

Diluting solutions In the laboratory, you may use concentrated solutions of standard molarities called stock solutions. For example, concentrated hydrochloric acid (HCl) is $12M$. Recall that a concentrated solution has a large amount of solute. You can prepare a less concentrated solution by diluting the stock solution with solvent. When you add solvent, you increase the number of solvent particles among which the solute particles move, as shown in **Figure 15-14**, thereby decreasing the solution's concentration. Would you still have the same number of moles of solute particles that were in the stock solution? Why?

How do you determine the volume of stock solution you must dilute? You know that

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

You can rearrange the expression of molarity to solve for moles of solute.

$$\text{Moles of solute} = \text{molarity} \times \text{liters of solution}$$

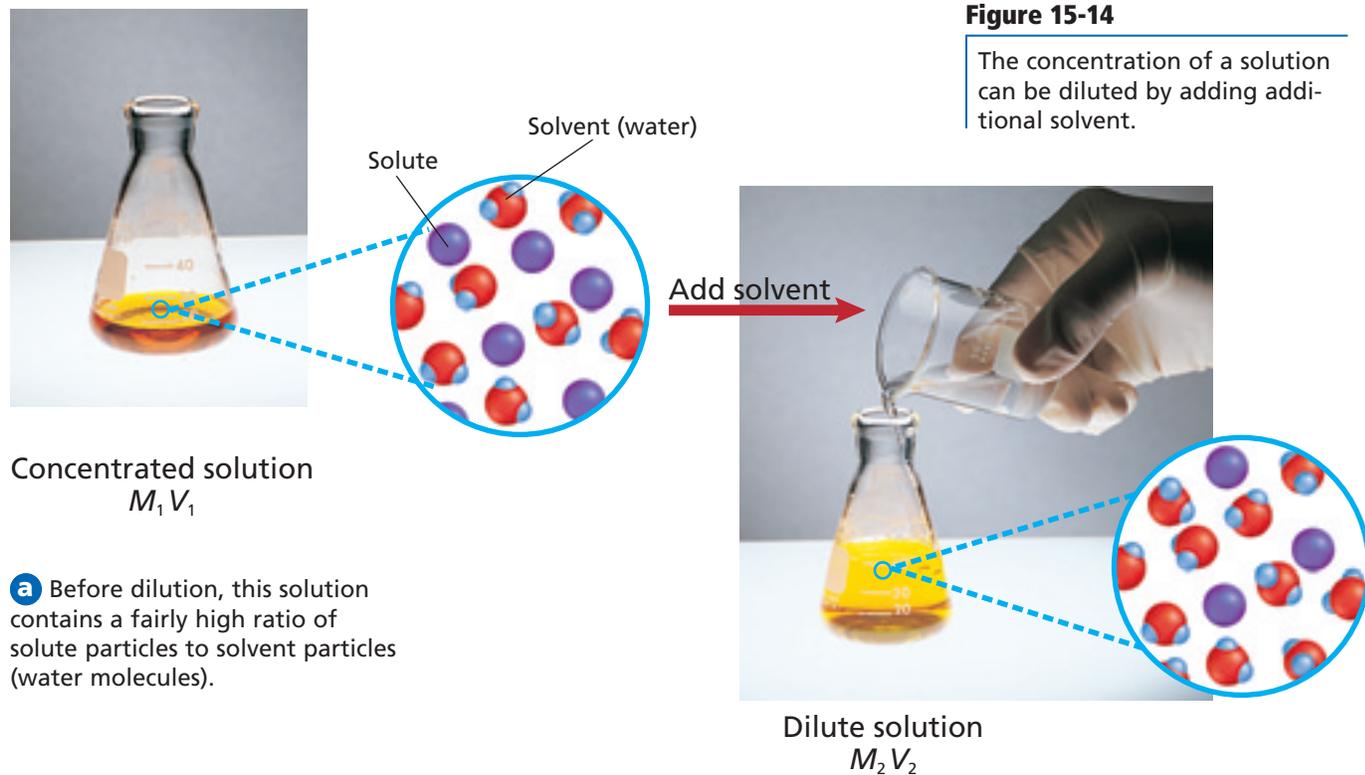
Because the total number of moles of solute does not change during dilution,

$$\text{Moles of solute in the stock solution} = \text{moles of solute after dilution}$$

You can write this relationship as the expression

$$M_1V_1 = M_2V_2$$

where M_1 and V_1 represent the molarity and volume of the stock solution and M_2 and V_2 represent the molarity and volume of the dilute solution.



EXAMPLE PROBLEM 15-4

Diluting Stock Solutions

What volume, in milliliters of 2.00M calcium chloride (CaCl_2) stock solution would you use to make 0.50 L of 0.300M calcium chloride solution?

1. Analyze the Problem

You are given the molarity of a stock solution of CaCl_2 and the volume and molarity of a dilute solution of CaCl_2 . Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then convert the volume to milliliters.

Known

$$\begin{aligned}M_1 &= 2.00M \text{ CaCl}_2 \\M_2 &= 0.300M \\V_2 &= 0.50 \text{ L}\end{aligned}$$

Unknown

$$V_1 = ? \text{ L } 2.00M \text{ CaCl}_2$$



Knowing how to apply the equation $M_1V_1 = M_2V_2$ makes the preparation of dilute solutions easy.

2. Solve for the Unknown

Solve the molarity–volume relationship for the volume of the stock solution, V_1 .

$$M_1V_1 = M_2V_2$$

Dividing both sides of the equation yields,

$$V_1 = V_2 \left(\frac{M_2}{M_1} \right)$$

Substitute the known values into the equation and solve.

$$V_1 = (0.50 \text{ L}) \left(\frac{0.300M}{2.000M} \right)$$

$$V_1 = 0.075 \text{ L}$$

Use the conversion factor $\frac{1000 \text{ mL}}{1 \text{ L}}$ to convert the volume from liters to milliliters.

$$V_1 = (0.075 \text{ L}) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 75 \text{ mL}$$

To make the dilution, measure out 75 mL of the stock solution and dilute it with enough water to make the final volume 0.50 L.

3. Evaluate the Answer

The volume V_1 was calculated and then its value was converted to milliliters. Of the given information, V_2 had the fewest number of significant figures with two. Thus, the volume V_1 should also have two significant figures, as it does.

PRACTICE PROBLEMS

21. What volume of a 3.00M KI stock solution would you use to make 0.300 L of a 1.25M KI solution?
22. How many milliliters of a 5.0M H_2SO_4 stock solution would you need to prepare 100.0 mL of 0.25M H_2SO_4 ?
23. If you dilute 20.0 mL of a 3.5M solution to make 100.0 mL of solution, what is the molarity of the dilute solution?

Practice!

For more practice with dilution problems, go to **Supplemental Practice Problems** in Appendix A.

Molality and Mole Fraction

The volume of a solution changes with temperature as it expands or contracts. This change in volume alters the molarity of the solution. Masses, however, do not change with temperature. Because of this, it is sometimes more useful to describe solutions in terms of how many moles of solute are dissolved in a specific mass of solvent. Such a description is called **molality**—the ratio of the number of moles of solute dissolved in one kilogram of solvent. The unit *m* is read as molal. A solution containing one mole of solute per kilogram of solvent is a one molal solution.

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kilogram of solvent}} = \frac{\text{moles of solute}}{1000 \text{ g of solvent}}$$

EXAMPLE PROBLEM 15-5

Calculating Molality

In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

1. Analyze the Problem

You are given the mass of solute and solvent. The molar mass of the solute can be used to determine the number of moles of solute in solution. Then, the molality can be calculated.

Known

mass of water (H₂O) = 100.0 g

mass of sodium chloride (NaCl) = 4.5 g

Unknown

$m = ? \text{ mol/kg}$

2. Solve for the Unknown

Use molar mass to calculate the number of moles of NaCl.

$$4.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ mol NaCl}$$

Convert the mass of H₂O from grams to kilograms.

$$100.0 \text{ g H}_2\text{O} \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.1000 \text{ kg H}_2\text{O}$$

Substitute the known values into the expression for molality and solve.

$$m = \frac{\text{moles of solute}}{\text{kilogram of solvent}} = \frac{0.077 \text{ mol NaCl}}{0.1000 \text{ kg H}_2\text{O}}$$
$$m = 0.77 \text{ mol/kg}$$

3. Evaluate the Answer

Because there was less than one-tenth mole of solute present in one-tenth kilogram of water, the molality should be less than one, as it is. The mass of sodium chloride was given with two significant figures; therefore the molality also is expressed with two significant figures.



The concentration of the resulting solution can be expressed in terms of moles of solute per volume (molarity) or moles of solute per mass (molality).

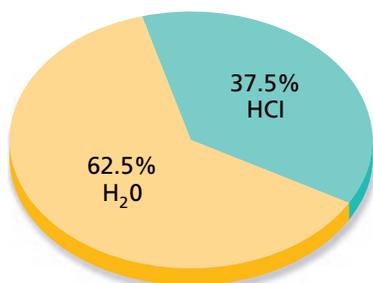
PRACTICE PROBLEMS

24. What is the molality of a solution containing 10.0 g Na₂SO₄ dissolved in 1000.0 g of water?
25. What is the molality of a solution containing 30.0 g of naphthalene (C₁₀H₈) dissolved in 500.0 g of toluene?



For more practice with molality problems, go to **Supplemental Practice Problems** in Appendix A.

Hydrochloric Acid in Aqueous Solution



$$X_{\text{HCl}} + X_{\text{H}_2\text{O}} = 1.000$$

$$0.229 + 0.771 = 1.000$$

Figure 15-15

The mole fraction expresses the number of moles of solvent and solute relative to the total number of moles of solution. Each mole fraction can be thought of as a percent. For example, the mole fraction of water ($X_{\text{H}_2\text{O}}$) is 0.771, which is equivalent to saying the solution contains 77.1% water (on a mole basis).

Mole fraction If you know the number of moles of solute and solvent, you can also express the concentration of a solution in what is known as a **mole fraction**—the ratio of the number of moles of solute in solution to the total number of moles of solute and solvent.

The symbol X is commonly used for mole fraction, with a subscript to indicate the solvent or solute. The mole fraction for the solvent (X_A) and the mole fraction for the solute (X_B) can be expressed as follows

$$X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_B}{n_A + n_B}$$

where n_A is the number of moles of solvent and n_B is the number of moles of solute. Why must the sum of the mole fractions for all components in a solution equal one?

Consider as an example the mole fraction of hydrochloric acid (HCl) in the aqueous solution shown in **Figure 15-15**. For every 100 grams of solution, 37.5 g would be HCl and 62.5 g would be H_2O . To convert these masses to moles, you would use the molar masses as conversion factors.

$$n_{\text{HCl}} = 37.5 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 1.03 \text{ mol HCl}$$

$$n_{\text{H}_2\text{O}} = 62.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = 3.47 \text{ mol H}_2\text{O}$$

Thus, the mole fractions of hydrochloric acid and water can be expressed as

$$X_{\text{HCl}} = \frac{n_{\text{HCl}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{1.03 \text{ mol HCl}}{1.03 \text{ mol HCl} + 3.47 \text{ mol H}_2\text{O}}$$

$$X_{\text{HCl}} = 0.229$$

$$X_{\text{H}_2\text{O}} = \frac{n_{\text{H}_2\text{O}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{3.47 \text{ mol H}_2\text{O}}{1.03 \text{ mol HCl} + 3.47 \text{ mol H}_2\text{O}}$$

$$X_{\text{H}_2\text{O}} = 0.771$$

PRACTICE PROBLEMS

26. What is the mole fraction of NaOH in an aqueous solution that contains 22.8% NaOH by mass?
27. An aqueous solution of NaCl has a mole fraction of 0.21. What is the mass of NaCl dissolved in 100.0 mL of solution?



For more practice with mole fraction problems, go to **Supplemental Problems** in Appendix A.

Section 15.2 Assessment

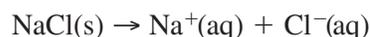
28. Distinguish between a dilute solution and a concentrated solution.
29. Compare and contrast five quantitative ways to describe the composition of solutions.
30. Describe the laboratory procedure for preparing a specific volume of a dilute solution from a concentrated stock solution.
31. **Thinking Critically** Explain the similarities and differences between a 1M solution of NaOH and a 1m solution of NaOH.
32. **Using Numbers** A can of chicken broth contains 450 mg of sodium chloride in 240.0 g of broth. What is the percent by mass of sodium chloride in the broth?

Solutes affect some of the physical properties of their solvents. Early researchers were puzzled to discover that the effects of a solute on a solvent depended only on how many solute particles were in solution, not on the specific solute dissolved. Physical properties of solutions that are affected by the number of particles but not the identity of dissolved solute particles are called **colligative properties**. The word colligative means “depending on the collection.” Colligative properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

Electrolytes and Colligative Properties

In Chapter 8, you learned that ionic compounds are called electrolytes because they dissociate in water to form a solution that conducts electric current. Some molecular compounds ionize in water and also are electrolytes. Electrolytes that produce many ions in solution are called strong electrolytes; those that produce only a few ions in solution are called weak electrolytes.

Sodium chloride is a strong electrolyte. It almost completely dissociates in solution, producing Na^+ and Cl^- ions.



Dissolving one mole of NaCl in a kilogram of water does not yield a $1m$ solution of ions. Rather, there would be almost two moles of solute particles in solution—approximately one mole each of Na^+ and Cl^- ions. How many moles of ions would you expect to find in a $1m$ aqueous solution of HCl ?

Nonelectrolytes in aqueous solution Many molecular compounds dissolve in solvents but do not ionize. Such solutions do not conduct an electric current, and the solutes are called nonelectrolytes. Sucrose is an example of a nonelectrolyte. A $1m$ sucrose solution contains only one mole of sucrose particles. **Figure 15-16** compares the conductivity of a solution containing an electrolyte solute with one containing a nonelectrolyte solute. Which compound would have the greater effect on colligative properties, sodium chloride or sucrose?

Objectives

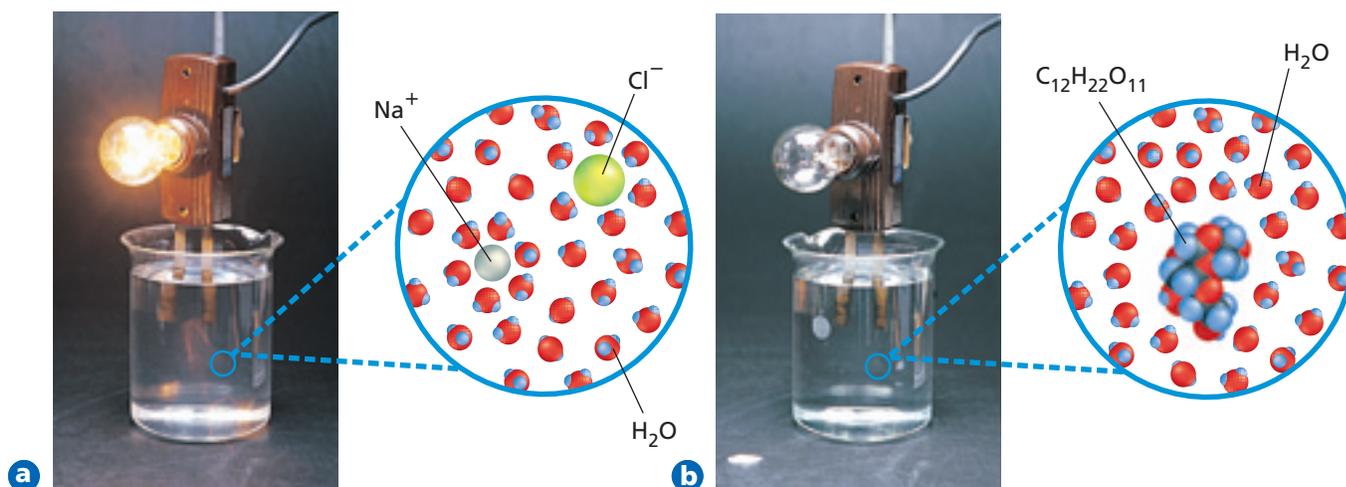
- **Explain** the nature of colligative properties.
- **Describe** four colligative properties of solutions.
- **Calculate** the boiling point elevation and the freezing point depression of a solution.

Vocabulary

colligative property
vapor pressure lowering
boiling point elevation
freezing point depression
osmosis
osmotic pressure

Figure 15-16

a Sodium chloride is a strong electrolyte and conducts electricity well. **b** Sucrose, while soluble in water, does not ionize and therefore does not conduct electricity. Which solute, sodium chloride or sucrose, produces more particles in solution per mole?



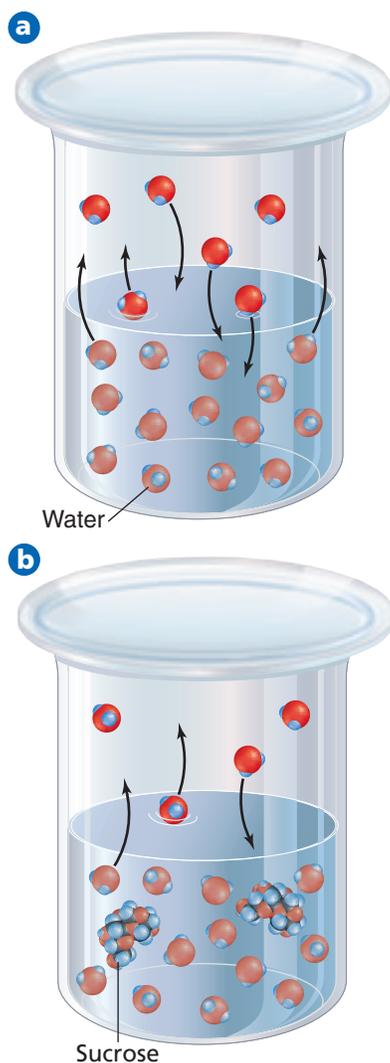


Figure 15-17

The vapor pressure of a pure solvent **(a)** is greater than the vapor pressure of a solution containing a nonvolatile solute **(b)**.

Vapor Pressure Lowering

In Chapter 13, you learned that vapor pressure is the pressure exerted in a closed container by liquid particles that have escaped the liquid's surface and entered the gaseous state. In a closed container at constant temperature and pressure, the solvent particles reach a state of dynamic equilibrium, escaping and reentering the liquid state at the same rate. The vapor pressure for a closed container of pure water is shown in **Figure 15-17a**.

Experiments show that adding a nonvolatile solute (one that has little tendency to become a gas) to a solvent lowers the solvent's vapor pressure. The particles that produce vapor pressure escape the liquid phase at its surface. When a solvent is pure, as shown in **Figure 15-17a**, its particles occupy the entire surface area. However, when the solvent contains solute, as shown in **Figure 15-17b**, a mix of solute and solvent particles occupies the surface area. With fewer solvent particles at the surface, fewer particles enter the gaseous state, and the vapor pressure is lowered. The greater the number of solute particles in a solvent, the lower the resulting vapor pressure. Thus, **vapor pressure lowering** is due to the number of solute particles in solution and is a colligative property of solutions.

You can predict the relative effect of a solute on vapor pressure based on whether the solute is an electrolyte or a nonelectrolyte. For example, one mole each of the solvated nonelectrolytes glucose, sucrose, and ethanol molecules has the same relative effect on the vapor pressure. However, one mole each of the solvated electrolytes phosphoric acid, sodium sulfate, and sodium chloride has an increasingly greater effect on vapor pressure because of the increasing number of ions each produces in solution.

Boiling Point Elevation

Because a solute lowers a solvent's vapor pressure, it also affects the boiling point of the solvent. Recall from Chapter 13 that the liquid in a pot on your stove boils when its vapor pressure equals atmospheric pressure. When the temperature of a solution containing a nonvolatile solute is raised to the boiling point of the pure solvent, the resulting vapor pressure is still less than the atmospheric pressure and the solution will not boil. Thus, the solution must be heated to a higher temperature to supply the additional kinetic energy needed to raise the vapor pressure to atmospheric pressure. The temperature difference between a solution's boiling point and a pure solvent's boiling point is called the **boiling point elevation**.

For nonelectrolytes, the value of the boiling point elevation, which is symbolized ΔT_b , is directly proportional to the solution's molality.

$$\Delta T_b = K_b m$$

The molal boiling point elevation constant, K_b , is the difference in boiling points between a $1m$ nonvolatile, nonelectrolyte solution and a pure solvent. It is expressed in units of $^{\circ}\text{C}/m$ and varies for different solvents. Values of K_b for several common solvents are found in **Table 15-4**. Note that water's K_b value is $0.512^{\circ}\text{C}/m$. This means that a $1m$ aqueous solution containing a nonvolatile, nonelectrolyte solute boils at 100.512°C , a temperature 0.512°C higher than pure water's boiling point of 100.0°C .

Like vapor pressure lowering, boiling point elevation is a colligative property. The value of the boiling point elevation is directly proportional to the solution's solute molality, that is, the greater the number of solute particles in the solution, the greater the boiling point elevation.

Table 15-4

Molal Boiling-Point Elevation Constants (K_b)		
Solvent	Boiling point ($^{\circ}\text{C}$)	K_b ($^{\circ}\text{C}/m$)
Water	100.0	0.512
Benzene	80.1	2.53
Carbon tetrachloride	76.7	5.03
Ethanol	78.5	1.22
Chloroform	61.7	3.63

Freezing Point Depression

The freezing point depression of a solution is another colligative property of solutions. At a solvent's freezing point temperature, the particles no longer have sufficient kinetic energy to overcome the interparticle attractive forces; the particles form into a more organized structure in the solid state. In a solution, the solute particles interfere with the attractive forces among the solvent particles. This prevents the solvent from entering the solid state at its normal freezing point. The freezing point of a solution is always lower than that of a pure solvent. **Figure 15-18** shows the differences in boiling and melting points of pure water and an aqueous solution. By comparing the solid and dashed lines, you can see that the temperature range over which the aqueous solution exists as a liquid is greater than that of pure water. You can observe the effect of freezing point depression by doing the **miniLAB** on this page.

A solution's **freezing point depression**, ΔT_f , is the difference in temperature between its freezing point and the freezing point of its pure solvent. For nonelectrolytes, the value of the freezing point depression, which is symbolized as ΔT_f , is directly proportional to the solution's molality.

$$\Delta T_f = K_f m$$

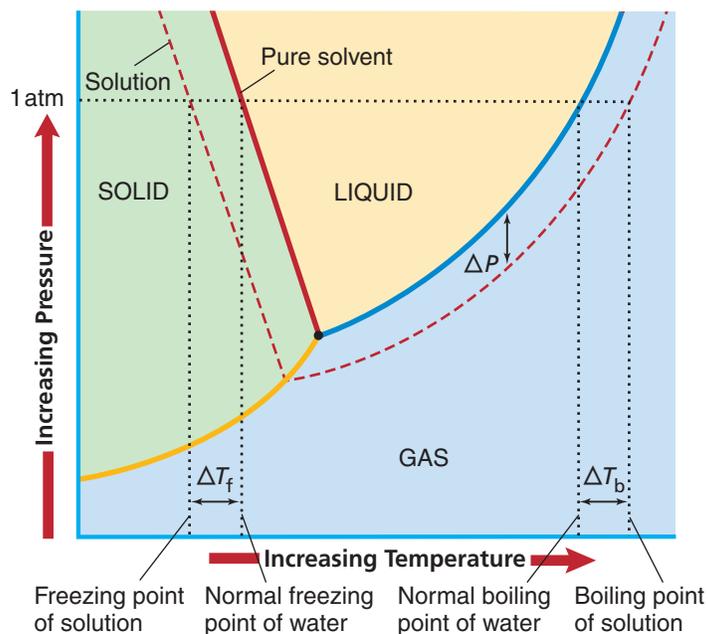


Figure 15-18

This phase diagram shows how temperature and pressure affect the solid, liquid, and gas phases of a pure solvent (solid lines) and a solution (dashed lines). The difference between the solid and dashed lines corresponds to vapor pressure lowering (ΔP), boiling point elevation (ΔT_b), and freezing point depression (ΔT_f).

miniLAB

Freezing Point Depression

Measuring The colligative property of freezing point depression can be observed in a simple laboratory investigation. You will measure the temperatures of two beakers and their contents.

Materials 400-mL beakers (2), crushed ice, rock salt (NaCl), water, stirring rods (2), graduated cylinder, thermometers (2), balance

Procedure

1. Fill two 400-mL beakers with crushed ice. Add 50 mL of cold tap water to each beaker.
2. Stir the contents of each beaker with a stirring rod until both beakers are at a constant temperature, approximately one minute.
3. Measure the temperature of each beaker using a thermometer and record the readings.
4. Add 75 g of rock salt to one of the beakers. Continue stirring both beakers. Some of the salt will dissolve.

5. When the temperature in each beaker is constant, record the readings.
6. To clean up, flush the contents of each beaker down the drain with excess water.

Analysis

1. Compare your readings taken for the ice water and the salt water. How do you explain the observed temperature change?
2. Why was salt only added to one of the beakers?
3. Salt is a strong electrolyte that produces two ions, Na^+ and Cl^- , when it dissociates in water. Why is this important to consider when calculating the colligative property of freezing point depression?
4. Predict if it would be better to use coarse rock salt or fine table salt when making homemade ice cream. Explain.

Values of K_f for several common solvents are found in **Table 15-5**. As with K_b values, K_f values are specific to their solvents. With water's K_f value of $1.86^\circ\text{C}/m$, a $1m$ aqueous solution containing a nonvolatile, nonelectrolyte solute freezes at -1.86°C rather than at pure water's freezing point of 0.0°C .

However, in using the ΔT_b and ΔT_f relationships with electrolytes, you must make sure to use the effective molality of the solution. For example, one mole of the electrolyte sodium chloride (NaCl) forms one mole of Na^+ ions and one mole of Cl^- ions in solution. Thus, a $1m$ aqueous NaCl solution produces two moles of solute particles in a kilogram of water—effectively acting as a $2m$ solution. Because of this, $2m$, and not $1m$, must be used in the ΔT_b and ΔT_f relationships for a $1m$ solution of sodium chloride. The following Example Problem illustrates this point.

Table 15-5

Molal Freezing Point Depression Constants (K_f)		
Solvent	Freezing point ($^\circ\text{C}$)	K_f ($^\circ\text{C}/m$)
Water	0.0	1.86
Benzene	5.5	5.12
Carbon tetrachloride	-23	29.8
Ethanol	-114.1	1.99
Chloroform	-63.5	4.68

EXAMPLE PROBLEM 15-6

Changes in Boiling and Freezing Points

What are the boiling point and freezing point of a $0.029m$ aqueous solution of sodium chloride (NaCl)?

1. Analyze the Problem

You are given the molality of an aqueous sodium chloride solution. First, calculate ΔT_b and ΔT_f based on the number of particles in solution. Then, to determine the elevated boiling point and the depressed freezing point, add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point.

Known

solute = sodium chloride (NaCl)
 molality of sodium chloride solution = $0.029m$

Unknown

boiling point = ? $^\circ\text{C}$
 freezing point = ? $^\circ\text{C}$

2. Solve for the Unknown

Each mole of the electrolyte sodium chloride dissociates in solution to produce two moles of particles. Calculate the effective number of solute particles in solution.

$$\text{particle molality} = 2 \times 0.029m = 0.058m$$

Substitute the known values for K_b , K_f , and particle molality into the ΔT_b and ΔT_f equations and solve.

$$\Delta T_b = K_b m = (0.512^\circ\text{C}/m)(0.058m) = 0.030^\circ\text{C}$$

$$\Delta T_f = K_f m = (1.86^\circ\text{C}/m)(0.058m) = 0.11^\circ\text{C}$$

Add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point to determine the elevated boiling point and depressed freezing point of the solution.

$$\text{boiling point} = 100.0^\circ\text{C} + 0.030^\circ\text{C} = 100.030^\circ\text{C}$$

$$\text{freezing point} = 0.0^\circ\text{C} - 0.11^\circ\text{C} = -0.11^\circ\text{C}$$

3. Evaluate the Answer

The boiling point is higher and the freezing point is lower, as expected. Because the molality of the solution has two significant figures, both ΔT_b and ΔT_f also have two significant figures. Because the normal boiling point and freezing point are exact values, they do not affect the number of significant figures in the final answer.

PRACTICE PROBLEMS

33. What are the boiling point and freezing point of a 0.625*m* aqueous solution of any nonvolatile, nonelectrolyte solute?
34. What are the boiling point and freezing point of a 0.40*m* solution of sucrose in ethanol?
35. A lab technician determines the boiling point elevation of an aqueous solution of a nonvolatile, nonelectrolyte to be 1.12°C. What is the solution's molality?
36. A student dissolves 0.500 mol of a nonvolatile, nonelectrolyte solute in one kilogram of benzene (C₆H₆). What is the boiling point elevation of the resulting solution?



For more practice with boiling point elevation and freezing point depression problems, go to **Supplemental Practice Problems** in Appendix A.

Osmosis and Osmotic Pressure

In Chapter 13, you learned that diffusion is the mixing of gases or liquids resulting from their random motions. **Osmosis** is the diffusion of solvent particles across a semipermeable membrane from an area of higher solvent concentration to an area of lower solvent concentration. Semipermeable membranes are barriers with tiny pores that allow some but not all kinds of particles to cross. The membranes surrounding all living cells are semipermeable membranes. Osmosis plays an important role in many biological systems such as kidney dialysis and the uptake of nutrients by plants.

Let's look at a simple system in which a sucrose-water solution is separated from its solvent—pure water—by a semipermeable membrane. During osmosis, water molecules move in both directions across the membrane, but the sugar molecules cannot cross it. Both water and sugar molecules contact the membrane on the solution side, but only water molecules contact the membrane on the pure solvent side. Thus, more water molecules cross the membrane from the pure solvent side than from the solution side.

The additional water molecules on the solution side of the membrane create pressure and push some water molecules back across the membrane. The amount of additional pressure caused by the water molecules that moved into the solution is called the **osmotic pressure**. Osmotic pressure depends upon the number of solute particles in a given volume of solution. Therefore, osmotic pressure is another colligative property of solutions.

Careers Using Chemistry

Renal Dialysis Technician

People with poorly functioning kidneys must routinely undergo dialysis in order to survive. The process, which removes impurities from the blood through osmosis, can be stressful for the patient. If you enjoy working with and helping people, and want to work in a medical setting, consider being a renal dialysis technician.

Dialysis technicians prepare the dialysis equipment and the patient for treatment, start the treatment, and monitor the process. They also make necessary adjustments, keep records, and respond to emergencies. Technicians spend more time with the patients than the doctors or nurses. They work in hospitals, outpatient facilities, and home-based dialysis programs.

Section 15.3 Assessment

37. Explain the nature of colligative properties.
38. Describe four colligative properties of solutions.
39. Explain why a solution has a lower boiling point than the pure solvent.
40. **Thinking Critically** Explain why the colligative properties described in this section may not apply to solutions containing volatile solutes. Hint: Volatile solutes are able to leave the liquid phase and enter the gas phase.
41. **Using Numbers** Calculate the boiling point elevation and freezing point depression of a solution containing 50.0 g of glucose (C₆H₁₂O₆) dissolved in 500.0 g of water.

Objectives

- **Identify** the properties of suspensions and colloids.
- **Describe** different types of colloids.
- **Explain** the electrostatic forces in colloids.

Vocabulary

suspension
colloid
Brownian motion
Tyndall effect

As you learned in Chapter 3, most of the forms of matter that you encounter are mixtures. A mixture is a combination of two or more substances that keep their basic identity. Components of a mixture come in contact with each other but do not undergo chemical change. You have been studying homogeneous mixtures called solutions so far in this chapter. Not all mixtures are solutions, however. Heterogeneous mixtures contain substances that exist in distinct phases. Two types of heterogeneous mixtures are suspensions and colloids.

Suspensions

Look at the mixture shown in **Figure 15-19a**. Although it resembles milk, it is actually a freshly-made mixture of cornstarch stirred into water. If you let this mixture stand undisturbed for a while, it separates into two distinct layers: a thick, white, pastelike substance on the bottom and the water on top. Cornstarch in water is a **suspension**, a mixture containing particles that settle out if left undisturbed. As shown in **Figure 15-19b**, pouring a liquid suspension through a filter also separates out the suspended particles. Other examples of suspensions include fine sand in water and muddy water.

Suspended particles are large compared to solvated particles, with diameters greater than 1000 nm (10^{-6} m) compared to diameters less than 1 nm (1^{-9} m) for solvated particles. Gravity acts on suspended particles in a short time, causing them to settle out of the mixture. Interestingly, the settled out cornstarch particles form a solidlike state on the bottom of the container. However, when stirred, the solidlike state quickly begins flowing like a liquid. Substances that behave this way are called thixotropic. One of the most common applications of a thixotropic mixture is house paint. The paint flows rather easily when applied with a brush, but quickly thickens to a solidlike state. The quickly thickening paint helps it stick to the house and prevents runs in the paint.

Figure 15-19

A suspension is a type of heterogeneous mixture. Suspension particles settle out over time **a**, and can be separated from the mixture by filtration **b**.



Table 15-6

Types of Colloids			
Category	Dispersed particles	Dispersing medium	Example
Solid sol	Solid	Solid	Colored gems
Sol	Solid	Liquid	Blood, gelatin
Solid emulsion	Liquid	Solid	Butter, cheese
Emulsion	Liquid	Liquid	Milk, mayonnaise
Solid foam	Gas	Solid	Marshmallow, soaps that float
Foam	Gas	Liquid	Whipped cream, beaten egg white
Aerosol	Solid	Gas	Smoke, dust in air
Aerosol	Liquid	Gas	Spray deodorant, clouds

Colloids

Particles in a suspension are much larger than atoms. In contrast, particles in a solution are atomic-scale in size. A heterogeneous mixture of intermediate size particles (between the size of solution particles and suspension particles) is a **colloid**. Colloid particles are between 1 nm and 1000 nm in diameter. The most abundant substance in the mixture is the dispersion medium. Milk is a colloid. Although homogenized milk resembles the cornstarch mixture in **Figure 15-19a**, you cannot separate its components by settling or by filtration.

Colloids are categorized according to the phases of their dispersed particles and dispersing mediums. Milk is a colloidal emulsion because liquid particles are dispersed in a liquid medium. Other types of colloids are described in **Table 15-6** and shown in **Figure 15-20**. How many of them are familiar to you? Can you name others?

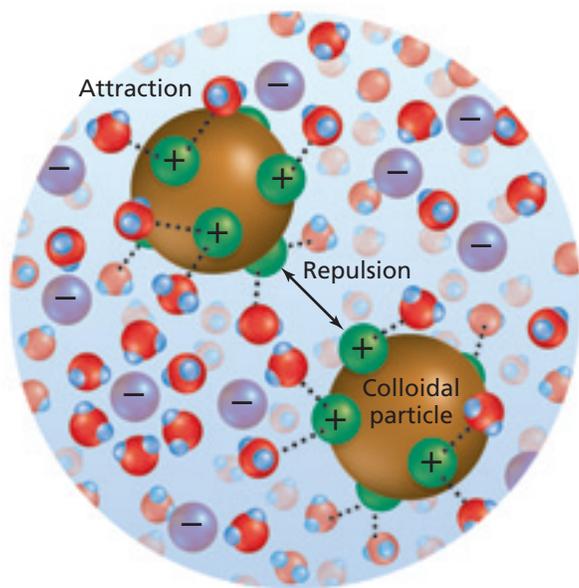
Figure 15-20

All of the photos shown here are examples of colloids. **a** Fog, which is a liquid aerosol, is formed when small liquid particles are dispersed in a gas. **b** Many paints are sols, a fluid colloidal system of fine solid particles in a liquid medium. **c** The bagel has a coarse texture with lots of small holes throughout. These small holes are formed by a type of foam produced by yeast in the dough.



Figure 15-21

The dispersing medium particles form charged layers around the colloid particles. These charged layers repel each other and keep the particles from settling out.



Brownian motion If you were to observe a liquid colloid under the magnification of a microscope, you would see that the dispersed particles make jerky, random movements. This erratic movement of colloid particles is called **Brownian motion**. It was first observed by and later named for the Scottish botanist Robert Brown, who noticed the random movements of pollen grains dispersed in water. Brownian motion results from collisions of particles of the dispersion medium with the dispersed particles. These collisions prevent the colloid particles from settling out of the mixture.

The reason the dispersed particles do not settle out is because they have polar or charged atomic groups on their surfaces. These areas on their surfaces attract the positively or negatively charged areas of the dispersing-medium particles. This results in the formation of electrostatic layers around the particles, as shown in **Figure 15-21**. The layers repel each other when the dispersed particles collide and, thus, the particles remain in the colloid. If you interfere with the electrostatic layering, colloid particles will settle out of the mixture. For example, if you stir an electrolyte into a colloid, the dispersed particles increase in size, destroying the colloid. Heating also destroys a colloid because it gives colliding particles enough kinetic energy to overcome the electrostatic forces and settle out.

problem-solving LAB

How can you measure the turbidity of a colloid?

Designing an Experiment The presence of cloudy water, or *turbidity*, in a lake or river results from the suspension of solids in the water. Most of these colloid particles come from erosion, industrial and human waste, algae blooms from fertilizers, and decaying organic matter. The degree of turbidity can greatly affect the health of a body of water in a number of ways, but its greatest impact is that it prevents light from reaching vegetation.

Analysis

The Tyndall effect can be used to measure the turbidity of water. Your goal is to plan a procedure and develop a scale to interpret the data.

Thinking Critically

Your plan should address the following.

1. What materials would you need to measure the Tyndall effect? Can you select CBLs, computers, or graphing calculators to collect or interpret data?
2. What variables can be used to relate the ability of light to pass through the liquid and the number of the colloid particles present?
3. How would you change the turbidity in the experiment?
4. What variable would change and by how much?
5. What will you use as a control?
6. How could you relate the variables used in the experiment to the actual number of colloid particles that are present?
7. Are there any safety precautions you must consider?





Figure 15-22

Two examples of the Tyndall effect are shown here. **a** The scattering of light by fog greatly reduces visibility. **b** The beam of light is visible in the colloid because of light scattering.

The Tyndall effect Whereas concentrated colloids are often cloudy or opaque, dilute colloids sometimes appear as clear as solutions. Dilute colloids appear to be homogeneous because their dispersed particles are so small that they cannot be seen by the unaided eye. However, dispersed colloid particles are large enough to scatter light, a phenomenon known as the **Tyndall effect**. In **Figure 15-22b**, a beam of light is shone through two unknown mixtures. Which mixture is the colloid? Which is the solution? You can see that dispersed colloid particles scatter the light, unlike solvated particles in the solution. Solutions never exhibit the Tyndall effect. Suspensions, such as the cornstarch shown in **Figure 15-19a**, also exhibit the Tyndall effect. You have observed the Tyndall effect if you have observed rays of sunlight passing through smoke-filled air, or viewed lights through fog at night, as shown in **Figure 15-22a**. Do the **problem-solving LAB** on the previous page to see how the Tyndall effect can be used to determine the amount of colloid particles in suspension.

Section 15.4 Assessment

42. Distinguish between suspensions and colloids.
43. Describe different types of colloids.
44. Why do dispersed colloid particles stay dispersed?
45. **Thinking Critically** Use the Tyndall effect to explain why it is more difficult to drive through fog using high beams than using low beams.
46. **Comparing and Contrasting** Make a table that compares the properties of solutions, suspensions, and colloids.

Beer's Law

Finding the concentration of an unknown solution is an important procedure in laboratory work. One method commonly used to determine solution concentration is to measure how much of a single wavelength of light is absorbed by the solution and compare it to known values of concentration and wavelength. Light absorbance is directly related to the concentration of a solution. This relationship is called Beer's law.

Problem

How is light absorbance used to find the concentration of a blue dye solution?

Objectives

- **Prepare** solutions of known concentration from a blue dye stock solution.
- **Measure** the absorbance of known and unknown aqueous solutions.
- **Infer** the relationship between light absorbance and concentration of a solution.

Materials

CBL system
graphing calculator
Vernier colorimeter
DIN adaptor and cable
TI Graph-Link (optional)
cuvette
cotton swabs
tissues for wiping
cuvette

blue dye stock solution
unknown solution
distilled water
50-mL graduated cylinder
100-mL beaker (2)
small test tube (5)
test tube rack
pipette (2)
pipette bulb
stirring rod

Safety Precautions



- Always wear safety goggles and a lab apron.
- The food-coloring solution can stain clothes.

Pre-Lab

1. Read the entire CHEMLAB procedure.
2. What is the total volume of solution in each test tube? Calculate the percent by volume of the solutions in test tubes 1 through 5. Prepare a data table.
3. Review with your teacher how a colorimeter works. How are absorbance (A) and transmittance ($%T$) related?

Known and Unknown Solutions Data			
Test tube	Concentration (%)	$%T$	A
1			
2			
3			
4			
5			
Unknown			

4. What is occurring during step 3 of the procedure? Why is a cuvette of water used?

Procedure

1. Transfer 30 mL of blue dye stock solution into a beaker. Transfer 30 mL of distilled water into another beaker.
2. Label five clean, dry test tubes 1 through 5.
3. Pipette 2 mL of blue dye stock solution from the beaker into test tube 1, 4 mL into test tube 2, 6 mL into test tube 3, and 8 mL into test tube 4.
CAUTION: Always pipette using a pipette bulb.
4. With another pipette, transfer 8 mL of distilled water from the beaker into test tube 1, 6 mL into tube 2, 4 mL into tube 3, and 2 mL into tube 4.
5. Mix the solution in test tube 1 with a stirring rod. Rinse and dry the stirring rod. Repeat this procedure with each test tube.

6. Pipette 10 mL of blue dye stock into test tube 5.
7. Load the ChemBio program into the calculator. Connect the CBL to the colorimeter using a DIN adaptor. Connect the CBL to the calculator using a link cable. Begin the ChemBio program on the calculator. Select "1" probe. Select 4: COLORIMETER. Enter Channel "1."
8. Fill a cuvette about three-fourths full with distilled water and dry its outside with a tissue. To calibrate the colorimeter, place the cuvette in the colorimeter and close the lid. Turn the wavelength knob to 0%T. Press TRIGGER on the CBL and enter 0 into the calculator. Turn the wavelength knob to Red (635 nm). Press TRIGGER on the CBL and enter 100 into the calculator. Leave the colorimeter set on Red for the rest of the lab. Remove the cuvette from the colorimeter. Empty the distilled water from the cuvette. Dry the inside of the cuvette with a clean cotton swab.
9. Select COLLECT DATA from the MAIN MENU. Select TRIGGER/PROMPT from the DATA COLLECTION menu. Fill the cuvette about three-fourths full with the solution from test tube 1. Dry the outside of the cuvette with a tissue and place the cuvette in the colorimeter. Close the lid. After 10 to 15 seconds, press TRIGGER and enter the concentration in percent from your data table into the calculator. Remove the cuvette and pour out the solution. Rinse the inside of the cuvette with distilled water and dry it with a clean cotton swab. Repeat this step for test tubes 2 through 5.
10. Select STOP AND GRAPH from the DATA COLLECTION menu when you have finished with data collection. Draw the graph, or use the TI Graph-Link to make a copy of the graph from the calculator screen. You also will want to copy the data from the STAT list into your data table (or you can print it from a screen print using Graph-Link).
11. Clean the cuvette with a cotton swab and fill it about three-fourths full with the unknown dye solution. Place the cuvette in the colorimeter and close the lid. From the MAIN MENU, select COLLECT DATA (do not select SET UP PROBES as this will erase your data lists). Select

MONITOR INPUT from the DATA COLLECTION MENU. Press ENTER to monitor the absorbance value of the colorimeter. After about 10-15 seconds, record the absorbance value and record it in your data table.

Cleanup and Disposal

1. All of the blue dye solutions can be rinsed down the drain.
2. Turn off the colorimeter. Clean and dry the cuvette. Return all equipment to its proper place.

Analyze and Conclude

1. **Analyzing Data** Evaluate how close your graph is to the direct relationship exhibited by Beer's law by doing a linear-regression line. Select FIT CURVE from the MAIN MENU (do not select SET UP PROBES as this will erase your data lists). Select LINEAR L1,L2. The calculator will give you an equation in the form of $y = ax + b$. One indicator of the fit of your graph is the size of b . A small value of b means the graph passes close to the origin. The closer the correlation coefficient r reported by the program is to 1.00, the better the fit of the graph.
2. **Drawing Conclusions** Use the graph of your absorbance and concentration data to determine the concentration of your unknown solution.
3. **Form a Hypothesis** Would you obtain the same data if red dye was used? Explain.
4. **Error Analysis** Analyze your b and r values. How closely do your results match Beer's law? Reexamine the procedure and suggest reasons why the correlation coefficient from your data does not equal 1.00.

Real-World Chemistry

1. Explain how Beer's law can be applied in food, drug, and medical testing.
2. The reaction of alcohol with orange dichromate ions to produce blue-green chromium(III) ions is used in the Breathalyzer test, a test that measures the presence of alcohol in a person's breath. How could a colorimeter be used in this analysis?

CHEMISTRY and Society

Sickle Cell Disease

There are millions of red blood cells in a single drop of blood. Red blood cells play a crucial role, transporting oxygen throughout the body.

Hemoglobin in Red Blood Cells

The cells pictured at the right appear red because they contain a bright red molecule—a protein called hemoglobin. One molecule of hemoglobin contains four iron(II) ions. Four oxygen molecules bind to each molecule of hemoglobin as blood passes through the lungs. As the blood circulates through the body, hemoglobin releases oxygen to the cells that make up various body tissues.

Healthy red blood cells are round and flexible, passing easily through narrow capillaries. But, hemoglobin in people with sickle cell disease differs from normal hemoglobin.

What makes sickle cells different?

Each hemoglobin molecule consists of four chains. The entire molecule contains 574 amino acid groups. In people with sickle cell disease, two non-polar amino acid groups are substituted for two polar amino acid groups. As a result, there is less repulsion between hemoglobin molecules, which allows abnormal hemoglobin molecules to clump together. This causes the abnormal hemoglobin molecules to be less soluble in the red blood cells, especially after oxygen has been absorbed through the lungs. The abnormal hemoglobin forces the red blood cells to become rigid and C-shaped, resembling the farming tool called a sickle.

These unusually shaped sickle cells clog the circulatory system, reducing blood flow. Therefore, oxygen supply to nearby tissues is reduced. Sickle cell disease causes pain, anemia, stroke, and susceptibility to infection.

Searching for a Cure

While most individuals with sickle cell disease have Mediterranean or African ancestry, it is common practice to screen all newborns in the United States for the condition. Research shows that early intervention can reduce the risk of serious infection, the leading cause of death in children with sickle cell disease. Intensive chemotherapy and stem cell



transfusions are being studied as ways to treat and possibly cure sickle cell disease.

Investigating the Issue

- 1. Communicating Ideas** Learn about the latest advances in sickle cell disease research. Create a chart that shows major symptoms, their causes, and their treatment. Share your findings with other students in your class.
- 2. Debating the Issue** The intensive chemotherapy required to treat sickle cell disease drastically weakens the immune system, leaving the patient vulnerable to overwhelming infection. Five to eight percent of the children who undergo this treatment do not survive. Should doctors be allowed to use this procedure to treat sickle cell disease?



Visit the Chemistry Web site at science.glencoe.com to find links to more information about sickle cell disease.

Summary

15.1 What are solutions?

- A solute dissolves in a solvent during a process called solvation. When the solvent is water, the process also is called hydration.
- Every substance has a characteristic solubility in a given solvent.
- Factors that affect solubility include the nature of the solute and solvent, temperature, and pressure.
- Henry's law states that the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid at a given temperature.

15.2 Solution Concentration

- The concentration of a solution is a quantitative measure of the amount of solute in a given amount of solvent or solution.
- Measures of concentration include mass and volume percentages, molarity, molality, and mole fraction.

- A dilute solution can be prepared from a more concentrated standard stock solution.

15.3 Colligative Properties of Solutions

- Physical properties affected by the concentration of the solute but not the nature of the solute are called colligative properties.
- Colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

15.4 Heterogeneous Mixtures

- One of the key differences between solutions, colloids, and suspensions is particle size.
- The random motion of colloidal dispersions due to molecular collisions is called Brownian motion.
- The scattering of light by colloidal particles is called the Tyndall effect. The Tyndall effect can be used to distinguish colloids from solutions.

Key Equations and Relationships

- Henry's law: $\frac{S_1}{P_1} = \frac{S_2}{P_2}$ (p. 460)
- Percent by mass = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$ (p. 463)
- Percent by volume = $\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$ (p. 464)
- Molarity (M) = $\frac{\text{moles of solute}}{\text{liters of solution}}$ (p. 464)
- Molarity-volume relationship: $M_1V_1 = M_2V_2$ (p. 467)
- Molality (m) = $\frac{\text{moles of solute}}{\text{kilogram of solvent}}$ (p. 469)
- Mole fractions: $X_A = \frac{n_A}{n_A + n_B}$ $X_B = \frac{n_B}{n_A + n_B}$ (p. 470)
- Boiling point elevation: $\Delta T_b = K_b m$ (p. 472)
- Freezing point depression: $\Delta T_f = K_f m$ (p. 473)

Vocabulary

- boiling point elevation (p. 472)
- Brownian motion (p. 478)
- colligative property (p. 471)
- colloid (p. 477)
- concentration (p. 462)
- freezing point depression (p. 473)
- heat of solution (p. 457)
- Henry's law (p. 460)
- immiscible (p. 454)
- insoluble (p. 454)
- miscible (p. 454)
- molality (p. 469)
- molarity (p. 464)
- mole fraction (p. 470)
- osmosis (p. 475)
- osmotic pressure (p. 475)
- saturated solution (p. 458)
- solubility (p. 457)
- soluble (p. 454)
- solvation (p. 455)
- supersaturated solution (p. 459)
- suspension (p. 476)
- Tyndall effect (p. 479)
- unsaturated solution (p. 458)
- vapor pressure lowering (p. 472)

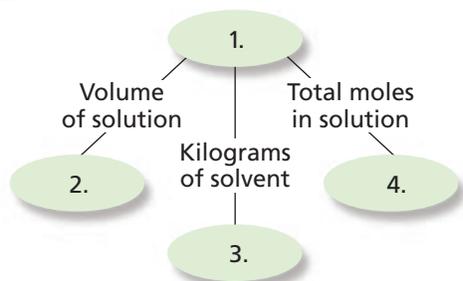


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 15 Assessment.

Concept Mapping

47. Complete the following concept map using the following terms: molarity, mole fraction, molality, moles of solute.



Mastering Concepts

48. What is the difference between solute and solvent? (15.1)
49. What determines whether a solute will be soluble in a given solvent? (15.1)
50. Explain the difference between saturated and unsaturated solutions. (15.1)
51. What does it mean if two liquids are said to be miscible? (15.1)
52. What are three ways to increase the rate of solvation? (15.1)
53. Why are gases less soluble at higher temperatures? (15.1)
54. What is the difference between percent by mass and percent by volume? (15.2)
55. What is the difference between molarity and molality? (15.2)
56. Explain on a particle basis why the vapor pressure of a solution is lower than a pure solvent. (15.3)
57. How does a solute affect the boiling point of a solution? (15.3)
58. How does a solute affect the freezing point of a solution? (15.3)
59. Describe osmosis. (15.4)
60. What is a colligative property? (15.4)

61. What is a suspension and how does it differ from a colloid? (15.4)
62. Name a colloid formed from a gas dispersed in a liquid. (15.4)
63. How can the Tyndall effect be used to distinguish between a colloid and a solution? Why? (15.4)

Mastering Problems

Henry's Law (15.1)

64. The solubility of a gas in water is 0.22 g/L at 20.0 kPa of pressure. What is the solubility when the pressure is increased to 115 kPa?
65. The solubility of a gas in water is 0.66 g/L at 15 kPa of pressure. What is the solubility when the pressure is increased to 40.0 kPa?
66. The solubility of a gas is 2.0 g/L at 50.0 kPa of pressure. How much gas will dissolve in 1 L at a pressure of 10.0 kPa?
67. The solubility of a gas is 4.5 g/L at a pressure of 1.0 atm. At what pressure will there be 45 g of gas in 1.0 L of solution?
68. The partial pressure of CO₂ inside a bottle of soft drink is 4.0 atm at 25°C. The solubility of CO₂ is 0.12 mol/L. When the bottle is opened, the partial pressure drops to 3.0×10^{-4} atm. What is the solubility of CO₂ in the open drink? Express your answer in grams per liter.

Percent Solutions (15.2)

69. Calculate the percent by mass of 3.55 g NaCl dissolved in 88 g water.
70. Calculate the percent by mass of benzene in a solution containing 14.2 g of benzene in 28.0 g of carbon tetrachloride.
71. What is the percent by volume of 25 mL of methanol in 75 mL of water?
72. A solution is made by adding 1.23 mol KCl to 1000.0 g of water. What is the percent by mass of KCl in this solution?
73. What mass of water must be added to 255.0 g NaCl to make a 15.00 percent by mass aqueous solution?
74. The label on a 250-mL stock bottle reads "21.5% alcohol by volume." What volume of alcohol does it contain?
75. A 14.0 percent by mass solution of potassium iodide dissolved in water has a density of 1.208 g/mL. How many grams of KI are in 25.0 mL of the solution?

Molarity (15.2)

- 76.** What is the molarity of the following solutions?
- 2.5 mol KCl in 1.0 L of solution
 - 1.35 mol H_2SO_4 in 245 mL of solution
 - 0.875 mol of ammonia in 155 mL of solution
- 77.** What is the molarity of the following solutions?
- 0.96 g MgCl_2 in 500 mL of solution
 - 9.33 g Na_2S in 450 mL solution
 - 2.48 g CaF_2 in 375 mL of solution
- 78.** How many moles of solute are contained in the following solutions?
- 15.25 mL 2.10M CaCl_2
 - 125 mL 0.0500M $\text{Ba}(\text{OH})_2$
 - 53.1 mL 12.2M HCl
- 79.** How many grams of solute are contained in the following solutions?
- 64.3 mL 0.0238M KOH
 - 142 mL 1.40M K_2SO_4
 - 750.0 mL 0.225M NH_4OH

Molar Dilution (15.2)

- 80.** How many milliliters of 2.55M NaOH is needed to make 125 mL 0.75M NaOH solution?
- 81.** How many milliliters of 0.400M HBr solution can be made from 50.0 mL of 8.00M HBr solution?
- 82.** What is the molarity of each resulting solution when the following mixtures are prepared?
- 500.0 mL H_2O is added to 20.0 mL 6.00M HNO_3
 - 30.0 mL 1.75M HCl is added to 80.0 mL 0.450M HCl

Molality and Mole Fraction (15.2)

- 83.** Calculate the molality of the following solutions.
- 15.7 g NaCl in 100.0 g H_2O
 - 20.0 g CaCl_2 in 700.0 g H_2O
 - 3.76 g NaOH in 0.850 L H_2O
- 84.** Calculate the mole fraction of NaCl, CaCl_2 , and NaOH in the solutions listed in the previous problem.
- 85.** What are the molality and mole fraction of solute in a 35.5 percent by mass aqueous solution of formic acid (CH_3OH)?

Colligative Properties (15.3)

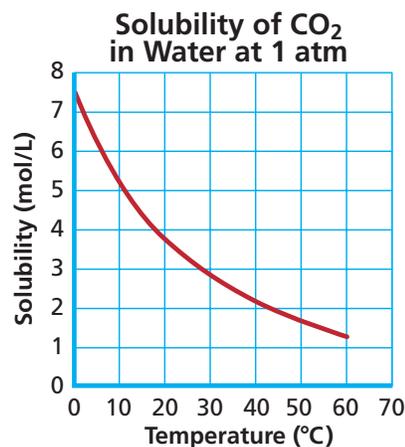
- 86.** Using the information in Tables 15-4 and 15-5, calculate the freezing point and boiling point of 12.0 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 50.0 g H_2O .

- 87.** Using the information in Tables 15-4 and 15-5, calculate the freezing point and boiling point of each of the following solutions.
- 2.75m NaOH in water
 - 0.586m of water in ethanol
 - 1.26m of naphthalene (C_{10}H_8) in benzene
- 88.** A rock salt (NaCl), ice, and water mixture is used to cool milk and cream to make homemade ice cream. How many grams of rock salt must be added to water to lower the freezing point 10.0°C ?
- 89.** Calculate the freezing point and boiling point of a solution that contains 55.4 g NaCl and 42.3 g KBr dissolved in 750.3 mL H_2O .

Mixed Review

Sharpen your problem-solving skills by answering the following.

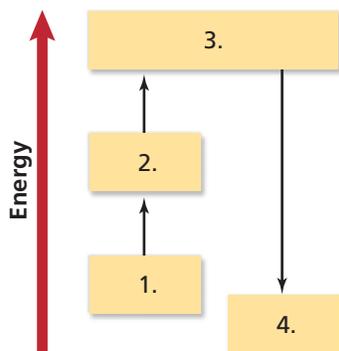
- 90.** If you prepared a saturated aqueous solution of potassium chloride at 25°C and then heated it to 50°C , would you describe the solution as unsaturated, saturated, or supersaturated? Explain.
- 91.** Use the graph below to explain why a carbonated beverage does not go flat as quickly when it contains ice.



- 92.** Which of the following substances will be soluble in the nonpolar solvent carbon tetrachloride (CCl_4): Br_2 , C_6H_{14} , NaNO_3 , HCl? Explain.
- 93.** How many grams of calcium nitrate ($\text{Ca}(\text{NO}_3)_2$) would you need to prepare 3.00 L of a 0.500M solution?
- 94.** What would be the molality of the solution described in the previous problem?

Thinking Critically

- 95. Inferring** Why not spread a nonelectrolyte on a road to help ice melt?
- 96. Using Scientific Diagrams** Complete the diagram below using the following phrases: solution, separated solvent + solute, separated solvent + separated solute, solvent + solute. Is the process described exothermic or endothermic?



- 97. Designing an Experiment** You are given a sample of a solid solute and three aqueous solutions containing that solute. How would you determine which solution is saturated, unsaturated, and supersaturated?
- 98. Using Graphs** The following solubility data was collected in an experiment. Plot a graph of the molarity of KI versus temperature. What is the solubility of KI at 55°C?


Table 15-7

Solubility of KI Data	
Temperature (°C)	Grams of KI per 100.0 g solution
20	144
40	162
60	176
80	192
100	206

- 99. Comparing** Which of the following solutions has the highest concentration? Rank the solutions from the greatest to the smallest boiling point depression. Explain your answer.
- 0.10 mol NaBr in 100.0 mL solution
 - 2.1 mol KOH in 1.00 L solution
 - 1.2 mol KMnO_4 in 3.00 L solution

Writing in Chemistry

- 100.** Investigate the total amount of salt used in the U.S. Construct a circle graph showing the different uses and amounts. Discuss each of these areas in detail. Salt was once used as a currency of high value. Find out why this was the case.
- 101.** Look up the various electrolytes in the human blood stream and discuss the importance of each.
- 102.** Research the contents of the tank scuba divers typically use. How does its composition differ from the air that you breathe? What is the condition known as the bends? How is it treated?

Cumulative Review

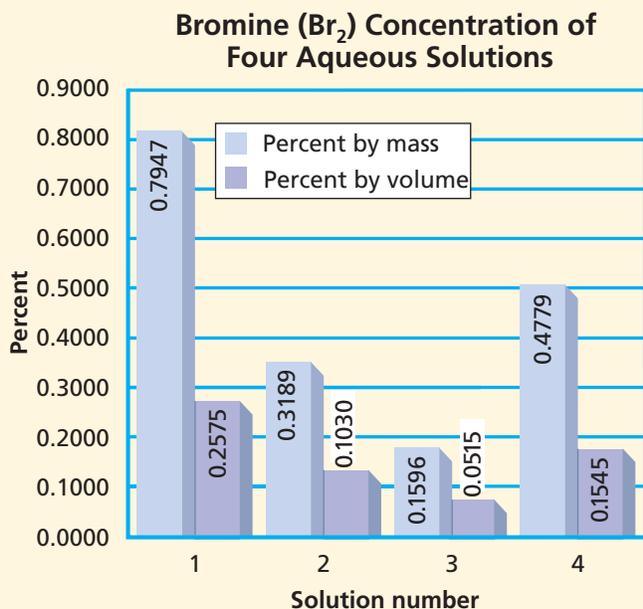
Refresh your understanding of previous chapters by answering the following.

- 103.** The radius of an argon atom is 94 pm. Assuming the atom is spherical, what is the volume of an argon atom in nm^3 ? $V = 4/3\pi r^3$ (Chapter 2)
- 104.** Identify which of the following molecules is polar. (Chapter 9)
- SiH_4
 - NO_2
 - H_2S
 - NCl_3
- 105.** Name the following compounds. (Chapter 8)
- NaBr
 - $\text{Pb}(\text{CH}_3\text{COO})_2$
 - $(\text{NH}_4)_2\text{CO}_3$
- 106.** A 12.0-g sample of an element contains 5.94×10^{22} atoms. What is the unknown element? (Chapter 11)
- 107.** Pure bismuth can be produced by the reaction of bismuth oxide with carbon at high temperatures.
- $$2\text{Bi}_2\text{O}_3 + 3\text{C} \rightarrow 4\text{Bi} + 3\text{CO}_2$$
- How many moles of Bi_2O_3 reacted to produce 12.6 moles of CO_2 ? (Chapter 12)
- 108.** A gaseous sample occupies 32.4 mL at -23°C and 0.75 atm. What volume will it occupy at STP? (Chapter 14)

Use these questions and the test-taking tip to prepare for your standardized test.

- How much water must be added to 6.0 mL of a 0.050M stock solution to dilute it to 0.020M?
 - 15 mL
 - 9.0 mL
 - 6.0 mL
 - 2.4 mL
- At a pressure of 1.00 atm and a temperature of 20°C, 1.72 g CO₂ will dissolve in 1 L of water. How much CO₂ will dissolve if the pressure is raised to 1.35 atm and the temperature stays the same?
 - 2.32 g/L
 - 1.27 g/L
 - 0.785 g/L
 - 0.431 g/L
- What is the molality of a solution containing 0.25 g of dichlorobenzene (C₆H₄Cl₂) dissolved in 10.0 g of cyclohexane (C₆H₁₂)?
 - 0.17 mol/kg
 - 0.014 mol/kg
 - 0.025 mol/kg
 - 0.00017 mol/kg
- If 1 mole of each of the solutes listed below is dissolved in 1 L of water, which solute will have the greatest effect on the vapor pressure of its respective solution?
 - KBr
 - C₆H₁₂O₆
 - MgCl₂
 - CaSO₄
- What volume of a 0.125M NiCl₂ solution contains 3.25 g NiCl₂?
 - 406 mL
 - 201 mL
 - 38.5 mL
 - 26.0 mL

Interpreting Graphs Use the graph to answer questions 6–8.



- The volume of bromine (Br₂) in 7.000 L of Solution 1 is _____.
 - 55.63 mL
 - 8.808 mL
 - 18.03 mL
 - 27.18 mL
- How many grams of Br₂ are in 55.00 g of Solution 4?
 - 3.560 g
 - 0.084 98 g
 - 1.151 g
 - 0.2628 g
- Which of the following relationships is true?
 - 2 × Concentration of solution 2 = Concentration of solution 3
 - 0.5 × Concentration of solution 2 = Concentration of solution 3
 - Concentration of solution 2 = 0.25 × Concentration of solution 3
 - Concentration of solution 2 = 3 × Concentration of solution 3
- All of the following are colligative properties EXCEPT _____.
 - boiling point elevation
 - freezing point depression
 - vapor pressure increase
 - osmotic pressure
- Colloids can be distinguished from solutions because _____.
 - dilute colloids have particles that can be seen with the naked eye
 - colloid particles are much smaller than solvated particles
 - dispersed colloid particles will settle out of the mixture in time
 - colloids will scatter light beams that are shone through them



TEST-TAKING TIP

Take a Break! If you have a chance to take a break or get up from your desk during a test, take it! Getting up and moving around will give you extra energy and help you clear your mind. During your stretch break, think about something other than the test so you'll be able to get back to the test with a fresh start.

Energy and Chemical Change

What You'll Learn

- ▶ You will measure and calculate the energy involved in chemical changes.
- ▶ You will write thermochemical equations and use them to calculate changes in enthalpy.
- ▶ You will explain how changes in enthalpy, entropy, and free energy affect the spontaneity of chemical reactions and other processes.

Why It's Important

Energy enables you to live, move from place to place, and stay comfortably warm or cool. Almost all of the energy you use comes from chemical reactions, including those that take place in your own body.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about energy in chemical reactions.

Each time the roller coaster zooms up and down the track, its energy changes back and forth between kinetic energy of motion and potential energy of position.



DISCOVERY LAB



Materials

test tubes (2)
thermometer
25-mL graduated cylinder
test tube rack
distilled water
5M HCl
5M NaOH

Temperature of a Reaction

Does a temperature change occur when a chemical reaction takes place?

Safety Precautions



Use care when handling HCl and NaOH solutions.

Procedure

1. Measure about 5 mL 5M NaOH solution and pour it into a large test tube. Determine the temperature of the solution.
2. Rinse the thermometer and the graduated cylinder with distilled water as directed by your teacher. Repeat step 1 using 5M HCl.
3. Pour the HCl solution into the NaOH solution. Immediately insert the thermometer and record the temperature of the mixture. How does it compare to the temperatures of NaOH and HCl solutions?

Analysis

What evidence do you have that a chemical reaction has occurred? Write the balanced chemical equation for the reaction between aqueous sodium hydroxide and aqueous hydrochloric acid.

Section

16.1

Energy

Objectives

- **Explain** what energy is and distinguish between potential and kinetic energy.
- **Relate** chemical potential energy to the heat lost or gained in chemical reactions.
- **Calculate** the amount of heat absorbed or released by a substance as its temperature changes.

Vocabulary

energy
law of conservation of energy
chemical potential energy
heat
calorie
joule
specific heat

You're familiar with the term *energy*. Perhaps you've heard someone say, "I just ran out of energy," after a strenuous game or a difficult day. Solar energy, nuclear energy, energy-efficient automobiles, and other energy-related topics are often discussed in the media. In the **DISCOVERY LAB**, you observed a temperature change associated with the release of energy during a chemical reaction. But does energy affect your everyday life?

The Nature of Energy

Energy cooks the food you eat and propels the vehicles that transport you. If the day is especially hot or cold, energy from burning fuels helps keep your home and school comfortable. Electrical energy provides light and operates devices from computers and TV sets to cellular phones, wristwatches, and calculators. Energy helped manufacture and deliver every material and device in your home, including the clothes you wear. Your every movement and thought requires energy. In fact, you can think of each cell in your body as a miniature factory that runs on energy derived from the food you eat. These examples only begin to define the role that energy plays in your life.

What is energy? **Energy** is the ability to do work or produce heat. It exists in two basic forms, potential energy and kinetic energy. Potential energy is energy due to the composition or position of an object. A macroscopic example of potential energy of position is water stored behind a dam above the turbines of a hydroelectric generating plant. When the dam gates are opened,

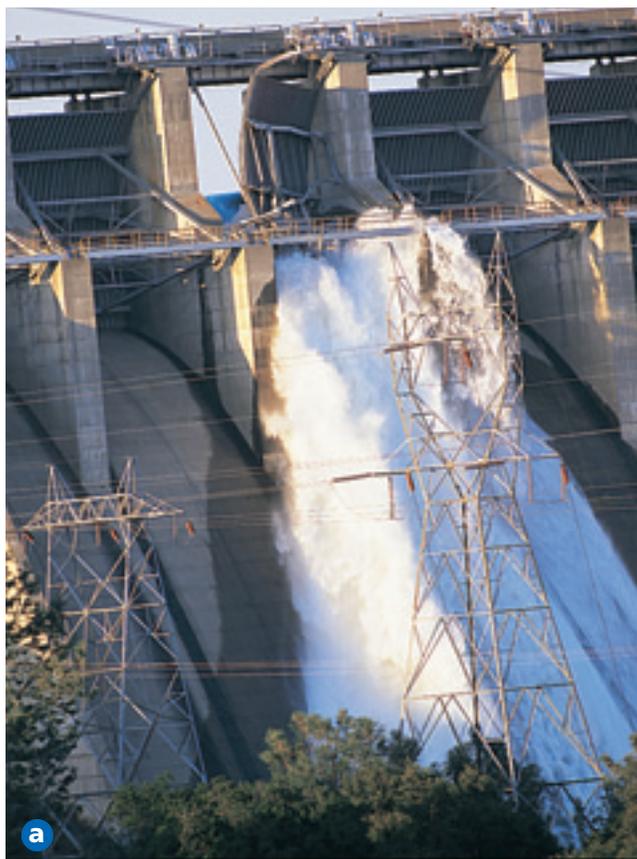


Figure 16-1

Energy is conserved in these energy transformations. In **a**, some of the potential energy of water stored behind Folsom Dam in California is converted to electrical energy. In **b**, the chemical potential energy stored in wood is converted to heat.

the water rushes down and does work by turning the turbines to produce electrical energy.

Kinetic energy is energy of motion. You can observe kinetic energy in the motion of people and objects all around you. The potential energy of the dammed water is converted to kinetic energy as the dam gates are opened and the water flows out.

Chemical systems contain both kinetic energy and potential energy. Recall from Chapter 13 that the kinetic energy of a substance is directly related to the constant random motion of its atoms or molecules and is proportional to temperature. As temperature increases, the motion of submicroscopic particles increases. The potential energy of a substance depends upon its composition: the type of atoms in the substance, the number and type of chemical bonds joining the atoms, and the particular way the atoms are arranged.

Law of conservation of energy When water rushes through turbines in the hydroelectric plant in **Figure 16-1a**, some of the water's potential energy is converted to electrical energy. When wood burns in a fireplace, as shown in **Figure 16-1b**, potential energy is liberated as heat. In both of these examples, energy changes from one form to another. But does the amount of energy change? No. As energy changes from one form to another, the total amount of energy remains constant. Energy is conserved. To better understand the conservation of energy, suppose you have money in two accounts at a bank and you transfer funds from one account to the other. Although the amount of money in each account has changed, the total amount of your money in the bank remains the same. When applied to energy, this analogy embodies the law of conservation of energy. The **law of conservation of energy** states that in any chemical reaction or physical process, energy can be converted from one form to another, but it is neither created nor destroyed.

Chemical potential energy The energy stored in a substance because of its composition is called **chemical potential energy**. Chemical potential energy plays an important role in chemical reactions. For example, consider octane (C_8H_{18}), one of the principal components of gasoline. The chemical potential energy of octane results from the arrangement of the carbon and



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

hydrogen atoms and the strength of the bonds that join them. When gasoline burns in an automobile's engine, some of octane's stored energy is converted to work in moving the pistons, which ultimately move the wheels and propel the automobile and its occupants from place to place. However, much of the potential energy of octane is released as heat. **Heat**, which is represented by the symbol q , is energy that is in the process of flowing from a warmer object to a cooler object. When the warmer object loses heat, its temperature decreases. When the cooler object absorbs heat, its temperature rises.

Measuring heat The flow of energy and the resulting change in temperature are clues to how heat is measured. In the metric system of units, the amount of heat required to raise the temperature of one gram of pure water by one degree Celsius (1°C) is defined as a **calorie** (cal). You've heard much about the caloric content of various foods. When your body breaks down sugars and fats to form carbon dioxide and water, these exothermic reactions generate heat that can be measured in Calories. Note that the nutritional Calorie is capitalized. That's because one nutritional Calorie, also known as one kilocalorie (kcal), equals 1000 calories. Suppose you eat a tablespoon of butter. One tablespoon of butter "has" 100 Calories. That means that if the butter was burned completely to produce carbon dioxide and water, 100 kcal (100 000 cal) of heat would be released.

The SI unit of heat and energy is the **joule** (J). One joule is the equivalent of 0.2390 calories, or one calorie equals 4.184 joules. **Table 16-1** shows the relationships among calories, nutritional Calories, joules, and kilojoules (kJ) and the conversion factors you can use to convert from one unit to another.

Table 16-1

Relationships Among Energy Units	
Relationship	Conversion factors
$1 \text{ J} = 0.2390 \text{ cal}$	$\frac{1 \text{ J}}{0.2390 \text{ cal}}$ $\frac{0.2390 \text{ cal}}{1 \text{ J}}$
$1 \text{ cal} = 4.184 \text{ J}$	$\frac{1 \text{ cal}}{4.184 \text{ J}}$ $\frac{4.184 \text{ J}}{1 \text{ cal}}$
$1 \text{ kJ} = 1000 \text{ J}$	$\frac{1 \text{ kJ}}{1000 \text{ J}}$ $\frac{1000 \text{ J}}{1 \text{ kJ}}$
$1 \text{ Calorie} = 1 \text{ kcal}$	$\frac{1 \text{ Calorie}}{1000 \text{ cal}}$
$1 \text{ kcal} = 1000 \text{ cal}$	$\frac{1000 \text{ cal}}{1 \text{ kcal}}$

EXAMPLE PROBLEM 16-1

Converting Energy Units

The breakfast shown in the photograph contains 230 nutritional Calories. How much energy in joules will this healthy breakfast supply?

1. Analyze the Problem

You are given an amount of energy in nutritional Calories. You must convert nutritional Calories to calories and then calories to joules.

Known

amount of energy = 230 Calories

Unknown

amount of energy = ? J

2. Solve for the Unknown

Use a conversion factor from **Table 16-1** to convert nutritional Calories to calories.

$$230 \text{ Calories} \times \frac{1000 \text{ cal}}{1 \text{ Calorie}} = 2.3 \times 10^5 \text{ cal}$$

Use a conversion factor to convert calories to joules.

$$2.3 \times 10^5 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 9.6 \times 10^5 \text{ J}$$

3. Evaluate the Answer

The minimum number of significant figures used in the conversion is two, so the answer correctly has two digits. A value of the order of 10^5 or 10^6 is expected because the given number of kilocalories is of the order of 10^2 and it must be multiplied by 10^3 to convert it to calories. Then, the calories must be multiplied by a factor of approximately 4. Therefore, the answer is reasonable.



It's important to eat the appropriate number of Calories. It's also important that the foods you select provide the nutrients your body needs.



For more practice converting from one energy unit to another, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

1. A fruit and oatmeal bar contains 142 nutritional Calories. Convert this energy to calories.
2. An exothermic reaction releases 86.5 kJ. How many kilocalories of energy are released?
3. If an endothermic process absorbs 256 J, how many kilocalories are absorbed?

Table 16-2

Specific Heats of Common Substances at 298 K (25°C)	
Substance	Specific heat J/(g·°C)
Water(l) (liquid)	4.184
Water(s) (ice)	2.03
Water(g) (steam)	2.01
Ethanol(l) (grain alcohol)	2.44
Aluminum(s)	0.897
Granite(s)	0.803
Iron(s)	0.449
Lead(s)	0.129
Silver(s)	0.235
Gold(s)	0.129

Specific Heat

You've learned that one calorie, or 4.184 J, is required to raise the temperature of one gram of pure water by one degree Celsius (1°C). That quantity, 4.184 J/(g·°C), is defined as the specific heat (c) of water. The **specific heat** of any substance is the amount of heat required to raise the temperature of one gram of that substance by one degree Celsius. Because different substances have different compositions, each substance has its own specific heat. The specific heats of several common substances are listed in **Table 16-2**.

Note how different the specific heats of the various substances are. If the temperature of water is to rise by one degree, 4.184 joules must be absorbed by each gram of water. But only 0.129 joule is required to raise the temperature of an equal mass of gold by one degree. Because of its high specific heat, water can absorb and release large quantities of heat. Have you ever noticed that vineyards and orchards are often planted near large bodies of water, as shown in **Figure 16-2**? During hot weather, the water in the lake or ocean absorbs heat from the air and thereby cools the surrounding area. During a cold snap, the water releases heat, warming the air in the surrounding area so that fruit trees and fruit are less susceptible to frost damage.

The high specific heat of water is also the basis for the use of water-filled plastic enclosures by gardeners in northern climates. Usually the planting of tender seedlings must wait until all danger of frost is past. However, the growing season can be extended into early spring if each plant is surrounded by a clear plastic enclosure filled with water. The water absorbs the heat of the Sun during the day. As the temperature drops at night, the water releases the absorbed

Figure 16-2

Ocean and lake fronts are preferred places for growing fruit because the water absorbs heat during sunny days and releases heat when the air cools, thus moderating temperatures.



heat. The plant is kept warm even when the temperature of the air drops to 0°C.

Calculating heat evolved and absorbed

If you've taken an early morning dive into a swimming pool similar to the one in **Figure 16-3**, you know that the water might be cold at that time of the day. Later in the day, especially if the Sun shines, the temperature of the water will be warmer. How much

warmer depends upon the specific heat of water, but other factors also are important. For example, suppose an architect designs a house that is to be partially heated by solar energy. Heat from the Sun will be stored in a solar pond similar to the swimming pool. The pond is to be made of 14 500 kg of granite rock and contain 22 500 kg of water. Both the granite and the water will absorb energy from the Sun during daylight hours. At night, the energy will not be released to the air, as happens with the swimming pool, but harnessed for use in the home. After conducting several experiments, the architect finds that the temperature of the water and granite increases an average of 22°C during the daylight hours and decreases the same amount during the night. Given these data, how much heat will the pond absorb and release during an average 24-hour period?

The specific heats of water and granite indicate how much heat one gram of each substance absorbs or releases when its temperature changes by 1°C. However, the pond contains much more than one gram of water and one gram of granite. And the temperatures of the two substances will increase and decrease an average of 22°C each day. How do mass and change in temperature affect the architect's calculations?

For the same change in temperature, 100 grams of water or granite absorb or release 100 times as much heat as one gram. Also, for the same 100 grams, increasing the temperature by 20 degrees Celsius requires 20 times as much heat as increasing the temperature by one degree Celsius. Therefore, the heat absorbed or released by a substance during a change in temperature depends not only upon the specific heat of the substance, but also upon the mass of the substance and the amount by which the temperature changes. You can express these relationships in an equation.

$$q = c \times m \times \Delta T$$

In the equation, q = the heat absorbed or released, c = the specific heat of the substance, m = the mass of the sample in grams, and ΔT is the change in temperature in °C. ΔT is the difference between the final temperature and the initial temperature or, $T_{\text{final}} - T_{\text{initial}}$.

You can use this equation to calculate the total amount of heat the solar pond will absorb and release on a typical day. A mass of 22 500 kg of water equals 2.25×10^7 g and 14 500 kg of granite equals 1.45×10^7 g. The change in temperature for both the water and the granite is 22°C. The specific heat of water is 4.184 J/(g·°C) and the specific heat of granite is 0.803 J/(g·°C). Because each substance has its own specific heat, the amount of heat absorbed or released by water and granite must be calculated separately.



Figure 16-3

The cold water of a swimming pool in the morning demonstrates that heat is lost by the water during the night.

$$q = c \times m \times \Delta T$$

$$q_{\text{water}} = 4.184 \text{ J/(g}\cdot\text{°C)} \times (2.25 \times 10^7 \text{ g}) \times 22\text{°C} = 2.1 \times 10^9 \text{ J}$$

$$q_{\text{granite}} = 0.803 \text{ J/(g}\cdot\text{°C)} \times (1.45 \times 10^7 \text{ g}) \times 22\text{°C} = 2.6 \times 10^8 \text{ J}$$

The total amount of heat absorbed or released by the pond is the sum of the two quantities.

$$q_{\text{total}} = q_{\text{water}} + q_{\text{granite}}$$

$$q_{\text{total}} = (2.1 \times 10^9 \text{ J}) + (2.6 \times 10^8 \text{ J}) = 2.4 \times 10^9 \text{ J or } 2.4 \times 10^6 \text{ kJ}$$

The solar pond will absorb 2.4 million kJ of heat during a sunny day and release 2.4 million kJ during the night.

EXAMPLE PROBLEM 16-2

Calculating Specific Heat

In the construction of bridges and skyscrapers, gaps must be left between adjoining steel beams to allow for the expansion and contraction of the metal due to heating and cooling. The temperature of a sample of iron with a mass of 10.0 g changed from 50.4°C to 25.0°C with the release of 114 J heat. What is the specific heat of iron?

1. Analyze the Problem

You are given the mass of the sample, the initial and final temperatures, and the quantity of heat released. The specific heat of iron is to be calculated. The equation that relates these variables can be rearranged to solve for c .

Known

joules of energy released = 114 J

$\Delta T = 50.4\text{°C} - 25.0\text{°C} = 25.4\text{°C}$

mass of iron = 10.0 g Fe

Unknown

specific heat of iron, $c = ? \text{ J/(g}\cdot\text{°C)}$

2. Solve for the Unknown

Rearrange the equation $q = c \times m \times \Delta T$ to isolate c by dividing each side of the equation by m and ΔT .

$$\frac{c \times m \times \Delta T}{m \times \Delta T} = \frac{q}{m \times \Delta T}$$

$$c = \frac{q}{m \times \Delta T}$$

Solve the equation using the known values.

$$c = \frac{114 \text{ J}}{(10.0 \text{ g})(25.4\text{°C})} = 0.449 \text{ J/(g}\cdot\text{°C)}$$

3. Evaluate the Answer

The values used in the calculation have three significant figures. Therefore, the answer is correctly stated with three digits. The value of the denominator of the equation is approximately two times the value of the numerator, so the final result (0.449), which is approximately 0.5, is reasonable. The calculated value is the same as that recorded for iron in **Table 16-2**.



Architects must make provision for the expansion and contraction of the steel frameworks of buildings because metals expand as they absorb heat and contract as they release heat.

PRACTICE PROBLEMS

- If the temperature of 34.4 g of ethanol increases from 25.0°C to 78.8°C, how much heat has been absorbed by the ethanol?
- A 4.50-g nugget of pure gold absorbed 276 J of heat. What was the final temperature of the gold if the initial temperature was 25.0°C? The specific heat of gold is 0.129 J/(g·°C).
- A 155-g sample of an unknown substance was heated from 25.0°C to 40.0°C. In the process, the substance absorbed 5696 J of energy. What is the specific heat of the substance? Identify the substance among those listed in **Table 16-2**.



For more practice calculating and using specific heat, go to **Supplemental Practice Problems** in Appendix A.

Using the Sun's energy The Sun is a virtually inexhaustible source of energy. Radiation from the Sun could supply all the energy needs of the world and reduce or eliminate the use of carbon dioxide-producing fuels, but practical problems have delayed the development of solar energy. The Sun shines for only a fraction of the day. Clouds often reduce the amount of available radiation. Because of this variability, effective storage of energy is critical. Solar ponds, such as the one you have been reading about, take up large land areas and can lose heat to the atmosphere.

Another method for storage involves the use of hydrates such as sodium sulfate decahydrate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$), which you learned about in Chapter 11. When heated by the Sun, the hydrate undergoes an endothermic process in which the sodium sulfate dissolves in the water of hydration. When the temperature drops at night, the hydrate re-crystallizes with the release of the absorbed solar energy.

A more promising approach to the harnessing of solar energy is the development of photovoltaic cells. These electronic devices convert solar radiation directly to electricity. Photovoltaic cells are now being used to provide the energy needs of space vehicles, as shown in **Figure 16-4**. At present, the cost of supplying electricity by means of photovoltaic cells is high compared to the cost of burning coal and oil. Therefore, although photovoltaic cells are clean and efficient, they remain a choice for the future.

The specific heat of a substance is a measure of how efficiently that substance absorbs heat. Water is particularly efficient and, for this reason, plays an important role in calorimetry, an experimental procedure you will learn about in Section 16.2.



Figure 16-4

Photovoltaic cells, positioned to absorb energy from the Sun, are the source of energy for the operation of this artist's conception of a space station.

Section 16.1 Assessment

- Explain what is meant by energy and list two units in which energy is measured.
- Distinguish between kinetic and potential energy in the following examples: two separated magnets; an avalanche of snow; books on library shelves; a mountain stream; a stock car race; separation of charge in a battery.
- What is the relationship between a calorie and a joule?
- Thinking Critically** One lawn chair is made of aluminum and another is made of iron. Both chairs are painted the same color. On a sunny day, which chair will be hottest to sit in? Explain why.
- Using Numbers** What is the specific heat of an unknown substance if a 2.50-g sample releases 12.0 cal as its temperature changes from 25.0°C to 20.0°C?

Objectives

- **Describe** how a calorimeter is used to measure energy absorbed or released.
- **Explain** the meaning of enthalpy and enthalpy change in chemical reactions and processes.

Vocabulary

calorimeter
thermochemistry
system
surroundings
universe
enthalpy
enthalpy (heat) of reaction

You have learned that some chemical reactions and processes must absorb energy in order to occur. These are called endothermic reactions. Others release energy and are called exothermic reactions. Whenever you cook food using methane or propane gas in your kitchen range, you utilize the heat released in the combustion of these fuels. But how do you measure the amount of heat released or absorbed when chemical reactions such as these occur?

Measuring Heat

Heat changes that occur during chemical and physical processes can be measured accurately and precisely using a calorimeter. A **calorimeter** is an insulated device used for measuring the amount of heat absorbed or released during a chemical or physical process. A known mass of water is placed in an insulated chamber in the calorimeter to absorb the energy released from the reacting system or to provide the energy absorbed by the system. The data to be collected is the change in temperature of this mass of water. Good results can be obtained in your calorimetry experiments by using a calorimeter made from nested plastic foam cups similar to the one shown in **Figure 16-5**. Because the foam cup is not tightly sealed, it is, in effect, open to the atmosphere. Reactions carried out in this type of calorimeter, therefore, occur at constant pressure.

Determining specific heat You can use a calorimeter to determine the specific heat of an unknown metal. Suppose you put 125 g of water into a foam-cup calorimeter and find that its initial temperature is 25.6°C, as shown in **Figure 16-5a**. Then, you heat a 50.0-g sample of the unknown metal to a temperature of 115.0°C and put the metal sample into the water. Heat flows from the hot metal to the cooler water and the temperature of the water rises. The flow of heat stops only when the temperature of the metal and the water are equal. In **Figure 16-5b**, the temperature of the water is constant. Both water and metal have attained a final temperature of 29.3°C. Assuming no heat is lost to the surroundings, the heat gained by the water is equal to the heat lost by the metal. This quantity of heat can be calculated using the equation you learned in Section 16.1.

$$q = c \times m \times \Delta T$$

First, calculate the heat gained by the water. For this you need the specific heat of water, 4.184 J/(g·°C).

$$q_{\text{water}} = 4.184 \text{ J/(g}\cdot\text{°C)} \times 125 \text{ g} \times (29.3\text{°C} - 25.6\text{°C})$$

$$q_{\text{water}} = 4.184 \text{ J/(g}\cdot\text{°C)} \times 125 \text{ g} \times 3.7\text{°C}$$

$$q_{\text{water}} = 1900 \text{ J}$$

The heat gained by the water, 1900 J, equals the heat lost by the metal, q_{metal} , so you can write this equation.

$$q_{\text{metal}} = 1900 \text{ J} = c_{\text{metal}} \times m \times \Delta T$$

$$1900 \text{ J} = c_{\text{metal}} \times m \times \Delta T$$

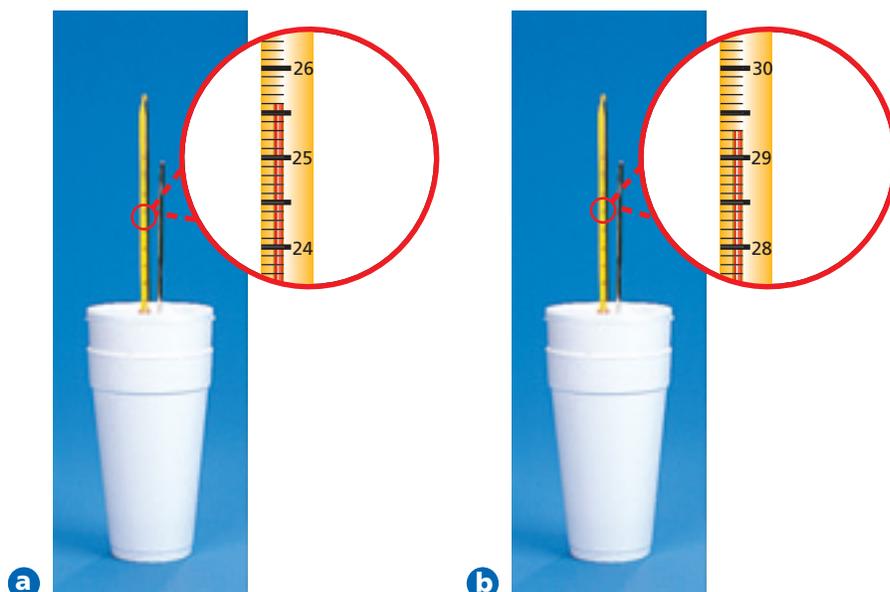


Figure 16-5

In **(a)**, the measured mass of water has an initial temperature of 25.6°C. A piece of hot metal is added. The metal transfers heat to the water until metal and water attain the same temperature. That final temperature is shown in **(b)** as 29.3°C.

Now, solve the equation for the specific heat of the metal, c_{metal} , by dividing both sides of the equation by $m \times \Delta T$.

$$c_{\text{(metal)}} = \frac{1900 \text{ J}}{m \times \Delta T}$$

The change in temperature for the metal, ΔT , is the difference between the final temperature of the water and the initial temperature of the metal ($115.0^\circ\text{C} - 29.3^\circ\text{C} = 85.7^\circ\text{C}$). Substitute the known values of m and ΔT (50.0 g and 85.7°C) into the equation and solve.

$$c_{\text{metal}} = \frac{1900 \text{ J}}{(50.0 \text{ g})(85.7^\circ\text{C})} = 0.44 \text{ J}/(\text{g}\cdot^\circ\text{C})$$

The unknown metal has a specific heat of $0.44 \text{ J}/(\text{g}\cdot^\circ\text{C})$. From **Table 16-2** on page 492, you can infer that the metal could be iron. The **CHEMLAB** at the end of this chapter will give you practice in calorimetry.

EXAMPLE PROBLEM 16-3

Using Data from Calorimetry

A piece of metal with a mass of 4.68 g absorbs 256 J of heat when its temperature increases by 182°C . What is the specific heat of the metal? Could the metal be one of the alkaline earth metals listed in **Table 16-3**?

1. Analyze the Problem

You are given the mass of the metal, the amount of heat it absorbs, and the temperature change. You must calculate the specific heat. The equation for q , the quantity of heat, should be used, but it must be solved for specific heat, c .

Known

mass of metal = 4.68 g metal
 quantity of heat absorbed, $q = 256 \text{ J}$
 $\Delta T = 182^\circ\text{C}$

Unknown

specific heat, $c = ? \text{ J}/(\text{g}\cdot^\circ\text{C})$

Continued on next page

Table 16-3

Periodic Trend in Specific Heats	
Alkaline earth elements	Specific heat ($\text{J}/\text{g}\cdot^\circ\text{C}$)
Beryllium	1.825
Magnesium	1.023
Calcium	0.647
Strontium	0.301
Barium	0.204

2. Solve for the Unknown

Solve for c by dividing both sides of the equation by $m \times \Delta T$.

$$q = c \times m \times \Delta T$$

$$c = \frac{q}{m \times \Delta T}$$

Substitute the known values into the equation.

$$c = \frac{256 \text{ J}}{(4.68 \text{ g})(182^\circ\text{C})} = 0.301 \text{ J/(g}\cdot^\circ\text{C)}$$

Table 16-3 indicates that the metal could be strontium.

3. Evaluate the Answer

The three quantities used in the calculation have three significant figures, so the answer is correctly stated with three digits. The calculations are correct and yield the expected unit. The calculated specific heat is the same as that of strontium.

Practice!

For more practice using calorimetry data, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- If 335 g water at 65.5°C loses 9750 J of heat, what is the final temperature of the water?
- The temperature of a sample of water increases from 20.0°C to 46.6°C as it absorbs 5650 J of heat. What is the mass of the sample?

Figure 16-6

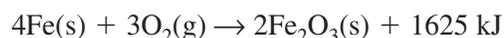
The energy released in the reaction of iron with oxygen provides comforting warmth for cold hands.



Chemical Energy and the Universe

Virtually every chemical reaction and change of physical state either releases or absorbs heat. Recall that an exothermic reaction is one in which energy is released and an endothermic reaction is one in which energy is absorbed. What happens to the heat released by an exothermic chemical reaction? What is the source of the heat absorbed by an endothermic reaction? Thermochemistry provides answers to these questions. **Thermochemistry** is the study of heat changes that accompany chemical reactions and phase changes.

Think about what happens when you warm your hands on a cold day by using a heat pack similar to the one shown in **Figure 16-6**. When you remove the plastic wrap, oxygen from the air enters the pack. The oxygen reacts with iron in the pack in an exothermic reaction described by the following equation. Note that energy is shown as a product of the reaction, which means that heat is released.

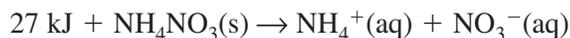


Because you are interested in the heat evolved by the chemical reaction going on inside the pack, it's convenient to think of the pack and its contents as the system. In thermochemistry, the **system** is the specific part of the universe that contains the reaction or process you wish to study. Everything in the universe other than the system is considered the **surroundings**. Therefore, the **universe** is defined as the system plus the surroundings.

$$\text{universe} = \text{system} + \text{surroundings}$$

What kind of energy transfer occurs during the exothermic heat-pack reaction? Heat produced by the reaction flows from the heat pack (the system) to your cold hands (part of the surroundings).

What happens in an endothermic reaction or process? The flow of heat is reversed. Heat flows from the surroundings to the system. For example, if you dissolve ammonium nitrate (NH_4NO_3) in water, the temperature of the water decreases. This is the equation for that process. Note that energy is on the reactant side of the equation, which means energy is absorbed.



The athletic trainer in **Figure 16-7** can cause this reaction to occur in a cold pack by breaking a membrane in the pack and allowing ammonium nitrate to mix with water. When the cold pack is placed on an injured knee, heat flows from the knee (part of the surroundings) into the cold pack (the system).

Enthalpy and enthalpy changes The total amount of energy a substance contains depends upon many factors, some of which are not totally understood today. Therefore, it's impossible to know the total heat content of a substance. Fortunately, chemists are usually more interested in changes in energy during reactions than in the absolute amounts of energy contained in the reactants and products.

For many reactions, the amount of energy lost or gained can be measured conveniently in a calorimeter at constant pressure, as in the experiment illustrated in **Figure 16-5** on page 497. The foam cup is not sealed, so the pressure is constant. Many reactions of interest take place at constant atmospheric pressure; for example, those that occur in living organisms on Earth's surface, in lakes and oceans, and in open beakers and flasks in the laboratory. To more easily measure and study the energy changes that accompany such reactions, chemists have defined a property called enthalpy. **Enthalpy** (H) is the heat content of a system at constant pressure.

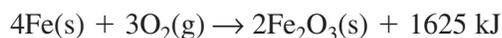
But why define a property if you can't know its absolute value? Although you can't measure the actual energy or enthalpy of a substance, you can measure the change in enthalpy, which is the heat absorbed or released in a chemical reaction. The change in enthalpy for a reaction is called the **enthalpy (heat) of reaction** (ΔH_{rxn}). You have already learned that a symbol preceded by the Greek letter Δ means a change in the property. Thus, ΔH_{rxn} is the difference between the enthalpy of the substances that exist at the end of the reaction and the enthalpy of the substances present at the start.

$$\Delta H_{\text{rxn}} = H_{\text{final}} - H_{\text{initial}}$$

Because the reactants are present at the beginning of the reaction and the products are present at the end, ΔH_{rxn} is defined by this equation.

$$\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$$

The sign of the enthalpy of reaction Recall the heat-pack reaction of iron with oxygen.



According to the equation, the reactants in this exothermic reaction lose heat. Therefore, $H_{\text{products}} < H_{\text{reactants}}$. When $H_{\text{reactants}}$ is subtracted from the smaller H_{products} , a negative value for ΔH_{rxn} is obtained. Enthalpy changes for exothermic reactions are always negative. The equation for the heat-pack reaction and its enthalpy change are usually written like this.



Figure 16-7

Heat is absorbed when ammonium nitrate dissolves in water in a cold pack. In this photograph, heat is transferred from the injured knee to the chemical process.

Careers Using Chemistry

Heating and Cooling Specialist

Are you fascinated by complex mechanical systems? Can you quickly figure out how they work? Then consider a career as a heating and cooling specialist.

Heating and cooling system mechanics install, maintain, and repair refrigeration and heating equipment. They must know how to make the equipment work as efficiently as possible. They must be able to read blueprints and use a wide range of tools, from pipe cutters to computerized diagnostic devices. Mechanics might specialize in one aspect of this field—or do it all.

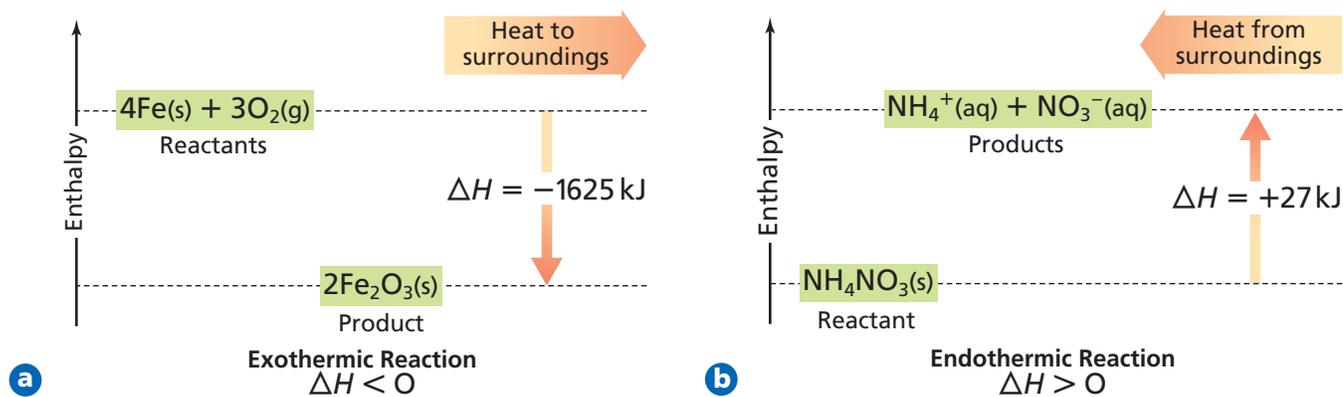
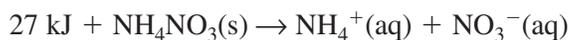


Figure 16-8

The downward arrow in **a** shows that 1625 kJ of heat is released to the surroundings in the reaction between Fe and O₂ to form Fe₂O₃. In contrast, the upward arrow in **b** indicates that 27 kJ of heat is absorbed from the surroundings in the process of dissolving NH₄NO₃.

Figure 16-8a is a diagram of the energy change for the exothermic reaction between iron and oxygen. You can see that the enthalpy of the product, Fe₂O₃, is 1625 kJ less than the enthalpy of the reactants Fe and O₂ because energy is released.

Similarly, recall the cold-pack reaction.



For this endothermic reaction, $\Delta H_{\text{products}} > \Delta H_{\text{reactants}}$. Therefore, when $\Delta H_{\text{reactants}}$ is subtracted from the larger $\Delta H_{\text{products}}$, a positive value for ΔH_{rxn} is obtained. Enthalpy changes for endothermic reactions are always positive. Chemists write the equation for the cold-pack reaction and its enthalpy change in this way.

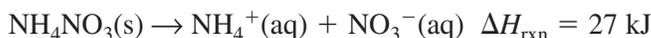


Figure 16-8b shows the energy change for the cold-pack reaction. In this endothermic reaction, the enthalpy of the products, aqueous NH₄⁺ and NO₃⁻ ions, is 27 kJ greater than the enthalpy of the reactant, NH₄NO₃ because energy is absorbed in the reaction. Compare **Figure 16-8a** and **Figure 16-8b** and then study **Table 16-4**, which shows the sign of ΔH_{rxn} for exothermic and endothermic reactions.

Recall that q was defined as the heat gained or lost in a chemical reaction or process. When the reaction takes place at constant pressure, the subscript p is sometimes added to the symbol q . The enthalpy change, ΔH , is equal to q_p , the heat gained or lost in a reaction or process carried out at constant pressure. Because all reactions presented in this textbook occur at constant pressure, you may assume that $q = \Delta H_{\text{rxn}}$.

Table 16-4

Enthalpy Changes for Exothermic and Endothermic Reactions	
Type of reaction	Sign of ΔH_{rxn}
Exothermic	Negative
Endothermic	Positive

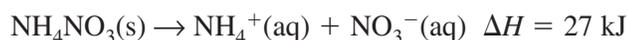
Section 16.2 Assessment

- Describe how you would calculate the amount of heat absorbed or released by a substance when its temperature changes.
- Why does ΔH for an exothermic reaction have a negative value?
- Why is a measured volume of water an essential part of a calorimeter?
- Explain the meaning of ΔH_{rxn} . Why is ΔH_{rxn} sometimes positive and sometimes negative?
- Thinking Critically** Could another liquid be used just as effectively as water in a calorimeter? Why or why not?
- Designing an Experiment** Explain how you would design an experiment to determine the specific heat of a 45-g piece of metal.

Chemical reactions either release energy (exothermic reactions) or absorb energy (endothermic reactions). The change in energy is an important part of chemical reactions so chemists include ΔH as part of the chemical equation.

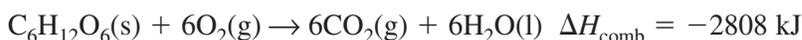
Writing Thermochemical Equations

The equations for the heat-pack and cold-pack reactions that you learned about in Section 16.2 are called thermochemical equations when they are written like this.



A **thermochemical equation** is a balanced chemical equation that includes the physical states of all reactants and products and the energy change, usually expressed as the change in enthalpy, ΔH .

The nature of the reaction or process described by a thermochemical equation is often written as a subscript of ΔH . For example, the highly exothermic combustion (comb) of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) occurs in the body as food is metabolized to produce energy for activities such as the one shown in **Figure 16-9**. The thermochemical equation for the combustion of glucose is



The energy released (-2808 kJ) is the enthalpy of combustion. The **enthalpy (heat) of combustion** (ΔH_{comb}) of a substance is the enthalpy change for the complete burning of one mole of the substance. Standard enthalpies of combustion for several common substances are given in **Table 16-5**. Standard enthalpy changes have the symbol ΔH° . The zero superscript tells you that the reactions were carried out under standard conditions. Standard conditions are one atmosphere pressure and 298 K (25°C) and should not be confused with standard temperature and pressure (STP).

Table 16-5

Standard Enthalpies of Combustion		
Substance	Formula	$\Delta H^\circ_{\text{comb}}$ (kJ/mol)
Sucrose (table sugar)	$\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(s)}$	-5644
Octane (a component of gasoline)	$\text{C}_8\text{H}_{18}\text{(l)}$	-5471
Glucose (a simple sugar found in fruit)	$\text{C}_6\text{H}_{12}\text{O}_6\text{(s)}$	-2808
Propane	$\text{C}_3\text{H}_8\text{(g)}$	-2219
Ethanol	$\text{C}_2\text{H}_5\text{OH(l)}$	-1367
Methane (the major component of natural gas)	$\text{CH}_4\text{(g)}$	-891
Methanol (wood alcohol)	$\text{CH}_3\text{OH(l)}$	-726
Carbon (graphite)	C(s)	-394
Hydrogen	$\text{H}_2\text{(g)}$	-286

Objectives

- **Write** thermochemical equations for chemical reactions and other processes.
- **Describe** how energy is lost or gained during changes of state.
- **Calculate** the heat absorbed or released in a chemical reaction.

Vocabulary

thermochemical equation
enthalpy (heat) of
combustion
molar enthalpy (heat) of
vaporization
molar enthalpy (heat) of
fusion

Figure 16-9

The energy expended by both horse and rider is obtained through the exothermic combustion of glucose in cells.



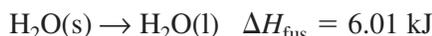
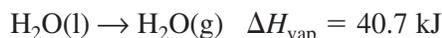
Changes of State

Many processes other than chemical reactions absorb or release heat. For example, think about what happens when you step out of a hot shower. You shiver as water evaporates from your skin. That's because your skin provides the heat needed to vaporize the water. As heat is taken from your skin to vaporize the water, you cool down. The heat required to vaporize one mole of a liquid is called its **molar enthalpy (heat) of vaporization** (ΔH_{vap}). Similarly, if you want a glass of cold water, you might drop an ice cube into it. The water cools as it provides the heat to melt the ice. The heat required to melt one mole of a solid substance is called its **molar enthalpy (heat) of fusion** (ΔH_{fus}). Because vaporizing a liquid and melting a solid are endothermic processes, their ΔH values are positive. Standard molar enthalpies of vaporization and fusion for four common compounds are shown in **Table 16-6**.

Table 16-6

Standard Enthalpies of Vaporization and Fusion			
Substance	Formula	$\Delta H_{\text{vap}}^{\circ}$ (kJ/mol)	$\Delta H_{\text{fus}}^{\circ}$ (kJ/mol)
Water	H ₂ O	40.7	6.01
Ethanol	C ₂ H ₅ OH	38.6	4.94
Methanol	CH ₃ OH	35.2	3.22
Ammonia	NH ₃	23.3	5.66

Thermochemical equations for changes of state The vaporization of water and the melting of ice can be described by the following equations.



The first equation indicates that 40.7 kJ of energy is absorbed when one mole of water is converted to one mole of water vapor. The second equation shows that when one mole of ice melts to form one mole of liquid water, 6.01 kJ of energy is absorbed.

What happens in the reverse processes, when water vapor condenses to liquid water or liquid water freezes to ice? The same amounts of energy are released in these exothermic processes as are absorbed in the endothermic processes of vaporization and melting. Thus, the molar enthalpy (heat) of condensation (ΔH_{cond}) and the molar enthalpy of vaporization have the same numerical value but opposite signs. Similarly, the molar enthalpy (heat) of solidification (ΔH_{solid}) and the molar enthalpy of fusion have the same numerical value but differ in sign.

$$\Delta H_{\text{vap}} = -\Delta H_{\text{cond}}$$

$$\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$$

Compare the following equations for the condensation and freezing of water with the equations above for the vaporization and melting of water. How would you summarize your observations? The relationships are illustrated in **Figure 16-10**.

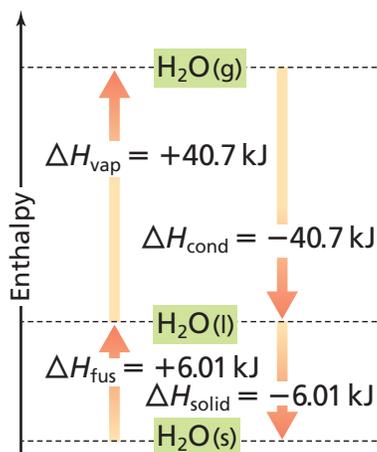


Figure 16-10

The left arrows show that the energy of the system increases as water melts and then vaporizes. The right arrows show that the energy of the system decreases as gaseous water condenses and then solidifies. What would be the energy of the process called sublimation in which ice is converted directly to water vapor?

Some farmers make use of the heat of fusion of water to protect fruit from freezing. They flood their orchards or fields with water, as shown in **Figure 16-11**. The energy released as the water freezes (ΔH_{fus}) often increases the temperature of the air enough to prevent frost damage to the fruit.

In the **problem-solving LAB**, you will interpret the heating curve of water using the heats of fusion and vaporization. The **How It Works** feature, at the end of the chapter, describes a practical application of the vaporization of a liquid.



Figure 16-11

If the temperature should drop to freezing, the water that floods this field will release heat (ΔH_{fus}) as it freezes and warm the surrounding air.

problem-solving LAB

How much energy is needed to heat water from a solid to a vapor?

Making and Using Graphs Water molecules have a strong attraction for one another because they are polar. The polarity of water accounts for its high specific heat and relatively high enthalpies of fusion and vaporization. Water's high specific heat and the presence of an enormous amount of water on Earth's surface have a large influence on the weather.

Analysis

Use the data in the table to plot a heating curve of temperature versus time for a 180-g sample of water as it is heated at a constant rate from -20°C to 120°C . Draw a best-fit line through the points. Note the time required for water to pass through each segment of the graph.

Thinking Critically

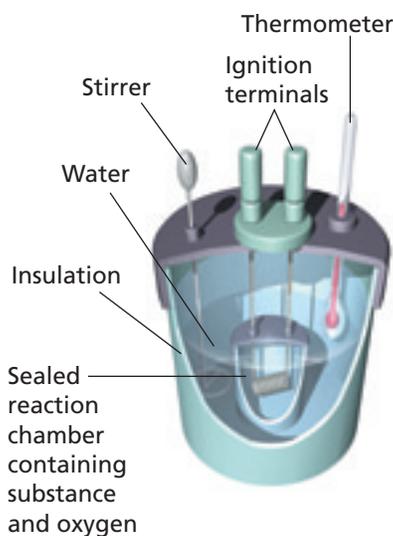
- For each of the five regions of the graph, indicate how the absorption of heat changes the energy (kinetic or potential) of the water molecules.
- Calculate the amount of heat required to pass through each region of the graph (180 g H_2O = 10 mol H_2O , $\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$, $\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$, $c = 4.184 \text{ J/g}\cdot^{\circ}\text{C}$). How does the length of time needed to pass through each region relate to the amount of heat absorbed?
- What would the heating curve of ethanol look like? Make a rough sketch of ethanol's curve from -120°C to 90°C . Ethanol melts at -114°C and boils at 78°C . What factors determine the

Time and Temperature Data for Water

Time (min)	Temperature $^{\circ}\text{C}$	Time (min)	Temperature $^{\circ}\text{C}$
0.0	-20	13.0	100
0.5	-11	13.5	100
1.0	0	14.0	100
1.5	0	14.5	100
2.0	0	15.0	100
2.5	0	15.5	100
3.0	9	16.0	100
3.5	18	16.5	100
4.0	26	17.0	100
4.5	34	17.5	100
5.0	42	18.0	100
5.5	51	18.5	100
6.0	58	19.0	100
6.5	65	19.5	100
7.0	71	20.0	100
7.5	77	20.5	100
8.0	83	21.0	100
8.5	88	21.5	100
9.0	92	22.0	100
9.5	95	22.5	100
10.0	98	23.0	100
10.5	100	23.5	100
11.0	100	24.0	100
11.5	100	24.5	108
12.0	100	25.0	120
12.5	100		

lengths of the flat regions of the graph and the slope of the curve between the flat regions?

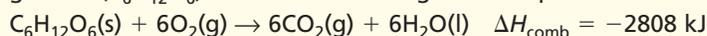
EXAMPLE PROBLEM 16-4



Inside the central chamber or bomb of a bomb calorimeter, reactions are carried out at constant volume. A sample of known mass is ignited by an electric spark and burned in an excess of oxygen. The heat released is transferred to a known mass of water in the well insulated outer chamber.

Calculating Energy Released in a Reaction

A bomb calorimeter is useful for measuring the energy released in combustion reactions. The reaction is carried out in a constant volume bomb with a high pressure of oxygen. How much heat is evolved when 54.0 g glucose ($C_6H_{12}O_6$) is burned according to this equation?



1. Analyze the Problem

You are given a mass of glucose, the equation for the combustion of glucose, and ΔH_{comb} . Grams of glucose must be converted to moles of glucose. Because the molar mass of glucose is more than three times the given mass of glucose, you can predict that the energy evolved will be less than one-third ΔH_{comb} .

Known

mass of glucose = 54.0 g $C_6H_{12}O_6$

$\Delta H_{\text{comb}} = -2808 \text{ kJ}$

Unknown

$q = ? \text{ kJ}$

2. Solve for the Unknown

Convert grams of $C_6H_{12}O_6$ to moles of $C_6H_{12}O_6$ by multiplying by the conversion factor that relates moles and mass, the inverse of molar mass.

$$54.0 \text{ g } C_6H_{12}O_6 \times \frac{1 \text{ mol } C_6H_{12}O_6}{180.18 \text{ g } C_6H_{12}O_6} = 0.300 \text{ mol } C_6H_{12}O_6$$

Multiply moles of $C_6H_{12}O_6$ by the conversion factor that relates kJ and moles, the enthalpy of combustion, ΔH_{comb} .

$$\frac{\text{kJ released}}{1 \text{ mole substance burned}} \\ 0.300 \text{ mol } C_6H_{12}O_6 \times \frac{2808 \text{ kJ}}{1 \text{ mol } C_6H_{12}O_6} = 842 \text{ kJ}$$

3. Evaluate the Answer

All values in the calculation have at least three significant figures, so the answer is correctly stated with three digits. As predicted, the released energy is less than one-third ΔH_{comb} .

PRACTICE PROBLEMS

20. Calculate the heat required to melt 25.7 g of solid methanol at its melting point.
21. How much heat is evolved when 275 g of ammonia gas condenses to a liquid at its boiling point?
22. What mass of methane must be burned in order to liberate 12 880 kJ of heat?

Example Problem 16-4 shows that when glucose is burned in a bomb calorimeter, a significant amount of energy is released. The same amount of energy is produced in your body when an equal mass of glucose is metabolized (converted to carbon dioxide and water). The reaction, which occurs in every cell of your body, may not occur in the same way as the combustion in Example Problem 16-4, but the overall result is the same. The metabolism of glucose and other sugars provides the energy you need to breathe, move, think, and grow.

Practice!

For more practice calculating the energy released in a reaction, go to **Supplemental Practice Problems** in Appendix A.

Enthalpy of Fusion for Ice

Applying Concepts When ice is added to water at room temperature, the water provides the energy for two processes. The first process is the melting of the ice. The energy required to melt ice is the enthalpy of fusion (ΔH_{fus}). The second process is raising the temperature of the melted ice from its initial temperature of 0.0°C to the final temperature of the liquid water. In this experiment, you will collect data to calculate the enthalpy of fusion for ice.

Materials foam cup, thermometer, stirring rod, ice, water, balance

Procedure

1. Measure the mass of an empty foam cup and record it in your data table.
2. Fill the foam cup about one-third full of water. Measure and record the mass.
3. Place the thermometer in the cup. Read and record the initial temperature of the water.

4. Quickly place a small quantity of ice in the plastic cup. Gently stir the water with a stirring rod until the ice melts. Record the lowest temperature reached as the final temperature.
5. Measure the mass of the cup and water.

Analysis

1. The heat lost by the liquid water equals the heat needed to melt the ice plus the heat needed to increase the temperature of the melted ice from 0.0°C to the final temperature. Calculate the heat lost by the water.
2. Calculate the heat gained by the melted ice as its temperature rose from 0.0°C to the final temperature.
3. The difference between the heat lost by the water and the heat gained by the melted ice equals the heat of fusion. Calculate the heat of fusion in joules per gram of ice.
4. Calculate ΔH_{fus} in kJ/mol.
5. Calculate the percent error of your experimental ΔH_{fus} . Compare your value to the actual value 6.01 kJ/mol .

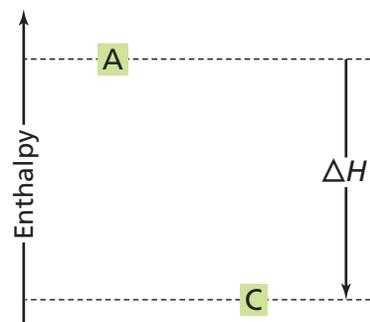
As in the **miniLAB** above, you can use calorimetry to measure the energy released or absorbed in a chemical reaction or change of state. However, sometimes carrying out an experiment is difficult or even impossible. In the next section, you'll see that there are ways to calculate energy changes.

Section 16.3 Assessment

23. List the information contained in a thermochemical equation.
24. Which of the following processes are exothermic? Endothermic?
 - a. $\text{C}_2\text{H}_5\text{OH}(\text{l}) \rightarrow \text{C}_2\text{H}_5\text{OH}(\text{g})$
 - b. $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{l})$
 - c. $\text{Br}_2(\text{l}) \rightarrow \text{Br}_2(\text{s})$
 - d. $\text{NaCl}(\text{s}) \rightarrow \text{NaCl}(\text{l})$
 - e. $\text{C}_5\text{H}_{12}(\text{g}) + 8\text{O}_2(\text{g}) \rightarrow 5\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l})$
25. Explain how you could calculate the heat released in freezing 0.250 mol water.
26. **Thinking Critically** The freezing of water at 0.0°C is an exothermic process. Explain how the positions and thus, the chemical potential energy of the water molecules change as water goes from

a liquid to a solid. Does the kinetic energy of the water molecules change?

27. **Interpreting Scientific Illustrations** The reaction $\text{A} \rightarrow \text{C}$ is shown in the enthalpy diagram. Is the reaction exothermic or endothermic? Explain your answer.



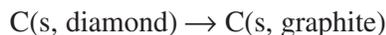
Objectives

- **Use** Hess's law of summation of enthalpies of reaction to calculate the enthalpy change for a reaction.
- **Explain** the basis for the table of standard enthalpies of formation.
- **Calculate** ΔH_{rxn} using thermochemical equations.
- **Determine** the enthalpy change for a reaction using standard enthalpies of formation data.

Vocabulary

Hess's law
standard enthalpy (heat) of formation

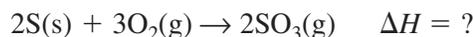
In principle, you can determine ΔH for any chemical reaction by using a calorimeter to measure the heat evolved or absorbed during the reaction. However, consider the reaction involving the conversion of carbon in its allotropic form diamond to carbon in its allotropic form graphite.



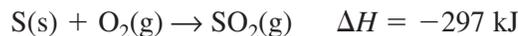
This reaction occurs over millions of years—so slowly that measuring the enthalpy change is virtually impossible. Other reactions occur under conditions difficult to duplicate in a laboratory. Still others don't occur cleanly; that is, products other than the desired ones may be formed. For reactions such as these, chemists use a theoretical way to determine ΔH .

Hess's Law

Suppose you are studying the formation of acid rain that results from the reaction of water in the atmosphere with sulfur trioxide expelled during a volcanic eruption such as the one shown in **Figure 16-12**. You would need to determine ΔH for this reaction.

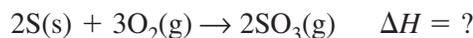


Unfortunately, when you try to duplicate the reaction in the laboratory by burning sulfur in air, only small quantities of sulfur trioxide are formed. What results is a mixture containing mostly sulfur dioxide produced according to this equation.



In situations such as this, you can calculate ΔH for the formation of sulfur trioxide using Hess's law of heat summation. **Hess's law** states that if you can add two or more thermochemical equations to produce a final equation for a reaction, then the sum of the enthalpy changes for the individual reactions is the enthalpy change for the final reaction. Hess's law enables you to calculate enthalpy changes for an enormous number of chemical reactions by imagining that each reaction occurs through a series of steps for which the enthalpy changes are known.

Applying Hess's law How can Hess's law be used to calculate the energy change for the reaction that produces SO_3 ?



First, chemical equations are needed that contain the substances found in the desired equation and have known enthalpy changes. The following equations contain S, O_2 , and SO_3 .

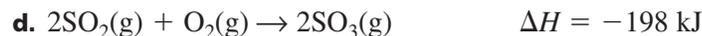
- $\text{S(s)} + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) \quad \Delta H = -297 \text{ kJ}$
- $2\text{SO}_3(\text{g}) \rightarrow 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = 198 \text{ kJ}$

The desired equation shows two moles of sulfur reacting, so Equation **a** must be rewritten for two moles of sulfur by multiplying the coefficients by 2. The enthalpy change, ΔH , must also be doubled because twice the energy will be

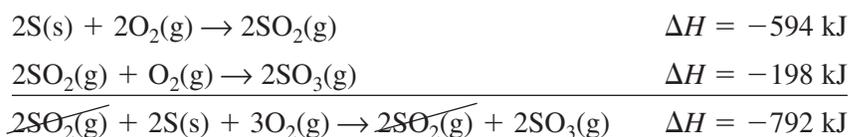
released if two moles of sulfur react. When these changes are made, Equation **a** becomes the following (Equation **c**).



Because you want to determine ΔH for a reaction in which SO_3 is a product rather than a reactant, Equation **b** must be reversed. Recall that when you reverse an equation, the sign of ΔH changes. The reverse of Equation **b** is Equation **d**.



Now, add Equations **c** and **d** to obtain the equation for the desired reaction. Add the ΔH values for the two equations to determine ΔH for the desired reaction. Any terms that are common to both sides of the combined equation should be canceled.



The sum of the two equations is the equation for the burning of sulfur to form SO_3 , and the sum of the ΔH values is the enthalpy change for the reaction.



The diagram in **Figure 16-13** will help you visualize the calculation.

Sometimes thermochemical equations are written with fractional coefficients because they are balanced for one mole of product. For example, the thermochemical equation for the reaction between sulfur and oxygen to form one mole of sulfur trioxide is the following.



What factor would you need to multiply this equation and its enthalpy change by to obtain the equation you worked with above?

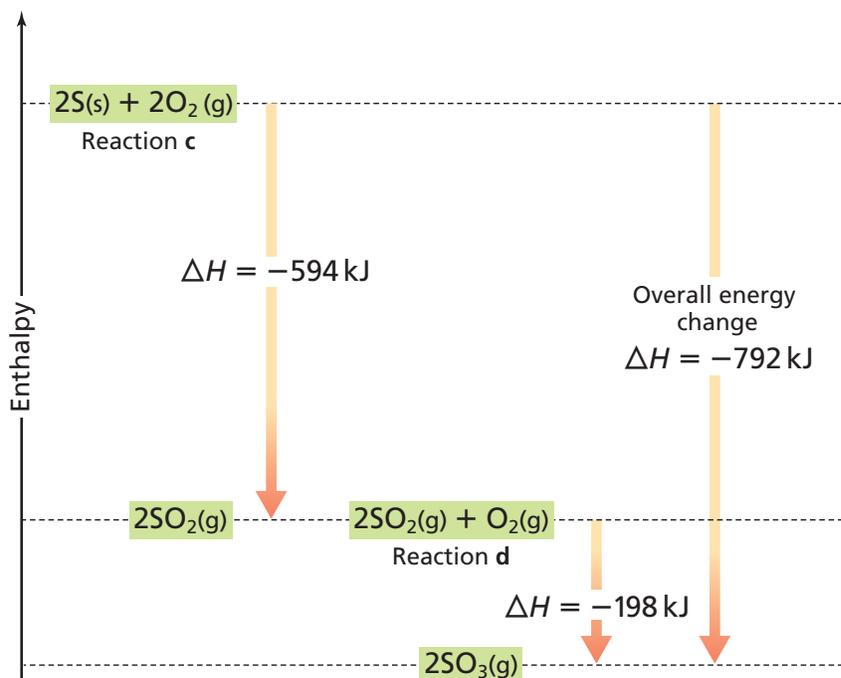


Figure 16-12

A volcanic eruption releases solid materials, gases, and heat into the atmosphere. Carbon dioxide and water account for more than 90% of the gases emitted. An eruption can send as much as 20 million tons of sulfur dioxide and sulfur trioxide into the stratosphere.

Figure 16-13

The arrow on the left indicates the release of 594 kJ as S and O_2 react to form SO_2 (Reaction **c**). Then, SO_2 and O_2 react to form SO_3 (Reaction **d**) with the release of 198 kJ (middle arrow). The overall energy change (the sum of the two processes) is shown by the arrow on the right. What is the enthalpy change for the conversion of S and O_2 to SO_3 ?

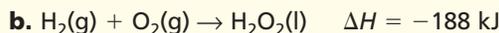
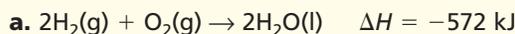
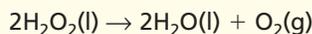
EXAMPLE PROBLEM 16-5



The photo shows an X-15 rocket-powered aircraft used to test the effects of high-speed, high-altitude flight on humans and materials. The rocket engines, mounted under the wing of a B52 bomber, are powered by the reaction of H_2 and O_2 . Data from 199 runs made by such aircraft were used in the development of the space shuttle.

Applying Hess's Law

Use thermochemical Equations **a** and **b** to determine ΔH for the decomposition of hydrogen peroxide (H_2O_2), a compound that has uses ranging from bleaching hair to energizing rocket engines.



1. Analyze the Problem

You have been given two chemical equations and their enthalpy changes. These two equations contain all the substances found in the desired equation.

Known

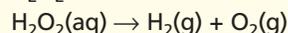
Equations **a** and **b** and their enthalpy changes

Unknown

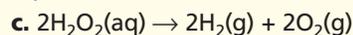
$$\Delta H = ? \text{ kJ}$$

2. Solve for the Unknown

H_2O_2 is a reactant, so reverse Equation **b**.



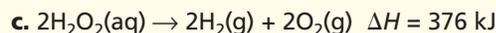
Two moles of H_2O_2 are needed. Multiply the equation by 2 to obtain Equation **c**.



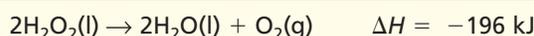
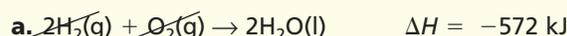
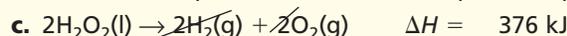
When you reverse an equation, you must change the sign of ΔH . When you double an equation, you must double ΔH .

$$\Delta H \text{ for Equation c} = -(\Delta H_{\text{Equation b}})(2)$$

$$\Delta H \text{ for Equation c} = -(-188 \text{ kJ})(2) = 376 \text{ kJ}$$



Add Equations **a** and **c** canceling any terms common to both sides of the combined equation. Add the enthalpies of Equations **a** and **c**.

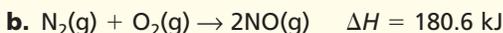


3. Evaluate the Answer

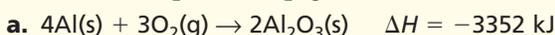
All values are accurate to the units place, so ΔH is correctly stated. The two equations produce the desired equation.

PRACTICE PROBLEMS

28. Use reactions **a** and **b** to determine ΔH for the following reaction.



29. Use reactions **a** and **b** to determine ΔH for the following reaction.



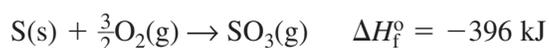
Practice!

For more practice using Hess's law, go to **Supplemental Practice Problems** in Appendix A.

Standard Enthalpy (Heat) of Formation

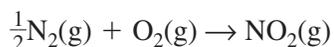
You have seen the usefulness of Hess's law in allowing you to calculate unknown ΔH values using known reactions and their experimentally determined ΔH values. However, recording and storing ΔH values for all known chemical reactions would be a huge and constantly evolving task. To avoid this problem, scientists have chosen to record and use enthalpy changes for one type of reaction—a reaction in which a compound is formed from its constituent elements in their standard states. The standard state of a substance means the normal physical state of the substance at one atmosphere pressure and 298 K (25°C). For example, in its standard state, iron is a solid, mercury is a liquid, and oxygen is a diatomic gas.

The ΔH value for such reactions is called the standard enthalpy (heat) of formation of the compound. The **standard enthalpy (heat) of formation** (ΔH_f°) is defined as the change in enthalpy that accompanies the formation of one mole of the compound in its standard state from its constituent elements in their standard states. A typical standard heat of formation reaction is the formation of one mole of SO_3 from its elements.



Where do standard heats of formation come from? When you state the height of a mountain, you do so relative to some point of reference—usually sea level. In a similar way, standard enthalpies of formation are stated based on the following arbitrary standard: Every free element in its standard state is assigned a ΔH_f° of exactly 0.0 kJ. With zero as the starting point, the experimentally determined enthalpies of formation of compounds can be placed on a scale above and below the elements in their standard states. Think of the zero of the enthalpy scale as being similar to the arbitrary assignment of 0.0°C to the freezing point of water. All substances warmer than freezing water have a temperature above zero. All substances colder than freezing water have a temperature below zero.

Standard enthalpies of formation of many compounds have been measured experimentally. For example, consider the equation for the formation of nitrogen dioxide.



The elements nitrogen and oxygen are diatomic gases in their standard states, so their standard enthalpies of formation are zero. When nitrogen and oxygen gases react to form one mole of nitrogen dioxide, the experimentally determined ΔH for the reaction is 33.2 kJ. That means that 33.2 kJ of energy is absorbed in this endothermic reaction. Thus, the energy content of the product, NO_2 , is 33.2 kJ greater than the energy content of the reactants. On a scale on which the ΔH_f° of reactants is defined as zero, ΔH_f° of $\text{NO}_2(\text{g})$ is +33.2 kJ. **Figure 16-14** shows that on the scale of standard enthalpies of formation, NO_2 is placed 33.2 kJ above the elements from which it was formed. Sulfur trioxide (SO_3) is placed 396 kJ below zero on the scale because the formation of $\text{SO}_3(\text{g})$ is an exothermic reaction. The energy content of the product $\text{SO}_3(\text{g})$ is 396 kJ less than the energy content of the elements from which it was formed.

Table 16-7 on the next page lists standard enthalpies of formation for some common compounds. A more complete list is in Appendix C, **Table C-13**.

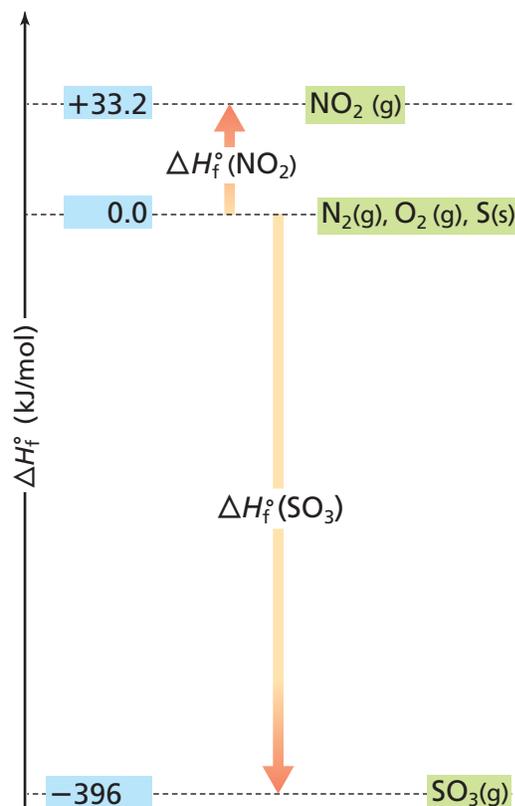


Figure 16-14

Note ΔH_f° for the elements N_2 , O_2 , and S is 0.0 kJ. When N_2 and O_2 react to form 1 mole of NO_2 , 33.2 kJ is absorbed. One mole of NO_2 has 33.2 kJ/mol more energy than the elements from which it is formed. Thus, ΔH_f° for NO_2 is +33.2 kJ/mol. When S and O_2 react to form one mole of SO_3 , 396 kJ is released. Thus, ΔH_f° for SO_3 is -396 kJ/mol.

Table 16-7

Standard Enthalpies of Formation for Selected Compounds		
Compound	Formation equation	ΔH_f° (kJ/mol)
NO ₂ (g)	$\frac{1}{2}\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$	33
H ₂ S(g)	$\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g})$	-21
NH ₃ (g)	$\frac{1}{2}\text{N}_2(\text{g}) + \frac{3}{2}\text{H}_2(\text{g}) \rightarrow \text{NH}_3(\text{g})$	-46
CH ₄ (g)	$\text{C}(\text{s, graphite}) + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_4(\text{g})$	-75
HF(g)	$\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{F}_2(\text{g}) \rightarrow \text{HF}(\text{g})$	-273
CCl ₄ (l)	$\text{C}(\text{s, graphite}) + 2\text{Cl}_2(\text{g}) \rightarrow \text{CCl}_4(\text{l})$	-128
H ₂ O(l)	$\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$	-286
SO ₂ (g)	$\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$	-297
CO ₂ (g)	$\text{C}(\text{s, graphite}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$	-394
Fe ₂ O ₃ (s)	$2\text{Fe}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s})$	-824
SF ₆ (g)	$\text{S}(\text{s}) + 3\text{F}_2(\text{g}) \rightarrow \text{SF}_6(\text{g})$	-1220
Al ₂ O ₃ (s)	$2\text{Al}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{Al}_2\text{O}_3(\text{s})$	-1680



Figure 16-15

Sulfur hexafluoride, one of the heaviest known gases, is used in this electrical substation to insulate circuit breakers, transformers, and other electrical equipment including high voltage transmission lines.

Using standard enthalpies of formation What is the significance of standard enthalpies of formation? How are they used? Standard enthalpies of formation provide useful data for calculating the enthalpies of reactions under standard conditions ($\Delta H_{\text{rxn}}^\circ$) using Hess's law. Suppose you want to calculate $\Delta H_{\text{rxn}}^\circ$ for this reaction, in which sulfur hexafluoride is produced. Sulfur hexafluoride is a stable, unreactive gas with some interesting applications, one of which is shown in **Figure 16-15**.



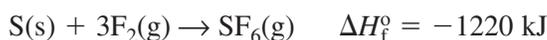
Recall that Hess's law allows you to combine equations and their ΔH values to produce the desired equation and its ΔH value. To apply Hess's law using standard enthalpies of formation data, you must have one equation for the formation of each compound in the desired equation. You can find these in **Table 16-7**. Three compounds are in the equation, HF, SF₆, and H₂S, so the three equations are these.

- $\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{F}_2(\text{g}) \rightarrow \text{HF}(\text{g}) \quad \Delta H_f^\circ = -273 \text{ kJ}$
- $\text{S}(\text{s}) + 3\text{F}_2(\text{g}) \rightarrow \text{SF}_6(\text{g}) \quad \Delta H_f^\circ = -1220 \text{ kJ}$
- $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g}) \quad \Delta H_f^\circ = -21 \text{ kJ}$

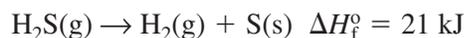
Equations **a** and **b** are for the formation of the products HF and SF₆. Therefore, they can be used in the direction in which they are written. But because two moles of HF are required, Equation **a** and its enthalpy must be multiplied by 2.



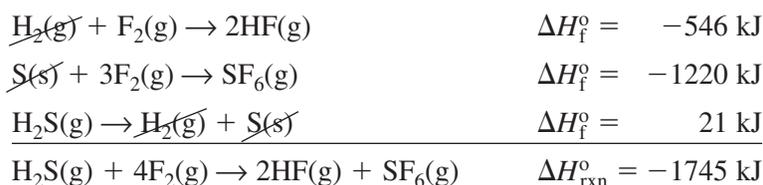
Equation **b** can be used as it is.



In the desired equation, H₂S is a reactant rather than a product. Therefore, Equation **c** must be reversed. Recall that reversing a reaction changes the sign of ΔH , so the sign of ΔH for Equation **c** becomes positive.



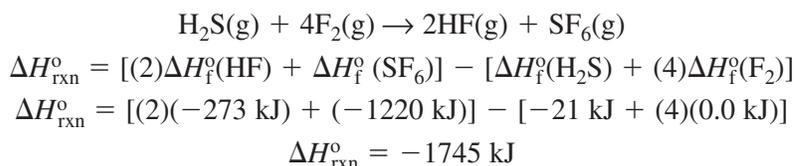
The three equations and their enthalpies can now be added. The elements H_2 and S cancel.



This example shows how standard heats of formation equations combine to produce the desired equation and its $\Delta H_{\text{rxn}}^\circ$. The entire procedure is summed up in the following formula.

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$$

The symbol Σ means “to take the sum of the terms.” The formula says to subtract the sum of heats of formation of the reactants from the sum of the heats of formation of the products. You can see how this formula applies to the reaction between hydrogen sulfide and fluorine.



Note that the heat of formation of HF is multiplied by 2 because two moles of HF are formed. Also note that it is not necessary to change the sign of $\Delta H_f^\circ(\text{H}_2\text{S})$ and that $\Delta H_f^\circ(\text{F}_2)$ is 0.0 kJ because the standard heat of formation of an element in its standard state is zero.

EXAMPLE PROBLEM 16-6

Enthalpy Change from Standard Enthalpies of Formation

Use standard enthalpies of formation to calculate $\Delta H_{\text{rxn}}^\circ$ for the combustion of methane. Methane is found in the atmosphere of the planet Pluto. $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$

1. Analyze the Problem

You are given an equation and asked to calculate the change in enthalpy. The formula $\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$ can be used with data from Table 16-7.

Known

$$\begin{aligned} \Delta H_f^\circ(\text{CO}_2) &= -394 \text{ kJ} \\ \Delta H_f^\circ(\text{H}_2\text{O}) &= -286 \text{ kJ} \\ \Delta H_f^\circ(\text{CH}_4) &= -75 \text{ kJ} \\ \Delta H_f^\circ(\text{O}_2) &= 0.0 \text{ kJ} \end{aligned}$$

Unknown

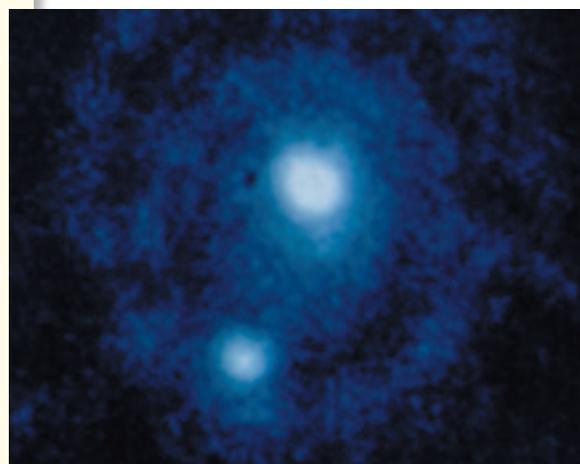
$$\Delta H_{\text{rxn}}^\circ = ? \text{ kJ}$$

2. Solve for the Unknown

Use the formula $\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$.

Expand the formula to include a term for each reactant and product. Multiply each term by the coefficient of the substance in the balanced chemical equation.

Continued on next page



Pluto, the smallest of the Sun's planets and the farthest, on average, from the Sun, is shown with its moon Charon. Methane has been found in the thin atmosphere of Pluto.

$$\Delta H_{\text{rxn}}^{\circ} = [\Delta H_f^{\circ}(\text{CO}_2) + (2)\Delta H_f^{\circ}(\text{H}_2\text{O})] - [\Delta H_f^{\circ}(\text{CH}_4) + (2)\Delta H_f^{\circ}(\text{O}_2)]$$

Substitute the ΔH_f° values into the formula.

$$\Delta H_{\text{rxn}}^{\circ} = [(-394 \text{ kJ}) + (2)(-286 \text{ kJ})] - [(-75 \text{ kJ}) + (2)(0.0 \text{ kJ})]$$

$$\Delta H_{\text{rxn}}^{\circ} = [-966 \text{ kJ}] - [-75 \text{ kJ}] = -966 \text{ kJ} + 75 \text{ kJ} = -891 \text{ kJ}$$

The combustion of 1 mol CH_4 releases 891 kJ.

3. Evaluate the Answer

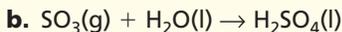
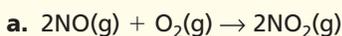
All values are accurate to the units place. Therefore, the answer is correct as stated. The calculated value is the same as that given in **Table 16-5** on page 501.



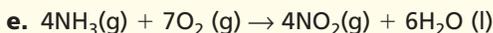
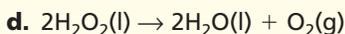
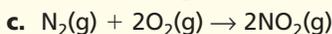
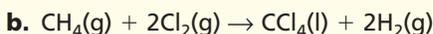
For more practice using standard enthalpies of formation, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

30. Show how the sum of enthalpy of formation equations produces each of the following reactions. You need not look up and include ΔH values.



31. Use standard enthalpies of formation from Appendix C, **Table C-13** to calculate $\Delta H_{\text{rxn}}^{\circ}$ for each of the following reactions.



You can see how valuable enthalpy of formation data are to a chemist. The $\Delta H_{\text{rxn}}^{\circ}$ for any reaction can be calculated if ΔH_f° is known for all of the compounds among the reactants and products.

Section 16.4 Assessment

- 32.** Explain what is meant by Hess's law and how it is used.
- 33.** What formula can be applied to determining $\Delta H_{\text{rxn}}^{\circ}$ when using Hess's law? Explain the formula in words.
- 34.** On the scale of standard enthalpies of formations, how are the elements in their standard states defined?
- 35.** Low energy is associated with stability. Examine the data in **Table 16-7**. What conclusion can you draw about the stabilities of the compounds listed relative to the elements in their standard states?
- 36. Thinking Critically** Could the absolute enthalpy or heat content of the elements at 298°C and one atmosphere pressure actually be 0.0 kJ? Explain why or why not.
- 37. Interpreting Scientific Illustrations** Using the data below, draw a diagram similar to **Figure 16-14** and use the diagram to determine the heat of vaporization of water at 298 K.
- Liquid water: $\Delta H_f^{\circ} = -285.8 \text{ kJ/mol}$
 Gaseous water: $\Delta H_f^{\circ} = -241.8 \text{ kJ/mol}$

In **Figure 16-16**, you can see a familiar picture of what happens to an iron object when it's left outdoors in moist air. Iron rusts slowly according to the following equation. It's the same chemical reaction that occurs in the heat pack you learned about earlier in the chapter.

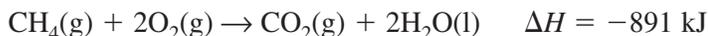


The heat pack goes into action the moment you activate it by ripping off the plastic covering. Similarly, unprotected iron objects rust whether you want them to or not.

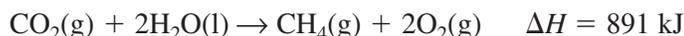
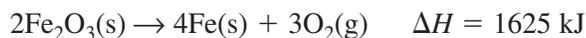
Spontaneous Processes

Rusting, like many other processes, is spontaneous. A **spontaneous process** is a physical or chemical change that occurs with no outside intervention. However, for many spontaneous processes, some energy must be supplied to get the process started. For example, you may have used a sparker or a match to light a Bunsen burner in your school lab. Or perhaps you are familiar with the continuously burning pilot lights that are used in gas stoves and furnaces to start these appliances immediately. Once the gas has been ignited, the combustion process can proceed spontaneously.

You saw in an earlier problem that the thermochemical equation for the combustion of methane, the major component of natural gas, is this

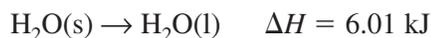


Suppose you reverse the direction of this equation and the equation for the rusting of iron. Recall that when you change the direction of a reaction, the sign of ΔH changes. Both reactions become endothermic.



Reversing the equation will not make rust decompose spontaneously into iron and oxygen under ordinary conditions. And carbon dioxide will not react with water to form methane and oxygen. These two equations represent reactions that are not spontaneous.

Do you notice a correlation? Iron rusting and methane burning are exothermic and spontaneous. The reverse reactions are endothermic and nonspontaneous. Based upon reactions such as these, some nineteenth-century scientists concluded that all exothermic processes are spontaneous and all endothermic processes are nonspontaneous. However, you need not look far for evidence that this conclusion is incorrect. For example, you know that ice melts at room temperature. That's a spontaneous, endothermic process.



It's evident that something other than ΔH plays a role in determining whether a chemical process occurs spontaneously under a given set of conditions. That something is called entropy.

Objectives

- **Differentiate** between spontaneous and nonspontaneous processes.
- **Explain** how changes in entropy and free energy determine the spontaneity of chemical reactions and other processes.

Vocabulary

spontaneous process
entropy
law of disorder
free energy

Figure 16-16

Year by year, the iron in this abandoned railroad equipment in the Alaskan tundra combines with oxygen to form rust, Fe_2O_3 .



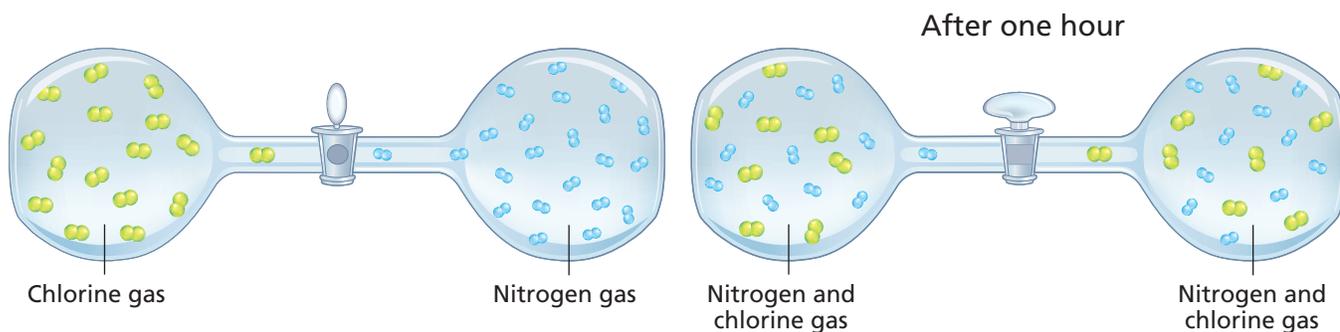


Figure 16-17

At the beginning of the experiment, chlorine and nitrogen molecules occupy separate bulbs. But when the stopcock is opened, the gas molecules move back and forth between the two bulbs and become thoroughly mixed. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

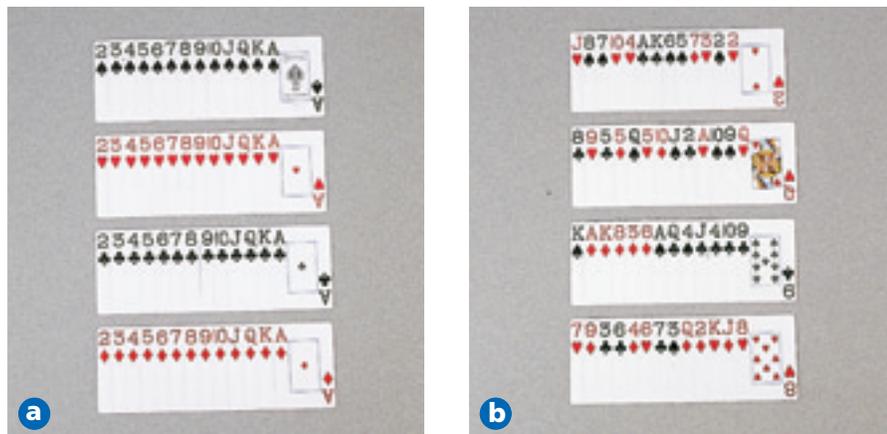
What is entropy? Think about what might happen when a glass bulb containing chlorine gas is connected to a similar glass bulb containing nitrogen gas, as shown in **Figure 16-17**. When the stopcock in the tube connecting the two bulbs is opened, the gases can pass freely from one bulb to the other. After an hour, you can see that the chlorine and nitrogen molecules are distributed uniformly between the two bulbs. You're probably not surprised at that result. After all, the delicious smell of brownies baking in the kitchen doesn't stay in the kitchen. It wafts to wherever you are. So, whether it's the aroma of brownies or the chlorine and nitrogen in the experiment, gases tend to mix. Why? The answer is that molecules are more likely to exist in a high state of disorder (mixed) than in a low state of disorder (unmixed). In thermodynamics, the term for disorder is *entropy*. **Entropy** (S) is a measure of the disorder or randomness of the particles that make up a system.

The tendency toward disorder or randomness is summarized in the **law of disorder**, which states that spontaneous processes always proceed in such a way that the entropy of the universe increases. This law also is called the second law of thermodynamics. You can observe the law of disorder in many everyday situations. For example, suppose you shuffle a deck of 52 playing cards thoroughly and lay them down one after another to form four rows of 13 cards each. The cards might form the highly ordered arrangement shown in **Figure 16-18a**. However, the probability that you will lay down this exact, low-entropy sequence is one chance in 8.07×10^{67} . In other words, it's nearly impossible. Because there are so many possible highly disordered arrangements, it's almost infinitely more likely that you'll produce a high-entropy sequence such as the one shown in **Figure 16-18b**.

Predicting changes in entropy Recall that the change in enthalpy for a reaction is equal to the enthalpy of the products minus the enthalpy of the reactants. The change in entropy (ΔS) during a reaction or process is similar.

Figure 16-18

What are the chances that a shuffled deck of cards will lay out into four perfectly ordered rows as shown in **a**? Experience tells you that the chances for an orderly arrangement are almost negligible, and that the random arrangement in **b** is far more likely.

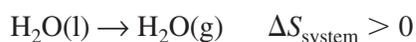


$$\Delta S_{\text{system}} = S_{\text{products}} - S_{\text{reactants}}$$

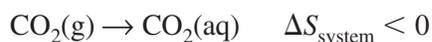
Therefore, if the entropy of a system increases during a reaction or process, $S_{\text{products}} > S_{\text{reactants}}$ and ΔS_{system} is positive. Conversely, if the entropy of a system decreases during a reaction or process, $S_{\text{products}} < S_{\text{reactants}}$ and ΔS_{system} is negative.

Can you predict if ΔS_{system} is positive or negative by examining the equation for a reaction or process? In some cases you can.

1. *Entropy changes associated with changes in state, such as those shown for water in Figure 16-19, can be predicted.* In solids, molecules are tightly packed and have limited movement, but they have some freedom to move in liquids. In gases, molecules are unrestricted in their movements except by the container in which they are held. Therefore, entropy increases as a substance changes from a solid to a liquid and from a liquid to a gas. For example, the entropy of the system increases (ΔS_{system} is positive) as water vaporizes and as methanol melts.



2. *The dissolving of a gas in a solvent always results in a decrease in entropy.* Gas particles have more entropy when they can move freely in the gaseous state than when they are dissolved in a liquid or solid that limits their movements and randomness. For example, ΔS_{system} is negative for the dissolving of carbon dioxide in water.



3. *Assuming no change in physical state, the entropy of a system usually increases when the number of gaseous product particles is greater than the number of gaseous reactant particles.* That's because the larger the number of gaseous particles, the more random arrangements are available. For the following reaction, ΔS_{system} is positive because two gaseous molecules react and three gaseous molecules are produced.

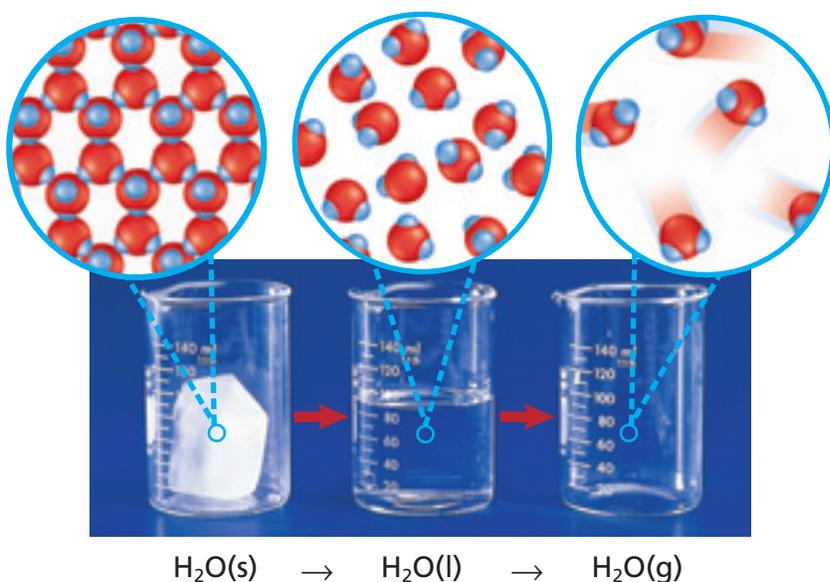
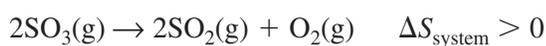
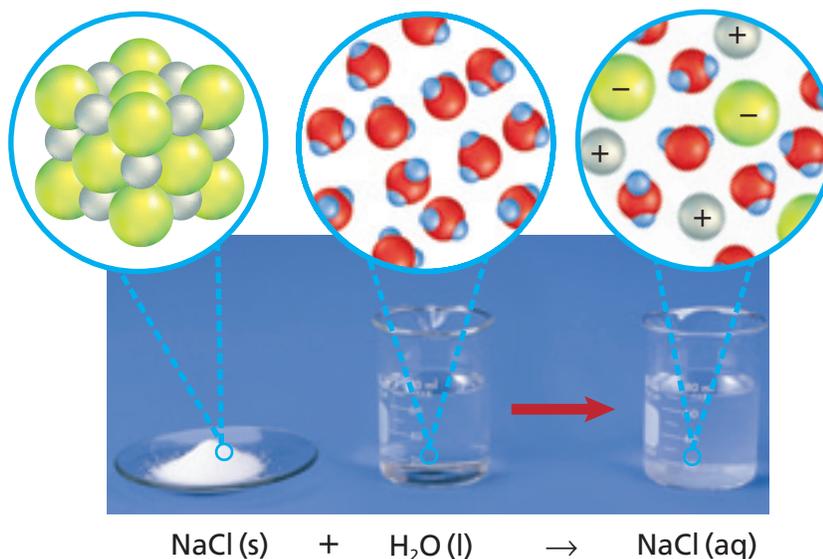


Figure 16-19

The molecules in an ice cube are in a rigid, orderly arrangement. The molecules in liquid water have some freedom to move and create different arrangements. Water molecules in the gas phase are completely separated and can create an almost infinite number of arrangements.

Figure 16-20

Sodium chloride and liquid water are pure substances each with a degree of orderliness. When sodium chloride dissolves in water, the entropy of the system increases because Na^+ and Cl^- ions and water molecules mix together to create a large number of random arrangements.



4. *With some exceptions, you can predict the change in entropy when a solid or a liquid dissolves to form a solution.* The solute particles, which are separate and pure before dissolving, become dispersed throughout the solvent. Therefore, dissolution usually increases the randomness and disorder of the particles, as shown in **Figure 16-20**, and the entropy of the system increases. For the dissolving of sodium chloride in water, ΔS_{system} is positive.



5. *An increase in the temperature of a substance is always accompanied by an increase in the random motion of its particles.* Recall that the kinetic energy of molecules increases with temperature. Increased kinetic energy means faster movement, more possible arrangements, and increased disorder. Therefore, the entropy of any substance increases as its temperature increases, and $\Delta S_{\text{system}} > 0$.

Practice!

For more practice predicting the sign of ΔS , go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

38. Predict the sign of ΔS_{system} for each of the following changes.
- | | |
|--|--|
| a. $\text{ClF(g)} + \text{F}_2(\text{g}) \rightarrow \text{ClF}_3(\text{g})$ | c. $\text{CH}_3\text{OH(l)} \rightarrow \text{CH}_3\text{OH(aq)}$ |
| b. $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{aq})$ | d. $\text{C}_{10}\text{H}_8(\text{l}) \rightarrow \text{C}_{10}\text{H}_8(\text{s})$ |

Entropy, the Universe, and Free Energy

You may be wondering whether entropy has an effect on reaction spontaneity. Recall that the law of disorder states that the entropy of the universe must increase as a result of a spontaneous reaction or process. Therefore, the following is true for any spontaneous process.

$$\Delta S_{\text{universe}} > 0$$

Because the universe equals the system plus the surroundings, any change in the entropy of the universe is the sum of changes occurring in the system and surroundings.

