Gases

*The density of a gas decreases
as the temperature of the gas increases.*

Hot-Air Balloons

Gases and Pressure

*O*BJECTIVES

- Define *pressure,* give units of pressure, and describe how pressure is measured.

- State the standard conditions of temperature and pressure and convert units of pressure.

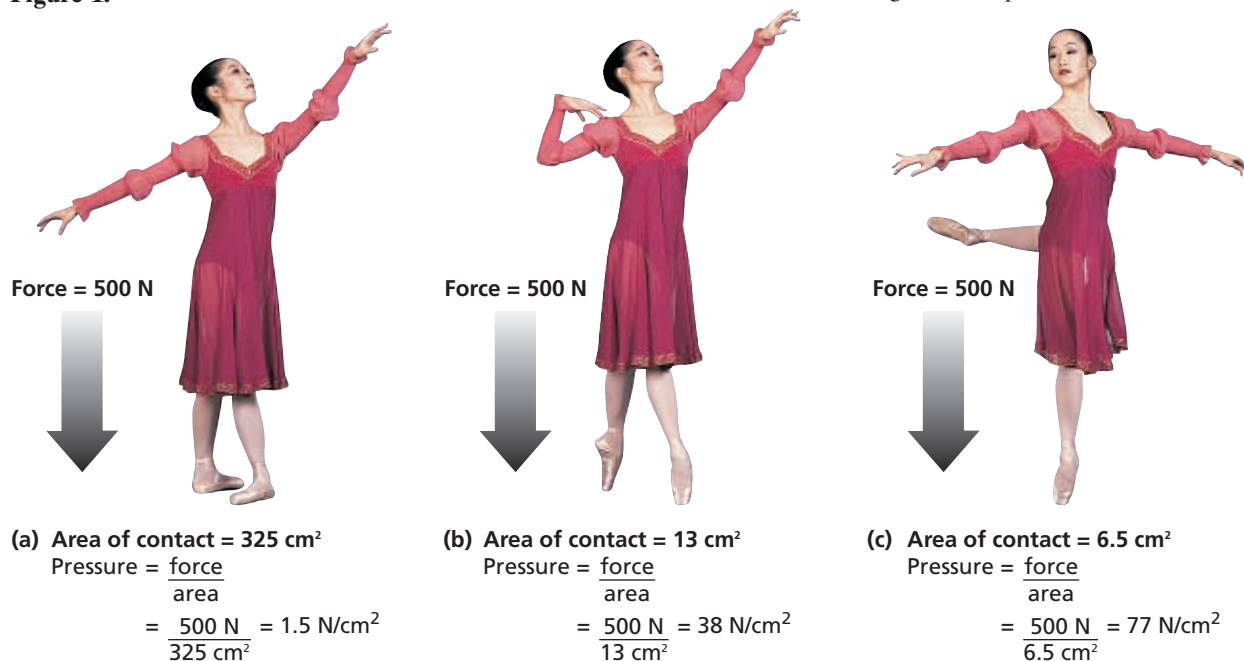- Use Dalton's law of partial pressures to calculate partial pressures and total pressures.

In the chapter "States of Matter," you read about the kinetic-molecular theory, which is based on the idea that particles of matter are always in motion. In this section, you will study the implications of the kinetic-molecular theory of gases.

You have learned that the temperature of a gas is related to the kinetic energy of the gas molecules. In this chapter, you will learn about other properties of gases, including pressure, volume, and amount of gas present, and the relationship between these properties.

Pressure and Force

If you pump air into an automobile tire, the pressure in the tire will increase. The pressure increase is caused by the increase in the number of collisions of molecules of air with the inside walls of the tire. The collisions cause an outward push, or force, against the inside walls. Gas molecules exert pressure on any surface with which they collide. The pressure exerted by a gas depends on volume, temperature, and the number of molecules present. **Pressure** *(P) is defined as the force per unit area on a surface.* The equation defining pressure is shown in **Figure 1.**

FIGURE 1 The pressure the ballet dancer exerts against the floor depends on the area of contact. The smaller the area of contact, the greater the pressure.
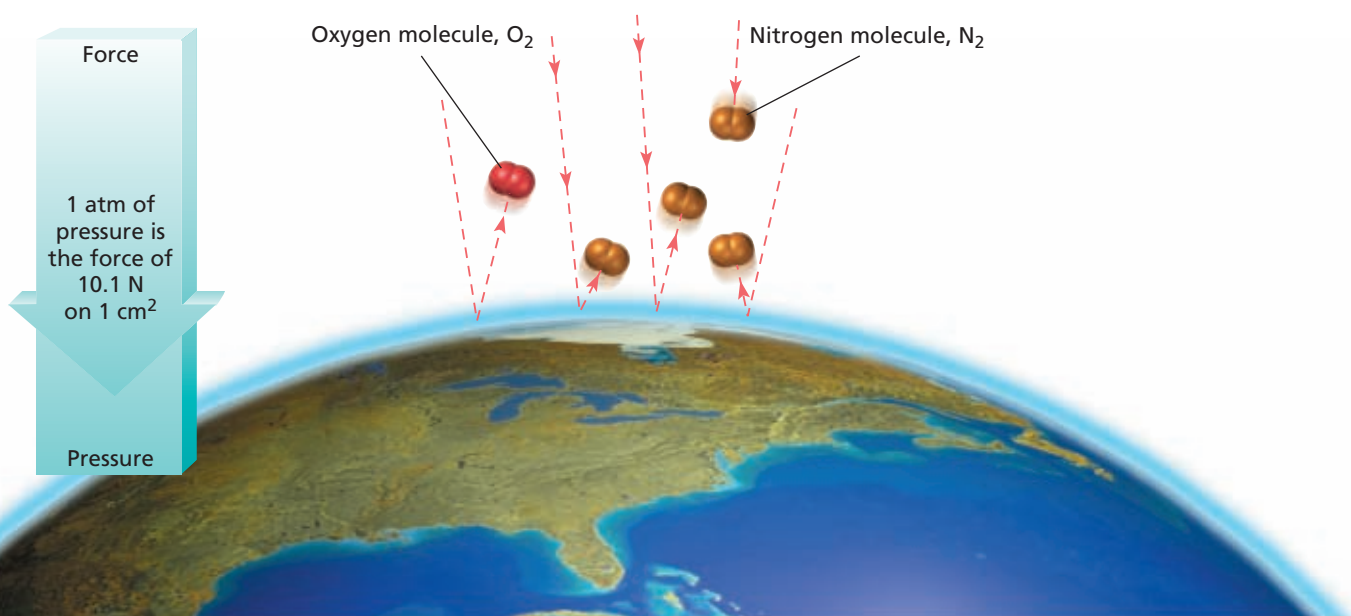

(a) Area of contact = 325 cm²
Pressure = $\dfrac{\text{force}}{\text{area}}$
$= \dfrac{500 \text{ N}}{325 \text{ cm}^2} = 1.5 \text{ N/cm}^2$

(b) Area of contact = 13 cm²
Pressure = $\dfrac{\text{force}}{\text{area}}$
$= \dfrac{500 \text{ N}}{13 \text{ cm}^2} = 38 \text{ N/cm}^2$

(c) Area of contact = 6.5 cm²
Pressure = $\dfrac{\text{force}}{\text{area}}$
$= \dfrac{500 \text{ N}}{6.5 \text{ cm}^2} = 77 \text{ N/cm}^2$

Force = 500 N

Force = 500 N

Force = 500 N

The SI unit for force is the **newton,** (N). *It is the force that will increase the speed of a one-kilogram mass by one meter per second each second that the force is applied.* At Earth's surface, gravity has an acceleration of 9.8 m/s^2. Consider a ballet dancer with a mass of 51 kg, as shown in **Figure 1.** A mass of 51 kg exerts a force of 500 N (51 kg × 9.8 m/s^2) on Earth's surface. No matter how the dancer stands, she exerts that much force against the floor. But the pressure exerted against the floor depends on the area of contact. When the dancer rests her weight on the soles of both feet, as shown in **Figure 1a,** the area of contact with the floor is about 325 cm^2. The pressure, or force per unit area, when she stands in this manner is 500 N/325 cm^2, or roughly 1.5 N/cm^2. When she stands on her toes, as in **Figure 1b,** the total area of contact with the floor is only 13 cm^2. The pressure exerted is then equal to 500 N/13 cm^2—roughly 38 N/cm^2. And when she stands on one toe, as in **Figure 1c,** the pressure exerted is twice that, or about 77 N/cm^2. Thus, the same force applied to a smaller area results in a greater pressure.

The atmosphere—the shell of air surrounding Earth—exerts pressure. **Figure 2** shows that atmospheric pressure at sea level is about equal to the weight of a 1.03 kg mass per square centimeter of surface, or 10.1 N/cm^2. The pressure of the atmosphere can be thought of as caused by the weight of the gases that compose the atmosphere. The atmosphere contains about 78% nitrogen, 21% oxygen, and 1% other gases, including argon and carbon dioxide. Atmospheric pressure is the sum of the individual pressures of the various gases in the atmosphere.

Your ears "pop" when you fly in an airplane because the density—and therefore the pressure—of the air is lower at higher altitudes. The popping happens when the air pressure inside your ears changes to reach the same pressure as the air inside the cabin.

FIGURE 2 Air molecules—most of which are nitrogen or oxygen—collide with Earth's surface, creating a pressure of 10.1 N/cm^2.

Force

1 atm of pressure is the force of 10.1 N on 1 cm^2

Pressure

Oxygen molecule, O$_2$ Nitrogen molecule, N$_2$

Measuring Pressure

A **barometer** *is a device used to measure atmospheric pressure.* The first type of barometer, illustrated in **Figure 3,** was introduced by Evangelista Torricelli during the early 1600s. Torricelli wondered why water pumps could raise water to a maximum height of only about 34 feet. He thought that the height must depend somehow on the weight of water compared with the weight of air. He reasoned that liquid mercury, which is about 14 times as dense as water, could be raised only 1/14 as high as water. To test this idea, Torricelli sealed a long glass tube at one end and filled it with mercury. Holding the open end with his thumb, he inverted the tube into a dish of mercury without allowing any air to enter the tube. When he removed his thumb, the mercury column in the tube dropped to a height of about 30 in. (760 mm) above the surface of the mercury in the dish. He repeated the experiment with tubes of different diameters and lengths longer than 760 mm. In every case, the mercury dropped to a height of about 760 mm.

The space above the mercury in such a tube is nearly a vacuum. The mercury in the tube pushes downward because of gravitational force. The column of mercury in the tube is stopped from falling beyond a certain point because the atmosphere exerts a pressure on the surface of the mercury outside the tube. This pressure is transmitted through the fluid mercury and is exerted upward on the column of mercury. The mercury in the tube falls only until the pressure exerted by its weight is equal to the pressure exerted by the atmosphere.

The exact height of the mercury in the tube depends on the atmospheric pressure, or force per unit area. The pressure is measured in terms of the mercury column's height in the barometer tube.

From experiments like Torricelli's, it is known that at sea level and at 0°C, the average pressure of the atmosphere can support a 760 mm column of mercury. The atmospheric pressure at any given place depends on the elevation and the weather conditions. If the atmospheric pressure is greater than the average at sea level, the height of the mercury column in a barometer will be greater than 760 mm. If the atmospheric pressure is less, the height of the mercury column will be less than 760 mm.

All gases, not only those in the atmosphere, exert pressure. A device called a manometer can be used to measure the pressure of an enclosed gas sample, as shown in **Figure 4.** The difference in the height of mercury in the two arms of the U-tube is a measure of the oxygen gas pressure in the container.

To understand gas pressure, consider a can that is filled with air. The atmosphere exerts a pressure against the outside of the can. The air inside the can pushes outward and balances the atmosphere's inward-pushing force. If a vacuum pump is used to remove the air from the can, the balancing outward force is removed. As a result, the unbalanced force due to atmospheric pressure immediately crushes the can.
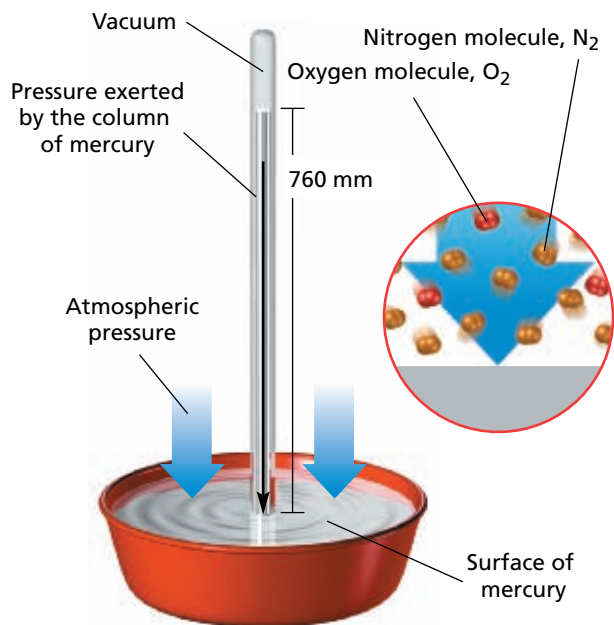

FIGURE 3 Torricelli discovered that the pressure of the atmosphere supports a column of mercury about 760 mm above the surface of the mercury in the dish.
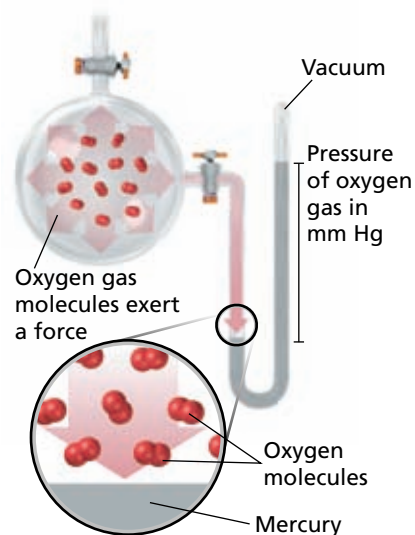

FIGURE 4 In the manometer above, the pressure of the oxygen gas in the flask pushes on the mercury column. The difference in the height of the mercury in the two arms of the U-tube indicates the oxygen gas pressure.

Units of Pressure

A number of different units are used to measure pressure. Because atmospheric pressure is often measured by a mercury barometer, pressure can be expressed in terms of the height of a mercury column. *Thus, the common unit of pressure is* **millimeters of mercury,** *symbolized mm Hg.* A pressure of 1 mm Hg is also called 1 torr in honor of Torricelli for his invention of the barometer. The average atmospheric pressure at sea level at 0°C is 760 mm Hg.

Pressures are often measured in units of atmospheres. *One* **atmosphere of pressure** *(atm) is defined as being exactly equivalent to 760 mm Hg.*

In SI, pressure is expressed in derived units called pascals. The unit is named for Blaise Pascal, a French mathematician and philosopher who studied pressure during the seventeenth century. *One* **pascal** *(Pa) is defined as the pressure exerted by a force of one newton (1 N) acting on an area of one square meter.*

In many cases, it is more convenient to express pressure in kilopascals (kPa). The standard atmosphere (1 atm) is equal to $1.013\ 25 \times 10^5$ Pa, or 101.325 kPa. The pressure units used in this book are summarized in **Table 1.**

Standard Temperature and Pressure

To compare volumes of gases, one must know the temperature and pressure at which the volumes are measured. *For purposes of comparison, scientists have agreed on standard conditions of exactly 1 atm pressure and 0°C.* These conditions are called *standard temperature and pressure* and are commonly abbreviated STP.

TABLE 1 *Units of Pressure*		
Unit	**Symbol**	**Definition/relationship**
pascal	Pa	SI pressure unit $1\ \text{Pa} = \dfrac{1\ \text{N}}{\text{m}^2}$
millimeter of mercury	mm Hg	pressure that supports a 1 mm mercury column in a barometer
torr	torr	1 torr = 1 mm Hg
atmosphere	atm	average atmospheric pressure at sea level and 0°C $\\$ 1 atm = 760 mm Hg $\\$ = 760 torr $\\$ = $1.013\ 25 \times 10^5$ Pa $\\$ = 101.325 kPa
pounds per square inch	psi	1 psi = $6.892\ 86 \times 10^3$ Pa $\\$ 1 atm = 14.700 psi

The average atmospheric pressure in Denver, Colorado, is 0.830 atm. Express this pressure in (a) millimeters of mercury (mm Hg) and (b) kilopascals (kPa).

SOLUTION

1 ANALYZE

Given: P of atmosphere = 0.830 atm

760 mm Hg = 1 atm (definition);   101.325 kPa = 1 atm (definition)

Unknown: a. P of atmosphere in mm Hg;   **b.** P of atmosphere in kPa

2 PLAN

a. atm $\longrightarrow$ mm Hg; $\text{atm} \times \dfrac{760 \text{ mm Hg}}{\text{atm}} = \text{mm Hg}$

b. atm $\longrightarrow$ kPa; $\text{atm} \times \dfrac{101.325 \text{ kPa}}{\text{atm}} = \text{kPa}$

3 COMPUTE

a. $0.830 \text{ atm} \times \dfrac{760 \text{ mm Hg}}{\text{atm}} = 631 \text{ mm Hg}$

b. $0.830 \text{ atm} \times \dfrac{101.325 \text{ kPa}}{\text{atm}} = 84.1 \text{ kPa}$

4 EVALUATE

Units have canceled to give the desired units, and answers are expressed to the correct number of significant figures. The known pressure and the calculated pressures are about 80% of the atmospheric pressure, as expressed in the new units.

PRACTICE

Answers in Appendix E

1. Convert a pressure of 1.75 atm to kPa and to mm Hg.

2. The critical pressure of carbon dioxide is 72.7 atm. What is this value in units of pascals?

extension
Go to **go.hrw.com** for more practice problems that ask you to convert between units of pressure.

Keyword: HC6GASX

Dalton's Law of Partial Pressures

John Dalton, the English chemist who proposed the atomic theory, also studied gas mixtures. The pressure exerted by each gas in an unreactive mixture is independent of that exerted by other gases present. *The pressure of each gas in a mixture is called the* **partial pressure** *of that gas.* **Dalton's law of partial pressures** *states that the total pressure of a gas mixture is the sum of the partial pressures of the component gases.* The law is true regardless of the number of different gases that are present. Dalton's law may be expressed as follows.

$$P_T = P_1 + P_2 + P_3 + \ldots$$

P_T is the total pressure of the mixture. P_1, P_2, P_3, and so on are the partial pressures of component gases 1, 2, 3, and so on.

FIGURE 5 Hydrogen can be collected by water displacement by reacting zinc with sulfuric acid. The hydrogen gas produced displaces the water in the gas collecting bottle. It now contains some water vapor.
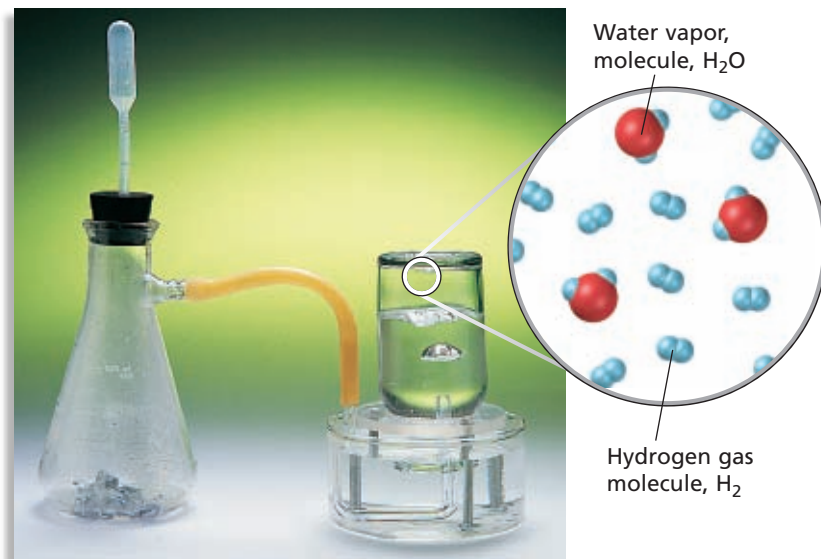
You can understand Dalton's law in terms of the kinetic-molecular theory. Each of the rapidly moving particles of gases in a mixture has an equal chance to collide with the container walls. Therefore, each gas exerts a pressure independent of that exerted by the other gases present. The total pressure is the result of the total number of collisions per unit of wall area in a given time.

Gases Collected by Water Displacement

Gases produced in the laboratory are often collected over water, as shown in **Figure 5.** The gas produced by the reaction displaces the water, which is more dense, in the collection bottle. You can apply Dalton's law of partial pressures in calculating the pressures of gases collected in this way. A gas collected by water displacement is not pure but is always mixed with water vapor. That is because water molecules at the liquid surface evaporate and mix with the gas molecules. Water vapor, like other gases, exerts a pressure known as *vapor pressure*.

Suppose you wished to determine the total pressure of the gas and water vapor inside a collection bottle. You would raise the bottle until the water levels inside and outside the bottle were the same. At that point, the total pressure inside the bottle would be the same as the atmospheric pressure, P_{atm}. According to Dalton's law of partial pressures, the following is true.

$$P_{atm} = P_{gas} + P_{H_2O}$$

Suppose you then needed to calculate the partial pressure of the dry gas collected. You would read the atmospheric pressure, P_{atm}, from a barometer in the laboratory. To make the calculation, subtract the vapor pressure of the water at the given temperature from the total pressure. The vapor pressure of water varies with temperature. You need to look up the value of P_{H_2O} at the temperature of the experiment in a standard reference table, such as Table A-8 in Appendix A of this book.

Oxygen gas from the decomposition of potassium chlorate, KClO₃, was collected by water displacement. The barometric pressure and the temperature during the experiment were 731.0 torr and 20.0°C, respectively. What was the partial pressure of the oxygen collected?

SOLUTION

1 ANALYZE

Given: $P_T = P_{atm} = 731.0$ torr
$P_{H_2O} = 17.5$ torr (vapor pressure of water at 20.0°C, from Table A-8)
$P_{atm} = P_{O_2} + P_{H_2O}$
Unknown: P_{O_2} in torr

2 PLAN

The partial pressure of the collected oxygen is found by subtracting the partial pressure of water vapor from the atmospheric pressure, according to Dalton's law of partial pressures.

$$P_{O_2} = P_{atm} - P_{H_2O}$$

3 COMPUTE

Substituting values for P_{atm} and P_{H_2O} gives P_{O_2}.

$$P_{O_2} = 731.0 \text{ torr} - 17.5 \text{ torr} = 713.5 \text{ torr}$$

4 EVALUATE

As expected, the oxygen partial pressure is less than atmospheric pressure. It is reasonably close to an estimated value of 713, calculated as 730 – 17.

PRACTICE

Answers in Appendix E

1. Some hydrogen gas is collected over water at 20.0°C. The levels of water inside and outside the gas-collection bottle are the same. The partial pressure of hydrogen is 742.5 torr. What is the barometric pressure at the time the gas is collected?

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate partial pressure.

Keyword: HC6GASX

SECTION REVIEW

1. Define *pressure.*
2. What units are used to express pressure measurements?
3. What are standard conditions for gas measurements?
4. Convert the following pressures to pressures in standard atmospheres:
   a. 151.98 kPa   **b.** 456 torr
5. A sample of nitrogen gas is collected over water at a temperature of 23.0°C. What is the pressure of the nitrogen gas if atmospheric pressure is 785 mm Hg?
6. Why can you calculate the total pressure of a mixture of gases by adding together the partial pressures of the component gases?

Critical Thinking

7. **EVALUATING METHODS** Clean rooms used for sterile biological research are sealed and operate at slightly above atmospheric pressure. Explain why.
8. **INFERRING RELATIONSHIPS** Explain why helium-filled balloons deflate over time faster than air-filled balloons do.

Chemistry in Action

The Gas Laws and Scuba Diving

An understanding of Dalton's law and Henry's law is essential to safe scuba diving. Dalton's law states that the total pressure of a gas mixture is equal to the sum of the partial pressures of the component gases. Henry's law predicts that the solubility of a gas in a liquid is a direct function of the partial pressure of that gas.

For every 33 ft of sea water that a diver descends, he or she feels one additional atmosphere of pressure because of the increasing weight of water overhead. Most divers use compressed air tanks to breathe underwater. The air in these tanks, which contains approximately 78% nitrogen and 21% oxygen, is the same as the air that we breathe. Once the compressed air enters the diver's lungs, it is subjected to the pressure caused by the water. The

increase in the air pressure leads to an increase in the partial pressures of the nitrogen and oxygen in air, as predicted by Dalton's law. Henry's law predicts that this increase in partial pressures will increase the solubility of nitrogen and oxygen in the diver's bloodstream.

The increase in the partial pressure of oxygen is not problematic under typical diving conditions because a diver's body can metabolize the extra oxygen that is present in the bloodstream. The body does not metabolize nitrogen, however, so it can accumulate in a diver's body during a dive. The extra nitrogen can affect the nerve cells of the diver, causing *nitrogen narcosis.* Divers suffering from nitrogen narcosis become disoriented and experience symptoms similar to intoxication. This impairment of a diver's judgment can be very dangerous. To decrease the probability of contracting nitrogen narcosis, divers can use gas mixtures that contain less nitrogen than compressed air does.

Dissolved nitrogen can also be harmful if a diver ascends too quickly. As Henry's law predicts, nitrogen becomes less soluble in the blood as the pressure decreases. This decrease in solubility causes nitrogen to leave the diver's tissues and blood. Normally,

the excess nitrogen is discharged through the lungs. However, if the diver comes up too rapidly, the nitrogen will form bubbles in the tissues and veins. This condition is known as *decompression sickness,* or "the bends." If the bubbles block blood flow, a wide range of effects, including severe joint pain, may occur. For this reason, divers are very careful to ascend slowly after diving.

At extreme depths, even the increased pressure of oxygen becomes a problem for scuba divers. Although oxygen is necessary for life, extended exposure to large concentrations of oxygen can result in *oxygen toxicity.* This condition can cause damage to the lungs and nervous system. Divers can avoid oxygen toxicity by breathing gas mixtures that contain more helium and less oxygen than compressed air does.

Questions

1. Why is the concentration of helium increased in some gas mixtures that divers breathe in order to reduce the partial pressure of oxygen?

2. Although carbon monoxide, CO, is normally present in trace quantities in air, compressed air scuba tanks that are filled improperly can contain elevated levels of CO. Why is this particularly dangerous for divers?

◄ *In order to dive safely, scuba divers must be aware of some of the basic gas laws.*

The Gas Laws

Objectives

- Use the kinetic-molecular the-ory to explain the relation-ships between gas volume, temperature, and pressure.

- Use Boyle's law to calculate volume-pressure changes at constant temperature.

- Use Charles's law to calculate volume-temperature changes at constant pressure.

- Use Gay-Lussac's law to cal-culate pressure-temperature changes at constant volume.

- Use the combined gas law to calculate volume-tempera-ture-pressure changes.

S cientists have been studying physical properties of gases for hundreds of years. In 1662, Robert Boyle discovered that gas pressure and volume are related mathematically. The observations of Boyle and others led to the development of the gas laws. The *gas laws* are simple mathematical rela-tionships between the volume, temperature, pressure, and amount of a gas.

Boyle's Law: Pressure-Volume Relationship

Robert Boyle discovered that doubling the pressure on a sample of gas at constant temperature reduces its volume by one-half. Reducing the pressure on a gas by one-half allows the volume of the gas to double. As one variable increases, the other decreases. **Figure 6** shows that as the volume of gas in the syringe decreases, the pressure of the gas increases.

You can use the kinetic-molecular theory to understand why this pressure-volume relationship holds. The pressure of a gas is caused by moving molecules hitting the container walls. Suppose the volume of a container is decreased but the same number of gas molecules is present at the same temperature. There will be more molecules per unit volume. The number of collisions with a given unit of wall area per unit time will increase as a result. Therefore, pressure will also increase.
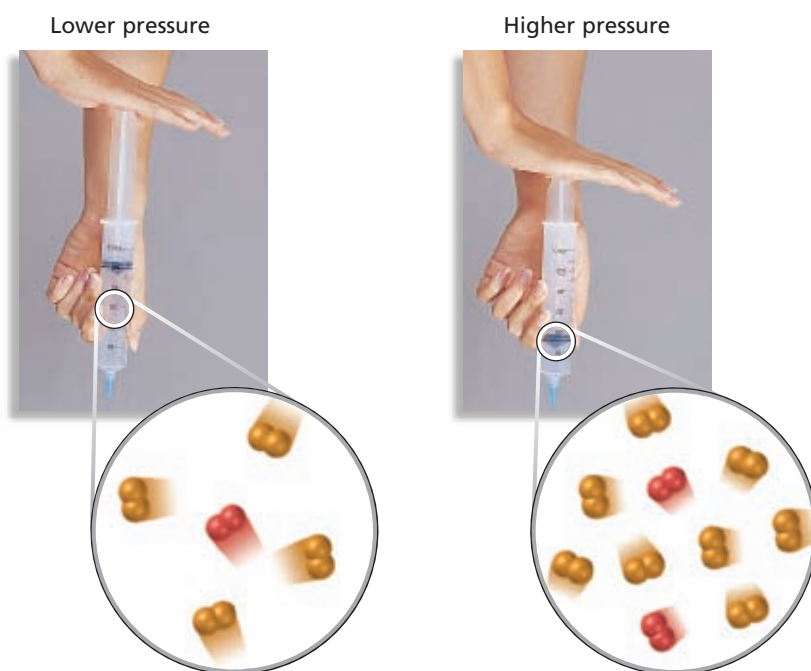
Lower pressure   Higher pressure

FIGURE 6   The volume of gas in the syringe shown in the photo is reduced when the plunger is pushed down. The gas pressure increases as the volume is reduced because the molecules collide more frequently with the walls of the container in a smaller volume.

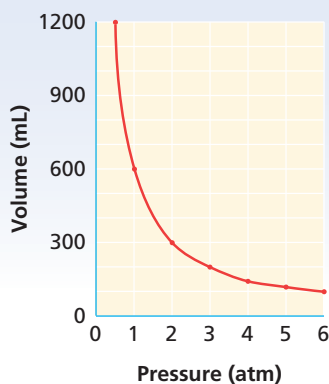## Volume Vs. Pressure for a Gas at Constant Temperature

FIGURE 7 This graph shows that there is an inverse relationship between volume and pressure.

Plotting the values of volume versus pressure for a gas at constant temperature gives a curve like that in **Figure 7.** The general volume-pressure relationship that is illustrated is called Boyle's law. **Boyle's law** *states that the volume of a fixed mass of gas varies inversely with the pressure at constant temperature.*

Mathematically, Boyle's law can be expressed as follows:

$$PV = k$$

In the equation above, P is the pressure, V is the volume, and k is a constant. Since P and V vary inversely, their product remains the same. Because two quantities that are equal to the same thing are equal to each other, the relationship between changes of pressure and volume can be expressed as shown below.

$$P_1V_1 = P_2V_2$$

P_1 and V_1 represent initial conditions. P_2 and V_2 represent a different set of conditions. Given three of the four values P_1, V_1, P_2, and V_2, you can use this equation to calculate the fourth value for a system at constant temperature.

SAMPLE PROBLEM C

For more help, go to the **Math Tutor** *at the end of this chapter.*

A sample of oxygen gas has a volume of 150.0 mL when its pressure is 0.947 atm. What will the volume of the gas be at a pressure of 0.987 atm if the temperature remains constant?

SOLUTION

1 ANALYZE
Given: V_1 of $O_2 = 150.0$ mL; P_1 of $O_2 = 0.947$ atm; P_2 of $O_2 = 0.987$ atm
Unknown: V_2 of O_2 in mL

2 PLAN
Rearrange the equation for Boyle's law $(P_1V_1 = P_2V_2)$ to obtain V_2.

$$V_2 = \frac{P_1V_1}{P_2}$$

3 COMPUTE
Substitute values for P_1, V_1, and P_2 to obtain the new volume, V_2.

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(0.947 \text{ atm})(150.0 \text{ mL O}_2)}{0.987 \text{ atm}} = 144 \text{ mL O}_2$$

4 EVALUATE
When the pressure is increased slightly at constant temperature, the volume decreases slightly, as expected. Units cancel to give milliliters, a volume unit.

PRACTICE

Answers in Appendix E

1. A balloon filled with helium gas has a volume of 500 mL at a pressure of 1 atm. The balloon is released and reaches an altitude of 6.5 km, where the pressure is 0.5 atm. If the temperature has remained the same, what volume does the gas occupy at this height?

extension
Go to **go.hrw.com** for more practice problems that ask you to use Boyle's Law.

Keyword: HC6GASX

Charles's Law: Volume-Temperature Relationship

Balloonists, such as those in the photo at the beginning of this chapter, are making use of a physical property of gases: if pressure is constant, gases expand when heated. When the temperature increases, the volume of a fixed number of gas molecules must increase if the pressure is to stay constant. At the higher temperature, the gas molecules move faster. They collide with the walls of the container more frequently and with more force. The volume of a flexible container must then increase in order for the pressure to remain the same.

The quantitative relationship between volume and temperature was discovered by the French scientist Jacques Charles in 1787. Charles's experiments showed that all gases expand to the same extent when heated through the same temperature interval. Charles found that the volume changes by 1/273 of the original volume for each Celsius degree, at constant pressure and an initial temperature of 0°C. For example, raising the temperature to 1°C causes the gas volume to increase by 1/273 of the volume it had at 0°C. If the temperature is increased by 273°C, the volume increases by 273/273 of the original, that is, the volume doubles. The same regularity of volume change occurs if a gas is cooled at constant pressure, as the balloons in **Figure 8** show.

The Kelvin temperature scale is a scale that starts at a temperature corresponding to –273.15°C. That temperature is the lowest one possible. *The temperature –273.15°C is referred to as* **absolute zero** *and is given a value of zero in the Kelvin scale.* This fact gives the following relationship between the two temperature scales.

$$K = 273.15 + °C$$

For calculations in this book, 273.15 is rounded off to 273.

FIGURE 8   As air-filled balloons are exposed to liquid nitrogen, they shrink greatly in volume because the air molecules in the balloon get closer together. When the balloons are removed from the liquid nitrogen and the air inside them is warmed to room temperature, they expand to their original volume.

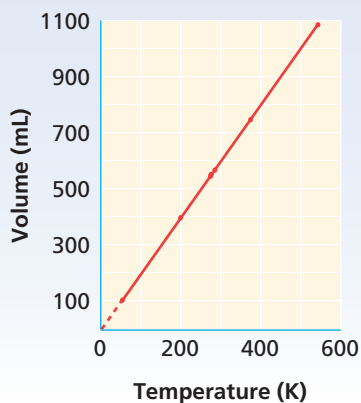## Volume Vs. Temperature for a Gas at Constant Pressure

FIGURE 9 This graph shows the plot of the volume versus the Kelvin temperature data of a sample of gas at constant pressure.

Gas volume and Kelvin temperature are directly proportional to each other at constant pressure as shown in **Figure 9.**

The relationship between Kelvin temperature and gas volume is known as Charles's law. **Charles's law** *states that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.* Charles's law may be expressed as follows:

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

The value of T is the Kelvin temperature, and k is a constant. The ratio V/T for any set of volume-temperature values always equals the same k. The form of Charles's law that can be applied directly to most volume-temperature problems involving gases is as follows:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V_1 and T_1 represent initial conditions. V_2 and T_2 represent a different set of conditions. When three of the four values T_1, V_1, T_2, and V_2 are known, you can use this equation to calculate the fourth value for a system at constant pressure.

SAMPLE PROBLEM D   *For more help, go to the* **Math Tutor** *at the end of this chapter.*

A sample of neon gas occupies a volume of 752 mL at 25°C. What volume will the gas occupy at 50°C if the pressure remains constant?

SOLUTION

1 ANALYZE   **Given:** V_1 of Ne = 752 mL; T_1 of Ne = 25°C + 273 = 298 K; T_2 of Ne = 50°C + 273 = 323 K
Unknown: V_2 of Ne in mL

2 PLAN   Because the gas remains at constant pressure, an increase in temperature will cause an increase in volume. To obtain V_2, rearrange the equation for Charles's law.

3 COMPUTE   Substitute values for V_1, T_1, and T_2 to obtain the new volume, V_2.

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(752 \text{ mL Ne})(323 \text{ K})}{298 \text{ K}} = 815 \text{ mL Ne}$$

4 EVALUATE   As expected, the volume of the gas increases as the temperature increases. Units cancel to yield milliliters, as desired. The answer contains the appropriate number of significant figures.

PRACTICE   *Answers in Appendix E*

1. A sample of neon gas has a volume of 752 mL at 25.0°C. What will the volume at 100.0°C be if pressure is constant?

2. A sample of nitrogen gas is contained in a piston with a freely moving cylinder. At 0.0°C, the volume of the gas is 375 mL. To what temperature must the gas be heated to occupy a volume of 500.0 mL?

extension
Go to **go.hrw.com** for more practice problems that ask you to use Charles's law.

