CHAPTER 12

States of Matter

BIG Idea **Kinetic-molecular theory explains the different properties of solids, liquids, and gases.**

12.1 Gases

MAIN Idea Gases expand, diffuse, exert pressure, and can be compressed because they are in a low-density state consisting of tiny, constantly-moving particles.

12.2 Forces of Attraction

MAIN Idea Intermolecular forces—including dispersion forces, dipole-dipole forces, and hydrogen bonds—determine a substance's state at a given temperature.

12.3 Liquids and Solids

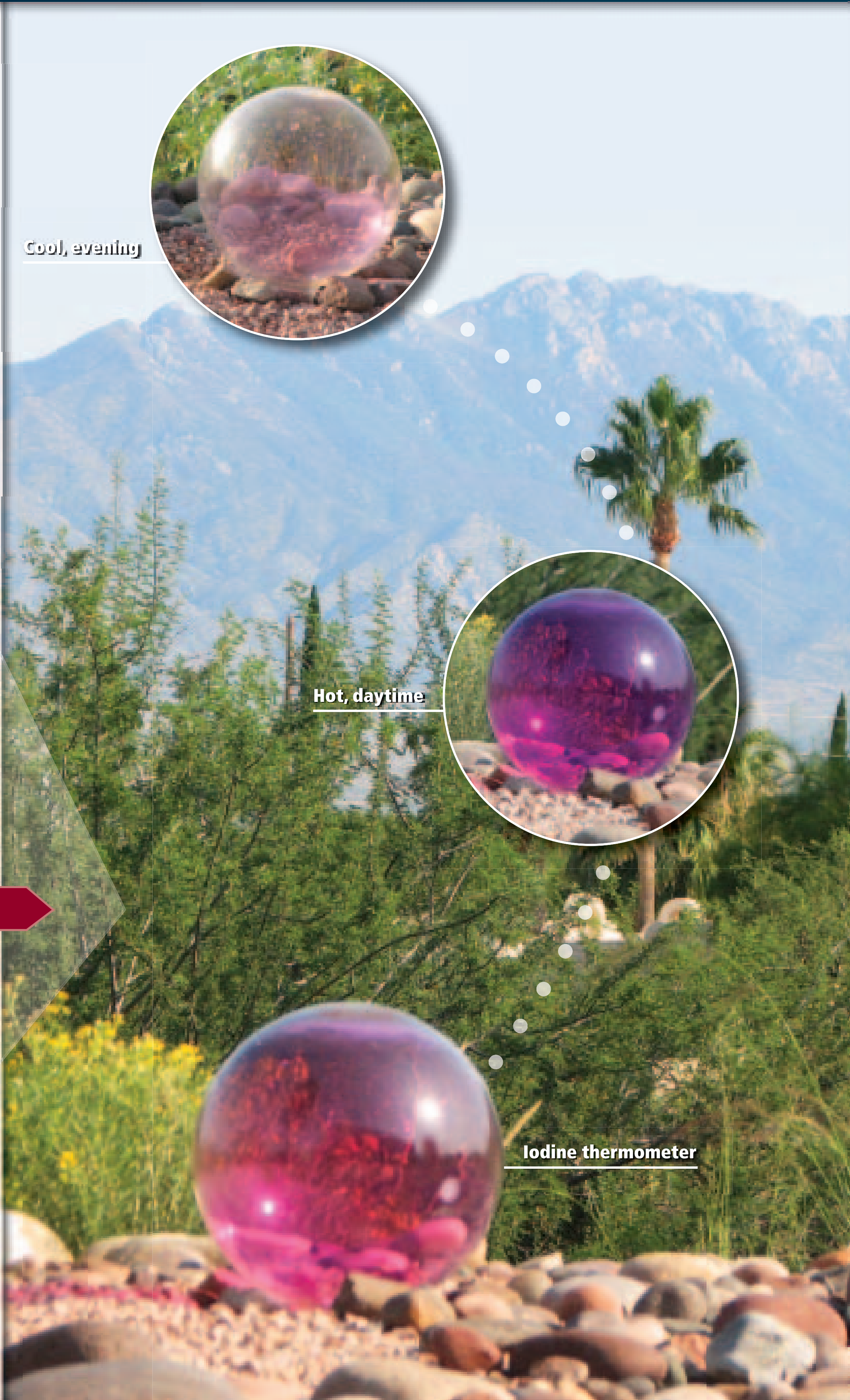
MAIN Idea The particles in solids and liquids have a limited range of motion and are not easily compressed.

12.4 Phase Changes

MAIN Idea Matter changes phase when energy is added or removed.

ChemFacts

- The iodine thermometer contains a few grams of iodine inside a sealed, round-bottom flask.
- As the outdoor temperature increases, the iodine changes from a solid directly to a gas.
- The deeper the violet color, the higher the temperature.



LAUNCH Lab

How do different liquids affect the speed of a sinking ball bearing?

You've probably noticed that different liquids might have vastly different properties. For example, liquids such as maple syrup, corn oil, and vegetable oil are much thicker than liquids such as water.

Procedure

1. Read and complete the lab safety form.
2. Fill a **100-mL graduated cylinder** with **water.** Be sure to fill it exactly to the 100-mL mark.
3. Place the end of a **ruler** on the tabletop. Drop a **ball bearing** (or other small, round object) from a mark on the ruler just above the surface of the water. Use a **stopwatch** to time the ball bearing as it sinks to the bottom. Record this time in a data table.
4. Repeat Steps 2 and 3 two more times, dropping the object from the same height each time. Calculate the average drop time of your three trials.
5. Repeat Steps 2–4 using **vegetable oil** instead of water.

Analysis

1. **Compare** the average drop time for the two liquids.
2. **Infer** the relationship between the times that you recorded and how easily the liquid flows as you pour it.

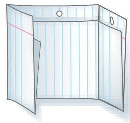
Inquiry How does temperature affect the speed with which a ball bearing sinks in a liquid? Develop a hypothesis, and design an experiment to test your hypothesis.

FOLDABLES™ Study Organizer

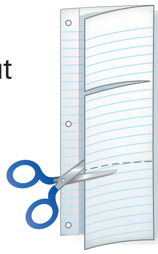
States of Matter Make the following Foldable to help you summarize information about three common states of matter.

STEP 1 Fold a sheet of paper in half lengthwise. Make the back edge about 2 cm longer than the front edge.

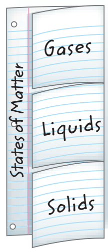
STEP 2 Fold into thirds.



STEP 3 Unfold and cut along the folds of the top flap to make three tabs.



STEP 4 Label the tabs as follows: *Gases*, *Liquids*, and *Solids*.



FOLDABLES **Use this Foldable with Sections 12.1 and 12.3.** As you read the sections, summarize information about three common states of matter in your own words.

Chemistry Online

Visit glencoe.com to:

- study the entire chapter online
- explore Concepts In Motion
- take Self-Check Quizzes
- use the Personal Tutor to work Example Problems step-by-step
- access Web links for more information, projects, and activities
- Find the Try at Home Lab, Viscosity Race

Matt Meadows

Section 12.1

Objectives

- **Use** the kinetic-molecular theory to explain the behavior of gases.
- **Describe** how mass affects the rates of diffusion and effusion.
- **Explain** how gas pressure is measured and calculate the partial pressure of a gas.

Review Vocabulary

kinetic energy: energy due to motion

New Vocabulary

kinetic-molecular theory
elastic collision
temperature
diffusion
Graham's law of effusion
pressure
barometer
pascal
atmosphere
Dalton's law of partial pressures

Gases

MAIN Idea Gases expand, diffuse, exert pressure, and can be compressed because they are in a low density state consisting of tiny, constantly-moving particles.

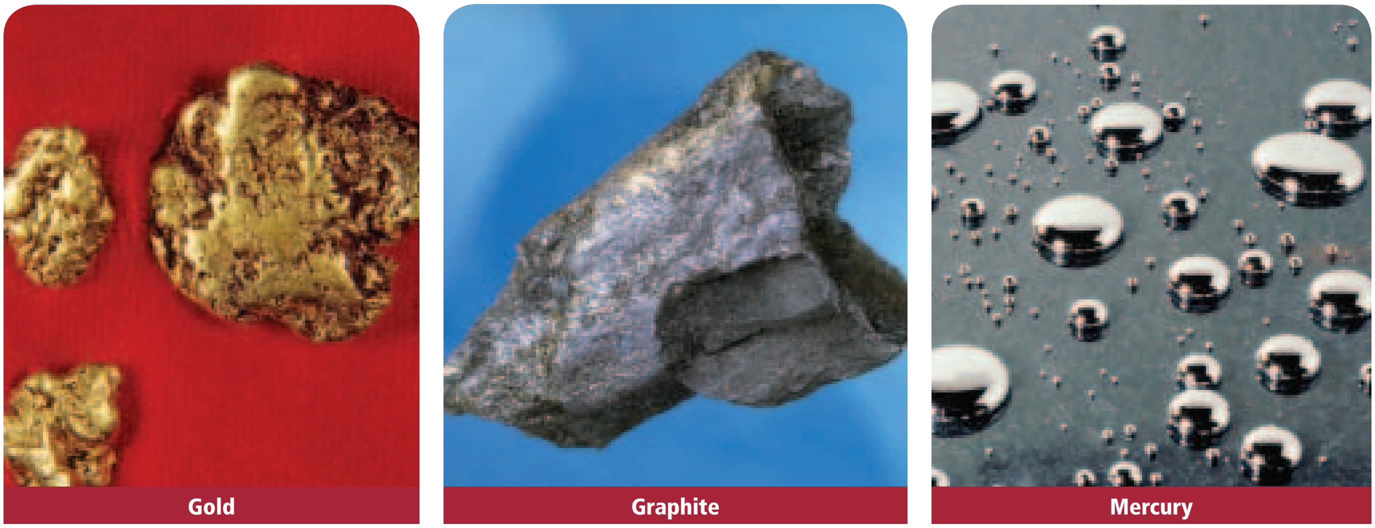
Real-World Reading Link If you have gone camping, you might have slept on an air-filled mattress. How did the mattress compare to lying on the ground? It was probably warmer and more comfortable. The properties of the air mattress are due to the particles that make up the air inside it.

The Kinetic-Molecular Theory

You have learned that composition—the types of atoms present—and structure—their arrangement—determine the chemical properties of matter. Composition and structure also affect the physical properties of matter. Based solely on physical appearance, you can distinguish between gold, graphite, and mercury, as shown in **Figure 12.1.** By contrast, substances that are gases at room temperature usually display similar physical properties despite their different compositions. Why is there so little variation in behavior among gases? Why are the physical properties of gases different from those of liquids and solids?

By the eighteenth century, scientists knew how to collect gaseous products by displacing water. Now, they could observe and measure properties of individual gases. About 1860, chemists Ludwig Boltzmann and James Maxwell, who were working in different countries, each proposed a model to explain the properties of gases. That model is the kinetic-molecular theory. Because all of the gases known to Boltzmann and Maxwell contained molecules, the name of the model refers to molecules. The word *kinetic* comes from a Greek word meaning *to move*. Objects in motion have energy called kinetic energy. The **kinetic-molecular theory** describes the behavior of matter in terms of particles in motion. The model makes several assumptions about the size, motion, and energy of gas particles.

Figure 12.1 You can distinguish some materials by looking at them, but this is not true for many gases.



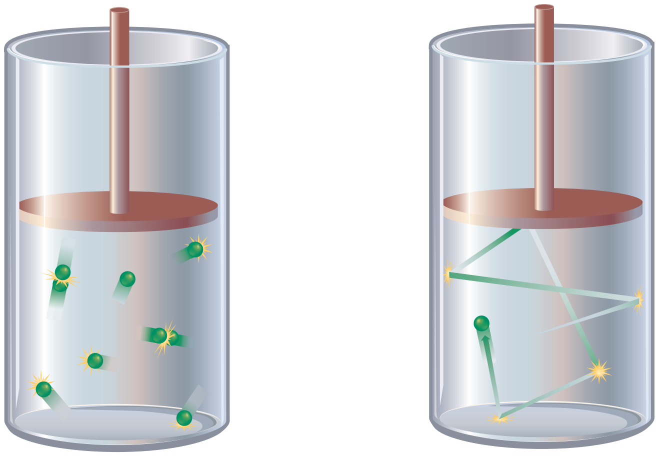


Figure 12.2 Kinetic energy can be transferred between gas particles during an elastic collision.

Explain ***the influence that gas particles have on each other, both in terms of collisions and what happens to particles between collisions.***

Particle size Gases consist of small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space. Because gas particles are far apart, they experience no significant attractive or repulsive forces.

Particle motion Gas particles are in constant, random motion. Particles move in a straight line until they collide with other particles or with the walls of their container, as shown in **Figure 12.2.** Collisions between gas particles are elastic. An **elastic collision** is one in which no kinetic energy is lost. Kinetic energy can be transferred between colliding particles, but the total kinetic energy of the two particles does not change.

Particle energy Two factors determine the kinetic energy of a particle: mass and velocity. The kinetic energy of a particle can be represented by the following equation.

$$KE = \frac{1}{2}mv^2$$

KE is kinetic energy, m is the mass of the particle, and v is its velocity. Velocity reflects both the speed and the direction of motion. In a sample of a single gas, all particles have the same mass, but all particles do not have the same velocity. Therefore, all particles do not have the same kinetic energy. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter.