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

CONNECTION

Volcanoes, fumaroles, hot springs, geysers, and boiling mud pools are manifestations of the great amount of geothermal energy in Earth's interior. Volcanoes are vents in Earth's crust from which flow molten rock (magma), steam, and other materials. When surface water moves downward through Earth's crust, it can interact with magma and/or hot rocks. Water that comes back to the surface in hot springs is heated to temperatures much higher than the surrounding air temperatures. Geysers are actually hot springs that spout hot water and steam into the air. Fumaroles emit steam and other gases such as hydrogen sulfide.

Geothermal energy provides an inexhaustible supply of energy that has been harnessed for a variety of applications. Buildings, homes, and greenhouses are warmed with energy from these sources. Geothermal energy also is used to warm soil to increase crop production in cooler climates. In several countries, electricity is generated by steam from geysers.



How do changes in a system's enthalpy and entropy affect $\Delta S_{\text{universe}}$? In nature, $\Delta S_{\text{universe}}$ tends to be positive for reactions and processes under the following conditions.

1. The reaction or process is exothermic, which means ΔH_{system} is negative. The heat released by an exothermic reaction raises the temperature of the surroundings and thereby increases the entropy of the surroundings. $\Delta S_{\text{surroundings}}$ is positive.
2. The entropy of the system increases, so ΔS_{system} is positive.

Thus, exothermic chemical reactions accompanied by an increase in entropy are all spontaneous.

Free energy Can you definitely determine if a reaction is spontaneous? In 1878, J. Willard Gibbs, a physicist at Yale University, defined a combined enthalpy-entropy function called Gibbs free energy that answers that question. For reactions or processes that take place at constant pressure and temperature, Gibbs free energy (G_{system}), commonly called **free energy**, is energy that is available to do work. Thus, free energy is useful energy. In contrast, some entropy is associated with energy that is spread out into the surroundings as, for example, random molecular motion and cannot be recovered to do useful work. The free energy change (ΔG_{system}) is the difference between the system's change in enthalpy (ΔH_{system}) and the product of the Kelvin temperature and the change in entropy ($T\Delta S_{\text{system}}$).

$$\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

When a reaction or process occurs under standard conditions (298 K and one atmosphere pressure) the standard free energy change can be expressed by this equation.

$$\Delta G_{\text{system}}^{\circ} = \Delta H_{\text{system}}^{\circ} - T\Delta S_{\text{system}}^{\circ}$$

The sign of the free energy change of the system, ΔG_{system} , tells you whether or not a reaction or process is spontaneous at a constant specified temperature and pressure. If the sign of the free energy change of the system is negative, the reaction is spontaneous. If the sign of the free energy change is positive, the reaction is nonspontaneous. The relationship between ΔG_{system} , $\Delta S_{\text{universe}}$, and reaction spontaneity is summarized in **Table 16-8**.

Calculating free energy change How do changes in enthalpy and entropy affect free energy change and spontaneity for the reaction between nitrogen and hydrogen to form ammonia?



Note that the entropy of the system decreases because four moles of gaseous molecules react and only two moles of gaseous molecules are produced. Therefore, $\Delta S_{\text{system}}^{\circ}$ is negative. A decrease in the entropy of the system tends

Table 16-8

ΔG_{system} and Reaction Spontaneity		
Type of reaction or process	ΔG_{system}	$\Delta S_{\text{universe}}$
Spontaneous	negative	positive
Nonspontaneous	positive	negative

to make the reaction nonspontaneous. But the reaction is exothermic ($\Delta H_{\text{system}}^{\circ}$ is negative), which tends to make the reaction spontaneous. To determine which of the two conflicting tendencies predominates, you must calculate $\Delta G_{\text{system}}^{\circ}$ for the reaction. First, convert $\Delta H_{\text{system}}^{\circ}$ to joules.

$$\Delta H_{\text{system}}^{\circ} = -91.8 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = -91\,800 \text{ J}$$

Now, substitute $\Delta H_{\text{system}}^{\circ}$, T , and $\Delta S_{\text{system}}^{\circ}$ into the defining equation for $\Delta G_{\text{system}}^{\circ}$.

$$\begin{aligned}\Delta G_{\text{system}}^{\circ} &= \Delta H_{\text{system}}^{\circ} - T\Delta S_{\text{system}}^{\circ} \\ \Delta G_{\text{system}}^{\circ} &= -91\,800 \text{ J} - (298 \text{ K})(-197 \text{ J/K}) \\ \Delta G_{\text{system}}^{\circ} &= -91\,800 \text{ J} + 58\,700 \text{ J} = -33\,100 \text{ J}\end{aligned}$$

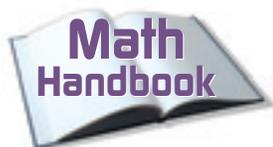
Because $\Delta G_{\text{system}}^{\circ}$ for this reaction is negative, the reaction is spontaneous under standard conditions.

The reaction between nitrogen and hydrogen demonstrates that the entropy of a system may actually decrease during a spontaneous process. However, it can do so only if the entropy of the surroundings increases more than the entropy of the system decreases. The situation is analogous to that of a company that manufactures two products, A and B. The company can operate profitably if it loses money on product A, but only if it earns more money on product B than it loses on product A. **Table 16-9** shows how reaction spontaneity depends on the signs of ΔH_{system} and ΔS_{system} .

If $\Delta G_{\text{system}}^{\circ} = 0$, both reactants and products are present in a state known as chemical equilibrium. Chemical equilibrium describes a state in which the rate of the forward reaction equals the rate of the reverse reaction. You will learn more about chemical equilibrium in Chapter 18.

Table 16-9

How ΔH_{system} and ΔS_{system} Affect Reaction Spontaneity		
	$-\Delta H_{\text{system}}$	$+\Delta H_{\text{system}}$
$+\Delta S_{\text{system}}$	Always spontaneous	Spontaneity depends upon temperature
$-\Delta S_{\text{system}}$	Spontaneity depends upon temperature	Never spontaneous



Review solving algebraic equations in the **Math Handbook** on page 897 of this text.

EXAMPLE PROBLEM 16-7

Determining Reaction Spontaneity

For a process, $\Delta H_{\text{system}} = 145 \text{ kJ}$ and $\Delta S_{\text{system}} = 322 \text{ J/K}$. Is the process spontaneous at 382 K?

1. Analyze the Problem

You are given ΔH_{system} and ΔS_{system} for a process and must calculate ΔG_{system} to determine its sign.

Known

$$T = 382 \text{ K}$$

$$\Delta H_{\text{system}} = 145 \text{ kJ}$$

$$\Delta S_{\text{system}} = 322 \text{ J/K}$$

Unknown

$$\text{sign of } \Delta G_{\text{system}} = ?$$



2. Solve for the Unknown

Convert ΔH_{system} , given in kJ, to joules.

$$145 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 145\,000 \text{ J}$$

Solve the free energy equation using the known values.

$$\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

$$\Delta G_{\text{system}} = 145\,000 \text{ J} - (382 \text{ K})(322 \text{ J/K})$$

$$\Delta G_{\text{system}} = 145\,000 \text{ J} - 123\,000 \text{ J} = 22\,000 \text{ J}$$

Because ΔG_{system} is positive, the reaction is nonspontaneous.

3. Evaluate the Answer

The given data have three significant figures. When ΔH_{system} and ΔS_{system} are expressed in joules, they are accurate to the nearest 1000. Therefore, ΔG_{system} is also accurate to the nearest 1000, so the answer is correctly stated. When both ΔH_{system} and ΔS_{system} are positive, ΔG_{system} is positive only if the temperature is low enough. The temperature (109°C) is not high enough to make the second term of the equation greater than the first and so ΔG_{system} is positive.

PRACTICE PROBLEMS

39. Given ΔH_{system} , T , and ΔS_{system} , determine if each of the following processes or reactions is spontaneous or nonspontaneous.

- $\Delta H_{\text{system}} = -75.9 \text{ kJ}$, $T = 273 \text{ K}$, $\Delta S_{\text{system}} = 138 \text{ J/K}$
- $\Delta H_{\text{system}} = -27.6 \text{ kJ}$, $T = 535 \text{ K}$, $\Delta S_{\text{system}} = -55.2 \text{ J/K}$
- $\Delta H_{\text{system}} = 365 \text{ kJ}$, $T = 388 \text{ K}$, $\Delta S_{\text{system}} = -55.2 \text{ J/K}$



For more practice predicting the spontaneity of a reaction, go to **Supplemental Practice Problems** in Appendix A.

Coupled reactions Many of the chemical reactions that enable plants and animals to live and grow are nonspontaneous, that is, ΔG_{system} is positive. Based on that information, you might ask why these nonspontaneous reactions occur so readily in nature. In living systems, nonspontaneous reactions often occur in conjunction with other reactions that are spontaneous (reactions for which ΔG_{system} is negative). Reactions of this type that occur together are called coupled reactions. In coupled reactions, free energy released by one or more spontaneous reactions is used to drive a nonspontaneous reaction.

Section 16.5 Assessment

- In terms of energy, explain the difference between a spontaneous and a nonspontaneous reaction.
- If a system becomes more disordered during a process, how does the system's entropy change?
- When you dissolve a teaspoonful of sugar in a cup of tea, does the entropy of the system increase or decrease? Define the system and explain your answer.
- Thinking Critically** Evaluate the following statement and explain why it is true or false: The law of disorder means that the entropy of a system can never decrease during a spontaneous reaction or process.
- Predicting** Predict the sign of ΔS_{system} for the following reaction. Explain the basis for your prediction.
$$2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$$

Calorimetry

In this laboratory investigation, you will use the methods of calorimetry to approximate the amount of energy contained in a potato chip. The burning of a potato chip releases heat stored in the substances contained in the chip. The heat will be absorbed by a mass of water.

Problem

How many Calories of energy does the potato chip contain? How can the experiment be improved to provide a more accurate answer?

Objectives

- **Identify** the reactants and products in the reaction.
- **Measure** mass and temperature in order to calculate the amount of heat released in the reaction.
- **Propose** changes in the procedure and design of the equipment to decrease the percent error.

Materials

large potato chip
250-mL beaker
100-mL graduated cylinder
evaporating dish
thermometer
ring stand with ring
wire gauze
matches
stirring rod
balance

Safety Precautions



- Always wear safety goggles and a lab apron.
- Tie back long hair.
- Hot objects may not appear to be hot.
- Do not heat broken, chipped, or cracked glassware.
- Do not eat any items used in the lab.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Form a hypothesis about how the quantity of heat produced by the combustion reaction will compare with the quantity of heat absorbed by the water.
4. What formula will you use to calculate the quantity of heat absorbed by the water?
5. Assuming that the potato chip contains compounds made up of carbon and hydrogen, what gases will be produced in the combustion reaction?

Observations of the Burning of a Potato Chip

Mass of beaker and 50 mL of water	
Mass of empty beaker	
Mass of water in beaker	
Mass of potato chip	
Highest temperature of water	
Initial temperature of water	
Change in temperature	

Procedure

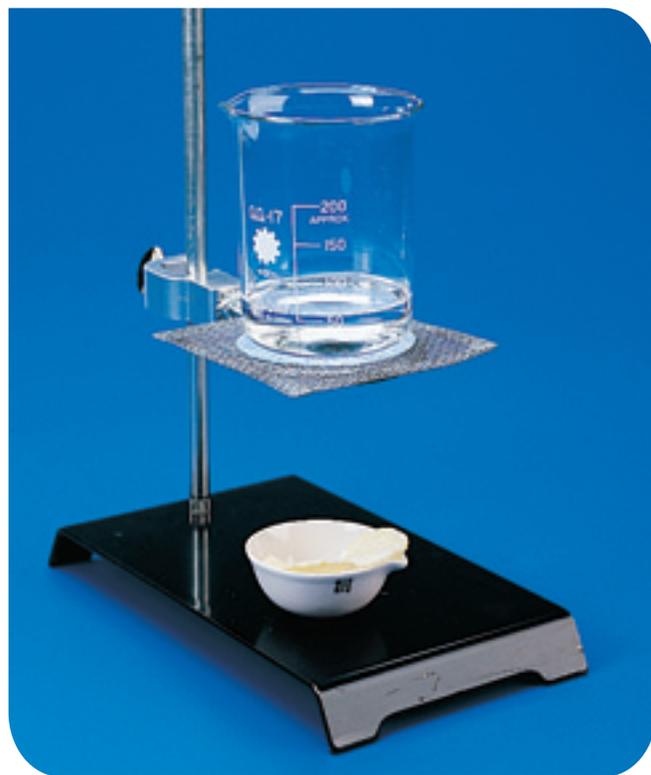
1. Measure the mass of a potato chip and record it in the data table.
2. Place the potato chip in an evaporating dish on the metal base of the ring stand. Position the ring and wire gauze so that they will be 10 cm above the top of the potato chip.
3. Measure the mass of an empty 250-mL beaker and record it in the data table.
4. Using the graduated cylinder, measure 50 mL of water and pour it into the beaker. Measure the mass of the beaker and water and record it in the data table.
5. Place the beaker on the wire gauze on the ring stand.
6. Measure and record the initial temperature of the water.
7. Use a match to ignite the bottom of the potato chip.
8. With a stirring rod, stir the water in the beaker while the chip burns. Measure the highest temperature of the water and record it in the data table.

Cleanup and Disposal

1. Clean all lab equipment and return it to its proper place.
2. Wash your hands thoroughly after all lab work and cleanup is complete.

Analyze and Conclude

1. **Classifying** Is the reaction exothermic or endothermic? Explain how you know.
2. **Observing and Inferring** Describe the reactant and products of the chemical reaction. Was the reactant (potato chip) completely consumed? What evidence supports your answer?
3. **Using Numbers** Calculate the mass of water in the beaker and the temperature change of the water. Use $q = c \times m \times \Delta T$ to calculate how much heat in joules was transferred to the water in the beaker by the burning of one chip.
4. **Using Numbers** Convert the quantity of heat in joules/chip to Calories/chip.



5. **Using Numbers** From the information on the chip's container, determine the mass in grams of one serving. Using your data, calculate the number of Calories that would be released by the combustion of one serving of chips.
6. **Error Analysis** Use the chip's container to determine how many Calories are contained in one serving. Compare your calculated Calories per serving with the value on the chip's container. Calculate the percent error.
7. **Observing and Inferring** Was all of the heat that was released collected by the water in the beaker? How can the experimental equipment be improved to decrease the percent error?

Real-World Chemistry

1. From the ingredients identified on the potato chip container, list the actual substances that burned to produce energy. Are there any ingredients that did not produce energy? Explain.
2. You have discovered that potato chips provide a significant number of Calories per serving. Would it be advisable to make potato chips a substantial part of your diet? Explain.

How It Works

Refrigerator

When you put soft drinks into a cooler containing ice, heat is absorbed from the soft drinks to melt the ice. The energy involved in the phase change from solid water to liquid water is responsible for cooling the soft drinks.

A refrigerator also uses the phase changes (evaporation and condensation) of a substance to move heat from the inside of the refrigerator to

the outside. A liquid called a refrigerant circulates through coils within the refrigerator and absorbs heat from the interior. The absorbed heat changes the refrigerant from a liquid to a gas. The gas is then compressed and circulated in condenser coils outside the refrigerator where it releases heat to the environment and is converted to a liquid.

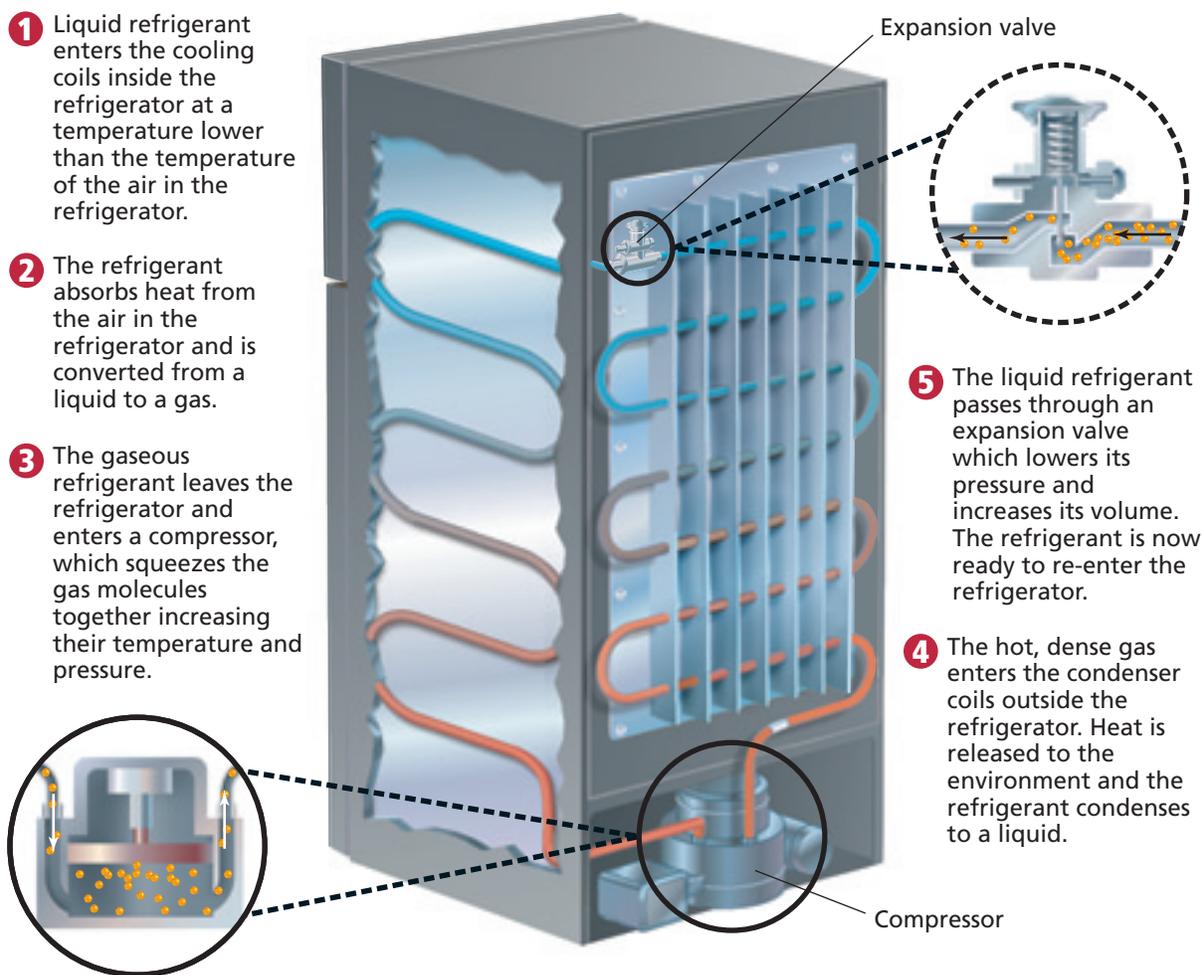
1 Liquid refrigerant enters the cooling coils inside the refrigerator at a temperature lower than the temperature of the air in the refrigerator.

2 The refrigerant absorbs heat from the air in the refrigerator and is converted from a liquid to a gas.

3 The gaseous refrigerant leaves the refrigerator and enters a compressor, which squeezes the gas molecules together increasing their temperature and pressure.

5 The liquid refrigerant passes through an expansion valve which lowers its pressure and increases its volume. The refrigerant is now ready to re-enter the refrigerator.

4 The hot, dense gas enters the condenser coils outside the refrigerator. Heat is released to the environment and the refrigerant condenses to a liquid.



Thinking Critically

1. Predicting What would happen if too much refrigerant was present in the system? Too little? Use phase changes to explain your predictions.

2. Hypothesizing A heat pump is a device that can either heat or cool a house. It is constructed in a fashion similar to a refrigerator. Explain how you think a heat pump works.

Summary

16.1 Energy

- Energy is the capacity to do work or to produce heat. Energy may change from one form to another, but the amount of energy does not change.
- Chemical potential energy is energy stored in chemical bonds of a substance by virtue of the arrangement of the atoms and molecules.
- Chemical potential energy is released or absorbed as heat during chemical processes or reactions.

16.2 Heat in Chemical Reactions and Processes

- In thermochemistry, the universe is defined as the system plus the surroundings. The system is the reaction or process that is being studied, and the surroundings include everything in the universe except the system.
- The heat lost or gained by a system during a reaction or process carried out at constant pressure is called the change in enthalpy (ΔH).
- When ΔH is positive, the reaction is endothermic. When ΔH is negative, the reaction is exothermic.

16.3 Thermochemical Equations

- A thermochemical equation includes the physical states of the reactants and products and specifies the change in enthalpy.

- The molar enthalpy (heat) of vaporization, ΔH_{vap} , is the amount of energy required to evaporate one mole of a liquid.
- The molar enthalpy (heat) of fusion, ΔH_{fus} , is the amount of energy needed to melt one mole of a solid.

16.4 Calculating Enthalpy Change

- Using Hess's law, the enthalpy change for a reaction can be calculated by adding two or more thermochemical equations and their enthalpy changes.
- Standard enthalpies of formation are based upon assigning a standard enthalpy of 0.0 kJ to all elements in their standard states.

16.5 Reaction Spontaneity

- Exothermic reactions tend to be spontaneous because they increase the entropy of the surroundings.
- Entropy is a measure of the disorder or randomness of the particles of a system. Spontaneous processes always result in an increase in the entropy of the universe.
- Free energy is the energy available to do work.
- If the change in a system's free energy is negative, the reaction is spontaneous. If the change in free energy is positive, the reaction is nonspontaneous.

Key Equations and Relationships

- $q = c \times m \times \Delta T$
(p. 493)
- $\Delta H_{\text{rxn}}^{\circ} = \sum \Delta H_{\text{f}}^{\circ}(\text{products}) - \sum \Delta H_{\text{f}}^{\circ}(\text{reactants})$
(p. 511)
- $\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$
(p. 517)

Vocabulary

- calorie (p. 491)
- calorimeter (p. 496)
- chemical potential energy (p. 490)
- energy (p. 489)
- enthalpy (p. 499)
- enthalpy (heat) of combustion (p. 501)
- enthalpy (heat) of reaction (p. 499)
- entropy (p. 514)
- free energy (p. 517)
- heat (p. 491)
- Hess's law (p. 506)
- joule (p. 491)
- law of conservation of energy (p. 490)
- law of disorder (p. 514)
- molar enthalpy (heat) of fusion (p. 502)
- molar enthalpy (heat) of vaporization (p. 502)
- specific heat (p. 492)
- spontaneous process (p. 513)
- standard enthalpy (heat) of formation (p. 509)
- surroundings (p. 498)
- system (p. 498)
- thermochemical equation (p. 501)
- thermochemistry (p. 498)
- universe (p. 498)

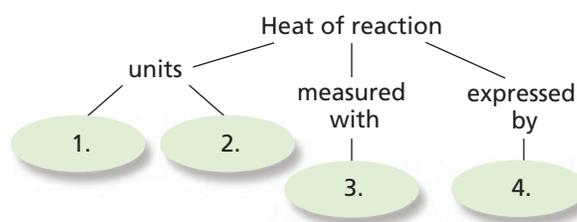


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 16 Assessment.

Concept Mapping

45. Fill in the boxes with the following terms: calorimeter, calorie, thermochemical equation, joule.



Mastering Concepts

46. Differentiate between potential energy and kinetic energy. (16.1)
47. How does the chemical potential energy of a system change during an endothermic reaction? (16.1)
48. How does the nutritional Calorie compare with the calorie? (16.1)
49. What quantity has the units $J/(g \cdot ^\circ C)$? (16.1)
50. Describe a situation that illustrates potential energy changing to kinetic energy. (16.1)
51. In describing a chemical reaction, what is meant by the system? The surroundings? (16.2)
52. Under what condition is the heat (q) evolved or absorbed in a chemical reaction equal to a change in enthalpy (ΔH)? (16.2)
53. The enthalpy change for a reaction, ΔH , is negative. What does this indicate about the chemical potential energy of the system before and after the reaction? (16.2)
54. How does the enthalpy of the products compare with the enthalpy of the reactants in an exothermic reaction? An endothermic reaction? (16.2)
55. What is the sign of ΔH for an exothermic reaction? An endothermic reaction? (16.2)
56. Explain why you need to know the specific heat of a substance in order to calculate how much heat is gained or lost by the substance as a result of a temperature change. (16.2)
57. How is the quantity of heat lost by the system related to the quantity of heat gained by the surroundings during an exothermic process? (16.2)
58. How is a thermochemical equation different from a balanced chemical equation? (16.3)
59. Which has the higher heat content, a mole of liquid water or a mole of ice? (16.3)
60. Write the correct sign of ΔH_{system} for each of the following changes in physical state. (16.3)
- $C_2H_5OH(s) \rightarrow C_2H_5OH(l)$
 - $H_2O(g) \rightarrow H_2O(l)$
 - $CH_3OH(l) \rightarrow CH_3OH(g)$
 - $NH_3(l) \rightarrow NH_3(s)$
61. How are the chemical elements in their standard states used as references for standard enthalpies of formation? (16.4)
62. Write the formula that can be used to calculate the enthalpy of a reaction from the enthalpies of formation of the reactants and products. (16.4)
63. Compare and contrast enthalpy and entropy. (16.5)
64. What does the entropy of a substance measure? (16.5)
65. What must happen to the entropy of the universe during a spontaneous reaction or process? (16.5)
66. From each pair, pick the one with the greater entropy. (16.5)
- $NH_3(g)$ and $NH_3(l)$
 - $NH_3(g)$ and $NH_3(aq)$
 - $CO_2(s)$ and $CO_2(g)$
 - $KBr(s)$ and $KBr(l)$
67. What is meant by the free energy of a system? (16.5)
68. What is the equation that defines free energy? (16.5)
69. How is the free energy change for a reaction related to its spontaneity? (16.5)
70. Explain how an exothermic reaction changes the entropy of the surroundings. Does such a reaction increase or decrease ΔG_{system} ? Explain your answer. (16.5)
71. Explain how an endothermic reaction changes the entropy of the surroundings. How does such a reaction affect ΔG_{system} ? (16.5)
72. Under what circumstance might a nonspontaneous reaction become spontaneous when the temperature increases? Decreases? (16.5)

- 73.** Explain why the gaseous state of a substance has a greater entropy than the liquid state of the same substance. (16.5)
- 74.** Predict the sign of ΔS_{system} for each of these chemical reactions. Explain your predictions. (16.5)
- $2\text{SO}_3(\text{g}) + \text{CO}_2(\text{g}) \rightarrow \text{CS}_2(\text{g}) + 4\text{O}_2(\text{g})$
 - $\text{PCl}_5(\text{g}) \rightarrow \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
 - $2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g})$
 - $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) \rightarrow 2\text{C}_2\text{H}_5\text{OH}(\text{l}) + 2\text{CO}_2(\text{g})$
 - $\text{H}_2\text{SO}_4(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{SO}_3(\text{g})$
- 86.** How much heat is required to warm 225 g of ice from -46.8°C to 0.0°C , melt the ice, warm the water from 0.0°C to 100.0°C , boil the water, and heat the steam to 173.0°C ?
- 87.** How much energy is involved in the dissolving of sodium hydroxide in water? The table shows data from an experiment in which a measured amount of NaOH is dissolved in water in a foam cup calorimeter. Calculate ΔH for the process in kJ/mol. If the enthalpy change for the solution of NaOH is -44.51 kJ/mol, what is the percent error of the experiment?

Mastering Problems

Energy (16.1)

- 75.** A fast-food item contains 544 nutritional Calories. Convert this energy to calories and to joules.
- 76.** An endothermic process absorbs 138 kJ. How many calories of heat are absorbed?
- 77.** An exothermic reaction releases 325 000 calories. Convert this energy to kJ.

Measuring Heat (16.2)

- 78.** How many joules of heat are lost by 3580 kg granite as it cools from 41.2°C to -12.9°C ?
- 79.** How much heat is absorbed by a 2000 kg granite boulder as energy from the sun causes its temperature to change from 10°C to 29°C ?
- 80.** A sample of silver with a mass of 63.3 g is heated to a temperature of 384.4 K and placed in a container of water at 290.0 K. The final temperature of the silver and water is 292.4 K. Assuming no heat loss, what mass of water was in the container? The specific heat of water is 4.184 J/(g \cdot $^\circ\text{C}$) and of silver, 0.24 J/(g \cdot $^\circ\text{C}$).
- 81.** A swimming pool, 20.0 m \times 12.5 m, is filled with water to a depth of 3.75 m. If the initial temperature of the water is 18.4°C , how much heat must be added to the water to raise its temperature to 29.0°C ? Assume that the density of water is 1.000 g/mL.

Calculating Energy Change (16.3)

- 82.** How much heat is required to vaporize 343 g of liquid ethanol at its boiling point? $\Delta H_{\text{vap}} = 38.6$ kJ/mol
- 83.** How much heat is evolved when 1255 g of water condenses to a liquid at 100°C ? $\Delta H_{\text{cond}} = -40.7$ kJ/mol
- 84.** How much heat is liberated by the combustion of 206 g of hydrogen? $\Delta H_{\text{comb}} = -286$ kJ/mol
- 85.** A sample of ammonia liberates 5.66 kJ of heat as it solidifies at its melting point. What is the mass of the sample? $\Delta H_{\text{solid}} = -5.66$ kJ/mol.

Table 16-10

Data for the Dissolving of NaOH	
Mass of NaOH + weighing paper	4.71 g
Mass of weighing paper	0.70 g
Mass of NaOH	
Mass of foam cup + water	109.85 g
Mass of foam cup	10.25 g
Mass of water	
Final temperature of water	35.9°C
Initial temperature of water	25.5°C
Change in temperature, ΔT	

Using Hess's Law (16.4)

- 88.** You are given these two equations.
- $$\text{Sn}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow \text{SnCl}_2(\text{s}) \quad \Delta H = -325 \text{ kJ}$$
- $$\text{SnCl}_2(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow \text{SnCl}_4(\text{l}) \quad \Delta H = -186 \text{ kJ}$$
- Calculate ΔH for this reaction.
- $$\text{Sn}(\text{s}) + 2\text{Cl}_2(\text{g}) \rightarrow \text{SnCl}_4(\text{l})$$
- 89.** Use standard enthalpies of formation from Table C-13 in Appendix C to calculate $\Delta H_{\text{rxn}}^\circ$ for each of these reactions.
- $2\text{NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
 - $\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\text{l})$
 - $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$
 - $2\text{H}_2\text{S}(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) + 2\text{SO}_2(\text{g})$
 - $4\text{FeS}(\text{s}) + 7\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s}) + 4\text{SO}_2(\text{g})$

Reaction Spontaneity (16.5)

- 90.** Calculate ΔG_{system} for each process and state if the process is spontaneous or nonspontaneous.
- $\Delta H_{\text{system}} = 145 \text{ kJ}$, $T = 293 \text{ K}$,
 $\Delta S_{\text{system}} = 195 \text{ J/K}$
 - $\Delta H_{\text{system}} = -232 \text{ kJ}$, $T = 273 \text{ K}$,
 $\Delta S_{\text{system}} = 138 \text{ J/K}$
 - $\Delta H_{\text{system}} = -15.9 \text{ kJ}$, $T = 373 \text{ K}$,
 $\Delta S_{\text{system}} = -268 \text{ J/K}$
- 91.** Calculate the temperature at which $\Delta G_{\text{system}} = -34.7 \text{ kJ}$ if $\Delta H_{\text{system}} = -28.8 \text{ kJ}$ and $\Delta S_{\text{system}} = 22.2 \text{ J/K}$?
- 92.** Under certain conditions, iron ore (Fe_3O_4) can be converted to iron by the following reaction.
- $$\text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g}) \rightarrow 3\text{Fe}(\text{s}) + 4\text{H}_2\text{O}(\text{g})$$
- $\Delta H_{\text{system}} = 149.8 \text{ kJ}$
 $\Delta S_{\text{system}} = 610.0 \text{ J/K}$
- Is the reaction spontaneous at 298 K? Explain why or why not based upon how the entropy of the system, surroundings, and universe change as a result of the reaction.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 93.** What mass of octane must be burned in order to liberate 5340 kJ of heat? $\Delta H_{\text{comb}} = -5471 \text{ kJ/mol}$.
- 94.** How much heat is released to the surroundings when 200 g of water at 96.0°C cools to 25.0°C? The specific heat of water is 4.184 J/(g·°C).
- 95.** What is the final temperature of 1280 g of water, originally at 20.0°C, if it absorbs 47.6 kJ of heat?
- 96.** Is the following reaction spontaneous at 456 K? If not, is it spontaneous at some other temperature? Explain your answer.
- $$\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$$
- $\Delta H_{\text{system}} = 68 \text{ kJ}$
 $\Delta S_{\text{system}} = -122 \text{ J/K}$
- 97.** Use Hess's law to determine ΔH for the reaction $\text{NO}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{NO}_2(\text{g})$ $\Delta H = ?$ given the following reactions. Show your work.
- $$\text{O}_2(\text{g}) \rightarrow 2\text{O}(\text{g}) \quad \Delta H = +495 \text{ kJ}$$
- $$2\text{O}_3(\text{g}) \rightarrow 3\text{O}_2(\text{g}) \quad \Delta H = -427 \text{ kJ}$$
- $$\text{NO}(\text{g}) + \text{O}_3(\text{g}) \rightarrow \text{NO}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = -199 \text{ kJ}$$

Thinking Critically

- 98. Using Numbers** A 133-g piece of granite rock is heated to 65.0°C, then placed in 643 g ethanol at 12.7°C. Assuming no heat loss, what is the final temperature of the granite and ethanol? See **Table 16-2**.
- 99. Applying Concepts** Write the thermochemical equation for the decomposition of liquid hydrogen peroxide (H_2O_2) to water vapor and oxygen gas. Calculate ΔH_{system} for the reaction using standard enthalpies of formation. Analyze the reaction and explain why NASA found this reaction suitable for providing thrust in the control jets of some space vehicles.
- 100. Recognizing Cause and Effect** Explain why the equation $\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$ is especially valuable in recognizing how reactions and processes affect the entropy of the universe.

Writing in Chemistry

- 101.** Research and explain how hydrogen might be produced, transported, and used as a fuel for automobiles. Summarize the benefits and drawbacks of using hydrogen as an alternative fuel for internal combustion engines.
- 102.** Research the use of wind as a source of electrical power. Explain the possible benefits, disadvantages, and limitations of its use.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 103.** Why is it necessary to perform repeated experiments in order to support a hypothesis? (Chapter 1)
- 104.** Phosphorus has the atomic number 15 and an atomic mass of 31 amu. How many protons, neutrons, and electrons are in a neutral phosphorus atom? (Chapter 4)
- 105.** What element has the electron configuration $[\text{Ar}]4s^13d^5$? (Chapter 5)
- 106.** Name the following molecular compounds. (Chapter 8)
- S_2Cl_2
 - CS_2
 - SO_3
 - P_4O_{10}
- 107.** Determine the molar mass for the following compounds. (Chapter 11)
- $\text{Co}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$
 - $\text{Fe}(\text{OH})_3$

Use these questions and the test-taking tip to prepare for your standardized test.

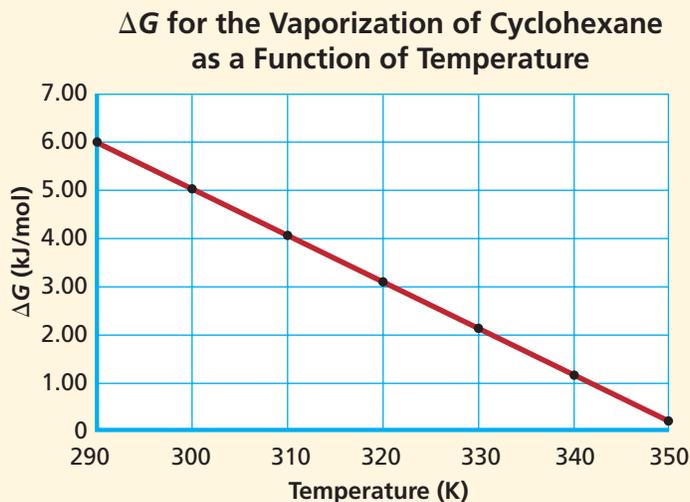
- The specific heat of ethanol is 2.44 J/g·°C. How many kilojoules of energy are required to heat 50.0 g of ethanol from -20.0°C to 68.0°C?
 - 10.7 kJ
 - 8.30 kJ
 - 2.44 kJ
 - 1.22 kJ
- When a reaction takes place at constant pressure, q for the reaction equals _____.
 - ΔS_{rxn}
 - ΔH_{rxn}
 - ΔG_{rxn}
 - c_{rxn}
- Determine ΔH for the reaction of aluminum and sulfur dioxide.

$$4\text{Al(s)} + 3\text{SO}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)} + 3\text{S(s)}$$
 Use the following equations:

$$4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)} \quad \Delta H = -3352 \text{ kJ/mol}$$

$$\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(g)} \quad \Delta H = -297 \text{ kJ/mol}$$
 - 4243 kJ
 - 3649 kJ
 - 3055 kJ
 - 2461 kJ

Interpreting Graphs Use the graph to answer questions 4–6.



- In the range of temperatures shown, the vaporization of cyclohexane _____.
 - does not occur at all
 - will occur spontaneously
 - is not spontaneous
 - occurs only at high temperatures

- The standard free energy, enthalpy, and entropy of vaporization are often written with the subscript *vap* to indicate that they are associated with a vaporization reaction. What is the $\Delta G_{\text{vap}}^\circ$ of cyclohexane at 300 K?
 - 5.00 kJ/mol
 - 4.00 kJ/mol
 - 3.00 kJ/mol
 - 2.00 kJ/mol
- When $\Delta G_{\text{vap}}^\circ$ is plotted as a function of temperature, the slope of the line equals $\Delta S_{\text{vap}}^\circ$ and the y-intercept of the line equals $\Delta H_{\text{vap}}^\circ$. The standard entropy of vaporization of cyclohexane is approximately _____.
 - 50.0 J/mol·K
 - 10.0 J/mol·K
 - 5.0 J/mol·K
 - 100 J/mol·K
- 3.00 g of aluminum foil is placed in an oven and heated from 20.0°C to 662.0°C. If it absorbs 1728 J of heat, what is the specific heat of aluminum?
 - 0.131 J/g·°C
 - 0.870 J/g·°C
 - 0.897 J/g·°C
 - 2.61 J/g·°C
- $\text{AB(s)} + \text{C}_2\text{(l)} \rightarrow \text{AC(g)} + \text{BC(g)}$

All of the following can be predicted about this reaction EXCEPT _____.

- the entropy of the system decreases
 - the entropy of the products is higher than that of the reactants
 - the change in entropy for this reaction, ΔS_{rxn} , is positive
 - the disorder of the system increases
- $\text{Co(s)} + \text{S(s)} + 2\text{O}_2\text{(g)} \rightarrow \text{CoSO}_4\text{(s)}$
 $\Delta H_f^\circ = -888.3 \text{ kJ/mol}$, $\Delta S_f^\circ = 118.0 \text{ J/mol}\cdot\text{K}$
 Given the above thermochemical data for the formation of cobalt(II) sulfate from its elements, what is ΔG_f° for CoSO_4 at 25°C?
 - 853.1 kJ/mol
 - 885.4 kJ/mol
 - 891.3 kJ/mol
 - 923.5 kJ/mol



TEST-TAKING TIP

Stock Up On Supplies Bring all your test-taking tools: number two pencils, black and blue pens, erasers, correction fluid, a sharpener, a ruler, a calculator, and a protractor. Bring munchies, too. You might not be able to eat them in the testing room, but they come in handy for the break, if you're allowed to go outside.

Reaction Rates

What You'll Learn

- ▶ You will investigate a model describing how chemical reactions occur as a result of collisions.
- ▶ You will compare the rates of chemical reactions under varying conditions.
- ▶ You will calculate the rates of chemical reactions.

Why It's Important

Perhaps someday you'll pilot a space shuttle mission. You and your crew will be propelled into space by the result of a chemical reaction. An understanding of reaction rates is the tool that allows us to control chemical reactions and use them effectively.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about reaction rates.

The launch of the space shuttle *Columbia* in 1996 thrust commander Eileen Collins into history as the first female commander of a shuttle mission.



DISCOVERY LAB



Materials

hydrogen peroxide
beaker or cup
baker's yeast
toothpicks

Speeding Reactions

Many chemical reactions occur so slowly that you don't even know they are happening. For some reactions, it is possible to alter the reaction speed using another substance.

Safety Precautions



Always use safety goggles and an apron in the lab.

Procedure

1. Create a "before and after" table and record your observations.
2. Pour about 10 mL of hydrogen peroxide into a small beaker or cup. Observe the hydrogen peroxide.
3. Add a "pinch" (1/8 tsp) of yeast to the hydrogen peroxide. Stir gently with a toothpick and observe the mixture again.

Analysis

Into what two products does hydrogen peroxide decompose? Why aren't bubbles produced in step 1? What is the function of the yeast?

Section

17.1

A Model for Reaction Rates

Objectives

- **Calculate** average rates of chemical reactions from experimental data.
- **Relate** rates of chemical reactions to collisions between reacting particles.

Vocabulary

reaction rate
collision theory
activated complex
transition state
activation energy

One of the most spectacular chemical reactions, the one between liquid hydrogen and liquid oxygen, provides the energy to launch the space shuttle *Columbia*, shown on the opposite page. This reaction is fast and exothermic. Yet other reactions and processes you're familiar with, such as the hardening of concrete or the formation of fossil fuels, occur at considerably slower rates. The **DISCOVERY LAB** for this chapter emphasized that the speed at which a reaction occurs can vary if other substances are introduced into the reaction. In this section, you'll learn about a model that scientists use to describe and calculate the rates at which chemical reactions occur.

Expressing Reaction Rates

As you know, some chemical reactions are fast and others are slow; however, *fast* and *slow* are inexact, relative terms. Chemists, engineers, medical researchers, and others often need to be more specific.

Think about how you express speed or rate in the situations shown in **Figure 17-1** on the next page. The speed of the sprinter on the track team may be expressed as meters per second. The speed at which the hiker moves might be expressed differently, perhaps as meters per minute. We generally define the average rate of an action or process to be the change in a given quantity during a specific period of time. Recall from your study of math that the Greek letter delta (Δ) before a quantity indicates a change in the quantity. In equation form, average rate or speed is written as

$$\text{Average rate} = \frac{\Delta \text{quantity}}{\Delta t}$$



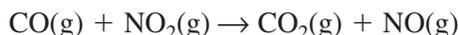
Figure 17-1

The average rate of travel for these activities is based on the change in distance over time. Similarly, the reaction rate is based on the change in concentration over time.

For chemical reactions, this equation defines the average rate at which reactants produce products, which is the amount of change of a reactant in a given period of time. Most often, chemists are concerned with changes in the molar concentration (mol/L, M) of a reactant or product during a reaction. Therefore, the **reaction rate** of a chemical reaction is stated as the change in concentration of a reactant or product per unit time, expressed as mol/(L·s). Brackets around the formula for a substance denote the molar concentration. For example, $[\text{NO}_2]$ represents the molar concentration of NO_2 .

It's important to understand that reaction rates are determined experimentally by measuring the concentrations of reactants and/or products in an actual chemical reaction. Reaction rates cannot be calculated from balanced equations as stoichiometric amounts can.

Suppose you wish to express the average rate of the reaction



during the time period beginning at time t_1 and ending at time t_2 . Calculating the rate at which the products of the reaction are produced results in a reaction rate having a positive value. The rate calculation based on the production of NO will have the form

$$\text{Average reaction rate} = \frac{[\text{NO}] \text{ at time } t_2 - [\text{NO}] \text{ at time } t_1}{t_2 - t_1} = \frac{\Delta[\text{NO}]}{\Delta t}$$

For example, if the concentration of NO is $0.00M$ at time $t_1 = 0.00 \text{ s}$ and $0.010M$ two seconds after the reaction begins, the following calculation gives the average rate of the reaction expressed as moles of NO produced per liter per second.

$$\begin{aligned} \text{Average reaction rate} &= \frac{0.010M - 0.000M}{2.00 \text{ s} - 0.00 \text{ s}} \\ &= \frac{0.010M}{2.00 \text{ s}} = 0.0050 \text{ mol/(L}\cdot\text{s)} \end{aligned}$$

Notice how the units calculate:

$$\frac{M}{s} = \frac{\text{mol/L}}{s} = \frac{\text{mol}}{L} \cdot \frac{1}{s} = \frac{\text{mol}}{L\cdot s}$$

Although $\text{mol/L} = M$, chemists typically reserve M to express concentration and $\text{mol/(L}\cdot\text{s)}$ to express rate.

Alternatively, you can choose to state the rate of the reaction as the rate at which CO is consumed, as shown below:

$$\text{Average reaction rate} = \frac{[\text{CO}] \text{ at time } t_2 - [\text{CO}] \text{ at time } t_1}{t_2 - t_1} = \frac{\Delta[\text{CO}]}{\Delta t}$$

Do you predict a positive or negative value for this reaction? In this case, a negative value indicates that the concentration of CO decreases as the reaction proceeds.

Actually, reaction rates must always be positive. For reactions in which reactants are consumed, scientists commonly apply a negative sign to the calculation to get a positive reaction rate. Therefore, the following form of the average rate equation is used to calculate the rate of consumption:

$$\text{Average reaction rate} = - \frac{\Delta \text{quantity}}{\Delta t}$$



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.



EXAMPLE PROBLEM 17-1

Calculating Average Reaction Rates

Reaction data for the reaction between butyl chloride (C_4H_9Cl) and water (H_2O) is given in Table 17-1. Calculate the average reaction rate over this time period expressed as moles of C_4H_9Cl consumed per liter per second.

1. Analyze the Problem

You are given the initial and final concentrations of C_4H_9Cl and the initial and final times. You can calculate the average reaction rate of the chemical reaction using the change in concentration of butyl chloride in four seconds. In this problem, the reactant butyl chloride is consumed.

Known

$$t_1 = 0.00 \text{ s}$$

$$t_2 = 4.00 \text{ s}$$

$$[C_4H_9Cl] \text{ at } t_1 = 0.220M$$

$$[C_4H_9Cl] \text{ at } t_2 = 0.100M$$

Unknown

$$\text{Average reaction rate} = ? \text{ mol}/(\text{L}\cdot\text{s})$$

2. Solve for the Unknown

Write the equation for the average reaction rate, insert the known quantities, and perform the calculation.

$$\begin{aligned} \text{Average reaction rate} &= - \frac{[C_4H_9Cl] \text{ at time } t_2 - [C_4H_9Cl] \text{ at time } t_1}{t_2 - t_1} \\ &= - \frac{0.100M - 0.220M}{4.00 \text{ s} - 0.00 \text{ s}} \\ (\text{Substitute units}) &= - \frac{0.100 \text{ mol/L} - 0.220 \text{ mol/L}}{4.00 \text{ s} - 0.00 \text{ s}} \\ &= - \frac{-0.120 \text{ mol/L}}{4.00 \text{ s}} \\ &= 0.0300 \text{ mol}/(\text{L}\cdot\text{s}) \end{aligned}$$

3. Evaluate the Answer

The average reaction rate of 0.0300 moles C_4H_9Cl consumed per liter per second seems reasonable based on start and end amounts. The answer is correctly expressed in three significant figures.

Table 17-1

Molar Concentration of C_4H_9Cl	
$[C_4H_9Cl]$ at $t = 0.00 \text{ s}$	$[C_4H_9Cl]$ at $t = 4.00 \text{ s}$
0.220M	0.100M

PRACTICE PROBLEMS

Use the data in the following table to calculate the average reaction rates.

Experimental Data for $H_2 + Cl_2 \rightarrow 2HCl$			
Time (s)	$[H_2]$ (M)	$[Cl_2]$ (M)	$[HCl]$ (M)
0.00	0.030	0.050	0.000
4.00	0.020	0.040	0.020

1. Calculate the average reaction rate expressed in moles H_2 consumed per liter per second.
2. Calculate the average reaction rate expressed in moles Cl_2 consumed per liter per second.
3. Calculate the average reaction rate expressed in moles HCl produced per liter per second.



For more practice calculating average reaction rates, go to **Supplemental Practice Problems** in Appendix A.

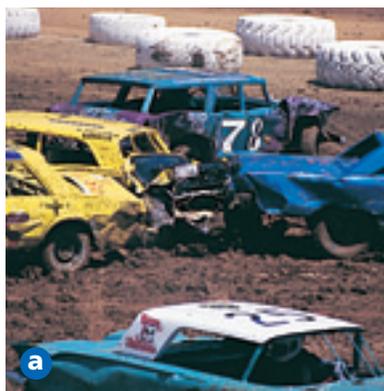
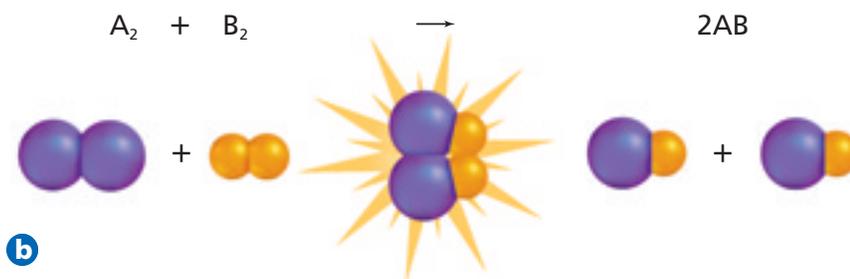


Figure 17-2

a The product of a demolition derby, the crunch of metal, is the result of collisions between the cars. **b** The product of a chemical reaction is the result of collisions between atoms, ions, and molecules. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

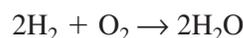


The Collision Theory

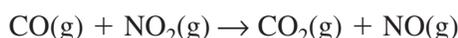
You have learned that reaction rates are calculated from experimental data. But what are we actually measuring with these calculations? Looking at chemical reactions from the molecular level will provide a clearer picture of exactly what reaction rates measure.

Have you ever seen a demolition derby in which the competing vehicles are constantly colliding? Each collision may result in the demolition of one or more vehicles as shown in **Figure 17-2a**. The reactants in a chemical reaction must also come together in order to form products, as shown in **Figure 17-2b**. The **collision theory** states that atoms, ions, and molecules must collide in order to react. The collision theory, summarized in **Table 17-2**, explains why reactions occur and how the rates of chemical reactions can be modified.

Consider the reaction between hydrogen gas (H_2) and oxygen gas (O_2).



According to the collision theory, H_2 and O_2 atoms must collide in order to react and produce H_2O . Now look at the reaction between carbon monoxide (CO) gas and nitrogen dioxide (NO_2) gas at a temperature above 500 K.



These molecules collide to produce carbon dioxide (CO_2) gas and nitrogen monoxide (NO) gas. However, detailed calculations of the number of molecular collisions per second yield a puzzling result—only a small fraction of collisions produce reactions. In this case, other factors must be considered.

Orientation and the activated complex Why don't the NO_2 and CO molecules shown in **Figure 17-3a** and **b** react when they collide? Analysis of this example indicates that in order for a collision to lead to a reaction, the carbon atom in a CO molecule must contact an oxygen atom in an NO_2 molecule at the instant of impact. Only in that way can a temporary bond between the carbon atom and an oxygen atom form. The collisions shown in **Figure 17-3a** and **b** do not lead to reactions because the molecules collide with unfavorable orientations. That is, because a carbon atom does not contact an oxygen atom at the instant of impact, the molecules simply rebound.

When the orientation of colliding molecules is correct, as **Figure 17-3c** illustrates, a reaction occurs as an oxygen atom is transferred from an NO_2 molecule to a CO molecule. A short-lived, intermediate substance is formed. The intermediate substance, in this case OCONO, is called an activated complex. An **activated complex** is a temporary, unstable arrangement of atoms that may form products or may break apart to re-form the reactants. Because the activated complex is as likely to form reactants as it is to form products, it is sometimes referred to as the **transition state**. An activated complex is the first step leading to the resulting chemical reaction.

Table 17-2

Collision Theory Summary

1. Reacting substances (atoms, ions, or molecules) must collide.
2. Reacting substances must collide with the correct orientation.
3. Reacting substances must collide with sufficient energy to form the activated complex.

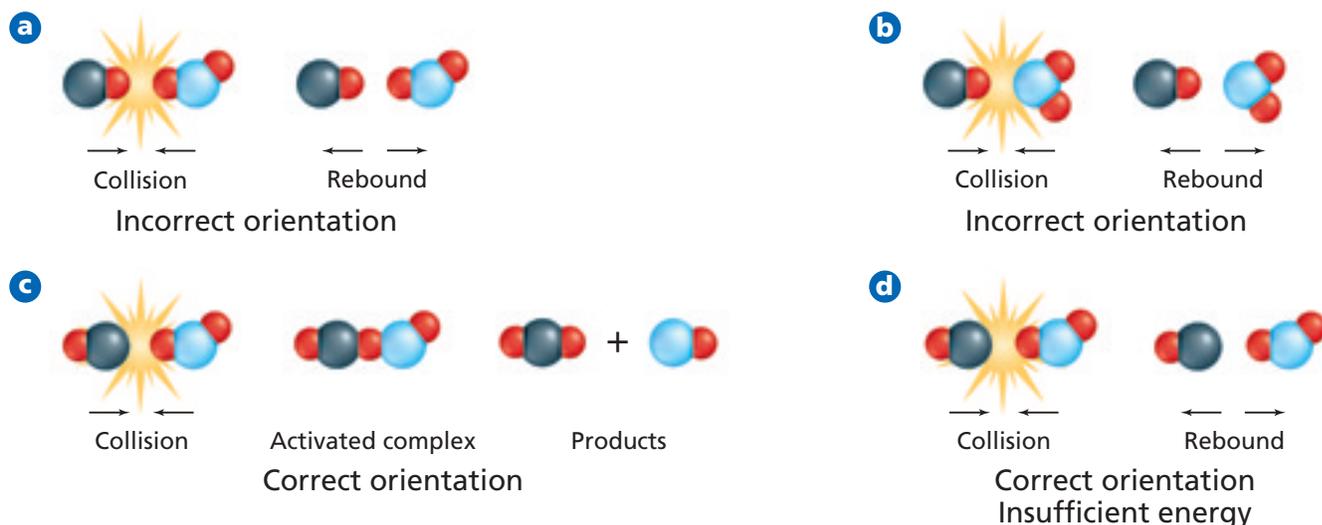


Figure 17-3

The collisions in **a** and **b** do not result in a reaction because the molecules are not in position to form bonds. The molecules in **c** are in the correct orientation when they collide, and a reaction occurs. The molecules in **d** are also in the correct positions on collision, but insufficient energy at the point of collision prevents a chemical reaction.

Activation energy and reaction The collision depicted in **Figure 17-3d** does not lead to a reaction for a different reason. Although the CO and NO₂ molecules collide with a favorable orientation, no reaction occurs because they collide with insufficient energy to form the activated complex. The minimum amount of energy that reacting particles must have to form the activated complex and lead to a reaction is called the **activation energy**, E_a .

Activation energy has a direct influence on the rate of a reaction. A high E_a means that relatively few collisions will have the required energy to produce the activated complex and the reaction rate will be low. On the other hand, a low E_a means that more collisions will have sufficient energy to react, and the reaction rate will be higher. The **problem-solving LAB** below emphasizes this fact. It might be helpful to think of this relationship in terms of a person pushing a heavy cart up a hill. If the hill is high, a substantial amount of energy and effort will be required to move the cart and it will take a long time to get it to the top. If the hill is low, less energy will be required and the task will be accomplished faster.

problem-solving LAB

Speed and Energy of Collision

Designing an Experiment Controlling the rate of a reaction is common in your everyday experience. Consider two of the major appliances in your kitchen: a refrigerator slows down chemical processes that cause food to spoil and an oven speeds up chemical processes that cause foods to cook. Petroleum and natural gases require a spark to run your car's engine and heat your home.

Analysis

According to collision theory, reactants must collide with enough energy in order to react. Recall from the kinetic-molecular theory that mass and velocity determine the kinetic energy of a parti-

cle. In addition, temperature is a measure of the average kinetic energy of the particles in a sample of matter. According to kinetic-molecular theory, how are the frequency and energy of collisions between gas particles related to temperature?

Thinking Critically

How would you design an experiment with a handful of marbles and a shoe box lid to simulate the range of speeds among a group of particles at a given temperature? How would you model an increase in temperature? Describe how your marble model would demonstrate the relationship between temperature and the frequency and energy of the collisions among a group of particles.

Figure 17-4

In an exothermic reaction, molecules collide with enough energy to overcome the activation energy barrier, form an activated complex, then release energy and form products at a lower energy level.

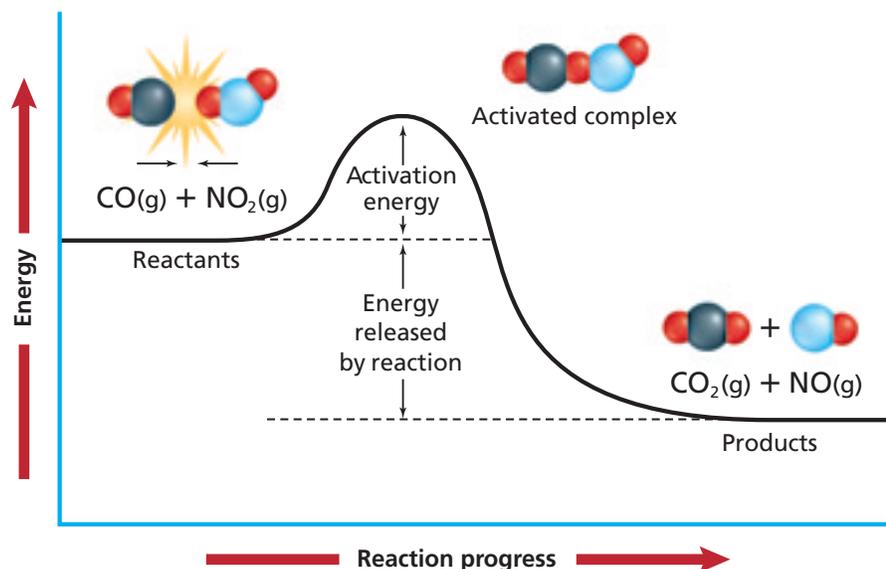
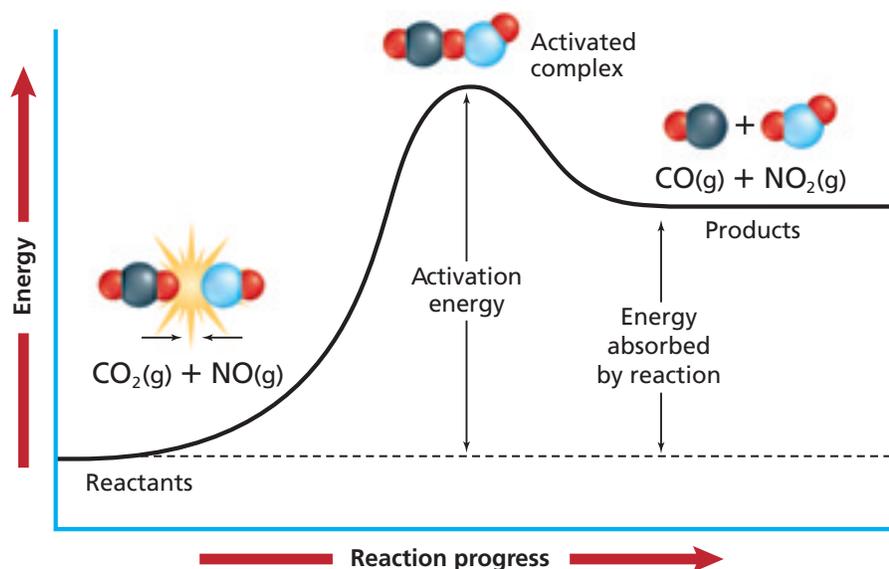


Figure 17-4 shows the energy diagram for the progress of the reaction between carbon monoxide and nitrogen dioxide. Does this energy diagram look somewhat different from those you studied in Chapter 16? Why? In addition to the energies of the reactants and products, this diagram shows the activation energy of the reaction. Activation energy can be thought of as a barrier the reactants must overcome in order to form the products. In this case, the CO and NO_2 molecules collide with enough energy to overcome the barrier, and the products formed lie at a lower energy level. Do you recall that reactions that lose energy, such as this example, are called exothermic reactions?

For many reactions, the process from reactants to products is reversible. **Figure 17-5** shows the reverse endothermic reaction between CO_2 and NO to reform CO and NO_2 . In this reverse reaction, the reactants, which are the molecules that were formed in the exothermic forward reaction, lie at a low energy level and must overcome a significant activation energy to reform CO and NO_2 . This requires an input of energy. If this reverse reaction is achieved, CO and NO_2 again lie at a high energy level.

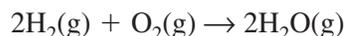
Figure 17-5

In the reverse endothermic reaction, the reactant molecules lying at a low energy level must absorb energy to overcome the activation energy barrier and form high-energy products.



The influence of spontaneity Recall from Chapter 16 that reaction spontaneity is related to change in free energy, ΔG . If ΔG is negative, the reaction is spontaneous under the conditions specified. If ΔG is positive, the reaction is not spontaneous. Let's now consider whether spontaneity has any effect on reaction rates. Are more spontaneous reactions faster than less spontaneous ones?

To investigate the relationship between spontaneity and reaction rate, consider the following gas-phase reaction between hydrogen and oxygen.



Here, $\Delta G = -458 \text{ kJ}$ at 298 K (25°C) and 1 atm pressure. Because ΔG is negative, the reaction is spontaneous. For the same reaction, $\Delta H = -484 \text{ kJ}$, which means that the reaction is highly exothermic. You can examine the speed of this reaction by filling a tape-wrapped soda bottle with stoichiometric quantities of the two gases—two volumes hydrogen and one volume oxygen. A thermometer in the stopper allows you to monitor the temperature inside the bottle. As you watch for evidence of a reaction, the temperature remains constant for hours. Have the gases escaped? Or have they simply failed to react? If you remove the stopper and hold a burning splint to the mouth of the bottle, a reaction occurs explosively. Clearly, the hydrogen and oxygen gases have not escaped from the bottle. Yet they did not react noticeably until you supplied additional energy in the form of a lighted splint. The example shown in **Figure 17-6** illustrates this same phenomenon. The soap bubbles you see billowing from the bowl are filled with hydrogen. When the lighted splint introduces additional energy, an explosive reaction occurs between the hydrogen that was contained in the bubbles and oxygen in the air.

As these examples show, reaction spontaneity in the form of ΔG implies *absolutely nothing* about the speed of the reaction; ΔG merely indicates the natural tendency for a reaction or process to proceed. Factors other than spontaneity, however, do affect the rate of a chemical reaction. These factors are keys in controlling and using the power of chemistry.

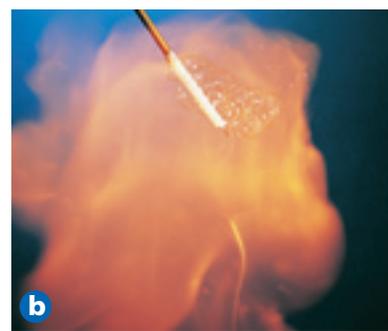


Figure 17-6

- a** Although the reaction of hydrogen and oxygen is spontaneous, the combination does not produce an obvious reaction when mixed.
- b** If an additional form of energy is introduced to the same spontaneous reaction, the reaction rate increases significantly.

Section 17.1 Assessment

4. What does the reaction rate indicate about a particular chemical reaction?
5. How is the rate of a chemical reaction usually expressed?
6. What is the collision theory, and how does it relate to reaction rates?
7. According to the collision theory, what must happen in order for two molecules to react?
8. How is the speed of a chemical reaction related to the spontaneity of the reaction?
9. **Thinking Critically** How would the rate of the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$ stated as the consumption of hydrogen compare with the rate stated as the consumption of oxygen?
10. **Interpreting Scientific Illustrations** Based on your analysis of **Figures 17-4** and **5**, how does E_a for the reaction $\text{CO} + \text{NO}_2 \leftarrow \text{CO}_2 + \text{NO}$ (the reverse reaction) compare with that of the reaction $\text{CO} + \text{NO}_2 \rightarrow \text{CO}_2 + \text{NO}$ (the forward reaction)?

Factors Affecting Reaction Rates

Objectives

- **Identify** factors that affect the rates of chemical reactions.
- **Explain** the role of a catalyst.

Vocabulary

catalyst
inhibitor
heterogeneous catalyst
homogeneous catalyst

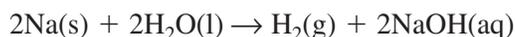
According to the collision theory, chemical reactions occur when molecules collide in a particular orientation and with sufficient energy to achieve activation energy. You can probably identify chemical reactions that occur fast, such as gasoline combustion, and others that occur more slowly, such as iron rusting. But can the reaction rate of any single reaction vary, or is the reaction rate constant regardless of the conditions? In this section you will learn that the reaction rate for almost any chemical reaction can be modified by varying the conditions of the reaction.

The Nature of Reactants

An important factor that affects the rate of a chemical reaction is the reactive nature of the reactants. As you know, some substances react more readily than others. For example, calcium and sodium are both reactive metals; however, what happens when each metal is added to water is distinctly different. When a small piece of calcium is placed in cold water, as shown in **Figure 17-7a**, the calcium and water react slowly to form hydrogen gas and aqueous calcium hydroxide.



When a small piece of sodium is placed in cold water, the sodium and water react quickly to form hydrogen gas and aqueous sodium hydroxide, as shown in **Figure 17-7b**.



Comparing the two equations, it's evident that the reactions are similar. However, the reaction between sodium and water occurs much faster because sodium is more reactive with water than calcium is. In fact, the reaction releases so much heat so quickly that the hydrogen gas ignites as it is formed.

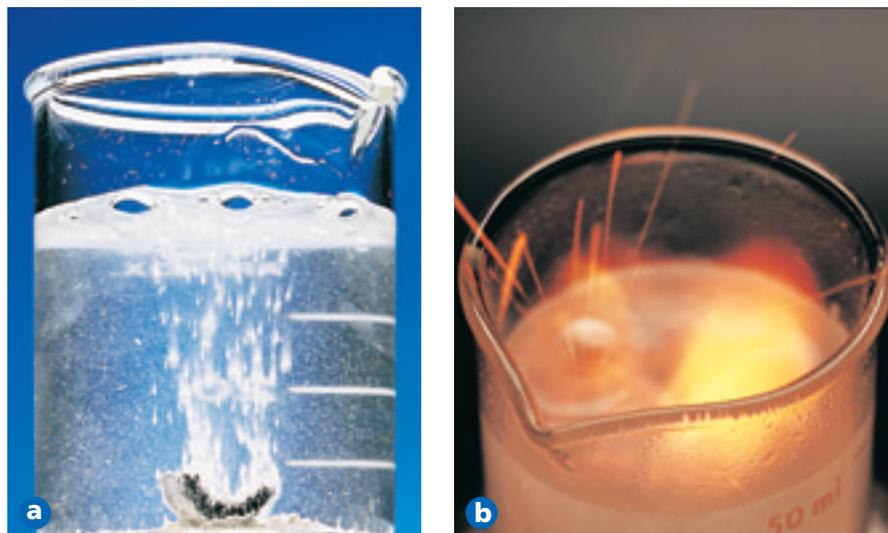


Figure 17-7

The tendency of a substance to react influences the rate of a reaction involving the substance. The more reactive a substance is, the faster the reaction rate.

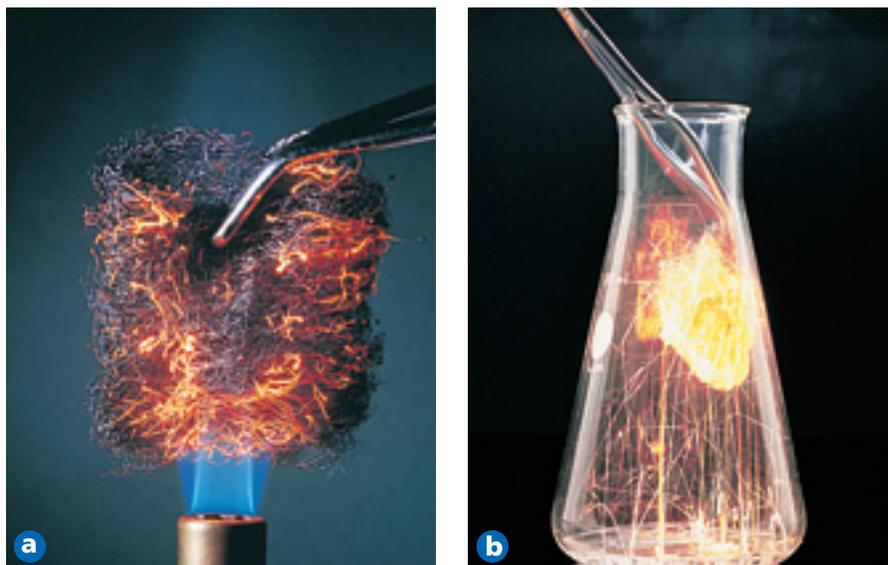


Figure 17-8

- a** The concentration of oxygen in the air surrounding the steel wool is much less than that of the pure oxygen in the flask.
- b** The higher oxygen concentration accounts for the faster reaction.

Concentration

Reactions speed up when the concentrations of reacting particles are increased. One of the fundamental principles of the collision theory is that particles must collide to react. The number of particles in a reaction makes a difference in the rate at which the reaction takes place. Think about a reaction where reactant A combines with reactant B. At a given concentration of A and B, the molecules of A collide with B to produce AB at a particular rate. What happens if the amount of B is increased? Increasing the concentration of B makes more molecules available with which A can collide. Reactant A “finds” reactant B more easily because there are more B molecules in the area, which increases probability of collision, and ultimately increases the rate of reaction between A and B.

Look at the two reactions shown in **Figure 17-8**. Steel wool is first heated over a burner until it is red hot. In **Figure 17-8a**, the hot steel wool reacts with oxygen in the air. How does this reaction compare with the one in **Figure 17-8b**, in which the hot steel wool is lowered immediately into a flask containing nearly 100 percent oxygen—approximately five times the concentration of oxygen in air? Applying the collision theory to this reaction, the higher concentration of oxygen increases the collision frequency between iron atoms and oxygen molecules, which increases the rate of the reaction.

If you have ever been near a person using bottled oxygen or seen an oxygen generator, you may have noticed a sign cautioning against smoking or using open flames. Can you now explain the reason for the caution notice? The high concentration of oxygen could cause a combustion reaction to occur at an explosive rate. The **CHEMLAB** at the end of this chapter gives you an opportunity to further investigate the effect of concentration on reaction rates.

Surface Area

Now suppose you were to lower a red-hot chunk of steel instead of steel wool into a flask of oxygen gas. The oxygen would react with the steel much more slowly than it would with the steel wool. Using what you know about the collision theory, can you explain why? You are correct if you said that, for the same mass of iron, steel wool has much more surface area than the chunk of steel. The greater surface area of the steel wool allows the oxygen molecules to collide with many more iron atoms per unit of time.



Figure 17-9

Increasing the surface area of reactants provides more opportunity for collisions with other reactants, thereby increasing the reaction rate.

Pulverizing (or grinding) a substance is one way to increase its rate of reaction. This is because, for the same mass, many small particles possess more total surface area than one large particle. So, a spoonful of granulated sugar placed in a cup of water dissolves faster than the same mass of sugar in a single chunk, which has less surface area, as shown in **Figure 17-9**. This example illustrates that increasing the surface area of a reactant does not change its concentration, but it does increase the rate of reaction by increasing the collision rate between reacting particles.

Temperature

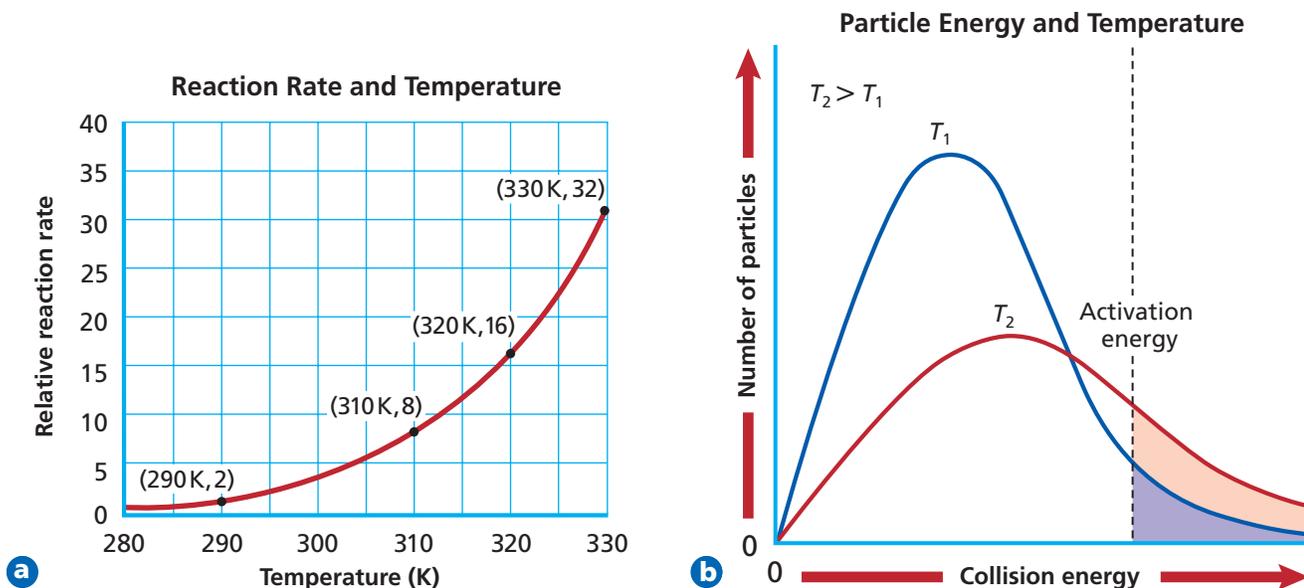
Generally, increasing the temperature at which a reaction occurs increases the reaction rate. For example, you know that the reactions that cause foods to spoil occur much faster at room temperature than when the foods are refrigerated. The graph in **Figure 17-10a** illustrates that increasing the temperature by 10 K can approximately double the rate of a reaction. How can a small increase in temperature have such a significant effect?

As you learned in Chapter 13, increasing the temperature of a substance increases the average kinetic energy of the particles that make up the substance. For that reason, reacting particles collide more frequently at higher temperatures than at lower temperatures. However, that fact alone doesn't account for the increase in reaction rate with increasing temperature. To better understand how reaction rate varies with temperature, examine the graph shown in **Figure 17-10b**. This graph compares the numbers of particles having sufficient energy to react at temperatures T_1 and T_2 , where T_2 is greater than T_1 . The shaded area under each curve represents the number of collisions having energy equal to or greater than the activation energy. The dotted line indicates the activation energy (E_a) for the reaction. How do the shaded areas compare? The number of high-energy collisions at the higher temperature, T_2 , is much greater than at the lower temperature, T_1 . Therefore, as the temperature increases more collisions result in a reaction.

As you can see, increasing the temperature of the reactants increases the reaction rate because raising the kinetic energy of the reacting particles raises both the collision frequency and the collision energy. You can investigate the relationship between reaction rate and temperature by performing the **miniLAB** for this chapter.

Figure 17-10

- a** Increasing the temperature of a reaction increases the frequency of collisions and therefore the rate of reaction.
- b** Increasing the temperature also raises the kinetic energy of the particles, thus more of the collisions at high temperatures have enough energy to overcome the activation energy barrier and react.



Examining Reaction Rate and Temperature

Recognizing Cause and Effect Several factors affect the rate of a chemical reaction. This lab allows you to examine the effect of temperature on a common chemical reaction.

Materials small beaker, thermometer, hot plate, 250-mL beaker, balance, water, effervescent (bicarbonate) tablet, stopwatch or clock with second hand

Procedure



1. Take a single effervescent tablet and break it into four roughly equal pieces.
2. Measure the mass of one piece of the tablet. Measure 50 mL of room-temperature water (approximately 20°C) into a small cup or beaker. Measure the temperature of the water.
3. With a stopwatch ready, add the piece of tablet to the water. Record the amount of time elapsed between when the tablet hits the water and when you see the first large bubbles appear as a result of the reaction.
4. Repeat steps 2 and 3 twice more, except use water temperatures of about 50°C and 65°C. Be sure to raise the temperature gradually and maintain the desired temperature (equilibrate) throughout the run.



Analysis

1. Calculate the reaction rate by finding the mass/time for each run.
2. Graph the reaction rate (mass/time) versus temperature for the runs.
3. What is the relationship between reaction rate and temperature for this reaction?
4. Using your graphed data, predict the reaction rate for the reaction carried out at 40°C. Heat and equilibrate the water to 40°C and use the last piece of tablet to test your prediction.
5. How did your prediction for the reaction rate at 40°C compare to the actual reaction rate?

Catalysts

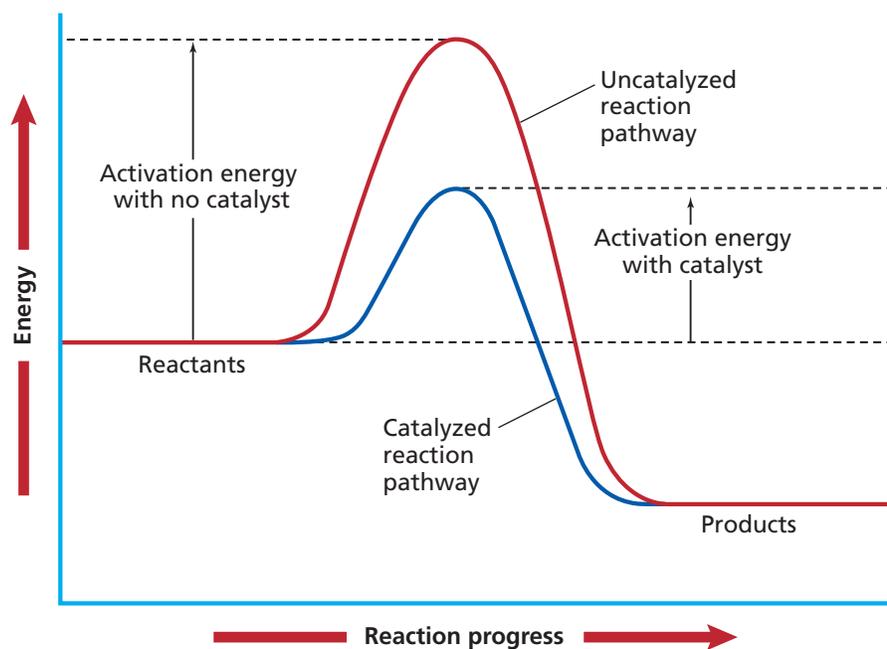
You've seen that increasing the temperature and/or the concentration of reactants can dramatically increase the rate of a reaction. However, an increase in temperature is not always the best (or most practical) thing to do. For example, suppose that you want to increase the rate of a reaction such as the decomposition of glucose in a living cell. Increasing the temperature and/or the concentration of reactants is not a viable alternative because it might harm or kill the cell.

It is a fact that many chemical reactions in living organisms would not occur quickly enough to sustain life at normal living temperatures if it were not for the presence of enzymes. An enzyme is a type of **catalyst**, a substance that increases the rate of a chemical reaction without itself being consumed in the reaction. Although catalysts are important substances in a chemical reaction, a catalyst does not yield more product and is not included in the product(s) of the reaction. In fact, catalysts are not included in the chemical equation.

How does a catalyst increase the reaction rate? **Figure 17-11** on the next page shows the energy diagram for an exothermic chemical reaction. The red line represents the uncatalyzed reaction pathway—the reaction pathway with no catalyst present. The blue line represents the catalyzed reaction pathway.

Figure 17-11

This energy diagram shows how the activation energy of the catalyzed reaction is lower and therefore the reaction produces the products at a faster rate than the uncatalyzed reaction does.



Note that the activation energy for the catalyzed reaction is much lower than for the uncatalyzed reaction. The lower activation energy for the catalyzed reaction means that more collisions have sufficient energy to initiate reaction. It might help you to visualize this relationship by thinking of the reaction's activation energy as a mountain range to be crossed, as shown in **Figure 17-12**. In this analogy, the tunnel, representing the catalyzed pathway, provides an easier and therefore quicker route to the other side of the mountains.

Another type of substance that affects reaction rates is called an inhibitor. Unlike a catalyst, which speeds up reaction rates, an **inhibitor** is a substance that slows down, or inhibits, reaction rates. Some inhibitors, in fact, actually prevent a reaction from happening at all.

In our fast-paced world, it might seem unlikely that it would be desirable to slow a reaction. But if you think about your environment, you can probably come up with several uses for inhibitors. One of the primary applications for inhibitors is in the food industry, where inhibitors are called preservatives. **Figure 17-13** shows how a commercially available fruit freshener

Figure 17-12

Compare the amount of time it would take this train to travel over or around the barrier with the time it takes to travel through the barrier. Obviously, travel is faster through the mountain with the aid of the tunnel.





Figure 17-13

The preservative that was applied to the apple on the left is an inhibitor that reacted with substances in the apple to slow the chemical reactions that cause the apple to brown.

inhibits fruit from browning once it is cut. These preservatives are safe to eat and give food a longer shelf-life. Another inhibitor is the compound maleic hydrazide ($C_4N_2H_4O_2$), which is used as a plant growth inhibitor and weed killer. Inhibitors also are important in biology. For example, a class of inhibitors called monoamine oxidase inhibitors blocks a chemical reaction that can cause depression.

Heterogeneous and homogeneous catalysts In order to reduce harmful engine emissions, automobiles manufactured today must have catalytic converters similar to the one described in **How It Works** at the end of the chapter. The most effective catalysts for this application are transition metal oxides and metals such as palladium and platinum. These substances catalyze reactions that convert nitrogen monoxide to nitrogen and oxygen, carbon monoxide to carbon dioxide, and unburned gasoline to carbon dioxide and water. Because the catalysts in a catalytic converter are solids and the reactions they catalyze are gaseous, the catalysts are called heterogeneous catalysts. A **heterogeneous catalyst** exists in a physical state different than that of the reaction it catalyzes. A catalyst that exists in the same physical state as the reaction it catalyzes is called a **homogeneous catalyst**. For example, if both an enzyme and the reaction it catalyzes are in aqueous solution, the enzyme is a homogeneous catalyst.

Section 17.2 Assessment

- How do temperature, concentration, and surface area affect the rate of a chemical reaction?
- How does the collision model explain the effect of concentration on the reaction rate?
- How does the activation energy of an uncatalyzed reaction compare with that of the catalyzed reaction?
- Thinking Critically** For a reaction of A and B that proceeds at a specific rate, $x \text{ mol}/(\text{L}\cdot\text{s})$, what is the effect of decreasing the amount of one of the reactants?
- Using the Internet** Conduct Internet research on how catalysts are used in industry, in agriculture, or in the treatment of contaminated soil, waste, or water. Write a short report summarizing your findings about one use of catalysts.

Objectives

- **Express** the relationship between reaction rate and concentration.
- **Determine** reaction orders using the method of initial rates.

Vocabulary

rate law
specific rate constant
reaction order
method of initial rates

In Section 17.1, you learned how to calculate the average rate of a chemical reaction given the initial and final times and concentrations. The word *average* is important because most chemical reactions slow down as the reactants are consumed. To understand why most reaction rates slow over time, recall that the collision theory states that chemical reactions can occur only when the reacting particles collide and that reaction rate depends upon reactant concentration. As reactants are consumed, fewer particles collide and the reaction slows. Chemists use the concept of rate laws to quantify the results of the collision theory in terms of a mathematical relationship between the rate of a chemical reaction and the reactant concentration.

Reaction Rate Laws

The equation that expresses the mathematical relationship between the rate of a chemical reaction and the concentration of reactants is called a **rate law**. For example, the reaction $A \rightarrow B$, which is a one-step reaction, has only one activated complex between reactants and products. The rate law for this reaction is expressed as

$$\text{Rate} = k[A]$$

where k is the **specific rate constant**, or a numerical value that relates reaction rate and concentration of reactants at a given temperature. Units for the rate constant include $L/(\text{mol}\cdot\text{s})$, $L^2/(\text{mol}^2\cdot\text{s})$, and s^{-1} . Depending on the reaction conditions, especially temperature, k is unique for every reaction.

The rate law means that the reaction rate is directly proportional to the molar concentration of A. Thus, doubling the concentration of A will double the reaction rate. Increasing the concentration of A by a factor of 5 will increase the reaction rate by a factor of 5. The specific rate constant, k , does not change with concentration; however, k does change with temperature. A large value of k means that A reacts rapidly to form B. What does a small value of k mean?



Figure 17-14

Specific rate constants are determined experimentally. Scientists have a number of methods at their disposal that can be used to establish k for a given reaction.

a A spectrophotometer measures the absorption of specific wavelengths of light by a reactant or product as a reaction progresses to determine the specific rate constant for the reaction.

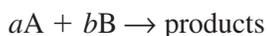
Reaction order In the expression $\text{Rate} = k[A]$, it is understood that the notation $[A]$ means the same as $[A]^1$. In other words, for reactant A, the understood exponent 1 is called the reaction order. The **reaction order** for a reactant defines how the rate is affected by the concentration of that reactant. For example, the rate law for the decomposition of H_2O_2 is expressed by the following equation.

$$\text{Rate} = k[\text{H}_2\text{O}_2]$$

Because the reaction rate is directly proportional to the concentration of H_2O_2 raised to the first power, $[\text{H}_2\text{O}_2]^1$, the decomposition of H_2O_2 is said to be first order in H_2O_2 . Because the reaction is first order in H_2O_2 , the reaction rate changes in the same proportion that the concentration of H_2O_2 changes. So if the H_2O_2 concentration decreases to one-half its original value, the reaction rate is halved as well.

Recall from Section 17.1 that reaction rates are determined from experimental data. Because reaction order is based on reaction rates, it follows that reaction order also is determined experimentally. Finally, because the rate constant describes the reaction order, k , too, must be determined experimentally. **Figure 17-14** illustrates two of several experimental methods that are commonly used to measure reaction rates.

Other reaction orders The overall reaction order of a chemical reaction is the sum of the orders for the individual reactants in the rate law. Many chemical reactions, particularly those having more than one reactant, are not first order. Consider the general form for a chemical reaction with two reactants. In this chemical equation, a and b are coefficients.



The general rate law for such a reaction is

$$\text{Rate} = k[A]^m[B]^n$$

where m and n are the reaction orders for A and B, respectively. Only if the reaction between A and B occurs in a single step (and with a single activated complex) does $m = a$ and $n = b$. That's unlikely, however, because single-step reactions are uncommon.



b A manometer measures pressure changes that result from the production of gas as a reaction progresses. The reaction rate is directly proportional to the rate at which the pressure increases.

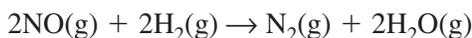


Review direct and inverse relationships in the **Math Handbook** on page 905 of this text.

CONNECTION

The time it takes for bonds to form and break during a chemical reaction is measured in femtoseconds. A femtosecond is one thousandth of a trillionth of a second. Until recently, scientists could only calculate and imagine the actual atomic activity of chemical bonding. In 1999, Dr. Ahmed Zewail of the California Institute of Technology won a Nobel Prize for his achievements in the field of femtochemistry. Zewail has developed an ultrafast laser device that can monitor and record chemical reactions in real time. Zewail's laser "flashes" every ten femtoseconds, allowing chemists to record changes in the wavelengths (colors) that vibrating molecules emitted during the course of a reaction. The changes correspond to bond formation and breakage and are mapped to the various intermediates and products that are formed during a reaction and that were previously impossible to witness.

For example, the reaction between nitrogen monoxide (NO) and hydrogen (H₂) is described by the following equation.



This reaction, which occurs in more than one step, has the following rate law.

$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

This rate law was determined experimentally. The data tell that the rate depends on the concentration of the reactants as follows. If [NO] doubles, the rate quadruples; if [H₂] doubles, the rate doubles. The reaction is described as second order in NO, first order in H₂, and third order overall because the sum of the orders for the individual reactants (the sum of the exponents) is (2 + 1), or 3.

Determining Reaction Order

One common experimental method of evaluating reaction order is called the method of initial rates. The **method of initial rates** determines reaction order by comparing the initial rates of a reaction carried out with varying reactant concentrations. To understand how this method works, let's use the general reaction $aA + bB \rightarrow \text{products}$. Suppose that this reaction is carried out with varying concentrations of A and B and yields the initial reaction rates shown in **Table 17-3**.

Table 17-3

Experimental Initial Rates for $aA + bB \rightarrow \text{products}$			
Trial	Initial [A] (M)	Initial [B] (M)	Initial Rate (mol/(L·s))
1	0.100	0.100	2.00×10^{-3}
2	0.200	0.100	4.00×10^{-3}
3	0.200	0.200	16.0×10^{-3}

Recall that the general rate law for this type of reaction is

$$\text{Rate} = k[\text{A}]^m[\text{B}]^n$$

To determine m , the concentrations and reaction rates in Trials 1 and 2 are compared. As you can see from the data, while the concentration of B remains constant, the concentration of A in Trial 2 is twice that of Trial 1. Note that the initial rate in Trial 2 is twice that of Trial 1. Because doubling [A] doubles the rate, the reaction must be first order in A. That is, because $2^m = 2$, m must equal 1. The same method is used to determine n , only this time Trials 2 and 3 are compared. Doubling the concentration of B causes the rate to increase by four times. Because $2^n = 4$, n must equal 2. This information suggests that the reaction is second order in B, giving the following overall rate law.

$$\text{Rate} = k[\text{A}]^1[\text{B}]^2$$

The overall reaction order is third order (sum of exponents $2 + 1 = 3$).

PRACTICE PROBLEMS

16. Write the rate law for the reaction $aA \rightarrow bB$ if the reaction is third order in A. [B] is not part of the rate law.
17. Given the following experimental data, use the method of initial rates to determine the rate law for the reaction $aA + bB \rightarrow \text{products}$. Hint: Any number to the zero power equals one. For example, $(0.22)^0 = 1$ and $(55.6)^0 = 1$.

Practice Problem 17 Experimental Data

Trial	Initial [A] (M)	Initial [B] (M)	Initial Rate (mol/(L·s))
1	0.100	0.100	2.00×10^{-3}
2	0.200	0.100	2.00×10^{-3}
3	0.200	0.200	4.00×10^{-3}

18. Given the following experimental data, use the method of initial rates to determine the rate law for the reaction $\text{CH}_3\text{CHO}(\text{g}) \rightarrow \text{CH}_4(\text{g}) + \text{CO}(\text{g})$.

Practice Problem 18 Experimental Data

Trial	Initial $[\text{CH}_3\text{CHO}]$ (M)	Initial rate (mol/(L·s))
1	2.00×10^{-3}	2.70×10^{-11}
2	4.00×10^{-3}	10.8×10^{-11}
3	8.00×10^{-3}	43.2×10^{-11}

In summary, the rate law for a reaction relates reaction rate, the rate constant k , and the concentration of the reactants. Although the equation for a reaction conveys a great deal of information, it is important to remember that the actual rate law and order of a complex reaction can be determined only by experiment.



For more practice determining reaction orders, go to **Supplemental Practice Problems** in Appendix A.

Section 17.3 Assessment

19. What does the rate law for a chemical reaction tell you about the reaction?
20. Use the rate law equations to show the difference between a first-order reaction having a single reactant and a second-order reaction having a single reactant.
21. What relationship is expressed by the specific rate constant for a chemical reaction?
22. **Thinking Critically** When giving the rate of a chemical reaction, explain why it is significant to know that the reaction rate is an average reaction rate.
23. **Designing an Experiment** Explain how you would design an experiment to determine the rate law for the general reaction $aA + bB \rightarrow \text{products}$ using the method of initial rates.

Objectives

- **Calculate** instantaneous rates of chemical reactions.
- **Understand** that many chemical reactions occur in steps.
- **Relate** the instantaneous rate of a complex reaction to its reaction mechanism.

Vocabulary

instantaneous rate
 complex reaction
 reaction mechanism
 intermediate
 rate-determining step

The average reaction rate you learned to calculate in Section 17.3 gives important information about the reaction over a period of time. However, chemists also may need to know at what rate the reaction is proceeding at a specific time. A pharmacist developing a new drug treatment might need to know the progress of a reaction at an exact instant. This information makes it possible to adjust the product for maximum performance. Can you think of other situations where it might be critical to know the specific reaction rate at a given time?

Instantaneous Reaction Rates

The decomposition of hydrogen peroxide (H_2O_2) is represented as follows.



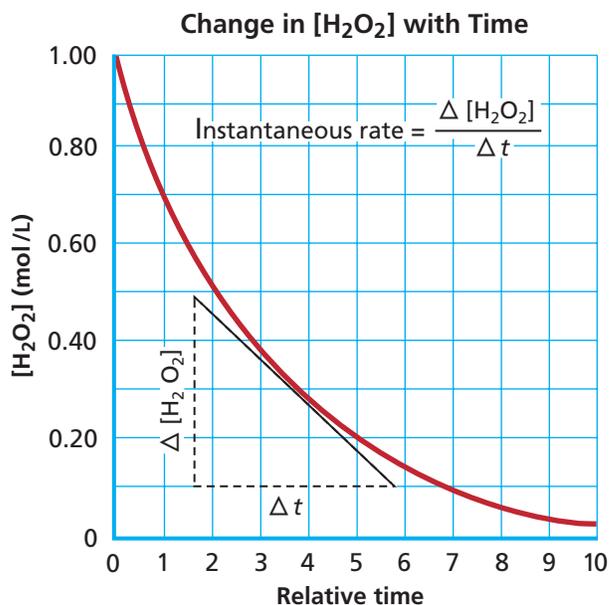
For this reaction, the decrease in H_2O_2 concentration over time is shown in **Figure 17-15**. The curved line shows how the reaction rate decreases as the reaction proceeds. The **instantaneous rate**, or the rate of decomposition at a specific time, can be determined by finding the slope of the straight line tangent to the curve at that instant. This is because the slope of the tangent to the curve at a particular point is $\Delta[\text{H}_2\text{O}_2]/\Delta t$, which is one way to express the reaction rate. In other words, the rate of change in H_2O_2 concentration relates to one specific point (or instant) on the graph.

Another way to determine the instantaneous rate for a chemical reaction is to use the experimentally determined rate law, given the reactant concentrations and the specific rate constant for the temperature at which the reaction occurs. For example, the decomposition of dinitrogen pentoxide (N_2O_5) into nitrogen dioxide (NO_2) and oxygen (O_2) is given by the following equation.



Figure 17-15

The instantaneous rate for a specific point in the reaction progress can be determined from the tangent to the curve that passes through that point. The equation for the slope of the line ($\Delta y/\Delta x$) is the equation for instantaneous rate in terms of $\Delta[\text{H}_2\text{O}_2]$ and Δt :

$$\text{instantaneous rate} = \frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t}$$


The experimentally determined rate law for this reaction is

$$\text{Rate} = k[\text{N}_2\text{O}_5]$$

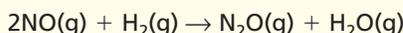
where $k = 1.0 \times 10^{-5} \text{ s}^{-1}$. If $[\text{N}_2\text{O}_5] = 0.350M$, the instantaneous reaction rate would be calculated as

$$\begin{aligned}\text{Rate} &= (1.0 \times 10^{-5} \text{ s}^{-1})(0.350 \text{ mol/L}) \\ &= 3.5 \times 10^{-6} \text{ mol/(L}\cdot\text{s)}\end{aligned}$$

EXAMPLE PROBLEM 17-2

Calculating Instantaneous Reaction Rates

The following equation is first order in H_2 and second order in NO with a rate constant of $2.90 \times 10^2 \text{ (L}^2\text{/(mol}^2\cdot\text{s))}$.



Calculate the instantaneous rate when the reactant concentrations are $[\text{NO}] = 0.00200M$ and $[\text{H}_2] = 0.00400M$.

1. Analyze the Problem

The rate law can be expressed by $\text{Rate} = k[\text{NO}]^2[\text{H}_2]$. Therefore, the instantaneous reaction rate can be determined by inserting reactant concentrations and the specific rate constant into the rate law equation.

Known

$$[\text{NO}] = 0.00200M$$

$$[\text{H}_2] = 0.00400M$$

$$k = 2.90 \times 10^2 \text{ (L}^2\text{/(mol}^2\cdot\text{s))}$$

Unknown

$$\text{Rate} = ? \text{ mol/(L}\cdot\text{s)}$$

2. Solve for the Unknown

Insert the known quantities into the rate law equation.

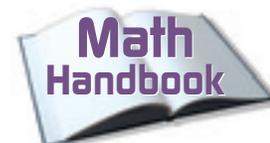
$$\text{Rate} = k[\text{NO}]^2[\text{H}_2]$$

$$\text{Rate} = 2.90 \times 10^2 \text{ (L}^2\text{/(mol}^2\cdot\text{s))}(0.00200 \text{ mol/L})^2(0.00400 \text{ mol/L})$$

$$\text{Rate} = 4.64 \times 10^{-6} \text{ mol/(L}\cdot\text{s)}$$

3. Evaluate the Answer

Are the units correct? Units in the calculation cancel to give $\text{mol/(L}\cdot\text{s)}$, which is the common unit for the expression of reaction rates. Is the magnitude reasonable? A magnitude of approximately $10^{-6} \text{ mol/(L}\cdot\text{s)}$ fits with the quantities given and the rate law equation.



Review solving algebraic equations in the **Math Handbook** on page 897 of this text.

PRACTICE PROBLEMS

Use the rate law in Example Problem 17-2 and the concentrations given in each practice problem to calculate the instantaneous rate for the reaction between NO and H_2 .

24. $[\text{NO}] = 0.00500M$ and $[\text{H}_2] = 0.00200M$

25. $[\text{NO}] = 0.0100M$ and $[\text{H}_2] = 0.00125M$

26. $[\text{NO}] = 0.00446M$ and $[\text{H}_2] = 0.00282M$



For more practice calculating instantaneous reaction rates, go to **Supplemental Practice Problems** in Appendix A.

Careers Using Chemistry

Food Technologist

Everyone needs to eat, and in today's fast-paced world, we've become accustomed to having safe, convenient, and nutritious food at our fingertips, ready to toss in the microwave. Food technologists help make nutrition in our fast-paced lifestyle more convenient by developing new ways to process, preserve, package, and store food.

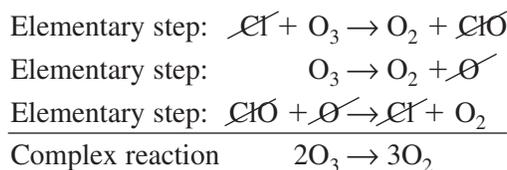
Food technologists are also called food chemists, food scientists, and food engineers. They find ways to make food products more healthful, more convenient, and better tasting. Part of their job is searching for ways to stop or slow chemical reactions and lengthen the shelf life of food. Food technologists might find careers in the food industry, universities, or the federal government.

Reaction Mechanisms

Most chemical reactions consist of a sequence of two or more simpler reactions. For example, recent evidence indicates that the reaction $2\text{O}_3 \rightarrow 3\text{O}_2$ occurs in three steps after intense ultraviolet radiation from the sun liberates chlorine atoms from certain compounds in Earth's stratosphere. Steps 1 and 2 in this reaction may occur simultaneously or in reverse order.

1. Chlorine atoms decompose ozone according to the equation $\text{Cl} + \text{O}_3 \rightarrow \text{O}_2 + \text{ClO}$.
2. Ultraviolet radiation causes the decomposition reaction $\text{O}_3 \rightarrow \text{O}_2 + \text{O}$.
3. ClO produced in the reaction in step 1 reacts with O produced in step 2 according to the equation $\text{ClO} + \text{O} \rightarrow \text{Cl} + \text{O}_2$.

Each of the reactions described in steps 1 through 3 is called an elementary step. These three elementary steps comprise the complex reaction $2\text{O}_3 \rightarrow 3\text{O}_2$. A **complex reaction** is one that consists of two or more elementary steps. The complete sequence of elementary steps that make up a complex reaction is called a **reaction mechanism**. Adding the elementary steps in steps 1 through 3 and canceling formulas that occur in equal amounts on both sides of the reaction arrow produce the net equation for the complex reaction as shown below:



Because chlorine atoms react in step 1 of the reaction mechanism and are re-formed in step 3, chlorine is said to catalyze the decomposition of ozone. Do you know why? Because ClO and O are formed in the reactions in steps 1 and 2, respectively, and are consumed in the reaction in step 3, they are called intermediates. In a complex reaction, an **intermediate** is a substance produced in one elementary step and consumed in a subsequent elementary step. Catalysts and intermediates do not appear in the net chemical equation.

Rate-determining step You have probably heard the expression, “A chain is no stronger than its weakest link.” Chemical reactions, too, have a “weakest link” in that a complex reaction can proceed no faster than the slowest of its elementary steps. In other words, the slowest elementary step in a reaction mechanism limits the instantaneous rate of the overall reaction.

Figure 17-16

In an automobile assembly plant, the process that is most time consuming limits the production rate. For example, because it takes longer to install the onboard computer system than it takes to attach the hood assembly, the computer installation is a rate-determining step.



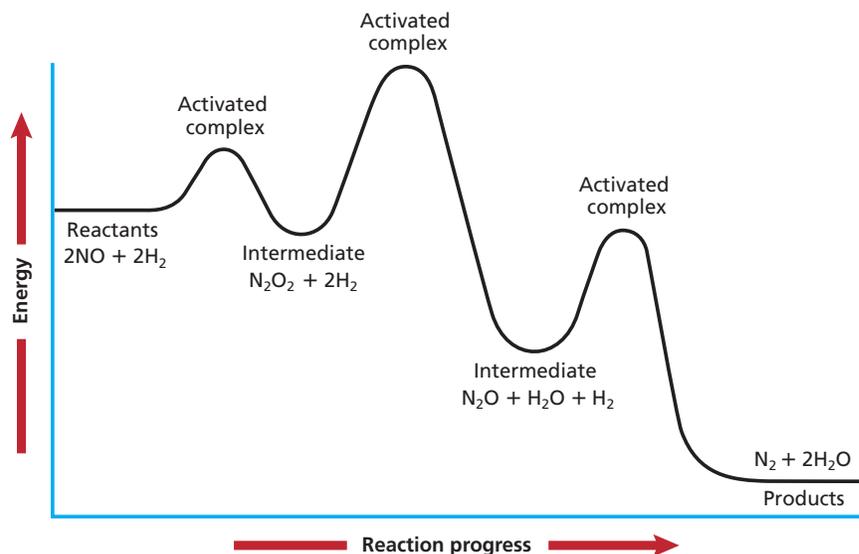
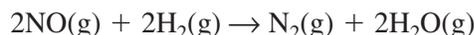


Figure 17-17

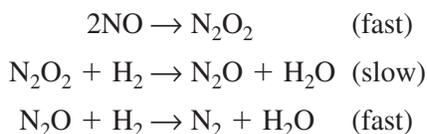
The three humps in this energy diagram correspond to activation energies for the intermediate steps of the reaction. The middle hump represents the highest energy barrier to overcome; therefore, the reaction involving $\text{N}_2\text{O}_2 + 2\text{H}_2$ is the rate-determining step.

For that reason, the slowest of the elementary steps in a complex reaction is called the **rate-determining step**. **Figure 17-16** illustrates this concept in a manufacturing plant.

To see how the rate-determining step affects reaction rate, again consider the gas-phase reaction between hydrogen and nitrogen monoxide discussed earlier in the chapter.



A proposed mechanism for this reaction consists of the following elementary steps.



Although the first and third elementary steps occur relatively fast, the middle step is the slowest, and it limits the overall reaction rate. It is the rate-determining step. The relative energy levels of reactants, intermediates, products, and activated complex are illustrated in **Figure 17-17**.

Chemists can make use of instantaneous reaction rates, reaction orders, and rate-determining steps to develop efficient ways to manufacture products such as pharmaceuticals. By knowing which step of the reaction is slowest, a chemist can work to speed up that rate-determining step and thereby increase the rate of the overall reaction.

Section 17.4 Assessment

27. How can the rate law for a chemical reaction be used to determine an instantaneous reaction rate?
28. Compare and contrast an elementary chemical reaction with a complex chemical reaction.
29. What is a reaction mechanism? An intermediate?
30. **Thinking Critically** How can you determine whether a product of one of the elementary steps in a complex reaction is an intermediate?
31. **Communicating** How would you explain the significance of the rate-determining step in a chemical reaction?

Concentration and Reaction Rate

The collision theory describes how the change in concentration of one reactant affects the rate of chemical reactions. In this laboratory experiment you will observe how concentration affects the reaction rate.

Problem

How does the concentration of a reactant affect the reaction rate?

Objectives

- **Sequence** the acid concentrations from the most to the least concentrated.
- **Observe** which concentration results in the fastest reaction rate.

Materials

graduated pipette	magnesium ribbon
10-mL	emery cloth or fine sandpaper
safety pipette filler	scissors
6M hydrochloric acid	plastic ruler
distilled water	tongs
25 mm × 150 mm test tubes (4)	watch with second hand
test-tube rack	stirring rod

Safety Precautions



- Always wear safety goggles and an apron in the lab.
- Never pipette any chemical by mouth.
- No open flame.

Pre-Lab

1. Read the entire CHEMLAB. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
2. Use emery paper or sand paper to polish the magnesium ribbon until it is shiny. Use scissors to cut the magnesium into four 1-cm pieces.
3. Place the four test tubes in the test-tube rack. Label the test tubes #1 (6.0M HCl), #2 (3M HCl), #3 (1.5M HCl), and #4 (0.75M HCl).
4. Form a hypothesis about how the chemical reaction rate is related to reactant concentration.
5. What reactant quantity is held constant? What are the independent and dependent variables?
6. What gas is produced in the reaction between magnesium and hydrochloric acid? Write the balanced formula equation for the reaction.
7. Why is it important to clean the magnesium ribbon? If one of the four pieces is not thoroughly polished, how will the rate of the reaction involving that piece be affected?

Reaction Time Data

Test tube	[HCl] (M)	Time (s)
1		
2		
3		
4		

Procedure

1. Use a safety pipette to draw 10 mL of 6.0M hydrochloric acid (HCl) into a 10-mL graduated pipette.
2. Dispense the 10 mL of 6.0M HCl into test tube #1.
3. Draw 5.0 mL of the 6.0M HCl from test tube #1 with the empty pipette. Dispense this acid into test tube #2 and use the pipette to add an additional

5.0 mL of distilled water to the acid. Use the stirring rod to mix thoroughly. This solution is 3.0M HCl.

4. Draw 5.0 mL of the 3.0M HCl from test tube #2 with the empty pipette. Dispense this acid into test tube #3 and use the pipette to add an additional 5.0 mL of distilled water to the acid. Use the stirring rod to mix thoroughly. This solution is 1.5M HCl.
5. Draw 5.0 mL of the 1.5M HCl from test tube #3 with the empty pipette. Dispense this acid into test tube #4 and use the pipette to add an additional 5.0 mL of distilled water to the acid. Use the stirring rod to mix thoroughly. This solution is 0.75M HCl.
6. Draw 5.0 mL of the 0.75M HCl from test tube #4 with the empty pipette. Neutralize and discard it in the sink.
7. Using the tongs, place a 1-cm length of magnesium ribbon into test tube #1. Record the time in seconds that it takes for the bubbling to stop.
8. Repeat step 7 using the remaining three test tubes of HCl and the three remaining pieces of magnesium ribbon. Record in your data table the time (in seconds) it takes for the bubbling to stop.

Cleanup and Disposal

1. Place acid solutions in an acid discard container. Your teacher will neutralize the acid for proper disposal.
2. Wash thoroughly all test tubes and lab equipment.
3. Discard other materials as directed by your teacher.
4. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Analyzing** In step 6, why is 5.0 mL HCl discarded?
2. **Making and Using Graphs** Plot the concentration of the acid on the x -axis and time it takes for the bubbling to stop on the y -axis. Draw a smooth curve through the data points.
3. **Interpreting Graphs** Is the curve in question 3 linear or nonlinear? What does the slope tell you?
4. **Drawing a Conclusion** Based on your graph, what do you conclude about the relationship between the acid concentration and the reaction rate?



5. **Hypothesizing** Write an hypothesis using the collision theory, reaction rate, and reactant concentration to explain your results.
6. **Designing an Experiment** Write a brief statement of how you would set up an experiment to test your hypothesis.
7. **Error Analysis** Compare your experimental results with those of several other students in the laboratory. Explain the differences.

Real-World Chemistry

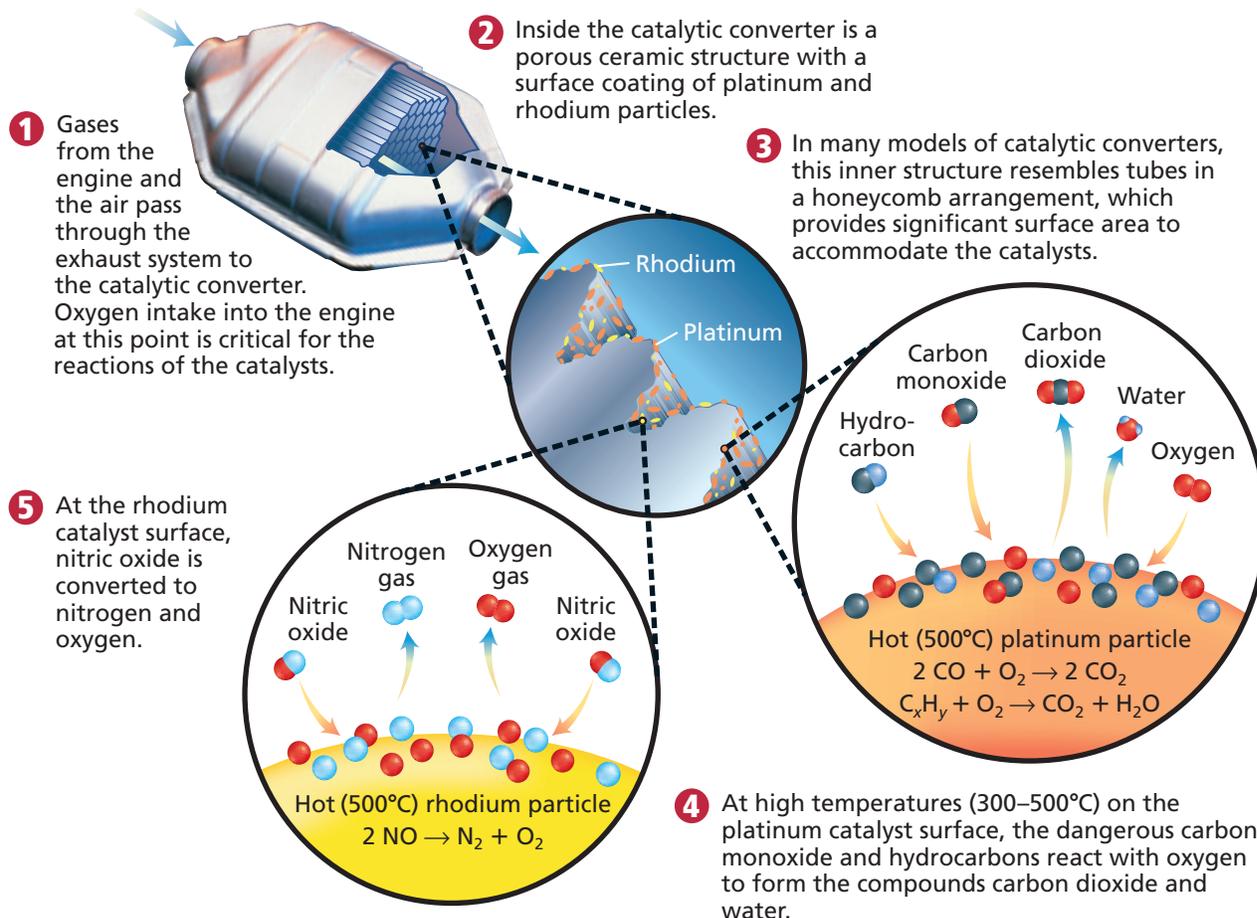
1. Describe a situation that may occur in your daily life that exemplifies the effect of concentration on the rate of a reaction.
2. Some hair-care products, such as hot-oil treatments, must be heated before application. Explain in terms of factors affecting reaction rates why heat is required.

How It Works

Catalytic Converter

Concerns for our environment have made a huge impact in the products and processes we use every day. The catalytic converter is one of those impacts.

When a combustion engine converts fuel into energy, the reactions of the combustion process are incomplete. Incomplete combustion results in the production of poisonous carbon monoxide and undesirable nitrogen oxides. Since 1975, catalytic converters have reduced the exhaust emissions that contribute to air pollution by approximately 90%.



Thinking Critically

1. Inferring Why is it necessary to have additional air in exhaust before it enters the catalytic converter?

2. Hypothesizing A catalytic converter is not effective when cold. Use the concept of activation energy to explain why.

Summary

17.1 A Model for Reaction Rates

- The rate of a chemical reaction is expressed as the rate at which a reactant is consumed or the rate at which a product is formed.
- Reaction rates are generally calculated and expressed in moles per liter per second (mol/(L·s)).
- In order to react, the particles in a chemical reaction must collide in a correct orientation and with sufficient energy to form the activated complex.
- The rate of a chemical reaction is unrelated to the spontaneity of the reaction.

17.2 Factors Affecting Reaction Rates

- Key factors that influence the rate of chemical reactions include reactivity, concentration, surface area, temperature, and catalysts.
- Catalysts increase the rates of chemical reactions by lowering activation energies.
- Raising the temperature of a reaction increases the rate of the reaction by increasing the collision frequency and the number of collisions forming the activated complex.

17.3 Reaction Rate Laws

- The mathematical relationship between the rate of a chemical reaction at a given temperature and the concentrations of reactants is called the rate law.
- The rate law for a chemical reaction is determined experimentally using the method of initial rates.

17.4 Instantaneous Reaction Rates and Reaction Mechanisms

- The instantaneous rate for a chemical reaction is calculated from the rate law, the specific rate constant, and the concentrations of all reactants.
- Most chemical reactions are complex reactions consisting of two or more elementary steps.
- A reaction mechanism is the complete sequence of elementary steps that make up a complex reaction.
- For a complex reaction, the rate-determining step limits the instantaneous rate of the overall reaction.

Key Equations and Relationships

- Average rate = $\frac{\Delta \text{quantity}}{\Delta t}$
(p. 529)

- Rate = $k[A]$
(p. 542)

- Rate = $k[A]^m[B]^n$
(p. 543)

Vocabulary

- activated complex (p. 532)
- activation energy (p. 533)
- catalyst (p. 539)
- collision theory (p. 532)
- complex reaction (p. 548)
- heterogeneous catalyst (p. 541)
- homogeneous catalyst (p. 541)
- inhibitor (p. 540)
- instantaneous rate (p. 546)
- intermediate (p. 548)
- method of initial rates (p. 544)
- rate law (p. 542)
- rate-determining step (p. 549)
- reaction mechanism (p. 548)
- reaction order (p. 543)
- reaction rate (p. 530)
- specific rate constant (p. 542)
- transition state (p. 532)

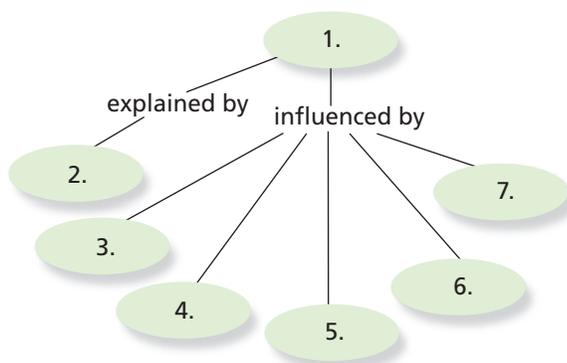


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 17 Assessment.

Concept Mapping

32. Complete the following concept map using the following terms: surface area, collision theory, temperature, reaction rates, concentration, reactivity, catalyst.



Mastering Concepts

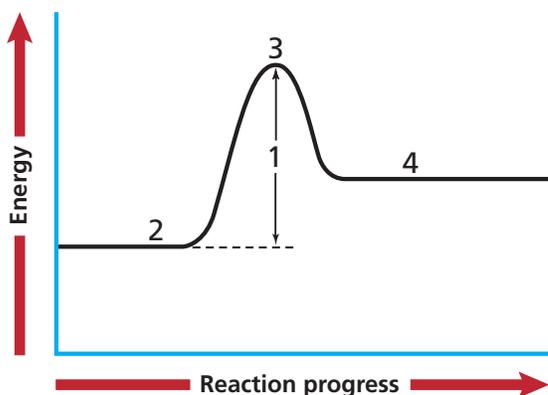
33. For a specific chemical reaction, assume that the change in free energy (ΔG) is negative. What does this information tell you about the rate of the reaction? (17.1)
34. How would you express the rate of the chemical reaction $A \rightarrow B$ based on the concentration of reactant A? How would that rate compare with the reaction rate based on the product B? (17.1)
35. What does the activation energy for a chemical reaction represent? (17.1)
36. What is the role of the activated complex in a chemical reaction? (17.1)
37. Suppose two molecules that can react collide. Under what circumstances do the colliding molecules not react? (17.1)
38. How is the activation energy for a chemical reaction related to whether or not a collision between molecules initiates a reaction? (17.1)
39. In the activated complex for a chemical reaction, what bonds are broken and what bonds are formed? (17.1)
40. If $A \rightarrow B$ is exothermic, how does the activation energy for the forward reaction compare with the activation energy for the reverse reaction ($A \leftarrow B$)? (17.1)
41. What role does the reactivity of the reactants play in determining the rate of a chemical reaction? (17.2)
42. Explain why a crushed solid reacts with a gas more quickly than a large chunk of the same solid. (17.2)
43. What do you call a substance that increases the rate of a chemical reaction without being consumed in the reaction? (17.2)
44. In general, what is the relationship between reaction rate and reactant concentration? (17.2)
45. In general, what is the relationship between reaction rate and temperature? (17.2)
46. Distinguish between a homogeneous catalyst and a heterogeneous catalyst. (17.2)
47. Explain how a catalyst affects the activation energy for a chemical reaction. (17.2)
48. Use the collision theory to explain why increasing the concentration of a reactant usually increases the reaction rate. (17.2)
49. Use the collision theory to explain why increasing the temperature usually increases the reaction rate. (17.2)
50. In a chemical reaction, what relationship does the rate law describe? (17.3)
51. What is the name of the proportionality constant in the mathematical expression that relates reaction rate and reactant concentration? (17.3)
52. What does the order of a reactant tell you about the way the concentration of that reactant appears in the rate law? (17.3)
53. Why does the specific rate constant for a chemical reaction often double for each increase of 10 K? (17.3)
54. Explain why the rates of most chemical reactions decrease over time. (17.3)
55. In the method of initial rates used to determine the rate law for a chemical reaction, what is the significance of the word initial? (17.3)
56. If a reaction has three reactants and is first order in one, second order in another, and third order in the third, what is the overall order of the reaction? (17.3)
57. What do you call the slowest of the elementary steps that make up a complex reaction? (17.4)
58. What is an intermediate in a complex reaction? (17.4)
59. Distinguish between an elementary step, a complex reaction, and a reaction mechanism. (17.4)
60. Under what circumstances is the rate law for the reaction $2A + 3B \rightarrow \text{products}$ correctly written as $\text{Rate} = k[A]^2[B]^3$? (17.3)

61. How does the activation energy of the rate-determining step in a complex reaction compare with the activation energies of the other elementary steps? (17.4)

Mastering Problems

A Model for Reaction Rates (17.1)

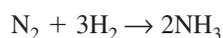
62. In the gas-phase reaction $I_2 + Cl_2 \rightarrow 2ICl$, the $[I_2]$ changes from $0.400M$ at time = 0 to $0.300M$ at time = 4.00 min. Calculate the average reaction rate in moles I_2 consumed per liter per minute.
63. If a chemical reaction occurs at a rate of 2.25×10^{-2} moles per liter per second at 322 K, what is the rate expressed in moles per liter per minute?
64. On the accompanying energy level diagram, match the appropriate number with the quantity it represents.
- a. reactants c. products
b. activated complex d. activation energy



65. Given the following data for the decomposition of hydrogen peroxide, calculate the average reaction rate in moles H_2O_2 consumed per liter per minute for each time interval.

Decomposition of H_2O_2	
Time (min)	$[H_2O_2]$ (M)
0	2.50
2	2.12
5	1.82
10	1.48
20	1.00

66. At a given temperature and for a specific time interval, the average rate of the following reaction is 1.88×10^{-4} moles N_2 consumed per liter per second.



Express the reaction rate in moles H_2 consumed per liter per second and in moles NH_3 produced per liter per second.

Factors Affecting Reaction Rates (17.2)

67. Estimate the rate of the reaction described in problem 63 at 332 K. Express the rate in moles per liter per second.
68. Estimate the rate of the reaction described in problem 63 at 352 K and with $[I_2]$ doubled (assume the reaction is first order in I_2).

Reaction Rate Laws (17.3)

69. Nitrogen monoxide gas and chlorine gas react according to the equation $2NO + Cl_2 \rightarrow 2NOCl$. Use the following data to determine the rate law for the reaction by the method of initial rates. Also, calculate the value of the specific rate constant.

2NO + Cl₂ Reaction Data

Initial [NO] (M)	Initial [Cl ₂] (M)	Initial rate (mol/(L·min))
0.50	0.50	1.90×10^{-2}
1.00	0.50	7.60×10^{-2}
1.00	1.00	15.20×10^{-2}

70. Use the following data to determine the rate law and specific rate constant for the reaction $2ClO_2(aq) + 2OH^-(aq) \rightarrow ClO_3^-(aq) + ClO_2^-(aq) + H_2O(l)$.

2ClO₂(aq) + 2OH⁻(aq) Reaction Data

Initial [ClO ₂] (M)	Initial [OH ⁻] (M)	Initial rate (mol/(L·min))
0.0500	0.200	6.90
0.100	0.200	27.6
0.100	0.100	13.8

Instantaneous Reaction Rates and Reaction Mechanisms (17.4)

71. The gas-phase reaction $2HBr + NO_2 \rightarrow H_2O + NO + Br_2$ is thought to occur by the following mechanism.
- $$HBr + NO_2 \rightarrow HOBr + NO \quad \Delta H = 4.2 \text{ kJ (slow)}$$
- $$HBr + HOBr \rightarrow H_2O + Br_2 \quad \Delta H = -86.2 \text{ kJ (fast)}$$

Draw the energy diagram that depicts this reaction mechanism. On the diagram, show the energy of the reactants, energy of the products, and relative activation energies of the two elementary steps.

- 72.** Are there any intermediates in the complex reaction described in problem 71? Explain why or why not. If any intermediates exist, what are their formulas?
- 73.** Given the rate law $\text{Rate} = k[\text{A}][\text{B}]^2$ for the generic reaction $\text{A} + \text{B} \rightarrow \text{products}$, the value for the specific rate constant ($4.75 \times 10^{-7} \text{ L}^2/(\text{mol}^2 \cdot \text{s})$), the concentration of A ($0.355M$), and the concentration of B ($0.0122M$), calculate the instantaneous reaction rate.

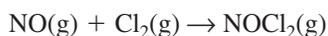
Mixed Review

Sharpen your problem-solving skills by answering the following.

- 74.** Use the method of initial rates and the following data to determine and express the rate law for the reaction $\text{A} + \text{B} \rightarrow 2\text{C}$.

A + B → 2C Reaction Data		
Initial [A] (M)	Initial [B] (M)	Initial rate (mol C/(L·s))
0.010	0.010	0.0060
0.020	0.010	0.0240
0.020	0.020	0.0960

- 75.** The concentration of reactant A decreases from 0.400 mol/L at time = 0 to 0.384 mol/L at time = 4.00 min . Calculate the average reaction rate during this time period. Express the rate in $\text{mol}/(\text{L} \cdot \text{min})$.
- 76.** It is believed that the following two elementary steps make up the mechanism for the reaction between nitrogen monoxide and chlorine:



Write the equation for the overall reaction and identify any intermediates in the reaction mechanism.

- 77.** One reaction that takes place in an automobile's engine and exhaust system is described by the equation $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{CO}_2(\text{g})$. This reaction's rate law at a particular temperature is given by the relationship $\text{rate} = 0.50 \text{ L}/(\text{mol} \cdot \text{s})[\text{NO}_2]^2$. What is the reaction's initial, instantaneous rate at $[\text{NO}_2] = 0.0048 \text{ mol/L}$?
- 78.** At 232 K , the rate of a certain chemical reaction is $3.20 \times 10^{-2} \text{ mol}/(\text{L} \cdot \text{min})$. Predict the reaction's approximate rate at 252 K .
- 79. Using Numbers** Draw a diagram that shows all of the possible collision combinations between two molecules of reactant A and two molecules of reactant B. Now, increase the number of molecules of A from two to four and sketch each possible A–B collision combination. By what factor did the number of collision combinations increase? What does this imply about the reaction rate?
- 80. Applying Concepts** Use the collision theory to explain two reasons why increasing the temperature of a reaction by 10 K often doubles the reaction rate.
- 81. Formulating Models** Create a table of concentrations, starting with $0.100M$ concentrations of all reactants, that you would propose in order to establish the rate law for the reaction $a\text{A} + b\text{B} + c\text{D} \rightarrow \text{products}$ using the method of initial rates.
- 82.** Research the way manufacturers in the United States produce nitric acid from ammonia. Write the reaction mechanism for the complex reaction. If catalysts are used in the process, explain how they are used and how they affect any of the elementary steps.
- 83.** Write an advertisement that explains why Company A's lawn care product (fertilizer or weed killer) works better than the competition's because of the smaller sized granules. Include applicable diagrams.

Thinking Critically

Writing in Chemistry

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 84.** Classify each of the following elements as a metal, nonmetal, or metalloid. (Chapter 6)
- molybdenum
 - bromine
 - arsenic
 - neon
 - cerium
- 85.** Balance the following equations. (Chapter 10)
- $\text{Sn}(\text{s}) + \text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SnO}_2 + \text{H}_2$
 - $\text{C}_8\text{H}_{18}(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{Al}(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Al}_2(\text{SO}_4)_3(\text{aq}) + \text{H}_2(\text{g})$
- 86.** What mass of iron(III) chloride is needed to prepare 1.00 L of a $0.255M$ solution? (Chapter 15)
- 87.** ΔH for a reaction is negative. Compare the energy of the products and the reactants. Is the reaction endothermic or exothermic? (Chapter 16)

Use these questions and the test-taking tip to prepare for your standardized test.

- The rate of a chemical reaction is all of the following EXCEPT _____.
 - the speed at which a reaction takes place
 - the change in concentration of a reactant per unit time
 - the change in concentration of a product per unit time
 - the amount of product formed in a certain period of time
- The complete dissociation of acid H_3A takes place in three steps:

$$\text{H}_3\text{A}(\text{aq}) \rightarrow \text{H}_2\text{A}^-(\text{aq}) + \text{H}^+(\text{aq})$$

$$\text{Rate} = k_1[\text{H}_3\text{A}] \quad k_1 = 3.2 \times 10^2 \text{ s}^{-1}$$

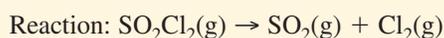
$$\text{H}_2\text{A}^-(\text{aq}) \rightarrow \text{HA}^{2-}(\text{aq}) + \text{H}^+(\text{aq})$$

$$\text{Rate} = k_2[\text{H}_2\text{A}^-] \quad k_2 = 1.5 \times 10^2 \text{ s}^{-1}$$

$$\text{HA}^{2-}(\text{aq}) \rightarrow \text{A}^{3-}(\text{aq}) + \text{H}^+(\text{aq})$$

$$\text{Rate} = k_3[\text{HA}^{2-}] \quad k_3 = 0.8 \times 10^2 \text{ s}^{-1}$$
 overall reaction: $\text{H}_3\text{A}(\text{aq}) \rightarrow \text{A}^{3-}(\text{aq}) + 3\text{H}^+(\text{aq})$
 When the reactant concentrations are $[\text{H}_3\text{A}] = 0.100M$, $[\text{H}_2\text{A}^-] = 0.500M$, and $[\text{HA}^{2-}] = 0.200M$, which reaction is the rate-determining step?
 - $\text{H}_3\text{A}(\text{aq}) \rightarrow \text{H}_2\text{A}^-(\text{aq}) + \text{H}^+(\text{aq})$
 - $\text{H}_2\text{A}^-(\text{aq}) \rightarrow \text{HA}^{2-}(\text{aq}) + \text{H}^+(\text{aq})$
 - $\text{HA}^{2-}(\text{aq}) \rightarrow \text{A}^{3-}(\text{aq}) + \text{H}^+(\text{aq})$
 - $\text{H}_3\text{A}(\text{aq}) \rightarrow \text{A}^{3-}(\text{aq}) + 3\text{H}^+(\text{aq})$
- Which of the following is NOT an acceptable unit for expressing rate?
 - M/min
 - L/s
 - $\text{mol}/mL \cdot h$
 - $\text{mol}/L \cdot \text{min}$

Interpreting Tables Use the table to answer questions 4–6.



Experimental Data Collected for Reaction			
Time (min)	$[\text{SO}_2\text{Cl}_2]$ (M)	$[\text{SO}_2]$ (M)	$[\text{Cl}_2]$ (M)
0.0	1.00	0.00	0.00
100.0	0.87	0.13	0.13
200.0	0.74	?	?

- The average reaction rate for this reaction, expressed in moles SO_2Cl_2 consumed per liter per minute, is _____.
 - $1.30 \times 10^{-3} \text{ mol}/L \cdot \text{min}$
 - $2.60 \times 10^{-1} \text{ mol}/L \cdot \text{min}$
 - $7.40 \times 10^{-3} \text{ mol}/L \cdot \text{min}$
 - $8.70 \times 10^{-3} \text{ mol}/L \cdot \text{min}$
- On the basis of the average reaction rate, the concentrations of SO_2 and Cl_2 at 200.0 min will be _____.
 - 0.130M
 - 0.260M
 - 0.39M
 - 0.52M
- How long will it take for half of the original amount of SO_2Cl_2 to decompose at the average reaction rate?
 - 285 min
 - 335 min
 - 385 min
 - 500 min
- Which of the following does NOT affect reaction rate?
 - catalysts
 - surface area of reactants
 - concentration of reactants
 - reactivity of products
- The reaction between persulfate ($\text{S}_2\text{O}_8^{2-}$) and iodide (I^-) ions is often studied in student laboratories because it occurs slowly enough for its rate to be measured:

$$\text{S}_2\text{O}_8^{2-}(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow 2\text{SO}_4^{2-}(\text{aq}) + \text{I}_2(\text{aq})$$
 This reaction has been experimentally determined to be first order in $\text{S}_2\text{O}_8^{2-}$ and first order in I^- . Therefore, the overall rate law for this reaction is
 - $\text{Rate} = k[\text{S}_2\text{O}_8^{2-}]^2[\text{I}^-]$
 - $\text{Rate} = k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]$
 - $\text{Rate} = k[\text{S}_2\text{O}_8^{2-}][\text{I}^-]^2$
 - $\text{Rate} = k[\text{S}_2\text{O}_8^{2-}]^2[\text{I}^-]^2$
- The rate law for the reaction $\text{A} + \text{B} + \text{C} \rightarrow \text{products}$ is:

$$\text{Rate} = k[\text{A}]^2[\text{C}]$$
 If $k = 6.92 \times 10^{-5} \text{ L}^3/\text{mol}^3 \cdot \text{s}$, $[\text{A}] = 0.175M$, $[\text{B}] = 0.230M$, and $[\text{C}] = 0.315M$, what is the instantaneous reaction rate?
 - $6.68 \times 10^{-7} \text{ mol}/L \cdot \text{s}$
 - $8.77 \times 10^{-7} \text{ mol}/L \cdot \text{s}$
 - $1.20 \times 10^{-6} \text{ mol}/L \cdot \text{s}$
 - $3.81 \times 10^{-6} \text{ mol}/L \cdot \text{s}$



TEST-TAKING TIP

Watch the Little Words Underline words like *least*, *not*, and *except* when you see them in test questions. They change the meaning of the question!

Chemical Equilibrium

What You'll Learn

- ▶ You will discover that many reactions and processes reach a state of equilibrium.
- ▶ You will use Le Châtelier's principle to explain how various factors affect chemical equilibria.
- ▶ You will calculate equilibrium concentrations of reactants and products using the equilibrium constant expression.
- ▶ You will determine the solubilities of sparingly soluble ionic compounds.

Why It's Important

The concentrations of substances called acids and bases in your blood are crucial to your health. These substances continuously enter and leave your bloodstream, but the chemical equilibria among them maintain the balance needed for good health.

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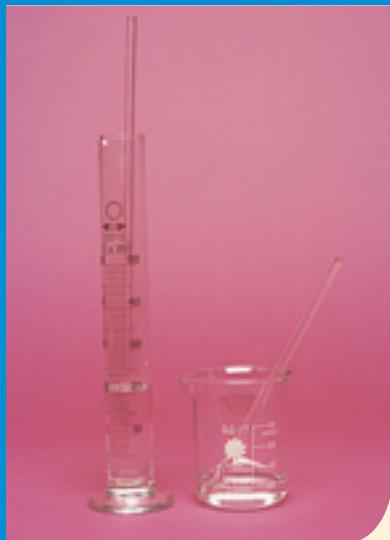


Visit the Chemistry Web site at science.glencoe.com to find links about chemical equilibrium.

One of the most important of ammonia's many uses is as a fertilizer.



DISCOVERY LAB



Materials

100-mL beaker
graduated cylinder
glass tubes, equal diameter,
open at both ends (2)

What's equal about equilibrium?

Does equilibrium mean that the amounts of reactants and products are equal?

Safety Precautions



Always wear safety goggles and a lab apron.

Procedure

1. Measure 20 mL of water in a graduated cylinder and pour it into a 100-mL beaker. Fill the graduated cylinder to the 20-mL mark. Place a glass tube in the graduated cylinder and another glass tube in the beaker. The tubes should reach the bottoms of the containers.
2. Cover the open ends of both glass tubes with your index fingers. Simultaneously, transfer the water from the cylinder to the beaker, and from the beaker to the cylinder.
3. Repeat the transfer process about 25 times. Record your observations.

Analysis

How can you explain your observations during the transfer process? What does this tell you about the concept of equilibrium?

Section

18.1

Equilibrium: A State of Dynamic Balance

Objectives

- **Recognize** the characteristics of chemical equilibrium.
- **Write** equilibrium expressions for systems that are at equilibrium.
- **Calculate** equilibrium constants from concentration data.

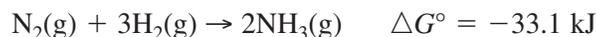
Vocabulary

reversible reaction
chemical equilibrium
law of chemical equilibrium
equilibrium constant
homogeneous equilibrium
heterogeneous equilibrium

When you get off a whirling amusement park ride, you probably pause a minute to “get your equilibrium.” If so, you are talking about getting your balance back after the ride exerted rapidly changing forces on you. But soon you are balanced steadily on your feet once more. Often, chemical reactions also reach a point of balance or equilibrium. The **DISCOVERY LAB** is an analogy for chemical equilibrium. You found that a point of balance was reached in the transfer of water from the beaker to the graduated cylinder and from the graduated cylinder to the beaker.

What is equilibrium?

Consider the reaction for the formation of ammonia from nitrogen and hydrogen that you learned about in Chapter 16.

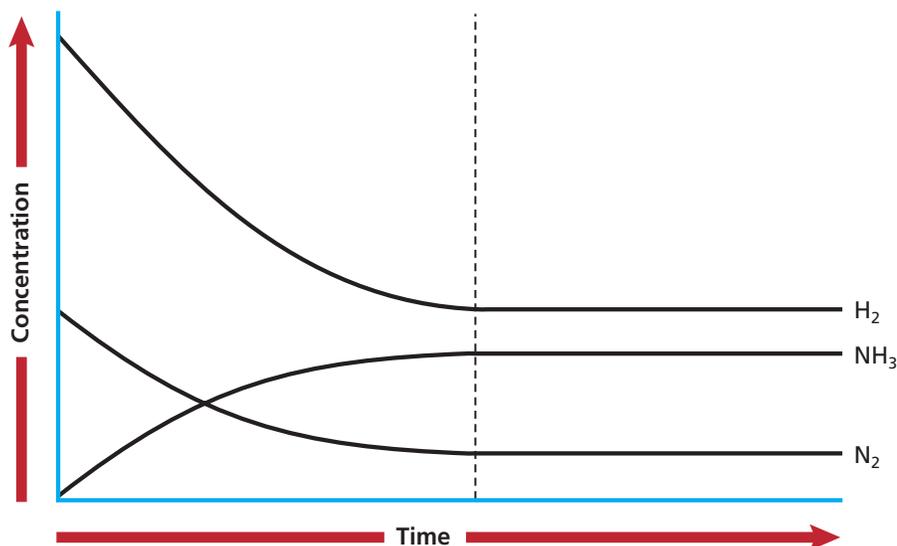


This reaction is important to agriculture because ammonia is used widely as a source of nitrogen for fertilizing corn and other farm crops. The photo on the opposite page shows ammonia being “knifed” into the soil.

Note that the equation for the production of ammonia has a negative standard free energy, ΔG° . Recall that a negative sign for ΔG° indicates that the

Figure 18-1

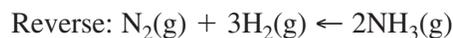
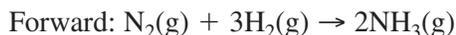
The concentrations of the reactants (H_2 and N_2) decrease at first while the concentration of the product (NH_3) increases. Then, before the reactants are used up, all concentrations become constant.



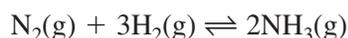
reaction is spontaneous under standard conditions. Standard conditions are defined as 298 K and one atmosphere pressure. But spontaneous reactions are not always fast. When carried out under standard conditions, this ammonia-forming reaction is much too slow. To produce ammonia at a rate that is practical, the reaction must be carried out at a much higher temperature than 298 K and a higher pressure than one atmosphere.

What happens when one mole of nitrogen and three moles of hydrogen, the amounts shown in the equation, are placed in a closed reaction vessel at 723 K? Because the reaction is spontaneous, nitrogen and hydrogen begin to react. **Figure 18-1** illustrates the progress of the reaction. Note that the concentration of the product, NH_3 , is zero at the start and gradually increases with time. The reactants, H_2 and N_2 , are consumed in the reaction, so their concentrations gradually decrease. After a period of time, however, the concentrations of H_2 , N_2 , and NH_3 no longer change. All concentrations become constant, as shown by the horizontal lines on the right side of the diagram. The concentrations of H_2 and N_2 are not zero, so not all of the reactants were converted to product even though ΔG° for this reaction is negative.

Reversible reactions When a reaction results in almost complete conversion of reactants to products, chemists say that the reaction goes to completion. But most reactions, including the ammonia-forming reaction, do not go to completion. They appear to stop. The reason is that these reactions are reversible. A **reversible reaction** is one that can occur in both the forward and the reverse directions.



Chemists combine these two equations into a single equation that uses a double arrow to show that both reactions occur.



When you read the equation, the reactants in the forward reaction are on the left. In the reverse reaction, the reactants are on the right. In the forward reaction, hydrogen and nitrogen combine to form the product ammonia. In the reverse



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

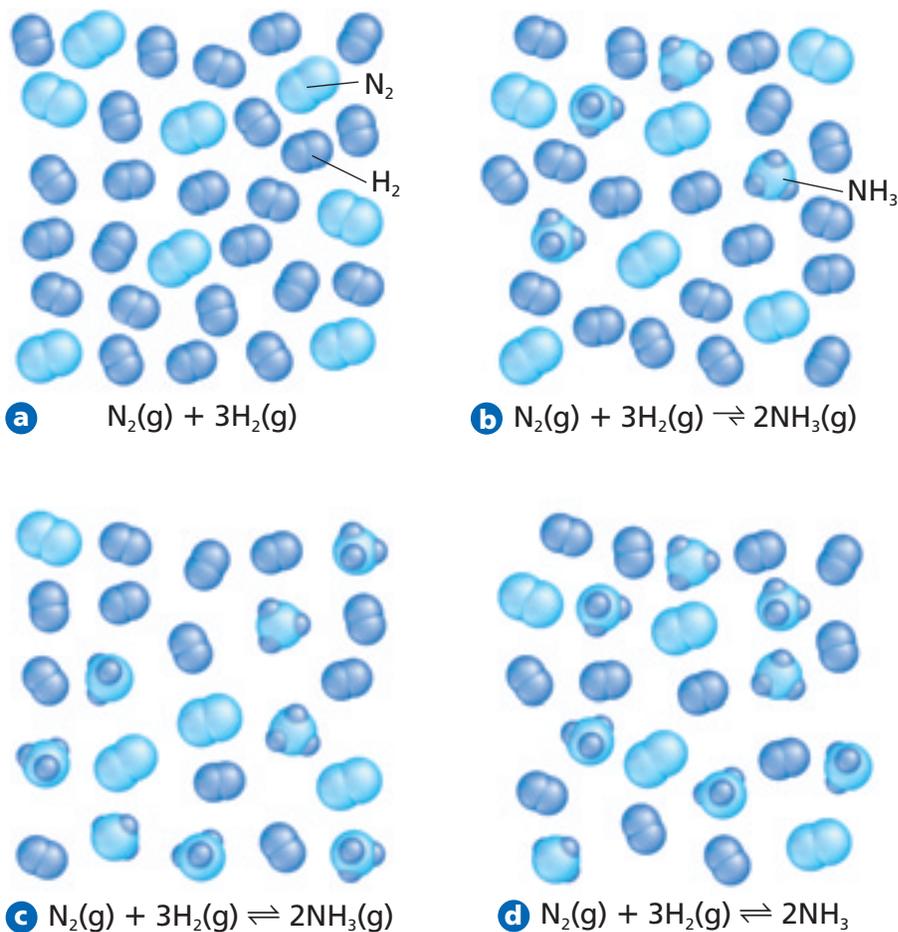
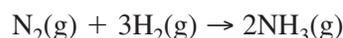


Figure 18-2

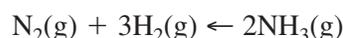
Only the forward reaction can occur in **a** because only the reactants are present. In **b**, the product, ammonia, is present and the reverse reaction begins. Compare **c** and **d**. Has equilibrium been established? Refer to **Table C-1** in Appendix C for a key to atom color correlations.

reaction, ammonia decomposes into the products hydrogen and nitrogen. How does the reversibility of this reaction affect the production of ammonia?

Figure 18-2a shows a mixture of nitrogen and hydrogen just as the reaction begins at a definite, initial rate. No ammonia is present so only the forward reaction can occur.



As hydrogen and nitrogen combine to form ammonia, their concentrations decrease, as shown in **Figure 18-2b**. Recall from Chapter 17 that the rate of a reaction depends upon the concentration of the reactants. The decrease in the concentration of the reactants causes the rate of the forward reaction to decrease. As soon as ammonia is present, the reverse reaction can occur, slowly at first, but at an increasing rate as the concentration of ammonia increases.



As the reaction proceeds, the rate of the forward reaction continues to decrease and the rate of the reverse reaction continues to increase until the two rates are equal. At that point, ammonia is being produced as fast as it is being decomposed, so the concentrations of nitrogen, hydrogen, and ammonia remain constant, as shown in **Figures 18-2c** and **18-2d**. The system has reached a state of balance or equilibrium. The word *equilibrium* means that opposing processes are in balance. **Chemical equilibrium** is a state in which the forward and reverse reactions balance each other because they take place at equal rates.

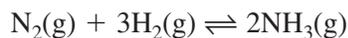
$$\text{Rate}_{\text{forward reaction}} = \text{Rate}_{\text{reverse reaction}}$$



Figure 18-3

If the only way in or out of San Francisco and Sausalito is across the Golden Gate Bridge, which joins the two cities, then the number of vehicles in the two cities will remain constant if the number of vehicles per hour crossing the bridge in one direction equals the number of vehicles crossing in the opposite direction. Will the same vehicles always be in the same city?

You can recognize that the ammonia-forming reaction reaches a state of chemical equilibrium because its chemical equation is written with a double arrow like this.



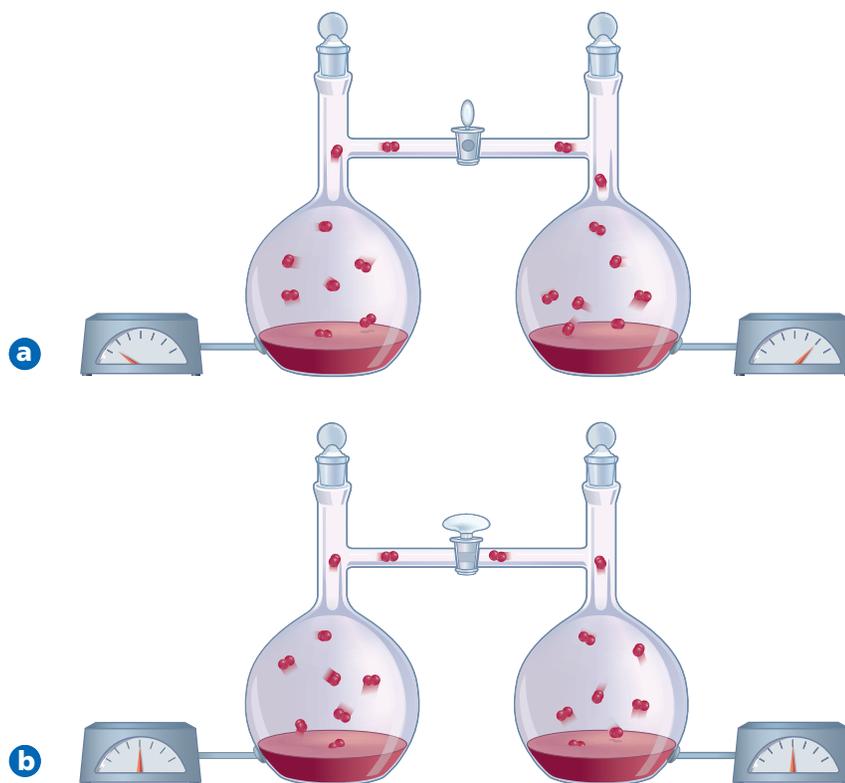
At equilibrium, the concentrations of reactants and products are constant, as you saw in **Figures 18-2c** and **18-2d**. However, that doesn't mean that the amounts or concentrations of reactants and products are equal. That is seldom the case. In fact, it's not unusual for the equilibrium concentrations of a reactant and product to differ by a factor of one million or more.

The dynamic nature of equilibrium Equilibrium is a state of action, not inaction. For example, consider this analogy. The Golden Gate Bridge, shown in **Figure 18-3**, connects two California cities, San Francisco and Sausalito. Suppose that all roads leading into and out of the two cities are closed for a day—except the Golden Gate Bridge. In addition, suppose that the number of vehicles per hour crossing the bridge in one direction equals the number of vehicles per hour traveling in the opposite direction. Given these circumstances, the number of vehicles in each of the two cities remains constant even though vehicles continue to cross the bridge. In this analogy, note that the total numbers of vehicles in the two cities do not have to be equal. Equilibrium requires only that the number of vehicles crossing the bridge in one direction is equal to the number crossing in the opposite direction.

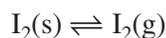
The dynamic nature of chemical equilibrium can be illustrated by placing equal masses of iodine crystals in two interconnected flasks, as shown in **Figure 18-4a**. The crystals in the flask on the left contain iodine molecules made up entirely of the nonradioactive isotope I-127. The crystals in the flask on the right contain iodine molecules made up of the radioactive isotope I-131. The Geiger counters indicate the radioactivity within each flask.

Figure 18-4

Radioactive iodine molecules from the solid in the flask on the right could not appear in the solid in the flask on the left unless iodine molecules were changing back and forth between the solid and gaseous phases.



Each flask is a closed system. No reactant or product can enter or leave. At 298 K and one atmosphere pressure, this equilibrium is established in both flasks.



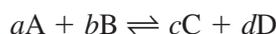
In the forward process, called sublimation, iodine molecules change directly from the solid phase to the gas phase. In the reverse process, gaseous iodine molecules return to the solid phase. A solid-vapor equilibrium is established in each flask.

When the stopcock in the tube connecting the two flasks is opened, iodine vapor can travel back and forth between the two flasks. After a period of time, the readings on the Geiger counters indicate that the flask on the left contains as many radioactive I-131 molecules as the flask on the right *in both the vapor and the solid phases*. See **Figure 18-4b**. How could radioactive I-131 molecules that were originally in the crystals in the flask on the right become part of the crystals in the flask on the left? The evidence suggests that iodine molecules constantly change from the solid phase to the gas phase according to the forward process, and that gaseous iodine molecules convert back to the solid phase according to the reverse process.

Equilibrium Expressions and Constants

You have learned that some chemical systems have little tendency to react and others go readily to completion. In between these two extremes are the majority of reactions that reach a state of equilibrium with varying amounts of reactants unconsumed. If the reactants are not consumed, then not all the product predicted by the balanced chemical equation will be produced. According to the equation for the ammonia-producing reaction, two moles of ammonia should be produced when one mole of nitrogen and three moles of hydrogen react. Because the reaction reaches a state of equilibrium, however, fewer than two moles of ammonia will actually be obtained. Chemists need to be able to predict the yield of a reaction.

In 1864, the Norwegian chemists Cato Maximilian Guldberg and Peter Waage proposed the **law of chemical equilibrium**, which states that at a given temperature, a chemical system may reach a state in which a particular ratio of reactant and product concentrations has a constant value. For example, the general equation for a reaction at equilibrium can be written as follows.



A and B are the reactants; C and D the products. The coefficients in the balanced equation are a , b , c , and d . If the law of chemical equilibrium is applied to this reaction, the following ratio is obtained.

$$K_{\text{eq}} = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$$

This ratio is called the equilibrium constant expression. The square brackets indicate the molar concentrations of the reactants and products at equilibrium in mol/L. The **equilibrium constant**, K_{eq} , is the numerical value of the ratio of product concentrations to reactant concentrations, with each concentration raised to the power corresponding to its coefficient in the balanced equation. The value of K_{eq} is constant only at a specified temperature.

How can you interpret the size of the equilibrium constant? Recall that ratios, or fractions, with large numerators are larger numbers than fractions

Physics

CONNECTION

In physics, a push or pull on an object is known as a force. Whenever someone pushes on a door, pulls on a dog's leash, or throws a ball, that person exerts a force. When two or more forces are exerted on the same object they add together. If the forces are exerted in opposite directions, one is essentially subtracted from the other. Thus, in a game of tug of war when two teams pull on a rope with equal force, the resulting force has a magnitude of zero. As a result, the object does not move and the system is said to be in equilibrium. Similarly, the acrobats below represent a system in equilibrium. Equal and opposite forces are called balanced forces. If, instead, one force is greater in magnitude than the other, the resulting force is greater than zero when the smaller force is subtracted from the greater one. The resulting force is called an unbalanced force, and an unbalanced force will cause an object to move.

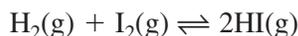


with large denominators. For example, compare the ratio 5/1 with 1/5. Five is a larger number than one-fifth. Because the product concentrations are in the numerator of the equilibrium expression, a numerically large K_{eq} means that the equilibrium mixture contains more products than reactants. Similarly, a numerically small K_{eq} means that the equilibrium mixture contains more reactants than products.

$K_{\text{eq}} > 1$: More products than reactants at equilibrium.

$K_{\text{eq}} < 1$: More reactants than products at equilibrium.

Constants for homogeneous equilibria How would you write the equilibrium constant expression for this reaction in which hydrogen and iodine react to form hydrogen iodide?



This reaction is a **homogeneous equilibrium**, which means that all the reactants and products are in the same physical state. All participants are gases. To begin writing the equilibrium constant expression, place the product concentration in the numerator and the reactant concentrations in the denominator.

$$\frac{[\text{HI}]}{[\text{H}_2][\text{I}_2]}$$

The expression becomes equal to K_{eq} when you add the coefficients from the balanced chemical equation as exponents.

$$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

K_{eq} for this homogeneous equilibrium at 731 K is 49.7. Note that 49.7 has no units. In writing equilibrium constant expressions, it's customary to omit units. Considering the size of K_{eq} , are there more products than reactants present at equilibrium?

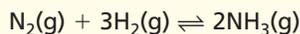
This plant for the manufacture of ammonia employs the Haber process, which maximizes the yield of ammonia by adjusting the temperature and pressure of the reaction.



EXAMPLE PROBLEM 18-1

Equilibrium Constant Expressions for Homogeneous Equilibria

Write the equilibrium constant expression for the reaction in which ammonia gas is produced from hydrogen and nitrogen.



1. Analyze the Problem

You have been given the equation for the reaction, which provides the information needed to write the equilibrium constant expression. The equilibrium is homogeneous because the reactants and product are in the same physical state. The form of the equilibrium constant expression is $K_{\text{eq}} = \frac{[\text{C}]^c}{[\text{A}]^a[\text{B}]^b}$.

Known

$$[\text{C}] = [\text{NH}_3]$$

$$\text{coefficient NH}_3 = 2$$

$$[\text{A}] = [\text{N}_2]$$

$$\text{coefficient N}_2 = 1$$

$$[\text{B}] = [\text{H}_2]$$

$$\text{coefficient H}_2 = 3$$

Unknown

$$K_{\text{eq}} = ?$$

2. Solve for the Unknown

Place the product concentration in the numerator and the reactant concentrations in the denominator.

$$\frac{[\text{NH}_3]}{[\text{N}_2][\text{H}_2]}$$

Raise the concentration of each reactant and product to a power equal to its coefficient in the balanced chemical equation and set the ratio equal to K_{eq} .

$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

3. Evaluate the Answer

The product concentration is in the numerator and the reactant concentrations are in the denominator. Product and reactant concentrations are raised to powers equal to their coefficients.

PRACTICE PROBLEMS

1. Write equilibrium constant expressions for these equilibria.

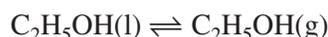
- $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$
- $\text{CO}(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g})$
- $2\text{H}_2\text{S}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{S}_2(\text{g})$



For more practice writing homogeneous equilibrium constant expressions, go to **Supplemental Practice Problems** in Appendix A.