Keyword: HC6GASX

Gay-Lussac's Law: Pressure-Temperature Relationship

You have just learned about the quantitative relationship between volume and temperature at constant pressure. What would you predict about the relationship between pressure and temperature at constant volume? You have seen that pressure is the result of collisions of molecules with container walls. The energy and frequency of collisions depend on the average kinetic energy of molecules. For a fixed quantity of gas at constant volume, the pressure should be directly proportional to the Kelvin temperature, which depends directly on average kinetic energy.

That prediction turns out to be correct. For every kelvin of temperature change, the pressure of a confined gas changes by 1/273 of the pressure at 0°C. Joseph Gay-Lussac is given credit for recognizing this in 1802. The data plotted in **Figure 10** illustrate **Gay-Lussac's law:** *The pressure of a fixed mass of gas at constant volume varies directly with the Kelvin temperature.* Mathematically, Gay-Lussac's law is expressed as follows.

$$P = kT \quad \text{or} \quad \frac{P}{T} = k$$

The value of T is the temperature in kelvins, and k is a constant that depends on the quantity of gas and the volume. For a given mass of gas at constant volume, the ratio P/T is the same for any set of pressure-temperature values. Unknown values can be found using the following form of Gay-Lussac's law.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

P_1 and T_1 represent initial conditions. P_2 and T_2 represent a different set of conditions. When values are known for three of the four quantities, the fourth value can be calculated for a system at constant volume.

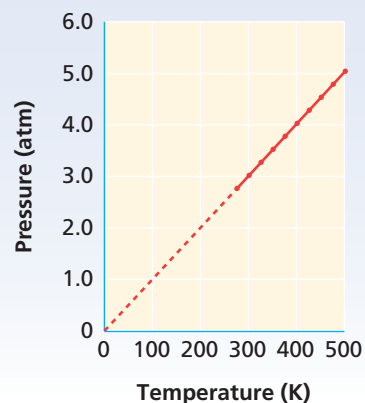## Pressure Vs. Temperature for a Gas at Constant Volume

FIGURE 10 A graph of pressure versus temperature of a sample of gas at constant volume is a straight line, indicating a direct relationship between pressure and temperature.

SAMPLE PROBLEM E

For more help, go to the Math Tutor at the end of this chapter.

The gas in a container is at a pressure of 3.00 atm at 25°C. Directions on the container warn the user not to keep it in a place where the temperature exceeds 52°C. What would the gas pressure in the container be at 52°C?

SOLUTION

1 ANALYZE

Given: P_1 of gas = 3.00 atm
T_1 of gas = 25°C + 273 = 298 K; T_2 of gas = 52°C + 273 = 325 K
Unknown: P_2 of gas in atm

2 PLAN

Because the gaseous contents remain at the constant volume of the container, an increase in temperature will cause an increase in pressure. Rearrange Gay-Lussac's law to obtain P_2.

$$P_2 = \frac{P_1 T_2}{T_1}$$

3 COMPUTE Substitute values for P_1, T_2, and T_1 to obtain the new pressure, P_2.

$$P_2 = \frac{(3.00 \text{ atm})(325 \text{ K})}{298 \text{ K}} = 3.27 \text{ atm}$$

4 EVALUATE As expected, a temperature increase at constant volume causes the pressure of the contents in the container to increase. Units cancel correctly. The answer contains the proper number of significant figures. It is close to an estimated value of 3.25, calculated as $(3 \times 325)/300$.

PRACTICE   *Answers in Appendix E*

1. At 120.0°C, the pressure of a sample of nitrogen is 1.07 atm. What will the pressure be at 205°C, assuming constant volume?

2. At 122°C the pressure of a sample of nitrogen gas is 1.07 atm. What will the pressure be at 205°C, assuming constant volume?

3. A sample of helium gas has a pressure of 1.20 atm at 22°C. At what Celsius temperature will the helium reach a pressure of 2.00 atm, assuming constant volume?

The Combined Gas Law

A gas sample often undergoes changes in temperature, pressure, and volume all at the same time. When this happens, three variables must be dealt with at once. Boyle's law, Charles's law, and Gay-Lussac's law can be combined into a single expression that is useful in such situations. *The* **combined gas law** *expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.* The combined gas law can be expressed as follows.

$$\frac{PV}{T} = k$$

In the equation, k is constant and depends on the amount of gas. The combined gas law can also be written as follows.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

The subscripts in the equation above indicate two different sets of conditions, and T represents Kelvin temperature.

From this expression, any value can be calculated if the other five are known. Note that each of the gas laws can be obtained from the combined gas law when the proper variable is constant. For example, Boyle's law is obtained when the temperature is constant. Because $T_1 = T_2$, T_1 and T_2 will cancel out on both sides of the combined gas law equation, giving Boyle's law.

$$P_1 V_1 = P_2 V_2$$

A helium-filled balloon has a volume of 50.0 L at 25°C and 1.08 atm. What volume will it have at 0.855 atm and 10.0°C?

SOLUTION

1 ANALYZE

Given: V_1 of He = 50.0 L; T_1 of He = 25°C + 273 = 298 K; T_2 of He = 10°C + 273 = 283 K
P_1 of He = 1.08 atm; P_2 of He = 0.855 atm
Unknown: V_2 of He in L

2 PLAN

Because the gas changes in both temperature and pressure, the combined gas law is needed. Rearrange the combined gas law to solve for the final volume, V_2.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \longrightarrow V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

3 COMPUTE

Substitute the known values into the equation to obtain a value for V_2.

$$V_2 = \frac{(1.08 \text{ atm})(50.0 \text{ L He})(283 \text{ K})}{(0.855 \text{ atm})(298 \text{ K})} = 60.0 \text{ L He}$$

4 EVALUATE

Here the pressure decreases much more than the temperature decreases. As expected, the net result of the two changes gives an increase in the volume, from 50.0 L to 60.0 L. Units cancel appropriately. The answer is correctly expressed to three significant figures. It is also reasonably close to an estimated value of 50, calculated as $(50 \times 300)/300$.

PRACTICE   *Answers in Appendix E*

1. The volume of a gas is 27.5 mL at 22.0°C and 0.974 atm. What will the volume be at 15.0°C and 0.993 atm?

2. A 700.0 mL gas sample at STP is compressed to a volume of 200.0 mL, and the temperature is increased to 30.0°C. What is the new pressure of the gas in Pa?

SECTION REVIEW

1. Relate the effect of temperature and pressure on a gas to the model of a gas given by the kinetic-molecular theory.

2. A sample of helium gas has a volume of 200.0 mL at 0.960 atm. What pressure, in atmospheres, is needed to reduce the volume at constant temperature to 50.0 mL?

3. A sample of nitrogen gas occupies 1.55 L at 27.0°C and 1.00 atm. What will the volume be at −100.0°C and the same pressure?

4. A gas occupies 2.0 m³ at 100.0 K and exerts a pressure of 100.0 kPa. What volume will the gas occupy if the temperature is increased to 400.0 K and the pressure is increased to 200.0 kPa?

Critical Thinking

5. **ANALYZING RESULTS** A student has the following data: V_1 = 822 mL, T_1 = 75°C, and T_2 = −25°C. He calculates V_2 and gets −274 mL. Is this value correct? Explain why or why not.

6. **APPLYING MODELS** Explain Charles's law in terms of the kinetic-molecular theory.

HISTORICAL CHEMISTRY

Chemistry's First Law

Historical Perspective

The notion that "nature abhors a vacuum"—meaning that there is no such thing as "empty space"—was proposed by the Greek philosopher Aristotle, whose word was unchallenged for nearly 2000 years. Then, in the mid-1600s, a new breed of thinkers known as *natural philosophers*—what we now know as "scientists"—began testing the long-held assumption that all space must contain matter. These investigations were some of the earliest experiments with gases, and they led to the discovery of the first empirical principle of chemistry, Boyle's law.

Overturning an Ancient Assumption

The first scientist to demonstrate the existence of a vacuum was Evangelista Torricelli. In 1643, he showed that when a glass tube that was 3 ft. long and about 1 in. in diameter was sealed at one end, filled with mercury, and inverted in a container full of mercury, the mercury in the tube fell to a height of about 30 in. above the level of mercury in the container. Some thinkers remained skeptical, but it was generally accepted that the space between the mercury and the sealed end of the tube was indeed a vacuum.

▲ *Evangelista Torricelli invented the mercury barometer.*

Torricelli then turned his attention to how the mercury in the glass tube of his apparatus was supported. The fact that liquids exert a pressure on objects immersed in them inspired him to hypothesize that a "sea of air" surrounded Earth. He further hypothesized that the air exerted pressure on the mercury in the container and thus supported the mercury in the column.

Support for the New Theory

Although the idea of an atmosphere that has weight and exerts a pressure on the objects within it seems obvious today, it was a radical theory at the time.

To test the effects of the atmosphere, Robert Boyle, one of the period's great scientists, had his talented assistant, Robert Hooke, create a piece of equipment that would revolutionize the study of air. The apparatus was an improved version of a pump designed by the German experimenter Otto von Guericke; the pump had a large receptacle in which a partial vacuum could be created.

Boyle placed Torricelli's setup, known today as a *barometer*, in the receptacle of the pump and observed the mercury column as he reduced the pressure around it. The height of the mercury decreased as the pressure surrounding the mercury in the container dropped, strongly supporting Torricelli's atmospheric theory.

Using Hooke's pump, Boyle performed additional studies that verified the idea that air exerted pressure and had weight. Boyle's experiments also led to the important conclusion that air was elastic: that is, it could expand and contract. Boyle discovered the fundamental law that bears his name during an investigation into air's elasticity.

An Ingenious Experiment

In response to a criticism of his findings, Boyle performed an experiment to show that air could be compressed to a pressure greater than that of the atmosphere. First, he prepared a glass J-tube with the short end sealed off and the long end left open. Then, he poured mercury into the tube, making sure that the levels in the two ends were the same and letting air travel freely between the ends, to ensure that each column was at atmospheric pressure.

Then, by pouring more mercury into the long end of the tube until it was about 30 in. above the level of mercury in the short end, Boyle exposed the trapped air to about twice as much atmospheric pressure. He observed that the

◄We can measure the difference between the atmospheric pressure on a mountaintop and the atmospheric pressure at sea level thanks to the work of Torricelli and Boyle.
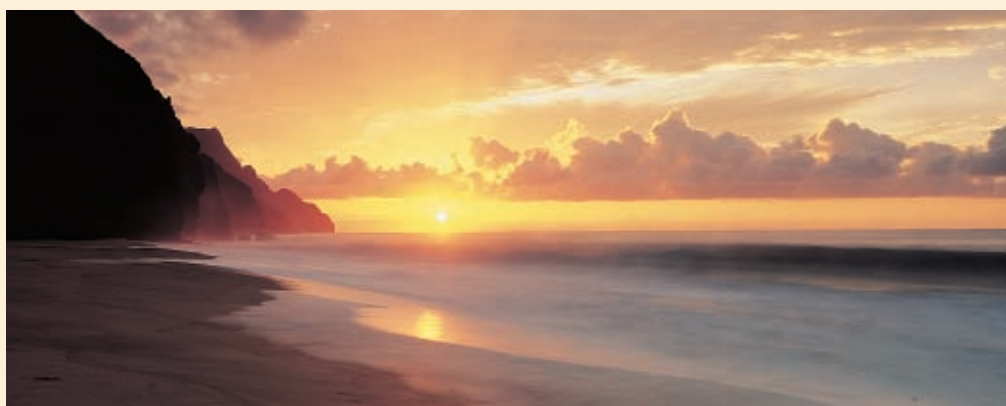
volume of the trapped air was halved. He continued to add mercury until the total pressure on the trapped air was about 4 times that of the atmosphere. Noting that the air had been compressed to a volume about one-fourth of what it originally was, Boyle discovered the inverse relationship between air's pressure and volume.

A Long-Standing Contribution

Boyle went on to show that the relationship between air pressure and volume, $P \propto 1/V$ (at constant temperature), held not only when the gas was compressed but also when it was allowed to expand. Future investigators would show that the law is a principle that applies to gases in general. Along with the findings of other researchers, such as Jacques Charles, Joseph Gay-Lussac, and Amadeo Avogadro, Boyle's discovery led chemists to the famous ideal gas law, $PV = nRT$, which serves as a starting point in the study of chemistry today.

Questions

1. Why was it necessary for Boyle to seal one end of his J-tube?

2. What would have happened if Boyle had allowed the temperature of the trapped air sample to change as he added mercury to the tube?

SC**L**INKS.

www.scilinks.org
Topic: Torricelli
Code: HC61539

Topic: Robert Boyle
Code: HC61317

Gas Volumes and the Ideal Gas Law

OBJECTIVES

- State the law of combining volumes.

- State Avogadro's law and explain its significance.

- Define *standard molar volume of a gas* and use it to calculate gas masses and volumes.

- State the ideal gas law.

- Using the ideal gas law, calculate pressure, volume, temperature, or amount of gas when the other three quantities are known.

In this section, you will study the relationships between the volumes of gases that react with each other. You will also learn about the relationship between molar amount of gas and volume, and a single gas law that unifies all the basic gas laws into a single equation.
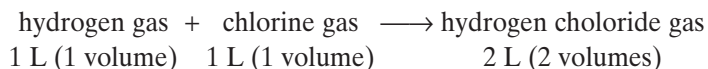
Measuring and Comparing the Volumes of Reacting Gases

In the early 1800s, French chemist Joseph Gay-Lussac studied gas volume relationships involving a chemical reaction between hydrogen and oxygen. He observed that 2 L of hydrogen can react with 1 L of oxygen to form 2 L of water vapor at constant temperature and pressure.

<div align="center">

hydrogen gas + oxygen gas ⟶ water vapor

2 L (2 volumes)   1 L (1 volume)   2 L (2 volumes)

</div>

In other words, this reaction shows a simple and definite 2:1:2 relationship between the volumes of the reactants and the product. Two volumes of hydrogen react with 1 volume of oxygen to produce 2 volumes of water vapor. The 2:1:2 relationship for this reaction applies to any proportions for volume—for example, 2 mL, 1 mL, and 2 mL; 600 L, 300 L, and 600 L; or 400 cm^3, 200 cm^3, and 400 cm^3.

Gay-Lussac also noticed simple and definite proportions by volume in other reactions of gases, such as in the reaction between hydrogen gas and chlorine gas.

<div align="center">

hydrogen gas + chlorine gas ⟶ hydrogen choloride gas

1 L (1 volume)   1 L (1 volume)   2 L (2 volumes)

</div>

In 1808, Gay-Lussac summarized the results of his experiments in a statement known today as **Gay-Lussac's law of combining volumes of gases.** The law states that *at constant temperature and pressure, the volumes of gaseous reactants and products can be expressed as ratios of small whole numbers.* This simple observation, combined with the insight of Avogadro, provided more understanding of how gases react and combine with each other.
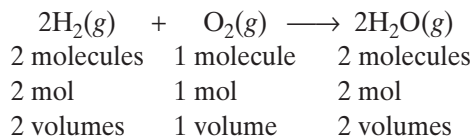
Avogadro's Law

Recall an important point of Dalton's atomic theory: atoms are indivisible. Dalton also thought that the particles of gaseous elements exist in the form of isolated single atoms. He believed that one atom of one element always combines with one atom of another element to form a single particle of the product. However, some of the volume relationships observed by Gay-Lussac could not be accounted for by Dalton's theory. For example, in reactions such as the formation of water vapor, mentioned on the preceding page, it would seem that the oxygen atoms involved would have to divide into two parts.

In 1811, Avogadro found a way to explain Gay-Lussac's simple ratios of combining volumes without violating Dalton's idea of indivisible atoms. He did this by rejecting Dalton's idea that reactant elements are always in monatomic form when they combine to form products. He reasoned that these molecules could contain more than one atom. Avogadro also put forth an idea known today as **Avogadro's law.** The law states that *equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.* **Figure 11** illustrates Avogadro's law. It follows that at the same temperature and pressure, the volume of any given gas varies directly with the number of molecules. Avogadro's law also indicates that gas volume is directly proportional to the amount of gas, at a given temperature and pressure. Note the equation for this relationship.

$$V = kn$$

Here, n is the amount of gas, in moles, and k is a constant.

Avogadro's reasoning applies to the combining volumes for the reaction of hydrogen and oxygen to form water vapor. Dalton had guessed that the formula of water was HO, because this formula seemed to be the most likely formula for such a common compound. But Avogadro's reasoning established that water must contain twice as many H atoms as O atoms, consistent with the formula H_2O. As shown below, the coefficients in a chemical reaction involving gases indicate the relative numbers of molecules, the relative numbers of moles, and the relative volumes.

$$2H_2(g) \; + \; O_2(g) \; \longrightarrow \; 2H_2O(g)$$

2 molecules	1 molecule	2 molecules
2 mol	1 mol	2 mol
2 volumes	1 volume	2 volumes

The simplest hypothetical formula for oxygen indicated two oxygen atoms, which turns out to be correct. The simplest possible molecule of water indicated two hydrogen atoms and one oxygen atom per molecule, which is also correct. Experiments eventually showed that all elements that are gases near room temperature, except the noble gases, normally exist as diatomic molecules.
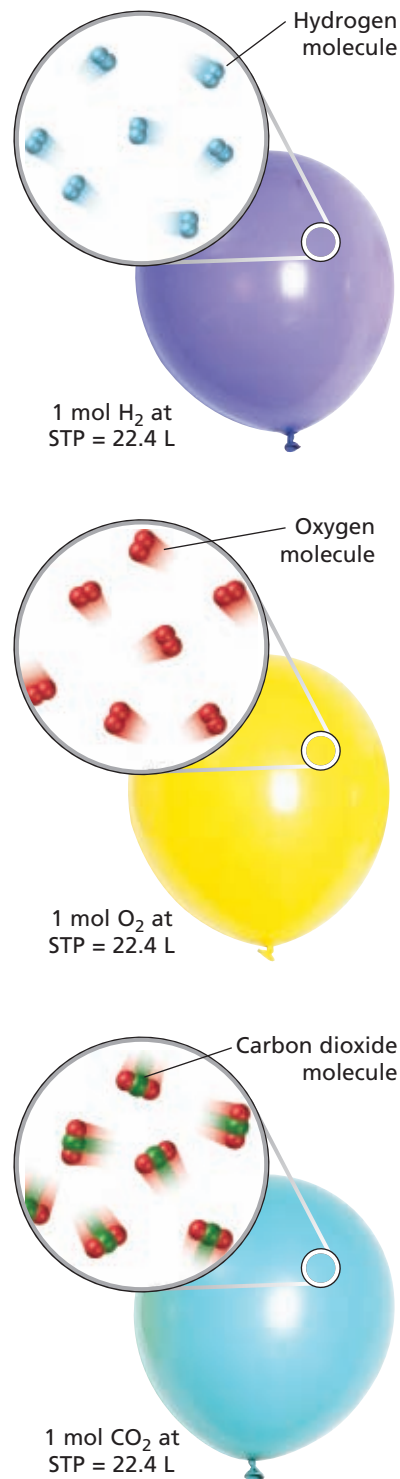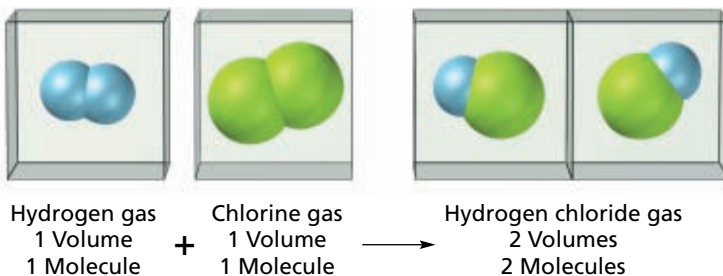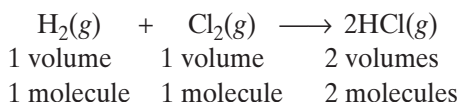

Hydrogen molecule

1 mol H_2 at STP = 22.4 L

Oxygen molecule

1 mol O_2 at STP = 22.4 L

Carbon dioxide molecule

1 mol CO_2 at STP = 22.4 L

FIGURE 11 At the same temperature and pressure, balloons of equal volume have equal numbers of molecules, regardless of which gas they contain.

Hydrogen gas
1 Volume
1 Molecule

+

Chlorine gas
1 Volume
1 Molecule

Hydrogen chloride gas
2 Volumes
2 Molecules

Consider the reaction of hydrogen and chlorine to produce hydrogen chloride, illustrated in **Figure 12.** According to Avogadro's law, equal volumes of hydrogen and chlorine contain the same number of molecules. Avogadro's idea of diatomic gases applies to this reaction also. He concluded that the hydrogen and chlorine components must each consist of two or more atoms joined together. The simplest assumption was that hydrogen and chlorine molecules are composed of two atoms each. That assumption leads to the following balanced equation for the reaction of hydrogen with chlorine.

$$H_2(g) \quad + \quad Cl_2(g) \quad \longrightarrow \quad 2HCl(g)$$

1 volume   1 volume   2 volumes

1 molecule   1 molecule   2 molecules

The simplest hypothetical formula for hydrogen chloride, HCl, indicates that the molecule contains one hydrogen atom and one chlorine atom. Given the ratios of the combined volumes, the simplest formulas for hydrogen and chlorine must be H_2 and Cl_2, respectively.

Molar Volume of a Gas

Recall that one mole of a molecular substance contains a number of molecules equal to Avogadro's constant (6.022×10^{23}). One mole of oxygen, O_2, contains 6.022×10^{23} diatomic oxygen molecules and has a mass of 31.9988 g. One mole of helium, a monatomic gas, contains the same number of helium atoms and has a mass of 4.002 602 g.

According to Avogadro's law, one mole of any gas will occupy the same volume as one mole of any other gas at the same temperature and pressure, despite mass differences. *The volume occupied by one mole of a gas at STP is known as the* **standard molar volume of a gas.** *It has been found to be 22.414 10 L.* For calculations in this book, we use 22.4 L as the standard molar volume.

Knowing the volume of a gas, you can use 1 mol/22.4 L as a conversion factor to find the number of moles, and therefore the mass, of a given volume of a given gas at STP. You can also use the molar volume of a gas to find the volume, at STP, of a known number of moles or a known mass of a gas.

Chemistry in Action
Automobile Air Bags

Since the late 1980s, air bags have been offered as a safety feature in cars to minimize injuries in the event of a high-speed collision. Modern automobile air bags use a series of very rapid chemical reactions to inflate the bag. When a collision is detected by sensors, an igniter triggers decomposition of solid sodium azide, Na_3N, to form N_2 gas. The hazardous sodium metal that also forms reacts with KNO_3 to form Na_2O, also producing additional N_2. Finally, the highly reactive Na_2O is removed by reaction with SiO_2 to form harmless silicate glass. This entire sequence of reactions occurs to inflate the air bag with nitrogen gas in as few as 40 milliseconds (0.04 s) after a collision is detected by sensors.

a. **What volume does 0.0685 mol of gas occupy at STP?**
b. **What quantity of gas, in moles, is contained in 2.21 L at STP?**

SOLUTION

a. Multiply the amount in moles by the conversion factor, $\dfrac{22.4 \text{ L}}{1 \text{ mol}}$.

$$0.0685 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 1.53 \text{ L}$$

b. Multiply the volume in liters by the conversion factor, $\dfrac{1 \text{ mol}}{22.4 \text{ L}}$.

$$2.21 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.0987 \text{ mol}$$

PRACTICE   *Answers in Appendix E*

1. At STP, what is the volume of 7.08 mol of nitrogen gas?

2. A sample of hydrogen gas occupies 14.1 L at STP. How many moles of the gas are present?

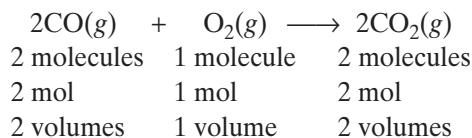# Gas Stoichiometry

You can apply the discoveries of Gay-Lussac and Avogadro to calculate the stoichiometry of reactions involving gases. For gaseous reactants or products, the coefficients in chemical equations not only indicate molar amounts and mole ratios but also reveal volume ratios, assuming conditions remain the same. For example, consider the reaction of carbon monoxide with oxygen to give carbon dioxide.

$$2CO(g) \quad + \quad O_2(g) \quad \longrightarrow \quad 2CO_2(g)$$

2 molecules	1 molecule	2 molecules
2 mol	1 mol	2 mol
2 volumes	1 volume	2 volumes

The possible volume ratios can be expressed in the following ways.

a. $\dfrac{2 \text{ volumes CO}}{1 \text{ volume O}_2}$ or $\dfrac{1 \text{ volume O}_2}{2 \text{ volumes CO}}$

b. $\dfrac{2 \text{ volumes CO}}{2 \text{ volumes CO}_2}$ or $\dfrac{2 \text{ volumes CO}_2}{2 \text{ volumes CO}}$

c. $\dfrac{1 \text{ volume O}_2}{2 \text{ volumes CO}_2}$ or $\dfrac{2 \text{ volumes CO}_2}{1 \text{ volume O}_2}$

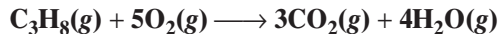**Propane, C_3H_8, is a gas that is sometimes used as a fuel for cooking and heating. The complete combustion of propane occurs according to the following balanced equation.**

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

(a) What will be the volume, in liters, of oxygen required for the complete combustion of 0.350 L of propane? (b) What will be the volume of carbon dioxide produced in the reaction? Assume that all volume measurements are made at the same temperature and pressure.

SOLUTION

1 ANALYZE

Given: balanced chemical equation
V of propane = 0.350 L

Unknown: **a.** V of O_2 in L
b. V of CO_2 in L

2 PLAN

a. V of $C_3H_8 \longrightarrow V$ of O_2

b. V of $C_3H_8 \longrightarrow V$ of CO_2

All volumes are to be compared at the same temperature and pressure. Therefore, volume ratios can be used like mole ratios to find the unknowns.

3 COMPUTE

a. $0.350 \text{ L } C_3H_8 \times \dfrac{5 \text{ L } O_2}{1 \text{ L } C_3H_8} = 1.75 \text{ L } O_2$

b. $0.350 \text{ L } C_3H_8 \times \dfrac{3 \text{ L } CO_2}{1 \text{ L } C_3H_8} = 1.05 \text{ L } CO_2$

4 EVALUATE

Each result is correctly given to three significant figures. The answers are reasonably close to estimated values of 2, calculated as 0.4×5, and 1.2, calculated as 0.4×3, respectively.

PRACTICE   *Answers in Appendix E*

1. Assuming all volume measurements are made at the same temperature and pressure, what volume of hydrogen gas is needed to react completely with 4.55 L of oxygen gas to produce water vapor?

2. What volume of oxygen gas is needed to react completely with 0.626 L of carbon monoxide gas, CO, to form gaseous carbon dioxide? Assume all volume measurements are made at the same temperature and pressure.

3. Nitric acid can be produced by the reaction of gaseous nitrogen dioxide with water, according to the following balanced chemical equation.

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(l) + NO(g)$$
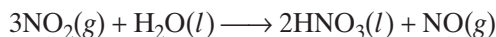
If 708 L of NO_2 gas react with water, what volume of NO gas will be produced? Assume the gases are measured under the same conditions before and after the reaction.

extension
Go to **go.hrw.com** for more practice problems that ask you to use mole ratios in stoichiometry.

Keyword: HC6GASX

The Ideal Gas Law

You have learned about equations describing the relationships between two or three of the four variables—pressure, volume, temperature, and moles—needed to describe a gas sample. All the gas laws you have learned thus far can be combined into a single equation. *The **ideal gas law** is the mathematical relationship among pressure, volume, temperature, and the number of moles of a gas.* It is the equation of state for an ideal gas, because the state of a gas can be defined by its pressure, volume, temperature, and number of moles. It is stated as shown below, where R is a constant.

$$PV = nRT$$

The ideal gas law reduces to Boyle's law, Charles's law, Gay-Lussac's law, or Avogadro's law when the appropriate variables are held constant.

The number of molecules or moles present will always affect at least one of the other three quantities. The collision rate of molecules per unit area of container wall depends on the number of molecules present. If the number of molecules is increased for a sample at constant volume and temperature, the collision rate increases. Therefore, the pressure increases, as shown by the model in **Figure 13a.** Consider what would happen if the pressure and temperature were kept constant while the number of molecules increased. According to Avogadro's law, the volume would increase. As **Figure 13b** shows, an increase in volume keeps the pressure constant at constant temperature. Increasing the volume keeps the collision rate per unit of wall area constant.
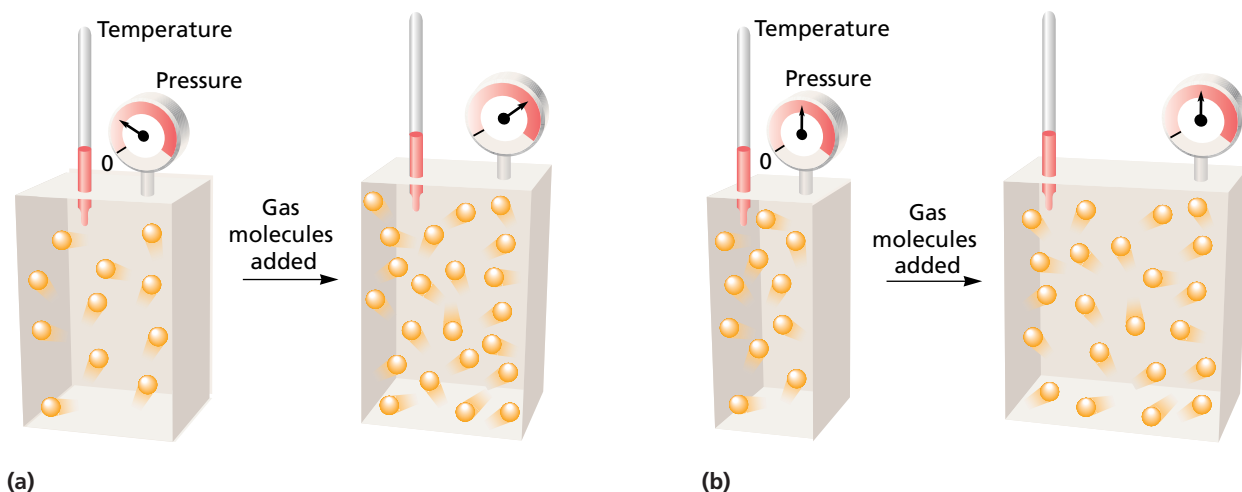
(a)
(b)

FIGURE 13   (a) When volume and temperature are constant, gas pressure increases as the number of molecules increases. (b) When pressure and temperature are constant, gas volume increases as the number of molecules increases.

The Ideal Gas Constant

In the equation representing the ideal gas law, *the constant R is known as the* **ideal gas constant.** Its value depends on the units chosen for pressure, volume, and temperature. Measured values of *P, V, T,* and *n* for a gas at near-ideal conditions can be used to calculate *R*. Recall that the volume of one mole of an ideal gas at STP (1 atm and 273.15 K) is 22.414 10 L. Substituting these values and solving the ideal gas law equation for *R* gives the following.

$$R = \frac{PV}{nT} = \frac{(1 \text{ atm})(22.414\ 10 \text{ L})}{(1 \text{ mol})(273.15 \text{ K})} = 0.082\ 058\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

This calculated value of *R* is usually rounded to 0.0821 L • atm/(mol•K). Use this value in ideal gas law calculations when the volume is in liters, the pressure is in atmospheres, and the temperature is in kelvins. See **Table 2** for the value of *R* when other units for *n, P, V,* and *T* are used.

Finding P, V, T, or n from the Ideal Gas Law

The ideal gas law can be applied to determine the existing conditions of a gas sample when three of the four variables, *P, V, T,* and *n,* are known. It can also be used to calculate the molar mass or density of a gas sample.

Be sure to match the units of the known quantities and the units of *R.* In this book, you will be using *R* = 0.0821 L•atm/(mol•K). Your first step in solving any ideal gas law problem should be to check the known values to be sure you are working with the correct units. If necessary, you must convert volumes to liters, pressures to atmospheres, temperatures to kelvins, and masses to numbers of moles before using the ideal gas law.

TABLE 2 *Numerical Values of Gas Constant, R*

Unit of *R*	Numerical value of *R*	Unit of *P*	Unit of *V*	Unit of *T*	Unit of *n*
$\frac{\text{L}\cdot\text{mm Hg}}{\text{mol}\cdot\text{K}}$	62.4	mm Hg	L	K	mol
$\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$	0.0821	atm	L	K	mol
$\frac{\text{J}}{\text{mol}\cdot\text{K}}$ *	8.314	Pa	m^3	K	mol
$\frac{\text{L}\cdot\text{kPa}}{\text{mol}\cdot\text{K}}$	8.314	kPa	L	K	mol

Note: 1 L•atm = 101.325 J; 1 J = 1 Pa•m^3
* SI units

What is the pressure in atmospheres exerted by a 0.500 mol sample of nitrogen gas in a 10.0 L container at 298 K?

SOLUTION

1 ANALYZE

Given: V of $N_2 = 10.0$ L
n of $N_2 = 0.500$ mol
T of $N_2 = 298$ K

Unknown: P of N_2 in atm

2 PLAN

$$n, V, T \longrightarrow P$$

The gas sample undergoes no change in conditions. Therefore, the ideal gas law can be rearranged and used to find the pressure as follows.

$$P = \frac{nRT}{V}$$

3 COMPUTE

$$P = \frac{(0.500 \text{ mol})\left(\dfrac{0.0821 \text{ L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298 \text{ K})}{10.0 \text{ L}} = 1.22 \text{ atm}$$

4 EVALUATE

All units cancel correctly to give the result in atmospheres. The answer is properly limited to three significant figures. It is also close to an estimated value of 1.5, computed as $(0.5 \times 0.1 \times 300)/10$.

PRACTICE   *Answers in Appendix E*

1. What pressure, in atmospheres, is exerted by 0.325 mol of hydrogen gas in a 4.08 L container at 35°C?

2. A gas sample occupies 8.77 L at 20°C. What is the pressure, in atmospheres, given that there are 1.45 mol of gas in the sample?

**SECTION REVIEW**

1. State Avogadro's law, and explain its significance.

2. What volume (in milliliters) at STP will be occupied by 0.0035 mol of methane, CH_4?

3. State the ideal gas law equation, and tell what each term means.

4. What would be the units for R if P is in pascals, T is in kelvins, V is in liters, and n is in moles?

5. A 4.44 L container holds 15.4 g of oxygen at 22.55°C. What is the pressure?

6. A tank of hydrogen gas has a volume of 22.9 L and holds 14.0 mol of the gas at 12°C. What is the pressure of the gas in atmospheres?

Critical Thinking

7. **ANALYZING DATA** Nitrous oxide is sometimes used as a source of oxygen gas:

$$2N_2O(g) \longrightarrow 2N_2(g) + O_2(g)$$
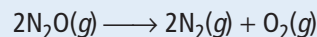
What volume of each product will be formed from 2.22 L N_2O? At STP, what is the density of the product gases when they are mixed?

Diffusion and Effusion

OBJECTIVES

- Describe the process of diffusion.

- State Graham's law of effusion.

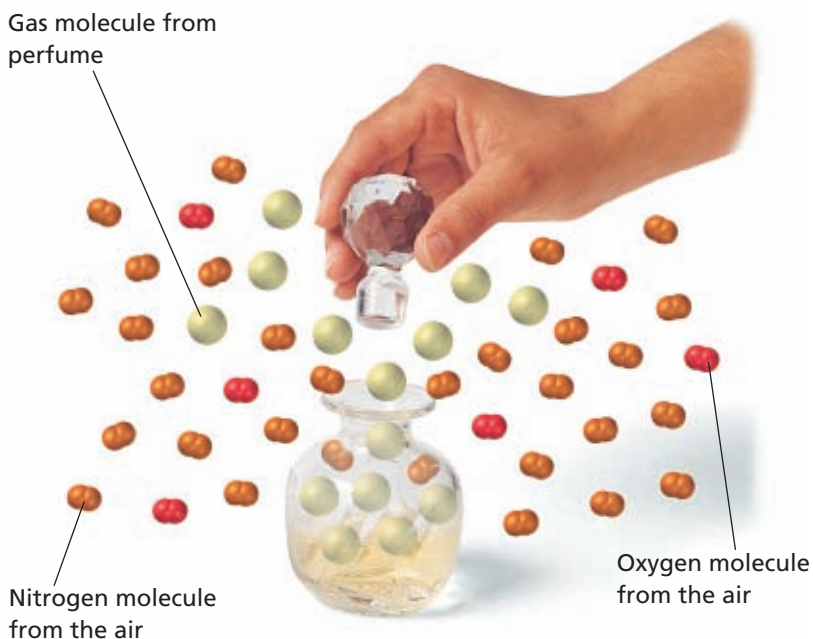- State the relationship between the average molecular velocities of two gases and their molar masses.