Explaining the Behavior of Gases

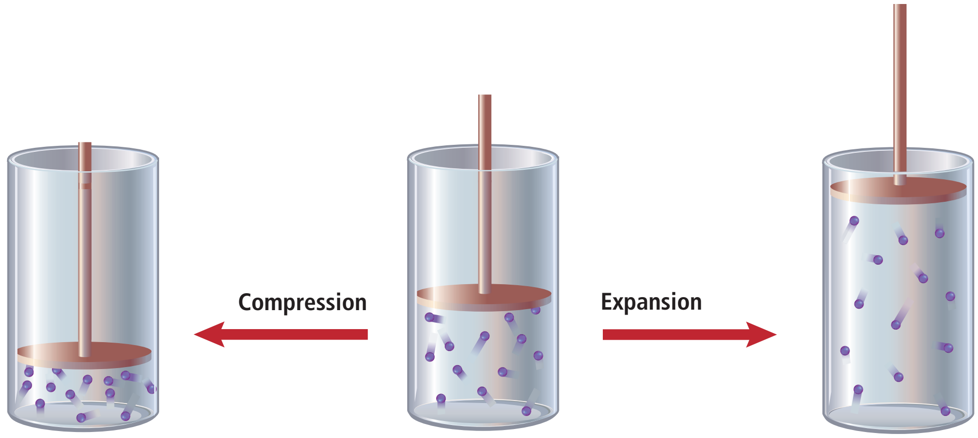
The kinetic-molecular theory helps explain the behavior of gases. For example, the constant motion of gas particles allows a gas to expand until it fills its container, such as when you blow up a beach ball. As you blow air into the ball, the air particles spread out and fill the inside of the container—the beach ball.

Low density Remember that density is mass per unit volume. The density of chlorine gas is 2.95×10^{-3} g/mL at 20°C; the density of solid gold is 19.3 g/mL. Gold is more than 6500 times as dense as chlorine. This large difference cannot be due only to the difference in mass between gold atoms and chlorine molecules (about 3:1). As the kinetic-molecular theory states, a great deal of space exists between gas particles. Thus, there are fewer chlorine molecules than gold atoms in the same volume.

VOCABULARY

WORD ORIGIN

Gas
comes from the Latin word *chaos*, which means *space*



■ **Figure 12.3** In a closed container, compression and expansion change the volume occupied by a constant mass of particles.
Relate *the change in volume to the density of the gas particles in each cylinder.*

Compression and expansion If you squeeze a pillow made of foam, you can compress it; that is, you can reduce its volume. The foam contains air pockets. The large amount of empty space between the particles in the air in those pockets allows the air to be pushed easily into a smaller volume. When you stop squeezing, the random motion of the particles fills the available space, and the pillow expands to its original shape. **Figure 12.3** illustrates what happens to the density of a gas in a container as it is compressed and as it is allowed to expand.

Diffusion and effusion According to the kinetic-molecular theory, there are no significant forces of attraction between gas particles. Thus, gas particles can flow easily past each other. Often, the space into which a gas flows is already occupied by another gas. The random motion of the gas particles causes the gases to mix until they are evenly distributed. **Diffusion** is the term used to describe the movement of one material through another. The term might be new, but you are probably familiar with the process. If food is cooking in the kitchen, you can smell it throughout the house because the gas particles diffuse. Particles diffuse from an area of high concentration (the kitchen) to one of low concentration (the other rooms in the house).

Effusion is a process related to diffusion. During effusion, a gas escapes through a tiny opening. What happens when you puncture a container, such as a balloon or a tire? In 1846, Thomas Graham conducted experiments to measure the rates of effusion for different gases at the same temperature. Graham designed his experiments so that the gases effused into a vacuum—space containing no matter. He discovered an inverse relationship between effusion rates and molar mass. **Graham's law of effusion** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

Graham's Law

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

The rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass.

The rate of diffusion depends mainly on the mass of the particles involved. Lighter particles diffuse more rapidly than heavier particles. Recall that different gases at the same temperature have the same average kinetic energy as described by the equation $KE = \frac{1}{2} mv^2$. However, the mass of gas particles varies from gas to gas. For lighter particles to have the same average kinetic energy as heavier particles, they must have, on average, a greater velocity.

Graham's law also applies to rates of diffusion, which is logical because heavier particles diffuse more slowly than lighter particles at the same temperature. Using Graham's law, you can set up a proportion to compare the diffusion rates for two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

Reading Check **Explain** why the rate of diffusion depends on the mass of the particles.

EXAMPLE Problem 12.1

Math Handbook
Square and Cube Roots
page 949

Graham's Law Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

1 Analyze the Problem

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known

molar mass$_{HCl}$ = 36.5 g/mol

molar mass$_{HCl}$ = 17.0 g/mol

Unknown

ratio of diffusion rates = ?

2 Solve for the Unknown

$$\frac{\text{Rate}_{NH_3}}{\text{Rate}_{HCl}} = \sqrt{\frac{\text{molar mass}_{HCl}}{\text{molar mass}_{NH_3}}}$$

State the ratio derived from Graham's law.

$$= \sqrt{\frac{36.5 \text{ g/mol}}{17.0 \text{ g/mol}}} = \mathbf{1.47}$$

Substitute molar mass$_{HCl}$ = 36.5 g/mol and molar mass$_{NH_3}$ = 17.0 g/mol.

The ratio of diffusion rates is 1.47.

3 Evaluate the Answer

A ratio of roughly 1.5 is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer also does. Note that the units cancel, and the answer is stated correctly without any units.

PRACTICE Problems

Extra Practice Page 984 and glencoe.com

1. Calculate the ratio of effusion rates for nitrogen (N_2) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.
3. **Challenge** What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

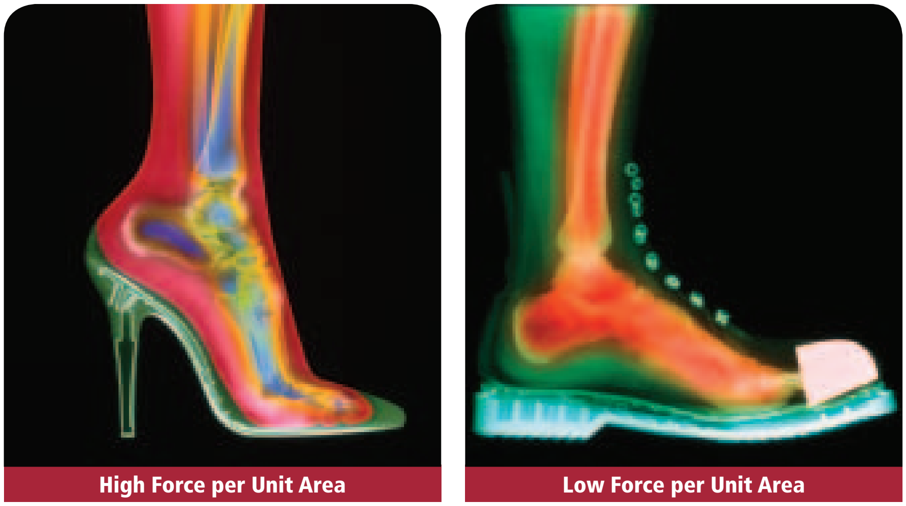


Figure 12.4 High-heeled shoes increase the pressure on a surface because the area touching the floor is reduced. In flatter-heeled shoes, such as boots, the force is applied over a larger area.

Infer *where the highest pressure is located between the floor and high-heel shoe.*

Gas Pressure

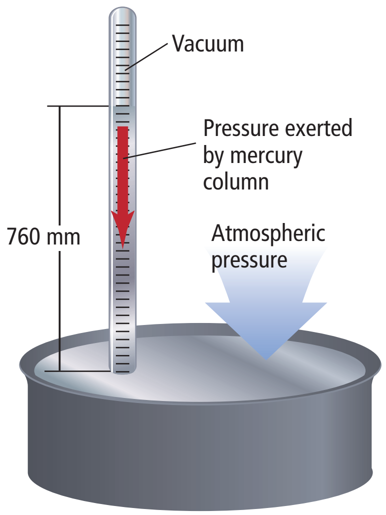
Have you watched someone try to walk across snow, mud, or hot asphalt in high heels? If so, you might have noticed that the heels sank into the soft surface. **Figure 12.4** shows why a person sinks when wearing high heels but does not sink when wearing boots. In each case, the force pressing down on the soft surface is related to the person's mass. With boots, the force is spread out over a larger area. **Pressure** is defined as force per unit area. The area of the bottom of a boot is much larger than the area of the bottom of a high heeled shoe. So, the pressure on the soft surface is less with a boot than it is with high heels.

Gas particles also exert pressure when they collide with the walls of their container. Because an individual gas particle has little mass, it can exert little pressure. However, a liter-sized container could hold 10^{22} gas particles. With this many particles colliding, the pressure can be high.

Air pressure Earth is surrounded by an atmosphere that extends into space for hundreds of kilometers. Because the particles in air move in every direction, they exert pressure in all directions. This pressure is called atmospheric pressure, or air pressure. Air pressure varies at different points on Earth. Because gravity is greater at the surface of Earth, there are more particles than at higher altitudes where the force of gravity is less. Fewer particles at higher elevations exert less force than the greater concentration of particles at lower altitudes. Therefore, air pressure is less at higher altitudes than it is at sea level. At sea level, atmospheric pressure is about one-kilogram per square centimeter.

Measuring air pressure Italian physicist Evangelista Torricelli (1608–1647) was the first to demonstrate that air exerted pressure. He noticed that water pumps were unable to pump water higher than about 10 m. He hypothesized that the height of a column of liquid would vary with the density of the liquid. To test this idea, Torricelli designed the equipment shown in **Figure 12.5.** He filled a thin glass tube that was closed at one end with mercury. While covering the open end so that air could not enter, he inverted the tube and placed it (open end down) in a dish of mercury. The open end was below the surface of the mercury in the dish. The height of the mercury in the tube fell to about 75 cm, which validated Terricelli's hypothesis because mercury is approximately 13.6 times more dense than water.

Figure 12.5 Torricelli was the first to show that the atmosphere exerted pressure.



Barometers The device that Torricelli invented is called a barometer. A **barometer** is an instrument used to measure atmospheric pressure. As Torricelli demonstrated, the height of the mercury in a barometer is always about 760 mm. The exact height of the mercury is determined by two forces. Gravity exerts a constant downward force on the mercury. This force is opposed by an upward force exerted by air pressing down on the surface of the mercury. Changes in air temperature or humidity cause air pressure to vary.

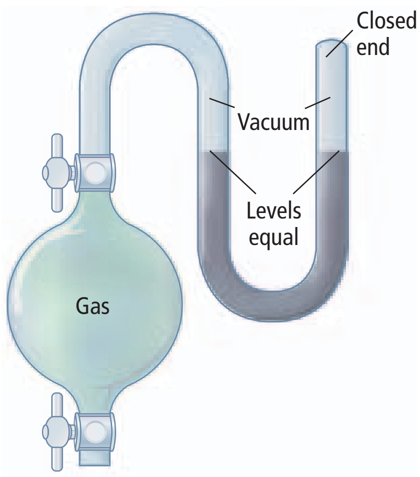
Manometers A manometer is an instrument used to measure gas pressure in a closed container. In a manometer, a flask is connected to a U-tube that contains mercury, as shown in **Figure 12.6.** When the valve between the flask and the U-tube is opened, gas particles diffuse out of the flask into the U-tube. The released gas particles push down on the mercury in the tube. The difference in the height of the mercury in the two arms is used to calculate the pressure of the gas in the flask.

Units of pressure The SI unit of pressure is the pascal (Pa). It is named for Blaise Pascal, a French mathematician and philosopher. The pascal is derived from the SI unit of force, the newton (N). One **pascal** is equal to a force of one newton per square meter: 1 Pa equals 1 N/m^2. Many fields of science still use more traditional units of pressure. For example, engineers often report pressure as pounds per square inch (psi). The pressures measured by barometers and manometers can be reported in millimeters of mercury (mm Hg). There is also a unit called the torr and another unit called a bar.

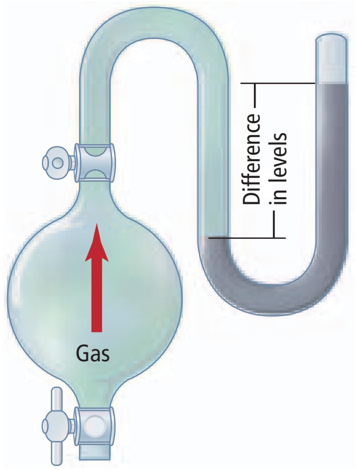
At sea level, the average air pressure is 101.3 kPa when the temperature is 0°C. Air pressure is often reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mm Hg or 760 torr or 101.3 kilopascals (kPa). **Table 12.1** compares different units of pressure. Because the units 1 atm, 760 mm Hg, and 760 torr are defined units, they should have as many significant figures as needed when used in calculations.

Table 12.1 Comparison of Pressure Units

Unit	Number Equivalent to 1 atm	Number Equivalent to 1 kPa
Kilopascal (kPa)	101.3 kPa	—
Atmosphere (atm)	—	0.009869 atm
Millimeters of mercury (mm Hg)	760 mm Hg	7.501 mm Hg
Torr	760 torr	7.501 torr
Pounds per square inch (psi or lb/in^2)	14.7 psi	0.145 psi
Bar	1.01 bar	100 kPa



Before gas is released into the U-tube, the mercury is at the same height in each arm.