Constants for heterogeneous equilibria You have learned to write K_{eq} expressions for homogeneous equilibria, those in which all reactants and products are in the same physical state. When the reactants and products of a reaction are present in more than one physical state, the equilibrium is called a **heterogeneous equilibrium**.

When ethanol is placed in a closed flask, a liquid-vapor equilibrium is established, as illustrated in **Figure 18-5**.



To write the equilibrium constant expression for this process, you would form a ratio of the product to the reactant. At a given temperature, the ratio would have a constant value K .

$$K = \frac{[\text{C}_2\text{H}_5\text{OH}(\text{g})]}{[\text{C}_2\text{H}_5\text{OH}(\text{l})]}$$

Note that the term in the denominator is the concentration of liquid ethanol. Because liquid ethanol is a pure substance, its concentration is constant at a given temperature. That's because the concentration of a pure substance is its density in moles per liter. At any given temperature, density does not change. No matter how much or how little $\text{C}_2\text{H}_5\text{OH}$ is present, its concentration remains constant. Therefore, the term in the denominator is a constant and can be combined with K .

$$K[\text{C}_2\text{H}_5\text{OH}(\text{l})] = [\text{C}_2\text{H}_5\text{OH}(\text{g})] = K_{\text{eq}}$$

The equilibrium constant expression for this phase change is

$$K_{\text{eq}} = [\text{C}_2\text{H}_5\text{OH}(\text{g})]$$

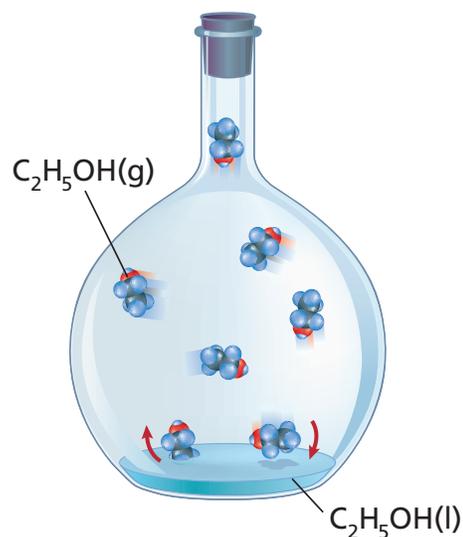
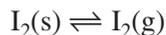


Figure 18-5

The rate of evaporation of ethanol equals the rate of condensation. This two-phase equilibrium is called a heterogeneous equilibrium. K_{eq} depends only on $[\text{C}_2\text{H}_5\text{OH}(\text{g})]$.

Solids also are pure substances with unchanging concentrations, so equilibria involving solids can be simplified in the same way. For example, recall the experiment involving the sublimation of iodine crystals in **Figure 18-4** on page 562.



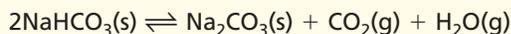
$$K_{\text{eq}} = [\text{I}_2(\text{g})]$$

The equilibrium depends only on the concentration of gaseous iodine in the system.

EXAMPLE PROBLEM 18-2

Equilibrium Constant Expressions for Heterogeneous Equilibria

Write the equilibrium constant expression for the decomposition of baking soda (sodium hydrogen carbonate).



1. Analyze the Problem

You are given a heterogeneous equilibrium involving gases and solids. The general form of the equilibrium constant expression for this reaction is

$$K_{\text{eq}} = \frac{[\text{C}]^c[\text{D}]^d[\text{E}]^e}{[\text{A}]^a[\text{B}]^b}$$

Because the reactant and one of the products are solids with constant concentrations, they can be omitted from the equilibrium constant expression.

Known

$$[\text{C}] = [\text{Na}_2\text{CO}_3]$$

$$\text{coefficient Na}_2\text{CO}_3 = 1$$

$$[\text{D}] = [\text{CO}_2]$$

$$\text{coefficient CO}_2 = 1$$

$$[\text{E}] = [\text{H}_2\text{O}]$$

$$\text{coefficient H}_2\text{O} = 1$$

$$[\text{A}] = [\text{NaHCO}_3]$$

$$\text{coefficient NaHCO}_3 = 2$$

Unknown

equilibrium constant expression = ?

2. Solve for the Unknown

Write a ratio with the concentrations of the products in the numerator and the concentration of the reactant in the denominator.

$$\frac{[\text{Na}_2\text{CO}_3][\text{CO}_2][\text{H}_2\text{O}]}{[\text{NaHCO}_3]^2}$$

Leave out $[\text{NaHCO}_3]$ and $[\text{Na}_2\text{CO}_3]$ because they are solids.

$$[\text{CO}_2][\text{H}_2\text{O}]$$

Because the coefficients of $[\text{CO}_2]$ and $[\text{H}_2\text{O}]$ are 1, the expression is complete.

$$K_{\text{eq}} = [\text{CO}_2][\text{H}_2\text{O}]$$

3. Evaluate the Answer

The expression correctly applies the law of chemical equilibrium to the equation.



Baking soda, or sodium hydrogen carbonate, makes this Irish soda bread light and airy. Baking soda is also useful as an antacid, a cleaner, and a deodorizer.

PRACTICE PROBLEMS

2. Write equilibrium constant expressions for these heterogeneous equilibria.

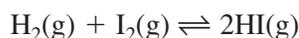
- $\text{C}_{10}\text{H}_8(\text{s}) \rightleftharpoons \text{C}_{10}\text{H}_8(\text{g})$
- $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
- $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{O}(\text{g})$
- $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{H}_2(\text{g})$
- $\text{FeO}(\text{s}) + \text{CO}(\text{g}) \rightleftharpoons \text{Fe}(\text{s}) + \text{CO}_2(\text{g})$



For more practice writing heterogeneous equilibrium constant expressions, go to **Supplemental Practice Problems** in Appendix A.

Determining the Value of Equilibrium Constants

For a given reaction at a given temperature, K_{eq} will always be the same regardless of the initial concentrations of reactants and products. To test the truth of this statement, three experiments were carried out to investigate this reaction.



The results are summarized in **Table 18-1**. In trial 1, 1.0000 mol H_2 and 2.0000 mol I_2 are placed in a 1.0000-L vessel. These initial concentrations have the symbols $[\text{H}_2]_0$ and $[\text{I}_2]_0$. No HI is present at the beginning of trial 1. In trial 2, only HI is present at the start of the experiment. In trial 3, each of the three participants has the same initial concentration.

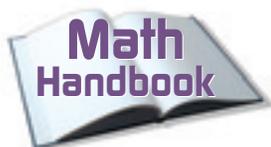
When equilibrium is established, the concentration of each substance is determined experimentally. In **Table 18-1**, the symbol $[\text{HI}]_{\text{eq}}$ represents the concentration of HI at equilibrium. Note that the equilibrium concentrations are not the same in the three trials, yet when each set of equilibrium concentrations is put into the equilibrium constant expression, the value of K_{eq} is the same. Each set of equilibrium concentrations represents an equilibrium position. Although an equilibrium system has only one value for K_{eq} at a particular temperature, it has an unlimited number of equilibrium positions. Equilibrium positions depend upon the initial concentrations of the reactants and products.

The large value of K_{eq} for the reaction $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ means that at equilibrium the product is present in larger amount than the reactants. Many equilibria, however, have small K_{eq} values. Do you remember what this means? K_{eq} for the equilibrium $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ equals 4.6×10^{-31} at 298 K. A K_{eq} this small means that the product, NO, is practically nonexistent at equilibrium.

Table 18-1

Three Experiments for an Equilibrium System							
$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ at 731 K							
Trial	$[\text{H}_2]_0$ (M)	$[\text{I}_2]_0$ (M)	$[\text{HI}]_0$ (M)	$[\text{H}_2]_{\text{eq}}$ (M)	$[\text{I}_2]_{\text{eq}}$ (M)	$[\text{HI}]_{\text{eq}}$ (M)	$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$
1	1.0000	2.0000	0.0	0.06587	1.0659	1.8682	$\frac{[1.8682]^2}{[0.06587][1.0659]} = 49.70$
2	0.0	0.0	5.0000	0.5525	0.5525	3.8950	$\frac{[3.8950]^2}{[0.5525][0.5525]} = 49.70$
3	1.0000	1.0000	1.0000	0.2485	0.2485	1.7515	$\frac{[1.7515]^2}{[0.2485][0.2485]} = 49.70$





Review solving algebraic equations in the **Math Handbook** on page 897 of this text.

EXAMPLE PROBLEM 18-3

Calculating the Value of Equilibrium Constants

Calculate the value of K_{eq} for the equilibrium constant expression

$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \text{ given concentration data at one equilibrium position:}$$
$$[\text{NH}_3] = 0.933 \text{ mol/L, } [\text{N}_2] = 0.533 \text{ mol/L, } [\text{H}_2] = 1.600 \text{ mol/L.}$$

1. Analyze the Problem

You have been given the equilibrium constant expression and the concentration of each reactant and product. You must calculate the equilibrium constant. Because the reactant, H_2 , has the largest concentration and is raised to the third power in the denominator, K_{eq} is likely to be less than 1.

Known

$$K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

$$[\text{NH}_3] = 0.933 \text{ mol/L}$$

$$[\text{N}_2] = 0.533 \text{ mol/L}$$

$$[\text{H}_2] = 1.600 \text{ mol/L}$$

Unknown

$$K_{\text{eq}} = ?$$

2. Solve for the Unknown

Substitute the known values into the equilibrium constant expression and calculate its value.

$$K_{\text{eq}} = \frac{[0.933]^2}{[0.533][1.600]^3} = 0.399$$

3. Evaluate the Answer

The smallest number of significant figures in the given data is three. Therefore, the answer is correctly stated with three digits. The calculation is correct and, as predicted, the value of K_{eq} is less than 1.

PRACTICE PROBLEMS

- Calculate K_{eq} for the equilibrium in Practice Problem 1a on page 565 using the data $[\text{N}_2\text{O}_4] = 0.0185 \text{ mol/L}$ and $[\text{NO}_2] = 0.0627 \text{ mol/L}$.
- Calculate K_{eq} for the equilibrium in Practice Problem 1b on page 565 using the data $[\text{CO}] = 0.0613 \text{ mol/L}$, $[\text{H}_2] = 0.1839 \text{ mol/L}$, $[\text{CH}_4] = 0.0387 \text{ mol/L}$, and $[\text{H}_2\text{O}] = 0.0387 \text{ mol/L}$.



For more practice evaluating quotients, go to **Supplemental Practice Problems** in Appendix A.

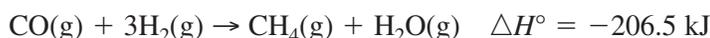
Section 18.1 Assessment

- How does the concept of reversibility explain the establishment of equilibrium?
- What characteristics define a system at equilibrium?
- When you write an equilibrium constant expression, how do you decide what goes in the numerator and in the denominator?
- Thinking Critically** Determine whether the following statement is correct: When chemical equilibrium exists, the reactant and product concentrations remain constant and the forward and reverse reactions cease. Explain your answer.
- Using Numbers** Determine the value of K_{eq} at 400 K for the decomposition of phosphorus pentachloride if $[\text{PCl}_5] = 0.135 \text{ mol/L}$, $[\text{PCl}_3] = 0.550 \text{ mol/L}$, and $[\text{Cl}_2] = 0.550 \text{ mol/L}$. The equation for the reaction is: $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$

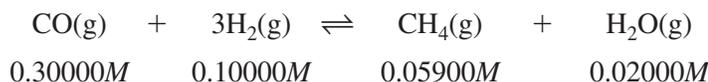
Manufacturers are aware that it makes good sense to minimize waste by finding ways of using leftover materials. Manufacturing new products from byproducts adds to profits and eliminates the problem of disposing of waste without causing environmental damage. What part does equilibrium play in promoting cost-cutting efficiency?

Le Châtelier's Principle

Suppose that the byproducts of an industrial process are the gases carbon monoxide and hydrogen and a company chemist believes these gases can be combined to produce the fuel methane (CH_4) using this reaction.



When the industrial chemist places CO and H_2 in a closed reaction vessel at 1200 K, the reaction establishes this equilibrium position (equilibrium position 1).



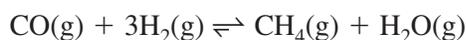
Inserting these concentrations into the equilibrium expression gives an equilibrium constant equal to 3.933.

$$K_{\text{eq}} = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{CO}][\text{H}_2]^3} = \frac{(0.05900)(0.02000)}{(0.30000)(0.10000)^3} = 3.933$$

Unfortunately, a methane concentration of 0.05900 mol/L in the equilibrium mixture is too low to be of any practical use. Can the chemist change the equilibrium position and thereby increase the amount of methane?

In 1888, the French chemist Henri-Louis Le Châtelier discovered that there are ways to control equilibria to make reactions, including this one, more productive. He proposed what is now called **Le Châtelier's principle**: If a stress is applied to a system at equilibrium, the system shifts in the direction that relieves the stress. A stress is any kind of change in a system at equilibrium that upsets the equilibrium. You can use Le Châtelier's principle to predict how changes in concentration, volume (pressure), and temperature affect equilibrium. Changes in volume and pressure are interrelated because decreasing the volume of a reaction vessel at constant temperature increases the pressure inside. Conversely, increasing the volume decreases the pressure.

Changes in concentration What happens if the industrial chemist injects additional carbon monoxide into the reaction vessel, raising the concentration of carbon monoxide from 0.30000M to 1.00000M? The higher carbon monoxide concentration immediately increases the number of effective collisions between CO and H_2 molecules and unbalances the equilibrium. The rate of the forward reaction increases, as indicated by the longer arrow to the right.



Objectives

- **Describe** how various factors affect chemical equilibrium.
- **Explain** how Le Châtelier's principle applies to equilibrium systems.

Vocabulary

Le Châtelier's principle

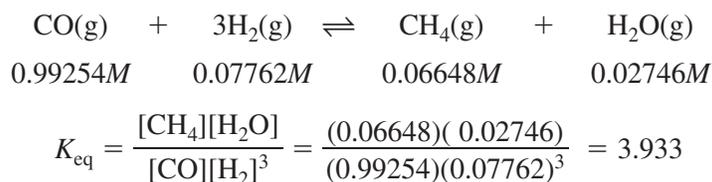
Careers Using Chemistry

Nurse Anesthetist

Would you like to work in a hospital operating room, helping patients remain pain-free during surgery? If so, consider a career as a nurse anesthetist.

In two-thirds of rural hospitals, nurse anesthetists provide all anesthesia services. In city hospitals, they work under anesthesiologists. These nurses also work in outpatient clinics and physicians' offices. As they regulate the anesthetic, they monitor body functions to make sure they stay in balance and the patient stays out of danger.

In time, the rate of the forward reaction slows down as the concentrations of CO and H₂ decrease. Simultaneously, the rate of the reverse reaction increases as more and more CH₄ and H₂O molecules are produced. Eventually, a new equilibrium position (position 2) is established with these concentrations.



Note that although K_{eq} has not changed, the new equilibrium position results in the desired effect—an increased concentration of methane. The results of this experiment are summarized in **Table 18-2**.

Table 18-2

Two Equilibrium Positions for the Equilibrium					
$\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$					
Equilibrium position	$[\text{CO}]_{\text{eq}}$ <i>M</i>	$[\text{H}_2]_{\text{eq}}$ <i>M</i>	$[\text{CH}_4]_{\text{eq}}$ <i>M</i>	$[\text{H}_2\text{O}]_{\text{eq}}$ <i>M</i>	K_{eq} <i>M</i>
1	0.30000	0.10000	0.05900	0.02000	3.933
2	0.99254	0.07762	0.06648	0.02746	3.933

Could you have predicted this result using Le Châtelier's principle? Yes. Think of the increased concentration of CO as a stress on the equilibrium. The equilibrium system reacts to the stress by consuming CO at an increased rate. This response, called a shift to the right, forms more CH₄ and H₂O. Any increase in the concentration of a reactant results in a shift to the right and additional product.

Suppose that rather than injecting more reactant, the chemist decides to remove a product (H₂O) by adding a desiccant to the reaction vessel. Recall from Chapter 11 that a desiccant is a substance that absorbs water. What does Le Châtelier's principle predict the equilibrium will do in response to a decrease in the concentration of water (the stress)? You are correct if you said that the equilibrium shifts in the direction that will tend to bring the concentration of water back up. That is, the equilibrium shifts to the right and results in additional product.

It might be helpful to think about the produce vendor in **Figure 18-6** who wants to keep a display of vegetables looking neat and tempting. The vendor must constantly add new vegetables to fill the empty spaces created when customers buy the vegetables. Similarly, the equilibrium reaction attempts to restore the lost water by producing more water. As a result, more water and more methane are produced. In any equilibrium, the removal of a product results in a shift to the right and the production of more product.

The equilibrium position also can be shifted to the left, toward the reactants. Do you have an idea how? Le Châtelier's principle predicts that if additional product is added to a reaction at equilibrium, the reaction will shift to the left. The stress is relieved by converting products to reactants. If one of the reactants is removed, a similar shift to the left will occur.

When predicting the results of a stress on an equilibrium using Le Châtelier's principle, it is important to have the equation for the reaction in

Figure 18-6

A neat arrangement of produce can be thought of as being at equilibrium—nothing is changing. But when a customer buys some of the vegetables, the equilibrium is disturbed. The produce vendor then restores equilibrium by filling the empty spots.



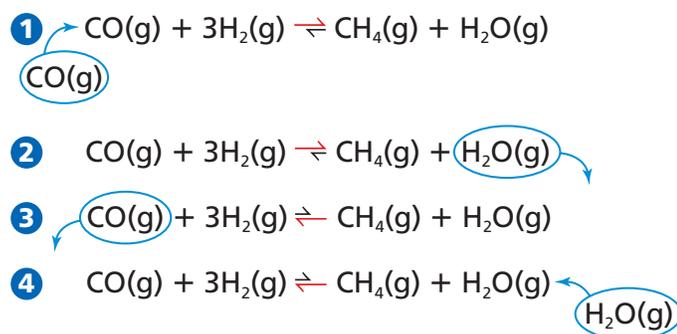
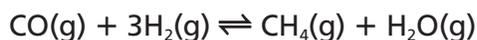
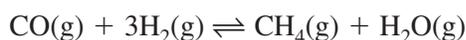


Figure 18-7

The addition or removal of a reactant or product shifts the equilibrium in the direction that relieves the stress. Note the unequal arrows, which indicate the direction of the shift. How would the reaction shift if you added H₂? Removed CH₄?

view. The effects of changing concentration are summarized in **Figure 18-7**. Refer to this figure as you practice applying Le Châtelier's principle.

Changes in volume Consider again the reaction for making methane from byproduct gases.



Can this reaction be forced to produce more methane by changing the volume of the reaction vessel? Suppose the vessel's volume can be changed using a piston-like device similar to the one shown in **Figure 18-8**. If the piston is forced downward, the volume of the system decreases. You have learned that Boyle's law says that decreasing the volume at constant temperature increases the pressure. The increased pressure is a stress on the reaction at equilibrium. How does the equilibrium respond to the disturbance and relieve the stress?

Recall that the pressure exerted by an ideal gas depends upon the number of gas particles that collide with the walls of the vessel. The more gas particles contained in the vessel, the greater the pressure. If the number of gas particles is increased at constant temperature, the pressure of the gas increases. Similarly, if the number of gas particles is decreased, the pressure also decreases. How does this relationship between numbers of gas particles and pressure apply to the reaction for making methane? Compare the number of moles of gaseous reactants in the equation with the number of moles of gaseous products. For every two moles of gaseous products (1 mol CH₄ and 1 mol H₂O), four moles of gaseous reactants are consumed (1 mol CO and 3 mol H₂), a net decrease of two moles. If you apply Le Châtelier's principle, you can see that the equilibrium can relieve the stress of increased pressure by shifting to the right. This shift decreases the total number of moles of gas and thus, the pressure inside the reaction vessel decreases. Although the shift to the right does not reduce the pressure to its original value, it has the desired effect—the equilibrium produces more methane. See **Figure 18-8**.

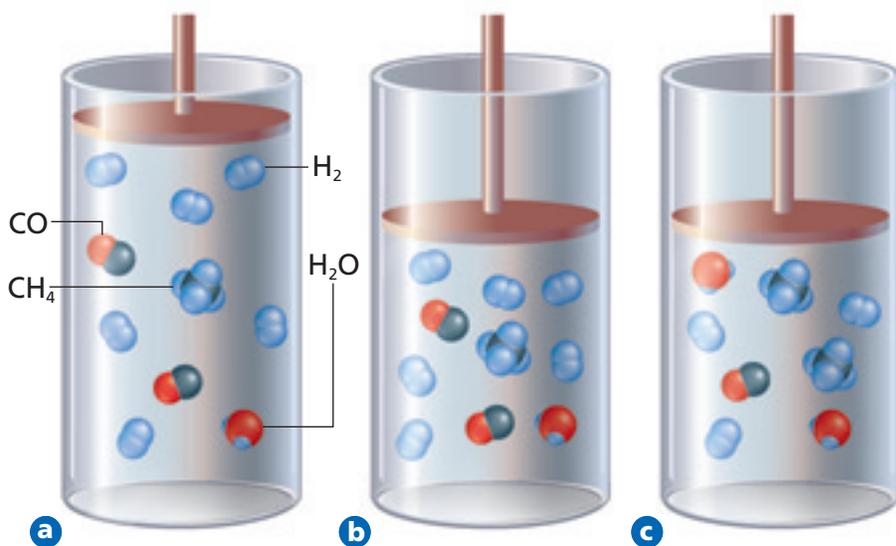
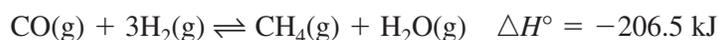


Figure 18-8

In **a**, the reaction between CO and H₂ is at equilibrium. In **b**, the piston has been lowered decreasing the volume and increasing the pressure. The outcome is seen in **c**. More molecules of the products have formed. Their formation helped relieve the stress on the system. How do the numbers of particles in **a** and **c** compare?

Changing the volume (and pressure) of an equilibrium system shifts the equilibrium only if the number of moles of gaseous reactants is different from the number of moles of gaseous products. Changing the volume of a reaction vessel containing the equilibrium $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ would cause no shift in the equilibrium because the number of moles of gas is the same on both sides of the equation.

Changes in temperature You have now learned that changes in concentration and volume change the equilibrium position by causing shifts to the right or left, but they do not change the equilibrium constant. A change in temperature, however, alters both the equilibrium position and the equilibrium constant. To understand how a change in temperature affects an equilibrium, recall that virtually every chemical reaction is either endothermic or exothermic. For example, the reaction for making methane has a negative ΔH° , which means that the forward reaction is exothermic and the reverse reaction is endothermic.



Heat can be thought of as a product in the forward reaction and a reactant in the reverse reaction. You can see this by reading the equation forward and backward.



How could an industrial chemist regulate the temperature to increase the amount of methane in the equilibrium mixture? According to Le Châtelier's principle, if heat is added, the reaction shifts in the direction in which heat is used up; that is, the reaction shifts to the left. A shift to the left means a decrease in the concentration of methane because methane is a reactant in the reverse reaction. However, lowering the temperature shifts the equilibrium to the right because the forward reaction liberates heat and relieves the stress. In shifting to the right, the equilibrium produces more methane.

Any change in temperature results in a change in K_{eq} . Recall that the larger the value of K_{eq} , the more product is found in the equilibrium mixture. Thus, for the methane-producing reaction, K_{eq} increases in value when the temperature is lowered and decreases when the temperature is raised.

Figure 18-9 shows how another equilibrium responds to changes in temperature. The endothermic equilibrium is described by this equation.

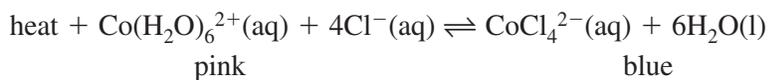


Figure 18-9

At room temperature, the solution in **a** is an equilibrium mixture of pink reactants and blue products. When the temperature is lowered in **b**, this endothermic reaction shifts toward the pink reactants. In **c**, the stress of higher temperature causes the reaction to shift toward the blue products.



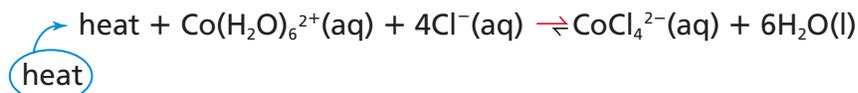
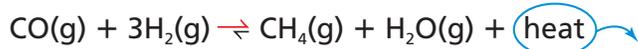
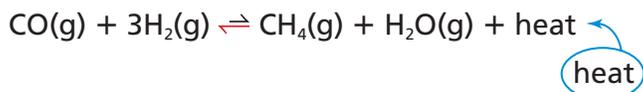


Figure 18-10

For the exothermic reaction between CO and H₂, raising the temperature shifts the equilibrium to the left (top equation). Lowering the temperature results in a shift to the right (second equation). The opposite is true for the endothermic reaction involving cobalt and chloride ions (third and last equations).

At room temperature, this mixture of aqueous ions appears violet because it contains significant amounts of pink $\text{Co(H}_2\text{O)}_6^{2+}\text{(aq)}$ and blue $\text{CoCl}_4^{2-}\text{(aq)}$. When cooled in an ice bath, the equilibrium mixture turns pink. The removal of heat is a stress on the equilibrium. The stress is relieved by an equilibrium shift to the left, generating more heat and producing more pink $\text{Co(H}_2\text{O)}_6^{2+}$ ions. When heat is added, the equilibrium mixture appears blue because the equilibrium shifts to the right to absorb the additional heat. As a result, more blue CoCl_4^{2-} ions are created. You can investigate this equilibrium system further in the **miniLAB** below. The diagram in **Figure 18-10** shows the effect of heating and cooling on exothermic and endothermic reactions.

Changes in concentration, volume, and temperature make a difference in the amount of product formed in a reaction. Can a catalyst also affect product concentration? A catalyst speeds up a reaction, but it does so equally in both directions. Therefore, a catalyzed reaction reaches equilibrium more quickly, but with no change in the amount of product formed. Read the **Chemistry and Technology** feature at the end of this chapter to learn how Le Châtelier's principle is applied to an important industrial process.

miniLAB

Shifts in Equilibrium

Observing and Inferring Le Châtelier's principle states that if a stress is placed on a reaction at equilibrium, the system will shift in a way that will relieve the stress. In this experiment, you will witness an equilibrium shift in a colorful way.

Materials test tubes (2); 10-mL graduated cylinder; 250-mL beaker; concentrated hydrochloric acid; 0.1M CoCl_2 solution; ice bath; table salt; hot plate

Procedures

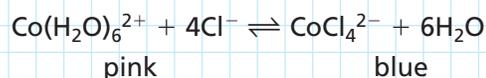


- Place about 2 mL of 0.1M CoCl_2 solution in a test tube. Record the color of the solution.
- Add about 3 mL of concentrated HCl to the test tube. Record the color of the solution.
CAUTION: HCl can burn skin and clothing.
- Add enough water to the test tube to make a color change occur. Record the color.

- Add about 2 mL of 0.1M CoCl_2 to another test tube. Add concentrated HCl dropwise until the solution turns purple. If the solution becomes blue, add water until it turns purple.
- Place the test tube in an ice bath that has had some salt sprinkled into the ice water. Record the color of the solution in the test tube.
- Place the test tube in a hot water bath that is at least 70°C. Record the color of the solution.

Analysis

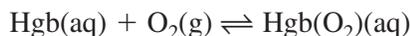
- The equation for the reversible reaction in this experiment is



Use the equation to explain your observations of color in steps 1–3.

- Explain how the equilibrium shifts when energy is added or removed.

A biological equilibrium If you have ever traveled to the mountains for a strenuous activity such as skiing, hiking, or mountain climbing, it is likely that you have felt tired and lightheaded for a time. This feeling is a result of the fact that at high altitudes, the air is thinner and contains fewer oxygen molecules. An important equilibrium in your body is disturbed. That equilibrium is represented by the following equation.



Hgb and Hgb(O₂) are greatly simplified formulas for hemoglobin and oxygenated hemoglobin, respectively. Hemoglobin is the blood protein that transports oxygen from your lungs to your muscles and other tissues where it is used in the metabolic processes that produce energy. What happens to this equilibrium at an altitude where both the atmospheric pressure and oxygen concentration are lower than normal for your body? Applying Le Châtelier's principle, the equilibrium shifts to the left to produce more oxygen. This shift reduces the amount of Hgb(O₂) in your blood and, therefore, the supply of oxygen to your muscles and other tissues.

After spending some time in the mountains, you probably noticed that your fatigue lessened. That's because your body adapted to the reduced oxygen concentration by producing more Hgb, which shifts the equilibrium back to the right and increases the amount of Hgb(O₂) in your blood. The Sherpas in **Figure 18-11**, who live and work in the mountains, do not experience discomfort because the equilibrium systems in their blood are adapted to high-altitude conditions.

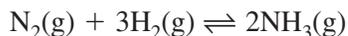
Figure 18-11

The equilibrium in the blood of Sherpas is adjusted to the lower level of oxygen in the air at high altitudes. Do you think Sherpas would experience a period of adjustment if they moved to sea level?



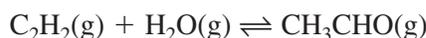
Section 18.2 Assessment

- 10.** How does a system at equilibrium respond to a stress? What factors are considered to be stresses on an equilibrium system?
- 11.** Use Le Châtelier's principle to predict how each of these changes would affect the ammonia equilibrium system.



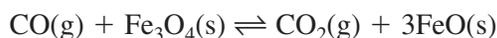
- removing hydrogen from the system
 - adding ammonia to the system
 - adding hydrogen to the system
- 12.** How would decreasing the volume of the reaction vessel affect each of these equilibria?
- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$
 - $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{HCl}(\text{g})$
 - $2\text{NOBr}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Br}_2(\text{g})$

- 13.** In the following equilibrium, would you raise or lower the temperature to obtain these results?



$$\Delta H^\circ = -151 \text{ kJ}$$

- an increase in the amount of CH₃CHO
 - a decrease in the amount of C₂H₂
 - an increase in the amount of H₂O
- 14. Thinking Critically** Why does changing the volume of the reaction vessel have no effect on this equilibrium?



- 15. Predicting** Predict how this equilibrium would respond to a simultaneous increase in both temperature and pressure.

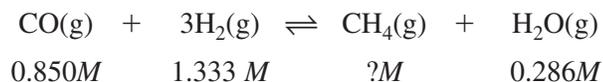


When a reaction has a large K_{eq} , the products are favored at equilibrium. That means that the equilibrium mixture contains more products than reactants. Conversely, when a reaction has a small K_{eq} , the reactants are favored at equilibrium, which means that the equilibrium mixture contains more reactants than products. Knowing the size of the equilibrium constant can help a chemist decide whether a reaction is practical for making a particular product.

Calculating Equilibrium Concentrations

The equilibrium constant expression can be useful in another way. Knowing the equilibrium constant expression, a chemist can calculate the equilibrium concentration of any substance involved in a reaction if the concentrations of all other reactants and products are known.

Suppose the industrial chemist that you read about earlier knows that at 1200 K, K_{eq} equals 3.933 for the reaction that forms methane from H_2 and CO. How much methane would actually be produced? If the concentrations of H_2 , CO, and H_2O are known, the concentration of CH_4 can be calculated.



The first thing the chemist would do is write the equilibrium constant expression.

$$K_{\text{eq}} = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{CO}][\text{H}_2]^3}$$

The equation can be solved for the unknown $[\text{CH}_4]$ by multiplying both sides of the equation by $[\text{CO}][\text{H}_2]^3$ and dividing both sides by $[\text{H}_2\text{O}]$.

$$[\text{CH}_4] = K_{\text{eq}} \times \frac{[\text{CO}][\text{H}_2]^3}{[\text{H}_2\text{O}]}$$

All the known concentrations and the value of K_{eq} (3.933) can now be substituted into the equilibrium constant expression.

$$[\text{CH}_4] = 3.933 \times \frac{(0.850)(1.333)^3}{(0.286)} = 27.7 \text{ mol/L}$$

The equilibrium concentration of CH_4 is 27.7 mol/L.

At this point an industrial chemist would evaluate whether an equilibrium concentration of 27.7 mol/L was sufficient to make the conversion of waste CO and H_2 to methane practical. Methane is becoming the fuel of choice for heating homes and cooking food. It is also the raw material for the manufacture of many products including acetylene and formic acid. Increasingly, it is being used as the energy source of fuel cells. At present, methane is relatively inexpensive but as demand grows, the cost will increase. Can the cost of a process which produces 27.7 mol/L CH_4 compete with the cost of obtaining methane from underground deposits? This is an important question for the manufacturer. **Figure 18-12** shows a surprising source of methane in the atmosphere.

Objectives

- **Determine** equilibrium concentrations of reactants and products.
- **Calculate** the solubility of a compound from its solubility product constant.
- **Explain** the common ion effect.

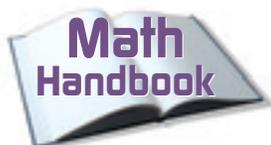
Vocabulary

solubility product constant
common ion
common ion effect

Figure 18-12

As termites relentlessly digest cellulose (wood) throughout the world, they produce methane gas, which enters the atmosphere. Although now present in comparatively small amounts, methane is counted as one of the greenhouse gases.



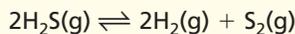


Review solving algebraic equations in the **Math Handbook** on page 897 of this text.

EXAMPLE PROBLEM 18-4

Calculating Equilibrium Concentrations

At 1405 K, hydrogen sulfide, also called rotten egg gas because of its bad odor, decomposes to form hydrogen and a diatomic sulfur molecule, S_2 . The equilibrium constant for the reaction is 2.27×10^{-3} .



What is the concentration of hydrogen gas if $[S_2] = 0.0540 \text{ mol/L}$ and $[H_2S] = 0.184 \text{ mol/L}$?

1. Analyze the Problem

You have been given K_{eq} and two of the three variables in the equilibrium constant expression. The equilibrium expression can be solved for $[H_2]$. K_{eq} is less than one, so more reactants than products are in the equilibrium mixture. Thus, you can predict that $[H_2]$ will be less than 0.184 mol/L , the concentration of the reactant H_2S .

Known

$$K_{eq} = 2.27 \times 10^{-3}$$
$$[S_2] = 0.0540 \text{ mol/L}$$
$$[H_2S] = 0.184 \text{ mol/L}$$

Unknown

$$[H_2] = ? \text{ mol/L}$$

2. Solve for the Unknown

Write the equilibrium constant expression.

$$\frac{[H_2]^2[S_2]}{[H_2S]^2} = K_{eq}$$

Solve the equation for $[H_2]^2$ by dividing both sides of the equation by $[S_2]$ and multiplying both sides by $[H_2S]^2$.

$$[H_2]^2 = K_{eq} \times \frac{[H_2S]^2}{[S_2]}$$

Substitute the known quantities into the expression and solve for $[H_2]$.

$$[H_2]^2 = 2.27 \times 10^{-3} \times \frac{(0.184)^2}{(0.0540)} = 1.42 \times 10^{-3}$$

$$[H_2] = \sqrt{1.42 \times 10^{-3}} = 0.0377 \text{ mol/L}$$

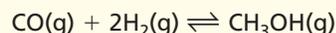
The equilibrium concentration of H_2 is 0.0377 mol/L .

3. Evaluate the Answer

All of the data for the problem have three significant figures, so the answer is correctly stated with three digits. As predicted, the equilibrium concentration of H_2 is less than 0.184 mol/L .

PRACTICE PROBLEMS

16. At a certain temperature, $K_{eq} = 10.5$ for the equilibrium



Calculate these concentrations:

- $[CO]$ in an equilibrium mixture containing $0.933 \text{ mol/L } H_2$ and $1.32 \text{ mol/L } CH_3OH$
- $[H_2]$ in an equilibrium mixture containing $1.09 \text{ mol/L } CO$ and $0.325 \text{ mol/L } CH_3OH$
- $[CH_3OH]$ in an equilibrium mixture containing $0.0661 \text{ mol/L } H_2$ and $3.85 \text{ mol/L } CO$.



For more practice calculating concentrations from the equilibrium constant expression, go to **Supplemental Practice Problems** in Appendix A.

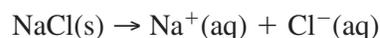


Figure 18-13

The briny taste of a pickle is a great accompaniment for a deli sandwich. But the main purpose of the salt is to preserve the pickle. Can you think of other foods that are preserved by salting?

Solubility Equilibria

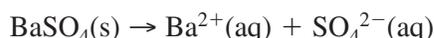
Some ionic compounds dissolve readily in water and some barely dissolve at all. Sodium chloride, or table salt, is typical of the soluble ionic compounds. On dissolving, all ionic compounds dissociate into ions.



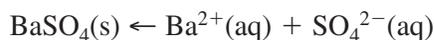
Approximately 36 g NaCl dissolves in 100 mL of water at 273 K. Without this high solubility, sodium chloride couldn't flavor and preserve foods like the pickles shown in **Figure 18-13**. Sodium chloride's vital role as an electrolyte in blood chemistry also depends upon its high solubility.

Although high solubility in water is often beneficial, low solubility also is important in many applications. For example, although barium ions are toxic to humans, patients are required to ingest barium sulfate prior to having an X ray of the digestive tract taken. X rays taken without barium sulfate in the digestive system are not well defined. Why can patients safely ingest barium sulfate?

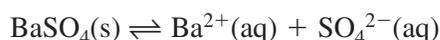
In water solution, barium sulfate dissociates according to this equation.



As soon as the first product ions form, the reverse reaction begins to re-form the reactants according to this equation.



With time, the rate of the reverse reaction becomes equal to the rate of the forward reaction and equilibrium is established.



For sparingly soluble compounds, such as BaSO₄, the rates become equal when the concentrations of the aqueous ions are exceedingly small. Nevertheless, the solution at equilibrium is a saturated solution.

The equilibrium constant expression for the dissolving of BaSO₄ is

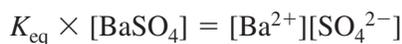
$$K_{\text{eq}} = \frac{[\text{Ba}^{2+}][\text{SO}_4^{2-}]}{[\text{BaSO}_4]}$$

In the equilibrium expression, [BaSO₄] is constant because barium sulfate is a solid. This constant value is combined with K_{eq} by multiplying both sides of the equation by [BaSO₄].



Figure 18-14

The presence of barium ions in the gastrointestinal system made the sharp definition of this X ray possible.

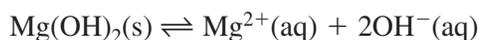


The product of K_{eq} and the concentration of the undissolved solid creates a new constant called the solubility product constant, K_{sp} . The **solubility product constant** is an equilibrium constant for the dissolving of a sparingly soluble ionic compound in water. The solubility product constant expression is

$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10} \text{ at } 298 \text{ K}$$

The solubility product constant expression is the product of the concentrations of the ions each raised to the power equal to the coefficient of the ion in the chemical equation. The small value of K_{sp} indicates that products are not favored at equilibrium. Thus, few barium ions are present at equilibrium ($1.0 \times 10^{-5} M$) and a patient can safely ingest a barium sulfate solution to obtain a clear X ray like the one shown in **Figure 18-14**.

Here is another example.



$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^{-}]^2$$

K_{sp} depends only on the concentrations of the ions in the saturated solution. However, to establish an equilibrium system, some undissolved solid, no matter how small the amount, must be present in the equilibrium mixture.

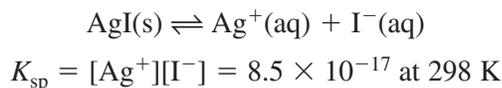
The solubility product constants for some ionic compounds are listed in **Table 18-3**. Note that they are all small numbers. Solubility product constants are measured and recorded only for sparingly soluble compounds.

Using solubility product constants The solubility product constants in **Table 18-3** have been determined through careful experiments. K_{sp} values are important because they can be used to determine the solubility of a sparingly soluble compound. Recall that the solubility of a compound in water is the amount of the substance that will dissolve in a given volume of water.

Table 18-3

Solubility Product Constants at 298 K					
Compound	K_{sp}	Compound	K_{sp}	Compound	K_{sp}
Carbonates		Halides		Hydroxides	
BaCO ₃	2.6×10^{-9}	CaF ₂	3.5×10^{-11}	Al(OH) ₃	4.6×10^{-33}
CaCO ₃	3.4×10^{-9}	PbBr ₂	6.6×10^{-6}	Ca(OH) ₂	5.0×10^{-6}
CuCO ₃	2.5×10^{-10}	PbCl ₂	1.7×10^{-5}	Cu(OH) ₂	2.2×10^{-20}
PbCO ₃	7.4×10^{-14}	PbF ₂	3.3×10^{-8}	Fe(OH) ₂	4.9×10^{-17}
MgCO ₃	6.8×10^{-6}	PbI ₂	9.8×10^{-9}	Fe(OH) ₃	2.8×10^{-39}
Ag ₂ CO ₃	8.5×10^{-12}	AgCl	1.8×10^{-10}	Mg(OH) ₂	5.6×10^{-12}
ZnCO ₃	1.5×10^{-10}	AgBr	5.4×10^{-13}	Zn(OH) ₂	3×10^{-17}
Hg ₂ CO ₃	3.6×10^{-17}	AgI	8.5×10^{-17}	Sulfates	
Chromates		Phosphates		BaSO ₄	1.1×10^{-10}
BaCrO ₄	1.2×10^{-10}	AlPO ₄	9.8×10^{-21}	CaSO ₄	4.9×10^{-5}
PbCrO ₄	2.3×10^{-13}	Ca ₃ (PO ₄) ₂	2.1×10^{-33}	PbSO ₄	2.5×10^{-8}
Ag ₂ CrO ₄	1.1×10^{-12}	Mg ₃ (PO ₄) ₂	1.0×10^{-24}	Ag ₂ SO ₄	1.2×10^{-5}

Suppose you wish to determine the solubility of silver iodide (AgI) in mol/L at 298 K. The equilibrium equation and solubility product constant expression are



The first thing you should do is let the symbol s represent the solubility of AgI; that is, the number of moles of AgI that dissolves in a liter of solution. The equation indicates that for every mole of AgI that dissolves, an equal number of moles of Ag^+ ions forms in solution. Therefore, $[\text{Ag}^+]$ equals s . Every Ag^+ has an accompanying I^- ion, so $[\text{I}^-]$ also equals s . Substituting s for $[\text{Ag}^+]$ and $[\text{I}^-]$, the K_{sp} expression becomes

$$[\text{Ag}^+][\text{I}^-] = (s)(s) = s^2 = 8.5 \times 10^{-17}$$

$$s = \sqrt{8.5 \times 10^{-17}} = 9.2 \times 10^{-9} \text{ mol/L.}$$

The solubility of AgI is 9.2×10^{-9} mol/L at 298 K.

EXAMPLE PROBLEM 18-5

Calculating Molar Solubility from K_{sp}

Use the K_{sp} value from Table 18-3 to calculate the solubility in mol/L of copper(II) carbonate (CuCO_3) at 298 K.

1. Analyze the Problem

You have been given the solubility product constant for CuCO_3 . The copper and carbonate ion concentrations are in a one-to-one relationship with the molar solubility of CuCO_3 . Use the solubility product constant expression to solve for the solubility. Because K_{sp} is of the order of 10^{-10} , you can predict that the solubility will be the square root of K_{sp} , or about 10^{-5} .

Known

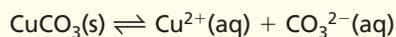
$$K_{\text{sp}} (\text{CuCO}_3) = 2.5 \times 10^{-10}$$

Unknown

$$\text{solubility of } \text{CuCO}_3 = ? \text{ mol/L}$$

2. Solve for the Unknown

Write the balanced chemical equation for the solubility equilibrium and the solubility product constant expression.



$$K_{\text{sp}} = [\text{Cu}^{2+}][\text{CO}_3^{2-}] = 2.5 \times 10^{-10}$$

Relate the solubility to $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$.

$$s = [\text{Cu}^{2+}] = [\text{CO}_3^{2-}]$$

Substitute s for $[\text{Cu}^{2+}]$ and $[\text{CO}_3^{2-}]$ and solve for s .

$$(s)(s) = s^2 = 2.5 \times 10^{-10}$$

$$s = \sqrt{2.5 \times 10^{-10}} = 1.6 \times 10^{-5} \text{ mol/L}$$

The molar solubility of CuCO_3 in water at 298 K is 1.6×10^{-5} mol/L.

3. Evaluate the Answer

The K_{sp} value has two significant figures, so the answer is correctly expressed with two digits. As predicted, the molar solubility of CuCO_3 is approximately 10^{-5} mol/L.



Finely ground copper carbonate is added to cattle and poultry feed to supply the necessary element, copper, to animal diets.



For more practice using the solubility product constant expression, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

17. Use the data in Table 18-3 to calculate the solubility in mol/L of these ionic compounds at 298 K.
- | | |
|---------------------|--------------------|
| a. PbCrO_4 | c. CaCO_3 |
| b. AgCl | d. CaSO_4 |

You have learned that the solubility product constant can be used to determine the molar solubility of an ionic compound. You can apply this information as you do the **CHEMLAB** at the end of this chapter. K_{sp} also can be used to find the concentrations of the ions in a saturated solution.

EXAMPLE PROBLEM 18-6

Calculating Ion Concentration from K_{sp}

Magnesium hydroxide is a white solid that is processed from seawater. Determine the hydroxide ion concentration at 298 K in a saturated solution of $\text{Mg}(\text{OH})_2$ if the K_{sp} equals 5.6×10^{-12} .

1. Analyze the Problem

You have been given the K_{sp} for $\text{Mg}(\text{OH})_2$. The moles of Mg^{2+} ions in solution equal the moles of $\text{Mg}(\text{OH})_2$ that dissolved, but the moles of OH^- ions in solution are two times the moles of $\text{Mg}(\text{OH})_2$ that dissolved. You can use these relationships to write the solubility product constant expression in terms of one unknown. Because the equilibrium expression is a third power equation, you can predict that $[\text{OH}^-]$ will be approximately the cube root of 10^{-12} , or approximately 10^{-4} .

Known

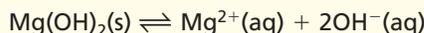
$$K_{\text{sp}} = 5.6 \times 10^{-12}$$

Unknown

$$[\text{OH}^-] = ? \text{ mol/L}$$

2. Solve for the Unknown

Write the equation for the solubility equilibrium and the K_{sp} expression.



$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 5.6 \times 10^{-12}$$

Let x equal $[\text{Mg}^{2+}]$. Because there are two OH^- ions for every Mg^{2+} ion, $[\text{OH}^-] = 2x$. Substitute these terms into the K_{sp} expression and solve for x .

$$(x)(2x)^2 = 5.6 \times 10^{-12}$$

$$(x)(4)(x)^2 = 5.6 \times 10^{-12}$$

$$4x^3 = 5.6 \times 10^{-12}$$

$$x^3 = \frac{5.6 \times 10^{-12}}{4} = 1.4 \times 10^{-12}$$

$$x = [\text{Mg}^{2+}] = \sqrt[3]{1.4 \times 10^{-12}} = 1.1 \times 10^{-4} \text{ mol/L}$$

Multiply $[\text{Mg}^{2+}]$ by 2 to obtain $[\text{OH}^-]$.

$$[\text{OH}^-] = 2[\text{Mg}^{2+}] = 2(1.1 \times 10^{-4} \text{ mol/L}) = 2.2 \times 10^{-4} \text{ mol/L}$$

3. Evaluate the Answer

The given K_{sp} has two significant figures, so the answer is correctly stated with two digits. As predicted, $[\text{OH}^-]$ is about 10^{-4} mol/L.

This thick white suspension of magnesium hydroxide ($\text{Mg}(\text{OH})_2$) is the antacid milk of magnesia. Because $\text{Mg}(\text{OH})_2$ is only sparingly soluble, few formula units are in solution. These will be neutralized by stomach acid and then others will dissolve. Can you explain that statement using Le Châtelier's principle?



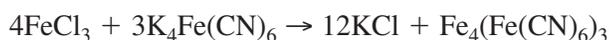
PRACTICE PROBLEMS

18. Use K_{sp} values from Table 18-3 to calculate the following.
- $[Ag^+]$ in a solution of AgBr at equilibrium
 - $[F^-]$ in a saturated solution of CaF_2
 - $[Ag^+]$ in a solution of Ag_2CrO_4 at equilibrium
 - the solubility of PbI_2



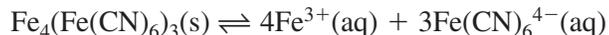
For more practice using the solubility product constant expression, go to **Supplemental Practice Problems** in Appendix A.

Predicting precipitates You have seen that solubility product constants are useful for finding the solubility of an ionic compound and for calculating the concentrations of ions in a saturated solution. You also can use K_{sp} to predict whether a precipitate will form when two ionic solutions are mixed. For example, will a precipitate form when equal volumes of 0.10M aqueous solutions of iron(III) chloride ($FeCl_3$) and potassium hexacyanoferrate(II) ($K_4Fe(CN)_6$) are poured together? A double-replacement reaction might occur according to this equation.



A precipitate is likely to form only if either product, KCl or $Fe_4(Fe(CN)_6)_3$, has low solubility. You are probably aware that KCl is a soluble compound and would be unlikely to precipitate. However, the K_{sp} for $Fe_4(Fe(CN)_6)_3$ is a very small number, 3.3×10^{-41} , which suggests that $Fe_4(Fe(CN)_6)_3$ might be expected to precipitate if the concentrations of its ions are large enough. Perhaps you are wondering how large is large enough.

The following equilibrium is possible between solid $Fe_4(Fe(CN)_6)_3$ (a precipitate) and its ions in solution, Fe^{3+} and $Fe(CN)_6^{4-}$.



When the two solutions are mixed, if the concentrations of the ions Fe^{3+} and $Fe(CN)_6^{4-}$ are greater than those that can exist in a saturated solution of $Fe_4(Fe(CN)_6)_3$, the equilibrium will shift to the left and $Fe_4(Fe(CN)_6)_3(s)$ will precipitate. To make a prediction, then, the ion concentrations must be calculated.

Table 18-4 shows the concentrations of the ions of reactants and products in the original solutions (0.10M $FeCl_3$ and 0.10M $K_4Fe(CN)_6$) and in the mixture immediately after equal volumes of the two solutions were mixed. Note that $[Cl^-]$ is three times as large as $[Fe^{3+}]$ because the ratio of Cl^- to Fe^{3+} in $FeCl_3$ is 3:1. Also note that $[K^+]$ is four times as large as $[Fe(CN)_6^{4-}]$ because the ratio of K^+ to $Fe(CN)_6^{4-}$ in $K_4Fe(CN)_6$ is 4:1. In addition, note that the concentration of each ion in the mixture is one-half its original concentration. This is because when equal volumes of two solutions are mixed, the same number of ions are dissolved in twice as much solution, so the concentration is reduced by one-half.

You can now use the data in the table to make a trial to see if the concentrations of Fe^{3+} and $Fe(CN)_6^{4-}$ in the mixed solution exceed the value of K_{sp} when substituted into the solubility product constant expression.

$$K_{sp} = [Fe^{3+}]^4[Fe(CN)_6^{4-}]^3$$

But first, remember that you have not determined whether the solution is saturated. When you make this substitution, it will not necessarily give the solubility product constant. Instead, it provides a number called the ion product, Q_{sp} . Q_{sp} is a trial value that can be compared with K_{sp} .

Table 18-4

Ion Concentrations in Original and Mixed Solutions	
Original solutions (mol/L)	Mixture (mol/L)
$[Fe^{3+}] = 0.10$	$[Fe^{3+}] = 0.050$
$[Cl^-] = 0.30$	$[Cl^-] = 0.15$
$[K^+] = 0.40$	$[K^+] = 0.20$
$[Fe(CN)_6^{4-}] = 0.10$	$[Fe(CN)_6^{4-}] = 0.050$



Figure 18-15

Because the ion product constant (Q_{sp}) is greater than K_{sp} , you could predict that this precipitate of $\text{Fe}_4(\text{Fe}(\text{CN})_6)_3$ would form.



The peeling lead-based paint on this building is a hazard to small children who sometimes ingest small pieces of paint. Lead can build up in the body and cause many physical symptoms as well as brain damage and mental retardation.

$$Q_{sp} = [\text{Fe}^{3+}]^4[\text{Fe}(\text{CN})_6^{4-}]^3 = (0.050)^4(0.050)^3 = 7.8 \times 10^{-10}$$

You can now compare Q_{sp} and K_{sp} . This comparison can have one of three outcomes: Q_{sp} can be less than K_{sp} , equal to K_{sp} , or greater than K_{sp} .

1. If $Q_{sp} < K_{sp}$, the solution is unsaturated. No precipitate will form.
2. If $Q_{sp} = K_{sp}$, the solution is saturated and no change will occur.
3. If $Q_{sp} > K_{sp}$, a precipitate will form reducing the concentrations of the ions in the solution until the product of their concentrations in the K_{sp} expression equals the numerical value of K_{sp} . Then the system is in equilibrium and the solution is saturated.

In the case of the $\text{Fe}_4(\text{Fe}(\text{CN})_6)_3$ equilibrium, Q_{sp} (7.8×10^{-10}) is larger than K_{sp} (3.3×10^{-41}) so a deeply colored blue precipitate of $\text{Fe}_4(\text{Fe}(\text{CN})_6)_3$ forms, as you can see in **Figure 18-15**. Example Problem 18-7 and the **problem-solving LAB** on the following page will give you practice in calculating Q_{sp} to predict precipitates.

EXAMPLE PROBLEM 18-7

Predicting a Precipitate

Predict whether a precipitate of PbCl_2 will form if 100 mL of 0.0100M NaCl is added to 100 mL of 0.0200M $\text{Pb}(\text{NO}_3)_2$.

1. Analyze the Problem

You have been given equal volumes of two solutions with known concentrations. The concentrations of the initial solutions allow you to calculate the concentrations of Pb^{2+} and Cl^- ions in the mixed solution. The initial concentrations, when multiplied together in the solubility product constant expression, give an ion product of the order of 10^{-6} , so it is probable that after dilution, Q_{sp} will be less than K_{sp} (1.7×10^{-5}) and PbCl_2 will not precipitate.

Known

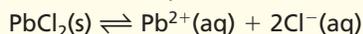
100 mL 0.0100M NaCl
100 mL 0.0200M $\text{Pb}(\text{NO}_3)_2$
 $K_{sp} = 1.7 \times 10^{-5}$

Unknown

$Q_{sp} > K_{sp}?$

2. Solve for the Unknown

Write the equation for the dissolving of PbCl_2 and the ion product expression, Q_{sp} .



$$Q_{sp} = [\text{Pb}^{2+}][\text{Cl}^{-}]^2$$

Divide the concentrations of Cl^- and Pb^{2+} in half because on mixing, the volume doubles.

$$[\text{Cl}^{-}] = \frac{0.0100M}{2} = 0.00500M$$

$$[\text{Pb}^{2+}] = \frac{0.0200M}{2} = 0.0100M$$

Substitute these values into the ion product expression.

$$Q_{sp} = (0.0100)(0.00500)^2 = 2.5 \times 10^{-7}$$

Compare Q_{sp} with K_{sp} .

$$Q_{sp} (2.5 \times 10^{-7}) < K_{sp} (1.7 \times 10^{-5})$$

A precipitate will not form.

3. Evaluate the Answer

As predicted, Q_{sp} is less than K_{sp} . The Pb^{2+} and Cl^- ions are not present in high enough concentrations in the mixed solution to cause precipitation to occur.

PRACTICE PROBLEMS

19. Use K_{sp} values from Table 18-3 to predict whether a precipitate will form when equal volumes of the following aqueous solutions are mixed.
- 0.10M $Pb(NO_3)_2$ and 0.030M NaF
 - 0.25M K_2SO_4 and 0.010M $AgNO_3$
 - 0.20M $MgCl_2$ and 0.0025M NaOH



For more practice predicting precipitates, go to **Supplemental Practice Problems** in Appendix A.

Common Ion Effect

The solubility of lead chromate ($PbCrO_4$) in water is 4.8×10^{-7} mol/L at 298 K. That means you can dissolve 4.8×10^{-7} mol $PbCrO_4$ in 1.00 L of pure water. However, you can't dissolve 4.8×10^{-7} mol $PbCrO_4$ in 1.00 L of 0.10M aqueous potassium chromate (K_2CrO_4) solution at that temperature. Why is $PbCrO_4$ less soluble in an aqueous K_2CrO_4 solution than in pure water?

problem-solving LAB

How does fluoride prevent tooth decay?

Using Numbers During the last half century, tooth decay has decreased significantly because minute quantities of fluoride ion ($6 \times 10^{-5}M$) are being added to most public drinking water and many people are using toothpastes containing sodium fluoride or tin(II) fluoride. Use what you know about the solubility of ionic compounds and reversible reactions to explore the role of the fluoride ion in maintaining cavity-free teeth.

Analysis

Tooth enamel, the hard, protective outer layer of the tooth, is 98% hydroxyapatite ($Ca_5(PO_4)_3OH$). Although quite insoluble ($K_{sp} = 6.8 \times 10^{-37}$ in water), demineralization, which is the dissolving of hydroxyapatite, does occur especially when the saliva contains acids. The reverse reaction, remineralization, also occurs. Remineralization is the re-depositing of tooth enamel. When hydroxyapatite is in solution with fluoride ions, a double-replacement reaction can occur. Fluoride ion replaces the hydroxide ion to form fluoroapatite ($Ca_5(PO_4)_3F$, $K_{sp} = 1 \times 10^{-60}$). Fluoroapatite remineralizes the tooth enamel, thus partially displacing hydroxyapatite. Because fluoroapatite is less soluble than hydroxyapatite, destructive demineralization is reduced.



Thinking Critically

- Write the equation for the dissolving of hydroxyapatite and its equilibrium constant expression. How do the conditions in the mouth differ from those of a true equilibrium?
- Write the equation for the double-replacement reaction that occurs between hydroxyapatite and sodium fluoride.
- Calculate the solubility of hydroxyapatite and fluoroapatite in water. Compare the solubilities.
- What is the ion product constant (Q_{sp}) for the reaction if 0.00050M NaF is mixed with an equal volume of 0.000015M $Ca_5(PO_4)_3OH$? Will a precipitate form (re-mineralization)?

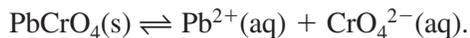
Figure 18-16

In a solution of PbCrO_4 in pure water, $[\text{Pb}^{2+}]$ and $[\text{CrO}_4^{2-}]$ are equal. But in a solution with a common CrO_4^{2-} ion, $[\text{Pb}^{2+}]$ is much less than $[\text{CrO}_4^{2-}]$. In each case, however, the product of the two ions equals K_{sp} and the solubility of PbCrO_4 equals $[\text{Pb}^{2+}]$.

$$[\text{Pb}^{2+}][\text{CrO}_4^{2-}] = 2.3 \times 10^{-13} \text{ (in water)}$$

$$[\text{Pb}^{2+}][\text{CrO}_4^{2-}] = 2.3 \times 10^{-13} \text{ (in } 0.10M \text{ K}_2\text{CrO}_4\text{)}$$

The equation for the PbCrO_4 solubility equilibrium and the solubility product constant expression are

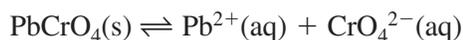


$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{CrO}_4^{2-}] = 2.3 \times 10^{-13}$$

Recall that K_{sp} is a constant at any given temperature, so if the concentration of either Pb^{2+} or CrO_4^{2-} increases when the system is at equilibrium, the concentration of the other ion must decrease. The product of the concentrations of the two ions must always equal K_{sp} . The K_2CrO_4 solution contains CrO_4^{2-} ions before any PbCrO_4 dissolves. In this example, the CrO_4^{2-} ion is called a common ion because it is part of both PbCrO_4 and K_2CrO_4 . A **common ion** is an ion that is common to two or more ionic compounds. The lowering of the solubility of a substance by the presence of a common ion is called the **common ion effect**.

Figure 18-16 illustrates how the balance between the ion concentrations differs in a solution of PbCrO_4 in pure water and in a $0.10M$ solution of K_2CrO_4 . In each case, $[\text{Pb}^{2+}]$ represents the solubility of PbCrO_4 .

The common ion effect and Le Châtelier's principle A saturated solution of lead chromate (PbCrO_4) is shown in **Figure 18-17a**. Note the solid yellow PbCrO_4 in the bottom of the test tube. The solution and solid are in equilibrium according to this equation.



When a solution of $\text{Pb}(\text{NO}_3)_2$ is added to the saturated PbCrO_4 solution, more solid PbCrO_4 precipitates, as you can see in **Figure 18-17b**. The Pb^{2+} ion, common to both $\text{Pb}(\text{NO}_3)_2$ and PbCrO_4 , reduces the solubility of PbCrO_4 . Can this precipitation of PbCrO_4 be explained by Le Châtelier's principle? Adding Pb^{2+} ion to the solubility equilibrium stresses the equilibrium. To relieve the stress, the equilibrium shifts to the left to form more solid PbCrO_4 .

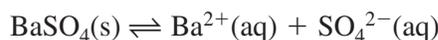
The common ion effect also plays a role in the use of barium sulfate (BaSO_4) when X rays are taken of the digestive system. Recall that patients who need such X rays must drink a mixture containing BaSO_4 . The low

solubility of BaSO_4 helps ensure that the amount of the toxic barium ion absorbed into patient's system is small enough to be harmless. The procedure is further safe-guarded by the addition of sodium sulfate (Na_2SO_4), a soluble ionic compound that provides a common ion, SO_4^{2-} . How does the additional SO_4^{2-} affect the concentration of barium ion in the mixture that patients must drink?

Figure 18-17

In **a**, $\text{PbCrO}_4(\text{s})$ is in equilibrium with its ions in solution. Then, in **b**, the equilibrium is stressed by the addition of $\text{Pb}(\text{NO}_3)_2$ and more PbCrO_4 precipitate forms. Use the equation and Le Châtelier's principle to convince yourself that a shift to the left occurred. Then write the K_{sp} for this reaction as in **Figure 18-16**.





Le Châtelier's principle predicts that additional SO_4^{2-} from the Na_2SO_4 will shift the equilibrium to the left to produce more solid BaSO_4 and reduce the number of harmful Ba^{2+} ions in solution.

Solubility Equilibria in the Laboratory

Suppose you are given a clear aqueous solution and told that it could contain almost any of the ions formed by the common metallic elements. Your job is to find out which ions the solution contains. It seems an impossible task, but chemists have worked out a scheme that allows you to separate and identify many common metal ions. The scheme is based upon the differing solubilities of the ions. For example, only three ions form insoluble chlorides— Ag^+ , Pb^{2+} , and Hg_2^{2+} . That may give you an idea of how you could proceed. If you add HCl to your solution, those three ions will precipitate as a white solid that would contain $\text{AgCl}(\text{s})$, $\text{PbCl}_2(\text{s})$, and $\text{Hg}_2\text{Cl}_2(\text{s})$ if all three of the metal ions are present. The ions that do not form insoluble chlorides will be left in the clear solution, which can be separated from the white solid.

Is PbCl_2 present in the white solid? To find out, you can use the fact that PbCl_2 is soluble in hot water but AgCl and Hg_2Cl_2 are not. After adding water and heating the mixture of chlorides, you can separate the liquid (that could contain dissolved PbCl_2) from AgCl and Hg_2Cl_2 , which are still in solid form. What test could you use to show that the separated liquid contains Pb^{2+} ? Remember that PbCrO_4 is a sparingly soluble compound. You could add K_2CrO_4 . If a yellow precipitate of PbCrO_4 forms, you have proof of the presence of Pb^{2+} .

Are Hg_2Cl_2 , or AgCl , or both compounds present in the remaining precipitate? Silver chloride is soluble in ammonia; mercury(I) chloride forms a black precipitate with ammonia. Suppose you add ammonia and get a black precipitate. You can conclude that Hg_2^{2+} is present. But did AgCl dissolve in ammonia, or was all the white precipitate Hg_2Cl_2 ? To find out, you can add HCl to the solution. HCl will interact with the ammonia in the solution and cause AgCl to precipitate again.

The procedure you have been reading is just one step in the complete analysis of the initial solution of ions. **Figure 18-18** is a flow chart of the steps in the procedure. Follow the chart step by step through each identification. Note that as each reactant is added (HCl, hot water, NH_3), a separation is made between those ions that are soluble and those that are not.

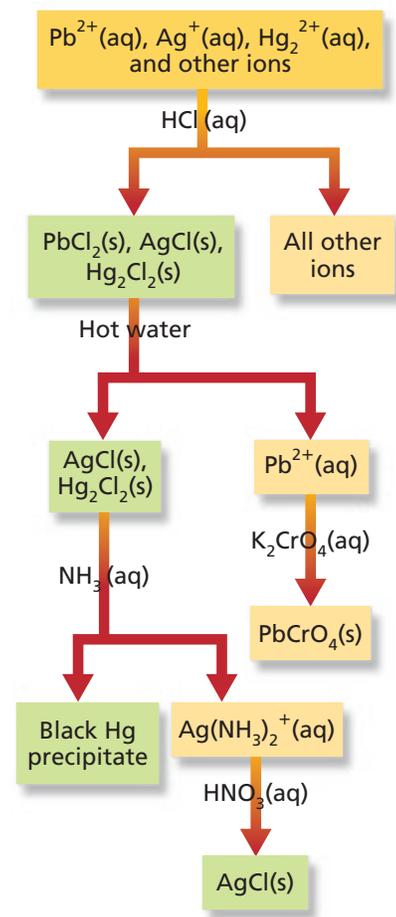


Figure 18-18

Follow the process down the chart. The reactants added at each step are written on the downward arrows. Which reactions are separations? Which reactions confirm the presence of an ion?

Section 18.3 Assessment

- List the information you would need in order to calculate the concentration of a product in a reaction mixture at equilibrium.
- How can you use the solubility product constant to calculate the solubility of a sparingly soluble ionic compound?
- What is a common ion? Explain how a common ion reduces the solubility of an ionic compound.
- Thinking Critically** When aqueous solutions of two ionic compounds are mixed, how does Q_{sp} relate to K_{sp} for a possible precipitate?
- Designing an Experiment** An aqueous solution is known to contain either Mg^{2+} or Pb^{2+} . Design an experiment based on solubilities that would help you determine which of the two ions is present. The solubilities of many ionic compounds are given in **Table C-10** in Appendix C.

Comparing Two Solubility Product Constants

Le Châtelier's principle is a powerful tool for explaining how a reaction at equilibrium shifts when a stress is placed on the system. In this experiment, you can use Le Châtelier's principle to evaluate the relative solubilities of two precipitates. By observing the formation of two precipitates in the same system, you can infer the relationship between the solubilities of the two ionic compounds and the numerical values of their solubility product constants (K_{sp}). You will be able to verify your own experimental results by calculating the molar solubilities of the two compounds using the K_{sp} for each compound.

Problem

How can a saturated solution of one ionic compound react with another ionic compound to form another precipitate? What is the relationship between solubility and the K_{sp} value of a saturated solution?

Objectives

- **Observe** evidence that a precipitate is in equilibrium with its ions in solution.
- **Infer** the relative solubilities of two sparingly soluble ionic compounds.
- **Compare** the values of the K_{sp} for two different compounds and relate them to your observations.
- **Explain** your observations of the two precipitates by using Le Châtelier's principle.
- **Calculate** the molar solubilities of the two ionic compounds from their K_{sp} values.

Materials

AgNO₃ solution
NaCl solution
Na₂S solution
24-well microplate
thin-stem pipettes (3)
wash bottle

Safety Precautions



- Always wear safety goggles, gloves, and a lab apron.
- Silver nitrate is highly toxic and will stain skin and clothing.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your observations.
3. State Le Châtelier's principle.
4. Identify the control and the independent variable in the experiment.
5. When a solid dissolves to form two ions and the solid's K_{sp} is known, what is the mathematical formula you can use to calculate the molar solubility?

Procedure

1. Place 10 drops of AgNO₃ solution in well A1 of a 24-well microplate. Place 10 drops of the same solution in well A2.
2. Add 10 drops of NaCl solution to well A1 and 10 drops to well A2.

Precipitate Formation	
	Observations
Step 3	
Step 5	
Step 6	



3. Allow the precipitate to form in each well. Record your observations.
4. To well A2, add 10 drops of Na_2S solution.
5. Allow the precipitate to form. Record your observations of the precipitate.
6. Compare the contents of wells A1 and A2 and record your observations in your table.

Cleanup and Disposal

1. Use a wash bottle to transfer the contents of the well plate into a large waste beaker.
2. Wash your hands thoroughly after all lab work and cleanup are complete.

Analyze and Conclude

1. **Analyzing Information** Write the complete equation for the double-replacement reaction that occurred when NaCl and AgNO_3 were mixed in wells A1 and A2 in step 2. Write the net ionic equation.
2. **Analyzing Information** Write the solubility product constant expression for the equilibrium established in wells A1 and A2 in step 2.
 $K_{\text{sp}}(\text{AgCl}) = 1.8 \times 10^{-10}$.
3. **Analyzing Information** Write the equation for the equilibrium that was established in well A2 when you added Na_2S . $K_{\text{sp}}(\text{Ag}_2\text{S}) = 8 \times 10^{-48}$
4. **Inferring** Identify the two precipitates by color.

5. Comparing and Contrasting Compare the K_{sp} values for the two precipitates. Infer which of the two ionic compounds is more soluble.

6. Recognizing Cause and Effect Use Le Châtelier's principle to explain how the addition of Na_2S in procedure step 4 affected the equilibrium established in well A2.

7. Using Numbers Calculate the molar solubilities of the two precipitates using the K_{sp} values. Which of the precipitates is more soluble?

8. Thinking Critically What evidence from this experiment supports your answer to question 7? Explain.

9. Error Analysis Did you observe the well plate from the side as well as from the top? What did you notice?

10. Developing General Rules The solubility of an ionic compound depends upon the nature of the cations and anions that make up the compound. The reactants you used in this **CHEMLAB** are all soluble ionic compounds, whereas, the precipitates are insoluble. How does soluble Na_2S differ from insoluble Ag_2S ? How does soluble NaCl differ from insoluble AgCl ? Use this information and K_{sp} data from **Table 18-3** and the *Handbook of Chemistry and Physics* to develop general rules for solubility. What group of metal ions is not found in sparingly soluble compounds? What polyatomic ions, positive and negative, form only soluble ionic compounds? How does K_{sp} relate to a compound's relative solubility?

Real-World Chemistry

1. Research how industries use precipitation to remove hazardous chemicals from wastewater before returning it to the water cycle.
2. *Hard water* is the name given to water supplies that contain significant concentrations of Mg^{2+} and Ca^{2+} ions. Check on the solubility of ionic compounds formed with these ions and predict what problems they may cause.
3. Explain what would happen if you lost the stopper for a bottle of a saturated solution of lead sulfate (PbSO_4) and the bottle stood open to the air for a week. Would your answer be different if it were an unsaturated solution? Explain.

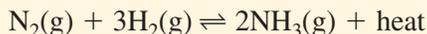
CHEMISTRY and Technology

The Haber Process

Diatomic nitrogen makes up about 79 percent of Earth's atmosphere. A few species of soil bacteria can use atmospheric nitrogen to produce ammonia (NH_3). Other species of bacteria then convert the ammonia into nitrite and nitrate ions, which can be absorbed and used by plants. Ammonia also can be synthesized.

Applying Le Châtelier's Principle

The process of synthesizing large amounts of ammonia from nitrogen and hydrogen gases was first demonstrated in 1909 by Fritz Haber, a German research chemist, and his English research assistant, Robert LeRossignol. The Haber process involves this reaction.



The process produces high yields of ammonia by manipulating three factors that influence the reaction—pressure, temperature, and catalytic action.

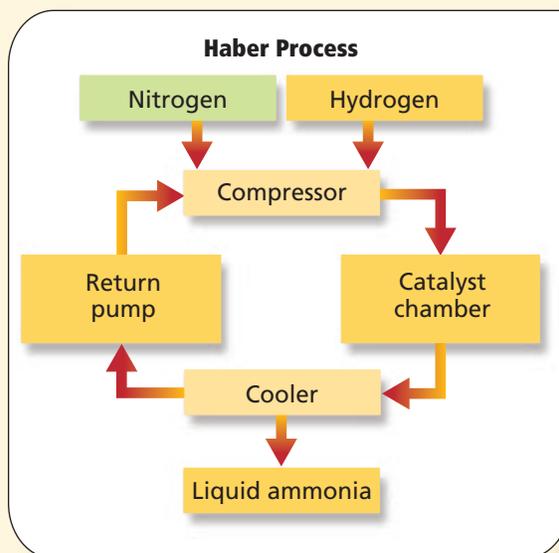
During the synthesis of ammonia, four molecules of reactant produce two molecules of product. According to Le Châtelier's principle, if the pressure on this reaction is increased, the forward reaction will speed up to reduce the stress because two molecules exert less pressure than four molecules. Increased pressure will also cause the reactants to collide more frequently, thus increasing the reaction rate. Haber's apparatus used a pressure of 2×10^5 kPa.

The forward reaction is favored by a low temperature because the stress caused by the heat generated by the reaction is reduced. But low temperature decreases the number of collisions between reactants, thus decreasing the rate of reaction. Haber compromised by using an intermediate temperature of about 450°C .

A catalyst is used to decrease the activation energy and thus, increase the rate at which equilibrium is reached. Haber used iron as a catalyst in his process.

The Industrial Process

Haber's process incorporated several operations that increased the yield of ammonia. The reactant



gases entering the chamber were warmed by heat produced by the reaction. The reactant-product mixture was allowed to cool slowly after reacting over the catalyst. Ammonia gas was removed from the process by liquefaction, and the unreacted nitrogen and hydrogen were recycled back into the process.

Carl Bosch, a German industrial chemist improved Haber's process by designing new reaction chambers, improving pressurizing pumps, and finding inexpensive catalysts. By 1913, Bosch had built the first plant for synthesizing ammonia in Oppau, Germany.

Investigating the Technology

- Thinking Critically** How does removing ammonia from the process affect equilibrium?
- Using Resources** Research how ammonia is used in the production of many agricultural fertilizers and other products.



Visit the Chemistry Web site at science.glencoe.com to find links to more information about the Haber process and ammonia.

Summary

18.1 Equilibrium: A State of Dynamic Balance

- A reversible reaction is one that can take place in both the forward and reverse directions.
- A reversible reaction leads to an equilibrium state in which the forward and reverse reactions take place at equal rates and the concentrations of reactants and products remain constant.
- You can write the equilibrium constant expression for an equilibrium system using the law of chemical equilibrium.
- The equilibrium constant expression is a ratio of the molar concentrations of the products divided by the molar concentrations of the reactants with all concentrations in the ratio raised to a power equal to their coefficients in the balanced chemical equation.
- The value of the equilibrium constant expression, K_{eq} , is a constant for a given temperature.
- A large value for K_{eq} means that the products are favored at equilibrium; a small K_{eq} value means that the reactants are favored.
- You can calculate K_{eq} by substituting known equilibrium concentrations into the equilibrium constant expression.

18.2 Factors Affecting Chemical Equilibrium

- Le Châtelier's principle describes how an equilibrium system shifts in response to a stress or disturbance. A stress is any change in the system at equilibrium.

- An equilibrium can be forced in the direction of the products by adding a reactant or by removing a product. It can be forced in the direction of the reactants by adding a product or removing a reactant.
- When an equilibrium shifts in response to a change in concentration or volume, the equilibrium position changes but K_{eq} remains constant. A change in temperature, however, alters both the equilibrium position and the value of K_{eq} .

18.3 Using Equilibrium Constants

- Given K_{eq} , the equilibrium concentration of a substance can be calculated if you know the equilibrium concentrations of all other reactants and products.
- The solubility product constant expression, K_{sp} , describes the equilibrium between a sparingly soluble ionic compound and its ions in solution.
- You can calculate the molar solubility of an ionic compound using the solubility product constant expression.
- The ion product, Q_{sp} , can be calculated from the molar concentrations of the ions in a solution and compared with the K_{sp} to determine whether a precipitate will form when two solutions are mixed.
- The solubility of a substance is lower when the substance is dissolved in a solution containing a common ion. This is called the common ion effect.

Key Equations and Relationships

- $K_{\text{eq}} = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}$ (p. 563)

Vocabulary

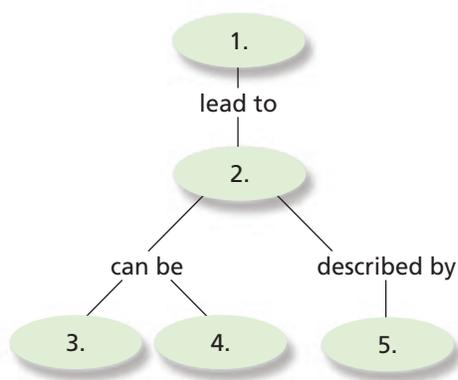
- chemical equilibrium (p. 561)
- common ion (p. 584)
- common ion effect (p. 584)
- equilibrium constant (p. 563)
- heterogeneous equilibrium (p. 565)
- homogeneous equilibrium (p. 564)
- law of chemical equilibrium (p. 563)
- Le Châtelier's principle (p. 569)
- reversible reaction (p. 560)
- solubility product constant (p. 578)



Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 18 Assessment.

Concept Mapping

25. Fill in the spaces on the concept map with the following phrases: equilibrium constant expressions, reversible reactions, heterogeneous equilibria, homogeneous equilibria, chemical equilibria.



Mastering Concepts

26. Describe an equilibrium in everyday life that illustrates a state of balance between two opposing processes. (18.1)
27. Given the fact that the concentrations of reactants and products are not changing, why is the word *dynamic* used for describing chemical equilibrium? (18.1)
28. How can you indicate in a chemical equation that a reaction is reversible? (18.1)
29. Although the general equation for a chemical reaction is reactants \rightarrow products, explain why this equation is not complete for a system at equilibrium. (18.1)
30. Explain the difference between a homogeneous equilibrium and a heterogeneous equilibrium. (18.1)
31. What is an equilibrium position? (18.1)
32. Explain how to use the law of chemical equilibrium in writing an equilibrium constant expression. (18.1)
33. Why does a numerically large K_{eq} mean that the products are favored in an equilibrium system? (18.1)
34. Why should you pay attention to the physical states of all reactants and products when writing equilibrium constant expressions? (18.1)
35. How can an equilibrium system contain small and unchanging amounts of products yet have large amounts of reactants? What can you say about the relative size of K_{eq} for such an equilibrium? (18.1)
36. Describe the opposing processes in the physical equilibrium that exists in a closed container half-filled with liquid ethanol. (18.1)
37. What is meant by a stress on a reaction at equilibrium? (18.2)
38. How does Le Châtelier's principle describe an equilibrium's response to a stress? (18.2)
39. Why does removing a product cause an equilibrium to shift in the direction of the products? (18.2)
40. When an equilibrium shifts toward the reactants in response to a stress, how is the equilibrium position changed? (18.2)
41. Use Le Châtelier's principle to explain how a shift in the equilibrium $\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ causes a soft drink to go flat when its container is left open to the atmosphere. (18.2)
42. How is K_{eq} changed when heat is added to an equilibrium in which the forward reaction is exothermic? Explain using Le Châtelier's principle. (18.2)
43. Changing the volume of the system alters the equilibrium position of this equilibrium.

$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$$
 But a similar change has no effect on this equilibrium.

$$\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons 2\text{HCl}(\text{g})$$
 Explain. (18.2)
44. How might the addition of a noble gas to the reaction vessel affect this equilibrium?

$$2\text{N}_2\text{H}_4(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 3\text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$$
 Assume that the volume of the reaction vessel does not change. (18.2)
45. When an equilibrium shifts to the right, what happens to the following? (18.2)
- the concentrations of the reactants
 - the concentrations of the products
46. How would each of the following changes affect the equilibrium position of the system used to produce methanol from carbon monoxide and hydrogen? (18.2)

$$\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g}) + \text{heat}$$
- adding CO to the system
 - cooling the system
 - adding a catalyst to the system
 - removing CH_3OH from the system
 - decreasing the volume of the system

47. Why is the concentration of a solid not included as part of the solubility product constant? (18.3)
48. What does it mean to say that two solutions have a common ion? Give an example that supports your answer. (18.3)
49. Explain the difference between Q_{sp} and K_{sp} . (18.3)
50. Explain why a common ion lowers the solubility of an ionic compound. (18.3)
51. Describe the solution that results when two solutions are mixed and Q_{sp} is found to equal K_{sp} . Does a precipitate form?

Mastering Problems

The Equilibrium Constant Expression (18.1)

52. Write equilibrium constant expressions for these homogeneous equilibria.
- $2\text{N}_2\text{H}_4(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 3\text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$
 - $2\text{NbCl}_4(\text{g}) \rightleftharpoons \text{NbCl}_3(\text{g}) + \text{NbCl}_5(\text{g})$
 - $\text{I}_2(\text{g}) \rightleftharpoons 2\text{I}(\text{g})$
 - $2\text{SO}_3(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{CS}_2(\text{g}) + 4\text{O}_2(\text{g})$
53. Write equilibrium constant expressions for these heterogeneous equilibria.
- $2\text{NaHCO}_3(\text{s}) \rightleftharpoons \text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g})$
 - $\text{C}_6\text{H}_6(\text{l}) \rightleftharpoons \text{C}_6\text{H}_6(\text{g})$
 - $\text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g}) \rightleftharpoons 3\text{Fe}(\text{s}) + 4\text{H}_2\text{O}(\text{g})$
54. Pure water has a density of 1.00 g/mL at 297 K. Calculate the molar concentration of pure water at this temperature.
55. Calculate K_{eq} for the following equilibrium when $[\text{SO}_3] = 0.0160 \text{ mol/L}$, $[\text{SO}_2] = 0.00560 \text{ mol/L}$, and $[\text{O}_2] = 0.00210 \text{ mol/L}$.
- $$2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$$
56. K_{eq} for this reaction is 3.63.
- $$\text{A} + 2\text{B} \rightleftharpoons \text{C}$$

The data in the table shows the concentrations of the reactants and product in two different reaction mixtures at the same temperature. Does the data provide evidence that both reactions are at equilibrium?

Table 18-5

Concentrations of A, B, and C		
A (mol/L)	B (mol/L)	C (mol/L)
0.500	0.621	0.700
0.250	0.525	0.250

57. When solid ammonium chloride is put in a reaction vessel at 323 K, the equilibrium concentrations of both ammonia and hydrogen chloride are found to be 0.0660 mol/L. $\text{NH}_4\text{Cl}(\text{s}) \rightleftharpoons \text{NH}_3(\text{g}) + \text{HCl}(\text{g})$. Calculate K_{eq} .
58. Suppose you have a cube of pure manganese metal measuring 5.25 cm on each side. You find that the mass of the cube is 1076.6 g. What is the molar concentration of manganese in the cube?

Le Châtelier's Principle (18.2)

59. Use Le Châtelier's principle to predict how each of the following changes would affect this equilibrium.
- $$\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$$
- adding $\text{H}_2\text{O}(\text{g})$ to the system
 - removing $\text{CO}(\text{g})$ from the system
 - adding $\text{H}_2(\text{g})$ to the system
 - adding something to the system to absorb $\text{CO}_2(\text{g})$
60. How would increasing the volume of the reaction vessel affect these equilibria?
- $\text{NH}_4\text{Cl}(\text{s}) \rightleftharpoons \text{NH}_3(\text{g}) + \text{HCl}(\text{g})$
 - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$
61. How would decreasing the volume of the reaction vessel affect these equilibria?
- $2\text{N}_2\text{H}_4(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 3\text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$
 - $2\text{H}_2\text{O}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
62. How would these equilibria be affected by increasing the temperature?
- $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) + \text{heat}$
 - $\text{heat} + \text{NaCl}(\text{s}) \rightleftharpoons \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
63. Ethylene (C_2H_4) reacts with hydrogen to form ethane (C_2H_6).
- $$\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightleftharpoons \text{C}_2\text{H}_6(\text{g}) + \text{heat}$$
- How would you regulate the temperature of this equilibrium in order to do the following?
- increase the yield of ethane
 - decrease the concentration of ethylene
 - increase the amount of hydrogen in the system
64. How would simultaneously decreasing the temperature and volume of the system affect these equilibria?
- $\text{heat} + \text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
 - $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) + \text{heat}$

Calculations Using K_{eq} (18.3)

65. K_{eq} is 1.60 at 933 K for this reaction.
- $$\text{H}_2(\text{g}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g})$$
- Calculate the equilibrium concentration of hydrogen when $[\text{CO}_2] = 0.320 \text{ mol/L}$, $[\text{H}_2\text{O}] = 0.240 \text{ mol/L}$, and $[\text{CO}] = 0.280 \text{ mol/L}$.

66. At 2273 K, $K_{\text{eq}} = 6.2 \times 10^{-4}$ for the reaction
 $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$
 If $[\text{N}_2] = 0.05200 \text{ mol/L}$ and $[\text{O}_2] = 0.00120 \text{ mol/L}$,
 what is the concentration of NO at equilibrium?

Calculations Using K_{sp} (18.3)

67. Calculate the ion product to determine if a precipitate will form when 125 mL 0.00500M sodium chloride is mixed with 125 mL 0.00100M silver nitrate solution.
68. Calculate the molar solubility of strontium chromate in water at 298 K if $K_{\text{sp}} = 3.5 \times 10^{-5}$.
69. Will a precipitate form when 1.00 L of 0.150M iron(II) chloride solution is mixed with 2.00 L of 0.0333M sodium hydroxide solution? Explain your reasoning and show your calculations.

Mixed Review

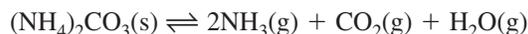
Sharpen your problem-solving skills by answering the following.

70. How many moles per liter of silver chloride will be in a saturated solution of AgCl? $K_{\text{sp}} = 1.8 \times 10^{-10}$
71. A 6.00-L vessel contains an equilibrium mixture of 0.0222 mol PCl_3 , 0.0189 mol PCl_5 , and 0.1044 mol Cl_2 . Calculate K_{eq} for the following reaction.
 $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
72. How would simultaneously increasing the temperature and volume of the system affect these equilibria?
 a. $2\text{O}_3(\text{g}) \rightleftharpoons 3\text{O}_2(\text{g}) + \text{heat}$
 b. $\text{heat} + \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$
73. The solubility product constant for lead(II) arsenate ($\text{Pb}_3(\text{AsO}_4)_2$), is 4.0×10^{-36} at 298 K. Calculate the molar solubility of the compound at this temperature.
74. How would these equilibria be affected by decreasing the temperature?
 a. $2\text{O}_3(\text{g}) \rightleftharpoons 3\text{O}_2(\text{g}) + \text{heat}$
 b. $\text{heat} + \text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{HF}(\text{g})$

Thinking Critically

75. **Predicting** Suppose you're thinking about using the following reaction to produce hydrogen from hydrogen sulfide.
 $2\text{H}_2\text{S}(\text{g}) + \text{heat} \rightleftharpoons 2\text{H}_2(\text{g}) + \text{S}_2(\text{g})$
 Given that K_{eq} for the equilibrium is 2.27×10^{-4} , would you expect a high yield of hydrogen? Explain how you could regulate the volume of the reaction vessel and the temperature to increase the yield.

76. **Applying Concepts** Smelling salts, sometimes used to revive a groggy or unconscious person, are made of ammonium carbonate. The equation for the endothermic decomposition of ammonium carbonate is



Would you expect smelling salts to work as well on a cold winter day as on a warm summer day? Explain your answer.

77. **Comparing and Contrasting** Which of the two solids, calcium phosphate or iron(III) phosphate, has the greater molar solubility? $K_{\text{sp}}(\text{Ca}_3(\text{PO}_4)_2) = 1.2 \times 10^{-29}$; $K_{\text{sp}}(\text{FePO}_4) = 1.0 \times 10^{-22}$. Which compound has the greater solubility expressed in grams per liter?
78. **Recognizing Cause and Effect** You have 12.56 g of a mixture made up of sodium chloride and barium chloride. Explain how you could use a precipitation reaction to determine how much of each compound the mixture contains.

Writing in Chemistry

79. Research the role that solubility plays in the formation of kidney stones. Find out what compounds are found in kidney stones and their K_{sp} values. Summarize your findings in a report.
80. The presence of magnesium and calcium ions in water makes the water "hard." Explain in terms of solubility why the presence of these ions is often undesirable. Find out what measures can be taken to eliminate them.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

81. How are electrons shared differently in H_2 , O_2 , and N_2 ? (Chapter 9)
82. How can you tell if a chemical equation is balanced? (Chapter 10)
83. What mass of carbon must burn to produce 4.56 L CO_2 gas at STP? (Chapter 14)
 $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
84. When you reverse a thermochemical equation, why must you change the sign of ΔH ? (Chapter 16)

Use these questions and the test-taking tip to prepare for your standardized test.

- A system reaches chemical equilibrium when _____.
 - no new product is formed by the forward reaction
 - the reverse reaction no longer occurs in the system
 - the concentration of reactants in the system is equal to the concentration of products
 - the rate at which the forward reaction occurs equals the rate of the reverse reaction
- A value of K_{eq} greater than 1 means that _____.
 - more reactants than products exist at equilibrium
 - more products than reactants exist at equilibrium
 - the rate of the forward reaction is high at equilibrium
 - the rate of the reverse reaction is high at equilibrium
- The hydrogen sulfide produced as a byproduct of petroleum refinement can be used to produce elemental sulfur: $2\text{H}_2\text{S}(\text{g}) + \text{SO}_2(\text{g}) \rightarrow 3\text{S}(\text{l}) + 2\text{H}_2\text{O}(\text{g})$
The equilibrium constant expression for this reaction is _____.
 - $K_{eq} = \frac{[\text{H}_2\text{O}]}{[\text{H}_2\text{S}][\text{SO}_2]}$
 - $K_{eq} = \frac{[\text{H}_2\text{S}]^2[\text{SO}_2]}{[\text{H}_2\text{S}]^2}$
 - $K_{eq} = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$
 - $K_{eq} = \frac{[\text{S}]^3[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$
- The following system is in equilibrium:
 $2\text{S}(\text{s}) + 5\text{F}_2(\text{g}) \rightleftharpoons \text{SF}_4(\text{g}) + \text{SF}_6(\text{g})$
The equilibrium will shift to the right if _____.
 - the concentration of SF_4 is increased
 - the concentration of SF_6 is increased
 - the pressure on the system is increased
 - the pressure on the system is decreased

Interpreting Tables Use the table to answer questions 5–7.

- The K_{sp} for MnCO_3 is _____.
 - 2.24×10^{-11}
 - 4.00×10^{-11}
 - 1.12×10^{-9}
 - 5.60×10^{-9}
- What is the molar solubility of MnCO_3 at 298 K?
 - $4.73 \times 10^{-6}M$
 - $6.32 \times 10^{-2}M$
 - $7.48 \times 10^{-5}M$
 - $3.35 \times 10^{-5}M$
- A 50.0-mL volume of $3.00 \times 10^{-6}M$ K_2CO_3 is mixed with 50.0 mL of MnCl_2 . A precipitate of MnCO_3 will form only when the concentration of the MnCl_2 solution is greater than _____.
 - $7.47 \times 10^{-6}M$
 - $1.49 \times 10^{-5}M$
 - $2.99 \times 10^{-5}M$
 - $1.02 \times 10^{-5}M$

Concentration Data for the Equilibrium System
 $\text{MnCO}_3(\text{s}) \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$ at 298 K

Trial	$[\text{Mn}^{2+}]_0$ (M)	$[\text{CO}_3^{2-}]_0$ (M)	$[\text{Mn}^{2+}]_{eq}$ (M)	$[\text{CO}_3^{2-}]_{eq}$ (M)
1	0.0000	0.00400	5.60×10^{-9}	4.00×10^{-3}
2	0.0100	0.0000	1.00×10^{-2}	2.24×10^{-9}
3	0.0000	0.0200	1.12×10^{-9}	2.00×10^{-2}

- Which of the following statements about the common ion effect is NOT true?
 - The effects of common ions on an equilibrium system can be explained by Le Châtelier's principle.
 - The decreased solubility of an ionic compound due to the presence of a common ion is called the common ion effect.
 - The addition of NaCl to a saturated solution of AgCl will produce the common ion effect.
 - The common ion effect is due to a shift in equilibrium towards the aqueous products of a system.
- If the forward reaction of a system in equilibrium is endothermic, increasing the temperature of the system will _____.
 - shift the equilibrium to the left
 - shift the equilibrium to the right
 - decrease the rate of the forward reaction
 - increase the rate of the reverse reaction
- $\text{Cl}_2(\text{g}) + 3\text{O}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{ClO}_3\text{F}(\text{g})$
The formation of perchloryl fluoride (ClO_3F) from its elements has an equilibrium constant of 3.42×10^{-9} at 298 K. If $[\text{Cl}_2] = 0.563M$, $[\text{O}_2] = 1.01M$, and $[\text{ClO}_3\text{F}] = 1.47 \times 10^{-5}M$ at equilibrium, what is the concentration of F_2 ?
 - 9.18×10^0M
 - $3.73 \times 10^{-10}M$
 - $1.09 \times 10^{-1}M$
 - $6.32 \times 10^{-2}M$



TEST-TAKING TIP

Maximize Your Score If possible, find out how your standardized test will be scored. In order to do your best, you need to know if there is a penalty for guessing, and if so, what the penalty is. If there is no random-guessing penalty at all, you should always fill in an answer, even if you haven't read the question!

Acids and Bases

What You'll Learn

- ▶ You will compare acids and bases and understand why their strengths vary.
- ▶ You will define pH and pOH and calculate the pH and pOH of aqueous solutions.
- ▶ You will calculate acid and base concentrations and determine concentrations experimentally.
- ▶ You will explain how buffers resist changes in pH.

Why It's Important

Acids and bases are present in the soil of Earth, the foods you eat, the products you buy. Amino acids make up the fabric of every organ in your body and are crucial to your existence.

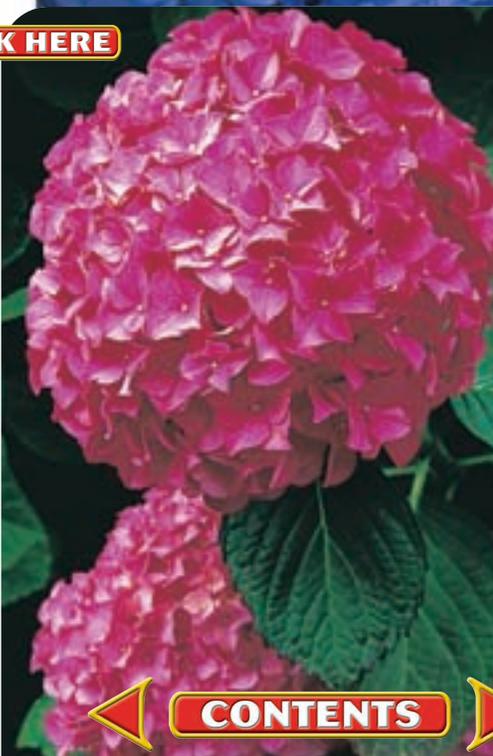


CLICK HERE

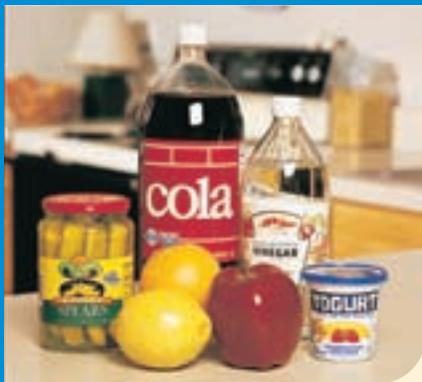


Visit the Chemistry Web site at science.glencoe.com to find links about acids and bases.

The color of the big-leaf hydrangea can vary from pink to blue depending upon the acidity of the soil in which it is grown.



DISCOVERY LAB



Materials

red litmus paper
blue litmus paper
microplate
household products (6-8)
phenolphthalein

Investigating What's in Your Cupboards

You can learn something about the properties of products in your household by testing them with strips of paper called litmus paper. Can you separate household products into two groups?

Safety Precautions



Always wear safety goggles and an apron.

Procedure

1. Place three or four drops of each liquid into separate wells of a microplate. Draw a chart to show the position of each liquid.
2. Test each product with litmus paper. Place 2 drops of phenolphthalein in each sample. Record your observations.

Analysis

Separate the products into two groups based upon your observations. How do the groups differ? What can you conclude?

Section

19.1

Acids and Bases: An Introduction

Objectives

- **Identify** the physical and chemical properties of acids and bases.
- **Classify** solutions as acidic, basic, or neutral.
- **Compare** the Arrhenius and Brønsted-Lowry models of acids and bases.

Vocabulary

acidic solution
basic solution
Arrhenius model
Brønsted-Lowry model
conjugate acid
conjugate base
conjugate acid-base pair
amphoteric

When ants sense danger to the ant colony, they emit a substance called formic acid that alerts the entire colony. Acids in rainwater hollow out enormous limestone caverns and destroy valuable buildings and statues. Acids flavor many of the beverages and foods you like, and it's an acid in your stomach that helps digest what you eat. Bases also play a role in your life. The soap you use and the antacid tablet you may take for an upset stomach are bases. Perhaps you have already concluded that the household products you used in the **DISCOVERY LAB** are acids and bases.

Properties of Acids and Bases

Acids and bases are some of the most important industrial compounds on Earth. In the U.S. alone, industries use 30 to 40 billion kilograms of sulfuric acid each year in the manufacture of products such as plastics, detergents, batteries, and metals.

You are probably already familiar with some of the physical properties of acids and bases. For example, you may know that acidic solutions taste sour. Carbonic and phosphoric acids give many carbonated beverages their sharp taste; citric and ascorbic acids give lemons and grapefruit their mouth-puckering tartness; and acetic acid makes vinegar taste sour. You may also know that basic solutions taste bitter and feel slippery. Just think about the bar of soap that slips from your hand in the shower. **CAUTION:** *You should never attempt to identify an acid or base (or any other substance in the laboratory) by its taste or feel.* The photo on the opposite page shows how the color of a hydrangea depends upon the presence of acids in the soil.

Figure 19-1

The blue litmus paper turned pink when it was dipped into the HCl solution, confirming that HCl is an acid. The red litmus paper turned blue after being dipped into the NaOH solution, so NaOH is a base.



Acids can also be identified by their reaction with some metals. Aluminum, magnesium, and zinc react with aqueous solutions of acids to produce hydrogen gas. The reaction between zinc and hydrochloric acid is described by the following equation.



Metal carbonates and hydrogen carbonates also react with aqueous solutions of acids to produce carbon dioxide gas. When vinegar is added to baking soda, a foaming, effervescent reaction occurs between acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) dissolved in the vinegar solution, and sodium hydrogen carbonate (NaHCO_3). The production of CO_2 gas accounts for the effervescence.



Geologists identify rocks as limestone (CaCO_3) by using an HCl solution. If a few drops of the acid produce bubbles of carbon dioxide, the rock is limestone.

Practice!

For more practice writing equations for acid reactions, go to **Supplemental Practice Problems** in Appendix A.

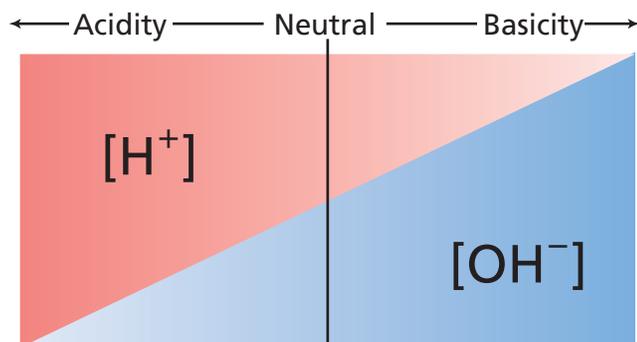
PRACTICE PROBLEMS

- Write a balanced formula equation for the reaction that occurs between each of the following pairs of reactants.
 - magnesium and nitric acid
 - aluminum and sulfuric acid
 - calcium carbonate and hydrobromic acid
 - potassium hydrogen carbonate and hydrochloric acid

The litmus in litmus paper is one of the dyes commonly used to distinguish solutions of acids and bases, as shown in **Figure 19-1**. Aqueous solutions of acids cause blue litmus paper to turn pink. Aqueous solutions of bases cause red litmus paper to turn blue. With this information you can now identify the two groups of household products you used in the **DISCOVERY LAB**.

Another property of acid and base solutions is their ability to conduct electricity. Pure water is a nonconductor of electricity, but the addition of an acid or base to water causes the resulting solution to become a conductor.

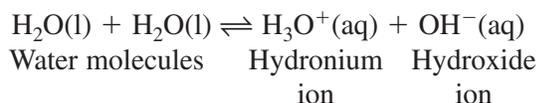
Ions in solution Why are some aqueous solutions acidic, others basic, and still others neutral? Neutral solutions are neither acidic nor basic. Scientists have learned that all water (aqueous) solutions contain hydrogen ions (H^+)



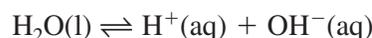
and hydroxide ions (OH^-). The relative amounts of the two ions determine whether an aqueous solution is acidic, basic, or neutral.

An **acidic solution** contains more hydrogen ions than hydroxide ions. A **basic solution** contains more hydroxide ions than hydrogen ions. What do you think a neutral solution contains? You are correct if you said a neutral solution contains equal concentrations of hydrogen ions and hydroxide ions. These relationships are illustrated in **Figure 19-2**.

The usual solvent for acids and bases is water. Water produces equal numbers of H^+ ions and OH^- ions in a process known as self-ionization. In self-ionization, two water molecules react to form a hydronium ion (H_3O^+) and a hydroxide ion according to this equilibrium.



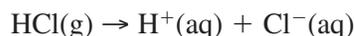
The hydronium ion is a hydrated hydrogen ion, which means that a water molecule is attached to a hydrogen ion by a covalent bond. However, the symbols H^+ and H_3O^+ can be used interchangeably in chemical equations to represent a hydrogen ion in aqueous solution. Thus, a simplified version of the equation for the self-ionization of water is



From the equation, you can infer that pure water is neutral because equal numbers of H^+ ions and OH^- ions are always present.

The Arrhenius model of acids and bases If pure water itself is neutral, how does an aqueous solution become acidic or basic? The first person to answer this question was the Swedish chemist Svante Arrhenius, who in 1883 proposed what is now called the Arrhenius model of acids and bases. The **Arrhenius model** states that an acid is a substance that contains hydrogen and ionizes to produce hydrogen ions in aqueous solution. A base is a substance that contains a hydroxide group and dissociates to produce a hydroxide ion in aqueous solution. Some household acids and bases are shown in **Figure 19-3**.

As an example of the Arrhenius model of acids and bases, consider what happens when hydrogen chloride gas dissolves in water. HCl molecules ionize to form H^+ ions, which make the solution acidic.



When the ionic compound NaOH dissolves in water, it dissociates to produce OH^- ions, which make the solution basic.

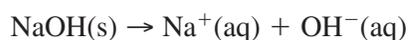


Figure 19-2

Note how $[\text{H}^+]$ and $[\text{OH}^-]$ change simultaneously. As $[\text{H}^+]$ decreases to the right, $[\text{OH}^-]$ increases to the right. At what point in the diagram are the two ion concentrations equal?



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.



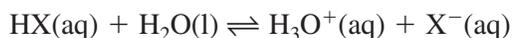
Figure 19-3

You probably have some of these products in your home. Examine their labels and the labels of other household products and list the acids and bases you find.

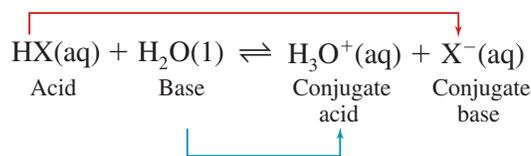
Although the Arrhenius model is useful in explaining many acidic and basic solutions, it has some shortcomings. For example, ammonia (NH₃) does not contain a hydroxide group, yet ammonia produces hydroxide ions in solution and is a well known base. Clearly, a model that includes all bases is needed.

The Brønsted-Lowry model The Danish chemist Johannes Brønsted and the English chemist Thomas Lowry independently proposed a more inclusive model of acids and bases—a model that focuses on the hydrogen ion (H⁺). In the **Brønsted-Lowry model** of acids and bases, an acid is a hydrogen-ion donor and a base is a hydrogen-ion acceptor.

What does it mean to be a hydrogen-ion donor or a hydrogen-ion acceptor? The symbols X and Y may be used to represent nonmetallic elements or negative polyatomic ions. Thus the general formula for an acid can be written as HX or HY. When a molecule of acid, HX, dissolves in water, it donates a H⁺ ion to a water molecule. The water molecule acts as a base and accepts the H⁺ ion.



On accepting the H⁺ ion, the water molecule becomes an acid H₃O⁺. The hydronium ion (H₃O⁺) is an acid because it has an extra H⁺ ion that it can donate. On donating its H⁺ ion, the acid HX becomes a base, X⁻. Why? You are correct if you said X⁻ is a base because it has a negative charge and can readily accept a positive hydrogen ion. Thus, an acid-base reaction in the reverse direction can occur. The acid H₃O⁺ can react with the base X⁻ to form water and HX and the following equilibrium is established.



The forward reaction is the reaction of an acid and a base. The reverse reaction also is the reaction of an acid and a base. The acid and base that react in the reverse reaction are identified under the equation as a conjugate acid and a conjugate base. A **conjugate acid** is the species produced when a base accepts a hydrogen ion from an acid. The base H₂O accepts a hydrogen ion from the acid HX and becomes the conjugate acid, H₃O⁺. A **conjugate base** is the species that results when an acid donates a hydrogen ion to a base. The acid HX donates its hydrogen ion and becomes the conjugate base X⁻. In the reaction above, the hydronium ion (H₃O⁺) is the conjugate acid of the base H₂O. The X⁻ ion is the conjugate base of the acid HX. Every Brønsted-Lowry interaction involves conjugate acid-base pairs. A **conjugate acid-base pair** consists of two substances related to each other by the donating and accepting of a single hydrogen ion.

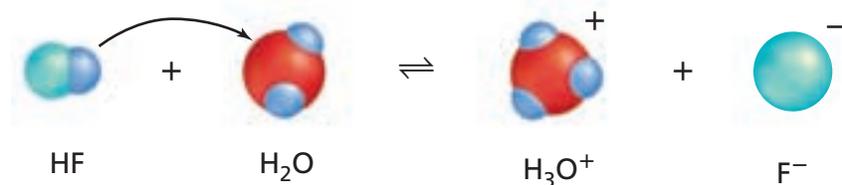
An analogy for conjugate acid-base pairs is shown in **Figure 19-4**. When you have the footbag (Hacky Sack), you are an acid. You pass the footbag (hydrogen ion) to your friend. Now your friend is an acid because she has the footbag (hydrogen ion) to give away and you are a base because you are able to accept the footbag (hydrogen ion). You, with the footbag, and your friend are the acid and base in the forward reaction. In the reverse reaction, your friend with the footbag is the conjugate acid and you are the conjugate base.



Figure 19-4

In a game of freestyle footbag, the players stand in a circle. If the footbag is a hydrogen ion, which person is an acid? If the footbag is passed to another person, which person is the acid? Is the other person a conjugate base or a conjugate acid?

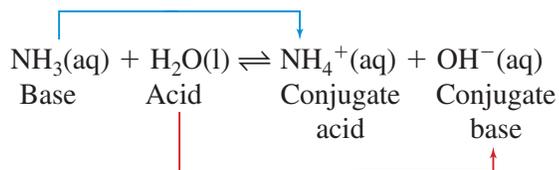
Now, consider the equation for the ionization of hydrogen fluoride in water. According to the Brønsted-Lowry definition, the equation is written this way.



What are the conjugate acid-base pairs? Hydrogen fluoride (HF), the acid in the forward reaction, produces its conjugate base F^- , the base in the reverse reaction. Water, the base in the forward reaction, produces its conjugate acid H_3O^+ , the acid in the reverse reaction.

Hydrogen fluoride is an acid according to both the Arrhenius and Brønsted-Lowry definitions. All of the acids and bases that fit the Arrhenius definition of acids and bases also fit the Brønsted-Lowry definition. But what about bases such as ammonia, that cannot be considered bases according to the Arrhenius definition because they lack a hydroxide group? Does the Brønsted-Lowry model explain why they are bases?

When ammonia dissolves in water, water is a Brønsted-Lowry acid in the forward reaction. Because the NH_3 molecule accepts a H^+ ion to form the ammonium ion (NH_4^+), ammonia is a Brønsted-Lowry base in the forward reaction.



Now look at the reverse reaction. The ammonium ion gives up a H^+ ion to form the molecule ammonia and thus acts as a Brønsted-Lowry acid. The ammonium ion is the conjugate acid of the base ammonia. The hydroxide ion accepts a H^+ ion to form a water molecule and is thus a Brønsted-Lowry base. The hydroxide ion is the conjugate base of the acid water.

Recall that when HF dissolves in water, water acts a base; when NH_3 dissolves in water, water acts as an acid. Depending upon what other substances are in the solution, water can act as either an acid or a base. Water and other substances that can act as both acids and bases are said to be **amphoteric**.

Compare what you have learned about the Arrhenius model and the Brønsted-Lowry model of acids and bases. It should be clear to you that all substances classified as acids and bases by the Arrhenius model are classified as acids and bases by the Brønsted-Lowry model. In addition, some substances *not* classified as bases by the Arrhenius model *are* classified as bases by the Brønsted-Lowry model.

PRACTICE PROBLEMS

2. Identify the conjugate acid-base pairs in the following reactions.
- $\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{HBr}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - $\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

Careers Using Chemistry

Agricultural Technician

Would you like to work on a large farm, checking soil pH and deciding when to use fertilizers and pesticides? If so, become an agricultural technician.

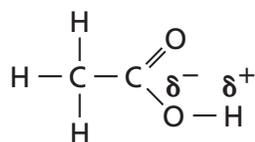
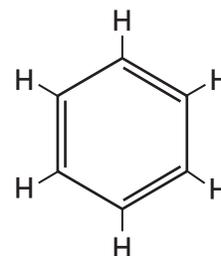
Agricultural technicians also work for the government, research groups, and companies that produce fertilizers and other farm chemicals. They know how to modify soil conditions, increase crop yields, help crops resist disease and insects, and store produce. Some agricultural technicians help keep farm animals healthy and productive.



For more practice identifying conjugate acid-base pairs, go to **Supplemental Practice Problems** in Appendix A.

Figure 19-5

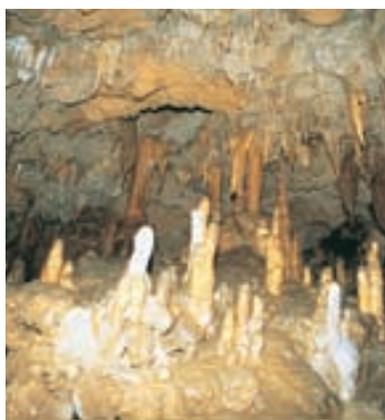
A partial separation of charge is shown in **a** for the O—H bond and in **b** for the H—F bond. These bonds are polar. In water solution, the H⁺ ionizes, so solutions of HF and ethanoic acid are acidic. The benzene molecule in **c** contains no polar bonds so benzene is not an acid.

**a** Ethanoic acid**b** Hydrogen fluoride**c** Benzene

Earth Science

CONNECTION

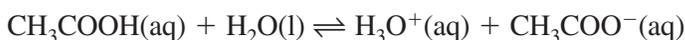
As rain falls, carbon dioxide gas in the air dissolves in the falling water to form a weak acid called carbonic acid. When the acidic rainwater reaches the ground, some water sinks into the pores of the soil to become groundwater. If that groundwater reaches bedrock consisting of limestone, the carbonic acid will slowly dissolve the limestone. Over the course of thousands of years, the limestone is dissolved to the point that huge underground tunnels known as caverns are produced. In many places throughout the cavern, groundwater may drip from the ceiling. As it does so, it deposits some of the dissolved limestone. Thin deposits shaped like icicles on the ceiling are called stalactites. Rounded masses on the floor are called stalagmites. In some cases, stalactites and stalagmites eventually meet in a column or pillar.



Monoprotic and Polyprotic Acids

You now know that HCl and HF are acids because they can donate a hydrogen ion in an acid-base reaction. From their chemical formulas, you can see that each acid can donate only one hydrogen ion per molecule. An acid that can donate only one hydrogen ion is called a monoprotic acid. Other monoprotic acids are perchloric acid (HClO₄), nitric acid (HNO₃), hydrobromic acid (HBr), and acetic acid (CH₃COOH). The formula for acetic acid is sometimes written HC₂H₃O₂ and the compound is often called ethanoic acid.

In **Figure 19-5a**, you can see that each ethanoic acid molecule contains four hydrogen atoms. Can ethanoic acid donate more than one hydrogen ion? No; each CH₃COOH molecule contains only one ionizable hydrogen atom.

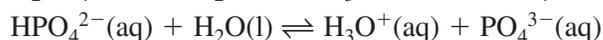
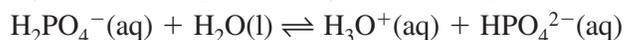
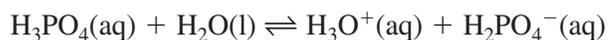


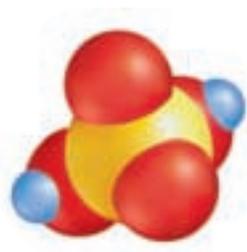
Only one of the four hydrogen atoms in the CH₃COOH molecule can be donated because only those hydrogen atoms bonded to electronegative elements by polar bonds are ionizable.

In an HF molecule, the hydrogen atom is bonded to a fluorine atom, which has the highest electronegativity of all the elements. In **Figure 19-5b**, you can see that the bond linking hydrogen and fluorine is polar; in solution, water facilitates the release of the H⁺ ion. In the CH₃COOH molecule shown in **Figure 19-5a**, three of the four hydrogen atoms are joined to a carbon atom by covalent bonds that are nonpolar because carbon and hydrogen have almost equal electronegativities. Only the hydrogen atom bonded to the electronegative oxygen atom can be released as a H⁺ ion in water. Although a molecule of benzene (C₆H₆) contains six hydrogen atoms, it is not an acid at all. As **Figure 19-5c** shows, none of the hydrogen atoms are ionizable because they are all joined to carbon atoms by nonpolar bonds.

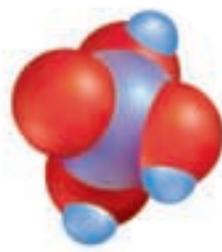
Some acids, however, do donate more than one hydrogen ion. For example, sulfuric acid (H₂SO₄) and carbonic acid (H₂CO₃) can donate two hydrogen ions. In each compound, both hydrogen atoms are attached to oxygen atoms by polar bonds. Acids that contain two ionizable hydrogen atoms per molecule are called diprotic acids. In a similar way, phosphoric acid (H₃PO₄) and boric acid (H₃BO₃) contain three ionizable hydrogen atoms per molecule. Acids with three hydrogen ions to donate are called triprotic acids. The term polyprotic acid can be used for any acid that has more than one ionizable hydrogen atom. **Figure 19-6** shows models of two polyprotic acids.

All polyprotic acids ionize in steps. The three ionizations of phosphoric acid are described by these equations.





H_2SO_4
Sulfuric acid



H_3PO_4
Phosphoric acid

Figure 19-6

Refer to **Table C-1** in Appendix C for a key to atom color conventions. Then, identify the two hydrogen atoms bonded to oxygen atoms in the sulfuric acid model and the three hydrogen atoms bonded to oxygen atoms in the phosphoric acid model.

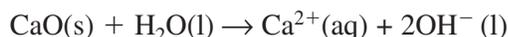
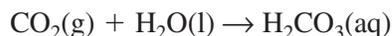
PRACTICE PROBLEMS

3. Write the steps in the complete ionization of the following polyprotic acids.
- H_2Se
 - H_3AsO_4
 - H_2SO_3



For more practice writing ionization equations for polyprotic acids, go to **Supplemental Practice Problems** in Appendix A.

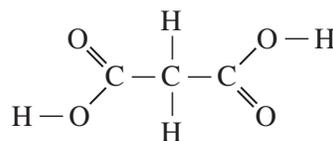
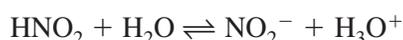
Anhydrides Some oxides can become acids or bases by adding the elements contained in water. These compounds are called anhydrides. Oxides of nonmetallic elements, such as carbon, sulfur, or nitrogen, produce an acid in aqueous solution. Oxides of metallic elements usually form basic solutions. For example, carbon dioxide, an oxide of a nonmetal, forms a solution of carbonic acid, whereas, calcium oxide (CaO) forms a basic solution of calcium hydroxide.



Similarly, SO_3 is the anhydride of H_2SO_4 and MgO is the anhydride of $\text{Mg}(\text{OH})_2$. What acid will form from N_2O_5 ?

Section 19.1 Assessment

- Compare the properties of acidic solutions and basic solutions.
- How do the concentrations of hydrogen ion and hydroxide ion determine whether a solution is acidic, basic, or neutral?
- Based on their formulas, which of the following compounds *may* be Arrhenius acids: CH_4 , SO_2 , H_2S , $\text{Ca}_3(\text{PO}_4)_2$? Explain your reasoning.
- Identify the conjugate acid-base pairs in the following equation.
- Thinking Critically** Methylamine (CH_3NH_2) forms hydroxide ions in aqueous solution. Why is methylamine a Brønsted-Lowry base but not an Arrhenius base?
- Interpreting Scientific Illustrations** In the accompanying structural formula, identify any hydrogen atoms that are likely to be ionizable.



Objectives

- **Relate** the strength of an acid or base to its degree of ionization.
- **Compare** the strength of a weak acid with the strength of its conjugate base and the strength of a weak base with the strength of its conjugate acid.
- **Explain** the relationship between the strengths of acids and bases and the values of their ionization constants.

Vocabulary

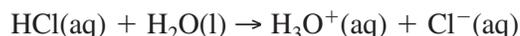
strong acid
 weak acid
 acid ionization constant
 strong base
 weak base
 base ionization constant

In the previous section, you learned that one of the properties of acidic and basic solutions is that they conduct electricity. What can electrical conductivity tell you about the hydrogen ions and hydroxide ions in these aqueous solutions?

Strengths of Acids

To answer this question, suppose you test the electrical conductivities of 0.10M aqueous solutions of hydrochloric acid and acetic acid using a conductivity apparatus similar to the one shown in **Figure 19-7**. When the electrodes are placed in the solutions, both bulbs glow, indicating that both solutions conduct electricity. However, if you compare the brightness of the bulb connected to the hydrochloric acid solution with that of the bulb connected to the acetic acid solution, you can't help but notice a significant difference. The 0.10M HCl solution conducts electricity better than the 0.10M HC₂H₃O₂ solution. Why is this true if the concentrations of the two acids are both 0.10M?

The answer is that it is ions that carry electricity through the solution and all the HCl molecules that make up the solution are in the form of hydrogen ions and chloride ions. Acids that ionize completely are called **strong acids**. Because strong acids produce the maximum number of ions, they are good conductors of electricity. Other strong acids include perchloric acid (HClO₄), nitric acid (HNO₃), hydroiodic acid (HI), and sulfuric acid (H₂SO₄). The ionization of hydrochloric acid in water may be represented by the following equation, which has a single arrow pointing to the right. What does a single arrow to the right mean?



Because hydrochloric acid is virtually 100% ionized, you can consider that the reaction goes to completion and essentially no reaction occurs in the reverse direction.

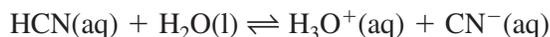
If the brightly lit bulb of the hydrochloric acid apparatus is due to the large number of ions in solution, then the weakly lit bulb of the acetic acid apparatus must mean that the acetic acid solution has fewer ions. Because the two solutions have the same molar concentrations, you can conclude that acetic acid

Figure 19-7

Both solutions have the same molar concentration, yet HCl produces a brighter light than HC₂H₃O₂. Because the amount of electrical current depends upon the number of ions in solution, the HCl solution must contain more ions.



Acid ionization constants As you have learned, a weak acid produces an equilibrium mixture of molecules and ions in aqueous solution. Thus, the equilibrium constant, K_{eq} , provides a quantitative measure of the degree of ionization of the acid. Consider hydrocyanic acid (HCN), a deadly poison with applications in the steel industry and in the processing of metal ores. See **Figure 19-9**. The ionization equation and equilibrium constant expression for hydrocyanic acid are



$$K_{\text{eq}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}][\text{H}_2\text{O}]}$$

The concentration of liquid water in the denominator of the expression is considered to be constant in dilute aqueous solutions, so it can be combined with K_{eq} to give a new equilibrium constant, K_{a} .

$$K_{\text{eq}} [\text{H}_2\text{O}] = K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10}$$

K_{a} is called the acid ionization constant. The **acid ionization constant** is the value of the equilibrium constant expression for the ionization of a weak acid. Like all equilibrium constants, the value of K_{a} indicates whether reactants or products are favored at equilibrium. For weak acids, the concentrations of the ions (products) in the numerator tend to be small compared to the concentration of un-ionized molecules (reactant) in the denominator. The weakest acids have the smallest K_{a} values because their solutions have the lowest concentrations of ions and the highest concentrations of un-ionized acid molecules. K_{a} values and ionization equations for several weak acids are listed in **Table 19-2**. Note that for polyprotic acids there is a K_{a} value for each ionization and the values decrease for each successive ionization.

Table 19-2

Ionization Constants for Weak Acids at 25°C		
Acid	Ionization equation	K_{a} (298 K)
Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$	8.9×10^{-8}
	$\text{HS}^- \rightleftharpoons \text{H}^+ + \text{S}^{2-}$	1×10^{-19}
Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$	6.3×10^{-4}
Methanoic (Formic)	$\text{HCOOH} \rightleftharpoons \text{H}^+ + \text{HCOO}^-$	1.8×10^{-4}
Ethanoic (Acetic)	$\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$	1.8×10^{-5}
Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$	4.5×10^{-7}
	$\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$	4.7×10^{-11}
Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$	4.0×10^{-8}

PRACTICE PROBLEMS

10. Write ionization equations and acid ionization constant expressions for the following acids.

- HClO_2
- HNO_2
- HIO

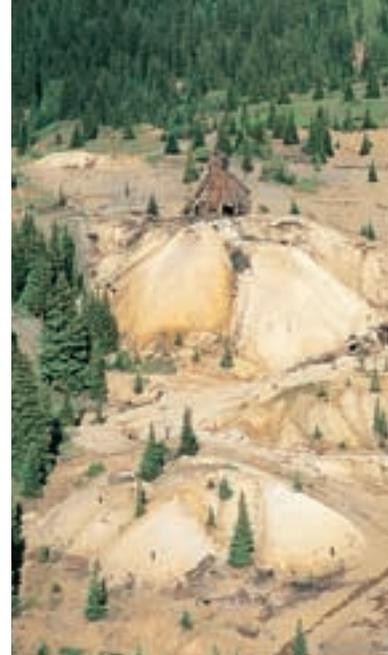


Figure 19-9

Mine tailings are crushed rock discarded as waste. Hydrocyanic acid and its compounds can be used to extract useable materials from mine tailings.



For more practice writing acid ionization constant expressions, go to **Supplemental Practice Problems** in Appendix A.

Table 19-4

Ionization Constants of Weak Bases		
Base	Ionization equation	K_b (298 K)
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	5.0×10^{-4}
Methylamine	$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	4.3×10^{-4}
Ammonia	$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	2.5×10^{-5}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_6\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	4.3×10^{-10}

Figure 19-10

The photos show a strong acid and a weak acid, a strong base and a weak base. Can you distinguish them? Which of the acids is more concentrated? Which is stronger? Which of the bases is more concentrated? Which is stronger?

Strong or weak, concentrated or dilute You have been learning about acids and bases that are often described as weak or strong. When you use the words, weak and strong, do they mean the same as when you describe your cup of tea as being weak or strong? No, there is a difference. When talking about your tea, you could substitute the words *dilute* and *concentrated* for weak and strong. But in talking about acids and bases, the words *weak* and *dilute* have different meanings. Similarly, the words *strong* and *concentrated* are not interchangeable. The terms *dilute* and *concentrated* refer to the number of the acid or base molecules dissolved in a volume of solution. The molarity of the solution is a measure of how dilute or concentrated a solution is. The words *weak* or *strong* refer to the degree to which the acid or base separates into ions. You have already learned that solutions of weak acids and bases contain few ions because few of the molecules are ionized. Solutions of strong acids and bases are completely separated into ions. It's possible to have a dilute solution of a strong acid such as hydrochloric acid, or a concentrated solution of a weak acid such as acetic acid. Which is strong, which is more concentrated, a solution of 0.6M hydrochloric acid or a solution of 6M acetic acid? See Figure 19-10.



Section 19.2 Assessment

- An acid is highly ionized in aqueous solution. Is the acid strong or weak? Explain your reasoning.
- How is the strength of a weak acid related to the strength of its conjugate base?
- Identify the acid-base pairs in the following.
 - $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
 - $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
- K_b for aniline is 4.3×10^{-10} . Explain what this tells you about aniline.
- Thinking Critically** Why is a strong base such as sodium hydroxide generally not considered to have a conjugate acid?
- Predicting** Use Table 19-2 to predict which aqueous solution would have the greater electrical conductivity: 0.1M HClO or 0.1M HF. Explain.

Objectives

- **Explain** the meaning of pH and pOH.
- **Relate** pH and pOH to the ion product constant for water.
- **Calculate** the pH and pOH of aqueous solutions.

Vocabulary

ion product constant for water

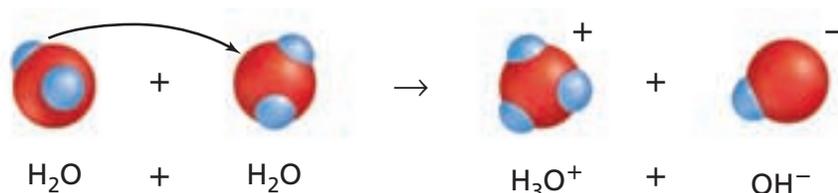
pH

pOH

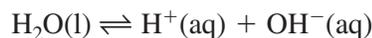
Water not only serves as the solvent in solutions of acids and bases, it also plays a role in the formation of the ions. In aqueous solutions of acids and bases, water sometimes acts as an acid and sometimes as a base. You can think of the self-ionization of water as an example of water assuming the role of an acid and a base in the same reaction.

Ion Product Constant for Water

Recall from Section 19.1 that pure water contains equal concentrations of H^+ and OH^- ions produced by self-ionization. One molecule of water acts as a Brønsted-Lowry acid and donates a hydrogen ion to a second water molecule. The second molecule of water accepts the hydrogen ion and becomes a hydronium ion. The 1:1 ratio between the products means that equal numbers of hydronium ions and hydroxide ions are formed.



The equation for the equilibrium can be simplified in this way.



The double arrow indicates that this is an equilibrium. Recall that the equilibrium constant expression is written by placing the concentrations of the products in the numerator and the concentrations of the reactants in the denominator. In this example, all terms are to the first power because all the coefficients are 1.

$$K_{\text{eq}} = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

The concentration of pure water is constant so it can be combined with K_{eq} by multiplying both sides of the equation by $[\text{H}_2\text{O}]$.

$$K_{\text{eq}}[\text{H}_2\text{O}] = K_{\text{w}} = [\text{H}^+][\text{OH}^-]$$

The result is a special equilibrium constant expression that applies only to the self-ionization of water. The constant, K_{w} , is called the ion product constant for water. The **ion product constant for water** is the value of the equilibrium constant expression for the self-ionization of water. Experiments show that in pure water at 298 K, $[\text{H}^+]$ and $[\text{OH}^-]$ are both equal to $1.0 \times 10^{-7} M$. Therefore, at 298 K, the value of K_{w} is 1.0×10^{-14} .

$$K_{\text{w}} = [\text{H}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7})$$

$$K_{\text{w}} = 1.0 \times 10^{-14}$$

The product of $[H^+]$ and $[OH^-]$ always equals 1.0×10^{-14} at 298 K. This means that if the concentration of H^+ ion increases, the concentration of OH^- ion must decrease. Similarly, an increase in the concentration of OH^- ion causes a decrease in the concentration of H^+ ion. You can think about these changes in terms of Le Châtelier's principle, which you learned about in Chapter 18. Adding extra hydrogen ions to the self-ionization of water at equilibrium is a stress on the system. The system reacts in a way to relieve the stress. The added H^+ ions react with OH^- ions to form more water molecules. Thus, the concentration of OH^- ion decreases. Example Problem 19-1 shows how you can use K_w to calculate the concentration of either the hydrogen ion or the hydroxide ion if you know the concentration of the other ion.

EXAMPLE PROBLEM 19-1

Using K_w to Calculate $[H^+]$ and $[OH^-]$

At 298 K, the H^+ ion concentration of an aqueous solution is $1.0 \times 10^{-5}M$. What is the OH^- ion concentration in the solution? Is the solution acidic, basic, or neutral?

1. Analyze the Problem

You are given the concentration of H^+ ion and you know that K_w equals 1.0×10^{-14} . You can use the ion product constant expression to solve for $[OH^-]$. Because $[H^+]$ is greater than 1.0×10^{-7} , you can predict that $[OH^-]$ will be less than 1.0×10^{-7} .

Known

$$[H^+] = 1.0 \times 10^{-5}M$$

$$K_w = 1.0 \times 10^{-14}$$

Unknown

$$[OH^-] = ? \text{ mol/L}$$

2. Solve for the Unknown

Write the ion product constant expression.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

Isolate $[OH^-]$ by dividing both sides of the equation by $[H^+]$.

$$[OH^-] = \frac{K_w}{[H^+]}$$

Substitute K_w and $[H^+]$ into the expression and solve.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} \text{ mol/L}$$

Because $[H^+] > [OH^-]$, the solution is acidic.

3. Evaluate the Answer

The answer is correctly stated with two significant figures because $[H^+]$ and K_w each have two. As predicted, the hydroxide ion concentration, $[OH^-]$, is less than $1.0 \times 10^{-7} \text{ mol/L}$.



The hydrogen ion and hydroxide ion concentrations of these familiar vegetables are the same as the concentrations calculated in this Example Problem.

PRACTICE PROBLEMS

18. The concentration of either the H^+ ion or the OH^- ion is given for three aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.
- $[H^+] = 1.0 \times 10^{-13}M$
 - $[OH^-] = 1.0 \times 10^{-7}M$
 - $[OH^-] = 1.0 \times 10^{-3}M$



For more practice calculating $[H^+]$ and $[OH^-]$ from K_w , go to **Supplemental Practice Problems** in Appendix A.

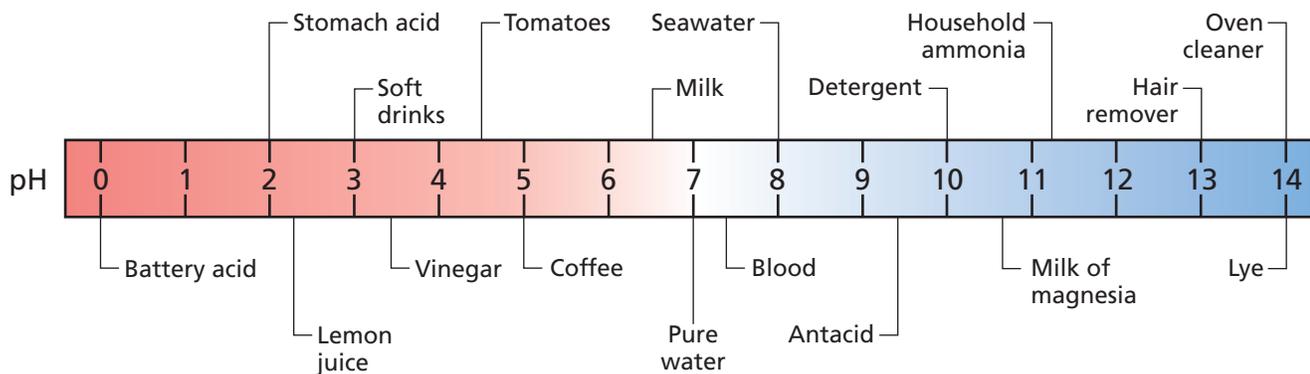


Figure 19-11

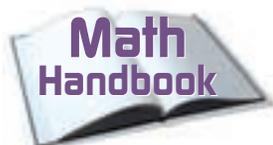
Which has the higher concentration of hydrogen ion, sea water or detergent? How many times higher?

pH and pOH

As you have probably noted, concentrations of H^+ ions are often small numbers expressed in exponential notation. Because these numbers are cumbersome, chemists adopted an easier way to express H^+ ion concentrations using a pH scale based on common logarithms. The **pH** of a solution is the negative logarithm of the hydrogen ion concentration.

$$\text{pH} = -\log [\text{H}^+]$$

At 298 K, the values of pH range from 0 to 14. A solution having a pH of 0.0 is strongly acidic; a solution having a pH of 14.0 is strongly basic; and a solution with $\text{pH} = 7.0$ is neutral. Thus, acidic solutions have pH values between 0 and 7. Basic solutions have pH values between 7 and 14. The logarithmic nature of the pH scale means that a change of one pH unit represents a tenfold change in ion concentration. A solution having a pH of 3.0 has ten times the hydrogen ion concentration of a solution with a pH of 4.0. The pH scale and pH values of some common substances are shown in **Figure 19-11**.



Review logarithms in the **Math Handbook** on page 910 of this text.

EXAMPLE PROBLEM 19-2

Calculating pH from $[\text{H}^+]$

What is the pH of a neutral solution at 298 K?

1. Analyze the Problem

You know that in a neutral solution at 298 K, $[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$. You need to find the negative log of $[\text{H}^+]$. Because the solution is neutral, you can predict that the pH will be 7.00. You will need a log table or a calculator with a log function.

Known

$$[\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$$

Unknown

$$\text{pH} = ?$$

2. Solve for the Unknown

$$\text{pH} = -\log [\text{H}^+]$$

Substitute $1.0 \times 10^{-7} \text{ M}$ for $[\text{H}^+]$ in the equation.

$$\text{pH} = -\log (1.0 \times 10^{-7})$$

$$\text{pH} = -(\log 1.0 + \log 10^{-7})$$

A log table or calculator shows that $\log 1.0 = 0$ and $\log 10^{-7} = -7$. Substitute these numbers in the pH equation.

$$\text{pH} = -[0 + (-7)] = 7.00$$

The pH of the neutral solution at 298 K is 7.00.

3. Evaluate the Answer

Values for pH are expressed with as many decimal places as the number of significant figures in the H^+ ion concentration. Thus, the pH is correctly stated with two decimal places. As predicted, the pH value is 7.00.

PRACTICE PROBLEMS

19. Calculate the pH of solutions having the following ion concentrations at 298 K.
- $[H^+] = 1.0 \times 10^{-2} M$
 - $[H^+] = 3.0 \times 10^{-6} M$
 - $[OH^-] = 8.2 \times 10^{-6} M$



For more practice calculating pH, go to **Supplemental Practice Problems** in Appendix A.

Using pOH Sometimes chemists find it convenient to express the basicity, or alkalinity, of a solution on a pOH scale that mirrors the relationship between pH and $[H^+]$. The **pOH** of a solution is the negative logarithm of the hydroxide ion concentration.

$$pOH = -\log [OH^-]$$

At 298 K, a solution having a pOH less than 7.0 is basic; a solution having a pOH of 7.0 is neutral; and a solution having a pOH greater than 7.0 is acidic. As with the pH scale, a change of one pOH unit expresses a tenfold change in ion concentration. For example, a solution with a pOH of 2.0 has 100 times the hydroxide ion concentration of a solution with a pOH of 4.0.

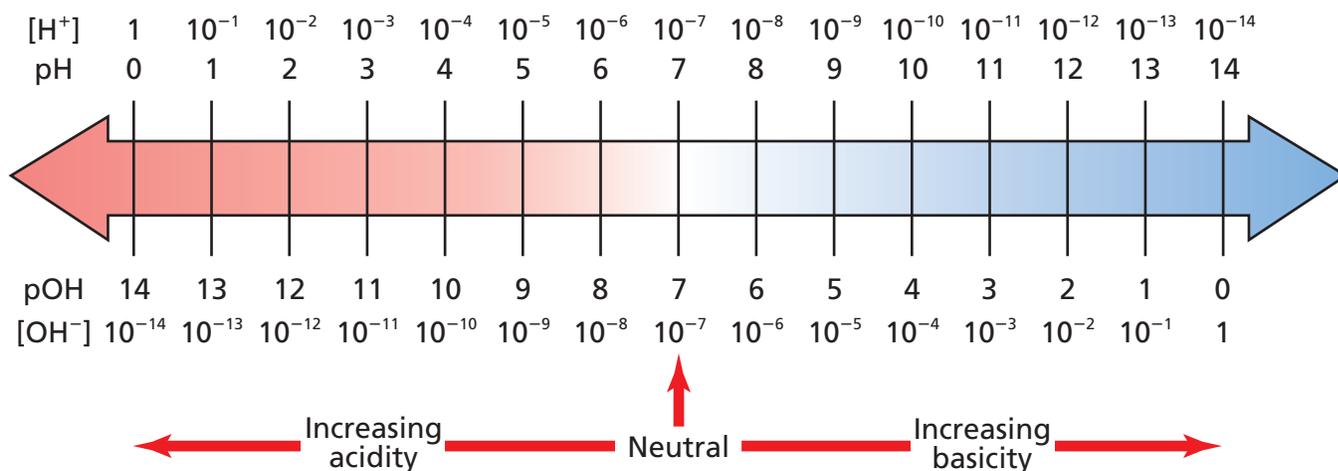
A simple relationship between pH and pOH makes it easy to calculate either quantity if the other is known.

$$pH + pOH = 14.00$$

Figure 19-12 illustrates the relationship between pH and $[H^+]$ and the relationship between pOH and $[OH^-]$ at 298 K. Use this diagram as a reference until you become thoroughly familiar with these relationships.

Figure 19-12

Study this diagram to sharpen your understanding of pH and pOH. Note that at each vertical position, the sum of pH (above the arrow) and pOH (below the arrow) equals 14. Also note that at every position the product of $[H^+]$ and $[OH^-]$ equals 10^{-14} .



EXAMPLE PROBLEM 19-3

Calculating pOH and pH from $[\text{OH}^-]$

An ordinary household ammonia cleaner is an aqueous solution of ammonia gas with a hydroxide-ion concentration of $4.0 \times 10^{-3}M$. Calculate pOH and pH of a typical cleaner at 298 K.

1. Analyze the Problem

You have been given the concentration of hydroxide ion and must calculate pOH and pH. First, you must calculate pOH using the definition of pOH. Then, pH can be calculated using the relationship $\text{pH} + \text{pOH} = 14.00$. The negative log of 10^{-3} (the power of 10 of the hydroxide ion concentration) is 3. Therefore, pOH should be close to 3 and pH should be close to 11 or 12.

Known

$$[\text{OH}^-] = 4.0 \times 10^{-3}M$$

Unknown

$$\text{pOH} = ?$$

$$\text{pH} = ?$$

2. Solve for the Unknown

$$\text{pOH} = -\log [\text{OH}^-]$$

Substitute $4.0 \times 10^{-3}M$ for $[\text{OH}^-]$ in the equation.

$$\text{pOH} = -\log (4.0 \times 10^{-3})$$

$$\text{pOH} = -(\log 4.0 + \log 10^{-3})$$

Log tables or your calculator indicate $\log 4.0 = 0.60$ and $\log 10^{-3} = -3$. Substitute these values in the equation.

$$\text{pOH} = -[0.60 + (-3)] = -(0.60 - 3) = 2.40$$

The pOH of the solution is 2.40.

Solve the equation $\text{pH} + \text{pOH} = 14.00$ for pH by subtracting pOH from both sides of the equation.

$$\text{pH} = 14.00 - \text{pOH}$$

Substitute the value of pOH.

$$\text{pH} = 14.00 - 2.40 = 11.60$$

The pH of the solution is 11.60.

3. Evaluate the Answer

The values of pH and pOH are correctly expressed with two decimal places because the given concentration has two significant figures. As predicted, pOH is close to 3 and pH is close to 12.



The pH paper shows that this common cleaner containing ammonia has a pH greater than 11.

PRACTICE PROBLEMS

20. Calculate the pH and pOH of aqueous solutions having the following ion concentrations.

a. $[\text{OH}^-] = 1.0 \times 10^{-6}M$

b. $[\text{OH}^-] = 6.5 \times 10^{-4}M$

c. $[\text{H}^+] = 3.6 \times 10^{-9}M$

d. $[\text{H}^+] = 0.025M$

Calculating ion concentrations from pH Suppose the pH of a solution is 3.50 and you must determine the concentrations of H^+ and OH^- . The definition of pH relates pH and H^+ ion concentration and can be solved for $[\text{H}^+]$.

$$\text{pH} = -\log [\text{H}^+]$$

Practice!

For more practice calculating pH and pOH from $[\text{OH}^-]$, go to **Supplemental Practice Problems** in Appendix A.

First you need to multiply both sides of the equation by -1 .

$$-\text{pH} = \log [\text{H}^+]$$

To calculate $[\text{H}^+]$ using this equation you must take the antilog of both sides of the equation.

$$\text{antilog} (-\text{pH}) = [\text{H}^+]$$

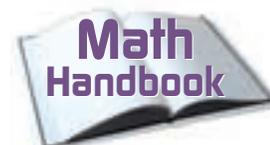
To calculate $[\text{H}^+]$, substitute 3.50 for pH in the equation.

$$\text{antilog} (-3.50) = [\text{H}^+]$$

Use a log table or your calculator to determine the antilog of -3.50 . The antilog is 3.2×10^{-4} .

$$[\text{H}^+] = 3.2 \times 10^{-4} \text{ mol/L}$$

You can calculate $[\text{OH}^-]$ using the relationship $[\text{OH}^-] = \text{antilog} (-\text{pOH})$.



Review antilogs in the **Math Handbook** on page 910 of this text.

EXAMPLE PROBLEM 19-4

Calculating $[\text{H}^+]$ and $[\text{OH}^-]$ from pH

What are $[\text{H}^+]$ and $[\text{OH}^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature is 298 K.

1. Analyze the Problem

You have been given the pH of a solution and must calculate $[\text{H}^+]$ and $[\text{OH}^-]$. You can obtain $[\text{H}^+]$ using the equation that defines pH. Then, subtract the pH from 14.00 to obtain pOH. The pH is close to 7 but greater than 7, so $[\text{H}^+]$ should be slightly less than 10^{-7} and $[\text{OH}^-]$ should be greater than 10^{-7} .

Known

$$\text{pH} = 7.40$$

Unknown

$$[\text{H}^+] = ? \text{ mol/L}$$

$$[\text{OH}^-] = ? \text{ mol/L}$$

2. Solve for the Unknown

Write the equation that defines pH and solve for $[\text{H}^+]$.

$$\text{pH} = -\log [\text{H}^+]$$

$$[\text{H}^+] = \text{antilog} (-\text{pH})$$

Substitute the known value of pH.

$$[\text{H}^+] = \text{antilog} (-7.40)$$

Use a log table or your calculator to find the antilog. The antilog of -7.40 is 4.0×10^{-8} .

$$[\text{H}^+] = 4.0 \times 10^{-8} \text{ M}$$

The concentration of hydrogen ion in blood is $4.0 \times 10^{-8} \text{ M}$.

To determine $[\text{OH}^-]$, calculate pOH using the equation

$$\text{pH} + \text{pOH} = 14.00.$$

Solve for pOH by subtracting pH from both sides of the equation.

$$\text{pOH} = 14.00 - \text{pH}$$

Substitute the known value of pH.

$$\text{pOH} = 14.00 - 7.40 = 6.60$$

Substitute 6.60 for pOH in the equation $[\text{OH}^-] = \text{antilog} (-\text{pOH})$

$$[\text{OH}^-] = \text{antilog} (-6.60)$$

Continued on next page



Blood banks collect blood from healthy people to hold in reserve for persons who need transfusions.

Use a log table or your calculator to find the antilog. The antilog of -6.60 is 2.5×10^{-7} .

$$[\text{OH}^-] = 2.5 \times 10^{-7}M.$$

The concentration of hydroxide ion in blood is $2.5 \times 10^{-7}M$.

3. Evaluate the Answer

The concentrations of the hydrogen ion and hydroxide ion are correctly stated with two significant figures because the given pH has two decimal places. As predicted, $[\text{H}^+]$ is less than 10^{-7} and $[\text{OH}^-]$ is greater than 10^{-7} .

PRACTICE PROBLEMS

21. The pH is given for three solutions. Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ in each solution.

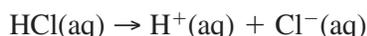
a. pH = 2.37

c. pH = 6.50

b. pH = 11.05

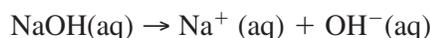
Calculating the pH of solutions of strong acids and strong bases

Look at the bottles of acid and base solutions in **Figure 19-13**. They are labeled with the number of moles of molecules or formula units that were dissolved in a liter of water (M) when the solutions were made. Each of the bottles contains a strong acid or base. Recall from Section 19.2 that strong acids and bases are essentially 100% ionized. That means that this reaction for the ionization of HCl goes to completion.



Every HCl molecule produces one H^+ ion. The bottle labeled $0.1M$ HCl contains 0.1 mole of H^+ ions per liter and 0.1 mole of Cl^- ions per liter. For all strong monoprotic acids, the concentration of the acid is the concentration of H^+ ion. Thus, you can use the concentration of the acid for calculating pH.

Similarly, the $0.1M$ solution of the strong base NaOH in **Figure 19-13** is fully ionized.



One formula unit of NaOH produces one OH^- ion. Thus, the concentration of the hydroxide ion is the same as the concentration of the solution, $0.1M$.

Some strong bases contain two or more hydroxide ions in each formula unit. Calcium hydroxide ($\text{Ca}(\text{OH})_2$) is an example. The concentration of hydroxide ion in a solution of $\text{Ca}(\text{OH})_2$ is twice the concentration of the ionic compound. Thus, the concentration of OH^- in a $7.5 \times 10^{-4}M$ solution of $\text{Ca}(\text{OH})_2$ is $7.5 \times 10^{-4}M \times 2 = 1.5 \times 10^{-3}M$.

PRACTICE PROBLEMS

22. Calculate the pH of the following solutions.

a. $1.0M$ HI

c. $1.0M$ KOH

b. $0.050M$ HNO_3

d. $2.4 \times 10^{-5}M$ $\text{Mg}(\text{OH})_2$



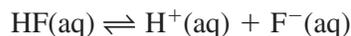
For more practice calculating $[\text{H}^+]$ and $[\text{OH}^-]$ from pH, go to **Supplemental Practice Problems** in Appendix A.



Figure 19-13

The label on a bottle of a strong acid or a strong base tells you immediately the concentration of hydrogen ions or hydroxide ions in the solution. That's because, in solution, strong acids and bases exist entirely as ions. What is $[\text{H}^+]$ in $0.1M$ HCl? What is $[\text{OH}^-]$ in $0.1M$ NaOH?

Using pH to calculate K_a Suppose you measured the pH of a 0.100M solution of the weak acid HF and found it to be 3.20. Would you have enough information to calculate K_a for HF?



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

From the pH, you could calculate $[\text{H}^+]$. Then, remember that for every mole per liter of H^+ ion there must be an equal concentration of F^- ion. That means that you know two of the variables in the K_a expression. What about the third, $[\text{HF}]$? The concentration of HF at equilibrium is equal to the initial concentration of the acid (0.100M) minus the moles per liter of HF that dissociated ($[\text{H}^+]$). Example Problem 19-5 illustrates a similar calculation for formic acid.

EXAMPLE PROBLEM 19-5

Calculating K_a from pH

The pH of a 0.100M solution of formic acid is 2.38. What is K_a for HCOOH?

1. Analyze the Problem

You are given the pH of the solution which allows you to calculate the concentration of the hydrogen ion. You know that the concentration of HCOO^- equals the concentration of H^+ . The concentration of un-ionized HCOOH is the difference between the initial concentration of the acid and $[\text{H}^+]$.

Known

pH = 2.38
concentration of the solution = 0.100M

Unknown

$K_a = ?$

2. Solve for the Unknown

Write the acid ionization constant expression.

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

Use the pH to calculate $[\text{H}^+]$.

$$\text{pH} = -\log[\text{H}^+]$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

Substitute the known value of pH.

$$[\text{H}^+] = \text{antilog}(-2.38)$$

Use a log table or calculator to find the antilog. The antilog of -2.38 is 4.2×10^{-3} .

$$[\text{H}^+] = 4.2 \times 10^{-3}\text{M}$$

$$[\text{HCOO}^-] = [\text{H}^+] = 4.2 \times 10^{-3}\text{M}$$

$[\text{HCOOH}]$ equals the initial concentration minus $[\text{H}^+]$.

$$[\text{HCOOH}] = 0.100\text{M} - 4.2 \times 10^{-3}\text{M} = 0.096\text{M}$$

Substitute the known values into the K_a expression.

$$K_a = \frac{(4.2 \times 10^{-3})(4.2 \times 10^{-3})}{(0.096)} = 1.8 \times 10^{-4}$$

The acid ionization constant for HCOOH is 1.8×10^{-4} .

Continued on next page



Natural rubber is an important agricultural export in Southeast Asia. Formic acid is used during the process that converts the milky latex fluid tapped from rubber trees into natural rubber.

3. Evaluate the Answer

The calculations are correct and the answer is correctly reported with two significant figures. The K_a is in the range of reasonably weak acids.

Practice!

For more practice calculating K_a , go to **Supplemental Practice Problems** in Appendix A.

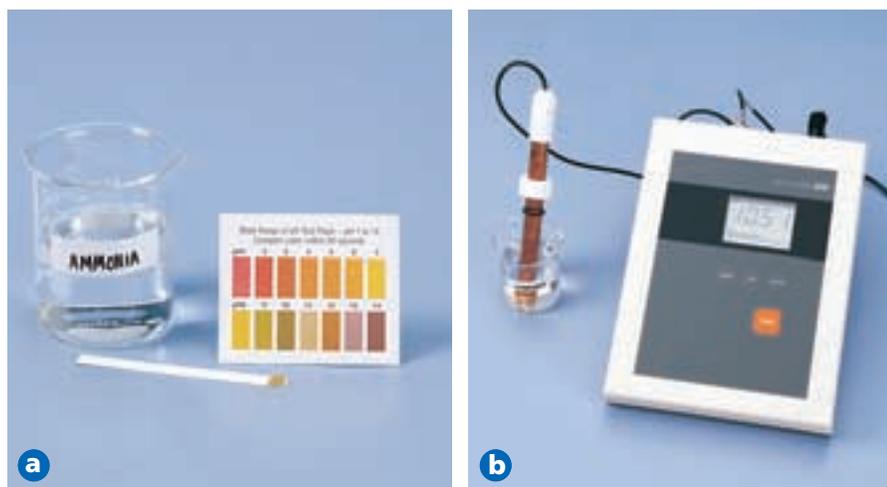
PRACTICE PROBLEMS

23. Calculate the K_a for the following acids using the given information.
- 0.220M solution of H_3AsO_4 , pH = 1.50
 - 0.0400M solution of $HClO_2$, pH = 1.80

Measuring pH Perhaps in an earlier science course, you used indicator paper to measure the pH of a solution. All pH paper is impregnated with one or more substances called indicators that change color depending upon the concentration of hydrogen ion in a solution. When a strip of pH paper is dipped into an acidic or basic solution, the color of the paper changes. To determine the pH, the new color of the paper is compared with standard pH colors on a chart, as shown in **Figure 19-14a**. The pH meter in **Figure 19-14b** provides a more accurate measure of pH. When electrodes are placed in a solution, the meter gives a direct analog or digital readout of pH.

Figure 19-14

The approximate pH of a solution can be obtained by wetting a piece of pH paper with the solution and comparing the color of the wet paper with a set of standard colors as shown in **a**. The pH meter in **b** provides a more accurate measurement in the form of a digital display of the pH.



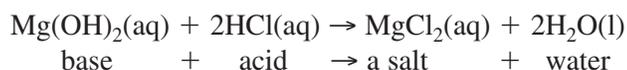
Section 19.3 Assessment

- What is the relationship between the pH of a solution and the concentration of hydrogen ions in the solution?
- If you know the pOH of a solution, how can you determine its pH?
- How does the ion product constant for water relate to the concentrations of H^+ and OH^- in aqueous solutions?
- Thinking Critically** Why is it logical to assume that the hydrogen ion concentration in an aqueous solution of a strong monoprotic acid equals the molarity of the acid?
- Applying Concepts** Would it be possible to calculate the pH of a weak acid solution if you knew the molarity of the solution and its K_a ? Explain.

If you were to experience heartburn or indigestion, you might take an antacid to relieve your discomfort. What kind of reaction occurs when magnesium hydroxide, the active ingredient in the common antacid called milk of magnesia, contacts hydrochloric acid produced by the stomach?

The Reaction Between Acids and Bases

When magnesium hydroxide and hydrochloric acid react, the resulting solution has properties characteristic of neither an acid nor a base. This type of reaction is called a neutralization reaction. A **neutralization reaction** is a reaction in which an acid and a base react in aqueous solution to produce a salt and water. A **salt** is an ionic compound made up of a cation from a base and an anion from an acid. Neutralization is a double-replacement reaction. In the reaction between magnesium hydroxide and hydrochloric acid, magnesium replaces hydrogen in HCl and hydrogen replaces magnesium in Mg(OH)₂. The reaction may be described by this balanced formula equation.

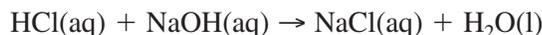


Note that the cation from the base (Mg²⁺) is combined with the anion from the acid (Cl⁻) in the salt MgCl₂.

PRACTICE PROBLEMS

- 29.** Write balanced formula equations for the following acid-base neutralization reactions.
- nitric acid and cesium hydroxide
 - hydrobromic acid and calcium hydroxide
 - sulfuric acid and potassium hydroxide
 - acetic acid and ammonium hydroxide

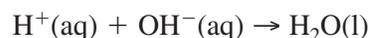
When considering neutralization reactions, it is important to determine whether all of the reactants and products exist in solution as molecules or formula units. For example, examine the formula equation and complete ionic equation for the reaction between hydrochloric acid and sodium hydroxide.



Because HCl is a strong acid, NaOH a strong base, and NaCl a soluble salt, all three compounds exist as ions in aqueous solution.