The constant motion of gas molecules causes them to spread out to fill any container in which they are placed. The gradual mixing of two or more gases due to their spontaneous, random motion is known as *diffusion,* illustrated in **Figure 14.** *Effusion* is the process whereby the molecules of a gas confined in a container randomly pass through a tiny opening in the container. In this section, you will learn how effusion can be used to estimate the molar mass of a gas.

Graham's Law of Effusion

The rates of effusion and diffusion depend on the relative velocities of gas molecules. The velocity of a gas varies inversely with the square root of its molar mass. Lighter molecules move faster than heavier molecules at the same temperature.

Recall that the average kinetic energy of the molecules in any gas depends only on the temperature and equals $\frac{1}{2}mv^2$. For two different gases, A and B, at the same temperature, the following relationship is true.

$$\frac{1}{2}M_A v_A{}^2 = \frac{1}{2}M_B v_B{}^2$$

Gas molecule from perfume

Oxygen molecule from the air

Nitrogen molecule from the air

FIGURE 14 When a bottle of perfume is opened, some of its molecules diffuse into the air and mix with the molecules in the air. At the same time, molecules from the air, such as nitrogen and oxygen, diffuse into the bottle and mix with the gaseous scent molecules.

From the equation relating the kinetic energy of two different gases at the same conditions, one can derive an equation relating the rates of effusion of two gases with their molecular mass. This equation is shown below.

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

In the mid-1800s, the Scottish chemist Thomas Graham studied the effusion and diffusion of gases. The above equation is a mathematical statement of some of Graham's discoveries. It describes the rates of effusion. It can also be used to find the molar mass of an unknown gas. **Graham's law of effusion** *states that the rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.*

QuickLAB

Wear safety goggles and an apron.

Diffusion

Question

Do different gases diffuse at different rates?

Procedure

Record all of your results in a data table.

1. Outdoors or in a room separate from the one in which you will carry out the rest of the investigation, pour approximately 10 mL of the household ammonia into one of the 250 mL beakers, and cover it with a watch glass. Pour roughly the same amount of perfume or cologne into the second beaker. Cover it with a watch glass also.

2. Take the two samples you just prepared into a large, draft-free room. Place the samples about 12 to 15 ft apart and at the same height. Position someone as the observer midway between the two beakers. Remove both watch glass covers at the same time.

3. Note whether the observer smells the ammonia or the perfume first. Record how long this takes. Also, record how long it takes the vapor of the other substance to reach the observer. Air the room after you have finished.

Discussion

1. What do the times that the two vapors took to reach the observer show about the two gases?

2. What factors other than molecular mass (which determines diffusion rate) could affect how quickly the observer smells each vapor?

Materials

- **household ammonia**
- **perfume or cologne**
- **two 250 mL beakers**
- **two watch glasses**
- **10 mL graduated cylinder**
- **clock or watch with second hand**

Compare the rates of effusion of hydrogen and oxygen at the same temperature and pressure.

SOLUTION

1 ANALYZE   **Given:** identities of two gases, H_2 and O_2
Unknown: relative rates of effusion

2 PLAN   molar mass ratio $\longrightarrow$ ratio of rates of effusion

The ratio of the rates of effusion of two gases at the same temperature and pressure can be found from Graham's law.

$$\frac{\text{rate of effusion of A}}{\text{rate of effusion of B}} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

3 COMPUTE $\dfrac{\text{rate of effusion of } H_2}{\text{rate of effusion of } O_2} = \dfrac{\sqrt{M_{O_2}}}{\sqrt{M_{H_2}}} = \dfrac{\sqrt{32.00 \text{ g/mol}}}{\sqrt{2.02 \text{ g/mol}}} = \sqrt{\dfrac{32.00 \text{ g/mol}}{2.02 \text{ g/mol}}} = 3.98$

Hydrogen effuses 3.98 times faster than oxygen.

4 EVALUATE   The result is correctly reported to three significant figures. It is also approximately equivalent to an estimated value of 4, calculated as $\sqrt{32} / \sqrt{2}$.

PRACTICE   *Answers in Appendix E*

1. Compare the rate of effusion of carbon dioxide with that of hydrogen chloride at the same temperature and pressure.

2. A sample of hydrogen effuses through a porous container about 9 times faster than an unknown gas. Estimate the molar mass of the unknown gas.

3. If a molecule of neon gas travels at an average of 400. m/s at a given temperature, estimate the average speed of a molecule of butane gas, C_4H_{10}, at the same temperature.

extension
Go to **go.hrw.com** for additional problems that ask you to compare rates of effusion.

Keyword: HC6GASX

SECTION REVIEW

1. Compare diffusion with effusion.

2. State Graham's law of effusion.

3. Estimate the molar mass of a gas that effuses at 1.6 times the effusion rate of carbon dioxide.

4. Determine the molecular mass ratio of two gases whose rates of effusion have a ratio of 16:1.

5. List the following gases in order of increasing average molecular velocity at 25°C: H_2O, He, HCl, BrF, and NO_2.

Critical Thinking

6. **ANALYZING INFORMATION** An unknown gas effuses at one-half the speed of oxygen. What is the molar mass of the unknown? The gas is known to be either HBr or HI. Which gas is it?

CHAPTER HIGHLIGHTS

Gases and Pressure

Vocabulary

pressure
newton
barometer
millimeters of mercury
atmosphere of pressure
pascal
partial pressure
Dalton's law of partial pressures

- The kinetic-molecular theory of gases describes an ideal gas. The behavior of most gases is nearly ideal except at very high pressures and low temperatures.
- A barometer measures atmospheric pressure.
- Dalton's law of partial pressure states that in a mixture of unreacting gases, the total pressure equals the sum of the partial pressures of each gas.

The Gas Laws

Vocabulary

Boyle's law
absolute zero
Charles's law
Gay-Lussac's law
combined gas law

- Boyle's law states the inverse relationship between the volume and the pressure of a gas:
$$PV = k$$
- Charles's law illustrates the direct relationship between a gas's volume and its temperature in kelvins:
$$V = kT$$
- Gay-Lussac's law represents the direct relationship between a gas's pressure and its temperature in kelvins:
$$P = kT$$
- The combined gas law, as its name implies, combines the previous relationships into the following mathematical expression:
$$\frac{PV}{T} = k$$

Gas Volumes and the Ideal Gas Law

Vocabulary

Gay-Lussac's law of combining volumes of gases
Avogadro's law
standard molar volume of a gas
ideal gas law
ideal gas constant

- Gay-Lussac's law of combining volumes states that the volumes of reacting gases and their products at the same temperature and pressure can be expressed as ratios of whole numbers.
- Avogadro's law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
- The volume occupied by one mole of an ideal gas at STP is called the *standard molar volume,* which is 22.414 10 L.
- Charles's law, Boyle's law, and Avogadro's law can be combined to create the ideal gas law:
$$PV = nRT$$

Diffusion and Effusion

Vocabulary

Graham's law of effusion

- Gases *diffuse,* or become more spread out, due to their constant random molecular motion.
- Graham's law of effusion states that the relative rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.

CHAPTER REVIEW

Gases and Pressure

SECTION 1 REVIEW

1. State the assumptions that the kinetic-molecular theory makes about the characteristics of gas particles.
2. What is an ideal gas?
3. a. Why does a gas in a closed container exert pressure?
   b. What is the relationship between the area a force is applied to and the resulting pressure?
4. a. Why does a column of mercury in a tube that is inverted in a dish of mercury have a height of about 760 mm at sea level?
   b. The density of water is approximately 1/13.5 the density of mercury. What height would be maintained by a column of water inverted in a dish of water at sea level?
   c. What accounts for the difference in the heights of the mercury and water columns?
5. a. Identify three units used to express pressure.
   b. Convert one atmosphere to millimeters of mercury.
   c. What is a pascal?
   d. What is the SI equivalent of one standard atmosphere of pressure?
6. a. Explain what is meant by the partial pressure of each gas within a mixture of gases.
   b. How do the partial pressures of gases in a mixture affect each other?

PRACTICE PROBLEMS

7. If the atmosphere can support a column of mercury 760 mm high at sea level, what height of a hypothetical liquid whose density is 1.40 times the density of mercury could be supported?
8. Convert each of the following into a pressure reading expressed in torrs.
   a. 1.25 atm
   b. 2.48×10^{-3} atm
   c. 4.75×10^4 atm
   d. 7.60×10^6 atm

9. Convert each of the following into the unit specified.
   a. 125 mm Hg into atmospheres
   b. 3.20 atm into pascals
   c. 5.38 kPa into millimeters of mercury
10. Three of the primary components of air are carbon dioxide, nitrogen, and oxygen. In a sample containing a mixture of only these gases at exactly 1 atm, the partial pressures of carbon dioxide and nitrogen are given as $P_{CO_2} = 0.285$ torr and $P_{N_2} = 593.525$ torr. What is the partial pressure of oxygen?
11. A sample of gas is collected over water at a temperature of 35.0°C when the barometric pressure reading is 742.0 torr. What is the partial pressure of the dry gas?

The Gas Laws

SECTION 2 REVIEW

12. How are the volume and pressure of a gas at constant temperature related?
13. Explain why pressure increases as a gas is compressed into a smaller volume.
14. How are the absolute temperature and volume of a gas at constant pressure related?
15. How are the pressure and absolute temperature of a gas at constant volume related?
16. Explain Gay-Lussac's law in terms of the kinetic-molecular theory.
17. State the combined gas law.

PRACTICE PROBLEMS

18. Use Boyle's law to solve for the missing value in each of the following:
    a. $P_1 = 350.0$ torr, $V_1 = 200.0$ mL, $P_2 = 700.0$ torr, $V_2 = ?$
    b. $V_1 = 2.4 \times 10^5$ L, $P_2 = 180$ mm Hg, $V_2 = 1.8 \times 10^3$ L, $P_1 = ?$
19. Use Charles's law to solve for the missing value in each of the following:
    a. $V_1 = 80.0$ mL, $T_1 = 27$°C, $T_2 = 77$°C, $V_2 = ?$
    b. $V_1 = 125$ L, $V_2 = 85.0$ L, $T_2 = 127$°C, $T_1 = ?$
    c. $T_1 = -33$°C, $V_2 = 54.0$ mL, $T_2 = 160.0$°C, $V_1 = ?$

20. A sample of air has a volume of 140.0 mL at 67°C. At what temperature would its volume be 50.0 mL at constant pressure?

21. The pressure exerted on a 240.0 mL sample of hydrogen gas at constant temperature is increased from 0.428 atm to 0.724 atm. What will the final volume of the sample be?

22. A sample of hydrogen at 47°C exerts a pressure of 0.329 atm. The gas is heated to 77°C at constant volume. What will its new pressure be?

23. A sample of gas at 47°C and 1.03 atm occupies a volume of 2.20 L. What volume would this gas occupy at 107°C and 0.789 atm?

24. The pressure on a gas at –73°C is doubled, but its volume is held constant. What will the final temperature be in degrees Celsius?

25. A flask containing 155 cm^3 of hydrogen was collected under a pressure of 22.5 kPa. What pressure would have been required for the volume of the gas to have been 90.0 cm^3, assuming the same temperature?

26. A gas has a volume of 450.0 mL. If the temperature is held constant, what volume would the gas occupy if the pressure were
a. doubled? (Hint: Express P_2 in terms of P_1.)
b. reduced to one-fourth of its original value?

27. A sample of oxygen that occupies 1.00×10^6 mL at 575 mm Hg is subjected to a pressure of 1.25 atm. What will the final volume of the sample be if the temperature is held constant?

28. To what temperature must a sample of nitrogen at 27°C and 0.625 atm be taken so that its pressure becomes 1.125 atm at constant volume?

29. A gas has a volume of 1.75 L at –23°C and 150.0 kPa. At what temperature would the gas occupy 1.30 L at 210.0 kPa?

30. A gas at 7.75×10^4 Pa and 17°C occupies a volume of 850.0 cm^3. At what temperature, in degrees Celsius, would the gas occupy 720.0 cm^3 at 8.10×10^4 Pa?

31. A meteorological balloon contains 250.0 L He at 22°C and 740.0 mm Hg. If the volume of the balloon can vary according to external conditions, what volume would it occupy at an altitude at which the temperature is –52°C and the pressure is 0.750 atm?

32. The balloon in the previous problem will burst if its volume reaches 400.0 L. Given the initial conditions specified in that problem, determine at what temperature, in degrees Celsius, the balloon will burst if its pressure at that bursting point is 0.475 atm.

33. The normal respiratory rate for a human being is 15.0 breaths per minute. The average volume of air for each breath is 505 cm^3 at 20.0°C and 9.95×10^4 Pa. What is the volume of air at STP that an individual breathes in one day? Give your answer in cubic meters.

Gas Volumes and the Ideal Gas Law

SECTION 3 REVIEW

34. a. What are the restrictions on the use of Gay-Lussac's law of combining volumes?
b. At the same temperature and pressure, what is the relationship between the volume of a gas and the number of molecules present?

35. a. In a balanced chemical equation, what is the relationship between the molar ratios and the volume ratios of gaseous reactants and products?
b. What restriction applies to the use of the volume ratios in solving stoichiometry problems?

36. According to Avogadro,
a. what is the relationship between gas volume and number of moles at constant temperature and pressure?
b. what is the mathematical expression denoting this relationship?

37. What is the relationship between the number of molecules and the mass of 22.4 L of different gases at STP?

38. a. In what situations is the ideal gas law most suitable for calculations?
b. When using this law, why do you have to pay particular attention to units?

39. a. Write the equation for the ideal gas law.
b. What relationship is expressed in the ideal gas law?

PRACTICE PROBLEMS

40. Suppose a 5.00 L sample of O_2 at a given temperature and pressure contains 1.08×10^{23} molecules. How many molecules would be contained in each of the following at the same temperature and pressure?
   a. 5.0 L H_2
   b. 5.0 L CO_2
   c. 10.0 L NH_3

41. How many moles are contained in each of the following at STP?
   a. 22.4 L N_2
   b. 5.60 L Cl_2
   c. 0.125 L Ne
   d. 70.0 mL NH_3

42. Find the mass, in grams, of each of the following at STP.
   a. 11.2 L H_2
   b. 2.80 L CO_2
   c. 15.0 mL SO_2
   d. 3.40 cm^3 F_2

43. Find the volume, in liters, of each of the following at STP.
   a. 8.00 g O_2
   b. 3.50 g CO
   c. 0.0170 g H_2S
   d. 2.25×10^5 kg NH_3

44. Acetylene gas, C_2H_2, undergoes combustion to produce carbon dioxide and water vapor. If 75.0 L CO_2 is produced,
   a. how many liters of C_2H_2 are required?
   b. what volume of H_2O vapor is produced?
   c. what volume of O_2 is required?

45. Assume that 5.60 L H_2 at STP reacts with excess CuO according to the following equation:

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

   Make sure the equation is balanced before beginning your calculations.
   a. How many moles of H_2 react?
   b. How many moles of Cu are produced?
   c. How many grams of Cu are produced?

46. If 29.0 L of methane, CH_4, undergoes complete combustion at 0.961 atm and 140°C, how many liters of each product would be present at the same temperature and pressure?

47. If air is 20.9% oxygen by volume,
   a. how many liters of air are needed for complete combustion of 25.0 L of octane vapor, C_8H_{18}?
   b. what volume of each product is produced?

48. Methanol, CH_3OH, is made by causing carbon monoxide and hydrogen gases to react at high temperature and pressure. If 4.50×10^2 mL CO and 8.25×10^2 mL H_2 are mixed,
   a. which reactant is present in excess?
   b. how much of that reactant remains after the reaction?
   c. what volume of CH_3OH is produced, assuming the same pressure?

49. Calculate the pressure, in atmospheres, exerted by each of the following:
   a. 2.50 L HF containing 1.35 mol at 320.0 K
   b. 4.75 L NO_2 containing 0.86 mol at 300.0 K
   c. 5.50×10^4 mL CO_2 containing 2.15 mol at 57°C

50. Calculate the volume, in liters, occupied by each of the following:
   a. 2.00 mol H_2 at 300.0 K and 1.25 atm
   b. 0.425 mol NH_3 at 37°C and 0.724 atm
   c. 4.00 g O_2 at 57°C and 0.888 atm

51. Determine the number of moles of gas contained in each of the following:
   a. 1.25 L at 250.0 K and 1.06 atm
   b. 0.80 L at 27°C and 0.925 atm
   c. 7.50×10^2 mL at –50.0°C and 0.921 atm

52. Find the mass of each of the following.
   a. 5.60 L O_2 at 1.75 atm and 250.0 K
   b. 3.50 L NH_3 at 0.921 atm and 27°C
   c. 125 mL SO_2 at 0.822 atm and –5°C

Diffusion and Effusion

SECTION 4 REVIEW

53. Describe in your own words the process of diffusion.

54. At a given temperature, what factor determines the rates at which different molecules undergo diffusion and effusion?

55. Ammonia, NH_3, and alcohol, C_2H_6O, are released together across a room. Which will you smell first?

PRACTICE PROBLEMS

56. Quantitatively compare the rates of effusion for the following pairs of gases at the same temperature and pressure:
a. hydrogen and nitrogen
b. fluorine and chlorine

57. What is the ratio of the average velocity of hydrogen molecules to that of neon atoms at the same temperature and pressure?

58. At a certain temperature and pressure, chlorine molecules have an average velocity of 324 m/s. What is the average velocity of sulfur dioxide molecules under the same conditions?

MIXED REVIEW

59. A mixture of three gases, A, B, and C, is at a total pressure of 6.11 atm. The partial pressure of gas A is 1.68 atm; that of gas B is 3.89 atm. What is the partial pressure of gas C?

60. A child receives a balloon filled with 2.30 L of helium from a vendor at an amusement park. The temperature outside is 311 K. What will the volume of the balloon be when the child brings it home to an air-conditioned house at 295 K? Assume that the pressure stays the same.

61. A sample of argon gas occupies a volume of 295 mL at 36°C. What volume will the gas occupy at 55°C, assuming constant pressure?

62. A sample of carbon dioxide gas occupies 638 mL at 0.893 atm and 12°C. What will the pressure be at a volume of 881 mL and a temperature of 18°C?

63. At 84°C, a gas in a container exerts a pressure of 0.503 atm. Assuming the size of the container has not changed, at what temperature in Celsius degrees would the pressure be 1.20 atm?

64. A weather balloon at Earth's surface has a volume of 4.00 L at 304 K and 755 mm Hg. If the balloon is released and the volume reaches 4.08 L at 728 mm Hg, what is the temperature?

65. A gas has a pressure of 4.62 atm when its volume is 2.33 L. If the temperature remains constant, what will the pressure be when the volume is changed to 1.03 L? Express the final pressure in torrs.

66. At a deep-sea station that is 200. m below the surface of the Pacific Ocean, workers live in a highly pressurized environment. How many liters of gas at STP must be compressed on the surface to fill the underwater environment with 2.00×10^7 L of gas at 20.0 atm? Assume that temperature remains constant.

67. An unknown gas effuses at 0.850 times the effusion rate of nitrogen dioxide, NO_2. Estimate the molar mass of the unknown gas.

68. A container holds 265 mL of chlorine gas, Cl_2. If the gas sample is at STP, what is its mass?

69. Suppose that 3.11 mol of carbon dioxide is at a pressure of 0.820 atm and a temperature of 39°C. What is the volume of the sample, in liters?

70. Compare the rates of diffusion of carbon monoxide, CO, and sulfur trioxide, SO_3.

71. A gas sample that has a mass of 0.993 g occupies 0.570 L. Given that the temperature is 281 K and the pressure is 1.44 atm, what is the molar mass of the gas?

72. How many moles of helium gas would it take to fill a balloon with a volume of 1000.0 cm^3 when the temperature is 32°C and the atmospheric pressure is 752 mm Hg?

73. A gas sample is collected at 16°C and 0.982 atm. If the sample has a mass of 7.40 g and a volume of 3.96 L, find the volume of the gas at STP and the molar mass.

CRITICAL THINKING

74. Applying Models
a. Why do we say the graph in **Figure 7** illustrates an inverse relationship?
b. Why do we say the data plotted in **Figure 9** indicates a direct relationship?

75. Inferring Conclusions   If all gases behaved as ideal gases under all conditions of temperature and pressure, solid or liquid forms of these substances would not exist. Explain.

76. Relating Ideas Pressure is defined as force per unit area. Yet Torricelli found that the diameter of the barometer dish and the surface area of contact between the mercury in the tube and in the dish did not affect the height of mercury that was supported. Explain this seemingly inconsistent observation in view of the relationship between pressure and surface area.

77. Evaluating Methods In solving a problem, what types of conditions involving temperature, pressure, volume, or number of moles would allow you to use
a. the combined gas law?
b. the ideal gas law?

78. Evaluating Ideas Gay-Lussac's law of combining volumes holds true for relative volumes at any proportionate size. Use Avogadro's law to explain why this proportionality exists.

79. Interpreting Graphics The graph below shows velocity distribution curves for the same gas under two different conditions, A and B. Compare the behavior of the gas under conditions A and B in relation to each of the following:
a. temperature
b. average kinetic energy
c. average molecular velocity
d. gas volume
e. gas pressure
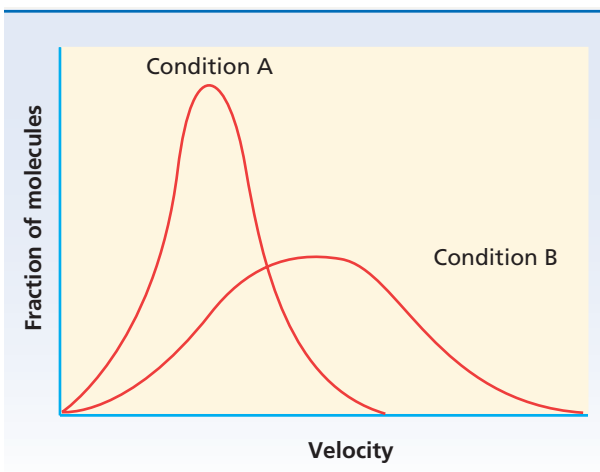
80. Interpreting Concepts The diagrams below represent equal volumes of four different gases.
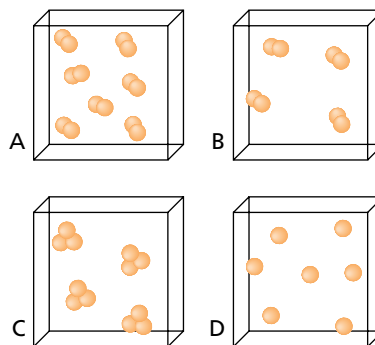
Use the diagrams to answer the following questions:
a. Are these gases at the same temperature and pressure? How do you know?
b. If the molar mass of gas B is 38 g/mol and that of gas C is 46 g/mol, which gas sample is denser?
c. To make the densities of gas samples B and C equal, which gas should expand in volume?
d. If the densities of gas samples A and C are equal, what is the relationship between their molar masses?

RESEARCH & WRITING

81. Design and conduct a meteorological study to examine the interrelationships between barometric pressure, temperature, humidity, and other weather variables. Prepare a report explaining your results.

82. Conduct library research on attempts made to approach absolute zero and on the interesting properties that materials exhibit near that temperature. Write a report on your findings.

83. How do scuba divers use the laws and principles that describe the behavior of gases to their advantage? What precautions do they take to prevent the bends?

84. Explain the processes involved in the liquefaction of gases. Name some substances that are gases under normal room conditions and that are typically used in the liquid form. Explain why this is so.

85. Write a summary describing how Gay-Lussac's work on combining volumes relates to Avogadro's study of gases. Explain how certain conclusions about gases followed logically from consideration of the work of both scientists.

USING THE HANDBOOK

86. Review the melting point data in the properties tables for each group of the *Elements Handbook*. What elements on the periodic table exist as gases at room temperature?
87. Review in the *Elements Handbook* the listing of the top 10 chemicals produced in the United States. Which of the top 10 chemicals are gases?
88. Most elements from Groups 1, 2, and 13 will react with water, acids, or bases to produce hydrogen gas. Review the common reactions information in the *Elements Handbook* and answer the following questions:
  a. What is the equation for the reaction of barium with water?
  b. What is the equation for the reaction between cesium and hydrochloric acid?
  c. What is the equation for the reaction of gallium with hydrofluoric acid?
  d. What mass of barium would be needed to react with excess water to produce 10.1 L H_2 at STP?
  e. What masses of cesium and hydrochloric acid would be required to produce 10.1 L H_2 at STP?
89. Group 1 metals react with oxygen to produce oxides, peroxides, or superoxides. Review the equations for these common reactions in the *Elements Handbook,* and answer the following:
  a. How do oxides, peroxides, and superoxides differ?
  b. What mass of product will be formed from a reaction of 5.00 L O_2 with excess sodium? The reaction occurs at 27°C and 1 atm.

ALTERNATIVE ASSESSMENT

90. The air pressure of car tires should be checked regularly for safety reasons and for prevention of uneven tire wear. Find out the units of measurement on a typical tire gauge, and determine how gauge pressure relates to atmospheric pressure.
91. During a typical day, record every instance in which you encounter the diffusion or effusion of gases (for example, when smelling perfume).
92. Performance Qualitatively compare the molecular masses of various gases by noting how long it takes you to smell them from a fixed distance. Work only with materials that are not dangerous, such as flavor extracts, fruit peels, and onions.
93. Performance Design an experiment to gather data to verify the ideal gas law. If your teacher approves of your plan, carry it out. Illustrate your data with a graph, and determine if the data are consistent with the ideal gas law.

extension

Graphing Calculator
Using the Ideal Gas Law
Go to **go.hrw.com** for a graphing calculator exercise that asks you to use the ideal gas law.

Keyword: HC6GASX

Math Tutor ALGEBRAIC REARRANGEMENTS OF GAS LAWS

When you solve problems in chemistry, it's usually a bad idea to just start entering numbers into a calculator. Instead, doing a little pencil-and-paper work beforehand will help you eliminate errors. When using the gas laws, you do not need to memorize all of the equations because they are easily derived from the equation for the combined gas law, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$. Study the table below. In each of Boyle's, Charles's, and Gay-Lussac's laws, one of the quantities—T, P, or V—does not change. By simply eliminating that factor from the equation, you obtain the equation for one particular gas law.

Gas law	Held constant	Cancellation	Result
Combined gas law	none	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Boyle's law	temperature	$\frac{P_1V_1}{\cancel{T_1}} = \frac{P_2V_2}{\cancel{T_2}}$	$P_1V_1 = P_2V_2$
Charles's law	pressure	$\frac{\cancel{P_1}V_1}{T_1} = \frac{\cancel{P_2}V_2}{T_2}$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay-Lussac's law	volume	$\frac{P_1\cancel{V_1}}{T_1} = \frac{P_2\cancel{V_2}}{T_2}$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

The conditions stated in the problem should make clear which factors change and which are held constant. This information will tell you which law's equation you need to use.

SAMPLE

A cylinder of nitrogen gas has a volume of 35.00 L at a pressure of 11.50 atm. What pressure will the nitrogen have if the contents of the cylinder are allowed to flow into a sealed reaction chamber whose volume is 140.0 L, and if the temperature remains constant?

Start with the combined gas law, and cancel the temperature, which does not change.

$$\frac{P_1V_1}{\cancel{T_1}} = \frac{P_2V_2}{\cancel{T_2}}; P_1V_1 = P_2V_2$$

You want to know the new pressure in the chamber, so solve for P_2.

$$\frac{P_1V_1}{V_2} = \frac{P_2\cancel{V_2}}{\cancel{V_2}}; \frac{P_1V_1}{V_2} = P_2$$

The resulting equation to use in solving the problem is $P_2 = \frac{P_1V_1}{V_2} = \frac{(11.50\ \text{atm})(35.00\ \text{L})}{140.0\ \text{L}}$
$= 2.875$ atm.

PRACTICE PROBLEMS

1. A sample of gas has a pressure P_1 at a temperature T_1. Write the equation that you would use to find the temperature, T_2 at which the gas has a pressure of P_2.

2. An ideal gas occupies a volume of 785 mL at a pressure of 0.879 atm. What volume will the gas occupy at a pressure of 0.994 atm?

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Pressure can be measured in
   A. grams.   **C.** pascals.
   B. meters.   **D.** liters.

2. A sample of oxygen gas has a volume of 150 mL when its pressure is 0.923 atm. If the pressure is increased to 0.987 atm and the temperature remains constant, what will the new volume be?
   A. 140 mL   **C.** 200 mL
   B. 160 mL   **D.** 240 mL

3. What is the pressure exerted by a 0.500 mol sample of nitrogen in a 10.0 L container at 20°C?
   A. 1.2 kPa   **C.** 0.10 kPa
   B. 10 kPa   **D.** 120 kPa

4. A sample of gas in a closed container at a temperature of 100.0°C and 3.0 atm is heated to 300.0°C. What is the pressure of the gas at the higher temperature?
   A. 35 atm   **C.** 59 atm
   B. 4.6 atm   **D.** 9.0 atm

5. An unknown gas effuses twice as fast as CH_4. What is the molar mass of the gas?
   A. 64 g/mol   **C.** 8 g/mol
   B. 32 g/mol   **D.** 4 g/mol

6. If 3 L N_2 and 3 L H_2 are mixed and react according to the equation below, how many liters of unreacted gas remain? Assume temperature and pressure remain constant.
$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$
   A. 4 L   **C.** 2 L
   B. 3 L   **D.** 1 L

7. Avogadro's law states that
   A. equal numbers of moles of gases at the same conditions occupy equal volumes, regardless of the identity of the gases.
   B. at constant pressure, gas volume is directly proportional to absolute temperature.
   C. the volume of a gas is inversely proportional to its amount in moles.
   D. at constant temperature, gas volume is inversely proportional to pressure.

SHORT ANSWER

8. Give a molecular explanation for the observation that the pressure of a gas increases when the gas volume is decreased.

9. The graph below shows a plot of volume versus pressure for a particular gas sample at constant temperature. Answer the following questions by referring to the graph. No calculation is necessary.
   a. What is the volume of this gas sample at standard pressure?
   b. What is the volume of this gas sample at 4.0 atm pressure?
   c. At what pressure would this gas sample occupy a volume of 5.0 L?

V Vs. P for a Gas at Constant Temperature
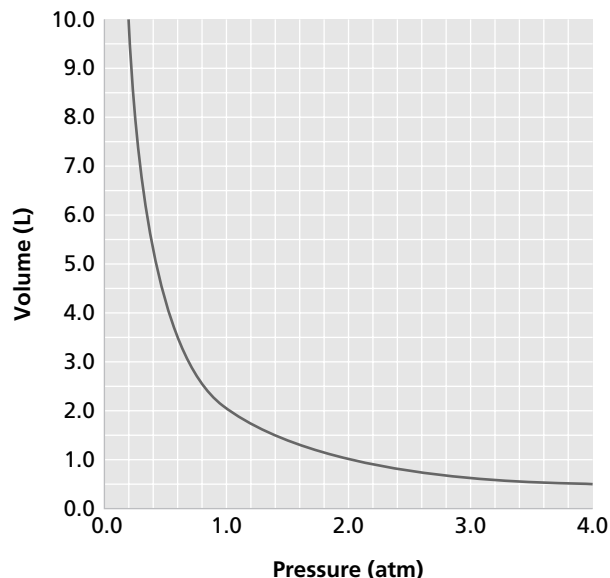
EXTENDED RESPONSE

10. Refer to the plot in question 9. Suppose the same gas sample were heated to a higher temperature and a new graph of V versus P were plotted. Would the new plot be identical to this one? If not, how would it differ?

Test TIP If you are permitted to, draw a line through each incorrect answer choice as you eliminate it.

Mass and Density of Air at Different Pressures

OBJECTIVES

- *Measure* the pressure exerted by a gas.
- *Measure* the mass of a gas sample at different pressures.
- *Graph* the relationship between the mass and pressure of a gas sample.
- *Calculate* the mass of an evacuated bottle.
- *Calculate* the volume of a bottle.
- *Calculate* the density of air at different pressures.

MATERIALS

- **automobile tire valve**
- **balance, centigram**
- **barometer**
- **cloth towel**
- **plastic soda bottle (2 or 3 L) or other heavy plastic bottle**
- **tire pressure gauge**

BACKGROUND

You have learned that the amount of gas present, the volume of the gas, the temperature of the gas sample, and the gas pressure are related to one another. If the volume and temperature of a gas sample are held constant, the mass of the gas and the pressure that the gas exerts are related in a simple way.

In this investigation, you will use an automobile tire pressure gauge to measure the mass of a bottle and the air that the bottle contains for several air pressures. A tire pressure gauge measures "gauge pressure," meaning the added pressure in the tire in addition to normal atmospheric air pressure. Gauge pressure is often expressed in the units pounds per square inch, gauge (psig) to distinguish them from absolute pressures in pounds per square inch (psi). You will graph the mass of the bottle plus air against the gas pressure and observe what kind of plot results. Extrapolating this plot in the proper way will let you determine both the mass and the volume of the empty bottle. This information will also allow you to calculate the density of air at various pressures.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Make a data table with five columns labeled "Gauge pressure (psig)," "Mass of bottle + air (g)," "Corrected gas pressure (psi)," "Mass of air (g)," and "Density of air (g/cm^3)."

PROCEDURE

Your teacher will provide you with a bottle. This bottle contains air under considerable pressure, so handle it carefully. Do not unscrew the cap of the bottle.

1. Use the tire pressure gauge to measure the gauge pressure of the air in the bottle, as accurately as you can read the gauge. It might be convenient for one student to hold the bottle securely, wrapped in a cloth towel, while another student makes the pressure measurement. Record this pressure in your data table.

2. Measure the mass of the bottle plus the air it contains, to the nearest 0.01 g. Record this mass in your data table.

3. With one student holding the wrapped bottle, depress the tire stem valve carefully to allow some air to escape from the bottle until the observed gauge pressure has decreased by 5 to 10 psig. Then, repeat the measurements in steps 1 and 2.

4. Repeat the steps of releasing some pressure (step 3) and then measuring gauge pressure (step 1) and measuring the mass (step 2) until no more air comes out.

5. Now repeat steps 1 and 2 one last time. The gauge pressure should be zero; if it is not, then you probably have not released enough air, and you should depress the valve for a longer time. You should have at least five measurements of gauge pressure and mass, including this final set.

6. Read the atmospheric pressure in the room from the barometer, and record the reading in your notebook.

CLEANUP AND DISPOSAL

7. Return all equipment to its proper place. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. **Organizing Data:** Correct each gauge pressure in your data table to the actual gas pressure in psi, by adding the barometric pressure (in psi) to each measured gauge pressure. Enter these results in the column "Corrected gas pressure."

2. **Analyzing Data:** Make a graph of your data. Plot corrected gas pressure on the x-axis and mass of bottle plus air on the y-axis. The x-axis should run from 10 psi to at least 80 psi. The y-axis scale should allow extrapolation to corrected gas pressure of zero.

3. **Analyzing Data:** If your graph is a straight line, write an equation for the line in the form $y = mx + b$. If the graph is not a straight line, explain why, and draw the straight line that comes closest to including all of your data points. Give the equation of this line.

4. **Interpreting Data:** What is the mass of the empty bottle? (Hint: When no more air escapes from the bottle in steps 4 and 5, the bottle is *not* empty; it still contains air at 1 atm.)

5. **Analyzing Data:** For each of your readings, calculate the mass of air in the bottle. Enter these masses in your data table.

6. **Interpreting Data:** The density of air at typical laboratory conditions is 1.19 g/L. Find the volume of the bottle.

7. **Interpreting Data:** Calculate the density of air at each pressure for which you made measurements. Enter these density values in your data table.

CONCLUSIONS

1. **Inferring Relationships:** Based on your results in this experiment, state the relationship between the mass of a gas sample and the gas pressure. Be sure to include limitations (that is, the quantities that must be kept constant).

2. **Interpreting Graphics:** Using your graph from item 2 of Analysis and Interpretation, predict the mass of the bottle plus air at a gauge reading of 60.0 psig. Estimate the mass of the gas in the bottle at that pressure.

Solutions

Solutions are homogeneous mixtures of two or more substances in a single phase.

Solution in a Pipette Being Dropped into a Test Tube

Types of Mixtures

It is easy to determine that some materials are mixtures because you can see their component parts. For example, soil is a mixture of substances, including small rocks and decomposed animal and plant matter. You can see this by picking up some soil in your hand and looking at it closely. Milk, on the other hand, does not appear to be a mixture, but in fact it is. Milk is composed principally of fats, proteins, milk sugar, and water. If you look at milk under a microscope, it will look something like **Figure 1a.** You can see round lipid (fat) droplets that measure from 1 to 10 μm in diameter. Irregularly shaped casein (protein) particles that are about 0.2 μm wide can also be seen. Both milk and soil are examples of heterogeneous mixtures because their composition is not uniform.

Salt (sodium chloride) and water form a homogeneous mixture. The sodium and chloride ions are interspersed among the water molecules, and the mixture appears uniform throughout. A model for a homogeneous mixture such as salt water is shown in **Figure 1b.**

OBJECTIVES

- Distinguish between heterogeneous and homogeneous mixtures.