After gas is released into the U-tube, the heights in the two arms are no longer equal.

Figure 12.6 A manometer measures the pressure of an enclosed gas.

DATA ANALYSIS LAB

*Based on Real Data

Make and Use Graphs

How are the depth of a dive and altitude related? Most divers dive at locations that are at or near sea level in altitude. However, divers in Saskatchewan, Alberta, and British Columbia, Canada, as well as much of the northwestern United States, dive at higher altitudes.

Think Critically

1. **Compare** Use the data in the table to make a graph of atmospheric pressure versus altitude.
2. **Calculate** What is your actual diving depth if your depth gauge reads 18 m, but you are at an altitude of 1800 m and your gauge does not compensate for altitude?
3. **Infer** Dive tables are used to determine how long it is safe for a diver to stay under water at a specific depth. Why is it important to know the correct depth of the dive?

Data and Observations

The table shows the pressure gauge correction factor for high altitude underwater diving.

Altitude Diving Correction Factors

Altitude (m)	Atmospheric Pressure (atm)	Pressure Gauge Correction Factor (m)
0	1.000	0.0
600	0.930	0.7
1200	0.864	1.4
1800	0.801	2.0
2400	0.743	2.7
3000	0.688	3.2

*Data obtained from: Sawatzky, D. 2000. Diving at Altitude Part I. *Diver Magazine*. June 2000.

Dalton's law of partial pressures When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. Illustrated in **Figure 12.7**, **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of 1 mol of any gas is the same. Dalton's law of partial pressures can be summarized by the equation at the top of the next page.

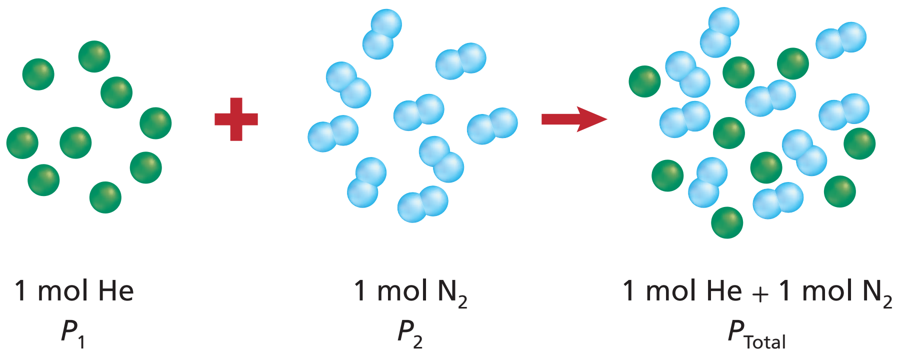


Figure 12.7 When gases mix, the total pressure of the mixture is equal to the sum of the partial pressures of the individual gases.

Determine *How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container?*

Dalton's Law of Partial Pressures

$$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$$

P_{total} represents total pressure. P_1, P_2, and P_3 represent the partial pressures of each gas up to the final gas, P_n.

To calculate the total pressure of a mixture of gases, add the partial pressures of each of the gases in the mixture.

Look again at **Figure 12.7.** What happens when 1 mol of helium and 1 mol of nitrogen are combined in a single closed container? Because neither the volume nor the number of particles changed, the pressures exerted by the two separate gases combined.

Chemistry Online

Personal Tutor For an online tutorial on Using Dalton's Law of Partial Pressures, visit glencoe.com.

EXAMPLE Problem 12.2

Math Handbook
Significant Figures
pages 949–951

The Partial Pressure of a Gas A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2 if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1 Analyze the Problem

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$P_{N_2} = 0.12$ atm

$P_{CO_2} = 0.70$ atm

$P_{total} = 0.97$ atm

Unknown

$P_{O_2} = ?$ atm

2 Solve for the Unknown

$P_{total} = P_{N_2} + P_{CO_2} + P_{O_2}$ — State Dalton's law of partial pressures.

$P_{O_2} = P_{total} - P_{CO_2} - P_{N_2}$ — Solve for P_{O_2}.

$P_{O_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$ — Substitute $P_{N_2} = 0.12$ atm, $P_{CO_2} = 0.70$ atm, and $P_{total} = 0.97$ atm.

$P_{O_2} = 0.15$ atm

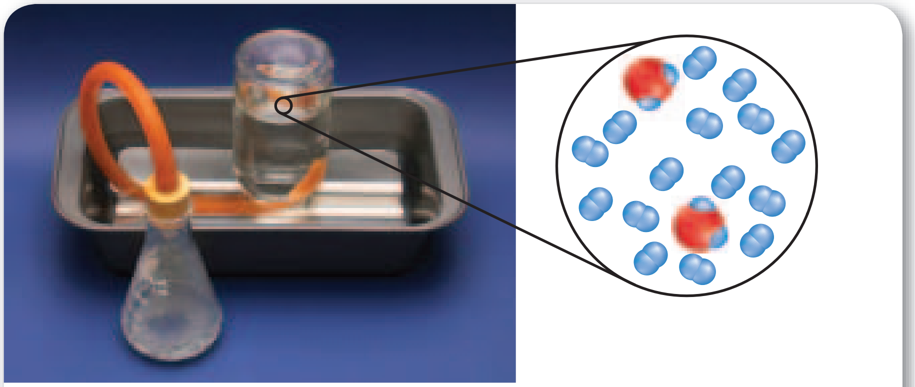
3 Evaluate the Answer

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure, 0.97 atm. The answer has two significant figures to match the data.

PRACTICE Problems

Extra Practice Page 984 and glencoe.com

4. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mm Hg and the partial pressure of helium is 439 mm Hg?
5. Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.
6. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.
7. **Challenge** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mm Hg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?



■ **Figure 12.8** In the flask, sulfuric acid (H_2SO_4) reacts with zinc to produce hydrogen gas. The hydrogen is collected at 20°C.

Calculate ***the partial pressure of hydrogen at 20°C if the total pressure of the hydrogen and water vapor mixture is 100.0 kPa.***

FOLDABLES Incorporate information from this section into your Foldable.

Using Dalton's law Partial pressures can be used to determine the amount of gas produced by a reaction. The gas produced is bubbled into an inverted container of water, as shown in **Figure 12.8.** As the gas collects, it displaces the water. The gas collected in the container will be a mixture of hydrogen and water vapor. Therefore, the total pressure inside the container will be the sum of the partial pressures of hydrogen and water vapor.

The partial pressures of gases at the same temperature are related to their concentration. The partial pressure of water vapor has a fixed value at a given temperature. You can look up the value in a reference table. At 20°C, the partial pressure of water vapor is 2.3 kPa. You can calculate the partial pressure of hydrogen by subtracting the partial pressure of water vapor from the total pressure.

As you will read in Chapter 13, knowing the pressure, volume, and temperature of a gas allows you to calculate the number of moles of the gas. Temperature and volume can be measured during an experiment. Once the temperature is known, the partial pressure of water vapor is used to calculate the pressure of the gas. The known values for volume, temperature, and pressure are then used to find the number of moles.

Section 12.1 Assessment

Section Summary

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Dalton's law of partial pressures is used to determine the pressures of individual gases in gas mixtures.
- Graham's law is used to compare the diffusion rates of two gases.

8. **MAIN Idea** **Explain** Use the kinetic theory to explain the behavior of gases.
9. **Describe** how the mass of a gas particle affects its rate of effusion and diffusion.
10. **Explain** how gas pressure is measured.
11. **Explain** why the container of water must be inverted when a gas is collected by displacement of water.
12. **Calculate** Suppose two gases in a container have a total pressure of 1.20 atm. What is the pressure of Gas B if the partial pressure of Gas A is 0.75 atm?
13. **Infer** whether or not temperature has any effect on the diffusion rate of a gas. Explain your answer.

Chemistry Online Self-Check Quiz glencoe.com

Section 12.2

Objectives

- **Describe** intramolecular forces.
- **Compare and contrast** intermolecular forces.

Review Vocabulary

polar covalent: a type of bond that forms when electrons are not shared equally

New Vocabulary

dispersion force
dipole-dipole force
hydrogen bond

Forces of Attraction

MAIN Idea Intermolecular forces—including dispersion forces, dipole-dipole forces, and hydrogen bonds—determine a substance's state at a given temperature.

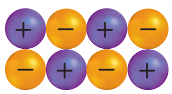
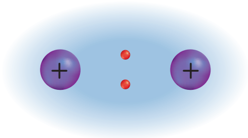
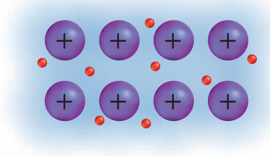
Real-World Reading Link You might be aware that water is one of the rare substances that is found as a solid, a liquid, and a gas at atmospheric conditions. This unique property, along with others that enable life as we understand it to exist, stems from the forces that exist between water molecules.

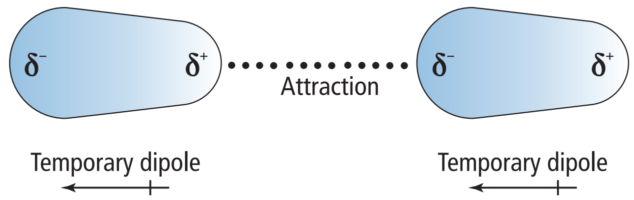
Intermolecular Forces

If all particles of matter at room temperature have the same average kinetic energy, why are some materials gases while others are liquids or solids? The answer lies with the attractive forces within and between particles. The attractive forces that hold particles together in ionic, covalent, and metallic bonds are called intramolecular forces. The prefix *intra-* means *within*. For example, intramural sports are competitions among teams from within a single school or district. The term *molecular* can refer to atoms, ions, or molecules. **Table 12.2** summarizes what you read about intramolecular forces in Chapters 7 and 8.

Intramolecular forces do not account for all attractions between particles. There are forces of attraction called intermolecular forces. The prefix *inter-* means *between* or *among*. For example, an interview is a conversation between two people. These forces can hold together identical particles, such as water molecules in a drop of water, or two different types of particles, such as carbon atoms in graphite and the cellulose particles in paper. The three intermolecular forces that will be discussed in this section are dispersion forces, dipole-dipole forces, and hydrogen bonds. Although some intermolecular forces are stronger than others, all intermolecular forces are weaker than the intramolecular forces involved in bonding.

Table 12.2 Comparison of Intramolecular Forces

Force	Model	Basis of Attraction	Example
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H_2
Metallic		metal cations and mobile electrons	Fe



■ **Figure 12.9** When two molecules are close together, the electron clouds repel each other, creating temporary dipoles. The δ sign represents an area of partial charge on the molecule.

Explain ***what the δ+ and δ− signs on a temporary dipole represent.***

Dispersion forces Recall that oxygen molecules are nonpolar because electrons are evenly distributed between the equally electronegative oxygen atoms. Under the right conditions, however, oxygen molecules can be compressed into a liquid. For oxygen to condense, there must be some force of attraction between its molecules.

The force of attraction between oxygen molecules is called a dispersion force. **Dispersion forces** are weak forces that result from temporary shifts in the density of electrons in electron clouds. Dispersion forces are sometimes called London forces after the German-American physicist who first described them, Fritz London.

Remember that the electrons in an electron cloud are in constant motion. When two molecules are in close contact, especially when they collide, the electron cloud of one molecule repels the electron cloud of the other molecule. The electron density around each nucleus is, for a moment, greater in one region of each cloud. Each molecule forms a temporary dipole. When temporary dipoles are close together, a weak dispersion force exists between oppositely charged regions of the dipoles, as shown in **Figure 12.9.**

Reading Check **Explain** why dispersion forces form.

Dispersion forces exist between all particles. Dispersion forces are weak for small particles, and these forces have an increasing effect as the number of electrons involved increases. Thus, dispersion forces tend to become stronger as the size of the particles increase. For example, fluorine, chlorine, bromine, and iodine exist as diatomic molecules. Recall that the number of nonvalence electrons increases from fluorine to chlorine to bromine to iodine. Because the larger halogen molecules have more electrons, there can be a greater difference between the positive and negative regions of their temporary dipoles and, thus, stronger dispersion forces. This difference in dispersion forces explains why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature.

Reading Check **Infer** the physical state of the element astatine at room temperature and explain your reasoning.

Dipole-dipole forces Polar molecules contain permanent dipoles; that is, some regions of a polar molecule are always partially negative and some regions of the molecule are always partially positive. These attractions between oppositely charged regions of polar molecules are called **dipole-dipole forces.** Neighboring polar molecules orient themselves so that oppositely charged regions align.

VOCABULARY

ACADEMIC VOCABULARY

Orient
to arrange in a specific position; to align in the same direction
The blooms of the flowers were all oriented toward the setting Sun.