The chloride ion and the sodium ion appear on both sides of the equation so they are spectator ions. They can be eliminated to obtain the net ionic equation for the neutralization of a strong acid by a strong base.



Objectives

- **Write** chemical equations for neutralization reactions.
- **Explain** how neutralization reactions are used in acid-base titrations.
- **Compare** the properties of buffered and unbuffered solutions.

Vocabulary

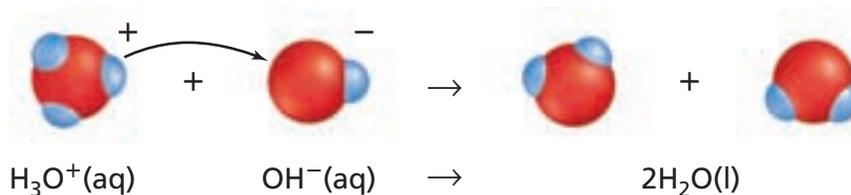
neutralization reaction
salt
titration
equivalence point
acid-base indicator
end point
salt hydrolysis
buffer
buffer capacity



For more practice writing neutralization equations, go to **Supplemental Practice Problems** in Appendix A.

Figure 19-15

A hydronium ion transfers a hydrogen ion to a hydroxide ion. The loss of the hydrogen ion by H_3O^+ results in a water molecule. The gain of a hydrogen ion by OH^- also results in a water molecule.

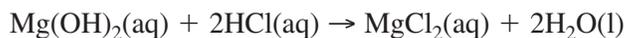


Recall that in an aqueous solution, a H^+ ion exists as a H_3O^+ ion, so the net ionic equation for an acid-base neutralization reaction is

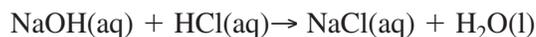


This neutralization reaction is illustrated in **Figure 19-15**. The **How It Works** feature at the end of this chapter shows that this equation does not apply to the neutralization of a strong acid by a weak base.

Acid-base titration The stoichiometry of an acid-base neutralization reaction is the same as that of any other reaction that occurs in solution. In the antacid reaction you just read about, one mole of magnesium hydroxide neutralizes two moles of hydrochloric acid.



In the reaction of sodium hydroxide and hydrogen chloride, one mole of sodium hydroxide neutralizes one mole of hydrogen chloride.

**Figure 19-16**

In the titration of an acid by a base, the pH meter measures the pH of the acid solution in the beaker as a solution of base with a known concentration is added from the buret.

Stoichiometry provides the basis for a procedure called titration, which is used to determine the concentrations of acidic and basic solutions. **Titration** is a method for determining the concentration of a solution by reacting a known volume of the solution with a solution of known concentration. If you wished to find the concentration of an acid solution, you would titrate the acid solution with a solution of a base of known concentration. You also could titrate a base of unknown concentration with an acid of known concentration. How is an acid-base titration carried out? **Figure 19-16** illustrates the equipment used for the following titration procedure using a pH meter.

1. A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker. The electrodes of a pH meter are immersed in this solution and the initial pH of the solution is read and recorded.
2. A buret is filled with the titrating solution of known concentration. This solution is called the standard solution.
3. Measured volumes of the standard solution are added slowly and mixed into the solution in the beaker. The pH is read and recorded after each addition. This process continues until the reaction reaches the stoichiometric point, which is the point at which moles of H^+ ion from the acid equal moles of OH^- ion from the base. The stoichiometric point is known as the **equivalence point** of the titration.

Figure 19-17 shows how the pH of the solution changes during the titration of 50.0 mL of 0.100M HCl, a strong acid, with 0.100M NaOH, a strong base. The initial pH of the 0.100M HCl is 1.00. As NaOH is added, the acid is neutralized and the solution's pH increases gradually. However, when nearly all of the H^+ ions from the acid have been used up, the pH increases dramatically with the addition of an exceedingly small volume of NaOH. This

abrupt increase in pH occurs at the equivalence point of the titration. Beyond the equivalence point, the addition of more NaOH again results in a gradual increase in pH.

For convenience, chemists often use a chemical dye rather than a pH meter to detect the equivalence point of an acid-base titration. Chemical dyes whose colors are affected by acidic and basic solutions are called **acid-base indicators**. The chart in **Figure 19-18** shows

the colors of several common acid-base indicators at various positions on the pH scale. The point at which the indicator used in a titration changes color is called the **end point** of the titration. The color change of the indicator selected for an acid-base titration should coincide closely with the equivalence point of the titration because the role of the indicator is to indicate to you, by means of a color change, that just enough of the titrating solution has been added to neutralize the unknown solution. In **Figure 19-18** you can see that bromthymol blue is an excellent choice for an equivalence point near pH 7. Bromthymol blue turns from yellow to blue as the pH of the solution changes from acidic to basic. Thus, the indicator's green transition color can mark an equivalence point near pH 7. What indicator might you choose for a titration that has its equivalence point at pH 5?

You might think that all titrations must have an equivalence point at pH 7 because that's the point at which concentrations of hydrogen ions and hydroxide ions are equal and the solution is neutral. This is not the case, however. Some titrations have equivalence points at $\text{pH} < 7$ and some have equivalence

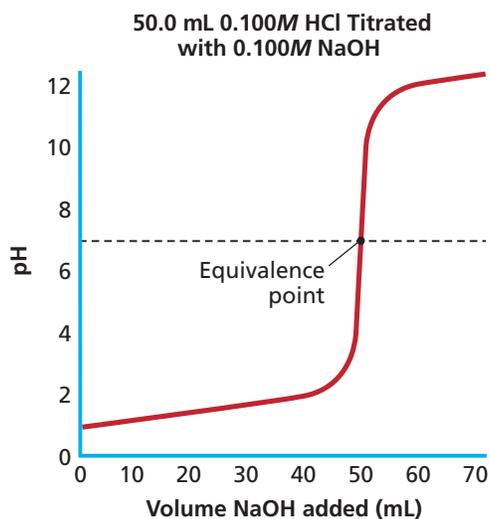


Figure 19-17

A steep rise in the pH of the acid solution indicates that all the H^+ ions from the acid have been neutralized by the OH^- ions of the base. The point at which the curve flexes (at its intersection with the dashed line) is the equivalence point of the titration. What is equivalent at this point?

Figure 19-18

Choosing the right indicator is important. The indicator must change color at the equivalence point of the titration and the equivalence point is not always at pH 7. Would methyl red be a good choice for the titration graphed in **Figure 19-17**?

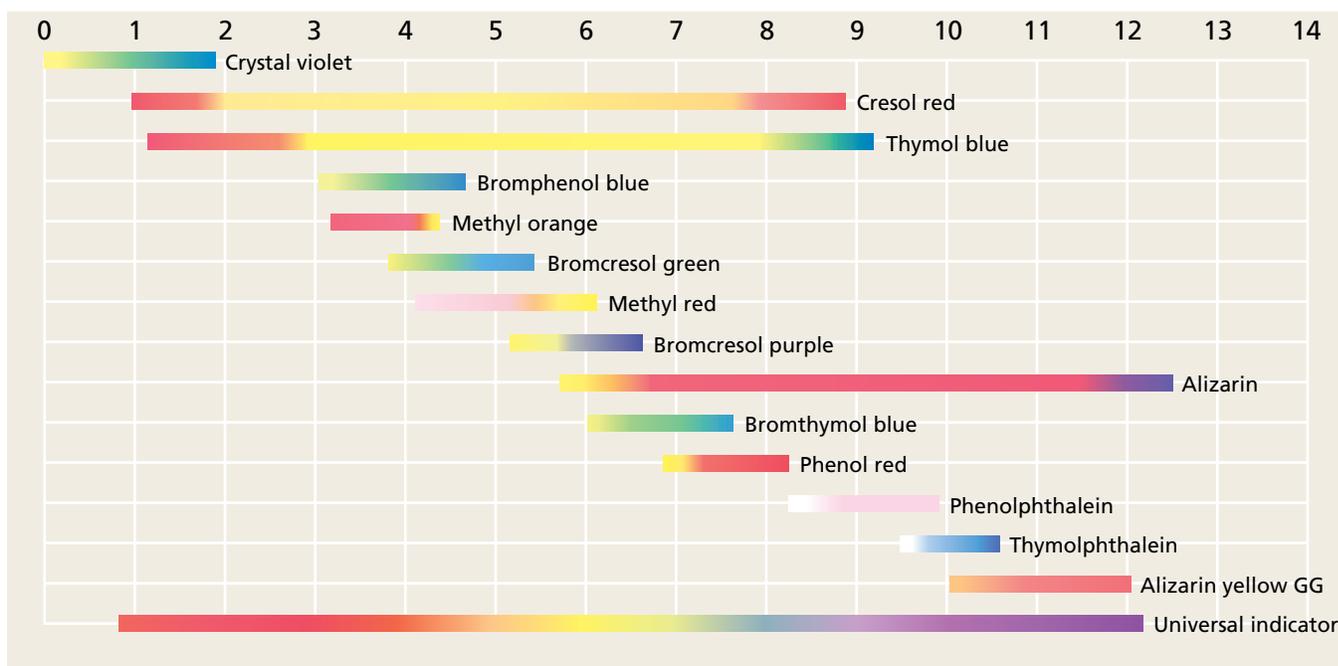
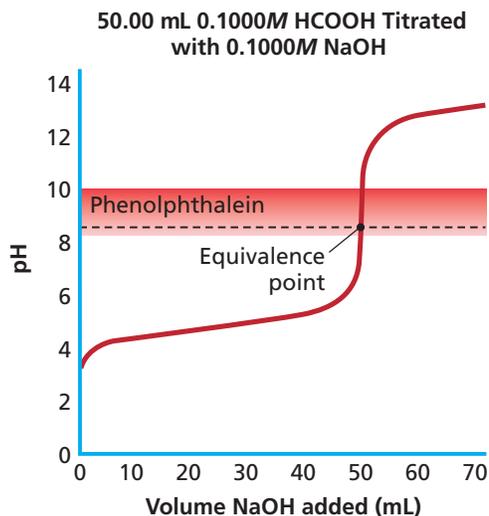


Figure 19-19

In this titration of a weak acid (HCOOH) by a strong base (NaOH), the equivalence point is not at pH 7. Check **Figure 19-18**. Do you agree that phenolphthalein is the best choice of indicator for this titration? Why?

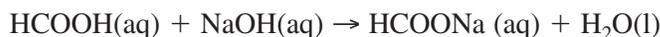


points at pH values > 7 . The pH at the equivalence point of a titration depends upon the relative strengths of the reacting acid and base. **Figure 19-19** shows that the equivalence point for the titration of methanoic acid (a weak acid) with sodium hydroxide (a strong base) lies between pH 8 and pH 9. Therefore, phenolphthalein is a good indicator for this titration. **Figure 19-20** illustrates and describes the titration of a solution of

methanoic acid (HCOOH) of unknown concentration with 0.1000M sodium hydroxide solution.

Calculating molarity From the experimental data, the molarity of the unknown HCOOH solution can be calculated by following these steps.

1. Write the balanced formula equation for the acid-base reaction.



2. Calculate the number of moles of NaOH contained in the volume of standard solution added. First, convert milliliters of NaOH to liters by multiplying the volume by a conversion factor that relates milliliters and liters.

$$18.28 \text{ mL NaOH} \times \frac{1 \text{ L NaOH}}{1000 \text{ mL NaOH}} = 0.01828 \text{ L NaOH}$$

Determine the moles of NaOH used by multiplying the volume by a conversion factor that relates moles and liters, the molarity of the solution.

$$0.01828 \text{ L NaOH} \times \frac{0.1000 \text{ mol NaOH}}{1 \text{ L NaOH}} = 1.828 \times 10^{-3} \text{ mol NaOH}$$

Figure 19-20

Titration is a precise procedure requiring practice. The white paper under the flask provides a background for viewing the indicator color change.



a The buret contains the standard solution (0.1000M NaOH) and the flask contains 25.00 mL HCOOH solution along with a small amount of phenolphthalein indicator.



b The standard solution is added slowly to the acid solution in the flask. Although the phenolphthalein indicator turns pink, the color disappears upon mixing until the end point is reached.



c The end point of the titration is marked by a permanent but very light pink color. A careful reading of the buret reveals that 18.28 mL 0.1000M NaOH has been added.

3. Use the mole ratio in the balanced equation to calculate the moles of reactant in the unknown solution.

$$1.828 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} = 1.828 \times 10^{-3} \text{ mol HCOOH}$$

4. Because molarity is defined as moles of solute per liter of solution, calculate the molarity of the unknown solution by dividing the moles of unknown (HCOOH) by the volume of the unknown solution expressed in liters. (25.00 mL = 0.02500 L)

$$M_{\text{HCOOH}} = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.02500 \text{ L HCOOH}} = 7.312 \times 10^{-2} M$$

The molarity of the HCOOH solution is $7.312 \times 10^{-2} M$, or $0.07312 M$. In the **CHEMLAB** at the end of this chapter, you will use titration to standardize a base.

PRACTICE PROBLEMS

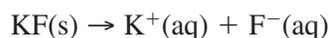
30. What is the molarity of a CsOH solution if 30.0 mL of the solution is neutralized by 26.4 mL 0.250M HBr solution?
31. What is the molarity of a nitric acid solution if 43.33 mL 0.1000M KOH solution is needed to neutralize 20.00 mL of unknown solution?
32. What is the concentration of a household ammonia cleaning solution if 49.90 mL 0.5900M HCl is required to neutralize 25.00 mL solution?

Practice!
For more practice calculating molarity, go to **Supplemental Practice Problems** in Appendix A.

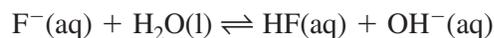
Salt Hydrolysis

You just learned that a salt is an ionic compound made up of a cation from a base and an anion from an acid. What reaction, if any, occurs when a salt is dissolved in pure water? In **Figure 19-21**, several drops of bromthymol blue indicator solution have been added to 0.10M aqueous solutions of the salts ammonium chloride (NH_4Cl), sodium nitrate (NaNO_3), and potassium fluoride (KF). Sodium nitrate turns the indicator green, which means that a solution of sodium nitrate is neutral. However, the blue color of the KF solution means that a solution of potassium fluoride is basic, and the yellow color of the ammonium chloride solution indicates that the NH_4Cl solution is acidic. Why are some aqueous salt solutions neutral, some basic, and some acidic?

The answer is that many salts react with water in a process known as salt hydrolysis. In **salt hydrolysis**, the anions of the dissociated salt accept hydrogen ions from water or the cations of the dissociated salt donate hydrogen ions to water. Does a reaction occur when potassium fluoride dissolves in water? Potassium fluoride is the salt of a strong base (KOH) and a weak acid (HF) and dissociates into potassium ions and fluoride ions.



Some fluoride ions react with water molecules to establish this equilibrium.



Note that the fluoride ion acts as a Brønsted-Lowry base and accepts a hydrogen ion from H_2O . Hydrogen fluoride molecules and OH^- ions are produced. Although the resulting equilibrium lies far to the left, the potassium fluoride solution is basic because some OH^- ions were produced



Figure 19-21

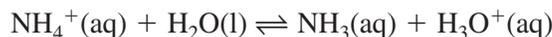
The indicator bromthymol blue provides surprising results when added to three solutions of ionic salts. An ammonium chloride solution is acidic, a sodium nitrate solution is neutral, and a potassium fluoride solution is basic. The explanation has to do with the strengths of the acid and base from which each salt was formed.

making the concentration of OH^- ions greater than the concentration of H^+ ions.

What about ammonium chloride? NH_4Cl is the salt of a weak base (NH_3) and a strong acid (HCl). When dissolved in water, the salt dissociates into ammonium ions and chloride ions.



The ammonium ions then react with water molecules to establish this equilibrium.



The ammonium ion is a Brønsted-Lowry acid in the forward reaction, which produces NH_3 molecules and hydronium ions. This equilibrium also lies far to the left. However, an ammonium chloride solution is acidic because the solution contains more hydronium ions than hydroxide ions.

The hydrolysis of KF produces a basic solution; the hydrolysis of NH_4Cl produces an acidic solution. What about sodium nitrate (NaNO_3)? Sodium nitrate is the salt of a strong acid (nitric acid) and a strong base (sodium hydroxide). Little or no salt hydrolysis occurs and a solution of sodium nitrate is neutral.

Practice!

For more practice writing equations for salt hydrolysis reactions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- 33.** Write equations for the salt hydrolysis reactions that occur when the following salts are dissolved in water. Classify each solution as acidic, basic, or neutral.
- | | |
|---------------------|----------------------|
| a. ammonium nitrate | c. potassium sulfate |
| b. rubidium acetate | d. calcium carbonate |

Buffered Solutions

If you maintain a tropical-fish aquarium similar to the one shown in **Figure 19-22**, you know that the pH of the water must be kept relatively constant if the fish are to survive. Control of pH is important in your body, too. The pH of your blood must be maintained at an average of 7.4. A potentially fatal problem will develop if the pH rises or falls as much as 0.3 pH unit. The gastric juices in your stomach must have a pH between 1.6 and 1.8 to promote digestion of certain foods. How does your body maintain pH values within such narrow limits? It does so by producing buffers.

Figure 19-22

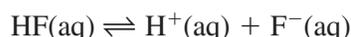
This freshwater aquarium is home to a variety of cichlid. To provide a healthy environment for this fish, it's necessary to know whether it is an African cichlid or a South American cichlid. The African cichlid requires water at a pH of 8.0 to 9.2; the South American cichlid requires a pH of 6.4 to 7.0. What would happen if you put a South American cichlid and some plants into a tank with water at pH of 8.5?



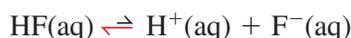
Buffers are solutions that resist changes in pH when limited amounts of acid or base are added. For example, adding 0.01 mole of HCl to 1 L of pure water lowers the pH by 5.0 units, from 7.0 to 2.0. Similarly, adding 0.01 mole of NaOH to 1 L of pure water increases the pH from 7.0 to 12.0. But if you add the same amount of either HCl or NaOH to 1 L of a buffered solution, the pH might change by no more than 0.1 unit.

How do buffers work? A buffer is a mixture of a weak acid and its conjugate base or a weak base and its conjugate acid. The mixture of ions and molecules in a buffer solution resists changes in pH by reacting with any hydrogen ions or hydroxide ions added to the buffered solution.

Suppose that a buffer solution contains 0.1M concentrations of hydrofluoric acid, and 0.1M fluoride ion (NaF). See **Figure 19-23**. HF is the acid and F⁻ is its conjugate base. The following equilibrium would be established.

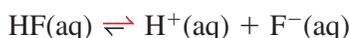


When an acid is added to this buffered solution, the equilibrium shifts to the left. The added H⁺ ions react with F⁻ ions to form additional HF molecules.



The pH of the solution remains fairly constant because the additional HF molecules do not ionize appreciably.

When a base is added to the hydrofluoric acid/fluoride ion buffer system, the added OH⁻ ions react with H⁺ ions to form water. This decreases the concentration of hydrogen ions and the equilibrium shifts to the right to replace the H⁺ ions.



Although the shift to the right consumes HF molecules and produces additional F⁻ ions, the pH remains fairly constant because the H⁺ ion concentration has not changed appreciably.

A buffer solution's capacity to resist pH change can be exceeded by the addition of too much acid or base. Excessive acid overwhelms the hydrofluoric acid/fluoride ion buffer system by using up almost all of the F⁻ ions. Similarly, too much base overwhelms the same system by using up almost all of the HF molecules. The amount of acid or base a buffer solution can absorb without a significant change in pH is called the **buffer capacity** of the solution. The greater the concentrations of the buffering molecules and ions in the solution, the greater the solution's buffer capacity. Do the **problem-solving LAB** on the next page to find out how buffers work in your blood.

Choosing a buffer A buffer can best resist both increases and decreases in pH when the concentrations of the conjugate acid-base pair are equal or nearly equal. Consider the H₂PO₄⁻/HPO₄²⁻ buffer system made by mixing equal molar amounts of NaH₂PO₄ and Na₂HPO₄.



What is the pH of such a buffer solution? The acid ionization constant expression for the equilibrium can provide the answer.

$$K_a = 6.2 \times 10^{-8} = \frac{[\text{H}^{\text{+}}][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$



Figure 19-23

Hydrofluoric acid is used to etch decorative designs on glass. Beeswax is coated on the glass and the design is drawn in the wax with a metal stylus. When the glass is dipped into hydrofluoric acid solution, the acid etches the glass in the grooves left by the stylus.

Table 19-5

Buffer Systems with Equimolar Components		
Buffer equilibrium	Conjugate acid-base pair in buffered solution	Buffer pH
$\text{HF}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{F}^-(\text{aq})$	HF/F^-	3.20
$\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.76
$\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.35
$\text{H}_2\text{PO}_4^-(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{NH}_4^+/\text{NH}_3$	9.4
$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$	10.70

Because the solution has been made up with equal molar amounts of Na_2HPO_4 and NaH_2PO_4 , $[\text{HPO}_4^{2-}]$ is equal to $[\text{H}_2\text{PO}_4^-]$. Thus, the two terms in the acid ionization expression cancel.

$$6.2 \times 10^{-8} = \frac{[\text{H}^+][\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

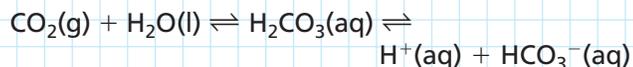
$$6.2 \times 10^{-8} = [\text{H}^+]$$

Recall that $\text{pH} = -\log [\text{H}^+]$.

problem-solving LAB

How does your blood maintain its pH?

The pH of human blood must be kept within the narrow range of 7.1 to 7.7. Outside this range, proteins in the body lose their structure and ability to function. Fortunately, a number of buffers maintain the necessary acid/base balance. The carbonic acid/hydrogen carbonate buffer is the most important.



As acids and bases are dumped into the bloodstream as a result of normal activity, the blood's buffer systems shift to effectively maintain a healthy pH.

Analysis

Depending upon the body's metabolic rate and other factors, the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift according to Le Châtelier's principle. In addition, the lungs can alter the rate at which CO_2 is

expelled from the body by breathing and the kidneys can alter the rate of removal of hydrogen carbonate ion.

Thinking Critically

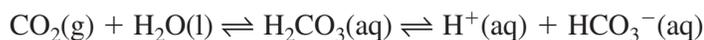
- Calculate the molarity of H^+ in blood at a normal $\text{pH} = 7.4$ and at $\text{pH} = 7.1$, the lower limit of a healthy pH range. If the blood's pH changes from 7.4 to 7.1, how many times greater is the concentration of hydrogen ion?
- The ratio of HCO_3^- to CO_2 is 20:1. Why is this imbalance favorable for maintaining a healthy pH?
- In the following situations, predict whether the pH of the blood will rise or fall, and which way the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift.
 - A person with severe stomach flu vomits many times during a 24-hour period.
 - To combat heartburn, a person foolishly takes a large amount of sodium hydrogen carbonate.

$$\text{pH} = -\log (6.2 \times 10^{-8}) = 7.21$$

Thus, when equimolar amounts of each of the components are present in the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer system, the system can maintain a pH close to 7.21. Note that the pH is the negative log of K_a . **Table 19-5** lists several buffer systems with the pH at which each is effective. The pH values were calculated as the pH for the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer system was calculated above.

Buffers in Blood Your blood is a slightly alkaline solution with a pH of approximately 7.4. To be healthy, that pH must be maintained within narrow limits. A condition called acidosis occurs if the pH falls more than 0.3 units below 7.4. An equally serious condition called alkalosis exists if the pH rises 0.3 units. You may have experienced a mild case of acidosis if you have overexerted and developed a cramp in your leg. Cramping results from the formation of lactic acid in muscle tissue.

Your body employs three principle strategies for maintaining proper blood pH. The first is excretion of excess acid or base in the urine. The second is the elimination of CO_2 , the anhydride of carbonic acid, by breathing. The third involves a number of buffer systems. You have already learned about the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ buffer in the **problem-solving LAB**.



The $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer system described above is also present in blood. Because the concentration of phosphates in the blood is not as high as the concentration of carbonates, the phosphate buffer plays a lesser role in maintaining blood pH even though it is a more efficient buffer. Other buffer systems, including some associated with the hemoglobin molecule work together with the carbonate and phosphate buffers to maintain pH.

Recall that if acid levels begin to rise, the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium shifts to the left to consume the added H^+ . At the same time, the respiratory system is signaled to increase the rate of breathing to eliminate the higher level of CO_2 . If levels of base begin to rise, the equilibrium shifts to the right as H^+ neutralizes the base and the respiratory system slows the removal of CO_2 by depressing the breathing rate. Can you explain why?

Suppose that during a concert a fan becomes overexcited and begins to hyperventilate. Excessive loss of CO_2 shifts the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium to the left and increases the pH. The situation can be controlled by helping the person calm down and breathe into a paper bag to recover exhaled CO_2 as shown in **Figure 19-24**.



Figure 19-24

Look at the equation for the carbonate buffer as you think about this situation. As CO_2 is eliminated through rapid breathing, the equilibrium shifts to the left. The $[\text{H}^+]$ decreases and pH increases. Re-breathing CO_2 from a paper bag helps restore the pH balance.

Section 19.4 Assessment

- 34.** Write the formula equation and the net ionic equation for the neutralization reaction between hydroiodic acid and potassium hydroxide.
- 35.** Explain the difference between the equivalence point and the end point of a titration.
- 36.** Predict the results of two experiments: A small amount of base is added to an unbuffered solution with a pH of 7 and the same amount of base is added to a buffered solution with a pH of 7.
- 37. Thinking Critically** When a salt is dissolved in water, how can you predict whether or not a salt hydrolysis reaction occurs?
- 38. Designing an Experiment** Describe how you would design and carry out a titration in which you use $0.250M$ HNO_3 to determine the molarity of a cesium hydroxide solution.

Standardizing a Base Solution by Titration

The procedure called titration can be used to standardize a solution of a base, which means determine its molar concentration. To standardize a base, a solution of the base with unknown molarity is gradually added to a solution containing a known mass of an acid. The procedure enables you to determine when the number of moles of added OH^- ions from the base equals the number of moles of H^+ ion from the acid.

Problem

How can you determine the molar concentration of a base solution? How do you know when the neutralization reaction has reached the equivalence point?

Objectives

- **Recognize** the color change of the indicator that shows that the equivalence point has been reached.
- **Measure** the mass of the acid and the volume of the base solution used.
- **Calculate** the molar concentration of the base solution.

Materials

50-mL buret	250-mL Erlenmeyer flask
buret clamp	500-mL Florence flask and rubber stopper
ring stand	250-mL beaker
sodium hydroxide pellets (NaOH)	centigram balance
potassium hydrogen phthalate ($\text{KHC}_8\text{H}_4\text{O}_4$)	wash bottle
distilled water	phenolphthalein solution
weighing bottle	dropper
spatula	

Safety Precautions



- Always wear safety goggles and a lab apron.

Pre-Lab

1. What is the equivalence point of a titration?
2. Read the entire CHEMLAB.
3. What is the independent variable? The dependent variable? Constant?
4. When the solid acid dissolves to form ions, how many moles of H^+ ions are produced for every mole of acid used?
5. What is the formula used to calculate molarity?
6. Prepare a data table that will accommodate multiple titration trials.
7. List safety precautions that must be taken.

Titration Data	
	Trial 1
mass of weighing bottle and acid	
mass of weighing bottle	
mass of solid acid	
moles of acid	
moles of base required	
final reading of base buret	
initial reading of base buret	
volume of base used in mL	
molarity of base	

Procedure

1. Place approximately 4 g NaOH in a 500-mL Florence flask. Add enough water to dissolve the pellets and bring the volume of the NaOH solution to about 400 mL. **CAUTION:** *The solution will get hot.* Keep the stopper in the flask.
2. Use the weighing bottle to mass by difference about 0.40 g of potassium hydrogen phthalate (molar mass = 204.32 g/mol) into a 250-mL Erlenmeyer flask. Record this data.
3. Using a wash bottle, rinse down the insides of the flask and add enough water to make about 50 mL of solution. Add two drops of phenolphthalein indicator solution.
4. Set up the buret as shown. Rinse the buret with about 10 mL of your base solution. Discard the rinse solution in a discard beaker.



5. Fill the buret with NaOH solution. To remove any air trapped in the tip, allow a small amount of the base to flow from the tip of the buret into the discard beaker. Read the buret to the nearest 0.02 mL and record this initial reading. The meniscus of the solution in the buret should be at eye level when you make a reading.
6. Place a piece of white paper under the buret. Allow the NaOH solution to flow slowly from the buret into the flask containing the acid. Control the flow of the base solution with your left hand, and gently swirl the flask with your right hand.
7. The NaOH solution may be added in a rapid stream of drops until the pink color begins to last

longer as the flask is swirled. At this stage, begin adding the base drop by drop.

8. The equivalence point is reached when one additional drop of base turns the acid in the flask pink. The pink color should persist as the flask is swirled. Record the final volume in the buret.
9. Calculate the molarity of your base using steps 2–5 below.
10. Refill your buret with base. Rinse your Erlenmeyer flask with water. Repeat the titration with additional samples of acid until you get three trials that show close agreement between the calculated values of the molarity.

Cleanup and Disposal

1. The neutralized solutions can be washed down the sink using plenty of water.

Analyze and Conclude

1. **Observing and Inferring** Identify the characteristics of this neutralization reaction.
2. **Collecting and Interpreting Data** Complete the data table. Calculate the number of moles of acid used in each trial by dividing the mass of the sample by the molar mass of the acid.
3. **Using Numbers** How many moles of base are required to react with the moles of acid you used?
4. **Using Numbers** Convert the volume of base used from milliliters to liters.
5. **Analyze and Conclude** For each trial, calculate the molar concentration of the base by dividing the moles of base by the volume of base in liters.
6. **Error Analysis** How well did your calculated molarities agree? Explain any irregularities.

Real-World Chemistry

1. Use what you have learned about titration to design a field investigation to determine whether your area is affected by acid rain. Research the factors that affect the pH of rain, such as location, prevailing winds, and industries. Form a hypothesis about the pH of rain in your area. What equipment will you need to collect samples? To perform the titration? What indicator will you use?

How It Works

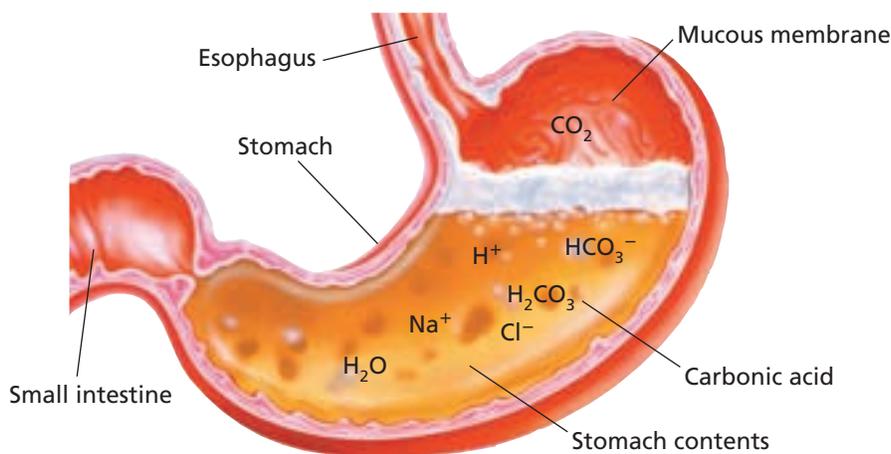
Antacids

Your stomach is a hollow organ where the food you eat is broken down into a usable form. The stomach contains powerful enzymes and hydrochloric acid, which are responsible for the breakdown process. Once the food is processed, it is released into the small intestine.

The pain and discomfort of indigestion is an indication that normal digestion has been interrupted. Heartburn is an irritation of the esophagus that is caused by stomach acid. Millions of people use antacids to treat indigestion and heartburn. Antacids are bases that neutralize digestive acids.



- 1 Acids and enzymes in the stomach help digest food. The pH in the stomach is about 2.5.
- 2 A basic mucous membrane lines the stomach and protects it from corrosion.



- 4 Carbonic acid decomposes into carbon dioxide gas and water.
$$\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$
- 3 A sodium hydrogen carbonate (sodium bicarbonate) antacid reacts with hydrochloric acid in the stomach and forms carbonic acid.
$$\text{NaHCO}_3(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq})$$

Thinking Critically

1. **Applying** Milk of magnesia is a suspension of magnesium hydroxide in water. Write the net ionic equation for the neutralization reaction between magnesium hydroxide and hydrochloric acid.
2. **Comparing and Contrasting** Why is sodium hydrogen carbonate an effective antacid but sodium hydroxide is not?

Summary

19.1 Acids and Bases: An Introduction

- Acidic solutions contain more hydrogen ions than hydroxide ions; neutral solutions contain equal concentrations of hydrogen ions and hydroxide ions; basic solutions contain more hydroxide ions than hydrogen ions.
- In the Arrhenius model, an acid is a substance that contains hydrogen and ionizes in aqueous solution to produce hydrogen ions. An Arrhenius base is a substance that contains a OH group and dissociates in aqueous solution to produce hydroxide ions.
- According to the Brønsted-Lowry model, an acid is a hydrogen ion donor and a base is a hydrogen ion acceptor.
- When a Brønsted-Lowry acid donates a hydrogen ion, a conjugate base is formed; when a Brønsted-Lowry base accepts a hydrogen ion, a conjugate acid is formed.

19.2 Strengths of Acids and Bases

- Strong acids and strong bases are completely ionized in dilute aqueous solution. Weak acids and weak bases are partially ionized.
- For weak acids and weak bases, the value of the acid or base ionization constant is a measure of the strength of the acid or base.

19.3 What is pH?

- The pH of a solution is the negative log of the hydrogen ion concentration and the pOH is the negative log of the hydroxide ion concentration.
- A neutral solution has a pH of 7.0 and a pOH of 7.0 because the concentrations of hydrogen ion and hydroxide ion are equal. The pH of a solution decreases as the solution becomes more acidic and increases as the solution becomes more basic. The pOH of a solution decreases as the solution becomes more basic and increases as the solution becomes more acidic.
- The ion product constant for water, K_w , equals the product of the hydrogen ion concentration and the hydroxide ion concentration.

19.4 Neutralization

- The general equation for an acid-base neutralization reaction is acid + base \rightarrow salt + water.
- The net ionic equation for the neutralization of a strong acid by a strong base is $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.
- Titration is the process in which an acid-base neutralization reaction is used to determine the concentration of a solution of unknown concentration.
- Buffered solutions contain mixtures of molecules and ions that resist changes in pH.

Key Equations and Relationships

- $pH = -\log [H^+]$ (p. 610)
- $pOH = -\log [OH^-]$ (p. 611)
- $pH + pOH = 14.00$ (p. 611)
- $K_w = [H^+][OH^-]$ (p. 608)

Vocabulary

- acid-base indicator (p. 619)
- acid ionization constant (p. 605)
- acidic solution (p. 597)
- amphoteric (p. 599)
- Arrhenius model (p. 597)
- base ionization constant (p. 606)
- basic solution (p. 597)
- Brønsted-Lowry model (p. 598)
- buffer (p. 623)
- buffer capacity (p. 623)
- conjugate acid (p. 598)
- conjugate acid-base pair (p. 598)
- conjugate base (p. 598)
- end point (p. 619)
- equivalence point (p. 618)
- ion product constant for water (p. 608)
- neutralization reaction (p. 617)
- pH (p. 610)
- pOH (p. 611)
- salt (p. 617)
- salt hydrolysis (p. 621)
- strong acid (p. 602)
- strong base (p. 606)
- titration (p. 618)
- weak acid (p. 603)
- weak base (p. 606)

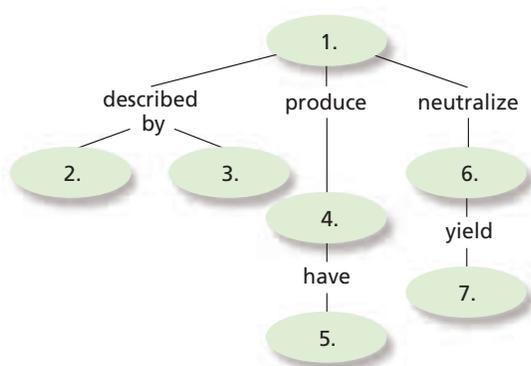


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Concept Mapping

39. Use the following words and phrases to complete the concept map: acidic solutions, acids, bases, Arrhenius model, $\text{pH} < 7$, a salt plus water, Brønsted-Lowry model.



Mastering Concepts

40. An aqueous solution tastes bitter and turns litmus blue. Is the solution acidic or basic? (19.1)
41. An acidic solution reacts with magnesium carbonate to produce a gas. What is the formula for the gas? (19.1)
42. In terms of ion concentrations, distinguish between acidic, neutral, and basic solutions. (19.1)
43. Write a balanced chemical equation that represents the self-ionization of water. (19.1)
44. How did Arrhenius describe acids and bases? (19.1)
45. Table sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) contains 22 hydrogen atoms per molecule. Does this make table sugar an Arrhenius acid? Explain your answer. (19.1)
46. Classify each of the following compounds as an Arrhenius acid or an Arrhenius base. (19.1)
- | | |
|-------------------------|-----------------------------|
| a. H_2S | c. $\text{Mg}(\text{OH})_2$ |
| b. RbOH | d. H_3PO_4 |
47. Explain the difference between a monoprotic acid, a diprotic acid, and a triprotic acid. Give an example of each. (19.1)
48. Why does acid rain dissolve statues made of marble (CaCO_3)? Write the formula equation for the reaction between sulfuric acid and calcium carbonate. (19.1)
49. Ammonia contains three hydrogen atoms per molecule. However, an aqueous ammonia solution is basic. Explain using the Brønsted-Lowry model of acids and bases. (19.1)
50. Identify the conjugate acid-base pairs in the equilibrium equation.
 $\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^-$ (19.1)
51. Gaseous HCl molecules interact with gaseous NH_3 molecules to form a white smoke made up of solid NH_4Cl particles. Explain whether or not this is an acid-base reaction according to both the Arrhenius model and the Brønsted-Lowry model. (19.1)
52. Explain the difference between a strong acid and a weak acid. (19.2)
53. Why are strong acids and bases also strong electrolytes? (19.2)
54. State whether each of the following acids is strong or weak. (19.2)
- | | |
|--------------------|----------------------|
| a. acetic acid | c. hydrofluoric acid |
| b. hydroiodic acid | d. phosphoric acid |
55. State whether each of the following bases is strong or weak. (19.2)
- | | |
|-----------------------|---------------------|
| a. rubidium hydroxide | c. ammonia |
| b. methylamine | d. sodium hydroxide |
56. How would you compare the strengths of two weak acids (19.2)
- | |
|--|
| a. experimentally? |
| b. by looking up information in a table or a handbook? |
57. **Figure 19-7** shows the conductivity of two acids. Explain how you could distinguish between solutions of two bases, one containing NaOH and the other NH_3 . (19.2)
58. Explain why the base ionization constant (K_b) is a measure of the strength of a base. (19.2)
59. Explain why a weak acid has a strong conjugate base. Give an equation that illustrates your answer. (19.2)
60. Explain why a weak base has a strong conjugate acid. Give an equation that illustrates your answer. (19.2)
61. Explain how you can calculate K_a for a weak acid if you know the concentration of the acid and its pH . (19.2)
62. What is the relationship between the pOH and the hydroxide-ion concentration of a solution? (19.3)

- 63.** Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the hydrogen-ion concentrations in the two solutions, how many times more acidic? (19.3)
- 64.** If the concentration of hydrogen ions in an aqueous solution decreases, what must happen to the concentration of hydroxide ions? Why? (19.3)
- 65.** Explain why pure water has a very slight electrical conductivity. (19.3)
- 66.** Why is the pH of pure water 7.0 at 298 K? (19.3)
- 67.** What is a standard solution in an acid-base titration? (19.4)
- 68.** How do you recognize the end point in an acid-base titration? (19.4)
- 69.** Give the name and formula of the acid and the base from which each salt was formed. (19.4)
- | | |
|--------------------|-----------------------------|
| a. NaCl | c. NH_4NO_2 |
| b. KClO_3 | d. CaS |
- 70.** How does buffering a solution change the solution's behavior when a base is added? When an acid is added? (19.4)
- 71.** Why are some aqueous salt solutions acidic or basic? (19.4)
- 72.** How does knowing the equivalence point in an acid-base titration help you choose an indicator for the titration? (19.4)
- 73.** An aqueous solution causes bromthymol blue to turn blue and phenolphthalein to turn colorless. What is the approximate pH of the solution? (19.4)
- 74.** In the net ionic equation for the neutralization reaction between nitric acid and magnesium hydroxide, what ions are left out of the equation? (19.4)
- 75.** Describe two ways you might detect the end point of an acid-base titration experimentally. (19.4)
- 76.** What is the approximate pH of the equivalence point in the titration pH curve? (19.4)
- 77.** Define both the equivalence point and the end point of an acid-base titration. Why should you choose an indicator so that the two points are nearly the same pH? (19.4)
- 78.** Explain how you can predict whether an aqueous salt solution is acidic, basic, or neutral by evaluating the strengths of the salt's acid and base parents. (19.4)
- 79.** In a hypochlorous acid/hypochlorite-ion buffer, what chemical species reacts when an acid is added to the solution? (19.4)
- 80.** Arrange the three buffers in order of increasing pH values. In order of increasing buffer capacity.
- 1.0M HClO/1.0M NaClO
 - 0.10M HClO/0.10M NaClO
 - 0.010M HClO/0.010M NaClO

Mastering Problems

Equations for Acid and Base Reactions (19.1)

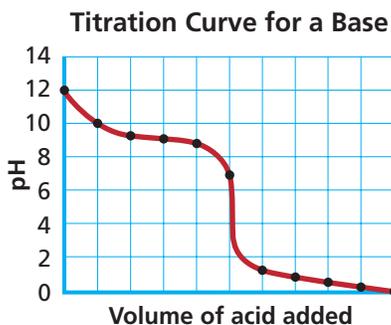
- 81.** Write a balanced formula equation for the reaction between sulfuric acid and calcium metal.
- 82.** Write a balanced formula equation for the reaction between potassium hydrogen carbonate and chlorous acid (HClO_2).
- 83.** Write the balanced chemical equation for the ionization of perchloric acid (HClO_4) in water.
- 84.** Write the balanced chemical equation for the dissociation of solid magnesium hydroxide in water.

Weak Acids and Bases (19.2)

- 85.** Write the equation for the ionization reaction and the acid ionization constant expression for the HS^- ion in water.
- 86.** Write the equation for the ionization reaction and the acid ionization constant expression for the third ionization of phosphoric acid in water.

K_w , pH, and pOH (19.3)

- 87.** Given the concentration of either hydrogen ion or hydroxide ion, use the ion product constant of water to calculate the concentration of the other ion at 298 K.
- $[\text{H}^+] = 1.0 \times 10^{-4}\text{M}$
 - $[\text{OH}^-] = 1.3 \times 10^{-2}\text{M}$
- 88.** Calculate the pH at 298 K of solutions having the following ion concentrations.
- $[\text{H}^+] = 1.0 \times 10^{-4}\text{M}$
 - $[\text{H}^+] = 5.8 \times 10^{-11}\text{M}$



- 89.** Calculate the pOH and pH at 298 K of solutions having the following ion concentrations.
- $[\text{OH}^-] = 1.0 \times 10^{-12}M$
 - $[\text{OH}^-] = 1.3 \times 10^{-2}M$
- 90.** Calculate the pH of each of the following strong acid or strong base solutions at 298 K.
- $2.6 \times 10^{-2}M$ HCl
 - $0.28M$ HNO_3
 - $7.5 \times 10^{-3}M$ NaOH
 - $0.44M$ KOH

Calculations Using K_a (19.3)

- 91.** A $8.6 \times 10^{-3}M$ solution of H_3PO_4 has a $\text{pH} = 2.30$. What is K_a for H_3PO_4 ?
- 92.** What is K_a for a solution of chloroacetic acid ($\text{C}_2\text{H}_3\text{ClO}_2$) which has a concentration of $0.112M$ and a pH of 1.92?

Neutralization Reactions (19.4)

- 93.** Write formula equations for the following acid-base neutralization reactions.
- sulfuric acid + sodium hydroxide
 - methanoic acid + potassium hydroxide
- 94.** Write formula equations and net ionic equations for the hydrolysis of the following salts in water.
- sodium carbonate
 - ammonium bromide
- 95.** In a titration, 33.21 mL $0.3020M$ rubidium hydroxide solution is required to exactly neutralize 20.00 mL hydrofluoric acid solution. What is the molarity of the hydrofluoric acid solution?
- 96.** A 35.00 mL-sample of NaOH solution is titrated to an endpoint by 14.76 mL $0.4122M$ HBr solution. What is the molarity of the NaOH solution?

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 97.** Calculate $[\text{H}^+]$ and $[\text{OH}^-]$ in each of the following solutions at 298 K.
- $\text{pH} = 3.00$
 - $\text{pH} = 5.24$
- 98.** Write the equation for the ionization reaction and the base ionization constant expression for ethylamine ($\text{C}_2\text{H}_5\text{NH}_2$) in water.
- 99.** Write net ionic equations for the three ionizations of boric acid (H_3BO_3) in water. Include H_2O in the three equations. Identify the conjugate acid-base pairs.

- 100.** How many milliliters of $0.225M$ HCl would be required to titrate 6.00 g KOH?
- 101.** What is the pH of a $0.200M$ solution of hypobromous acid (HBrO), $K_a = 2.8 \times 10^{-9}$.

Thinking Critically

- 102. Analyzing and Concluding** Is it possible that an acid according to the Arrhenius model is not a Brønsted-Lowry acid? Is it possible that an acid according to the Brønsted-Lowry model is not an Arrhenius acid? Explain and give examples.
- 103. Applying Concepts** Use the ion product constant of water at 298 K to explain why a solution with a pH of 3.0 must have a pOH of 11.0.
- 104. Using Numbers** The ion product constant of water rises with temperature. What is the pH of pure water at 313 K if K_w at that temperature is 2.917×10^{-14} ?
- 105. Interpreting Scientific Illustrations** Sketch the shape of the approximate pH vs. volume curve that would result from titrating a diprotic acid with a $0.10M$ NaOH solution.
- 106. Recognizing Cause and Effect** Illustrate how a buffer works using the $\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$ buffer system. Show with equations how the weak base/conjugate acid system is affected when small amounts of acid and base are added to a solution containing this buffer system.

Writing in Chemistry

- 107.** Examine the labels of at least two brands of shampoo and record any information regarding pH. Research the pH of skin, hair, and the pH ranges of various shampoos. Write a report summarizing your findings.
- 108.** The twenty amino acids combine to form proteins in living systems. Research the structures and K_a values for five amino acids. Compare the strengths of these acids with the weak acids in **Table 19-2**.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 109.** What factors determine whether a molecule is polar or nonpolar? (Chapter 9)
- 110.** When 5.00 g of a compound was burned in a calorimeter, the temperature of 2.00 kg of water increased from 24.5°C to 40.5°C . How much heat would be released by the combustion of 1.00 mol of the compound (molar mass = 46.1 g/mol)? (Chapter 16)

Use these questions and the test-taking tip to prepare for your standardized test.

- A carbonated soft drink has a pH of 2.5. What is the concentration of H^+ ions in the soft drink?
 - $3 \times 10^{-12}M$
 - $3 \times 10^{-3}M$
 - $4.0 \times 10^{-1}M$
 - 1.1×10^1M
- At the equivalence point of a strong acid-strong base titration, what is the approximate pH?
 - 3
 - 5
 - 7
 - 9
- Hydrogen bromide (HBr) is a strong, highly corrosive acid. What is the pOH of a 0.0375M HBr solution?
 - 12.57
 - 12.27
 - 1.73
 - 1.43
- A compound that accepts H^+ ions is
 - an Arrhenius acid
 - an Arrhenius base
 - a Brønsted-Lowry acid
 - a Brønsted-Lowry base

Interpreting Tables Use the table to answer questions 5–7.

- Which of the following acids is the strongest?
 - formic acid
 - cyanoacetic acid
 - lutidinic acid
 - barbituric acid
- What is the acid dissociation constant of propanoic acid?
 - 1.4×10^{-5}
 - 2.43×10^0
 - 3.72×10^{-3}
 - 7.3×10^4
- What is the pH of a 0.400M solution of cyanoacetic acid?
 - 2.059
 - 1.22
 - 2.45
 - 1.42

- Which of the following is NOT a characteristic of a base?
 - bitter taste
 - ability to conduct electricity
 - reactivity with some metals
 - slippery feel
- Diprotic succinic acid ($H_2C_4H_4O_4$) is an important part of the process that converts glucose to energy in the human body. What is the K_a expression for the second ionization of succinic acid?
 - $K_a = \frac{[H_3O^+][HC_4H_4O_4^-]}{[H_2C_4H_4O_4]}$
 - $K_a = \frac{[H_3O^+][C_4H_4O_4^{2-}]}{[HC_4H_4O_4^-]}$
 - $K_a = \frac{[H_2C_4H_4O_4]}{[H_3O^+][HC_4H_4O_4^-]}$
 - $K_a = \frac{[H_2C_4H_4O_4]}{[H_3O^+][C_4H_4O_4^{2-}]}$
- A solution of 0.600M HCl is used to titrate 15.00 mL of KOH solution. The endpoint of the titration is reached after the addition of 27.13 mL of HCl. What is the concentration of the KOH solution?
 - 9.000M
 - 1.09M
 - 0.332M
 - 0.0163M



TEST-TAKING TIP

Slow Down! Check to make sure you're answering the question that each problem is posing. Read the questions and every answer choice very carefully. Remember that doing most of the problems and getting them right is always preferable to doing all the problems and getting lots of them wrong.

Ionization Constants and pH Data for Several Weak Organic Acids

Acid	Ionization equation	pH of 1.000M solution	K_a
Formic	$HCHO_2 \rightleftharpoons H^+ + CHO_2^-$	1.87	1.78×10^{-4}
Cyanoacetic	$HC_3H_2NO_2 \rightleftharpoons H^+ + C_3H_2NO_2^-$?	3.55×10^{-3}
Propanoic	$HC_3H_5O_2 \rightleftharpoons H^+ + C_3H_5O_2^-$	2.43	?
Lutidinic	$H_2C_7H_3NO_4 \rightleftharpoons H^+ + C_7H_4NO_4^-$	1.09	7.08×10^{-3}
Barbituric	$HC_4H_3N_2O_3 \rightleftharpoons H^+ + C_4H_3N_2O_3^-$	2.01	9.77×10^{-5}

Redox Reactions

What You'll Learn

- ▶ You will examine the processes of oxidation and reduction in electron-transfer reactions.
- ▶ You will discover how oxidation numbers of elements in compounds are determined and how they relate to electron transfer.
- ▶ You will separate redox reactions into their oxidation and reduction processes.
- ▶ You will use two different methods to balance oxidation–reduction equations.

Why It's Important

Oxidation and reduction reactions are among the most prevalent in chemistry. From natural phenomena to commercial manufacturing, redox reactions play a major role in your daily life.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about the chemistry of oxidation–reduction reactions.

When threatened, the bombardier beetle sprays chemicals from its abdomen that, when combined, undergo an oxidation–reduction reaction. The result is a boiling-hot, foul-smelling "bomb" that allows the beetle to escape predators.



DISCOVERY LAB



Materials

test tube
iron nail
steel wool or sandpaper
1M copper(II) sulfate (CuSO_4)

Observing an Oxidation–Reduction Reaction

Rust is the result of a reaction of iron and oxygen. Iron nails can also react with substances other than oxygen, as you will find out in this experiment.

Safety Precautions



Always wear safety goggles and an apron in the laboratory.

Procedure

1. Use a piece of steel wool to polish the end of an iron nail.
2. Add about 3 mL 1.0M CuSO_4 to a test tube. Place the polished end of the nail into the CuSO_4 solution. Let stand and observe for about 10 minutes. Record your observations.

Analysis

What is the substance found clinging to the nail? What happened to the color of the copper(II) sulfate solution? Write the balanced chemical equation for the reaction you observed.

Section

20.1

Oxidation and Reduction

Objectives

- **Describe** the processes of oxidation and reduction.
- **Identify** oxidizing and reducing agents.
- **Determine** the oxidation number of an element in a compound.
- **Interpret** redox reactions in terms of change in oxidation state.

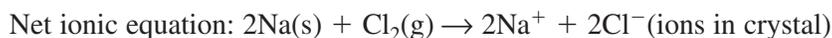
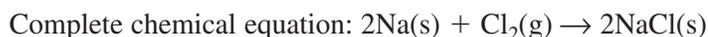
Vocabulary

oxidation–reduction reaction
redox reaction
oxidation
reduction
oxidizing agent
reducing agent

In Chapter 10, you learned that a chemical reaction can usually be classified as one of five types—synthesis, decomposition, combustion, single-replacement, or double-replacement. In this chapter, you'll investigate a special characteristic of many of these reactions—the ability of elements to gain or lose electrons when they react with other elements. You experimented with this characteristic when you did the **DISCOVERY LAB**.

Electron Transfer and Redox Reactions

One of the defining characteristics of single-replacement and combustion reactions is that they always involve the transfer of electrons from one atom to another. So do many, but not all, synthesis and decomposition reactions. For example, you studied the synthesis reaction in which sodium and chlorine react to form the ionic compound sodium chloride.



In this reaction, an electron from each of two sodium atoms is transferred to the Cl_2 molecule to form two Cl^- ions. An example of a combustion reaction is the burning of magnesium in air.

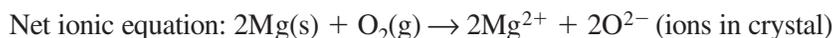
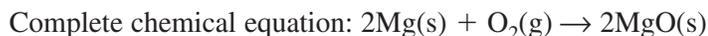
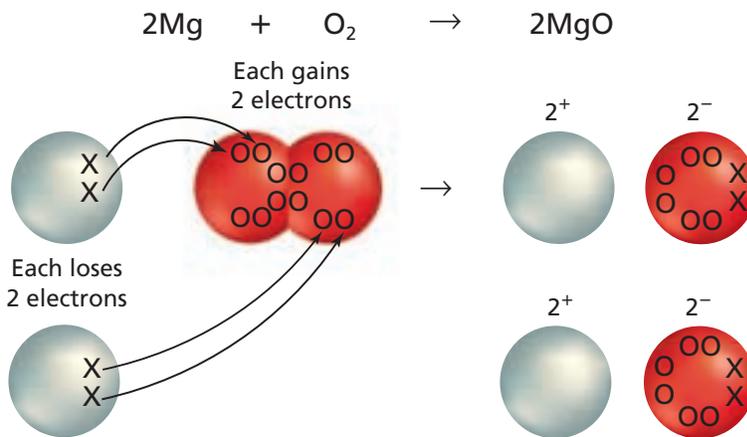




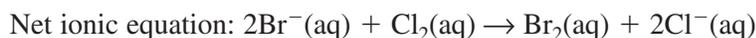
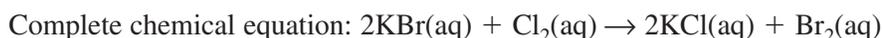
Figure 20-1

The reaction of magnesium and oxygen involves a transfer of electrons from magnesium to oxygen. Therefore, this reaction is an oxidation–reduction reaction. Using the classifications given in Chapter 10, this redox reaction also is classified as a combustion reaction.

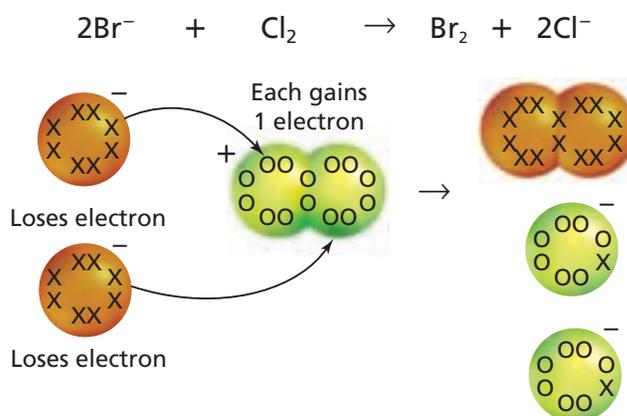


When magnesium reacts with oxygen, as illustrated in **Figure 20-1**, each magnesium atom transfers two electrons to each oxygen atom. What is the result of this electron transfer? The two magnesium atoms become Mg^{2+} ions and the two oxygen atoms become O^{2-} ions (oxide ions). If you compare this reaction with the reaction of sodium and chlorine, you will see that they are alike in that both involve the transfer of electrons between atoms. A reaction in which electrons are transferred from one atom to another is called an **oxidation–reduction reaction**. For simplicity, chemists often refer to oxidation–reduction reactions as **redox reactions**.

Now consider the single-replacement reaction in which chlorine in an aqueous solution replaces bromine from an aqueous solution of potassium bromide, which is shown in **Figure 20-2**.



Note that chlorine “steals” electrons from bromide ions to become chloride ions. When the bromide ions lose their extra electrons, the two bromine atoms form a covalent bond with each other to produce Br_2 molecules. The result of this reaction, the characteristic color of elemental bromine in solution, is shown in **Figure 20-2**. The formation of the covalent bond by sharing of electrons also is an oxidation–reduction reaction.

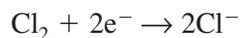


How do oxidation and reduction differ? Originally, the word *oxidation* referred only to reactions in which a substance combined with oxygen, such as the burning of magnesium in air or the burning of natural gas (methane, CH₄) in air. Today, **oxidation** is defined as the loss of electrons from atoms of a substance. Look again at the net ionic equation for the reaction of sodium and chlorine. Sodium is oxidized because it loses an electron. To state this reaction more clearly,



For oxidation to take place, the electrons lost by the substance that is oxidized must be accepted by atoms or ions of another substance. In other words, there must be an accompanying process that involves the gain of electrons.

Reduction is defined as the gain of electrons by atoms of a substance. Following our sodium chloride example further, the reduction reaction that accompanies the oxidation of sodium is the reduction of chlorine.



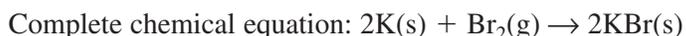
Can oxidation occur without reduction? By our definitions, oxidation and reduction are complementary processes; oxidation cannot occur unless reduction also occurs.

It is important to recognize and distinguish between oxidation and reduction. The following memory aid may help.

LEO the lion says **GER** or, for short, **LEO GER**

This phrase will help you remember that **L**oss of **E**lectrons is **O**xidation, and **G**ain of **E**lectrons is **R**eduction.

Changes in oxidation number You may recall from previous chapters that the *oxidation number* of an atom in an ionic compound is the number of electrons lost or gained by the atom when it forms ions. For example, look at the following equation for the redox reaction of potassium metal with bromine vapor.



Potassium, a group 1A element that tends to lose one electron in reactions because of its low electronegativity, is assigned an oxidation number of +1. On the other hand, bromine, a group 7A element that tends to gain one electron in reactions because of its high electronegativity, is assigned an oxidation number of -1. In redox terms, you would say that potassium atoms are oxidized from 0 to the +1 state because each loses an electron, and bromine atoms are reduced from 0 to the -1 state because each gains an electron. Can you see why the term reduction is used? When an atom or ion is reduced, the numerical value of its oxidation number is lowered.

Oxidation numbers are tools that scientists use in written chemical equations to help them keep track of the movement of electrons in a redox reaction. Like some of the other tools you have learned about in chemistry, oxidation numbers have a specific notation. Oxidation numbers are written with the positive or negative sign before the number (+3, +2), whereas ionic charge is written with the sign after the number (3+, 2+).

Oxidation number: +3

Ionic charge: 3+

Biology

CONNECTION

What do the bacteria *Xenorhabdus luminescens*, fireflies, and many deep-sea fish have in common? These and other organisms emit light. Bioluminescence is the conversion of potential energy in chemical bonds into light during a redox reaction. Bioluminescent light is a cold light that gives off little heat. Depending on the species, bioluminescence is produced by different chemicals and by different means. In fireflies, for example, light results from the oxidation of the molecule luciferin.

Scientists are still unraveling the mystery of bioluminescence. Some luminescent organisms emit light all the time, whereas others emit light only when they are disturbed, such as when churned by ocean waves. Deep-sea fish and some jellyfish appear to be able to control the light they emit, and one species of mushroom is known to emit light of two different colors. Zoologists have also determined that some light-emitting organisms do not produce light themselves, but produce light by harboring bioluminescent bacteria.

Organisms appear to use bioluminescence for different purposes. In the ocean depths, bioluminescence probably aids vision and recognition. Other purposes might include mating and defense against prey.



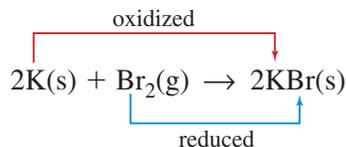


Figure 20-3

A redox reaction involving silver and environmental sulfides deposited tarnish on the discolored cup. Another redox reaction, essentially the reverse of the first, reduces silver ions in the tarnish to silver atoms that can be rinsed away, leaving the shiny silver cup.

Oxidizing and Reducing Agents

Chemists also describe the potassium–bromine reaction in another way by saying that “potassium is oxidized by bromine.” This description is useful because it clearly identifies both the substance that is oxidized and the substance that does the oxidizing. The substance that oxidizes another substance by accepting its electrons is called an **oxidizing agent**. This term is another way of saying “the substance that is reduced.” The substance that reduces another substance by losing electrons is called a **reducing agent**. A reducing agent supplies electrons to the substance getting reduced (gaining electrons), and is itself oxidized because it loses electrons. By this definition, the reducing agent in the potassium–bromine reaction is potassium, the substance that is oxidized.



A common application of redox chemistry is to remove tarnish from metal objects, such as the silver cups in **Figure 20-3**. The **miniLAB** below describes this tarnish removal technique. Other oxidizing and reducing agents such as those shown in **Figure 20-4** play significant roles in your daily life. For example, when you add chlorine bleach to your laundry to whiten clothes, you are using an aqueous solution of sodium hypochlorite (NaClO), an oxidizing agent. It oxidizes dyes, stains, and other materials that discolor clothes. Hydrogen peroxide (H_2O_2) can be used as an antiseptic because it oxidizes some of the vital biomolecules of germs, or as an agent to lighten hair because it oxidizes the dark pigment of the hair.

miniLAB

Cleaning by Redox

Applying Concepts The tarnish on silver is silver sulfide, which is formed when the silver reacts with sulfide compounds in the environment. In this miniLAB you will use an oxidation–reduction reaction to remove the tarnish from silver or a silver-plated object.

Materials aluminum foil, steel wool, small tarnished silver object, 400-mL beaker (or size large enough to hold the tarnished object), baking soda, table salt, hot plate, beaker tongs

Procedure



1. Buff a piece of aluminum foil lightly with steel wool to remove any oxide coating.
2. Wrap the tarnished object in the aluminum foil, making sure that the tarnished area makes firm contact with the foil.
3. Place the wrapped object in the beaker and add sufficient tap water to cover.

4. Add about 1 spoonful of baking soda and about 1 spoonful of table salt.
5. Set the beaker and contents on a hot plate and heat until the water is nearly boiling. Maintain the heat approximately 15 min until the tarnish disappears.

Analysis

1. Write the equation for the reaction of silver with hydrogen sulfide, yielding silver sulfide and hydrogen.
2. Write the equation for the reaction of the tarnish (silver sulfide) with the aluminum foil, yielding aluminum sulfide and silver.
3. Which metal, aluminum or silver, is more reactive? How do you know this from your results?
4. Why should you not use an aluminum pan to clean silver objects by this method?



Figure 20-4

Each of these images illustrates a common use of redox chemistry.

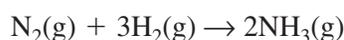
a Hydrogen peroxide sanitizes wounds by a redox reaction that kills germs and bacteria.

b Photography also uses a series of redox reactions in the image capture and development processes. The feature **How It Works** at the end of this chapter describes how redox reactions are used in photography.

c Chlorine is a strong oxidizer that is used in chlorine bleach to whiten laundry.

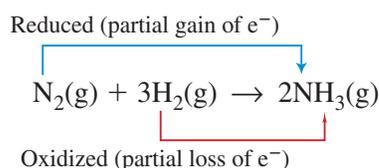
Redox and Electronegativity

The chemistry of oxidation–reduction reactions is not limited to atoms of an element changing to ions or the reverse. Some redox reactions involve changes in molecular substances or polyatomic ions in which atoms are covalently bonded to other atoms. For example, the following equation represents the redox reaction used to manufacture ammonia.



This process involves neither ions nor any obvious transfer of electrons. The reactants and products are all molecular compounds. Still, it is a redox reaction in which nitrogen is the oxidizing agent and hydrogen is the reducing agent.

In situations such as the formation of NH_3 where two atoms share electrons, how is it possible to say that one atom lost electrons and was oxidized while the other atom gained electrons and was reduced? The answer is that you have to know which atom attracts electrons more strongly, or, in other words, which atom is more electronegative. You might find it helpful to review the discussion of electronegativity trends in Chapters 6 and 8. **Table 20-1** gives the specific electronegativities for hydrogen and nitrogen.



For the purpose of studying oxidation–reduction reactions, the more electronegative atom (nitrogen) is treated as if it had been reduced by gaining electrons from the other atom. The less electronegative atom (hydrogen) is treated as if it had been oxidized by losing electrons.

Table 20-1

Electronegativity of Hydrogen and Nitrogen	
Hydrogen	2.20
Nitrogen	3.04

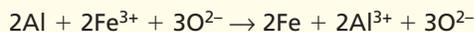


Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

EXAMPLE PROBLEM 20-1

Identifying Oxidation–Reduction Reactions

The following equation represents the redox reaction of aluminum and iron.



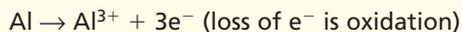
Identify what is oxidized and what is reduced in this reaction. Identify the oxidizing agent and the reducing agent.

1. Analyze the Problem

You are given the ions in the reaction. Using this information, you must determine the electron transfers that take place. Then you can apply the definitions of oxidizing agent and reducing agent to answer the question.

2. Solve for the Unknown

Identify the oxidation process and the reduction process by evaluating the electron transfer. In this case, aluminum loses three electrons and becomes an aluminum ion in the oxidation process. The iron ion accepts the three electrons lost from aluminum in the reduction process.



Aluminum is oxidized and therefore is the reducing agent. Iron is reduced and therefore is the oxidizing agent.

3. Evaluate the Answer

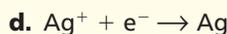
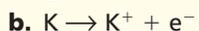
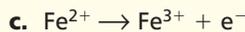
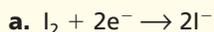
In this process aluminum lost electrons and is oxidized, whereas iron gained electrons and is reduced. The definitions of oxidation and reduction and oxidizing agent and reducing agent apply. Note that the oxidation number of oxygen is unchanged in this reaction; therefore, oxygen is not a key factor in this problem.

Practice!

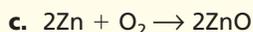
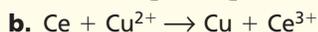
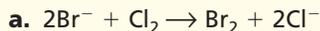
For more practice with identifying oxidation-reduction reactions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

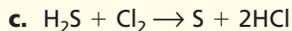
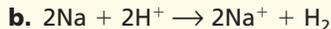
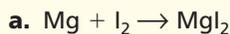
1. Identify each of the following changes as either oxidation or reduction. Recall that e^- is the symbol for an electron.



2. Identify what is oxidized and what is reduced in the following processes.



3. Identify the oxidizing agent and the reducing agent in each of the following reactions.



Determining Oxidation Numbers

In order to understand all kinds of redox reactions, you must have a way to determine the oxidation number of the atoms involved in the reaction. Chemists use a set of rules to make this determination easier.

Rules for determining oxidation numbers

1. The oxidation number of an uncombined atom is zero. This is true for elements that exist as polyatomic molecules such as O_2 , Cl_2 , H_2 , N_2 , S_8 .
2. The oxidation number of a monatomic ion is equal to the charge on the ion. For example, the oxidation number of a Ca^{2+} ion is +2, and the oxidation number of a Br^- ion is -1.
3. The oxidation number of the more electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion. In ammonia (NH_3), for example, nitrogen is more electronegative than hydrogen, meaning that it attracts electrons more strongly than does hydrogen. So nitrogen is assigned an oxidation number of -3, as if it had gained three electrons to complete an octet. In the compound silicon tetrachloride ($SiCl_4$), chlorine is more electronegative than silicon, so each chlorine has an oxidation number of -1 as if it had taken an electron from silicon. The silicon atom is given an oxidation number of +4 as if it had lost electrons to the four chlorine atoms.
4. The most electronegative element, fluorine, always has an oxidation number of -1 when it is bonded to another element.
5. The oxidation number of oxygen in compounds is always -2, except in peroxides, such as hydrogen peroxide (H_2O_2), where it is -1. When it is bonded to fluorine, the only element more electronegative than oxygen, the oxidation number of oxygen is +2.
6. The oxidation number of hydrogen in most of its compounds is +1. The exception to this rule occurs when hydrogen is bonded to less electronegative metals to form hydrides such as LiH , NaH , CaH_2 , and AlH_3 . In these compounds, hydrogen's oxidation number is -1 because it attracts electrons more strongly than does the metal atom.
7. The metals of groups 1 and 2 and aluminum in group 3A form compounds in which the metal atom always has a positive oxidation number equal to the number of its valence electrons (+1, +2, and +3, respectively).
8. The sum of the oxidation numbers in a neutral compound is zero. Notice how the oxidation numbers add up to zero in the following examples.



9. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge on the ion. The following examples illustrate.



Many elements other than those specified in the rules above, including most of the transition metals, metalloids, and nonmetals, can be found with different oxidation numbers in different compounds. For example, the two copper compounds and the two chromium compounds shown in **Figure 20-5 a** and **b**, respectively, have different oxidation numbers.

Careers Using Chemistry

Photochemical Etching Artist

Does the idea of using photographic equipment and chemicals as art appeal to you? Then you might want to investigate a career in photochemical etching.

Photochemical etching was first used to produce precise machine parts, such as microchips and other electronic devices. Now artists use it to create metal art. In this process, ultraviolet light is used to transfer a pattern onto a piece of metal. Then chemicals are applied to remove certain areas in the pattern, creating an intricate design on the metal.

Figure 20-5

The difference in oxidation number gives compounds different chemical and physical properties.

a The oxidation number of copper in copper(I) oxide (orange color) is +1, whereas the oxidation number of copper in copper(II) oxide (black color) is +2.

b Likewise, the oxidation number of chromium in chromium(II) chloride tetrahydrate (blue solution) differs from that in chromium(III) chloride hexahydrate (green solution).



EXAMPLE PROBLEM 20-2

Determining Oxidation Numbers

Determine the oxidation number of each element in the following compounds or ions.

- KClO_3 (potassium chlorate)
- SO_3^{2-} (sulfite ion)

1. Analyze the Problem

From the rules for determining oxidation numbers, you are given the oxidation number of oxygen (-2) and potassium (group 1 metal, $+1$). You are also given the overall charge of the compound or ion. Using this information and applying the rules, you can determine the oxidation numbers of chlorine and sulfur.

2. Solve for the Unknown

Strategy

Assign the known oxidation numbers to their elements, set the sum of all oxidation numbers to zero or to the ion charge, and solve for the unknown oxidation number. (Let n_{element} = oxidation number of the element in question.)

Solution

- Potassium chlorate is a neutral salt, so oxidation numbers must add up to zero. According to rule 5, the oxidation number of oxygen in compounds is -2 . Rule 7 states that Group 1 metals have a $+1$ oxidation number in compounds.

$$(+1) + (n_{\text{Cl}}) + 3(-2) = 0$$



$$1 + n_{\text{Cl}} + (-6) = 0$$

$$n_{\text{Cl}} = +5$$

- Sulfite ion has a charge of $2-$, so oxidation numbers must add up to -2 . According to rule 5, the oxidation number of oxygen in compounds is -2 .

$$(n_{\text{S}}) + 3(-2) = -2$$



$$n_{\text{S}} + (-6) = -2$$

$$n_{\text{S}} = +4$$

3. Evaluate the Answer

The rules for determining oxidation numbers have been correctly applied. All of the oxidation numbers in each substance add up to the proper value.



Among other applications, potassium chlorate is used in explosives, fireworks, and matches.

Practice!

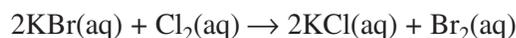
For more practice with determining oxidation numbers, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

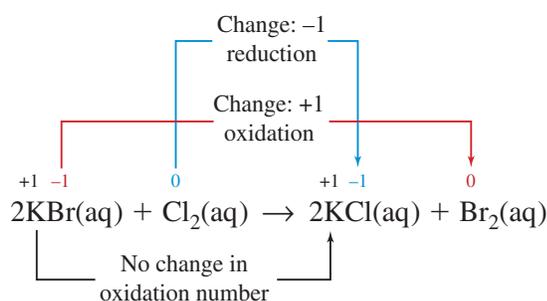
- Determine the oxidation number of the boldface element in the following formulas for compounds.
 - NaClO_4
 - AlPO_4
 - HNO_2
- Determine the oxidation number of the boldface element in the following formulas for ions.
 - NH_4^+
 - AsO_4^{3-}
 - CrO_4^{2-}

Oxidation Number in Redox Reactions

Having studied oxidation numbers, you are now able to relate oxidation–reduction reactions to changes in oxidation number. Look again at the equation for a reaction that you saw at the beginning of this section, the replacement of bromine in aqueous KBr by Cl₂.



To see how oxidation numbers change, start by assigning numbers, shown in **Table 20-2**, to all elements in the balanced equation. Then review the changes as shown in the accompanying diagram.



As you can see, the oxidation number of bromine changed from -1 to 0 , an increase of 1 . At the same time, the oxidation number of chlorine changed from 0 to -1 , a decrease of 1 . Therefore, chlorine is reduced and bromine is oxidized. All redox reactions follow the same pattern. When an atom is oxidized, its oxidation number increases. When an atom is reduced, its oxidation number decreases. Note that there is no change in the oxidation number of potassium. The potassium ion takes no part in the reaction and is called a spectator ion. How would the reaction differ if you used zinc bromide (ZnBr₂) instead of potassium bromide?

Table 20-2

Oxidation Number Assignment		
Element	Oxidation number	Rule
K in KBr	+1	7
Br in KBr	-1	8
Cl in Cl ₂	0	1
K in KCl	+1	7
Cl in KCl	-1	8
Br in Br ₂	0	1



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Section 20.1 Assessment

- Describe the processes of oxidation and reduction.
- Explain the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?
- Determine the oxidation number of the boldface element in these compounds.
 - Sb**₂O₅
 - H**NO₃
 - Ca**N**₂
 - Cu**W**O₄ (copper(II) tungstate)
- Determine the oxidation number of the boldface element in these ions.
 - I**O₄⁻
 - Mn**O**₄⁻
 - B**₄O₇²⁻
 - N**H₂⁻
- Thinking Critically** Write the equation for the reaction of iron metal with hydrobromic acid to form iron(III) bromide and hydrogen gas. Determine which element is reduced and which element is oxidized in this reaction.
- Making Predictions** Alkali metals are strong reducing agents. Would you predict that their reducing ability would increase or decrease as you move down the family from sodium to francium? Give reasons for your prediction.

Objectives

- **Relate** changes in oxidation numbers to the transfer of electrons.
- **Use** changes in oxidation number to balance redox equations.
- **Balance** net ionic redox equations by the oxidation-number method.

Vocabulary

oxidation-number method

You already know that chemical equations are written to represent chemical reactions by showing what substances react and what products are formed. You also know that chemical equations must be balanced to show the correct quantities of reactants and products. Equations for oxidation–reduction reactions are no different. In this section, you’ll learn a specific method to balance redox equations.

The Oxidation-Number Method

Many equations for redox reactions are easy to balance. For example, try to balance the following unbalanced equation for the redox reaction that occurs when potassium chlorate is heated and decomposes to produce potassium chloride and oxygen gas.



Did you get the correct result shown below?



Equations for other redox reactions are not as easy to balance. Study the following unbalanced equation for the reaction that occurs when copper metal is placed in concentrated nitric acid. This reaction is shown in **Figure 20-6**. The brown gas you see is nitrogen dioxide (NO_2), produced by the reduction of nitrate ions (NO_3^-), and the blue solution is the result of the oxidation of copper to copper(II) ions.



Note that oxygen appears in only one reactant, HNO_3 , but in all three products. Nitrogen appears in HNO_3 and in two of the products. Redox equations such as this one, in which the same element appears in several reactants and products, can be hard to balance by the conventional method. For this reason, you need to learn a different balancing technique for redox equations that is based on the fact that the number of electrons transferred from atoms must equal the number of electrons accepted by other atoms.

As you know, when an atom loses electrons, its oxidation number increases; when an atom gains electrons, its oxidation number decreases. Therefore, the total increase in oxidation numbers (oxidation) must equal the total decrease in oxidation numbers (reduction) of the atoms involved in the reaction. The balancing technique called the **oxidation-number method** is based on these principles. The **CHEMLAB** at the end of this chapter gives you the opportunity to perform and balance the copper–nitric acid redox reaction.



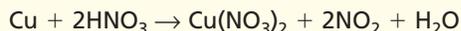
Figure 20-6

Some chemical equations for redox reactions, such as this reaction between copper and nitric acid, can be difficult to balance because one or more elements may appear several times on both sides of the equation. The oxidation-number method makes it easier to balance the equation for this redox reaction.

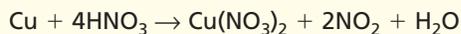
Steps for balancing redox equations by the oxidation-number method

1. Assign oxidation numbers to all atoms in the equation.
2. Identify the atoms that are oxidized and the atoms that are reduced.
3. Determine the change in oxidation number for the atoms that are oxidized and for the atoms that are reduced.
4. Make the change in oxidation numbers equal in magnitude by adjusting coefficients in the equation.

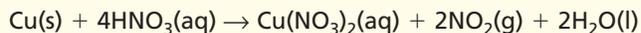
Step 5. Use the conventional method to balance the number of atoms and the charges in the remainder of the equation. This equation shows no charged species, so charge balance need not be considered here.



The coefficient of HNO_3 must be increased from 2 to 4 to balance the four nitrogen atoms in the products.



Add a coefficient of 2 to H_2O to balance the four hydrogen atoms on the left.



3. Evaluate the Answer

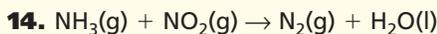
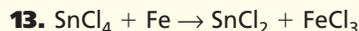
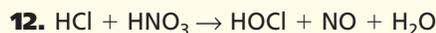
The number of atoms of each element is equal on both sides of the equation. No subscripts have been changed.



For more practice balancing redox reactions, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Use the oxidation-number method to balance these redox equations.

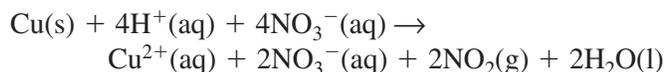


Balancing Net Ionic Redox Equations

Sometimes chemists prefer to express redox reactions in the simplest possible terms—as an equation showing only the oxidation and reduction processes and nothing else. Look again at the balanced equation for the oxidation of copper by nitric acid.



Note that the reaction takes place in aqueous solution, so HNO_3 , which is a strong acid, will be ionized. Likewise, copper(II) nitrate ($\text{Cu}(\text{NO}_3)_2$) will be dissociated into ions. Therefore, the equation can also be written in ionic form.



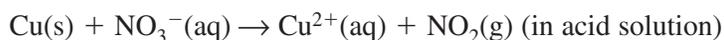
You can see that there are four nitrate ions among the reactants, but only two of them undergo change to form two nitrogen dioxide molecules. What is the role of the other nitrate ions? The other two are only spectator ions and can be eliminated from the equation. To simplify things when writing redox equations in ionic form, chemists usually indicate hydrogen ions by H^+ with the understanding that they exist in hydrated form as hydronium ions (H_3O^+). The equation can then be rewritten showing only the substances that undergo change.



Now look at the equation in unbalanced form.



You might even see this same reaction expressed in a way that shows only the substances that are oxidized and reduced.



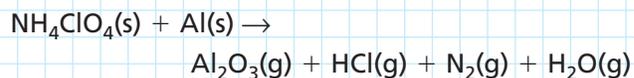
In this case, the hydrogen ion and the water molecule are eliminated because neither is oxidized nor reduced. The only additional information needed is that the reaction takes place in acid solution. In acid solution, hydrogen ions (H^+) and water molecules are abundant and free to participate in redox reactions as either reactants or products. Some redox reactions can occur only in basic solution. When you balance equations for these reactions, you may add hydroxide ions (OH^-) and water molecules to either side of the equation. Basic solutions have an abundance of OH^- ions instead of H_3O^+ ions.

If you try to balance the equation given above, you will see that it appears impossible. Net ionic equations can still be balanced, though, by applying the oxidation-number method, as you will see in Example Problem 20-4. The **problem-solving LAB** below shows how the oxidation-number method can be used for a real-world application.

problem-solving LAB

How does redox lift a space shuttle?

Using Numbers The space shuttle gains nearly 72% of its lift from its solid rocket boosters (SRBs) during the first two minutes of launch. The two pencil-shaped SRB tanks are attached to both sides of the liquid hydrogen and oxygen fuel tank. Each SRB contains 495 000 kg of an explosive mixture of ammonium perchlorate and aluminum. The unbalanced equation for the reaction is given below.



Use the oxidation-number method to balance this redox equation.

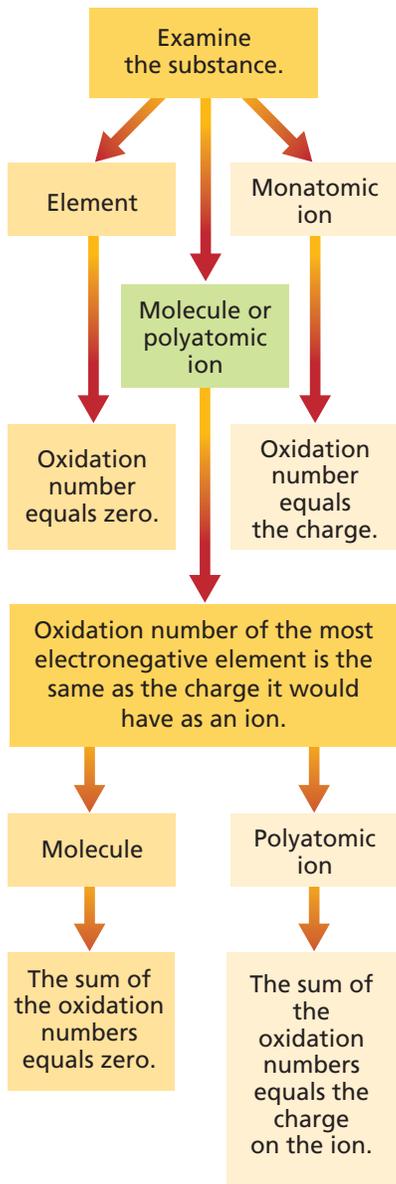
Analysis

The SRB reaction can be balanced using the rules for assigning oxidation numbers. In this way, you can ensure that the oxidation reaction balances the reduction reaction. Write the oxidation number above each element in the equation. Which elements are reduced and which are oxidized? What coefficients are required to balance the reaction?

Thinking Critically

What are some of the benefits of using SRBs for the first two minutes of launch?



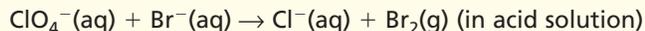


Use this flow chart to review the rules for assigning oxidation numbers listed on page 641.

EXAMPLE PROBLEM 20-4

Balancing a Net Ionic Redox Equation

Use the oxidation-number method to balance this net ionic redox equation for the reaction between the perchlorate ion and the bromide ion in acid solution.

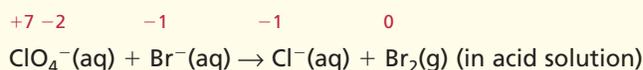


1. Analyze the Problem

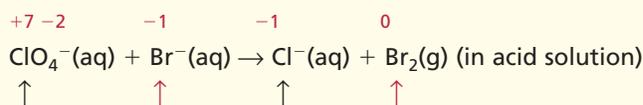
You are given the formulas for the reactants and products, and you have the rules for determining oxidation number. You also know that the increase in oxidation number of the oxidized atoms must equal the decrease in oxidation number of the reduced atoms. You are told that the reaction takes place under acidic conditions. With this information you can adjust the coefficients to balance the equation.

2. Solve for the Unknown

Step 1. Apply rules to assign oxidation numbers to all atoms in the equation using the flow chart on the left.

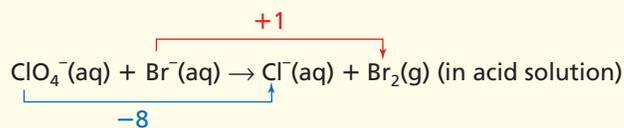


Step 2. Identify which atoms are oxidized and which are reduced.

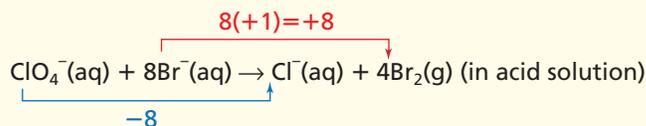


The oxidation number of bromine increases from -1 to 0 as it is oxidized. The oxidation number of chlorine decreases from $+7$ to -1 as it is reduced. Note that no oxygen atoms appear in the products. This deficiency will be fixed shortly.

Step 3. Draw a line connecting the atoms involved in oxidation and another line connecting the atoms involved in reduction. Write the net change in oxidation number above or below each line.

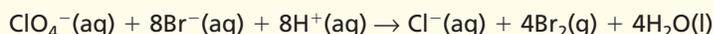


Step 4. Make the changes in oxidation number equal in magnitude by placing the appropriate coefficients in front of the formulas in the equation.



Note that 4Br_2 represents eight bromine atoms to balance the 8Br^- on the left side.

Step 5. Add enough hydrogen ions and water molecules to the equation to balance the oxygen atoms on both sides. This is the reason that you must know the conditions of the reaction.



Four oxygen atoms are present in one ClO_4^- ion, enough to form four H_2O molecules. Therefore, a total of 8H^+ ions will also be needed to form the four water molecules.

3. Evaluate the Answer

Review the balanced equation, counting the number of atoms to be sure there are equal numbers of each element. As with any ionic equation, you must also check to see if the net charge on the right equals the net charge on the left. Check to be sure that you did not change any subscripts in the equation.

PRACTICE PROBLEMS

Use the oxidation-number method to balance the following net ionic redox equations.

15. $\text{H}_2\text{S}(\text{g}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{S}(\text{s}) + \text{NO}(\text{g})$ (in acid solution)
16. $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{I}_2(\text{s})$ (in acid solution)
17. $\text{I}^-(\text{aq}) + \text{MnO}_4^-(\text{aq}) \rightarrow \text{I}_2(\text{s}) + \text{MnO}_2(\text{s})$ (in basic solution) Hint: Hydroxide ions will appear on the right, and water molecules on the left.



For more practice balancing redox reactions, go to **Supplemental Practice Problems** in Appendix A.

The oxidation-number method is convenient for balancing most redox equations, but you will see in the next section that there are occasions when you must balance the net ionic charge on both sides of the equation in addition to balancing the atoms.

Section 20.2 Assessment

18. A reactant in an oxidation–reduction reaction loses four electrons when it is oxidized. How many electrons must be gained by the reactant that is reduced?
19. Why is it important to know the conditions under which an aqueous redox reaction takes place in order to balance the ionic equation for the reaction?
20. Balance these equations for redox reactions by using the oxidation-number method.
 - a. $\text{HClO}_3(\text{aq}) \rightarrow \text{ClO}_2(\text{g}) + \text{HClO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - b. $\text{H}_2\text{O}_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) + \text{FeSO}_4(\text{aq}) \rightarrow \text{Fe}_2(\text{SO}_4)_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
(Hint: Review rule 5 to learn how to determine the oxidation numbers in H_2O_2 .)
 - c. $\text{H}_2\text{SeO}_3(\text{aq}) + \text{HClO}_3(\text{aq}) \rightarrow \text{H}_2\text{SeO}_4(\text{aq}) + \text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
21. Balance these net ionic equations for redox reactions.
 - a. $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Cr}^{3+}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$ (in acid solution)
 - b. $\text{Zn}(\text{s}) + \text{V}_2\text{O}_5(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{V}_2\text{O}_4(\text{aq})$ (in acid solution)
 - c. $\text{N}_2\text{O}(\text{g}) + \text{ClO}^-(\text{aq}) \rightarrow \text{Cl}^-(\text{aq}) + \text{NO}_2^-(\text{aq})$ (in basic solution)
22. **Thinking Critically** Explain how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?
23. **Applying Concepts** The processes used to remove metals from their ores usually involve a reduction process. For example, mercury may be obtained by roasting the ore mercury(II) sulfide with calcium oxide to produce mercury metal, calcium sulfide, and calcium sulfate. Write and balance the redox equation for this process.

Objectives

- **Recognize** the interdependence of oxidation and reduction processes.
- **Derive** oxidation and reduction half-reactions from a redox equation.
- **Balance** redox equations by the half-reaction method.

Vocabulary

species
half-reaction

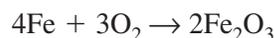
Throughout this chapter, you have read about oxidation–reduction reactions. You know that redox reactions involve the loss and gain of electrons. Thus, the “pairing” or complementary nature of redox reactions is probably apparent to you. So, let’s consider the two halves of redox reactions.

Identifying Half-Reactions

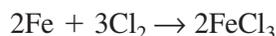
As you know, oxidation–reduction reactions can involve molecules, ions, free atoms, or combinations of all three. To make it easier to discuss redox reactions without constantly specifying the kind of particle involved, chemists use the term **species**. In chemistry, a **species** is any kind of chemical unit involved in a process. For example, a solution of sugar in water contains two major species. In the equilibrium equation $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$, there are four species: the two molecules NH_3 and H_2O and the two ions NH_4^+ and OH^- .

Oxidation–reduction reactions take place whenever a species that can give up electrons (reducing agent) comes in contact with another species that can accept them (oxidizing agent). In a sense, the species being reduced doesn’t “care” where the electrons are coming from as long as they are readily available. Likewise, the species undergoing oxidation doesn’t “care” where its electrons go as long as a willing receiver is available. The cartoon electron donor/acceptor machine shown in **Figure 20-7** illustrates this principle.

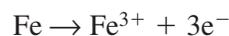
For example, active metals can be oxidized by a variety of oxidizing agents. When iron rusts, it is oxidized by oxygen. You can also say that iron reduces oxygen because iron acts as a reducing agent.



But iron can reduce many other oxidizing agents, including chlorine.



In this reaction, each iron atom is oxidized by losing three electrons to become an Fe^{3+} ion. This is the oxidation half of the oxidation–reduction reaction. Recall that e^- represents one electron.



At the same time, each chlorine atom in Cl_2 is reduced by accepting one electron to become a Cl^- ion. This is the reduction half of the oxidation–reduction reaction.

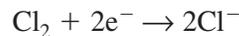


Figure 20-7

This imaginary electron machine illustrates the point that the electrons that are involved in redox reactions can come from any source, as long as they are available to transfer. It also points out that reduction, the complimentary process of oxidation, requires something to accept the electrons. These two processes can be identified and separated to help understand redox reactions and balance redox equations.

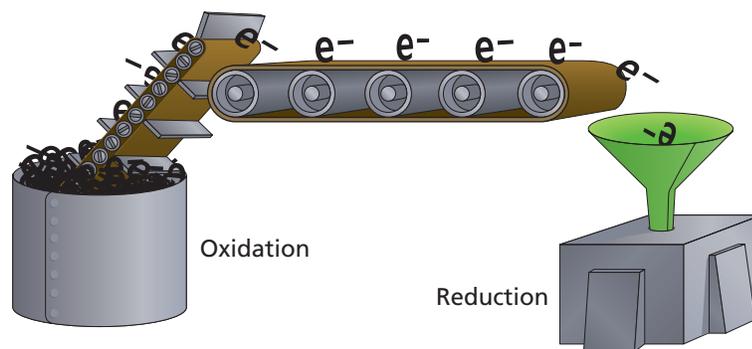


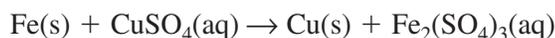
Table 20-3

Various Oxidation–Reduction Reactions in which Iron Is Oxidized		
Overall reaction (unbalanced)	Oxidation half-reaction	Reduction half-reaction
$\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{O}_2 + 4\text{e}^- \rightarrow 2\text{O}^{2-}$
$\text{Fe} + \text{Cl}_2 \rightarrow \text{FeCl}_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$
$\text{Fe} + \text{F}_2 \rightarrow \text{FeF}_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$
$\text{Fe} + \text{HBr} \rightarrow \text{H}_2 + \text{FeBr}_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
$\text{Fe} + \text{AgNO}_3 \rightarrow \text{Ag} + \text{Fe}(\text{NO}_3)_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$
$\text{Fe} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{Fe}_2(\text{SO}_4)_3$	$\text{Fe} \rightarrow \text{Fe}^{3+} + 3\text{e}^-$	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

Equations such as these represent half-reactions. A **half-reaction** is one of the two parts of a redox reaction—the oxidation half alone or the reduction half alone. **Table 20-3** shows a variety of reduction half-reactions that can accept electrons from iron, oxidizing it from Fe to Fe^{3+} .

Balancing Redox Equations by Half-Reactions

You will learn more about the importance of half-reactions when you study electrochemistry in Chapter 21. For now, however, you can learn to use half-reactions to balance a redox equation. First, look at an unbalanced equation taken from **Table 20-3** to see how to separate a redox equation into half-reactions. For example, the following unbalanced equation represents the reaction that occurs when you put an iron nail into a solution of copper(II) sulfate, as shown in **Figure 20-8**. Iron atoms are oxidized as they lose electrons to the copper(II) ions.



Steps for balancing by half-reactions

- Write the net ionic equation for the reaction, omitting spectator ions.

$$\text{Fe} + \text{Cu}^{2+} + \cancel{3\text{SO}_4^{2-}} \rightarrow \text{Cu} + 2\text{Fe}^{3+} + \cancel{3\text{SO}_4^{2-}}$$

$$\text{Fe} + \text{Cu}^{2+} \rightarrow \text{Cu} + 2\text{Fe}^{3+}$$
- Write the oxidation and reduction half-reactions for the net ionic equation.

$$\text{Fe} \rightarrow 2\text{Fe}^{3+} + 6\text{e}^-$$

$$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$$
- Balance the atoms and charges in each half-reaction.

$$2\text{Fe} \rightarrow 2\text{Fe}^{3+} + 6\text{e}^-$$

$$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$$
- Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.

$$2\text{Fe} \rightarrow 2\text{Fe}^{3+} + \cancel{6\text{e}^-}$$

$$3\text{Cu}^{2+} + \cancel{6\text{e}^-} \rightarrow 3\text{Cu}$$
- Add the balanced half-reactions and return spectator ions.

$$2\text{Fe} + 3\text{Cu}^{2+} \rightarrow 3\text{Cu} + 2\text{Fe}^{3+}$$

$$2\text{Fe}(\text{s}) + 3\text{CuSO}_4(\text{aq}) \rightarrow 3\text{Cu}(\text{s}) + \text{Fe}_2(\text{SO}_4)_3(\text{aq})$$

Figure 20-8

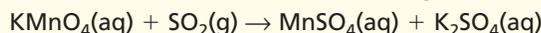
As a result of this redox reaction between iron and copper sulfate solution, solid copper metal is deposited on the iron nail. To balance the equation given in the text for this reaction you could use the method of half-reactions.



EXAMPLE PROBLEM 20-5

Balancing a Redox Equation by Half-Reactions

The permanganate ion (MnO_4^-) is widely used in the chemistry laboratory as a strong oxidizing agent. Permanganate is usually sold as potassium salt KMnO_4 . When sulfur dioxide gas is bubbled into an acidic solution of potassium permanganate, it reacts to form sulfate ions. Balance the redox equation for this reaction using half-reactions.

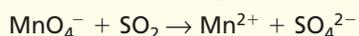


1. Analyze the Problem

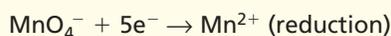
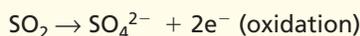
You are given the skeleton equation for the reaction of permanganate and sulfur dioxide. You also know that the reaction takes place in an acid solution. With this information, the rules for determining oxidation numbers, and the steps for balancing by half-reactions, you can write a complete balanced equation.

2. Solve for the Unknown

Step 1. Write the net ionic equation for the reaction. You can eliminate coefficients, spectator ions, and state symbols.

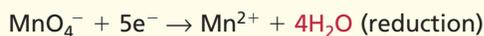
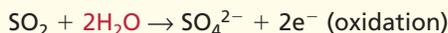


Step 2. Write the oxidation and reduction half-reactions for the net ionic equation, including oxidation numbers. Recall the rules for assigning oxidation numbers from page 641.

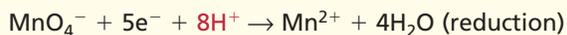
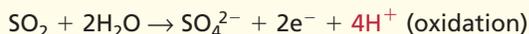


Step 3. Balance the atoms and charges in the half-reactions.

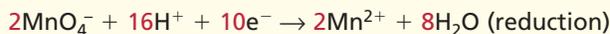
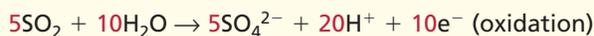
- a. In this case, the sulfur atoms are balanced on both sides of the equation and the manganese atoms are balanced on both sides of the equation. Recall that in acid solution, H_2O molecules are available in abundance and can be used to balance oxygen atoms in the half-reactions.



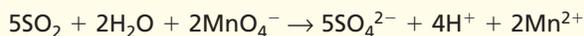
- b. Recall also that in acid solution, H^+ ions are readily available and can be used to balance the charge in the half-reactions.



Step 4. Adjust the coefficients so that the number of electrons lost in oxidation (2) equals the number of electrons gained in reduction (5). In this case, the least common multiple of 2 and 5 is 10. Cross-multiplying gives the balanced oxidation and reduction half-reactions.



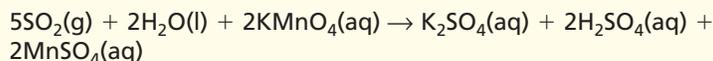
Step 5. Add the balanced half-reactions and simplify by canceling or reducing like terms on both sides of the equation.



Potassium permanganate (KMnO_4) is used to disinfect decorative fish ponds and the water in fish hatcheries.



Finally, return spectator ions (K^+) and restore the state descriptions. In this case, two K^+ ions go with the two MnO_4^- anions on the left side of the equation. Therefore, you can add only two K^+ ions to the right side, which will form K_2SO_4 . The remaining SO_4^{2-} ions recombine with Mn^{2+} and H^+ to form two molecules of $MnSO_4$ and H_2SO_4 , respectively.



3. Evaluate the Answer

A review of the balanced equation indicates that the number of atoms of each element is equal on both sides of the equation. No subscripts have been changed.

PRACTICE PROBLEMS

Use the half-reaction method to balance the following redox equations. Begin with step 2 of Example Problem 20-5, and leave the balanced equation in ionic form.

24. $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)
25. $Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acid solution)
26. $N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution). Hint: Add O and H in the form of OH^- ions and H_2O molecules.



For more practice solving problems using the half-reaction method, go to **Supplemental Practice Problems** in Appendix A.

You have learned two methods by which redox equations can be balanced. Both methods will provide the same results; however, the half-reaction method may be more useful to your study of electrochemistry.

Section 20.3 Assessment

27. Explain why an oxidation process must always accompany a reduction process.
28. Write the oxidation and reduction half-reactions for this redox equation.
 $Pb(s) + Pd(NO_3)_2(aq) \rightarrow Pb(NO_3)_2(aq) + Pd(s)$
29. Use the half-reaction method to balance this redox equation. Begin with step 2 of Example Problem 20-5, and leave the balanced equation in ionic form.
 $AsO_4^{3-}(aq) + Zn(s) \rightarrow AsH_3(g) + Zn^{2+}(aq)$ (in acid solution)
30. **Thinking Critically** The oxidation half-reaction of a redox reaction is $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$ and the reduction half-reaction is $Au^{3+} + 3e^- \rightarrow Au$. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have no electrons left over?
31. **Calculating** The concentration of thallium(I) ions in solution may be determined by oxidizing to thallium(III) ions with an aqueous solution of potassium permanganate ($KMnO_4$) under acidic conditions. Suppose that a 100.00 mL sample of a solution of unknown Tl^+ concentration is titrated to the endpoint with 28.23 mL of a 0.0560M solution of potassium permanganate. What is the concentration of Tl^+ ions in the sample? You must first balance the redox equation for the reaction to determine its stoichiometry.
 $Tl^+(aq) + MnO_4^-(aq) \rightarrow Tl^{3+}(aq) + Mn^{2+}(aq)$

Redox Reactions

In Section 20.2, a redox reaction involving copper and nitric acid is discussed. This reaction is balanced by a method called the oxidation-number method. In this lab, you will carry out this reaction, along with another redox reaction that involves a common household substance. You will practice balancing various redox reactions using both the oxidation-number method (from Section 20.2) and the half-reaction method (from Section 20.3).

Problem

What are some examples of redox reactions and how can the equations describing them be balanced?

Objectives

- **Observe** various redox reactions.
- **Balance** redox reactions using the oxidation-number method.
- **Balance** redox reactions using the half-reaction method.

Materials

copper metal	household ammonia
6M nitric acid	crystal drain cleaner
evaporating dish	thermometer
forceps	250-mL beaker
distilled water	
dropper pipette	
spoon	

Safety Precautions



- The reaction of copper with nitric acid should be done in a ventilation hood. Do not breathe the fumes from this reaction.
- Nitric acid and ammonia can cause burns. Avoid contact with skin and eyes.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.

Data Table	
Step 1	
Step 2	
Step 3	
Step 4	
Step 5	

3. Review what a redox reaction is.
4. Read the label of the crystal drain cleaner package. Understand that the compound is solid sodium hydroxide that contains aluminum. When the material is added to water, sodium hydroxide dissolves rapidly, producing heat. Aluminum reacts with water in the basic solution to produce $\text{Al}(\text{OH})_4^-$ ions

and hydrogen gas. Is aluminum oxidized or reduced in the reaction? Is hydrogen oxidized or reduced in the reaction? Explain your answers.

Procedure

1. In a ventilation hood, place a piece of copper metal in a clean, dry evaporating dish. Add enough 6M nitric acid to cover the metal. **CAUTION:** Nitric acid can cause burns. The reaction of nitric acid with copper generates dangerous fumes. Use a ventilation hood. Observe what happens and record your observations in the data table.
2. Pour about 2 mL of the solution from the evaporating dish into a test tube that contains about 2 mL of distilled water. Add ammonia until a change occurs. Record your observation in the data table.
3. Add about 50 mL of tap water to a 250-mL beaker. Use a thermometer to measure the temperature of the water. Record your observations in the data table.

4. Pour approximately 1 cm³ of dry drain cleaner onto a watchglass. **CAUTION:** *Drain cleaner is caustic and will burn skin. Use forceps to move the crystals and observe their composition. Record your observations in the data table.*



5. Carefully pour about one-half spoonful of the crystals into the water in the beaker. As the crystals react with the water, watch the thermometer in the water for a few minutes and record in the data table the highest temperature reached and any other observations you make.

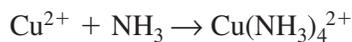
Cleanup and Disposal

- After step 1 is completed, use forceps to remove any excess pieces of copper metal. Rinse the copper metal with tap water and dispose of the metal as your teacher instructs.
- After step 2 is finished, pour the solution down the drain and flush with a lot of water.
- After step 5 is finished, pour the solution down the drain and flush with a lot of water.

Analyze and Conclude

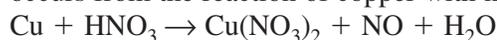
- Applying Concepts** The reaction between copper and nitric acid is discussed in Section 20.2. Write the half-reaction for the substance that is oxidized.
- Applying Concepts** Write the half-reaction for the substance that is reduced.

3. **Thinking Critically** In step 2, a deep blue copper–ammonia complex is formed according to the following reaction.



Is this a redox reaction? Why or why not?

4. **Using Numbers** The following side reaction occurs from the reaction of copper with nitric acid.

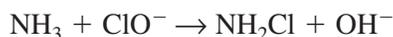


Balance this redox reaction using both the oxidation-number method and the half-reaction method.

5. **Using Numbers** Write and balance the redox reaction of sodium hydroxide with aluminum and water.
6. **Error Analysis** Give possible reasons why you might not have been able to balance the equation for the redox reaction you performed in this experiment.

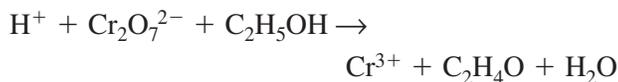
Real-World Chemistry

- Using your observations in this lab, how do drain cleaning crystals remove clogs?
- Ammonia and bleach are two common household chemicals that should never be mixed. One product of this reaction is chloramine, a poisonous, volatile compound. The reaction is as follows.



What is the balanced redox reaction?

3. One type of breathalyzer detects whether ethanol is in the breath of a person. Ethanol is oxidized to acetaldehyde by dichromate ions in acidic solution. The dichromate ion in solution is orange, while the Cr³⁺ aqueous ion is green. The appearance of a green color in the breathalyzer test shows that the breath exceeds the legal limit of alcohol. The equation is



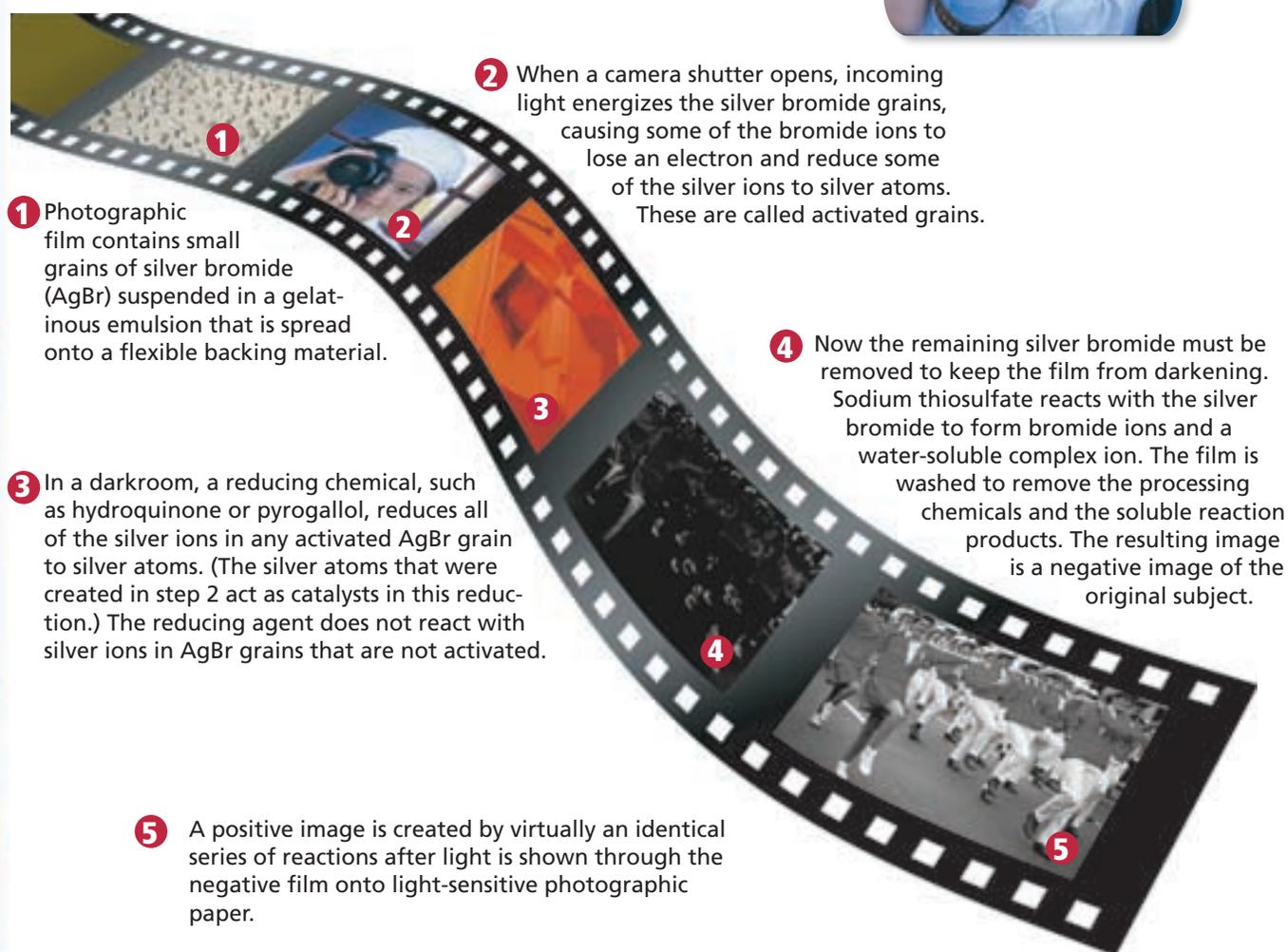
Balance this redox reaction.

4. Diluted hydrochloric acid can be used to remove limestone (calcium carbonate) surrounding phosphate and silicate fossils. The reaction produces carbon dioxide, water, and aqueous calcium chloride. Write the balanced chemical equation. Is it a redox reaction? Explain.

How It Works

Photographic Film

Black and white photography is a popular art form. The spectacular images that are captured on and printed from black and white film are the product of a series of redox reactions. The first redox reaction captures the image on the film inside the camera. The second is a reaction to produce a negative image of the exposed film. The third redox reaction creates the positive print from the negative film image.



Thinking Critically

1. Predicting Time, temperature, and solution concentration are all important factors in the developing process. What would be the consequence of leaving the film too long in the hydroquinone developer? (*Hint: the developer is a reducing agent.*)

2. Hypothesizing Color photographs are composed of dyes formed during the development of the silver image. The silver metal must be removed because it is opaque. Use the redox principle to explain how the silver could be removed from the color photo.

Summary

20.1 Oxidation and Reduction

- An oxidation–reduction (redox) reaction is any chemical reaction in which electrons are transferred from one atom to another. Most chemical reactions other than double replacement reactions are oxidation–reduction reactions.
- Oxidation occurs when an atom loses electrons. Reduction occurs when an atom gains electrons.
- Oxidation increases an atom’s oxidation number and reduction decreases an atom’s oxidation number.
- Oxidation and reduction must always occur together because the reduction process requires a supply of electrons available only from the oxidation process. At the same time, the oxidation process must have a receiver for lost electrons.
- The species that is oxidized in a redox reaction is the reducing agent and loses electrons. The species that is reduced is the oxidizing agent and gains electrons.
- The oxidation number of an element in a compound can be determined by applying a series of rules to the atoms in the compound.

Summary of Oxidation and Reduction Processes

Process	Oxidation	Reduction
Examples	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$ $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$ $\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$
Electron transfer	Atom loses electrons	Atom gains electrons
Change in oxidation number	Increases	Decreases
Function	Reducing agent	Oxidizing agent

20.2 Balancing Redox Equations

- Many simple redox equations may be balanced by inspection. The oxidation-number method can be used to balance more difficult reactions.
- Redox equations are often balanced by examining the change in oxidation number that occurs in the oxidized and reduced species. The following principle is then applied.

e^- lost in oxidation = e^- gained in reduction

 Therefore, the total increase in oxidation numbers equals the total decrease in oxidation numbers.
- Redox reactions that take place in acidic solution often use water molecules and hydrogen ions in the process. Redox reactions that take place in basic solution often use water molecules and hydroxide ions in the process. Therefore, it is appropriate to add these species to the equation in order to balance the numbers of oxygen and/or hydrogen atoms.
- Redox reactions involving ionic species may be represented by net ionic equations, leaving out spectator ions.

20.3 Half-Reactions

- The oxidation and reduction processes of a redox reaction can be represented as half-reactions.
- An oxidation half-reaction shows the number of electrons lost when a species is oxidized. A reduction half-reaction shows the number of electrons gained when a species is reduced.
- The fact that an oxidation process must always be coupled with a reduction process provides a basis for using half-reactions to balance redox equations.

Vocabulary

- half-reaction (p. 651)
- oxidation (p. 637)
- oxidation-number method (p. 644)
- oxidation–reduction reaction (p. 636)
- oxidizing agent (p. 638)
- redox reaction (p. 636)
- reducing agent (p. 638)
- reduction (p. 637)
- species (p. 650)

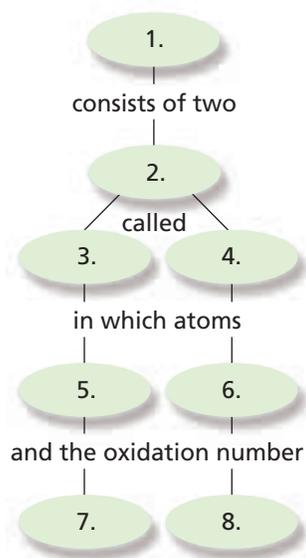


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 20 Assessment.

Concept Mapping

32. Complete the concept map using the following terms: decreases, half-reactions, gain electrons, reduction, lose electrons, redox reaction, oxidation, increases.



Mastering Concepts

33. What is the main characteristic of oxidation–reduction reactions? (20.1)
34. In terms of electrons, what happens when an atom is oxidized? When an atom is reduced? (20.1)
35. What is the oxidation number of alkaline earth metals in their compounds? Of alkali metals? (20.1)
36. How does the oxidation number change when an element is oxidized? When it is reduced? (20.1)
37. Explain why oxidation and reduction must occur simultaneously. (20.1)
38. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer. (20.2)
- $$\text{Fe(s)} + \text{Ag}^+(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Ag(s)}$$
39. How can you tell that the redox equation in question 38 is not balanced? (20.2)
40. How does the change in oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained? (20.2)
41. Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution? (20.3)
42. Does the following equation represent a reduction or an oxidation process? Explain your answer. (20.3)
- $$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$$
43. What term is used for the type of reaction represented in question 42? (20.3)

Mastering Problems

Oxidation and Reduction (20.1)

44. Identify the species oxidized and the species reduced in each of these redox equations.
- $3\text{Br}_2 + 2\text{Ga} \rightarrow 2\text{GaBr}_3$
 - $\text{HCl} + \text{Zn} \rightarrow \text{ZnCl}_2 + \text{H}_2$
 - $\text{Mg} + \text{N}_2 \rightarrow \text{Mg}_3\text{N}_2$
45. Identify the oxidizing agent and the reducing agent in each of these redox equations.
- $\text{H}_2\text{S} + \text{Cl}_2 \rightarrow 2\text{HCl} + \text{S}$
 - $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$
 - $2\text{Na} + \text{I}_2 \rightarrow 2\text{NaI}$
46. Identify each of these half-reactions as either oxidation or reduction.
- $\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$
 - $\text{NO}_2 \rightarrow \text{NO}_3^- + \text{e}^-$
 - $\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$
47. Determine the oxidation number of the bold element in these substances and ions.
- CaCrO_4
 - NaHSO_4
 - NO_2^-
 - BrO_3^-
48. Determine the net change of oxidation number of each of the elements in these redox equations.
- $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$
 - $\text{Cl}_2 + \text{ZnI}_2 \rightarrow \text{ZnCl}_2 + \text{I}_2$
 - $\text{CdO} + \text{CO} \rightarrow \text{Cd} + \text{CO}_2$
49. Which of these equations does **not** represent a redox reaction? Explain your answer.
- $\text{LiOH} + \text{HNO}_3 \rightarrow \text{LiNO}_3 + \text{H}_2\text{O}$
 - $\text{Ag} + \text{S} \rightarrow \text{Ag}_2\text{S}$
 - $\text{MgI}_2 + \text{Br}_2 \rightarrow \text{MgBr}_2 + \text{I}_2$

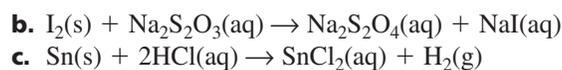
- 50.** Identify each of these half-reactions as either oxidation or reduction.
- $\text{Fe}^{3+} \rightarrow \text{Fe}^{2+}$
 - $\text{HPO}_3^{2-} \rightarrow \text{HPO}_4^{2-}$
 - $\text{BCl}_3 \rightarrow \text{B}_2\text{Cl}_4$
- 51.** Determine the oxidation number of nitrogen in each of these molecules or ions.
- | | | |
|------------------|---------------------------|-------------------------|
| a. NH_3 | c. N_2H_4 | e. N_2O |
| b. KCN | d. NO_3^- | f. NF_3 |
- 52.** Determine the oxidation number of each element in these compounds or ions.
- FeCr_2O_4 (iron(II) chromite)
 - $\text{Au}_2(\text{SeO}_4)_3$ (gold(III) selenate)
 - $\text{Ni}(\text{CN})_2$ (nickel(II) cyanide)
- 53.** Explain how the sulfite ion (SO_3^{2-}) differs from sulfur trioxide (SO_3).

Balancing Redox Equations (20.2)

- 54.** Use the oxidation-number method to balance these redox equations.
- $\text{CO} + \text{I}_2\text{O}_5 \rightarrow \text{I}_2 + \text{CO}_2$
 - $\text{Cl}_2 + \text{NaOH} \rightarrow \text{NaCl} + \text{HOCl}$
 - $\text{SO}_2 + \text{Br}_2 + \text{H}_2\text{O} \rightarrow \text{HBr} + \text{H}_2\text{SO}_4$
 - $\text{HBrO}_3 \rightarrow \text{Br}_2 + \text{H}_2\text{O} + \text{O}_2$
- 55.** Use the oxidation-number method to balance the following ionic redox equations.
- $\text{Al} + \text{I}_2 \rightarrow \text{Al}^{3+} + \text{I}^-$
 - $\text{MnO}_2 + \text{Br}^- \rightarrow \text{Mn}^{2+} + \text{Br}_2$ (in acid solution)
 - $\text{Cu} + \text{NO}_3^- \rightarrow \text{Cu}^{2+} + \text{NO}$ (in acid solution)
 - $\text{Zn} + \text{NO}_3^- \rightarrow \text{Zn}^{2+} + \text{NO}_2$ (in acid solution)
- 56.** Use the oxidation-number method to balance these redox equations.
- $\text{PbS} + \text{O}_2 \rightarrow \text{PbO} + \text{SO}_2$
 - $\text{NaWO}_3 + \text{NaOH} + \text{O}_2 \rightarrow \text{Na}_2\text{WO}_4 + \text{H}_2\text{O}$
 - $\text{NH}_3 + \text{CuO} \rightarrow \text{Cu} + \text{N}_2 + \text{H}_2\text{O}$
 - $\text{Al}_2\text{O}_3 + \text{C} + \text{Cl}_2 \rightarrow \text{AlCl}_3 + \text{CO}$
- 57.** Use the oxidation-number method to balance these ionic redox equations.
- $\text{MoCl}_5 + \text{S}^{2-} \rightarrow \text{MoS}_2 + \text{Cl}^- + \text{S}$
 - $\text{Al} + \text{OH}^- + \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{AlO}_2^-$
 - $\text{TiCl}_6^{2-} + \text{Zn} \rightarrow \text{Ti}^{3+} + \text{Cl}^- + \text{Zn}^{2+}$

Half-Reactions (20.3)

- 58.** Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.
- $\text{PbO}(\text{s}) + \text{NH}_3(\text{g}) \rightarrow \text{N}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + \text{Pb}(\text{s})$



- 59.** Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed. Keep balanced equations in net ionic form.
- $\text{Cl}^-(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{ClO}^-(\text{aq}) + \text{NO}(\text{g})$ (in acid solution)
 - $\text{IO}_3^-(\text{aq}) + \text{Br}^-(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + \text{IBr}(\text{s})$ (in acid solution)
 - $\text{I}_2(\text{s}) + \text{Na}_2\text{S}_2\text{O}_3(\text{aq}) \rightarrow \text{Na}_2\text{S}_2\text{O}_4(\text{aq}) + \text{NaI}(\text{aq})$ (in acid solution)
- 60.** Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed.
- $\text{NH}_3(\text{g}) + \text{NO}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{Mn}^{2+}(\text{aq}) + \text{BiO}_3^-(\text{aq}) \rightarrow \text{Bi}^{2+}(\text{aq}) + \text{MnO}_4^-(\text{aq})$ (in acid solution)
 - $\text{Br}_2 \rightarrow \text{Br}^- + \text{BrO}_3^-$ (in basic solution)
- 61.** Balance the following redox chemical equation. Rewrite the equation in full ionic form, then derive the net ionic equation and balance by the half-reaction method. Give the final answer as it is shown below but with the balancing coefficients.
- $$\text{KMnO}_4(\text{aq}) + \text{FeSO}_4(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Fe}_2(\text{SO}_4)_3(\text{aq}) + \text{MnSO}_4(\text{aq}) + \text{K}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$$
- 62.** Balance this equation in the same manner as in question 61 above.
- $$\text{HNO}_3(\text{aq}) + \text{K}_2\text{Cr}_2\text{O}_4(\text{aq}) + \text{Fe}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{Fe}(\text{NO}_3)_3(\text{aq}) + \text{Cr}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 63.** Determine the oxidation number of the bold element in each of the following examples.
- | | |
|-----------------------|----------------------------|
| a. OF_2 | c. RuO_4 |
| b. UO_2^{2+} | d. Fe_2O_3 |
- 64.** Identify the reducing agents in these equations.
- $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$
 - $\text{Na}_2\text{SO}_4 + 4\text{C} \rightarrow \text{Na}_2\text{S} + 4\text{CO}$
 - $4\text{IrF}_5 + \text{Ir} \rightarrow 5\text{IrF}_4$

65. Write a balanced ionic redox equation using the following pairs of redox half-reactions.

- $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^{-}$
 $\text{Te}^{2+} + 2\text{e}^{-} \rightarrow \text{Te}$
- $\text{IO}_4^{-} + 2\text{e}^{-} \rightarrow \text{IO}_3^{-}$
 $\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^{-}$ (in acid solution)
- $\text{I}_2 + 2\text{e}^{-} \rightarrow 2\text{I}^{-}$
 $\text{N}_2\text{O} \rightarrow \text{NO}_3^{-} + 4\text{e}^{-}$ (in acid solution)

66. Balance these redox equations by any method.

- $\text{P} + \text{H}_2\text{O} + \text{HNO}_3 \rightarrow \text{H}_3\text{PO}_4 + \text{NO}$
- $\text{KClO}_3 + \text{HCl} \rightarrow \text{Cl}_2 + \text{ClO}_2 + \text{H}_2\text{O} + \text{KCl}$

67. Balance these ionic redox equations by any method.

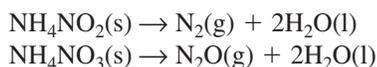
- $\text{Sb}^{3+} + \text{MnO}_4^{-} \rightarrow \text{SbO}_4^{3-} + \text{Mn}^{2+}$ in acid solution
- $\text{N}_2\text{O} + \text{ClO}^{-} \rightarrow \text{Cl}^{-} + \text{NO}_2^{-}$ in basic solution

68. Balance these ionic redox equations by any method.

- $\text{Mg} + \text{Fe}^{3+} \rightarrow \text{Mg}^{2+} + \text{Fe}$
- $\text{ClO}_3^{-} + \text{SO}_2 \rightarrow \text{Cl}^{-} + \text{SO}_4^{2-}$ (in acid solution)

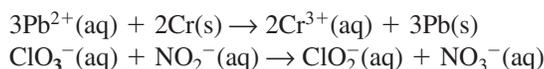
Thinking Critically

69. Applying Concepts The following equations show redox reactions that are sometimes used in the laboratory to generate pure nitrogen gas and pure dinitrogen monoxide gas (nitrous oxide, N_2O).



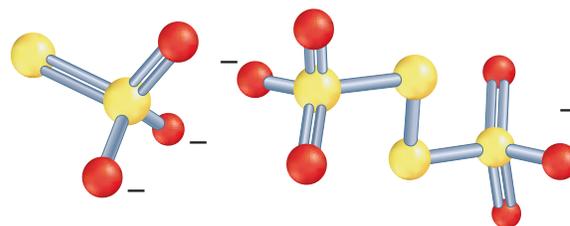
- Determine the oxidation number of each element in the two equations and then make diagrams as in Example Problem 20-2 showing the changes in oxidation numbers that occur in each reaction.
- Identify the atom that is oxidized and the atom that is reduced in each of the two reactions.
- Identify the oxidizing and reducing agents in each of the two reactions.
- Write a sentence telling how the electron transfer taking place in these two reactions differs from that taking place here.
 $2\text{AgNO}_3 + \text{Zn} \rightarrow \text{Zn}(\text{NO}_3)_2 + 2\text{Ag}$

70. Comparing and Contrasting All redox reactions involve a transfer of electrons, but this transfer can take place between different types of atoms as illustrated by the two equations below. How does electron transfer in the first reaction differ from electron transfer in the second reaction?



71. Using Numbers Examine the net ionic equation below for the reaction that occurs when the thiosulfate ion ($\text{S}_2\text{O}_3^{2-}$) is oxidized to the tetrathionate ion

($\text{S}_4\text{O}_6^{2-}$). Balance the equation using the half-reaction method. The structures of the two ions will help you to determine the oxidation numbers to use.



Thiosulfate ion ($\text{S}_2\text{O}_3^{2-}$) Tetrathionate ion ($\text{S}_4\text{O}_6^{2-}$)

72. Predicting Consider the fact that all of the following are stable compounds. What can you infer about the oxidation state of phosphorus in its compounds?



Writing in Chemistry

- Research the role of oxidation–reduction reactions in the manufacture of steel. Write a summary of your findings, including appropriate diagrams and equations representing the reactions.
- Practice your technical writing skills by writing (in your own words) a procedure for cleaning tarnished silverware by a redox chemical process. Be sure to include background information describing the process as well as logical steps that would enable anyone to accomplish the task.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- How do the following characteristics apply to the electron configurations of transition metals? (Chapter 7)
 - Ions vary in charge.
 - Many elements have high melting points.
 - Many of their solids are colored.
 - Some elements are hard solids.
- When iron(III) chloride reacts in an atmosphere of pure oxygen the following occurs:
 $4\text{FeCl}_3(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s}) + 6\text{Cl}_2(\text{g})$
If 45.0 g of iron(III) chloride reacts and 20.5 g of iron(III) oxide is recovered, determine the percent yield. (Chapter 12)

Use these questions and the test-taking tip to prepare for your standardized test.

- The reducing agent in a redox reaction is all of the following EXCEPT _____.
 - the substance oxidized
 - the electron acceptor
 - the reducer of another substance
 - the electron donor
- The oxidation numbers of the elements in CuSO_4 are _____.
 - $\text{Cu} = +2$, $\text{S} = +6$, $\text{O} = -2$
 - $\text{Cu} = +3$, $\text{S} = +5$, $\text{O} = -2$
 - $\text{Cu} = +2$, $\text{S} = +2$, $\text{O} = -1$
 - $\text{Cu} = +2$, $\text{S} = 0$, $\text{O} = -2$
- For the reaction $\text{X} + \text{Y} \rightarrow \text{XY}$, the element that will be reduced is the one that is _____.
 - more reactive
 - more massive
 - more electronegative
 - more radioactive
- The net ionic reaction between iodine and lead(IV) oxide is shown below:
 $\text{I}_2(\text{s}) + \text{PbO}_2(\text{s}) \rightarrow \text{IO}_3^-(\text{aq}) + \text{Pb}^{2+}(\text{aq})$
 If the reaction takes place in acidic solution, the balanced equation is _____.
 - $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 8\text{OH}^-(\text{aq})$
 - $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + \text{H}^+(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 8\text{H}^+(\text{aq})$
 - $\text{I}_2(\text{s}) + 5\text{PbO}_2(\text{s}) + 8\text{H}^+(\text{aq}) \rightarrow 2\text{IO}_3^-(\text{aq}) + 5\text{Pb}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$
- The reaction between sodium iodide and chloride is shown below:
 $2\text{NaI}(\text{aq}) + \text{Cl}_2(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{I}_2(\text{aq})$
 The oxidation state of Na remains unchanged because _____.
 - Na^+ is a spectator ion
 - Na^+ cannot be reduced
 - Na is an uncombined element
 - Na^+ is a monatomic ion
- The reaction between nickel and copper(II) chloride is shown below:
 $\text{Ni}(\text{s}) + \text{CuCl}_2(\text{aq}) \rightarrow \text{Cu}(\text{s}) + \text{NiCl}_2(\text{aq})$
 The half reactions for this redox reaction are _____.
 - $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $\text{Cl}_2 \rightarrow 2\text{Cl}^- + 2\text{e}^-$

- $\text{Ni} \rightarrow \text{Ni}^+ + \text{e}^-$, $\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$
- $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
- $\text{Ni} \rightarrow \text{Ni}^{2+} + 2\text{e}^-$, $2\text{Cu}^+ + 2\text{e}^- \rightarrow \text{Cu}$

Interpreting Tables Use the table to answer questions 7–9.

Data for Elements in the Redox Reaction
 $\text{Zn} + \text{HNO}_3 \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{NO}_2 + \text{H}_2\text{O}$

Element	Oxidation Number	Complex ion of which element is a part
Zn	0	none
Zn in $\text{Zn}(\text{NO}_3)_2$	+2	none
H in HNO_3	+1	none
H in H_2O	?	none
N in HNO_3	?	NO_3^-
N in NO_2	+4	none
N in $\text{Zn}(\text{NO}_3)_2$?	NO_3^-
O in HNO_3	-2	NO_3^-
O in NO_2	?	none
O in $\text{Zn}(\text{NO}_3)_2$?	NO_3^-
O in H_2O	-2	none

- Which of these elements forms a monatomic ion that is a spectator in the redox reaction?
 - Zn
 - O
 - N
 - H
- The oxidation number of N in $\text{Zn}(\text{NO}_3)_2$ is _____.
 - +3
 - +5
 - +1
 - +6
- The element that is oxidized in this reaction is _____.
 - Zn
 - O
 - N
 - H



TEST-TAKING TIP

Write It Down! Most tests ask you a large number of questions in a small amount of time. Write down your work wherever possible. Write out the half reactions for a redox problem, and make sure they add up. Do math on paper, not in your head. Underline and reread important facts in passages and diagrams—don't try to memorize them.

Electrochemistry

What You'll Learn

- ▶ You will learn how oxidation–reduction reactions produce electric current.
- ▶ You will determine the voltage of the current produced by pairs of redox half-reactions.
- ▶ You will determine the direction of current flow for a particular pair of redox half-reactions.
- ▶ You will investigate how electric current can be used to carry out redox reactions.

Why It's Important

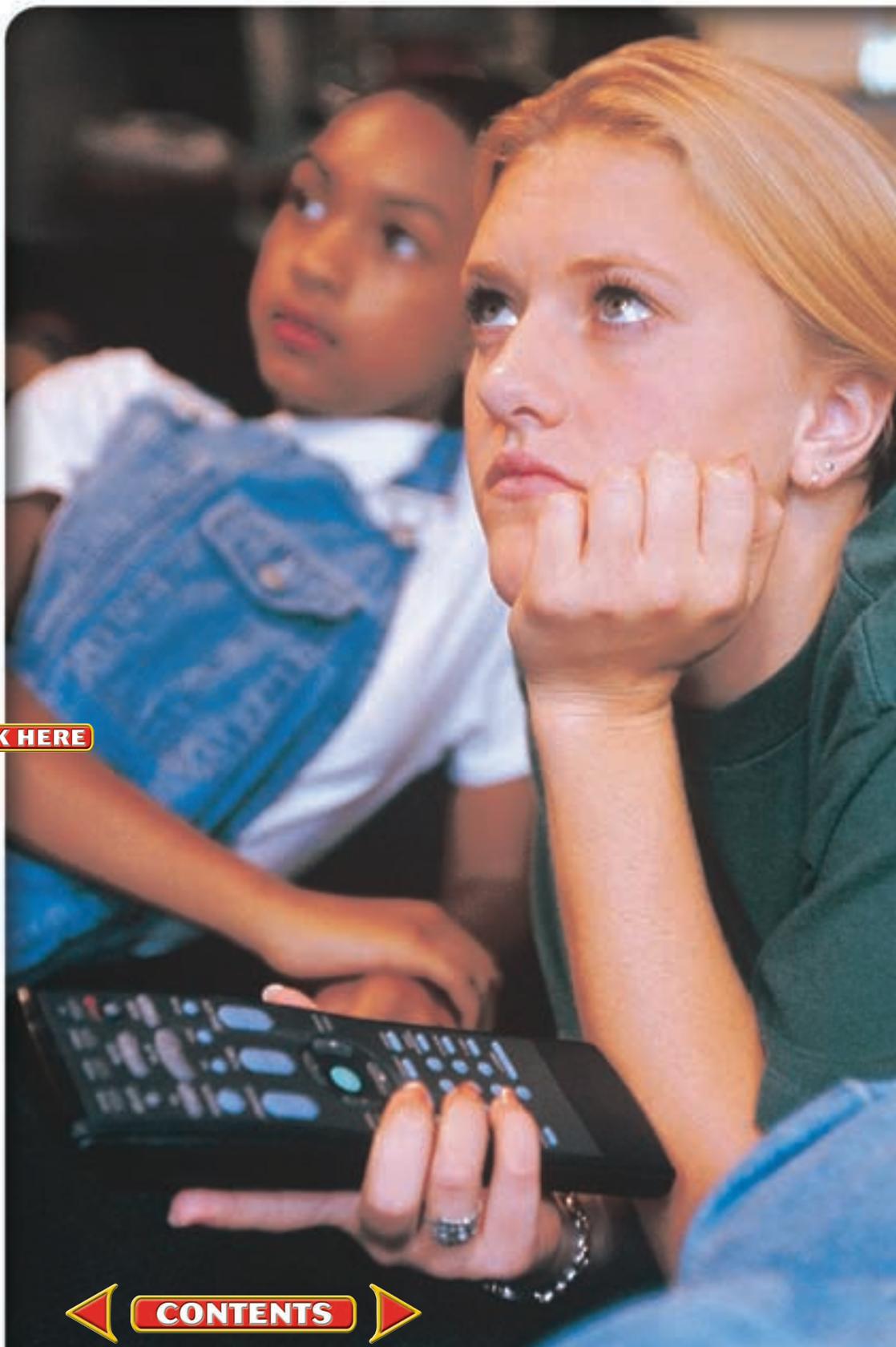
You've probably used aluminum foil to cover food, and most kids have hidden under a blanket to read by flashlight. The manufacturing process for aluminum and the light from your flashlight involve electrochemistry.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about electrochemistry.

You can look around you to see how important electricity is, from the lights in the room to the batteries that power the remote control, to the manufacturing processes that produce consumer goods. Electrochemistry is one method by which electricity can be generated for use.



DISCOVERY LAB



Materials

lemon pieces
zinc metal strip
copper metal strip
voltmeter with leads

A Lemon Battery?

You can purchase a handy package of portable power at any convenience store—a battery. You also can craft a working battery from a lemon. How are these power sources alike?

Safety Precautions



Always wear an apron and safety goggles in the lab. Use caution with electricity.

Procedure

1. Insert the zinc and copper strips into the lemon, about 2 cm apart from each other.
2. Attach the black lead from a voltmeter to the zinc and the red lead to the copper. Read and record the potential difference (voltage) from the voltmeter.
3. Remove one of the metals from the lemon and observe what happens to the potential difference on the voltmeter.

Analysis

What is the purpose of the zinc and copper metals? What is the purpose of the lemon?

Section

21.1

Voltaic Cells

Objectives

- **Describe** a way to obtain electrical energy from a redox reaction.
- **Identify** the parts of a voltaic cell and explain how each part operates.
- **Calculate** cell potentials and **determine** the spontaneity of redox reactions.

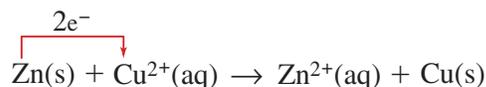
Vocabulary

salt bridge
electrochemical cell
voltaic cell
half-cell
anode
cathode
reduction potential
standard hydrogen electrode
battery

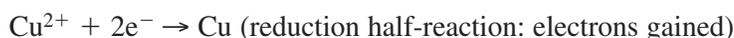
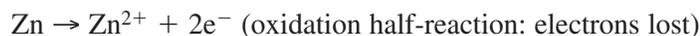
You now know much of the basic chemistry on which advanced chemistry is based. This chapter introduces an important branch of chemistry called electrochemistry. Electrochemistry is the study of the process by which chemical energy is converted to electrical energy and vice versa.

Redox in Electrochemistry

In Chapter 20, you learned that all redox reactions involve a transfer of electrons from the species that is oxidized to the species that is reduced. The **DISCOVERY LAB** illustrates the simple redox reaction in which zinc atoms are oxidized to form Zn^{2+} ions. The two electrons donated from each zinc atom are accepted by a Cu^{2+} ion, which changes to an atom of copper metal. The following net ionic equation illustrates the electron transfer that occurs.



Two half-reactions make up this redox process:



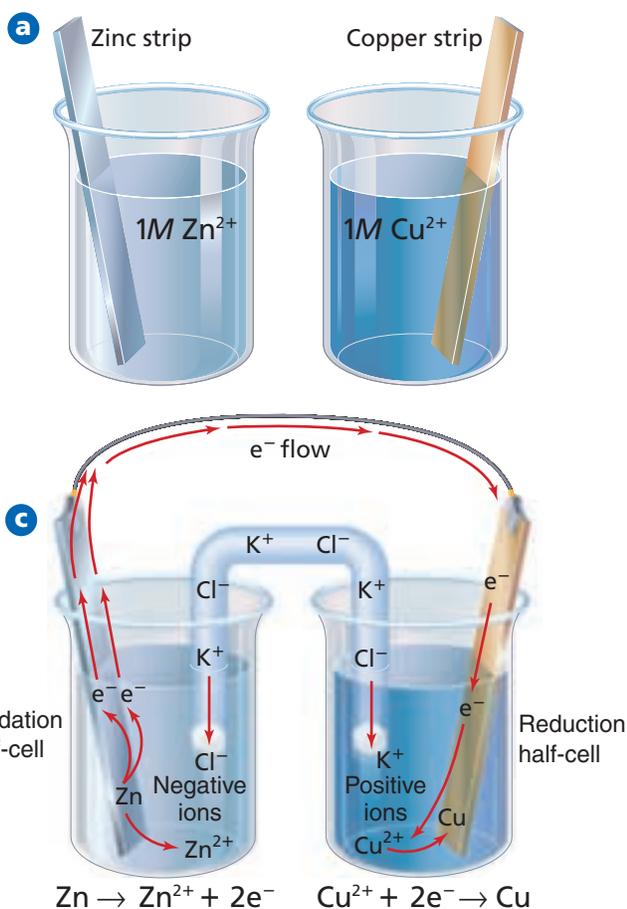
What do you think would happen if you separated the oxidation half-reaction from the reduction half-reaction? Can redox occur? Consider **Figure 21-1a** in which a zinc strip is immersed in a solution of zinc sulfate and a copper strip is immersed in a solution of copper(II) sulfate. Two problems prohibit a redox reaction in this situation. First, with this setup there is no way for zinc atoms to transfer electrons to copper(II) ions. This problem can be solved by connecting a metal wire between the zinc and copper strips, as shown in **Figure 21-1b**. The wire serves as a pathway for electrons to flow from the zinc strip to the copper strip.

Even with the conducting wire, another problem exists that prohibits the redox reaction. A positive charge builds up in one solution and a negative charge builds up in the other. The buildup of positive zinc ions on the left prohibits the oxidation of zinc atoms. On the other side, the buildup of negative ions prohibits the reduction of copper ions. To solve this problem a salt bridge must be built into the system. A **salt bridge** is a pathway constructed to allow the passage of ions from one side to another, as shown in **Figure 21-1c**. A salt bridge consists of a tube containing a conducting solution of a soluble salt, such as KCl, held in place by an agar gel or other form of plug. The ions can move through the plug but the solutions in the two beakers cannot mix.

When the metal wire and the salt bridge are in place, electrons flow through the wire from the oxidation half-reaction to the reduction half-reaction while positive and negative ions move through the salt bridge. A flow of charged particles, such as electrons, is called an electric current. Take a minute to identify the electric current in **Figure 21-1**. The flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current. The energy

Figure 21-1

- a** These containers are constructed and arranged so that zinc is oxidized on one side while copper ions are reduced on the other. **b** A wire connected between the zinc and copper strips provides a pathway for the flow of electrons.
- c** However, it takes a salt bridge to complete the system and allow the redox reaction to produce an electric current.



of the flowing electrons can be put to use in lighting a bulb or running an electric motor.

The completed device shown in **Figure 21-1c** is an electrochemical cell. An **electrochemical cell** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction. A **voltaic cell** is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction. **Figure 21-2** shows a version of the original voltaic cell as devised by its inventor Alessandro Volta.

Chemistry of Voltaic Cells

An electrochemical cell consists of two parts, called **half-cells**, in which the separate oxidation and reduction reactions take place. Each half-cell contains an electrode, which is the object that conducts electrons to or from another substance, usually a solution of ions. In **Figure 21-1**, the beaker with the zinc electrode is where the oxidation part of the redox reaction takes place. The beaker with the copper electrode is where the reduction part of the reaction takes place. The reaction that takes place in each half-cell is the half-reaction, sometimes called half-cell reaction, that you studied in Chapter 20. The electrode where oxidation takes place is called the **anode** of the cell. The electrode where reduction takes place is called the **cathode** of the cell. Which beaker in **Figure 21-1** contains the anode and which contains the cathode?

Recall from Chapter 16 that an object's potential energy is due to its position or composition. In electrochemistry, electrical potential energy is a measure of the amount of current that can be generated from a voltaic cell to do work. Electric charge can flow between two points only when a difference in electrical potential energy exists between the two points. In an electrochemical cell, these two points are the two electrodes. The potential difference of a voltaic cell is an indication of the energy that is available to move electrons from the anode to the cathode.

To better understand this concept, consider the analogy illustrated in **Figure 21-3**. A golf ball that lands on a hillside will roll downhill into a low spot because a difference in gravitational potential energy exists between the two points. The kinetic energy attained by a golf ball rolling down a hill is determined by the difference in height (potential energy) between the high point and the low point. Similarly, the energy of the electrons flowing from anode to cathode in a voltaic cell is determined by the

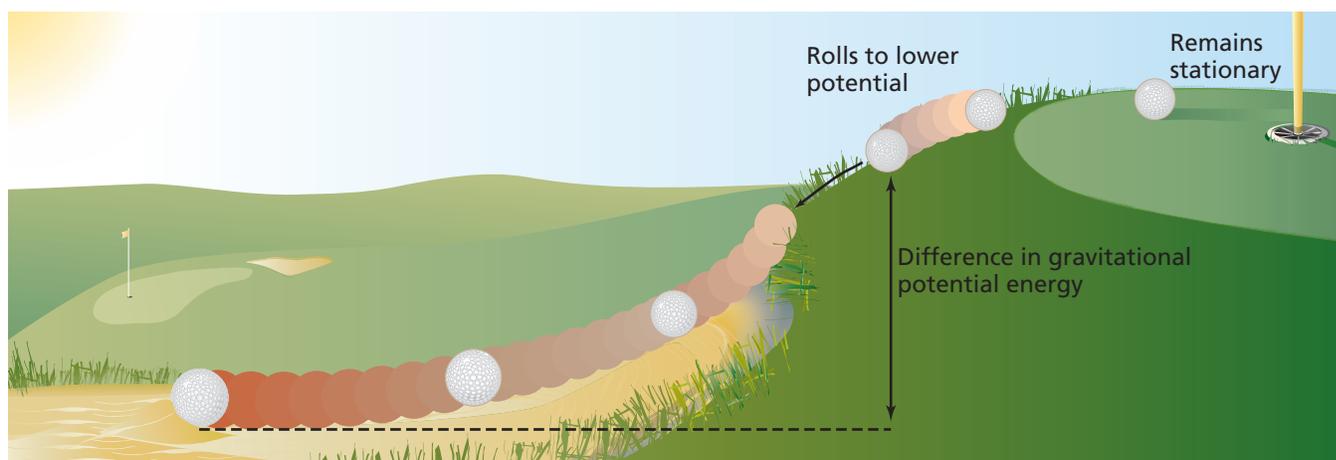


Figure 21-2

The voltaic cell is named for Alessandro Volta (1745–1827), the Italian physicist who is credited with its invention in 1800.

Figure 21-3

In this illustration of gravitational potential, we can say that the golf ball has potential energy relative to the bottom of the hill because there is a difference in the height position of the ball from the top of the hill to the bottom. Similarly, an electrochemical cell has potential energy to produce a current because there is a difference in the ability of the electrodes to move electrons from the anode to the cathode.





Go to the **Chemistry Interactive CD-Rom** to find additional resources for this chapter.

difference in electrical potential between the two electrodes. In redox terms, the voltage of a cell is determined by the difference in the tendency of the electrode material to accept electrons.

Thinking again of the golf ball analogy, the force of gravity causes the ball always to roll downhill to a lower energy state, not uphill to a higher energy state. In other words, the golf ball rolling process occurs spontaneously only in the downhill direction. In the zinc–copper cell under standard conditions, copper(II) ions at the cathode accept electrons more readily than the zinc ions at the anode. In other words, the redox reaction occurs spontaneously only when the electrons flow from the zinc to the copper.

Calculating Electrochemical Cell Potential

You know from the previous chapter that the gain of electrons is called reduction. Building on this fact, we can call the tendency of a substance to gain electrons its **reduction potential**. The reduction potential of an electrode, measured in volts, cannot be determined directly because the reduction half-reaction must be coupled with an oxidation half-reaction. When these half-reactions are coupled, the voltage corresponds to the *difference* in potential between the half-reactions. The electrical potential difference between two points is expressed in volts (V). Long ago, chemists decided to measure the reduction potential of all electrodes against one standard electrode, the standard hydrogen electrode. The **standard hydrogen electrode** consists of a small sheet of platinum immersed in an HCl solution that has a hydrogen ion concentration of 1M. Hydrogen gas at a pressure of 1 atm is bubbled in and the temperature is maintained at 25°C, as shown in **Figure 21-4**. The potential (also called the standard reduction potential) of this standard hydrogen electrode is defined as 0 V. This electrode can act as an oxidation half-reaction or a reduction half-reaction, depending upon the half-cell to which it is connected. The reactions at the hydrogen electrode are

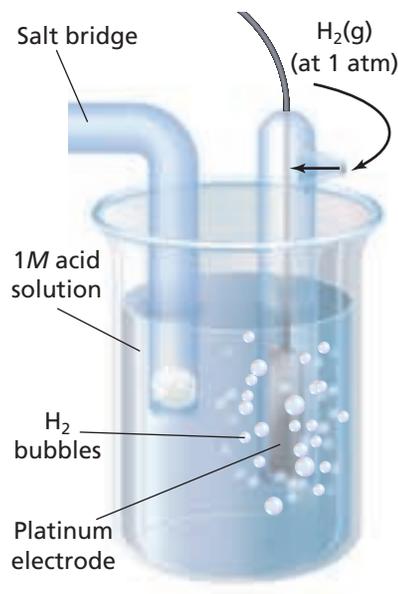
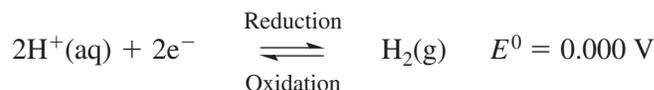


Figure 21-4

A standard hydrogen electrode consists of a platinum electrode with hydrogen gas at 1 atm pressure bubbling into an acidic solution that is 1M in hydrogen ions. The reduction potential for this configuration is 0 V.

Over the years, chemists have measured and recorded the standard reduction potentials, abbreviated E^0 , of many different half-cells. **Table 21-1** lists some common half-cell reactions in order of increasing reduction potential. The values in the table are based on using the half-cell reaction that is being measured as the cathode and the standard hydrogen electrode as the anode. All of the half-reactions in **Table 21-1** are written as reductions. However, in any voltaic cell, which always contains two half-reactions, the half-reaction with the lower reduction potential will proceed in the opposite direction and will be an oxidation reaction. In other words, the half-reaction that is more positive will proceed as a reduction and the half-reaction that is more negative will proceed as an oxidation.

Table 21-1

Standard Reduction Potentials at 25°C, 1 atm, and 1M Ion Concentration				
Half-reaction	E^0 (V)	Half-reaction	E^0 (V)	
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.0401	$\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$	0.151	
$\text{Cs}^+ + \text{e}^- \rightarrow \text{Cs}$	-3.026	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	0.153	
$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.931	$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{SO}_3 + \text{H}_2\text{O}$	0.172	
$\text{Ba}^{2+} + 2\text{e}^- \rightarrow \text{Ba}$	-2.912	$\text{Bi}^{3+} + 3\text{e}^- \rightarrow \text{Bi}$	0.308	
$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.868	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	0.3419	
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	0.401	
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.372	$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	0.521	
$\text{Ce}^{3+} + 3\text{e}^- \rightarrow \text{Ce}$	-2.336	$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	0.5355	
$\text{H}_2 + 2\text{e}^- \rightarrow 2\text{H}^-$	-2.323	$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2$	0.695	
$\text{Nd}^{3+} + 3\text{e}^- \rightarrow \text{Nd}$	-2.1	$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	0.771	
$\text{Be}^{2+} + 2\text{e}^- \rightarrow \text{Be}$	-1.847	$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$	0.775	
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.662	$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	0.7973	
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.185	$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.7996	
$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.913	$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	0.851	
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.8277	$2\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	0.920	
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.7618	$\text{Pd}^{2+} + 2\text{e}^- \rightarrow \text{Pd}$	0.951	
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.744	$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	0.957	
$\text{Ga}^{3+} + 3\text{e}^- \rightarrow \text{Ga}$	-0.549	$\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-$	1.066	
$2\text{CO}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{C}_2\text{O}_4$	-0.49	$\text{Ir}^{3+} + 3\text{e}^- \rightarrow \text{Ir}$	1.156	
$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.47627	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$	1.18	
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.447	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.229	
$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.407	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	1.35827	
$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.4030	$\text{Au}^{3+} + 2\text{e}^- \rightarrow \text{Au}^+$	1.401	
$\text{PbI}_2 + 2\text{e}^- \rightarrow \text{Pb} + 2\text{I}^-$	-0.365	$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	1.498	
$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.3588	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.507	
$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28	$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	1.692	
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.257	$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	1.776	
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.1375	$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	1.92	
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262	$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	2.010	
$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.037	$\text{O}_3 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{O}_2 + \text{H}_2\text{O}$	2.076	
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.0000	$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	2.866	

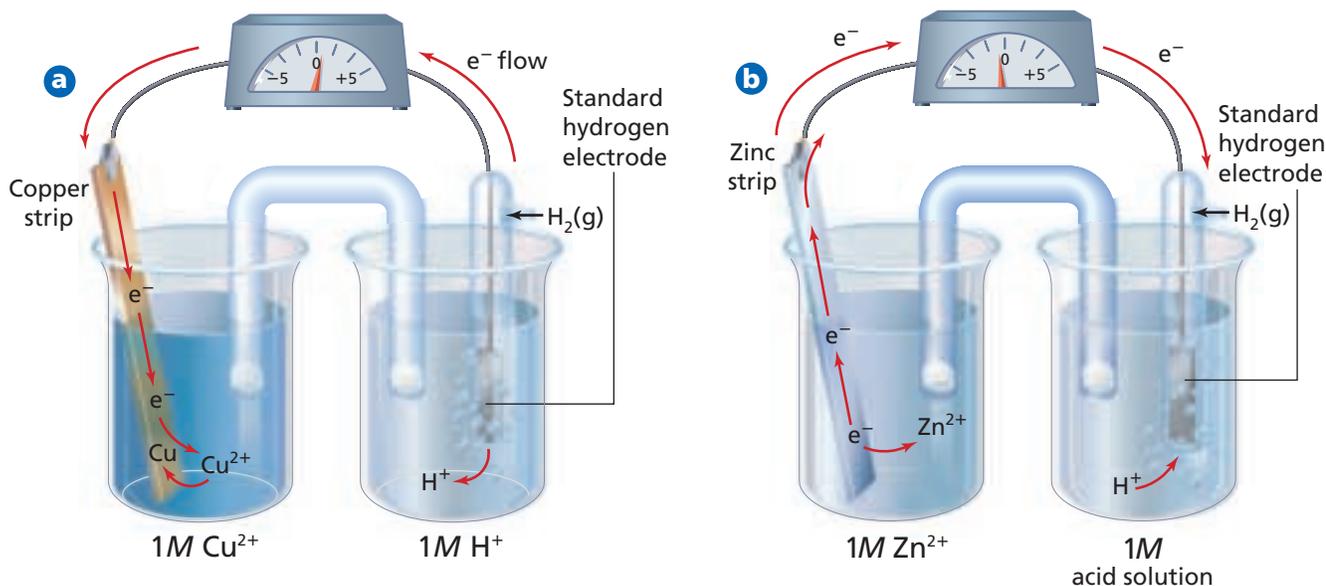


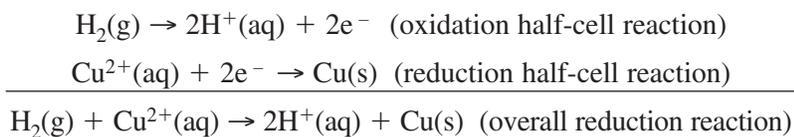
Figure 21-5

a When a $\text{Cu}|\text{Cu}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce Cu^{2+} ions to Cu atoms. The voltage of this reaction is $+0.342\text{ V}$.

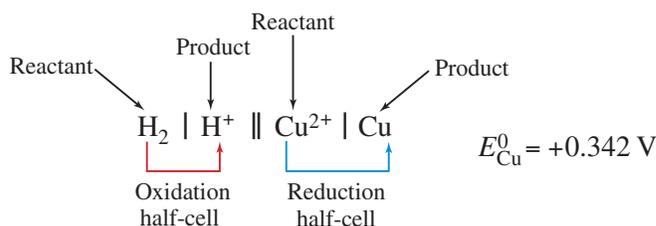
b When a $\text{Zn}|\text{Zn}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to Zn^{2+} ions. The voltage of this reaction is -0.762 V .

The electrode being measured also must be under standard conditions; that is, immersed in a 1 M solution of its ions at 25°C and 1 atm pressure. The superscript zero in the notation E^0 is a shorthand way of indicating “measured under standard conditions.”