- List three different solute-solvent combinations.

- Compare the properties of suspensions, colloids, and solutions.

- Distinguish between electrolytes and nonelectrolytes.

Solutions

Suppose a sugar cube is dropped into a glass of water. You know from experience that the sugar will dissolve. Sugar is described as "soluble in water." *By **soluble** we mean capable of being dissolved*.

What happens as sugar dissolves? The lump gradually disappears as sugar molecules leave the surface of their crystals and mix with water molecules. Eventually all the sugar molecules become uniformly distributed among the water molecules, as indicated by the equally sweet taste of any part of the mixture. All visible traces of the solid sugar are
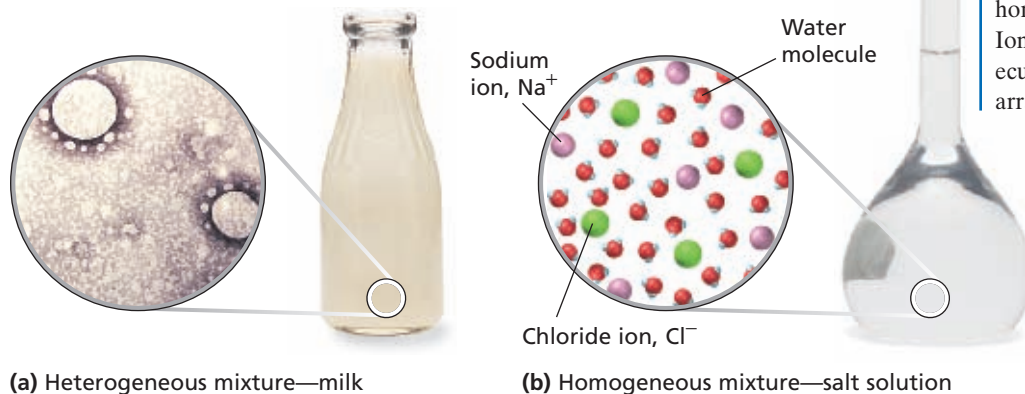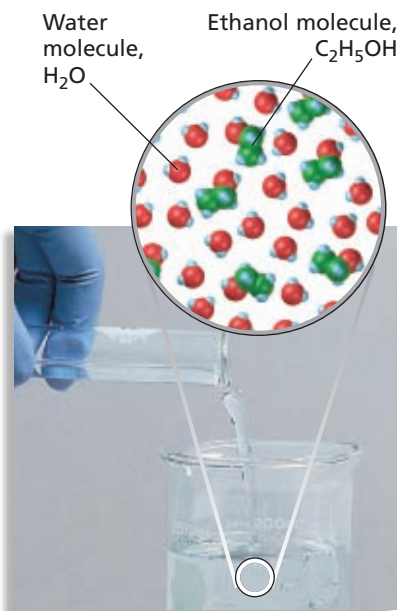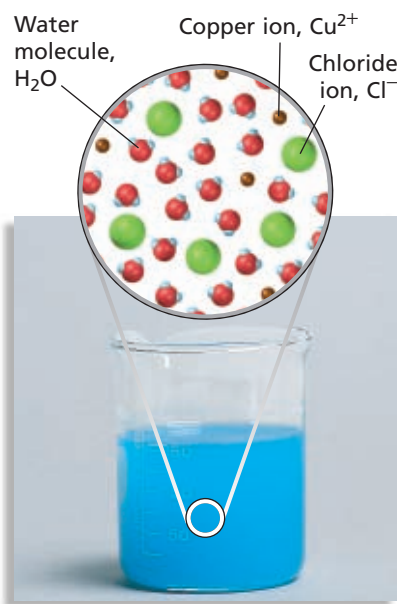
FIGURE 1 (a) Milk consists of visible particles in a nonuniform arrangement. (b) Salt water is an example of a homogeneous mixture. Ions and water molecules are in a random arrangement.

Water molecule

Sodium ion, Na$^+$

Chloride ion, Cl$^-$

(a) Heterogeneous mixture—milk

(b) Homogeneous mixture—salt solution

Water molecule, H_2O

Ethanol molecule, C_2H_5OH

(a)

Water molecule, H_2O

Copper ion, Cu^{2+}

Chloride ion, Cl^-

(b)

FIGURE 2 The solute in a solution can be a solid, liquid, or gas. (a) The ethanol-water solution is made from a liquid solute in a liquid solvent. (b) The copper(II) chloride–water solution is made from a solid solute in a liquid solvent. Note that the composition of each solution is uniform.

gone. Such a mixture is called a solution. *A **solution** is a homogeneous mixture of two or more substances in a single phase.* In a solution, atoms, molecules, or ions are thoroughly mixed, resulting in a mixture that has the same composition and properties throughout.

Components of Solutions

In the simplest type of solution, such as a sugar-water solution, the particles of one substance are randomly mixed with the particles of another substance. *The dissolving medium in a solution is called the **solvent,** and the substance dissolved in a solution is called the **solute.*** The solute is generally designated as that component of a solution that is of lesser quantity. In the ethanol-water solution shown in **Figure 2,** ethanol is the solute and water is the solvent. Occasionally, these terms have little meaning. For example, in a 50%-50% solution of ethanol and water, it would be difficult, and in fact unnecessary, to say which is the solvent and which is the solute.

In a solution, the dissolved solute particles are so small that they cannot be seen. They remain mixed with the solvent indefinitely, so long as the existing conditions remain unchanged. If the solutions in **Figure 2** are poured through filter paper, both the solute and the solvent will pass through the paper. The solute-particle dimensions are those of atoms, molecules, and ions—which range from about 0.01 to 1 nm in diameter.

Types of Solutions

Solutions may exist as gases, liquids, or solids. Some possible solute-solvent combinations of gases, liquids, and solids in solutions are summarized in **Table 1.** In each example, one component is designated as the solvent and one as the solute.

Many alloys, such as brass (made from zinc and copper) and sterling silver (made from silver and copper), are solid solutions in which the atoms of two or more metals are uniformly mixed. By properly choosing the proportions of each metal in the alloy, many desirable properties can be obtained. For example, alloys can have higher strength and

TABLE 1 *Some Solute-Solvent Combinations for Solutions*

Solute state	Solvent state	Example
Gas	gas	oxygen in nitrogen
Gas	liquid	carbon dioxide in water
Liquid	liquid	alcohol in water
Liquid	solid	mercury in silver and tin (dental amalgam)
Solid	liquid	sugar in water
Solid	solid	copper in nickel (Monel™ alloy)

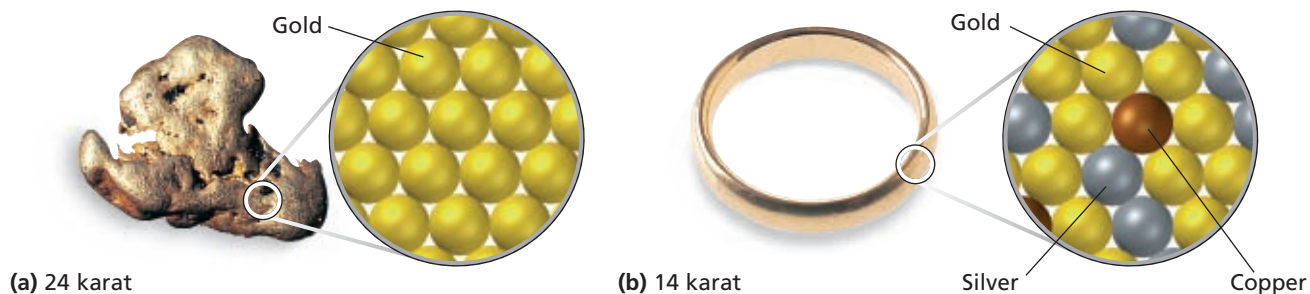**(a)** 24 karat   **(b)** 14 karat

FIGURE 3 (a) 24-karat gold is pure gold. (b) 14-karat gold is an alloy of gold with silver and copper. 14-karat gold is 14/24, or 58.3%, gold.

greater resistance to corrosion than the pure metals. Pure gold (24K), for instance, is too soft to use in jewelry. Alloying it with silver and copper greatly increases its strength and hardness while retaining its appearance and corrosion resistance. **Figure 3** shows a comparison between pure gold and a gold alloy. 14-karat gold is a solution because the gold, silver, and copper are evenly mixed at the atomic level.

Suspensions

If the particles in a solvent are so large that they settle out unless the mixture is constantly stirred or agitated, the mixture is called a **suspension.** Think of a jar of muddy water. If left undisturbed, particles of soil collect on the bottom of the jar. The soil particles are denser than the solvent, water. Gravity pulls them to the bottom of the container. Particles over 1000 nm in diameter—1000 times as large as atoms, molecules, or ions—form suspensions. The particles in suspension can be separated from heterogeneous mixtures by passing the mixture through a filter.

Colloids

Particles that are intermediate in size between those in solutions and suspensions form mixtures known as colloidal dispersions, or simply **colloids.** Particles between 1 nm and 1000 nm in diameter may form colloids. After large soil particles settle out of muddy water, the water is often still cloudy because colloidal particles remain dispersed in the water. If the cloudy mixture is poured through a filter, the colloidal particles will pass through, and the mixture will remain cloudy. The particles in a colloid are small enough to be suspended throughout the solvent by the constant movement of the surrounding molecules. The colloidal particles make up the *dispersed phase,* and water is the *dispersing medium*. Examples of the various types of colloids are given in **Table 2.** Note that some familiar terms, such as *emulsion* and *foam,* refer to specific types of colloids. For example, mayonnaise is an emulsion of oil
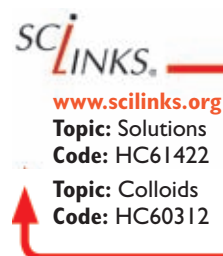
TABLE 2 *Classes of Colloids*

Class of colloid	Phases	Example
Sol	solid dispersed in liquid	paints, mud
Gel	solid network extending throughout liquid	gelatin
Liquid emulsion	liquid dispersed in a liquid	milk, mayonnaise
Foam	gas dispersed in liquid	shaving cream, whipped cream
Solid aerosol	solid dispersed in gas	smoke, airborne particulate matter, auto exhaust
Liquid aerosol	liquid dispersed in gas	fog, mist, clouds, aerosol spray
Solid emulsion	liquid dispersed in solid	cheese, butter

FIGURE 4 A beam of light distinguishes a colloid from a solution. The particles in a colloid will scatter light, making the beam visible. The mixture of gelatin and water in the jar on the right is a colloid. The mixture of water and sodium chloride in the jar on the left is a true solution.

droplets in water; the egg yolk in it acts as an emulsifying agent, which helps to keep the oil droplets dispersed.

Tyndall Effect

Many colloids appear homogeneous because the individual particles cannot be seen. The particles are, however, large enough to scatter light. You have probably noticed that a headlight beam is visible from the side on a foggy night. This effect, known as the *Tyndall effect,* occurs when light is scattered by colloidal particles dispersed in a transparent medium. The Tyndall effect is a property that can be used to distinguish between a solution and a colloid, as demonstrated in **Figure 4.**

The distinctive properties of solutions, colloids, and suspensions are summarized in **Table 3.** The individual particles of a colloid can be detected under a microscope if a bright light is cast on the specimen at a right angle. The particles, which appear as tiny specks of light, are seen to move rapidly in a random motion. This motion is due to collisions of rapidly moving molecules and is called *Brownian motion,* after its discoverer, Robert Brown.

TABLE 3 *Properties of Solutions, Colloids, and Suspensions*

Solutions	Colloids	Suspensions
Homogeneous	Heterogeneous	Heterogeneous
Particle size: 0.01–1 nm; can be atoms, ions, molecules	Particle size: 1–1000 nm, dispersed; can be aggregates or large molecules	Particle size: over 1000 nm, suspended; can be large particles or aggregates
Do not separate on standing	Do not separate on standing	Particles settle out
Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
Do not scatter light	Scatter light (Tyndall effect)	May scatter light, but are not transparent

Observing Solutions, Suspensions, and Colloids

Procedure

1. Prepare seven mixtures, each containing 250 mL of water and one of the following substances.
   a. 12 g of sucrose
   b. 3 g of soluble starch
   c. 5 g of clay
   d. 2 mL of food coloring
   e. 2 g of sodium borate
   f. 50 mL of cooking oil
   g. 3 g of gelatin

Making the gelatin mixture:
Soften the gelatin in 65 mL of cold water, and then add 185 mL of boiling water.

2. Observe the seven mixtures and their characteristics. Record the appearance of each mixture after stirring.

3. Transfer to individual test tubes 10 mL of each mixture that does not separate after stirring. Shine a flashlight on each mixture in a dark room. Make note of the mixtures in which the path of the light beam is visible.

Discussion

1. Using your observations, classify each mixture as a solution, suspension, or colloid.

2. What characteristics did you use to classify each mixture?

Materials

- balance
- 7 beakers, 400 mL
- clay
- cooking oil
- flashlight
- gelatin, plain
- hot plate (to boil H_2O)
- red food coloring
- sodium borate ($Na_2B_4O_7 \cdot 10H_2O$)
- soluble starch
- stirring rod
- sucrose
- test-tube rack
- water

Solutes: Electrolytes Versus Nonelectrolytes

Substances that dissolve in water are classified according to whether they yield molecules or ions in solution. When an ionic compound dissolves, the positive and negative ions separate from each other and are surrounded by water molecules. These solute ions are free to move, making it possible for an electric current to pass through the solution. *A substance that dissolves in water to give a solution that conducts electric current is called an* **electrolyte.** Sodium chloride, NaCl, is an electrolyte, as is any soluble ionic compound. Certain highly polar molecular compounds, such as hydrogen chloride, HCl, are also electrolytes because HCl molecules form the ions H_3O^+ and Cl^- when dissolved in water.

By contrast, a solution containing neutral solute molecules does not conduct electric current because it does not contain mobile charged

Chloride ion, Cl^- — Water molecule, H_2O — Sodium ion, Na^+

Sugar molecule, $C_{12}H_{22}O_{11}$ — Water molecule, H_2O

Hydronium ion, H_3O^+ — Water molecule, H_2O — Chloride ion, Cl^-

(a) Salt solution—
electrolyte solute

(b) Sugar solution—
nonelectrolyte solute

(c) Hydrochloric acid solution—
electrolyte solute
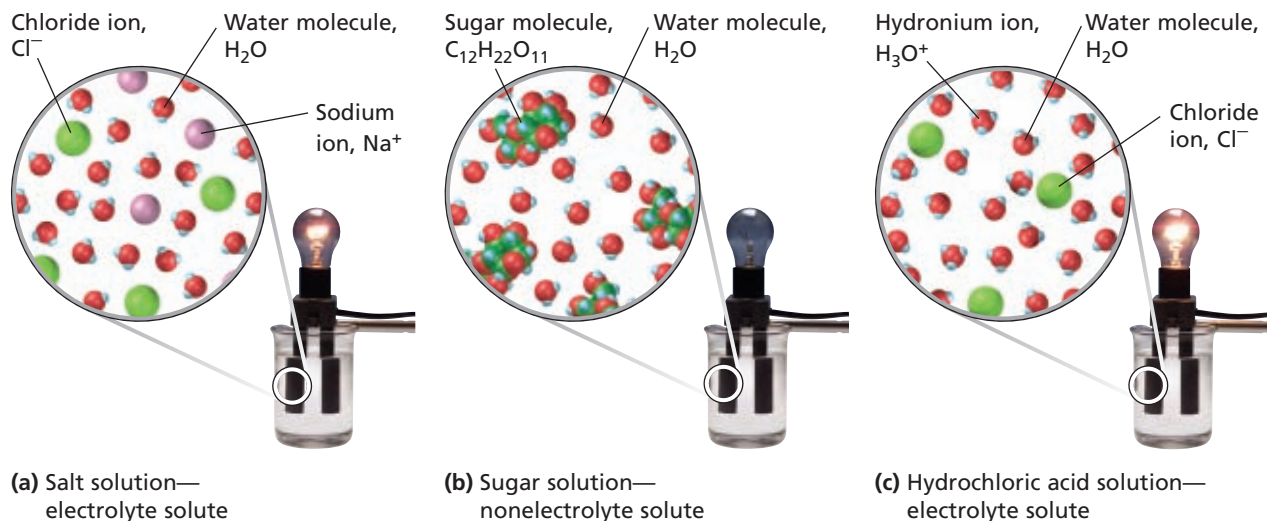
FIGURE 5 (a) Sodium chloride dissolves in water to produce a salt solution that conducts electric current. NaCl is an electrolyte. (b) Sucrose dissolves in water to produce a sugar solution that does not conduct electricity. Sucrose is a nonelectrolyte. (c) Hydrogen chloride dissolves in water to produce a solution that conducts current. HCl is an electrolyte.

particles. *A substance that dissolves in water to give a solution that does not conduct an electric current is called a* **nonelectrolyte.** Sugar is a nonelectrolyte. **Figure 5** shows an apparatus for testing the conductivity of solutions. The electrodes are conductors that are attached to a power supply and that make electric contact with the test solution. For a current to pass through the light-bulb filament, the test solution must provide a conducting path between the two electrodes. A nonconducting solution is like an open switch between the electrodes, and there is no current in the circuit.

The light bulb glows brightly if a solution that is a good conductor is tested. Such solutions contain solutes that are electrolytes. For a moderately conductive solution, however, the light bulb is dim. If a solution is a poor conductor, the light bulb does not glow at all. Such solutions contain solutes that are nonelectrolytes. You will learn more about the strengths and behavior of electrolytes in Chapter 13.

SECTION REVIEW

1. Classify the following as either a heterogeneous or homogeneous mixture, and explain your answers.

   a. orange juice   **b.** tap water

2. **a.** What are substances called whose water solutions conduct electricity? **b.** Why does a salt solution conduct electricity? **c.** Why does a sugar-water solution not conduct electricity?

3. Make a drawing of the particles in an NaCl solution to show why this solution conducts electricity. Make a drawing of the particles in an NaCl crystal to show why pure salt does not conduct.

4. Describe one way to prove that a mixture of sugar and water is a solution and that a mixture of sand and water is not a solution.

5. Name the solute and solvent in the following:

   a. 14-karat gold
   b. corn syrup
   c. carbonated, or sparkling, water

Critical Thinking

6. **ANALYZING INFORMATION** If you allow a container of sea water to sit in the sun, the liquid level gets lower and lower, and finally crystals appear. What is happening?

The Solution Process

Factors Affecting the Rate of Dissolution

If you have ever tried to dissolve sugar in iced tea, you know that temperature has something to do with how quickly a solute dissolves. What other factors affect how quickly you can dissolve sugar in iced tea?

Increasing the Surface Area of the Solute

Sugar dissolves as sugar molecules leave the crystal surface and mix with water molecules. The same is true for any solid solute in a liquid solvent: molecules or ions of the solute are attracted by the solvent.

Because the dissolution process occurs at the surface of the solute, it can be speeded up if the surface area of the solute is increased. Crushing sugar that is in cubes or large crystals increases its surface area. In general, the more finely divided a substance is, the greater the surface area per unit mass and the more quickly it dissolves. **Figure 6** shows a model of solutions that are made from the same solute but have a different amount of surface area exposed to the solvent.

Agitating a Solution

Very close to the surface of a solute, the concentration of dissolved solute is high. Stirring or shaking helps to disperse the solute particles

Small surface area exposed to solvent—slow rate

Large surface area exposed to solvent—faster rate

Solvent particle
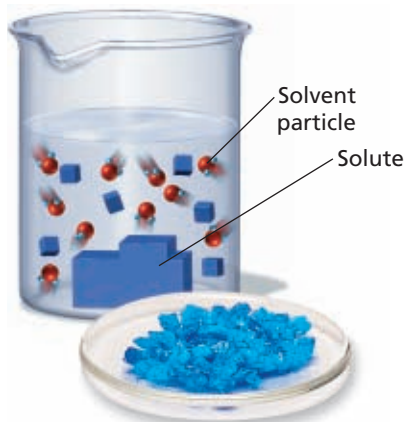
Solute

$CuSO_4 \cdot 5H_2O$ large crystals

$CuSO_4 \cdot 5H_2O$ powdered
Increased surface area

FIGURE 6 The rate at which a solid solute dissolves can be increased by increasing the surface area. A powdered solute has a greater surface area exposed to solvent particles and therefore dissolves faster than a solute in large crystals.

SCI LINKS.

www.scilinks.org
Topic: Solubility
Code: HC61421

and bring fresh solvent into contact with the solute surface. Thus, the effect of stirring is similar to that of crushing a solid—contact between the solvent and the solute surface is increased.

Heating a Solvent

You have probably noticed that sugar and many other materials dissolve more quickly in warm water than in cold water. As the temperature of the solvent increases, solvent molecules move faster, and their average kinetic energy increases. Therefore, at higher temperatures, collisions between the solvent molecules and the solute are more frequent and are of higher energy than at lower temperatures. This helps to separate solute molecules from one another and to disperse them among the solvent molecules.

Solubility

If you add spoonful after spoonful of sugar to tea, eventually no more sugar will dissolve. For every combination of solvent with a solid solute at a given temperature, there is a limit to the amount of solute that can be dissolved. The point at which this limit is reached for any solute-solvent combination is difficult to predict precisely and depends on the nature of the solute, the nature of the solvent, and the temperature.

The following model describes why there is a limit. When solid sugar is first added to water, sugar molecules leave the solid surface and move about at random in the solvent. Some of these dissolved molecules may collide with the crystal and remain there (recrystallize). As more solid dissolves and the concentration of dissolved molecules increases, these collisions become more frequent. Eventually, molecules are returning to the crystal at the same rate at which they are going into solution, and a dynamic equilibrium is established between dissolution and crystallization. Ionic solids behave similarly, as shown in **Figure 7.**

Solution equilibrium *is the physical state in which the opposing processes of dissolution and crystallization of a solute occur at equal rates.*
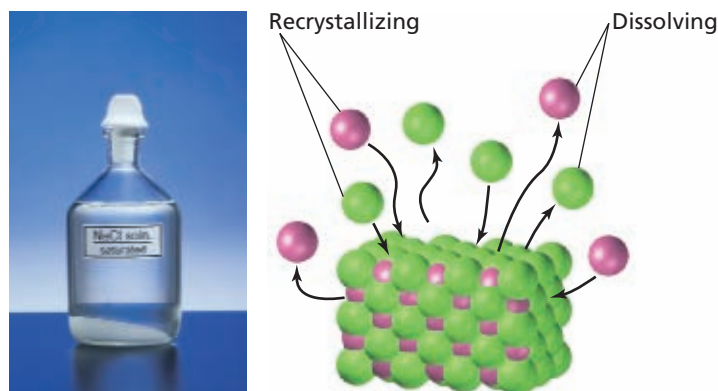
FIGURE 7 A saturated solution in a closed system is at equilibrium. The solute is recrystallizing at the same rate that it is dissolving, even though it appears that there is no activity in the system.
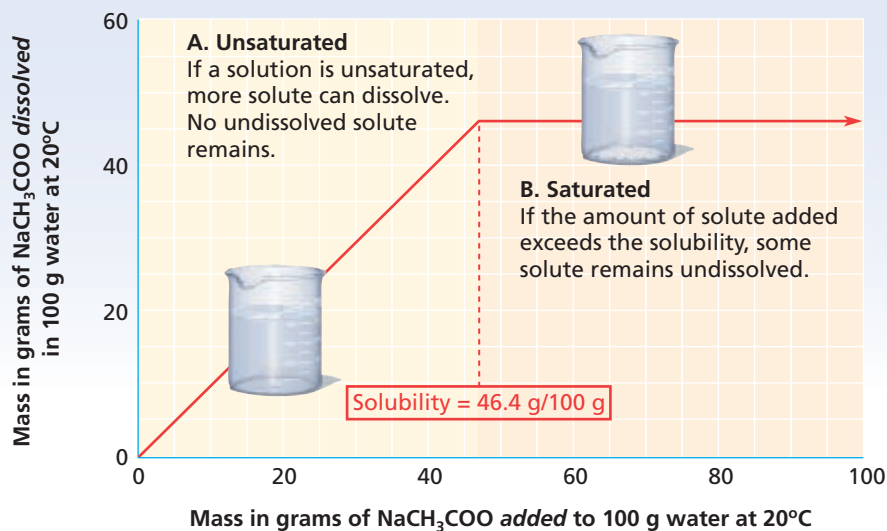
Mass of Solute Added Vs. Mass of Solute Dissolved

Mass in grams of NaCH₃COO dissolved in 100 g water at 20°C (y-axis, 0 to 60)

A. Unsaturated
If a solution is unsaturated, more solute can dissolve. No undissolved solute remains.

B. Saturated
If the amount of solute added exceeds the solubility, some solute remains undissolved.

Solubility = 46.4 g/100 g

Mass in grams of NaCH₃COO *added* to 100 g water at 20°C (x-axis, 0 to 100)

FIGURE 8 The graph shows the range of solute masses that will produce an unsaturated solution. Once the saturation point is exceeded, the system will contain undissolved solute.

Saturated Versus Unsaturated Solutions

A solution that contains the maximum amount of dissolved solute is described as a **saturated solution.** How can you tell that the NaCH₃COO solution pictured in **Figure 8** is saturated? If more sodium acetate is added to the solution, it falls to the bottom and does not dissolve because an equilibrium has been established between ions leaving and entering the solid phase. If more water is added to the saturated solution, then more sodium acetate will dissolve in it. At 20°C, 46.4 g of NaCH₃COO is the maximum amount that will dissolve in 100. g of water. *A solution that contains less solute than a saturated solution under the existing conditions is an* **unsaturated solution.**

Supersaturated Solutions

When a saturated solution of a solute whose solubility increases with temperature is cooled, the excess solute usually comes out of solution, leaving the solution saturated at the lower temperature. But sometimes, if the solution is left to cool undisturbed, the excess solute does not separate and a supersaturated solution is produced. *A* **supersaturated solution** *is a solution that contains more dissolved solute than a saturated solution contains under the same conditions*. A supersaturated solution may remain unchanged for a long time if it is not disturbed, but once crystals begin to form, the process continues until equilibrium is reestablished at the lower temperature. An example of a supersaturated solution is one prepared from a saturated solution of sodium thiosulfate, $Na_2S_2O_3$, or sodium acetate, $NaCH_3COO$. Solute is added to hot water until the solution is saturated, and the hot solution is filtered. The filtrate is left to stand undisturbed as it cools. Dropping a small crystal of the solute into the supersaturated solution ("seeding") or disturbing the solution causes a rapid formation of crystals by the excess solute.

Solubility Values

The **solubility** *of a substance is the amount of that substance required to form a saturated solution with a specific amount of solvent at a specified temperature.* The solubility of sugar, for example, is 204 g per 100. g of water at 20.°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubilities must be determined experimentally, and they vary widely, as illustrated in **Table 4.** Solubility values can be found in chemical handbooks and are usually given as grams of solute per 100. g of solvent or per 100. mL of solvent at a given temperature.

The rate at which a solid dissolves is unrelated to its solubility at that temperature. The maximum amount of a given solute that dissolves and reaches equilibrium is always the same under the same conditions.

Solute-Solvent Interactions

Lithium chloride is highly soluble in water, but gasoline is not. On the other hand, gasoline mixes readily with benzene, C_6H_6, but lithium chloride does not. Why are there such differences in solubility?

"Like dissolves like" is a rough but useful rule for predicting whether one substance will dissolve in another. What makes substances similar depends on the type of bonding, the polarity or nonpolarity of molecules, and the intermolecular forces between the solute and solvent.

TABLE 4 *Solubility of Solutes as a Function of Temperature (in g solute/100. g H_2O)*

Substance	Temperature (°C)					
	0	20	40	60	80	100
$AgNO_3$	122	216	311	440	585	733
$Ba(OH)_2$	1.67	3.89	8.22	20.94	101.4	—
$C_{12}H_{22}O_{11}$	179	204	238	287	362	487
$Ca(OH)_2$	0.189	0.173	0.141	0.121	—	0.07
$Ce_2(SO_4)_3$	20.8	10.1	—	3.87	—	—
KCl	28.0	34.2	40.1	45.8	51.3	56.3
KI	128	144	162	176	192	206
KNO_3	13.9	31.6	61.3	106	167	245
LiCl	69.2	83.5	89.8	98.4	112	128
Li_2CO_3	1.54	1.33	1.17	1.01	0.85	0.72
NaCl	35.7	35.9	36.4	37.1	38.0	39.2
$NaNO_3$	73	87.6	102	122	148	180
CO_2 (gas at SP)	0.335	0.169	0.0973	0.058	—	—
O_2 (gas at SP)	0.00694	0.00537	0.00308	0.00227	0.00138	0.00
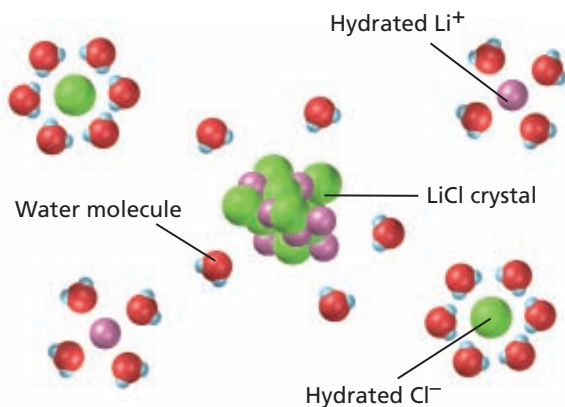
Dissolving Ionic Compounds in Aqueous Solution

The polarity of water molecules plays an important role in the formation of solutions of ionic compounds in water. The slightly charged parts of water molecules attract the ions in the ionic compounds and surround them to keep them separated from the other ions in the solution. Suppose we drop a few crystals of lithium chloride into a beaker of water. At the crystal surfaces, water molecules come into contact with Li^+ and Cl^- ions. The positive ends of the water molecules are attracted to Cl^- ions, while the negative ends are attracted to Li^+ ions. The attraction between water molecules and the ions is strong enough to draw the ions away from the crystal surface and into solution, as illustrated in **Figure 9.** *This solution process with water as the solvent is referred to as* **hydration.** The ions are said to be *hydrated.* As hydrated ions diffuse into the solution, other ions are exposed and are drawn away from the crystal surface by the solvent. The entire crystal gradually dissolves, and hydrated ions become uniformly distributed in the solution.

When crystallized from aqueous solutions, some ionic substances form crystals that incorporate water molecules. These crystalline compounds, known as *hydrates,* retain specific ratios of water molecules and are represented by formulas such as $CuSO_4 \cdot 5H_2O$. Heating the crystals of a hydrate can drive off the water of hydration and leave the anhydrous salt. When a crystalline hydrate dissolves in water, the water of hydration returns to the solvent. The behavior of a solution made from a hydrate is no different from the behavior of one made from the anhydrous form. Dissolving either form results in a system containing hydrated ions and water.

Nonpolar Solvents

Ionic compounds are generally not soluble in nonpolar solvents such as carbon tetrachloride, CCl_4, and toluene, $C_6H_5CH_3$. The nonpolar solvent molecules do not attract the ions of the crystal strongly enough to overcome the forces holding the crystal together.

Would you expect lithium chloride to dissolve in toluene? No, LiCl is not soluble in toluene. LiCl and $C_6H_5CH_3$ differ widely in bonding, polarity, and intermolecular forces.
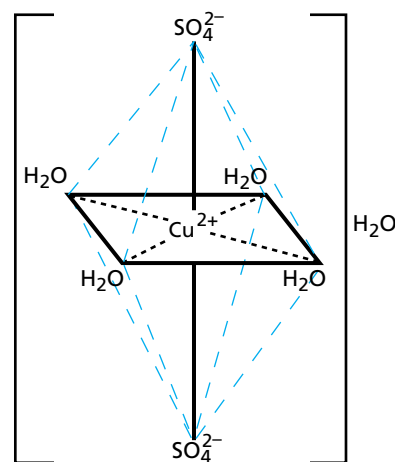
Liquid Solutes and Solvents
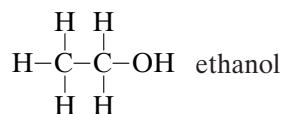
When you shake a bottle of salad dressing, oil droplets become dispersed in the water. As soon as you stop shaking the bottle, the strong attraction of hydrogen bonding between the water molecules squeezes out the oil droplets, forming separate layers. *Liquids that are not soluble in each other are* **immiscible.** Toluene and water, shown in **Figure 11,** are another example of immiscible substances.

Nonpolar substances, such as fats, oils, and greases, are generally quite soluble in nonpolar liquids, such as carbon tetrachloride, toluene, and gasoline. The only attractions between the nonpolar molecules are London forces, which are quite weak. The intermolecular forces existing in the solution are therefore very similar to those in pure substances. Thus, the molecules can mix freely with one another.

Liquids that dissolve freely in one another in any proportion are said to be **miscible.** Benzene and carbon tetrachloride are miscible. The nonpolar molecules of these substances exert no strong forces of attraction or repulsion, so the molecules mix freely. Ethanol and water, shown in **Figure 12,** also mix freely, but for a different reason. The $-OH$ group on an ethanol molecule is somewhat polar. This group can form hydrogen bonds with water as well as with other ethanol molecules. The intermolecular forces in the mixture are so similar to those in the pure liquids that the liquids are mutually soluble in all proportions.

$$\text{H}-\overset{\overset{\displaystyle H}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-\overset{\overset{\displaystyle H}{|}}{\underset{\underset{\displaystyle H}{|}}{C}}-\text{OH}\quad \text{ethanol}$$

Gasoline is a solution composed mainly of nonpolar hydrocarbons and is also an excellent solvent for fats, oils, and greases. The major intermolecular forces acting between the nonpolar molecules are weak London forces.

Ethanol is intermediate in polarity between water and carbon tetrachloride. It is not as good a solvent for polar or ionic substances as water is. Sodium chloride is only slightly soluble in ethanol. On the other hand, ethanol is a better solvent than water is for less-polar substances because the molecule has a nonpolar region.
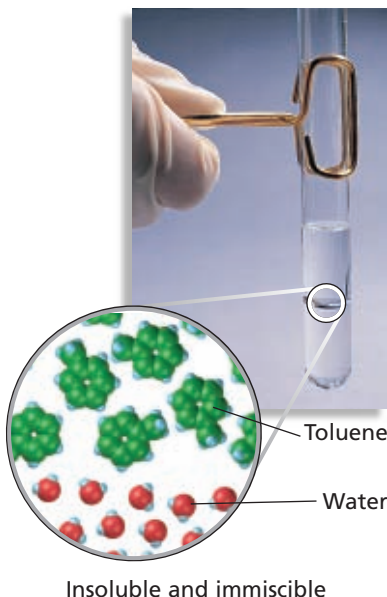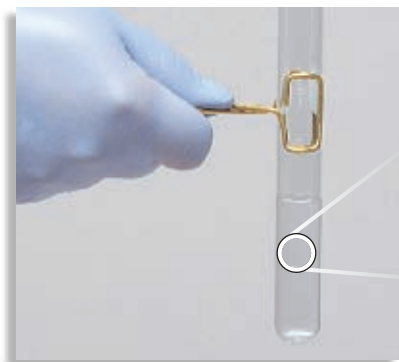

FIGURE 11 Toluene and water are immiscible. The components of this system exist in two distinct phases.

Insoluble and immiscible

FIGURE 12 (a) Water and ethanol are miscible. The components of this system exist in a single phase with a uniform arrangement. (b) Hydrogen bonding between the solute and solvent enhances the solubility of ethanol in water.

(a) Miscible

Water molecule, H_2O

Ethanol molecule, C_2H_5OH

Hydrogen bond

$\delta-$ $\delta+$ $\delta-$ $\delta+$ $\delta-$ $\delta+$

(b)
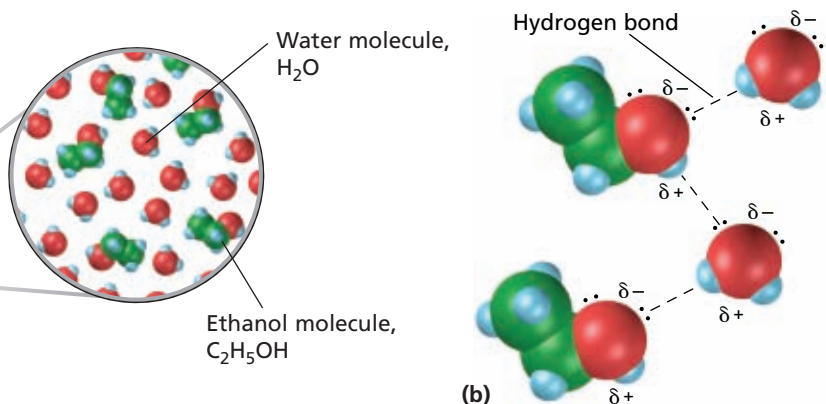
Effects of Pressure on Solubility

Changes in pressure have very little effect on the solubilities of liquids or solids in liquid solvents. However, increases in pressure increase gas solubilities in liquids.