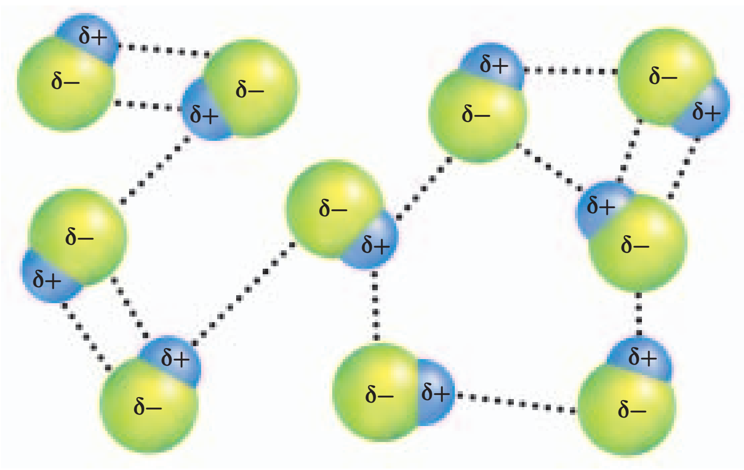


Figure 12.10 Neighboring polar molecules orient themselves so that oppositely charged regions align.

Identify ***the types of forces that are represented in this figure.***

When hydrogen-chloride gas molecules approach, the partially positive hydrogen atom in one molecule is attracted to the partially negative chlorine atom in another molecule. **Figure 12.10** shows multiple attractions among hydrogen-chloride molecules. Because the dipoles are permanent, you might expect dipole-dipole forces to be stronger than dispersion forces. This prediction holds true for small polar molecules with large dipoles. However, for many polar molecules, including the HCl molecules in **Figure 12.10,** dispersion forces dominate dipole-dipole forces.

Reading Check **Compare** dipole-dipole forces and dispersion forces.

Hydrogen bonds One special type of dipole-dipole attraction is called a hydrogen bond. A **hydrogen bond** is a dipole-dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair. Hydrogen bonds typically dominate both dispersion forces and dipole-dipole forces. For a hydrogen bond to form, hydrogen must be bonded to either a fluorine, oxygen, or nitrogen atom. These atoms are electronegative enough to cause a large partial positive charge on the hydrogen atom, yet small enough that their lone pairs of electrons can come close to hydrogen atoms. For example, in a water molecule, the hydrogen atoms have a large partial positive charge and the oxygen atom has a large partial negative charge. When water molecules approach, a hydrogen atom on one molecule is attracted to the lone pair of electrons on the oxygen atom on the other molecule, as shown in **Figure 12.11.**

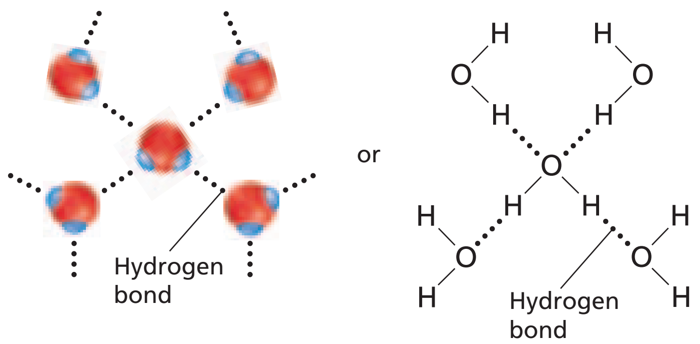
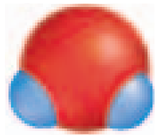
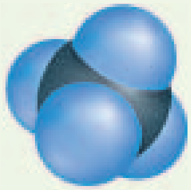
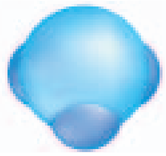


Figure 12.11 The hydrogen bonds between water molecules are stronger than typical dipole-dipole attractions because the bond between hydrogen and oxygen is highly polar.

Table 12.3	**Properties of Three Molecular Compounds**		
Compound	**Molecular Structure**	**Molar Mass (g)**	**Boiling Point (°C)**
Water (H_2O)		18.0	100
Methane (CH_4)		16.0	−33.4
Ammonia (NH_3)		17.0	−164

Hydrogen bonds explain why water is a liquid at room temperature, while compounds of comparable mass are gases. Look at the data in **Table 12.3.** The difference between methane and water is easy to explain. Because methane molecules are nonpolar, the only forces holding the molecules together are relatively weak dispersion forces. The difference between ammonia and water is not as obvious. Molecules of both compounds can form hydrogen bonds. Yet, ammonia is a gas at room temperature, which indicates that the attractive forces between ammonia molecules are not as strong. Because oxygen atoms are more electronegative than nitrogen atoms, the O–H bonds in water are more polar than the N–H bonds in ammonia. As a result, the hydrogen bonds between water molecules are stronger than those between ammonia molecules.

Section 12.2 Assessment

Section Summary

- Intramolecular forces are stronger than intermolecular forces.
- Dispersion forces are intermolecular forces between temporary dipoles.
- Dipole-dipole forces occur between polar molecules.

14. MAIN Idea **Explain** what determines a substance's state at a given temperature.

15. Compare and contrast intermolecular forces and describe intramolecular forces.

16. Evaluate Which of the molecules listed below can form hydrogen bonds? For which of the molecules would dispersion forces be the only intermolecular force? Give reasons for your answers.

a. H_2 **b.** H_2S **c.** HCl **d.** HF

17. Intepret Data In a methane molecule (CH_4), there are four single covalent bonds. In an octane molecule (C_8H_{18}), there are 25 single covalent bonds. How does the number of bonds affect the dispersion forces in samples of methane and octane? Which compound is a gas at room temperature? Which is a liquid?

Chemistry Online Self-Check Quiz glencoe.com

Section 12.3

Objectives

- **Contrast** the arrangement of particles in liquids and solids.
- **Describe** the factors that affect viscosity.
- **Explain** how the unit cell and crystal lattice are related.

Review Vocabulary

meniscus: the curved surface of a column of liquid

New Vocabulary

viscosity
surface tension
surfactant
crystalline solid
unit cell
allotrope
amorphous solid

Liquids and Solids

MAIN Idea The particles in solids and liquids have a limited range of motion and are not easily compressed.

Real-World Reading Link Did you ever wonder why syrup that is stored in the refrigerator is harder to pour than syrup stored in the pantry? You probably know that warming syrup makes it pour more easily. But why does an increase in temperature help?

Liquids

Although the kinetic-molecular theory was developed to explain the behavior of gases, the model also applies to liquids and solids. When applying the kinetic-molecular theory to the solid and liquid states of matter, you must consider the forces of attraction between particles as well as their energy of motion.

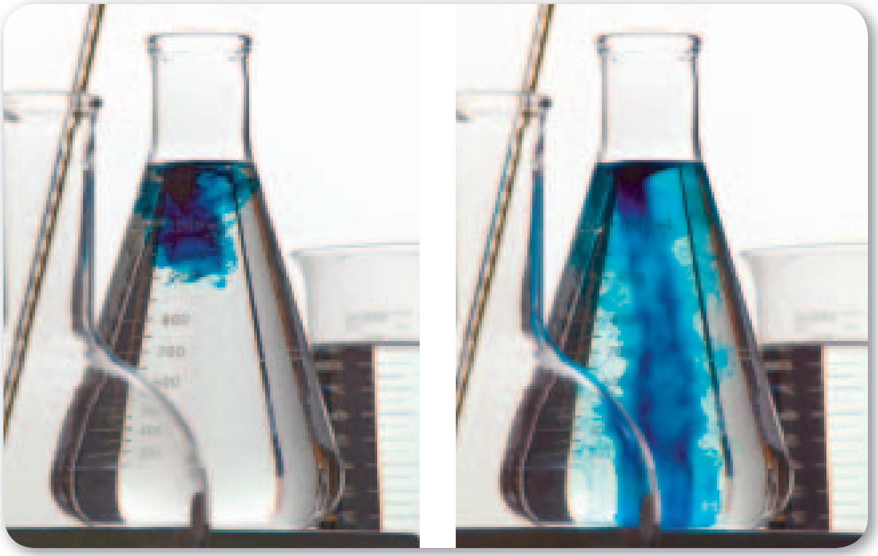
In Chapter 3, you read that a liquid can take the shape of its container but its volume is fixed. In other words, the particles can flow to adjust to the shape of a container, but the liquid cannot expand to fill its container, as shown in **Figure 12.12.** According to the kinetic-molecular theory, individual particles do not have fixed positions in the liquid. Forces of attraction between particles in the liquid limit their range of motion so that the particles remain closely packed in a fixed volume.

Density and compression At 25°C and 1 atm of air pressure, liquids are much denser than gases. The density of a liquid is much greater than that of its vapor at the same conditions. For example, liquid water is about 1250 times denser than water vapor at 25°C and 1 atm of pressure. Because they are at the same temperature, both gas and liquid particles have the same average kinetic energy. Thus, the higher density of liquids is due to the intermolecular forces that hold particles together.

Unlike gases, liquids are considered incompressible in many applications. The change in volume for liquids is much smaller because liquid particles are already tightly packed. An enormous amount of pressure must be applied to reduce the volume of a liquid by a very small amount.

Figure 12.12 Liquids flow and take the shape of their container, but they do not expand to fill their container like gases.
Infer *the reason that the liquid is at the same level in each of the interconnected tubes.*

■ **Figure 12.13** Gases and liquids have the ability to flow and diffuse. These photos show one liquid diffusing through another liquid.



Fluidity Gases and liquids are classified as fluids because they can flow and diffuse. **Figure 12.13** shows one liquid diffusing through another liquid. Liquids usually diffuse more slowly than gases at the same temperature, because intermolecular attractions interfere with the flow. Thus, liquids are less fluid than gases. A comparison between water and natural gas can illustrate this difference. When there is a leak in a basement water pipe, the water remains in the basement unless the amount of water released exceeds the volume of the basement.

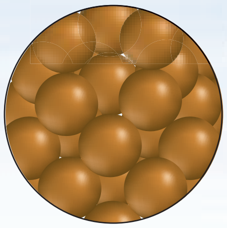
A gas will not stay in the basement. For example, natural gas, or methane, is a fuel burned in gas furnaces, hot-water heaters, and stoves. Gas that leaks from a gas pipe diffuses throughout the house. Because natural gas is odorless, companies that supply the fuel include a compound with a distinct odor. Adding odor to natural gas warns the homeowner of the leak. The customer has time to shut off the gas supply, open windows to allow the gas to diffuse, and call the gas company to report the leak.

■ **Figure 12.14**

Studying States of Matter

Scientific discoveries have led to a greater understanding of the states of matter.

460 B.C. Philosophers theorize that all matter is composed of four elements: fire, air, water, and earth.



400 B.C. The philosopher Democritus develops the theory that all matter is made of tiny, indivisible pieces called *atomos.*

360 B.C. Aristotle, a philosopher with great influence, rejects Democritus's theory and supports the belief that all matter is composed of fire, air, water, and earth.

1643 A.D. The invention of the barometer proves that air has weight.

1734 Daniel Bernoulli proposes that gas pressure results from gas particles colliding with the walls of the container.

Viscosity You are already familiar with viscosity if you have ever tried to get honey out of a bottle. **Viscosity** is a measure of the resistance of a liquid to flow. The particles in a liquid are close enough for attractive forces to slow their movement as they flow past one another. The viscosity of a liquid is determined by the type of intermolecular forces in the liquid, the size and shape of the particles, and the temperature.

You should note that not all liquids have viscosity. Scientists discovered superfluids in 1937. Scientists cooled liquid helium below −270.998°C and discovered that the properties of the liquid changed. The superfluid helium lost viscosity—the resistance to flow. The discovery of superfluidity and other milestones in our understanding of states of matter are shown in **Figure 12.14.**

Attractive forces In typical liquids, the stronger the intermolecular attractive forces, the higher the viscosity. If you have used glycerol in the laboratory to help insert a glass tube into a rubber stopper, you know that glycerol is a viscous liquid. **Figure 12.15** uses structural formulas to show the hydrogen bonding that makes glycerol so viscous. The hydrogen atoms attached to the oxygen atoms in each glycerol molecule are able to form hydrogen bonds with other glycerol molecules. The red dots in **Figure 12.15** show where the hydrogen bonds form between molecules.

Particle size and shape The size and shape of particles also affect viscosity. Recall that the overall kinetic energy of a particle is determined by its mass and velocity. Suppose the attractive forces between molecules in Liquid A and Liquid B are similar. If the molecules in Liquid A are more massive than the molecules in Liquid B, Liquid A will have a greater viscosity. Liquid A's molecules will, on average, move more slowly than the molecules in Liquid B. Molecules with long chains, such as cooking oils and motor oil, have a higher viscosity than shorter, more-compact molecules, assuming the molecules exert the same type of attractive forces. Within the long chains, there is less distance between atoms on neighboring molecules and, thus, a greater chance for attractions between atoms.

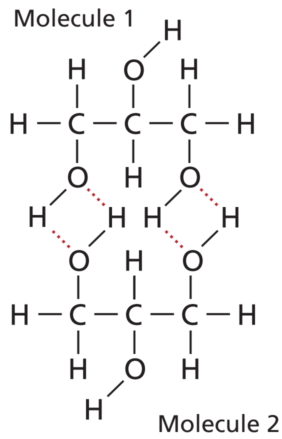
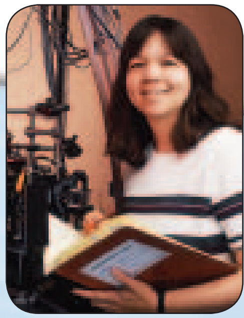


Figure 12.15 This diagram shows two glycerol molecules and the hydrogen bonds between them.