With **Table 21-1**, it is possible to calculate the electrical potential of a voltaic cell having a copper electrode and a zinc electrode under standard conditions. The first step is to determine the standard reduction potential for the copper half-cell (E_{Cu}^0). When the copper electrode is attached to a standard hydrogen electrode, as in **Figure 21-5a**, electrons flow *from* the hydrogen electrode *to* the copper electrode, and copper ions are reduced to copper metal. The E^0 , measured by a voltmeter, is $+0.342\text{ V}$. The positive voltage indicates that Cu^{2+} ions at the copper electrode being measured accept electrons more readily than do H^+ ions at the standard hydrogen electrode. From this information, you can determine where oxidation and reduction take place. In this case, oxidation takes place at the hydrogen electrode, and reduction takes place at the copper electrode. The oxidation and reduction half-cell reactions and the overall reaction are



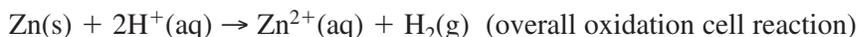
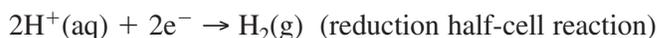
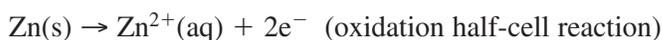
This reaction can be written in a form called cell notation to clearly describe the copper reduction reaction.



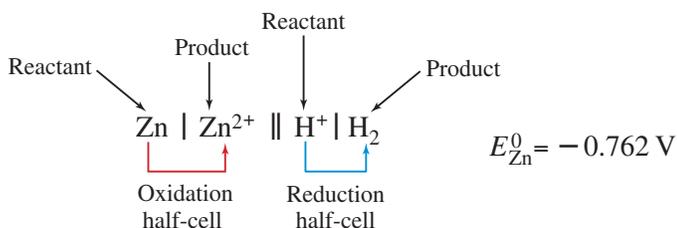
Note that the two participants in the oxidation reaction are written first and in the order they appear in the oxidation half-reaction—*reactant* | *product*.

They are followed by a double vertical line (||) indicating the wire and salt bridge connecting the half-cells. Then, the two participants in the reduction reaction are written in the same *reactant* | *product* order. Note in this example that it is customary to place a plus sign before a positive voltage to avoid confusion.

The next step is to determine the standard reduction potential for the zinc half-cell (E_{Zn}^0). When the zinc electrode is measured against the standard hydrogen electrode under standard conditions, as in **Figure 21-5b**, electrons flow from the zinc electrode to the hydrogen electrode. The E^0 of the zinc half-cell, measured by a voltmeter, is -0.762 V . This means that the H^+ ions at the hydrogen electrode accept electrons more readily than do the zinc ions, and thus the hydrogen ions have a *higher* reduction potential than the zinc ions. If you remember that the hydrogen electrode is assigned a zero potential, you will realize that the reduction potential of the zinc electrode must have a negative value. Electrodes at which oxidation is carried out when connected to a hydrogen electrode have negative standard reduction potentials. How would you write the oxidation and reduction reactions and the overall zinc oxidation reaction? The reactions are written as follows.



This reaction can be written in cell notation to clearly describe the zinc oxidation reaction.



The final step is to combine the Cu and Zn half-cells as a voltaic cell, which means calculating the voltaic cell's standard potential using the following formula

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

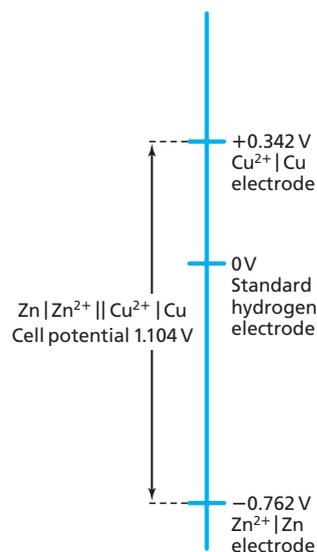
Because reduction occurred at the copper electrode and oxidation occurred at the zinc electrode, the E^0 values are substituted as follows.

$$\begin{aligned} E_{\text{cell}}^0 &= E_{\text{Cu}^{2+}|\text{Cu}}^0 - E_{\text{Zn}^{2+}|\text{Zn}}^0 \\ &= 0.342 \text{ V} - (-0.762 \text{ V}) \\ &= +1.104 \text{ V} \end{aligned}$$

Figure 21-6 describes this potential calculation. The graph shows the zinc half-cell with the lower reduction potential (the oxidation half-reaction) and the copper half-cell with the higher reduction potential (the reduction half-reaction). You can see that the space between the two (E_{cell}^0) is the difference between the potentials of the individual half-cells. The Example Problem that follows gives a step-by-step description of calculating cell potentials.

Figure 21-6

This simple graph illustrates how the overall cell potential is derived from the difference in reduction potential of two electrodes.



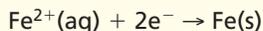
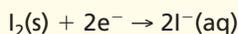
EXAMPLE PROBLEM 21-1



Cloud seeding using iodine compounds to promote precipitation may be one way to ease severe drought conditions.

Calculating Cell Potential

The two reduction half-reactions in this example represent the half-cells of a voltaic cell.



Determine the overall cell reaction and the standard cell potential. Write the cell chemistry using cell notation with vertical lines separating components.

1. Analyze the Problem

You are given the half-cell descriptions for a voltaic cell and standard reduction potentials in **Table 21-1**. In any voltaic cell, the half-reaction with the lower reduction potential will proceed as an oxidation. With this information, you can write the overall cell reaction and calculate the standard cell potential.

Known

Standard reduction potentials for the half-cells

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

Unknown

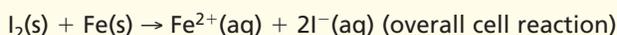
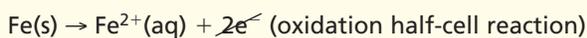
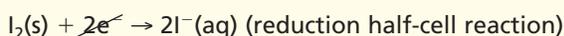
$$E_{\text{cell}}^0 = ?$$

2. Solve for the Unknown

Find the standard reduction potentials of each half-reaction in a reference source such as **Table 21-1**.



Note that reduction of iodine has the higher reduction potential. This half-reaction will proceed in the forward direction as a reduction. The iron half-reaction will proceed in the reverse direction as an oxidation. Rewrite the half-reactions in the correct direction.



Balance the reaction if necessary. Note that this reaction is balanced as written.

Calculate cell standard potential.

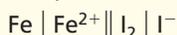
$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

$$E_{\text{cell}}^0 = E_{\text{I}_2|\text{I}^-}^0 - E_{\text{Fe}^{2+}|\text{Fe}}^0$$

$$E_{\text{cell}}^0 = +0.536 \text{ V} - (-0.447 \text{ V})$$

$$E_{\text{cell}}^0 = +0.983 \text{ V}$$

Write the reaction using cell notation. When representing a reaction in cell notation, the species in the oxidation half-reaction are written first in the following order: reactant | product, or in this case, $\text{Fe} | \text{Fe}^{2+}$. The species in the reduction half-reaction are written next in the order reactant | product, or in this case, $\text{I}_2 | \text{I}^-$. Therefore, the complete cell is represented as

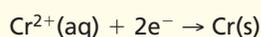
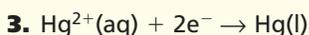
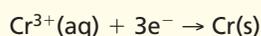
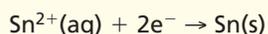
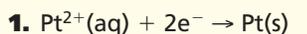


3. Evaluate the Answer

The standard potential for this voltaic cell seems reasonable given the reduction potentials of the half-cells that comprise it. The mathematical operations with negative numbers are correct and the answer is correct to the thousandths place.

PRACTICE PROBLEMS

For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction and calculate the standard cell potential. Express the reaction using cell notation. You may wish to refer to Chapter 20 to review writing and balancing redox equations.



For more practice calculating cell potential, go to **Supplemental Practice Problems** in Appendix A.

The **CHEMLAB** at the end of this chapter offers an opportunity to create voltaic cells and calculate cell potentials.

Using Standard Reduction Potentials

You just practiced using the data from **Table 21-1** to calculate the standard potential (voltage) of voltaic cells by determining the difference in reduction potential of the half-cell reactions. Another important use of standard reduction potentials is to determine if a proposed reaction under standard conditions will be spontaneous. How can standard reduction potentials indicate spontaneity? Electrons in a voltaic cell always flow from the half-cell with the lower standard reduction potential to the half-cell with higher reduction potential, giving a positive cell voltage. To predict whether any proposed redox reaction will occur spontaneously, simply write the process in the form of half-reactions and look up the reduction potential of each. Use the values to calculate the potential of a voltaic cell operating with these two half-cell reactions. If the calculated potential is positive, the reaction is spontaneous. If the value is negative, the reaction is not spontaneous. However, the reverse reaction will occur because it will give a positive cell voltage and, therefore, the reverse reaction is spontaneous.

For example, use the standard reduction potentials in **Table 21-1** to calculate the potential of the cell $\text{Li} | \text{Li}^{+} || \text{Ag}^{+} | \text{Ag}$.

$$\begin{aligned} E_{\text{cell}}^0 &= E_{\text{Ag}^+|\text{Ag}}^0 - E_{\text{Li}^+|\text{Li}}^0 \\ &= 0.7996 \text{ V} - (-3.0401 \text{ V}) \\ &= 3.8397 \text{ V} \end{aligned}$$

The calculated potential for the cell $\text{Li} | \text{Li}^{+} || \text{Ag}^{+} | \text{Ag}$ is a positive number. Thus, the redox reaction in this cell will proceed spontaneously if the system is at 25°C, 1 atm, and 1M concentration.

Practice!

For more practice determining if redox reactions are spontaneous, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Examining the reactions in **Table 21-1** can help you determine the correct half-reactions.

- $\text{Sn(s)} + 2\text{Cu}^+(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + 2\text{Cu(s)}$
- $\text{Mg(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Pb(s)} + \text{Mg}^{2+}(\text{aq})$
- $2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O(l)} + 10\text{Hg}^{2+}(\text{aq}) \rightarrow 2\text{MnO}_4^-(\text{aq}) + 16\text{H}^+(\text{aq}) + 5\text{Hg}_2^{2+}(\text{aq})$
- $2\text{SO}_4^{2-}(\text{aq}) + \text{Co}^{2+}(\text{aq}) \rightarrow \text{Co(s)} + \text{S}_2\text{O}_8^{2-}(\text{aq})$

Now that you know the chemistry of voltaic cells, you may begin to see their value as a power source. It turns out that Alessandro Volta also realized the value of this chemistry. From his experimentation, Volta deduced that if one cell generates a current, several cells connected together should produce a larger current. Volta connected a series of cells by making a pile of alternating zinc plates serving as anodes and silver plates serving as cathodes separated by layers of cloth soaked with salt solution. The cloth had the same function as the salt bridge. This arrangement became known as a “voltaic pile” and was literally a “pile”, as you saw in **Figure 21-2**. Volta’s pile was the first battery. A **battery** consists of one or more electrochemical cells in a single package that generates electrical current. You’ll learn more about batteries in the next section of this chapter.

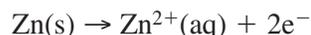
Section 21.1 Assessment

- Under what conditions can a redox reaction be used to cause an electric current to flow through a wire?
- What are the components of a voltaic cell? What is the role of each component in the operation of the cell?
- Write the balanced equation for the spontaneous cell reaction that will occur in a cell with these reduction half-reactions.
 - $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag(s)}$
 $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni(s)}$
 - $\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Mg(s)}$
 $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$
 - $\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}^{2+}(\text{aq})$
 $\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Cr(s)}$
- These equations represent overall cell reactions. Determine the standard potential for each cell and identify the reactions as spontaneous or nonspontaneous as written.
 - $2\text{Al}^{3+}(\text{aq}) + 3\text{Cu(s)} \rightarrow 3\text{Cu}^{2+}(\text{aq}) + 2\text{Al(s)}$
 - $\text{Hg}^{2+}(\text{aq}) + 2\text{Cu}^+(\text{aq}) \rightarrow 2\text{Cu}^{2+}(\text{aq}) + \text{Hg}$
 - $\text{Cd(s)} + 2\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O(l)}$
- Thinking Critically** The reduction half-reaction $\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$ has a lower standard reduction potential than the reduction half-reaction $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$. In terms of electron transfer, what is the significance of this difference in reduction potentials? Which of these reactions would produce the higher voltage in a cell in which the oxidation half-reaction is $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$? Explain your choice.
- Predicting** Suppose you have a half-cell of unknown composition, but you know the cell is at standard conditions. You connect it to a copper–copper(II) sulfate half-cell, also at standard conditions, and the voltmeter in the circuit reads 0.869 V. Is it possible to predict the probable composition of the unknown half-cell? Explain your answer.

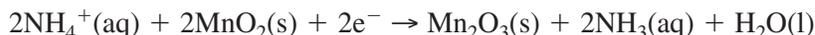
How many different uses of batteries can you identify? Batteries power flashlights, remote controls, calculators, hearing aids, portable CD players, heart pacemakers, smoke and carbon monoxide detectors, and video cameras, to name a few. Batteries in cars and trucks provide electric energy to start the engine and power the vehicle's many electric lights, sound system, and other accessories, even when the engine is not running.

Dry Cells

In Section 21.1, you learned that the simplest form of a battery is a single voltaic cell. The word *battery* once referred only to a group of single cells in one package, such as the 9 V battery. Today, however, the word *battery* refers to both single cells and packages of several cells, such as the batteries shown in **Figure 21-7**. From the time of its invention in the 1860s until recently, the most commonly used voltaic cell was the zinc–carbon dry cell, shown in **Figure 21-8** on the next page. A **dry cell** is an electrochemical cell in which the electrolyte is a moist paste. The paste in a zinc–carbon dry cell consists of zinc chloride, manganese(IV) oxide, ammonium chloride, and a small amount of water inside a zinc case. The zinc shell is the cell's anode, where the oxidation of zinc metal occurs. The following equation describes the oxidation half-cell reaction for this dry cell.



A carbon (graphite) rod in the center of the dry cell serves as the cathode, but the reduction half-cell reaction takes place in the paste. An electrode made of a material that does not participate in the redox reaction is called an inactive electrode. The carbon rod in this type of dry cell is an inactive cathode. (Contrast this with the zinc case, which is an active anode because the zinc is oxidized.) The reduction half-cell reaction for this dry cell follows.



Chemists know that the reduction reaction is more complex than this equation shows, but they still do not know precisely what happens in the reaction. A spacer, made of a porous material and damp from the liquid in the paste, separates the paste from the zinc anode. The spacer acts as a salt bridge to

Objectives

- **Describe** the structure, composition, and operation of the typical carbon–zinc dry cell battery.
- **Distinguish** between primary and secondary batteries and give two examples of each type.
- **Explain** the structure and operation of the hydrogen–oxygen fuel cell.
- **Describe** the process of corrosion of iron and methods to prevent corrosion.

Vocabulary

dry cell
primary battery
secondary battery
fuel cell
corrosion
galvanizing



Figure 21-7

The small batteries (AA, AAA, C, and D) that power household appliances and small electrical devices consist of single electrochemical cells. The large battery of a car, however, consists of six electrochemical cells connected together.



Figure 21-8

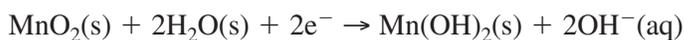
The so-called dry cell actually contains a moist paste in which the cathode half-reaction takes place. **a** In the zinc-carbon dry cell, the zinc case acts as the anode. **b** The alkaline battery uses powdered zinc and is contained in a steel case. **c** Although it looks different, the mercury battery is much like the alkaline battery, except that the mercury battery uses mercury(II) oxide (HgO) in the cathode instead of manganese(IV) oxide (MnO₂).

allow the transfer of ions, much like the model voltaic cell you studied in Section 21.1. The zinc-carbon dry cell produces a voltage of 1.5 V, until the reduction product ammonia comes out of its aqueous solution as a gas. At that point, the voltage drops to a level that makes the battery useless.

As new research and development bring about newer and more efficient products, the standard zinc-carbon dry cell has begun to give way to the alkaline dry cell, also shown in **Figure 21-8b**. The alkaline cell oxidizes zinc, but the zinc is in a powdered form. Why might the powdered zinc be an advantage? The powdered form, as you learned in Chapter 17, provides more surface area for reaction. The zinc is mixed in a paste with potassium hydroxide, a strong alkali base, and the paste is contained in a steel case. The cathode mixture is manganese(IV) oxide, also mixed with potassium hydroxide. The zinc anode half-cell reaction is

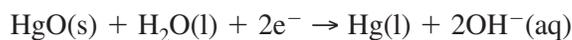


The cathode half-cell reaction is



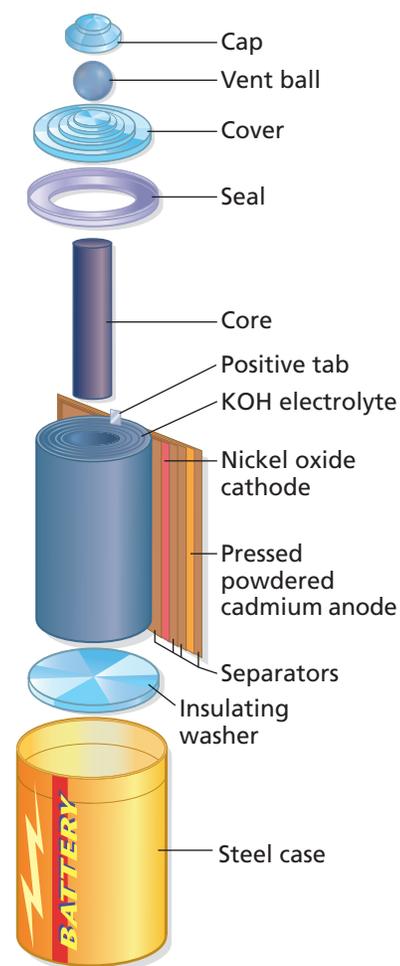
Alkaline batteries do not need the carbon rod cathode and are, therefore, smaller and more compatible with smaller devices.

The mercury battery shown in **Figure 21-8c** is smaller yet and is used to power devices such as hearing aids and calculators. The mercury battery uses the same anode half-reaction as the alkaline battery, with this cathode half-reaction.

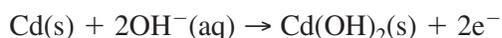


Primary and secondary batteries

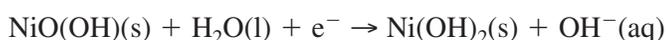
Batteries are divided into two types depending on their chemical processes. The zinc–carbon, alkaline–zinc, and mercury cells are classified as primary batteries. **Primary batteries** produce electric energy by means of redox reactions that are not easily reversed. These cells deliver current until the reactants are gone, and then the battery is discarded. Batteries that are rechargeable depend on reversible redox reactions. They are called **secondary batteries**. What devices can you identify that use secondary batteries? A car battery and the battery in a laptop computer are examples of secondary batteries, sometimes called *storage batteries*.



The storage batteries that power devices such as cordless drills and screwdrivers, shavers, and camcorders are usually nickel–cadmium rechargeable batteries, sometimes called NiCad batteries, as shown in **Figure 21-9**. For maximum efficiency, the anode and cathode are long, thin ribbons of material separated by a layer that ions can pass through. The ribbons are wound into a tight coil and packaged into a steel case. The anode reaction that occurs when the battery is used to generate electric current is the oxidation of cadmium in the presence of a base.



The cathode reaction is the reduction of nickel from the +3 to the +2 oxidation state.

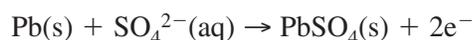


When the battery is recharged, these reactions are reversed.

Lead–Acid Storage Battery

Another common storage battery is the lead–acid battery. The standard automobile battery is an example of a lead–acid battery. Most auto batteries of this type contain six cells that generate about 2 V each for a total output of 12 V. The anode of each cell consists of two or more grids of porous lead, and the cathode consists of lead grids filled with lead(IV) oxide. This type of battery probably should be called a lead–lead(IV) oxide battery, but the term *lead–acid* is commonly used because the battery's electrolyte is a solution of sulfuric acid.

The following equation represents the oxidation half-cell reaction at the anode where lead is oxidized from the zero oxidation state to the +2 oxidation state.



The reduction of lead from the +4 to the +2 oxidation state takes place at the cathode. The half-cell reaction for the cathode is

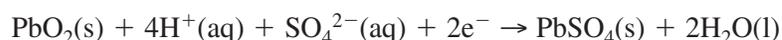


Figure 21-9

Cordless tools and phones often use rechargeable batteries, such as the NiCad type, for power. The battery pack in the phone is recharged when the handset is replaced on the base. In this case, the base is plugged into an electrical outlet, which supplies the power to drive the nonspontaneous recharge reaction.

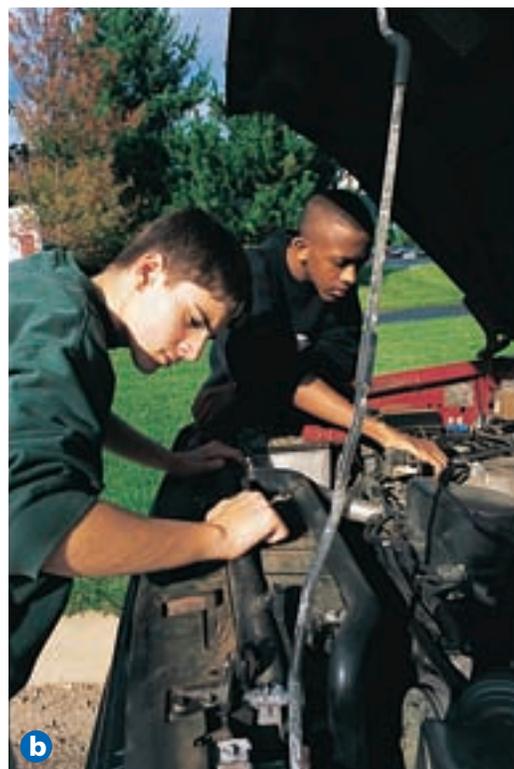
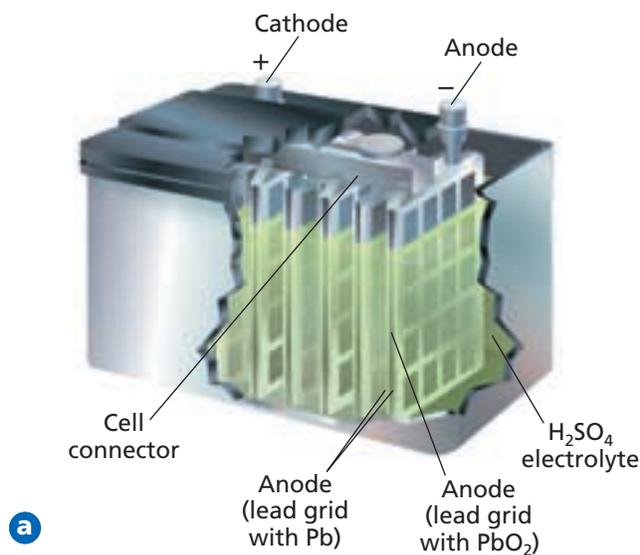


Figure 21-10

a Lead–acid batteries contain lead plates and lead(IV) oxide plates. The electrolyte is a solution of sulfuric acid. When the battery is in use, the sulfuric acid is depleted and the electrolyte becomes less dense. **b** The condition of the battery can be checked by measuring the density of the electrolyte solution.

The overall reaction is



By looking at the half-cell reactions, you can see that lead(II) sulfate (PbSO_4) is the reaction product in both oxidation and reduction. Also, Pb , PbO_2 , and PbSO_4 are solid substances, so they stay in place where they are formed. Thus, whether the battery is discharging or charging, the reactants are available where they are needed. Sulfuric acid serves as the electrolyte in the battery, but, as the overall cell equation shows, it is depleted as the battery generates electric current. For this reason, it is possible to measure a battery's state of charge by measuring the density of the electrolyte. Sulfuric acid (H_2SO_4) has almost twice the density of water. As the H_2SO_4 is used up, the solution becomes less dense. What happens when the battery is recharging? In this case, the reactions reverse, forming lead and lead(IV) oxide and releasing sulfuric acid, shown as $4\text{H}^+ + 2\text{SO}_4^{2-}$ in the equation.

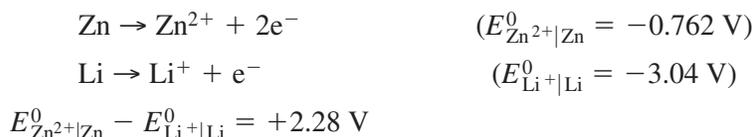
The lead-storage battery shown in **Figure 21-10** is a good choice for motor vehicles because it provides a large initial supply of energy to start the engine, has a long shelf-life, and is reliable at low temperatures.

Lithium Batteries

Although lead–acid batteries are reliable and suitable for many applications, the recent trend has been, and continues to be, the development of batteries with less mass and higher capacity to power devices from wristwatches to electric cars. For applications where a battery is the key component and must provide a significant amount of power, such as for the operation of electric cars, lead–acid batteries are too heavy to be feasible.

The solution is to develop lightweight batteries that store a large amount of energy for their size. Scientists and engineers have focused a lot of attention on the element lithium for two main reasons: lithium is the lightest

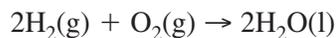
known metal and lithium has the lowest standard reduction potential of the metallic elements, -3.04 V (see **Table 21-1**). This means that a battery that oxidizes lithium at the anode can generate almost 2.3 volts more than a similar battery in which zinc is oxidized. Compare the two oxidation half-reactions and their standard reduction potentials.



Lithium batteries can be either primary or secondary, depending on which reduction reactions are coupled to the oxidation of lithium. For example, some lithium batteries use the same cathode reaction as zinc–carbon dry cells, the reduction of manganese(IV) oxide (MnO_2) to manganese(III) oxide (Mn_2O_3). These batteries produce electric current at about 3 V compared to 1.5 V for zinc–carbon cells. Lithium batteries last much longer than other kinds of batteries. As a result, they are often used in watches, computers, and cameras to maintain time, date, memory, and personal settings—even when the device is turned off.

Fuel Cells

The dynamic reaction in which a fuel, such as hydrogen gas or methane, burns may seem quite different from the relatively calm redox reactions that take place in batteries. However, the burning of fuel shown in **Figure 21-11** also is an oxidation–reduction reaction. What happens when hydrogen burns in air?



In this reaction, hydrogen is oxidized from zero oxidation state in H_2 to a +1 oxidation state in water. Oxygen is reduced from zero oxidation state in O_2 to -2 in water. When hydrogen burns in air, hydrogen atoms form polar covalent bonds with oxygen, in effect sharing electrons directly with the highly electronegative oxygen atom. This reaction is analogous to the reaction that occurs when a zinc strip is immersed in a solution of copper(II) sulfate. Electrons from the oxidation of zinc transfer directly to copper(II) ions, reducing them to atoms of copper metal. Because the zinc atoms and copper ions are in close contact, there is no reason for electrons to flow through an external circuit.

You saw in Section 21.1 how the oxidation and reduction processes can be separated to cause electrons to flow through a wire. The same can be done so that a fuel “burns” in a highly controlled way while generating electric current. A **fuel cell** is a voltaic cell in which the oxidation of a fuel is used to produce electric energy. Although many people believe the fuel cell to be a modern invention, the first one was demonstrated in 1839 by William Grove (1811–1896), a British electrochemist. He called his cell a “gas battery.” It was not until the 1950s, when scientists began working earnestly on the space program, that efficient, practical fuel cells were developed. As in other voltaic



Careers Using Chemistry

Electrochemist

Do you like to experiment with batteries and other devices that generate electric current? Then consider becoming an electrochemist.

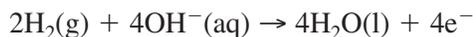
Electrochemists study electrons. The presence, absence, and movement of electrons often determine whether a chemical reaction will take place. Electrochemists work with semiconductors, medical sensors, metal plating, and anti-corrosion technology in the automobile, medical, and computer industries and in other fields.

Figure 21-11

This oxyhydrogen torch uses hydrogen as its fuel and oxidizes hydrogen to water in a vigorous combustion reaction. Like the torch, fuel cells also oxidize hydrogen to water, but fuel cells operate at a much more controlled rate.

cells, an electrolyte is required so that ions can migrate between electrodes. In the case of the fuel cell, a common electrolyte is an alkaline solution of potassium hydroxide.

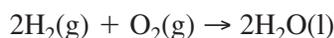
In a fuel cell, each electrode is a hollow chamber of porous carbon walls that allow contact between the inner chamber and the electrolyte surrounding it. The walls of the chamber also contain catalysts, such as powdered platinum or palladium, which speed up the reactions. These catalysts are similar to those in an automobile's catalytic converter, which you read about in Chapter 17. The following oxidation half-reaction takes place at the anode.



The reaction uses the hydroxide ions abundant in the alkaline electrolyte and releases electrons to the anode. Electrons from the oxidation of hydrogen flow through the external circuit to the cathode where the following reduction half-reaction takes place



The electrons allow the reduction of oxygen in the presence of water to form four hydroxide ions, which replenish the hydroxide ions used up at the anode. By combining the two half-reactions, the overall cell reaction is determined. As you can see, it is the same as the equation for the burning of hydrogen in oxygen.



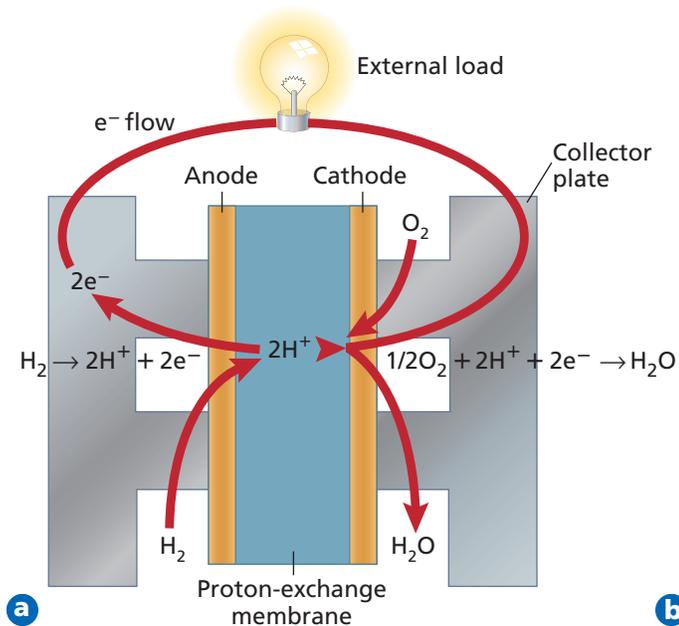
The hydrogen–oxygen fuel cell also burns hydrogen in a sense, but the burning is controlled so that most of the chemical energy is converted to electrical energy, not to heat.

In understanding fuel cells, it might be helpful to think of them as reaction chambers into which oxygen and hydrogen are fed from outside sources and from which the reaction product, water, is removed. Because the fuel for the cell is provided from an outside source, fuel cells never “run down.” They keep on producing electricity as long as fuel is fed to them.

Newer fuel cells, such as the cell in the **problem-solving LAB** and the cell shown in **Figure 21-12**, use a plastic sheet called a proton-exchange membrane,

Figure 21-12

a In this fuel cell, hydrogen is the fuel. The half-reactions are separated by a proton-exchange membrane so that the electrons lost in oxidation flow through an external circuit to reach the site of reduction. As electrons travel through the external circuit, they can do useful work, such as running electric motors. The byproduct of this redox reaction is water, a harmless substance. **b** A “stack” of PEM-type cells can generate enough energy to power an electric car.



Interpreting Scientific Illustrations

Can the simplest atom power a car?

Daimler–Chrysler’s NECAR IV may be a preview of the future of automobiles. The NECAR IV is a compact car that is powered by a hydrogen fuel cell (HFC). This model can reach speeds of 90 mph (145 km/h), carry up to five passengers and cargo, and travel 280 miles (450 km) before refueling. It runs on oxygen from the air and pure hydrogen supplied in a tank. Its exhaust is water, making it pollution free.

Analysis

To power a 1200-kg automobile, a hydrogen fuel cell must produce about 144 volts. Although the

cell potential for a hydrogen fuel cell is 1.229 V, approximately 43% of that voltage is lost to the production of heat, meaning that each cell’s actual potential is closer to 0.7 V. To generate enough voltage to power a car, the HFCs must be stacked, similar to the stack shown in **Figure 21-12b**.

When the electrodes are connected to a device, such as an electric motor, the electrons, which cannot travel through the PEM, travel through the external circuit and turn the motor.

Thinking Critically

1. How many cells would be needed to power a 1200-kg automobile? What would be the length of the stack if one cell is 1.2 mm thick? Is it small enough to fit under the hood?
2. Compare and contrast a stack of cells with a voltaic pile.

or PEM, instead of a liquid electrode to separate the reactions. PEMs are not corrosive and they are safer and lighter in weight than liquid electrodes. Hydrogen ions (H^+) are protons, and they can pass directly through the membrane from the anode (where hydrogen is oxidized) to the cathode. There, they combine with oxygen molecules and electrons returning from the external circuit to form water molecules, which are released as steam.

One potential application for fuel cells is as an alternative power source for automobiles. However, as the **Chemistry and Technology** feature at the end of this chapter explains, scientists must resolve some fundamental challenges before fuel cells power the cars we drive.

Corrosion

So far in this chapter, you have examined the spontaneous redox reactions in voltaic cells. Spontaneous redox reactions also occur in nature. A prime example is corrosion, usually called rusting, of iron. **Corrosion** is the loss of metal resulting from an oxidation–reduction reaction of the metal with substances in the environment. Although rusting is usually thought of as a reaction between iron and oxygen, it is more complex. Can you recall ever seeing or reading about corrosion happening in perfectly dry air or in water that contains no dissolved oxygen? Probably not because both water and oxygen must be present for rusting to take place. For this reason, the steel hulls of ships, such as that shown in **Figure 21-13**, are especially susceptible to corrosion in the form of rust.

Rusting usually begins where there is a pit or small break in the surface of the iron. This

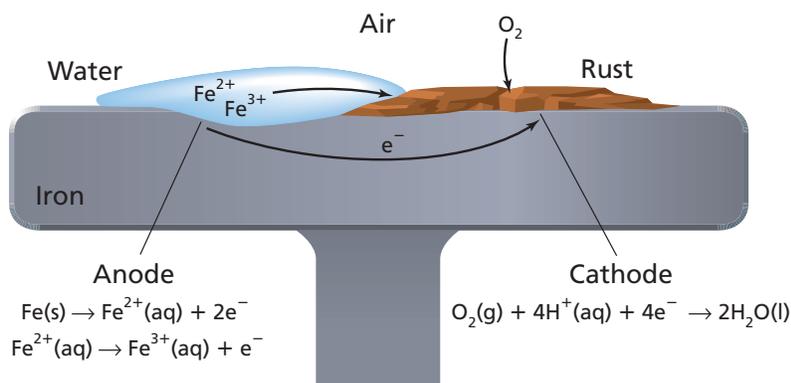
Figure 21-13

The steel hull of oceangoing ships offers strength, reliability, and durability, but steel is susceptible to corrosion. Ships such as this one in St. John Harbour, New Brunswick, Canada, often show the telltale sign of corrosion, which is rust.

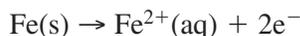


Figure 21-14

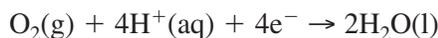
Corrosion occurs when air, water, and iron set up a voltaic cell similar to the conditions shown at the surface of this iron I-beam.



region becomes the anode of the cell as iron atoms begin to lose electrons, as illustrated in **Figure 21-14**.



The iron(II) ions become part of the water solution while the electrons move through the iron to the cathode region. In effect, the piece of iron becomes the external circuit as well as the anode. The cathode is usually located at the edge of the water drop where water, iron, and air come in contact. Here, the electrons reduce oxygen from the air in this half-reaction.

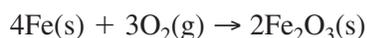


The supply of H^{+} ions is probably furnished by carbonic acid formed when CO_2 from the air dissolves in water.

Next, the Fe^{2+} ions in solution are oxidized to Fe^{3+} ions by reaction with oxygen dissolved in the water. The Fe^{3+} ions combine with oxygen to form insoluble Fe_2O_3 , rust.



Combining the three equations yields the overall cell reaction for the corrosion of iron.

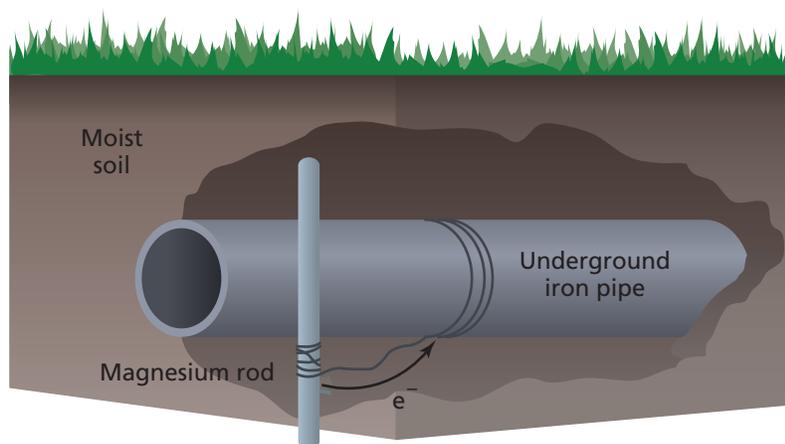
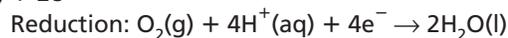
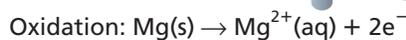


You can observe the process of corrosion first-hand by performing the **miniLAB** that results in the corrosion of an iron nail.

**a****Figure 21-15**

Because corrosion can cause considerable damage, it is important to investigate ways to prevent rust and deterioration.

- a** Paint or another protective coating is one way to protect steel structures from corrosion.
- b** Sacrificial anodes of magnesium or other active metals are also used to prevent corrosion.

**b**

Rusting is a rather slow process because water droplets have few ions and therefore are not good electrolytes. However, if the water contains abundant ions, as in seawater or in regions where roads are salted in winter, corrosion occurs much faster because the solutions are excellent electrolytes.

It has been estimated that corrosion of cars, bridges, ships, the structures of buildings, and other metallic objects costs more than 100 billion dollars a year in the U.S. alone. For this reason, people have devised several means to minimize corrosion. One is simply to apply a coat of paint to seal out both air and moisture, but, because paint deteriorates, objects such as the ship's hull shown in **Figure 21-15a** must be repainted often.

The steel hulls of ships are constantly in contact with saltwater, so the prevention of corrosion is vital. Although the hull may be painted, another method is used to minimize corrosion. Blocks of metals, such as magnesium, aluminum, or titanium, that oxidize more easily than iron are placed in contact with the steel hull. These blocks rather than the iron in the hull become the anode of the corrosion cell. As a result, these blocks, called sacrificial anodes, are corroded while the iron in the hull is spared. Of course, the sacrificial anodes must be replaced before they corrode away completely, leaving the ship's hull unprotected. A similar technique is used to protect iron pipes that are run underground. Magnesium bars are attached to the pipe by wires, and these bars corrode instead of the pipe, as shown in **Figure 21-15b**.

Another approach to preventing corrosion is to coat iron with another metal that is more resistant to corrosion. In the **galvanizing** process, iron is

miniLAB

Corrosion

Comparing and Contrasting A lot of money is spent every year correcting and preventing the effects of corrosion. Corrosion is a real-world concern of which everyone needs to be aware.

Materials iron nails (4); magnesium ribbon (2 pieces, each about 5 cm long); copper metal (2 pieces, each about 5 cm long); 150-mL beakers (4); distilled water; saltwater solution; sandpaper.

Procedures



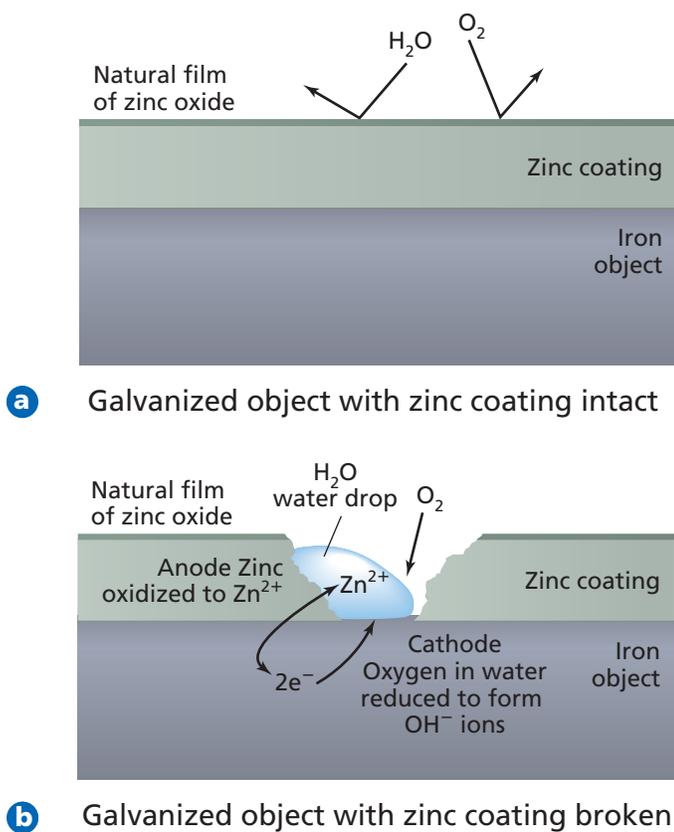
1. Use the sandpaper to buff the surfaces of each nail. Wrap two nails with the magnesium ribbon and two nails with the copper. Wrap the metals tightly enough so that the nails do not slip out.
2. Place each of the nails in a separate beaker. Add distilled water to one of the beakers containing a copper-wrapped nail and one of the beakers containing a magnesium-wrapped nail. Add enough distilled water to just cover the wrapped nails. Add saltwater to the other two beakers. Record your observations for each of the beakers.



3. Let the beakers stand overnight in the warmest place available. Examine the nails and solutions the next day and record your observations.

Analysis

1. Describe the difference between copper-wrapped nails in the distilled water and saltwater after standing overnight.
2. Describe the difference between the magnesium-wrapped nails in the distilled water and saltwater.
3. In general, what is the difference between a copper-wrapped nail and a magnesium-wrapped nail?



c

Figure 21-16

Galvanizing helps prevent corrosion in two ways. **a** The zinc coating seals the iron from air and water by forming a barrier of zinc oxide that repels water and oxygen. **b** If the zinc coating breaks, the zinc acts as a sacrificial anode. **c** Metal objects that are left outside are often galvanized to prevent rust and corrosion caused by the elements.

coated with a layer of zinc by either dipping the iron object into molten zinc or by electroplating the zinc onto it. Although zinc is more readily oxidized than iron, it is one of the *self-protecting metals*, a group that also includes aluminum and chromium. When exposed to air, these metals oxidize at the surface, but the thin metal oxide coating clings tightly to the metal and seals it from further oxidation.

Galvanizing protects iron in two ways. As long as the zinc layer is intact, water and oxygen cannot reach the iron's surface. Inevitably, though, the zinc coating cracks. When this happens, zinc protects iron from rapid corrosion by becoming the anode of the voltaic cell set up when water and oxygen contact iron and zinc at the same time. **Figure 21-16** illustrates how these two forms of corrosion protection work.

Section 21.2 Assessment

- What is reduced and what is oxidized in the ordinary zinc–carbon dry cell battery? What features make the alkaline dry cell an improvement over the earlier type of dry cell battery?
- Explain how primary and secondary batteries differ. Give an example of each type.
- Explain why a fuel cell does not run down like other batteries.
- What is a sacrificial anode? How is one used?
- Thinking Critically** Standard dry cell batteries have a relatively short shelf life. Explain why older dry-cell batteries may not have the same power output as newer batteries. How can the shelf life of these batteries be extended?
- Calculating** Use data from **Table 21-1** to calculate the cell potential of the hydrogen–oxygen fuel cell described in this section.

In Section 21.2, you learned that rechargeable batteries can regain their electrical potential and be reused. You learned that a current can be introduced to cause the reverse reaction, which is not spontaneous, to happen. This aspect of electrochemistry—the use of an external force, usually in the form of electricity, to drive a chemical reaction—is important and has many practical applications.

Reversing Redox Reactions

When a battery generates electricity, electrons given up at the anode flow through an external circuit to the cathode where they are used in a reduction reaction. A secondary battery is one that can be recharged by passing a current through it in the opposite direction. The current source can be a generator or even another battery. The energy of this current reverses the cell redox reaction and regenerates the battery's original substances.

To understand this process, look at **Figure 21-17**, which shows an electrochemical cell powering a light bulb by a spontaneous redox reaction. The beakers on the left contain a zinc strip in a solution of zinc ions. The beakers on the right contain a copper strip in a solution of copper ions. You saw in Section 21.1 how electrons in this system flow from the zinc side to the copper side, creating an electric current. As the reaction continues, the zinc strip deteriorates, while copper from the copper ion solution is deposited as copper metal on the copper strip. Over time, the flow of electrons decreases and the strength of the electrical output diminishes to the point that the bulb will not light. However, the cell can be regenerated if current is applied in the reverse direction using an external voltage source. The reverse reaction is nonspontaneous, which is why the voltage source is required. If the voltage source remains long enough, eventually the cell will return to near its original strength and can again produce electrical energy to light the bulb.

The use of electrical energy to bring about a chemical reaction is called **electrolysis**. An electrochemical cell in which electrolysis occurs is called an **electrolytic cell**. For example, when a secondary battery is recharged, it is acting as an electrolytic cell.

Objectives

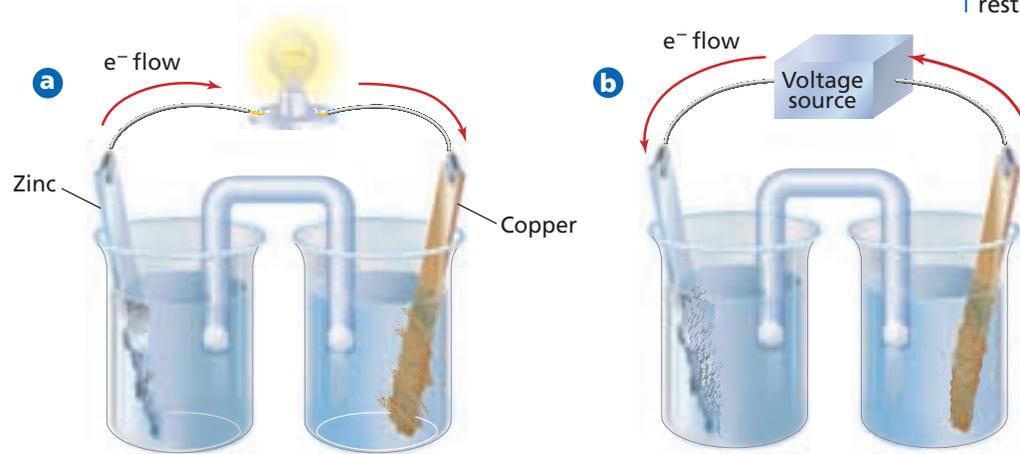
- **Describe** how it is possible to reverse a spontaneous redox reaction in an electrochemical cell.
- **Compare** the reactions involved in the electrolysis of molten sodium chloride with those in the electrolysis of brine.
- **Discuss** the importance of electrolysis in the smelting and purification of metals.

Vocabulary

electrolysis
electrolytic cell

Figure 21-17

a Electrons in this zinc–copper electrochemical cell flow from the zinc strip to the copper strip, causing an electric current that powers the light bulb. As the spontaneous reaction continues, much of the zinc strip is oxidized to zinc ions and the copper ions are reduced to copper metal, which is deposited on the copper strip. **b** If an outside voltage source is applied to reverse the flow of electrons, the original conditions of the cell are restored.



History

CONNECTION

Electrolysis can be useful to clean historic objects recovered from shipwrecks. Coatings of salts from the seawater on metal objects are removed by an electrochemical process. A voltaic cell is set up with a cathode that is the object itself and a stainless steel anode in a basic solution. Chloride ions are removed when the electric current is turned on.

In another process, bacteria convert sulfate ions to hydrogen sulfide gas and cause silver coins and bars to become coated with silver sulfide after long periods of time at the bottom of the ocean. In an electrolytic cell, the silver in silver sulfide can be reduced to silver metal and reclaimed.



Applications of Electrolysis

Recall that the voltaic cells convert chemical energy to electrical energy as a result of a spontaneous redox reaction. Electrolytic cells do just the opposite: they use electrical energy to drive a nonspontaneous reaction. A common example is the electrolysis of water. In this case, an electric current decomposes water into hydrogen and oxygen.



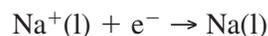
The electrolysis of water is one method by which hydrogen gas can be generated for commercial use. Aside from the decomposition of water, electrolysis has many other practical applications.

Electrolysis of molten NaCl Just as electrolysis can decompose water into its elements, it also can separate molten sodium chloride into sodium metal and chlorine gas. This process, the only practical way to obtain elemental sodium, is carried out in a chamber called a Down's cell, as shown in **Figure 21-18**. The electrolyte in the cell is the molten sodium chloride itself. Remember that ionic compounds can conduct electricity only when their ions are free to move, such as when they are dissolved in water or are in the molten state.

The anode reaction of the Down's cell is the oxidation of chloride ions.



At the cathode, sodium ions are reduced to atoms of sodium metal.



The net cell reaction is

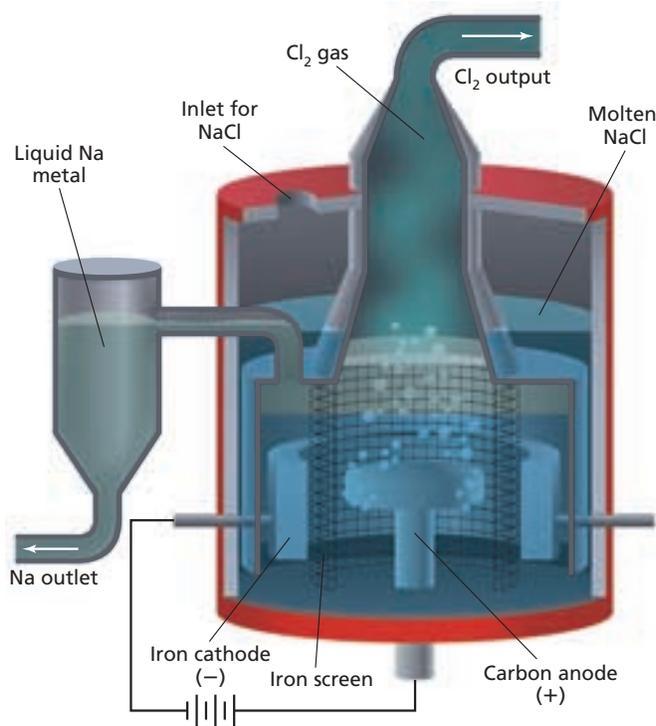
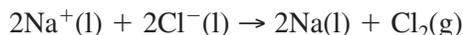


Figure 21-18

In a Down's cell, electrons supplied by a generator are used to reduce sodium ions. As electrons are removed from the anode, chloride ions are oxidized to chlorine gas.

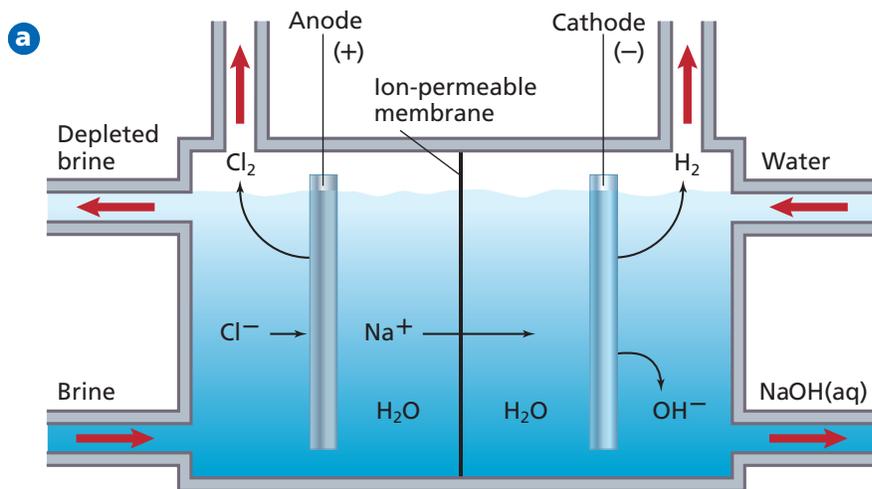
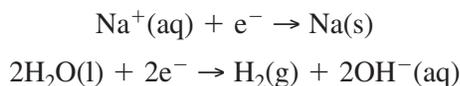


Figure 21-19

- a** Commercial facilities use an electrolytic process to derive hydrogen gas, chlorine gas, and sodium hydroxide from brine. **b** Chlorine gas is used to manufacture polyvinyl chloride products such as these irrigation pipes. **c** Sodium hydroxide is the key ingredient in drain cleaners.

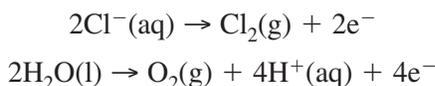


Electrolysis of brine The decomposition of brine, an aqueous solution of sodium chloride, is another process that is accomplished by electrolysis. **Figure 21-19** illustrates the electrolytic cell and products of the electrolysis of brine. Two reactions are possible at the cathode, the reduction of sodium ions and the reduction of hydrogen in water molecules.



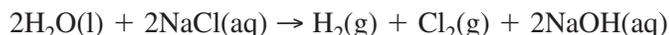
Here, a key product of reduction is hydrogen gas.

Two reactions are possible at the anode, the oxidation of chloride ions and the oxidation of oxygen in water molecules.



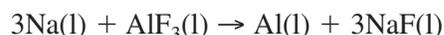
When the concentration of chloride ions is high, the primary product of oxidation is chlorine gas. However, as the brine solution becomes dilute, oxygen gas becomes the primary product.

To appreciate the commercial importance of this process, look at the overall cell reaction.



All three products of the electrolysis of brine—hydrogen gas, chlorine gas, and sodium hydroxide—are substances that are commercially important to industry.

Aluminum manufacture Aluminum is the most abundant metallic element in Earth's crust, but until the late nineteenth century, aluminum metal was more precious than gold. Aluminum was expensive because no one knew how to purify it in large quantities. Instead, it was produced by a tedious and expensive small-scale process in which metallic sodium was used to reduce aluminum ions in molten aluminum fluoride to metallic aluminum.



In 1886, 22-year-old Charles Martin Hall (1863–1914) developed a process to produce aluminum by electrolysis using heat from a blacksmith

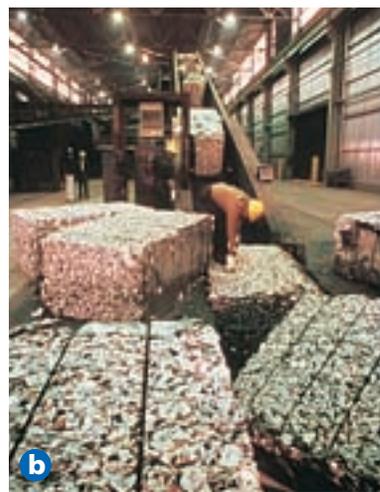
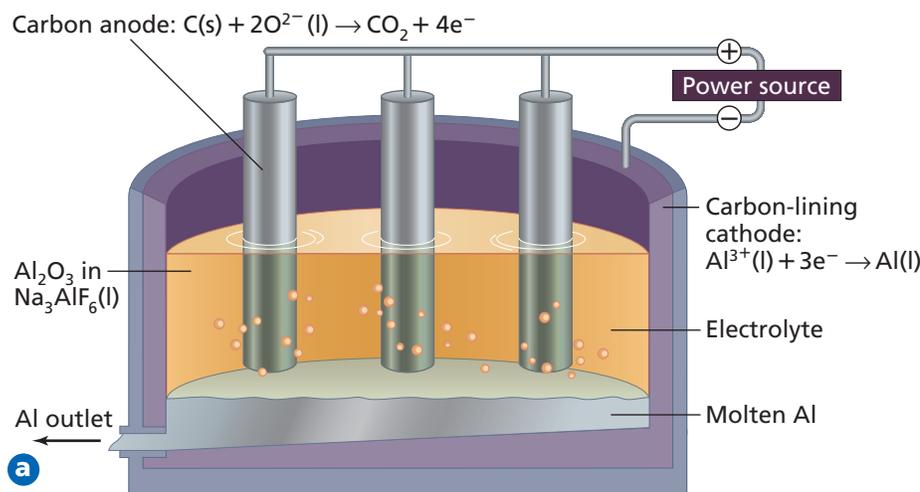
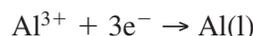


Figure 21-20

a Aluminum is produced by the Hall–Héroult process in cells similar to this one. Note that the cathode is the carbon (graphite) lining of the cell itself. **b** Every ton of aluminum that is recycled saves huge quantities of electrical energy that would be spent to produce new aluminum from ore.

forge, electricity from home-made batteries, and his mother’s iron skillet as electrodes. At almost the same time, one of Le Châtelier’s students, Paul L. T. Héroult (1863–1914), also 22 years old, discovered the same process. Today, it is called the Hall–Héroult process, which is illustrated in **Figure 21-20**.

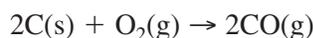
In the modern version of this process, aluminum metal is obtained by electrolysis of aluminum oxide, which is refined from bauxite ore ($Al_2O_3 \cdot 2H_2O$). The aluminum oxide is dissolved at $1000^\circ C$ in molten synthetic cryolite (Na_3AlF_6), another aluminum compound. The cell is lined with graphite, which forms the cathode for the reaction. Another set of graphite rods is immersed in the molten solution as an anode. The following half-reaction occurs at the cathode.



The molten aluminum settles to the bottom of the cell and is drawn off periodically. Oxide ions are oxidized at the cathode in this half-reaction.

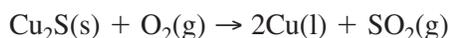


Because temperatures are high, the liberated oxygen reacts with the carbon of the anode to form carbon monoxide.



The Hall–Héroult process uses huge amounts of electric energy. For this reason, aluminum often is produced in plants built close to large hydroelectric power stations, where electric energy is abundant and less expensive. The vast amount of electricity needed to produce aluminum from ore is the prime reason that it is especially important to recycle aluminum. Recycled aluminum already has undergone electrolysis, so the only energy required to make it usable again is the heat used to melt it in a furnace.

Purification of ores Another application of electrolysis is in the purification of metals such as copper. Most copper is mined in the form of the ores chalcopyrite ($CuFeS_2$), chalcocite (Cu_2S), and malachite ($Cu_2CO_3(OH)_2$). The sulfides are most abundant and yield copper metal when heated strongly in the presence of oxygen.



The copper from this process contains many impurities and must be refined, so the molten copper is cast into large, thick plates. These plates are then used as an anode in an electrolytic cell containing a solution of copper(II) sulfate. The cathode of the cell is a thin sheet of pure copper. As current is passed through the cell, copper atoms in the impure anode are oxidized to copper(II) ions, which migrate through the solution to the cathode where they are reduced to copper atoms. These atoms become part of the cathode while impurities fall to the bottom of the cell.

Electroplating Objects can be electroplated with a metal such as silver in a method similar to that used to refine copper. The object to be silver plated is the cathode of an electrolytic cell that has a silver anode, as shown in **Figure 21-21**. At the cathode, silver ions present in the electrolyte solution are reduced to silver metal by electrons from an external power source. The silver forms a thin coating over the object being plated. The anode consists of a silver bar or sheet, which is oxidized to silver ions as electrons are removed by the power source. Current passing through the cell must be carefully controlled in order to get a smooth, even metal coating.



Figure 21-21

In an electrolytic cell used for silver plating, the object to be plated is the cathode where silver ions in the electrolyte solution are reduced to silver metal and deposited on the object.

Section 21.3 Assessment

20. Define electrolysis and relate the definition to spontaneity of redox reactions.
21. What are the products of the electrolysis of brine? Of the electrolysis of molten sodium chloride? Explain why the reaction products differ.
22. Describe the process by which the copper that results from smelting of ore is purified by electrolysis.
23. **Thinking Critically** Suppose you want to plate an object with gold by electrolysis. What sort of substances would you need to have in the electrolyte solution? What would you use as the cathode and the anode of the cell?
24. **Inferring** Producing a kilogram of silver from its ions by electrolysis requires much less electrical energy than producing a kilogram of aluminum from its ions. Give a reason for this difference.

Voltaic Cell Potentials

A voltaic cell converts chemical energy into electrical energy. It consists of two parts called half-cells. When two different metals, one in each half-cell, are used in the voltaic cell, a potential difference is produced. In this experiment, you will measure the potential difference of various combinations of metals used in voltaic cells and compare these values to the values found in the standard reduction potentials table.

Problem

How can you measure the potential of voltaic cells?

Objectives

- **Construct** voltaic cells using various combinations of metals for electrodes.
- **Design** the arrangement of the voltaic cells in a microplate in such a way as to use materials efficiently.
- **Determine** which metals are the anode and cathode in voltaic cells.
- **Compare** the experimental cell potential to the theoretical value found in **Table 21-1**.

Materials

metal strips (approximately 0.6 cm by 1.3 cm) of copper, aluminum, zinc, and magnesium
 1M copper(II) nitrate
 1M aluminum nitrate
 1M zinc nitrate
 1M magnesium nitrate
 24-well microplate
 Beral-type pipette (5)

CBL System
 voltage probe
 filter paper (6 pieces size 0.6 cm by 2.5 cm)
 1M potassium nitrate
 forceps
 steel wool or sandpaper
 table of standard reduction potentials

Safety Precautions



- Always wear goggles and an apron in the lab.
- The chemicals used in this experiment are eye and skin irritants. Wash thoroughly if they are spilled on the skin.

Pre-Lab

1. Read the entire CHEMLAB.
2. Plan and organize how you will arrange voltaic cells in the 24-well microplate using the four metal combinations so that your time and materials will be used in the most efficient manner possible. Have your instructor approve your plan before you begin the experiment.
3. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table similar to the example below in which to record your observations.
4. Review the definition of a voltaic cell.

Voltaic Cell Potential Data

Anode metal (black)	Cathode metal (red)	Actual cell potential (V)	Anode half-reaction and theoretical potential	Cathode half-reaction and theoretical potential	Theoretical cell potential	% Error

- Review the purpose of a salt bridge in the voltaic cell. In this experiment, the filter paper strips soaked in potassium nitrate are the salt bridges.
- Review the equation to calculate cell potential.
- For the voltaic cell $\text{Mg} | \text{Mg}^{2+} || \text{Hg}^{2+} | \text{Hg}$, identify which metal is the anode and which metal is the cathode. Which metal is being oxidized and which metal is being reduced? What is the theoretical potential for this voltaic cell?
- Review the equation to calculate percent error.

Procedure

- Prepare the CBL to read potential differences (voltage). Plug the voltage probe into Channel 1. Turn the CBL on. Push the MODE button once to activate the voltmeter function.
- Soak the strips of filter paper in 2 mL of potassium nitrate solution. These are the salt bridges for the experiment. Use forceps to handle the salt bridges.
- Using the plan from your Pre-Lab, construct voltaic cells using the four metals and 1 mL of each of the solutions. Remember to minimize the use of solutions. Put the metals in the wells that contain the appropriate solution (for example, put the zinc metal in the solution with zinc nitrate). Use a different salt bridge for each voltaic cell. If you get a negative value for potential difference, switch the leads of the probe on the metals.



- Record which metals are the anode and cathode in each cell in the data table. The black lead of the probe will be attached to the metal that acts as the anode. The red lead will be attached to the cathode.
- Record the cell potential of each cell.

Cleanup and Disposal

- Use forceps to remove the metals from the microplate.
- Rinse the solution off the metal pieces with water, then use steel wool or sandpaper to clean them.
- Rinse the wells of the microplate.
- Return each metal to its correct container.

Analyze and Conclude

- Applying Concepts** Write the half-reactions for the anode and cathode in each of the voltaic cells in the data table. Look up the half-reaction potentials from the standard reduction potentials table (**Table 21-1**) and record these in the data table.
- Using Numbers** Calculate the theoretical potential for each voltaic cell and record it in the data table.
- Predicting** Using your data, rank the metals you used in order of most active to least active.
- Using Models** Calculate the percent error of the voltaic cell potential.
- Error Analysis** Why is the percent error calculated in step 4 large for some voltaic cells and small for others?

Real-World Chemistry

- Why is lithium metal becoming a popular electrode in modern batteries? Use the standard reduction potentials table to help you answer this question.
- What type of battery is used in pacemakers to regulate a patient's heartbeat? What are some of the benefits of this battery?

CHEMISTRY and Technology

Fuel Cells

An important consideration for auto manufacturers today is fuel economy and pollution reduction. Fuel cells represent one option to achieve these goals.

Like a battery, a fuel cell produces electricity from a redox reaction. Unlike a battery, a fuel cell can generate electric current indefinitely because it oxidizes a continuous stream of fuel from an outside source.

Fuel Cell Basics

Fuel cells can use several kinds of fuel, including natural gas and petroleum products. However, these are fossil fuels that produce carbon dioxide, an undesirable byproduct, when oxidized. Another drawback of common fuel cells is that they operate at temperatures from 200°C to 1000°C, and some contain hot, caustic, liquid electrolyte.

To avoid these problems, engineers have focused on a cell in which the fuel is hydrogen gas, the oxidant is oxygen from the air, and the product is water vapor. One of the more promising hydrogen fuel cells is one in which the half-cell reactions are separated by a thin polymer sheet called a proton-exchange membrane (PEM). The PEM fuel cell operates at approximately 100°C, and the moist membrane itself is the electrolyte.

On the anode side of the membrane, a platinum catalyst causes H₂ molecules to dissociate into atoms. The PEM allows hydrogen protons to pass, while the electrons must travel an electrical circuit. On the cathode side of the membrane, the protons combine with oxygen from the air and the electrons from the circuit to form water.

Making Fuel Cells Practical

Before PEM hydrogen fuel cells become practical, several issues must be resolved. PEM cells are expensive, in part because of the platinum powder that catalyzes the reaction. Safe storage and delivery of hydrogen is another key issue. The vehicle must carry enough hydrogen to power the vehicle over an acceptable range. Chemists are investigating ways to store and deliver hydrogen to fuel the cells. For example, carbon cage molecules in the form of balls or tubes can trap large quantities of hydrogen. When pressure is reduced



and the temperature is raised, these cages release the hydrogen in gaseous form.

Probably the most critical consideration is the hydrogen source. Unfortunately, much of the hydrogen gas produced commercially comes from the hydrocarbons of fossil fuels, which defeats the goal of fossil fuel conservation. Hydrogen can be produced from water by electrolysis, but this requires electrical energy that is produced by processes involving the combustion of fossil fuels. Clearly, a suitable method of hydrogen production is required. However, even with their present limitations, fuel cells are a viable energy alternative for the future.

Investigating the Technology

- 1. Thinking Critically** If electrolysis is a safe and technologically sound method to generate hydrogen gas, how might the process be made more environmentally safe and economically practical?
- 2. Using Resources** Investigate the technology of carbon cage molecules, including fullerenes and nanotubes, and elaborate on their utility in hydrogen fuel cells.



Visit the Chemistry Web site at science.glencoe.com to find links about hydrogen fuel cells.

Summary

21.1 Voltaic Cells

- In a voltaic cell, the oxidation and reduction half-reactions of a redox reaction are separated and ions flow through a salt-bridge conductor.
- In a voltaic cell, oxidation takes place at the anode, and reduction takes place at the cathode.
- The standard potential of a half-cell reaction is the voltage it generates when paired with a standard hydrogen electrode. Standard potentials are measured at 25°C and 1 atm pressure with a 1M concentration of ions in the half-cells.
- The reduction potential of a half-cell is negative if it undergoes oxidation connected to a standard hydrogen electrode. The reduction potential of a half-cell is positive if it undergoes reduction when connected to a standard hydrogen electrode.
- The standard potential of a voltaic cell is the difference between the standard reduction potentials of the half-cell reactions.

21.2 Types of Batteries

- Batteries are voltaic cells packaged in a compact, usable form.
- A battery can consist of a single cell or multiple cells.
- Primary batteries can be used only once, whereas secondary batteries can be recharged.
- When a battery is recharged, electrical energy supplied to the battery reverses the direction of the redox reaction that takes place when the battery is delivering current. Thus, the original reactants are restored.

- Fuel cells are batteries in which the substance oxidized is a fuel such as hydrogen.
- Iron can be protected from corrosion: by applying a coating of another metal or paint to keep out air and water, or by attaching a piece of metal (a sacrificial anode) that is oxidized more readily than iron.

21.3 Electrolysis

- Electrical energy can be used to bring about non-spontaneous redox reactions that produce useful products. This process is called electrolysis and takes place in an electrolytic cell.
- Metallic sodium and chlorine gas may be obtained by the electrolysis of molten sodium chloride.
- Electrolysis of strong aqueous sodium chloride solution (brine) yields hydrogen gas and hydroxide ions at the cathode while producing chlorine gas at the anode.
- Metals such as copper can be purified by making them the anode of an electrolytic cell where they are oxidized to ions, which are then reduced to pure metal at the cathode.
- Objects can be electroplated by making them the cathode of an electrolytic cell in which ions of the desired plating metal are present.
- Aluminum is produced by the electrolysis of aluminum oxide. The process uses a great amount of electrical energy.

Key Equations and Relationships

- Potential of a voltaic cell (p. 669) $E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$

Vocabulary

- anode (p. 665)
- battery (p. 672)
- cathode (p. 665)
- corrosion (p. 679)
- dry cell (p. 673)
- electrochemical cell (p. 665)
- electrolysis (p. 683)
- electrolytic cell (p. 683)
- fuel cell (p. 677)
- galvanizing (p. 681)
- half-cell (p. 665)
- primary battery (p. 675)
- reduction potential (p. 666)
- salt bridge (p. 664)
- secondary battery (p. 675)
- standard hydrogen electrode (p. 666)
- voltaic cell (p. 665)

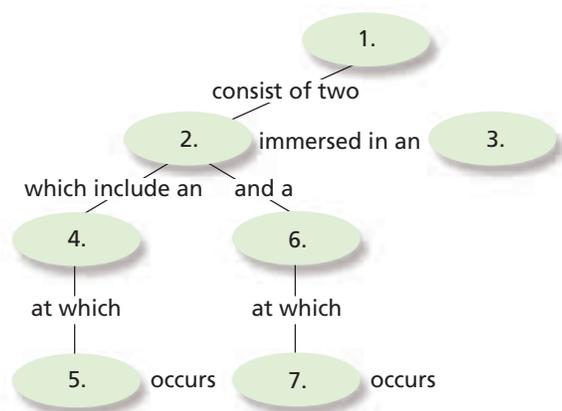


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 21 Assessment.

Concept Mapping

25. Complete the concept map using the following terms: reduction, electrodes, electrochemical cells, anode, oxidation, cathode, electrolyte.



Mastering Concepts

26. What feature of an oxidation–reduction reaction allows it to be used to generate an electric current? (21.1)
27. Describe the process that releases electrons in a zinc–copper voltaic cell. (21.1)
28. What is the function of a salt bridge in a voltaic cell? (21.1)
29. What information do you need in order to determine the standard voltage of a voltaic cell? (21.1)
30. In a voltaic cell represented by $\text{Al}|\text{Al}^{3+}||\text{Cu}^{2+}|\text{Cu}$, what is oxidized and what is reduced as the cell delivers current? (21.1)
31. Under what conditions are standard reduction potentials measured? (21.1)
32. What part of a zinc–carbon dry cell is the anode? Describe the reaction that takes place there. (21.2)
33. How do primary and secondary batteries differ? (21.2)
34. What substance is reduced in a lead–acid storage battery? What substance is oxidized? What substance, other than water, is produced in each reaction? (21.2)
35. List two ways that a fuel cell differs from an ordinary battery. (21.2)
36. What is galvanizing? How does galvanizing protect iron from corrosion? (21.2)
37. How can the spontaneous redox reaction of a voltaic cell be reversed? (21.3)
38. Why is an outside source of energy needed for electrolysis? (21.3)
39. Where does oxidation take place in an electrolytic cell? (21.3)
40. What reaction takes place at the cathode when molten sodium chloride is electrolyzed? (21.3)
41. Explain why the electrolysis of brine is done on a large scale at many sites around the world. (21.3)
42. Explain how recycling aluminum conserves energy. (21.3)

Mastering Problems

Voltaic Cells (21.1)

Use data from **Table 21-1** in the following problems. Assume that all half-cells are under standard conditions.

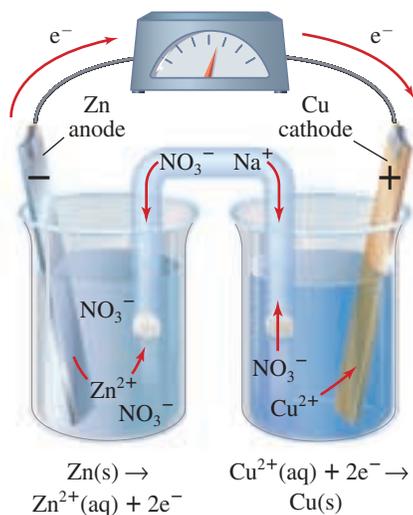
43. Write the standard cell notation for the following cells in which the half-cell listed is connected to the standard hydrogen electrode. An example is $\text{Na}|\text{Na}^+||\text{H}^+|\text{H}_2$.
- $\text{Zn}|\text{Zn}^{2+}$
 - $\text{Hg}|\text{Hg}^{2+}$
 - $\text{Cu}|\text{Cu}^{2+}$
 - $\text{Al}|\text{Al}^{3+}$
44. Determine the voltage of the cell formed by pairing each of the following half-cells with the standard hydrogen electrode.
- $\text{Cr}|\text{Cr}^{2+}$
 - $\text{Br}_2|\text{Br}^-$
 - $\text{Ga}|\text{Ga}^{3+}$
 - $\text{NO}_2|\text{NO}_3^-$
45. Determine whether each of the following redox reactions is spontaneous or nonspontaneous as written.
- $\text{Mn}^{2+} + 2\text{Br}^- \rightarrow \text{Br}_2 + \text{Mn}$
 - $\text{Fe}^{2+} + \text{Sn}^{2+} \rightarrow \text{Fe}^{3+} + \text{Sn}$
 - $\text{Ni}^{2+} + \text{Mg} \rightarrow \text{Mg}^{2+} + \text{Ni}$
 - $\text{Pb}^{2+} + \text{Cu}^+ \rightarrow \text{Pb} + \text{Cu}^{2+}$

46. Calculate the cell potential of voltaic cells that contain the following pairs of half-cells.
- Chromium in a solution of Cr^{3+} ions; copper in a solution of Cu^{2+} ions
 - Zinc in a solution of Zn^{2+} ions; platinum in a solution of Pt^{2+} ions
 - A half-cell containing both HgCl_2 and Hg_2Cl_2 ; lead in a solution of Pb^{2+} ions
 - Tin in a solution of Sn^{2+} ions; iodine in a solution of I^- ions

Mixed Review

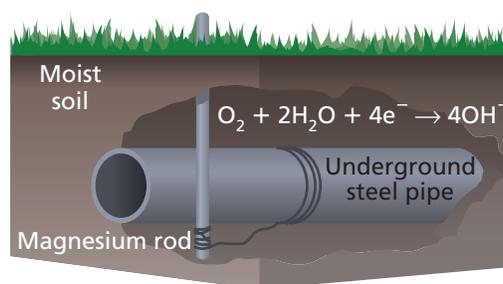
Sharpen your problem-solving skills by answering the following.

- Why do electrons flow from one electrode to the other in a voltaic cell?
- What substance is electrolyzed to produce aluminum metal?
- Write the oxidation and reduction half-reactions for a silver-chromium voltaic cell. Identify the anode, cathode, and electron flow.



- Primary cells are cells like the ordinary dry cell that cannot be recharged. Storage cells are cells like the nickel-cadmium cell or the lead-storage battery that can be recharged. Is the Zn-Cu^{2+} cell shown above a primary cell or a storage cell?
- Determine the voltage of the cell in which each of the following half-cells is connected to a $\text{Ag}|\text{Ag}^+$ half-cell.
 - $\text{Be}^{2+}|\text{Be}$
 - $\text{S}|\text{S}^{2-}$
 - $\text{Au}^+|\text{Au}$
 - $\text{I}_2|\text{I}^-$

- Explain why water is necessary for the corrosion of iron.
- In the electrolytic refining of copper, what factor determines which piece of copper is the anode and which is the cathode?
- Explain how the oxidation of hydrogen in a fuel cell differs from the oxidation of hydrogen when it burns in air.
- Lead-acid batteries and other rechargeable batteries are sometimes called *storage batteries*. Precisely what is being stored in these batteries?
- Buried steel pipeline can be protected against corrosion by cathodic protection. In this process, the steel pipe is connected to a more active metal, such as magnesium, that would corrode instead of the steel. Use the diagram to answer the following questions.

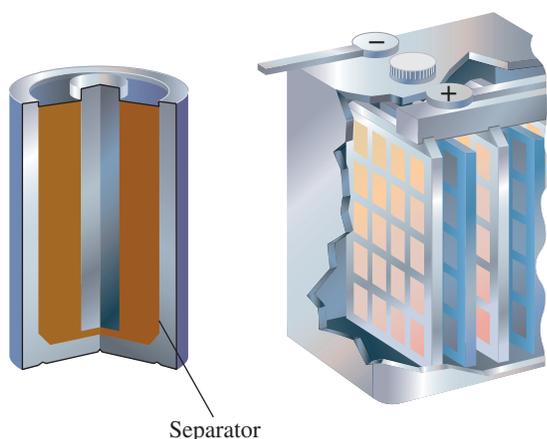


- What is the cathode? What is the anode?
- Describe briefly how the magnesium metal protects the steel.

Thinking Critically

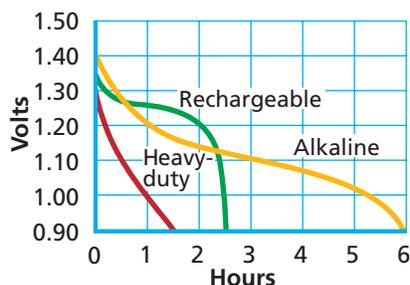
- Communicating** Write two or three sentences describing the processes that take place in a voltaic cell and account for the direction of electron flow. Use the words *cathode*, *anode*, *oxidation*, *reduction*, and *potential difference* in your sentences.
- Predicting** How would standard reduction potentials be different if scientists had chosen the $\text{Cu}^{2+}-\text{Cu}$ half-cell as a standard instead of the H^+-H_2 half-cell? What would the potential of the hydrogen electrode be if the copper electrode were the standard? How would the relationships among the standard reduction potentials change?
- Applying Concepts** Suppose that you have a voltaic cell in which one half-cell is made up of a strip of tin immersed in a solution of tin(II) ions.
 - How could you tell by measuring voltage whether the tin strip was acting as a cathode or an anode in the cell?
 - How could you tell by simple observation whether the tin strip was acting as a cathode or an anode?

- 60. Hypothesizing** The potential of a half-cell varies with concentration of reactants and products. For this reason, standard potentials are measured at $1M$ concentration. Maintaining a pressure of 1 atm is especially important in half-cells that involve gases as reactants or products. Suggest a reason that gas pressure is especially critical in these cells.
- 61. Analyze and Conclude** Zinc-carbon, alkaline, mercury, lithium, and NiCad batteries all contain a separator between the anode and cathode that allows exchange of ions but keeps the anode and cathode reactants from mixing. No such separator is present (or needed) in a lead-acid battery.



Study the structures of this battery carefully and suggest a reason why no separator is required.

- 62. Interpreting Graphs** Various commercial batteries were tested to determine how each would perform when operating a motorized toy. Use the graph below to determine which battery would be the best choice to operate the toy for a long period of time. Which battery goes dead abruptly?



- 63. Hypothesizing** The density of a sample of electrolyte from a lead-acid battery is measured and found to be 1.09 g/mL . Do you think this density indicates a fully charged battery, partially discharged battery, or a “dead” battery? Give a reason for your choice.

- 64. Drawing a Conclusion** The black tarnish that appears on silver is silver sulfide (Ag_2S). A way to get rid of silver tarnish is to place the silver object into a pan lined with wrinkled aluminum foil and water in which some salt or baking soda has been dissolved. As the silver sits in contact with the aluminum foil, the black tarnish disappears where the object is immersed in the solution. Using what you have learned about electrochemistry, propose an explanation for this observation.
- 65. Thinking Critically** Lead-acid storage batteries produce current with a voltage of about 12 volts. If you have a voltmeter in your car, you see that the meter reads about 14 V while the car is running normally. Suggest an explanation for this higher voltage reading.
- 66. Applying Concepts** During electrolysis, an electrolytic cell gives off bromine vapor and hydrogen gas. After electrolysis, the cell is found to contain a concentrated solution of potassium hydroxide. What was the composition of the cell before electrolysis began?
- 67. Hypothesizing** Instead of galvanizing iron, it can be plated with copper to protect it. What would happen when the copper coating became broken or cracked? Would the copper continue to protect the iron as zinc does? Explain fully.

Writing in Chemistry

- 68.** Research the commercial process of electroplating silver in the costume jewelry industry. Write a summary of your findings, including a diagram.
- 69.** Several years ago, the supporting structure of the Statue of Liberty became so corroded that it had to be replaced entirely. Find out what the structure was and why it corroded so badly. Write a report that discusses the chemical processes involved and include a timeline of the statue, starting before 1886 in France.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 70.** If the volume of a sample of chlorine gas is 8.2 L at 1.2 atm and 306 K, what volume will the gas occupy at STP? (Chapter 14)
- 71.** According to the collision model of chemical reactions, how is it possible that two molecules may collide but not react? (Chapter 17)

Use these questions and the test-taking tip to prepare for your standardized test.

- A salt bridge is essential to a voltaic cell for all of the following reasons EXCEPT _____.
 - it allows ions to move from the solution of one cell to the other
 - it causes electric current to flow between the two electrodes of a cell
 - it relieves the buildup of positive charge on the anode side of the cell
 - it allows electrons to move from the solution of one cell to the other

Using Tables Use the table to answer questions 2–5.

Selected Standard Reduction Potentials at 25°C, 1 atm, and 1M Ion Concentration	
Half-Reaction	E^0 (V)
$\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}$	-2.372
$\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$	-1.662
$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}$	-0.1262
$\text{Ag}^+ + e^- \rightarrow \text{Ag}$	0.7996
$\text{Hg}^{2+} + 2e^- \rightarrow \text{Hg}$	0.851

- Which of the metal ions is most easily reduced?
 - Mg^{2+}
 - Hg^{2+}
 - Ag^+
 - Al^{3+}
- On the basis of the standard reduction potentials shown above, which of the following standard cell notations below correctly represents its voltaic cell?
 - $\text{Ag} \mid \text{Ag}^+ \parallel \text{Al}^{3+} \mid \text{Al}$
 - $\text{Mg} \mid \text{Mg}^{2+} \parallel \text{H}^+ \mid \text{H}_2$
 - $\text{H}_2 \mid \text{H}^+ \parallel \text{Pb}^{2+} \mid \text{Pb}$
 - $\text{Pb} \mid \text{Pb}^{2+} \parallel \text{Al}^{3+} \mid \text{Al}$
- A voltaic cell consists of a magnesium bar dipping into a 1M Mg^{2+} solution and a silver bar dipping into a 1M Ag^+ solution. What is the standard potential of this cell?
 - 1.572 V
 - 3.172 V
 - 0.773 V
 - 3.971 V

- Assuming standard conditions, which of the following cells will produce a potential of 2.513 V?
 - $\text{Al} \mid \text{Al}^{3+} \parallel \text{Hg}^{2+} \mid \text{Hg}$
 - $\text{H}_2 \mid \text{H}^+ \parallel \text{Hg}^{2+} \mid \text{Hg}$
 - $\text{Mg} \mid \text{Mg}^{2+} \parallel \text{Al}^{3+} \mid \text{Al}$
 - $\text{Pb} \mid \text{Pb}^{2+} \parallel \text{Ag}^+ \mid \text{Ag}$
- Which of the following statements is NOT true of batteries?
 - Batteries are compact forms of voltaic cells.
 - Secondary batteries also are known as storage batteries.
 - Batteries can consist only of a single cell.
 - The redox reaction in a rechargeable battery is reversible.
- The corrosion, or rusting, of iron is an example of a naturally occurring voltaic cell. To prevent corrosion, sacrificial anodes are sometimes attached to rust-susceptible iron. Sacrificial anodes must _____.
 - be more likely to be reduced than iron
 - have a higher reduction potential than iron
 - be more porous and abraded than iron
 - lose electrons more easily than iron
- A strip of metal X is immersed in a 1M solution of X^+ ions. When this half-cell is connected to a standard hydrogen electrode, a voltmeter reads a positive reduction potential. Which of the following is true of the X electrode?
 - It accepts electrons more readily than H^+ ions.
 - It is undergoing oxidation.
 - It is adding positive X^+ ions to its solution.
 - It acts as the anode in the cell.
- To electroplate an iron fork with silver, which of the following is true?
 - The silver electrode must have more mass than the fork.
 - The iron fork must act as the anode in the cell.
 - Electrical current must be applied to the iron fork.
 - Iron ions must be present in the cell solution.



TEST-TAKING TIP

Your Mistakes Can Teach You The mistakes you make before the test are helpful because they show you the areas in which you need more work. When calculating the standard potential of an electrochemical cell, remember that coefficients do not change the standard reduction potentials of the half-reactions.

Hydrocarbons

What You'll Learn

- ▶ You will compare the structures and properties of alkanes, alkenes, and alkynes.
- ▶ You will recognize and compare the properties of structural isomers and stereoisomers.
- ▶ You will describe how useful hydrocarbons are obtained from natural sources.

Why It's Important

Fuels, medicines, synthetic textiles, plastics, and dyes are just a few examples of hydrocarbon-derived organic chemicals we use every day.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about hydrocarbons and their structures.

Oil pipelines transport petroleum that remains a vital resource for the manufacture of fuels, plastics, solvents, pharmaceuticals, and other important carbon compounds.



DISCOVERY LAB



Materials

wide-mouth jars and lids (3)
lead weights, BB size (9)
screws, nuts, and washers
long forceps
10W-30 and 20W-50 motor
oil
stopwatch
paper towels

Viscosity of Motor Oil

The molecules of motor oil have long chains of carbon atoms. Oil's viscosity, a measure of resistance to flow, is related to the oil's weight numbers. How do two weights of oil differ in viscosity?

Safety Precautions



Procedure

1. Add a 50-mm depth of water to the first jar, the same depth of 10W-30 motor oil to the second jar, and an equal depth of 20W-50 oil to the third jar.
2. Drop a lead weight from just above the surface of the liquid in the first jar. Time the weight as it sinks to the bottom. Repeat the process twice with two other small metal objects.
3. Repeat step 2 with the jars of oil.
4. Use forceps to remove the weights. Dry them on paper towels. Save the oil for reuse.

Analysis

What do your results tell you about the relationship between viscosity and the weight numbers of motor oil? Which oil is more likely to be used in heavy machinery that requires a high-viscosity oil?

Section

22.1

Alkanes

Objectives

- **Describe** the structures of alkanes.
- **Name** an alkane by examining its structure.
- **Draw** the structure of an alkane given its name.

Vocabulary

organic compound
hydrocarbon
alkane
homologous series
parent chain
substituent group

The Alaskan pipeline shown in the photo on the opposite page was built to carry crude oil from the oil fields in the frozen north to an ice-free southern Alaskan seaport. Crude oil, also called petroleum, is a complex mixture of carbon compounds produced by heat and pressure acting on the remains of once-living organisms buried deep beneath Earth's surface.

Organic Chemistry

Chemists of the early nineteenth century knew that living things produce an immense variety of carbon compounds. Chemists referred to these compounds as "organic" compounds because they were produced by living organisms.

Once Dalton's atomic theory was accepted in the early nineteenth century, chemists began to understand that compounds, including those made by living organisms, consisted of arrangements of atoms bonded together in certain combinations. With this knowledge, they were able to go to their laboratories and synthesize many new and useful substances—but not the organic compounds made by living things. One of the main reasons for their failure was that chemists, like most other scientists of the time, accepted an idea called *vitalism*. According to vitalism, organisms possess a mysterious

4A	
Carbon 6 C 12.011	
Silicon 14 Si 28.086	
Germanium 32 Ge 72.61	
Tin 50 Sn 118.710	
Lead 82 Pb 207.2	

Figure 22-1

Carbon is a nonmetal element located in group 4A of the periodic table. Elements in this group have four valence electrons. Carbon usually shares its four valence electrons to form four covalent bonds.

“vital force” that enables them to assemble carbon compounds, and such compounds could never be produced in the laboratory.

Disproving vitalism Friedrich Wöhler (1800–1882) was a German chemist who questioned the idea of vitalism. While working in Berlin in 1828, he carried out a reaction that he thought would produce the compound ammonium cyanate (NH_4OCN). To his surprise, the product of the reaction turned out to be urea (NH_2CONH_2), a compound that had the same empirical formula as ammonium cyanate. Previously, urea was known only as a waste product found in the urine of humans and other animals. Wöhler wrote to the Swedish chemist Berzelius that he had “made urea without the kidney of an animal, either man or dog.”

Although Wöhler’s experiment did not immediately disprove vitalism, it started a chain of similar experiments by other European chemists. Eventually, the idea that the synthesis of organic compounds required a vital force was discredited. Today the term **organic compound** is applied to all carbon-containing compounds with the primary exceptions of carbon oxides, carbides, and carbonates, which are considered inorganic. An entire branch of chemistry, called organic chemistry, is devoted to the study of carbon compounds.

Why is an entire field of study focused on compounds containing carbon? To answer this question, recall that carbon is in the group 4A elements of the periodic table, as shown in **Figure 22-1**. With its electron configuration of $1s^22s^22p^2$, carbon nearly always shares its electrons, forming four covalent bonds. In organic compounds, carbon atoms are found bonded to hydrogen atoms or atoms of other elements that are near carbon in the periodic table—especially nitrogen, oxygen, sulfur, phosphorus, and the halogens.

Most importantly, carbon atoms also bond to other carbon atoms and can form chains from two to thousands of carbon atoms in length. Also, because carbon forms four bonds, it can form complex, branched-chain structures, ring structures, and even cage-like structures. With all of these bonding possibilities, millions of different organic compounds are known, and chemists are synthesizing more every day.

Hydrocarbons

The simplest organic compounds are the **hydrocarbons**, which contain only the elements carbon and hydrogen. How many different compounds do you think it is possible to make from two elements? You might guess one, or maybe a few more, say 10 or 12. Actually, thousands of hydrocarbons are known, each containing only the elements carbon and hydrogen.

As you know, carbon forms four bonds. Hydrogen, having only one valence electron, forms only one covalent bond by sharing this electron with another atom. Therefore, the simplest hydrocarbon molecule consists of a carbon atom bonded to four hydrogen atoms, CH_4 . This substance is called methane and is an excellent fuel and the main component of natural gas.

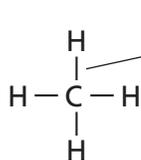
Figure 22-2

Methane (CH_4) consists of one carbon atom bonded to four hydrogen atoms in a tetrahedral arrangement. Here are four ways to represent a methane molecule. Refer to **Table C-1** in Appendix C for a key to atom color convention.

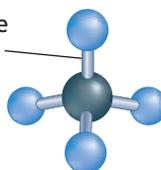
Models of Methane



Molecular formula



Structural formula



Ball-and-stick model



Space-filling model

In chemistry, covalent bonds in which two electrons are shared are represented by a single straight line, denoting a single covalent bond. **Figure 22-2** shows four different ways to represent a methane molecule.

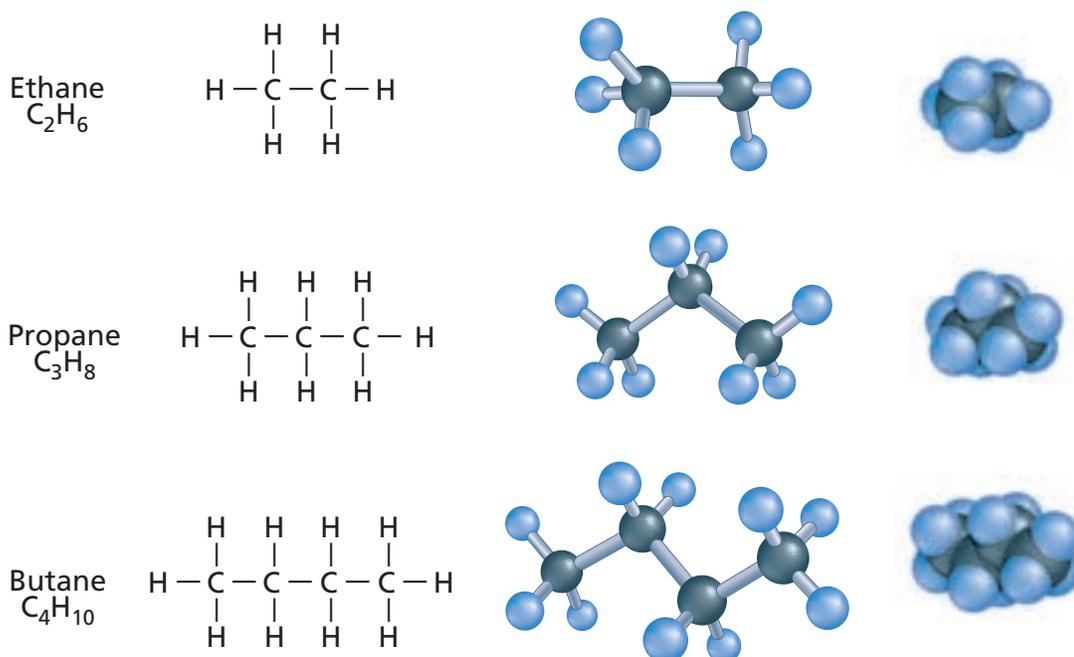
A review of models Chemists represent organic molecules in a variety of ways. Most often, they use the type of model that best shows the information they want to highlight. As you see in **Figure 22-2** on the previous page, molecular formulas give no information about the geometry of the molecule. A structural formula shows the general arrangement of atoms in the molecule but not the exact geometry. Space-filling models give a more realistic picture of what a molecule would look like if you could see it, but ball-and-stick models demonstrate the geometry of the molecule more clearly. Keep in mind, however, that the atoms are not at the ends of sticks but are held closely together by electron-sharing bonds.

Straight-Chain Alkanes

Methane is the smallest member of a series of hydrocarbons known as alkanes. **Alkanes** are hydrocarbons that have only single bonds between atoms. The next member of the series consists of two carbon atoms bonded together with a single bond and six hydrogen atoms sharing the remaining valence electrons of the carbon atoms. This molecule is called ethane (C_2H_6) and is shown in **Figure 22-3**.

The third member of the alkane series, propane, has three carbon atoms and eight hydrogen atoms, giving it the molecular formula C_3H_8 . The next member, butane, has four carbon atoms in a continuous chain and the formula C_4H_{10} . Compare the structures of ethane, propane, and butane in **Figure 22-3**.

Most propane and butane come from petroleum. Propane, known also as LP (liquified propane) gas, is sold as a fuel for cooking and heating. Another use of propane is illustrated in **Figure 22-4** on page 700. Butane is used as fuel in small lighters and in some torches. It is also used in the manufacture of synthetic rubber and gasoline. Carry out the **CHEMLAB** at the end of the chapter to determine which alkanes are found in your lab's gas supply.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Figure 22-3

Ethane makes up about nine percent of natural gas. Because the bonds around carbon are arranged in a tetrahedral fashion, propane and butane have a zig-zag geometry.



Figure 22-4

This taxi in Bangkok, Thailand is one of many cars, trucks, and buses that have been modified to burn propane. Burning propane benefits the environment because it produces less air pollution than does gasoline.

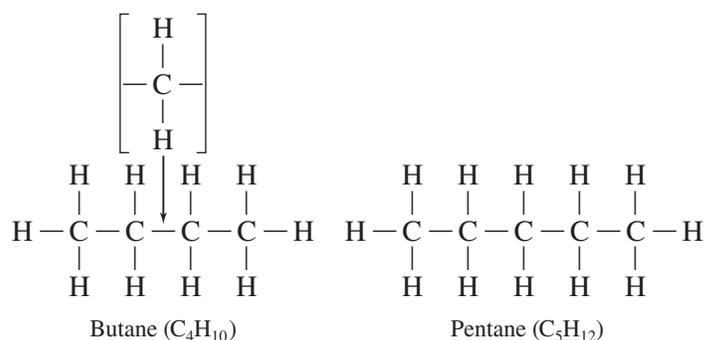
By now you have noticed that names of alkanes end in *-ane*. Also, alkanes with five or more carbons in a chain have names that use a prefix derived from the Greek or Latin word for the number of carbons in each chain. For example, *pentane* has five carbons just as a *pentagon* has five sides, and *octane* has eight carbons just as an *octopus* has eight tentacles. Because methane, ethane, propane, and butane were named before alkane structures were known, their names do not have numerical prefixes. **Table 22-1** shows the names and structures of the first ten alkanes. Notice the underlined prefix representing the number of carbon atoms in the molecule.

Table 22-1

First Ten of the Alkane Series				
Name	Molecular formula	Condensed structural formula	Melting point (°C)	Boiling point (°C)
Methane	CH ₄	CH ₄	-182	-162
Ethane	C ₂ H ₆	CH ₃ CH ₃	-183	-89
Propane	C ₃ H ₈	CH ₃ CH ₂ CH ₃	-188	-42
Butane	C ₄ H ₁₀	CH ₃ CH ₂ CH ₂ CH ₃	-138	-0.5
<u>Pentane</u>	C ₅ H ₁₂	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃	-130	36
<u>Hexane</u>	C ₆ H ₁₄	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-95	69
<u>Heptane</u>	C ₇ H ₁₆	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-91	98
<u>Octane</u>	C ₈ H ₁₈	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-57	126
<u>Nonane</u>	C ₉ H ₂₀	CH ₃ (CH ₂) ₇ CH ₃	-54	151
<u>Decane</u>	C ₁₀ H ₂₂	CH ₃ (CH ₂) ₈ CH ₃	-29	174

In **Table 22-1**, you can see that the structural formulas are written in a different way from those in **Figure 22-3**. These formulas, called condensed structural formulas, save space by not showing how the hydrogen atoms branch off the carbon atoms. Condensed formulas can be written in several ways. In **Table 22-1**, the lines between carbon atoms have been eliminated to save space.

Looking at the alkane series in **Table 22-1**, you can see that —CH₂— is a repeating unit in the chain of carbon atoms. Note, for example, in the diagram below that pentane has one more —CH₂— unit than butane.



You can further condense structural formulas by writing the —CH₂— unit in parentheses followed by a subscript to show the number of units, as is done with nonane and decane.

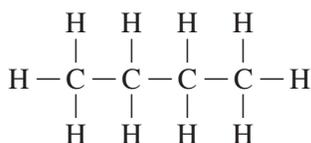
CONNECTION

By the middle of the 19th century, inorganic chemistry and physics were regarded as rigorous experimental sciences. However, the biological sciences were held back in part by the prevailing belief in vitalism—the idea that the processes and materials of living things could not be explained by the same laws and theories that applied to physics and chemistry. Vitalism was dealt its final blow in 1897, when Eduard Buchner, a German chemist, showed that extracts of yeast could carry out fermentation of sugar when no living cells were present.

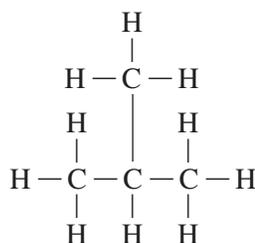
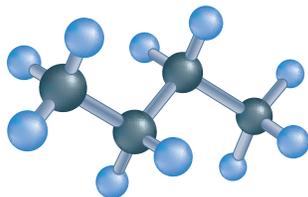
A series of compounds that differ from one another by a repeating unit is called a **homologous series**. A homologous series has a fixed numerical relationship among the numbers of atoms. For alkanes, the relationship between the numbers of carbon and hydrogen atoms can be expressed as C_nH_{2n+2} , where n is equal to the number of carbon atoms in the alkane. Given the number of carbon atoms in an alkane, you can write the molecular formula for any alkane. For example, heptane has seven carbons so its formula is $C_7H_{2(7)+2}$, or C_7H_{16} . What is the molecular formula for a 13-carbon alkane?

Branched-Chain Alkanes

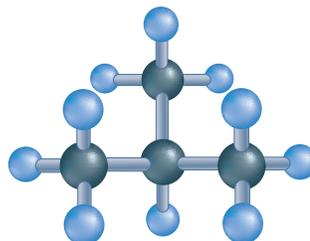
The alkanes you have studied so far are called straight-chain alkanes because the carbon atoms are bonded to each other in a single line. Now look at the two structures in the following diagram. If you count the carbon and hydrogen atoms, you will discover that both structures have the same molecular formula, C_4H_{10} . Do these structures represent the same substance?



Butane
Molecular formula: C_4H_{10}



Isobutane
Molecular formula: C_4H_{10}



If you think that the structures represent two different substances, you are right. The structure on the left represents butane, and the structure on the right represents a branched-chain alkane, known as isobutane, a substance whose chemical and physical properties are different from those of butane. As you see, carbon atoms can bond to one, two, three, or even four other carbon atoms. This property makes possible a variety of branched-chain alkanes. How do you name isobutane using IUPAC rules?

Naming Branched-Chain Alkanes

You've seen that both a straight-chain and a branched-chain alkane can have the same molecular formula. This fact illustrates a basic principle of organic chemistry: the order and arrangement of atoms in an organic molecule determine its identity. Therefore, the name of an organic compound also must describe the molecular structure of the compound accurately.

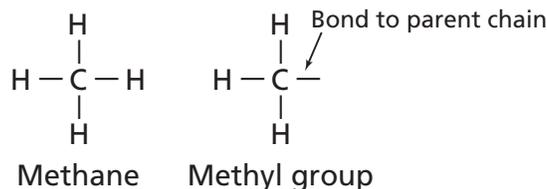
For the purpose of naming, branched-chain alkanes are viewed as consisting of a straight chain of carbon atoms with other carbon atoms or groups of carbon atoms branching off the straight chain. The longest continuous chain of carbon atoms is called the **parent chain**. All side branches are called **substituent groups** because they appear to substitute for a hydrogen atom in the straight chain.

Alternative Fuel Technician

Are you interested in improving the quality of the air you breathe and reducing our dependence on petroleum resources? Do you like to work with cars? Then consider a career as an alternative fuels technician.

Alternative fuels are being developed and used today to replace fossil fuels such as oil. As an alternative fuels technician, you may work in research developing new vehicles. Or, you may educate consumers and service and maintain their environmentally friendly vehicles.

Each alkane-based substituent group branching from the parent chain is named for the straight-chain alkane having the same number of carbon atoms as the substituent. The ending *-ane* is replaced with the letters *-yl*, as shown in the following diagram.



An alkane-based substituent group is called an alkyl group. Several common alkyl groups are shown in **Table 22-2**. What relationship exists between these groups and the alkanes in **Table 22-1**?

The naming process To name organic structures, chemists follow systematic rules approved by the International Union of Pure and Applied Chemistry (IUPAC). Here are the rules for naming branched-chain alkanes.

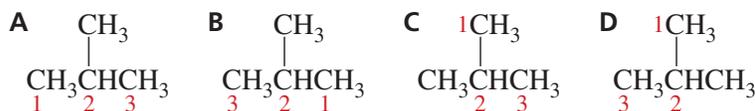
1. *Count the number of carbon atoms in the longest continuous chain.* Use the name of the straight-chain alkane with that number of carbons as the name of the parent chain of the structure.
2. *Number each carbon in the parent chain.* Locate the end carbon closest to a substituent group. Label that carbon position one. This step gives all the substituent groups the lowest position numbers possible.
3. *Name each alkyl group substituent.* The names of these groups are placed before the name of the parent chain.

Table 22-2

Common Alkyl Groups		
Name	Condensed structural formula	Structural formula
Methyl	CH ₃ –	$ \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}- \\ \\ \text{H} \end{array} $
Ethyl	CH ₃ CH ₂ –	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}- \\ \quad \\ \text{H} \quad \text{H} \end{array} $
Propyl	CH ₃ CH ₂ CH ₂ –	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}- \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array} $
Isopropyl	CH ₃ CHCH ₃	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $
Butyl	CH ₃ CH ₂ CH ₂ CH ₂ –	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}- \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $

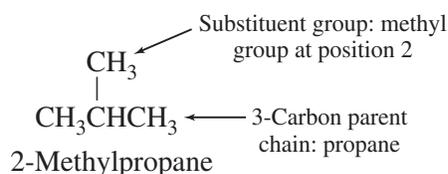
- If the same alkyl group occurs more than once as a branch on the parent structure, use a prefix (*di-*, *tri-*, *tetra-*, etc.) before its name to indicate how many times it appears. Then use the number of the carbon to which each is attached to indicate its position.
- Whenever different alkyl groups are attached to the same parent structure, place their names in alphabetical order. Do not consider the prefixes (*di-*, *tri-*, etc.) when determining alphabetical order.
- Write the entire name using hyphens to separate numbers from words and commas to separate numbers. No space is added between the substituent name and the name of the parent chain.

Now, try to name the branched-chain structure, isobutane.



- The longest chain in the structures above contains three carbons. This is true if you start on the left (A), right (B), or carbon in the branch (C, D), as you can see in the accompanying diagram. Therefore, the name of the parent chain will be *propane*.
- No matter where the numbering starts in this molecule, the alkyl group is at position 2. So, the four options are equivalent.
- The alkyl group here is a methyl group because it has only one carbon.
- No prefix is needed because only one alkyl group is present.
- Alphabetical order does not need to be considered because only one group is present.

After applying the rules, you can write the IUPAC name *2-methylpropane* for isobutane. See **Figure 22-5**. Note that the name of the alkyl group is added in front of the name of the parent chain with no space between them. A hyphen separates the number from the word.



Because structural formulas can be written with chains oriented in various ways, you need to be careful in finding the longest continuous carbon chain. The following examples are written as skeletal formulas. A skeletal formula shows only the carbon atoms of an organic molecule in order to emphasize the chain arrangement. Hydrogen atoms are not shown in these skeletal formulas. Study the correct numbering of carbon atoms in the following examples. Note that numbering either carbon chain starting on the left-most carbon would disobey rule 1 on the previous page.

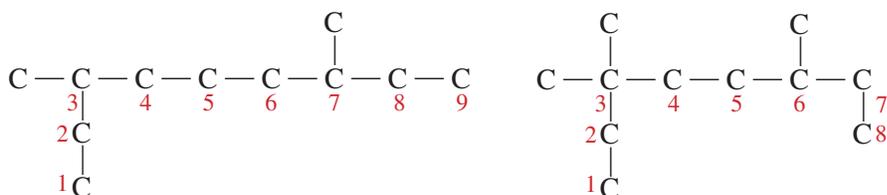
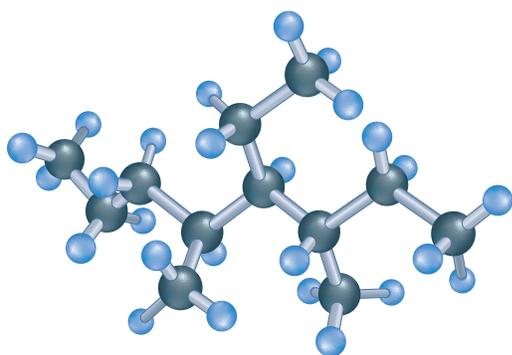


Figure 22-5

Many of today's automobile and truck air conditioners can use hydrocarbon refrigerant mixtures containing 2-methylpropane that are environmentally safe.

Example Problem 22-1 and the Practice Problems that follow it will help you develop skill at naming branched-chain alkanes.

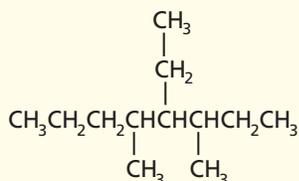


This ball-and-stick model of the molecule in Example Problem 22-1 shows how the molecule looks in three dimensions.

EXAMPLE PROBLEM 22-1

Naming Branched-Chain Alkanes

Name the following alkane.

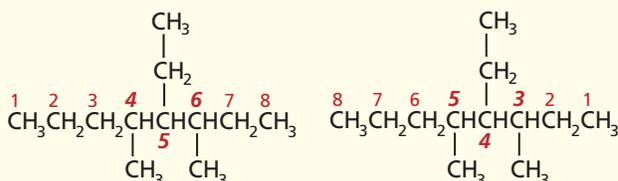


1. Analyze the Problem

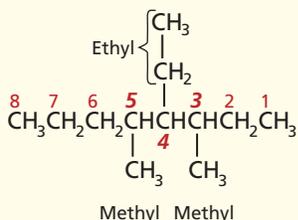
You are given a structure. To determine the name of the parent chain and the names and locations of branches, follow the IUPAC rules.

2. Solve for the Unknown

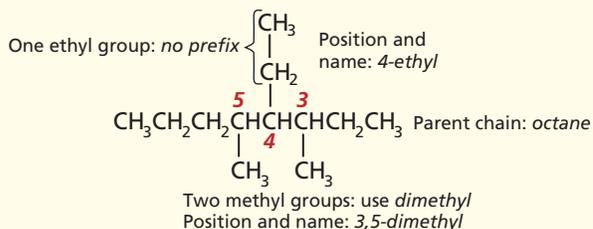
- Count carefully to find the longest chain. In this case, it is easy. The longest chain has eight carbon atoms, so the parent name is *octane*.
- Number the chain in both directions. Numbering from the left puts the alkyl groups at positions 4, 5, and 6. Numbering from the right puts alkyl groups at positions 3, 4, and 5. Therefore, 3, 4, and 5 are the lowest position numbers and will be used in the name.



- Identify and name the alkyl groups branching from the parent chain. There are one-carbon *methyl* groups at positions 3 and 5, and a two-carbon *ethyl* group at position 4.



- Look for and count the alkyl groups that occur more than once. Determine the prefix to use to show the number of times each group appears. In this example, the prefix *di-* will be added to the name *methyl* because two methyl groups are present. No prefix is needed for the one ethyl group. Then show the position of each group with the appropriate number.



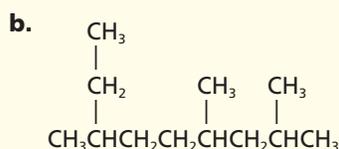
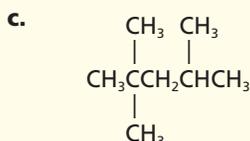
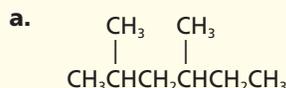
- f. Place the names of the alkyl branches in alphabetical order, *ignoring the prefixes*. Alphabetical order puts the name ethyl before dimethyl.
- g. Write the name of the structure using hyphens and commas as needed. The name should be written as *4-ethyl-3,5-dimethyloctane*.

3. Evaluate the Answer

The longest continuous carbon chain has been found and numbered correctly. All branches have been designated with correct prefixes and alkyl-group names. Alphabetical order and punctuation are correct.

PRACTICE PROBLEMS

1. Use IUPAC rules to name the following structures.



2. Draw the structures of the following branched-chain alkanes.

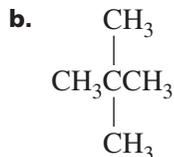
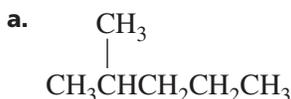
- a. 2,3-dimethyl-5-propyldecane
b. 3,4,5-triethyloctane



For more practice with naming branched-chain alkanes go to **Supplemental Practice Problems** in Appendix A.

Section 22.1 Assessment

3. Use IUPAC rules to name the following structures.



4. Write a condensed structural formula for each of the following.
- a. 3,4-diethylheptane
b. 4-isopropyl-3-methyldecane
5. Name two types of carbon-containing compounds that are considered inorganic rather than organic.
6. Write correct molecular formulas for pentadecane, a 15-carbon alkane, and icosane, a 20-carbon alkane.
7. Why is the name *3-butylpentane* incorrect? Based on this name, write the structural formula for the

compound. What is the correct IUPAC name for 3-butylpentane?

8. **Thinking Critically** Hexane is called a straight-chain alkane. Yet, a molecule of hexane has a zig-zag rather than a linear geometry. Explain this apparent paradox. Explain what characteristic of carbon atoms causes the zig-zag geometry of straight-chain alkanes.
9. **Graphing** Use data from **Table 22-1** to graph boiling point versus the number of carbon atoms in the alkane chain for the first ten alkanes. Use the graph to predict boiling points for straight-chain alkanes with 11 and 12 carbon atoms. For more help, refer to *Drawing and Using Line Graphs* in the **Math Handbook** on pages 903–907 of this textbook.



Cyclic Alkanes and Alkane Properties

Objectives

- **Name** a cyclic alkane by examining its structure.
- **Draw** the structure of a cyclic alkane given its name.
- **Describe** the properties of alkanes.
- **Distinguish** between saturated and unsaturated hydrocarbons.

Vocabulary

cyclic hydrocarbon
cycloalkane
saturated hydrocarbon
unsaturated hydrocarbon

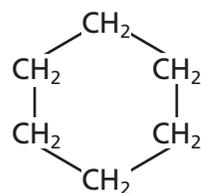
One of the reasons that such a variety of organic compounds exists is that carbon atoms can form ring structures. An organic compound that contains a hydrocarbon ring is called a **cyclic hydrocarbon**.

Cycloalkanes

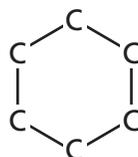
To indicate that a hydrocarbon has a ring structure, the prefix *cyclo-* is used with the hydrocarbon name. Thus cyclic hydrocarbons that contain single bonds only are called **cycloalkanes**. Cycloalkanes can have rings with three, four, five, six, or even more carbon atoms. Cyclopropane, the smallest cycloalkane, is a gas that was used for many years as an anesthetic for surgery. However, it is no longer used because it is highly flammable. The name for the six-carbon cycloalkane is cyclohexane. Obtained from petroleum, cyclohexane is used in paint and varnish removers and for extracting essential oils to make perfume. Note that cyclohexane (C_6H_{12}) has two fewer hydrogen atoms than straight-chain hexane (C_6H_{14}) because a valence electron from each of two carbon atoms is now forming a carbon-carbon bond rather than a carbon-hydrogen bond.

As you can see in **Figure 22-6**, cyclic hydrocarbons such as cyclohexane are represented by condensed, skeletal, and line structures. Line structures show only the carbon-carbon bonds with carbon atoms understood to be at each vertex of the structure. Hydrogen atoms are assumed to occupy the remaining bonding positions unless substituents are present.

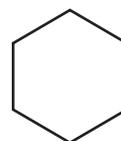
Naming substituted cycloalkanes Like other alkanes, cycloalkanes can have substituent groups. Substituted cycloalkanes are named by following the same IUPAC rules used for straight-chain alkanes, but with a few modifications. With cycloalkanes, there is no need to find the longest chain because the ring is always considered to be the parent chain. Because a cyclic structure has no ends, numbering is started on the carbon that is bonded to the substituent group. When there are two or more substituents, the carbons are numbered around the ring in a way that gives the lowest possible set of numbers for the substituents. If only one group is attached to the ring, no number is necessary. The following Example Problem will show you how a cycloalkane is named.



Condensed structural formula



Skeletal structure



Line structure

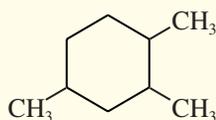
Figure 22-6

Cyclohexane can be represented in several ways. Chemists most often draw line structures for cyclic hydrocarbons because the molecules have distinct shapes that are easily identifiable.

EXAMPLE PROBLEM 22-2

Naming Cycloalkanes

Name the cycloalkane shown.



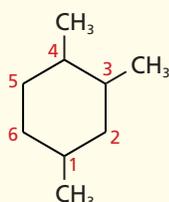
1. Analyze the Problem

You are given a structure. To determine the parent cyclic structure and the location of branches, follow the IUPAC rules.

2. Solve for the Unknown

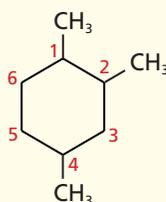
- Count the carbons in the ring and use the name of the parent cyclic hydrocarbon. In this case, the ring has six carbons, so the parent name is *cyclohexane*.
- Number the ring, starting from one of the CH_3 - branches. Find the numbering that gives the lowest possible set of numbers for the branches. Here are two ways of numbering the ring.

A



1,3,4-Trimethylcyclohexane

B



1,2,4-Trimethylcyclohexane

Numbering from the carbon atom at the bottom of the ring puts the CH_3 - groups at positions 1, 3, and 4 in structure A. Numbering from the carbon at the top of the ring gives the position numbers 1, 2, and 4. All other numbering schemes place the CH_3 - groups at higher position numbers. Thus, 1, 2, and 4 are the lowest possible position numbers and will be used in the name.

- Name the substituents. All three are the same—one-carbon methyl groups.
- Add the prefix to show the number of groups present. Three methyl groups are present, so the prefix *tri-* will be added to the name *methyl* to make *trimethyl*.
- Alphabetical order can be ignored because only one type of group is present.
- Put the name together using the name of the parent cycloalkane. Use commas to separate numbers and hyphens between numbers and words. Write the name as *1,2,4-trimethylcyclohexane*.

3. Evaluate the Answer

The parent ring structure is numbered to give the branches the lowest possible set of numbers. The prefix *tri-* indicates that three methyl groups are present. No alphabetization is necessary because all branches are methyl groups.



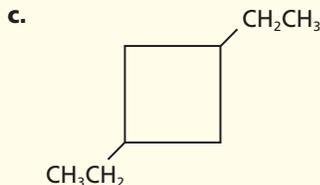
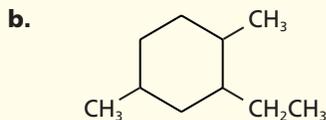
Stains and varnishes used to refinish wood often use cyclohexane as a solvent.

Practice!

For more practice with naming cycloalkanes go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

10. Use IUPAC rules to name the following structures.



11. Draw the structures of the following cycloalkanes.

- 1-ethyl-3-propylcyclopentane
- 1,2,2,4-tetramethylcyclohexane

Properties of Alkanes

You have learned that the structure of a molecule affects its properties. For example, ammonia (NH_3) can accept a proton from an acid to become an ammonium ion (NH_4^+) because the nitrogen atom has an unshared pair of electrons. As another example, the O—H bonds in a water molecule are polar, and because the H—O—H molecule has a bent geometry, the molecule itself is polar. Thus water molecules are attracted to each other and can form hydrogen bonds with each other. As a result, the boiling and melting points of water are much higher than those of other substances having similar molecular mass and size.

What properties would you predict for alkanes? All of the bonds in these hydrocarbons are between either a carbon atom and a hydrogen atom or two carbon atoms. Are these bonds polar? Remember, a bond between two atoms is polar only if the atoms differ by at least 0.5 in their Pauling electronegativity values. Carbon's electronegativity value is 2.55, and hydrogen's is 2.20, so a C—H bond has a difference of only 0.35. Thus, it is not polar. A bond between two identical atoms such as carbon can never be polar because the difference in their electronegativity values is zero. Because all of the bonds in alkanes are nonpolar, alkane molecules are nonpolar.

Physical properties of alkanes What types of physical properties do nonpolar compounds have? A comparison of two molecular substances—one

Table 22-3

Comparing Physical Properties of Water and Methane		
Substance and formula	Water (H_2O)	Methane (CH_4)
Model		
Molecular mass	18 amu	16 amu
State at room temperature	liquid	gas
Boiling point	100°C	-162°C
Melting point	0°C	-182°C



polar and the other nonpolar—will help you answer this question. **Table 22-3** compares the properties of water and methane.

Note that the molecular mass of methane (16 amu) is close to the molecular mass of water (18 amu). Also, water and methane molecules are similar in size. However, when you compare the melting and boiling points of methane to those of water, you can see evidence that the molecules differ in some significant way. These temperatures differ greatly because methane molecules have little intermolecular attraction compared to water molecules. This difference in attraction can be explained by the fact that methane molecules are nonpolar and do not form hydrogen bonds with each other, whereas water molecules are polar and freely form hydrogen bonds. What straight-chain alkane in **Table 22-1** has a boiling point nearest that of water?

The difference in polarity and hydrogen bonding also explains the immiscibility of alkanes and other hydrocarbons with water. If you try to dissolve alkanes, such as lubricating oils, in water, the two liquids separate almost immediately into two phases. This separation happens because the attractive forces between alkane molecules are stronger than the attractive forces between the alkane and water molecules. Therefore, alkanes are more soluble in solvents composed of nonpolar molecules like themselves than in water, a polar solvent. This is another example of the rule of thumb that “like dissolves like.”

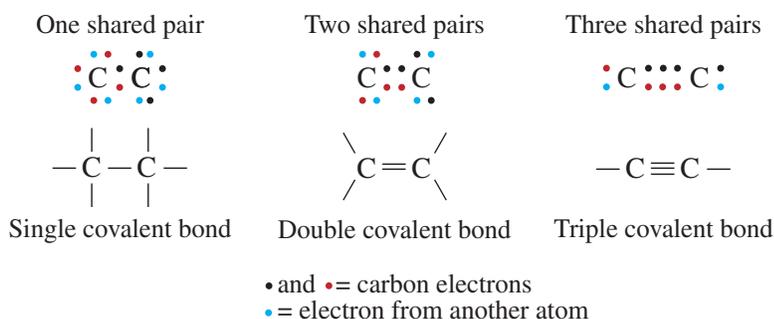
Chemical properties of alkanes The main chemical property of alkanes is their low reactivity. Recall that many chemical reactions occur when a reactant with a full electrical charge, such as an ion, or with a partial charge, such as a polar molecule, is attracted to another reactant with the opposite charge. Molecules such as alkanes, in which atoms are connected by nonpolar bonds, have no charge. As a result, they have little attraction for ions or polar molecules. The low reactivity of alkanes also can be attributed to the relatively strong C—C and C—H bonds. The low reactivity of alkanes limits their uses. As you can see from **Figure 22-7**, alkanes are commonly used as fuels because they readily undergo combustion in oxygen.

Figure 22-7

Alkane hydrocarbons are often used as fuels to provide heat and, sometimes, light. The smudge pots shown on the left produce heat that helps prevent cold-weather damage to citrus crops. The gas lantern gives off light, and the gas-log fireplace provides both heat and light.

Multiple Carbon-Carbon Bonds

Carbon atoms can bond to each other not only by single covalent bonds but also by double and triple covalent bonds. In a double bond, atoms share two pairs of electrons; in a triple bond, they share three pairs of electrons. The following diagram shows Lewis structures and structural formulas for single, double, and triple covalent bonds.



In the nineteenth century, before they understood bonding and the structure of organic substances, chemists experimented with hydrocarbons obtained from heating animal fats and plant oils. They classified these hydrocarbons according to a chemical test in which they mixed each hydrocarbon with bromine and then measured how much reacted with the hydrocarbon. Some hydrocarbons would react with a little bromine, some would react with more, and some would not react at all. Chemists called the hydrocarbons that reacted with bromine unsaturated hydrocarbons in the same sense that an unsaturated aqueous solution can dissolve more solute. Hydrocarbons that did not react with bromine were said to be saturated.

Modern chemists can explain the results of the chemists of 170 years ago. Hydrocarbons that reacted with bromine had double or triple covalent bonds. Those that took up no bromine had only single covalent bonds. Today, a **saturated hydrocarbon** is defined as a hydrocarbon having only single bonds—in other words, an alkane. An **unsaturated hydrocarbon** is a hydrocarbon that has at least one double or triple bond between carbon atoms. You will learn more about unsaturated hydrocarbons in Section 22.3.

Section 22.2 Assessment

12. Use IUPAC rules to name the following structures.
- a.
- b.
13. Write a condensed structural formula for each of the following.
- a. 1-ethyl-4-methylcyclohexane
- b. 1,2-dimethylcyclopropane
14. Describe the main structural characteristics of alkane molecules. Give two examples of how these characteristics determine the physical properties of alkanes.
15. What structural characteristic distinguishes saturated from unsaturated hydrocarbons?
16. **Thinking Critically** Some shortening is described as “hydrogenated vegetable oil.” This means that the oils reacted with hydrogen in the presence of a catalyst. Make a hypothesis to explain why hydrogen reacted with the oils.
17. **Modeling** Construct ball-and-stick molecular models of the following cyclic alkanes.
- a. isopropylcyclobutane
- b. 1,2,4-trimethylcyclopentane

You now know that alkanes are saturated hydrocarbons because they contain only single covalent bonds between carbon atoms, and that unsaturated hydrocarbons have at least one double or triple bond between carbon atoms.

Alkenes

Unsaturated hydrocarbons that contain one or more double covalent bonds between carbon atoms in a chain are called **alkenes**. Because an alkene must have a double bond between carbon atoms, there is no 1-carbon alkene. The simplest alkene has two carbon atoms double-bonded to each other. The remaining four electrons—two from each carbon atom—are shared with four hydrogen atoms to give the molecule ethene (C_2H_4).

Alkenes with only one double bond constitute a homologous series. If you study the molecular formulas for the substances shown in **Table 22-4**, you will see that each has twice as many hydrogen atoms as carbon atoms. The general formula for the series is C_nH_{2n} . Each alkene has two fewer hydrogen atoms than the corresponding alkane because two electrons now form the second covalent bond and are no longer available for bonding to hydrogen atoms. What are the molecular formulas for 6-carbon and 9-carbon alkenes?

Objectives

- **Compare** the properties of alkenes and alkynes with those of alkanes.
- **Describe** the molecular structures of alkenes and alkynes.
- **Name** an alkene or alkyne by examining its structure.
- **Draw** the structure of an alkene or alkyne by analyzing its name.

Vocabulary

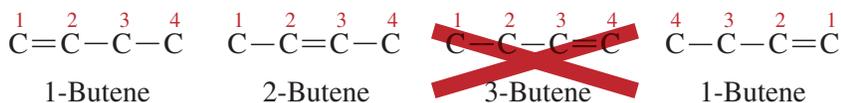
alkene
alkyne

Table 22-4

Examples of Alkenes					
Name	Molecular formula	Structural formula	Condensed structural formula	Melting point (°C)	Boiling point (°C)
Ethene	C_2H_4		$CH_2=CH_2$	-169	-104
Propene	C_3H_6		$CH_3CH=CH_2$	-185	-48
1-Butene	C_4H_8		$CH_2=CHCH_2CH_3$	-185	-6
2-Butene	C_4H_8		$CH_3CH=CHCH_3$	-106	0.8

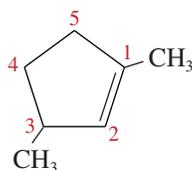
Naming alkenes Alkenes are named in much the same way as alkanes. Their names are formed by changing the *-ane* ending of the corresponding alkane to *-ene*. An alkane with two carbons is named *ethane*, and an alkene with two carbons is named *ethene*. Likewise, a three-carbon alkene is named propene. Ethene and propene have older, common names *ethylene* and *propylene*, respectively.

To name alkenes with four or more carbons in the chain, it is necessary to specify the location of the double bond. You do this by numbering the carbons in the parent chain starting at the end of the chain that will give the first carbon in the double bond the lowest number. Then you use only that number in the name.

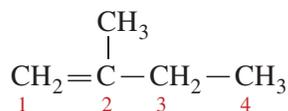


Note that the third structure is not “3-butene” because it is identical to the first structure, 1-butene. It is important to recognize that 1-butene and 2-butene are two different substances, each with its own properties.

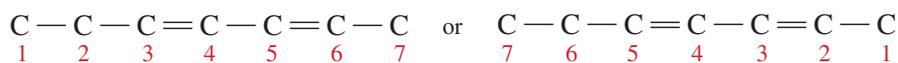
Cyclic alkenes are named in much the same way as cyclic alkanes; however, carbon number 1 must be one of the carbons connected by the double bond. Note the numbering in the compound shown below, 1,3-dimethylcyclopentene.



Naming branched-chain alkenes When naming branched-chain alkenes, follow the IUPAC rules for naming branched-chain alkanes—with two differences. First, in alkenes the parent chain is always the longest chain that contains the double bond, whether it is the longest chain of carbon atoms or not. Second, the position of the double bond, not the branches, determines how the chain is numbered. Note that there are two 4-carbon chains in the molecule shown below, but only the one with the double bond is used as a basis for naming. This branched-chain alkene is 2-methylbutene.



Some unsaturated hydrocarbons contain more than one double (or triple) bond. The number of double bonds in such molecules is shown by using a prefix (*di-*, *tri-*, *tetra-*, etc.) before the suffix *-ene*. The positions of the bonds are numbered in a way that gives the lowest set of numbers. Which numbering system would you use in the following example?

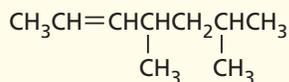


Because the molecule has a seven-carbon chain, you use the prefix *hepta-*. Because it has two double bonds, you use the prefix *di-* before *-ene*, giving the name *heptadiene*. Adding the numbers 2 and 4 to designate the positions of the double bonds, you arrive at the name *2,4-heptadiene*.

EXAMPLE PROBLEM 22-3

Naming Branched-Chain Alkenes

Name the alkene shown.

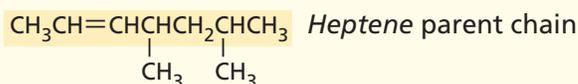


1. Analyze the Problem

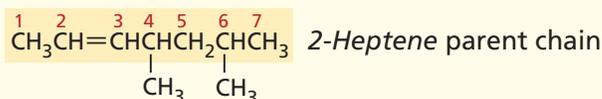
You are given a branched-chain alkene that contains one double bond and two alkyl groups. Follow the IUPAC rules to name the organic compound.

2. Solve for the Unknown

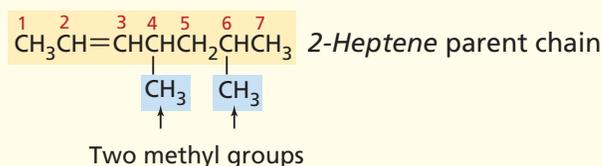
- a. The longest continuous carbon chain that includes the double bond contains seven carbons. The 7-carbon alkane is heptane, but the name is changed to heptene because a double bond is present.



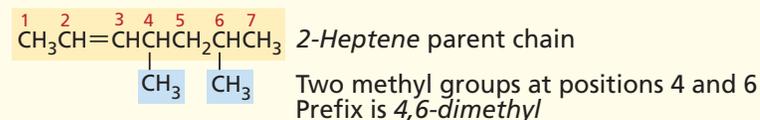
- b. Number the chain to give the lowest number to the double bond.



- c. Name each substituent.



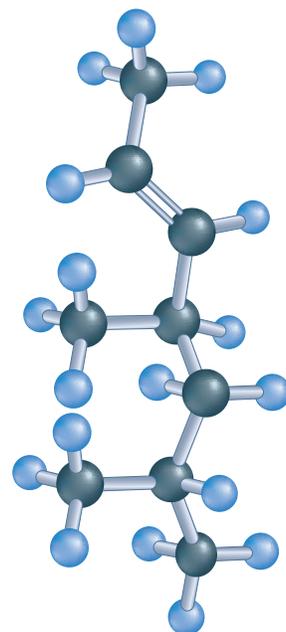
- d. Determine how many of each substituent is present, and assign the correct prefix to represent that number. Then, include the position numbers to get the complete prefix.



- e. The names of substituents do not have to be alphabetized because they are the same.
- f. Apply the complete prefix to the name of the parent alkene chain. Use commas to separate numbers and hyphens between numbers and words. Write the name *4,6-dimethyl-2-heptene*.

3. Evaluate the Answer

The longest carbon chain includes the double bond, and the position of the double bond has the lowest possible number. Correct prefixes and alkyl-group names designate the branches.



This is a ball-and-stick model of the structure in Example Problem 22-3.

Table 22-5

Examples of Alkynes					
Name	Molecular formula	Structural formula	Condensed structural formula	Melting point (°C)	Boiling point (°C)
Ethyne	C ₂ H ₂	H—C≡C—H	CH≡CH	−81	Sublimes at −85°C
Propyne	C ₃ H ₄	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}\equiv\text{C}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$	CH ₃ C≡CH	−103	−23
1-Butyne	C ₄ H ₆	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}\equiv\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$	CH≡CCH ₂ CH ₃	−126	8
2-Butyne	C ₄ H ₆	$\begin{array}{c} \text{H} \quad \quad \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}\equiv\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \quad \quad \text{H} \end{array}$	CH ₃ C≡CCH ₃	−32	27

The reactivities of alkynes make them useful starting materials in many synthesis reactions. In the **miniLAB** below, you will generate ethyne and investigate some of its properties.

miniLAB

Synthesis and Reactivity of Ethyne

Observing and Inferring Ethyne, often called *acetylene*, is used as a fuel in welding torches. In this lab, you will generate ethyne from the reaction of calcium carbide with water.

Materials 150-mL beaker, stirring rod, liquid dishwashing detergent, calcium carbide, forceps, wood splints, matches, ruler about 40 cm long, rubber band, phenolphthalein solution

Procedure

- Use a rubber band to attach a wood splint to one end of the ruler so that about 10 cm of the splint extends beyond the stick.
- Place 120 mL water in the beaker and add 5 mL dishwashing detergent. Mix thoroughly.
- Use forceps to pick up a pea-sized lump of calcium carbide (CaC₂). Do not touch the CaC₂ with your fingers. **CAUTION:** *If CaC₂ dust touches your skin, wash it away immediately with a lot of water.* Place the lump of CaC₂ in the beaker of detergent solution.

- Use a match to light the splint while holding the ruler at the opposite end. Immediately bring the burning splint to the bubbles that have formed from the reaction in the beaker. Extinguish the splint after observing the reaction.
- Use the stirring rod to dislodge a few large bubbles of ethyne and determine whether they float upward or sink in air.
- Rinse the beaker thoroughly, then add 25 mL distilled water and a drop of phenolphthalein solution. Use forceps to place a small piece of CaC₂ in the solution. Observe the results.

Analysis

- Describe your observations in steps 3 and 4. Could ethyne be used as a fuel?
- Based on your observations in step 5, what can you infer about the density of ethyne compared to the density of air?
- The reaction of calcium carbide with water yields two products. One is ethyne gas (C₂H₂). From your observation in step 6, suggest what the other product is, and write a balanced chemical equation for the reaction.

Figure 22-9

- a** An alkyne you may already be familiar with is ethyne, commonly called acetylene, used to produce an extremely hot flame needed for welding metals.
- b** Early automobiles burned acetylene in their headlamps to light the road ahead. Drivers had to get out to light the headlamps with a match.



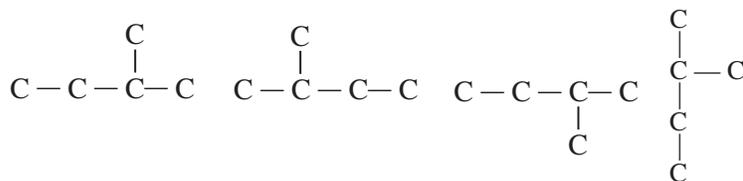
Properties of alkynes Alkynes have physical and chemical properties similar to those of the alkenes. Many of the reactions alkenes undergo, alkynes undergo as well. However, alkynes generally are more reactive than alkenes because the triple bond of alkynes has even greater electron density than the double bond of alkenes. This cluster of electrons is effective at inducing dipoles in nearby molecules, causing them to become unevenly charged and thus reactive.

Ethyne is a by-product of oil refining and also is made in industrial quantities by the reaction of calcium carbide, CaC_2 , with water. You learned about the production of ethyne by this method when you did the **miniLAB** on the previous page. When supplied with enough oxygen, ethyne burns with an intensely hot flame that can reach temperatures as high as 3000°C . Acetylene torches are commonly used in welding, as you see in **Figure 22-9**. Because the triple bond makes alkynes reactive, ethyne is used as a starting material in the manufacture of plastics and other organic chemicals used in industry.

Section 22.3 Assessment

- 20.** In what major way do the chemical properties of alkenes and alkynes differ from those of alkanes? What is responsible for this difference?
- 21.** Name the structures shown using IUPAC rules.
- a.**
$$\begin{array}{c} \text{CH}_3 \\ | \\ \text{CH} \equiv \text{CCH}_2 \\ | \\ \text{CH}_2\text{CH}_3 \end{array}$$
- b.**
$$\begin{array}{c} \text{CH}_3\text{CHCH} = \text{CHCH}_2\text{CH}_3 \\ | \\ \text{CH}_3 \end{array}$$
- c.**
$$\begin{array}{c} \text{CH}_3 \\ | \\ \text{CH}_3\text{C} = \text{CHCH} = \text{CH}_2 \end{array}$$
- d.**
$$\begin{array}{c} \text{CH}_3 \quad \text{CH}_3 \\ \diagdown \quad \diagup \\ \text{C} = \text{C} \\ \diagup \quad \diagdown \\ \text{CH}_3 \quad \text{CH}_3 \end{array}$$
- 22. Thinking Critically** Speculate on how the boiling and freezing points of alkynes compare with those of alkanes with the same number of carbon atoms. Explain your reasoning, then look up data to see if it supports your idea.
- 23. Making Predictions** A carbon atom in an alkane is bonded to four other atoms. In an alkene, a carbon in a double bond is bonded to three other atoms, and in an alkyne, a carbon in a triple bond is bonded to two other atoms. What geometric arrangement would you predict for the bonds surrounding the carbon atom in each of these cases? (Hint:) VSEPR theory can be used to predict shape.

Study the structural formulas in the following diagram. How many different organic compounds are represented by these structures?



If you said that all of these structures represent the same compound, you are correct. The diagram simply shows four ways of drawing a structural formula for 2-methylbutane. As you can see, the structural formula for a given hydrocarbon can be written in several ways. Before you continue reading, make sure that you understand why these structural formulas are all alike. You'll soon learn, however, that three distinctly different alkanes have the molecular formula C_5H_{12} .

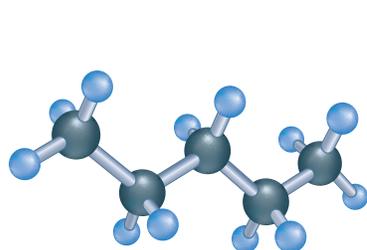
Structural Isomers

Now, examine the models of three alkanes in **Figure 22-10** to determine how they are similar and how they differ.

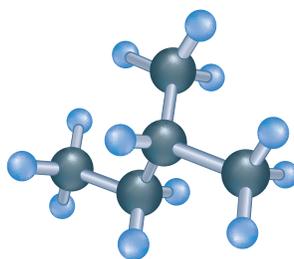
All three have five carbon atoms and 12 hydrogen atoms, so they have the molecular formula C_5H_{12} . However, as you can see, these models represent three different arrangements of atoms, pentane, 2-methylbutane, and 2,2-dimethylpropane. These three compounds are isomers. **Isomers** are two or more compounds that have the same molecular formula but different molecular structures. Note that cyclopentane and pentane are not isomers because cyclopentane's molecular formula is C_5H_{10} .

There are two main classes of isomers. **Figure 22-10** shows compounds that are examples of structural isomers. The atoms of **structural isomers** are bonded in different orders. The members of a group of structural isomers have different chemical and physical properties despite having the same formula. This observation supports one of the main principles of chemistry: The structure of a substance determines its properties. How does the trend in boiling points of C_5H_{12} isomers relate to their molecular structures?

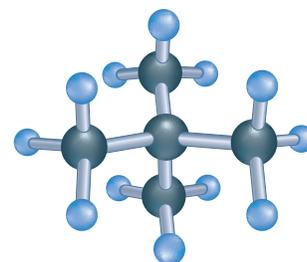
As the number of carbons in a hydrocarbon increases, the number of possible structural isomers increases. For example, nine alkanes having the molecular formula C_7H_{16} exist. $C_{20}H_{42}$ has 316 319 structural isomers.



Pentane
bp = 36°C



2-Methylbutane
bp = 28°C



2,2-Dimethylpropane
bp = 9°C

Objectives

- **Distinguish** between the two main categories of isomers, structural isomers and stereoisomers.
- **Differentiate** between *cis*- and *trans*-geometric isomers.
- **Recognize** different structural isomers given a structural formula.
- **Describe** the structural variation in molecules that results in optical isomers.

Vocabulary

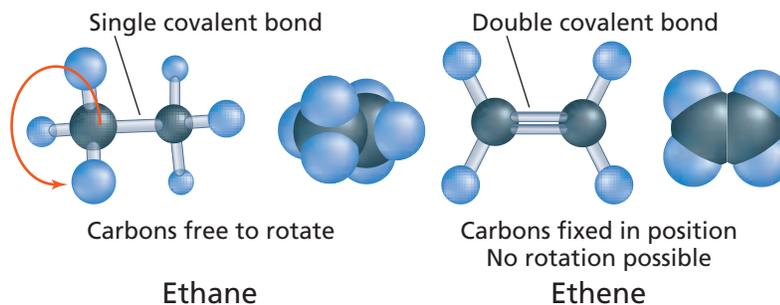
isomer
structural isomer
stereoisomer
geometric isomer
chirality
asymmetric carbon
optical isomer
polarized light
optical rotation

Figure 22-10

There are three different compounds that have the molecular formula C_5H_{12} . They are structural isomers. Note how their boiling points differ. Draw structural formulas for these three isomers.

Figure 22-11

The single-bonded carbons in ethane are free to rotate about the bond. The double-bonded carbons in ethene resist being rotated. How do you think this difference in ability to rotate would affect atoms or groups of atoms bonded to single-bonded and double-bonded carbon atoms?



Stereoisomers

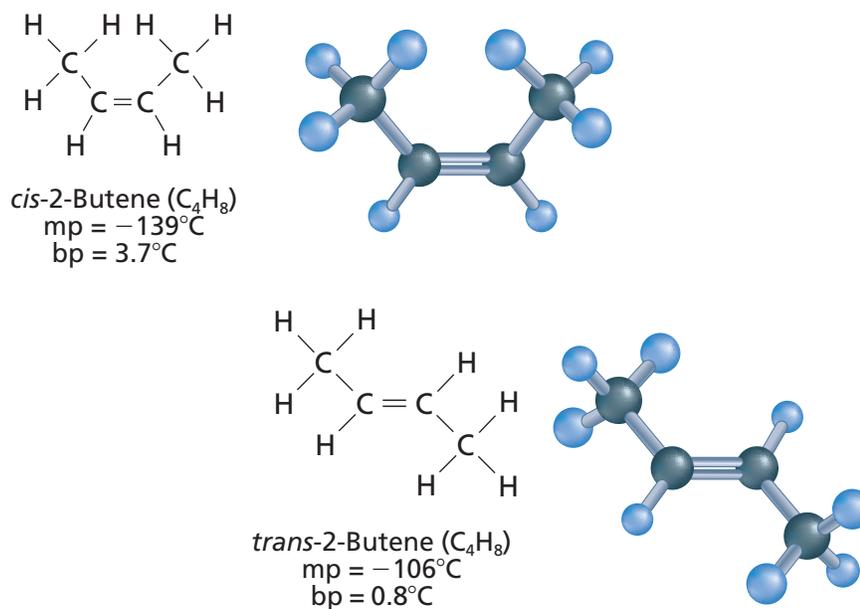
The second class of isomers involves a more subtle difference in bonding. **Stereoisomers** are isomers in which all atoms are bonded in the same order but are arranged differently in space. There are two types of stereoisomers. One type occurs in alkenes, which contain double bonds. Two carbon atoms with a single bond between them can rotate freely in relationship to each other. However, when a second covalent bond is present, the carbons can no longer rotate; they are locked in place, as shown in **Figure 22-11**.

To understand the consequences of this inability to rotate, compare the two possible structures of 2-butene shown in **Figure 22-12**. The arrangement in which the two methyl groups are on the same side of the molecule is indicated by the prefix *cis*-. The arrangement in which the two methyl groups are on opposite sides of the molecule is indicated by the prefix *trans*-. These terms derive from Latin: *cis* means *on the same side*, and *trans* means *across from*. Because the double-bonded carbons cannot rotate, the *cis*- form cannot easily change into the *trans*- form.

Isomers resulting from different arrangements of groups around a double bond are called **geometric isomers**. Note how the difference in geometry affects the isomers' physical properties such as melting point and boiling point. Geometric isomers differ in some chemical properties as well. If the compound is biologically active, such as a drug, the *cis*- and *trans*- isomers usually have greatly different effects. You may have read about the possible health concerns associated with *trans*- fatty acids in the diet. The *cis*- forms of the same acids seem not to be as harmful.

Figure 22-12

These isomers of 2-butene differ in the arrangement in space of the two methyl groups at the ends. The double-bonded carbon atoms cannot rotate with respect to each other, so the methyl groups must be in one of these two arrangements.



Chirality

In 1848, the young French chemist Louis Pasteur (1822–1895) reported his discovery that crystals of the organic compound tartaric acid, which is a by-product of the fermentation of grape juice to make wine, existed in two shapes that were not the same, but were mirror images of each other. Because a person's hands are like mirror images, as shown in **Figure 22-13a**, the crystals were called the right-handed and left-handed forms. The two forms of tartaric acid had the same chemical properties, melting point, density, and solubility in water, but only the left-handed form was produced by fermentation. In addition, bacteria were able to multiply when they were fed the left-handed form as a nutrient, but could not use the right-handed form.

Pasteur concluded that the two crystalline forms of tartaric acid exist because the tartaric acid molecules themselves exist in two arrangements, as shown in **Figure 22-13b**. Today, these two forms are called D-tartaric acid and L-tartaric acid. The letters D and L stand for the Latin prefixes *dextro-*, which means *to the right*, and *levo-*, which means *to the left*. Since Pasteur's time, chemists have discovered thousands of compounds that exist in right and left forms. This property is called **chirality**, a word derived from the Latin prefix *chiro-* for hand. Many of the substances found in living organisms, such as the amino acids that make up proteins, have this property. In general, living organisms make use of only one chiral form of a substance because only this form fits the active site of an enzyme.

Optical Isomers

In the 1860s, chemists realized that chirality occurs whenever a compound contains an asymmetric carbon. An **asymmetric carbon** is a carbon atom that has four different atoms or groups of atoms attached to it. The four groups always can be arranged in two different ways. Suppose that groups W, X, Y, and Z are attached to the same carbon atom in the two arrangements shown on the next page. Note that the structures differ in that groups X and Y have

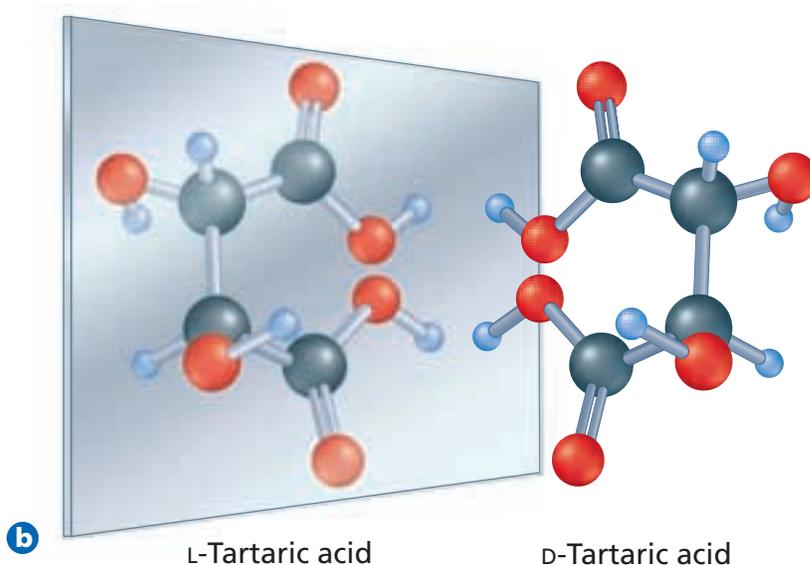


Figure 22-13

- Molecules of D-tartaric acid and L-tartaric acid are related to each other in the same way that your right hand and left hand are related. A mirror image of your right hand appears the same as your left hand.
- These models represent the two forms of tartaric acid that Pasteur studied. If the model of D-tartaric acid (right) is reflected in a mirror, its image is a model of L-tartaric acid.

been exchanged. You cannot rotate the two arrangements in any way that will make them identical to each other.



Now suppose that you build models of these two structures. Is there any way you could turn one structure so that it looks the same as the other? (Whether letters appear forward or backward does not matter.) You would discover that there is no way to accomplish the task without removing X and Y from the carbon atom and switching their positions. So, the molecules are different even though they look very much alike.

Isomers that result from different arrangements of four different groups about the same carbon atom represent another class of stereoisomers called **optical isomers**. Optical isomers have the same physical and chemical properties except in chemical reactions where chirality is important, such as enzyme-catalyzed reactions in biological systems. Human cells, for example, incorporate only L-amino acids into proteins. Only the L-form of ascorbic acid is active as Vitamin C. The chirality of a drug molecule can greatly affect its activity. For example, only the L-form of the drug methyldopa is effective in reducing high blood pressure.

Now try your hand at distinguishing among types of isomers in the **problem-solving LAB**.

Optical rotation Mirror-image isomers are called optical isomers because they affect light passing through them. Recall that light is a form of electromagnetic radiation—transverse waves that can travel through empty space. Normally, the light waves in a beam from the sun or a lightbulb move in all possible planes, as shown in **Figure 22-14**. However, light can be filtered or reflected in such a way that the resulting waves all lie in the same plane. This type of light is called **polarized light**.

problem-solving LAB

Identifying Structural, Geometric, and Optical Isomers

Interpreting Scientific Illustrations

Identifying isomers requires skill in visualizing a molecule in three dimensions. Building models of molecules helps clarify their geometry.

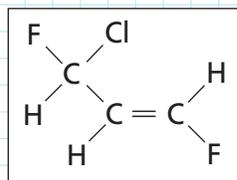
Analysis

Structure 1 represents an organic molecule. **Structures 2, 3, and 4** represent three different isomers of the first molecule. Study each of these three structures to determine how they are related to **Structure 1**.

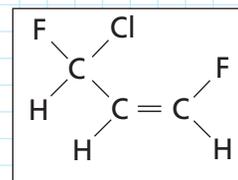
Thinking Critically

Write a sentence for each isomer describing how

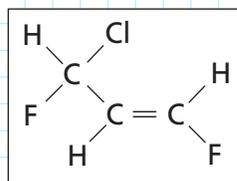
it differs from **Structure 1**. Which kind of isomer does each structure represent?



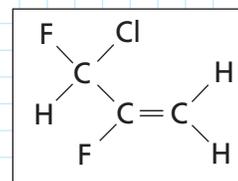
Structure 1



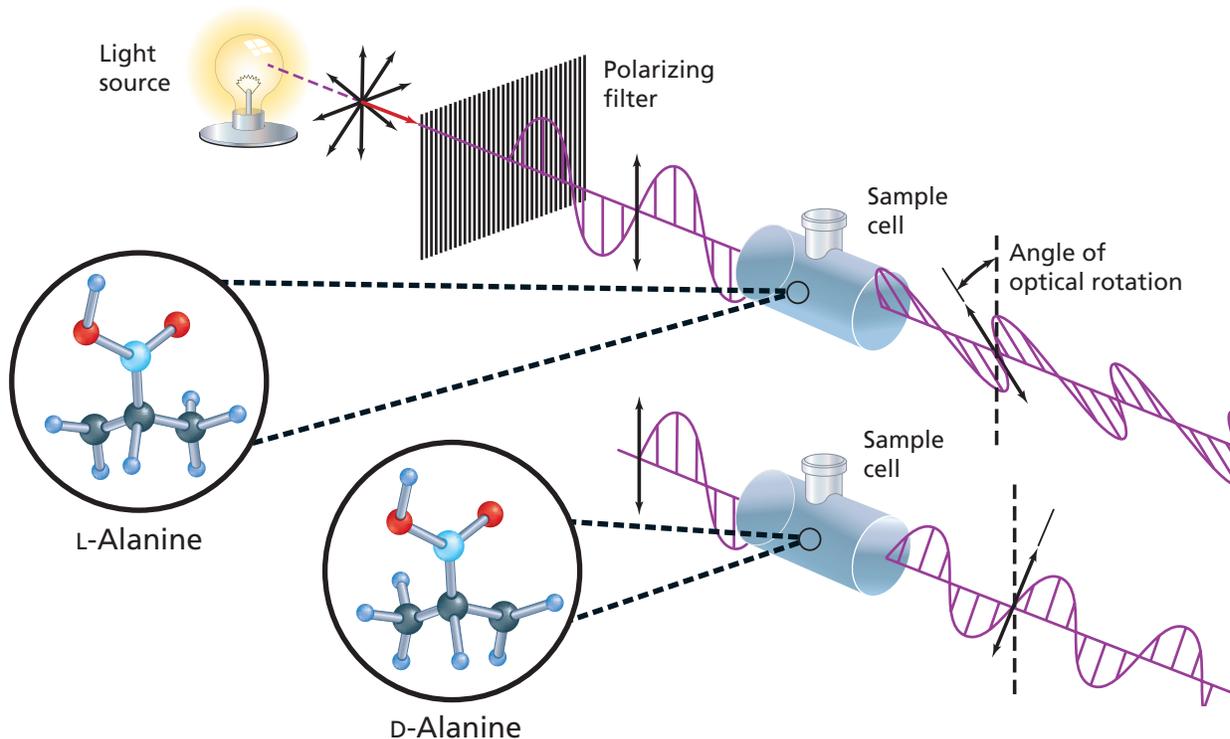
Structure 2



Structure 3



Structure 4



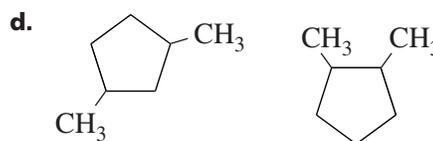
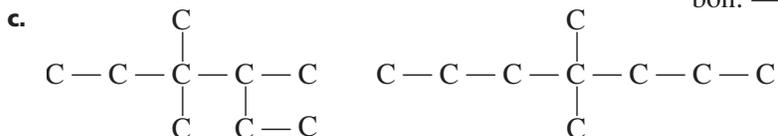
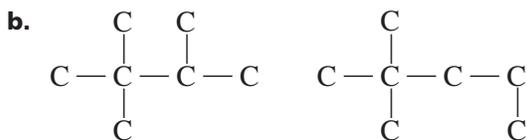
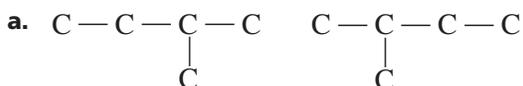
When polarized light passes through a solution containing an optical isomer, the plane of polarization is rotated to the right (clockwise, when looking toward the light source) by a D-isomer or to the left (counterclockwise) by an L-isomer. Rotation to the left is illustrated in the upper part of **Figure 22-14**. This effect is called **optical rotation**. What do you think happens to polarized light when it passes through a 50:50 mixture of the D-form and L-form of a chiral substance?