When a gas is in contact with the surface of a liquid, gas molecules can enter the liquid. As the amount of dissolved gas increases, some molecules begin to escape and reenter the gas phase. An equilibrium is eventually established between the rates at which gas molecules enter and leave the liquid phase. As long as this equilibrium is undisturbed, the solubility of the gas in the liquid is unchanged at a given pressure.

$$\text{gas} + \text{solvent} \rightleftarrows \text{solution}$$
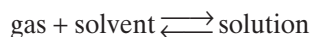
Increasing the pressure of the solute gas above the solution puts stress on the equilibrium. Molecules collide with the liquid surface more often. The increase in pressure is partially offset by an increase in the rate of gas molecules entering the solution. In turn, the increase in the amount of dissolved gas causes an increase in the rate at which molecules escape from the liquid surface and become vapor. Eventually, equilibrium is restored at a higher gas solubility. An increase in gas pressure causes the equilibrium to shift so that more molecules are in the liquid phase.
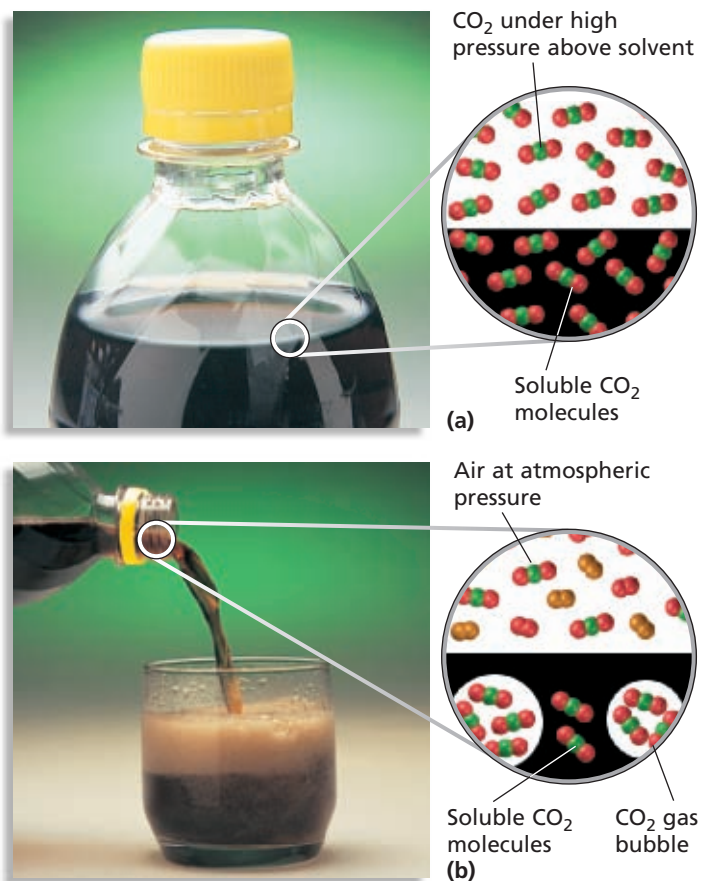
Henry's Law

The solubility of a gas in a liquid is directly proportional to the partial pressure of that gas on the surface of the liquid. This is a statement of **Henry's law,** named after the English chemist William Henry. Henry's law applies to gas-liquid solutions at constant temperature.

Recall that when a mixture of ideal gases is confined in a constant volume at a constant temperature, each gas exerts the same pressure it would exert if it occupied the space alone. Assuming that the gases do not react in any way, each gas dissolves to the extent it would dissolve if no other gases were present.

In carbonated beverages, the solubility of CO_2 is increased by increasing the pressure. At the bottling plant, carbon dioxide gas is forced into the solution of flavored water at a pressure of 5–10 atm. The gas-in-liquid solution is then sealed in bottles or cans. When the cap is removed, the pressure is reduced to 1 atm, and some of the carbon dioxide escapes as gas bubbles. *The rapid escape of a gas from a liquid in which it is dissolved is known as* **effervescence** *and is shown in* **Figure 13.**

FIGURE 13 (a) There are no gas bubbles in the unopened bottle of soda because the pressure of CO_2 applied during bottling keeps the carbon dioxide gas dissolved in the liquid. (b) When the cap on the bottle is removed, the pressure of CO_2 on the liquid is reduced, and CO_2 can escape from the liquid. The soda effervesces when the bottle is opened and the pressure is reduced.

CO_2 under high pressure above solvent

Soluble CO_2 molecules

(a)

Air at atmospheric pressure

Soluble CO_2 molecules

CO_2 gas bubble

(b)

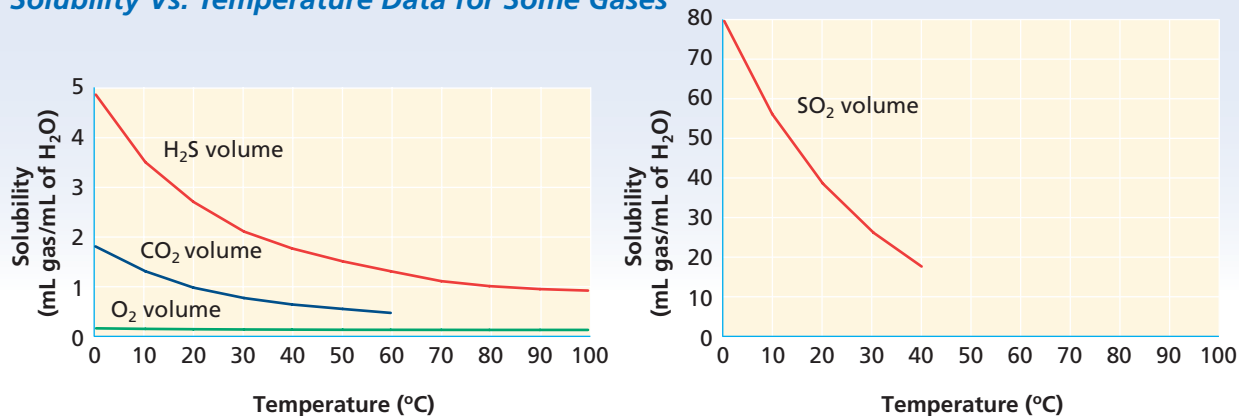## Solubility Vs. Temperature Data for Some Gases

FIGURE 14 The solubility of gases in water decreases with increasing temperature. Which gas has the greater solubility at 30°C—CO_2 or SO_2?

Effects of Temperature on Solubility

First, let's consider gas solubility. Increasing the temperature usually decreases gas solubility. As the temperature increases, the average kinetic energy of the molecules in solution increases. A greater number of solute molecules are able to escape from the attraction of solvent molecules and return to the gas phase. At higher temperatures, therefore, equilibrium is reached with fewer gas molecules in solution, and gases are generally less soluble, as shown in **Figure 14.**

The effect of temperature on the solubility of solids in liquids is more difficult to predict. Often, increasing the temperature increases the solubility of solids. However, an equivalent temperature increase can result

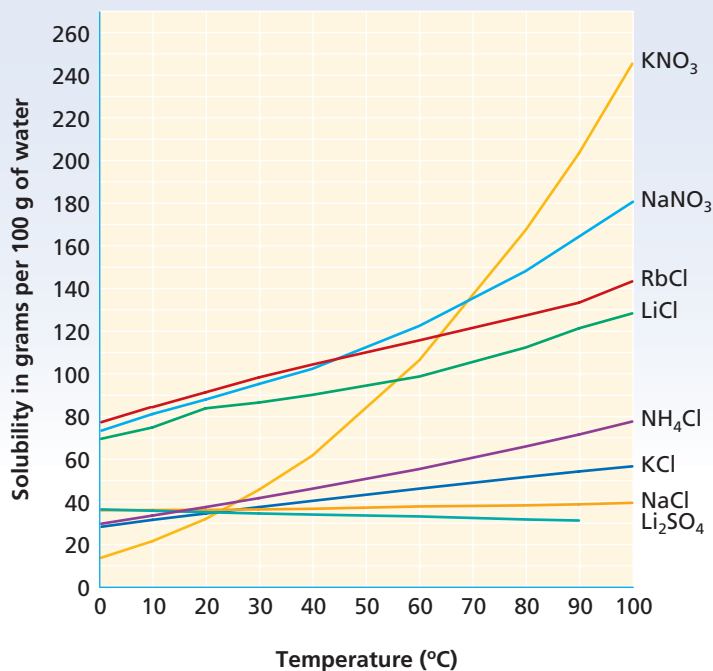## Solubility Vs. Temperature for Some Solid Solutes

FIGURE 15 Solubility curves for various solid solutes generally show increasing solubility with increases in temperature. From the graph, you can see that the solubility of $NaNO_3$ is affected more by temperature than is NaCl.

in a large increase in solubility for some solvents and only a slight change for others.

In **Table 4** and **Figure 15,** compare the effect of temperature on the solubilities of potassium nitrate, KNO_3, and sodium chloride, NaCl. About 14 g of potassium nitrate will dissolve in 100. g of water at 0.°C. The solubility of potassium nitrate increases by more than 150 g KNO_3 per 100. g H_2O when the temperature is raised to 80.°C. Under similar circumstances, the solubility of sodium chloride increases by only about 2 g NaCl per 100. g H_2O. In some cases, solubility of a solid *decreases* with an increase in temperature. For example, between 0.°C and 60.°C the solubility of cerium sulfate, $Ce_2(SO_4)_3$, decreases by about 17 g/100 g.

Enthalpies of Solution

The formation of a solution is accompanied by an energy change. If you dissolve some potassium iodide, KI, in water, you will find that the outside of the container feels cold to the touch. But if you dissolve some sodium hydroxide, NaOH, in the same way, the outside of the container feels hot. The formation of a solid-liquid solution can apparently either absorb energy (KI in water) or release energy as heat (NaOH in water).

During the formation of a solution, solvent and solute particles experience changes in the forces attracting them to other particles. Before dissolving begins, solvent molecules are held together by intermolecular forces (solvent-solvent attraction). In the solute, molecules are held together by intermolecular forces (solute-solute attraction). Energy is required to separate solute molecules and solvent molecules from their neighbors. *A solute particle that is surrounded by solvent molecules,* as shown by the model in **Figure 9,** *is said to be* **solvated.**

Solution formation can be pictured as the result of the three interactions summarized in **Figure 16.**

FIGURE 16 The graph shows the changes in the enthalpy that occur during the formation of a solution. How would the graph differ for a system with an endothermic heat of solution?

Enthalpy (*H*)

Components

Solute   Solvent

Step 1
Solute particles becoming separated from solid.
Energy absorbed

Step 2
Solvent particles being moved apart to allow solute particles to enter liquid.
Energy absorbed

Step 3
Solvent particles being attracted to and solvating solute particles.
Energy released

ΔH solution   Exothermic
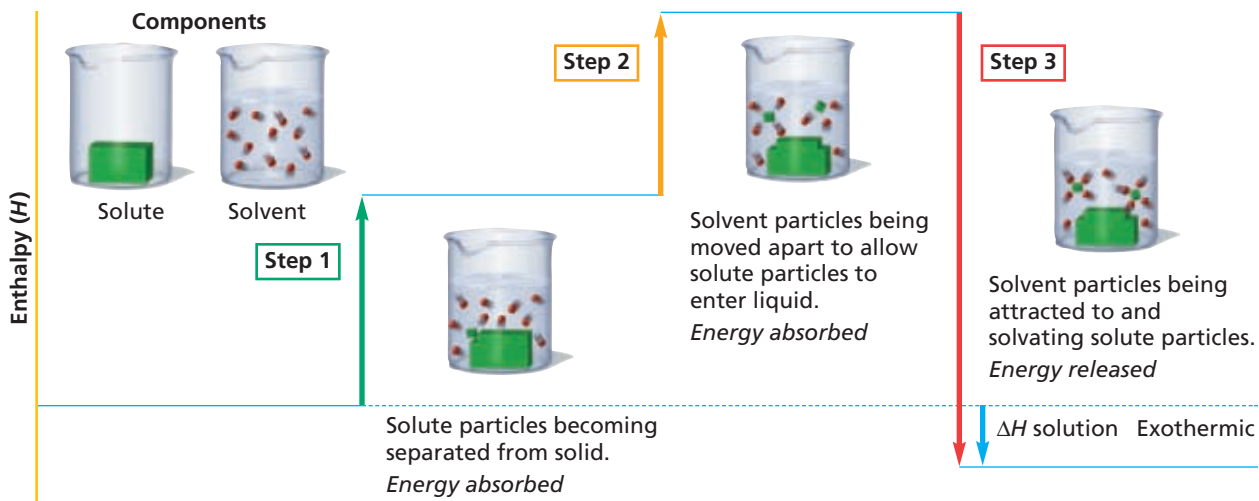
TABLE 5 Enthalpies of Solution (kJ/mol solute at 25°C)

Substance	Enthalpy of solution	Substance	Enthalpy of solution
$AgNO_3(s)$	+22.59	$KOH(s)$	–57.61
$CH_3COOH(l)$	–1.51	$MgSO_4(s)$	+15.9
$HCl(g)$	–74.84	$NaCl(s)$	+3.88
$HI(g)$	–81.67	$NaNO_3(s)$	+20.50
$KCl(s)$	+17.22	$NaOH(s)$	–44.51
$KClO_3(s)$	+41.38	$NH_3(g)$	–30.50
$KI(s)$	+20.33	$NH_4Cl(s)$	+14.78
$KNO_3(s)$	+34.89	$NH_4NO_3(s)$	+25.69

The net amount of energy absorbed as heat by the solution when a specific amount of solute dissolves in a solvent is the **enthalpy of solution.** From the model in **Figure 16,** you can see that the enthalpy of solution is negative (energy is released) when the sum of attractions from Steps 1 and 2 is less than Step 3. The enthalpy of solution is positive (energy is absorbed) when the sum of attractions from Steps 1 and 2 is greater than Step 3.

You know that heating decreases the solubility of a gas, so dissolution of gases is exothermic. How do the values for the enthalpies of solution in **Table 5** support this idea of exothermic solution processes for gaseous solutes?

In the gaseous state, molecules are so far apart that there are virtually no intermolecular forces of attraction between them. Therefore, the solute-solute interaction has little effect on the enthalpy of a solution of a gas. Energy is released when a gas dissolves in a liquid because attraction between solute gas and solvent molecules outweighs the energy needed to separate solvent molecules.

SECTION REVIEW

1. Why would you expect a packet of sugar to dissolve faster in hot tea than in iced tea?

2. a. Explain how you would prepare a saturated solution of sugar in water. b. How would you then make it a supersaturated solution?

3. Explain why ethanol will dissolve in water and carbon tetrachloride will not.

4. When a solute molecule is solvated, is energy released or absorbed?

5. If a warm bottle of soda and a cold bottle of soda are opened, which will effervesce more and why?

Critical Thinking

6. **PREDICTING OUTCOMES** You get a small amount of lubricating oil on your clothing. Which would work better to remove the oil—water or toluene? Explain your answer.

7. **INTERPRETING CONCEPTS** A commercial "fizz saver" pumps helium under pressure into a soda bottle to keep gas from escaping. Will this keep CO_2 in the soda bottle? Explain your answer.
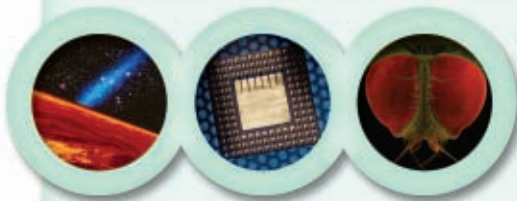
Artificial Blood

A patient lies bleeding on a stretcher. The doctor leans over to check the patient's wounds and barks an order to a nearby nurse: "Get him a unit of artificial blood, stat!" According to Dr. Peter Keipert, Program Director of Oxygen Carriers Development at Alliance Pharmaceutical Corp., this scenario may soon be commonplace thanks to a synthetic mixture that can perform one of the main functions of human blood—transporting oxygen.

The hemoglobin inside red blood cells collects oxygen in our lungs, transports it to all the tissues of the body, and then takes carbon dioxide back to the lungs. Dr. Keipert's blood substitute accomplishes the same task, but it uses nonpolar chemicals called *perfluorocarbons* instead of hemoglobin to transport the oxygen. The perfluorocarbons are carried in a water-based saline solution, but because nonpolar substances and water do not mix well, a bonding chemical called a *surfactant* is added to hold the mixture together. The perfluorocarbons are sheared into tiny droplets and then coated with the bonding molecules. One end of these molecules attaches to the perfluorocarbon, and the other end attaches to the water, creating a milky emulsion. The blood-substitute mixture, called Oxygent™, is administered to a patient in the same way regular blood is. The perfluorocarbons are eventually exhaled through the lungs.

Oxygent only functions to carry gases to and from tissues; it cannot clot or perform any of the immune-system functions that blood does. Still, the substitute has several advantages over real blood. Oxygent has a shelf life of more than a year. Oxygent also eliminates many of the risks associated with blood transfusions. Because the substitute can dissolve larger amounts of oxygen than real blood can, smaller amounts of the mixture are needed.

Oxygent is currently being tested in surgical patients.

"Once this product is approved and has been demonstrated to be safe and effective in elective surgery, I think you will see its use spread into the emergency critical-care arena," says Dr. Keipert. "A patient who has lost a lot of blood and who is currently being resuscitated with normal fluids like saline solutions would be given Oxygent as an additional oxygen-delivery agent in the emergency room."

Questions

1. How would the approval of Oxygent benefit the medical community?

2. How do scientists prevent the nonpolar perfluorocarbons in Oxygent from separating from the water?
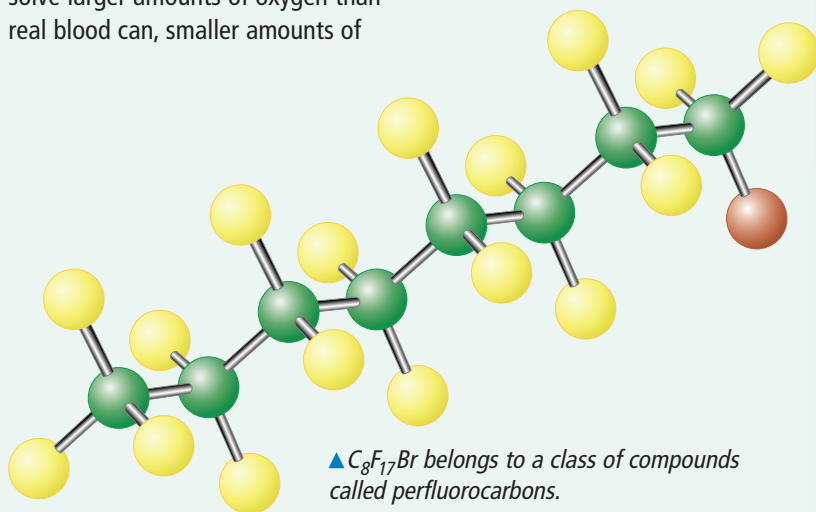
▲ $C_8F_{17}Br$ belongs to a class of compounds called perfluorocarbons.

Concentration of Solutions

OBJECTIVES

● Given the mass of solute and volume of solvent, calculate the concentration of a solution.

● Given the concentration of a solution, determine the amount of solute in a given amount of solution.

● Given the concentration of a solution, determine the amount of solution that contains a given amount of solute.

*T*he **concentration** *of a solution is a measure of the amount of solute in a given amount of solvent or solution.* Some medications are solutions of drugs—a one-teaspoon dose at the correct concentration might cure the patient, while the same dose in the wrong concentration might kill the patient.

In this section, we introduce two different ways of expressing the concentrations of solutions: molarity and molality.

Sometimes, solutions are referred to as "dilute" or "concentrated," but these are not very definite terms. "Dilute" just means that there is a relatively small amount of solute in a solvent. "Concentrated," on the other hand, means that there is a relatively large amount of solute in a solvent. Note that these terms are unrelated to the degree to which a solution is saturated. A saturated solution of a substance that is not very soluble might be very dilute.

Molarity

Molarity *is the number of moles of solute in one liter of solution.* To relate the molarity of a solution to the mass of solute present, you must know the molar mass of the solute. For example, a "one molar" solution of sodium hydroxide, NaOH, contains one mole of NaOH in every liter of solution. The symbol for molarity is M, and the concentration of a one molar solution of sodium hydroxide is written as 1 M NaOH.

One mole of NaOH has a mass of 40.0 g. If this quantity of NaOH is dissolved in enough water to make exactly 1.00 L of solution, the solution is a 1 M solution. If 20.0 g of NaOH, which is 0.500 mol, is dissolved in enough water to make 1.00 L of solution, a 0.500 M NaOH solution is produced. This relationship between molarity, moles, and volume may be expressed in the following ways.

$$\text{molarity (M)} = \frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}}$$

$$= \frac{0.500 \text{ mol NaOH}}{1.00 \text{ L}}$$

$$= 0.500 \text{ M NaOH}$$

If twice the molar mass of NaOH, 80.0 g, is dissolved in enough water to make 1 L of solution, a 2 M solution is produced. The molarity of any solution can be calculated by dividing the number of moles of solute by the number of liters of solution.

Note that a 1 M solution is *not* made by adding 1 mol of solute to 1 L of *solvent*. In such a case, the final total volume of the solution might not be 1 L. Instead, 1 mol of solute is first dissolved in less than 1 L of solvent. Then, the resulting solution is carefully diluted with more solvent to bring the *total volume* to 1 L, as shown in **Figure 17.** The following sample problem will show you how molarity is often used.

FIGURE 17 The preparation of a 0.5000 M solution of $CuSO_4 \cdot 5H_2O$ starts with calculating the mass of solute needed.

1

Start by calculating the mass of $CuSO_4 \cdot 5H_2O$ needed. Making a liter of this solution requires 0.5000 mol of solute. Convert the moles to mass by multiplying by the molar mass of $CuSO_4 \cdot 5H_2O$. This mass is calculated to be 124.8 g.

2

Add some solvent to the solute to dissolve it, and then pour it into a 1.0 L volumetric flask.

3

Rinse the weighing beaker with more solvent to remove all the solute, and pour the rinse into the flask. Add water until the volume of the solution nears the neck of the flask.

4

Put the stopper in the flask, and swirl the solution thoroughly.

5

Carefully fill the flask to the 1.0 L mark with water.

6

Restopper the flask, and invert it at least 10 times to ensure complete mixing.

7

The resulting solution has 0.5000 mol of solute dissolved in 1.000 L of solution, which is a 0.5000 M concentration.

For more help, go to the **Math Tutor** *at the end of this chapter.*

You have 3.50 L of solution that contains 90.0 g of sodium chloride, NaCl.
What is the molarity of that solution?

SOLUTION

1 *ANALYZE*   **Given:** solute mass = 90.0 g NaCl
                       solution volume = 3.50 L
              Unknown: molarity of NaCl solution

2 *PLAN*   Molarity is the number of moles of solute per liter of solution. The solute is described
             in the problem by mass, not the amount in moles. You need one conversion (grams to moles
             of solute) using the inverted molar mass of NaCl to arrive at your answer.

$$\text{grams of solute} \longrightarrow \text{number of moles of solute} \longrightarrow \text{molarity}$$

$$\text{g NaCl} \times \frac{1 \text{ mol NaCl}}{\text{g NaCl}} = \text{mol NaCl}$$

$$\frac{\text{amount of solute (mol)}}{V \text{ solution (L)}} = \text{molarity of solution (M)}$$

3 *COMPUTE*   You will need the molar mass of NaCl.
                NaCl = 58.44 g/mol

$$90.0 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 1.54 \text{ mol NaCl}$$

$$\frac{1.54 \text{ mol NaCl}}{3.50 \text{ L of solution}} = 0.440 \text{ M NaCl}$$

4 *EVALUATE*   Because each factor involved is limited to three significant digits, the answer should
                 have three significant digits, which it does. The units cancel correctly to give the desired
                 moles of solute per liter of solution, which is molarity.

For more help, go to the **Math Tutor** *at the end of this chapter.*

You have 0.8 L of a 0.5 M HCl solution. How many moles of HCl does this solution contain?

SOLUTION

1 *ANALYZE*   **Given:** volume of solution = 0.8 L
                       concentration of solution = 0.5 M HCl
              Unknown: moles of HCl in a given volume

2 *PLAN*   The molarity indicates the moles of solute that are in one liter of solution.
             Given the volume of the solution, the number of moles of solute
             can then be found.

$$\text{concentration (mol of HCl/L of solution)} \times \text{volume (L of solution)} = \text{mol of HCl}$$

3 COMPUTE

$$\frac{0.5 \text{ mol HCl}}{1.0 \text{ L of solution}} \times 0.8 \text{ L of solution} = 0.4 \text{ mol HCl}$$

4 EVALUATE

The answer is correctly given to one significant digit. The units cancel correctly to give the desired unit, mol. There should be less than 0.5 mol HCl, because less than 1 L of solution was used.

SAMPLE PROBLEM C   *For more help, go to the* ***Math Tutor*** *at the end of this chapter.*

To produce 40.0 g of silver chromate, you will need at least 23.4 g of potassium chromate in solution as a reactant. All you have on hand is 5 L of a 6.0 M K_2CrO_4 solution. What volume of the solution is needed to give you the 23.4 g K_2CrO_4 needed for the reaction?

SOLUTION

1 ANALYZE

Given: volume of solution = 5 L
concentration of solution = 6.0 M K_2CrO_4
mass of solute = 23.4 g K_2CrO_4
mass of product = 40.0 g Ag_2CrO_4
Unknown: volume of K_2CrO_4 solution in L

2 PLAN

The molarity indicates the moles of solute that are in 1 L of solution. Given the mass of solute needed, the amount in moles of solute can then be found. Use the molarity and the amount in moles of K_2CrO_4 to determine the volume of K_2CrO_4 that will provide 23.4 g.

$$\text{grams of solute} \longrightarrow \text{moles solute}$$
$$\text{moles solute and molarity} \longrightarrow \text{liters of solution needed}$$

3 COMPUTE

To get the moles of solute, you'll need to calculate the molar mass of K_2CrO_4.

$$1 \text{ mol } K_2CrO_4 = 194.2 \text{ g } K_2CrO_4$$

$$23.4 \text{ g } K_2CrO_4 \times \frac{1 \text{ mol } K_2CrO_4}{194.2 \text{ g } K_2CrO_4} = 0.120 \text{ mol } K_2CrO_4$$

$$6.0 \text{ M } K_2CrO_4 = \frac{0.120 \text{ mol } K_2CrO_4}{x \text{ L } K_2CrO_4 \text{ soln}}$$

$$x = 0.020 \text{ L } K_2CrO_4 \text{ soln}$$

4 EVALUATE

The answer is correctly given to two significant digits. The units cancel correctly to give the desired unit, liters of solution.

PRACTICE   *Answers in Appendix E*

1. What is the molarity of a solution composed of 5.85 g of potassium iodide, KI, dissolved in enough water to make 0.125 L of solution?

2. How many moles of H_2SO_4 are present in 0.500 L of a 0.150 M H_2SO_4 solution?

3. What volume of 3.00 M NaCl is needed for a reaction that requires 146.3 g of NaCl?

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate molarity.

Keyword: HC6SLNX

Molality

Molality *is the concentration of a solution expressed in moles of solute per kilogram of solvent.* A solution that contains 1 mol of solute, sodium hydroxide, NaOH, for example, dissolved in exactly 1 kg of solvent is a "one molal" solution. The symbol for molality is *m,* and the concentration of this solution is written as 1 *m* NaOH.

One mole of NaOH has a molar mass of 40.0 g, so 40.0 g of NaOH dissolved in 1 kg of water results in a one molal NaOH solution. If 20.0 g of NaOH, which is 0.500 mol of NaOH, is dissolved in exactly 1 kg of water, the concentration of the solution is 0.500 *m* NaOH.

$$\text{molality} = \frac{\text{moles solute}}{\text{mass of solvent (kg)}}$$

$$\frac{0.500 \text{ mol NaOH}}{1 \text{ kg H}_2\text{O}} = 0.500 \ m \text{ NaOH}$$

If 80.0 g of sodium hydroxide, which is 2 mol, is dissolved in 1 kg of water, a 2.00 *m* solution of NaOH is produced. The molality of any solution can be found by dividing the number of moles of solute by the mass in kilograms of the solvent in which it is dissolved. Note that if the amount of solvent is expressed in grams, the mass of solvent must be converted to kilograms by multiplying by the following conversion factor.

$$1 \text{ kg}/1000 \text{ g}$$

Figure 18 shows how a 0.5000 *m* solution of $CuSO_4 \cdot 5H_2O$ is prepared, in contrast with the 0.5000 M solution in **Figure 17.**

FIGURE 18 The preparation of a 0.5000 *m* solution of $CuSO_4 \cdot 5H_2O$ also starts with the calculation of the mass of solute needed.

1. Calculate the mass of $CuSO_4 \cdot 5H_2O$ needed. Making this solution will require 0.5000 mol of $CuSO_4 \cdot 5H_2O$ per kilogram of solvent (1000 g). This mass is calculated to be 124.8 g.

2. Add 1.000 kg of solvent to the solute in the beaker. Because the solvent is water, 1.000 kg will equal 1000 mL.

3. Mix thoroughly.

4. The resulting solution has 0.5000 mol of solute dissolved in 1.000 kg of solvent.

Concentrations are expressed as molalities when studying properties of solutions related to vapor pressure and temperature changes. Molality is used because it does not change with changes in temperature. Below is a comparison of the equations for molarity and molality.

$$\text{molarity, M} = \frac{\text{amount of A (mol)}}{\text{volume of solution (L)}}$$

$$\text{molality, } m = \frac{\text{amount of A (mol)}}{\text{mass of solvent (kg)}}$$

SAMPLE PROBLEM D

A solution was prepared by dissolving 17.1 g of sucrose (table sugar, $C_{12}H_{22}O_{11}$) in 125 g of water. Find the molal concentration of this solution.

SOLUTION

1 ANALYZE

Given: solute mass = 17.1 g $C_{12}H_{22}O_{11}$
solvent mass = 125 g H_2O
Unknown: molal concentration

2 PLAN

To find molality, you need moles of solute and kilograms of solvent. The given grams of sucrose must be converted to moles. The mass in grams of solvent must be converted to kilograms.

$$\text{mol } C_{12}H_{22}O_{11} = \frac{\text{g } C_{12}H_{22}O_{11}}{\text{molar mass } C_{12}H_{22}O_{11}}$$

$$\text{kg } H_2O = \text{g } H_2O \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$\text{molality } C_{12}H_{22}O_{11} = \frac{\text{mol } C_{12}H_{22}O_{11}}{\text{kg } H_2O}$$

3 COMPUTE

Use the periodic table to compute the molar mass of $C_{12}H_{22}O_{11}$.
$C_{12}H_{22}O_{11}$ = 342.34 g/mol

$$17.1 \text{ g } C_{12}H_{22}O_{11} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.34 \text{ g } C_{12}H_{22}O_{11}} = 0.0500 \text{ mol } C_{12}H_{22}O_{11}$$

$$\frac{125 \text{ g } H_2O}{1000 \text{ g/kg}} = 0.125 \text{ kg } H_2O$$

$$\frac{0.0500 \text{ mol } C_{12}H_{22}O_{11}}{0.125 \text{ kg } H_2O} = 0.400 \text{ } m \text{ } C_{12}H_{22}O_{11}$$

4 EVALUATE

The answer is correctly given to three significant digits. The unit mol solute/kg solvent is correct for molality.

A solution of iodine, I_2, in carbon tetrachloride, CCl_4, is used when iodine is needed for certain chemical tests. How much iodine must be added to prepare a 0.480 m solution of iodine in CCl_4 if 100.0 g of CCl_4 is used?

SOLUTION

1 ANALYZE

Given: molality of solution = 0.480 m I_2
mass of solvent = 100.0 g CCl_4
Unknown: mass of solute

2 PLAN

Your first step should be to convert the grams of solvent to kilograms. The molality gives you the moles of solute, which can be converted to the grams of solute using the molar mass of I_2.

3 COMPUTE

Use the periodic table to compute the molar mass of I_2.
I_2 = 253.8 g/mol

$$100.0 \text{ g } CCl_4 \times \frac{1 \text{ kg}}{1000 \text{ g } CCl_4} = 0.100 \text{ kg } CCl_4$$

$$0.480 \text{ } m = \frac{x \text{ mol } I_2}{0.1 \text{ kg } H_2O} \qquad x = 0.0480 \text{ mol } I_2$$

$$0.0480 \text{ mol } I_2 \times \frac{253.8 \text{ g } I_2}{\text{mol } I_2} = 12.2 \text{ g } I_2$$

4 EVALUATE

The answer has three significant digits and the units for mass of I_2.

PRACTICE

Answers in Appendix E

1. What is the molality of acetone in a solution composed of 255 g of acetone, $(CH_3)_2CO$, dissolved in 200. g of water?

2. What quantity, in grams, of methanol, CH_3OH, is required to prepare a 0.244 m solution in 400. g of water?

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate molality.

Keyword: HC6SLNX

SECTION REVIEW

1. What quantity represents the ratio of the number of moles of solute for a given volume of solution?

2. We dissolve 5.00 grams of sugar, $C_{12}H_{22}O_{11}$, in water to make 1.000 L of solution. What is the concentration of this solution expressed as a molarity?

Critical Thinking

3. **ANALYZING DATA** You evaporate all of the water from 100. mL of NaCl solution and obtain 11.3 grams of NaCl. What was the molarity of the NaCl solution?

4. **RELATING IDEAS** Suppose you know the molarity of a solution. What additional information would you need to calculate the molality of the solution?

CHAPTER HIGHLIGHTS

Types of Mixtures

Vocabulary

soluble
solution
solvent
solute
suspension
colloid
electrolyte
nonelectrolyte

- Solutions are homogeneous mixtures.
- Mixtures are classified as solutions, suspensions, or colloids, depending on the size of the solute particles in the mixture.
- The dissolved substance is the solute. Solutions that have water as a solvent are aqueous solutions.
- Solutions can consist of solutes and solvents that are solids, liquids, or gases.
- Suspensions settle out upon standing. Colloids do not settle out, and they scatter light that is shined through them.
- Most ionic solutes and some molecular solutes form aqueous solutions that conduct an electric current. These solutes are called electrolytes.
- Nonelectrolytes are solutes that dissolve in water to form solutions that do not conduct.

The Solution Process

Vocabulary

solution equilibrium
saturated solution
unsaturated solution
supersaturated solution
solubility
hydration
immiscible
miscible
Henry's law
effervescence
solvated
enthalpy of solution

- A solute dissolves at a rate that depends on the surface area of the solute, how vigorously the solution is mixed, and the temperature of the solvent.
- The solubility of a substance indicates how much of that substance will dissolve in a specified amount of solvent under certain conditions.
- The solubility of a substance depends on the temperature.
- The solubility of gases in liquids increases with increases in pressure.
- The solubility of gases in liquids decreases with increases in temperature.
- The overall energy absorbed as heat by the system when a specified amount of solute dissolved during solution formation is called the enthalpy of solution.

Concentration of Solutions

Vocabulary

concentration
molarity
molality

- Two useful expressions of concentration are molarity and molality.
- The molar concentration of a solution represents the ratio of moles of solute to liters of solution.
- The molal concentration of a solution represents the ratio of moles of solute to kilograms of solvent.

For more practice, go to the Problem Bank in Appendix D.

Types of Mixtures

SECTION 1 REVIEW

1. a. What is the Tyndall effect?
   b. Identify one example of this effect.
2. Given an unknown mixture consisting of two or more substances, explain how we could determine whether that mixture is a true solution, a colloid, or a suspension.
3. Explain why a suspension is considered a heterogeneous mixture.
4. Does a solution have to involve a liquid? Explain your answer.
5. What is the difference between an electrolyte and a nonelectrolyte?

The Solution Process

SECTION 2 REVIEW

6. a. What is solution equilibrium?
   b. What factors determine the point at which a given solute-solvent combination reaches equilibrium?
7. a. What is a saturated solution?
   b. What visible evidence indicates that a solution is saturated?
   c. What is an unsaturated solution?
8. a. What is meant by the solubility of a substance?
   b. What condition(s) must be specified when expressing the solubility of a substance?
9. a. What rule of thumb is useful for predicting whether one substance will dissolve in another?
   b. Describe what the rule means in terms of various combinations of polar and nonpolar solutes and solvents.
10. a. How does pressure affect the solubility of a gas in a liquid?
    b. What law is a statement of this relationship?
    c. If the pressure of a gas above a liquid is increased, what happens to the amount of the gas that will dissolve in the liquid, if all other conditions remain constant?

d. Two bottles of soda are opened. One is a cold bottle and the other is at room temperature. Which system would show more effervescence and why?