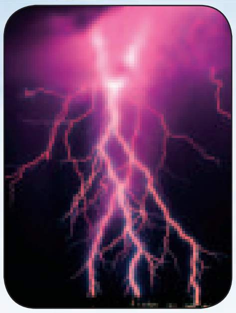
Determine *the possible number of hydrogen bonds a glycerol molecule can form with a second molecule.*

1808 John Dalton proposes that all matter is composed of tiny particles.

1937 Scientists discover superfluids—unusual fluids with properties not observed in ordinary matter.



2003 Deborah S. Jin creates the first fermionic condensate—a superfluid considered to be a sixth state of matter.



1927 The term *plasma* is first used to describe a fourth state of matter, which is found in lightning.

1995 A fifth state of matter, a gaseous superfluid called a Bose-Einstein condensate, is created and named after Satyendra Nath Bose and Albert Einstein.

Concepts In Motion

Interactive Time Line To learn more about these discoveries and others, visit glencoe.com.

Chemistry Online

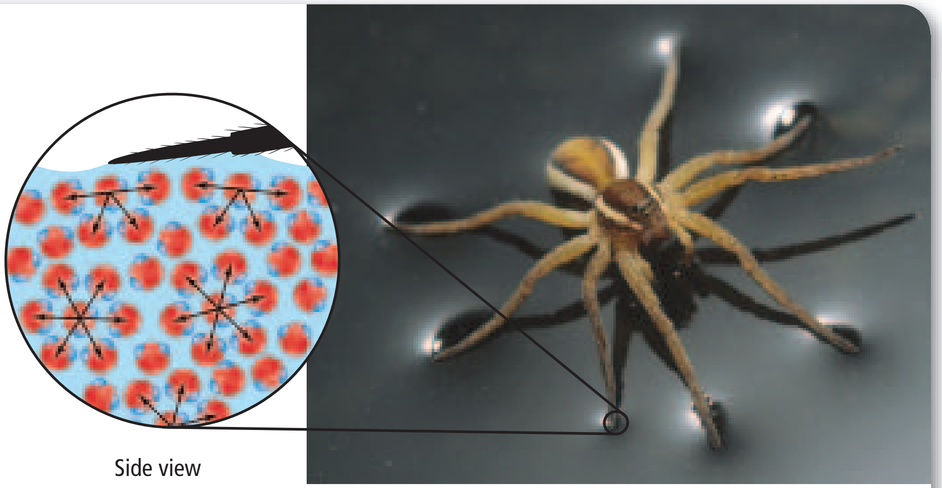
Temperature Viscosity decreases with temperature. When you pour a small amount of cooking oil into a frying pan, the oil tends not to spread across the bottom of the pan until you heat it. With the increase in temperature, there is an increase in the average kinetic energy of the oil molecules. The added energy makes it easier for the molecules to overcome the intermolecular forces that keep the molecules from flowing.

Another example of the effects of temperature on viscosity is motor oil. Motor oil keeps the moving parts of an internal combustion engine lubricated. Because temperature changes affect the viscosity of motor oil, people once used different motor-oil blends in winter and summer. The motor oil used in winter was designed to flow at low temperatures. The motor oil used in summer was more viscous so that it could maintain sufficient viscosity on extremely hot days or during long trips. Today, additives in motor oil help adjust the viscosity so that the same oil blend can be used all year. Molecules in the additives are compact spheres with relatively low viscosity at cool temperatures. At high temperatures, the shape of the additive molecules changes to long strands. These strands get tangled with the oil molecules, which increases the viscosity of the oil.

Reading Check **Infer** why it is important for motor oil to remain viscous.

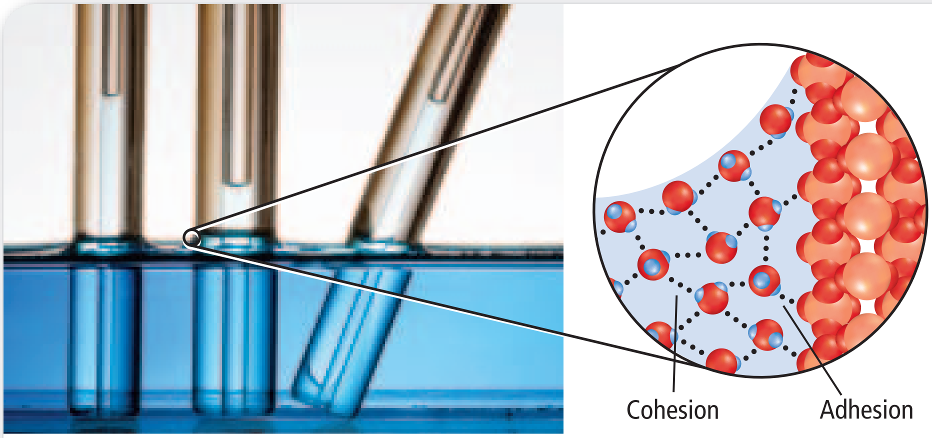
Surface tension Intermolecular forces do not have an equal effect on all particles in a liquid, as shown in **Figure 12.16.** Particles in the middle of the liquid can be attracted to particles above them, below them, and to either side. For particles at the surface of the liquid, there are no attractions from above to balance the attractions from below. Thus, there is a net attractive force pulling down on particles at the surface. The surface tends to have the smallest possible area and to act as though it is stretched tight like the head of a drum. For the surface area to increase, particles from the interior must move to the surface. It takes energy to overcome the attractions holding these particles in the interior. The energy required to increase the surface area of a liquid by a given amount is called **surface tension.** Surface tension is a measure of the inward pull by particles in the interior.

■ **Figure 12.16** At the surface of water, the particles are drawn toward the interior until attractive and repulsive forces are balanced.



Intermolecular forces just below the surface of the water create surface tension.

The surface tension of the water allows this spider to walk on the surface of the water.



The force of attraction between the water molecules and the silicon dioxide in the glass causes the water molecules to creep up the glass.

Water molecules are attracted to each other—cohesion—and to the silicon dioxide molecules in the glass—adhesion.

■ **Figure 12.17** Water molecules have cohesive and adhesive properties.
Infer ***why the water level is higher in the smaller diameter tube.***

In general, the stronger the attractions between particles, the greater the surface tension. Water has a high surface tension because its molecules can form multiple hydrogen bonds. Drops of water are shaped like spheres because the surface area of a sphere is smaller than the surface area of any other shape of similar volume. Water's high surface tension is what allows the spider in **Figure 12.16** to walk on the surface of the pond.

The same forces that allow the spider to stay dry on the surface of a pond also makes it difficult to use water alone to remove dirt from skin and clothing. Because dirt particles cannot penetrate the surface of the waterdrops, water alone cannot remove the dirt. Soaps and detergents decrease the surface tension of water by disrupting the hydrogen bonds between water molecules. When the hydrogen bonds are broken, the water spreads out allowing the dirt to be carried away by the water. Compounds that lower the surface tension of water are called surface-active agents or **surfactants.**

Cohesion and adhesion When water is placed into a narrow container, such as the glass tubes in **Figure 12.17.** you can see that the surface of the water is not straight. The surface forms a concave meniscus; that is, the surface dips in the center. **Figure 12.17** models what is happening to the water at the molecular level. There are two types of forces at work: cohesion and adhesion. Cohesion describes the force of attraction between identical molecules. Adhesion describes the force of attraction between molecules that are different. Because the adhesive forces between water molecules and the silicon dioxide in glass are greater than the cohesive forces between water molecules, the water rises along the inner walls of the cylinder.

Capillary action If the cylinder is extremely narrow, a thin film of water will be drawn upward. Narrow tubes are called capillary tubes. This movement of a liquid such as water is called capillary action, or capillarity. Capillary action helps explain how paper towels can absorb large amounts of water. The water is drawn into the narrow spaces between the cellulose fibers in paper towels by capillary action. In addition, the water molecules form hydrogen bonds with cellulose molecules.

VOCABULARY
SCIENCE USAGE V. COMMON USAGE

Force

Science usage: a push or a pull, having both magnitude and direction, that is exerted on an object
The gravitational force exists between any two objects with mass and is directly proportional to their masses.

Common usage: a group of people who have the power to work toward a desired outcome
The U.S. labor force increased its productivity last year.

Solids

Did you ever wonder why solids have a definite shape and volume? According to the kinetic-molecular theory, a mole of solid particles has as much kinetic energy as a mole of liquid or gas particles at the same temperature. By definition, the particles in a solid must be in constant motion. For a substance to be a solid rather than a liquid at a given temperature, there must be strong attractive forces acting between particles in the solid. These forces limit the motion of the particles to vibrations around fixed locations in the solid. Thus, there is more order in a solid than in a liquid. Because of this order, solids are not fluid. Only gases and liquids are classified as fluids.

Density of solids In general, the particles in a solid are more closely packed than those in a liquid. Thus, most solids are more dense than most liquids. When the liquid and solid states of a substance coexist, the solid almost always sinks in the liquid. Solid cubes of benzene sink in liquid benzene because solid benzene is more dense than liquid benzene. There is about a 10% difference in density between the solid and liquid states of most substances. Because the particles in a solid are closely packed, ordinary amounts of pressure will not change the volume of a solid.

You cannot predict the relative densities of ice and liquid water based on benzene. Ice cubes and icebergs float because water is less dense as a solid than it is as a liquid. **Figure 12.18** shows the reason for the exception. As water freezes, each H_2O molecule can form hydrogen bonds with up to four neighboring molecules. As a result, the water molecules in ice are less-closely packed together than in liquid water.

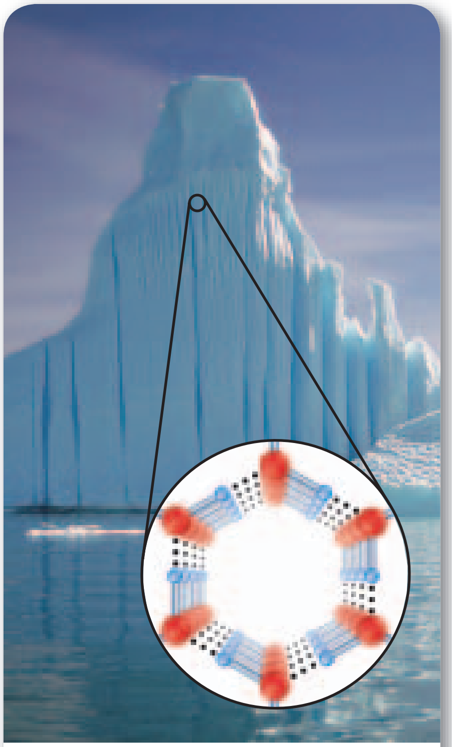
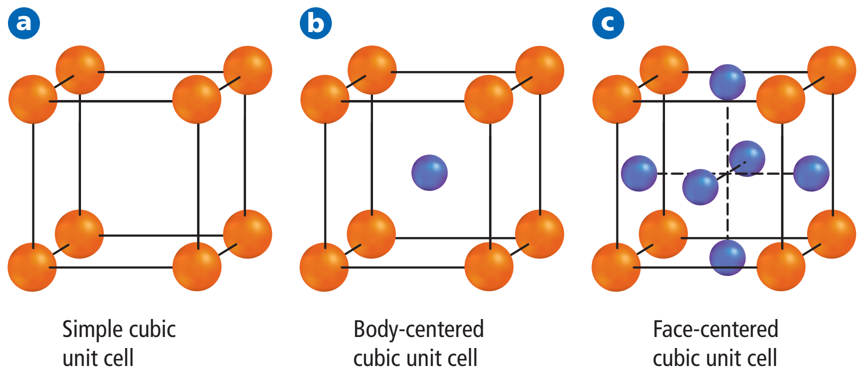


Figure 12.18 An iceberg can float because the rigid, three-dimensional structure of ice keeps water molecules farther apart than they are in liquid water. This open, symmetrical structure of ice results from hydrogen bonding.

Reading Check **Describe** in your own words why ice floats in water.

Crystalline solids Although ice is unusual in its density, ice is typical of most solids in that its molecules are packed together in a predictable way. A **crystalline solid** is a solid whose atoms, ions, or molecules are arranged in an orderly, geometric structure. The locations of particles in a crystalline solid can be represented as points on a framework called a crystal lattice. **Figure 12.19** shows three ways that particles in a crystal lattice can be arranged to form a cube.

Figure 12.19 These drawings show three of the ways particles are arranged in crystal lattices. Each sphere represents a particle. **a.** Particles are arranged only at the corners of the cube. **b.** There is a particle in the center of the cube. **c.** There are particles in the center of each of the six cubic faces but no particle in the center of the cube itself.



CAREERS IN CHEMISTRY

Metallurgist Metallurgists are engineers who are involved in all stages of processing metals, from extracting and refining to casting the final product. At each stage, metallurgists must understand the physical and chemical properties of metals. A college degree is necessary to become a metallurgist, and many go on to earn advanced degrees. For more information on chemistry careers, visit glencoe.com.

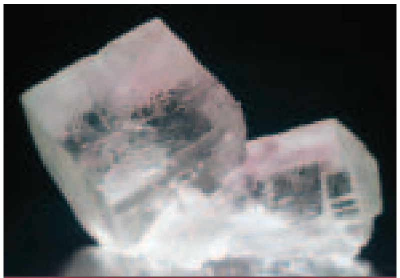
A **unit cell** is the smallest arrangement of atoms in a crystal lattice that has the same symmetry as the whole crystal. Like the formula unit that you read about in Chapter 7, a unit cell is a small, representative part of a larger whole. The unit cell can be thought of as a building block whose shape determines the shape of the crystal.