Figure 22-14

Polarized light can be produced by passing ordinary light through a filter that transmits light waves that lie in only one plane. Here, the filtered light waves are in a vertical plane before they pass through the sample cells.

Section 22.4 Assessment

- 24.** Explain the difference between structural isomers and stereoisomers.
- 25.** Draw all of the structural isomers possible for the alkane with a molecular formula of C_6H_{14} . Show only the carbon chains.
- 26.** Decide whether the carbon chains shown in each of the following pairs represent the same compound or pairs of isomers.



- 27.** Draw the structures of *cis*-3-hexene and *trans*-3-hexene.
- 28. Thinking Critically** A certain reaction yields 80% *trans*-2-pentene and 20% *cis*-2-pentene. Draw the structures of these two geometric isomers, and develop a hypothesis to explain why the isomers form in the proportions cited.
- 29. Formulating Models** Starting with a single carbon atom, draw two different optical isomers by attaching the following atoms or groups to the carbon: $-H$, $-CH_3$, $-CH_2CH_3$, $-CH_2CH_2CH_3$

Objectives

- **Compare** and **contrast** the properties of aromatic and aliphatic hydrocarbons.
- **Explain** what a carcinogen is and list some examples.
- **Describe** the processes used to separate petroleum into fractions and to balance each fraction's output with market demands.
- **Identify** the fractions into which petroleum can be separated.

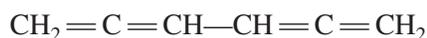
Vocabulary

aromatic compound
aliphatic compound
fractional distillation
cracking

By the middle of the nineteenth century, chemists had a basic understanding of the structures of hydrocarbons with single, double, and triple covalent bonds. However, a fourth class of hydrocarbon compounds remained a mystery. The simplest example of this class of hydrocarbon is benzene, which the English physicist Michael Faraday (1791–1867) had first isolated in 1825 from the gases given off when either whale oil or coal was heated.

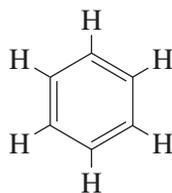
The Structure of Benzene

Although chemists had determined that benzene's molecular formula was C_6H_6 , it was hard for them to determine what sort of hydrocarbon structure would give such a formula. After all, the formula of the saturated hydrocarbon with six carbon atoms, hexane, was C_6H_{14} . Because the benzene molecule had so few hydrogen atoms, chemists reasoned that it must be unsaturated; that is, it must have several double or triple bonds, or a combination of both. They proposed many different structures, including this one suggested in 1860.



Although this structure has a molecular formula of C_6H_6 , such a hydrocarbon should be unstable and extremely reactive because of its many double bonds. However, benzene was fairly unreactive and, when it did react, it was not in the ways that alkenes and alkynes usually react. For that reason, chemists reasoned that structures such as the one shown above must be incorrect.

Kekulé's dream In 1865, the German chemist Friedrich August Kekulé (1829–1896) proposed a different kind of structure for benzene—a hexagon of carbon atoms with alternating single and double bonds. How does the molecular formula of this structure compare with that of benzene?



Kekulé claimed that benzene's structure came to him in a dream while he dozed in front of a fireplace in Ghent, Belgium. He said that he had dreamed of the Ouroboros, an ancient Egyptian emblem of a snake devouring its own tail, and that had made him think of a ring-shaped structure. The flat, hexagonal structure Kekulé proposed explained some of the properties of benzene, but it still could not explain benzene's lack of reactivity.

A modern model of benzene Since the time of Kekulé's proposal, research has confirmed that benzene's molecular structure is indeed hexagonal. However, an explanation of benzene's unreactivity had to wait until the 1930s when Linus Pauling proposed the theory of hybrid orbitals. When applied to benzene, this theory predicts that the pairs of electrons that form the second bond of each

of benzene's double bonds are not localized between only two specific carbon atoms as they are in alkenes. Instead, the electron pairs are delocalized, which means they are shared among all six carbons in the ring. **Figure 22-15** shows that this delocalization makes the benzene molecule chemically stable because electrons shared by six carbon nuclei are harder to pull away than electrons held by only two nuclei. The six hydrogen atoms are usually not shown, but you should remember that they are there. In this representation, the circle in the middle of the hexagon symbolizes the cloud formed by the three pairs of electrons.



Aromatic Compounds

Organic compounds that contain benzene rings as part of their structure are called **aromatic compounds**. The term *aromatic* was originally used because many of the benzene-related compounds known in the nineteenth century were found in pleasant-smelling oils that came from spices, fruits, and other plant parts. Hydrocarbons such as the alkanes, alkenes, and alkynes are called **aliphatic compounds** to distinguish them from aromatic compounds. The term *aliphatic* comes from the Greek word for fat, which is *aleiphatos*. Early chemists obtained aliphatic compounds by heating animal fats.

Structures of some aromatic compounds are shown in **Figure 22-16**. Note that naphthalene has a structure that looks like two benzene rings arranged side by side. Naphthalene is an example of a *fused ring system*, in which an organic compound has two or more cyclic structures with a common side. As in benzene, electrons in naphthalene are shared around all ten carbon atoms making up the double ring. Anthracene is another example of a fused ring system; it appears to be formed from three benzene rings.



Naphthalene

a Naphthalene is used in chemical manufacturing and in some kinds of moth repellent.

b *p*-Xylene is a starting material for the manufacture of polyester fabrics.



p-Xylene
(1,4-dimethylbenzene)

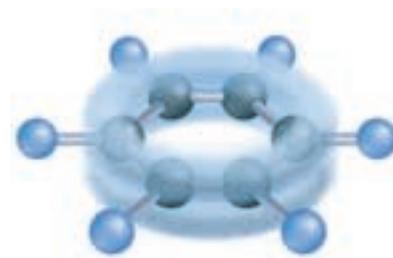


Figure 22-15

Benzene's bonding electrons spread evenly in a double-donut shape around the ring instead of remaining near individual atoms.

Figure 22-16

Shown here are a few of the many aromatic organic compounds that have practical uses. The common names of these compounds are used more frequently than their formal names.

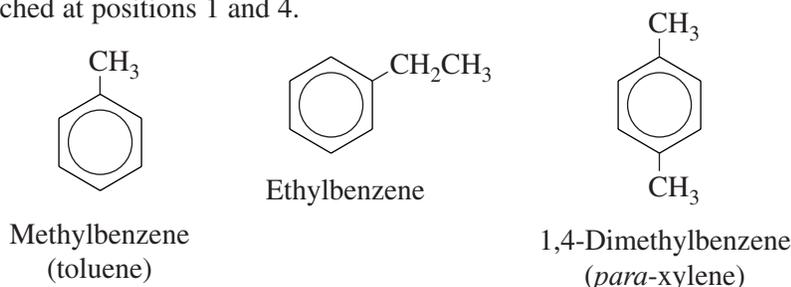


Anthracene

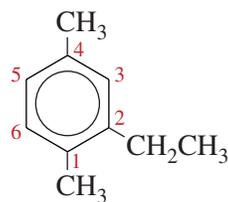
c Anthracene is important in the manufacture of richly colored dyes and pigments.

Substituted aromatic compounds Like other hydrocarbons, aromatic compounds may have different groups attached to their carbon atoms. For example, methylbenzene, also known as toluene, consists of a methyl group attached to a benzene ring in place of one hydrogen atom. Whenever you see something attached to an aromatic ring system, remember that the hydrogen atom is no longer there.

Substituted benzene compounds are named in the same way cyclic alkanes are. For example, ethylbenzene has a 2-carbon ethyl group attached, and 1,4-dimethylbenzene, also known as *para*-xylene, has methyl groups attached at positions 1 and 4.



Just as with substituted cycloalkanes, substituted benzene rings are numbered in a way that gives the lowest possible numbers for the substituents. In the following structure, numbering the ring as shown gives the numbers 1, 2, and 4 for the substituent positions. Because *ethyl* is lower in the alphabet than *methyl*, it is written first in the name 2-ethyl-1,4-dimethylbenzene.



2-Ethyl-1,4-dimethylbenzene

Carcinogens Many aromatic compounds, particularly benzene, toluene, and xylene, were once commonly used as industrial and laboratory solvents. However, tests have shown that the use of such compounds should be limited because they may affect the health of people who are exposed to them regularly. Health risks linked to aromatic compounds include respiratory ailments, liver problems, and damage to the nervous system. Beyond these hazards, some aromatic compounds are carcinogens, which are substances that can cause cancer.

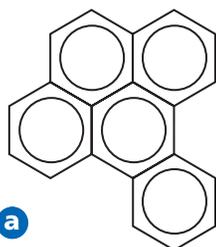
The first known carcinogen was an aromatic substance discovered around the turn of the twentieth century in chimney soot. Chimney sweeps in Great

Britain were known to have abnormally high rates of cancer of the scrotum, and the cause was found to be the aromatic compound benzopyrene, shown in **Figure 22-17a**. This compound is a by-product of the burning of complex mixtures of organic substances, such as wood and coal. Some aromatic compounds found in gasoline are also known to be carcinogenic, as you can tell from warning labels on gasoline pumps, **Figure 22-17b**.

Figure 22-17

a This is the structure of benzopyrene, produced when coal is burned for heat. It caused cancer in British chimney sweeps.

b Signs like this one warn consumers of the carcinogens in gasoline.



a
Benzopyrene



b

Natural Sources of Hydrocarbons

Today, benzene and other aromatic and aliphatic hydrocarbons are obtained from fossil fuels, which formed over millions of years from the remains of living things. The main source of these compounds is petroleum, a complex mixture of alkanes, some aromatic hydrocarbons, and organic compounds containing sulfur or nitrogen atoms.

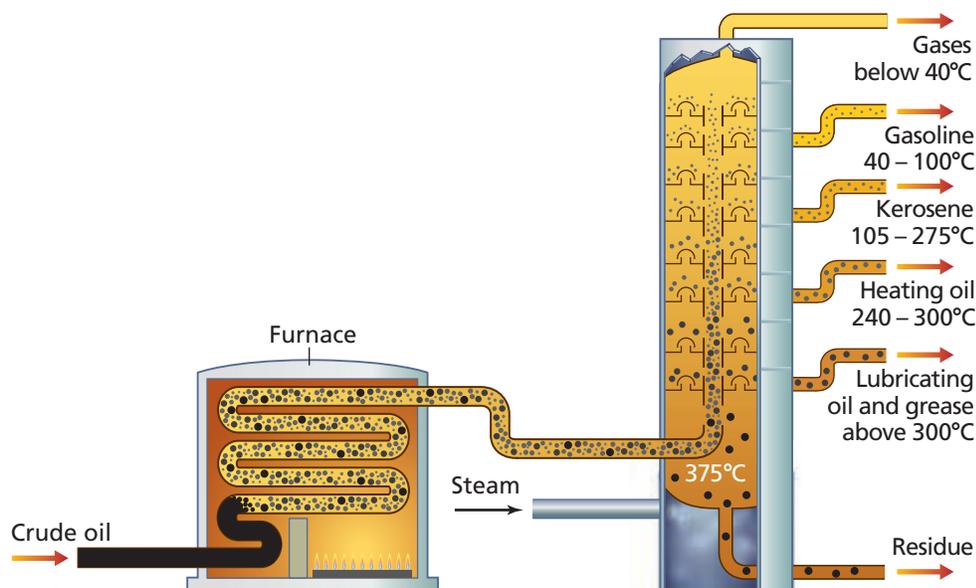
Petroleum formed from the remains of microorganisms that lived in Earth's oceans millions of years ago. Over time, the remains formed thick layers of mudlike deposits on the ocean bottom. Heat from Earth's interior and the tremendous pressure of overlying sediments transformed this mud into oil-rich shale and natural gas. In certain kinds of geological formations, the petroleum ran out of the shale and collected in pools deep in Earth's crust. Natural gas, which formed at the same time and in the same way as petroleum, is usually found with petroleum deposits. Natural gas is composed primarily of methane, but it also has small amounts of alkanes having two to five carbon atoms.

Fractional distillation Petroleum is a complex mixture containing more than a thousand different compounds. For this reason, raw petroleum, sometimes called crude oil, has little practical use. Petroleum is much more useful to humans when it is separated into simpler components, called fractions. Separation is carried out by boiling the petroleum and collecting the fractions as they condense at different temperatures. This process is called **fractional distillation** or sometimes just *fractionation*. Fractional distillation of petroleum is done in a fractionating tower similar to the one shown in **Figure 22-18**.

The temperature inside the fractionating tower is controlled so that it remains near 400°C at the bottom where the petroleum is boiling and gradually decreases moving toward the top. The condensation temperatures (boiling points) of alkanes generally decrease as molecular mass decreases. Therefore, as the vapors travel up through the column, hydrocarbons with more carbon atoms condense closer to the bottom of the tower and are drawn off. Hydrocarbons with fewer carbon atoms remain in the vapor phase until they reach regions of cooler temperatures farther up the column. By tapping

Figure 22-18

This diagram of a fractionating tower shows that fractions with lower boiling points, such as gasoline and gaseous products, are drawn off in the cooler regions near the top of the tower. Oils and greases, having much higher boiling points, stay near the bottom of the tower and are drawn off there.



CONNECTION

Throughout human history, people have collected petroleum to burn in lamps to provide light. They found petroleum seeping from cracks in rocks in certain locations. In fact, the word petroleum literally means "rock oil" and is derived from the Latin words for rock (*petra*) and oil (*oleum*). In the 19th century, as the U.S. entered the machine age and its population increased, the demand for petroleum to produce kerosene for lighting and as a machine lubricant also increased. Because there was no reliable petroleum supply, Edwin Drake drilled the first oil well in the United States near Titusville, Pennsylvania in 1859. The oil industry flourished for a time, but when Edison introduced the electric light in 1882, investors feared that the industry was doomed. However, the invention of the automobile in the 1890s soon revived the industry on a massive scale.

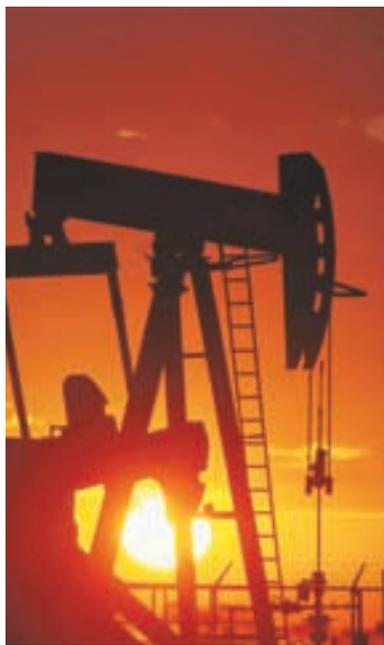


Table 22-6

Petroleum Components Separated by Fractional Distillation			
Fraction	Sizes of hydrocarbons	Boiling range (°C)	Common uses
Gases	CH_4 to C_4H_{10}	Below 40	Fuel gas, raw material for plastics manufacture
Gasoline	C_5H_{12} to $\text{C}_{12}\text{H}_{26}$	40–100	Fuel, solvents
Kerosene	$\text{C}_{12}\text{H}_{26}$ to $\text{C}_{16}\text{H}_{34}$	105–275	Home heating, jet fuel, diesel fuel
Heating oil	$\text{C}_{15}\text{H}_{32}$ to $\text{C}_{18}\text{H}_{38}$	240–300	Industrial heating
Lubricating oil	$\text{C}_{17}\text{H}_{36}$ and up	Above 300	Lubricants
Residue	$\text{C}_{20}\text{H}_{42}$ and up	Above 350	Tar, asphalt, paraffin

into the column at various heights, plant operators can collect the kinds of hydrocarbons they want.

Each group of components removed from the fractionating tower is called a fraction. **Table 22-6** is a list of the names given to the typical fractions separated from petroleum along with their boiling points, hydrocarbon size ranges, and common uses. Unfortunately, fractional distillation of petroleum does not yield these fractions in the same proportions that they are needed. For example, distillation seldom yields enough gasoline, yet it yields more of the heavier oils than the market demands.

Many years ago, petroleum chemists and engineers developed a process to help match the supply with the demand. This process in which heavier fractions are converted to gasoline by breaking their large molecules into smaller molecules is called **cracking**. Cracking is done in the absence of oxygen and in the presence of a catalyst. In addition to breaking heavier hydrocarbons into molecules of the size range needed for gasoline, cracking also produces starting materials for the synthesis of many different organic products.

Rating Gasolines

None of the petroleum fractions are pure substances. As you can see in **Table 22-6**, gasoline is not a pure substance but rather a mixture of hydrocarbons. Most alkane molecules in gasoline have 5 to 12 carbon atoms. However, the gasoline you put into your car today is different from the gasoline your great-grandparents put in their Model T in 1920. The gasoline fraction that is distilled from petroleum is modified by adjusting its composition and adding substances in order to improve its performance in today's automobile engines and to reduce pollution from car exhaust.

It is critical that the gasoline-air mixture in the cylinder of an automobile engine ignite at exactly the right instant and burn evenly. If it ignites too early or too late, much energy is wasted, fuel efficiency drops, and engines wear out prematurely. Most straight-chain hydrocarbons burn unevenly and tend to ignite from heat and pressure before the piston is in the proper position and the spark plug fires. This early ignition causes a rattling or pinging noise called *knocking*. Branched-chain alkanes, as well as alkenes and cyclic alkanes burn more evenly than alkanes with straight chains. Even burning helps prevent engine knocking.

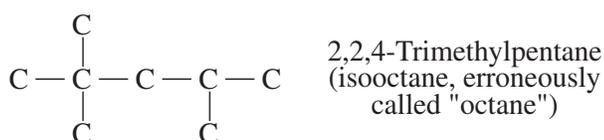


Figure 22-19

Gas stations provide a range of choices in octane ratings. Many cars manufactured today require mid-grade gasoline. However, cars with high-compression engines may need fuel with a higher octane rating.

Octane ratings In the late 1920s, an octane rating system for gasoline was established, resulting in the octane ratings posted on gasoline pumps like those shown in **Figure 22-19**. Mid-grade gasoline today has a rating of about 89, whereas premium gasolines have higher ratings of up to about 97. Several factors determine which octane rating a car needs, including how much the piston compresses the air-fuel mixture and the altitude at which the car is driven.

What do these ratings mean? You may be surprised to learn that they have almost nothing to do with the 8-carbon, straight-chain alkane called octane. They were first established by assigning a rating of zero to heptane, which was known to cause severe knocking, and a rating of 100 to 2,2,4-trimethylpentane, the compound that had the best anti-knock properties when tests were first performed. The compound 2,2,4-trimethylpentane was commonly called isooctane and erroneously called “octane” by technicians who tested gasoline.



A gasoline with a rating of 90 performs about the same as a mixture of 90% isooctane and 10% heptane. Today, compounds can be added to gasoline to produce octane ratings greater than 100.

Section 22.5 Assessment

- 30.** What properties of benzene made chemists think it was not an alkene with several double bonds?
- 31.** What feature accounts for the difference between aromatic and aliphatic hydrocarbons? Why should people avoid contact with aromatic hydrocarbons?
- 32.** Explain how the physical properties of hydrocarbons make fractional distillation possible.
- 33.** What is the purpose of cracking hydrocarbons?
- 34. Thinking Critically** In addition to adjusting octane rating, refiners also vary the volatility of gasoline, mainly by adding (or not adding) butane, C_4H_{10} . Where and when do you think refiners produce gasoline of higher volatility?
- 35. Interpreting Data** Look at the data in **Table 22-6**. What property of hydrocarbon molecules seems to correlate to the viscosity of a particular fraction when it is cooled to room temperature?

Analyzing Hydrocarbon Burner Gases

The fuel that makes a Bunsen burner work is a mixture of alkane hydrocarbons. One type of fuel is natural gas, whose primary component is methane (CH_4). The other type is called LP gas and consists primarily of propane (C_3H_8). In this experiment, you will use the ideal gas equation to help identify the main component of your classroom fuel supply.

Problem

What type of alkane gas is used in the burner fuel supplied to your laboratory?

Objectives

- **Measure** a volume of gas by water displacement.
- **Measure** the temperature, pressure, and mass at which the volume of the gas was measured.
- **Calculate** the molar mass of the burner gas using the ideal gas equation.

Materials

barometer
 thermometer
 1-L or 2-L plastic soda bottle with cap
 burner tubing
 pneumatic trough
 100-mL graduated cylinder
 balance (0.01g)
 paper towels

Safety Precautions



- Always wear safety goggles and an apron in the lab.
- Be certain that there are no open flames in the lab.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Include safety precautions, procedure notes, and a data table.
3. Use the formulas of methane, ethane, and propane to calculate the compounds' molar masses.
4. Given R and gas pressure, volume, and temperature, show how you will rearrange the ideal gas equation to solve for moles of gas.
5. Suppose that your burner gas contains a small amount of ethane (C_2H_6). How will the presence of this compound affect your calculated molar mass if the burner gas is predominantly:
 - a. methane
 - b. propane
6. Prepare your data table.

Mass and Volume Data

Mass of bottle + air (g)	
Mass of air (g)	
Mass of "empty" bottle (g)	
Mass of bottle + collected burner gas (g)	
Mass of collected burner gas (g)	
Barometric pressure (atm)	
Temperature ($^{\circ}\text{C}$)	
Temperature (K)	
Volume of gas collected (L)	

Procedure

1. Connect the burner tubing from the gas supply to the inlet of the pneumatic trough. Fill the trough with tap water. Open the gas valve slightly and let a little gas bubble through the tank in order to flush all of the air out of the tubing.



2. Measure the mass of the dry plastic soda bottle and cap. Record the mass in the data table (bottle + air). Record both the barometric pressure and the air temperature.
3. Fill the bottle to overflowing with tap water and screw on the cap. If some air bubbles remain, tap the bottle gently on the desktop until all air has risen to the top. Take off the cap, add more water, and recap the bottle.
4. Place the thermometer in the trough. Invert the capped bottle into the pneumatic trough and remove the cap while keeping the mouth of the bottle underwater. Hold the mouth of the bottle directly over the inlet opening of the trough.
5. Slowly open the gas valve and bubble gas into the inverted bottle until all of the water has been displaced. Close the gas valve immediately. Record the temperature of the water.
6. While the bottle is still inverted, screw on the cap. Remove the bottle from the water. Thoroughly dry the outside of the bottle.
7. Measure the mass of the bottle containing the burner gas and record the mass in the data table (bottle + burner gas).
8. Place the bottle in a fume hood and remove the cap. Compress the bottle several times to expel most of the gas. Refill the bottle to overflowing with water and determine the volume of the bottle by pouring the water into a graduated cylinder. Record the volume of the bottle.

Cleanup and Disposal

1. Be certain that all gas valves are closed firmly and dump water out of pneumatic troughs.
2. Clean up water spills and dispose of materials as directed by your teacher.
3. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Acquiring Information** Use the volume of the bottle and look up the density of air to compute the mass of the air the bottle contains. Use gas laws to compute the density of air at the temperature and pressure of your laboratory. The density of air at 1 atm and 20°C is 1.205 g/L.
2. **Using Numbers** Calculate the mass of the empty bottle by subtracting the mass of air from the mass of the bottle and air combined.
3. **Using Numbers** Determine the mass of the collected gas by subtracting the mass of the empty bottle from the mass of the bottle and gas.
4. **Interpreting Data** Use the volume of gas, water temperature, and barometric pressure along with the ideal gas law to calculate the number of moles of gas collected.
5. **Using Numbers** Use the mass of gas and the number of moles to calculate the molar mass of the gas.
6. **Drawing a Conclusion** How does your experimental molar mass compare with the molar masses of methane, ethane, and propane? Suggest which of these gases are in the burner gas in your laboratory.
7. **Error Analysis** If your experimental molar mass does not agree with that of any one of the three possible gases, suggest possible sources of error in the experiment. What factor other than error could produce such a result?

Real-World Chemistry

1. Substances called *odorants* are mixed with natural gas before it is distributed to homes, businesses, and institutions. Why must an odorant be used, and what substances are used as odorants?
2. At 1 atm and 20°C, the densities of methane and propane are 0.65 g/L and 1.83 g/L, respectively. Would either gas tend to settle in a low area such as the basement of a home? Explain.

Everyday Chemistry

Unlimited Alternative Energy

Fossil fuels are the primary source of energy to run our cars, heat our homes, and produce electricity. Decaying plants and animals millions of years ago produced coal, oil, and natural gas. These sources of energy are finite and need to be conserved.

The Costs of Fossil Fuels

Each time we pay an electricity bill or purchase gasoline, we pay for fossil fuels. The labor to mine coal or drill for oil, the labor and materials to build and operate power plants, and the transportation of coal and oil to these plants are just part of the costs of using fossil fuels. There also are indirect costs such as health problems caused by pollution, environmental problems such as acid rain, and the protection of foreign sources of oil. There is an alternative source of energy readily available for use in your everyday life—solar energy.

What is solar energy?

The sun has always been an energy source. Plants use sunlight to make food through a process called photosynthesis. Animals use energy from the Sun by eating plants. Throughout history, humans have used the heat from sunlight directly to cook food and heat water and homes. When you hang laundry outside to dry, you use solar energy to do work. Plants grow in greenhouses during winter months due to warmth from the Sun. Although the Sun radiates prodigious amounts of energy each day, it will continue to disperse solar energy for approximately five billion years.

By the late 1800s, the use of solar water heaters was common in sunny parts of the United States. When large deposits of oil and natural gas were discovered, these systems were replaced with heaters using cheaper fossil fuels.

Passive Solar Energy

Many different active solar techniques can be used to convert sunlight into useful forms of energy. Add-on devices such as photovoltaic cells that convert sunlight into electricity and rooftop solar panels that use sunlight to heat water use mechanical means to distribute solar energy. "Low-technology" passive solar techniques provide clean, inexpensive energy. Passive solar techniques make use of the building's components and design.

Each time you open the curtains to let in the Sun's rays for warmth or light, you are using passive solar energy. Every building can meet some of its heating requirements with passive solar energy. The use of triple-pane windows, heavily insulated walls and ceilings, and materials with high heat capacity such as adobe walls and clay tile floors increase energy storage.

Passive solar techniques also can be used to reduce use of electricity for air conditioning. Vegetation planted for shade, light colors that reflect sunlight, and careful attention to placement of windows for good airflow will keep a home cooler in summer.

The use of sunlight to replace electric lighting in a building is called daylighting. Daylighting can be useful in the home, and large office buildings that demand large amounts of lighting during the day benefit from large windows, skylights, and atria.



Testing Your Knowledge

- 1. Acquiring Information** Investigate and describe how a solar furnace works. Is it an active or passive use of solar energy? Explain your answer.
- 2. Hypothesizing** Geothermal energy, wind energy, and solar energy are forms of alternative energy. Which form might be feasible in your state? Explain.

Summary

22.1 Alkanes

- Organic compounds contain the element carbon, which is able to bond with other carbon atoms to form straight chains and branched chains.
- Hydrocarbons are organic substances composed of only the elements carbon and hydrogen.
- Alkanes contain only single bonds between carbon atoms.
- Alkanes and other organic compounds are best represented by structural formulas and can be named using systematic rules determined by the International Union of Pure and Applied Chemistry (IUPAC).

22.2 Cyclic Alkanes and Alkane Properties

- Alkanes that contain hydrocarbon rings are called cyclic alkanes.
- Alkanes are nonpolar compounds that have low reactivity and lower melting and boiling points than polar molecules of similar size and mass.

22.3 Alkenes and Alkynes

- Alkenes and alkynes are hydrocarbons that contain at least one double or triple bond, respectively.
- Alkenes and alkynes are nonpolar compounds with greater reactivity than alkanes but with other properties similar to those of the alkanes.
- Alkenes and alkynes, whether straight-chain, branched-chain, or cyclic, can be named using the systematic rules determined by IUPAC.

22.4 Isomers

- Isomers are two or more compounds with the same molecular formula but different molecular structures.

- Structural isomers differ in the order in which atoms are bonded to each other. A straight-chain hydrocarbon and a branched-chain hydrocarbon with the same molecular formula are structural isomers.
- Stereoisomers have all atoms bonded in the same order but arranged differently in space.
- Geometric isomers, a category of stereoisomers, result from different arrangements of groups about carbon atoms that are double bonded to each other.
- Optical isomers, another class of stereoisomers, result from the two possible arrangements of four different atoms or groups of atoms bonded to the same carbon atom. The two isomers are chiral because they are mirror images of each other.

22.5 Aromatic Hydrocarbons and Petroleum

- Aromatic hydrocarbons contain benzene rings as part of their molecular structures. Nonaromatic hydrocarbons are called aliphatic hydrocarbons.
- Aromatic hydrocarbons tend to be less reactive than alkenes or alkynes because they have no double bonds. Instead, electrons are shared evenly over the entire benzene ring.
- Some aromatic compounds, such as naphthalene, contain two or more benzene rings fused together.
- Some aromatic compounds are carcinogenic, which means they can cause cancer.
- The major sources of hydrocarbons are petroleum and natural gas.
- Petroleum can be separated into components of different boiling ranges by the process of fractional distillation.

Vocabulary

- | | | |
|-------------------------------|------------------------------------|------------------------------------|
| • aliphatic compound (p. 723) | • cycloalkane (p. 706) | • parent chain (p. 701) |
| • alkane (p. 699) | • fractional distillation (p. 725) | • polarized light (p. 720) |
| • alkene (p. 711) | • geometric isomer (p. 718) | • saturated hydrocarbon (p. 710) |
| • alkyne (p. 714) | • homologous series (p. 701) | • stereoisomer (p. 718) |
| • aromatic compound (p. 723) | • hydrocarbon (p. 698) | • structural isomer (p. 717) |
| • asymmetric carbon (p. 719) | • isomer (p. 717) | • substituent group (p. 701) |
| • chirality (p. 719) | • optical isomer (p. 720) | • unsaturated hydrocarbon (p. 710) |
| • cracking (p. 726) | • optical rotation (p. 721) | |
| • cyclic hydrocarbon (p. 706) | • organic compound (p. 698) | |

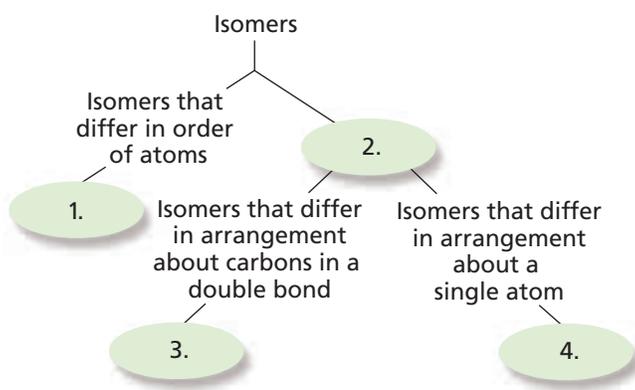


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 22 Assessment.

Concept Mapping

36. Complete the following concept map that shows how the following isomer types are related: stereoisomers, structural isomers, optical isomers, all isomers, geometric isomers.



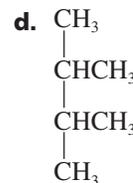
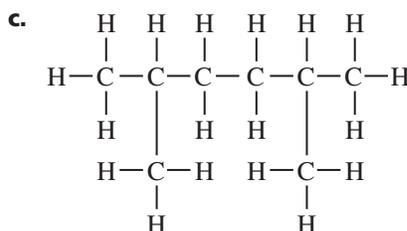
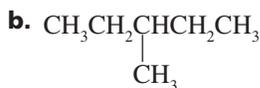
Mastering Concepts

37. What is the main characteristic of an organic compound? (22.1)
38. What characteristic of carbon accounts for the huge variety of organic compounds? (22.1)
39. Describe the characteristics of a homologous series of hydrocarbons. (22.1)
40. Draw the structural formula of each of the following. (22.1)
- a. ethane b. butane c. hexane
41. Write the condensed structural formulas for the alkanes in the previous question. (22.1)
42. Write the name and draw the structure of the alkyl group that corresponds to each of the following alkanes. (22.1)
- a. methane b. butane c. octane
43. How does the structure of a cycloalkane differ from that of a straight-chain or branched-chain alkane? (22.2)
44. Explain the difference between saturated hydrocarbons and unsaturated hydrocarbons. (22.2)
45. Explain how intermolecular attractions generally affect substances' boiling and freezing points.
46. Explain how alkenes differ from alkanes. How do alkynes differ from both alkenes and alkanes? (22.3)
47. The names of hydrocarbons are based on the name of the parent chain. Explain how the determination of the parent chain when naming alkenes differs from the same determination when naming alkanes. (22.3)
48. Name the most common alkyne. How is this substance used? (22.3)
49. How are two isomers alike and how are they different? (22.4)
50. Describe the difference between *cis*- and *trans*- isomers in terms of geometrical arrangement. (22.4)
51. What characteristics does a chiral substance have? (22.4)
52. How does polarized light differ from ordinary light, such as light from the Sun? (22.4)
53. How do optical isomers affect polarized light? (22.4)
54. What structural characteristic do all aromatic hydrocarbons share? (22.5)
55. Draw the structural formula of 1,2-dimethylbenzene. (22.5)
56. What are carcinogens? (22.5)
57. What does fractional distillation of petroleum accomplish? (22.5)
58. What physical property determines the height at which hydrocarbons condense in a fractionation tower? (22.5)
59. What is the cracking process and why is it necessary in petroleum processing? (22.5)

Mastering Problems

Alkanes (22.1)

60. Name the compound represented by each of the following structural formulas.



61. Draw full structural formulas for the following compounds.

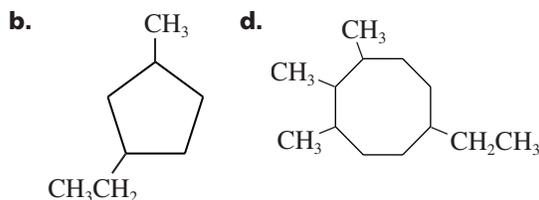
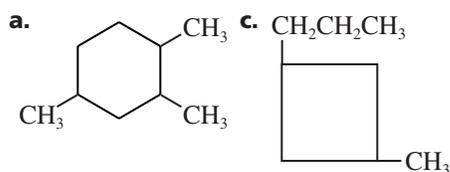
- a. heptane
 b. 2-methylhexane
 c. 2,3-dimethylpentane
 d. 2,2-dimethylpropane

Cyclic Alkanes and Alkane Properties (22.2)

62. Draw condensed structural formulas for the following compounds. Use line structures for rings.

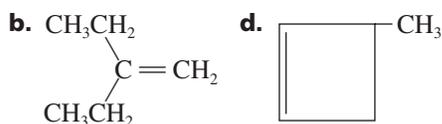
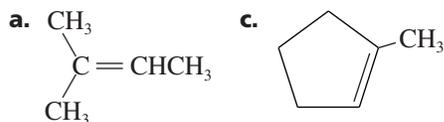
- a. 1,2-dimethylcyclopropane
 b. 1,1-diethyl-2-methylcyclopentane

63. Name the compound represented by each of the following structural formulas.



Alkenes and Alkynes (22.3)

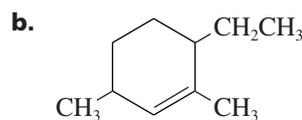
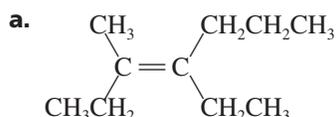
64. Name the compound represented by each of the following condensed structural formulas.



65. Draw condensed structural formulas for the following compounds. Use line structures for rings.

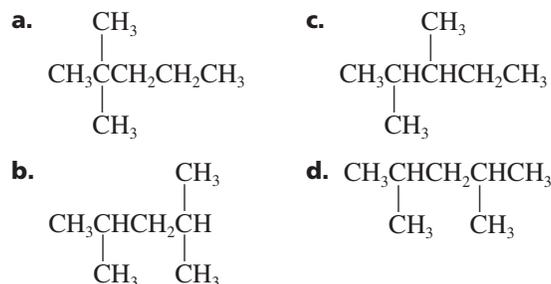
- a. 1,4-diethylcyclohexene
 b. 2,4-dimethyl-1-octene
 c. 2,2-dimethyl-3-hexyne

66. Name the compound represented by each of the following condensed structural formulas.

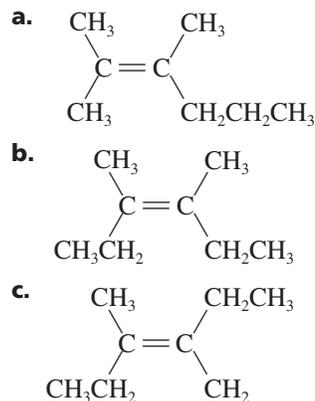


Isomers (22.4)

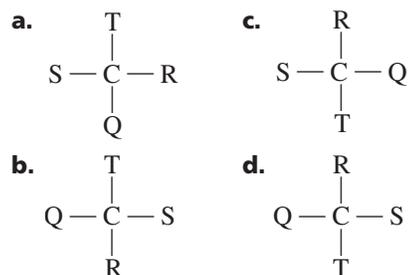
67. Identify the pair of structural isomers in the following group of condensed structural formulas.



68. Identify the pair of geometric isomers among the following structures. Explain your selections. Explain how the third structure is related to the other two.



69. Three of the following structures are exactly alike, but the fourth represents an optical isomer of the other three. Identify the optical isomer and explain how you made your choice.

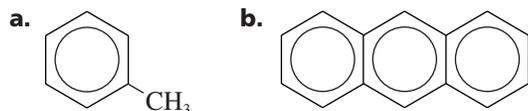


70. Draw condensed structural formulas for four different structural isomers with the molecular formula C_4H_8 .

71. Draw and label the *cis*- and *trans*- isomers of the molecule represented by the following condensed formula.
 $\text{CH}_3\text{CH} = \text{CHCH}_2\text{CH}_3$

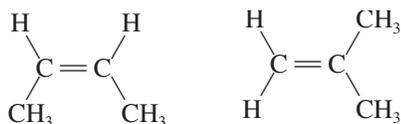
Aromatic Hydrocarbons and Petroleum (22.5)

72. Name the compound represented by each of the following structural formulas.


Mixed Review

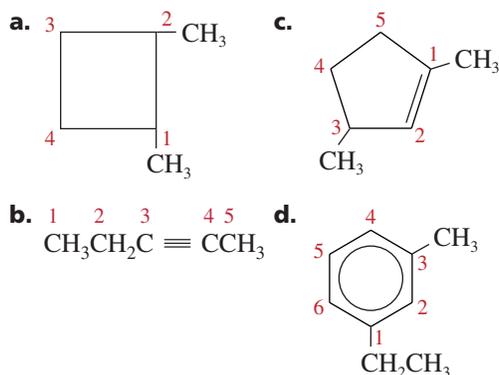
Sharpen your problem-solving skills by answering the following.

73. Do the following structural formulas represent the same molecule? Explain your answer.



74. How many hydrogen atoms are in an alkane molecule with nine carbon atoms? How many are in an alkene with nine carbon atoms and one double bond?

75. Determine whether or not each of the following structures represents the correct numbering. If the numbering is incorrect, redraw the structure with the correct numbering.



76. Why do chemists use structural formulas for organic compounds rather than molecular formulas such as C_5H_{12} ?

77. The general formula for alkanes is $\text{C}_n\text{H}_{2n+2}$. Determine the general formula for cycloalkanes.

78. Why are unsaturated hydrocarbons more useful than saturated hydrocarbons as starting materials in chemical manufacturing?

79. Is cyclopentane an isomer of pentane? Explain your answer.

Thinking Critically

80. **Thinking Critically** Determine which two of the following names cannot be correct and draw the structures of the molecules.

- 2-ethyl-2-butene
- 1,4-dimethylcyclohexene
- 1,5-dimethylbenzene

81. **Drawing a Conclusion** The sugar glucose is sometimes called *dextrose* because a solution of glucose is known to be dextrorotatory. Analyze the word *dextrorotatory*, and suggest what the word means.

82. **Interpreting Scientific Illustrations** Draw Kekulé's structure of benzene and explain why it does not truly represent the actual structure.

83. **Recognizing Cause and Effect** Explain why alkanes such as hexane and cyclohexane are effective at dissolving grease, whereas water is not.

84. **Hypothesizing** Do you think that, on average, structural isomers or stereoisomers will have a larger difference in their physical properties? Explain your reasoning. Research this question by comparing physical properties of pairs of isomers as given in the *CRC Handbook of Chemistry and Physics* or *The Merck Index* to see if your hypothesis is correct.

Writing in Chemistry

85. For many years, a principal antiknock ingredient in gasoline was the compound tetraethyllead. Do research to learn about the structure of this compound, the history of its development and use, and why its use was discontinued in the United States. Find out if it is still used as a gasoline additive elsewhere in the world.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

86. What element has the following ground-state electron configuration: $[\text{Ar}]4s^23d^6$? (Chapter 5)

87. What is the charge of an ion formed from the following families? (Chapter 7)

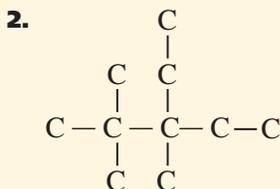
- alkali metals
- alkaline earth metals
- halogens

88. Write the chemical equations for the complete combustion of ethane, ethene, and ethyne into carbon dioxide and water. (Chapter 10)

Use these questions and test-taking tip to prepare for your standardized test.

1. The condensed structural formula of heptane is ____.

a. $(\text{CH}_3)_2(\text{CH}_2)_5$ c. $\text{CH}_3(\text{CH}_2)_5\text{CH}_3$
 b. $\text{CH}_3(\text{CH}_2)_6$ d. $\text{CH}_3\text{CH}_3(\text{CH}_2)_5$



The name of the compound whose skeletal formula is shown above is ____.

- a. 2,2,3-trimethyl-3-ethylpentane
 b. 3-ethyl-3,4,4-trimethylpentane
 c. 2-butyl-2-ethylbutane
 d. 3-ethyl-2,2,3-trimethylpentane
3. All of the following are structural isomers of $\text{CH}_2 = \text{CHCH}_2\text{CH} = \text{CHCH}_3$ EXCEPT ____.
- a. $\text{CH}_2 = \text{CHCH}_2\text{CH}_2\text{CH} = \text{CH}_2$
 b. $\text{CH}_3\text{CH} = \text{CHCH}_2\text{CH} = \text{CH}_2$
 c. $\text{CH}_3\text{CH} = \text{CHCH} = \text{CHCH}_3$
 d. $\text{CH}_2 = \text{C} = \text{CHCH}_2\text{CH}_2\text{CH}_3$

Interpreting Tables Use the table to answer questions 4–6.

Data for Various Hydrocarbons				
Name	Number of C atoms	Number of H atoms	Melting point (°C)	Boiling point (°C)
Heptane	7	16	-90.6	98.5
1-Heptene	7	14	-119.7	93.6
1-Heptyne	7	12	-81	99.7
Octane	8	18	-56.8	125.6
1-Octene	8	16	-101.7	121.2
1-Octyne	8	14	-79.3	126.3

4. Based on the information in the table, the type of hydrocarbon that becomes a gas at the lowest temperature is an ____.
- a. alkane c. alkyne
 b. alkene d. aromatic

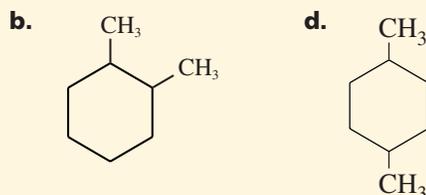
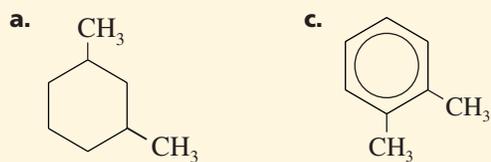
5. If n is the number of carbon atoms in the hydrocarbon, then the general formula for an alkyne with one triple bond is ____.

a. C_nH_{n+2} c. C_nH_{2n}
 b. $\text{C}_n\text{H}_{2n+2}$ d. $\text{C}_n\text{H}_{2n-2}$

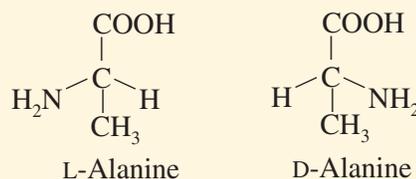
6. It can be predicted from the table that nonane will have a melting point that is ____.

a. greater than that of octane
 b. less than that of heptane
 c. greater than that of decane
 d. less than that of hexane

7. Which compound below is 1,2-dimethylcyclohexane?



8. Alanine, like all amino acids, exists in two forms:



Almost all of the amino acids found in living organisms are in the L-form. L-Alanine and D-alanine are ____.

- a. structural isomers c. optical isomers
 b. geometric isomers d. stereoisotopes



TEST-TAKING TIP

Beat the Clock ... And Then Go Back As you take a practice test, pace yourself to finish each section just a few minutes early so you can go back and check over your work. You'll sometimes find a mistake or two. Don't worry. To err is human. To catch it before you hand it in is better.

Substituted Hydrocarbons and Their Reactions

What You'll Learn

- ▶ You will recognize the names and structures of several important organic functional groups.
- ▶ You will classify reactions of organic substances as substitution, addition, elimination, oxidation-reduction, or condensation and predict products of these reactions.
- ▶ You will relate the structures of synthetic polymers to their properties.

Why It's Important

Whether you are removing a sandwich from plastic wrap, taking an aspirin, or shooting baskets, you're using organic materials made of substituted hydrocarbons. These compounds are in turn made of molecules whose atoms include carbon, hydrogen, and other elements.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about the chemistry of substituted hydrocarbons.

The spooled threads shown in the photo are made from large organic molecules called polymers.



DISCOVERY LAB



Making Slime

In addition to carbon and hydrogen, most organic substances contain other elements that give the substances unique properties. In this lab, you will work with an organic substance consisting of long carbon chains to which many $-OH$ groups are bonded. How will the properties of this substance change when these groups react to form bonds called crosslinks between the chains?

Safety Precautions



Do not allow solutions or product to contact eyes or exposed skin.

Materials

4% sodium tetraborate (borax) solution
4% polyvinyl alcohol solution
disposable plastic cup
stirring rod

Procedure

1. Pour 20 mL of 4% polyvinyl alcohol solution into a small disposable plastic cup. Note the viscosity of the solution as you stir it.
2. While stirring, add 6 mL of 4% sodium tetraborate solution to the polyvinyl alcohol solution. Continue to stir until there is no further change in the consistency of the product.
3. Use your gloved hand to scoop the material out of the cup. Knead the polymer into a ball.

Analysis

What physical property of the product differs markedly from those of the reactants?

Section

23.1

Functional Groups

Objectives

- **Describe** a functional group and give examples.
- **Name** and **draw** alkyl and aryl halide structures.
- **Discuss** the chemical and physical properties of organic halides.
- **Describe** how substitution reactions form alkyl and aryl halides.

Vocabulary

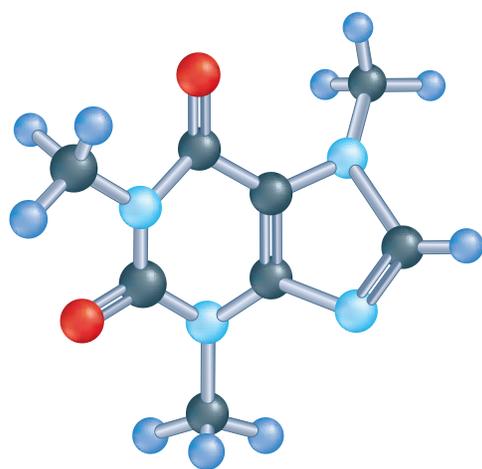
functional group
halocarbon
alkyl halide
aryl halide
substitution reaction
halogenation

All of the spools shown on the opposite page hold thread manufactured from large organic molecules called polymers. By weaving these polymer threads into fabrics, they can serve many different purposes. Most of the organic molecules used to make these polymers contain atoms of other elements in addition to carbon and hydrogen. The presence of these other elements gives fabric such strength that it can be used to make a bulletproof vest or such resistance to creasing that it never has to be ironed.

Functional Groups

You learned in Chapter 22 that thousands of different hydrocarbons exist because carbon atoms can link together to form straight and branched chains, rings of many sizes, and molecules with single, double, and triple bonds. In hydrocarbons, carbon atoms are linked only to other carbon atoms or hydrogen atoms. But carbon atoms also can form strong covalent bonds with other elements, the most common of which are oxygen, nitrogen, fluorine, chlorine, bromine, iodine, sulfur, and phosphorus.

Atoms of these elements occur in organic substances as parts of functional groups. A **functional group** in an organic molecule is an atom or group of atoms that always reacts in a certain way. The addition of a functional group



Caffeine

Figure 23-1

Many organic compounds contain atoms of elements in addition to carbon and hydrogen. For example, caffeine, a compound found in many beverages, contains both oxygen and nitrogen atoms. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

to a hydrocarbon structure always produces a substance with physical and chemical properties that differ from those of the parent hydrocarbon. Organic compounds containing several important functional groups are shown in **Table 23-1**. The symbols R and R' represent any carbon chains or rings bonded to the functional group. And * represents a hydrogen atom, carbon chain, or carbon ring.

Table 23-1

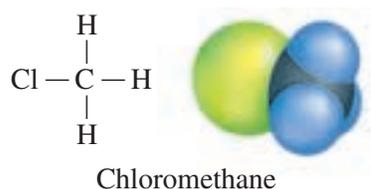
Organic Compounds and Their Functional Groups		
Compound type	General formula	Functional group
Halocarbon	R—X (X = F, Cl, Br, I)	Halogen
Alcohol	R—OH	Hydroxyl
Ether	R—O—R'	Ether
Amine	R—NH ₂	Amino
Aldehyde	$\begin{array}{c} \text{O} \\ \parallel \\ * - \text{C} - \text{H} \end{array}$	Carbonyl
Ketone	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R} - \text{C} - \text{R}' \end{array}$	Carbonyl
Carboxylic acid	$\begin{array}{c} \text{O} \\ \parallel \\ * - \text{C} - \text{OH} \end{array}$	Carboxyl
Ester	$\begin{array}{c} \text{O} \\ \parallel \\ * - \text{C} - \text{O} - \text{R} \end{array}$	Ester
Amide	$\begin{array}{c} \text{O} \quad \text{H} \\ \parallel \quad \\ * - \text{C} - \text{N} - \text{R} \end{array}$	Amido

Keep in mind that double and triple bonds between two carbon atoms are considered functional groups even though only carbon and hydrogen atoms are involved. By learning the properties associated with a given functional group, you can predict the properties of organic compounds for which you know the structure, even if you have never studied them. Examine the caffeine structure shown in **Figure 23-1** and identify the molecule's functional groups.

Organic Compounds Containing Halogens

The simplest functional groups can be thought of as substituent groups attached to a hydrocarbon. The elements in group 7A of the periodic table—fluorine, chlorine, bromine, and iodine—are the halogens. Any organic compound that contains a halogen substituent is called a **halocarbon**. If you replace any of the hydrogen atoms in an alkane with a halogen atom, you form an alkyl halide. An **alkyl halide** is an organic compound containing a halogen atom covalently bonded to an aliphatic carbon atom. The first four

halogens—fluorine, chlorine, bromine, and iodine—are found in many organic compounds. For example, chloromethane is the alkyl halide formed when a chlorine atom replaces one of methane's four carbon atoms.

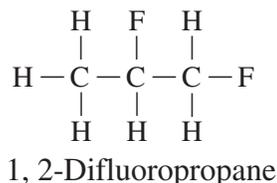
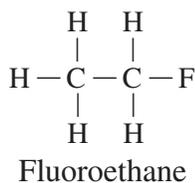


An **aryl halide** is an organic compound containing a halogen atom bonded to a benzene ring or other aromatic group. The structural formula for an aryl halide is created by first drawing the aromatic structure and then replacing its hydrogen atoms with the halogen atoms specified.

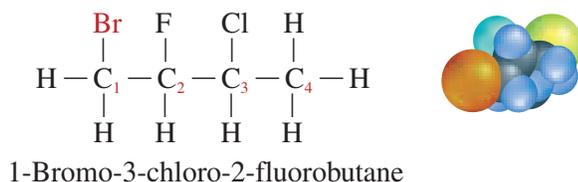


Naming Halocarbons

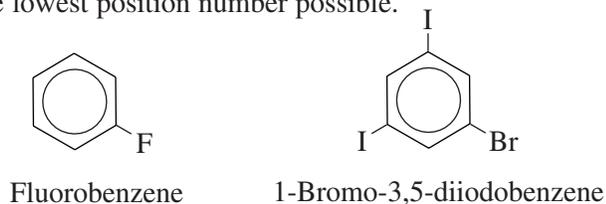
Organic molecules containing functional groups are given IUPAC names based on their main-chain alkane structures. For the alkyl halides, a prefix indicates which halogen is present. The prefixes are formed by changing the *-ine* at the end of each halogen name to *-o*. Thus, the prefix for fluorine is *fluoro-*, chlorine is *chloro-*, bromine is *bromo-*, and iodine is *iodo-*.



If more than one kind of halogen atom is present in the same molecule, the atoms are listed alphabetically in the name. The chain also must be numbered in a way that gives the lowest position number to the substituent that comes first in the alphabet. Note how the following alkyl halide is named.



Note that the benzene ring in an aryl halide is numbered to give each substituent the lowest position number possible.



CONNECTION

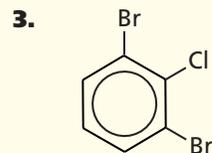
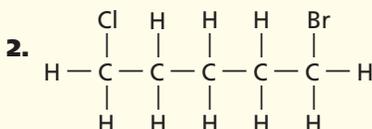
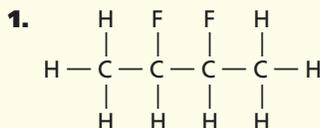
Naphthalene, anthracene, and similar hydrocarbons are termed polycyclic aromatic hydrocarbons (PAHs) because they are composed of multiple aromatic rings. PAHs have been found in meteorites and identified in the material surrounding dying stars. Scientists have mixed PAHs with water ice in a vacuum at -260°C to simulate the conditions found in interstellar clouds. To simulate radiation emitted by nearby stars, they shined ultraviolet light on the mixture. About ten percent of the PAHs were converted to alcohols, ketones, and esters—molecules that can be used to form compounds that are important in biological systems.



For more practice with naming alkyl and aryl halides, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Name the alkyl or aryl halide whose structure is shown.



Properties and Uses of Halocarbons

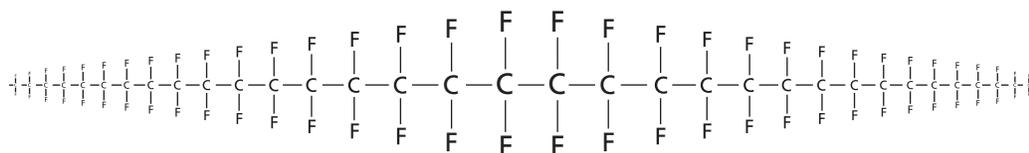
It is easiest to talk about properties of organic compounds containing functional groups by comparing those compounds with alkanes, whose properties were discussed in Chapter 22. **Table 23-2** lists some of the physical properties of certain alkanes and alkyl halides.

Table 23-2

A Comparison of Alkyl Halides with Their Parent Alkanes			
Structure	Name	Boiling point (°C)	Density (g/mL) in liquid state
CH ₄	methane	-162	0.423 at -162°C (boiling point)
CH ₃ Cl	chloromethane	-24	0.911 at 25°C (under pressure)
CH ₃ CH ₂ CH ₂ CH ₂ CH ₃	pentane	36	0.626
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ F	1-fluoropentane	62.8	0.791
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ Cl	1-chloropentane	108 <i>Increases</i>	0.882 <i>Increases</i>
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ Br	1-bromopentane	130	1.218
CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ I	1-iodopentane	155	1.516

Note that each alkyl chloride has a higher boiling point and a higher density than the alkane with the same number of carbon atoms. Note also that the boiling points and densities increase as the halogen changes from fluorine to chlorine, bromine, and iodine. This trend occurs primarily because the halogens from fluorine to iodine have increasing numbers of electrons that lie farther from the halogen nucleus. These electrons shift position easily and, as a result, the halogen-substituted hydrocarbons have an increasing tendency to form temporary dipoles. Because the dipoles attract each other, the energy needed to separate the molecules also increases. Thus, the boiling points of halogen-substituted alkanes increase as the size of the halogen atom increases.

Alkyl halides are used as solvents and cleaning agents because they readily dissolve nonpolar molecules such as greases. Teflon (polytetrafluoroethylene) is a plastic made from gaseous tetrafluoroethylene. Another plastic



Teflon

The chain units extend for hundreds of carbon atoms

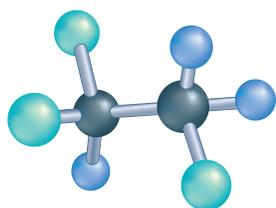
Figure 23-2

Teflon increases a carpet's resistance to staining and makes it easier to clean.

commonly called *vinyl* is polyvinyl chloride (PVC). It can be manufactured soft or hard, as thin sheets or molded into objects. Organic halides are seldom found in nature, although human thyroid hormones are organic iodides. One example of an organic halide is shown in **Figure 23-2**.

Alkyl halides are widely used as refrigerants. Until the late 1980s, alkyl halides called chlorofluorocarbons (CFCs) were widely used in refrigerators and air-conditioning systems. Because of their potential to damage Earth's ozone layer, CFCs have been replaced by HFCs, hydrofluorocarbons, which contain only hydrogen and fluorine atoms bonded to carbon. One of the more common HFCs is 1,1,2-trifluoroethane, also called R134a.

Halogen atoms bonded to carbon atoms are much more reactive than the hydrogen atoms they replace. For this reason, alkyl halides are often used as starting materials in the chemical industry.

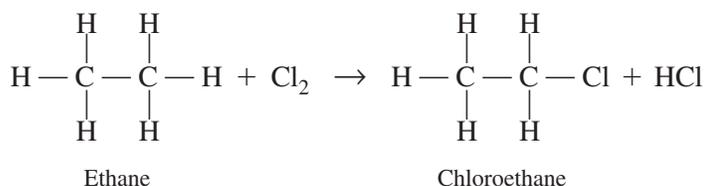


1,1,2-Trifluoroethane
HFC R134a

Substitution Reactions

Where does the immense variety of organic compounds come from? Amazingly enough, the ultimate source of nearly all synthetic organic compounds is petroleum. Petroleum is a fossil fuel that consists almost entirely of hydrocarbons, especially alkanes. How can alkanes be converted into compounds as different as alkyl halides, alcohols, and amines?

One way is to introduce a functional group through substitution. A **substitution reaction** is one in which one atom or a group of atoms in a molecule is replaced by another atom or group of atoms. With alkanes, hydrogen atoms may be replaced by atoms of halogens, typically chlorine or bromine, in a process called **halogenation**. One example of a halogenation reaction is the substitution of a chlorine atom for one of ethane's hydrogen atoms.



A substitution reaction

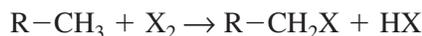




Figure 23-3

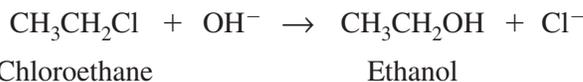
Moth balls are used in some closets to protect woolens from damage. 1,4-Dichlorobenzene (*p*-dichlorobenzene) is the active ingredient in some brands of moth balls.

Equations for organic reactions are sometimes shown in generic form. The following equation represents the halogenation of alkanes written in generic form.

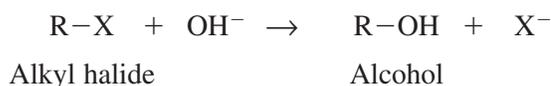


In this reaction, X can be fluorine, chlorine, or bromine, but not iodine. Iodine does not react well with alkanes. A common use of a halogenated hydrocarbon is shown in **Figure 23-3**.

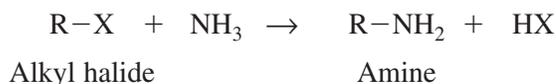
Once an alkane has been halogenated, the resulting alkyl halide can undergo other kinds of substitution reactions in which the halogen atom is replaced by another atom or group of atoms. For example, reacting an alkyl halide with an aqueous alkali results in the replacement of the halogen atom by an $-\text{OH}$ group, forming an alcohol.



In generic form, the reaction is as follows.



Reacting an alkyl halide with ammonia (NH_3) replaces the halogen atom with an amino group ($-\text{NH}_2$), forming an alkyl amine, as shown in the following equation.

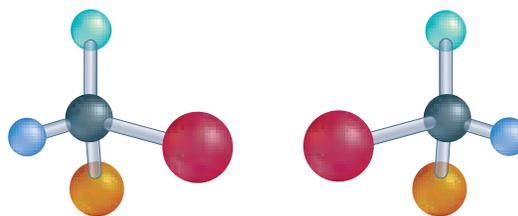


You will learn more about reactions and properties of amines in the next section.

Section 23.1 Assessment

- Draw structures for the following molecules.
 - 2-chlorobutane
 - 1,3-difluorohexane
 - 1,1,1-trichloroethane
 - 4-bromo-1-chlorobenzene
- Name the functional group present in each of the following structures. Name the type of organic compound each substance represents.
 - $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$
 - $\text{CH}_3\text{CH}_2\text{F}$
 - $\text{CH}_3\text{CH}_2\text{NH}_2$
 - $$\begin{array}{c} \text{O} \\ || \\ \text{CH}_3\text{C} - \text{OH} \end{array}$$
- How would you expect the boiling points of propane and 1-chloropropane to compare? Explain your answer.

- Thinking Critically** Examine the pair of structures shown below and decide whether it represents a pair of optical isomers. Explain your answer.



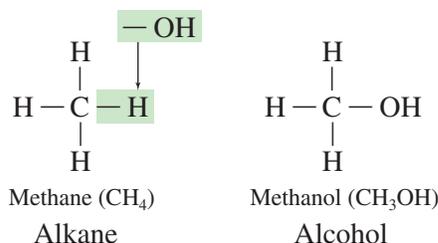
- Applying Concepts** Place the following substances in order of increasing boiling point. Do not look up boiling points, but use what you learned in this section to suggest the correct order.

2-chloropentane	1-bromohexane	butane
2-iodopentane	3-methylpentane	

In Section 23.1, you learned that hydrogen atoms bonded to carbon atoms in hydrocarbons can be replaced by halogen atoms. Many other kinds of atoms or groups of atoms also can bond to carbon in the place of hydrogen atoms. In addition to structural variations, replacement of hydrogen by other elements is a reason that such a wide variety of organic compounds is possible.

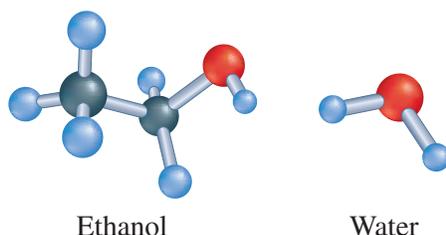
Alcohols

Many organic compounds contain oxygen atoms bonded to carbon atoms. Because an oxygen atom has six valence electrons, it commonly forms two covalent bonds to gain a stable octet. An oxygen atom can form a double bond with a carbon atom, replacing two hydrogen atoms, or it can form one single bond with a carbon atom and another single bond with another atom, such as hydrogen. An oxygen-hydrogen group covalently bonded to a carbon atom is called a **hydroxyl group** ($-\text{OH}$). An organic compound in which a hydroxyl group replaces a hydrogen atom of a hydrocarbon is called an **alcohol**. The general formula for an alcohol is ROH . The following diagram illustrates the relationship of the simplest alkane, methane, to the simplest alcohol, methanol.



Ethanol, a two-carbon alcohol, and carbon dioxide are produced by yeasts when they ferment sugars, such as those in grapes and bread dough. Ethanol is found in alcoholic beverages and medicinal products. Because it is an effective antiseptic, ethanol may be used to swab skin before an injection is given, as shown in **Figure 23-4**. It also is a gasoline additive and an important starting material for the synthesis of more complex organic compounds.

Models of an ethanol molecule and a water molecule are shown in the following diagram.



As you compare the models, you can see that the covalent bonds from oxygen are at approximately the same angle as they are in water. Therefore, the hydroxyl groups of alcohol molecules are moderately polar, as with water, and are able to form hydrogen bonds with the hydroxyl groups of other alcohol molecules. As a result, alcohols have much higher boiling points than

Objectives

- **Identify** the functional groups that characterize alcohols, ethers, and amines.
- **Draw** the structures of alcohols, ethers, and amines.
- **Discuss** the properties and uses of alcohols, ethers, and amines.

Vocabulary

hydroxyl group
alcohol
denatured alcohol
ether
amine

Figure 23-4

Ethanol is an effective antiseptic used to sterilize skin before an injection. It also is found in antiseptic hand cleaners.



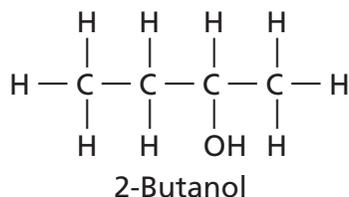


Figure 23-5

Because it evaporates more slowly than smaller alcohol molecules, 2-butanol is used as a solvent in some stains and varnishes. Draw the structures of two other alcohols that have the same molecular formula as 2-butanol, $\text{C}_4\text{H}_{10}\text{O}$.

hydrocarbons of similar shape and size. Also, because of polarity and hydrogen bonding, ethanol is completely miscible with water. In fact, once they are mixed, it is difficult to separate water and ethanol completely. Distillation is used to remove ethanol from water, but even after that process is complete, about 5% water remains in the ethanol-water mixture.

On the shelves of drugstores you can find bottles of ethanol labeled denatured alcohol. **Denatured alcohol** is ethanol to which small amounts of noxious materials such as aviation gasoline or other organic solvents have been added. Ethanol is denatured in order to make it unfit to drink. Because of their polar hydroxyl groups, alcohols make good solvents for other polar organic substances. For example, methanol, the smallest alcohol, is a common industrial solvent found in some paint strippers, and 2-butanol is found in some stains and varnishes, as shown in **Figure 23-5**. Perform the **CHEMLAB** at the end of this chapter to learn about some other properties of small-chain alcohols.

Note that the names of alcohols are based on alkane names, like the names of alkyl halides. For example, CH_4 is methane and CH_3OH is methanol; CH_3CH_3 is ethane and $\text{CH}_3\text{CH}_2\text{OH}$ is ethanol. When naming a simple alcohol based on an alkane carbon chain, the IUPAC rules call for naming the parent carbon chain or ring first and then changing the $-e$ at the end of the name to $-ol$ to indicate the presence of a hydroxyl group. In alcohols of three or more carbon atoms, the hydroxyl group can be at two or more positions. To indicate the position, a prefix consisting of a number followed by a dash is added. For example, examine the names and structures of two isomers of butanol shown below. Explain why the names *3-butanol* and *4-butanol* cannot represent real substances.

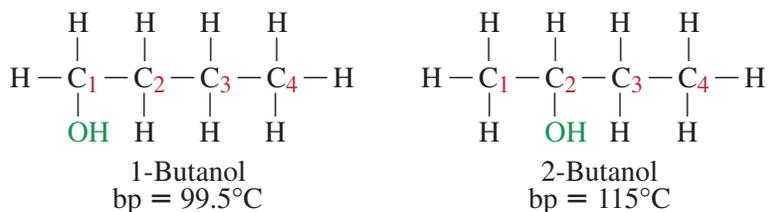
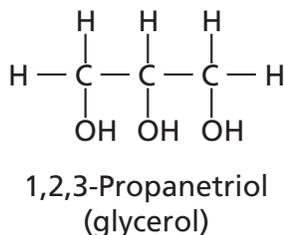
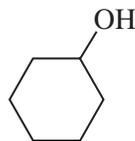


Figure 23-6

Glycerol is a liquid often used in automobile antifreeze and airplane deicing fluid.



Now look at the following structural formula. What is the name of this compound?

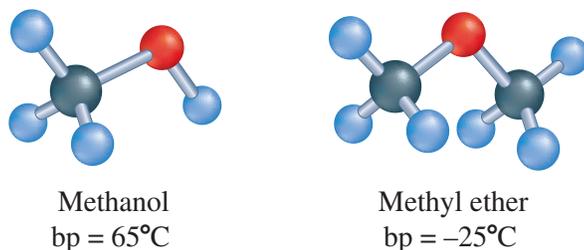


You are correct if you said *cyclohexanol*. The compound's ring structure contains six carbons with only single bonds so you know that the parent hydrocarbon is cyclohexane. Because an $-\text{OH}$ group is bonded to a carbon, it is an alcohol and the name will end in $-ol$. No number is necessary because all carbons in the ring are equivalent. Cyclohexanol is a poisonous compound used as a solvent for certain plastics and in the manufacture of insecticides.

A carbon chain also can have more than one hydroxyl group. To name these compounds, prefixes such as *di-*, *tri-*, and *tetra-* are used before the $-ol$ to indicate the number of hydroxyl groups present. The full alkane name, including $-ane$, is used before the prefix as in the following example. 1,2,3-propanetriol, commonly called glycerol, is another alcohol containing more than one hydroxyl group, **Figure 23-6**.

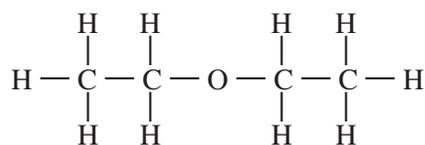
Ethers

Ethers are another group of organic compounds in which oxygen is bonded to carbon. An **ether** is an organic compound containing an oxygen atom bonded to two carbon atoms. Ethers have the general formula ROR'. The simplest ether is one in which oxygen is bonded to two methyl groups. Note the relationship between methanol and methyl ether in the following diagram.



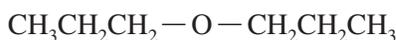
Because ethers have no hydrogen atoms bonded to the oxygen atom, their molecules cannot form hydrogen bonds with each other. Therefore, ethers generally are more volatile and have much lower boiling points than alcohols of similar size and mass. Ethers are much less soluble in water than alcohols because they have no hydrogen to donate to a hydrogen bond. However, the oxygen atom can act as a receptor for the hydrogen atoms of water molecules.

The term *ether* was first used in chemistry as a name for ethyl ether, the volatile, highly flammable substance that was commonly used as an anesthetic in surgery from 1842 into the twentieth century, **Figure 23-7**. As time passed, the term *ether* was applied to other organic substances having two hydrocarbon chains attached to the same oxygen atom.

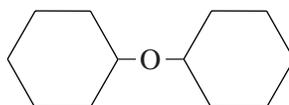


Ethyl ether

Name ethers that have two identical alkyl chains bonded to oxygen by first naming the alkyl group and then adding the word *ether*. Here are the names and structures of two of these symmetrical ethers.

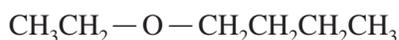


Propyl ether



Cyclohexyl ether

If the two alkyl groups are different, they are listed in alphabetical order and then followed by the word ether. Such ethers are asymmetrical, or uneven, in appearance.



Butylethyl ether



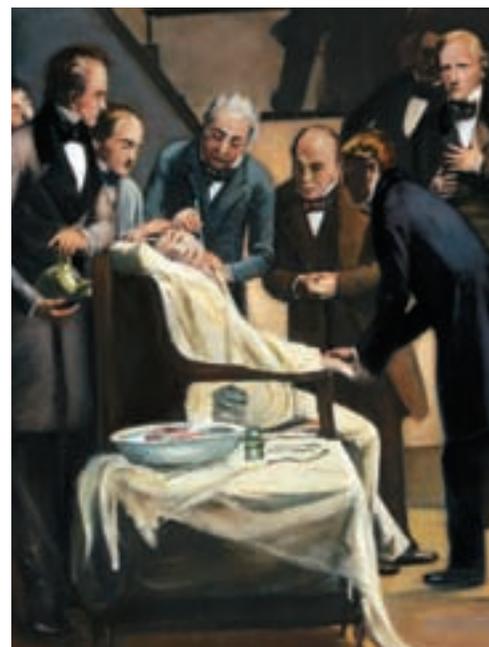
Ethylmethyl ether

Amines

Another class of organic compounds contains nitrogen. **Amines** contain nitrogen atoms bonded to carbon atoms in aliphatic chains or aromatic rings and have the general formula RNH₂.

Figure 23-7

An apparatus similar to the one held by the man on the left was used to administer ether to patients in the 1840s. Dr. Crawford W. Long of Georgia discovered the anesthetic properties of ethyl ether in 1842.



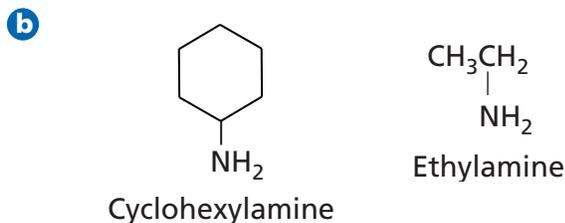
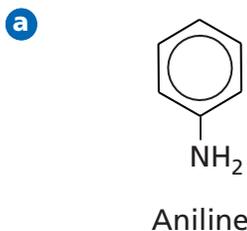


Figure 23-8

a Aniline is an important industrial substance, especially in the production of dyes with deep shades of color. **b** Cyclohexylamine and ethylamine are important in the manufacture of pesticides, plastics, pharmaceuticals, and the rubber used to make the tires for this race car.

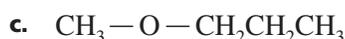
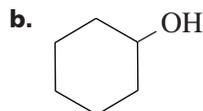
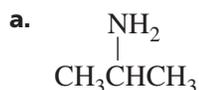
When naming amines, the $-\text{NH}_2$ (amino) group is indicated by the suffix *-amine*. When necessary, the position of the amino group is designated by a number. If more than one amino group is present, the prefixes *di-*, *tri-*, *tetra-*, and so on are used to indicate the number of groups.



Examples of amines used in industry are shown in **Figure 23-8**. All volatile amines have odors that humans find offensive, and amines are responsible for many of the odors characteristic of dead, decaying organisms.

Section 23.2 Assessment

- 9.** Draw structures for the following alcohols:
- 1-propanol
 - 2-propanol
 - 2-methyl-2-butanol
 - 1,3-cyclopentanediol
- 10.** Draw structures for the following molecules:
- propyl ether
 - ethylpropyl ether
 - 1,2-propanediamine
 - cyclobutylamine
- 11.** Identify the functional group present in each of the following structures. Name the substance represented by each structure.



- 12. Thinking Critically** Which of the following compounds would you expect to be more soluble in water? Explain your reasoning.



- 13. Applying Concepts** Alcohols and amines are used as starting materials in reactions that produce useful substances such as pharmaceuticals, plastics, and synthetic fibers, as well as other industrial chemicals. What properties of these compounds make them more useful than hydrocarbons for this purpose?

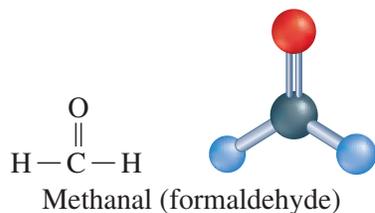
You have now learned that in the organic compounds known as alcohols and ethers, oxygen is bonded to two different atoms. In other kinds of organic compounds, an oxygen atom is double-bonded to a carbon atom.

Organic Compounds Containing the Carbonyl Group

The arrangement in which an oxygen atom is double-bonded to a carbon atom is called a **carbonyl group**. This group, which can be represented as shown below, is the functional group in organic compounds known as aldehydes and ketones.

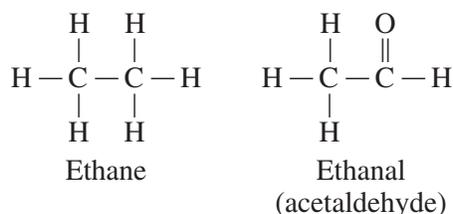


Aldehydes An **aldehyde** is an organic compound in which a carbonyl group located at the end of a carbon chain is bonded to a carbon atom on one side and a hydrogen atom on the other. Aldehydes have the general formula $*\text{CHO}$, where $*$ represents an alkyl group or a hydrogen atom.



Aldehydes are formally named by changing the final $-e$ of the name of the alkane with the same number of carbon atoms to the suffix $-al$. Thus, the formal name of the compound represented above is methanal, based on the one-carbon alkane methane. Methanal is commonly called formaldehyde. A water solution of formaldehyde was used in the past to preserve biological specimens. However, formaldehyde's use has been restricted in recent years because evidence shows it may cause cancer. Industrially, large quantities of formaldehyde are reacted with urea to manufacture a type of grease-resistant, hard plastic used to make buttons, appliance and automotive parts, and electrical outlets, as well as the glue that holds the layers of plywood together.

Note the relationship between the structures and names of ethane and ethanal in the following diagram.



Objectives

- **Draw and identify** the structures of carbonyl compounds including aldehydes, ketones, carboxylic acids, esters, and amides.
- **Discuss** the properties and uses of compounds containing the carbonyl group.

Vocabulary

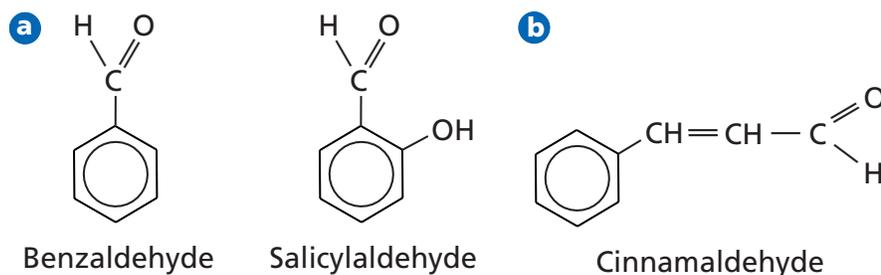
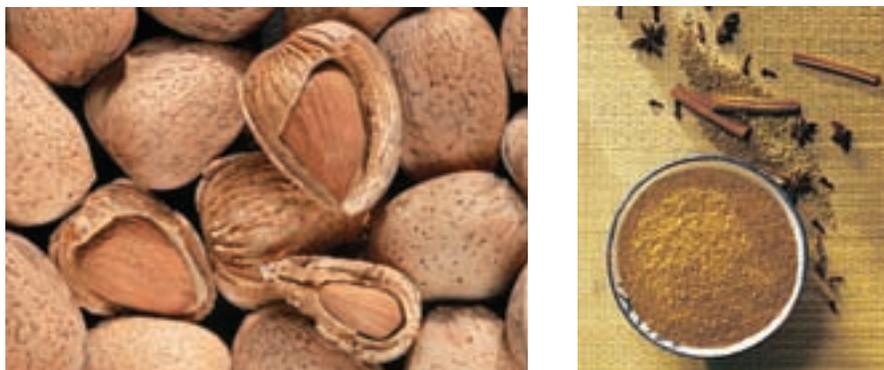
carbonyl group
aldehyde
ketone
carboxylic acid
carboxyl group
ester
amide
condensation reaction

Ethanal also has the common name *acetaldehyde*. Scientists often use the common names of organic compounds because they are so familiar to chemists. Because the carbonyl group in an aldehyde always occurs at the end of a carbon chain, no numbers are used in the name unless branches or additional functional groups are present.

Many aldehydes have characteristic odors and flavors. **Figure 23-9** shows examples of naturally occurring aldehydes.

Figure 23-9

a Benzaldehyde and salicylaldehyde are two components of the flavor of almonds. **b** The aroma and flavor of cinnamon are produced largely by cinnamaldehyde. Cinnamon is the ground bark of a tropical tree.



An aldehyde molecule contains a polar, reactive structure. However, like ethers, aldehyde molecules cannot form hydrogen bonds among themselves because the molecules have no hydrogen atoms bonded to an oxygen atom. Therefore, aldehydes have lower boiling points than alcohols with the same number of carbon atoms. Water molecules can form hydrogen bonds with the oxygen atom of aldehydes, so aldehydes are more soluble in water than alkanes but not as soluble as alcohols or amines.

Ketones A carbonyl group also can be located within a carbon chain rather than at the end. A **ketone** is an organic compound in which the carbon of the carbonyl group is bonded to two other carbon atoms. Ketones have the general formula



The carbon atoms on either side of the carbonyl group are bonded to other atoms. The simplest ketone has only hydrogen atoms bonded to the side carbons, as shown in the following diagram. The common name of this ketone is acetone.

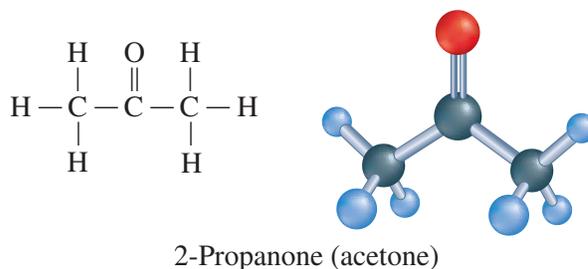
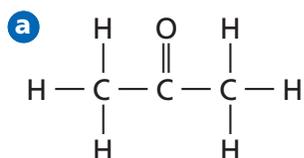


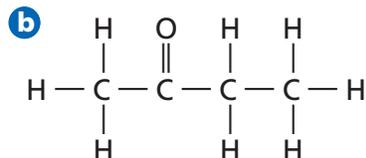


Figure 23-10

- a** The solvent 2-propanone (acetone) is used in fingernail polish; it can dissolve and remove the polish as well.
- b** Although the solvent 2-butanone is not familiar to most people, it is a popular solvent in industries that put coatings on other materials.



2-Propanone
(acetone)



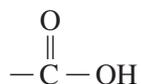
2-Butanone
(methyl ethyl ketone)

Ketones are formally named by changing the *-e* at the end of the alkane name to *-one*, and including a number before the name to indicate the position of the ketone group. In the previous example, the alkane name propane is changed to *propanone*. The carbonyl group can be located only in the center, but the prefix *2-* is usually added to the name for clarity. **Figure 23-10** shows representative ketones and their uses.

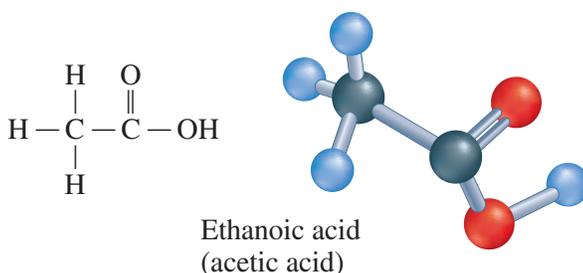
Ketones and aldehydes share many chemical and physical properties because their structures are so similar. Ketones are polar molecules and are less reactive than aldehydes. For this reason, ketones are popular solvents for other moderately polar substances, including waxes, plastics, paints, lacquers, varnishes, and glues. Like aldehydes, ketone molecules cannot form hydrogen bonds with each other but can form hydrogen bonds with water molecules. Therefore, ketones are somewhat soluble in water. Acetone is completely miscible with water.

Carboxylic Acids

A **carboxylic acid** is an organic compound that has a carboxyl group. A **carboxyl group** consists of a carbonyl group bonded to a hydroxyl group. Thus, carboxylic acids have the general formula



The following diagram shows the structure of a carboxylic acid familiar to you—acetic acid, the acid found in vinegar.



Although many carboxylic acids have common names, the formal name is formed by changing the *-ane* of the parent alkane to *-anoic acid*. Thus, the formal name of acetic acid is ethanoic acid.

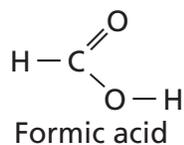
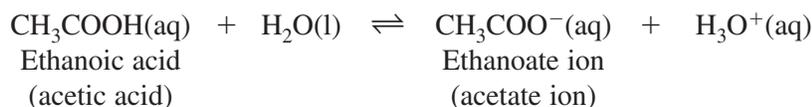


Figure 23-11

Formic acid is the simplest carboxylic acid. Stinging ants produce formic acid as a defense mechanism.

A carboxyl group is usually represented in condensed form by writing $-\text{COOH}$. For example, ethanoic acid can be written as CH_3COOH . The simplest carboxylic acid consists of a carboxyl group bonded to a single hydrogen atom, HCOOH . Its formal name is methanoic acid, but it is more commonly known as formic acid. See **Figure 23-11**.

Carboxylic acids are polar and reactive. Those that dissolve in water ionize weakly to produce hydronium ions and the anion of the acid in equilibrium with water and the un-ionized acid. The ionization of ethanoic acid is an example.



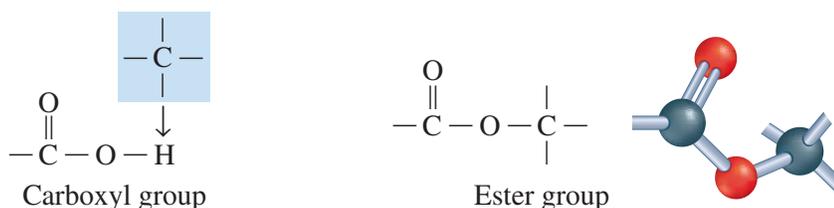
Carboxylic acids can ionize in water solution because the two oxygen atoms are highly electronegative and attract electrons away from the hydrogen atom in the $-\text{OH}$ group. As a result, the hydrogen proton can transfer to another atom that has a pair of electrons not involved in bonding, such as the oxygen atom of a water molecule. Because they ionize in water, soluble carboxylic acids turn blue litmus paper red and have a sour taste.

Some important carboxylic acids, such as oxalic acid and adipic acid, have two or more carboxyl groups. An acid with two carboxyl groups is called a dicarboxylic acid. Others have additional functional groups such as hydroxyl groups, as shown in **Figure 23-12**. Typically, these acids are more soluble in water and often more acidic than acids with only a carboxyl group.

Organic Compounds Derived From Carboxylic Acids

Several classes of organic compounds have structures in which the hydrogen or the hydroxyl group of a carboxylic acid is replaced by a different atom or group of atoms. The two most common classes are esters and amides.

Esters An **ester** is any organic compound with a carboxyl group in which the hydrogen of the hydroxyl group has been replaced by an alkyl group, producing the following arrangement



The name of an ester is formed by writing the name of the alkyl group followed by the name of the acid with the *-ic acid* ending replaced by *-ate*, as illustrated by the example shown below. Note how the name *propyl ethanoate*

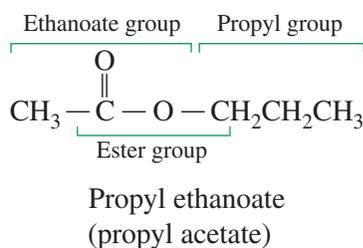
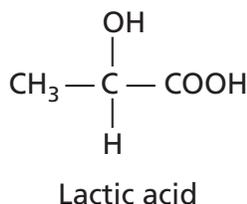
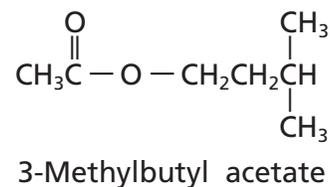
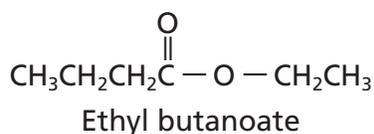
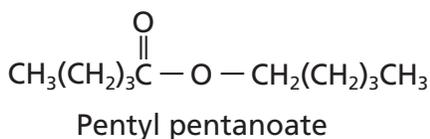


Figure 23-12

Lactic acid produces the tangy taste of buttermilk. It also accumulates in the muscles during heavy exercise.





results from the structural formula. The name shown in parentheses is based on the name *acetic acid*, the common name for ethanoic acid.

Esters are polar molecules and many are volatile and sweet-smelling. Many kinds of esters are found in the natural fragrances and flavors of flowers and fruits. Natural flavors such as apple or banana result from mixtures of many different organic molecules, including esters. See **Figure 23-13**. But some of these flavors can be imitated by a single ester structure. Consequently, esters are manufactured for use as flavors in many foods and beverages and as fragrances in candles, perfumes, and other scented items. You will prepare an ester in the following **miniLAB**.

Figure 23-13

Esters are responsible for the flavors and aromas of many fruits. Pentyl pentanoate smells like ripe apples. Ethyl butanoate has the aroma of pineapples, and 3-methylbutyl acetate smells like bananas although it imparts a pear flavor to foods. Most natural aromas and flavors are mixtures of esters, aldehydes, and alcohols.

miniLAB

Making an Ester

Observing and Inferring Flowers and fruits have pleasant odors partly because they contain substances called esters. Companies make blends of synthetic esters to mimic the flavors and fragrances of esters found in nature. In this experiment, you will make an ester that has a familiar smell.

Materials salicylic acid, methanol, distilled water, 10-mL graduated cylinder, Beral pipette, 250-mL beaker, concentrated sulfuric acid, top or bottom of a petri dish, cotton ball, small test tube, balance, weighing paper, hot plate, test-tube holder

Procedure



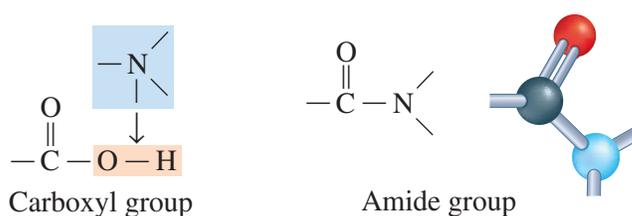
1. Prepare a hot-water bath by pouring 150 mL of tap water into a 250-mL beaker. Place the beaker on a hot plate set at medium.

2. Place 1.5 g of salicylic acid in a small test tube and add 3 mL of distilled water. Then add 3 mL of methanol and 3 drops of concentrated sulfuric acid to the test tube. **CAUTION: Sulfuric acid is corrosive. Handle with care.**
3. When the water is hot but not boiling, place the test tube in the bath for 5 minutes.
4. Place the cotton ball in the petri dish half. Pour the contents of the test tube onto the cotton ball. Record your observation of the odor of the product.

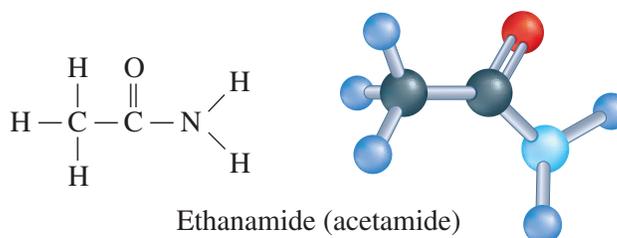
Analysis

1. The ester you produced has the common name oil of wintergreen. Write a chemical equation using names and structural formulas for the reaction that produced the ester.
2. What are the advantages and disadvantages of using synthetic esters in consumer products as compared to using natural esters?
3. Name some products that you think could contain the ester you made in this experiment.

Amides An **amide** is an organic compound in which the $-OH$ group of a carboxylic acid is replaced by a nitrogen atom bonded to other atoms. The general structure of an amide is shown below.



Amides are named by writing the name of the alkane with the same number of carbon atoms, and then replacing the final $-e$ with $-amide$. Thus, the amide shown below is called ethanamide, but it also may be named acetamide from the common name, acetic acid.



The amide functional group is found repeated many times in natural proteins and some synthetic materials. For example, you may have used a nonaspirin pain reliever containing acetaminophen. In the acetaminophen structure shown below, you can see that the amide ($-NH-$) group connects a carbonyl group and an aromatic group.

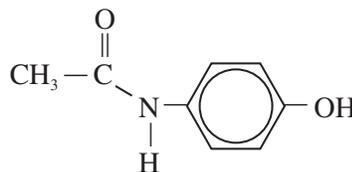


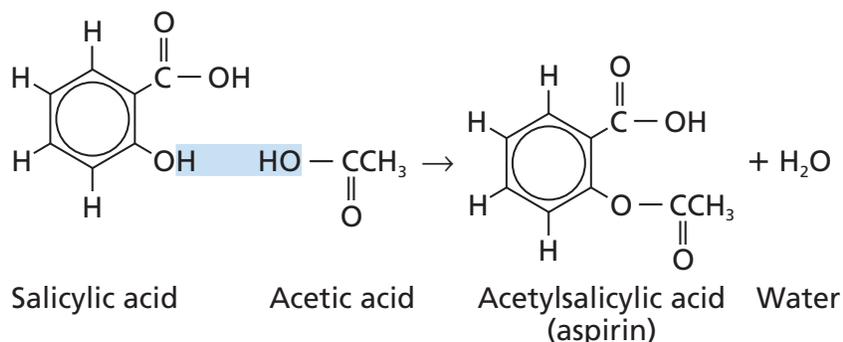
Figure 23-14

To synthesize aspirin, two organic molecules are combined in a condensation reaction to form a larger molecule.



Condensation Reactions

Many laboratory syntheses and industrial processes involve the reaction of two organic reactants to form a larger organic product, such as the aspirin shown in **Figure 23-14**. This type of reaction is known as a condensation reaction.



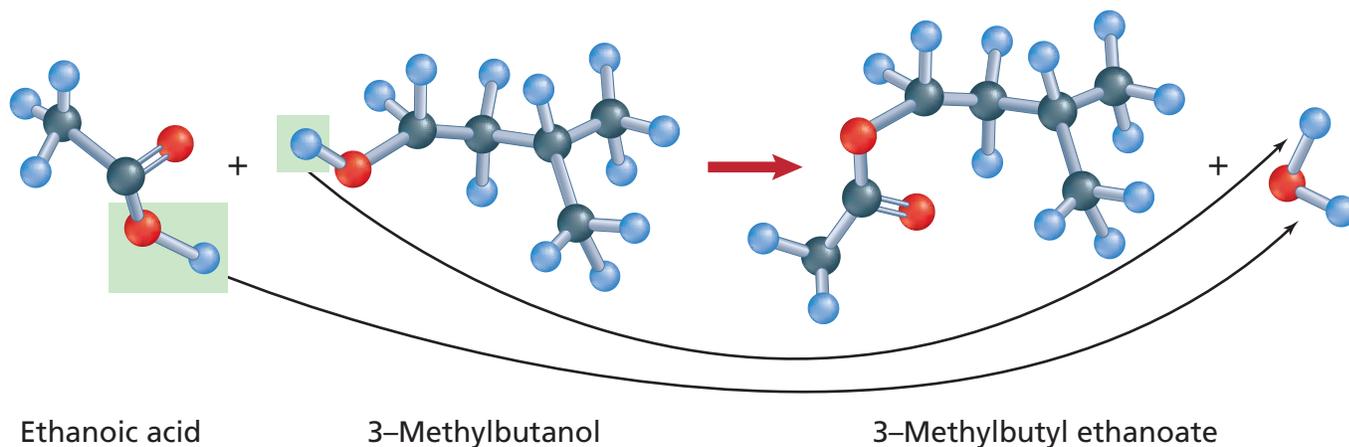
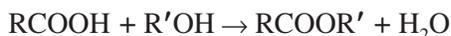


Figure 23-15

When ethanoic acid and 3-methylbutanol undergo a condensation reaction, the ester 3-methylbutyl ethanoate is formed. The water molecule eliminated is made up of the -OH group from the acid and the H- from the alcohol.

In a **condensation reaction**, two smaller organic molecules combine to form a more complex molecule, accompanied by the loss of a small molecule such as water. Typically, the molecule lost is formed from one particle from each of the reactant molecules. In essence, a condensation reaction is an elimination reaction in which a bond is formed between two atoms not previously bonded to each other.

The most common condensation reactions involve the combining of carboxylic acids with other organic molecules. A common way to synthesize an ester is by a condensation reaction between a carboxylic acid and an alcohol. Such a reaction can be represented by the general equation



For example, the reaction represented by the ball-and-stick models in **Figure 23-15** produces an ester that is a part of artificial banana flavoring.

Section 23.3 Assessment

- 14.** Draw the structure for a carbonyl group.
- 15.** Classify each of the following structures as one of the types of organic substances you have studied in this section.
- | | |
|--|--|
| <p>a.</p> $\text{CH}_3\text{CH}_2\text{CH}_2\overset{\text{O}}{\parallel}\text{CH}$ | <p>d.</p> $\text{CH}_3\text{CH}_2\text{-O-C(=O)-CH}_3$ |
| <p>b.</p> $\text{CH}_3\text{CH}_2\overset{\text{O}}{\parallel}\text{CCH}_3$ | <p>e.</p> $\text{CH}_3\text{CH}_2\text{CH}_2\overset{\text{O}}{\parallel}\text{C-NH}_2$ |
| <p>c.</p> $\text{CH}_3\text{CH}_2\text{CH}_2\overset{\text{O}}{\parallel}\text{C-OH}$ | <p>f.</p> |
- 16.** What are the products of a condensation reaction between a carboxylic acid and an alcohol?
- 17.** What features of the substances discussed in this section make many of them useful solvents? Explain how these features affect the properties of the molecules.
- 18. Thinking Critically** Suggest a reason for the observation that water-soluble organic compounds with carboxyl groups exhibit acidic properties in solutions, whereas similar compounds with aldehyde structures do not.
- 19. Formulating Models** You learned that the molecular formulas for alkanes follow the pattern $\text{C}_n\text{H}_{2n+2}$. Derive a general formula to represent an aldehyde, a ketone, and a carboxylic acid. Could you examine a molecular formula for one of these three types of compounds and determine which type the formula represents? Explain.

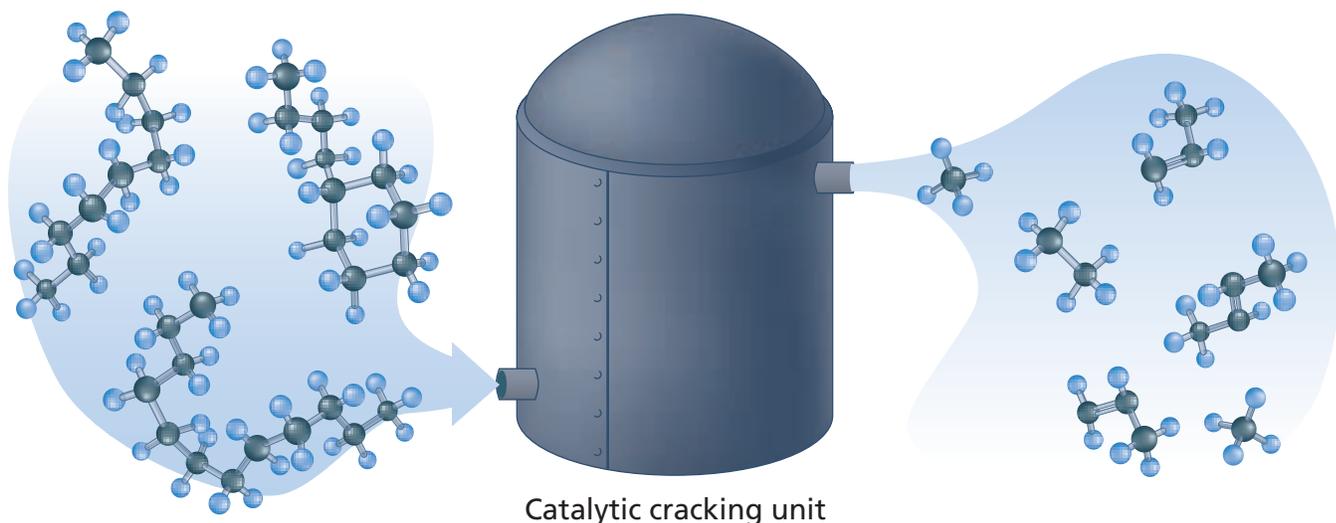


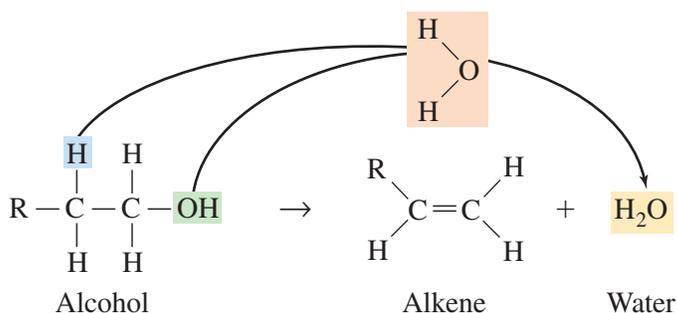
Figure 23-17

The process of catalytic cracking breaks long-chain alkanes into smaller alkanes and alkenes that are more valuable to industry.

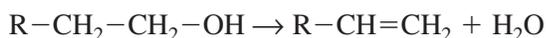
Alkyl halides can undergo elimination reactions to produce an alkene and a hydrogen halide, as shown here.



Likewise, alcohols also can undergo elimination reactions by losing a hydrogen atom and a hydroxyl group to form water. An elimination reaction in which the atoms removed form water is called a **dehydration reaction**. In the following dehydration reaction, the alcohol is broken down into an alkene and water.



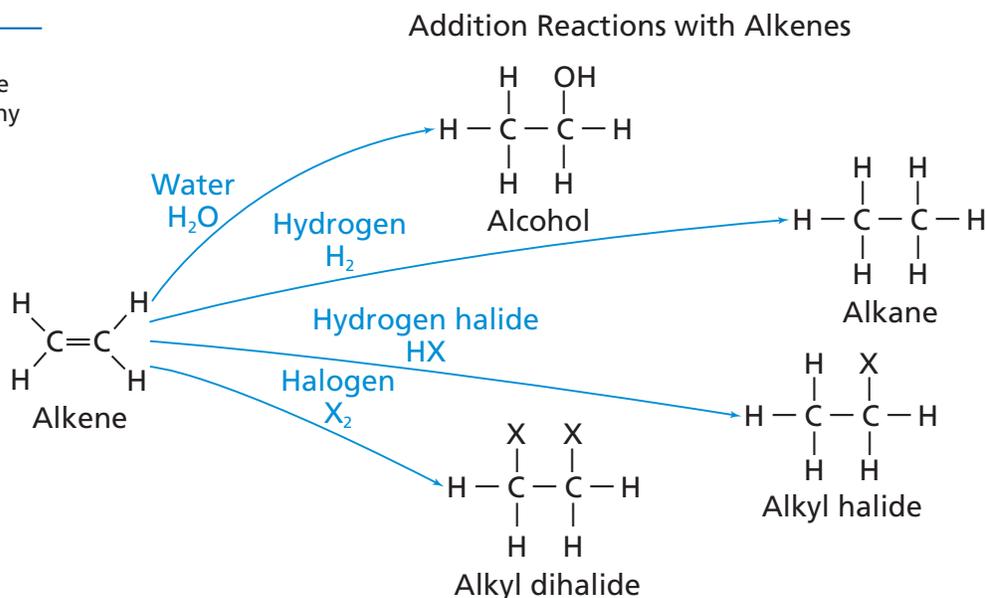
The generic form of this dehydration reaction can be written as follows.



Addition reactions Another type of organic reaction appears to be an elimination reaction in reverse. An **addition reaction** results when other atoms bond to each of two atoms bonded by double or triple covalent bonds. Addition reactions typically involve double-bonded carbon atoms in alkenes or triple-bonded carbon atoms in alkynes. Addition reactions occur because double and triple bonds have a rich concentration of electrons. Therefore, molecules and ions that attract electrons tend to form bonds that use some of the electrons from the multiple bonds. The most

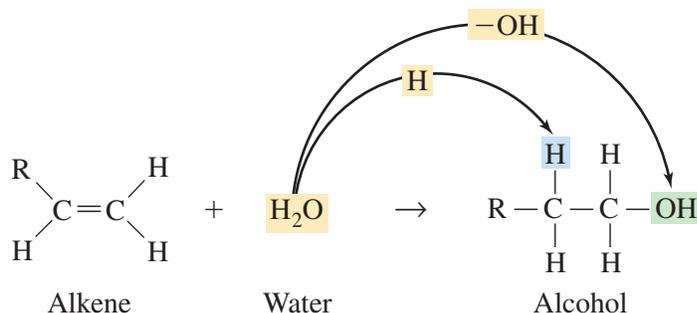
Figure 23-18

These examples are common addition reactions that can be carried out with alkenes. Many other addition reactions are possible.

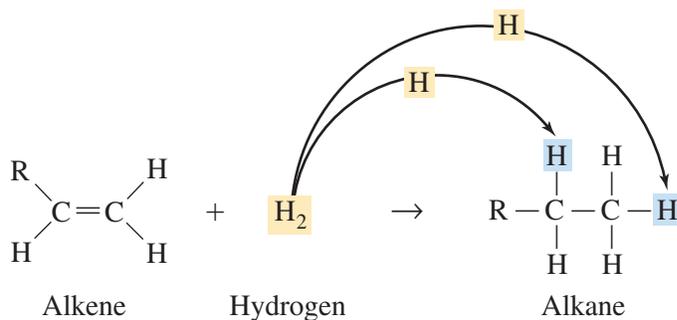


common addition reactions are those in which H_2O , H_2 , HX , or X_2 add to an alkene as shown in **Figure 23-18**.

A **hydration reaction** is an addition reaction in which a hydrogen atom and a hydroxyl group from a water molecule add to a double or triple bond. As shown in the following generic example, a hydration reaction is the opposite of a dehydration reaction.



A reaction that involves the addition of hydrogen to atoms in a double or triple bond is called a **hydrogenation reaction**. This reaction is the reverse of one of the reactions you studied earlier in this section. Which one? One molecule of H_2 reacts to fully hydrogenate each double bond in a molecule. When H_2 adds to the double bond of an alkene, the alkene is converted to an alkane.

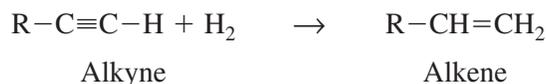


Catalysts are usually needed in the hydrogenation of alkenes because the reaction's activation energy is too large without them. Catalysts such as powdered platinum or palladium provide a surface that adsorbs the reactants and makes their electrons more available to bond to other atoms.

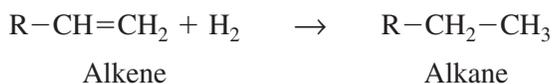
Hydrogenation reactions are commonly used to convert the liquid unsaturated fats found in oils from plants such as soybean, corn, and peanuts into saturated fats that are solid at room temperature. These hydrogenated fats are then used to make margarine and solid shortening. You will learn more about the composition of fats in Chapter 24.

Alkyne also may also be hydrogenated to produce alkenes or alkanes. Two molecules of H_2 must be added to each triple bond in order to convert an alkyne to an alkane, as shown in the following equations.

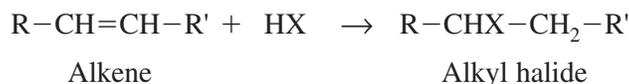
First H_2 molecule added



Second H_2 molecule added



The addition of hydrogen halides to alkenes obtained from petroleum or natural gas is an addition reaction useful to industry for the production of alkyl halides.



As you learned earlier, R' is used to represent a second alkyl group. Chemists use the symbols R' , R'' , R''' , and so on to represent different alkyl groups in organic molecules.

Do the **problem-solving LAB** below to learn how functional groups give organic compounds characteristic properties that may be used to identify a compound's general type.

problem-solving LAB

Categorizing Organic Compounds

Analyze and Conclude Functional groups give organic compounds distinct properties that may be used to identify the type of compound present. Suppose you examined the properties of several compounds and made the following observations.

Analysis

Compound 1 is a liquid that has a pungent odor. It is miscible with water, and the solution is a weak electrolyte.

Compound 2 is a crystalline solid with a melting point of 112°C and almost no odor. It is soluble in

water, and chemical analysis shows that it contains nitrogen in addition to carbon, hydrogen and oxygen.

Compound 3 is a liquid that has a strong aroma resembling apricots. When spilled on a wood tabletop, it damages the finish.

Compound 4 is an oily-looking liquid with a disagreeable odor similar to a combination of ammonia and dead fish. It is soluble in water.

Thinking Critically

Use what you have learned about the properties of organic compounds with functional groups to suggest the category to which each compound belongs.

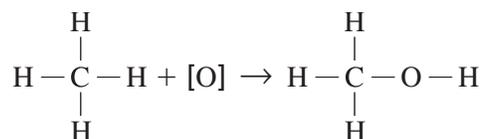




Figure 23-19

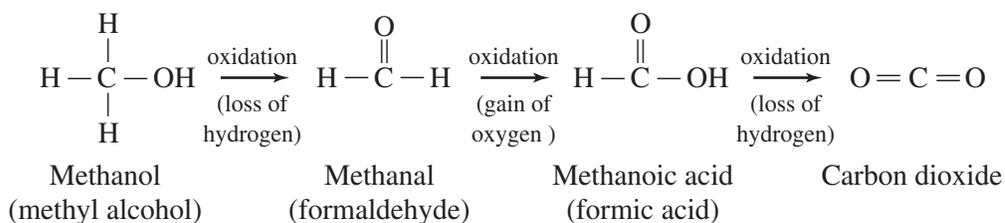
The oxidation of hydrocarbons provides the energy to run most of the vehicles on this crowded freeway.

Oxidation-reduction reactions Many organic compounds can be converted to other compounds by oxidation and reduction reactions. For example, suppose that you wish to convert methane, the main constituent of natural gas, to methanol, a common industrial solvent and raw material for making formaldehyde and methyl esters. The conversion of methane to methanol may be represented by the following equation, in which [O] represents oxygen from an agent such as copper(II) oxide, potassium dichromate, or sulfuric acid.



What happens to methane in this reaction? Before answering, it might be helpful to review the definitions of oxidation and reduction. Oxidation is the loss of electrons, and a substance is oxidized when it gains oxygen or loses hydrogen. Reduction is the gain of electrons, and a substance is reduced when it loses oxygen or gains hydrogen. Thus, methane is oxidized as it gains oxygen and is converted to methanol. Of course, every redox reaction involves both an oxidation and a reduction; however, organic redox reactions are described based on the change in the organic compound.

Oxidizing the methanol produced in the previous reaction is the first step in the sequence of reactions described by the following equations. For clarity, oxidizing agents are omitted.



Preparing an aldehyde by this method is not always a simple task because the oxidation may continue, forming the carboxylic acid. However, not all alcohols can be oxidized to aldehydes and, subsequently, carboxylic acids. To understand why, compare the oxidations of 1-propanol and 2-propanol shown in the following equations.

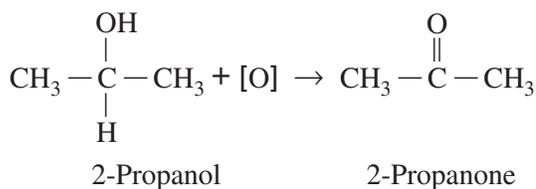
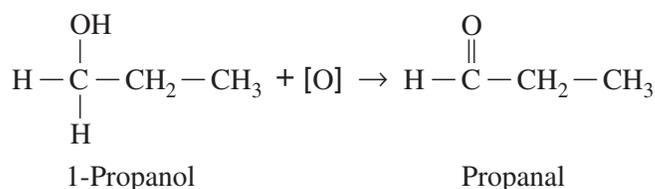


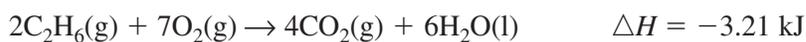


Figure 23-20

Organic oxidation-reduction reactions provide the energy that these students need to live and play active sports such as basketball.

Note that oxidizing 2-propanol yields a ketone, not an aldehyde. Unlike aldehydes, ketones resist further oxidation to carboxylic acids. Thus, while the propanal formed by oxidizing 1-propanol easily oxidizes to form propanoic acid, the 2-propanone formed by oxidizing 2-propanol does not react to form a carboxylic acid.

How important are organic oxidations and reductions? You've seen that oxidation and reduction reactions can change one functional group into another. That ability enables chemists to use organic redox reactions, in conjunction with the substitution and addition reactions you learned about earlier in the chapter, to synthesize a tremendous variety of useful products. On a personal note, living systems—including you—depend on the energy released by oxidation reactions. Of course, some of the most dramatic oxidation-reduction reactions are combustion reactions. All organic compounds that contain carbon and hydrogen burn in excess oxygen to produce carbon dioxide and water. For example, the highly exothermic combustion of ethane is described by the following thermochemical equation.

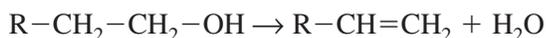


As you learned in Chapter 10, much of the world relies on the combustion of hydrocarbons as a primary source of energy. Our reliance on the energy from organic oxidation reactions is illustrated in **Figures 23-19** and **23-20**.

Predicting Products of Organic Reactions

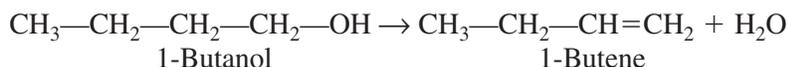
The generic equations representing the different types of organic reactions you have learned—substitution, elimination, addition, oxidation-reduction, and condensation—can be used to predict the products of other organic reactions of the same types. For example, suppose you were asked to predict the product of an elimination reaction in which 1-butanol is a reactant. You know that a common elimination reaction involving an alcohol is a dehydration reaction.

The generic equation for the dehydration of an alcohol is



To determine the actual product, first draw the structure of 1-butanol. Then use the generic equation as a model to see how 1-butanol would react.

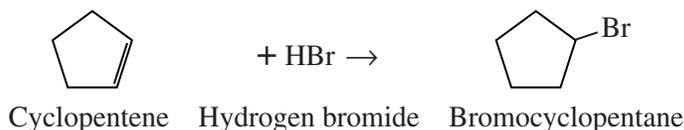
From the generic reaction, you can see that only the carbon bonded to the $-OH$ and the carbon next to it react. Finally, draw the structure of the likely products as shown in the following equation. Your result should be the following equation.



As another example, suppose that you wish to predict the product of the reaction between cyclopentene and hydrogen bromide. Recall that the generic equation for an addition reaction between an alkene and an alkyl halide is as follows.

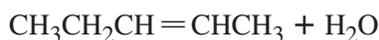
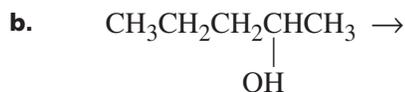
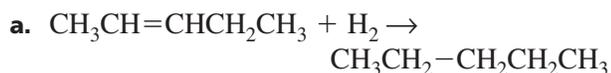


First, draw the structure for cyclopentene, the organic reactant, and add the formula for hydrogen bromide, the other reactant. From the generic equation, you can see that a hydrogen atom and a halide atom add across the double bond to form an alkyl halide. Finally, draw the formula for the likely product. If you are correct, you have written the following equation.



Section 23.4 Assessment

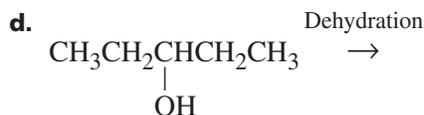
20. Classify each of the following reactions as either substitution, elimination, addition, or condensation.



21. Identify the type of organic reaction that would best accomplish each of the following conversions.

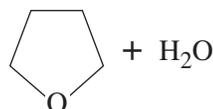
- alkyl halide \rightarrow alkene
- alkene \rightarrow alcohol
- alcohol + carboxylic acid \rightarrow ester

22. Complete each of the following equations by writing the condensed structural formula for the product that is most likely to form.



23. Thinking Critically Explain why the hydration reaction involving 1-butene may yield two distinct products whereas the hydration of 2-butene yields but one product.

24. Comparing and Contrasting Explain the difference between an elimination reaction and a condensation reaction. Which type is best represented by the following equation?



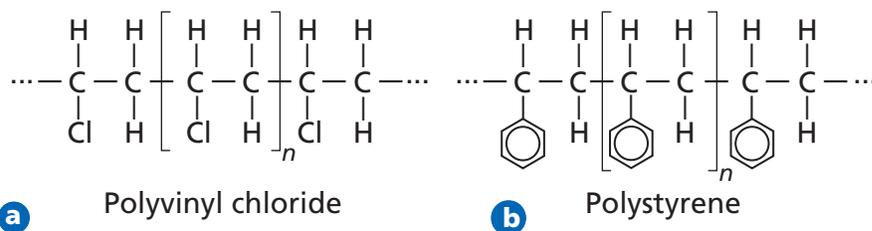
Think how different your life would be without plastic sandwich bags, plastic foam cups, nylon and polyester fabrics, vinyl siding on buildings, foam cushions, and a variety of other synthetic materials. Synthetic materials do not exist in nature; they are produced by the chemical industry.

The Age of Polymers

The structural formulas shown in **Figure 23-21** represent extremely long molecules having groups of atoms that repeat in a regular pattern. These molecules are examples of synthetic polymers. **Polymers** are large molecules consisting of many repeating structural units. The letter n represents the number of structural units in the polymer chain. Because polymer n values vary widely, molecular masses of polymers range from less than 10 000 amu to more than 1 000 000 amu. A typical Teflon chain has about 400 units, giving it a molecular mass of around 40 000 amu.

Before the development of synthetic polymers, people were limited to using natural substances such as stone, wood, metals, wool, and cotton. By the turn of the twentieth century, a few chemically treated natural polymers such as rubber and the first plastic, celluloid, had become available. Celluloid is made by treating cellulose from cotton or wood fiber with nitric acid.

The first synthetic polymer was Bakelite, synthesized in 1909. Bakelite is still used today in stove-top appliances because of its resistance to heat. Since 1909, hundreds of other synthetic polymers have been developed. In the future, people may refer to this time as the Age of Polymers.



Objectives

- **Describe** the relationship between a polymer and the monomers from which it forms.
- **Classify** polymerization reactions as addition or condensation.
- **Predict** polymer properties based on their molecular structures and the presence of functional groups.

Vocabulary

polymer
monomer
polymerization reaction
addition polymerization
condensation polymerization
plastic
thermoplastic
thermosetting

Figure 23-21

a The polymer polyvinyl chloride is commonly called “vinyl.” It is used to make flexible, waterproof objects. **b** Polystyrene plastic is inexpensive and easy to mold into parts for model cars and airplanes.

Reactions Used to Make Polymers

Polymers are relatively easy to manufacture. They usually can be synthesized in one step in which the major reactant is a substance consisting of small, simple organic molecules called monomers. A **monomer** is a molecule from which a polymer is made.

When a polymer is made, monomers bond together one after another in a rapid series of steps. A catalyst usually is required for the reaction to take place at a reasonable pace. With some polymers, such as Dacron and nylon, two or more kinds of monomers bond to each other in an alternating sequence. A reaction in which monomer units are bonded together to form a polymer is called a **polymerization reaction**. The repeating group of atoms formed by the bonding of the monomers is called the structural unit of the polymer. The structural unit of a polymer made from two different monomers has the components of both monomers.

Unbreakable children's toys are often made of polyethylene, which is synthesized by polymerizing ethene under pressure. Two monomers react to form polyethylene terephthalate (PET), a versatile plastic that is used to make bottles and recording tape. When made into fiber, it is called Dacron. Polyethylene and PET, examples of polymers made by two types of reactions, are shown in **Table 23-3**.

Addition polymerization In **addition polymerization**, all of the atoms present in the monomers are retained in the polymer product. When the monomer is ethene, an addition polymerization results in the polymer polyethylene, as shown in **Table 23-3**. Unsaturated bonds are broken in addition polymerization just as they are in addition reactions. The difference is that

Careers Using Chemistry

Dental Assistant

Would you like to mix polymers, process X rays, and in other ways help a dentist care for people's teeth? If so, you might become a dental assistant.

While dental hygienists clean teeth, dental assistants work directly with the dentist. They prepare and sterilize instruments, remove sutures, mix adhesives, prepare fillings, make casts of teeth, and create temporary crowns. Some assistants also keep treatment records and schedule patient appointments.

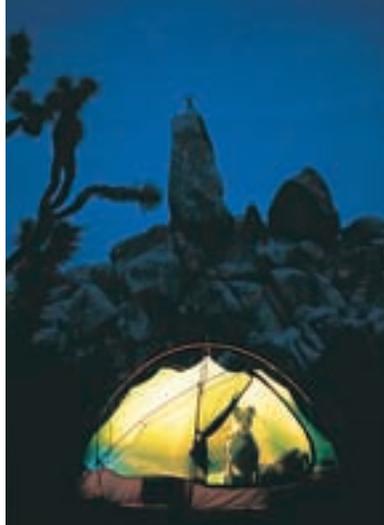
Table 23-3

Monomers and Polymers		
Monomer(s)	Structural unit of polymer	Application
$\begin{array}{c} \text{H} & & \text{H} \\ & \backslash & / \\ & \text{C} = \text{C} \\ & / & \backslash \\ \text{H} & & \text{H} \end{array}$ <p>Ethene (ethylene)</p>	$\begin{array}{ccccccc} \text{H} & \left[\begin{array}{cc} \text{H} & \text{H} \end{array} \right] & \text{H} & \text{H} & \text{H} & & \\ & & & & & & \\ -\text{C} & -\text{C} & -\text{C} & -\text{C} & -\text{C} & -\text{C}- & \\ & & & & & & \\ \text{H} & \left[\begin{array}{cc} \text{H} & \text{H} \end{array} \right] & \text{H} & \text{H} & \text{H} & & \\ & n & & & & & \end{array}$ <p>Polyethylene</p>	
$\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{HO}-\text{C}-\text{C}-\text{OH} \\ & \\ \text{H} & \text{H} \end{array}$ <p>1,2-Ethanediol (ethylene glycol)</p> $\begin{array}{c} \text{O} & & \text{O} \\ & & \\ \text{HO}-\text{C} & -\text{C}_6\text{H}_4 & -\text{C}-\text{OH} \end{array}$ <p>Terephthalic acid</p>	$\left[\text{O}-\overset{\text{O}}{\parallel}{\text{C}}-\text{C}_6\text{H}_4-\overset{\text{O}}{\parallel}{\text{C}}-\text{O}-\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} & -\text{C} \\ & \\ \text{H} & \text{H} \end{array} \right]_n$ <p>Polyethylene terephthalate (Dacron in fiber form) (Mylar in film form)</p>	

the molecule added is a second molecule of the same substance, ethene. Note that the addition polymers in **Table 23-4** are similar in structure to polyethylene. That is, the molecular structure of each is equivalent to polyethylene in which other atoms or groups of atoms are attached to the chain in place of hydrogen atoms. All of these polymers are made by addition polymerization.

Table 23-4

Common Polymers		
Polymer	Uses	Structural unit
Polyethylene (PE)	Plastic bags and wrap, food containers, children's toys, bottles	See Table 23-3 .
Polyvinyl chloride (PVC)	Plastic pipes, meat wrap, upholstery, rainwear, house siding, garden hose	See Figure 23-21 .
Polyacrylonitrile (Orlon)	Fabrics for clothing and upholstery, carpet	$\left[\text{CH}_2 - \underset{\text{C} \equiv \text{N}}{\text{CH}} \right]_n$
Polyvinylidene chloride (Saran)	Food wrap, fabrics	$\left[\text{CH}_2 - \underset{\text{Cl}}{\overset{\text{Cl}}{\text{C}}} \right]_n$
Polytetrafluoroethylene (Teflon, PTFE)	Nonstick coatings, bearings, lubricants	$\left[\underset{\text{F}}{\overset{\text{F}}{\text{C}}} - \underset{\text{F}}{\overset{\text{F}}{\text{C}}} \right]_n$
Polymethyl methacrylate (Lucite, Plexiglass)	"Nonbreakable" windows, inexpensive lenses, art objects	$\left[\text{CH}_2 - \underset{\text{CH}_3}{\overset{\text{O}}{\text{C}} - \text{OCH}_3} \right]_n$
Polypropylene (PP)	Beverage containers, rope, netting, kitchen appliances	$\left[\text{CH}_2 - \underset{\text{CH}_3}{\text{CH}} \right]_n$
Polystyrene styrene plastic	Foam packing and insulation, plant pots, disposable food containers, model kits	See Figure 23-21 .
Polyethylene terephthalate (PET, Dacron, Mylar)	Soft drink bottles, tire cord, clothing, recording tape, replacements for blood vessels	See Table 23-3 .
Nylon	Upholstery, clothing, carpet, fishing line, small gears, bearings	See Figure 23-22 .
Polyurethane	Foam furniture cushions, waterproof coatings, parts of shoes	$\left[\overset{\text{O}}{\parallel} \text{C} - \text{NH} - \text{CH}_2 - \text{CH}_2 - \text{NH} - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{CH}_2 - \text{CH}_2 - \text{O} \right]_n$



Condensation polymerization Condensation polymerization takes place when monomers containing at least two functional groups combine with the loss of a small by-product, usually water. Nylon and Kevlar are made this way. Nylon was first synthesized in 1931 and soon became popular because it is strong and can be drawn into thin strands resembling silk. Nylon-6,6 is the name of one type of nylon that is synthesized from two different six-carbon monomers. One monomer is a chain with the end carbon atoms being part of carboxyl groups. The other monomer is a chain having amino groups at both ends. These monomers undergo a condensation polymerization that forms amide groups linking the subunits of the polymer, as shown by the tinted -NH- group in **Figure 23-22**. Note that one water molecule is released for every new amide bond formed.

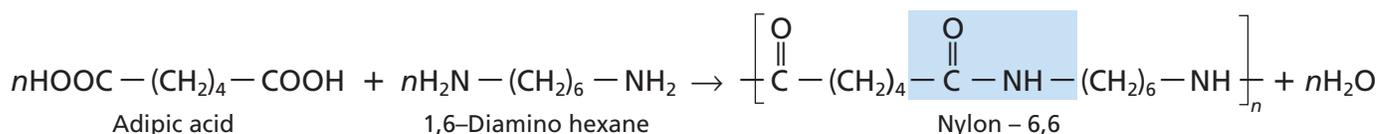


Figure 23-22

Adipic acid and 1,6-diamino-hexane are the monomers that polymerize by condensation to form nylon-6,6. You are probably familiar with nylon fabrics, such as those used to make tents, but nylon can be molded into solid objects too.

Materials Made from Polymers: Uses and Recycling

Why do we use so many different polymers today? One reason is that they are easy to synthesize. Another reason is that the starting materials used to make them are inexpensive. Still another, more important, reason is that polymers have a wide range of properties. Some polymers can be drawn into fine fibers that are softer than silk, while others are as strong as steel. Polymers don't rust like steel does, and many are more durable than natural materials such as wood. Objects made from lumber that is actually plastic, such as those shown in **Figure 23-23**, may be familiar to you.

Properties of polymers Another reason that polymers are in such great demand is that it is easy to mold them into different shapes or to draw them into thin fibers. It is not easy to do this with metals and other natural materials because either they must be heated to high temperatures, do not melt at all, or are too weak to be used to form small, thin items. A **plastic** is a polymer that can be heated and molded while relatively soft.

As with all substances, polymers have properties that result directly from their molecular structure. For example, polyethylene is a long-chain alkane. Thus, it has a waxy feel, does not dissolve in water, is nonreactive, and is a poor electrical conductor. These properties make it ideal for use in food and beverage containers and as an insulator on electrical wire and TV cable.

Polymers fall into two different categories based on their melting characteristics. A **thermoplastic** polymer is one that can be melted and molded repeatedly into shapes that are retained when it is cooled. Polyethylene and nylon are examples of thermoplastic polymers. A **thermosetting** polymer is one that can be molded when it is first prepared, but when cool cannot be remelted. This property is explained by the fact that thermosetting polymers begin to form networks of bonds in many directions when they are synthesized. By the time they have cooled, thermosetting polymers have become, in essence, a single large molecule. Bakelite is an example of a thermosetting polymer. Instead of melting, Bakelite decomposes or burns when overheated.



Figure 23-23

Plastic lumber is made from recycled plastic, especially soft drink bottles and polyethylene waste.



Recycling polymers The starting materials for the synthesis of most polymers are derived from fossil fuels. As the supply of fossil fuels is depleted, recycling plastics will become more important. Thermosetting polymers are more difficult to recycle than thermoplastic polymers because only thermoplastic materials can be melted and remolded repeatedly.

Currently, only about 1% of the plastic waste we produce in the United States is recycled. This figure contrasts with the 20% of paper waste and 30% of aluminum waste that are recycled. This low rate of plastics recycling is due in part to the large variety of different plastics found in products. The plastics must be sorted according to polymer composition. This task is time-consuming and expensive. The plastics industry and the government have tried to improve the process by providing standardized codes that indicate the composition of each plastic product. See **Figure 23-24**.

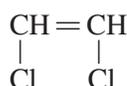
Figure 23-24

These codes tell recyclers what kind of plastic an object is made of.

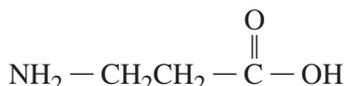
Section 23.5 Assessment

25. Draw the structure for the polymer that could be produced from each of the following monomers by the method stated.

a. Addition

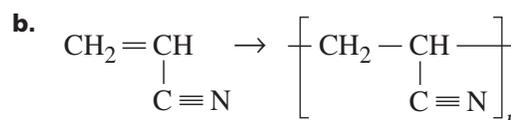
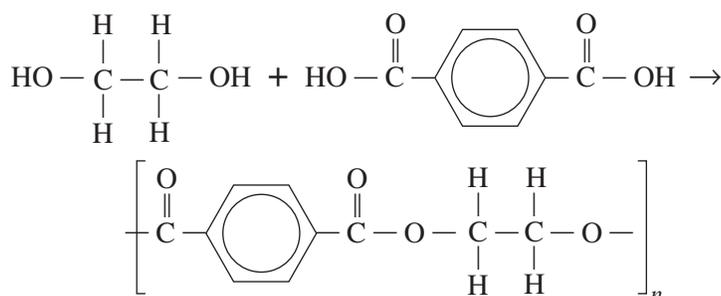


b. Condensation



26. Label each of the following polymerization reactions as addition or condensation. Write the formulas of the secondary product of the condensation reaction.

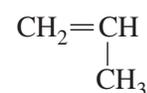
a.



27. Compare the properties of thermosetting and thermoplastic polymers.

28. Thinking Critically A chemical process called crosslinking forms covalent bonds between separate polymer chains. How do you think a polymer's properties will change as the number of crosslinks increases? What effect might additional crosslinking have on a thermoplastic polymer?

29. Predicting Predict the physical properties of the polymer that is made from the following monomer. Mention solubility in water, electrical conductivity, texture, and chemical reactivity. Do you think it will be thermoplastic or thermosetting? Give reasons for your predictions.



Properties of Alcohols

Alcohols are organic compounds that contain the $-OH$ functional group. In this experiment, you will determine the strength of intermolecular forces of alcohols by determining how fast various alcohols evaporate. The evaporation of a liquid is an endothermic process, absorbing energy from the surroundings. This means that the temperature will decrease as evaporation occurs.

Problem

How do intermolecular forces differ in three alcohols?

Objectives

- **Measure** the rate of evaporation for water and several alcohols.
- **Infer** the relative strength of intermolecular forces of alcohols from rate of evaporation data.

Materials

thermometer
stopwatch
facial tissue
cloth towel
Beral pipettes (5)
methanol
ethanol (95%)
2-propanol (99%)
wire twist tie or small rubber band
piece of cardboard for use as a fan

Safety Precautions



- Always wear safety goggles and a lab apron.
- The alcohols are flammable. Keep them away from open flames.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Draw structural formulas for the three alcohols you will use in this activity. Describe how the structures of these compounds are alike and how they are different.
4. What types of forces exist between these kinds of molecules? Suggest which alcohol may have the greatest intermolecular forces.

Procedure

1. Cut out five 2-cm by 6-cm strips of tissue.
2. Place a thermometer on a folded towel lying on a flat table so that the bulb of the thermometer extends over the edge of the table. Make sure the thermometer cannot roll off the table.
3. Wrap a strip of tissue around the bulb of the thermometer. Secure the tissue with a wire twist tie placed above the bulb of the thermometer.
4. Choose one person to control the stopwatch and read the temperature on the thermometer. A second person will put a small amount of the liquid to be tested into a Beral pipette.
5. When both people are ready, squeeze enough liquid onto the tissue to completely saturate it. At the same time, the other person starts the stopwatch, reads the temperature, and records it in the data table.
6. Fan the tissue-covered thermometer bulb with a piece of cardboard or other stiff paper. After one minute, read and record the final temperature in the data table. Remove the tissue and wipe the bulb dry.

Evaporation Data

Substance	Starting temp (°C)	Temp after one minute (°C)	ΔT (°C)
Water			
Methanol			
Ethanol			
2-Propanol			
Other alcohol			

7. Repeat steps 3 through 6, for each of the three alcohols: methanol, ethanol, and 2-propanol. If your teacher has another alcohol, use it also.



Cleanup and Disposal

- Place tissues in the trash. Pipettes can be reused.

Analyze and Conclude

- Communicating** Formulate a statement that summarizes your data, relating temperature change to the substances tested. Do not draw any conclusions yet.
- Acquiring and Analyzing Information** Explain why the temperatures changed during the experiment.
- Observing and Inferring** What can you conclude about the relationship between heat transfer and the differences in the temperature changes you observed?
- Drawing Conclusions** Assume that the three alcohols have approximately the same molar enthalpy of vaporization. What can you say about the relative rates of evaporation of the three alcohols?
- Drawing Conclusions** Consider your answer to question 4. What can you conclude about the relative strength of intermolecular forces existing in the three alcohols?

- Predicting** Suppose you also tested the alcohol 1-pentanol in this experiment. Where among the alcohols tested would you predict 1-pentanol to rank in rate of evaporation from fastest to slowest? Describe the temperature change you would expect to observe. Explain your reasoning.
- Thinking Critically** Molar enthalpies of vaporization for the three alcohols are given in the table below. Note that they are not the same.

Molar Enthalpies of Vaporization	
Substance	Enthalpy of vaporization at 25°C (kJ/mol)
Methanol	37.4
Ethanol	42.3
2-Propanol	45.4

In what way, if any, does this data change your conclusion about intermolecular forces?

- Observing and Inferring** Make a general statement comparing the molecular size of an alcohol in terms of the number of carbons in the carbon chain to the rate of evaporation of that alcohol.
- Error Analysis** Suggest a way to make this experiment more quantitative and controlled.

Real-World Chemistry

- How can this experiment help explain why small-chain alcohols have a warning label indicating that they are flammable?
- Would you expect to see such a warning label on a bottle of 1-decanol? Explain.
- A mixture of 70% 2-propanol (isopropanol) and 30% water is sold as rubbing alcohol, which may be used to help reduce a fever. Explain how this process works.
- Why do you suppose that 2-propanol is a component in some products used to soothe sunburned skin?

CHEMISTRY and Technology

Carbon: Stronger than Steel?

If you play golf or tennis, fish or ride a bicycle, chances are good that you have used carbon-fiber technology. In the 1960s, the aerospace industry began searching for structural materials that were as strong as metals but were light in weight because making space vehicles lighter meant that larger satellites or more experimental equipment could be lifted into space with the same amount of fuel. At about the same time, the petroleum industry showed that it could make fibers that were nearly pure carbon. Carbon fibers are stronger than steel when pulled at the ends, but they break when bent sharply. Engineers solved this problem by making a mat with carbon fibers oriented in all directions. Embedding the mat in a plastic matrix produced sheets of extremely strong but lightweight material. Such a material—a mixture of two or more materials that produces a combination stronger than the materials alone—is called a composite material.

From Spacecraft to Sports

Carbon fibers are expensive to make, but their cost has decreased steadily. In the 1980s, a few expensive consumer items were produced, beginning with golf clubs having carbon-fiber-reinforced shafts. Then came tennis rackets, fishing rods, skis, and snowboards. Today, bicycles with frames made of carbon fiber are becoming increasingly popular. These frames are about 40% lighter than similar frames made of metal tubing. With lighter frames, riders use less energy to go the same distance. In the aircraft industry, less weight means fuel savings, so more and more parts such as seat frames and structural panels are being manufactured from carbon fiber.

Making Carbon Fibers

Carbon fibers are polymers composed of long chains of carbon atoms in aromatic rings. Each carbon fiber is composed of narrow sheets that are extremely long. How do you polymerize carbon atoms?

Manufacturers first produce fibers of a synthetic polymer called polyacrylonitrile, or PAN. PAN



fibers are heated intensely in an oxygen-free atmosphere, driving off all hydrogen, oxygen, and nitrogen atoms. After this process, the fiber consists of only carbon atoms with strong, stable bonds that give carbon fibers their great strength.

The Future of Carbon Fibers

Other possible applications of carbon-fiber composites include prosthetic limbs, musical instruments, sports helmets and other protective gear, lightweight building materials, and reinforcing beams used to repair aging bridges. Because carbon fibers conduct electricity, scientists and engineers are working to find ways to incorporate electronic circuits into structural materials.

Investigating the Technology

1. **Using Resources** What advantages might carbon fiber have over metals in artificial limbs?



CLICK HERE

Visit the Chemistry Web site at science.glencoe.com to find links about carbon fiber composites.

Summary

23.1 Functional Groups

- Carbon forms bonds with atoms other than C and H, especially O, N, S, P, F, Cl, Br, and I.
- An atom or group of atoms that always react in a certain way in an organic molecule is referred to as a functional group.
- An alkyl halide is an organic compound that has one or more halogen atoms (F, Cl, Br, or I) bonded to a carbon atom.

23.2 Alcohols, Ethers, and Amines

- An alcohol is an organic compound that has an $-OH$ group bonded to a carbon atom.
- Because they readily form hydrogen bonds, alcohols have higher boiling points and higher water solubilities than other organic compounds.
- Alcohols are used as solvents and as starting materials in synthesis reactions, medicinal products, and the food and beverage industries.
- An amine is an organic compound that contains a nitrogen atom bonded to one or more carbon atoms.
- An ether is an organic compound having the $R-O-R'$ structure.

23.3 Carbonyl Compounds

- Carbonyl compounds are organic compounds that contain the $C=O$ group.
- Five important classes of organic compounds containing carbonyl compounds are aldehydes, ketones, carboxylic acids, esters, and amides.

23.4 Other Reactions of Organic Compounds

- Most reactions of organic compounds can be classified into one of five categories: substitution; elimination; addition; oxidation-reduction; condensation.
- Knowing the types of organic compounds reacting may enable you to predict the reaction products.

23.5 Polymers

- Polymers are large molecules formed by combining smaller molecules called monomers.
- Polymers are synthesized through addition or condensation reactions.
- The functional groups present in polymers can be used to predict polymer properties.

Vocabulary

- addition polymerization (p. 762)
- addition reaction (p. 755)
- alcohol (p. 743)
- aldehyde (p. 747)
- alkyl halide (p. 738)
- amide (p. 752)
- amine (p. 745)
- aryl halide (p. 739)
- condensation polymerization (p. 764)
- condensation reaction (p. 753)
- carbonyl group (p. 747)
- carboxylic acid (p. 749)
- carboxyl group (p. 749)
- dehydration reaction (p. 755)
- dehydrogenation reaction (p. 754)
- denatured alcohol (p. 744)
- elimination reaction (p. 754)
- ester (p. 750)
- ether (p. 745)
- functional group (p. 737)
- halocarbon (p. 738)
- halogenation (p. 741)
- hydration reaction (p. 756)
- hydrogenation reaction (p. 756)
- hydroxyl group (p. 743)
- ketone (p. 748)
- monomer (p. 762)
- plastic (p. 764)
- polymer (p. 761)
- polymerization reaction (p. 762)
- substitution reaction (p. 741)
- thermoplastic (p. 764)
- thermosetting (p. 764)

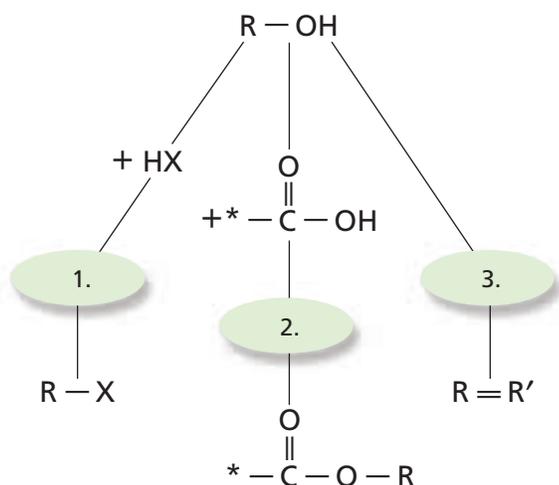


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 23 Assessment.

Concept Mapping

30. Identify the types of reactions used to convert alcohols into alkyl halides, esters, and alkenes.



Mastering Concepts

31. What is a functional group? (23.1)
32. Describe and compare the structures of alkyl halides and aryl halides. (23.1)
33. What reactant would you use to convert methane to bromomethane? (23.1)
34. Name the amines represented by the condensed formulas below. (23.2)
- $\text{CH}_3(\text{CH}_2)_3\text{CH}_2\text{NH}_2$
 - $\text{CH}_3(\text{CH}_2)_5\text{CH}_2\text{NH}_2$
 - $\text{CH}_3(\text{CH}_2)_2\text{CH}(\text{NH}_2)\text{CH}_3$
 - $\text{CH}_3(\text{CH}_2)_8\text{CH}_2\text{NH}_2$
35. How is ethanol denatured? (23.2)
36. Name one alcohol, amine, or ether that is used for each of the following purposes. (23.2)
- antiseptic
 - solvent in paint strippers
 - antifreeze
 - anesthetic
 - dye production
37. Explain why an alcohol molecule will always have a higher solubility in water than an ether molecule having an identical molecular mass. (23.2)
38. Explain why ethanol has a much higher boiling point than aminoethane even though their molecular masses are nearly equal. (23.2)
39. Draw the general structure for each of the following classes of organic compounds. (23.3)
- aldehyde
 - ketone
 - carboxylic acid
 - ester
 - amide
40. Name an aldehyde, ketone, carboxylic acid, ester, or amide used for each of the following purposes. (23.3)
- preserving biological specimens
 - solvent in fingernail polish
 - acid in vinegar
 - flavoring in foods and beverages
41. What type of reaction is used to produce aspirin from salicylic acid and acetic acid? (23.3)
42. What is the starting material for making most synthetic organic compounds? (23.4)
43. Explain the importance of classifying reactions. (23.4)
44. List the type of organic reaction needed to carry out each of the following transformations. (23.4)
- alkene \rightarrow alkane
 - alkyl halide \rightarrow alcohol
 - alkyl halide \rightarrow alkene
 - amine + carboxylic acid \rightarrow amide
 - alcohol \rightarrow alkyl halide
 - alkene \rightarrow alcohol
45. Explain the difference between addition polymerization and condensation polymerization. (23.5)
46. Which type of polymer is easier to recycle, thermosetting or thermoplastic? Explain your answer. (23.5)

Mastering Problems

Functional Groups (23.1)

47. Draw structures for these alkyl and aryl halides.
- chlorobenzene
 - 1-bromo-4-chlorohexane
 - 1,2-difluoro-3-iodocyclohexane
 - 1, 3-dibromobenzene
 - 1,1,2,2-tetrafluoroethane

61. Examine the structures of the following polymers in **Table 23-4**. Decide whether each is made by addition or condensation polymerization.

- a. nylon c. polyurethane
b. polyacrylonitrile d. polypropylene

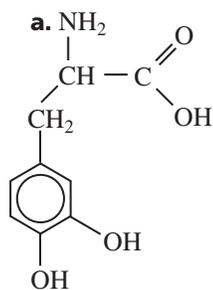
Mixed Review

Sharpen your problem-solving skills by answering the following.

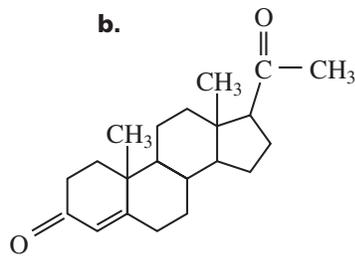
- 62.** Which halogen is found in hormones made by a normal human thyroid gland?
- 63.** Describe the properties of carboxylic acids.
- 64.** List two uses of esters.
- 65.** Draw structures of the following compounds. (23.3)
- a. butanone d. heptanoamide
b. propanal e. ethyl pentanoate
c. hexanoic acid f. benzoic acid
- 66.** Name the type of organic compound formed by each of the following reactions.
- a. elimination from an alcohol
b. addition of hydrogen chloride to an alkene
c. addition of water to an alkene
d. substitution of a hydroxyl group for a halogen atom
- 67.** List two uses for each of the following polymers.
- a. polypropylene c. polytetrafluoroethylene
b. polyurethane d. polyvinyl chloride
- 68.** Draw structures of and supply names for the organic compounds produced by reacting ethene with each of the following substances.
- a. water b. hydrogen
c. hydrogen chloride d. fluorine

Thinking Critically

69. Interpreting Scientific Illustrations List all the functional groups present in each of the following complex organic molecules.



Levodopa



Progesterone

70. Thinking Critically Ethanoic acid (acetic acid) is very soluble in water. However, naturally occurring long-chain carboxylic acids such as palmitic acid ($\text{CH}_3(\text{CH}_2)_{14}\text{COOH}$) are insoluble in water. Explain.

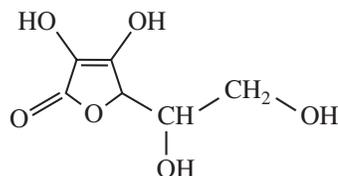
71. Communicating Write structural formulas for all structural isomers of molecules having the following formulas. Name each isomer.

- a. $\text{C}_3\text{H}_8\text{O}$ b. $\text{C}_2\text{H}_4\text{Cl}_2$ c. C_3H_6

72. Recognizing Cause and Effect Arrange the following compounds in order of increasing boiling point.

butanol, butane, 1-aminobutane, ethyl ether

73. Interpreting Scientific Illustrations Human cells require vitamin C to properly synthesize materials that make up connective tissue such as that found in ligaments. List the functional groups present in the Vitamin C molecule.



Vitamin C

Writing in Chemistry

74. While living organisms have made polymers like cotton, silk, wool, and latex rubber for thousands of years, the first laboratory synthesis of a polymer occurred only in the late 1800s. Imagine that you live in the 1880s, before society entered what some chemists refer to as the "Age of Polymers." Write a short story describing how your life would differ from its present form because of the absence of synthetic polymers.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

75. Why do the following characteristics apply to transition metals? (Chapter 7)

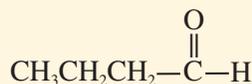
- a. Ions vary in charge.
b. Many of their solids are colored.
c. Many are hard solids.

76. What is a rate-determining step? (Chapter 17)

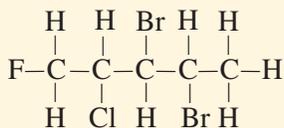
77. According to Le Châtelier's principle, how would increasing the volume of the reaction vessel affect the equilibrium $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$? (Chapter 18)

Use these questions and the test-taking tip to prepare for your standardized test.

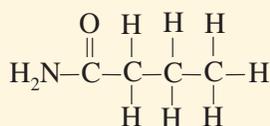
1. The compound pictured below is _____.



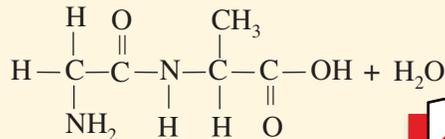
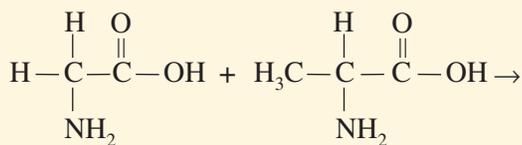
- a. an aldehyde c. a ketone
b. an ether d. an alcohol
2. The name of the compound shown is _____.



- a. 2-chloro-3,4-dibromo-1-fluoropentane
b. 4-chloro-2,3-dibromo-5-fluoropentane
c. 3,4-dibromo-2-chloro-1-fluoropentane
d. 2,3-dibromo-4-chloro-5-fluoropentane
3. The products of this reaction are _____.
- $$\text{CH}_3\text{CH}_2\text{CH}_2\text{Br} + \text{NH}_3 \rightarrow ?$$
- a. $\text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_2\text{Br}$ and H_2
b. $\text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_3$ and Br_2
c. $\text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_2$ and HBr
d. $\text{CH}_3\text{CH}_2\text{CH}_3$ and NH_2Br
4. What type of compound does this molecule represent?



- a. an amine c. an ester
b. an amide d. an ether
5. What kind of reaction is this?

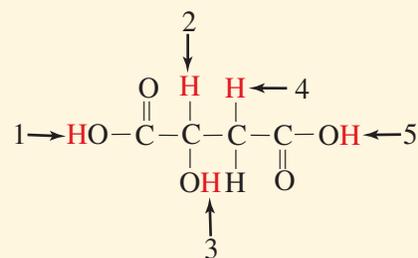


- a. substitution c. addition
b. condensation d. elimination

Interpreting Tables Use the table to answer the following questions.

Acid Ionization Constants for Selected Carboxylic Acids			
Common name	Formula	Hydrogen ionized	K_a
Succinic acid	$\text{C}_2\text{H}_4(\text{COOH})_2$	1st	6.92×10^{-5}
		2nd	2.45×10^{-6}
Oxaloacetic acid	$\text{C}_2\text{H}_2\text{O}(\text{COOH})_2$	1st	6.03×10^{-3}
		2nd	1.29×10^{-4}
Acrylic acid	$\text{C}_3\text{H}_4\text{O}_2$	1st	5.62×10^{-5}
Malic acid	$\text{C}_4\text{H}_6\text{O}_5$	1st	3.98×10^{-4}
		2nd	7.76×10^{-6}

6. Oxaloacetic acid has a higher K_a than succinic acid because it probably possesses _____.
- a. fewer carboxyl groups
b. a less polar structure
c. additional functional groups that make it more soluble in water
d. fewer hydrogen atoms that can transfer to the oxygen atom in H_2O
7. Which of the indicated hydrogen atoms will malic acid lose when it ionizes completely?



- a. 1 and 5 c. 2, 3, and 4
b. 1, 3, and 5 d. 2 and 4

Use As Much Time As You Can You will not get extra points for finishing early. Work slowly and carefully on any test and make sure you don't make careless errors because you are hurrying to finish.

The Chemistry of Life

What You'll Learn

- ▶ You will learn the functions of the four major classes of biological molecules: proteins, carbohydrates, lipids, and nucleic acids.
- ▶ You will identify the building blocks that form the major biological molecules.
- ▶ You will compare and contrast the metabolic processes of cellular respiration, photosynthesis, and fermentation.

Why It's Important

The large biological molecules in your body are essential to the organization and operation of its millions of cells. The structure of these molecules is directly related to their function, and how they function affects your health and survival.

CLICK HERE



Visit the Chemistry Web site at science.glencoe.com to find links about the chemistry of life.

The silk that makes up this spider's web is, gram for gram, stronger than steel, yet it is lightweight and stretchable. Spider silk is made of protein, a biological molecule.



DISCOVERY LAB



Materials

400-mL beaker
hot plate
10-mL graduated cylinder
boiling chip
10% glucose solution
test tube
tongs
Benedict's solution
stirring rod
other food solutions such as
10% starch or honey

Testing for Simple Sugars

Your body constantly uses energy. Many different food sources can supply that energy, which is stored in the bonds of molecules called simple sugars. What foods contain simple sugars?

Safety Precautions



Procedure

1. Fill a 400-mL beaker one-third full of water. Place this water bath on a hot plate and begin to heat it to boiling.
2. Place 5.0 mL 10% glucose solution in a test tube.
3. Add 3.0 mL Benedict's solution to the test tube. Mix the two solutions using a stirring rod. Add a boiling chip to the test tube.
4. Using tongs, place the test tube in the boiling water bath and heat for five minutes.
5. Record a color change from blue to yellow or orange as a positive test for a simple sugar.
6. Repeat the procedure using food samples such as a 10% starch solution, a 10% gelatin (protein) suspension, or a few drops of honey suspended in water.

Analysis

Was a color change observed? Which foods tested positive for the presence of a simple sugar?

Section 24.1

Proteins

Objectives

- **Describe** the structures of amino acids and proteins.
- **Explain** the roles of proteins in cells.

Vocabulary

protein
amino acid
peptide bond
peptide
denaturation
enzyme
substrate
active site

An amazing variety of chemical reactions take place in living organisms. At the forefront of coordinating the numerous and intricate reactions of life are the large molecules called proteins, whose name comes from the Greek root word *protos*, meaning first.

Protein Structure

You have learned that polymers are large molecules made of many repeating building blocks called monomers. **Proteins** are organic polymers made of amino acids linked together in a specific way. But proteins are not just large, randomly arranged chains of amino acids. To function properly, each protein must be folded into a specific three-dimensional structure. The spider silk shown on the opposite page would not be the incredibly strong yet lightweight protein that it is if it were not constructed in its specific way. You will learn in this section how proteins are made from their amino-acid building blocks and how different types of proteins function.

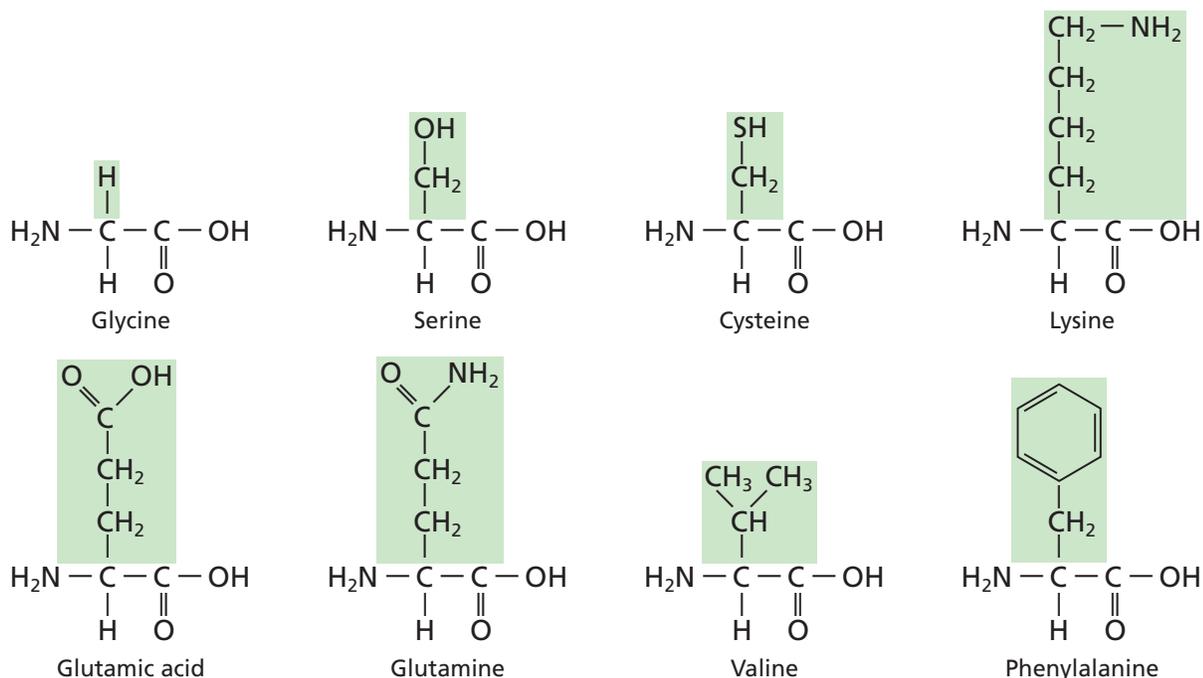
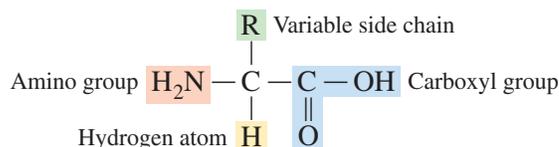


Figure 24-1

A large variety of side chains can be found on amino acids. Some of them are shown on these representative amino acids and are highlighted in green.