11. Based on **Figure 15,** determine the solubility of each of the following in grams of solute per 100. g H_2O.
    a. $NaNO_3$ at 10°C
    b. KNO_3 at 60°C
    c. NaCl at 50°C
12. Based on **Figure 15,** at what temperature would each of the following solubility levels be observed?
    a. 50 g KCl in 100 g H_2O
    b. 100 g $NaNO_3$ in 100 g H_2O
    c. 60 g KNO_3 in 100 g H_2O
13. The enthalpy of solution for $AgNO_3$ is +22.8 kJ/mol.
    a. Write the equation that represents the dissolution of $AgNO_3$ in water.
    b. Is the dissolution process endothermic or exothermic? Is the crystallization process endothermic or exothermic?
    c. As $AgNO_3$ dissolves, what change occurs in the temperature of the solution?
    d. When the system is at equilibrium, how do the rates of dissolution and crystallization compare?
    e. If the solution is then heated, how will the rates of dissolution and crystallization be affected? Why?
    f. How will the increased temperature affect the amount of solute that can be dissolved?
    g. If the solution is allowed to reach equilibrium and is then cooled, how will the system be affected?
14. What opposing forces are at equilibrium in the sodium chloride system shown in **Figure 7?**

Concentration of Solutions

SECTION 3 REVIEW

15. On which property of solutions does the concept of concentration rely?

16. In what units is molarity expressed?

17. Under what circumstances might we prefer to express solution concentrations in terms of
   a. molarity?
   b. molality?

18. If you dissolve 2.00 mol KI in 1.00 L of water, will you get a 2.00 M solution? Explain.

PRACTICE PROBLEMS

19. a. Suppose you wanted to dissolve 106 g of Na_2CO_3 in enough H_2O to make 6.00 L of solution.
      (1) What is the molar mass of Na_2CO_3?
      (2) What is the molarity of this solution?
   b. What is the molarity of a solution of 14.0 g NH_4Br in enough H_2O to make 150 mL of solution?

20. a. Suppose you wanted to produce 1.00 L of a 3.50 M aqueous solution of H_2SO_4.
      (1) What is the solute?
      (2) What is the solvent?
      (3) How many grams of solute are needed to make this solution?
   b. How many grams of solute are needed to make 2.50 L of a 1.75 M solution of $Ba(NO_3)_2$?

21. How many moles of NaOH are contained in 65.0 mL of a 2.20 M solution of NaOH in H_2O? (Hint: See Sample Problem B.)

22. A solution is made by dissolving 26.42 g of $(NH_4)_2SO_4$ in enough H_2O to make 50.00 mL of solution.
   a. What is the molar mass of $(NH_4)_2SO_4$?
   b. What is the molarity of this solution?

23. Suppose you wanted to find out how many milliliters of 1.0 M $AgNO_3$ are needed to provide 169.9 g of pure $AgNO_3$.
   a. What is step 1 in solving the problem?
   b. What is the molar mass of $AgNO_3$?
   c. How many milliliters of solution are needed?

24. a. Balance the equation:
      $$H_3PO_4 + Ca(OH)_2 \longrightarrow Ca_3(PO_4)_2 + H_2O$$
   b. What mass of each product results if 750 mL of 6.00 M H_3PO_4 reacts according to the equation?

25. How many milliliters of 0.750 M H_3PO_4 are required to react with 250. mL of 0.150 M $Ba(OH)_2$ if the products are barium phosphate and water?

26. 75.0 mL of an $AgNO_3$ solution reacts with enough Cu to produce 0.250 g of Ag by single displacement. What is the molarity of the initial $AgNO_3$ solution if $Cu(NO_3)_2$ is the other product?

27. Determine the number of grams of solute needed to make each of the following molal solutions:
   a. a 4.50 m solution of H_2SO_4 in 1.00 kg H_2O
   b. a 1.00 m solution of HNO_3 in 2.00 kg H_2O

28. A solution is prepared by dissolving 17.1 g of sucrose, $C_{12}H_{22}O_{11}$, in 275 g of H_2O.
   a. What is the molar mass of sucrose?
   b. What is the molality of that solution?

29. How many kilograms of H_2O must be added to 75.5 g of $Ca(NO_3)_2$ to form a 0.500 m solution?

30. A solution made from ethanol, C_2H_5OH, and water is 1.75 m in ethanol. How many grams of C_2H_5OH are contained per 250. g of water?

MIXED REVIEW

31. Na_2SO_4 is dissolved in water to make 450. mL of a 0.250 M solution.
   a. What is the molar mass of Na_2SO_4?
   b. How many moles of Na_2SO_4 are needed?

32. Citric acid is one component of some soft drinks. Suppose that 2.00 L of solution are made from 150. mg of citric acid, $C_6H_8O_7$.
   a. What is the molar mass of citric acid?
   b. What is the molarity of citric acid in the solution?

33. Suppose you wanted to know how many grams of KCl would be left if 350 mL of a 2.0 M KCl solution were evaporated to dryness.
   a. What is the molar mass of KCl?
   b. How would heating the solution affect the mass of KCl remaining?
   c. How many grams of KCl would remain?

34. Sodium metal reacts violently with water to form NaOH and release hydrogen gas. Suppose that 10.0 g of Na react completely with 1.00 L of water and the final solution volume is 1.00 L.
  a. What is the molar mass of NaOH?
  b. Write a balanced equation for the reaction.
  c. What is the molarity of the NaOH solution formed by the reaction?

35. In cars, ethylene glycol, $C_2H_6O_2$, is used as a coolant and antifreeze. A mechanic fills a radiator with 6.5 kg of ethylene glycol and 1.5 kg of water.
  a. What is the molar mass of ethylene glycol?
  b. What is the molality of the water in the solution?

36. Plot a solubility graph for $AgNO_3$ from the following data, with grams of solute (by increments of 50) per 100 g of H_2O on the vertical axis and with temperature in °C on the horizontal axis.

Grams solute per 100 g H_2O	Temperature (°C)
122	0
216	30
311	40
440	60
585	80
733	100

  a. How does the solubility of $AgNO_3$ vary with the temperature of the water?
  b. Estimate the solubility of $AgNO_3$ at 35°C, 55°C, and 75°C.
  c. At what temperature would the solubility of $AgNO_3$ be 275 g per 100 g of H_2O?
  d. If 100 g of $AgNO_3$ were added to 100 g of H_2O at 10°C, would the resulting solution be saturated or unsaturated? What would occur if 325 g of $AgNO_3$ were added to 100 g of H_2O at 35°C?

37. If a saturated solution of KNO_3 in 100. g of H_2O at 60°C is cooled to 20°C, approximately how many grams of the solute will precipitate out of the solution? (Use **Table 4.**)

38. a. Suppose you wanted to dissolve 294.3 g of H_2SO_4 in 1.000 kg of H_2O.
    (1) What is the solute?
    (2) What is the solvent?
    (3) What is the molality of this solution?
  b. What is the molality of a solution of 63.0 g HNO_3 in 0.250 kg H_2O?

CRITICAL THINKING

39. Predicting Outcomes You have been investigating the nature of suspensions, colloids, and solutions and have collected the following observational data on four unknown samples. From the data, infer whether each sample is a solution, suspension, or colloid.

DATA TABLE 1 *Samples*

Sample	Color	Clarity (clear or cloudy)	Settle out	Tyndall effect
1	green	clear	no	no
2	blue	cloudy	yes	no
3	colorless	clear	no	yes
4	white	cloudy	no	yes

Based on your inferences in Data Table 1, you decide to conduct one more test of the particles. You filter the samples and then reexamine the filtrate. You obtain the data found in Data Table 2. Infer the classifications of the filtrate based on the data in Data Table 2.

DATA TABLE 2 *Filtrate of Samples*

Sample	Color	Clarity (clear or cloudy)	On filter paper	Tyndall effect
1	green	clear	nothing	no
2	blue	cloudy	gray solid	yes
3	colorless	clear	none	yes
4	colorless	clear	white solid	no

USING THE HANDBOOK

40. Review the information on alloys in the Elements Handbook.
   a. Why is aluminum such an important component of alloys?
   b. What metals make up bronze?
   c. What metals make up brass?
   d. What is steel?
   e. What is the composition of the mixture called *cast iron?*

41. **Table 5A** of the *Elements Handbook* contains carbon monoxide concentration data expressed as parts per million (ppm). The OSHA (Occupational Safety and Health Administration) limit for worker exposure to CO is 200 ppm for an eight-hour period.
   a. At what concentration do harmful effects occur in less than one hour?
   b. By what factor does the concentration in item (a) exceed the maximum limit set by OSHA?

RESEARCH & WRITING

42. Find out about the chemistry of emulsifying agents. How do these substances affect the dissolution of immiscible substances such as oil and water? As part of your research on this topic, find out why eggs are an emulsifying agent for baking mixtures.

ALTERNATIVE ASSESSMENT

43. Make a comparison of the electrolyte concentration in various brands of sports drinks. Using the labeling information for sugar, calculate the molarity of sugar in each product or brand. Construct a poster to show the results of your analysis of the product labels.

44. Write a set of instructions on how to prepare a solution that is 1 M $CuSO_4$ using $CuSO_4 \cdot 5H_2O$ as the solute. How do the instructions differ if the solute is anhydrous $CuSO_4$? Your instructions should include a list of all materials needed.

extension

Graphing Calculator
Solubility vs. Temperature
Go to **go.hrw.com** for a graphing calculator exercise that asks you to graph the solubility versus temperatures for NaCl and KNO_3.

Keyword: HC6SLNX

You can use the relationship below to calculate the concentration in molarity of any solution.

$$\text{molarity of solution (M)} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

Suppose you dissolve 20.00 g of NaOH in some water and dilute the solution to a volume of 250.0 mL (0.2500 L). You don't know the molarity of this solution until you know how many moles of NaOH were dissolved. You know that the number of moles of a substance can be found by dividing the mass of the substance by the mass of 1 mol (molar mass) of the substance. The molar mass of NaOH is 40.00, so the number of moles of NaOH dissolved is

$$20.00 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 0.5000 \text{ mol NaOH}$$

Now you know that the solution has 0.5000 mol NaOH dissolved in 0.2500 L of solution, so you can calculate molarity.

$$\text{molarity of NaOH} \times \frac{\text{mol NaOH}}{\text{L solution}} = \frac{0.5000 \text{ mol NaOH}}{0.2500 \text{ L solution}} = 2.000 \text{ mol/L} = 2.000 \text{ M NaOH}$$

Problem-Solving TIPS

- Remember that balances measure mass, not moles, so you often have to convert between mass and moles of solute when making or using solutions.

SAMPLE

A 0.5000 L volume of a solution contains 36.49 g of magnesium chloride, MgCl$_2$. What is the molarity of the solution?

You know the volume of the solution, but you need to find the number of moles of the solute MgCl$_2$ by the following conversion.

$$\text{mass MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{\text{molar mass MgCl}_2} = \text{mol MgCl}_2$$

$$36.49 \text{ g MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{95.20 \text{ g MgCl}_2} = 0.3833 \text{ mol MgCl}_2$$

Now you can calculate mol MgCl$_2$ per liter of solution (molarity).

$$\frac{0.3833 \text{ mol MgCl}_2}{0.5000 \text{ L solution}} = 0.7666 \text{ M MgCl}_2$$

PRACTICE PROBLEMS

1. What is the molarity of a solution that contains 0.0350 mol of sodium sulfate, Na$_2$SO$_4$, dissolved in 50.0 mL of solution?

2. What is the molarity of a solution that contains 45.00 g of cadmium nitrate, Cd(NO$_3$)$_2$, dissolved in 400.0 mL of solution?

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Water is an excellent solvent because
   A. it is a covalent compound.
   B. it is a nonconductor of electricity.
   C. its molecules are quite polar.
   D. it is a clear, colorless liquid.

2. Two liquids are likely to be immiscible if
   A. both have polar molecules.
   B. both have nonpolar molecules.
   C. one is polar and the other is nonpolar.
   D. one is water and the other is methyl alcohol, CH_3OH.

3. The solubility of a gas in a liquid would be increased by an
   A. addition of an electrolyte.
   B. addition of an emulsifier.
   C. agitation of the solution.
   D. increase in its partial pressure.

4. Which of the following types of compounds is most likely to be a strong electrolyte?
   A. a polar compound
   B. a nonpolar compound
   C. a covalent compound
   D. an ionic compound

5. A saturated solution can become supersaturated under which of the following conditions?
   A. It contains electrolytes.
   B. The solution is heated and then allowed to cool.
   C. More solvent is added.
   D. More solute is added.

6. Molarity is expressed in units of
   A. moles of solute per liter of solution.
   B. liters of solution per mole of solute.
   C. moles of solute per liter of solvent.
   D. liters of solvent per mole of solute.

7. What mass of NaOH is contained in 2.5 L of a 0.010 M solution?
   A. 0.010 g   **C.** 2.5 g
   B. 1.0 g   **D.** 0.40 g

8. Which one of the following statements is false?
   A. Gases are generally more soluble in water under high pressures than under low pressures.
   B. As temperature increases, the solubilities of some solids in water increase and the solubilities of other solids in water decrease.
   C. Water dissolves many ionic solutes because of its ability to hydrate ions in solution.
   D. Many solids dissolve more quickly in a cold solvent than in a warm solvent.

SHORT ANSWER

9. Several experiments are carried out to determine the solubility of cadmium iodide, CdI_2, in water. In each experiment, a measured mass of CdI_2 is added to 100 g of water at 25°C and the mixture is stirred. Any undissolved CdI_2 is then filtered off and dried, and its mass is determined. Results for several such experiments are shown in the table below. What is the solubility of CdI_2 in water at this temperature?

Mass of CdI_2 added, g	Mass of undissolved CdI_2 recovered, g
17.9	0.0
38.2	0.0
53.6	0.0
79.3	0.0
93.6	7.4
104.3	18.1

EXTENDED RESPONSE

10. Explain why oil and water do not mix.

11. Write a set of instructions on how to prepare a solution that is 0.100 M KBr, using solid KBr (molar mass 119 g/mol) as the solute. Your instructions should include a list of all materials and equipment needed.

Test TIP Allow a few minutes at the end of the test-taking period to check for careless mistakes, such as marking two answers for a single question.

Separation of Pen Inks by Paper Chromatography

OBJECTIVES

- *Demonstrate* proficiency in qualitatively separating mixtures using paper chromatography.

- *Determine* the R_f factor(s) for each component of each tested ink.

- *Explain* how the inks are separated by paper chromatography.

- *Observe* the separation of a mixture by the method of paper chromatography.

MATERIALS

- **12 cm circular chromatography paper or filter paper**

- **distilled water**

- **filter paper wick, 2 cm equilateral triangle**

- **isopropanol**

- **numbered pens, each with a different black ink, 4**

- **pencil**

- **petri dish with lid**

- **scissors**

FIGURE A Paper chromatography reveals the different colored dyes that black ink contains.

BACKGROUND

Paper Chromatography
Details on this technique can be found in the Pre-Laboratory Procedure "Paper Chromatography" on page 848.

Writing Inks
Most ballpoint pen inks are complex mixtures, containing pigments or dyes that can be separated by paper chromatography, as shown in **Figure A.**

Black inks can contain three or more colors; the number of colors depends on the manufacturer. Each ink formulation has a characteristic pattern that uniquely identifies it.

In this experiment you will develop radial paper chromatograms for four black ballpoint pen inks, using water as solvent. You will then repeat this process using isopropanol as the solvent. You will then measure the distance traveled by each of the individual ink components and the distance traveled by the solvent front. Finally, you will use these measurements to calculate the R_f factor for each component.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Determine the formula, structure, polarity, density, and volatility at room temperature for water and isopropanol. The following titles are sources that provide general information on specific elements and compounds: *CRC Handbook of Chemistry and Physics, McGraw-Hill Dictionary of Chemical Terms,* and *Merck Index.*

2. Prepare two data tables, one for the chromatogram made with water and one for the chromatogram made with isopropanol. Record the pen number you are testing. In your data tables, create columns for each color in which to record the distance and R_f value. Your teacher can provide you with sample data tables. Leave room below each data table to record the distance that the solvent reaches.

PROCEDURE

Part A: Prepare a chromatogram using water as the solvent

1. Construct an apparatus for paper chromatography as described in the Pre-Laboratory Procedure on page 848. You will make only four dots. You will use ballpoint pens rather than micropipets to spot your paper.

2. After 15 min or when the water is about 1 cm from the outside edge of the paper, remove the paper from the Petri dish and allow the chromatogram to dry. Record in the data table the colors that have separated from each of the four different black inks.

Part B: Prepare a chromatogram using isopropanol as the solvent

3. Repeat Procedure steps 1 to 2, replacing the water in the Petri dish with isopropanol.

Part C: Determine R_f values for each component

4. After the chromatogram is dry, use a pencil to mark the point where the solvent front stopped.

5. With a ruler, measure the distance from the initial ink spot to your mark, and record this distance on your data table.

6. Make a small dot with your pencil in the center of each color band.

7. With a ruler, measure the distance from the initial ink spot to each dot separately, and record each distance on your data table.

8. Divide each value recorded in Procedure step 7 by the value recorded in Procedure step 5. The result is the R_f value for that component. Record the R_f values in your data table. Tape or staple the chromatogram to your data table.

CLEANUP AND DISPOSAL

9. The water may be poured down the sink. Chromatograms and other pieces of filter paper may be discarded in the trash. The isopropanol solution should be placed in the waste disposal container designated by your teacher. Clean up your equipment and lab station. Thoroughly wash your hands after completing the lab session and cleanup.

ANALYSIS AND INTERPRETATION

1. Evaluating Conclusions: Is the color in each pen the result of a single dye or multiple dyes? Justify your answer.

2. Relating Ideas: What can be said about the properties of a component ink that has an R_f value of 0.50?

3. Analyzing Methods: Suggest a reason for stopping the process when the solvent front is 1 cm from the edge of the filter paper rather than when it is even with the edge of the paper.

4. Predicting Outcomes: Predict the results of forgetting to remove the chromatogram from the water in the petri dish until the next day.

CONCLUSIONS

1. Analyzing Results: Compare the R_f values for the colors from pen number 2 when water was the solvent and the R_f values obtained when isopropanol was the solvent. Explain why they differ.

2. Evaluating Methods: Would you consider isopropanol a better choice for the solvent than water? Why or why not?

3. Analyzing Conclusions: Are the properties of the component that traveled the farthest in the water chromatogram likely to be similar to the properties of the component that traveled the farthest in the isopropanol chromatogram? Explain your reasoning.

4. Inferring Conclusions: What can you conclude about the composition of the inks in ballpoint pens from your chromatogram?

Ions in Aqueous Solutions and Colligative Properties

These formations were made by the precipitation of ionic compounds from an aqueous solution.

Cyanotrichite Crystals

# *Compounds in Aqueous Solutions*

OBJECTIVES

● Write equations for the dissolution of soluble ionic compounds in water.

● Predict whether a precipitate will form when solutions of soluble ionic compounds are combined, and write net ionic equations for precipitation reactions.

● Compare dissociation of ionic compounds with ionization of molecular compounds.

● Draw the structure of the hydronium ion, and explain why it is used to represent the hydrogen ion in solution.

● Distinguish between strong electrolytes and weak electrolytes.

As you have learned, solid compounds can be ionic or molecular. In an ionic solid, a crystal structure is made up of charged particles held together by ionic attractions. In a molecular solid, molecules are composed of covalently bonded atoms. The solid is held together by non-covalent, intermolecular forces. When they dissolve in water, ionic compounds and molecular compounds behave differently.

Dissociation

When a compound that is made of ions dissolves in water, the ions separate from one another, as shown in **Figure 1.** *This separation of ions that occurs when an ionic compound dissolves is called* **dissociation.** For example, dissociation of sodium chloride and calcium chloride in water can be represented by the following equations. (As usual, (*s*) indicates a solid species, and (*aq*) indicates a species in an aqueous solution. Note that each equation is balanced for charge as well as for atoms.)

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

$$CaCl_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2Cl^-(aq)$$
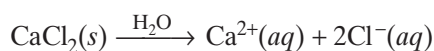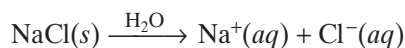
Notice the number of ions produced per formula unit in the equations above. One formula unit of sodium chloride gives two ions in solution, whereas one formula unit of calcium chloride gives three ions in solution.
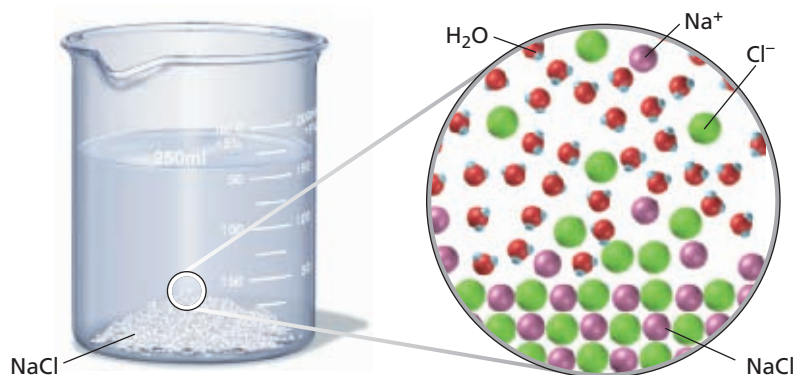

FIGURE 1 When NaCl dissolves in water, the ions separate as they leave the crystal.

Assuming 100% dissociation, a solution that contains 1 mol of sodium chloride contains 1 mol of Na^+ ions and 1 mol of Cl^- ions. In this book, you can assume 100% dissociation for all soluble ionic compounds. The dissociation of NaCl can be represented as follows.

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$
$$\text{1 mol} \qquad\qquad \text{1 mol} \quad\;\; \text{1 mol}$$
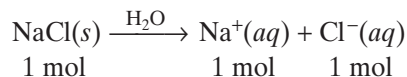
A solution that contains 1 mol of calcium chloride contains 1 mol of Ca^{2+} ions and 2 mol of Cl^- ions—a total of 3 mol of ions.

$$CaCl_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2Cl^-(aq)$$
$$\text{1 mol} \qquad\qquad \text{1 mol} \qquad \text{2 mol}$$
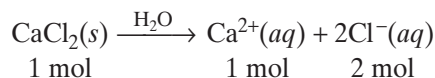
SAMPLE PROBLEM A

Write the equation for the dissolution of aluminum sulfate, $Al_2(SO_4)_3$, in water. How many moles of aluminum ions and sulfate ions are produced by dissolving 1 mol of aluminum sulfate? What is the total number of moles of ions produced by dissolving 1 mol of aluminum sulfate?

SOLUTION

1 ANALYZE

Given: amount of solute = 1 mol $Al_2(SO_4)_3$
solvent identity = water
Unknown: **a.** moles of aluminum ions and sulfate ions
b. total number of moles of solute ions produced

2 PLAN

The coefficients in the balanced dissociation equation will reveal the mole relationships, so you can use the equation to determine the number of moles of solute ions produced.

$$Al_2(SO_4)_3(s) \xrightarrow{H_2O} 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$$

3 COMPUTE

a. 1 mol $Al_2(SO_4)_3 \rightarrow$ 2 mol Al^{3+} + 3 mol SO_4^{2-}
b. 2 mol Al^{3+} + 3 mol SO_4^{2-} = 5 mol of solute ions

4 EVALUATE

The equation is correctly balanced. Because one formula unit of $Al_2(SO_4)_3$ produces 5 ions, 1 mol of $Al_2(SO_4)_3$ produces 5 mol of ions.

PRACTICE

Answers in Appendix E

1. Write the equation for the dissolution of each of the following in water, and then determine the number of moles of each ion produced as well as the total number of moles of ions produced.
   a. 1 mol ammonium chloride
   b. 1 mol sodium sulfide
   c. 0.5 mol barium nitrate
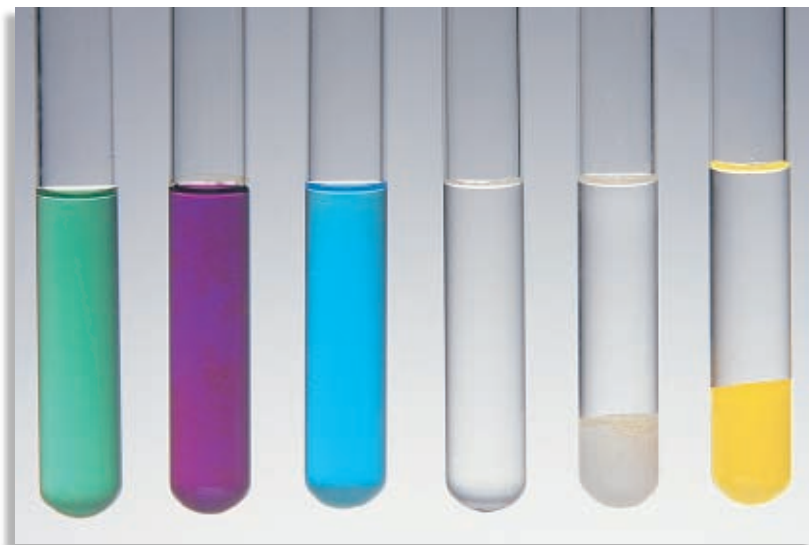
Precipitation Reactions

Although no ionic compound is completely insoluble, compounds of very low solubility can be considered insoluble for most practical purposes. Some examples of ionic compounds that are soluble and insoluble in water are shown in **Figure 2.** It is difficult to write solubility rules that cover all possible conditions. However, we can write some general guidelines to help predict whether a compound made of a certain combination of ions is soluble. These general solubility guidelines are given in **Table 1.**

By looking at the table you can tell that most sodium compounds are soluble. Sodium carbonate, Na_2CO_3, is soluble because it contains sodium. Its dissociation equation is as follows.

$$Na_2CO_3(s) \xrightarrow{H_2O} 2Na^+(aq) + CO_3^{2-}(aq)$$

TABLE 1 *General Solubility Guidelines*
1. Sodium, potassium, and ammonium compounds are soluble in water.
2. Nitrates, acetates, and chlorates are soluble.
3. Most chlorides are soluble, except those of silver, mercury(I), and lead. Lead(II) chloride is soluble in hot water.
4. Most sulfates are soluble, except those of barium, strontium, lead, calcium, and mercury.
5. Most carbonates, phosphates, and silicates are insoluble, except those of sodium, potassium, and ammonium.
6. Most sulfides are insoluble, except those of calcium, strontium, sodium, potassium, and ammonium.

Is calcium phosphate, $Ca_3(PO_4)_2$, soluble or insoluble? According to **Table 1,** most phosphates are insoluble. Calcium phosphate is not one of the exceptions listed, so it is insoluble. Dissociation equations cannot be written for insoluble compounds.

The information in **Table 1** is also useful in predicting what will happen if solutions of two different soluble compounds are mixed. If the mixing results in a combination of ions that forms an insoluble compound, a double-displacement reaction and precipitation will occur. Precipitation occurs when the attraction between the ions is greater than the attraction between the ions and surrounding water molecules.

Will a precipitate form when solutions of ammonium sulfide and cadmium nitrate are combined? By using the table, you can tell that cadmium nitrate, $Cd(NO_3)_2$, is soluble because it is a nitrate and all nitrates are soluble. You can also tell that ammonium sulfide, $(NH_4)_2S$, is soluble. It is one of the sulfides listed in the table as being soluble. Their dissociation equations are as follows.

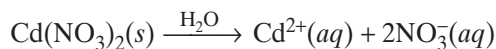$$(NH_4)_2S(s) \xrightarrow{H_2O} 2NH_4^+(aq) + S^{2-}(aq)$$

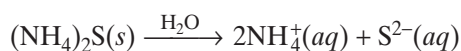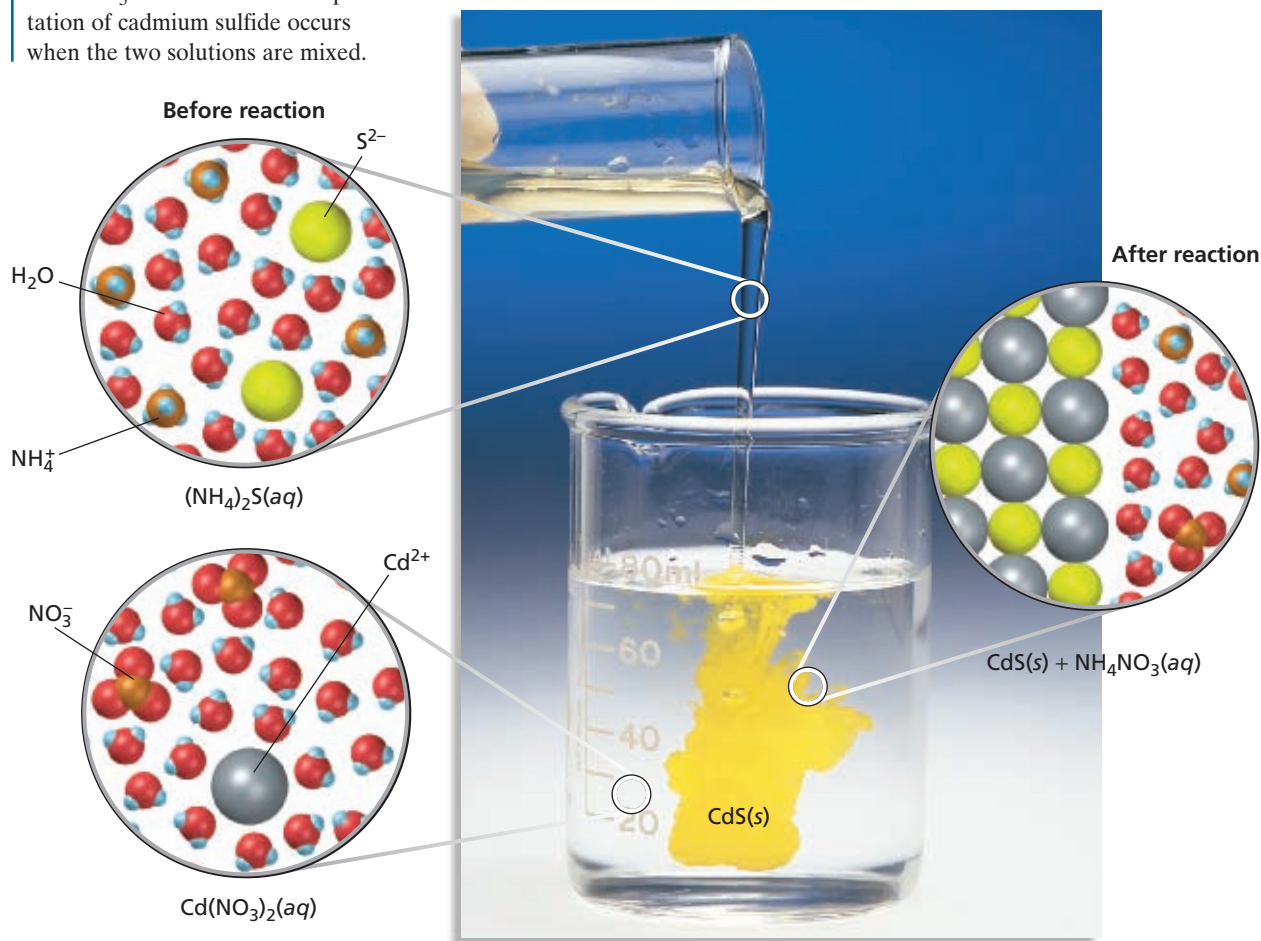$$Cd(NO_3)_2(s) \xrightarrow{H_2O} Cd^{2+}(aq) + 2NO_3^-(aq)$$

FIGURE 3 Ammonium sulfide is a soluble compound that dissociates in water to form NH_4^+ and S^{2-} ions. Cadmium nitrate is a soluble compound that dissociates in water to form NO_3^- and Cd^{2+} ions. Precipitation of cadmium sulfide occurs when the two solutions are mixed.

Before reaction

S^{2-}

H_2O

NH_4^+

$(NH_4)_2S(aq)$

Cd^{2+}

NO_3^-

$Cd(NO_3)_2(aq)$

After reaction

$CdS(s) + NH_4NO_3(aq)$

$CdS(s)$

The two possible products of a double-displacement reaction between $(NH_4)_2S$ and $Cd(NO_3)_2$ are ammonium nitrate, NH_4NO_3, and cadmium sulfide, CdS. The question marks indicate that the states are unknown.

$$(NH_4)_2S(aq) + Cd(NO_3)_2(aq) \longrightarrow 2NH_4NO_3(?) + CdS(?)$$
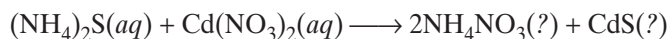
To decide whether a precipitate can form, you must know the solubilities of these two compounds. Consulting **Table 1,** you can see that NH_4NO_3 is soluble in water. However, CdS is insoluble. You can therefore predict that when solutions of ammonium sulfide and cadmium nitrate are combined, ammonium nitrate will not precipitate and cadmium sulfide will. As illustrated in **Figure 3,** crystals of CdS form when the solutions are mixed. In the following equation, the designations (aq) and (s) show that $NH_4NO_3(aq)$ remains in solution and CdS(s) precipitates.

$$(NH_4)_2S(aq) + Cd(NO_3)_2(aq) \longrightarrow 2NH_4NO_3(aq) + CdS(s)$$
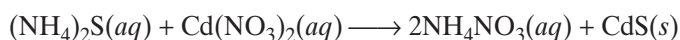
Net Ionic Equations

Reactions of ions in aqueous solution are usually represented by net ionic equations rather than formula equations. *A **net ionic equation** includes only those compounds and ions that undergo a chemical change in a reaction in an aqueous solution.* To write a net ionic equation, you first convert the chemical equation into an overall ionic equation. All soluble ionic compounds are shown as dissociated ions in solution. The precipitates are shown as solids. The precipitation of cadmium sulfide described previously can be shown by the following overall ionic equation.

$$Cd^{2+}(aq) + 2NO_3^-(aq) + 2NH_4^+(aq) + S^{2-}(aq) \longrightarrow$$
$$CdS(s) + 2NO_3^-(aq) + 2NH_4^+(aq)$$
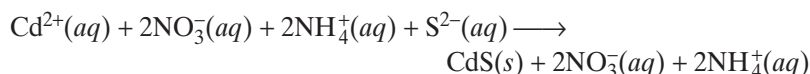
Notice that the aqueous ammonium ion, $NH_4^+(aq)$, and the aqueous nitrate ion, $NO_3^-(aq)$, appear on both sides of this equation. Therefore, they have not undergone any chemical change and are still present in their original form. *Ions that do not take part in a chemical reaction and are found in solution both before and after the reaction are **spectator ions.***

To convert an ionic equation into a net ionic equation, the spectator ions are canceled on both sides of the equation. Eliminating the NH_4^+ and NO_3^- ions from the overall ionic equation above gives the following net ionic equation.

$$Cd^{2+}(aq) + S^{2-}(aq) \longrightarrow CdS(s)$$
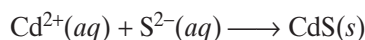
This net ionic equation applies not only to the reaction between $(NH_4)_2S$ and $Cd(NO_3)_2$ but also to *any* reaction in which a precipitate of cadmium sulfide forms when the ions are combined in solution. For example, it is also the net ionic equation for the precipitation of CdS when $CdSO_4$ and H_2S react.

*For more help, go to the **Math Tutor** at the end of Chapter 14.*

Identify the precipitate that forms when aqueous solutions of zinc nitrate and ammonium sulfide are combined. Write the equation for the possible double-displacement reaction. Then write the formula equation, overall ionic equation, and net ionic equation for the reaction.

SOLUTION

1 *ANALYZE*   **Given:** identity of reactants: zinc nitrate and ammonium sulfide
   reaction medium: aqueous solution

Unknown: **a.** equation for the possible double-displacement reaction **b.** identity of the
              precipitate **c.** formula equation **d.** overall ionic equation **e.** net ionic equation

2 *PLAN*   Write the possible double-displacement reaction between $Zn(NO_3)_2$ and $(NH_4)_2S$. Use
Table 1 to determine if any of the products are insoluble and will precipitate. Write a formula equation and an overall ionic equation, and then cancel the spectator ions to produce a net ionic equation.
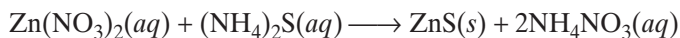
3 *COMPUTE*   **a.** The equation for the possible double-displacement reaction is as follows.

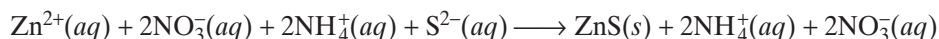$$Zn(NO_3)_2(aq) + (NH_4)_2S(aq) \longrightarrow ZnS(?) + 2NH_4NO_3(?)$$
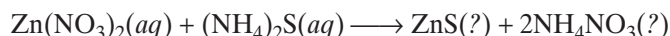
b. Table 1 reveals that zinc sulfide is not a soluble sulfide and is therefore a precipitate.
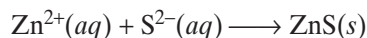Ammonium nitrate is soluble according to the table.

c. The formula equation is as follows.

$$Zn(NO_3)_2(aq) + (NH_4)_2S(aq) \longrightarrow ZnS(s) + 2NH_4NO_3(aq)$$

d. The overall ionic equation is as follows.

$$Zn^{2+}(aq) + 2NO_3^-(aq) + 2NH_4^+(aq) + S^{2-}(aq) \longrightarrow ZnS(s) + 2NH_4^+(aq) + 2NO_3^-(aq)$$

e. The ammonium and nitrate ions appear on both sides of the equation as spectator ions.
The net ionic equation is as follows.

$$Zn^{2+}(aq) + S^{2-}(aq) \longrightarrow ZnS(s)$$

PRACTICE   *Answers in Appendix E*

1. Will a precipitate form if solutions of potassium sulfate and barium nitrate are combined? If so, write the net ionic equation for the reaction.