Table 12.4 shows seven categories of crystals based on shape. Crystal shapes differ because the surfaces, or faces, of unit cells do not always meet at right angles, and the edges of the faces vary in length. In **Table 12.4,** the edges are labeled *a, b,* and *c;* the angles at which the faces meet are labeled α, β, and γ.

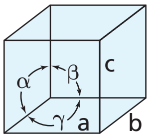
Concepts In Motion

Table 12.4 Unit Cells

Interactive Table Explore unit cells at glencoe.com.

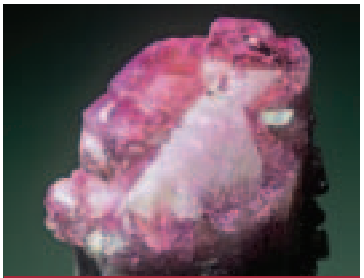


Halite (rock salt)

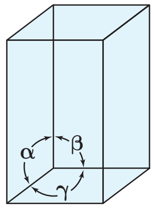


$a = b = c$
$\alpha = \beta = \gamma = 90°$

Cubic

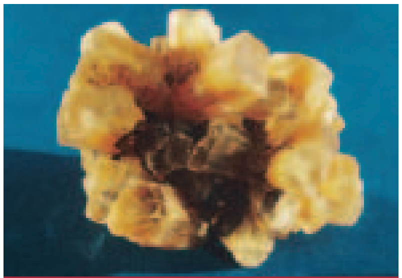


Vesuvianite

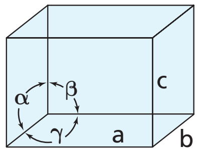


$a = b \neq c$
$\alpha = \beta = \gamma = 90°$

Tetragonal

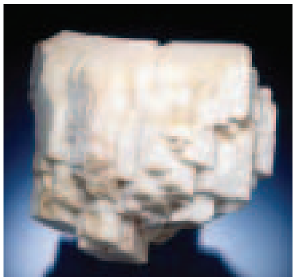


Aragonite

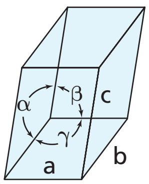


$a \neq b \neq c$
$\alpha = \beta = \gamma = 90°$

Orthorhombic



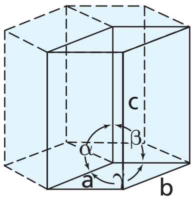
Microcline



$a \neq b \neq c$
$\alpha \neq \beta \neq \gamma \neq 90°$

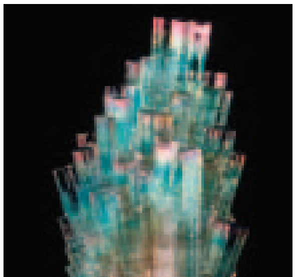
Triclinic

Beryl (emerald)

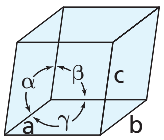


$a = b \neq c$
$\alpha = \beta = 90°, \gamma = 120°$

Hexagonal



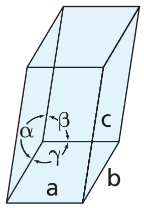
Tourmaline



$a = b \neq c$
$\alpha = \beta = \gamma \neq 90°$

Rhombohedral

Crocite



$a \neq b \neq c$
$\alpha = \gamma = 90° \neq \beta$

Monoclinic

Table 12.5 Types of Crystalline Solids

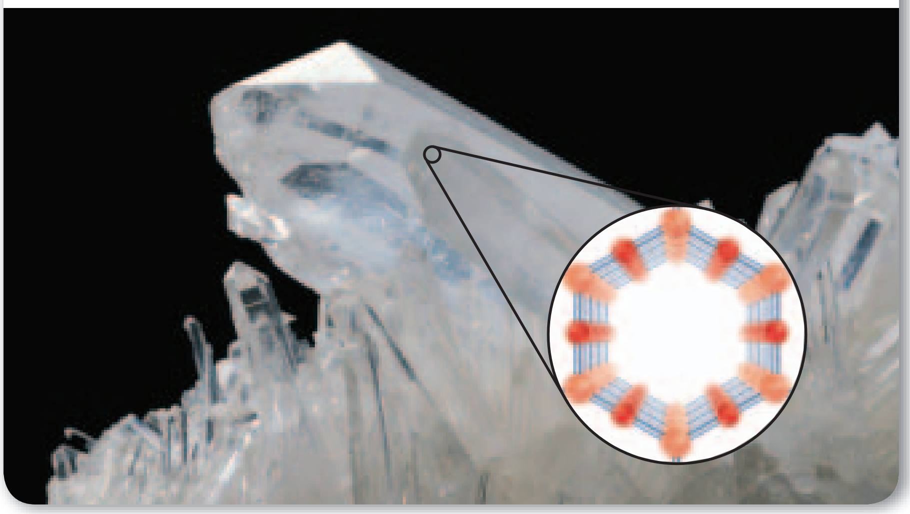
Interactive Table Explore types of crystalline solids at glencoe.com.

Type	Unit Particles	Characteristics of Solid Phase	Examples
Atomic	atoms	soft to very soft; very low melting points; poor conductivity	group 18 elements
Molecular	molecules	fairly soft; low to moderately high melting points; poor conductivity	I_2, H_2O, NH_3, CO_2, $C_{12}H_{22}O_{11}$ (table sugar)
Covalent network	atoms connected by covalent bonds	very hard; very high melting points; often poor conductivity	diamond (C) and quartz (SiO_2)
Ionic	ions	hard; brittle; high melting points; poor conductivity	NaCl, KBr, $CaCO_3$
Metallic	atoms surrounded by mobile valence electrons	soft to hard; low to very high melting points; malleable and ductile; excellent conductivity	all metallic elements

Categories of crystalline solids Crystalline solids can be classified into five categories based on the types of particles they contain and how thoses particles are bonded together: atomic solids, molecular solids, covalent network solids, ionic solids, and metallic solids. **Table 12.5** summarizes the general characteristics of each category and provides examples. The only atomic solids are noble gases. Their properties reflect the weak dispersion forces between the atoms.

Molecular solids In molecular solids, the molecules are held together by dispersion forces, dipole-dipole forces, or hydrogen bonds. Most molecular compounds are not solids at room temperature. Even water, which can form strong hydrogen bonds, is a liquid at room temperature. Molecular compounds such as sugar are solids at room temperature because of their large molar masses. With larger molecules, many weak attractions can combine to hold the molecules together. Because they contain no ions, molecular solids are poor conductors of heat and electricity.

Figure 12.20 The most common kind of quartz has a hexagonal crystal structure.



MiniLab

Model Crystal Unit Cells

How can you make physical models that illustrate the structures of crystals?

Procedure

1. Read and complete the lab safety form.
2. Cut **four soda straws** into thirds. Wire the straw pieces together to make a cube using **22- or 26-gauge wire.** Use **scissors** to cut the wire. Refer to **Table 12.4** for a guide to crystal shapes.
3. To model a rhombohedral crystal, deform the cube from Step 2 until no angles are 90°.
4. To model a hexagonal crystal, flatten the model from Step 3 until it looks like a pie with six slices.
5. To model a tetragonal crystal, cut **4 straws** in half. Cut 4 of the pieces in half again. Wire the 8 shorter pieces to make 4 square ends. Use the longer pieces to connect the square ends.
6. To model the orthorhombic crystal, cut **4 straws** in half. Cut one-third off 4 of the halves, creating 4 each of three different lengths. Connect the 4 long, 4 medium, and 4 short pieces so that each side is a rectangle.
7. To model the monoclinic crystal, deform the model from Step 6 along one axis. To model the triclinic crystal, deform the model from Step 6 until it has no 90° angles.

Analysis

1. **Evaluate** Which two models have three axes of equal length? How do these models differ?
2. **Determine** which model includes a square and a rectangle.
3. **Determine** which models have three unequal axes.
4. **Infer** Do you think crystals are perfect, or do they have defects? Explain your answer.

Covalent network solids Atoms such as carbon and silicon, which can form multiple covalent bonds, are able to form covalent network solids. The covalent network structure of quartz, which contains silicon, is shown in **Figure 12.20.** Carbon forms three types of covalent network solids—diamond, graphite, and buckminsterfullerene. An element, such as carbon, that exists in different forms at the same state—solid, liquid, or gas—is called an **allotrope.** For more information about carbon allotropes see the Elements Handbook.

Ionic solids Remember that each ion in an ionic solid is surrounded by ions of opposite charge. The type of ions and the ratio of ions determine the structure of the lattice and the shape of the crystal. The network of attractions that extends throughout an ionic crystal gives these compounds their high melting points and hardness. Ionic crystals are strong, but brittle. When ionic crystals are struck, the cations and anions are shifted from their fixed positions. Repulsions between ions of like charge cause the crystal to shatter.

Metallic solids Recall from Chapter 7 that metallic solids consist of positive metal ions surrounded by a sea of mobile electrons. The strength of the metallic bonds between cations and electrons varies among metals and accounts for their wide range of physical properties. For example, tin melts at 232°C, but nickel melts at 1455°C. The mobile electrons make metals malleable—easily hammered into shapes—and ductile—easily drawn into wires. When force is applied to a metal, the electrons shift and thereby keep the metal ions bonded in their new positions. Mobile electrons make metals good conductors of heat and electricity. Businesses, equipment, and homes, such as the one shown in **Figure 12.21,** use metal wiring to carry electricity.

Reading Check **Describe** the properties of metals that make them useful for making jewelry.

Figure 12.21 Homes, business, and equipment of all types use metal wiring to carry electricity. The metal is usually copper, but other metals are used in special applications.

Figure 12.22 Native Americans used the glass-like amorphous rock obsidian to make arrowheads and knives, because it can form sharp edges when broken. Obsidian rock forms when lava cools too quickly to form crystals.

FOLDABLES Incorporate information from this section into your Foldable.

Amorphous solids An **amorphous solid** is one in which the particles are not arranged in a regular, repeating pattern. It does not contain crystals. The term *amorphous* is derived from a Greek word that means *without shape*. An amorphous solid often forms when a molten material cools too quickly to allow enough time for crystals to form. **Figure 12.22** shows an example of an amorphous solid.

Glass, rubber, and many plastics are amorphous solids. Recent studies have shown that glass might have some structure. When X-ray diffraction is used to study glass, there appears to be no pattern to the distribution of atoms. When neutrons are used instead, an orderly pattern of silicate units can be detected in some regions. Researchers hope to use this new information to control the structure of glass for optical applications and to produce glass that can conduct electricity.

Section 12.3 Assessment

Section Summary

- The kinetic-molecular theory explains the behavior of solids and liquids.
- Intermolecular forces in liquids affect viscosity, surface tension, cohesion, and adhesion.
- Crystalline solids can be classified by their shape and composition.

18. MAIN Idea **Contrast** the arrangement of particles in solids and liquids.
19. **Describe** the factors that affect viscosity.
20. **Explain** why soap and water are used to clean clothing instead of water alone.
21. **Compare** a unit cell and a crystal lattice.
22. **Describe** the difference between a molecular solid and a covalent network solid.
23. **Explain** why water forms a meniscus when it is in a graduated cylinder.
24. **Infer** why the surface of mercury in a thermometer is convex; that is, the surface is higher at the center.
25. **Predict** which solid is more likely to be amorphous—one formed by allowing a molten material to cool slowly to room temperature or one formed by quickly cooling the same material in an ice bath.
26. **Design** an experiment to compare the relative abilities of water and isopropyl alcohol to support skipping stones. Include a prediction about which liquid will be better, along with a brief explanation of your prediction.

Chemistry Online Self-Check Quiz glencoe.com

Section 12.4

Objectives

- **Explain** how the addition and removal of energy can cause a phase change.
- **Interpret** a phase diagram.

Review Vocabulary

phase change: a change from one state of matter to another

New Vocabulary

melting point
vaporization
evaporation
vapor pressure
boiling point
freezing point
condensation
deposition
phase diagram
triple point

Phase Changes

MAIN Idea Matter changes phase when energy is added or removed.

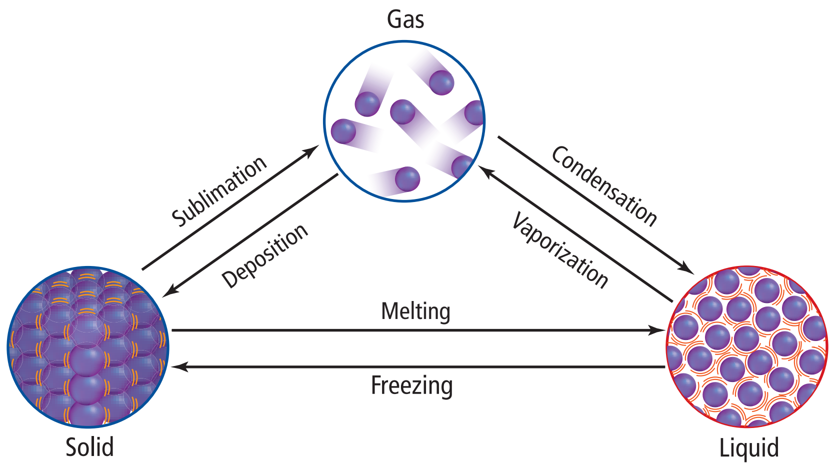
Real-World Reading Link Have you ever wondered where the matter in a solid air freshener goes? The day it is opened and put in a room, it is a solid, fragrant mass. Day-by-day, the solid gets smaller and smaller. Finally, almost nothing is left and it is time to put a new one out. You never observe a puddle of liquid like you would see if it had melted.