Amino acids As you saw in Chapter 23, many different functional groups are found in organic compounds. **Amino acids**, as their name implies, are organic molecules that have both an amino group and an acidic carboxyl group. The general structure of an amino acid is shown below.



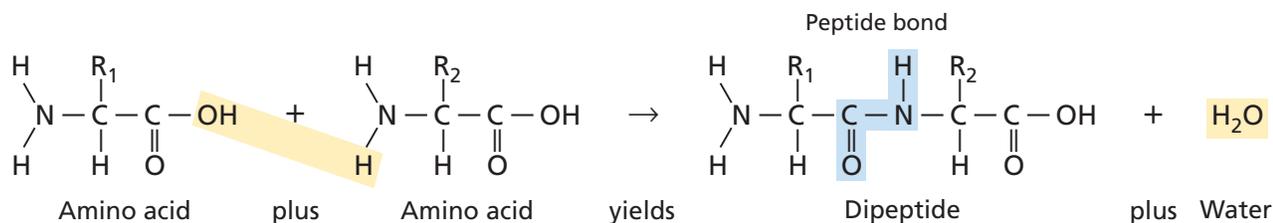
Each amino acid has a central carbon atom around which are arranged four groups: an amino group (—NH₂), a carboxyl group (—COOH), a hydrogen atom, and a variable side chain, *R*. The side chains range from a single hydrogen atom to a complex double-ring structure. Examine the different side chains of the amino acids shown in **Figure 24-1**. Identify the nonpolar alkanes, polar hydroxyl groups, acidic and basic groups such as carboxyl and amino groups, aromatic rings, and sulfur-containing groups. This wide range of side chains gives the different amino acids a large variety of chemical and physical properties and is an important reason why proteins can carry out so many different functions.

Figure 24-2

The amino group of one amino acid bonds to the carboxyl group of another amino acid to form a dipeptide. The organic functional group formed is an amide linkage and is called a peptide bond.

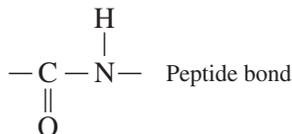
Twenty different amino acids are commonly found in the proteins of living things. The name of each amino acid and its three-letter abbreviation are listed in **Table 24-1**. What is the abbreviation for glycine?

The peptide bond The amino and carboxyl groups provide convenient bonding sites for linking amino acids together. Since an amino acid is both an amine and a carboxylic acid, two amino acids can combine to form an



amide, releasing water in the process. This reaction is a condensation reaction. As **Figure 24-2** shows, the amino group of one amino acid reacts with the carboxyl group of another amino acid to form an amide functional group. Where do the H and OH that form water come from?

The amide bond that joins two amino acids is referred to by biochemists as a **peptide bond**.



A molecule that consists of two amino acids bound together by a peptide bond is called a dipeptide. **Figure 24-3a** shows the structure of a dipeptide that is formed from the amino acids glycine (Gly) and phenylalanine (Phe). **Figure 24-3b** shows a different dipeptide, also formed by linking together glycine and phenylalanine. Is Gly-Phe the same compound as Phe-Gly? No, they're different. Examine these two dipeptides to see that the order in which amino acids are linked in a dipeptide is important.

Each end of the two-amino-acid unit in a dipeptide still has a free group—one end has a free amino group and the other end has a free carboxyl group. Each of those groups can be linked to the opposite end of yet another amino acid, forming more peptide bonds. A chain of two or more amino acids linked together by peptide bonds is called a **peptide**. Living cells always build peptides by adding amino acids to the carboxyl end of a growing chain.

Polypeptides As peptide chains increase in length, other ways of referring to them become necessary. A chain of ten or more amino acids joined by peptide bonds is referred to as a polypeptide. When a chain reaches a length of about 50 amino acids, it's called a protein.

Because there are only 20 different amino acids that form proteins, it might seem reasonable to think that only a limited number of different protein structures are possible. But a protein can have from 50 to a thousand or more amino acids, arranged in any possible sequence. To calculate the number of possible sequences these amino acids can have, you need to consider that each position on the chain can have any of 20 possible amino acids. For a peptide that contains n amino acids, there are 20^n possible sequences of the amino acids. So a dipeptide, with only two amino acids, can have 20^2 , or 400, different possible amino acid sequences. Even the smallest protein containing only 50 amino acids has 20^{50} , or more than 1×10^{65} , possible arrangements of amino acids! It is estimated that human cells make between 80 000 and 100 000 different proteins. You can see that this is only a very small fraction of the total number of proteins possible.

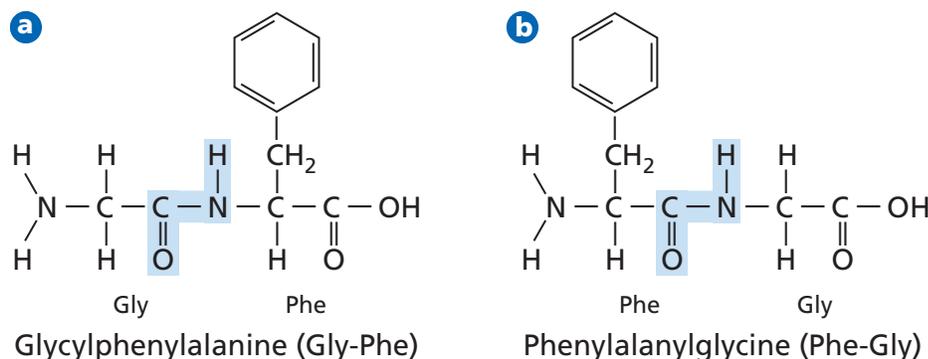


Table 24-1

The 20 Amino Acids	
Amino acid	Abbreviation
Alanine	Ala
Arginine	Arg
Asparagine	Asn
Aspartic acid	Asp
Cysteine	Cys
Glutamic acid	Glu
Glutamine	Gln
Glycine	Gly
Histidine	His
Isoleucine	Ile
Leucine	Leu
Lysine	Lys
Methionine	Met
Phenylalanine	Phe
Proline	Pro
Serine	Ser
Threonine	Thr
Tryptophan	Trp
Tyrosine	Tyr
Valine	Val

Figure 24-3

- a** Glycine and phenylalanine can be combined in this configuration.
- b** Glycine and phenylalanine can also be combined in this configuration. Why are these two structures different substances?

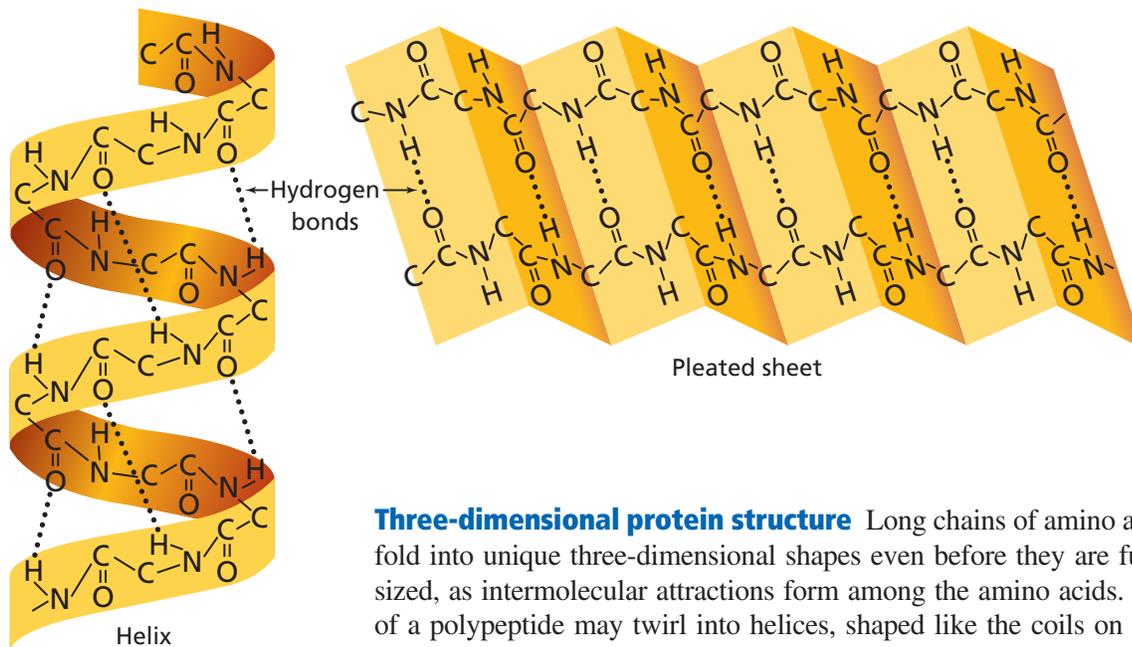


Figure 24-4

The folding of polypeptide chains into both helices (left) and sheets (right) involves amino acids that are fairly close together in the chain being held in position by hydrogen bonds. Other interactions among the various side chains are not shown here but play an important role in determining the three-dimensional shape of a polypeptide.

Three-dimensional protein structure Long chains of amino acids start to fold into unique three-dimensional shapes even before they are fully synthesized, as intermolecular attractions form among the amino acids. Some areas of a polypeptide may twirl into helices, shaped like the coils on a telephone cord. Other areas may bend back and forth again and again into a sheet structure, like the folds of an accordion. A polypeptide chain may also turn back on itself and change direction. A given protein may have several helices, sheets, and turns, or none at all. **Figure 24-4** shows the folding patterns of a typical helix and a sheet. The overall three-dimensional shape of many proteins is globular, or shaped like an irregular sphere. Other proteins have a long fibrous shape. The three-dimensional shape is determined by the interactions among the amino acids.

Changes in temperature, ionic strength, pH, and other factors can act to disrupt these interactions, resulting in the unfolding and uncoiling of a protein. **Denaturation** is the process in which a protein's natural three-dimensional structure is disrupted. Cooking often denatures the proteins in foods. When an egg is hard-boiled, the protein-rich egg white solidifies due to the denaturation of its protein. Because proteins function properly only when folded, denatured proteins generally are inactive.

The Many Functions of Proteins

Proteins play many roles in living cells. They are involved in catalyzing reactions, transporting substances, regulating cellular processes, forming structures, digesting foods, recycling wastes, and even serving as an energy source when other sources are scarce. See the **Everyday Chemistry** feature at the end of this chapter to learn how proteins play a role in the sense of smell.

Enzymes In most organisms, the largest number of different proteins function as enzymes, catalyzing the many reactions that go on in living cells. An **enzyme** is a biological catalyst. Recall that a catalyst speeds up a chemical reaction without being consumed in the reaction. A catalyst usually lowers the activation energy of a reaction by stabilizing the transition state.

How do enzymes function? The term **substrate** is used to refer to a reactant in an enzyme-catalyzed reaction. Substrates bind to specific sites on enzyme molecules, usually pockets or crevices. The pocket to which the substrates bind is called the **active site** of the enzyme. After the substrates bind to the active site, the active site changes shape slightly to fit more tightly around the substrates. This recognition process is called induced fit. In the diagram in **Figure 24-5**, you'll see that the shapes of the substrates must fit



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

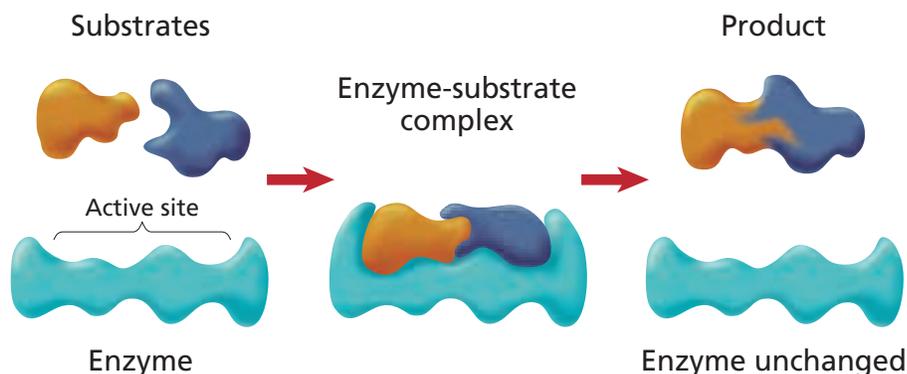


Figure 24-5

Substrates are brought close together as they fit into the uniquely shaped active site of an enzyme. The enzyme then changes shape as it “molds itself” to the substrates, forming an enzyme-substrate complex. Bonds are broken and new bonds form to produce the product in the reaction.

that of the active site, in the same way that puzzle pieces or a lock and key fit together. A molecule only slightly different in shape from an enzyme’s normal substrate doesn’t bind as well to the active site or undergo the catalyzed reaction.

The structure that forms when substrates are bound to an enzyme is called an enzyme-substrate complex. The large size of enzyme molecules allows them to form multiple bonds with their substrates, and the large variety of amino acid side chains in the enzyme allows a number of different intermolecular forces to form. These intermolecular forces lower the activation energy needed for the reaction in which bonds are broken and the substrates are converted to product.

An example of an enzyme you may have used is papain, found in papayas, pineapple, and other plant sources. This enzyme catalyzes a reaction that breaks down protein molecules into free amino acids. Papain is the active ingredient in many meat tenderizers. When you sprinkle the dried form of papain on moist meat, you activate the papain so that it breaks down the tough protein fibers in the meat, making the meat more tender.

Transport proteins Some proteins are involved in transporting smaller molecules throughout the body. The protein hemoglobin, modeled in **Figure 24-6**, carries oxygen in the blood, from your lungs to the rest of your body. Other proteins combine with biological molecules called lipids to transport them from one part of your body to another through the bloodstream. You will learn about lipids later in this chapter.

Figure 24-6

- a** Hemoglobin is a globular protein with four polypeptide chains. The red structure in each chain is heme, an organic group containing an iron ion to which oxygen binds.
- b** The pink skin of this pig is due to the hemoglobin in its blood.

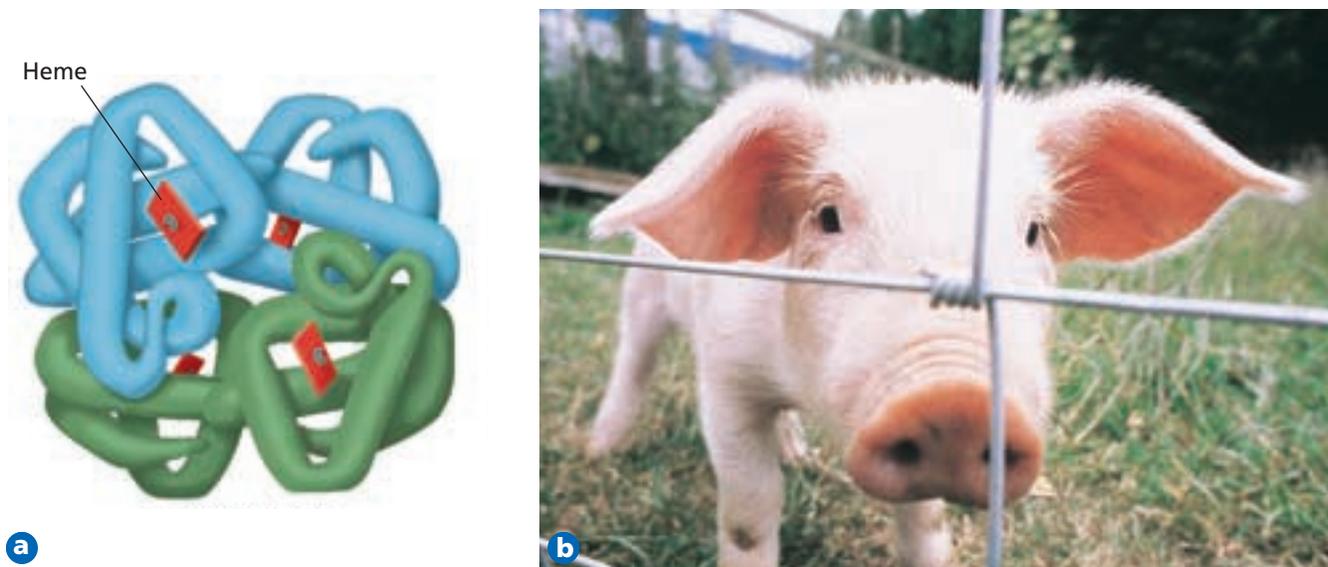
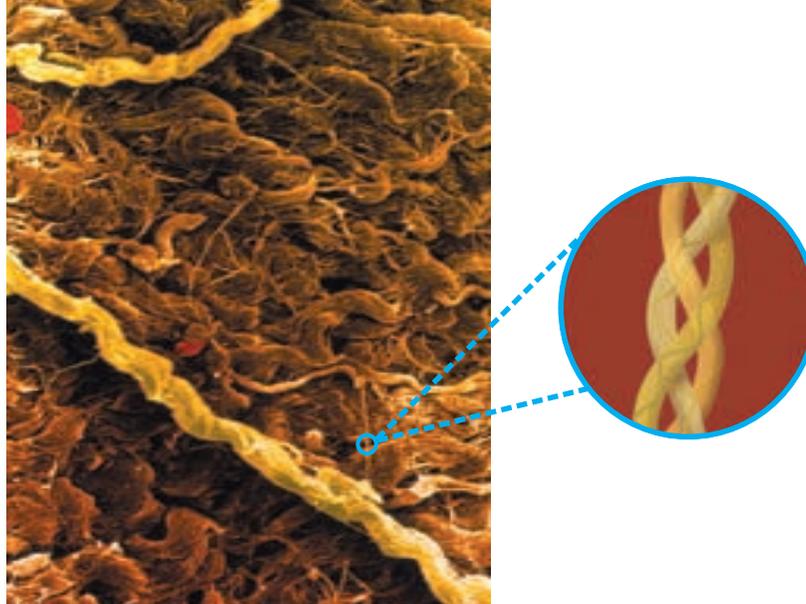


Figure 24-7

Collagen is a fibrous protein that consists of three helical polypeptides (right) that are coiled around one another to form the strong fibers of connective tissue (left). The large ropelike structures in the photo are bundles of collagen fibers, held together by cross-linkage. Collagen is an extremely large protein. Each of the three strands is about 1000 amino acids long.



Structural proteins The sole function of some proteins is to form structures vital to organisms. These molecules are known as structural proteins. The most abundant structural protein in most animals is collagen, which is part of skin, ligaments, tendons, and bones. **Figure 24-7** shows the structure of collagen. Other structural proteins make up hair, fur, wool, hooves, fingernails, cocoons, and feathers.

Hormones Hormones are messenger molecules that carry signals from one part of the body to another. Some hormones are proteins. Insulin, a familiar example, is a small (51 amino acids) protein hormone made by cells in the pancreas. When insulin is released into the bloodstream, it signals body cells that blood sugar is abundant and should be stored. A lack of insulin results in diabetes, a disease in which there is too much sugar in the bloodstream. Another protein hormone, chorionic gonadotropin, is synthesized by a developing embryo. Release of this hormone causes the development of a placenta that nourishes the embryo.

Because modern technology has made possible the laboratory synthesis of proteins, some protein hormones are being synthetically produced for use as medicines. Insulin, thyroid hormones, and growth hormones are some examples. Both natural and synthetic proteins are used in a variety of products—from meat tenderizer to cleaning solutions to health and beauty aids.

Section 24.1 Assessment

1. Compare the structures of amino acids, dipeptides, polypeptides, and proteins. Which has the largest molecular mass? The smallest?
2. Draw the structure of the dipeptide Gly-Ser, circling the peptide bond.
3. List four functions that different proteins have in living organisms.
4. **Thinking Critically** How do the properties of proteins make them such useful catalysts? How do they differ from other catalysts you have studied?
5. **Applying Concepts** Identify an amino acid from **Figure 24-1** that can be classified into each of the categories in the following pairs.
 - a. nonpolar versus polar
 - b. aromatic versus aliphatic
 - c. acidic versus basic

Analyzing the term *carbohydrate* offers a hint about the structure of this group of molecules. Early observations that these compounds have the general chemical formula $C_n(H_2O)_n$ and appear to be hydrates of carbon led to their being called carbohydrates. Although it is now known that there are no full water molecules attached to carbohydrates, the name has stayed.

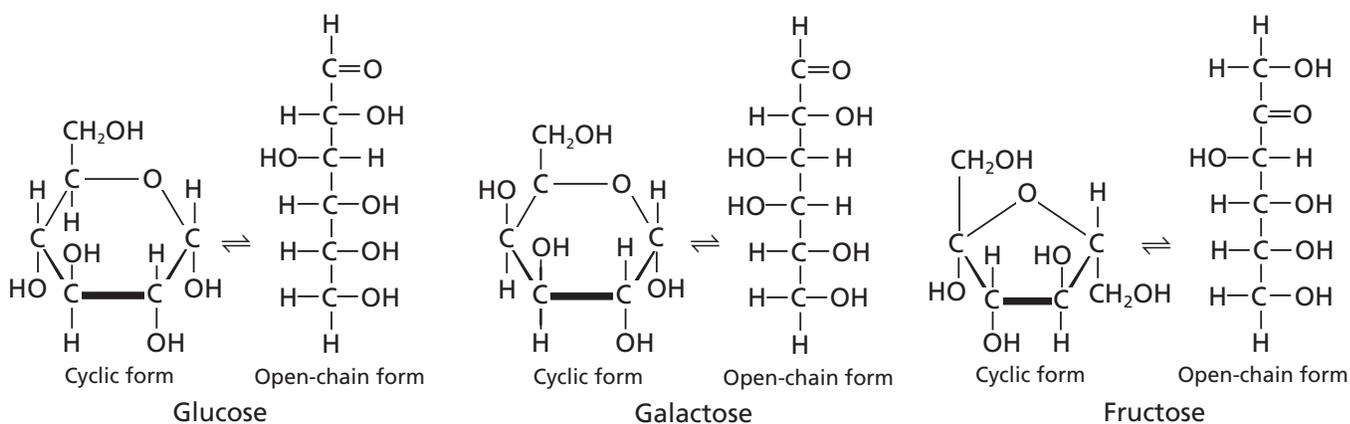
Kinds of Carbohydrates

Why do marathon runners eat large quantities of pasta before a big race? Pasta, as well as milk, fruit, bread, and potatoes, is rich in carbohydrates. The main function of carbohydrates in living organisms is as a source of energy, both immediate and stored. **Carbohydrates** are compounds that contain multiple hydroxyl groups ($—OH$) as well as a carbonyl functional group ($C=O$). These molecules range in size from single monomers to polymers made of hundreds or thousands of monomer units.

Monosaccharides The simplest carbohydrates, often called simple sugars, are **monosaccharides**. The most common monosaccharides have either five or six carbon atoms. They have a carbonyl group on one carbon and hydroxyl groups on most of the other carbons. The presence of a carbonyl group makes these compounds either aldehydes or ketones, depending upon the location of the carbonyl group. Multiple polar groups make monosaccharides water soluble and give them high melting points.

Glucose is a six-carbon sugar that has an aldehyde structure. Glucose is present in high concentration in blood because it serves as the major source of immediate energy in the body. For this reason, glucose is often called blood sugar. Closely related to glucose is galactose, which differs only in how a hydrogen and a hydroxyl group are oriented in space around one of the six carbon atoms. As you recall from Chapter 22, this relationship makes glucose and galactose stereoisomers; that is, their atoms are arranged differently in space. Fructose, also known as fruit sugar because it is the major carbohydrate in most fruits, is a six-carbon monosaccharide that has a ketone structure. Fructose is a structural isomer of glucose.

When monosaccharides are in aqueous solution, they exist in both open-chain and cyclic structures. The cyclic structures are more stable and are the predominant form of monosaccharides at equilibrium. Note in **Figure 24-8** that the carbonyl groups are present only in the open-chain structures. In the cyclic structures, they are converted to hydroxyl groups.



Objectives

- **Describe** the structures of monosaccharides, disaccharides, and polysaccharides.
- **Explain** the functions of carbohydrates in living things.

Vocabulary

carbohydrate
monosaccharide
disaccharide
polysaccharide

Figure 24-8

Glucose, galactose, and fructose are monosaccharides. In aqueous solutions, they exist in an equilibrium between their open-chain and cyclic forms, which are interconverted rapidly.

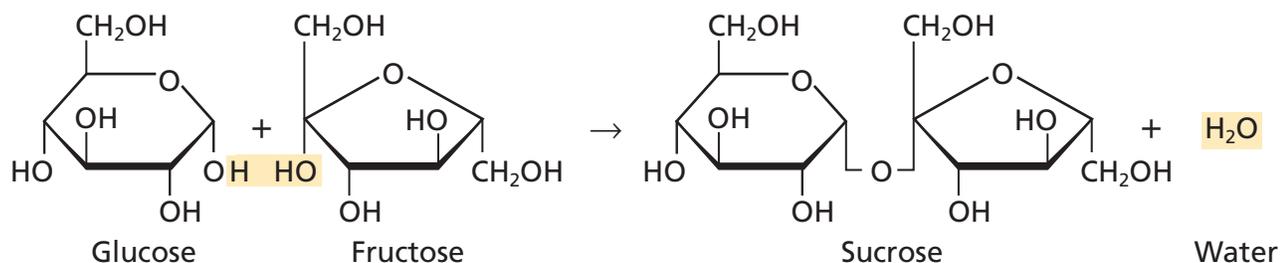


Figure 24-9

When glucose and fructose bond together, the disaccharide sucrose is formed. Note that water also is a product of this condensation reaction. Remember that each ring structure is made of carbon atoms, which are not shown for simplicity.

Disaccharides Like amino acids, monosaccharides can be linked together by a condensation reaction in which water is released. When two monosaccharides bond together, a **disaccharide** is formed. See **Figure 24-9**. The new bond formed is an ether functional group (C — O — C). Where does the water that is produced in this reaction come from?

One common disaccharide is sucrose, which is also known as table sugar because sucrose is used mainly as a sweetener. Sucrose is formed by linking glucose and fructose. Another common disaccharide is lactose, the most important carbohydrate in milk. It often is called milk sugar. Lactose is formed when glucose and galactose bond together. Some foods that contain common disaccharides are shown in **Figure 24-10a**.

Disaccharides are too large to be absorbed into the bloodstream from the human digestive system, so they must first be broken down into monosaccharides. This is accomplished by a number of enzymes in the digestive system, one of which is sucrase in the small intestine. Sucrase breaks down sucrose into glucose and fructose. Another enzyme, called lactase, breaks down lactose into glucose and galactose. Some people do not have an active form of the enzyme lactase. Unless they ingest lactase enzyme along with foods that contain milk, they experience gas, bloating, and much discomfort as lactose accumulates in their digestive systems. These people are said to be lactose intolerant.

Polysaccharides You may have seen large carbohydrate polymers referred to as complex carbohydrates in nutritional references. Another name for a complex carbohydrate is **polysaccharide**, which is a polymer of simple sugars that contains 12 or more monomer units. The same type of bond that joins two monosaccharides in a disaccharide links them together in a polysaccharide. Pasta is a good source of polysaccharides. The pasta shown in **Figure 24-10b** contains large amounts of starch, a polysaccharide from plants.

Figure 24-10

- a** When you have a snack of milk and cookies, you are ingesting the disaccharides lactose and sucrose.
- b** Pasta is rich in the polysaccharide starch.



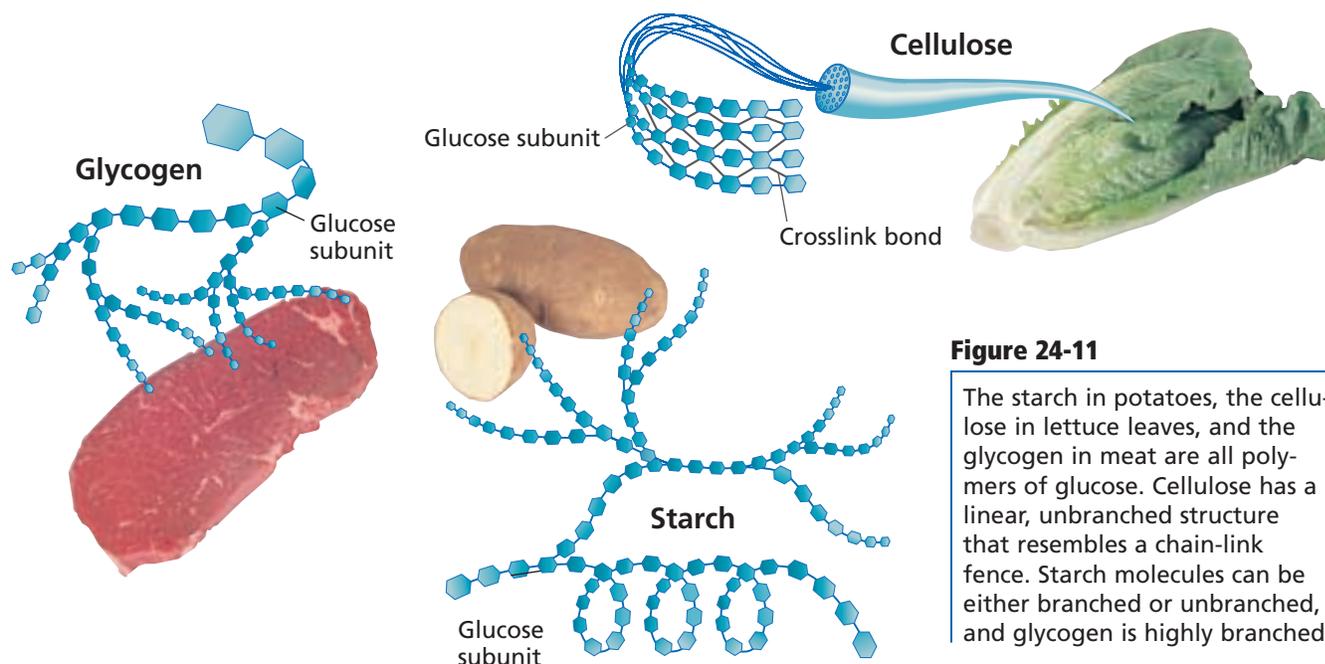


Figure 24-11

The starch in potatoes, the cellulose in lettuce leaves, and the glycogen in meat are all polymers of glucose. Cellulose has a linear, unbranched structure that resembles a chain-link fence. Starch molecules can be either branched or unbranched, and glycogen is highly branched.

Three important polysaccharides are starch, cellulose, and glycogen. All are composed solely of glucose monomers, as **Figure 24-11** illustrates. However, that's their only similarity, as all three have different properties and functions. Plants make both starch and cellulose. Starch is a soft, water-soluble molecule used to store energy, whereas cellulose is a water-insoluble polymer that forms rigid plant cell walls such as those found in wood. Glycogen is the animal counterpart of starch. It is made by animals to store energy, mostly in the liver and muscles.

How can these three polymers all be made solely of glucose monomers yet have such different properties? The answer lies in the way the bonds that link the monomers together are oriented in space. Because of this difference in bond shape, humans can digest starch but not cellulose. Digestive enzymes can't fit cellulose into their active sites due to the specific lock-and-key fit needed for enzyme action. As a result, the cellulose in the fruits, vegetables, and grains that we eat is labeled "dietary fiber" because it passes through the digestive system largely unchanged.

Section 24.2 Assessment

- Compare the structures of monosaccharides, disaccharides, and polysaccharides. Which has the largest molecular mass? The smallest?
- What is the main function of carbohydrates in living organisms?
- Compare and contrast the structures of starch and cellulose. How do the structural differences affect our ability to digest these two polysaccharides?
- Thinking Critically** If a carbohydrate has 2^n possible isomers, where n is equal to the number of carbon atoms in the carbohydrate structure that are chiral (meaning a carbon atom with four different groups bonded to it), calculate the number of possible isomers for each of the following monosaccharides. Use **Figure 24-8** to help you.
 - galactose
 - glucose
 - fructose
- Interpreting Scientific Illustrations** Draw the structure of the disaccharide sucrose, circling the ether functional group that bonds the monomer sugars together.

Objectives

- **Describe** the structures of fatty acids, triglycerides, phospholipids, and steroids.
- **Explain** the functions of lipids in living organisms.
- **Identify** some reactions that fatty acids undergo.
- **Relate** the structure and function of cell membranes.

Vocabulary

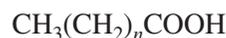
lipid
fatty acid
triglyceride
saponification
phospholipid
wax
steroid

The wax that you use to polish your car, the fat that drips out of your burger, and the vitamin D that fortifies the milk you drank for lunch—what do these diverse compounds have in common? They are all lipids.

What is a lipid?

A **lipid** is a large, nonpolar, biological molecule. Because lipids are nonpolar, they are insoluble in water. Lipids have two major functions in living organisms. They store energy efficiently, and they make up most of the structure of cell membranes. Unlike proteins and carbohydrates, lipids are not polymers with repeated monomer subunits.

Fatty acids Although lipids are not polymers, many lipids have a major building block in common. This building block is the **fatty acid**, a long-chain carboxylic acid. Most naturally occurring fatty acids contain between 12 and 24 carbon atoms. Their structure can be represented by the formula



Most fatty acids have an even number of carbon atoms, which is a result of their being constructed two carbons at a time in enzymatic reactions.

Fatty acids can be grouped into two main categories depending on the presence or absence of double bonds between carbon atoms. Fatty acids that contain no double bonds are referred to as saturated. Those that have one or more double bonds are called unsaturated. **Figure 24-12** shows the structures of two common fatty acids. Stearic acid is an 18-carbon saturated fatty acid; oleic acid is an 18-carbon unsaturated fatty acid. What makes oleic acid unsaturated?

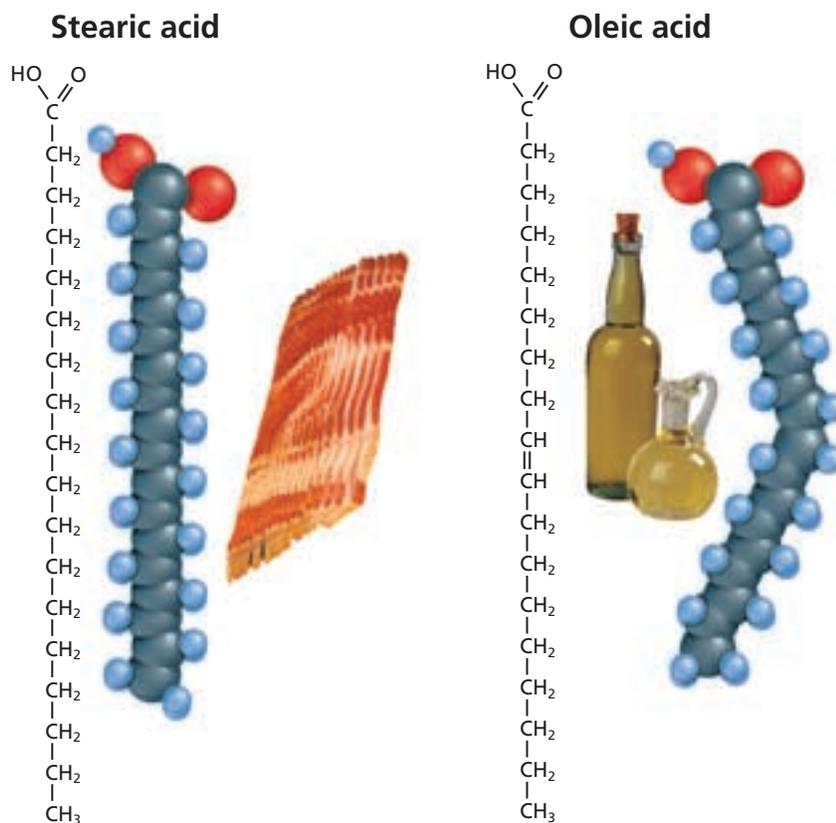
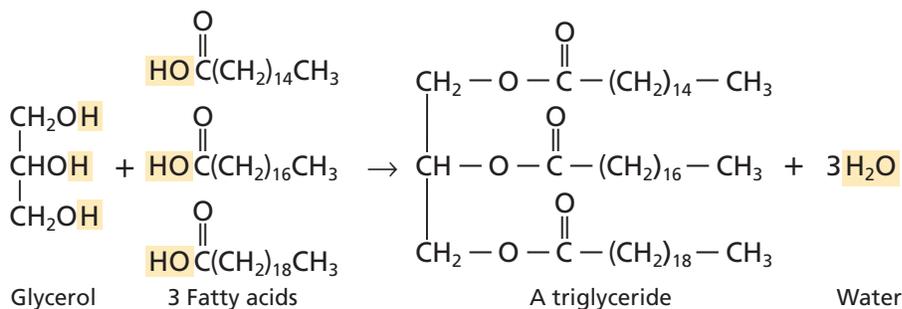


Figure 24-12

Two fatty acids abundant in our diets are the 18-carbon saturated stearic acid and the 18-carbon unsaturated oleic acid. How is the structure of the molecule affected by the presence of a double bond? Refer to **Table C-1** in Appendix C for a key to atom color conventions.



An unsaturated fatty acid can become saturated if it reacts with hydrogen. As you learned in Chapter 23, hydrogenation is an addition reaction in which hydrogen gas reacts with carbon atoms that are linked by multiple bonds. Each unsaturated carbon atom can pick up one hydrogen atom to become saturated. For example, oleic acid can be hydrogenated to form stearic acid.

The double bonds in naturally occurring fatty acids are almost all in the *cis* geometric isomer form. Recall from Chapter 22 that the *cis* isomer has identical groups oriented on the same side of the molecule around a double bond. Because of the *cis* orientation, unsaturated fatty acids have a kink, or bend, in their structure that prevents them from packing together. They don't form as many intermolecular attractions as saturated fatty acid molecules. As a result, unsaturated fatty acids have lower melting points.

Triglycerides Although fatty acids are abundant in living organisms, they are rarely found alone. They most often are found bonded to glycerol, a molecule with three carbons each containing a hydroxyl group. When three fatty acids are bonded to a glycerol backbone through ester bonds, a **triglyceride** is formed. The formation of a triglyceride is shown in **Figure 24-13**.

Triglycerides can be either solids or liquids at room temperature. If liquid, they are usually called oils. If solid at room temperature, they're called fats. Most mixtures of triglycerides from plant sources, such as corn, olive, and peanut oils, are liquids because the triglycerides contain unsaturated fatty acids that have fairly low melting points. Animal fats, such as butter, contain a larger proportion of saturated fatty acids. They have higher melting points and usually are solids at room temperature.

Fatty acids are stored in fat cells of your body as triglycerides. When energy is abundant, fat cells store the excess energy in the fatty acids of triglycerides. When energy is scarce, the cells break down the triglycerides, forming free fatty acids and glycerol. Further breakdown of the fatty acids releases the energy used to form them. Although enzymes break down triglycerides in living cells, the reaction can be duplicated outside of cells by using a strong base such as sodium hydroxide. This reaction—the hydrolysis of a triglyceride using an aqueous solution of a strong base to form carboxylate salts and glycerol—is **saponification**, shown below.

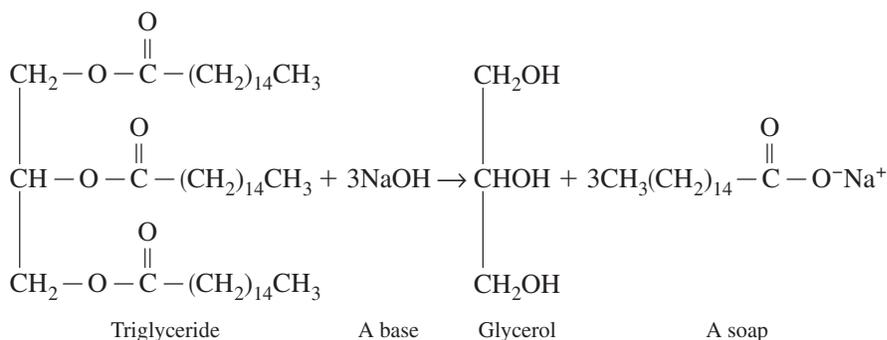


Figure 24-13

Ester bonds in a triglyceride are formed when the hydroxyl groups of glycerol combine with the carboxyl groups of the fatty acids.

Biology

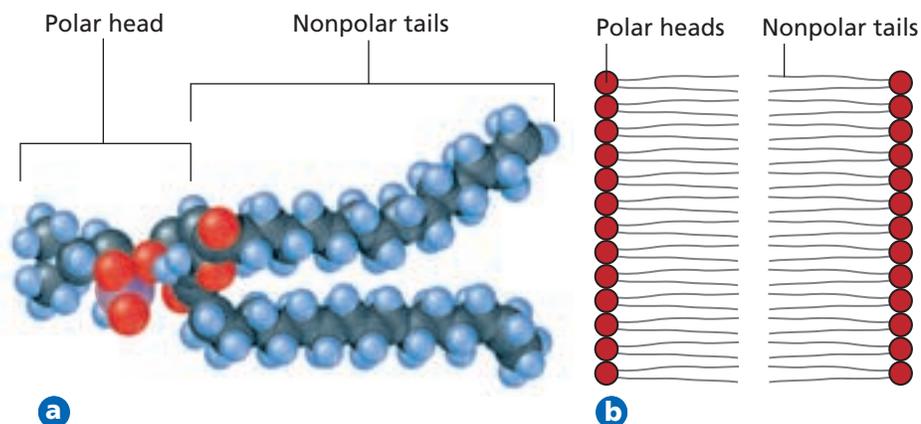
CONNECTION

The venom of poisonous snakes contains a class of enzymes known as phospholipases. These enzymes catalyze the breakdown of phospholipids, triglycerides in which one fatty acid has been replaced by a phosphate group. The venom of the eastern diamondback rattlesnake contains a phospholipase that hydrolyzes the ester bond at the middle carbon of phospholipids. If the larger of the two breakdown products of this reaction gets into the bloodstream, it dissolves the membranes of red blood cells, causing them to rupture. A bite from the eastern diamondback can lead to death if not treated immediately.



Figure 24-14

- a** A phospholipid has a polar head and two nonpolar tails.
- b** The membranes of living cells are formed by a double layer of lipids called a bilayer. The polar heads face out of both sides of the bilayer, where they are in contact with the watery environment inside and outside the cell. The nonpolar tails point into the center of the bilayer.



Saponification is used to make soaps, which usually are the sodium salts of fatty acids. A soap has both a polar end and a nonpolar end. Soaps can be used to clean nonpolar dirt and oil with water because the nonpolar dirt and oil bond to the nonpolar end of the soap molecules, and the polar end of the soap molecules is soluble in water. Thus, the dirt-laden soap molecules can be rinsed away with the water. You can make soap by doing the **miniLAB** on this page.

Phospholipids Another important type of triglyceride, phospholipids, are found in greatest abundance in cellular membranes. A **phospholipid** is a triglyceride in which one of the fatty acids is replaced by a polar phosphate group. As you can see in **Figure 24-14a**, the polar part of the molecule forms

miniLAB

A Saponification Reaction

Applying Concepts The reaction between a triglyceride and a strong base such as sodium hydroxide is called saponification. In this reaction, the ester bonds in the triglyceride are hydrolyzed by the base. The sodium salts of the fatty acids, called soaps, precipitate out, and glycerol is left in solution.

Materials solid vegetable shortening, 250-mL beaker, 600-mL beaker, 6.0M NaOH, ethanol, saturated NaCl solution, stirring rod, hot plate, tongs, 25-mL graduated cylinder, evaporating dish, cheesecloth (20 cm × 20 cm), funnel

Procedure



- Place a 250-mL beaker on the hot plate. Add 25 g solid vegetable shortening to the beaker. Turn the hot plate on at a medium setting.
- As the vegetable shortening melts, slowly add 12 mL ethanol and then 5 mL 6.0M NaOH to the beaker. **CAUTION:** Ethanol is flammable. NaOH causes skin burns. Wear gloves.

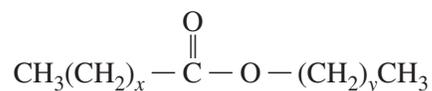
- Heat the mixture, stirring occasionally, for about 15 minutes, but do not allow it to boil.
- When the mixture begins to thicken, use tongs to remove the beaker from the heat. Allow the beaker to cool for five minutes, then place it in a cold water bath in the 600-mL beaker.
- Add 25 mL saturated NaCl solution to the mixture in the beaker. The soap is not very soluble and will appear as small clumps.
- Collect the solid soap clumps by filtering them through a cheesecloth-lined funnel.
- Using gloved hands, press the soap into an evaporating dish. Allow the soap to air dry for one or two days.
- Remove your gloves and wash your hands.

Analysis

- What type of bonds present in the triglycerides are broken during the saponification reaction?
- What is the common name for the sodium salt of a fatty acid?
- How does soap remove dirt from a surface?
- Write a word equation for the saponification reaction in this lab.

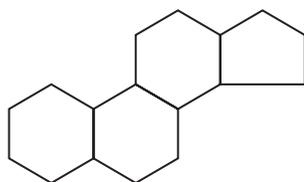
a “head,” and the nonpolar fatty acids look like tails. How are the polar and nonpolar parts of a phospholipid arranged in the membranes of cells? A typical cell membrane has two layers of phospholipids, which are arranged with their nonpolar tails pointing inward and their polar heads pointing outward. See **Figure 24-14b**. This arrangement is called a lipid bilayer. Because the lipid bilayer structure acts as a barrier, the cell is able to regulate the materials that enter and leave through the membrane.

Waxes Another type of lipid, waxes, also contain fatty acids. A **wax** is a lipid that is formed by combining a fatty acid with a long-chain alcohol. The general structure of these soft, solid fats with low melting points is shown below, with x and y representing variable numbers of CH_2 groups.



Both plants and animals make waxes. Plant leaves are often coated with wax, which prevents water loss. Notice in **Figure 24-15** how raindrops “bead up” on the leaves of a plant, indicating the presence of the waxy layer. The honeycomb of bees also is made of a wax, commonly called beeswax. Combining the 16-carbon fatty acid palmitic acid and a 30-carbon alcohol chain makes a common form of beeswax. Candles are sometimes made of beeswax because it burns slowly and evenly.

Steroids Not all lipids contain fatty acid chains. **Steroids** are lipids that have multiple cyclic rings in their structures. All steroids are built from the basic four-ring steroid structure shown below.



Some hormones, such as many sex hormones, are steroids that function to regulate metabolic processes. Cholesterol, another steroid, is an important structural component of cell membranes. Vitamin D also contains the four-ring steroid structure and plays a role in the formation of bones.

Figure 24-15

Plants produce wax that coats their leaves (top). The wax protects the leaves from drying out. The honeycomb of a beehive is constructed from beeswax (bottom).



Section 24.3 Assessment

11. Write the equation for the complete hydrogenation of the polyunsaturated fatty acid linoleic acid, $\text{CH}_3(\text{CH}_2)_4\text{CH}=\text{CHCH}_2\text{CH}=\text{CH}(\text{CH}_2)_7\text{COOH}$.
12. List an important function of each of these types of lipids.

a. triglycerides	c. waxes
b. phospholipids	d. steroids
13. Compare and contrast the structures of a steroid, a phospholipid, a wax, and a triglyceride.
14. **Thinking Critically** What possible solvent can be used to extract lipids from cell membranes?
15. **Interpreting Scientific Illustrations** Draw the general structure of a phospholipid. Label the polar and nonpolar portions of the structure.

Objectives

- **Identify** the structural components of nucleic acids.
- **Relate** the function of DNA to its structure.
- **Describe** the structure and function of RNA.

Vocabulary

nucleic acid
nucleotide

Nucleic acids are the fourth class of biological molecules that you will study. They are the information-storage molecules of the cell. This group of nitrogen-containing molecules got its name from the cellular location in which the molecules are primarily found—the nucleus. It is from this control center of cells that nucleic acids carry out their major functions.

Structure of Nucleic Acids

A **nucleic acid** is a nitrogen-containing biological polymer that is involved in the storage and transmission of genetic information. The monomer that makes up a nucleic acid is called a **nucleotide**. Each nucleotide has three parts: an inorganic phosphate group, a five-carbon monosaccharide sugar, and a nitrogen-containing structure called a nitrogen base. Examine each part of **Figure 24-16a**. Although the phosphate group is the same in all nucleotides, the sugar and the nitrogen base vary.

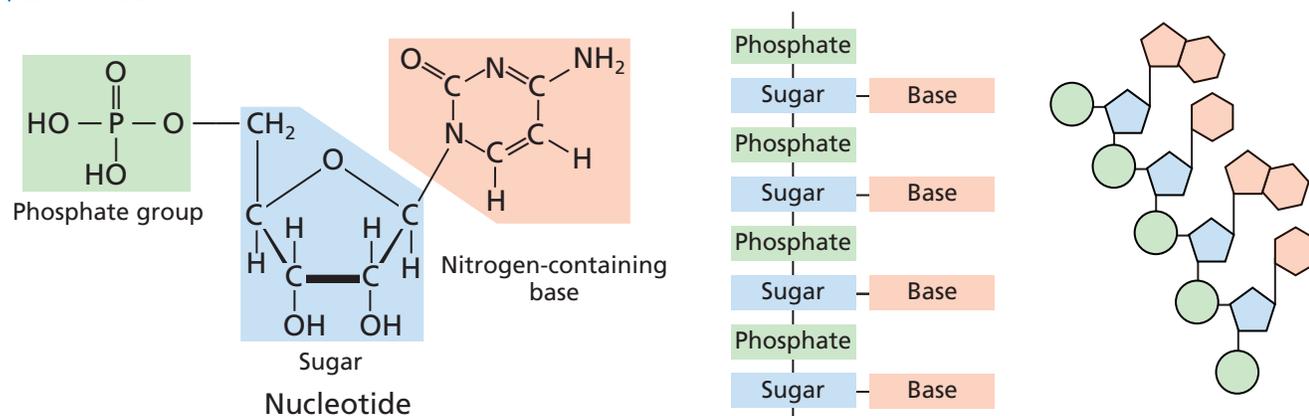
In a nucleic acid, the sugar of one nucleotide is bonded to the phosphate of another nucleotide, as shown in **Figure 24-16b**. Thus, the nucleotides are strung together in a chain, or strand, containing alternating sugar and phosphate groups. Each sugar is also bonded to a nitrogen base that sticks out from the chain. The nitrogen bases on adjoining nucleotide units are stacked one above the other in a slightly askew position, much like the steps in a staircase. You can see this orientation in **Figure 24-16c**. Intermolecular forces hold each nitrogen base close to the nitrogen bases above and below it.

DNA: The Double Helix

You've probably heard of DNA (deoxyribonucleic acid), one of the two kinds of nucleic acids found in living cells. DNA contains the master plans for building all the proteins in an organism's body.

Figure 24-16

Nucleotides are the monomers from which nucleic acid polymers are formed.



The structure of DNA The structure of DNA consists of two long chains of nucleotides wound together to form a spiral structure. Each nucleotide in DNA contains a phosphate group, the five-carbon sugar deoxyribose, and a nitrogen base. The alternating sugar and phosphate groups in each chain make up the outside, or backbone, of the spiral structure. The nitrogen bases are on the inside. Because the spiral structure is composed of two chains, it is known as a double helix. You can see from **Figure 24-17** that the DNA molecule is similar to a zipper in which the two ends have been twisted in opposite directions. The two sugar-phosphate backbones form the outsides of the zipper. What forms the zipper's teeth?

DNA contains four different nitrogen bases: adenine (A), thymine (T), cytosine (C), and guanine (G). As **Figure 24-18** shows, both adenine and guanine contain a double ring. Thymine and cytosine are single-ring structures. Looking back at **Figure 24-17**, you can see that each nitrogen base on one strand of the helix is oriented next to a nitrogen base on the opposite strand, in the same way that the teeth of a zipper are oriented. The side-by-side base pairs are close enough so that hydrogen bonds form between them. Because each nitrogen base has a unique arrangement of organic functional groups that can form hydrogen bonds, the nitrogen bases always pair in a specific way so that the optimum number of hydrogen bonds form. As **Figure 24-18** shows, guanine always binds to cytosine, and adenine always binds to thymine. The G—C and A—T pairs are called complementary base pairs.

It is because of complementary base pairing that the amount of adenine in a molecule of DNA always equals the amount of thymine, and the amount of cytosine always equals the amount of guanine. In 1953, James Watson and Francis Crick used this observation to make one of the greatest scientific discoveries of the twentieth century when they determined the double-helix structure of DNA. They accomplished this feat without actually carrying out many laboratory experiments themselves. Instead, they analyzed and synthesized the work of numerous scientists who had carefully carried out studies on DNA. The X-ray diffraction patterns of DNA fibers taken by Maurice Wilkins and Rosalind Franklin were of special importance because they clearly showed the dimensions of the DNA molecule and the molecule's helical structure. Watson and Crick used these results and molecular modeling techniques to build their DNA structure with balls and sticks. This discovery illustrates the importance of being able to visualize molecules in order to uncover patterns in bonding arrangements. How might such molecular modeling be done today?

The function of DNA Watson and Crick used their model to predict how DNA's chemical structure enables it to carry out its function. DNA stores the genetic information of a cell in the cell's nucleus. Before the cell divides, the DNA is copied so that the new generation of cells gets the same



Figure 24-17

The structure of DNA is a double helix that resembles a twisted zipper.

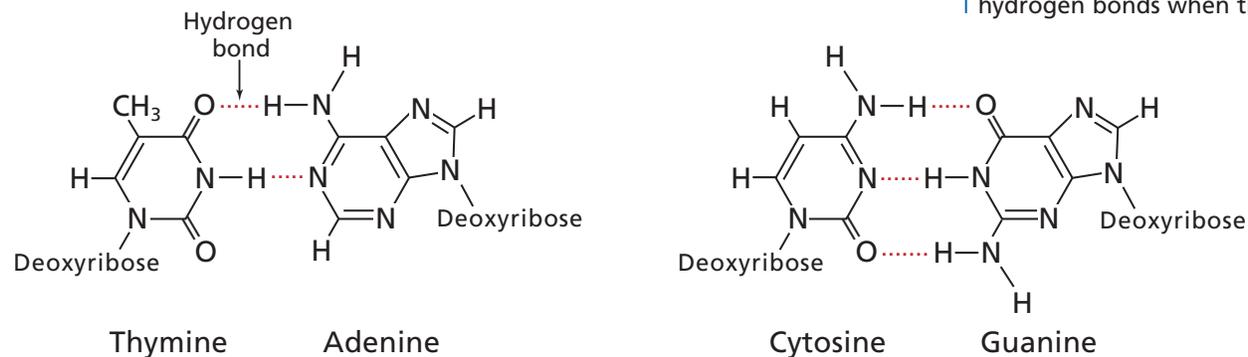


Figure 24-18

In DNA, base pairing exists between a double-ringed base and a single-ringed base. Adenine and thymine always pair, forming two hydrogen bonds between them. Guanine and cytosine always form three hydrogen bonds when they pair.

genetic information. Having determined that the two chains of the DNA helix are complementary, Watson and Crick realized that complementary base pairing provides a mechanism by which the genetic material of a cell is copied. The **problem-solving LAB** shows how DNA copies, or replicates, itself.

The four nitrogen bases of DNA serve as the letters of the alphabet in the information-storage language of living cells. The specific sequence of these letters represents an organism's master instructions, just as the sequence of letters in the words of this sentence convey special meaning. The sequence of bases is different in every species of organism, allowing for an enormous diversity of life forms—all from a language that uses only four letters. It is estimated that the DNA in a human cell has about three billion complementary base pairs, arranged in a sequence unique to humans.

problem-solving LAB

How does DNA replicate?

Formulating Models DNA must be copied, or replicated, before a cell can divide so that each of the two new cells that are formed by cell division has a complete set of genetic instructions. It is very important that the replicating process be accurate; the new DNA molecules must be identical to the original. Watson and Crick noticed that their model for the three-dimensional structure of DNA provided a mechanism for accurate replication.

When DNA begins to replicate, the two nucleotide strands start to unzip. An enzyme breaks the hydrogen bonds between the nitrogen bases, and the strands separate to expose the nitrogen bases. Other enzymes deliver free nucleotides from the surrounding medium to the exposed nitrogen bases, adenine hydrogen-bonding with thymine and cytosine bonding with guanine.

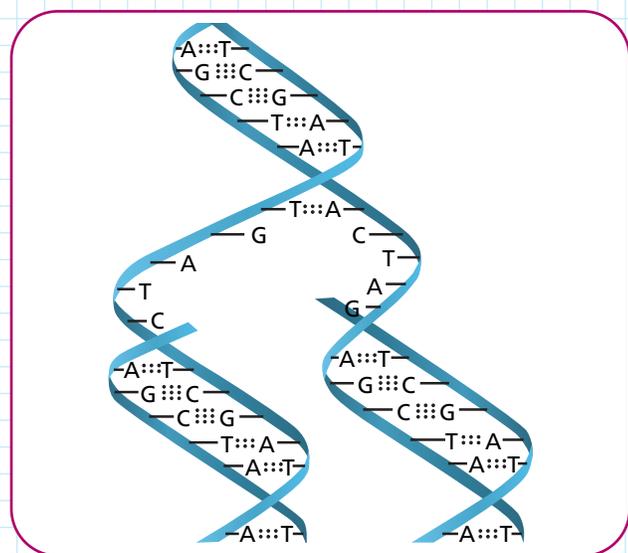
Thus, each strand builds a complementary strand by base pairing with free nucleotides. This process is shown in **Diagram a**. When the free nucleotides have been hydrogen-bonded into place, their sugars and phosphates bond covalently to those on adjacent nucleotides to form the new backbone. Each strand of the original DNA molecule is now bonded to a new strand.

Analysis

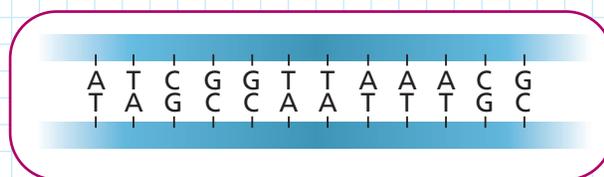
Diagram b shows a small segment of a DNA molecule. Copy the base sequence onto a clean sheet of paper, being careful not to make copying errors. Show the steps of replication to produce two segments of the DNA.

Thinking Critically

1. How does the base sequence of a newly synthesized strand compare with the original strand to which it is bonded?
2. If the original DNA segment is colored red and the free nucleotides are colored blue, what pattern of colors will the newly replicated DNA segments have? Will all new segments have the same color pattern? What does this pattern mean?
3. What would happen if you had made an error when copying the base sequence onto your sheet of paper? How might an organism be affected if this error had occurred during replication of its DNA?



a



b

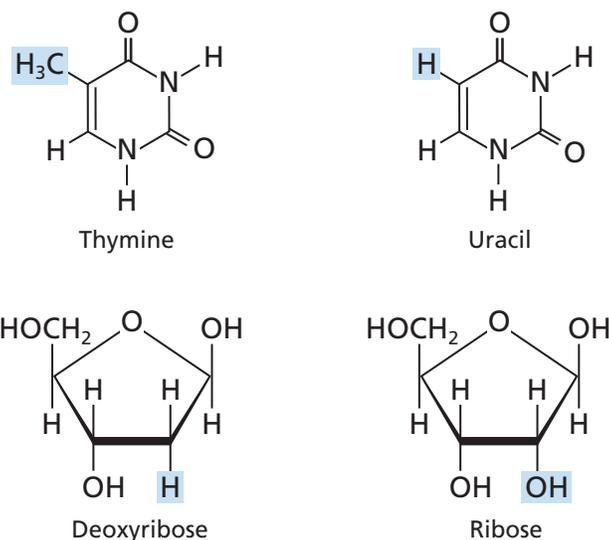


Figure 24-19

DNA and RNA are nucleic acids with some important differences in structure. DNA contains thymine; RNA contains uracil, which lacks a methyl group. DNA's sugar, deoxyribose, has a hydrogen in one position instead of the hydroxyl group that ribose has.

RNA

RNA (ribonucleic acid) also is a nucleic acid. Its general structure differs from that of DNA in three important ways. As you know, DNA contains the nitrogen bases adenine, cytosine, guanine, and thymine. RNA contains adenine, cytosine, guanine, and uracil; thymine is never found in RNA. Secondly, RNA contains the sugar ribose. DNA contains the sugar deoxyribose, which has a hydrogen atom in place of a hydroxyl group at one position. Compare these different structures shown in **Figure 24-19**.

The third difference between DNA and RNA arises as a result of these structural differences. DNA is normally arranged in a double helix in which hydrogen bonding links the two chains together through their bases. RNA is usually single stranded, with no such hydrogen bonds forming.

Whereas DNA functions to store genetic information, RNA allows cells to use the information found in DNA. You have learned that the genetic information of a cell is contained in the sequence of nitrogen bases in the DNA molecule. Cells use this base sequence to make RNA with a corresponding sequence. The RNA is then used to make proteins, each with an amino acid sequence that is determined by the order of nitrogen bases in RNA. The translation of the base-sequence language of nucleic acids into the amino acid-sequence language of proteins is known as the genetic code. Because proteins are the molecular tools that actually carry out most activities in a cell, the DNA double helix is ultimately responsible for controlling the thousands of chemical reactions that take place.

Section 24.4 Assessment

- List the three chemical structures that make up a nucleotide.
- Compare and contrast the structures and functions of DNA and RNA.
- A sample of nucleic acid was determined to contain adenine, uracil, cytosine, and guanine. What type of nucleic acid is this?
- Thinking Critically** Analyze the structure of nucleic acids to determine what structural feature makes them acidic.
- Predicting** What is the genetic code? Predict what might happen to a protein if the DNA that coded for the protein contained the wrong base sequence.

Objectives

- **Distinguish** between anabolism and catabolism.
- **Describe** the role of ATP in metabolism.
- **Compare** and **contrast** the processes of photosynthesis, cellular respiration, and fermentation.

Vocabulary

metabolism
catabolism
anabolism
ATP
photosynthesis
cellular respiration
fermentation

You've now studied the four major kinds of biological molecules and learned that they all are present in the food you eat. What happens to these molecules after they enter your body?

Anabolism and Catabolism

Many thousands of chemical reactions take place in the cells of a living organism. The set of reactions carried out by an organism is its **metabolism**. Why are so many reactions involved in metabolism? Living organisms must accomplish two major functions in order to survive. They have to extract energy from nutrients in forms that they can use immediately as well as store for future use. In addition, they have to use nutrients to make building blocks for synthesizing all of the molecules needed to carry out their life functions. These processes are summarized in **Figure 24-20**.

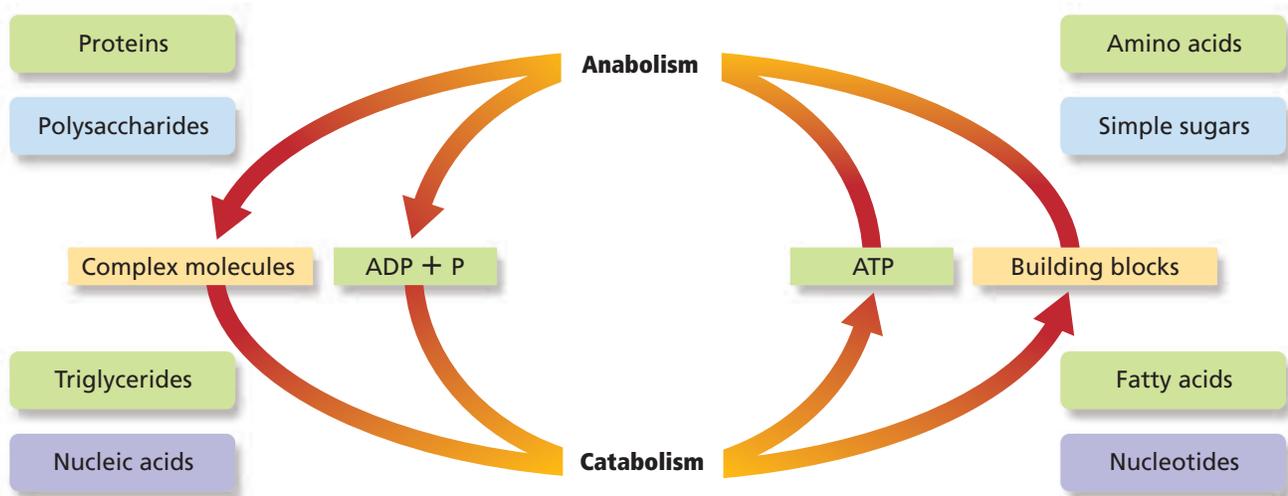
The term **catabolism** refers to the metabolic reactions that break down complex biological molecules such as proteins, polysaccharides, triglycerides, and nucleic acids for the purposes of forming smaller building blocks and extracting energy. After you eat a meal of spaghetti and meatballs, your body immediately begins to break down the starch polymer in the spaghetti into glucose. The glucose is then broken down into smaller molecules in a series of energy-releasing catabolic reactions. Meanwhile, the protein polymers in the meatballs are catabolized into amino acids.

The term **anabolism** refers to the metabolic reactions that use energy and small building blocks to synthesize the complex molecules needed by an organism. After your body has extracted the energy from the starch in pasta, it uses that energy and the amino-acid building blocks produced from the meat proteins to synthesize the specific proteins that allow your muscles to contract, catalyze metabolic reactions, and carry out many other functions in your body.

Figure 24-20 shows the relationship between catabolism and anabolism. The simple building-block molecules that are listed on the right side of the diagram are used to build the complex molecules that are listed on the left side of the diagram. As the arrow moves from right to left, anabolic reactions

Figure 24-20

A large number of different metabolic reactions take place in living cells. Some involve breaking down nutrients to extract energy; these are catabolic processes. Others involve using energy to build large biological molecules; these reactions are anabolic processes.



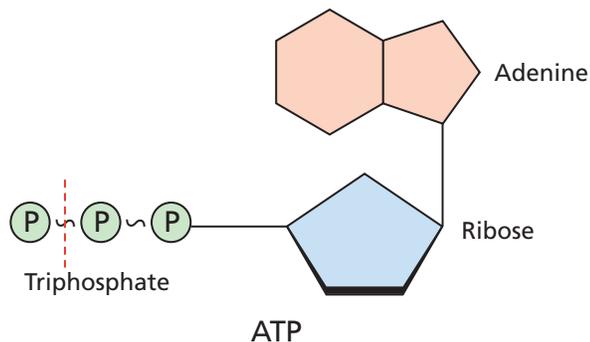


Figure 24-21

ATP is a nucleotide that contains an adenine nitrogen base, a ribose sugar, and three phosphate groups. When the final phosphate group is removed from ATP, as modeled by the red dotted line, ADP is formed and energy is released.

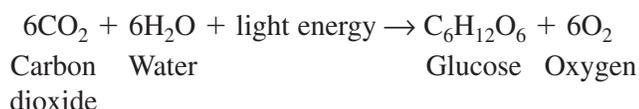
take place. As the lower arrow moves from left to right, catabolic processes take place and the complex molecules are broken down into their smaller building blocks.

ATP Catabolism and anabolism are linked by common building blocks that catabolic reactions produce and anabolic reactions use. A common form of potential chemical energy also links the two processes. **ATP** (adenosine triphosphate) is a nucleotide that functions as the universal energy-storage molecule in living cells. During catabolic reactions, cells harness the chemical energy of foods and store it in the bonds of ATP. When these bonds are broken, the chemical energy is released and used by cells to drive anabolic reactions that might not otherwise occur. Most cellular reactions have an efficiency of only about 40 percent at best; the remaining 60 percent of the energy in food is lost as heat, which is used to keep your body warm.

The structure of ATP is shown in **Figure 24-21**. During catabolic reactions, cells produce ATP by adding an inorganic phosphate group to the nucleotide adenosine diphosphate (ADP) in an endothermic reaction. One mole of ATP stores approximately 30.5 kJ of energy under normal cellular conditions. During anabolism, the reverse reaction occurs. ATP is broken down to form ADP and inorganic phosphate in an exothermic reaction. Approximately 30.5 kJ of energy is released from each mole of ATP.

Photosynthesis

What is the source of the energy that fuels metabolism? For most living things, certain wavelengths of sunlight provide all of this energy. Some bacteria and the cells of all plants and algae, including the brown algae shown in **Figure 24-22**, are able to capture light energy and convert some of it to chemical energy. Animals can't capture light energy, so they get energy by eating plants or by eating other animals that eat plants. The process that converts energy from sunlight to chemical energy in the bonds of carbohydrates is called **photosynthesis**. During the complex process of photosynthesis, carbon dioxide and water provide the carbon, hydrogen, and oxygen atoms that make up carbohydrates and oxygen gas, which also is formed. The following net reaction takes place during photosynthesis.



Photosynthesis results in the reduction of the carbon atoms in carbon dioxide as glucose is formed. During this redox process, oxygen atoms in water are oxidized to oxygen gas.

Figure 24-22

These giant seaweeds, called kelps, must grow close to the ocean's surface where light is available for photosynthesis. They cannot grow in the dark depths where sunlight does not penetrate the seawater.



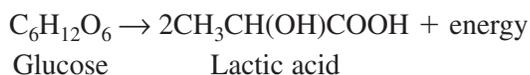


Figure 24-24

Bread dough is heavy before yeast cells begin to ferment the carbohydrates (left). After fermentation has produced carbon dioxide, the dough becomes lighter and fluffier as it rises (right).

This reaction, called alcoholic fermentation, is important to certain segments of the food industry, as you can see in **Figure 24-24**. Alcoholic fermentation is needed to make bread dough rise, form tofu from soybeans, and produce the ethanol in alcoholic beverages. Another use of the ethanol that is produced by yeast is as an additive to gasoline, called gasohol. You can observe yeast cells fermenting sugar in the **CHEMLAB** at the end of this chapter.

Lactic acid fermentation Have you ever gotten muscle fatigue while running a race? During strenuous activity, muscle cells often use oxygen faster than it can be supplied by the blood. When the supply of oxygen is depleted, cellular respiration stops. Although animal cells can't undergo alcoholic fermentation, they can produce lactic acid and a small amount of energy from glucose through lactic acid fermentation.



The lactic acid that is produced is moved from the muscles through the blood to the liver. There, it is converted back into glucose that can be used in catabolic processes to yield more energy once oxygen becomes available. However, if lactic acid builds up in muscle cells at a faster rate than the blood can remove it, muscle fatigue results. Buildup of lactic acid is what causes a burning pain in the muscle during strenuous exercise.

Section 24.5 Assessment

- 21.** Compare and contrast the processes of anabolism and catabolism.
- 22.** Explain the role ATP plays in the metabolism of living organisms.
- 23.** Decide whether each of the following processes is anabolic or catabolic.
 - a. photosynthesis
 - b. cellular respiration
 - c. fermentation
- 24. Thinking Critically** Why is it necessary to use sealed casks when making wine?
- 25. Calculating** How many moles of ATP would a yeast cell produce if six moles of glucose were oxidized completely in the presence of oxygen? How many moles of ATP would the yeast cell produce from six moles of glucose if the cell were deprived of oxygen? For more help, refer to *Arithmetic Operations* in the **Math Handbook** on page 887 of this textbook.

Alcoholic Fermentation in Yeast

Yeast cells are able to metabolize many types of sugars. In this experiment, you will observe the fermentation of sugar by baker's yeast. When yeast cells are mixed with a sucrose solution, they must first hydrolyze the sucrose to glucose and fructose. Then the glucose is broken down in the absence of oxygen to form ethanol and carbon dioxide. You can test for the production of carbon dioxide by using a CBL pressure sensor to measure an increase in pressure.

Problem

What is the rate of alcoholic fermentation of sugar by baker's yeast?

Objectives

- **Measure** the pressure of carbon dioxide produced by the alcoholic fermentation of sugar by yeast.
- **Calculate** the rate of production of carbon dioxide by the alcoholic fermentation of sugar by yeast.

Materials

CBL system	stirring rod
graphing calculator	600-mL beaker
ChemBio program	thermometer
Vernier pressure sensor	basting bulb
link cable	hot and cold water
CBL-DIN cable	yeast suspension
test tube with #5 rubber-stopper assembly	vegetable oil
5% sucrose solution	utility clamp
ring stand	10-mL graduated cylinders (2)
	pipette

Safety Precautions



- Always wear safety goggles and an apron in the lab.
- Do not use the thermometer as a stirring rod.

Pre-Lab

1. Reread the section of this chapter that describes alcoholic fermentation.
2. Write the chemical equation for the alcoholic fermentation of glucose.
3. Read the entire **CHEMLAB**.
4. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions and procedure notes.
5. Form a hypothesis about how the pressure inside the test tube is related to the production of carbon dioxide during the reaction. Refer to the ideal gas law in your explanation.
6. Why is temperature control an essential feature of the **CHEMLAB**?

Procedure

1. Load the ChemBio program into your graphing calculator. Connect the CBL and calculator with the link cable. Connect the pressure sensor to the CBL with a CBL-DIN cable.
2. Prepare a water bath using the 600-mL beaker. The beaker should be about two-thirds full of water. The water temperature should be between 36°C and 38°C.
3. Set up the test tube, ring stand, and utility clamp as shown in the figure. Obtain about 3 mL yeast suspension in a 10-mL graduated cylinder, and pour it into the test tube. Obtain about 3 mL 5% sucrose solution in a 10-mL graduated cylinder. Add the sucrose solution to the yeast in the test tube. Stir to mix. Pour enough vegetable oil on top of the mixture to completely cover the surface.

- Place the stopper assembly into the test tube. Make sure it has an airtight fit. Leave both valves of the assembly open to the atmosphere.
- While one lab partner does step 5, the other partner should do steps 6 and 7. Lower the test tube into the water bath and allow it to incubate for 10 minutes. Keep the temperature of the water bath between 36°C and 38°C by adding small amounts of hot or cold water with the basting bulb as needed.



- Start the ChemBio program. Choose 1:SET UP PROBES under MAIN MENU. Choose 1 for number of probes. Choose 3:PRESSURE under SELECT PROBE. Enter 1 for Channel. Choose 1:USE STORED for CALIBRATION. Choose 1:ATM for PRESSURE UNITS.
- Choose 2:COLLECT DATA under MAIN MENU. Choose 2:TIME GRAPH under DATA COLLECTION. Use time between sample seconds = 10. Use number of samples = 60. (This will give you 600 seconds or 10 minutes of data). Choose 1:USE TIME SETUP under CONTINUE? Set Ymin = 0.8, Ymax = 1.3, and Yscl = 0.1. Do not press ENTER until the test tube has finished incubating.
- After the test tube has incubated for ten minutes, close the valve attached to the stopper. Make sure the valve near the pressure sensor is open to the sensor. Start measuring the gas pressure by press-

ing ENTER. Monitor the pressure reading on the CBL unit. If the pressure exceeds 1.3 atm, the stopper can pop off. Open the air valve on the pressure sensor to release this excess gas pressure.

- After ten minutes, the data collection will stop. Open the air valve on the stopper. If needed, you can run a second trial by closing the air valve and choosing 2:YES to REPEAT? If you are finished, press 1:NO.

Cleanup and Disposal

- Rinse out and wash all items.
- Rinse the yeast suspension/sucrose/vegetable oil mixture down the sink with large amounts of water.
- Return all lab equipment to its proper place.

Analyze and Conclude

- Making and Using Graphs** Choose 3:VIEW GRAPH from the MAIN MENU. Make a sketch of the graph. (You also may want to record the data table by using 4:VIEW DATA.)
- Interpreting Data** The rate of carbon dioxide production by the yeast can be found by calculating the slope of the graph. Return to the MAIN MENU and choose 5:FIT CURVE. Choose 1:LINEAR L1, L2. The slope will be listed under LINEAR as "A" of $Y = A * X + B$. Record this value.
- Communicating** How does your rate of carbon dioxide production compare with the rates of other members of the class?
- Analyzing** Why did you add vegetable oil to the test tube in step 3?
- Error Analysis** Suppose that the pressure does not change during a trial. What might be some possible reasons for this?

Real-World Chemistry

- Yeast is used in baking bread because the carbon dioxide bubbles make the bread rise. The other product of alcoholic fermentation is ethanol. Why can't you taste this alcohol when you eat bread?
- How would the appearance of a loaf of bread be different if you used twice as much yeast as the recipe called for?

Everyday Chemistry

Sense of Smell

Take a deep breath. What do you smell? Maybe it's the eraser on your pencil, some flowers outside your window, an apple pie coming out of the oven, or even stinky gym shoes on the floor. Your sense of smell, otherwise known as olfaction, tells you a lot about the world around you.

Fitting In

How exactly do you smell odors? If you said that you smell with your nose, you are only partially correct. The nose you see when you look in the mirror doesn't actually detect odors. Its job is to pull odor molecules, known as odorants, into your nasal passages when you breathe.

The air you inhale carries odorants to a small region located high inside your nasal passage just below and between your eyes. The tissues of this region contain nerve cells that have hairlike fibers called cilia on one end.

When receptors located on the cilia are stimulated by odorants, the nerve cells send messages to the brain. The brain then identifies the odors you smell.

The exact manner in which receptors detect odorants is not entirely understood. However, researchers have found evidence to suggest that it involves the shapes of both the receptors and the odorants. Receptors are usually large, globular protein molecules that are similar in shape to enzymes. Their surfaces contain small crevices. An odorant must fit into and bind with a crevice on the surface of the receptor much as a substrate molecule fits into and binds with the active site on the surface of an enzyme. Each type of receptor has a uniquely shaped crevice on its surface that binds only with an odorant of a particular shape.

Not By Shape Alone

The process of smelling is not that simple, however. Scientists have discovered that shape is not the only property that determines whether an odorant will bind with a receptor. Odorant molecules must be able to travel through air and they must be able to dissolve in the fluids of the nasal passages. In addition, some research suggests that odor receptors can detect the energy levels of the odorants. An odorant with exactly the right energy will vibrate in such a way that it produces a response in the receptor. This may be what causes the nerve cell to send a signal to the brain.

So Many Odors

How many odors do you think you can recognize? If your olfactory system is at its best, you can distinguish about 10 000 different odors. Does this mean that you have more than 10 000 dif-

ferent types of olfactory receptors? The answer is no. In fact, you have fewer than 1000 different types of olfactory receptors.

As researchers searched for an explanation of how the olfactory receptors recognize so many different odors, they found that the olfactory system uses a combination of receptors for each odor. A single odorant can be recognized by more than one receptor because different receptors respond to different parts of the same odorant molecule. Rather than responding to a signal from a single receptor for each odor, the brain interprets a combination of signals from several receptors in order to identify a particular odor. In this way, a limited number of receptors can be used to identify a large number of odors.



Testing Your Knowledge

- 1. Relating Concepts** How is the receptor-odorant relationship similar to an enzyme-substrate relationship?
- 2. Applying** How might fragrance designers benefit from an understanding of how the shape of an odorant stimulates a receptor?

Summary

24.1 Proteins

- Proteins are biological polymers made of amino acids that are linked together by peptide bonds.
- Protein chains fold into intricate three-dimensional structures.
- Many proteins function as enzymes, which are highly specific and powerful biological catalysts. Other proteins function to transport important chemical substances, or provide structure in organisms.

24.2 Carbohydrates

- Monosaccharides, known as simple sugars, are aldehydes or ketones that also have multiple hydroxyl groups.
- Bonding two simple sugars together forms a disaccharide such as sucrose or lactose.
- Polysaccharides such as starch, cellulose, and glycogen are polymers of simple sugars.
- Carbohydrates function in living things to provide immediate and stored energy.

24.3 Lipids

- Fatty acids are long-chain carboxylic acids that usually have between 12 and 24 carbon atoms.
- Saturated fatty acids have no double bonds; unsaturated fatty acids have one or more double bonds.
- Fatty acids can be linked to glycerol backbones to form triglycerides.
- The membranes of living cells have a lipid bilayer structure.
- Steroids are lipids that have a multiple-ring structure.

24.4 Nucleic Acids

- Nucleic acids are polymers of nucleotides, which consist of a nitrogen base, a phosphate group, and a sugar.
- DNA contains deoxyribose sugar and the nitrogen bases adenine, cytosine, guanine, and thymine.
- RNA contains ribose sugar and the nitrogen bases adenine, cytosine, guanine, and uracil.
- DNA functions to store the genetic information in living cells and transmit it from one generation of cells to the next. RNA functions in protein synthesis.

24.5 Metabolism

- Metabolism is the sum of the many chemical reactions that go on in living cells.
- Catabolism refers to reactions that cells undergo to extract energy and chemical building blocks from large biological molecules.
- Anabolism refers to the reactions through which cells use energy and small building blocks to build the large biological molecules needed for cell structures and for carrying out cell functions.
- During photosynthesis, cells use carbon dioxide, water, and light energy to produce carbohydrates and oxygen.
- During cellular respiration, cells break down carbohydrates in the presence of oxygen gas to produce carbon dioxide and water. Energy released is stored as chemical potential energy in the molecule ATP.
- In the absence of oxygen, cells can carry out either alcoholic or lactic acid fermentation.

Vocabulary

- active site (p. 778)
- amino acid (p. 776)
- anabolism (p. 792)
- ATP (p. 793)
- carbohydrate (p. 781)
- catabolism (p. 792)
- cellular respiration (p. 794)
- denaturation (p. 778)
- disaccharide (p. 782)
- enzyme (p. 778)
- fatty acid (p. 784)
- fermentation (p. 794)
- lipid (p. 784)
- metabolism (p. 792)
- monosaccharide (p. 781)
- nucleic acid (p. 788)
- nucleotide (p. 788)
- peptide (p. 777)
- peptide bond (p. 777)
- phospholipid (p. 786)
- photosynthesis (p. 793)
- polysaccharide (p. 782)
- protein (p. 775)
- saponification (p. 785)
- steroid (p. 787)
- substrate (p. 778)
- triglyceride (p. 785)
- wax (p. 787)

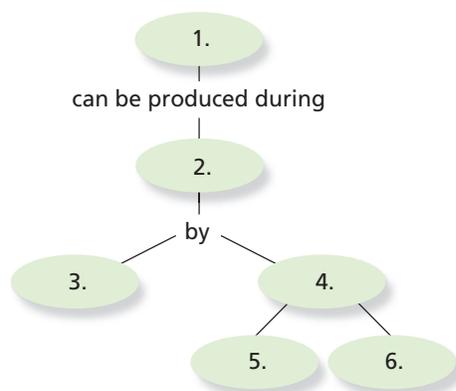


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Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 24 Assessment.

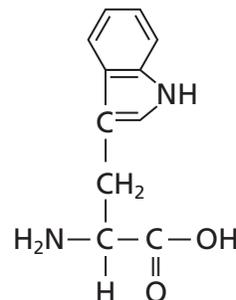
Concept Mapping

26. Complete the concept map using the following terms: lactic acid, ATP, cellular respiration, metabolism, alcoholic, fermentation.



Mastering Concepts

27. What should you call a chain of eight amino acids? A chain of 200 amino acids? (24.1)
28. Name five parts of your body that are made of structural proteins. (24.1)
29. Describe two common shapes found in the three-dimensional folding of proteins. (24.1)
30. Name the organic functional groups in the side chains of the following amino acids. (24.1)
- | | |
|--------------|------------------|
| a. glutamine | c. glutamic acid |
| b. serine | d. lysine |
31. Explain how the active site of an enzyme functions. (24.1)
32. Name the amino acids represented by each of the following abbreviations. (24.1)
- | | |
|--------|--------|
| a. Gly | d. Phe |
| b. Tyr | e. Glu |
| c. Trp | f. His |
33. Name an amino acid that has an aromatic ring in its side chain. (24.1)
34. Name two nonpolar and two polar amino acids. (24.1)
35. Is the dipeptide lysine-valine the same compound as the dipeptide valine-lysine? Explain. (24.1)
36. How do enzymes lower the activation energy for a reaction? (24.1)
37. The structure shown is tryptophan. Describe some of the properties you would expect tryptophan to have, based on its structure. In what class of large molecules do you think tryptophan is a building block? (24.1)



38. Most proteins with a globular shape are oriented so that they have mostly nonpolar amino acids on the inside and polar amino acids located on the outer surface. Does this make sense in terms of the nature of the cellular environment? Explain. (24.1)
39. Classify the following carbohydrates as monosaccharides, disaccharides, or polysaccharides. (24.2)
- | | |
|------------|--------------|
| a. starch | e. cellulose |
| b. glucose | f. glycogen |
| c. sucrose | g. fructose |
| d. ribose | h. lactose |
40. Name two isomers of glucose. (24.2)
41. What kind of bond is formed when two monosaccharides combine to form a disaccharide? (24.2)
42. Give a scientific term for each of the following. (24.2)
- | | |
|----------------|----------------|
| a. blood sugar | c. table sugar |
| b. fruit sugar | d. milk sugar |
43. Explain how the different bonding arrangements in cellulose and starch give them such different properties. (24.2)
44. The disaccharide maltose is formed from two glucose monomers. Draw its structure. (24.2)
45. Hydrolysis of cellulose, glycogen, and starch produces only one monosaccharide. Why is this so? What monosaccharide is produced? (24.2)
46. Digestion of disaccharides and polysaccharides cannot take place in the absence of water. Why do you think this is so? Include an equation in your answer. (24.2)
47. Draw the structure of the open-chain form of fructose. Circle all chiral carbons, and then calculate the number of stereoisomers with the same formula as fructose. (24.2)

48. Compare and contrast the structures of a triglyceride and a phospholipid. (24.3)
49. Predict whether a triglyceride from beef fat or a triglyceride from olive oil will have a higher melting point. Explain your reasoning. (24.3)
50. Explain how the structure of soaps makes them effective cleaning agents. (24.3)
51. Draw a portion of a lipid bilayer membrane, labeling the polar and nonpolar parts of the membrane. (24.3)
52. Where and in what form are fatty acids stored in the body? (24.3)
53. What type of lipid does not contain fatty acid chains? Why are these molecules classified as lipids? (24.3)
54. Draw the structure of the soap sodium palmitate (palmitate is the conjugate base of the 16-carbon saturated fatty acid, palmitic acid) and label its polar and nonpolar ends. (24.3)
55. What three structures make up a nucleotide? (24.4)
56. Name two nucleic acids found in organisms. (24.4)
57. Explain the roles of DNA and RNA in the production of proteins. (24.4)
58. Where in living cells is DNA found? (24.4)
59. Describe the types of bonds and attractions that link the monomers together in a DNA molecule. (24.4)
60. In the double helical structure of DNA, the base guanine is always bonded to cytosine and adenine is always bonded to thymine. What do you expect to be the relative proportional amounts of A, T, C, and G in a given length of DNA? (24.4)
61. One strand in a DNA molecule has the following base sequence. What is the base sequence of the other strand in the DNA molecule? (24.4)

$$\text{C-C-G-T-G-G-A-C-A-T-T-A}$$
62. Is digestion an anabolic or a catabolic process? Explain. (24.5)
63. Compare the net reactions for photosynthesis and cellular respiration with respect to reactants, products, and energy. (24.5)

Mastering Problems

Protein Structure (24.1)

64. How many peptide bonds are present in a peptide that has five amino acids?
65. How many peptides, each containing four different amino acids, can be made from Phe, Lys, Pro, and Asp? List the peptides, using the three-letter abbreviations for the amino acids.

66. The average molecular mass of an amino acid is 110 g/mol. Calculate the approximate number of amino acids in a protein that has a molecular mass of 36 500 g/mol.

Calculations with Lipids (24.3)

67. The fatty acid palmitic acid has a density of 0.853 g/mL at 62°C. What will be the mass of a 0.886-L sample of palmitic acid at that temperature?
68. How many moles of hydrogen gas are required for complete hydrogenation of one mole of linolenic acid, whose structure is shown below? Write a balanced equation for the hydrogenation reaction.

$$\text{CH}_3\text{CH}_2\text{CH}=\text{CHCH}_2\text{CH}=\text{CHCH}_2\text{CH}=\text{CH}(\text{CH}_2)_7\text{COOH}$$

Linolenic acid
69. A chicken egg contains about 213 mg of cholesterol. Calculate how many moles of cholesterol this represents if the molecular mass of cholesterol is 386.
70. Calculate the number of moles of sodium hydroxide needed in the saponification of 16 moles of a triglyceride.

Working with DNA (24.4)

71. The genetic code is a triplet code, that is, a sequence of three bases in RNA codes for each amino acid in a peptide chain or protein. How many RNA bases are required to code for a protein that contains 577 amino acids?
72. It has been calculated that the average length of a base pair in a DNA double helix is 3.4 Å. The human genome (the complete set of all DNA in the nucleus of a human cell) contains about three billion base pairs of DNA. In centimeters, how long is the DNA in the human genome? Assume that the DNA is stretched out and not coiled around proteins as it actually is in a living cell. (1 Å = 10⁻¹⁰ m)
73. A cell of the bacterium *Escherichia coli* has about 4.2 × 10⁶ base pairs of DNA, whereas each human cell has about 3 × 10⁹ base pairs of DNA. What percentage of the size of the human genome does the *E. coli* DNA represent?

Energy Calculations (24.5)

74. Every mole of glucose that undergoes alcoholic fermentation in yeast results in the net synthesis of two moles of ATP. How much energy in kJ is stored in two moles of ATP? Assume 100% efficiency.
75. How many moles of lactic acid are produced when three moles of glucose undergo fermentation in your muscle cells? Assume 100% completion of the process.

- 76.** The synthesis of one mole of the fatty acid palmitic acid from two-carbon building blocks requires seven moles of ATP. How many kJ of energy are required for the synthesis of one mole of palmitic acid?
- 77.** How many grams of glucose can be oxidized completely by 2.0 L of O₂ gas at STP during cellular respiration?
- 78.** Calculate and compare the total energy in kJ that is converted to ATP during the processes of cellular respiration and fermentation.

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 79.** Draw the carbonyl functional groups present in glucose and fructose. How are the groups similar? How are they different?
- 80.** List the names of the monomers that make up proteins, complex carbohydrates, and nucleic acids.
- 81.** Describe the functions of proteins, carbohydrates, lipids, and nucleic acids in living cells.
- 82.** Write a balanced equation for the hydrolysis of lactose.
- 83.** Write a balanced equation for photosynthesis.
- 84.** Write and balance the equation for cellular respiration.
- 85.** Write a balanced equation for the synthesis of sucrose from glucose and fructose.

Thinking Critically

- 86. Using Numbers** Approximately 38 moles of ATP are formed when glucose is completely oxidized during cellular respiration. If the heat of combustion for 1 mole of glucose is 2.82×10^3 kJ/mol and each mole of ATP stores 30.5 kJ of energy, what is the efficiency of cellular respiration in terms of the percentage of available energy that is stored in the chemical bonds of ATP?
- 87. Recognizing Cause and Effect** Some diets suggest severely restricting the intake of lipids. Why is it not a good idea to eliminate all lipids from the diet?
- 88. Making and Using Graphs**  A number of saturated fatty acids and values for some of their physical properties are listed in **Table 24-2**.
- Make a graph plotting number of carbon atoms versus melting point.
 - Graph the number of carbon atoms versus density.
 - Draw conclusions about the relationships between the number of carbon atoms in a saturated fatty acid and its density and melting point values.

- d.** Predict the approximate melting point of a saturated fatty acid that has 24 carbon atoms.

Table 24-2

Physical Properties of Saturated Fatty Acids			
Name	Number of carbon atoms	Melting point (°C)	Density (g/mL) (values at 60–80°C)
Palmitic acid	16	63	0.853
Myristic acid	14	58	0.862
Arachidic acid	20	77	0.824
Caprylic acid	8	16	0.910
Docosanoic acid	22	80	0.822
Stearic acid	18	70	0.847
Lauric acid	12	44	0.868

Writing in Chemistry

- 89.** Write a set of instructions that could be included in a package of contact-lens cleaning solution containing an enzyme. This enzyme catalyzes the breakdown of protein residues that adhere to the lenses. Include information about the structure and function of enzymes and the care that must be taken to avoid their denaturation during use.
- 90.** Use the library or the Internet to research cholesterol. Where is this molecule used in your body? What is its function? Why is too much dietary cholesterol considered to be bad for you?

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 91. a.** Write the balanced equation for the synthesis of ethanol from ethene and water.
- b.** If 448 L of ethene gas reacts with excess water at STP, how many grams of ethanol will be produced? (Chapter 14)
- 92.** Identify whether each of the reactants in these reactions is acting as an acid or a base. (Chapter 19)
- $\text{HBr} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Br}^-$
 - $\text{NH}_3 + \text{HCOOH} \rightarrow \text{NH}_4^+ + \text{HCOO}^-$
 - $\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{CO}_3^{2-} + \text{H}_3\text{O}^+$
- 93.** What is a voltaic cell? (Chapter 21)

Use these questions and the test-taking tip to prepare for your standardized test.

- Which of the following is NOT true of carbohydrates?
 - Monosaccharides in aqueous solutions interconvert continuously between an open-chain structure and a cyclic structure.
 - The monosaccharides in starch are linked together by the same kind of bond that links the monosaccharides in lactose.
 - All carbohydrates have the general chemical formula $C_n(H_2O)_n$.
 - Cellulose, made only by plants, is easily digestible by humans.
- All of the following are differences between RNA and DNA EXCEPT _____.
 - DNA contains the sugar deoxyribose, while RNA contains the sugar ribose
 - RNA contains the nitrogen base uracil, while DNA does not
 - RNA is usually single-stranded, while DNA is usually double-stranded
 - DNA contains the nitrogen base adenine, while RNA does not
- Cellular respiration produces about 38 moles of ATP for every mole of glucose consumed:

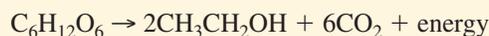
$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + 38ATP$$
 If each molecule of ATP can release 30.5 kJ of energy, how much energy can be obtained from a candy bar containing 130 g of glucose?
 - 27.4 kJ
 - 836 kJ
 - 1159 kJ
 - 3970 kJ
- The sequence of bases in RNA determines the sequence of amino acids in a protein. Three bases code for a single amino acid; for example, CAG is the code for glutamine. Therefore, a strand of RNA 2.73×10^4 bases long codes for a protein that has _____.
 - 8.19×10^4 amino acids
 - 9.10×10^3 amino acids
 - 2.73×10^4 amino acids
 - 4.55×10^3 amino acids



TEST-TAKING TIP

Use The Process of Elimination On any multiple-choice test, there are two ways to find the correct answer to each question. Either you can choose the right answer immediately or you can eliminate the answers that you know are wrong.

- The equation for the alcoholic fermentation of glucose is shown below:



The ethanol content of wine is about 12%. Therefore, in every 100 g of wine there are 12.0 g of ethanol. How many grams of glucose were catabolized to produce these 12.0 g of ethanol?

- 23.4 g
- 12.0 g
- 47.0 g
- 27.0 g

Analyzing Tables Use the table below to answer the following questions.

Note: # of X = # of molecules of X in one DNA molecule

$$\%X = \frac{\# \text{ of } X}{\# \text{ of A} + \# \text{ of G} + \# \text{ of C} + \# \text{ of T}}$$

where X is any nitrogen base

- What is the %T of DNA D?
 - 28.4%
 - 78.4%
 - 71.6%
 - 21.6%
- Every nitrogen base found in a DNA molecule is part of a nucleotide of that molecule. The A nucleotide, C nucleotide, G nucleotide, and T nucleotide have molar masses of 347.22 g/mol, 323.20 g/mol, 363.23 g/mol, and 338.21 g/mol respectively. What is the mass of one mole of DNA A?
 - 2.79×10^5 g
 - 2.7001×10^5 g
 - 2.6390×10^5 g
 - 2.72×10^5 g
- How many molecules of adenine are in one molecule of DNA B?
 - 402
 - 434
 - 216
 - 175

Nitrogen Base Data for Four Different Double-Stranded DNA Molecules

DNA Molecule	# of A	# of G	# of C	# of T	%A	%G	%C	%T
A	165	?	231	?	20.8	?	29.2	?
B	?	402	?	?	?	32.5	?	?
C	?	?	194	234	?	?	22.7	27.3
D	266	203	?	?	28.4	21.6	?	?

Nuclear Chemistry

What You'll Learn

- ▶ You will trace the history of nuclear chemistry from discovery to application.
- ▶ You will identify types of radioactive decay and solve decay rate problems.
- ▶ You will describe the reactions involved in nuclear fission and fusion.
- ▶ You will learn about applications of nuclear reactions and the effects of radiation exposure.

Why It's Important

From its role in shaping world politics to its applications that produce electrical power and diagnose and treat disease, nuclear chemistry has profound effects upon the world in which we live.



Visit the Chemistry Web site at science.glencoe.com to find links about nuclear chemistry.

Many medical diagnostic tests and treatments involve the use of radioactive substances. Here, a radioactive substance known as a radiotracer is used to illuminate the carotid artery, which runs through the neck into the skull.



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DISCOVERY LAB



Chain Reactions

When the products of one nuclear reaction cause additional nuclear reactions to occur, the resulting chain reaction can release large amounts of energy in a short period of time. Explore escalating chain reactions by modeling them with dominoes.

Procedure

1. Obtain a set of domino tiles.
2. Stand the individual dominoes on end and arrange them so that when the first domino falls, it causes the other dominoes to fall in series.
3. Practice using different arrangements until you determine how to cause the dominoes to fall in the least amount of time.
4. Time your domino chain reaction. Compare your time with those of your classmates.

Analysis

What arrangement caused the dominoes to fall in the least amount of time? Do the dominoes fall at a steady rate or an escalating rate? What happens to the domino chain reaction if a tile does not contact the next tile in the sequence?

Materials

28 domino tiles (1 set)
stopwatch

Section

25.1

Nuclear Radiation

Objectives

- **List** the founding scientists in the study of radioactivity and **state** their discoveries.
- **Identify** alpha, beta, and gamma radiation in terms of composition and key properties.

Vocabulary

radioisotope
X ray

You may recall from Chapter 4 that the nuclei of some atoms are unstable and undergo nuclear reactions. In this chapter you will study nuclear chemistry, which is concerned with the structure of atomic nuclei and the changes they undergo. An application of a nuclear reaction is shown in the photo of the human neck and skull. **Table 25-1** offers a comparison of chemical and nuclear reactions.

Table 25-1

Characteristics of Chemical and Nuclear Reactions	
Chemical reactions	Nuclear reactions
1. Occur when bonds are broken and formed.	1. Occur when nuclei emit particles and/or rays.
2. Atoms remain unchanged, though they may be rearranged.	2. Atoms are often converted into atoms of another element.
3. Involve only valence electrons.	3. May involve protons, neutrons, and electrons.
4. Associated with small energy changes.	4. Associated with large energy changes.
5. Reaction rate is influenced by temperature, pressure, concentration, and catalysts.	5. Reaction rate is not normally affected by temperature, pressure, or catalysts.

The Discovery of Radioactivity

In 1895, Wilhelm Roentgen (1845–1923) found that invisible rays were emitted when electrons bombarded the surface of certain materials. The emitted rays were discovered because they caused photographic plates to darken. Roentgen named these invisible high-energy emissions X rays. As is true in many fields, Roentgen's discovery of X rays created excitement within the scientific community and stimulated further research. At that time, French physicist Henri Becquerel (1852–1908) was studying minerals that emit light after being exposed to sunlight, a phenomenon called phosphorescence. Building on Roentgen's work, Becquerel wanted to determine whether phosphorescent minerals also emitted X rays. Becquerel accidentally discovered that phosphorescent uranium salts—even when not exposed to light—produced spontaneous emissions that darkened photographic plates. **Figure 25-1** shows the darkening of photographic film that is exposed to uranium-containing ore.

Marie Curie (1867–1934) and her husband Pierre (1859–1906) took Becquerel's mineral sample (called pitchblende) and isolated the components emitting the rays. They concluded that the darkening of the photographic plates was due to rays emitted specifically from the uranium atoms present in the mineral sample. Marie Curie named the process by which materials give off such rays radioactivity; the rays and particles emitted by a radioactive source are called radiation.

The work of Marie and Pierre Curie was extremely important in establishing the origin of radioactivity and the field of nuclear chemistry. In 1898, the Curies identified two new elements, polonium and radium, on the basis of their radioactivity. Henri Becquerel and the Curies shared the 1903 Nobel Prize in Physics for their work. Marie Curie also received the 1911 Nobel Prize in Chemistry for her work with polonium and radium. **Figure 25-2** shows the Curies at work in their laboratory.

Types of Radiation

While reading about the discovery of radioactivity, several questions may have occurred to you. Which atomic nuclei are radioactive? What types of radiation do radioactive nuclei emit? It is best to start with the second question first, and explore the types of radiation emitted by radioactive sources.

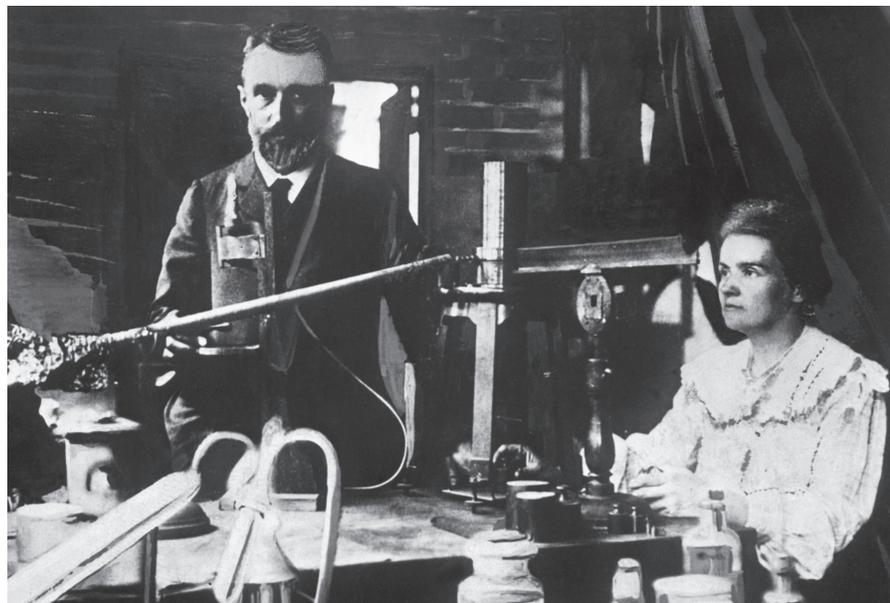


Figure 25-2

Both Pierre and Marie Curie played important roles in founding the field of nuclear chemistry. Marie Curie went on to show that unlike chemical reactions, radioactivity is not affected by changes in physical conditions such as temperature and pressure. She is the only person in history to receive Nobel Prizes in two different sciences—physics in 1903, and chemistry in 1911.

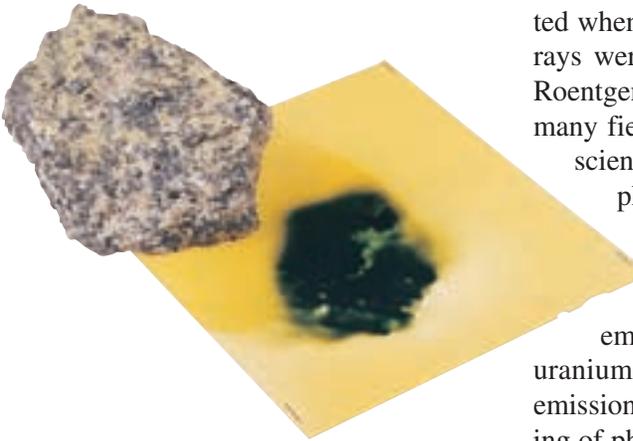


Figure 25-1

A sample of radioactive uranium-containing ore exposes photographic film.

As you may recall, isotopes are atoms of the same element that have different numbers of neutrons. Isotopes of atoms with unstable nuclei are called **radioisotopes**. These unstable nuclei emit radiation to attain more stable atomic configurations in a process called radioactive decay. During radioactive decay, unstable atoms lose energy by emitting one of several types of radiation. The three most common types of radiation are alpha (α), beta (β), and gamma (γ). **Table 25-2** summarizes some of their important properties. Later in this chapter you'll learn about other types of radiation that may be emitted in a nuclear reaction.

Table 25-2

Properties of Alpha, Beta, and Gamma Radiation			
Property	Alpha (α)	Beta (β)	Gamma (γ)
Composition	Alpha particles	Beta particles	High-energy electromagnetic radiation
Description of radiation	Helium nuclei ${}^4_2\text{He}$	Electrons ${}^0_{-1}\beta$	photons ${}^0_0\gamma$
Charge	2+	1-	0
Mass	6.64×10^{-24} kg	9.11×10^{-28} kg	0
Approximate energy*	5 MeV	0.05 to 1 MeV	1 MeV
Relative penetrating power	Blocked by paper	Blocked by metal foil	Not completely blocked by lead or concrete

*(1 MeV = 1.60×10^{-13} J)

Ernest Rutherford (1871–1937), whom you know of because of his famous gold foil experiment that helped define modern atomic structure, identified alpha, beta, and gamma radiation when studying the effects of an electric field on the emissions from a radioactive source. As you can see in **Figure 25-3**, alpha particles carry a 2+ charge and are deflected toward the negatively charged plate. Beta particles carry a 1- charge and are deflected toward the positively charged plate. In contrast, gamma rays carry no charge and are not affected by the electric field.

Figure 25-3

The effect of an electric field on three types of radiation is shown here. Positively charged alpha particles are deflected toward the negatively charged plate. Negatively charged beta particles are deflected toward the positively charged plate. Beta particles undergo greater deflection because they have considerably less mass than alpha particles. Gamma rays, which have no electrical charge, are not deflected.

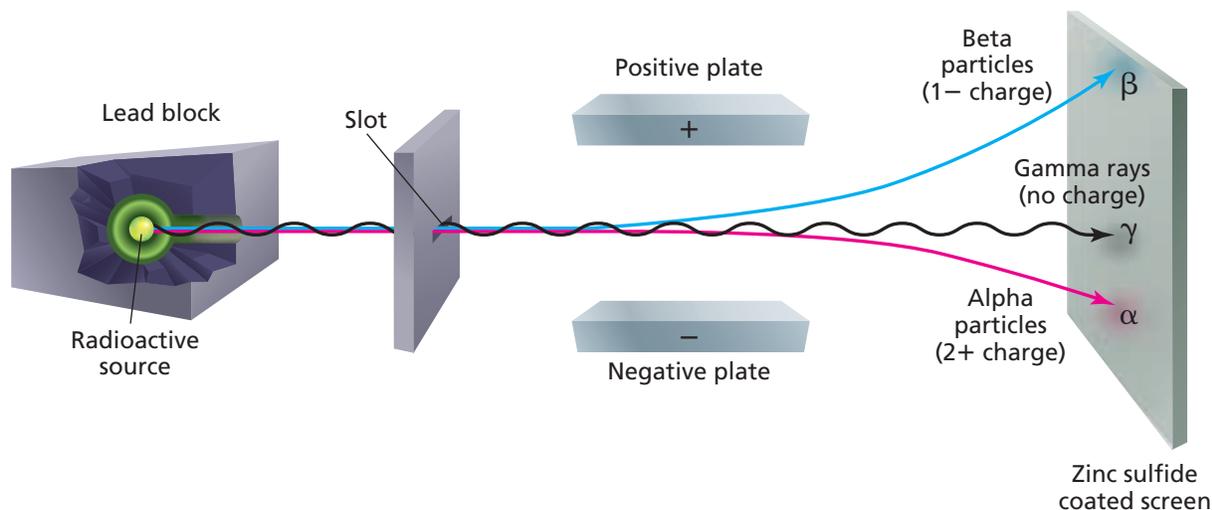


Figure 25-4

A radium-226 nucleus undergoes alpha decay to form radon-222. What is the charge on the alpha particle that is emitted?



Physics

CONNECTION

Exposing certain foods such as strawberries to small amounts of gamma radiation allows them to last longer before spoiling. The process, called irradiation, uses decaying cobalt-60 nuclei as a source of gamma radiation. By disrupting their cell metabolism, gamma radiation destroys the microorganisms that cause food to spoil. Irradiation can also be used to slow the ripening process, extend shelf life, and allow foods to be stored for long periods of time without refrigeration. The food does not become radioactive from the process. More than 30 nations around the world have approved irradiation as a means of preserving food.

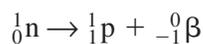


An alpha particle (α) has the same composition as a helium nucleus—two protons and two neutrons—and is therefore given the symbol ${}^4_2\text{He}$. The charge of an alpha particle is $2+$ due to the presence of the two protons. Alpha radiation consists of a stream of alpha particles. As you can see in **Figure 25-4**, radium-226, an atom whose nucleus contains 88 protons and 138 neutrons, undergoes alpha decay by emitting an alpha particle. Notice that after the decay, the resulting atom has an atomic number of 86, a mass number of 222, and is no longer radium. The newly formed radioisotope is radon-222.

In examining **Figure 25-4**, you should note that the particles involved are balanced. That is, the sum of the mass numbers (superscripts) and the sum of the atomic numbers (subscripts) on each side of the arrow are equal. Also note that when a radioactive nucleus emits an alpha particle, the product nucleus has an atomic number that is lower by two and a mass number that is lower by four. What particle is formed when polonium-210 (${}^{210}_{84}\text{Po}$) undergoes alpha decay?

Because of their mass and charge, alpha particles are relatively slow-moving compared with other types of radiation. Thus, alpha particles are not very penetrating—a single sheet of paper stops alpha particles.

A beta particle is a very-fast moving electron that has been emitted from a neutron of an unstable nucleus. Beta particles are represented by the symbol ${}^0_{-1}\beta$. The zero superscript indicates the insignificant mass of an electron in comparison with the mass of a nucleus. The -1 subscript denotes the negative charge of the particle. Beta radiation consists of a stream of fast-moving electrons. An example of the beta decay process is the decay of iodine-131 into xenon-131 by beta-particle emission, as shown in **Figure 25-5**. Note that the mass number of the product nucleus is the same as that of the original nucleus (they are both 131), but its atomic number has increased by 1 (54 instead of 53). This change in atomic number, and thus, change in identity, occurs because the electron emitted during the beta decay has been removed from a neutron, leaving behind a proton.



As you may recall, because the number of protons in an atom determines its identity, the formation of an additional proton results in the transformation from iodine to xenon. Because beta particles are both lightweight and fast

Figure 25-5

An iodine-131 nucleus undergoes beta decay to form xenon-131. How does beta decay affect the mass number of the decaying nucleus?

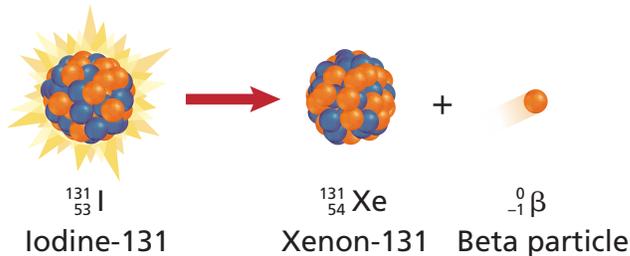


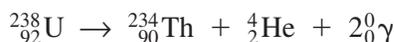


Figure 25-6

A heavy lead-lined vest absorbs errant radiation from a dental X-ray machine, limiting exposure to the rest of the body. Dentists use dental X rays as a diagnostic tool for determining the health of teeth and gums.

moving, they have greater penetrating power than alpha particles. A thin metal foil is required to stop beta particles.

Gamma rays are high-energy (short wavelength) electromagnetic radiation. They are denoted by the symbol ${}^0_0\gamma$. As you can see from the symbol, both the subscript and superscript are zero. Thus, the emission of gamma rays does not change the atomic number or mass number of a nucleus. Gamma rays almost always accompany alpha and beta radiation, as they account for most of the energy loss that occurs as a nucleus decays. For example, gamma rays accompany the alpha-decay reaction of uranium-238.



The 2 in front of the γ symbol indicates that two gamma rays of different frequencies are emitted. Because gamma rays have no effect on mass number or atomic number, it is customary to omit them from nuclear equations.

As you have learned, the discovery of X rays helped set off the events that led to the discovery of radioactivity. **X rays**, like gamma rays, are a form of high-energy electromagnetic radiation. Unlike gamma rays, X rays are not produced by radioactive sources. Instead, X rays are emitted from certain materials that are in an excited electron state. Both X rays and gamma rays are extremely penetrating and can be very damaging to living tissue. X rays and gamma rays are only partially blocked by lead and concrete. If your dentist has recently taken dental X rays of your teeth, you may recall wearing a heavy vest over your chest during the procedure. The vest, which can be seen in **Figure 25-6**, is lined with lead, and its purpose is to limit your body's exposure to the potentially damaging X rays.



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Section 25.1 Assessment

- Describe the contributions of Roentgen, Becquerel, Rutherford, and the Curies to the field of nuclear chemistry.
- What subatomic particles are involved in nuclear reactions? What subatomic particles are involved in chemical reactions?
- Using **Table 25-2** as a guide, qualitatively compare and contrast alpha, beta, and gamma radiation in terms of composition, energy, mass, and penetrating power.
- Thinking Critically** An atom undergoes a reaction and attains a more stable form. How do you know if the reaction was a chemical reaction or a nuclear reaction?
- Converting Units** **Table 25-2** gives Approximate Energy values in units of MeV. Convert each value into joules using the following conversion factor ($1 \text{ MeV} = 1.61 \times 10^{-13} \text{ J}$). For more help, refer to *Unit Conversion* in the **Math Handbook** on page 901 of this textbook.

Objectives

- **Explain** why certain nuclei are radioactive.
- **Apply** your knowledge of radioactive decay to write balanced nuclear equations.

Vocabulary

nucleon
 strong nuclear force
 band of stability
 positron emission
 positron
 electron capture
 radioactive decay series

It may surprise you to learn that of all the known isotopes, only about 17% are stable and don't decay spontaneously. In Chapter 4 you learned that the stability of an atom is determined by the neutron-to-proton ratio of its nucleus. You may be wondering if there is a way to know what type of radioactive decay a particular radioisotope will undergo. There is, and as you'll learn in this section, it is the neutron-to-proton ratio of the nucleus that determines the type of radioactive decay that will occur.

Nuclear Stability

Every atom has an extremely dense nucleus that contains most of the atom's mass. The nucleus contains positively charged protons and neutral neutrons, both of which are referred to as **nucleons**. You may have wondered how protons remain in the densely packed nucleus despite the strong electrostatic repulsion forces produced by the positively charged particles. The answer is that the **strong nuclear force**, a force that acts only on subatomic particles that are extremely close together, overcomes the electrostatic repulsion between protons.

The fact that the strong nuclear force acts on both protons and neutrons is important. Because neutrons are neutral, a neutron that is adjacent to a positively charged proton creates no repulsive electrostatic force. Yet these two adjacent particles are subject to the strong nuclear force that works to hold them together. Likewise, two adjacent neutrons create no attractive or repulsive electrostatic force, but they too are subject to the strong nuclear force holding them together. Thus, the presence of neutrons adds an attractive force within the nucleus. The number of neutrons in a nucleus is important because nuclear stability is related to the balance between electrostatic and strong nuclear forces.

To a certain degree, the stability of a nucleus can be correlated with its neutron-to-proton (n/p) ratio. For atoms with low atomic numbers (< 20), the most stable nuclei are those with neutron-to-proton ratios of 1 : 1. For example, helium (${}^4_2\text{He}$) has two neutrons and two protons, and a neutron-to-proton ratio of 1 : 1. As atomic number increases, more and more neutrons are needed to produce a strong nuclear force that is sufficient to balance the electrostatic repulsion forces. Thus, the neutron-to-proton ratio for stable atoms gradually increases, reaching a maximum of approximately 1.5 : 1 for the largest atoms. An example of this is lead (${}^{206}_{82}\text{Pb}$). With 124 neutrons and 82 protons, lead has a neutron-to-proton ratio of 1.51 : 1. You can see the calculation of lead's neutron-to-proton ratio in **Figure 25-7**.

Figure 25-7

The steps in calculating the neutron-to-proton ratio (the n/p ratio) for lead-206 are illustrated here.

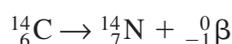
Mass number	—	206	Pb
Atomic number	—	82	
Atomic number = Number of protons = 82			
Mass number = Number of protons + Number of neutrons = 206			
Number of neutrons = Mass number – Atomic number			
Number of neutrons = 206 – 82 = 124			
Neutron-to-proton ratio = $\frac{\text{Number of neutrons}}{\text{Number of protons}}$			
Neutron-to-proton ratio = $\frac{124}{82}$			
Neutron-to-proton ratio = $\frac{1.51}{1}$			

Examine the plot of the number of neutrons versus the number of protons for all known stable nuclei shown in **Figure 25-8**. As you can see, the slope of the plot indicates that the number of neutrons required for a stable nucleus increases as the number of protons increases. This correlates with the increase in the neutron-to-proton ratio of stable nuclei with increasing atomic number. The area on the graph within which all stable nuclei are found is known as the **band of stability**. As shown in **Figure 25-8**, ${}^4_2\text{He}$ and ${}^{206}_{82}\text{Pb}$, with their very different neutron-to-proton ratios, are both positioned within the band of stability. Radioactive nuclei are found outside the band of stability—either above or below—and undergo decay in order to gain stability. After decay, the new atom is positioned more closely to, if not within, the band of stability. The band of stability ends at bismuth-209; all elements with atomic numbers greater than 83 are radioactive.

Types of Radioactive Decay

The type of radioactive decay a particular radioisotope undergoes depends to a large degree on the underlying causes for its instability. Atoms lying above the band of stability generally have too many neutrons to be stable, whereas atoms lying below the band of stability tend to have too many protons to be stable.

Beta decay A radioisotope that lies above the band of stability is unstable because it has too many neutrons relative to its number of protons. For example, unstable ${}^{14}_6\text{C}$ has a neutron-to-proton ratio of 1.33 : 1, whereas stable elements of similar mass, such as ${}^{12}_6\text{C}$ and ${}^{14}_7\text{N}$, have neutron-to-proton ratios of approximately 1 : 1. It is not surprising then that ${}^{14}_6\text{C}$ undergoes beta decay, as this type of decay decreases the number of neutrons in the nucleus.



Note that the atomic number of the product nucleus, ${}^{14}_7\text{N}$, has increased by one. The nitrogen-14 atom now has a stable neutron-to-proton ratio of 1 : 1. Thus, beta emission has the effect of increasing the stability of a neutron-rich atom by lowering its neutron-to-proton ratio. The resulting atom is closer to, if not within, the band of stability.

Alpha decay All nuclei with more than 83 protons are radioactive and decay spontaneously. Both the number of neutrons and the number of protons must be reduced in order to make these radioisotopes stable. These very heavy nuclei often decay by emitting alpha particles. For example, polonium-210 spontaneously decays by alpha emission.



The atomic number of ${}^{210}_{84}\text{Po}$ decreases by two and the mass number decreases by four as the nucleus decays into ${}^{206}_{82}\text{Pb}$. How does the n/p ratio change?

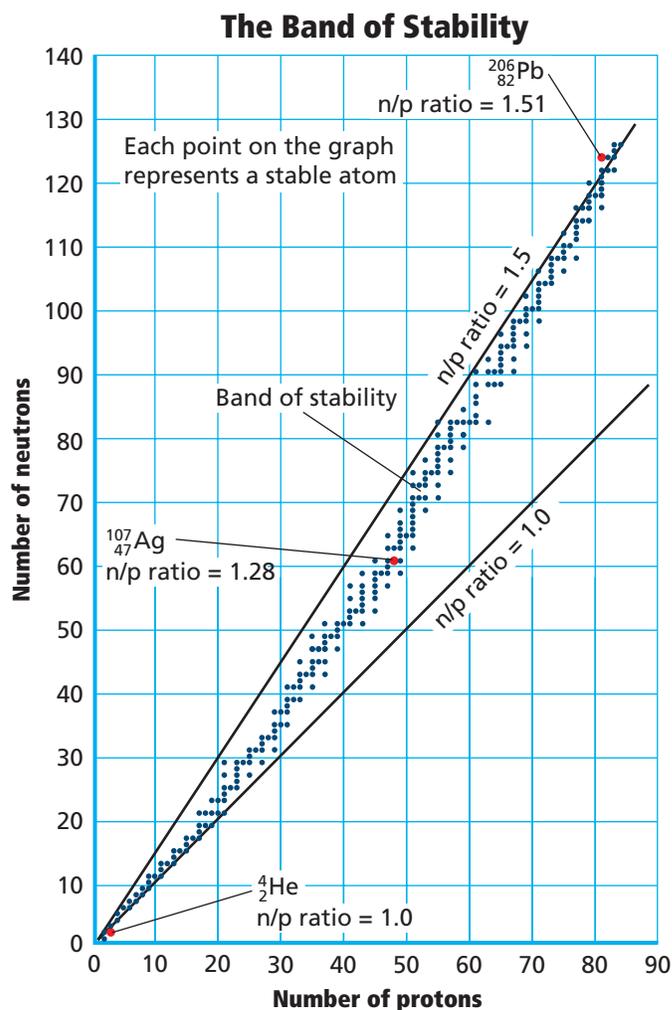


Figure 25-8

Plotting the number of neutrons versus the number of protons for all stable nuclei produces a region called the band of stability. As you move from low atomic number nuclei to high atomic number nuclei, the neutron-to-proton ratio for stable nuclei gradually increases from 1 : 1 to 1.5 : 1.

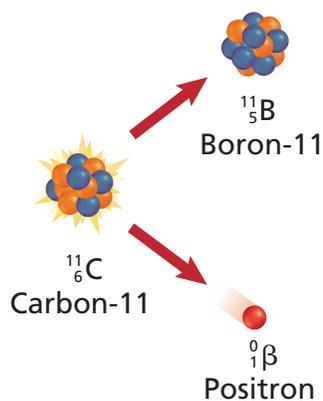


Figure 25-9

Positron emission from carbon-11 decreases the number of protons from six to five, and increases the number of neutrons from five to six.

Positron emission and electron capture For nuclei with low neutron-to-proton ratios lying below the band of stability, there are two common radioactive decay processes that occur, positron emission and electron capture. These two processes tend to increase the neutron-to-proton ratio of the neutron-poor atom. After an unstable atom undergoes electron capture or positron emission, the resulting atom is closer to, if not within, the band of stability.

Positron emission is a radioactive decay process that involves the emission of a positron from a nucleus. A **positron** is a particle with the same mass as an electron but opposite charge, thus it is represented by the symbol ${}_{1}^{0}\beta$. During positron emission, a proton in the nucleus is converted into a neutron and a positron, and then the positron is emitted.

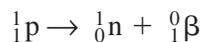
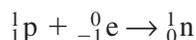


Figure 25-9 shows the positron emission of a carbon-11 nucleus. Carbon-11 lies below the band of stability and has a low neutron-to-proton ratio of 0.8 : 1. Carbon-11 undergoes positron emission to form boron-11. Positron emission decreases the number of protons from six to five, and increases the number of neutrons from five to six. The resulting atom, ${}_{5}^{11}\text{B}$, has a neutron-to-proton ratio of 1.2 : 1, which is within the band of stability.

Electron capture is the other common radioactive decay process that decreases the number of protons in unstable nuclei lying below the band of stability. **Electron capture** occurs when the nucleus of an atom draws in a surrounding electron, usually one from the lowest energy level. This captured electron combines with a proton to form a neutron.



The atomic number of the nucleus decreases by one as a consequence of electron capture. The formation of the neutron also results in an X-ray photon being emitted. These two characteristics of electron capture can be seen in the electron capture of rubidium-81 shown in **Figure 25-10**. The balanced nuclear equation for the reaction is shown below.



How is the neutron-to-proton ratio of the product, Kr-81, different from that of Rb-81?

The five types of radioactive decay you have read about in this chapter are summarized in **Table 25-3**. Which of the decay processes listed in the table result in an increased neutron-to-proton ratio? In a decrease?

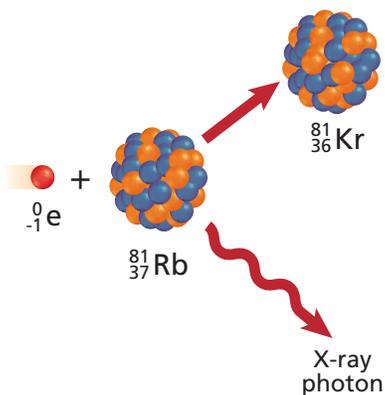


Figure 25-10

The electron capture of rubidium-81 decreases the number of protons in the nucleus while the mass number remains the same. Thus, the neutron-to-proton ratio increases from 1.19 : 1 for ${}_{37}^{81}\text{Rb}$ to 1.25 : 1 for ${}_{36}^{81}\text{Kr}$.

Table 25-3

Summary of Radioactive Decay Processes			
Type of radioactive decay	Particle emitted	Change in mass number	Change in atomic number
Alpha decay	${}_{2}^{4}\text{He}$	Decreases by 4	Decreases by 2
Beta decay	${}_{-1}^{0}\beta$	No change	Increases by 1
Positron emission	${}_{1}^{0}\beta$	No change	Decreases by 1
Electron capture	X-ray photon	No change	Decreases by 1
Gamma emission	${}_{0}^{0}\gamma$	No change	No change

Writing and Balancing Nuclear Equations

The radioactive decay processes you have just read about are all examples of nuclear reactions. As you probably noticed, nuclear reactions are expressed by balanced nuclear equations just as chemical reactions are expressed by balanced chemical equations. However, in balanced chemical equations, numbers and kinds of atoms are conserved; in balanced nuclear equations, mass numbers and atomic numbers are conserved.

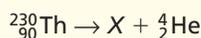
EXAMPLE PROBLEM 25-1

Balancing a Nuclear Equation

Using the information provided in Table 25-3, write a balanced nuclear equation for the alpha decay of thorium-230 (${}^{230}_{90}\text{Th}$).

1. Analyze the Problem

You are given that a thorium atom undergoes alpha decay and forms an unknown product. Thorium-230 is the initial reactant, while the alpha particle is one of the products of the reaction. The reaction is summarized below.



You must determine the unknown product of the reaction, X . This can be done through the conservation of atomic number and mass number. The periodic table can then be used to identify X .

Known

reactant: thorium-230 (${}^{230}_{90}\text{Th}$)

decay type: alpha particle emission (${}^4_2\text{He}$)

Unknown

reaction product $X = ?$

balanced nuclear equation = ?

2. Solve for the Unknown

Using each particle's mass number, make sure mass number is conserved on each side of the reaction arrow.

$$\text{mass number: } 230 = X + 4$$

$$X = 230 - 4 = 226$$

Thus, the mass number of X is 226.

Using each particle's atomic number, make sure atomic number is conserved on each side of the reaction arrow.

$$\text{atomic number: } 90 = X + 2$$

$$X = 90 - 2 = 88$$

Thus, the atomic number of X is 88. The periodic table identifies the element as radium (Ra).

Write the balanced nuclear equation.



3. Evaluate the Answer

The correct formula for an alpha particle is used. The sums of the superscripts on each side of the equation are equal. The same is true for the subscripts. Therefore, the atomic number and the mass number are conserved. The nuclear equation is balanced.



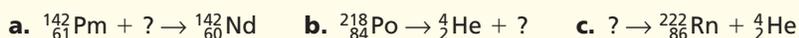
Thorium fluoride is used in the manufacture of propane lantern mantles and high-intensity searchlights.

Practice!

For more practice with balancing nuclear equations, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- Write a balanced nuclear equation for the reaction in which oxygen-15 undergoes positron emission.
- Determine what type of decay occurs when thorium-231 undergoes radioactive decay to form protactinium-231.
- Write a balanced nuclear equation for the reaction in which the transition metal zirconium-97 undergoes beta decay.
- Complete the following nuclear equations.

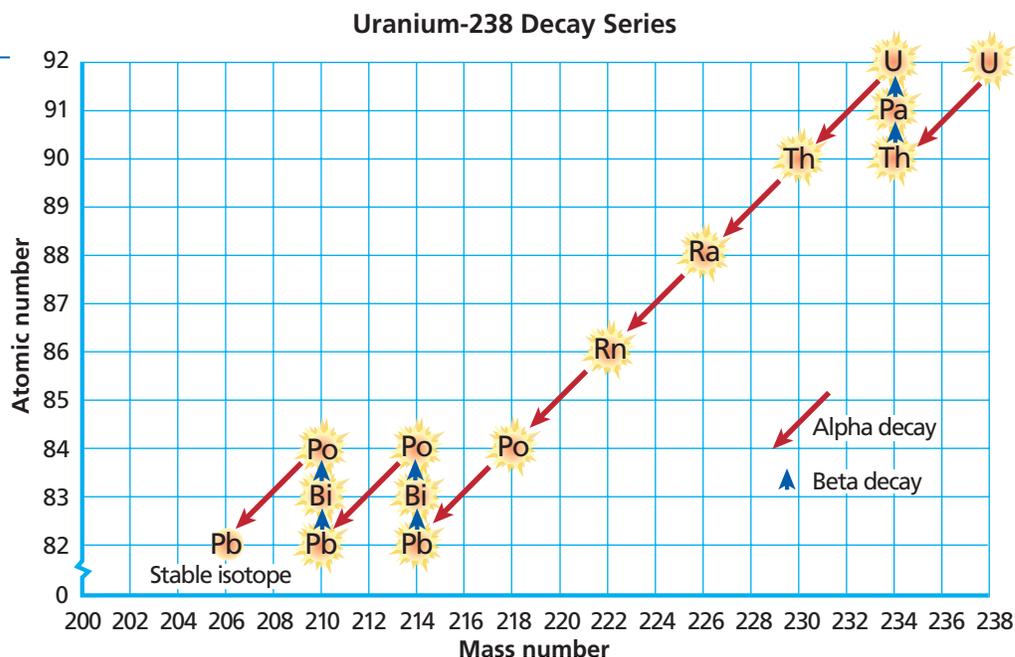


Radioactive Series

A series of nuclear reactions that begins with an unstable nucleus and results in the formation of a stable nucleus is called a **radioactive decay series**. As you can see in **Figure 25-11**, uranium-238 first decays to thorium-235, which in turn decays to protactinium-234. Decay reactions continue until a stable nucleus, lead-206, is formed.

Figure 25-11

Uranium-238 undergoes 14 different radioactive decay processes before forming stable lead-206.



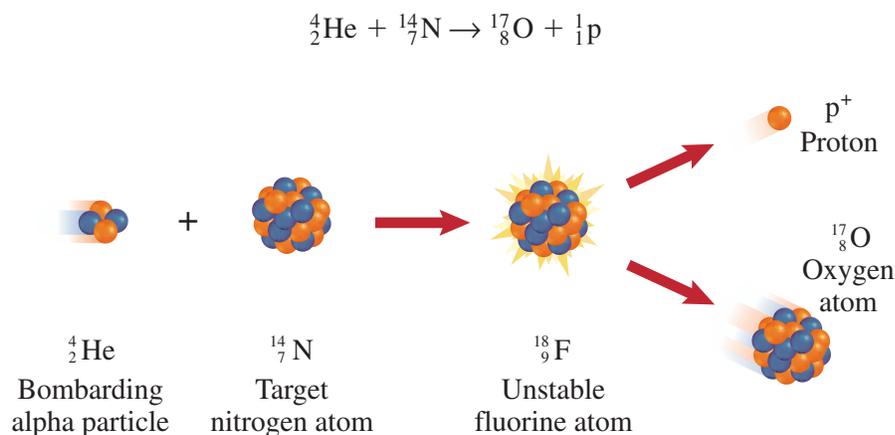
Section 25.2 Assessment

- Explain how you can predict whether or not an isotope is likely to be stable if you know the number of neutrons and protons in its nucleus.
- Write the nuclear equation for the alpha decay of astatine-213.
- Complete and balance the following.
 - ${}_{29}^{66}\text{Cu} \rightarrow {}_{30}^{66}\text{Zn} + ?$
 - $? \rightarrow {}_{77}^{181}\text{Ir} + {}_2^4\text{He}$
- Thinking Critically** A new element is synthesized in a laboratory. The element has a neutron-to-proton ratio of 1.6 : 1. Will the element be radioactive? If so, what type of radioactive decay will it most likely undergo?
- Communicating** Describe the forces acting on the particles within a nucleus. Explain why neutrons are the "glue" holding the nucleus together.

All the nuclear reactions that have been described thus far are examples of radioactive decay, where one element is converted into another element by the spontaneous emission of radiation. This conversion of an atom of one element to an atom of another element is called **transmutation**. Except for gamma emission, which does not alter an atom's atomic number, all nuclear reactions are transmutation reactions. Some unstable nuclei, such as the uranium salts used by Henri Becquerel, undergo transmutation naturally. However, transmutation may also be forced, or induced, by bombarding a stable nucleus with high-energy alpha, beta, or gamma radiation.

Induced Transmutation

In 1919, Ernest Rutherford performed the first laboratory conversion of one element into another element. By bombarding nitrogen-14 with high-speed alpha particles, an unstable fluorine-18 occurred, and then oxygen-17 was formed. This transmutation reaction is illustrated below.



Rutherford's experiments demonstrated that nuclear reactions can be induced, in other words, produced artificially. The process, which involves striking nuclei with high-velocity charged particles, is called **induced transmutation**. The charged particles, such as alpha particles used by Rutherford, must be moving at extremely high speeds to overcome the electrostatic repulsion between themselves and the target nucleus. Because of this, scientists have developed methods to accelerate charged particles to extreme speeds by using very strong electrostatic and magnetic fields. Particle accelerators, which are commonly called "atom smashers," are built to produce the high-speed particles needed to induce transmutation. The inside of the Fermi National Accelerator Laboratory's Tevatron is shown in **Figure 25-12**. The Tevatron is the world's highest-energy particle accelerator. The purpose of the facility, which opened in 1999, is to research high-energy nuclear reactions and particle physics. Since Rutherford's first experiments involving induced transmutation, scientists have used the technique to synthesize hundreds of new isotopes in the laboratory.

Transuranium elements The elements immediately following uranium in the periodic table—elements with atomic numbers 93 and greater—are known as the **transuranium elements**. All transuranium elements have been produced in the laboratory by induced transmutation and are radioactive.

Objectives

- **Describe** how induced transmutation is used to produce a transuranium element.
- **Solve** problems involving radioactive decay rates.

Vocabulary

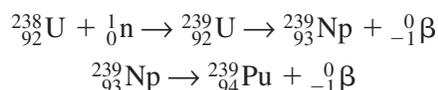
transmutation
 induced transmutation
 transuranium element
 half-life
 radiochemical dating

Figure 25-12

The main Tevatron ring has a diameter of more than 0.8 km and a circumference of about 6.4 km. The accelerator uses conventional and superconducting magnets to accelerate particles to high speeds and high energies.



First discovered in 1940, elements 93 (neptunium) and 94 (plutonium) are prepared by bombarding uranium-238 with neutrons.



If you read through the names of the transuranium elements, you'll notice that many of them have been named in honor of their discoverers or the laboratories at which they were created. There are ongoing efforts throughout the world's major scientific research centers to synthesize new transuranium elements and study their properties.

EXAMPLE PROBLEM 25-2

Induced Transmutation Reaction Equations

Write a balanced nuclear equation for the induced transmutation of aluminum-27 into phosphorus-30 by alpha particle bombardment. A neutron is emitted from the aluminum atom in the reaction.

1. Analyze the Problem

You are given all of the particles involved in an induced transmutation reaction, from which you must write the balanced nuclear equation. Because the alpha particle bombards the aluminum atom, they are reactants and must appear on the reactant side of the reaction arrow. Obtain the atomic number of aluminum and phosphorus from the periodic table. Write out the nuclear reaction, being sure to include the alpha particle (reactant) and the neutron (product).

Known

reactants: aluminum-27 and an alpha particle
products: phosphorus-30 and a neutron

Unknown

nuclear equation for the reaction = ?

2. Solve for the Unknown

Write the formula for each participant in the reaction.

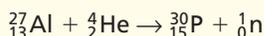
Aluminum is atomic number 13; its nuclear formula is ${}_{13}^{27}\text{Al}$.

Phosphorus is atomic number 15; its nuclear formula is ${}_{15}^{30}\text{P}$.

The formula for the alpha particle is ${}_{2}^4\text{He}$.

The formula for the neutron is ${}_0^1\text{n}$.

Write the balanced nuclear equation.



3. Evaluate the Answer

The sums of the superscripts on each side of the equation are equal. The same is true for the subscripts. Therefore, the atomic number and the mass number are conserved. The formula for each participant in the reaction is also correct. The nuclear equation is written correctly.

PRACTICE PROBLEMS

- Write the balanced nuclear equation for the induced transmutation of aluminum-27 into sodium-24 by neutron bombardment. An alpha particle is released in the reaction.
- Write the balanced nuclear equation for the alpha particle bombardment of ${}_{94}^{239}\text{Pu}$. One of the reaction products is a neutron.

Practice!

For more practice with transmutation reactions, go to **Supplemental Practice Problems** in Appendix A.

Radioactive Decay Rates

You may be wondering how it is that there are any naturally occurring radioisotopes found on Earth. After all, if radioisotopes undergo continuous radioactive decay, won't they eventually disappear? And with the exception of synthesized radioisotopes, few new radioisotopes are being formed. Furthermore, radioisotopes have been decaying for about 4.6 billion years (the span of Earth's existence). Yet naturally occurring radioisotopes are not uncommon on Earth. The explanation for this involves the differing decay rates of isotopes.

Radioactive decay rates are measured in half-lives. A **half-life** is the time required for one-half of a radioisotope's nuclei to decay into its products. For example, the half-life of the radioisotope strontium-90 is 29 years. If you had 10.0 g of strontium-90 today, 29 years from now you would have 5.0 g left. **Table 25-4** shows how this decay continues through four half-lives of strontium-90. **Figure 25-13** presents the data from the table in terms of the percent of strontium-90 remaining after each half-life.

Table 25-4

The Decay of Strontium-90		
Number of half-lives	Elapsed time	Amount of strontium-90 present
0	0	10.0 g
1	29 years	$10.0 \text{ g} \times \left(\frac{1}{2}\right) = 5.00 \text{ g}$
2	58 years	$10.0 \text{ g} \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 2.50 \text{ g}$
3	87 years	$10.0 \text{ g} \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 1.25 \text{ g}$
4	116 years	$10.0 \text{ g} \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 0.625 \text{ g}$

The decay continues until negligible strontium-90 remains.

The data in **Table 25-4** can be summarized in a simple equation representing the decay of any radioactive element.

$$\text{Amount remaining} = (\text{Initial amount})\left(\frac{1}{2}\right)^n$$

In the equation, n is equal to the number of half-lives that have passed. Note that the initial amount may be in units of mass or number of particles. A more versatile form of the equation can be written if the exponent n is replaced by the equivalent quantity t/T , where t is the elapsed time and T is the duration of the half-life.

$$\text{Amount remaining} = (\text{Initial amount})\left(\frac{1}{2}\right)^{t/T}$$

Note that both t and T must have the same units of time. This type of expression is known as an exponential decay function. **Figure 25-13** shows the graph of a typical exponential decay function—in this case, the decay curve for strontium-90.

Each radioisotope has its own characteristic half-life. Half-lives for several representative radioisotopes are given in **Table 25-5** on the following page. As the table shows, half-lives have an incredible range of values, from millionths of a second to billions of years! To gain a greater understanding of half-life, do the **miniLAB** on page 819.

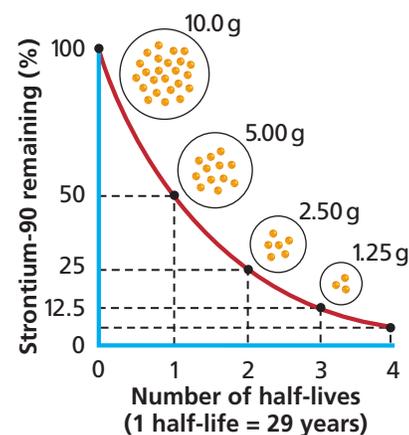


Figure 25-13

The graph shows the percent of a strontium-90 sample remaining over a period of four half-lives. With the passing of each half-life, half of the strontium-90 sample decays. Approximately what percent of strontium-90 remains after $1\frac{1}{2}$ half-lives have passed?

Table 25-5

Half-Lives of Several Radioisotopes		
Radioisotope	Symbol	Half-life
Polonium-214	${}_{84}^{214}\text{Po}$	163.7 microseconds
Cobalt-60	${}_{27}^{60}\text{Co}$	5.272 years
Radon-222	${}_{86}^{222}\text{Ra}$	3.8 days
Phosphorous-32	${}_{15}^{32}\text{P}$	14.28 days
Tritium	${}_{1}^3\text{H}$	12.32 years
Carbon-14	${}_{6}^{14}\text{C}$	5730 years
Uranium-238	${}_{92}^{238}\text{U}$	4.46×10^9 years

EXAMPLE PROBLEM 25-3



Iron-59 is used to produce this image of a patient's circulatory system.

Calculating Amount of Remaining Isotope

Iron-59 is used in medicine to diagnose blood circulation disorders. The half-life of iron-59 is 44.5 days. How much of a 2.000-mg sample will remain after 133.5 days?

1. Analyze the Problem

You are given a known mass of a radioisotope with a known half-life. You must first determine the number of half-lives that passed during the 133.5 day period. Then use the exponential decay equation to calculate the amount of the sample remaining.

Known

Initial amount = 2.000 mg
 Elapsed time (t) = 133.5 days
 Half-life (T) = 44.5 days

Unknown

Amount remaining = ? mg

2. Solve for the Unknown

Determine the number of half-lives passed during the 133.5 days.

$$\text{Number of half-lives } (n) = \frac{\text{Elapsed time}}{\text{Half-life}}$$

$$n = \frac{(133.5 \text{ days})}{(44.5 \text{ days/half-life})} = 3.00 \text{ half-lives}$$

Substitute the values for n and initial mass into the exponential decay equation and solve.

$$\text{Amount remaining} = (\text{Initial amount})\left(\frac{1}{2}\right)^n$$

$$\text{Amount remaining} = (2.000 \text{ mg})\left(\frac{1}{2}\right)^{3.00}$$

$$\text{Amount remaining} = (2.000 \text{ mg})\left(\frac{1}{8}\right)$$

$$\text{Amount remaining} = 0.2500 \text{ mg}$$

3. Evaluate the Answer

Three half-lives is equivalent to $\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)$, or $\frac{1}{8}$. The answer (0.2500 mg) is equal to $\frac{1}{8}$ of the original mass of 2.000 mg. The answer has four significant figures because the original mass was given with four significant figures. The values of n and $\frac{1}{2}$ do not affect the number of significant figures in the answer.

PRACTICE PROBLEMS

17. If gallium-68 has a half-life of 68.3 minutes, how much of a 10.0-mg sample is left after one half-life? Two half-lives? Three half-lives?
18. If the passing of five half-lives leaves 25.0 mg of a strontium-90 sample, how much was present in the beginning?
19. Using the half-life given in Table 25-5, how much of a 1.0-g polonium-214 sample remains after 818 microseconds?



For more practice with half-life problems, go to **Supplemental Practice Problems** in Appendix A.

Radiochemical Dating

Chemical reaction rates are greatly affected by changes in temperature, pressure, and concentration, and by the presence of a catalyst. In contrast, nuclear reaction rates remain constant regardless of such changes. In fact, the half-life of any particular radioisotope is constant. Because of this, radioisotopes can be used to determine the age of an object. The process of determining the age of an object by measuring the amount of a certain radioisotope remaining in that object is called **radiochemical dating**.

miniLAB

Modeling Radioactive Decay

Formulating Models Because of safety concerns, it is usually not possible to directly experiment with radioactive isotopes in the classroom. Thus, in this lab you will use pennies to model the half-life of a typical radioactive isotope. Each penny represents an individual atom of the radioisotope.

Materials 100 pennies, 5-oz. or larger plastic cup, graph paper, graphing calculator (optional)

Procedure

1. Place the pennies in the plastic cup.
2. Place your hand over the top of the cup and shake the cup several times.
3. Pour the pennies onto a table. Remove all the pennies that are "heads-up." These pennies represent atoms of the radioisotope that have undergone radioactive decay.
4. Count the number of pennies that remain ("tails-up" pennies) and record this number in the Decay Results data table as the Number of pennies remaining for trial 1.
5. Place all of the "tails-up" pennies back in the plastic cup.
6. Repeat steps 2 through 5 for as many times as needed until no pennies remain.

Decay Results

Trial number	Number of pennies remaining
0	100
1	
2	
3	
4	
5	
6	
7	
8	

Analysis

1. Make a graph of Trial number versus Number of pennies remaining from the Decay Results data table. Draw a smooth curve through the plotted points.
2. How many trials did it take for 50% of the sample to decay? 75%? 90%?
3. If the time between each trial is one minute, what is the half-life of the radioisotope?
4. Suppose that instead of using pennies to model the radioisotope you use 100 dice. After each toss, any die that comes up a "6" represents a decayed atom and is removed. How would the result using the dice compare with the result obtained from using the pennies?

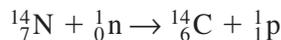


Figure 25-14

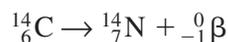
a Radiochemical dating is often used to determine the age of bones discovered at archaeological sites. Using this technique, these human bones frozen in a glacier were estimated to be from about 3000 B.C.

b Could carbon-14 dating be used to determine the age of a stone disk with a leather thong which was found with the skeleton in the glacier?

A type of radiochemical dating known as carbon dating is commonly used to measure the age of artifacts that were once part of a living organism, such as the human skeleton shown in **Figure 25-14**. Carbon dating, as its name implies, makes use of the radioactive decay of carbon-14. The procedure relies on the fact that unstable carbon-14 is formed by cosmic rays in the upper atmosphere at a fairly constant rate.



These carbon-14 atoms become evenly spread throughout Earth's biosphere, where they mix with stable carbon-12 and carbon-13 atoms. Plants then use carbon dioxide from the environment, which contains carbon-14, to build more complex molecules through the process of photosynthesis. When animals eat plants, the carbon-14 atoms that were part of the plant become part of the animal. Because organisms are constantly taking in carbon compounds, they contain the same ratio of carbon-14 to carbon-12 and carbon-13 found in the atmosphere. However, this all changes once the organism dies. After death, organisms no longer ingest new carbon compounds, and the carbon-14 they already contain continues to decay. The carbon-14 undergoes beta decay to form nitrogen-14.



This beta decay reaction has a half-life of 5730 years. Because the amount of stable carbon in the dead organism remains constant while the carbon-14 continues to decay, the ratio of unstable carbon-14 to stable carbon-12 and carbon-13 decreases. By measuring this ratio and comparing it to the nearly constant ratio present in the atmosphere, the age of an object can be estimated. For example, if an object's C-14 to (C-12 + C-13) ratio is one-quarter the ratio measured in the atmosphere, the object is approximately two half-lives, or 11 460 years, old. Carbon-14 dating is limited to accurately dating objects up to approximately 24 000 years of age.

The decay process of a different radioisotope, uranium-238 to lead-206, is commonly used to date objects such as rocks. Because the half-life of uranium-238 is 4.5×10^9 years, it can be used to estimate the age of objects that are too old to be dated using carbon-14. By radiochemical dating of meteorites, the age of the solar system has been estimated at 4.6×10^9 years of age.

Section 25.3 Assessment

- 20.** Describe the process of induced transmutation. Give two examples of induced transmutation reactions that produce transuranium elements.
- 21.** The initial mass of a radioisotope is 10.0 g. If the radioisotope has a half-life of 2.75 years, how much remains after four half-lives?
- 22.** After 2.00 years, 1.986 g of a radioisotope remains from a sample that had an original mass of 2.000 g.
 - a.** Calculate the half-life of the radioisotope.
 - b.** How much of the radioisotope remains after 10.00 years?
- 23. Thinking Critically** Compare and contrast how the half-life of a radioisotope is similar to a sporting tournament in which the losing team is eliminated.
- 24. Graphing**  A sample of polonium-214 originally has a mass of 1.0 g. Express the mass of polonium-214 remaining as a percent of the original sample after a period of one, two, and three half-lives. Graph the percent remaining versus the number of half-lives. Approximately how much time has elapsed when 20% of the original sample remains?

As you know, there are major differences between chemical and nuclear reactions. One such difference is the amount of energy the reactions produce. You already know that exothermic chemical reactions can be used to generate electricity. Coal-and oil-burning power plants are common examples. The amount of energy released by chemical reactions, however, is insignificant compared with that of certain nuclear reactions. In this section you will learn about the awesome amounts of energy produced by nuclear reactions. The energy producing capability of nuclear reactions leads naturally to their best-known application—nuclear power plants.

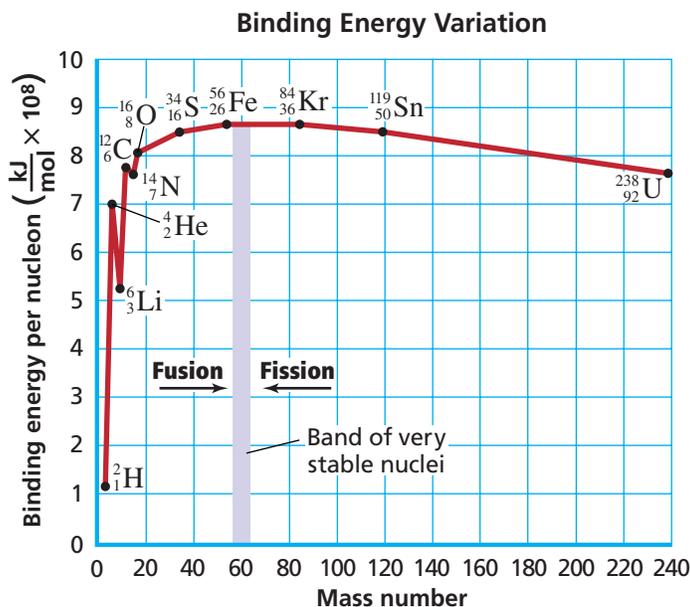
Nuclear Reactions and Energy

In your study of chemical reactions, you learned that mass is conserved. For most practical situations this is true—but, in the strictest sense, it is not. It has been discovered that energy and mass can be converted into each other. Mass and energy are related by Albert Einstein's most famous equation.

$$\Delta E = \Delta mc^2$$

In this equation, ΔE stands for change in energy (joules), Δm for change in mass (kg), and c for the speed of light (3.00×10^8 m/s). This equation is of major importance for all chemical and nuclear reactions, as it means a loss or gain in mass accompanies *any* reaction that produces or consumes energy.

The binding together or breaking apart of an atom's nucleons also involves energy changes. Energy is released when an atom's nucleons bind together. The nuclear binding energy is the amount of energy needed to break one mole of nuclei into individual nucleons. The larger the binding energy per nucleon is, the more strongly the nucleons are held together, and the more stable the nucleus is. Less stable atoms have lower binding energies per nucleon. **Figure 25-15** shows the average binding energy per nucleon versus mass number for



Objectives

- **Compare** and **contrast** nuclear fission and nuclear fusion.
- **Explain** the process by which nuclear reactors generate electricity.

Vocabulary

mass defect
nuclear fission
critical mass
breeder reactor
nuclear fusion
thermonuclear reaction

Figure 25-15

The graph shows the relationship between binding energy per nucleon and mass number. The greater the binding energy, the greater the stability of the nucleus. Light nuclei can gain stability by undergoing nuclear fusion, whereas heavy nuclei can gain stability by undergoing nuclear fission.

the elements. Note that the binding energy per nucleon reaches a maximum around a mass number of 60. This means that elements with a mass number near 60 are the most stable. You will see the importance of this relationship in the nuclear fission and fusion processes.

In typical chemical reactions, the energy produced or consumed is so small that the accompanying changes in mass are negligible. In contrast, the mass changes and associated energy changes in nuclear reactions are significant. For example, the energy released from the nuclear reaction of one kilogram of uranium is equivalent to the energy released during the chemical combustion of approximately four billion kilograms of coal!

The tremendous energy released by certain nuclear reactions is a measure of the energy required to bond the subatomic particles in nuclei together. You might be wondering where this energy comes from. The answer involves the $\Delta E = \Delta mc^2$ equation. Scientists have determined that the mass of the nucleus is always less than the sum of the masses of the individual protons and neutrons that comprise it. This difference in mass between a nucleus and its component nucleons is called the **mass defect**. Applying Einstein's $\Delta E = \Delta mc^2$ equation, you can see how the missing mass in the nucleus provides the tremendous energy required to bind the nucleus together.

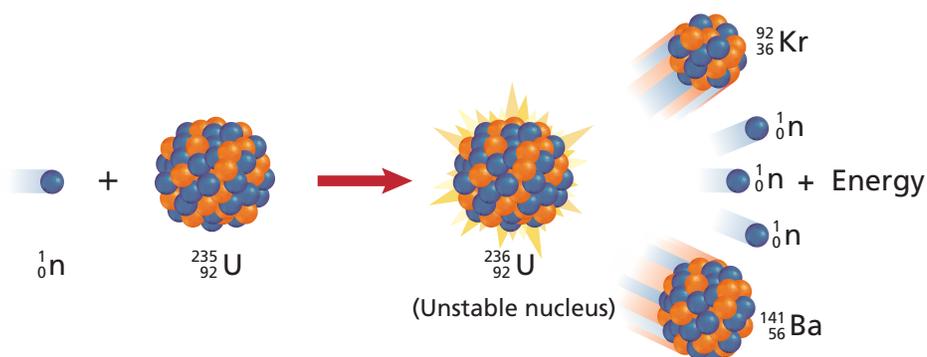
Nuclear Fission

Binding energies in **Figure 25-15** indicate that heavy nuclei would be more stable if they fragmented into several smaller nuclei. Because atoms with mass numbers around 60 are the most stable, heavy atoms (mass number greater than 60) tend to fragment into smaller atoms in order to increase their stability. The splitting of a nucleus into fragments is known as **nuclear fission**. The fission of a nucleus is accompanied by a very large release of energy.

Nuclear power plants use nuclear fission to generate power. The first nuclear fission reaction discovered involved uranium-235. As you can see in **Figure 25-16**, when a uranium-235 nucleus is struck by a neutron, it undergoes fission. Barium-141 and krypton-92 are just two of the many possible products of this fission reaction. In fact, scientists have identified more than 200 different product isotopes from the fission of a uranium-235 nucleus.

Figure 25-16

The figure shows the nuclear fission of uranium-235. When bombarded with a neutron, uranium-235 forms unstable uranium-236, which then splits into two lighter nuclei and additional neutrons. The splitting (fission) of the nucleus is accompanied by a large release of energy.