2. Will a precipitate form if solutions of potassium nitrate and magnesium sulfate are combined? If so, write the net ionic equation for the reaction.

3. Will a precipitate form if solutions of barium chloride and sodium sulfate are combined? If so, identify the spectator ions and write the net ionic equation.

4. Write the net ionic equation for the precipitation of nickel(II) sulfide.

extension

Go to **go.hrw.com** for more practice problems that ask you to identify precipitates that form when solutions are combined.

Keyword: HC6IONX

Ionization

Some molecular compounds can also form ions in solution. Usually such compounds are polar. *Ions are formed from solute molecules by the action of the solvent in a process called* **ionization.** The more general meaning of this term is the creation of ions where there were none. Note that *ionization* is different from *dissociation*. When an ionic compound dissolves, the ions that were already present separate from one another. When a molecular compound dissolves and ionizes in a polar solvent, ions are formed where none existed in the undissolved compound. Like all ions in aqueous solution, the ions formed by such a molecular solute are hydrated. The energy released as heat during the hydration of the ions provides the energy needed to break the covalent bonds.

In general, the extent to which a solute ionizes in solution depends on the strength of the bonds within the molecules of the solute and the strength of attraction between the solute and solvent molecules. If the strength of a bond within the solute molecule is weaker than the attractive forces of the solvent molecules, then the covalent bond of the solute breaks and the molecule is separated into ions. Hydrogen chloride, HCl, is a molecular compound that ionizes in aqueous solution. It contains a highly polar bond. The attraction between a polar HCl molecule and the polar water molecules is strong enough to break the HCl bond, forming hydrogen ions and chloride ions.

$$HCl \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$
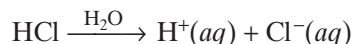
The Hydronium Ion

Many molecular compounds contain a hydrogen atom bonded by a polar covalent bond. Some of these compounds ionize in an aqueous solution to release H^+. The H^+ ion attracts other molecules or ions so strongly that it does not normally exist alone. The ionization of hydrogen chloride in water is better described as a chemical reaction in which a proton is transferred directly from HCl to a water molecule, where it becomes covalently bonded to oxygen and forms H_3O^+.

$$H_2O(l) + HCl(g) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$
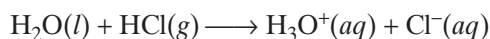
This process is represented in **Figure 4.** *The H_3O^+ ion is known as the* **hydronium ion.** The reaction of the H^+ ion to form the hydronium ion produces much of the energy needed to ionize a molecular solute.
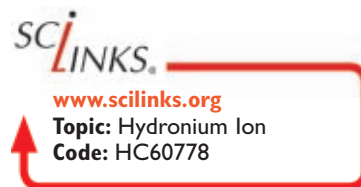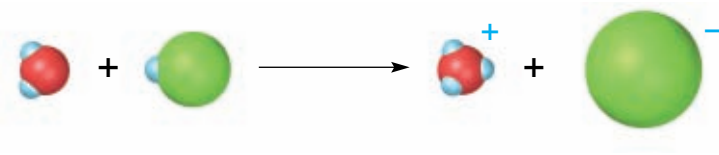
FIGURE 4 When hydrogen chloride gas dissolves in water, it ionizes to form an H^+ ion and a Cl^- ion. The H^+ ion immediately bonds to a water molecule, forming a hydronium ion. The aqueous solution of hydrogen chloride is called hydrochloric acid.
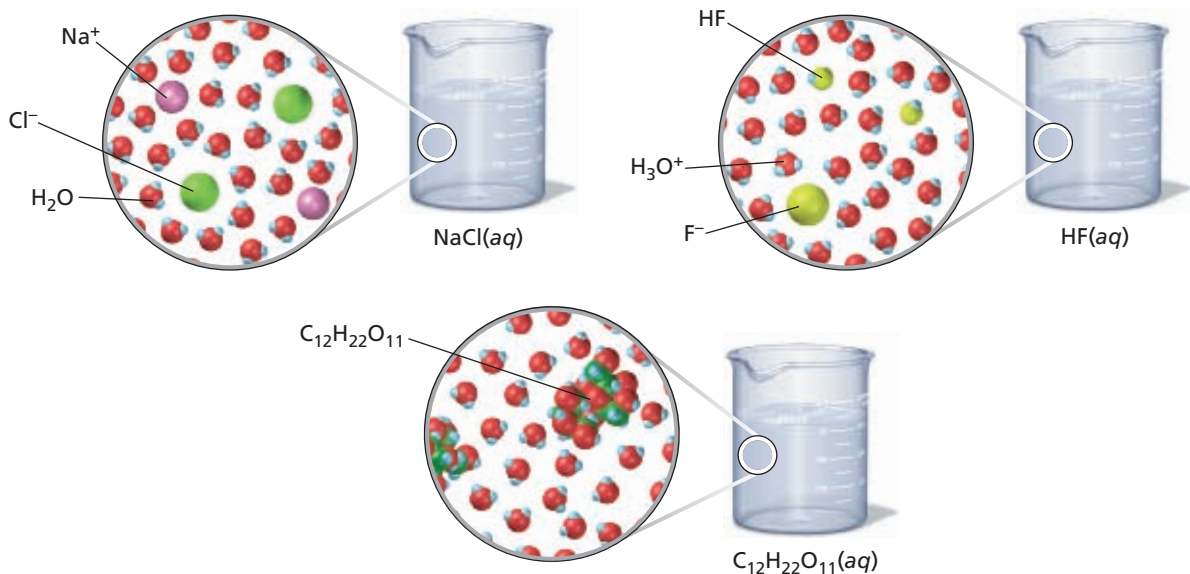
Strong electrolytes, such as NaCl, yield only ions when they dissolve in aqueous solution. Weak electrolytes, such as HF, exist as both ions and unionized molecules in aqueous solution. Nonelectrolytes, such as sucrose, $C_{12}H_{22}O_{11}$, do not form any ions in aqueous solution.

Strong and Weak Electrolytes

As discussed in Chapter 12, substances that yield ions and conduct an electric current in solution are electrolytes. Substances that do not yield ions and do not conduct an electric current in solution are non-electrolytes. Hydrogen chloride is one of a series of compounds composed of hydrogen and the members of Group 17 (known as the halogens). The hydrogen halides are all molecular compounds with single polar-covalent bonds. All are gases, all are very soluble in water, and all are electrolytes. Hydrogen chloride, hydrogen bromide, and hydrogen iodide strongly conduct an electric current in an aqueous solution. However, hydrogen fluoride only weakly conducts an electric current at the same concentration. The strength with which substances conduct an electric current is related to their ability to form ions in solution, as shown in **Figure 5.**

Strong Electrolytes

Hydrogen chloride, hydrogen bromide, and hydrogen iodide are 100% ionized in dilute aqueous solution. A **strong electrolyte** *is any compound whose dilute aqueous solutions conduct electricity well; this is due to the presence of all or almost all of the dissolved compound in the form of ions.* Hydrogen chloride, hydrogen bromide, and hydrogen iodide are all acids in aqueous solution. These acids, several other acids, and all soluble ionic compounds are strong electrolytes.

The distinguishing feature of strong electrolytes is that to whatever extent they dissolve in water, they yield only ions. For example, some strong electrolytes, such as NaCl, may be highly soluble in water and form ions in solution. Others may not dissolve much, but the amount that does dissolve exists solely as ions in solution.
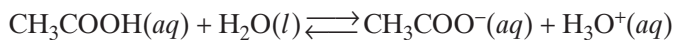
Weak Electrolytes

Some molecular compounds form aqueous solutions that contain not only dissolved ions but also some dissolved molecules that are not ionized. Hydrogen fluoride, HF, dissolves in water to give an acidic solution known as hydrofluoric acid. However, the hydrogen-fluorine bond is much stronger than the bonds between hydrogen and the other halogens. When hydrogen fluoride dissolves, some molecules ionize. But the reverse reaction—the transfer of H^+ ions back to F^- ions to form hydrogen fluoride molecules—also takes place.

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$
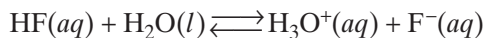
Thus, the concentration of dissolved intact HF is much greater than the concentration of H_3O^+ and F^- ions.

Hydrogen fluoride is an example of a *weak electrolyte. A* **weak electrolyte** *is any compound whose dilute aqueous solutions conduct electricity poorly; this is due to the presence of a small amount of the dissolved compound in the form of ions.* This is in contrast to a nonelectrolyte, such as the molecular compound sucrose, which dissolves but does not produce any ions in solution. Another example of a weak electrolyte is acetic acid, CH_3COOH. Only a small percentage of the acetic acid molecules ionize in aqueous solution.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

The description of an electrolyte as strong or weak must not be confused with the description of a solution as concentrated or dilute. Strong and weak electrolytes differ in the *degree of ionization or dissociation.* Concentrated and dilute solutions differ in the *amount of solute dissolved* in a given quantity of a solvent. Hydrochloric acid is always a strong electrolyte. This is true even in a solution that is 0.000 01 M—a very dilute solution. By contrast, acetic acid is always considered a weak electrolyte, even in a 10 M solution—a very concentrated solution.

SECTION REVIEW

1. Write the equation for the dissolution of $Sr(NO_3)_2$ in water. How many moles of strontium ions and nitrate ions are produced by dissolving 0.5 mol of strontium nitrate?

2. Will a precipitate form if solutions of magnesium acetate and strontium chloride are combined?

3. What determines whether a molecular compound will be ionized in a polar solvent?

4. Explain why HCl is a strong electrolyte and HF is a weak electrolyte.

Critical Thinking

5. **PREDICTING OUTCOMES** For each of the following pairs, tell which solution contains the larger total concentration of ions.

   a. 0.10 M HCl and 0.05 M HCl

   b. 0.10 M HCl and 0.10 M HF

   c. 0.10 M HCl and 0.10 M $CaCl_2$

The Riddle of Electrolysis

Historical Perspective

When Michael Faraday performed his electrochemical experiments, little was known about the relationship between matter and electricity. Chemists were still debating the existence of atoms, and the discovery of the electron was more than 50 years in the future. Combining his talents in electrical and chemical investigation, Faraday pointed researchers to the intimate connection between chemical reactions and electricity while setting the stage for the development of a new branch of chemistry.

Electrifying Experiments

In 1800, Italian physicist Alessandro Volta introduced his *voltaic pile,* better known as the battery. The stack of alternating zinc and silver disks provided scientists with a source of electric current for the first time.

That same year, chemists discovered a new phenomenon using Volta's device. They immersed the two poles of a battery at different locations in a container of water. The current caused the water to decompose into its elemental components, with hydrogen evolving at the positive pole of the battery and oxygen evolving at the negative pole. Similar experiments using solutions of certain solids dissolved in water resulted in the decom-position of the solids, with the two products of the solids' breakdown also evolving at opposite poles of the battery. This electrochemical decomposition was later named *electrolysis*.

The Roots of Electrolytic Theory

The discovery of electrolysis led two pioneering chemists to ponder the connection between chemical forces and electricity. One of the chemists was Humphry Davy, who thought that chemical bonding must be driven by the same forces that drive electrical attractions.

The Swedish chemist Jöns Jacob Berzelius took Davy's idea a step further. He postulated that matter consisted of combinations of "electropositive" and "electronegative" substances, classifying the parts by the pole at which they accumulated during electrolysis.

These ideas inspired two early electrolytic theories, each of which ultimately proved incorrect but contributed to our present understanding of the phenomenon. The *contact theory* proposed that electrolytic current was due merely to the contact of the battery's metals with the electrolytic solution. The *chemical theory,* on the other hand, attributed the current to undefined changes in the solution's components.

▼ *Chlorine being produced by electrolysis.*
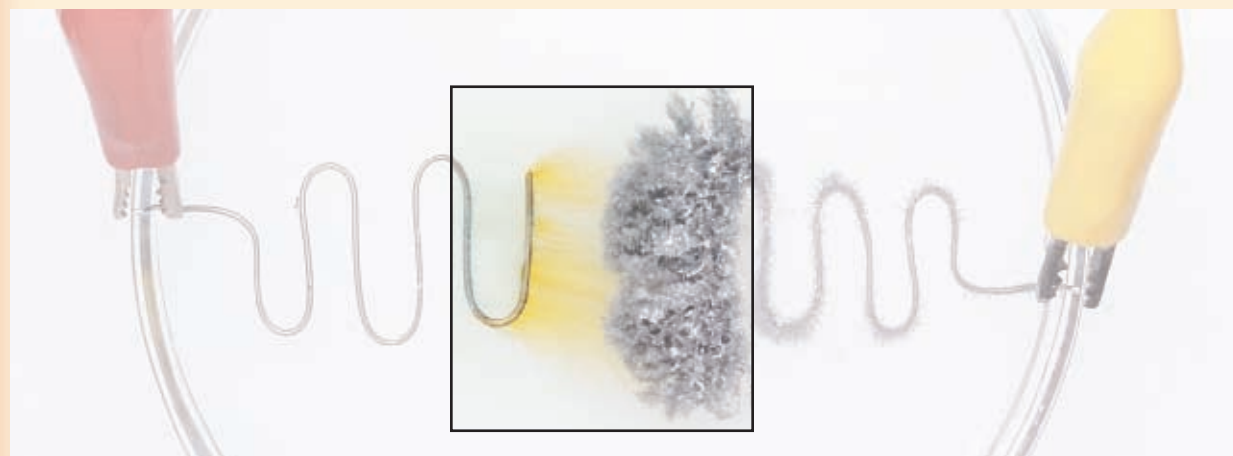
Faraday Provides a Spark

Although Michael Faraday is best remembered for his work in electromagnetism, he began his career as Humphry Davy's laboratory assistant at the Royal Institution in London and went on to be the professor of chemistry there for over 30 years. In the 1830s, Faraday devised several ingenious experiments to determine whether the current in an electrolytic solution is dependent solely on the contact of the battery's poles with the solution. In a typical setup, one of the poles was separated from the solution, and electricity was permitted to enter the solution by way of a spark. In all cases, Faraday observed current in the electrolytic cell despite one or both of the poles not being in direct contact with the electrolytic solution. In 1833, he made the hypothesis that the process of electrolysis was due to the intrinsic properties of the metals in solution and the effect of current on these properties.

Although the battery's poles were, in fact, later shown to play a part in the current, Faraday had established the active role of the electrolytic solution in electrolysis. And in realizing that electricity affected the chemical nature of the solution, he anticipated the ideas of oxidation and reduction despite that the concepts of electrons and ions were unknown at the time.

Faraday's Legacy

Faraday continued to study the role of the electrolytic solution, or electrolyte, as he named it, in electrolysis. He also coined most of the other modern terms of electrolysis, including *electrode, ion, anode, cathode, anion,* and *cation.* These investigations culminated in the discovery of his basic laws of electrolysis.

Still valid today, these principles put electrolysis on a quantitative footing, leading to our current understanding of the phenomenon. They also bolstered the atomic theory, which was still seriously contested by many chemists at the time. And perhaps most important, Faraday's experiments inspired his successors to further clarify the chemical nature of solutions. This ultimately led to Svante Arrhenius's theory of electrolytic dissociation and the evolution of a new division in the chemical field, known today as *physical chemistry*.
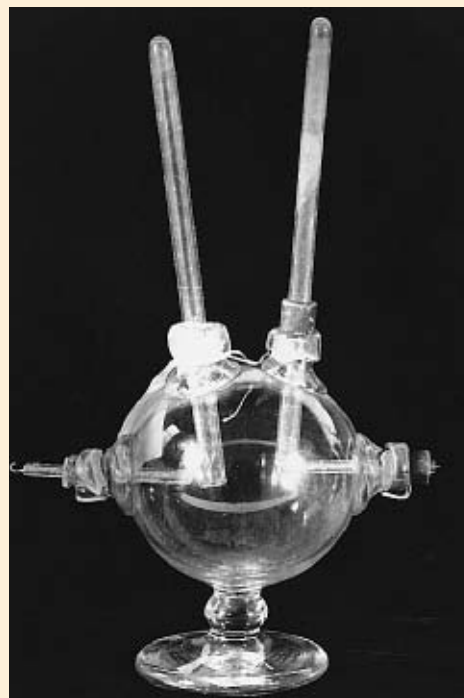
▲ *Michael Faraday used this instrument in his studies of electrolysis.*

Questions

1. To which scientist does this feature give the most credit for the initial understanding of the basic nature of electrolysis? Briefly summarize his contributions in this area.

2. How are amounts of hydrogen and oxygen gas made from the electrolytic decomposition of water related to the chemical formula of water?

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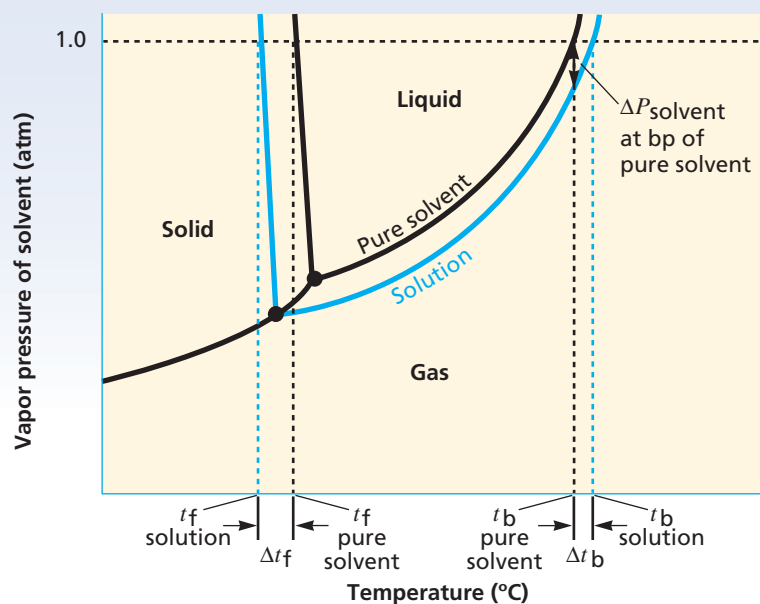**Code:** HC60479

Colligative Properties of Solutions

The presence of solutes affects the properties of the solutions. Some of these properties are not dependent on the nature of the dissolved substance but only on how many dissolved particles are present. *Properties that depend on the concentration of solute particles but not on their identity are called* **colligative properties.** In calculations involving some colligative properties, the concentration is given in terms of molality, *m*.

Vapor-Pressure Lowering

The boiling point and freezing point of a solution differ from those of the pure solvent. The graph in **Figure 6** shows that a nonvolatile solute raises the boiling point and lowers the freezing point. *A* **nonvolatile substance** *is one that has little tendency to become a gas under existing conditions*.

FIGURE 6   Vapor pressure as a function of temperature is shown for a pure solvent and for a solution of a nonvolatile solute in that solvent. The vapor pressure of the solution is lower than the vapor pressure of the pure solvent. This can be seen by noting the decrease in pressure between the pure solvent and the solution at the temperature that is the boiling point of the pure solvent. The solute thus reduces the freezing point and elevates the boiling point.

Vapor Pressure vs. Temperature for a Pure Solvent and a Solution with a Nonvolatile Solute

Pure water

Aqueous solution of nonvolatile solute

Used to represent $C_{12}H_{22}O_{11}$, sucrose
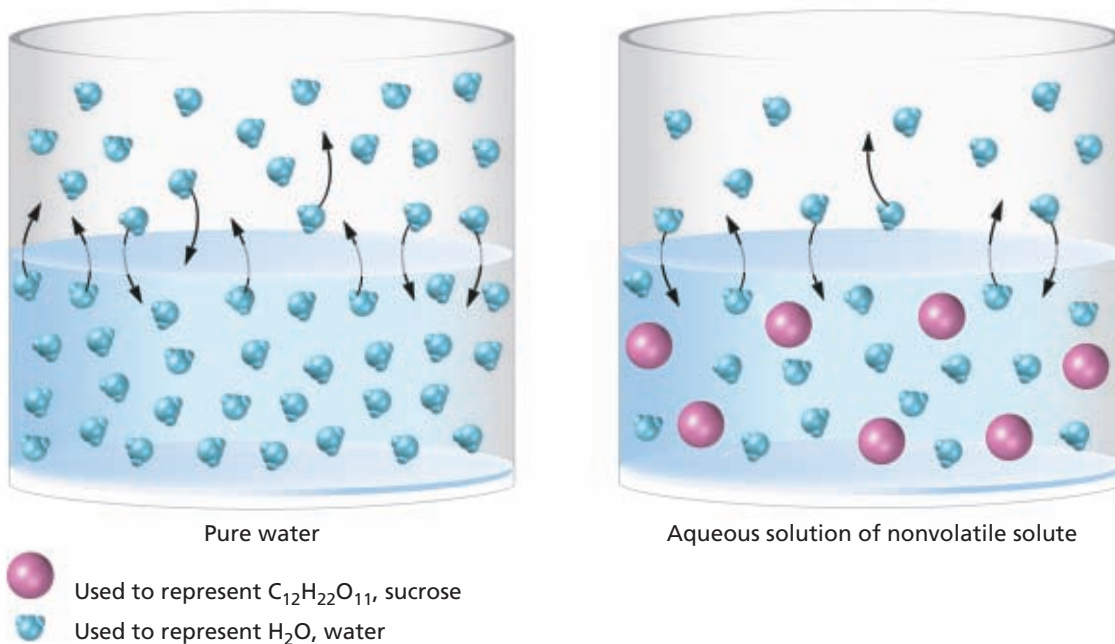
Used to represent H_2O, water

FIGURE 7 At a given temperature, the vapor pressure of water over pure water is greater than the vapor pressure of water over an aqueous solution containing a nonvolatile solute, such as sucrose.

To understand why a nonvolatile solute changes the boiling point and freezing point, you must consider equilibrium vapor pressure, which was discussed in Chapter 10. Vapor pressure is the pressure caused by molecules in the gas phase that are in equilibrium with the liquid phase. Experiments show that the vapor pressure of a solvent containing a nonvolatile solute is lower than the vapor pressure of the pure solvent at the same temperature, as shown in **Figure 7.** As the number of solute particles increases in a given volume of solution, the proportion of solvent (water) molecules decreases. Fewer water molecules will be available to escape from the liquid. As a result, the tendency of water molecules to leave the solution and enter the vapor phase decreases. Thus, the vapor pressure of the solution is less than the vapor pressure of pure water.

Nonelectrolyte solutions of the same molality have the same concentration of particles. Equally dilute molal solutions of any nonelectrolyte solutes in the same solvent lower the vapor pressure equally. For example, a 1 *m* aqueous solution of the nonelectrolyte glucose, $C_6H_{12}O_6$, lowers the vapor pressure of water by 5.5×10^{-4} atm at 25°C. A 1 *m* solution of sucrose, $C_{12}H_{22}O_{11}$, another nonelectrolyte, also lowers the vapor pressure by 5.5×10^{-4} atm. Because vapor-pressure lowering depends on the concentration of a nonelectrolyte solute and is independent of solute identity, it is a colligative property.

Refer to the graph in **Figure 6.** Because the vapor pressure has been lowered, the solution remains liquid over a larger temperature range. This lowers the freezing point and raises the boiling point. It follows that changes in boiling point and freezing point also depend on the concentration of solute and are therefore colligative properties.

Freezing-Point Depression

The freezing point of a 1 *m* solution of any nonelectrolyte solute in water is found by experiment to be 1.86°C lower than the freezing point of water. That is, when 1 mol of a nonelectrolyte solute is dissolved in 1 kg of water, the freezing point of the solution is −1.86°C instead of 0.00°C. When 2 mol of a nonelectrolyte solute is dissolved in 1 kg of water, the freezing point of the solution is −3.72°C. This is 2 × (−1.86°C). In fact, for any concentration of a nonelectrolyte solute in water, the decrease in freezing point can be estimated by using the value of −1.86°C/*m*. This value, called the **molal freezing-point constant** (K_f) *is the freezing-point depression of the solvent in a 1-molal solution of a nonvolatile, nonelectrolyte solute.*

Each solvent has its own characteristic molal freezing-point constant. The values of K_f for some common solvents are given in **Table 2.** These values are most accurate for dilute solutions at 1 atmosphere of pressure. Some variations are introduced in the value of K_f at other pressures and with more-concentrated solutions. The table also shows the values of a related quantity called K_b, which you will study next.

As stated earlier, the freezing point of a solution containing 1 mol of a nonelectrolyte solute in 1 kg water is 1.86°C lower than the normal freezing point of water. *The* **freezing-point depression,** Δt_f, *is the difference between the freezing points of the pure solvent and a solution of a nonelectrolyte in that solvent, and it is directly proportional to the molal concentration of the solution.* As shown by the previous example, if the molal concentration is doubled, the freezing-point depression is doubled. Freezing-point depression can be calculated by the following equation.

$$\Delta t_f = K_f m$$

K_f is expressed as °C/*m*, *m* is expressed in mol solute/kg solvent (molality), and Δt_f is expressed in °C. Sample Problems C and D show how this relationship can be used to determine the freezing-point depression and molal concentration of a solution.

TABLE 2 *Molal Freezing-Point and Boiling-Point Constants*

Solvent	Normal f.p. (°C)	Molal f.p. constant, K_f (°C/*m*)	Normal b.p. (°C)	Molal b.p. constant, K_b (°C/*m*)
Acetic acid	16.6	−3.90	117.9	3.07
Camphor	178.8	−39.7	207.4	5.61
Ether	−116.3	−1.79	34.6	2.02
Naphthalene	80.2	−6.94	217.7	5.80
Phenol	40.9	−7.40	181.8	3.60
Water	0.00	−1.86	100.0	0.51

For more help, go to the Math Tutor at the end of this chapter.

What is the freezing-point depression of water in a solution of 17.1 g of sucrose, $C_{12}H_{22}O_{11}$, in 200. g of water? What is the actual freezing point of the solution?

SOLUTION

1 ANALYZE   **Given:** solute mass and chemical formula = 17.1 g $C_{12}H_{22}O_{11}$
                solvent mass and identity = 200. g water
             Unknown: a. freezing-point depression
                        b. freezing point of the solution

2 PLAN   Find the molal freezing-point constant, K_f, for water in **Table 2.** To use the equation for freezing-point depression, $\Delta t_f = K_f m$, you need to determine the molality of the solution.

$$\text{mass of solute (g)} \times \frac{1 \text{ mol solute}}{\text{molar mass of solute (g)}} = \text{amount of solute (mol)}$$

$$\frac{\text{amount of solute (mol)}}{\text{mass of solvent (g)}} \times \frac{1000 \text{ g water}}{1 \text{ kg water}} = \text{molality}$$

$$\Delta t_f = K_f m$$
$$\text{f.p. solution} = \text{f.p. solvent} + \Delta t_f$$

3 COMPUTE $17.1 \text{ g } C_{12}H_{22}O_{11} \times \dfrac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.34 \text{ g } C_{12}H_{22}O_{11}} = 0.0500 \text{ mol } C_{12}H_{22}O_{11}$

$$\frac{0.0500 \text{ mol } C_{12}H_{22}O_{11}}{200. \text{ g water}} \times \frac{1000 \text{ g water}}{\text{kg water}} = \frac{0.250 \text{ mol } C_{12}H_{22}O_{11}}{\text{kg water}} = 0.250 \text{ } m$$

a. $\Delta t_f = 0.250 \text{ } m \times (-1.86°C/m) = -0.465°C$
b. f.p. solution $= 0.000°C + (-0.465°C) = -0.465°C$

For more help, go to the Math Tutor at the end of this chapter.

A water solution containing an unknown quantity of a nonelectrolyte solute is found to have a freezing point of −0.23°C. What is the molal concentration of the solution?

SOLUTION

1 ANALYZE   **Given:** freezing point of solution = −0.23°C
             Unknown: molality of the solution

2 PLAN   Water is the solvent, so you will need the value of K_f, the molal-freezing-point constant for water, from **Table 2.** The Δt_f for this solution is the difference between the f.p. of water and the f.p. of the solution. Use the equation for freezing-point depression to calculate molality.

$$\Delta t_f = \text{f.p. of solution} - \text{f.p. of pure solvent}$$
$$\Delta t_f = K_f m \qquad \text{Solve for molality, } m.$$

$$m = \frac{\Delta t_f}{K_f}$$

3 COMPUTE

$$\Delta t_f = -0.23°C - 0.00°C = -0.23°C$$

$$m = \frac{-0.23°C}{-1.86°C/m} = 0.12\ m$$

4 EVALUATE As shown by the unit cancellation, the answer gives the molality, as desired. The answer is properly limited to two significant digits.

PRACTICE *Answers in Appendix E*

1. A solution consists of 10.3 g of the nonelectrolyte glucose, $C_6H_{12}O_6$, dissolved in 250. g of water. What is the freezing-point depression of the solution?

2. In a laboratory experiment, the freezing point of an aqueous solution of glucose is found to be $-0.325°C$. What is the molal concentration of this solution?

3. If 0.500 mol of a nonelectrolyte solute are dissolved in 500.0 g of ether, what is the freezing point of the solution?

4. The freezing point of an aqueous solution that contains a nonelectrolyte is $-9.0°C$.
   a. What is the freezing-point depression of the solution?
   b. What is the molal concentration of the solution?

Boiling-Point Elevation

As discussed in Chapter 12, the boiling point of a liquid is the temperature at which the vapor pressure of the liquid is equal to the prevailing atmospheric pressure. Therefore, a change in the vapor pressure of the liquid will cause a corresponding change in the boiling point. As stated earlier, the vapor pressure of a solution containing a nonvolatile solute is lower than the vapor pressure of the pure solvent. This means that more energy as heat will be required to raise the vapor pressure of the solution to equal the atmospheric pressure. Thus, the boiling point of a solution is higher than the boiling point of the pure solvent.

*The **molal boiling-point constant** (K_b) is the boiling-point elevation of the solvent in a 1-molal solution of a nonvolatile, nonelectrolyte solute.* The boiling-point elevation of a 1-molal solution of any nonelectrolyte solute in water has been found by experiment to be $0.51°C$. Thus, the molal boiling-point constant for water is $0.51°C/m$.

For different solvents, the molal boiling-point constants have different values. Some other values for K_b are included in **Table 2.** Like the freezing-point constants, these values are most accurate for dilute solutions.

*The **boiling-point elevation,** Δt_b, is the difference between the boiling points of the pure solvent and a nonelectrolyte solution of that solvent,*

and it is directly proportional to the molal concentration of the solution. Boiling-point elevation can be calculated by the following equation.

$$\Delta t_b = K_b m$$

When K_b is expressed in °C/m and m is expressed in mol of solute/kg of solvent, Δt_b is the boiling-point elevation in °C.

SAMPLE PROBLEM E

For more help, go to the Math Tutor at the end of this chapter.

What is the boiling-point elevation of a solution made from 20.1 g of a nonelectrolyte solute and 400.0 g of water? The molar mass of the solute is 62.0 g.

SOLUTION

1 ANALYZE   **Given:** solute mass = 20.1 g
                        solute molar mass = 62.0 g
                        solvent mass and identity = 400.0 g of water
               Unknown: boiling-point elevation

2 PLAN   Find the molal boiling-point constant, K_b, for water in **Table 2.** To use the equation for boiling-point elevation, $\Delta t_b = K_b m$, you need to determine the molality of the solution.

$$\text{mass of solute (g)} \times \frac{1 \text{ mol solute}}{\text{molar mass of solute (g)}} = \text{amount of solute (mol)}$$

$$\frac{\text{amount of solute (mol)}}{\text{mass of solvent (g)}} \times \frac{1000 \text{ g water}}{1 \text{ kg water}} = \text{molality}$$

$$\Delta t_b = K_b m$$

3 COMPUTE

$$20.1 \text{ g of solute} \times \frac{1 \text{ mol solute}}{62.0 \text{ g of solute}} = 0.324 \text{ mol of solute}$$

$$\frac{0.324 \text{ mol of solute}}{400.0 \text{ g water}} \times \frac{1000 \text{ g water}}{1 \text{ kg water}} = 0.810 \frac{\text{mol solute}}{\text{kg water}} = 0.810 \, m$$

$$\Delta t_b = 0.51°C/m \times 0.810 \, m = 0.41°C$$

PRACTICE

Answers in Appendix E

1. A solution contains 50.0 g of sucrose, $C_{12}H_{22}O_{11}$, a nonelectrolyte, dissolved in 500.0 g of water. What is the boiling-point elevation?

2. A solution contains 450.0 g of sucrose, $C_{12}H_{22}O_{11}$, a nonelectrolyte, dissolved in 250.0 g of water. What is the boiling point of the solution?

3. If the boiling-point elevation of an aqueous solution containing a nonvolatile electrolyte is 1.02°C, what is the molality of the solution?

4. The boiling point of an aqueous solution containing a nonvolatile electrolyte is 100.75°C.
   a. What is the boiling-point elevation?
   b. What is the molality of the solution?

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate the boiling-point elevation.

Keyword: HC6IONX

IONS IN AQUEOUS SOLUTIONS AND COLLIGATIVE PROPERTIES   **451**

(a)

(b)

Pure water

Sucrose solution

Semipermeable membrane

⬤ Used to represent $C_{12}H_{22}O_{11}$, sucrose
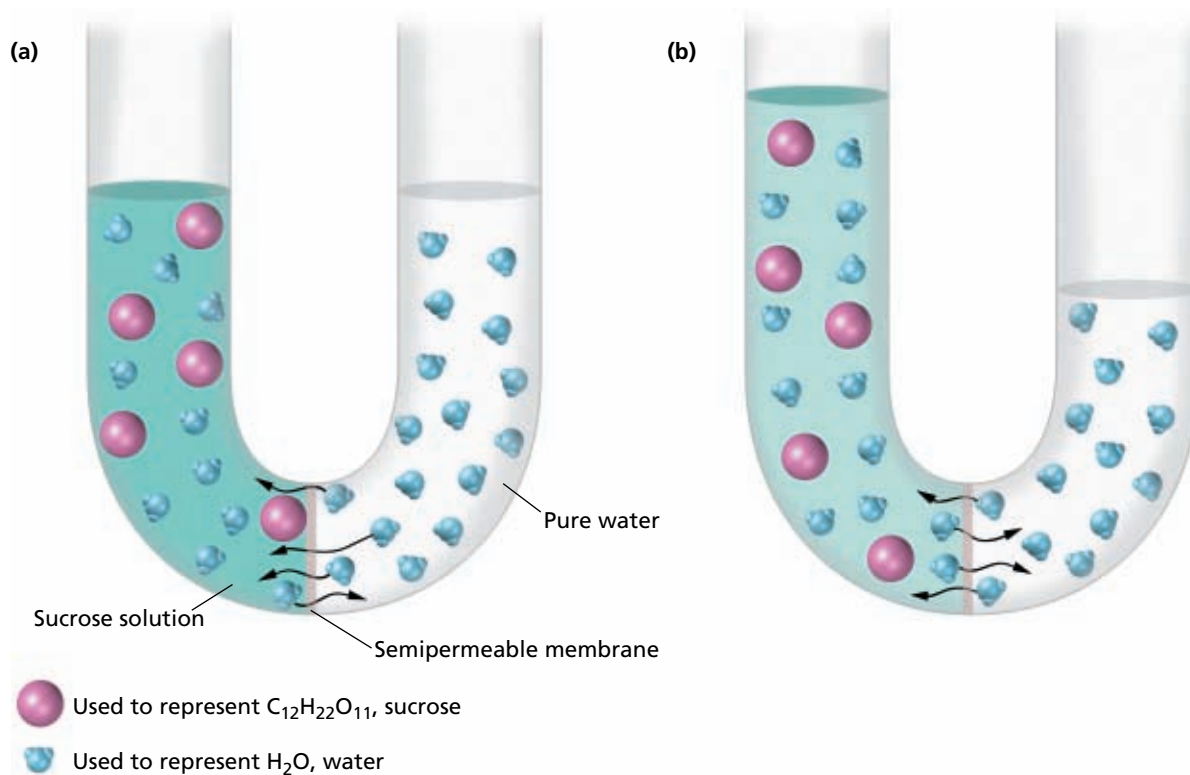
⬤ Used to represent H_2O, water

FIGURE 8 (a) When pure water and an aqueous sucrose solution are separated by a semipermeable membrane, the net movement of water molecules through the membrane is from the pure water side into the aqueous solution. (b) The level of the solution rises until pressure exerted by the height of the solution equals the osmotic pressure, at which point no net movement of water molecules occurs.