Phase Changes That Require Energy

Most substances can exist in three states depending on the temperature and pressure. A few substances, such as water, exist in all three states under ordinary conditions. States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture. Ice water is a heterogeneous mixture with two phases, solid ice and liquid water. When energy is added or removed from a system, one phase can change into another, as shown in **Figure 12.23.** Because you are familiar with the phases of water—ice, liquid water, and water vapor—and have observed changes between those phases, we can use water as the primary example in the discussion of phase changes.

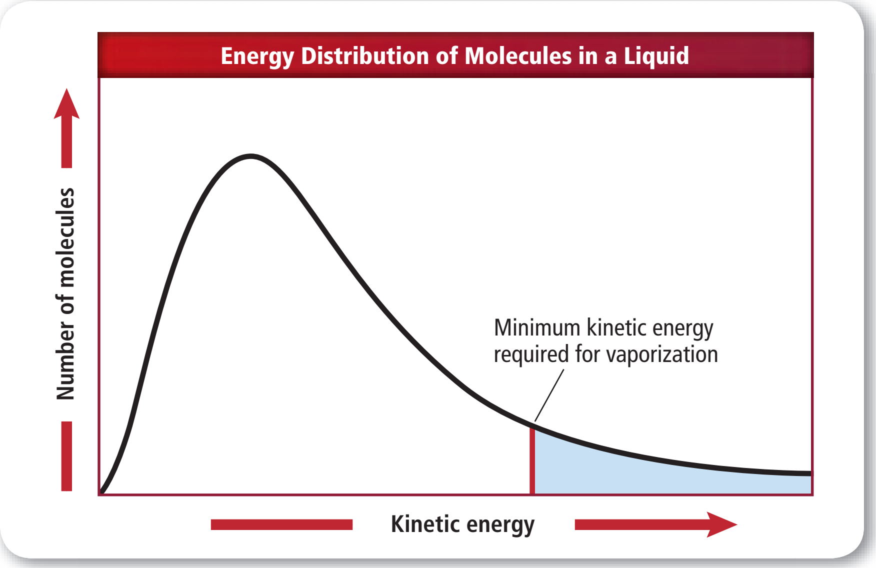
Melting What does happen to ice cubes in a glass of ice water? When ice cubes are placed in water, the water is at a higher temperature than the ice. Heat flows from the water to the ice. Heat is the transfer of energy from an object at a higher temperature to an object at a lower temperature. At ice's melting point, the energy absorbed by the ice is not used to raise the temperature of the ice. Instead, it disrupts the hydrogen bonds holding the water molecules together in the ice crystal. When molecules on the surface of the ice absorb enough energy to break the hydrogen bonds, they move apart and enter the liquid phase. As molecules are removed, the ice cube shrinks. The process continues until all of the ice melts. If a tray of ice cubes is left on a counter, where does the energy to melt the cubes come from?

Figure 12.23 The diagram shows the six possible transitions between phases.
Determine *what phase changes occur between solids and liquids.*



■ **Figure 12.24** This graph shows a typical distribution of kinetic energy of molecules in a liquid at 25°C. The most probable amount of kinetic energy for a molecule lies at the peak of the curve.

Describe *how the curve would look for the same liquid at 30°C.*



Real-World Chemistry
Evaporation

Perspiration Evaporation is one way your body controls its temperature. When you become hot, your body releases sweat from glands in your skin. Water molecules in sweat can absorb heat energy from your skin and evaporate. Excess heat is carried from all parts of your body to your skin by your blood.

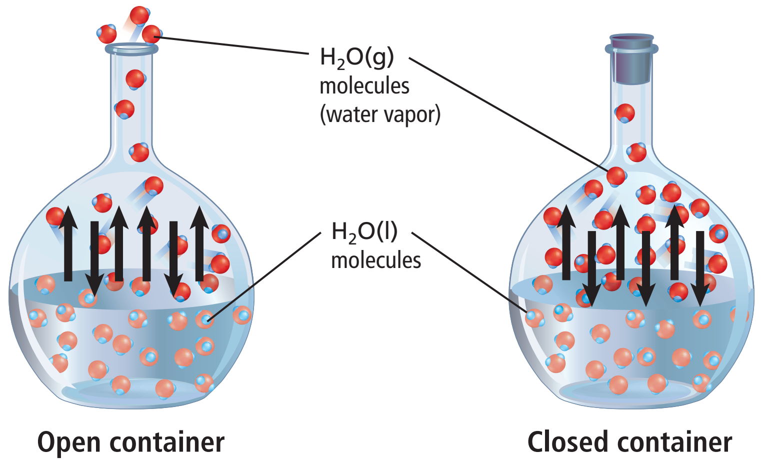
The amount of energy required to melt 1 mol of a solid depends on the strength of the forces keeping the particles together in the solid. Because hydrogen bonds between water molecules are strong, a relatively large amount of energy is required. However, the energy required to melt ice is much less than the energy required to melt table salt because the ionic bonds in sodium chloride are much stronger than the hydrogen bonds in ice.

The temperature at which the liquid phase and the solid phase of a given substance can coexist is a characteristic physical property of many solids. The **melting point** of a crystalline solid is the temperature at which the forces holding its crystal lattice together are broken and it becomes a liquid. It is difficult to specify an exact melting point for an amorphous solid because they tend to melt over a temperature range.

Vaporization While ice melts, the temperature of the ice-water mixture remains constant. Once all of the ice has melted, additional energy added to the system increases the kinetic energy of the liquid molecules. The temperature of the system begins to rise. In liquid water, some molecules will have more kinetic energy than other molecules. **Figure 12.24** shows how energy is distributed among the molecules in a liquid at 25°C. The shaded portion indicates those molecules that have the energy required to overcome the forces of attraction holding the molecules together in the liquid.

Graph Check Describe what happens to the particles in the shaded portion on the graph.

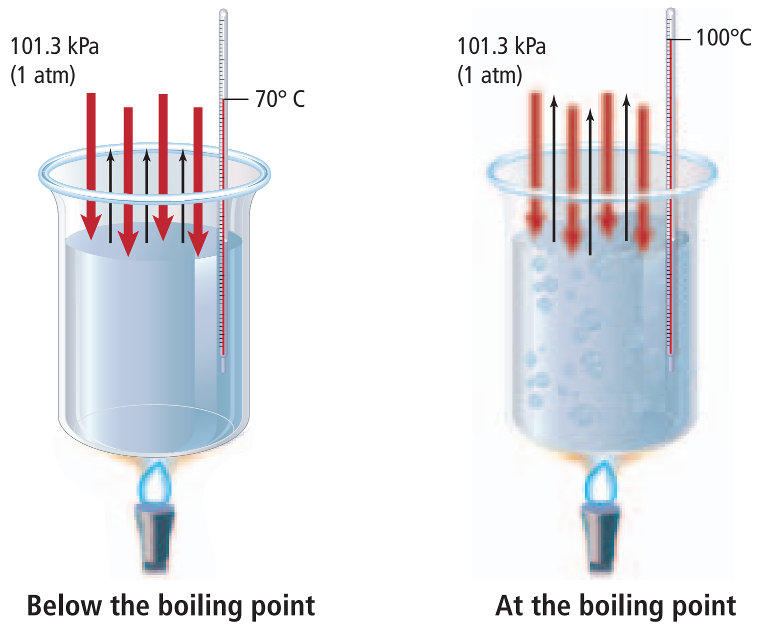
Particles that escape from the liquid enter the gas phase. For a substance that is ordinarily a liquid at room temperature, the gas phase is called a vapor. **Vaporization** is the process by which a liquid changes to a gas or vapor. If the input of energy is gradual, the molecules tend to escape from the surface of the liquid. Remember that molecules at the surface are attracted to fewer other molecules than are molecules in the interior. When vaporization occurs only at the surface of a liquid, the process is called **evaporation.** Even at cold temperatures, some water molecules have enough energy to evaporate. As the temperature rises, more and more molecules enter the gas phase.



■ **Figure 12.25** Evaporation occurs in both open and closed containers. In an open container, water molecules that evaporate can escape from the container. Water vapor collects above the liquid in a closed container.

Figure 12.25 compares evaporation in an open container with evaporation in a closed container. If water is in an open container, all the molecules will eventually evaporate. The time it takes for them to evaporate depends on the amount of water and the available energy. In a partially filled, closed container, the situation is different. Water vapor collects above the liquid and exerts pressure on the surface of the liquid. The pressure exerted by a vapor over a liquid is called **vapor pressure.**

Boiling The temperature at which the vapor pressure of a liquid equals the external or atmospheric pressure is called the **boiling point.** Use **Figure 12.26** to compare what happens to a liquid at temperatures below its boiling point with what happens to a liquid at its boiling point. At the boiling point, molecules throughout the liquid have enough energy to vaporize. Bubbles of vapor collect below the surface of the liquid and rise to the surface.



■ **Figure 12.26** As temperature increases, water molecules gain kinetic energy. Vapor pressure increases (black arrows) but is less than atmospheric pressure (red arrows). A liquid has reached its boiling point when its vapor pressure is equal to atmospheric pressure. At sea level, the boiling point of water is 100°C.

■ **Figure 12.27** These steaks are kept cold by dry ice.
Explain *why dry ice is preferred over regular ice for shipping steaks and other food products.*

■ **Figure 12.28** Normally, air becomes cooler as elevation increases. A temperature inversion occurs when the situation is reversed and the air becomes warmer at higher elevations. Inversions can trap smog over cities and fog in mountain valleys.

Sublimation Many substances have the ability to change directly from the solid phase to the gas phase. Recall from Chapter 3 that sublimation is the process by which a solid changes directly to a gas without first becoming a liquid. Solid iodine and solid carbon dioxide (dry ice) sublime at room temperature. Dry ice, shown in **Figure 12.27,** keeps objects that could be damaged by melting water cold during shipping. Mothballs, which contain the compounds naphthalene or *p*-dichlorobenzene, also sublime, as do solid air fresheners.

Phase Changes That Release Energy

Have you ever awakened on a chilly morning to see frost on your windows or the grass covered with water droplets? When you set a glass of ice water on a picnic table, do you notice beads of water on the outside of the glass? These events are examples of phase changes that release energy into the surroundings.

Freezing Suppose you place liquid water in an ice tray into a freezer. As heat is removed from the water, the molecules lose kinetic energy and their velocity decreases. The molecules are less likely to flow past one another. When enough energy has been removed, the hydrogen bonds between water molecules keep the molecules fixed, or frozen, into set positions. Freezing is the reverse of melting. The **freezing point** is the temperature at which a liquid is converted into a crystalline solid.

Condensation When a water vapor molecule loses energy, its velocity decreases. The water vapor molecule is more likely to form a hydrogen bond with another water molecule. The formation of a hydrogen bond releases thermal energy and indicates a change from the vapor phase to the liquid phase. The process by which a gas or a vapor becomes a liquid is called **condensation.** Condensation is the reverse of vaporization.

Different factors contribute to condensation. However, condensation always involves the transfer of thermal energy. For example, water vapor molecules can come in contact with a cold surface, such as the side of a glass of ice water. Thermal energy transfers from the water vapor molecules to the cool glass, causing condensation on the outside of the glass. A similar process can occur during the night when water vapor in the air condenses and dew forms on blades of grass.

Connection to Earth Science Precipitation, clouds, and fog all result from condensation. They form as air cools when it rises or passes over cooler land or water. Their formations require a second factor, microscopic particles suspended in the air called condensation nuclei. These can be particles, such as soot and dust, or aerosols, such as sulfur dioxide and nitrogen oxide, on which water vapor condenses. In some circumstances, warm air can settle on top of cooler air, which is called a temperature inversion. **Figure 12.28** shows fog trapped in a mountain valley by such an inversion.

Reading Check **Describe** the condensation of water vapor in the atmosphere.

Deposition When water vapor comes in contact with a cold window in winter, it forms a solid deposit on the window called frost. **Deposition** is the process by which a substance changes from a gas or vapor to a solid without first becoming a liquid. Deposition is the reverse of sublimation. Snowflakes form when water vapor high up in the atmosphere changes directly into solid ice crystals. Energy is released as the crystals form.

Phase Diagrams

There are two variables that combine to control the phase of a substance: temperature and pressure. These variables can have opposite effects on a substance. For example, a temperature increase causes more liquid to vaporize, but an increase in pressure causes more vapor to condense. A **phase diagram** is a graph of pressure versus temperature that shows in which phase a substance exists under different conditions of temperature and pressure.

Figure 12.29 shows the phase diagram for water. You can use this graph to predict what phase water will be in for any combination of temperature and pressure. Note that there are three regions representing the solid, liquid, and vapor phases of water and three curves that separate the regions from one another. At points that fall along the curves, two phases of water can coexist. The short, yellow curve shows the temperature and pressure conditions under which solid water and water vapor can coexist. The long, blue curve shows the temperature and pressure conditions under which liquid water and water vapor can coexist. The red curve shows the temperature and pressure conditions under which solid water and liquid water can coexist.