Each fission of uranium-235 releases additional neutrons. If one fission reaction produces two neutrons, these two neutrons can cause two additional fissions. If those two fissions release four neutrons, those four neutrons could then produce four more fissions, and so on, as shown in **Figure 25-17**. This self-sustaining process in which one reaction initiates the next is called a chain reaction. As you can imagine, the number of fissions and the amount of energy released can increase extremely rapidly. The explosion from an atomic bomb is an example of an uncontrolled chain reaction.

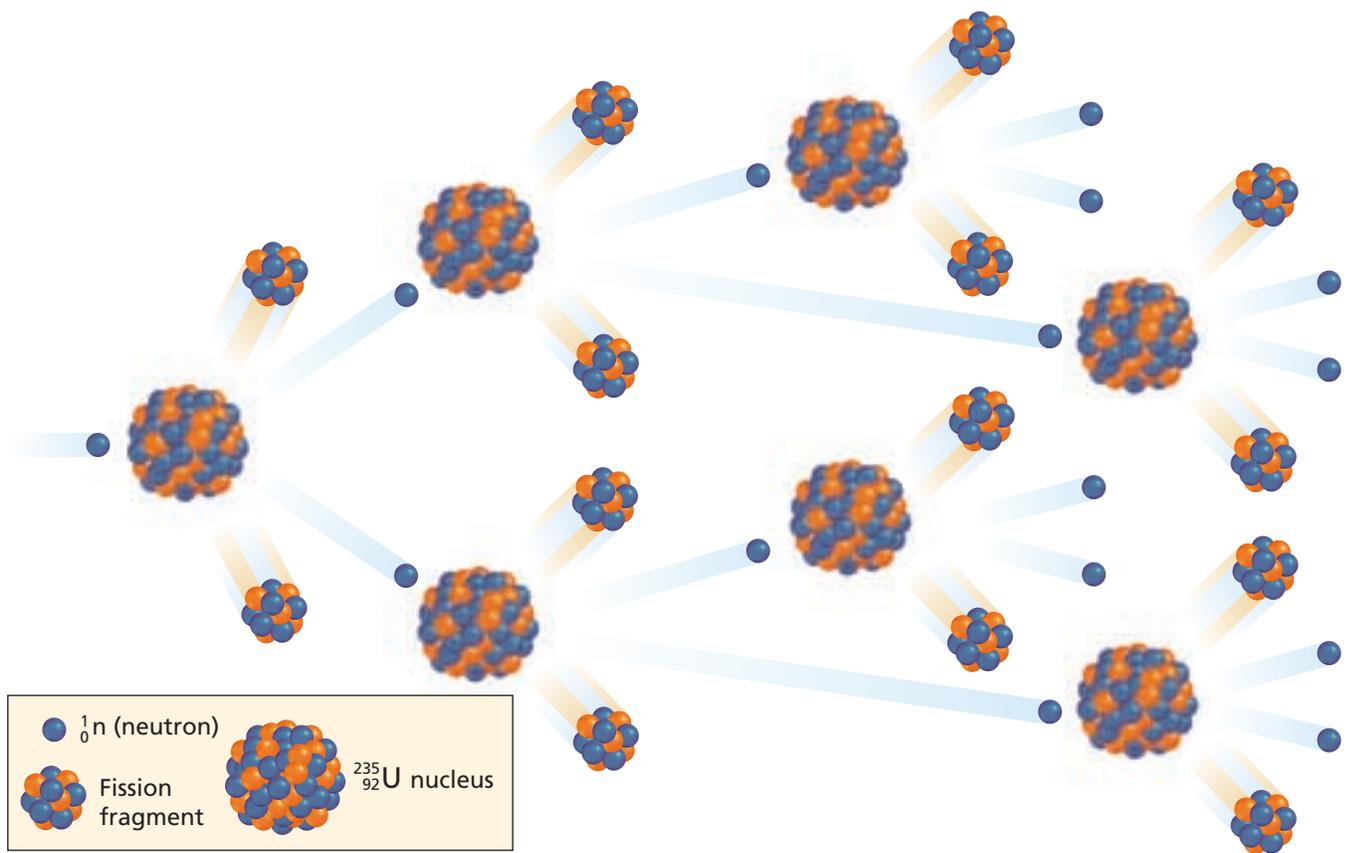


Figure 25-17

This figure illustrates the ongoing reactions characteristic of a nuclear fission chain reaction.

A sample of fissionable material must have sufficient mass in order for a fission chain reaction to occur. If it does not, neutrons escape from the sample before they have the opportunity to strike other nuclei and continue the chain reaction—the chain reaction never begins. A sample that is not massive enough to sustain a chain reaction is said to have subcritical mass. A sample that is massive enough to sustain a chain reaction has **critical mass**. When a critical mass is present, the neutrons released in one fission cause other fissions to occur. If much more mass than the critical mass is present, the chain reaction rapidly escalates. This can lead to a violent nuclear explosion. A sample of fissionable material with a mass greater than the critical mass is said to have supercritical mass. **Figure 25-18** shows the effect of mass on the initiation and progression of a fission reaction.

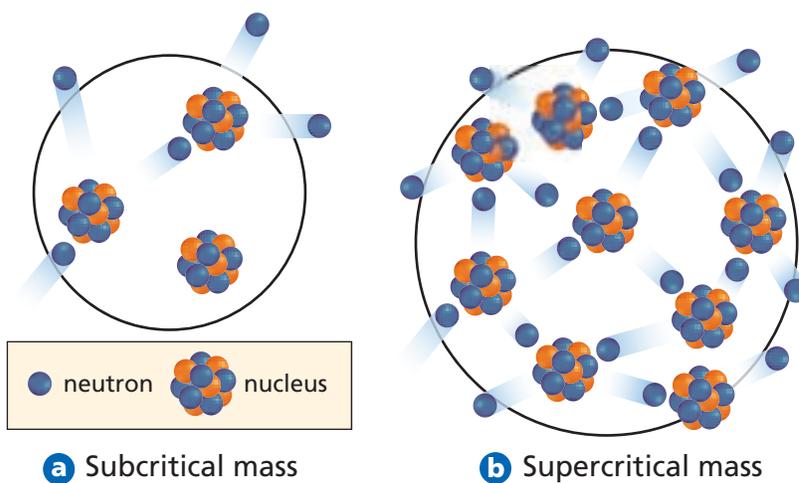


Figure 25-18

The amount of fissionable matter present determines whether a nuclear chain reaction can be sustained. **a** In a subcritical mass, the chain reaction stops because neutrons escape the sample before causing sufficient fissions to sustain the reaction. **b** In a supercritical mass, the chain reaction accelerates as neutrons cause more and more fissions to occur.

Figure 25-19

Worldwide, there are more than 400 operating nuclear power plants such as this one. Currently, 18 countries rely on nuclear power to supply at least 25% of their total electricity needs.



Nuclear Reactors

You may be familiar with the sight of a nuclear power plant, such as the one shown in **Figure 25-19**. Nuclear fission produces the energy generated by nuclear reactors. This energy is primarily used to generate electricity. What fuels the energy production of the reactor? A common fuel is fissionable uranium(IV) oxide (UO_2) encased in corrosion-resistant fuel rods. The fuel is enriched to contain 3% uranium-235, the amount required to sustain a chain reaction. Additional rods composed of cadmium or boron control the fission process inside the reactor by absorbing neutrons released during the reaction.

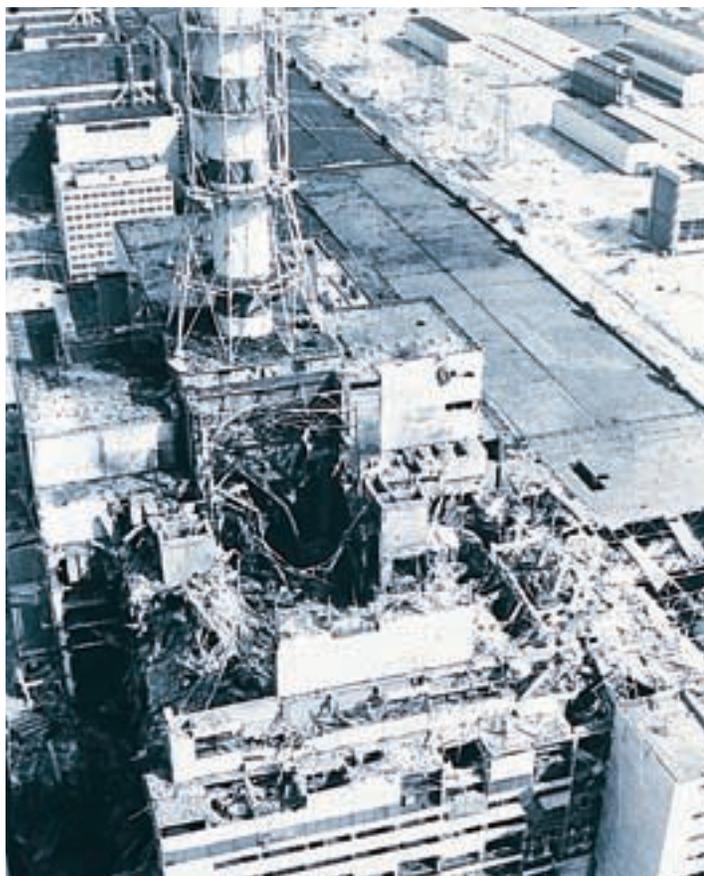
Keeping the chain reaction going while preventing it from racing out of control requires precise monitoring and continual adjusting of the control rods. Much of the concern about nuclear power plants focuses on the risk of losing control of the nuclear reactor, possibly resulting in the accidental release of harmful levels of radiation. The Three Mile Island nuclear accident in the United States in 1979 and the Chernobyl nuclear accident in Ukraine in 1986, shown in **Figure 25-20**, provide powerful examples of why controlling the reactor is critical.

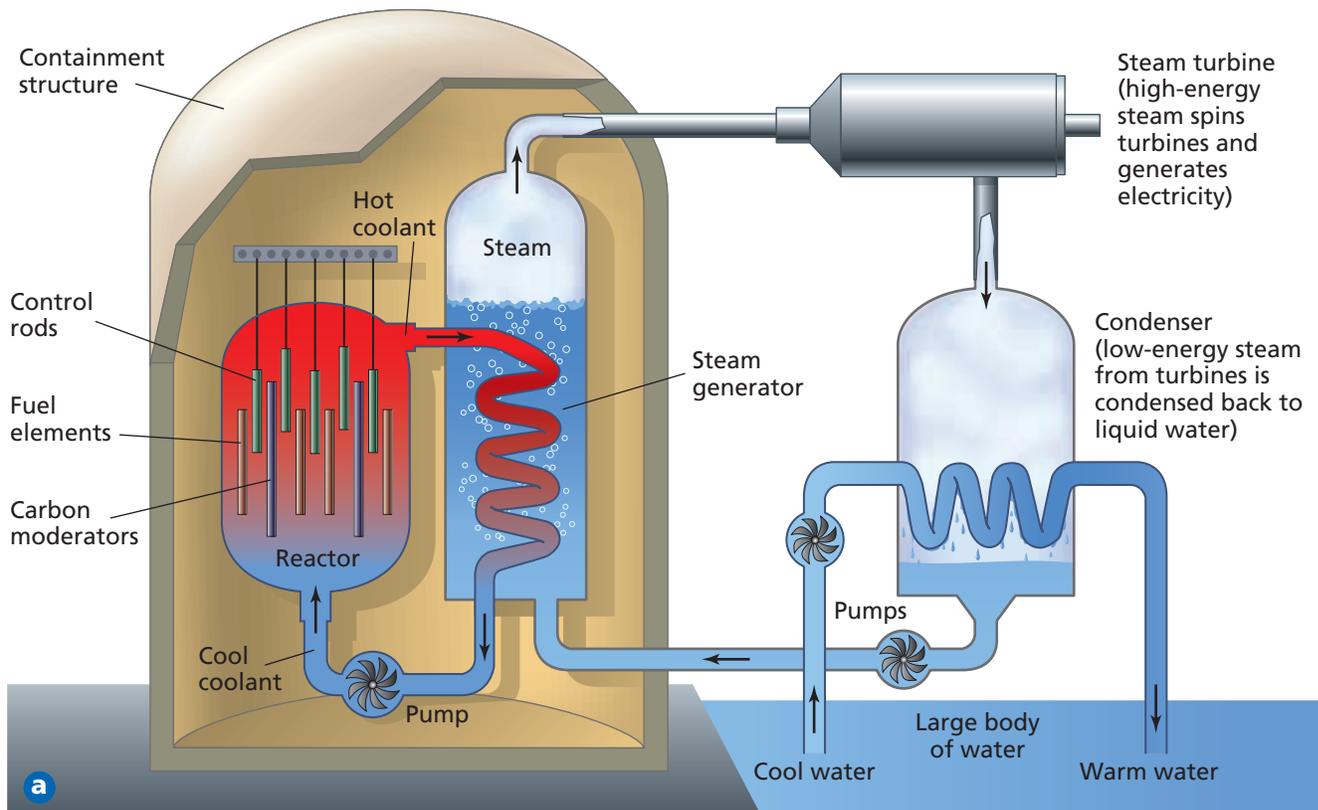
The fission within a nuclear reactor is started by a neutron-emitting source and is stopped by positioning the control rods to absorb virtually all of the neutrons produced in the reaction. The reactor core contains a reflector that acts to reflect neutrons back into the core where they will react with the fuel rods. A coolant, usually water, circulates through the reactor core to carry off the heat generated by the nuclear fission reaction. The hot coolant heats water that is used to power steam-driven turbines which produce electrical power.

In many ways, the design of a nuclear power plant and that of a fossil fuel burning power plant are very similar. In both cases heat from a reaction is used to generate steam. The steam is then used to drive turbines that produce electricity. In a typical fossil fuel power plant, the chemical combustion of coal, oil, or gas generates the heat, whereas in a nuclear power plant, a nuclear fission reaction generates the heat. Because of the hazardous radioactive fuels and fission products present at nuclear power plants, a dense concrete structure is usually built to enclose

Figure 25-20

The exploded fourth reactor of the Chernobyl nuclear power plant is seen in this aerial view taken in April, 1986. Radiation from the reactor caused great environmental damage to the surrounding areas.





the reactor. The main purpose of the containment structure is to shield personnel and nearby residents from harmful radiation. The major components of a nuclear power plant are illustrated in **Figure 25-21**.

As the reactor operates, the fuel rods are gradually depleted and products from the fission reaction accumulate. Because of this, the reactor must be serviced periodically. Spent fuel rods can be reprocessed and repackaged to make new fuel rods. Some fission products, however, are extremely radioactive and cannot be used again. These products must be stored as nuclear waste. The storage of highly radioactive nuclear waste is one of the major issues surrounding the debate over the use of nuclear power.

Approximately 20 half-lives are required for the radioactivity of nuclear waste materials to reach levels acceptable for biological exposure. For some types of nuclear fuels, the wastes remain substantially radioactive for thousands of years. A considerable amount of scientific research has been devoted to the disposal of radioactive wastes. How does improper disposal or storage of nuclear wastes affect the environment? Is this a short-term effect? Why?

Another issue is the limited supply of the uranium-235 used in the fuel rods. One option is to build reactors that produce new quantities of fissionable fuels. Reactors able to produce more fuel than they use are called **breeder reactors**. Although the design of breeder reactors poses many difficult technical problems, breeder reactors are in operation in several countries.



Figure 25-21

- a** The diagram illustrates the major parts of a nuclear power plant. The photos show
- b** the submerged reactor and
- c** the steam-driven turbines.

How does this containment structure protect the environment? What would be the effect on the environment if this structure leaked radiation?



Nuclear Fusion

Recall from the binding energy diagram in **Figure 25-15** that a mass number of about 60 has the most stable atomic configuration. Thus, it is possible to bind together two or more light (mass number less than 60) and less stable nuclei to form a single more stable nucleus. The combining of atomic nuclei is called **nuclear fusion**. Nuclear fusion reactions are capable of releasing very large amounts of energy. You already have some everyday knowledge of this fact—the Sun is powered by a series of fusion reactions as hydrogen atoms fuse to form helium atoms.



Figure 25-22

The characteristic circular-shape of the tokamak reactor is clearly seen here. The reactor uses strong magnetic fields to contain the intensely hot fusion reaction and keep it from direct contact with the interior reactor walls.

Scientists have spent several decades researching nuclear fusion. Why? One reason is that there is a tremendous abundance of lightweight isotopes, such as hydrogen, that can be used to fuel fusion reactions. Also, fusion reaction products are generally not radioactive. Unfortunately, there are major problems that have yet to be overcome on a commercially viable scale. One major problem is that fusion requires extremely high energies to initiate and sustain a reaction. The required energy, which is achieved only at extremely high temperatures, is needed to overcome the electrostatic repulsion between the nuclei in the reaction. Because of the energy requirements, fusion reactions are also known as **thermonuclear reactions**. Just how much energy is needed? The lowest temperature capable of producing a fusion reaction is 40 000 000 K! This temperature—and even higher temperatures—have been achieved using an atomic explosion to initiate the fusion process, but this approach is not practical for controlled electrical power generation.

Obviously there are many problems that must be resolved before fusion becomes a practical energy source. Another significant problem is confinement of the reaction. There are currently no materials capable of withstanding the tremendous temperatures that are required by a fusion reaction. Much of the current research centers around an apparatus called a tokamak reactor. A tokamak reactor, which you can see in **Figure 25-22**, uses strong magnetic fields to contain the fusion reaction. While significant progress has been made in the field of fusion, temperatures high enough for continuous fusion have not been sustained for long periods of time.

Section 25.4 Assessment

25. Compare and contrast nuclear fission and nuclear fusion reactions in terms of the particles involved and the changes they undergo.
26. Describe the process that occurs during a nuclear chain reaction.
27. Explain how nuclear fission can be used to generate electrical power.
28. **Thinking Critically** Present an argument supporting or opposing nuclear power as your state's primary power source. Assume that the primary source of power currently is the burning of fossil fuels.
29. **Calculating** What is the energy change (ΔE) associated with a change in mass (Δm) of 1.00 mg?

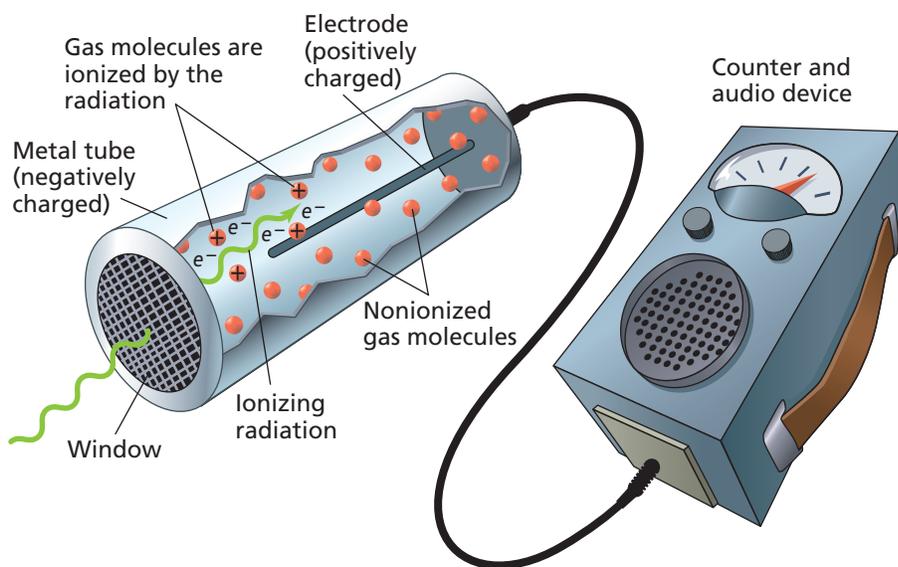
As you learned in the previous section, using nuclear fission reactions to generate electrical power is an important application of nuclear chemistry. Another very important application is in medicine, where the use of radioisotopes has made dramatic changes in the way some diseases are treated. This section explores the detection, uses, and effects of radiation.

Detecting Radioactivity

You learned earlier that Becquerel discovered radioactivity because of the effect of radiation on photographic plates. Since this discovery, several other methods have been devised to detect radiation. The effect of radiation on photographic film is the same as the effect of visible light on the film in your camera. With some care, film can be used to provide a quantitative measure of radioactivity. A film badge is a device containing a piece of radiation-sensitive film that is used to monitor radiation exposure. People who work with radioactive substances carry film badges to monitor the extent of their exposure to radiation.

Radiation energetic enough to ionize matter with which it collides is called **ionizing radiation**. The Geiger counter is a radiation detection device that makes use of ionizing radiation in its operation. As you can see in **Figure 25-23**, a Geiger counter consists of a metal tube filled with a gas. In the center of the tube is a wire that is connected to a power supply. When ionizing radiation penetrates the end of the tube, the gas inside the tube absorbs the radiation and becomes ionized. The ionized gas contains ions and free electrons. The free electrons are attracted to the wire, causing a current to flow. A current meter that is built into the Geiger counter measures the current flow through the ionized gas. This current measurement is used to determine the amount of ionizing radiation present.

Another detection device is a scintillation counter, which uses a phosphor-coated surface to detect radiation. Scintillations are bright flashes of light



Objectives

- **Describe** several methods used to detect and measure radiation.
- **Explain** an application of radiation used in the treatment of disease.
- **Describe** some of the damaging effects of radiation on biological systems.

Vocabulary

ionizing radiation
radiotracer

Figure 25-23

The Geiger counter is used to detect and measure radiation levels. The small device is usually hand-held. Ionizing radiation produces an electric current in the Geiger counter. The current is displayed on a scaled meter, while a speaker is used to produce audible sounds.

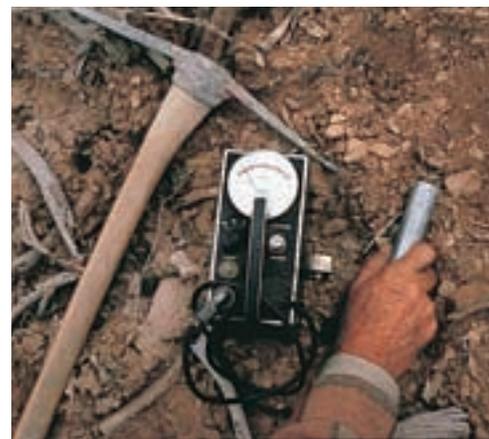


Figure 25-24

- a** Phosphors are used to aid in the night time readability of watches.
- b** Television and computer screens are phosphor coated. An electron gun aimed at the back of the screen illuminates the phosphors.



produced when ionizing radiation excites the electrons in certain types of atoms called phosphors. The number and brightness of the scintillations are detected and recorded, giving a measure of the amount of ionizing radiation present. As shown in **Figure 25-24a**, the dials of some watches are painted with a radium-containing phosphor that causes the watch to glow in the dark. Televisions and computer monitors, such as those shown in **Figure 25-24b**, also make use of phosphor screens that produce scintillations when struck by electrons.

Careers Using Chemistry

Radiation Therapist

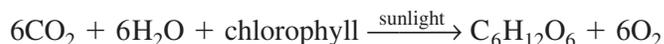
Would you like to work with a team of medical specialists that help people battle and recover from cancer? If so, consider a career as a radiation therapist.

Radiation therapists, working under the direction of physicians, administer doses of radiation to cancer patients in order to kill invasive cancer cells. Computers are used to help plan the course of treatment. The therapists also prepare the patients for treatment and explain the expected side effects. The administration of the radiation often involves the use of state-of-the-art medical equipment.

Uses of Radiation

With proper safety procedures, radiation can be very useful in many scientific experiments. Neutron activation analysis is used to detect trace amounts of elements present in a sample. Computer chip manufacturers use this technique to analyze the composition of highly purified silicon wafers. In the process, the sample is bombarded with a beam of neutrons from a radioactive source, causing some of the atoms in the sample to become radioactive. The type of radiation emitted by the sample is used to determine the types and quantities of elements present. Neutron activation analysis is a very sensitive measurement technique capable of detecting quantities of less than 1×10^{-9} g.

Radioisotopes can also be used to follow the course of an element through a chemical reaction. For example, CO_2 gas containing radioactive carbon-14 isotopes has been used to study glucose formation in photosynthesis.



Because the CO_2 containing carbon-14 is used to trace the progress of carbon through the reaction, it is referred to as a radiotracer. A **radiotracer** is a radioisotope that emits non-ionizing radiation and is used to signal the presence of an element or specific substance. The fact that all of an element's isotopes have the same chemical properties makes the use of radioisotopes possible. Thus, replacing a stable atom of an element in a reaction with one of its isotopes does not alter the reaction. Radiotracers are important in a number of areas of chemical research, particularly in analyzing the reaction mechanisms of complex, multi-step reactions.

Radiotracers also have important uses in medicine. Iodine-131, for example, is commonly used to detect diseases associated with the thyroid gland. One of the important functions of the thyroid gland is to extract iodine from the bloodstream to make the hormone thyroxine. If a thyroid problem is suspected, a doctor will give the patient a drink containing a small amount of iodine-131. The iodine-containing radioisotope is then used to monitor the functioning of the thyroid gland, as shown in **Figure 25-25**. After allowing

time for the iodine to be absorbed, the amount of iodide taken up by the thyroid is measured. This is just one example of the many ways radioisotopes are useful in diagnosing disease.

Another radiation-based medical diagnostic tool is called positron emission transaxial tomography (PET). In this procedure, a radiotracer that decays by positron emission is injected into the patient's bloodstream. Positrons emitted by the radiotracer cause gamma ray emissions that are then detected by an array of sensors surrounding the patient. The type of radiotracer injected depends on what biological function the doctor wants to monitor. For example, as shown in **Figure 25-26**, a fluorine-based radiotracer is commonly used in the PET analysis of the brain's glucose metabolism. Abnormalities in how glucose is metabolized by the brain can help in the diagnosis of brain cancer, schizophrenia, epilepsy, and other brain disorders.

Radiation can pose serious health problems for humans because of the damage it causes to living cells. Healthy cells can be badly damaged or completely destroyed by radiation. However, radiation can also destroy unhealthy cells, including cancer cells. All cancers are characterized by the runaway growth of abnormal cells. This growth can produce masses of abnormal tissue, called malignant tumors. Radiation therapy is used to treat cancer by destroying the fast-growing cancer cells. In fact, cancer cells are more susceptible to destruction by radiation than healthy ones. In the process of destroying unhealthy cells, some healthy cells are also damaged. Despite this major drawback, radiation therapy has become one of the most important treatment options used in the fight against cancer.

Biological Effects of Radiation

Although radiation has a number of medical and scientific applications, it can be very harmful. The damage produced from ionizing radiation absorbed by the body depends on several factors, such as the energy of the radiation, the type of tissue absorbing the radiation, and the distance from the source. Do the **problem-solving LAB** on the next page to see how radiation exposure is affected by distance. Gamma rays are particularly dangerous because they easily penetrate human tissue. In contrast, the skin usually stops alpha radiation, and beta radiation generally penetrates only about 1–2 cm beneath the skin.

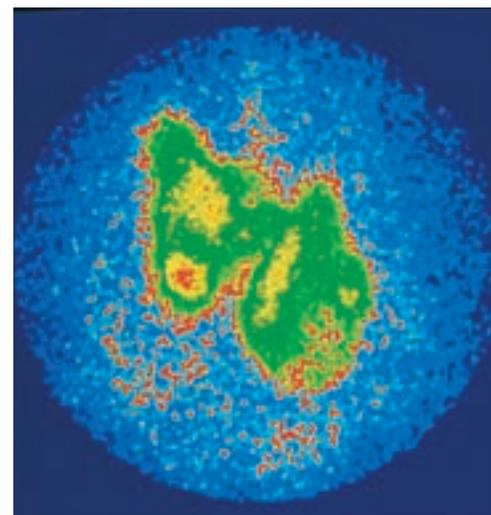


Figure 25-25

Doctors analyze radiotracer produced images such as this to determine whether a patient's thyroid gland is functioning properly.

Figure 25-26

Shown here are images produced from a fluorine-based PET scan of a patient's brain.

High-energy ionizing radiation is dangerous because it can readily fragment and ionize molecules within biological tissue. When ionizing radiation penetrates a living biological system, the ionized atoms and molecules that are generated are unstable and highly reactive. A free radical is an atom or molecule that contains one or more unpaired electrons and is one example of the highly reactive products of ionizing radiation. In a biological system, free radicals can affect a large number of other molecules and ultimately disrupt the operation of normal cells.

Ionizing radiation damage to living systems can be classified as either somatic or genetic. Somatic damage affects only nonreproductive body tissue and therefore affects the organism only during its own lifetime. Genetic damage, on the other hand, can affect offspring because it damages reproductive tissue, which may affect the genes and chromosomes in the sperm and the eggs of the organism. Somatic damage from radiation includes burns similar to those produced by fire, and cancer caused by damage to the cell's growth mechanism. Genetic damage is more difficult to study because it may not become apparent for several generations. Many scientists presently believe that exposure to any amount of radiation poses some risk to the body.

A dose of radiation refers to the amount of radiation your body absorbs from a radioactive source. Two units, the rad and rem, are commonly used to measure radiation doses. The rad, which stands for radiation-Absorbed Dose,

problem-solving LAB

How does distance affect radiation exposure?

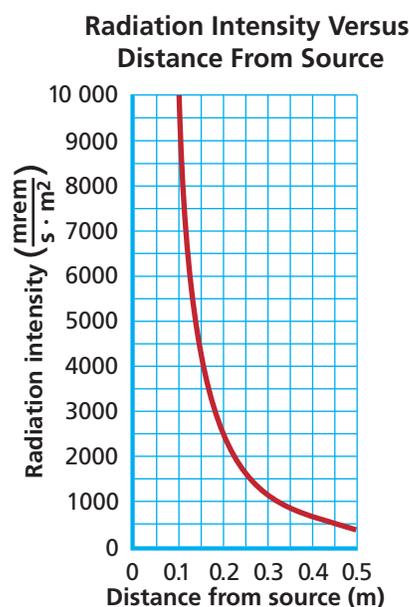
Interpreting Graphs Some people, such as X-ray technicians, must work near radioactive sources. They know the importance of keeping a safe distance from the source. The graph shows the intensity of a radioactive source versus the distance from the source.

Analysis

Note how the intensity of the radiation varies with the distance from the source. The unit of radiation intensity is millirem per second per square meter. (This is the amount of radiation striking a square meter of area each second.) Use the graph to answer the Thinking Critically questions that follow.

Thinking Critically

- How does the radiation exposure change as the distance doubles from 0.1 meter to 0.2 meter? As the distance quadruples from 0.1 meter to 0.4 meter?
- State in words the mathematical relationship described in your answer to question 1.
- The maximum radiation exposure intensity that is considered safe is $0.69 \text{ mrem/s}\cdot\text{m}^2$. At what



distance from the source has the radiation decreased to this level? (*Hint:* Use the equation

$$\frac{I_1}{d_1^2} = \frac{I_2}{d_2^2}$$

where I_1 is the radiation intensity at distance d_1 , and I_2 is the radiation intensity at distance d_2 .)

Table 25-6

Effects of Short-term Radiation Exposure	
Dose (rem)	Effects on humans
0–25	No detectable effects
25–50	Temporary decrease in white blood cell population
100–200	Nausea, substantial decrease in white blood cell population
500	50% chance of death within 30 days of exposure

is a measure of the amount of radiation that results in the absorption of 0.01 J of energy per kilogram of tissue. The dose in rads, however, does not account for the energy of the radiation, the type of living tissue absorbing the radiation, or the time of the exposure. To account for these factors, the dose in rads is multiplied by a numerical factor that is related to the radiation's effect on the tissue involved. The result of this multiplication is a unit called the rem. The rem, which stands for roentgen Equivalent for Man, is named after Wilhelm Roentgen, who, as you learned in Section 25.1, discovered X rays in 1895.

So how does radiation exposure affect you? You may be surprised to learn that you are exposed to radiation on a regular basis. A variety of sources—some naturally occurring, others not—constantly bombard your body with radiation. Read the **Chemistry and Society** at the end of this chapter to learn about one of the most common sources of exposure. Your exposure to these sources results in an average annual radiation exposure of 100–300 millirems of high-energy radiation. Remembering that the SI prefix *milli-* stands for one-thousandth, this dose of radiation is equivalent to 0.10 to 0.30 rem. To help you put this in perspective, a typical dental X ray exposes you to the virtually harmless dose of 0.0005 rem. However, both short exposures to high doses of radiation and long exposures to lower doses of radiation can be harmful. As you can see from the data in **Table 25-6**, short-term radiation exposures in excess of 500 rem can be fatal. **Table 25-7** shows your annual exposure to common radiation sources. You can perform the **CHEMLAB** at the end of this chapter to examine radiation emitted by a common substance.

Table 25-7

Average Annual Radiation exposure	
Source	Average exposure (mrem/y)
Cosmic Radiation	20–50
Radiation from ground	25–170
Radiation from buildings	10–160
Radiation from air	20–260
Human body (internal)	~20
Medical and dental X rays	50–75
Nuclear weapon testing	<1
Air travel	5
Total average	100–300

Section 25.5 Assessment

- Describe several methods used to detect and measure radiation.
- Explain one way in which nuclear chemistry is used to diagnose or treat disease.
- Compare and contrast somatic and genetic biological damage.
- Thinking Critically** Using what you know about the biological effects of radiation, explain why it is safe to use radioisotopes for the diagnosis of medical problems.
- Hypothesizing** The average person is exposed to 100–300 millirems of radiation per year. Airline crewmembers, however, are exposed to almost twice the average amount of radiation per year. Suggest a possible reason for their increased exposure to radiation.

Measuring Naturally Occurring Radiation

As you may know, some common everyday substances are radioactive. In this lab you will investigate the three naturally occurring potassium isotopes found in a common store-bought salt substitute. Two of potassium's isotopes, potassium-39 (93.1%) and potassium-41 (6.89%) are stable. However, potassium-40 (0.01%) decays by beta emission to form stable calcium-40. You will first measure the background radiation level, and then use that information to determine the radiation due to the beta decay of potassium-40. You will also measure radiation at various locations around your school.

Problem

How can you determine if a substance contains radioactive isotopes?

Objectives

- **Measure** background radiation and radiation emitted by a radioactive isotope.
- **Compare** the level of background radiation to the level of radiation emitted by a radioactive isotope.

Materials

CBL system	CBL-P adapter
RADIATIN software program	TI-Graphlink cable
graphing calculator	petri dish (with lid)
link-to-link cable	salt substitute or pure potassium chloride (KCl)
Student Radiation Monitor	balance

Safety Precautions



- Always wear safety goggles and a lab apron.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Include any necessary safety precautions, procedure notes, and a data table.
3. What is an isotope? A radioactive isotope?
4. Write the nuclear equation for the radioactive decay of potassium-40 by beta emission. Identify the "parent" and "daughter" nuclides in the decay.
5. Using nuclide-stability rules, form a hypothesis that explains why calcium-40 should be a more stable nuclide than potassium-40.

Radiation Level Data	
Data point	Counts/minute

Procedure

1. Load the program RADIATIN into the graphing calculator.
2. Connect the graphing calculator to the CBL system using the link-to-link cable. Connect the CBL system to the Student Radiation Monitor using the CBL-P adapter. Turn on all devices. Set the Student Radiation Monitor on the audio setting and place it on top of an empty petri dish.

3. Start the RADIATIN program. Go to MAIN MENU. Select 4:SET NO. SAMPLE. Choose 20 for the number of samples in each reading. Press ENTER.
4. Select 1:COLLECT DATA from the MAIN MENU. Select 4:TRIGGER/PROMPT from the COLLECTING MODE menu. Press ENTER to begin collecting data. After a few seconds, the calculator will ask you to enter a PROMPT. Enter 1 (because this is the first data point) and press ENTER. Choose 1:MORE DATA under TRIGGER/PROMPT.
5. Press ENTER to begin the next data point. A graph will appear. When asked to enter the next PROMPT, enter the number that appears at the top right corner of the calculator screen, and then press ENTER. Choose 1:MORE DATA under TRIGGER/PROMPT.
6. Repeat step 5 until you have at least five data points. This set of data is the background level of radiation from natural sources.
7. Use the balance to measure out 10.0 g salt substitute or pure potassium chloride (KCl). Pour the substance into the center of the petri dish so that it forms a small mound. Place the Student Radiation Monitor on top of the petri dish so that the Geiger Tube is positioned over the mound. Repeat step 5 until you have at least five data points.
8. When you are finished collecting data, choose 2:STOP AND GRAPH under TRIGGER/PROMPT. The data points (PROMPTED) are stored in L1, the counts per minute (CTS/MIN) are stored in L2. Press ENTER to view a graph of data.

Cleanup and Disposal

1. Return the salt substitute or potassium chloride (KCl) used in the experiment to the container prepared by your instructor.
2. Disconnect the lab setup and return all equipment to its proper place.

Analyze and Conclude

1. **Collecting Data** Record the data found in L1 and L2 (STAT, EDIT) in the Radiation Level Data table.
2. **Graphing Data** Graph the data from L1 and L2.  Use the graph from the graphing calculator as a guide.



3. **Interpreting Data** What is the average background radiation level in counts/minute?
4. **Interpreting Data** What is the average radiation level in counts/minute for the potassium-40 isotope found in the salt substitute?
5. **Observing and Inferring** How can you explain the difference between the background radiation level and the radiation level of the salt substitute?
6. **Thinking Critically** Is the data for the background radiation and the radiation from the potassium-containing sample consistent or random in nature? Propose an explanation for the pattern or lack of pattern seen in the data.
7. **Error Analysis** Describe several ways to improve the experimental procedure so it yields more accurate radiation level data.

Real-World Chemistry

1. Arrange with your teacher to plan and perform a field investigation using the experimental setup from this experiment to measure the background level radiation at various points around school or around town. Propose an explanation for your findings.
2. Using the procedure in this lab, determine if other consumer products contain radioisotopes. Report on your findings.

CHEMISTRY and Society

The Realities of Radon

In the late 1800s, doctors identified lung cancer as a prevalent cause of mine worker deaths throughout the world. It became clear that something in the air of underground mines was related to the deaths. Today scientists believe radon gas was the primary culprit. The heaviest known gas, with a density nine times greater than air, radon is naturally occurring and highly radioactive. Regardless of your location, radon, a gas you cannot see, smell, or taste, is an ever-present part of the air you breathe.

Why be concerned?

Radon is produced when uranium-238, an element present in many rock layers and soil, decays. Because buildings are built on top of soil, and sometimes constructed of stone and brick, radon gas often accumulates inside.

Radon exposure is the second leading cause of lung cancer in the United States, causing about 14 000 deaths annually. Because of the health threat posed by radon, Congress mandated the Environmental Protection Agency (EPA) make the reduction of indoor radon gas levels a national goal.

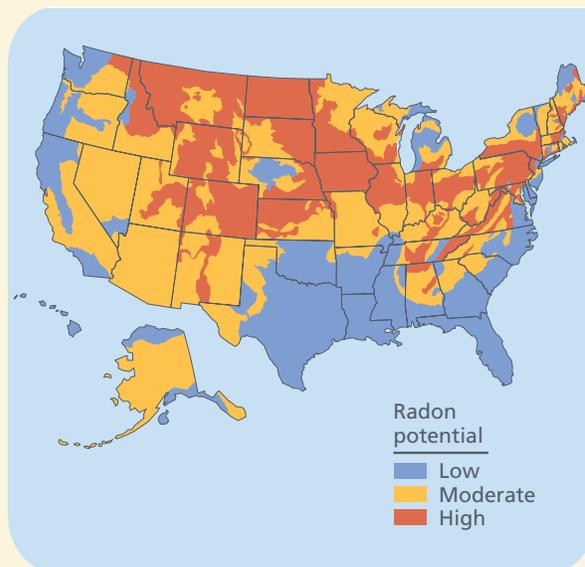
Although radon is an inert gas, its decay produces highly radioactive heavy metals. Polonium is one of the radioactive decay products. Breathing radon-containing air results in the collection of polonium in the tissues of the lungs. Inside the lungs, the polonium decays further, emitting alpha particles and gamma radiation. This radiation damages the cells of the lungs.

How much is too much?

The average outdoor radiation level due to radon is 0.04 pCi/L (picocuries per liter of air). The maximum level considered to be safe is 4.0 pCi/L.

What is your risk factor?

Examine the map showing the risk of radon in your area of the United States. Are you in a high-risk area? Fortunately, only 8% of U.S. homes are believed to have radon levels exceeding 4 pCi/L. Homes at highest risk include those built on uranium-rich rock and thin, dry, permeable soil. However, even buildings in high-risk areas may be



safe if there are few pathways into the building, or if the building's air is regularly exchanged with the outside air. People worried about radon levels can test the quality of their air. Do-it-yourself kits are available that test for high levels of radon.

Investigating the Issue

- 1. Communicating Ideas** Research radon levels in your area, as well as any radon-related regulations governing real estate transactions.
- 2. Debating the Issue** Knowing the radon levels in your area, should it be mandated that property inspections include radon testing prior to a sale? If radon levels above 4 pCi/L are discovered at your home or school, how should the problem be addressed?

CHEMISTRY
Online



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Visit the Chemistry Web site at science.glencoe.com to find links to more information about radon exposure.

Summary

25.1 Nuclear Radiation

- Wilhelm Roentgen discovered X rays in 1895. Henri Becquerel, Marie Curie and Pierre Curie pioneered the fields of radioactivity and nuclear chemistry.
- Radioisotopes, isotopes of atoms with unstable nuclei, emit radiation to attain more stable atomic configurations.

25.2 Radioactive Decay

- The strong nuclear force acts on protons and neutrons within a nucleus to hold the nucleus together.
- The neutron-to-proton (n/p) ratio of a nucleus affects its stability. Stable n/p ratios range from 1 : 1 for small atoms to 1.5 : 1 for the largest atoms.
- Atomic number and mass number are conserved in nuclear reactions and equations.
- **Table 25-3** summarizes the characteristics of the five primary types of radioactive decay.

25.3 Transmutation

- The conversion of an atom of one element to an atom of another by radioactive decay processes is called transmutation. Induced transmutation is the process of bombarding nuclei with high-velocity charged particles in order to create new elements.
- A half-life is the time required for half the atoms in a radioactive sample to decay. Every radioisotope has a characteristic half-life.

- Radiochemical dating is a technique for determining the age of an object by measuring the amount of certain radioisotopes remaining in the object.

25.4 Fission and Fusion of Atomic Nuclei

- Nuclear fission is the splitting of large nuclei into smaller more stable fragments. Fission reactions release large amounts of energy.
- In a chain reaction, one reaction induces others to occur. A sufficient mass of a fissionable material must be present for a fission chain reaction to occur.
- Nuclear reactors make use of nuclear fission reactions to generate steam. The steam is used to drive turbines that produce electrical power.
- Nuclear fusion is the process of binding smaller nuclei into a single larger and more stable nucleus. Fusion reactions release large amounts of energy, but require extremely high temperatures.

25.5 Applications and Effects of Nuclear Reactions

- Geiger counters, scintillation counters, and film badges are used to detect and measure radiation.
- Radiotracers, which emit non-ionizing radiation, are used to diagnose disease and to analyze complex chemical reaction mechanisms.
- Both short-term and long-term radiation exposure can cause damage to living cells.

Key Equations and Relationships

- Exponential decay function:
Amount remaining = (Initial amount) $\left(\frac{1}{2}\right)^n$ or Amount remaining = (Initial amount) $\left(\frac{1}{2}\right)^{t/T}$ (p. 817)

Vocabulary

- band of stability (p. 811)
- breeder reactor (p. 825)
- critical mass (p. 823)
- electron capture (p. 812)
- half-life (p. 817)
- induced transmutation (p. 815)
- ionizing radiation (p. 827)
- mass defect (p. 822)
- nuclear fission (p. 822)
- nuclear fusion (p. 826)
- nucleon (p. 810)
- positron (p. 812)
- positron emission (p. 812)
- radioactive decay series (p. 814)
- radiochemical dating (p. 819)
- radioisotope (p. 807)
- radiotracer (p. 828)
- thermonuclear reaction (p. 826)
- transmutation (p. 815)
- transuranium element (p. 815)
- strong nuclear force (p. 810)
- X ray (p. 809)

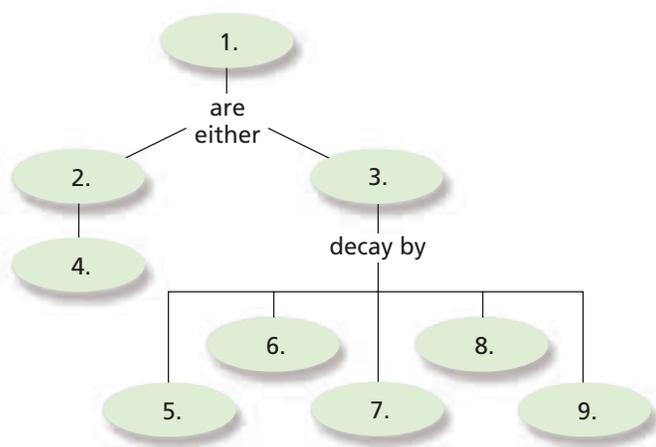


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Concept Mapping

35. Complete the concept map using the following terms: positron emission, alpha decay, atoms, unstable, do not decay, beta decay, stable, gamma emission, and electron capture.



Mastering Concepts

36. What did the Curies contribute to the field of radioactivity and nuclear chemistry? (25.1)
37. Compare and contrast chemical reactions and nuclear reactions in terms of energy changes and the particles involved. (25.1)
38. Match each numbered choice on the right with the correct radiation type on the left. (25.1)
- | | |
|----------|------------------------------|
| a. alpha | 1. high speed electrons |
| b. beta | 2. 2+ charge |
| c. gamma | 3. no charge |
| | 4. helium nucleus |
| | 5. blocked very easily |
| | 6. electromagnetic radiation |
39. What is a nucleon? (25.2)
40. What is the difference between a positron and an electron? (25.2)
41. Describe the differences between a balanced nuclear equation and a balanced chemical equation. (25.2)
42. What is the strong nuclear force? On what particles does it act? (25.2)
43. Explain the difference between positron emission and electron capture. (25.2)
44. Explain the relationship between an atom's neutron-to-proton ratio and its stability. (25.2)
45. What is the significance of the band of stability? (25.2)
46. What is a radioactive decay series? When does the decay series end? (25.2)
47. What scientist performed the first induced transmutation reaction? What element was synthesized in the reaction? (25.3)
48. Define transmutation. Are all nuclear reactions also transmutation reactions? Explain. (25.3)
49. Why are some radioisotopes found in nature, while others are not? (25.3)
50. What are some of the characteristics of transuranium elements? (25.3)
51. Using the band of stability diagram shown in Figure 25-8 would you expect $^{39}_{20}\text{Ca}$ to be radioactive? Explain. (25.3)
52. Carbon-14 dating makes use of a specific ratio of two different radioisotopes. Define the ratio used in carbon-14 dating. Why is this ratio constant in living organisms? (25.3)
53. Why is carbon-14 dating limited to objects that are approximately 24 000 years old or less? (25.3)
54. What is a mass defect? (25.4)
55. Describe some of the current limitations of fusion as a power source. (25.4)
56. Describe some of the problems of using fission as a power source. (25.4)
57. What is a chain reaction? Give an example of a nuclear chain reaction. (25.4)
58. How is binding energy per nucleon related to mass number? (25.4)
59. Explain how binding energy per nucleon is related to fission and fusion reactions. (25.4)
60. Discuss how the amount of a fissionable material present affects the likelihood of a chain reaction. (25.4)
61. Explain the purpose of control rods in a nuclear reactor. (25.4)
62. What is a breeder reactor? Why were breeder reactors developed? (25.4)
63. Why does nuclear fusion require so much heat? How is heat contained within a tokamak reactor? (25.4)
64. What is ionizing radiation? (25.5)

65. What is the difference between somatic and genetic damage? (25.5)
66. What property of isotopes allows radiotracers to be useful in studying chemical reactions? (25.5)
67. List several applications that involve phosphors. (25.5)

Mastering Problems

Radioactive Decay (25.2)

68. Calculate the neutron-to-proton ratio for each of the following atoms.
- | | |
|---------------|--------------|
| a. tin-134 | d. nickel-63 |
| b. silver-107 | e. carbon-14 |
| c. carbon-12 | f. iron-61 |
69. Complete the following equations:
- a. ${}_{83}^{214}\text{Bi} \rightarrow {}_2^4\text{He} + ?$
- b. ${}_{93}^{239}\text{Np} \rightarrow {}_{94}^{239}\text{Pu} + ?$
70. Write a balanced nuclear equation for the alpha decay of americium-241.
71. Write a balanced nuclear equation for the beta decay of bromine-84.
72. Write a balanced nuclear equation for the beta decay of selenium-75.

Induced Transmutation (25.3)

73. Write a balanced nuclear equation for the induced conversion of carbon-13 to carbon-14.
74. Write the balanced nuclear equation for the alpha particle bombardment of ${}_{99}^{253}\text{Es}$. One of the reaction products is a neutron.
75. Write a balanced nuclear equation for the induced transmutation of uranium-238 into californium-246 by bombardment with carbon-12.
76. Write the balanced nuclear equation for the alpha particle bombardment of plutonium-239. The reaction products include a hydrogen atom and two neutrons.

Half-Life (25.3)

77. The half-life of tritium (${}^3_1\text{H}$) is 12.3 years. If 48.0 mg of tritium is released from a nuclear power plant during the course of a mishap, what mass of the nuclide will remain after 49.2 years? After 98.4 years?
78. Technetium-104 has a half-life of 18.0 minutes. How much of a 165.0 g sample remains after 90.0 minutes?
79. Manganese-56 decays by beta emission and has a half-life of 2.6 hours. How many half-lives are there in 24 hours? How many mg of a 20.0 mg sample will remain after five half-lives?

80. A 20.0 g sample of thorium-234 has a half-life of 25 days. How much will remain as a percentage of the original sample after 90 days?
81. The half-life of polonium-218 is 3.0 minutes. If you start with 20.0 g, how long will it be before only 1.0 g remains?
82. A sample of an unknown radioisotope exhibits 8540 decays per second. After 350.0 minutes, the number of decays has decreased to 1250 per second. What is the half-life?
83. Phosphorous-32 has a half-life of 14.32 days. Write and graph an equation for the amount remaining of phosphorous-32 after t days if the sample initially contains 150.0 mg of phosphorous-32.
84. Plot the exponential decay curve for a period of five half-lives for the decay of thorium-234 given in Problem 80. How much time has elapsed when 30% of the original sample remains?
85. A rock once contained 1.0 mg of uranium-238, but now contains only 0.257 mg. Given that the half-life for uranium-238 is 4.5×10^9 y, how old is the rock?

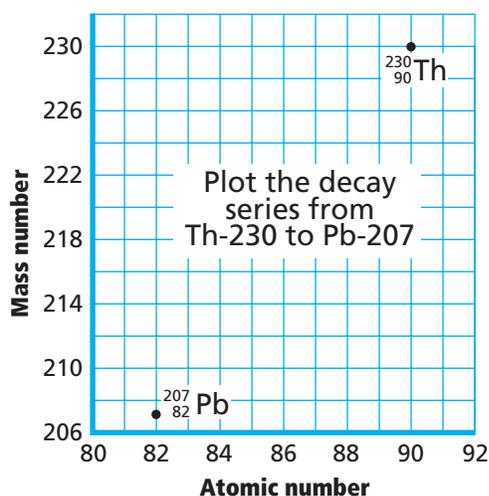
Mixed Review

Sharpen your problem-solving skills by answering the following.

86. A sample initially contains 150.0 mg of radon-222. After 11.4 days, the sample contains 18.7 mg of radon-222. Calculate the half-life.
87. Write a balanced nuclear equation for the positron emission of nitrogen-13.
88. Describe the penetration power of alpha, beta, and gamma radiation.
89. Plot the exponential decay curve for a period of six half-lives using the data given for technetium-104 in Problem 78. How much time has elapsed when 60% of the original sample remains?
90. What information about an atom can you use to predict whether or not it will be radioactive?
91. A bromine-80 nucleus can decay by gamma emission, positron emission, or electron capture. What is the product nucleus in each case?
92. Explain how a Geiger counter measures levels of ionizing radiation present.
93. The half-life of plutonium-239 is 24 000 years. What fraction of nuclear waste generated today will be present in the year 3000?

Thinking Critically

- 94. Making and Using Graphs** Thorium-231 decays to lead-207 in a stepwise fashion by emitting the following particles in successive steps: β , α , α , β , α , α , β , α . Plot each step of the decay series on a graph of mass number versus atomic number. Label each plotted point with the symbol of the radioisotope.



- 95. Analyzing Information** Scientists often use heavy ion bombardment to produce new elements. Balance the following nuclear reactions involving heavy ion bombardments.
- ${}^6_3\text{Li} + {}^{63}_{28}\text{Ni} \rightarrow ?$
 - ${}^{252}_{98}\text{Cf} + {}^{10}_5\text{B} \rightarrow ?$
- 96. Applying Concepts** Chemical treatment is often used to destroy harmful chemicals. For example, bases neutralize acids. Why can't chemical treatment be applied to destroy the fission products produced in a nuclear reactor?
- 97. Applying Concepts** A radioactive decay series that begins with ${}^{237}_{93}\text{Np}$ ends with the formation of stable ${}^{209}_{83}\text{Bi}$. How many alpha emissions and how many beta emissions are involved in the sequence of radioactive decays?

Writing in Chemistry

- 98.** Research and report on the lives of Marie Curie and her daughter, Irene Curie Joliot. What kind of scientific training did each receive? What was it like to be a female chemist in their time? What other discoveries did each make beyond those made in the field of nuclear chemistry?
- 99.** Research and write a report on the nuclear power plant accidents that occurred at Three Mile Island in Pennsylvania and Chernobyl in the former Soviet

Union. What went wrong in each case? How much radiation escaped and entered the surrounding environment? What were the health effects of the released radiation?

- 100.** Evaluate environmental issues associated with nuclear wastes. Research the Yucca Mountain nuclear waste disposal plan, the Hanford nuclear site, or a local nuclear facility. Prepare a poster or multimedia presentation on your findings.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 101.** Identify each of the following as a chemical property or a physical property. (Chapter 3)
- The element mercury has a high density.
 - Solid carbon dioxide sublimates at room temperature.
 - Zinc oxidizes when exposed to air.
 - Sucrose is a white crystalline solid.
- 102.** Why does the second period of the periodic table contain eight elements? (Chapter 6)
- 103.** Draw the following molecules and show the locations of hydrogen bonds between the molecules. (Chapter 9)
- two water molecules
 - two ammonia molecules
 - one water molecule and one ammonia molecule
- 104.** Name the process taking place in each situation described below. (Chapter 13)
- a solid air freshener cube getting smaller and smaller
 - dewdrops forming on leaves in the morning
 - steam rising from a hot spring
 - a crust of ice forming on top of a pond
- 105.** If the volume of a sample of chlorine gas is 4.5 L at 0.65 atm and 321 K, what volume will the gas occupy at STP? (Chapter 14)
- 106.** The temperature of 756 g of water in a calorimeter increases from 23.2°C to 37.6°C. How much heat was given off by the reaction in the calorimeter? (Chapter 16)
- 107.** Explain what a buffer is and why buffers are found in body fluids. (Chapter 19)
- 108.** Explain how the structure of benzene can be used to explain its unusually high stability compared to other unsaturated cyclic hydrocarbons. (Chapter 22)

Use these questions and the test-taking tip to prepare for your standardized test.

- All of the following are true of alpha particles EXCEPT _____.
 - they have the same composition as helium nuclei
 - they carry a charge of $2+$
 - they are more penetrating than β particles
 - they are represented by the symbol ${}^4_2\text{He}$
- In the first steps of its radioactive decay series, thorium-232 decays to radium-228, which then decays to actinium-228. The balanced nuclear equations describing these first two decay steps are _____.
 - ${}^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + {}^0_{-1}\beta$,
 ${}^{228}_{88}\text{Ra} \rightarrow {}^{228}_{89}\text{Ac} + {}^0_0\gamma$
 - ${}^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + {}^4_2\text{He}$,
 ${}^{228}_{88}\text{Ra} \rightarrow {}^{228}_{89}\text{Ac} + {}^0_{-1}\beta$
 - ${}^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + 2 {}^0_0\gamma$,
 ${}^{228}_{88}\text{Ra} \rightarrow {}^{228}_{89}\text{Ac} + {}^0_{-1}\beta$
 - ${}^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + {}^4_2\text{He}$,
 ${}^{228}_{88}\text{Ra} + {}^0_{-1}\beta \rightarrow {}^{228}_{89}\text{Ac}$
- In the early 1930s, van de Graaf generators were used to generate neutrons by bombarding stable beryllium atoms with deuterons (${}^2_1\text{H}$), the nuclei of deuterium atoms. A neutron is released in the reaction. The balanced nuclear equation describing this induced transmutation is _____.
 - ${}^9_4\text{Be} + {}^2_1\text{H} \rightarrow {}^{10}_5\text{B} + {}^1_0\text{n}$
 - ${}^9_4\text{Be} + {}^2_1\text{H} \rightarrow {}^8_3\text{B} + {}^1_0\text{n}$
 - ${}^9_4\text{Be} \rightarrow {}^{10}_5\text{B} + {}^2_1\text{H} + {}^1_0\text{n}$
 - ${}^9_4\text{Be} + {}^2_1\text{H} \rightarrow {}^{11}_5\text{B} + {}^1_0\text{n}$
- Geologists use the decay of potassium-40 in volcanic rocks to determine their age. Potassium-40 has a half-life of 1.26×10^9 years, so it can be used to date very old rocks. If a sample of rock 3.15×10^8 years old contains 2.73×10^{-7} g of potassium-40 today, how much potassium-40 was originally present in the rock?
 - 2.30×10^{-7} g
 - 1.71×10^{-8} g
 - 3.25×10^{-7} g
 - 4.37×10^{-6} g
- Based on its position relative to the band of stability, ${}^{80}\text{Zn}$ will undergo _____.
 - beta decay
 - positron emission
 - electron capture
 - all of the above
- Based on its position relative to the band of stability, phosphorous-26 will transmute into which of the following isotopes in the first step of its radioactive decay series?
 - ${}^{22}_{13}\text{Al}$
 - ${}^{26}_{16}\text{S}$
 - ${}^{26}_{17}\text{Cl}$
 - ${}^{26}_{14}\text{Si}$
- Which of the following is NOT characteristic of a sample of a fissionable element capable of sustaining a chain reaction?
 - The element has a mass number < 60 .
 - The sample of the element possesses critical mass.
 - The element's atoms have low binding energy per nucleon.
 - The element's nuclei are more stable if split into smaller nuclei.
- The immense amount of energy released by the Sun is due to which of the following reactions occurring within its core?
 - nuclear fission
 - gamma decay
 - nuclear fusion
 - alpha decay
- All of the following are beneficial applications of radioactivity EXCEPT _____.
 - tracing the path of an element through a complex reaction
 - mutating genetic material using ionizing radiation
 - diagnosing brain disorders using PET scans
 - destroying cancer cells with radiation therapy

Interpreting Graphs Use **Figure 25-8** on page 811 to answer questions 5–7.

- Calcium-35 will undergo positron emission because _____.
 - it lies above the line of stability
 - it lies below the line of stability
 - it has a high neutron-to-proton ratio
 - it has an overabundance of neutrons



TEST-TAKING TIP

Don't Be Afraid To Ask For Help If you're practicing for a test and you find yourself stuck, unable to understand why you got a question wrong, or unable to answer it in the first place, ask someone for help. As long as you ask for help before the test, you'll do fine!

Chemistry in the Environment

What You'll Learn

- ▶ You will apply the concepts you have learned to a study of the chemistry of Earth's environment.
- ▶ You will explore the ways in which human activities affect the chemical nature of the environment.

Why It's Important

A knowledge of chemistry will help you develop a deeper appreciation for the environment and an increased awareness of the effect of human activities on Earth's air, land, and water.



Visit the Chemistry Web site at science.glencoe.com to find links about chemistry in the environment.

This satellite photo of Earth shows some of our planet's unique features: features that allow it to sustain life.



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DISCOVERY LAB



Clarification of Water

Municipal water-treatment plants sometimes have to remove mud from the water after rainwater washes dirt into reservoirs. One way to do this is to use chemicals that cause the mud particles to clump together and settle to the bottom of the treatment tank, a process called sedimentation.

Safety Precautions



Lime is an irritant. It is harmful if inhaled or swallowed.

Procedure

1. Place two spoonfuls of soil into a 600-mL beaker. Add water and stir until the water is muddy.
2. Divide the muddy water into two equal samples in two 400-mL beakers. Label the beakers Lime and Control.
3. Slowly sprinkle a spoonful of powdered lime (CaO) over the surface of the muddy water in the beaker labeled Lime. Do not add lime to the control beaker. Allow both beakers to remain undisturbed. Observe the beakers every five minutes until 15 minutes have elapsed. Record your observations.
4. Filter the treated water through a funnel that contains a paper filter. Collect the filtrate in a 250-mL beaker. Observe whether or not the water samples become clearer after filtration.
5. Repeat step 4 for the control, using a clean filter and beaker.

Analysis

How long did it take for the treated water to begin to clear? How did the addition of lime speed up the process of sedimentation? Did the paper filter help the water become clearer?

Materials

soil
600-mL beaker
400-mL beakers (2)
spoons (3)
powdered lime (CaO)
funnel
filter paper (2)
250-mL beakers (2)
ring stand with ring
stirring rod

Section

26.1

Earth's Atmosphere

Objectives

- **Describe** the structure and composition of Earth's atmosphere.
- **Identify** common chemical reactions in the atmosphere.
- **Analyze** how human activities affect the atmosphere.

Vocabulary

atmosphere
troposphere
stratosphere

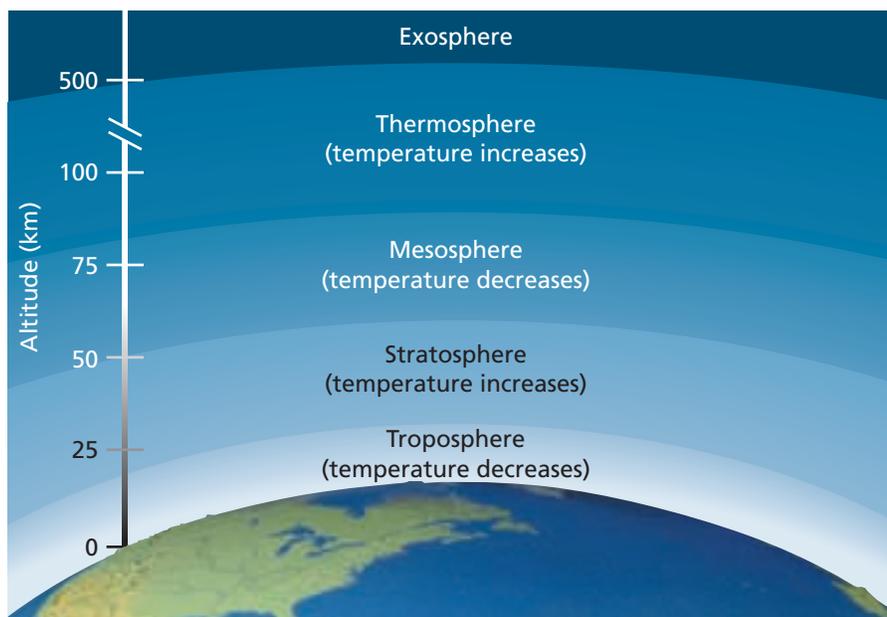
To the best of our knowledge, Earth is the only planet capable of supporting life as we know it. One glance at the photo on the opposite page helps explain why. See those wispy clouds? They are part of a protective envelope, the **atmosphere**, that blankets Earth and plays a key role in maintaining life.

A Balanced Atmosphere

Take a deep breath. You've just inhaled part of Earth's atmosphere. The atmosphere extends from Earth's surface to hundreds of kilometers into space. A largely gaseous zone, the atmosphere contains the air we breathe, the clouds overhead, and the all-important substances that protect Earth and its inhabitants from the Sun's most powerful radiation. Chemical reactions that occur in the atmosphere help maintain a balance among the different atmospheric gases, but human activities, such as burning fossil fuels, can change this balance.

Figure 26-1

Earth's atmosphere has five layers that vary in composition, temperature, altitude, and pressure. Which layer would you expect to have the greatest pressure?



Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Structure of Earth's Atmosphere

Earth's atmosphere is divided into five layers based on altitude and temperature variation. The lowest layer—the **troposphere**—extends from Earth's surface to a height of approximately 15 km, as shown in **Figure 26-1**. Temperatures in the troposphere generally decrease with increasing altitude, reaching a minimum of -58°C at 12 km. Rain, snow, wind, and other weather phenomena occur in this layer. We live our entire lives within the troposphere. Only astronauts in spacecraft go beyond its reach.

Above the troposphere, temperatures increase with altitude, reaching a maximum of nearly 2°C at about 50 km. This region of the atmosphere is called the **stratosphere**. The stratosphere contains a layer of ozone, a gas that helps shield Earth's surface from the Sun's harmful ultraviolet radiation. Ozone protects Earth by absorbing solar radiation, which raises the temperature of the stratosphere in the process. You read about the ozone layer in Chapter 1 as you began your study of chemistry.

Beyond the stratosphere lie the mesosphere and the thermosphere. Temperatures in the mesosphere decrease with altitude because there is little ozone in the air to absorb solar radiation. The thermosphere is a region of rapidly increasing temperatures. This is because the relatively few gas molecules in this region have extremely high kinetic energies. At an altitude of about 200 km, temperatures can reach 1000°C .

The outermost layer of the atmosphere is the exosphere. Extending from about 500 km outward, the exosphere marks the transition from Earth's atmosphere to outer space. There is no clear boundary between the two, however. There are simply fewer and fewer molecules of gas at increasingly higher altitudes. Eventually, there are so few molecules that, for all practical purposes, Earth's atmosphere has ended.

Composition of Earth's Atmosphere

Just as the temperature of the atmosphere varies by altitude, so does its composition. Roughly 75% of the mass of all atmospheric gases is found in the troposphere. Nitrogen and oxygen make up the vast majority of these gases. However, there are a number of minor components. The percent composition of dry air near sea level is summarized in **Table 26-1**.

Table 26-1

Composition of Dry Air Near Sea Level	
Component	Percent
Nitrogen	78
Oxygen	20.9
Argon	0.934
Carbon dioxide	0.033
Neon	0.00182
Helium	0.00052
Methane	0.00015
Krypton	0.00011
Hydrogen	0.00005
Nitrous oxide	0.00005
Xenon	0.000009





Figure 26-2

When sunlight strikes dust particles in the atmosphere, it is scattered in all directions. The blue portion of the light is scattered away from your eyes, leaving the reds, oranges, and yellows you see here.

In addition to gases, the troposphere contains solids in the form of dust, salts, and ice. Dust—tiny particles from Earth’s surface, ash, soot, and plant pollen—enters the atmosphere when it is lifted from Earth’s surface and carried by wind. The dust particles can cause sunsets to show spectacular colors, as you can see in **Figure 26-2**. Salts are picked up from ocean spray. Ice is present in the form of snowflakes and hailstones. The troposphere also contains liquids, the most common of which is water in the form of droplets found in clouds. Water is the only substance that exists as a solid, liquid, and gas in Earth’s atmosphere.

Because Earth has a strong gravitational field, most gases are held relatively close to the surface of the planet. Only lighter gases such as helium and hydrogen rise to the exosphere. These light gases, consisting of molecules that have absorbed radiation from the Sun, move so rapidly that they are able to escape Earth’s gravitational field. Thus, there’s a small but constant seepage of gases into outer space.

Chemistry in the Outer Atmosphere

Earth is constantly being bombarded with radiation and high-energy particles from outer space. The short-wavelength, high-energy ultraviolet (UV) radiation is the most damaging to living things. Because this radiation is capable of breaking the bonds in DNA molecules, it can cause cancer and genetic mutations. How can we exist here? Life as we know it is possible primarily because two processes, which occur in the thermosphere and the exosphere, shield us from most of this radiation.

Photodissociation Photodissociation is a process in which high-energy ultraviolet solar radiation is absorbed by molecules, causing their chemical bonds to break. In the upper atmosphere, the photodissociation of oxygen absorbs much of the high-energy UV radiation and produces atomic oxygen.



The amount of atomic oxygen in the atmosphere increases with increasing altitude. Why is this true? Because so much of the UV radiation that enters the atmosphere is absorbed in the photodissociation of oxygen, most of the oxygen above 150 km is in the form of atomic oxygen. Below this altitude, the percentage of atomic oxygen decreases, and most of the oxygen in the troposphere is in the form of O_2 molecules.

Photoionization The second process that absorbs high-energy solar radiation is photoionization, which occurs when a molecule or atom absorbs sufficient energy to remove an electron. Molecular nitrogen and oxygen, as well as atomic oxygen, undergo photoionization in the upper atmosphere. Note that a positively charged particle is produced for every negatively charged electron in the atmosphere, so neutrality of charge is maintained.



Ultraviolet radiation with the very highest energy is absorbed during photodissociation and photoionization in the upper atmosphere. Because most of this harmful radiation does not reach Earth's surface, life can exist.

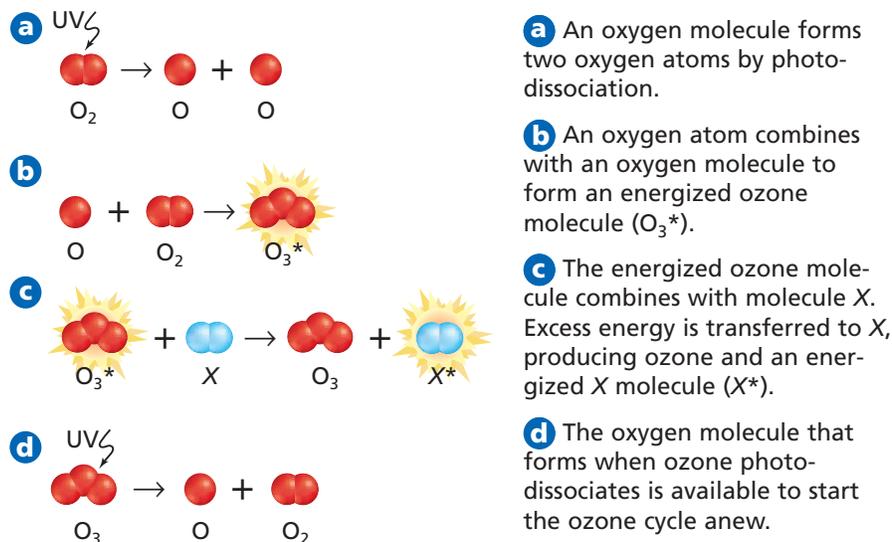
Chemistry in the Stratosphere

In addition to light gases, the upper atmosphere—more specifically, the stratosphere—contains a substance called ozone. In Chapter 1, you learned about the ozone layer and how it contributes to shielding Earth's surface from ultraviolet radiation. Now let's examine the chemical reactions that lead to the formation of ozone in the stratosphere.

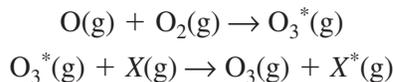
Formation of ozone Although the UV radiation with the very highest energy has been absorbed by photoionization reactions in the outer atmosphere, much of the UV radiation that has sufficient energy to cause photodissociation still reaches the stratosphere. In the stratosphere, these ultraviolet waves are absorbed by O_2 molecules, which are more plentiful here than in the upper atmosphere. The O_2 molecules are split into two atoms of oxygen. These highly reactive atoms immediately collide with other O_2 molecules, forming ozone (O_3). The O_3 molecule that forms is highly unstable because its bonds contain excess energy that was gained from the UV radiation. To achieve stability, the energized O_3 molecule must lose this excess energy by colliding with another atom or molecule, denoted in **Figure 26-3** as molecule X , and transferring energy to it. Usually, N_2 or O_2 molecules are most abundant and serve as energy-absorbing molecules for the reaction.

Figure 26-3

Ozone molecules are formed in the stratosphere. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



The formation of O_3 and the transfer of excess energy to molecule X are summarized below. Remember that X is most often N_2 or O_2 . An asterisk on a molecule indicates that the molecule is energized.

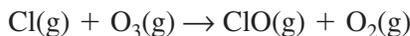


The ozone molecule does not last long after being formed. Ozone can absorb high-energy solar radiation and photodissociate back to O_2 and O , thus starting the ozone cycle anew. If high-energy radiation were not absorbed by ozone and oxygen molecules, it would penetrate the troposphere and possibly damage or kill living things.

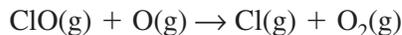
Thinning of the ozone layer As you read in earlier chapters, F. Sherwood Rowland and Mario Molina have proposed that chlorofluorocarbons (CFCs) are responsible for the thinning of the ozone layer that has been observed in the stratosphere. CFCs such as Freon-11 ($CFCl_3$) and Freon-12 (CF_2Cl_2), shown in **Figure 26-4**, have been used as coolants in refrigerators and air conditioners and as propellants in spray cans. They also have been used as foaming agents in the manufacture of some plastics. CFCs are highly stable molecules in the troposphere. After release, they eventually diffuse into the stratosphere, where they become unstable in the presence of the high-energy radiation found at this altitude. The CFCs absorb high-energy UV radiation, causing photodissociation in which the carbon-chlorine bonds in the molecules are broken.



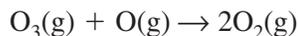
The atomic chlorine formed from this reaction reacts with stratospheric ozone to form chlorine monoxide (ClO) and O_2 .



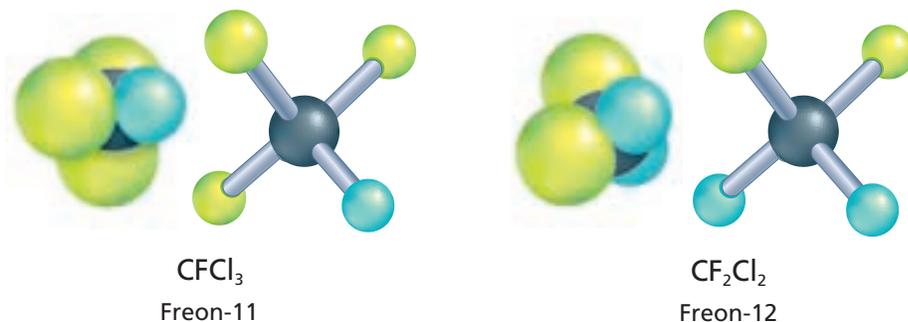
The chlorine monoxide then combines with free oxygen atoms to regenerate free chlorine atoms and oxygen molecules.



Because Cl atoms speed up the depletion of ozone and are first used and then re-formed, they function as a catalyst. The net result of these reactions is the conversion of ozone into O_2 .



Each Cl atom remains in the stratosphere for about two years before it is destroyed in other reactions. During this time, it is capable of catalyzing the breakdown of about 100 000 molecules of ozone.



Careers Using Chemistry

Environmental Health Inspector

Are you concerned about the quality of the air you breathe, the water you drink, and the food you eat? If so, you might become an environmental health inspector.

Also called sanitarians, many environmental health inspectors work for local, state, and federal governments. They analyze air, water, and food to identify possible contaminants. They might inspect dairies, restaurants, hospitals, processing plants, and industries. When they find contaminants, they look for the source, stop the pollution, and require the polluter to clean it up.

Figure 26-4

Freon-11 and Freon-12 are very stable compounds due to the strength of the $C-F$ and $C-Cl$ bonds. Because these bonds are so strong, Freon molecules are not easily broken down in the troposphere. In the stratosphere, however, high-energy radiation causes the bonds to break.

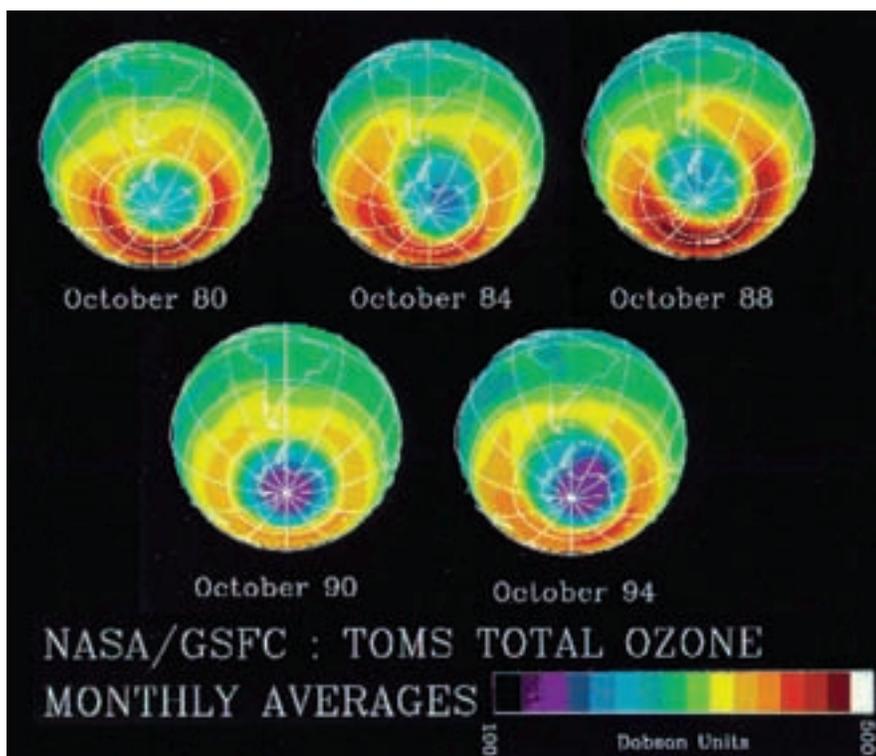


Figure 26-5

Ozone distribution in the southern hemisphere was mapped by a NASA satellite from 1980 to 1994. Observe the area of lowest ozone concentration (purple) over Antarctica. How has it changed?

Chemistry in the Troposphere

The composition of the troposphere varies from area to area, mostly as a result of human activities. Some compounds that may be present in only minute quantities in the air above a remote rain forest area may be present in much higher quantities over a large city. If you live in a city, you may have noticed that the air occasionally appears hazy, as it does in **Figure 26-6**. The haze that you see is a form of air pollution more commonly known as smog.

Photochemical smog In large cities such as Los Angeles, Denver, and Mexico City, a hazy, brown blanket of smog is created when sunlight reacts with pollutants in the air. Because the smog forms with the aid of light, it is called photochemical smog. The smog-producing pollutants enter the

Figure 26-6

Smog reduces visibility, irritates the eyes, and damages vegetation. This smog was photographed over New York City.



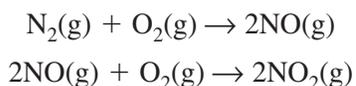


Figure 26-7

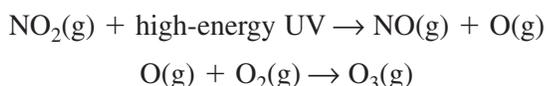
The combustion of petroleum in automobiles is a major source of smog. Sunlight acts on the exhaust gases of automobiles, factories, power plants, and homes to produce smog.

troposphere when fossil fuels such as coal, natural gas, and gasoline are burned. **Figure 26-7** shows one of the major sources of smog.

The burning of fossil fuels in internal combustion engines causes nitrogen and oxygen to react, forming nitrogen oxides such as NO and NO₂.



The NO₂, in turn, photodissociates in the presence of high energy UV that penetrates through the upper atmosphere to form atomic oxygen, which combines with O₂ to form ozone.



In the troposphere, ozone can irritate your eyes and lungs and increase your susceptibility to asthma and pneumonia.

Photochemical smog also contains unburned hydrocarbons and carbon monoxide, both of which come from the exhaust of automobile engines. These pollutants can be reduced or eliminated from the atmosphere in a variety of ways. Cleaner running engines and catalytic converters greatly reduce NO and hydrocarbon levels. Strict federal tailpipe emission standards are encouraging automobile manufacturers to develop new cars that are powered by electricity or alternative fuels such as natural gas.

Acid rain Sulfur-containing compounds are normally present in small quantities in the troposphere. However, human activities have greatly increased the concentration of these compounds in the air. Sulfur dioxide (SO₂) is the most harmful of the sulfur-containing compounds.

Most of the sulfur dioxide in the troposphere is produced when coal and oil that contain high concentrations of sulfur are burned in power plants. The sulfur dioxide that forms is oxidized to sulfur trioxide (SO₃) when it combines with either O₂ or O₃ in the atmosphere. When SO₃ reacts with moisture in the air, sulfuric acid is formed.



Earth Science

CONNECTION

Not all soils and bodies of water are equally affected by acid rain. Soils that contain calcium carbonate from limestone and lakes that are surrounded by and lie on top of calcium carbonate-rich soils are protected from much of the acid rain damage. The hydrogen ion in acid rain combines with the dissolved carbonate ions from the limestone to produce hydrogen carbonate ions. The hydrogen carbonate/carbonate solution in the water and soil serves as a buffer, absorbing additional hydrogen ions from acid rain and maintaining a stable pH. Areas that contain silicate rocks and soil are affected to a greater degree by acid rain. Southern Michigan and the Adirondack Mountain region of northern New York are especially susceptible to acid rain damage because of the large percentage of silicate rocks found there.



Figure 26-8

Acid-rain damage to buildings, statues, and trees amounts to billions of dollars a year.

to outdoor surfaces can be extensive. The acid in precipitation reacts with CaCO_3 , the major component of marble and limestone. What products are produced by this reaction?

Acidic air pollution is created also when nitrogen oxides from car exhausts combine with atmospheric moisture to form nitric acid. In either case, when this acidic moisture falls to Earth as rain or snow, it is known as acid rain. You can model the formation of acid rain in the **miniLAB** on this page.

Acid rain increases the acidity of some types of soil, resulting in the removal of essential nutrients from the soil. The loss of nutrients adversely affects the area's vegetation, leaving trees and other plants with less resistance to disease, insects, and bad weather. Acid rain also increases the acidity of streams, rivers, and lakes, which can kill or harm aquatic life. As

Figure 26-8 shows, damage to trees and

miniLAB

Acid Rain in a Bag

Making a Model Acid precipitation often falls to Earth hundreds of kilometers away from where the pollutant gases enter the atmosphere because the gases diffuse through the air and are carried by the wind. In this lab, you will model the formation of acid rain to observe how the damage caused by acid varies with the distance from the source of pollution. You also will observe another factor that affects the amount of damage caused by acid rain.

Materials plastic petri dish bottom; 1-gallon zipper-close, plastic bag; white paper; droppers; 0.04% bromocresol green indicator; 0.5M KNO_2 ; 1.0M H_2SO_4 ; clock or watch

Procedure



1. Place 25 drops of 0.04% bromocresol green indicator of varying sizes in the bottom half of a plastic petri dish so that they are about 1 cm apart. Be sure that there are both large and small drops at any given distance from the center. Leave the center of the petri dish empty.
2. Place a zipper-close plastic bag on a piece of white paper.
3. Carefully slide the petri dish containing the drops of indicator inside the plastic bag.

4. In the center of the petri dish, place one large drop of 0.5M KNO_2 . To this KNO_2 drop, add two drops of 1.0M H_2SO_4 . **CAUTION:** KNO_2 and H_2SO_4 are skin irritants. Carefully seal the bag. Observe whether the mixing of these two chemicals produces any bubbles of gas. This is the pollution source.
5. Observe and compare the color changes that take place in the drops of indicator of different sizes and distances from the pollution source. Record your observations every 15 seconds.
6. To clean up, carefully remove the petri dish from the bag, rinse it with water, then dry it.

Analysis

1. As the gas reacts with water in the drops, two acids form, $2\text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 + \text{HNO}_2$. What are these acids?
2. Did the small or large drops change color first? Why?
3. Did the distance of the indicator drops from the pollution source have an effect on how quickly the reaction occurred? Explain.
4. State two hypotheses that will explain your observations, and incorporate the answers from questions 2 and 3 in your hypotheses.
5. Based on your hypotheses in question 4, what can you infer about the damage done to plants by acid fog as compared with acid rain?

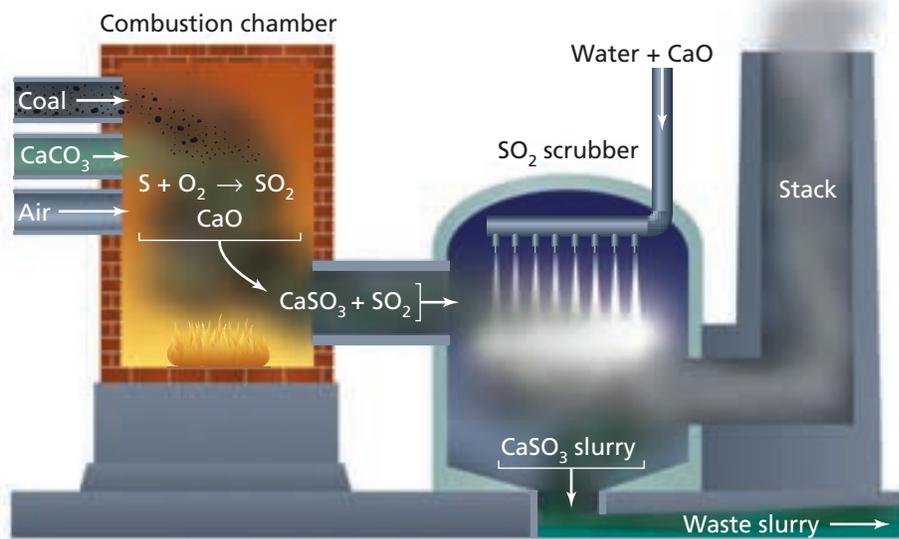


Figure 26-9

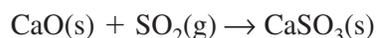
Follow the process as this power plant removes the sulfur from coal to reduce air pollution. Where is the sulfur found at the end of the process?

As human dependence on fossil fuels has increased, so has the acidity of rain. About 200 years ago, rain had a pH between 6 and 7.6, almost neutral. Today, it is common in many regions for rain to have a pH between 4 and 4.5, or even lower. In fact, the pH of rain in Wheeling, West Virginia, has been measured at 1.8, midway between the acidity of lemon juice and that of stomach acid. What do you think might account for this extreme acidity?

Given the damaging effects of acid rain, measures have been taken to reduce SO_2 emissions into the environment. High-sulfur coal may be washed before it is burned in power plants, or the SO_2 that is produced when coal burns may be removed during the burning process. This removal is accomplished by adding powdered limestone (CaCO_3) to the combustion chamber along with the coal, as shown in **Figure 26-9**. The limestone is decomposed to lime (CaO) and carbon dioxide.



The lime then reacts with SO_2 to form calcium sulfite.



About half of the SO_2 from coal is removed by adding solid limestone powder to the combustion material. The rest of the SO_2 must be removed by "scrubbing" the reaction gas with a shower of CaO and water. In this step, the remaining SO_2 is converted into solid CaSO_3 , which precipitates as a watery slurry. Most of the sulfur is removed from the coal-burning process.

Section 26.1 Assessment

1. Compare and contrast the major layers of the atmosphere in terms of temperature, altitude, and composition.
2. Write the equations for the chemical reactions that lead to the formation of ozone in the stratosphere.
3. What is the source of the sulfur that contributes to the formation of acid rain?
4. **Thinking Critically** Hydrogen is the most abundant element in the universe. Why is it so rare in Earth's atmosphere?
5. **Applying Concepts** Explain why the following statement is true: Ozone in the troposphere is considered a pollutant, but ozone in the stratosphere is essential for life on Earth.

Objectives

- **Trace** the cycle of water in the environment.
- **Identify** the chemical composition of seawater.
- **Describe** methods of desalination, and **relate** the shortage of freshwater in some regions to the development of desalination techniques.
- **Outline** the steps of a water-treatment process.

Vocabulary

hydrosphere
salinity
desalination

Look back at the photo of Earth in the chapter opener. The blue oceans clearly illustrate the abundance of water on Earth. Water is a rare substance on other planets. On Earth, however, water is found as a solid, liquid, and gas throughout the environment. You have read that the presence of an atmosphere is critical for the existence of living things on Earth. Because living things cannot exist without water, it, too, is essential for life.

The Hydrosphere

Water is the most abundant substance in the human body and the most common substance on Earth, covering approximately 72% of the surface of this planet. All the water found in and on Earth's surface and in the atmosphere is collectively referred to as the **hydrosphere**. More than 97% of this surface water is located in the oceans. Another 2.1% is frozen in glaciers and polar ice caps. That leaves a meager 0.6% available as liquid freshwater.

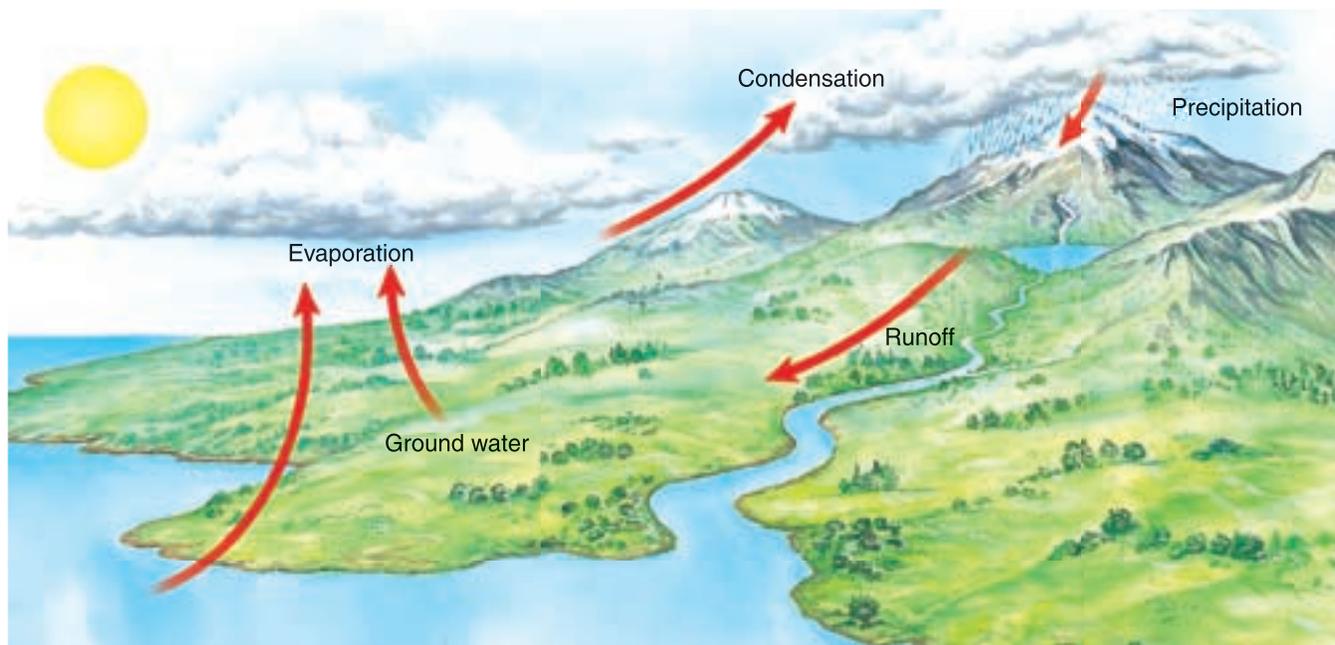
The Water Cycle

Both seawater and freshwater move through Earth's atmosphere, its surface, and below its surface in a process known as the water cycle. You may also see the water cycle referred to as the hydrologic cycle. In this cycle, water continually moves through the environment by the processes of evaporation, condensation, and precipitation. The Sun provides the energy for these processes. Follow **Figure 26-10** as you read about the water cycle.

Solar radiation causes liquid water to evaporate into a gaseous state. The resulting water vapor rises in the atmosphere and cools. As it cools, the water vapor again becomes a liquid when it condenses on dust particles in the air, forming clouds. Clouds are made of millions of tiny water droplets that collide with each other to form larger drops. When the drops grow so large that they can no longer stay suspended in the clouds, they fall to Earth in the form of precipitation—rain, snow, sleet, or hail.

Figure 26-10

The water cycle, powered by the Sun, circulates water through the atmosphere, Earth's surface, and below its surface.



Most of the falling precipitation soaks into the ground and becomes part of groundwater, the underground water that collects in small spaces between soil and rock particles. If the soil becomes saturated with water, the excess water flows along Earth's surface and into lakes and streams. This is called runoff.

Look again at **Figure 26-10**. As you can see, water cycles through the atmosphere, on Earth's surface, and under the surface. Can atmospheric processes affect the hydrosphere? The answer is most definitely yes. Processes that take place in the atmosphere, such as the formation of acid rain, can have a direct impact on the hydrosphere. The interrelatedness of the components of the environment is an important concept to keep in mind as you explore Earth's water, beginning with the vast and mighty seas.

Earth's Oceans

If you've ever swallowed a gulp of seawater, you know that it tastes quite different from tap water. Seawater contains dissolved salts, which give the water a salty taste. Where do the salts come from? Rivers and groundwater dissolve elements such as calcium, magnesium, and sodium from rocks and minerals. Flowing rivers then transport these elements to the oceans. Erupting volcanoes add sulfur and chlorine to the oceans as well.

Salinity is a measure of the mass of salts dissolved in seawater. It is usually measured in grams of salt per kilogram of seawater. The average salinity of ocean water is about 35 g per kg, so ocean water contains about 3.5% dissolved salts. Most of these salts dissociate in water and are present in the form of ions. **Table 26-2** lists the ions in seawater. Note that chlorine and sodium are the most abundant elements in seawater. Although Earth's oceans are vast, the proportions and quantities of dissolved salts are nearly constant in all areas. Indeed, they have stayed almost the same for hundreds of millions of years. Why is this so? As rivers, volcanoes, and atmospheric processes add new substances to seawater, elements are removed from the oceans by biological processes and sedimentation. Thus, the oceans are considered to be in a steady state with respect to salinity.

Desalination In some areas of the world, such as the Middle East, freshwater is scarce. Can the people in these areas drink the much more abundant ocean water? Because seawater has a high salinity, it can't be consumed by living organisms. If humans are to use ocean water for drinking and for irrigation of crops, the salts must first be removed. The removal of salts from seawater to make it usable by living things is called **desalination**.

Dissolved salts can be removed from seawater by distillation, a simple process in which the seawater is boiled to evaporate the volatile water. Pure water vapor is collected and condensed, leaving the nonvolatile salts behind. This process is quite energy intensive and is not practical for a large-scale operation. It is rarely used commercially.

You learned in Chapter 15 that osmosis is the flow of solvent molecules through a semipermeable membrane from a region of higher solvent concentration to a region of lower solvent concentration. If a high-enough pressure is applied to the system, osmosis can be reversed; that is, the solvent can be forced to flow from low to high concentrations of solvent. This process is called reverse osmosis. In a modern desalination plant, seawater is forced under pressure into cylinders containing hollow, semipermeable fibers. A cylinder holds more than three million fibers, each of which has the diameter of a human hair. The water molecules pass inward through the walls of the fibers, and the salts are held back. Desalinated water flows through the inside of the fibers and is collected.

Table 26-2

Ionic Composition of Seawater	
Dissolved ion	Percent in seawater
Chloride (Cl^-)	55.04
Sodium (Na^+)	30.61
Sulfate (SO_4^{2-})	7.68
Magnesium (Mg^{2+})	3.69
Calcium (Ca^{2+})	1.16
Potassium (K^+)	1.10
Hydrogen carbonate (HCO_3^-)	0.41
Bromide (Br^-)	0.19
All others	0.12

Biology

CONNECTION

What happens if you drink seawater? The ions in the water enter your blood from the digestive tract. Your kidneys can't remove the ions fast enough, so the ion concentration remains high in the blood. Water moves out of your cells by osmosis, causing dehydration. The cells shrivel and begin to malfunction. Meanwhile, your body begins to swell as the volume of blood and the fluid surrounding cells increases due to the additional water. Dehydration causes the thirst center of your brain to signal your kidneys to stop making urine, so the ions aren't removed from your body.

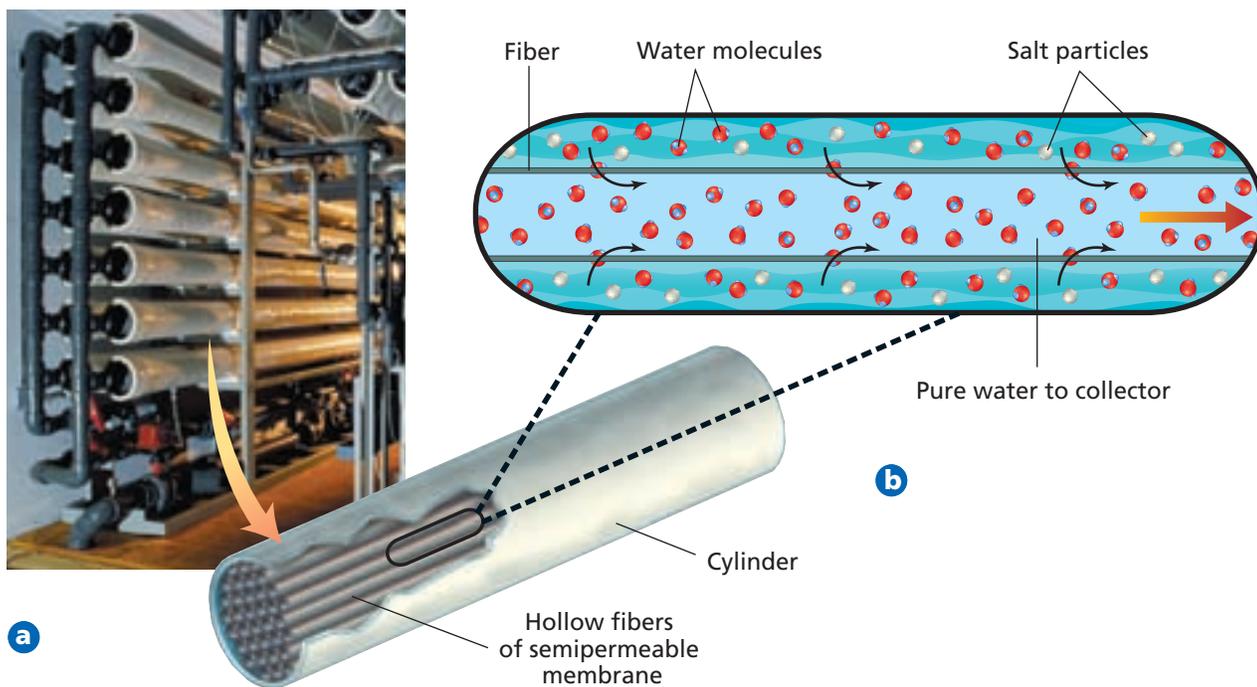


Figure 26-11

- a** This room in a desalination plant houses thousands of cylinders.
- b** Each cylinder contains millions of tiny fibers across which reverse osmosis takes place. Maintaining the high pressure necessary for this process makes it very energy intensive.

Figure 26-11 shows how seawater can be desalinated under pressure by reverse osmosis. The largest desalination plant in the world is located in Jubail, Saudi Arabia, where it produces 50% of the country's drinking water by the reverse osmosis of seawater from the Persian Gulf. Smaller desalination plants are in operation in Israel, California, and Florida.

Earth's Freshwater

How much water do you use in a day? The answer may surprise you. An average person needs to drink about 1.5 L per day for survival. But this is only a small fraction of the total amount of water used for daily activities. In the United States, about 7 L of water per person per day is used for cooking and drinking; about 120 L for bathing, laundering, and housecleaning; 80 L for flushing toilets; and 80 L for watering yards. Massive quantities of water are also used in agriculture and industry to produce food and other products. **Figure 26-12** shows the major consumers of water in the United States by daily percentage.

Recall that only 0.6% of the hydrosphere is made up of liquid freshwater. Almost all of this freshwater is found as groundwater and other underground sources. Very small amounts (0.01%) come also from surface water such as

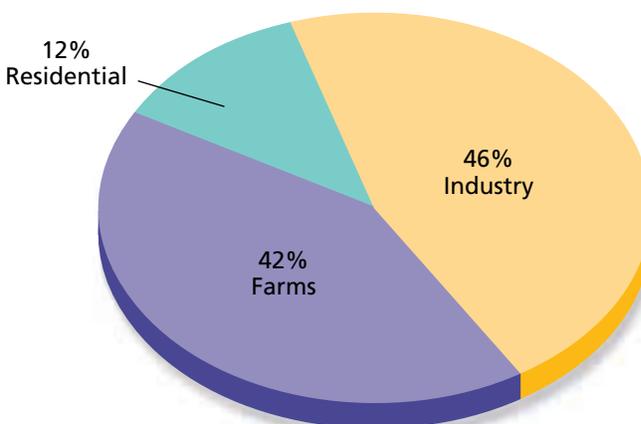


Figure 26-12

Industry and farming use most of the freshwater in the United States. Residential use accounts for about 12%.



Figure 26-13

Algae often experience a population explosion when fertilizer is carried into lakes and ponds. When the algae die and decay, oxygen in the pond is depleted, and fish may die as a result.

lakes and rivers, and the atmosphere provides a tiny amount (0.001%) as water vapor. Freshwater, essential for life on Earth, is our most precious resource. Unfortunately, human activities can affect the quality of freshwater.

Human Impact on the Hydrosphere

A freshwater stream may look sparkling and clean, but it's probably not safe for drinking. Many rivers and lakes in the United States are polluted. Bacteria and viruses enter water supplies through contamination by sewage and industrial wastes. Wastes from landfills and mines leak into groundwater reservoirs. Pesticides, herbicides, and fertilizers are picked up by rainwater and carried into streams. Streams flow into rivers, and rivers empty into oceans. In addition, coastal cities pump waste directly into the oceans. For this reason, much of the oceans' pollution is found along the coasts of continents.

A major cause of freshwater pollution, however, is caused by legal, everyday activities. Water is polluted when we flush toilets, wash hands, brush teeth, and water lawns. The main culprits are nitrogen and phosphorus—two elements found in household detergents, soaps, and fertilizers. Nitrogen and phosphorus are difficult to remove from sewage and wastewater. They contribute to water pollution by causing certain algae and bacteria to reproduce rapidly in the water. The result is an algal bloom, shown in **Figure 26-13**. When the algae die, they decompose, a process that consumes oxygen. The result is the depletion of oxygen in the water. Without sufficient oxygen, fish and other aquatic organisms cannot survive.

Municipal Water and Sewage Treatment

As you have learned, the water that most people use for their daily activities comes from lakes, rivers, reservoirs, and underground sources. Because it has been exposed to a variety of contaminants, it must be purified at a water treatment plant before it can be used.

The treatment of water in municipal treatment plants usually involves five steps: coarse filtration, sedimentation, sand filtration, aeration, and sterilization. The water is first passed through a screen to remove large solids. It then enters large settling tanks where sand and other medium-sized particles settle out. Often, CaO (powdered lime) and $\text{Al}_2(\text{SO}_4)_3$ (alum) are added to coagulate small particles and bacteria. These clumps of particles settle out of solution during a sedimentation step, as you saw in the **DISCOVERY LAB**. After settling, the water is filtered through a bed of sand. The filtered water is then sprayed into the air in a process called aeration. During aeration, oxygen combines with many of the harmful dissolved organic substances in the water, oxidizing them to harmless compounds.

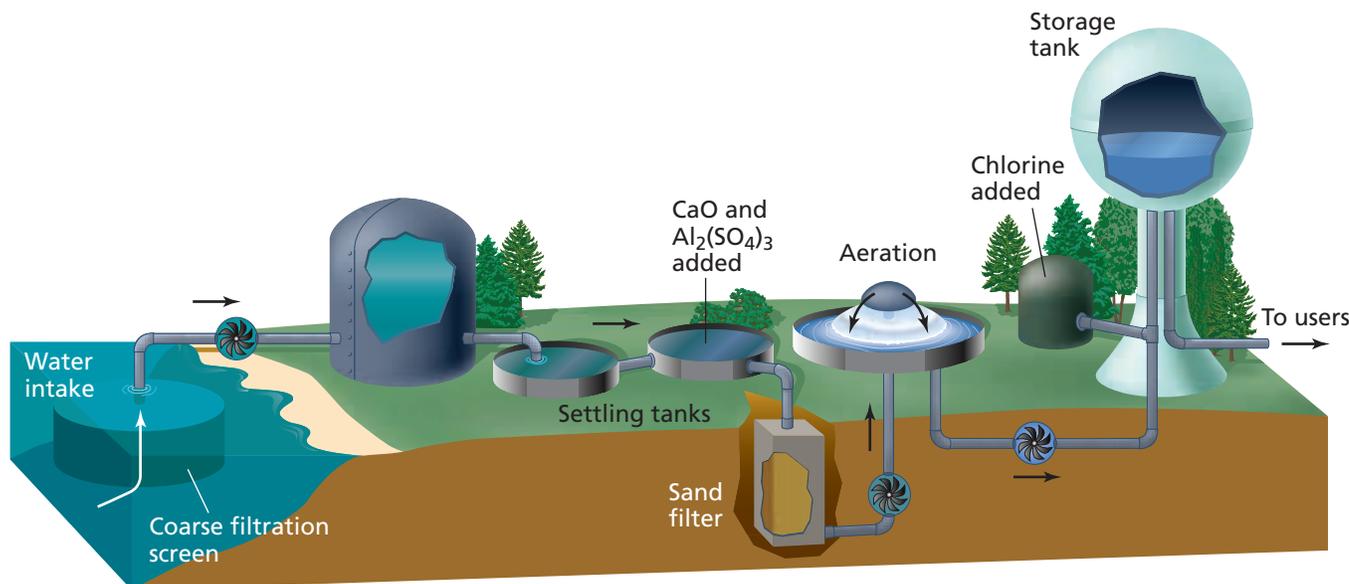


Figure 26-14

Most public water systems use these steps to treat water for human consumption.

The final step in water treatment, sterilization, is accomplished by treating the water with substances that kill bacteria. The most commonly used substance is liquefied chlorine gas. Chlorine reacts with water to form hypochlorous acid (HClO), which destroys any remaining microorganisms in the water.



Figure 26-14 summarizes the flow of water through a water treatment plant.

In addition to water treatment plants, cities and towns also have sewage treatment plants to maintain clean water supplies. The steps involved in sewage treatment are similar to those used to treat water. The incoming sewage is filtered to remove debris and larger suspended solids. Then, as in conventional water treatment, the sewage is passed into large settling tanks where suspended solids settle out. During aeration, the sewage water comes into contact with large amounts of air, and the increased oxygen levels promote the rapid growth of bacteria needed to biodegrade the wastes. After treatment, about 90% of solids and dissolved biodegradable wastes have been removed.

A final treatment removes many inorganic pollutants, such as the toxic heavy metals Cd^{2+} and Pb^{2+} . This treatment is usually expensive, but an increasing percentage of wastewater treatment facilities are taking this extra step to ensure that toxic substances do not enter water supplies. Before release into the environment, the processed water is treated with Cl_2 to kill any remaining bacteria.

Section 26.2 Assessment

- Use the water cycle to explain how the hydrosphere and atmosphere are interrelated.
- Sequence the five steps involved in municipal water treatment.
- Explain why much of the water that comes out of your faucet is “recycled” water.
- Thinking Critically** Why does desalination by reverse osmosis require the use of high pressure?
- Making and Using Graphs** Use the data in **Table 26-2, Ionic Composition of Seawater**, to make a bar graph showing the percentages of the most abundant ions in seawater.

Earth is a large, dynamic planet that has existed, according to best estimates, for approximately 4.6 billion years. When Earth was newly formed, it was a huge molten mass. As this mass cooled, it differentiated into regions of varying composition and density. These regions, or layers, include a dense central core, a thick mantle, and a thin crust. The core is further divided into a small, solid inner core and a larger, liquid outer core. **Figure 26-15** shows Earth's layers.

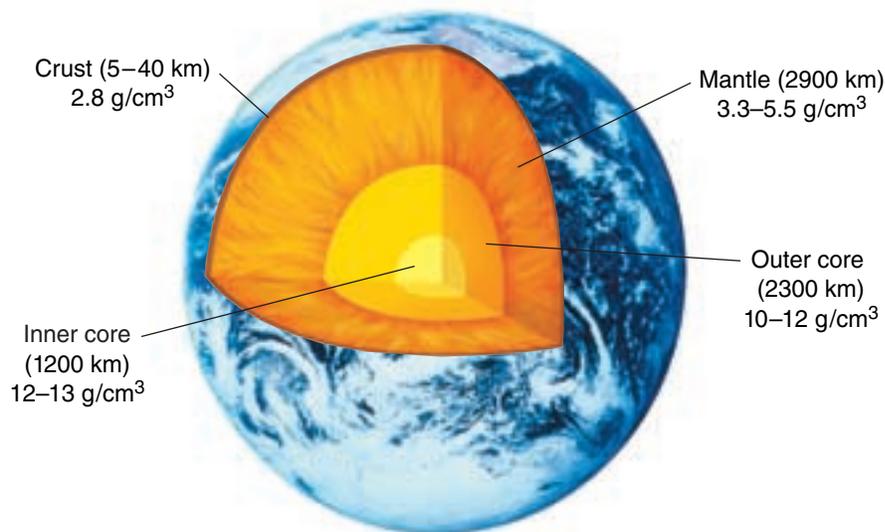
Gravity caused more-dense elements to sink beneath less-dense elements in molten Earth. Hence, the core is composed mostly of iron and small amounts of nickel. In contrast to the dense core, the less-dense outer region at first consisted mostly of rock as oxygen combined with silicon, aluminum, magnesium, and small amounts of iron. This lighter outer region later separated into the mantle and the crust as Earth cooled further. Many light elements such as hydrogen escaped Earth's gravitational force and are now found mostly in space.

Earth's crust makes up about one percent of Earth's mass. Oceanic crust lies beneath Earth's oceans. Continental crust is the part of the crust beneath landmasses.

The Lithosphere

You have already studied the liquid outer part of Earth—the hydrosphere—and the gaseous atmosphere. Each of these parts has a distinct composition and environmental chemistry, as you have learned.

The solid crust and the upper mantle make up the region called the **lithosphere**. Oxygen is the most abundant element in the lithosphere. Unlike the hydrosphere and the atmosphere, the lithosphere contains a large variety of other elements, including deposits of alkali, alkaline earth, and transition metal elements. **Table 26-3** lists the most abundant elements in the continental crust portion of the lithosphere. With the exception of gold, platinum, and a few other rare metals that are found free in nature, most metallic elements occur as compounds in minerals. A mineral is a solid, inorganic compound found in nature. Minerals have distinct crystalline structures and chemical compositions. Most are combinations of metals and nonmetals.



Objectives

- **Identify** Earth's major regions.
- **List** the major elements in Earth's crust.
- **Describe** the composition of minerals.

Vocabulary

lithosphere

Table 26-3

Abundance of Elements in the Continental Crust	
Element	Percent by mass
Oxygen	45.2
Silicon	27.2
Aluminum	8.2
Iron	5.8
Calcium	5.1
Magnesium	2.8
Sodium	2.3
Potassium	1.7
Other elements (Ti, H, Mn, Cu, Pb, Zn, Sn, etc.)	1.7

Figure 26-15

Earth's layers include a small inner core and large outer core of high density, a thick mantle of medium density, and a very thin crust of low density.



a Mercury is extracted from cinnabar and used in instruments that measure blood pressure and temperature.



b Bauxite is the major ore of aluminum, a metal with many uses.



c Galena is a major ore of lead. The material that holds a stained glass window together contains lead.



d Iron can be extracted from many ores, including hematite. Because iron is so strong, it is used as the framework for large buildings.

The metals in a mineral cannot always be extracted from the mineral in an economically feasible way. Sometimes, the concentration of the mineral in the surrounding rock is too low for the mineral to be mined at a reasonable cost. The cost of energy needed to mine, extract, or purify the metal also may be too high. If the metal can be extracted and purified from a mineral at a reasonable profit, the mineral is called an ore. Several ores and the metals extracted from them are shown in **Figure 26-17**.

Figure 26-17

Metals are extracted from their ores and used for many purposes.

Section 26.3 Assessment

- 11.** Name and describe Earth's major regions.
- 12.** Why does the composition of Earth's crust differ so greatly from that of its core?
- 13.** What is a mineral? List some common minerals that exist as metallic oxides, as carbonates, and as sulfides.
- 14. Thinking Critically** Nitrogen makes up 0.002% of Earth's lithosphere but 78% of the atmosphere. How can you explain this difference?
- 15. Applying Concepts** What is an ore? Under what conditions would a mineral not be an ore?

Objectives

- **Trace** the pathways of carbon and nitrogen through the environment.
- **Compare** and **contrast** the greenhouse effect and global warming.

Vocabulary

greenhouse effect
global warming
nitrogen fixation

Did you know that the atoms of carbon, nitrogen, and other elements in your body are far older than you? In fact, they've been around since before life began on Earth. The amount of matter on Earth never changes. As a result, it must be recycled constantly. You learned about the water cycle earlier in the chapter. A number of elements cycle through the environment in similar, distinct pathways.

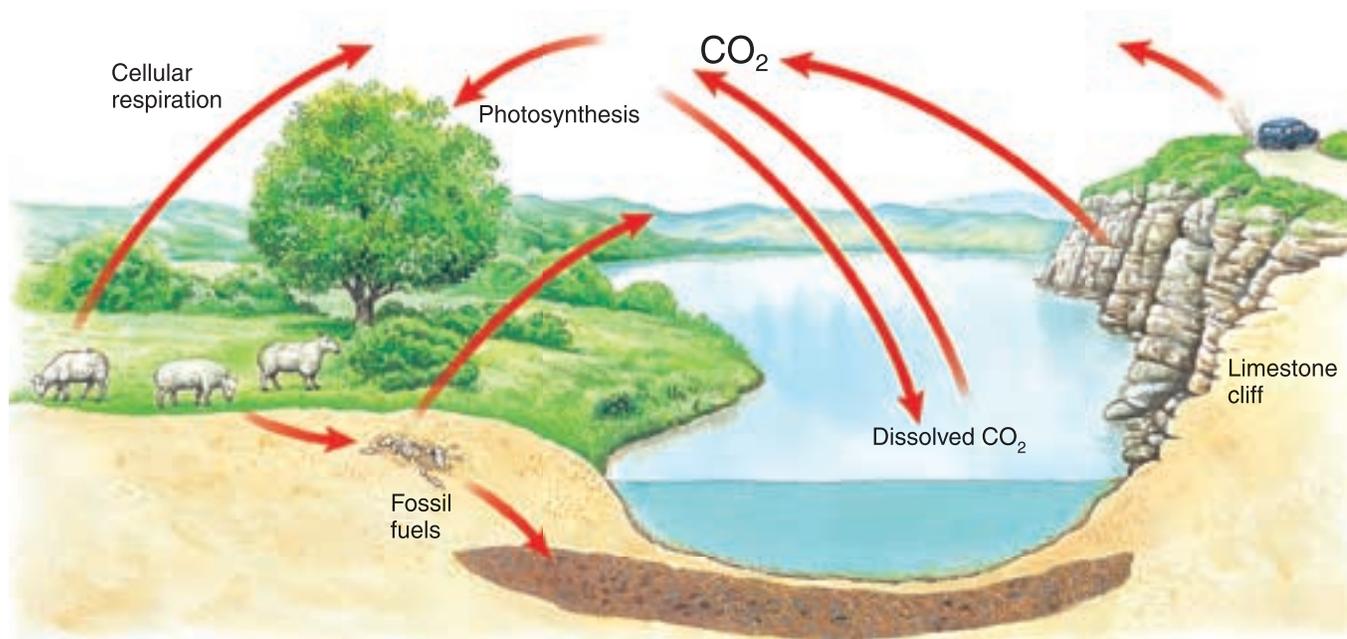
The Carbon Cycle

Carbon dioxide (CO_2) constitutes only about 0.03% of Earth's atmosphere. However, it plays a vital role in maintaining life on Earth. There is a fine balance in nature between the processes that produce carbon dioxide and those that consume it. You have learned that green plants, algae, and some bacteria remove carbon dioxide from the atmosphere during photosynthesis. Do you recall what products are formed during this process? Photosynthesis produces carbon-containing carbohydrates, which animals ingest when they eat plants and other animals. Both plants and animals convert the carbohydrates to CO_2 , which is released into the atmosphere as a waste product of cellular respiration. Once in the atmosphere, the CO_2 can be used again by plants.

Figure 26-18 shows the carbon cycle. Carbon dioxide in the atmosphere is in equilibrium with an enormous quantity that is dissolved in oceans, lakes, and streams. Some of this dissolved CO_2 was once in the form of calcium carbonate (CaCO_3), the main component of the shells of ancient marine animals. The shells were eventually converted into limestone, which represents a large store of carbon on Earth. When the limestone was exposed to the atmosphere by receding seas, it weathered under the action of rain and surface water, producing carbon dioxide. Some of this CO_2 was released into the atmosphere. This process continues today.

Figure 26-18

Carbon cycles in and out of the environment through many pathways.



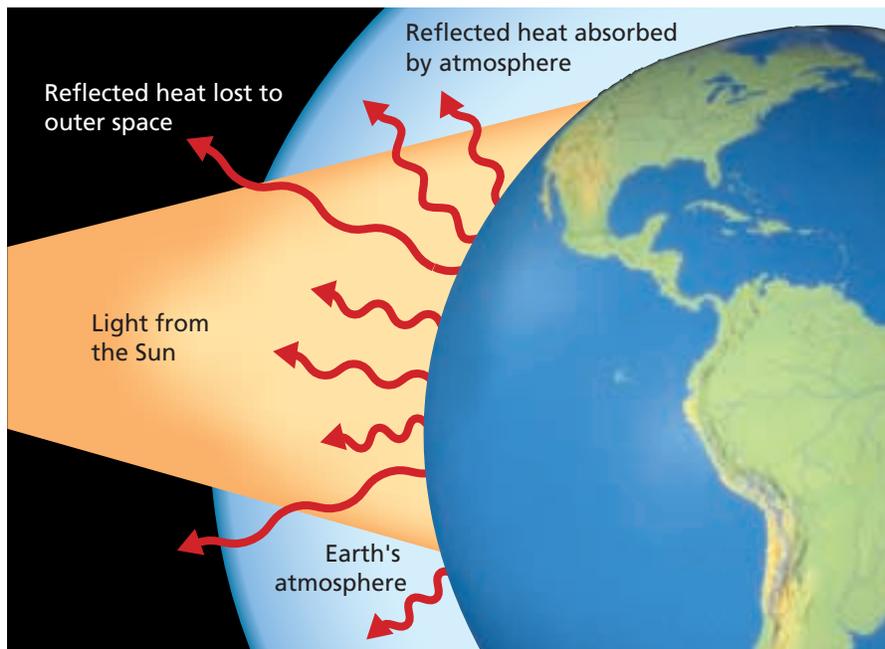


Figure 26-19

About 25% of the sunlight that strikes Earth's atmosphere is reflected back into space. Most of the remaining 75% is absorbed by atmospheric gases and Earth in the form of heat.

Carbon dioxide also enters the atmosphere when plants and animals decompose. Recall from Chapter 22 that the remains of ancient plants and animals were converted under pressure to fossil fuels. When fossil fuels are burned, the carbon is converted to CO_2 . As you'll learn, the burning of fossil fuels and other human activities may be disrupting the balance of the carbon cycle.

Upsetting the balance To understand the effect of human activities on the carbon cycle, it is first necessary to explore a phenomenon known as the greenhouse effect. The **greenhouse effect** is the natural warming of Earth's surface that occurs when certain gases in the atmosphere absorb some of the solar energy that is converted to heat and reflected from Earth's surface. As **Figure 26-19** shows, the process works much like a greenhouse. Sunlight reaches Earth and is converted to heat, but the heat can't easily escape through the "greenhouse gases" to travel back into space. Instead, the heat is absorbed by molecules of greenhouse gases and transferred to the atmosphere, where it warms Earth's surface. Without the greenhouse effect, the surface of our planet would be too cold to sustain life as we know it.

Carbon dioxide is a major greenhouse gas. Most CO_2 occurs naturally. But, when we burn fossil fuels, huge quantities of CO_2 —more than 5 billion metric tons a year—are added to the atmosphere. Moreover, the amount of CO_2 that is removed from the atmosphere by photosynthesis is being reduced by the continued destruction of vast forested areas, particularly rain forests. As a result of these activities, the level of atmospheric CO_2 has been increasing slowly over the past 300 years. **Diagram a** in the **problem-solving LAB** on the next page shows that the rate of increase is accelerating.

Increases in greenhouse gases such as CO_2 lead to corresponding increases in the greenhouse effect. Some scientists have predicted that these increases will cause global temperatures to rise, a condition known as **global warming**.

Scientists don't agree on the causes or the consequences of global warming, but they do know that average global temperatures are increasing slightly—about 0.5°C over the past 100 years. At present levels of fossil-fuel consumption, some scientists predict that global temperatures could increase by as much as 0.3°C each decade during the twenty-first century. Although this may seem like a small amount, it represents the largest increase in global temperatures since the end of the last ice age.

While scientists continue to study and debate the issue of global warming, most concede that it has the potential to change Earth's climate and that tampering with the climate could be dangerous. Thus, a drastic reduction in the use of fossil fuels is considered by many to be essential to slow and eventually stop global warming. For this reason, alternative energy resources such as solar power are now important areas of scientific research. You can learn about one type of alternative energy source—solar ponds—by doing the **CHEMLAB** at the end of the chapter.

The Nitrogen Cycle

Nitrogen is an essential component of many substances that make up living organisms. It is a key element in protein molecules, nucleic acids, and ATP. Like other elements, the supply of nitrogen on Earth is fixed. It must be recycled, as shown in **Figure 26-20**. Follow the diagram as you read about the nitrogen cycle.

Although nitrogen makes up 78% of Earth's atmosphere, most living things can't use nitrogen in its gaseous state. It must be fixed, or converted to a useful form, by a process called **nitrogen fixation**.

problem-solving LAB

Global warming: fact or fiction?

Drawing Conclusions Environmental issues affecting the entire planet usually involve so many variables that it is often difficult to pinpoint a primary source of the problem. Conflicting data from many experiments often divide the views of those in the scientific community. The effect of carbon dioxide emissions on global warming has been a particularly controversial issue because efforts to resolve the problem may require humans to make dramatic changes in their lifestyles and could exact enormous demands on the world economy.

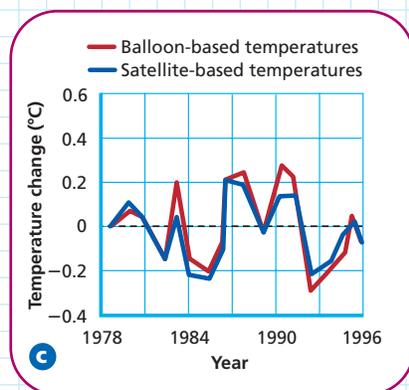
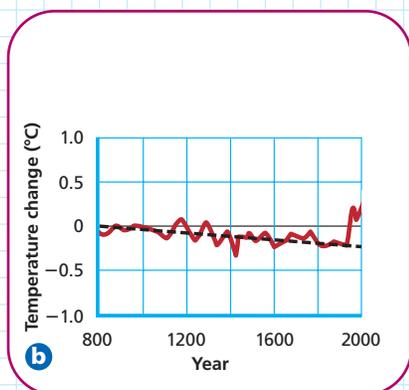
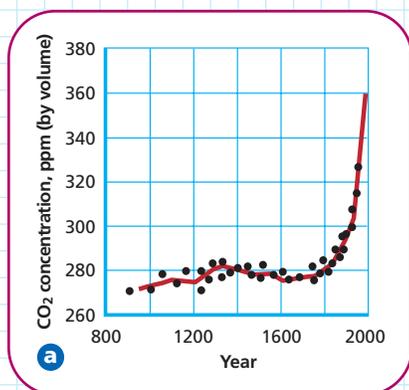
Analysis

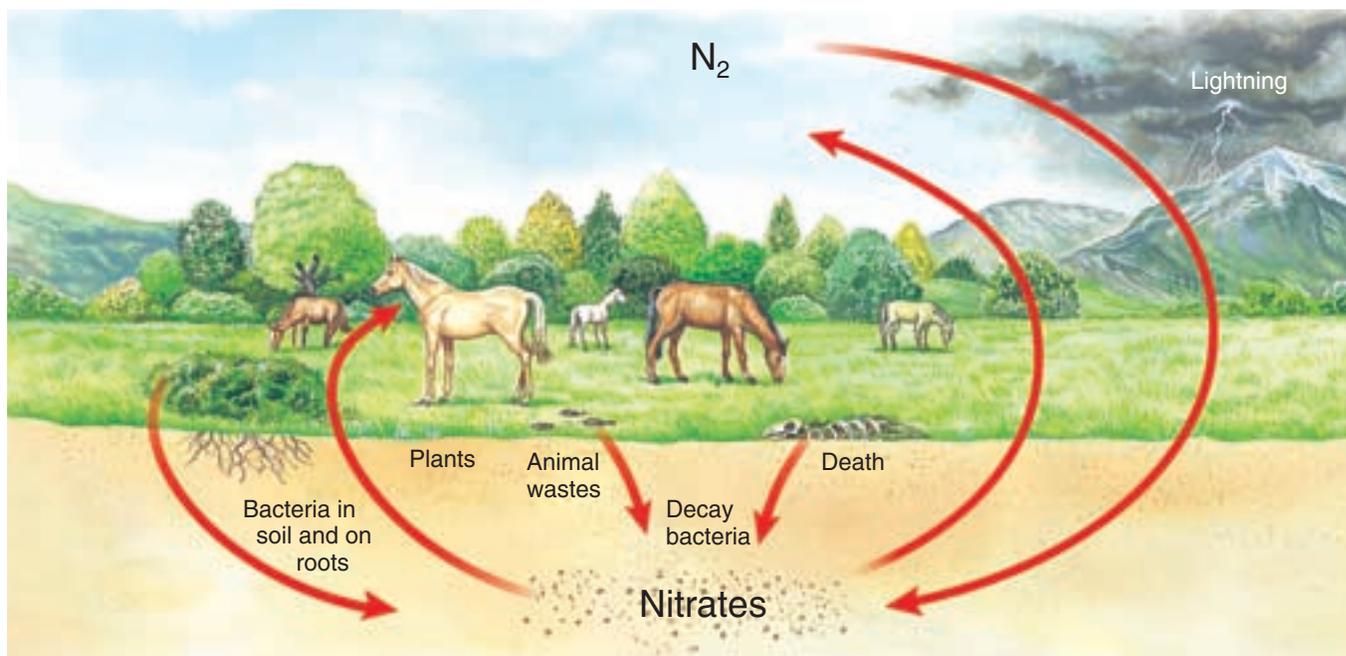
Each graph shows an indicator that is being used to monitor global warming. **Diagram a** shows the atmospheric CO₂ concentration for the past millennium. **Diagram b** shows the average temperature

change of Earth's surface during the past millennium. **Diagram c** shows the average temperature change of Earth's atmosphere since 1978.

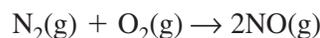
Thinking Critically

1. Before the twenty-first century, what was the trend (the dotted line in **diagram b**) for Earth's average surface temperature? What does this trend predict for the average surface temperature in the year 2000?
2. What do the atmospheric temperature data suggest? How does this compare to the surface data? What could cause this discrepancy?
3. In comparing these three graphs, how strong is the relationship between global warming and the concentration of carbon dioxide in the atmosphere? Is there a reason for concern? What other environmental influences would you consider to help answer this question?

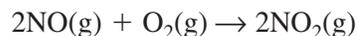




Two primary routes for nitrogen fixation exist in nature. In the atmosphere, lightning combines N_2 and O_2 to form NO .



Once NO is formed, it is oxidized to NO_2 .



In the atmosphere, rain converts NO_2 to HNO_3 , which then falls to Earth as aqueous NO_3^- .



Nitrogen fixation is also accomplished by nitrogen-fixing bacteria, which live in the soil and on the roots of certain legumes such as peas, beans, peanuts, and alfalfa. In this process, N_2 is first reduced to NH_3 and NH_4^+ , then oxidized to NO_3^- .

Plants absorb nitrate ions through their roots and use them to synthesize complex nitrogen compounds. Because animals can't synthesize these complex molecules, they must get them by eating plants or other animals. Nitrogen compounds that are unused by the animals' bodies are excreted as waste. Soil microorganisms convert this waste to N_2 and nitrogen is recycled back into the environment.

Figure 26-20

In the nitrogen cycle, atmospheric nitrogen gas is converted to nitrates, which plants use to make biological compounds. Eventually, nitrates are converted back to nitrogen gas.

Section 26.4 Assessment

- Diagram and label the important parts of the carbon cycle.
- Describe the two primary routes of nitrogen fixation.
- What two cellular processes are important parts of the carbon cycle?
- Thinking Critically** Levels of CO_2 , a major greenhouse gas, fluctuate on a seasonal basis. Recalling how carbon is cycled through the environment, explain the seasonal fluctuations of CO_2 .
- Applying Concepts** How might the greenhouse effect lead to global warming?

Solar Pond

If you made a list of popular types of alternative energy sources, solar energy probably would be near the top. Of course, the energy we use from all sources ultimately originates from the Sun. It would seem that solar energy would be the easiest to use. The problem is how to store solar energy when the Sun is not shining. In this experiment, you will investigate one method that could be used to trap and store solar energy.

Problem

Build a small-scale model of a solar pond and test how it traps and stores solar energy.

Objectives

- **Construct** a small-scale solar pond using simple materials.
- **Collect** temperature data as the solar pond model heats and cools.
- **Hypothesize** as to why a solar pond is able to trap and store energy.

Materials

CBL System	waterproof tape
graphing calculator	table salt
ChemBio program	hot plate
link-to-link cable	stirring rod
temperature probes (2)	250-mL beaker
150-watt light bulb	beaker tongs
socket and clamp for bulb	TI-Graph Link (optional)
black plastic frozen-dinner dish	ring stand and clamp
	250-mL graduated cylinder

Safety Precautions



- The light bulb will become hot when it is turned on.
- Do not touch the hot plate while it is on.

Pre-Lab

1. Read the entire CHEMLAB.
2. Prepare all written materials that you will take into the laboratory. Include safety precautions and procedure notes.
3. Water is transparent to visible light but opaque to infrared radiation. How do you think these properties will affect your solar pond model?
4. If you used only tap water in your model, convection currents would bring warmer, less dense water from the bottom to the surface. Do you think this will happen with your solar pond model? Explain your answer.
5. Predict which of the two layers of the model will have the higher final temperature. Explain your prediction.

Procedure

1. Prepare a saturated table salt (NaCl) solution by heating 100 mL of tap water in a beaker on a hot plate. When the water is boiling, slowly add enough table salt to saturate the solution while stirring with a stirring rod. Remove the beaker from the hot plate with beaker tongs and allow the solution to cool slowly overnight.
2. The next day, prepare the solar pond model. Place the black plastic dish on the lab bench where you want to run the experiment. Use a small piece of waterproof tape to attach one of the temperature probes to the bottom of the black plastic dish. Plug this probe into Channel 1 of the CBL System. Slowly pour the 100 mL of saturated salt solution into the dish.
3. Carefully add about 100 mL of tap water on top of the saturated salt-water layer in the dish. Use care not to mix the two layers. Suspend the end of the second temperature probe in the tap-water layer and plug it into Channel 2 of the CBL System.

4. Connect the graphing calculator to the CBL System using the link cable. Turn on both units. Run the ChemBio program. Choose 1:SET UP PROBES from MAIN MENU. Choose 2 probes. Under SELECT PROBE, choose 1:TEMPERATURE. Enter 1 for Channel. This is for the probe at the bottom of the salt water. Under SELECT PROBE, choose 1:TEMPERATURE. Enter 2 for Channel. This is for the probe in the tap-water layer.
5. Under MAIN MENU, choose 2: COLLECT DATA. Choose 2: TIME GRAPH. For time between samples in seconds, choose 30. For number of samples, choose 60. This will allow the experiment to run for 30 minutes. Set the calculator to use this time setup. Input the following: Ymin = 0, Ymax = 30, Yscl = 1. Do not start collecting data yet.
6. Position the 150-watt light bulb about 15 to 20 cm over the top of the solar pond model. Turn on the light. Press ENTER on the calculator to begin collecting data. After about 6 to 8 minutes, turn off the light bulb and move it away from the solar pond model. Do not disturb the experiment until the calculator is finished with its 30 minute run.

Cleanup and Disposal

Rinse the salt solution off the temperature probes.

Analyze and Conclude

1.  **Graphing Data** Make a copy of the graph from the graphing calculator. If you have TI-Graph Link and a computer, do a screen print.
2. **Interpreting Graphs** Describe the shape of each curve in the graph of time versus temperature before and after the light bulb was turned off. Explain the significance of the difference.
3. **Comparing and Contrasting** Which layer of your solar pond model did the best job of trapping and storing heat?
4. **Applying Concepts** Why does the graph of time versus temperature decrease more rapidly near the surface when the light bulb is turned off?
5. **Forming a Hypothesis** Make a hypothesis to explain what is happening in your model.



6. **Designing an Experiment** How would you test your hypothesis?
7. **Error Analysis** How might your results have been different if you had used a white dish by mistake instead of a black dish? Explain.

Real-World Chemistry

1. Water in a lake rises to the surface when heated and sinks to the bottom when cooled in a process called convection. Compare and contrast the density of the water as it rises with the density of the water as it sinks.
2. The El Paso Solar Pond was the first in the world to successfully use solar pond technology to store and supply heat for industrial processes. It was built with three main layers: a top layer that contains little salt, a middle layer with a salt content that increases with depth, and a very salty bottom layer that stores the heat. Which layer has the greatest density? The least density? Why doesn't the storage layer in the El Paso Solar Pond cool by convection?

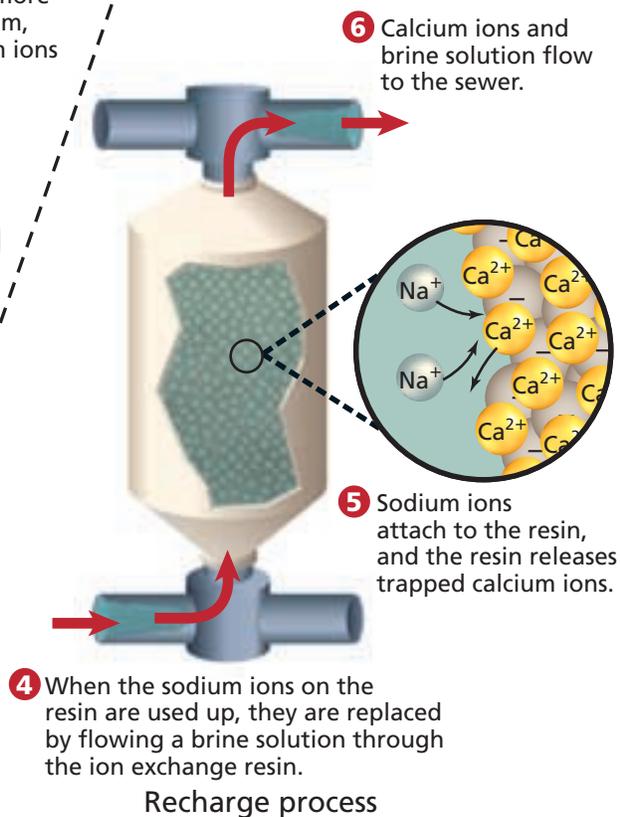
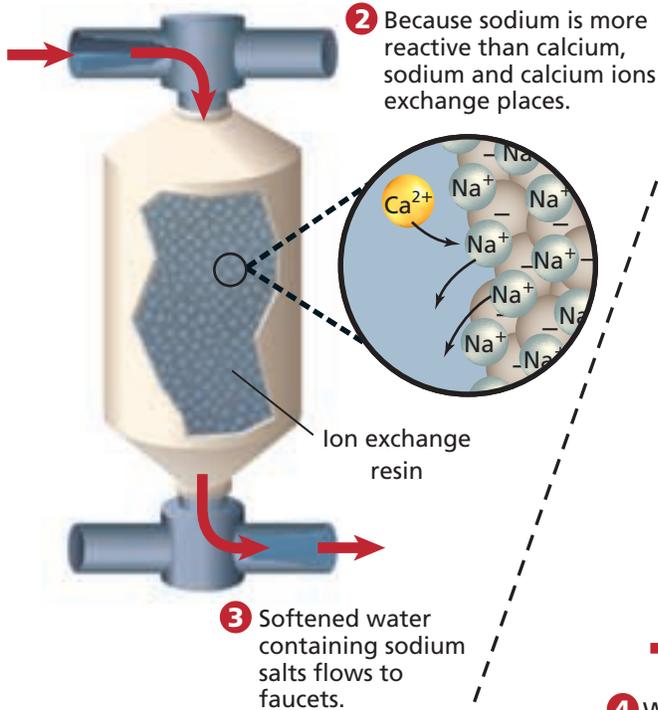
How It Works

Water Softener

Water that contains substantial amounts of dissolved calcium and magnesium salts is called hard water. Water acquires these salts when it passes through soil and rocks that contain calcium and magnesium carbonates. While these salts are not toxic, they react with soaps and detergents to form an insoluble scum on sinks, bathtubs, and showers. When clothes are washed in hard water, some of the insoluble material adheres to the clothes and alters the way they feel. A device called a water softener can solve the hard-water problem by removing the excess calcium and magnesium ions.



- 1 Hard water flows over an ion exchange column containing pellets coated with sodium ions.



Thinking Critically

1. **Predicting** Would you expect a water softener to be effective in removing materials such as sugar, alcohol, grease, and oil? Explain your answer.
2. **Using the Internet** Use the Internet to find information about water softening processes that use ions other than sodium ions. Report your findings to the class.

Summary

26.1 Earth's Atmosphere

- Earth's atmosphere is divided into the troposphere, stratosphere, mesosphere, thermosphere, and exosphere. The layers vary according to altitude and temperature.
- Nitrogen and oxygen make up the majority of the atmospheric gases.
- Photodissociation and photoionization are important processes that absorb high-energy ultraviolet radiation in the upper atmosphere.
- The ozone layer in the stratosphere forms a protective barrier against harmful high-energy ultraviolet radiation. Levels of ozone have been depleted by chemical reactions with CFCs in the upper atmosphere above the North and South Poles.
- Photochemical smog is a major pollutant in many urban centers and is formed from chemicals that are produced principally by internal combustion engines.
- Emissions of SO₂ from the burning of fossil fuels have led to the production of acid rain.

26.2 Earth's Water

- Water is continuously recycled in the environment by the processes of evaporation, condensation, and precipitation.
- Earth has massive oceans that contain large quantities of dissolved salts. Salinity is a measure of the mass of these dissolved salts.
- Salinity is expressed as grams of salt per kilogram of seawater.
- Reverse osmosis can be used to desalinate ocean water and make it fit for human use.

- Freshwater is a precious natural resource. Municipal water supplies must be treated to ensure that they are safe to use. Sewage treatment plants process water that has been used so that it can be returned safely to the environment.

26.3 Earth's Crust

- Earth is divided into a core, mantle, and crust. The core is further divided into an inner and outer core. The crust is where living things reside. The crust can be divided further into the solid lithosphere, liquid hydrosphere, and gaseous atmosphere.
- The lithosphere contains a large number of elements. Most of these occur as minerals.
- Minerals are solid, inorganic compounds found in nature. They have distinct crystalline structures and chemical compositions. Minerals of various elements are found in the lithosphere mainly as oxides, carbonates, and sulfides.
- An ore is a substance, commonly a mineral, from which the metal it contains can be extracted and purified at a reasonable profit.

26.4 Cycles in the Environment

- Carbon dioxide, one of the principal components in the carbon cycle, is taken up by plants during photosynthesis and given off by plants and animals as a product of cellular respiration.
- The CO₂ in the atmosphere is a major cause of the greenhouse effect. Increases in greenhouse gases may cause global warming. Scientists do not agree on the causes or consequences of global warming.
- Nitrogen gas is converted into biologically useful nitrates by nitrogen fixation. Nitrogen is returned to the atmosphere by a cycle in the environment.

Vocabulary

- atmosphere (p. 841)
- desalination (p. 851)
- global warming (p. 859)
- greenhouse effect (p. 859)
- hydrosphere (p. 850)
- lithosphere (p. 855)
- nitrogen fixation (p. 860)
- salinity (p. 851)
- stratosphere (p. 842)
- troposphere (p. 842)

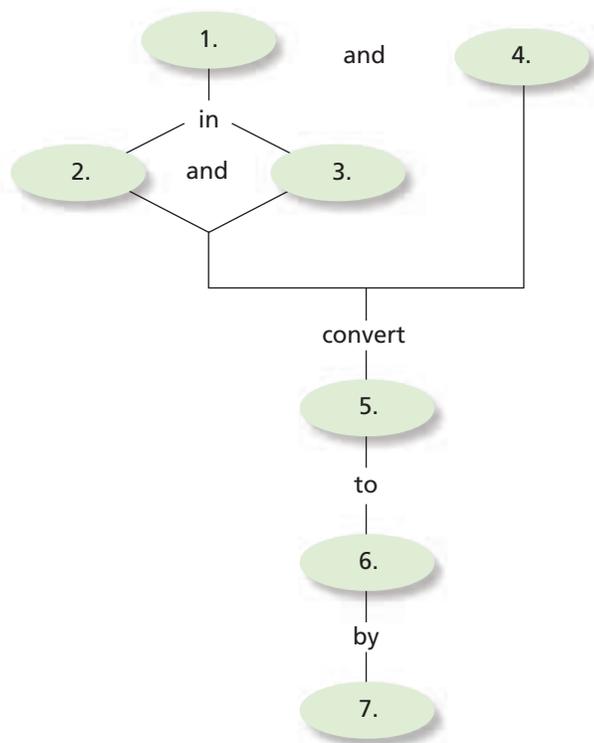


CLICK HERE

Go to the Chemistry Web site at science.glencoe.com or use the Chemistry CD-ROM for additional Chapter 26 Assessment.

Concept Mapping

21. Complete the concept map using the following terms: roots of legumes, nitrogen oxides, soil, bacteria, nitrogen fixation, nitrogen gas, lightning.



Mastering Concepts

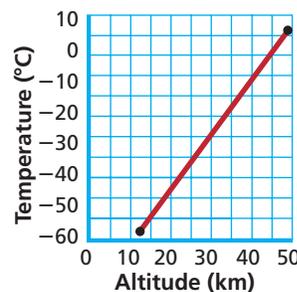
22. What is ozone? (26.1)
23. Why is smog more correctly termed photochemical smog? (26.1)
24. What kind of radiation causes molecules to be ionized in the upper atmosphere? (26.1)
25. What is the basis for the division of the atmosphere into layers? (26.1)
26. How do chlorine atoms act as a catalyst for ozone decomposition? (26.1)
27. What steps can be taken to reduce sulfur dioxide emissions into the environment? (26.1)
28. What provides the energy for the water cycle? (26.2)

29. What are the two most abundant ions in seawater? (26.2)
30. Why does the chemical composition of the oceans remain constant over time? (26.2)
31. What is the function of a semipermeable membrane in reverse-osmosis desalination? (26.2)
32. What are the four most abundant elements in the lithosphere? (26.3)
33. How does the position of a metal in the periodic table determine whether the element exists in nature as an oxide, a carbonate, or a sulfide? (26.3)
34. During what process is CO_2 from the troposphere converted into carbohydrates? (26.4)
35. What is the greenhouse effect? (26.4)
36. In what two ways is nitrogen fixed? (26.4)

Mastering Problems

Calculations About the Atmosphere (26.1)

37. How many liters of nitrogen are there in a 25.3-liter sample of air?
38. Identify the layer of the atmosphere that is represented by the following graph of temperature versus altitude.



39. Assuming a temperature of 20.0°C and an atmospheric pressure of 729 mm Hg, calculate the number of particles of carbon dioxide gas in one cubic meter of dry air.

Calculations About the Hydrosphere (26.2)

40. Water must have a salt concentration less than 500 ppm by mass to be fit for human use. Assuming that the salt is all NaCl, what is the molarity of this concentration?
41. Construct a circle graph to illustrate the amount of water used for various purposes by an average person each day.
42. What is the molarity of NaCl in a saltwater solution with a salinity of 6 g/kg? Assume that the only salt in the solution is NaCl and that the solution has a density of 1.0 g/mL.

Calculations About the Lithosphere (26.3)

- 43.** Table 26-5 shows the mass percent of iron in the universe, in Earth as a whole, and in various layers of Earth. Explain the differences in terms of the process that caused formation of Earth.

Table 26-5

Abundance of Iron in Various Locations	
Location	Percent by mass
Whole universe	0.19
Whole Earth	34.6
Atmosphere, Lithosphere, Hydrosphere	4.7
Mantle	13.3
Core	88.6

- 44.** Earth's crust is 0.87% hydrogen by mass. The lithosphere is only 0.15% hydrogen. Account for this difference.
- 45.** What is the mass percent of iron in each of the following iron minerals: FeS_2 , Fe_2O_3 , Fe_3O_4 ? If the iron can be extracted with equal ease from all three minerals, which would you use as a source of iron?

Calculations About Natural Cycles (26.4)

- 46.** About 30 million tons of HNO_3 are produced each year when lightning fixes nitrogen in the atmosphere. What is the mass percent of nitrogen in HNO_3 ? How much nitrogen does 30 million tons of HNO_3 contain?
- 47.** If a plant consumes 50 g carbon dioxide in the process of photosynthesis during a given time period, how many liters of oxygen would it produce at 1 atm and 25°C ? The balanced equation for photosynthesis is $6\text{CO}_2 + 6\text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2$

Mixed Review

Sharpen your problem-solving skills by answering the following.

- 48.** Why must living organisms rely on nitrogen fixation for their source of nitrogen?
- 49.** What is the major cause of the greenhouse effect?
- 50.** What is the purpose of adding CaO and $\text{Al}_2(\text{SO}_4)_3$ in municipal water treatment?
- 51.** Write balanced equations for the destruction of ozone by CFCs.
- 52.** Why does nitrogen cause freshwater pollution?
- 53.** In what form do plants use nitrogen? In what biological molecules is nitrogen found?
- 54.** Write balanced equations showing the formation of acid rain.
- 55.** Using balanced equations, compare photoionization and photodissociation of an oxygen molecule in the upper atmosphere.
- 56.** What two elements would you expect to be most abundant in Earth's hydrosphere? Why?

Thinking Critically

- 57. Recognizing Cause and Effect** Environmental chemistry has a number of cause-and-effect relationships. Describe these relationships in **a)** the formation of acid rain, **b)** ozone depletion.
- 58. Interpreting Data** Why are photodissociation and photoionization reactions more common in the upper atmosphere than in the lower atmosphere?
- 59. Applying Concepts** Why do oxygen atoms exist for a longer period of time in the upper atmosphere than in the stratosphere?

Writing in Chemistry

- 60.** Write a short story tracing the path that a carbon atom in a carbon dioxide molecule might follow. Assume that the CO_2 molecule is in the troposphere.
- 61.** Prepare an instruction sheet for a portable, reverse-osmosis desalination apparatus that can be used for camping. Research and include information on how efficient the apparatus is, how much pressure must be used, and how much drinking water can be produced in a given amount of time.
- 62.** Metallurgy is the study of extracting and purifying metals from their ores. Conduct research to learn how iron is extracted from its ores, how it is purified, and how steel is made. Make a poster showing the steps in these processes, and include a short summary of each step and the chemical equations involved.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 63.** Why is it necessary to perform repeated experiments in order to support a hypothesis? (Chapter 1)
- 64.** A 49.01-g sample of lead displaces 4.5 mL of water. What is the density of lead? (Chapter 2)
- 65.** Are the following physical or chemical changes? (Chapter 3)
- water boils
 - charcoal burns
 - sugar dissolves in tea
 - potassium reacts with water
 - an ice cube melts
- 66.** The isotope of carbon that is used to date artifacts contains six protons and eight neutrons. What is the atomic number of this isotope? How many electrons does it have? What is its mass number? (Chapter 4)
- 67.** Table 26-6 shows abundance of the two isotopes of silver found in nature. The more abundant isotope has an atomic mass of a little less than 107, but the average atomic mass of silver on the periodic table is about 107.9. Explain why it is higher. (Chapter 5)

Table 26-6

Abundance of Silver	
Isotope	Abundance
Ag-107	51.8%
Ag-109	48.2%

- 68.** For each of the following elements, tell how many electrons are in each energy level and write the electron dot structure. (Chapter 6)
- Ar-18
 - Mg-12
 - N-7
 - Al-13
 - F-9
 - S-16
- 69.** Use the periodic table to separate these 12 elements into six pairs of elements having similar properties. (Chapter 7)
- Ca, K, Ga, P, Si, Rb, B, Sr, Sn, Cl, Bi, Br
- 70.** Write the formula for the binary ionic compound that forms from each pair of elements. (Chapter 8)
- manganese(III) and iodine
 - calcium and oxygen
 - aluminum and fluorine
 - potassium and sulfur

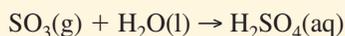
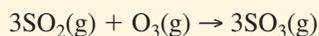
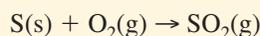
- zinc and bromine
- lead(IV) and oxygen

- 71.** Name the following molecular compounds. (Chapter 9)
- NO
 - IBr
 - N₂O₄
 - CO
 - SiO₂
 - ClF₃
- 72.** Classify each of the following reactions. (Chapter 10)
- N₂O₄(g) → 2NO₂(g)
 - 2Fe(s) + O₂(g) → 2FeO(s)
 - 2Al(s) + 3Cl₂(g) → 2AlCl₃(s)
 - BaCl₂(aq) + Na₂SO₄(aq) → BaSO₄(s) + 2NaCl(aq)
 - Mg(s) + CuSO₄(aq) → Cu(s) + MgSO₄(aq)
- 73.** Calculate the total number of ions in 10.8 g of magnesium bromide. (Chapter 11)
- 74.** Calculate the mass of AgCl formed when 85.6 g of silver sulfide (Ag₂S) react with excess hydrochloric acid (HCl). (Chapter 12)
- 75.** Why will water in a flask begin to boil at room temperature as air is pumped out of the flask? (Chapter 13)
- 76.** How many moles are contained in a 2.44-L sample of gas at 25.0°C and 202 kPa? (Chapter 14)
- 77.** How would you prepare 5.0 L of a 1.5M solution of glucose (C₆H₁₂O₆)? (Chapter 15)
- 78.** Compare the energy of the reactants and the products for a reaction in which ΔH is negative. Is the reaction endothermic or exothermic? (Chapter 16)
- 79.** Identify the following hydrocarbons as alkane, alkene, or alkyne. (Chapter 22)
- 1-hexyne
 - 3-methyldecane
 - propene
 - cis-2-butene
- 80.** Water is released in a reaction in which two different functional groups condense to form an ester. What two functional groups take part in this reaction? (Chapter 23)
- 81.** Name the two functional groups in an amino acid that become linked when a peptide bond is formed. (Chapter 24)
- 82.** What element is formed if magnesium-24 is bombarded with a neutron and then ejects a proton? Write the balanced nuclear equation. (Chapter 25)

Use these questions and the test-taking tip to prepare for your standardized test.

- Photodissociation, photoionization, and photochemical smog are all caused by the reaction of atmospheric gases with high energy _____.
 - infrared radiation
 - ultraviolet radiation
 - cosmic rays
 - microwaves
- The most abundant gas in Earth's atmosphere is _____.
 - O₂
 - N₂
 - CO₂
 - H₂

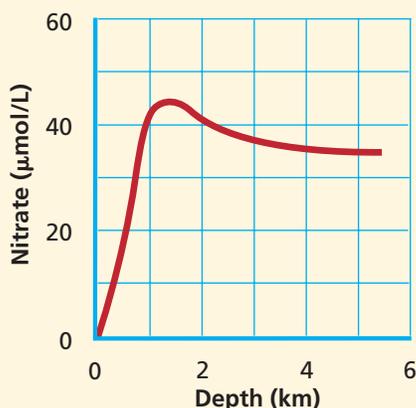
- When high-sulfur coal is burned, the following series of reactions occurs to produce acid rain:



A coal is classified as high-sulfur by the United States Department of Energy if it contains more than 763 g of sulfur per unit. A unit is defined as the amount of coal needed to produce 1.0×10^6 British thermal units (BTUs) of heat. If a unit of coal containing 768 g of sulfur is burned, how much sulfuric acid (H₂SO₄) will be produced?

- 2.35×10^3 g
- 4.08×10^3 g
- 7.06×10^3 g
- 7.53×10^4 g

Interpreting Graphs Use the graph to answer questions 4–6.



- The nitrogen cycle occurs in the oceans as well as in the soil. Nitrates (NO₃⁻) at the surface of the ocean are almost completely used up by living organisms. As these organisms die, they sink. As they sink, the nitrogen compounds inside their bodies are oxidized back to NO₃⁻. In very deep waters, ocean currents mix the

water and decrease NO₃⁻ concentration. These processes explain why the graph shows NO₃⁻ concentration _____.

- continuously decreasing with depth
 - continuously increasing with depth
 - decreasing and then increasing with depth
 - increasing and then decreasing with depth
- Nitrate concentration reaches a maximum at approximately what depth?
 - 5500 m
 - 500 m
 - 3000 m
 - 1500 m
 - At a depth of 4000 m, about how many grams of NO₃⁻ are found in each liter of seawater?
 - 1.6×10^{-3} g
 - 2.2×10^{-3} g
 - 3.5×10^{-7} g
 - 2.0×10^3 g
 - In the water cycle, evaporation must be balanced by the processes of _____.
 - condensation and sublimation
 - ionization and precipitation
 - precipitation and condensation
 - dissolution and precipitation
 - Because the denser elements sank to the center of Earth as the new planet cooled, all of the following elements are relatively rare in Earth's crust EXCEPT _____.
 - aluminum
 - gold
 - platinum
 - lead
 - The proportions and amounts of dissolved salts in the oceans remain almost constant because _____.
 - the sources from which the ocean receives its salts have constant compositions
 - the rates at which salts are added to and taken from the oceans balance each other
 - none of the dissolved salts are ever removed from the seawater
 - living organisms have no use for any of the salts dissolved in seawater



TEST-TAKING TIP

If It Looks Too Good To Be True Beware of answer choices in multiple-choice questions that seem ready-made and obvious. Remember that only one answer choice for each question is correct. The rest are made up by the test-makers to distract you. This means that they may look very appealing. Check each answer choice carefully before finally selecting it.