Osmotic Pressure

Figure 8 illustrates another colligative property. In the figure, an aqueous sucrose solution is separated from pure water by a semipermeable membrane. A **semipermeable membrane** *allows the passage of some particles while blocking the passage of others.* The level of the sucrose solution will rise until a certain height is reached. What causes the level of the solution to rise?

The semipermeable membrane allows water molecules, but not sucrose molecules, to pass through. The sucrose molecules on the solution side allow fewer water molecules to strike the membrane than strike on the pure water side in the same amount of time. Thus, the rate at which water molecules leave the pure water side is greater than the rate at which they leave the solution. This causes the level of the solution to rise. The level rises until the pressure exerted by the height of the solution is large enough to force water molecules back through the membrane from the solution at a rate equal to that at which they enter from the pure water side.

The movement of solvent through a semipermeable membrane from the side of lower solute concentration to the side of higher solute concentration is **osmosis.** Osmosis occurs whenever two solutions of different concentrations are separated by a semipermeable membrane. **Osmotic pressure** *is the external pressure that must be applied to stop osmosis.* In the example given above, osmosis caused the level of the solution to rise until the height of the solution provided the pressure necessary to stop osmosis. Because osmotic pressure is dependent on the concentration of
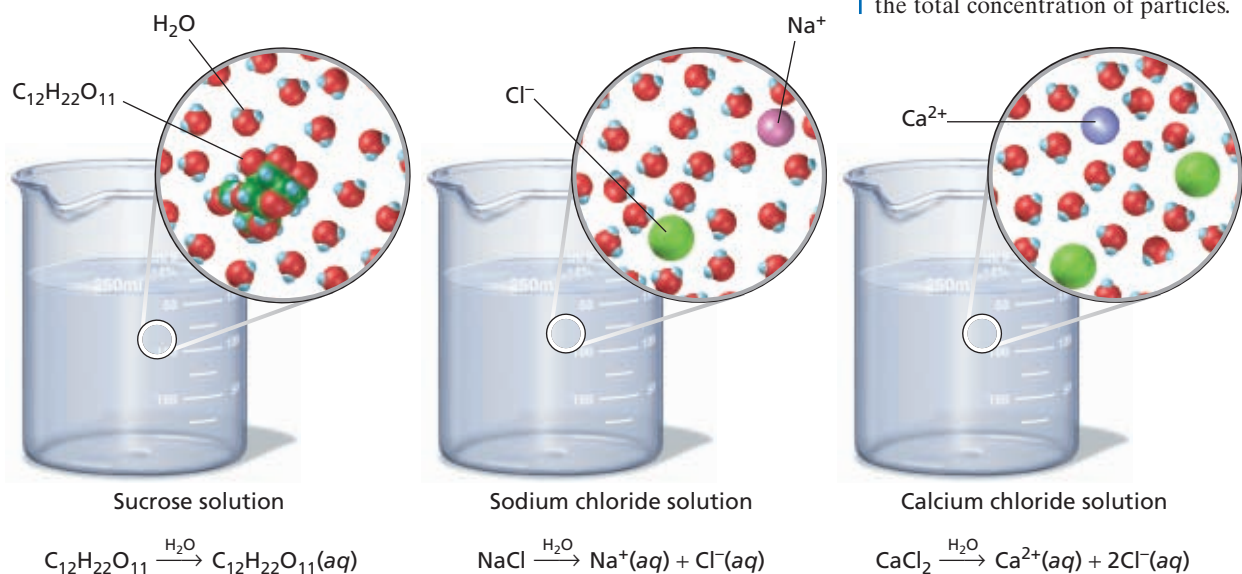
solute particles and not on the type of solute particles, it is a colligative property. The greater the concentration of a solution, the greater the osmotic pressure of the solution.

Regulation of osmosis is vital to the life of a cell because cell membranes are semipermeable. Cells lose water and shrink when placed in a solution of higher concentration. They gain water and swell when placed in a solution of lower concentration. In vertebrates, cells are protected from swelling and shrinking by blood and lymph that surround the cells. Blood and lymph are equal in concentration to the concentration inside the cell.

Electrolytes and Colligative Properties

Early investigators were puzzled by experiments in which certain substances depressed the freezing point or elevated the boiling point of a solvent more than expected. For example, a 0.1 m solution of sodium chloride, NaCl, lowers the freezing point of the solvent nearly twice as much as a 0.1 m solution of sucrose. A 0.1 m solution of calcium chloride, $CaCl_2$, lowers the freezing point of the solvent nearly three times as much as a 0.1 m solution of sucrose. The effect on boiling points is similar.

To understand why this is so, contrast the behavior of sucrose with that of sodium chloride in aqueous solutions. Sugar is a nonelectrolyte. Each sucrose molecule dissolves to produce only one particle in solution, so 1 mol of sucrose dissolves to produce only 1 mol of particles in solution. NaCl, however, is a strong electrolyte. Each mole of NaCl dissolves to produce 2 mol of particles in solution: 1 mol of sodium ions and 1 mol of chloride ions. **Figure 9** compares the production of particles in solution for three different solutes. As you can see, electrolytes produce more than 1 mol of solute particles for each mole of compound dissolved.

FIGURE 9 Compare the number of particles produced per formula unit for these three solutes. Colligative properties depend on the total concentration of particles.

Sucrose solution

$$C_{12}H_{22}O_{11} \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

Sodium chloride solution

$$NaCl \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

Calcium chloride solution

$$CaCl_2 \xrightarrow{H_2O} Ca^{2+}(aq) + 2Cl^-(aq)$$

Calculated Values for Electrolyte Solutions

Remember that colligative properties depend on the total concentration of solute particles regardless of their identity. The changes in colligative properties caused by electrolytes will be proportional to the total molality of all dissolved particles, not to formula units. For the same molal concentrations of sucrose and sodium chloride, you would expect the effect on colligative properties to be twice as large for sodium chloride as for sucrose. What about barium nitrate, $Ba(NO_3)_2$? Each mole of barium nitrate yields 3 mol of ions in solution.

$$Ba(NO_3)_2(s) \longrightarrow Ba^{2+}(aq) + 2NO_3^-(aq)$$

You would expect a $Ba(NO_3)_2$ solution of a given molality to lower the freezing point of its solvent three times as much as a nonelectrolytic solution of the same molality.

SAMPLE PROBLEM F *For more help, go to the Math Tutor at the end of this chapter.*

What is the expected change in the freezing point of water in a solution of 62.5 g of barium nitrate, $Ba(NO_3)_2$, in 1.00 kg of water?

SOLUTION

1 *ANALYZE*

Given: solute mass and formula = 62.5 g $Ba(NO_3)_2$
solvent mass and identity = 1.00 kg water
$\Delta t_f = K_f m$
Unknown: expected freezing-point depression

2 *PLAN*

The molality can be calculated by converting the solute mass to moles and then dividing by the number of kilograms of solvent. That molality is in terms of formula units of $Ba(NO_3)_2$ and must be converted to molality in terms of dissociated ions in solution. It must be multiplied by the number of moles of ions produced per mole of formula unit. This adjusted molality can then be used to calculate the freezing-point depression.

$$\frac{\text{mass of solute (g)}}{\text{mass of solvent (kg)}} \times \frac{1 \text{ mol solute}}{\text{molar mass of solute (g)}} = \text{molality of solution}\left(\frac{\text{mol}}{\text{kg}}\right)$$

$$\text{molality of solution}\left(\frac{\text{mol}}{\text{kg}}\right) \times \text{molality conversion}\left(\frac{\text{mol ions}}{\text{mol}}\right) \times K_f\left(\frac{°C \cdot \text{kg } H_2O}{\text{mol ions}}\right)$$

$$= \text{expected freezing-point depression (°C)}$$

This problem is similar to Sample Problem E, except that the solute is ionic rather than a nonionizing molecular solute. The number of particles in solution will therefore equal the number of ions of the solute.

3 *COMPUTE*

$$\frac{62.5 \text{ g } Ba(NO_3)_2}{1.00 \text{ kg } H_2O} \times \frac{\text{mol } Ba(NO_3)_2}{261.35 \text{ g } Ba(NO_3)_2} = \frac{0.239 \text{ mol } Ba(NO_3)_2}{\text{kg } H_2O}$$

$$Ba(NO_3)_2(s) \longrightarrow Ba^{2+}(aq) + 2NO_3^-(aq)$$

Each formula unit of barium nitrate yields three ions in solution.

$$\frac{0.239 \text{ mol Ba(NO}_3)_2}{\text{kg H}_2\text{O}} \times \frac{3 \text{ mol ions}}{\text{mol Ba(NO}_3)_2} \times \frac{-1.86°C \cdot \text{kg H}_2\text{O}}{\text{mol ions}} = -1.33°C$$

4   EVALUATE   The units cancel properly to give the desired answer in °C. The answer is correctly given to three significant digits. The mass of the solute is approximately one-fourth its molar mass and would give 0.75 mol of ions in the 1 kg of solvent, so the estimated answer of $0.75 \times (-1.86°C) = -1.4°C$ supports our computation.

PRACTICE   *Answers in Appendix E*

1. What is the expected freezing-point depression for a solution that contains 2.0 mol of magnesium sulfate dissolved in 1.0 kg of water?

2. What is the expected boiling-point elevation of water for a solution that contains 150 g of sodium chloride dissolved in 1.0 kg of water?

3. The freezing point of an aqueous sodium chloride solution is −0.20°C. What is the molality of the solution?

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate the freezing-point depression and boiling-point elevation for electrolyte solutions.

Keyword: HC6IONX

Actual Values for Electrolyte Solutions

It is important to remember that the values just calculated are only *expected* values. As stated earlier, a 0.1 *m* solution of sodium chloride lowers the freezing point *nearly* twice as much as a 0.1 *m* solution of sucrose. The actual values of the colligative properties for all strong electrolytes are *almost* what would be expected based on the number of particles they produce in solution. Some specific examples are given in **Table 3.** The freezing-point depression of a compound that produces two ions per formula unit is almost twice that of a nonelectrolytic solution. The freezing-point depression of a compound that produces three ions per formula unit is almost three times that of a nonelectrolytic solution.

TABLE 3 *Molal Freezing-Point Depressions for Aqueous Solutions of Ionic Solutes*

Solute	Concentration (*m*)	Δt_f, observed (°C)	Δt_f, nonelectrolyte solution (°C)	$\dfrac{\Delta t_f, \text{observed}}{\Delta t_f, \text{nonelectrolyte solution}}$
KCl	0.1	−0.345	−0.186	1.85
	0.01	−0.0361	−0.0186	1.94
	0.001	−0.00366	−0.00186	1.97
MgSO$_4$	0.1	−0.225	−0.186	1.21
	0.01	−0.0285	−0.0186	1.53
	0.001	−0.00338	−0.00186	1.82
BaCl$_2$	0.1	−0.470	−0.186	2.53
	0.01	−0.0503	−0.0186	2.70
	0.001	−0.00530	−0.00186	2.84

Look at the values given for KCl solutions in **Table 3.** The freezing-point depression of a 0.1 *m* KCl solution is only 1.85 times greater than that of a nonelectrolyte solution. However, as the concentration decreases, the freezing-point depression comes closer to the value that is twice that of a nonelectrolyte solution.

The differences between the expected and calculated values are caused by the attractive forces that exist between dissociated ions in aqueous solution. The attraction between the hydrated ions in the solution is small compared with those in the crystalline solid. However, forces of attraction do interfere with the movements of the aqueous ions. Only in very dilute solutions is the average distance between the ions large enough and the attraction between ions small enough for the solute ions to move about almost completely freely.

Peter Debye and Erich Hückel introduced a theory in 1923 to account for this attraction between ions in dilute aqueous solutions. According to this theory, the attraction between dissociated ions of ionic solids in dilute aqueous solutions is caused by an ionic atmosphere that surrounds each ion. This means that each ion is, on average, surrounded by more ions of opposite charge than of like charge. A cluster of hydrated ions can act as a single unit rather than as individual ions. Thus, the effective total concentration is less than expected, based on the number of ions known to be present.

Ions of higher charge attract other ions very strongly. They therefore cluster more and have lower effective concentrations than ions with smaller charge. For example, ions formed by $MgSO_4$ have charges of 2+ and 2−. Ions formed by KCl have charges of 1+ and 1−. Note in **Table 3** that $MgSO_4$ in a solution does not depress the freezing point as much as the same concentration of KCl.

FIGURE 10 The salts applied to icy roads are electrolytes. They lower the freezing point of water and melt the ice.

SECTION REVIEW

1. What colligative properties are displayed by each of the following situations?

   a. Antifreeze is added to a car's cooling system to prevent freezing when the air temperature is below 0°C.

   b. Ice melts on sidewalks after salt has been spread on them.

2. Two moles of a nonelectrolyte solute are dissolved in 1 kg of an unknown solvent. The solution freezes at 7.8°C *below* its normal freezing point. What is the molal freezing-point constant of the unknown solvent? Suggest a possible identity of the solvent.

3. If two solutions of equal amounts in a U-tube are separated by a semipermeable membrane, will the level of the more-concentrated solution or the less-concentrated solution rise?

4. **a.** Calculate the expected freezing-point depression of a 0.200 *m* KNO_3 solution.

   b. Will the value you calculated match the actual freezing-point depression for this solution? Why or why not?

Critical Thinking

5. **INFERRING RELATIONSHIPS** The freezing-point depressions of aqueous solutions A, B, and C are −2.3°C, −1.2°C, and −4.1°C, respectively. Predict the order of the boiling-point elevations of these solutions, from lowest to highest. Explain your ranking.

Compounds in Aqueous Solutions

Vocabulary
dissociation
net ionic equation
spectator ions
ionization
hydronium ion
strong electrolyte
weak electrolyte

- The separation of ions that occurs when an ionic solid dissolves is called *dissociation*.
- When two different ionic solutions are mixed, a precipitate may form if ions from the two solutions react to form an insoluble compound.
- A net ionic equation for a reaction in aqueous solution includes only compounds and ions that change chemically in the reaction. Spectator ions are ions that do not take part in such a reaction.
- Formation of ions from molecular compounds is called *ionization*. A molecular compound may ionize in an aqueous solution if the attraction of the polar water molecules is strong enough to break the polar-covalent bonds of the solute molecules.
- An H_3O^+ ion is called a *hydronium ion*.
- All, or almost all, of a dissolved strong electrolyte exists as ions in an aqueous solution, whereas a relatively small amount of a dissolved weak electrolyte exists as ions in an aqueous solution.

Colligative Properties of Solutions

Vocabulary
colligative properties
nonvolatile substance
molal freezing-point constant, K_f
freezing-point depression, Δt_f
molal boiling-point constant, K_b
boiling-point elevation, Δt_b
semipermeable membrane
osmosis
osmotic pressure

- Colligative properties of solutions depend only on the total number of solute particles present. Boiling-point elevation, freezing-point depression, vapor-pressure lowering, and osmotic pressure are colligative properties.
- The molal boiling-point and freezing-point constants are used to calculate boiling-point elevations and freezing-point depressions of solvents containing nonvolatile solutes.
- Electrolytes have a greater effect on the freezing and boiling points of solvents than nonelectrolytes do.
- Except in very dilute solutions, the values of colligative properties of electrolyte solutions are less than expected because of the attraction between ions in solution.

Compounds in Aqueous Solutions

SECTION 1 REVIEW

1. How many moles of ions are contained in 1 L of a 1 M solution of KCl? of $Mg(NO_3)_2$?
2. Use **Table 1** to predict whether each of the following compounds is considered soluble or insoluble:
   a. KCl
   b. $NaNO_3$
   c. AgCl
   d. $BaSO_4$
   e. $Ca_3(PO_4)_2$
   f. $Pb(ClO_3)_2$
   g. $(NH_4)_2S$
   h. $PbCl_2$ (in cold water)
   i. FeS
   j. $Al_2(SO_4)_3$
3. What is a net ionic equation?
4. a. What is ionization?
   b. Distinguish between ionization and dissociation.
5. a. Define and distinguish between strong electrolytes and weak electrolytes.
   b. Give two examples of each type.
6. What determines the strength of a solute as an electrolyte?
7. Distinguish between the use of the terms *strong* and *weak* and the use of the terms *dilute* and *concentrated* when describing electrolyte solutions.

PRACTICE PROBLEMS

8. Write the equation for the dissolution of each of the following ionic compounds in water. (Hint: See Sample Problem A.)
   a. KI   c. $MgCl_2$
   b. $NaNO_3$ d. Na_2SO_4
9. For the compounds listed in the previous problem, determine the number of moles of each ion produced as well as the total number of moles of ions produced when 1 mol of each compound dissolves in water.

10. Write the equation for the dissolution of each of the following in water, and then indicate the total number of moles of solute ions formed.
    a. 0.50 mol strontium nitrate
    b. 0.50 mol sodium phosphate
11. Using **Table 1,** write the balanced chemical equation, write the overall ionic equation, identify the spectator ions and possible precipitates, and write the net ionic equation for each of the following reactions. (Hint: See Sample Problem B.)
    a. mercury(II) chloride (aq) + potassium sulfide $(aq) \longrightarrow$
    b. sodium carbonate (aq) + calcium chloride $(aq) \longrightarrow$
    c. copper(II) chloride (aq) + ammonium phosphate $(aq) \longrightarrow$
12. Identify the spectator ions in the reaction between KCl and $AgNO_3$ in an aqueous solution.
13. Copper(II) chloride and lead(II) nitrate react in aqueous solutions by double displacement. Write the balanced chemical equation, the overall ionic equation, and the net ionic equation for this reaction. If 13.45 g of copper(II) chloride react, what is the maximum amount of precipitate that could be formed?

Colligative Properties of Solutions

SECTION 2 REVIEW

14. How does the presence of a nonvolatile solute affect each of the following properties of the solvent in which the solute is dissolved?
    a. vapor pressure
    b. freezing point
    c. boiling point
    d. osmotic pressure
15. Using **Figure 6** as a guide, make a graph of vapor pressure versus temperature that shows the comparison of pure water, a solution with x concentration of solute, and a solution with $2x$ concentration of solute. What is the relationship between Δt_f for the x solution and Δt_f for the $2x$ solution?

16. a. Why does the level of the more-concentrated solution rise when two solutions of different concentrations are separated by a semipermeable membrane?

b. When does the level of the solution stop rising?

c. When the level stops rising, what is the net movement of water molecules across the membrane?

17. a. Compare the effects of nonvolatile electrolytes with the effects of nonvolatile non-electrolytes on the freezing and boiling points of solvents in which they are dissolved.

b. Why are such differences observed?

18. Why does the actual freezing-point depression of an electrolytic solution differ from the freezing-point depression calculated on the basis of the concentration of particles?

PRACTICE PROBLEMS

19. Determine the freezing-point depression of H_2O in each of the following solutions. (Hint: See Sample Problem C.)

a. 1.50 m solution of $C_{12}H_{22}O_{11}$ (sucrose) in H_2O

b. 171 g of $C_{12}H_{22}O_{11}$ in 1.00 kg H_2O

c. 77.0 g of $C_{12}H_{22}O_{11}$ in 400. g H_2O

20. Given the following freezing-point depressions, determine the molality of each solution of an unknown nonelectrolyte in water. (Hint: See Sample Problem D.)

a. $-0.930°C$

b. $-3.72°C$

c. $-8.37°C$

21. A solution contains 20.0 g of $C_6H_{12}O_6$ (glucose) in 250. g of water.

a. What is the freezing-point depression of the solvent?

b. What is the freezing point of the solution?

22. How many grams of antifreeze, $C_2H_4(OH)_2$, would be required per 500. g of water to prevent the water from freezing at a temperature of $-20.0°C$?

23. Pure benzene, C_6H_6, freezes at 5.45°C. A solution containing 7.24 g $C_2Cl_4H_2$ in 115 g of benzene (specific gravity = 0.879) freezes at 3.55°C. Based on these data, what is the molal freezing-point constant for benzene?

24. If 1.500 g of a solute that has a molar mass of 125.0 g were dissolved in 35.00 g of camphor, what would be the resulting freezing point of the solution?

25. Determine the boiling-point elevation of H_2O in each of the following solutions. (Hint: See Sample Problem E.)

a. 2.5 m solution of $C_6H_{12}O_6$ (glucose) in H_2O

b. 3.20 g $C_6H_{12}O_6$ in 1.00 kg H_2O

c. 20.0 g $C_{12}H_{22}O_{11}$ (sucrose) in 500. g H_2O

26. Given the following boiling points, determine the molality of each water solution.

a. 100.25°C

b. 101.53°C

c. 102.805°C

27. Given 1.00 m aqueous solutions of each of the following electrolytic substances, determine the expected change in the freezing point of the solvent. (Hint: See Sample Problem F.)

a. KI

b. $CaCl_2$

c. $Ba(NO_3)_2$

28. What is the expected change in the freezing point of water for a 0.015 m aqueous solution of $AlCl_3$?

29. What is the expected freezing point of a solution containing 85.0 g NaCl dissolved in 450. g of water?

30. Determine the expected boiling point of a solution made by dissolving 25.0 g of barium chloride in 0.150 kg of water.

31. The change in the boiling point of water for an aqueous solution of potassium iodide is 0.65°C. Determine the molal concentration of potassium iodide.

32. The freezing point of an aqueous solution of barium nitrate is $-2.65°C$. Determine the molal concentration of barium nitrate.

33. Calculate the expected freezing point of a solution containing 1.00 kg H_2O and 0.250 mol NaCl.

34. Experimental data for a 1.00 m MgI_2 aqueous solution indicate an actual change in the freezing point of water of $-4.78°C$. Find the expected change in the freezing point of water. Suggest a possible reason for the discrepancy between the experimental and expected values.

MIXED REVIEW

35. Given 0.01 *m* aqueous solutions of each of the following, arrange the solutions in order of increasing change in the freezing point of the solution.
 a. NaI
 b. $CaCl_2$
 c. K_3PO_4
 d. $C_6H_{12}O_6$ (glucose)

36. What is the molal concentration of an aqueous calcium chloride solution that freezes at –2.43°C?

37. a. Write the balanced formula equation that shows the possible products of a double-displacement reaction between calcium nitrate and sodium chloride.
 b. Using **Table 1,** determine whether there is a precipitate.
 c. Does this reaction occur?

38. Write a balanced equation to show what occurs when hydrogen bromide dissolves and reacts with water. Include a hydronium ion in the equation.

39. Write the equation for the dissolution of each of the following in water, and then indicate the total number of moles of solute ions formed.
 a. 0.275 mol of potassium sulfide
 b. 0.15 mol of aluminum sulfate

40. Calculate the expected change in the boiling point of water in a solution made up of 131.2 g of silver nitrate, $AgNO_3$, in 2.00 kg of water.

41. Nitrous acid, HNO_2, is a weak electrolyte. Nitric acid, HNO_3, is a strong electrolyte. Write equations to represent the ionization of each in water. Include the hydronium ion, and show the appropriate kind of arrow in each equation.

42. Find the boiling point of an aqueous solution containing a nonelectrolyte that freezes at –6.51°C.

43. Write a balanced equation for the dissolution of sodium carbonate, Na_2CO_3, in water. Find the number of moles of each ion produced when 0.20 mol of sodium carbonate dissolves. Then, find the total number of moles of ions.

44. Given the reaction below and the information in **Table 1,** write the net ionic equation for the reaction.
 potassium phosphate (*aq*) + lead (II) nitrate (*aq*)

45. Find the expected freezing point of a water solution that contains 268 g of aluminum nitrate, $Al(NO_3)_3$, in 8.50 kg of water.

CRITICAL THINKING

46. Applying Models
 a. You are conducting a freezing-point determination in the laboratory by using an aqueous solution of KNO_3. The observed freezing point of the solution is –1.15°C. Using a pure water sample, you recorded the freezing point of the pure solvent on the same thermometer as 0.25°C. Determine the molal concentration of KNO_3. Assume that there are no forces of attraction between ions.
 b. You are not satisfied with the result in part (a) because you suspect that you should not ignore the effect of ion interaction. You take a 10.00 mL sample of the solution. After carefully evaporating the water from the solution, you obtain 0.415 g KNO_3. Determine the actual molal concentration of KNO_3 and the percentage difference between the concentration observed in the freezing-point determination and the actual concentration of KNO_3. Assume that the solution's density is 1.00 g/mL.

47. Analyzing Information The observed freezing-point depression for electrolyte solutions is sometimes less than the calculated value. Why does this occur? Is the difference greater for concentrated solutions or dilute solutions?

48. Analyzing Information The osmotic pressure of a dilute solution can be calculated as follows:

$\pi = MRT$
π = osmotic pressure
M = concentration in moles per liter
R = ideal gas constant
T = absolute temperature of the solution

How does the osmotic-pressure equation compare with the ideal gas law?

USING THE HANDBOOK

49. Common reactions for Group 13 elements are found in the *Elements Handbook*. Review this material, and answer the following:
  a. Write net ionic equations for each of the example reactions shown on page 809.
  b. Which reactions did not change when written in net ionic form? Why?

50. Common reactions for Group 14 elements are found in the *Elements Handbook*. Review this material, and answer the following:
  a. Write net ionic equations for each of the example reactions shown on page 813.
  b. Which reactions did not change when written in net ionic form? Why?

RESEARCH & WRITING

51. Find out how much salt a large northern city, such as New York City or Chicago, uses on its streets in a typical winter. What environmental problems result from this use of salt? What substitutes for salt are being used to melt ice and snow?

52. Research the role of electrolytes and electrolytic solutions in your body. Find out how electrolytes work in the functioning of nerves and muscles. What are some of the health problems that can arise from an imbalance of electrolytes in body fluids?

ALTERNATIVE ASSESSMENT

53. Performance Determine the freezing point of four different mixtures of water and ethylene glycol (use commercial antifreeze). What mixture has the lowest freezing point?

54. Performance Find the optimum mixture of salt and ice for reducing the temperature of the chilling bath for an ice-cream freezer. Use your data to write a set of instructions on how to prepare the chilling bath for making ice cream.

55. Performance Using a low-voltage dry cell, assemble a conductivity apparatus. Secure several unknown aqueous solutions of equal molality from your instructor, and use the apparatus to distinguish the electrolytes from the non-electrolytes. Among those identified as electrolytes, rank their relative strengths as conductors from good to poor.

56. Performance Using equal volumes of the unknown solutions from the preceding activity, explain how you could use the freezing-point depression concept to distinguish the electrolytes from the nonelectrolytes. Explain how you could determine the number of ions contained per molecule among the solutes identified as electrolytes. Design and conduct an experiment to test your theories.

Math Tutor BOILING AND FREEZING POINTS OF SOLUTIONS

The presence of a nonvolatile solute causes the freezing point of a solution to be lower and the boiling point to be higher than those of the pure solvent. The freezing-point depression, Δt_f, is the amount that the freezing point is lowered. It is calculated by using the formula $\Delta t_f = K_f m$. The boiling-point elevation, Δt_b, is the amount that the boiling point is elevated. It is found by using the equation $\Delta t_b = K_b m$. To determine Δt_f or Δt_b, you need to know its *molal* concentration, m (moles of solute per kilogram of solvent). You also need to know the molal freezing-point constant, K_f, or the molal boiling-point constant, K_b. The values for K_f and K_b depend on the solvent and are given in **Table 2.** When these equations are used for electrolytes, the molality represents the total number of moles of ions in solution.

Problem-Solving **TIPS**

- Make sure that you find the molal concentration, not the molar concentration.
- For electrolytes, calculate the total number of moles of ions in solution.

SAMPLE

What is the theoretical boiling point of a solution of 247 g of potassium chloride, KCl, dissolved in 2.90 kg of water? Potassium chloride is a strong electrolyte.

First, determine the molality of the KCl that dissolved.

$$\text{mass of solute (g)} \times \frac{1 \text{ mol solute}}{\text{molar mass of solute (g)}} = \text{amount of solute (mol)}$$

$$247 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} = 3.31 \text{ mol KCl}$$

$$\text{molality KCl}(m) = \frac{\text{mol solute particles}}{\text{mass of solvent (kg)}} = \frac{3.31 \text{ mol KCl}}{2.90 \text{ kg H}_2\text{O}} = 1.14 \, m \text{ KCl}$$

Because KCl is an electrolyte, the total moles of ions in solution must now be determined. The equation $\text{KCl}(s) \longrightarrow \text{K}^+(aq) + \text{Cl}^-(aq)$ shows that 1 mol KCl will yield 2 mol of ions.

To find the total molality of particles in solution, multiply the molality of the solute by the moles of ions produced by the dissociation: $2 \times 1.14 \, m \text{ KCl} = 2.28 \, m$.

Next, use the equation $\Delta t_b = K_b m$, where K_b for water is $0.51°\text{C}/m$ and the value for m is 2.28.

$$\Delta t_b = (0.51°\text{C}/m)(2.28 \, m) = 1.16°\text{C}$$

The new boiling point is equal to the sum of the boiling point of the solvent and Δt_b.

$$\text{boiling point of solution} = 100°\text{C} + 1.16°\text{C} = 101.16°\text{C}$$

PRACTICE PROBLEMS

1. What is the freezing point of a solution containing 28.0 g of the strong electrolyte calcium chloride, $CaCl_2$, dissolved in 295 g of water?

2. What is the boiling point of a solution composed of 850 g of ethylene glycol, $C_2H_6O_2$, mixed with 1100 g of water?

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Acetic acid is a weak electrolyte because it
   A. is miscible with water.
   B. forms hydronium and hydroxide ions in aqueous solution.
   C. lowers the freezing point of water.
   D. ionizes only slightly in aqueous solution.

2. Which of the following solutions would contain the highest concentration of hydronium ions, H_3O^+?
   A. 0.10 M HCl
   B. 0.10 M HF
   C. 0.10 M CH_3COOH
   D. 0.10 M NaCl

3. Which of the following is the best representation of the precipitation reaction that occurs when aqueous solutions of sodium carbonate and calcium chloride are mixed?
   A. $Na^+(aq) + Cl^-(aq) \longrightarrow 2NaCl(s)$
   B. $2Na^+(aq) + CO_3^{2-}(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \longrightarrow CaCO_3(s) + 2NaCl(s)$
   C. $Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$
   D. No precipitation reaction occurs.

4. Which of the following is *not* a colligative property?
   A. molality
   B. vapor-pressure lowering
   C. boiling-point elevation
   D. osmotic pressure

5. Solution A contains 0.1 mol of sucrose, $C_{12}H_{22}O_{11}$, dissolved in 500 g of water. Solution B contains 0.1 mol of sodium chloride, NaCl, in 500 g of water. Which of the following statements about these solutions is true?
   A. Both solutions have the same vapor pressure.
   B. Solution A would boil at a higher temperature than solution B would.
   C. Solution A would freeze at a higher temperature than solution B would.
   D. Both solutions would boil at the same temperature.

The table below shows the freezing points of solutions of three nonvolatile solutes, X, Y, and Z, in water. Refer to the table to answer items 6 and 7. The K_f for water is $-1.86°C/m$.

Solute	Solute (mol)	Water (g)	Freezing point (°C)
X	1.00	1000	−5.58
Y	1.00	1000	−1.86
Z	1.00	1000	−3.72

6. Which of the following statements is *not* true?
   A. All three solutes are nonelectrolytes.
   B. Solute X could be $Ca(NO_3)_2$.
   C. Solute Y could be sucrose.
   D. Solute Z could be KCl.

7. Which of the solutions described in the table would have the highest boiling point?
   A. the solution of solute X
   B. the solution of solute Y
   C. the solution of solute Z
   D. All three solutions would have the same boiling point.

SHORT ANSWER

8. An aqueous solution of an unknown quantity of a nonelectrolyte solute is found to have a freezing point of −0.58°C. What is the molal concentration of the solution?

9. Give the names and formulas of two strong electrolytes.

EXTENDED RESPONSE

10. Write the formula equation, the overall ionic equation, and the net ionic equation for the precipitation reaction that occurs when solutions of zinc chloride, $ZnCl_2$, and sodium sulfide, Na_2S, are mixed.

11. Distinguish between dissociation and ionization. Write one chemical equation to illustrate each of these terms.

Test TIP Before choosing an answer to a question, try to answer the question without looking at the answer choices on the test.

Testing Water for Ions

OBJECTIVES

- *Observe* chemical reactions involving aqueous solutions of ions.

- *Relate* observations of chemical properties to the presence of ions.

- *Infer* whether an ion is present in a water sample.

- *Apply* concepts concerning aqueous solutions of ions.

MATERIALS

- 24-well microplate lid

- fine-tipped dropper bulbs, labeled, with solutions, 10

- overhead projector (optional)

- paper towels

- solution 1: reference (all ions)

- solution 2: distilled water (no ions)

- solution 3: tap water (may have ions)

- solution 4: bottled spring water (may have ions)

- solution 5: local river or lake water (may have ions)

- solution 6: solution X, prepared by your teacher (may have ions)

- solution A: NaSCN solution (test for Fe^{3+})

- solution B: $Na_2C_2O_4$ solution (test for Ca^{2+})

- solution C: $AgNO_3$ solution (test for Cl^-)

- solution D: $Sr(NO_3)_2$ solution (test for SO_4^{2-})

- white paper

BACKGROUND

The physical and chemical properties of aqueous solutions are affected by small amounts of dissolved ions. For example, if a water sample has enough Mg^{2+} or Ca^{2+} ions, it does not create lather when soap is added. This is common in places where there are many minerals in the water (hard water). Other ions, such as Pb^{2+} and Co^{2+}, can accumulate in body tissues; therefore, solutions of these ions are poisonous.

Because some sources of water may contain harmful or unwanted substances, it is important to find out what ions are present. In this experiment, you will test various water samples for the presence of four common, nontoxic ions: Fe^{3+}, Ca^{2+}, Cl^-, and SO_4^{2-}. Some of the samples may contain these ions in very small concentrations, so make very careful observations.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Create a data table in your lab notebook following your teacher's instructions. Record all of your observations in your data table.

2. Place the 24-well microplate lid in front of you on a white background. Label the columns and rows as instructed by your teacher. The coordinates will designate the individual circles. For example, the circle in the top right corner would be 1-D.

PROCEDURE

1. Obtain labeled dropper bulbs containing the six different solutions from your teacher.

2. Place a drop of the solution from bulb 1 into circles 1-A, 1-B, 1-C, and 1-D (the top row). Solution 1 contains all four of the dissolved ions, so these drops will show what a **positive** test for each ion looks like. **Be careful to keep the solutions in the appropriate circles. Any spills will cause poor results.**

3. Place a drop of the solution from bulb 2 into each of the circles in row 2. This solution is distilled water and should not contain any of the ions. It will show what a **negative** test looks like.

4. Place a drop from bulb 3 into each of the circles in row 3 and a drop from bulb 4 into each of the circles in row 4. Follow the same procedure for bulb 5 (into row 5) and bulb 6 (into row 6). These solutions may or may not contain ions. The materials list gives contents of each bulb.

5. Now that each circle contains a solution to be analyzed, use the solutions in bulbs A–D to test for the presence of the ions. Bulb A contains NaSCN, sodium thiocyanate, which reacts with any Fe^{3+} to form the complex ion $Fe(SCN)^{2+}$, which results in a deep red solution. Bulb B contains $Na_2C_2O_4$, sodium oxalate, which reacts with Ca^{2+} ions. Bulb C contains $AgNO_3$, silver nitrate, which reacts with Cl^- ions. Bulb D contains $Sr(NO_3)_2$, strontium nitrate, which reacts with SO_4^{2-} ions. The contents of bulbs B–D react with the specified ion to yield insoluble precipitates.

6. **Holding the tip of bulb A 1 to 2 cm above the drop of water to be tested,** add one drop of solution A to the drop of reference solution in circle 1-A and one drop to the distilled water in circle 2-A. Circle 1-A should show a positive test, and circle 2-A should show a negative test. In your data table, record your observations about what the positive and negative tests look like.

7. Use the NaSCN solution in bulb A to test the rest of the water drops in column A to determine whether they contain the Fe^{3+} ion. Record your observations in your data table. For each of the tests in which the ion was present, specify whether it seemed to be at a high, moderate, or low concentration.

8. Follow the procedure used for bulb A with bulbs B, C, and D to test for the other ions. Record your observations about the test results. Specify whether the solutions contained Ca^{2+}, Cl^-, or SO_4^{2-} and whether the ions seemed to be present at a high, moderate, or low concentration. A black background may be useful for these three tests.

9. If some of the results are difficult to discern, place your microplate on an overhead projector. Examine the drops for signs of cloudiness. Looking at the drops from the side, keep your line of vision 10° to 15° above the plane of the lid. Compare each drop tested with the control drops in row 2. If any sign of cloudiness is detected in a test sample, it is due to the Tyndall effect and is a positive test result. Record your results.

CLEANUP AND DISPOSAL

10. Clean all equipment and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. Organizing Ideas: Describe what each positive test looked like. Write the balanced chemical equations and net ionic equations for each of the positive tests.

CONCLUSIONS

1. Organizing Conclusions: List the solutions that you tested and the ions that you found in each solution. Include notes on whether the concentration of each ion was high, moderate, or low based on your observations.

2. Predicting Outcomes: Using your test results, predict which water sample would be the "hardest." Explain your reasoning.