Point A on the phase diagram of water—the point where the yellow, blue, and red curves meet—is the triple point for water. The **triple point** is the point on a phase diagram that represents the temperature and pressure at which three phases of a substance can coexist. All six phase changes can occur at the triple point: freezing and melting; evaporation and condensation; sublimation and deposition. Point B is called the critical point. This point indicates the critical pressure and critical temperature above which water cannot exist as a liquid. If water vapor is at the critical temperature, an increase in pressure will not change the vapor into a liquid.

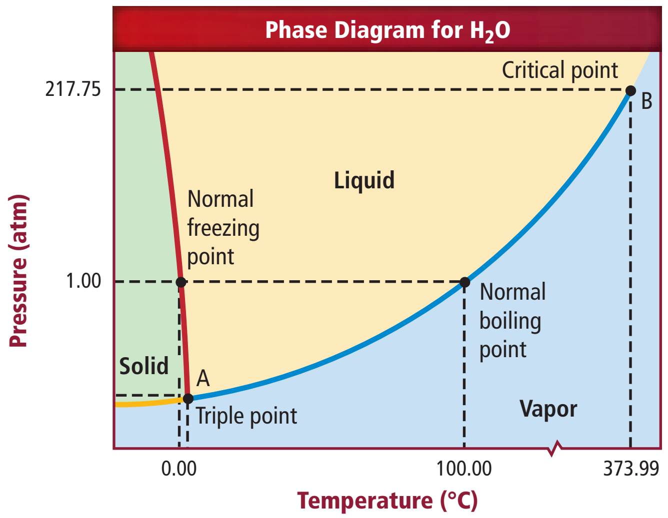


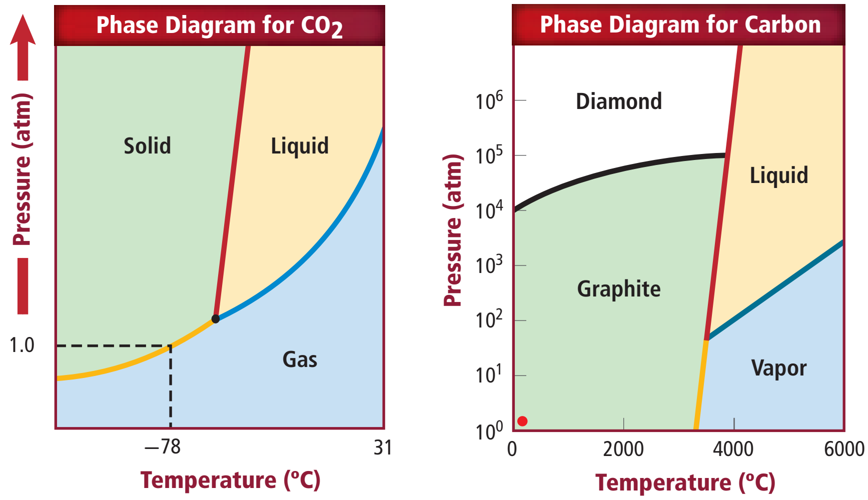
Figure 12.29 This phase diagram shows the phase of water at different temperatures and pressures.

Graph Check

Determine the phase of water at 2.00 atm and 100.00°C.

Concepts In Motion

Interactive Figure To see an animation of a phase diagram, visit glencoe.com.



■ **Figure 12.30** Phase diagrams show useful information, such as why carbon dioxide sublimes at normal conditions and the existence of two forms of solid carbon.

Graph Check

Contrast the slope of the red line in water's phase diagram with that of the red line in carbon dioxide's phase diagram. How do water and carbon dioxide differ in their reaction to increased pressure at the solid/liquid boundary?

The phase diagram for each substance is different because the normal boiling and freezing points of substances are different. However, each diagram will supply the same type of data for the phases, including a triple point. Of course, the range of temperatures chosen will vary to reflect the physical properties of the substance.

Phase diagrams can provide important information for substances. For example, the phase diagram for carbon dioxide in **Figure 12.30** shows why carbon dioxide sublimes at normal conditions. Find 1.0 atm on the carbon dioxide graph and follow the dashed line to the yellow line. The graph shows that carbon dioxide changes from a solid to a gas at 1 atm. If you extend the dashed line past the yellow line, the graph shows that carbon dioxide does not liquefy as temperature increases. It remains a gas.

The diagram on the right is a phase diagram for carbon. Notice that the graph contains two allotropes of carbon in the solid region. Graphite is the standard state of carbon at normal temperatures and pressures, designated by a red dot. Diamond is more stable at higher temperatures and pressures. Diamonds that exist at normal room conditions originally formed at high temperature and pressure.

Section 12.4 Assessment

Section Summary

- States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture.
- Energy changes occur during phase changes.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

27. MAIN Idea **Explain** how the addition or removal of energy can cause a phase change.
28. **Explain** the difference between the processes of melting and freezing.
29. **Compare** deposition and sublimation.
30. **Compare and contrast** sublimation and evaporation.
31. **Describe** the information that a phase diagram supplies.
32. **Explain** what the triple point and the critical point on a phase diagram represent.
33. **Determine** the phase of water at 75.00°C and 3.00 atm using **Figure 12.29.**

Chemistry Online Self-Check Quiz glencoe.com

Everyday Chemistry

Cocoa Chemistry

Chocolate is a food product that is native to Central America and Mexico. The Aztec ruler Montezuma served the bitter cocoa-bean drink to Hernan Cortéz in 1519. Cortéz took the cocoa beans and the recipe for the chocolate beverage to Spain where it became a very popular, but expensive beverage. Chocolate remained a food product for the wealthy until the mid-nineteenth century, when the price of chocolate became affordable and processing techniques improved. The chocolate served today bears little resemblance to the chocolate served in Montezuma's court. Processing techniques as well as additives create the smooth, sweet, delightful treat that you enjoy today.

Melts in your mouth Chocolate is a mixture of cocoa, cocoa butter, and other ingredients. This mixture is a solid at room temperature, but melts in your mouth. Why? Because one of the main ingredients in chocolate—cocoa butter—is a fat that melts at near body temperature.

Particle size Chocolate is a liquid during the mixing process. The cocoa butter in the melted chocolate coats the solid particles of cocoa, sugar, and milk solids. The solid particles in the mixture must not be too large, or the chocolate will have a gritty texture. Generally, the particles are ground to a maximum diameter of 2.0×10^{-5} to 3.0×10^{-5} m.

Controlling flow As you can see in **Figure 1,** a large number of small particles has a larger surface area than a single particle of the same mass.

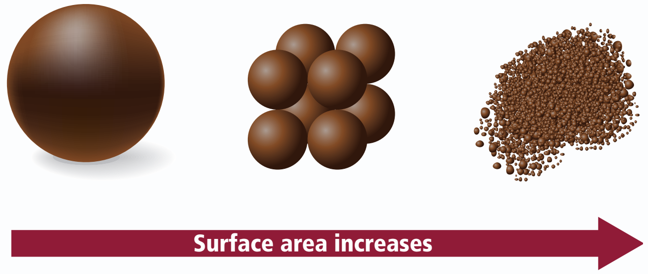


Figure 1 Although the mass of each particle or group of particles is the same, increasing the surface area allows more cocoa butter to coat the particles, which improves the flow of the chocolate.

Figure 2 Chocolate is carefully processed so that the proper crystal structure forms in the chocolate. These crystals give chocolate the characteristics found in popular chocolate bars.

Smaller particles in the chocolate requires more cocoa butter to coat the solid surfaces. It is the excess cocoa butter *between* the solid particles that allows chocolate to flow.

Smooth texture If the chocolate contains too little cocoa butter between the particles, the chocolate will be too thick to flow into a mold. To improve the flow of the chocolate without increasing particle size, manufacturers can either add more fat to the mixture or add an emulsifier, such as lecithin. Lecithin is a fat often obtained from soybeans that helps keep the fat molecules evenly suspended, or emulsified, in the chocolate.

Crystallization Another important process in chocolate manufacturing is tempering. During the tempering of the chocolate, the temperature of the chocolate is carefully controlled to ensure that the desired crystals form. When chocolate is not properly tempered, crystals form that create poor-quality chocolate. The desired crystals make the chocolate in **Figure 2** glossy and firm, and allow it to snap well and melt near body temperature.

WRITING in Chemistry

Research to find out more about chocolate and write a short report. For more information about chocolate, visit glencoe.com.

CHEMLAB

SMALL SCALE

INTERNET: COMPARE RATES OF EVAPORATION

Background: Several factors determine how fast a sample of liquid will evaporate. The volume of the sample is a key factor. A drop of water takes less time to evaporate than a liter of water. The amount of energy supplied to the sample is another factor.

Question: *How do intermolecular forces affect the evaporation rates of liquids?*

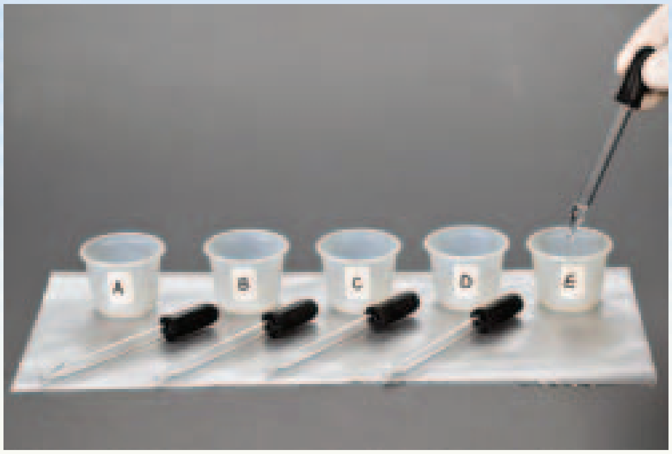
Materials

distilled water
ethanol
isopropyl alcohol
acetone
household ammonia
droppers (5)
small plastic cups (5)
grease pencil or masking tape and a marking pen
paper towel
square of waxed paper
stopwatch

Safety Precautions

Procedure

1. Read and complete the lab safety form.
2. Make a data table to record data.
3. Use a grease pencil or masking tape to label each of 5 small plastic cups. Use *A* for distilled water, *B* for ethanol, *C* for isopropyl alcohol, *D* for acetone, and *E* for household ammonia. Place the plastic cups on a paper towel.
4. Use a dropper to collect about 1 mL of distilled water and place the water in the cup labeled *A*. Place the dropper on the paper towel directly in front of the cup. Repeat with the other liquids.
5. Place a square of waxed paper on your lab surface. Plan where on the waxed paper you will place each of the five drops that you will test to avoid mixing.
6. Have your stopwatch ready. Collect some water in your water dropper and place a single drop on the waxed paper. Begin timing. Time how long it takes for the drop to completely evaporate. While you wait, make a top-view and side-view drawing of the drop. If the drop takes longer than 5 min to evaporate, record > 300 min in your data table.
7. Repeat Step 6 with the four other liquids.
8. Use the above procedure to design an experiment in which you can observe the effect of temperature on the rate of evaporation of ethanol. Your teacher will provide a sample of warm ethanol.



9. **Cleanup and Disposal** Clean up lab materials as instructed by your teacher.

Analyze and Conclude

1. **Classify** Which liquids evaporated quickly? Which liquids were slow to evaporate?
2. **Evaluate** Based on your data, in which liquid(s) are the attractive forces between molecules most likely dispersion forces?
3. **Consider** What is the relationship between surface tension and the shape of a liquid drop? What are the attractive forces that increase surface tension?
4. **Assess** The isopropyl alcohol you used was a mixture of isopropyl alcohol and water. Would pure isopropyl alcohol evaporate more quickly or more slowly compared to the alcohol and water mixture? Give a reason for your answer.
5. **Evaluate** Household ammonia is a mixture of ammonia and water. Based on the data you collected, is there more ammonia or more water in the mixture? Explain.
6. **Evaluate** How does the rate of evaporation of warm ethanol compare to ethanol at room temperature?
7. **Share** your data at glencoe.com.
8. **Error Analysis** How could you change the procedure to make it more precise?

INQUIRY EXTENSION

Design an Experiment How would different surfaces affect your results? Design an experiment to test your hypothesis.

Matt Meadows

Study Guide

BIG Idea Kinetic-molecular theory explains the different properties of solids, liquids, and gases.

Section 12.1 Gases

MAIN Idea Gases expand, diffuse, exert pressure, and can be compressed because they are in a low-density state consisting of tiny, constantly-moving particles.

Vocabulary

- atmosphere (p. 407)
- barometer (p. 407
- Dalton's law of partial pressures (p. 408)
- diffusion (p. 404)
- elastic collision (p. 403)
- Graham's law of effusion (p. 404)
- kinetic-molecular theory (p. 402)
- pascal (p. 407)
- pressure (p. 406)
- temperature (p. 403)

Key Concepts

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Dalton's law of partial pressures is used to determine the pressures of individual gases in gas mixtures.
- Graham's law is used to compare the diffusion rates of two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

Section 12.2 Forces of Attraction

MAIN Idea Intermolecular forces—including dispersion forces, dipole-dipole forces, and hydrogen bonds—determine a substance's state at a given temperature.

Vocabulary

- dipole-dipole force (p. 412)
- dispersion force (p. 412)
- hydrogen bond (p. 413)

Key Concepts

- Intramolecular forces are stronger than intermolecular forces.
- Dispersion forces are intermolecular forces between temporary dipoles.
- Dipole-dipole forces occur between polar molecules.

Section 12.3 Liquids and Solids

MAIN Idea The particles in solids and liquids have a limited range of motion and are not easily compressed.

Vocabulary

- allotrope (p. 423)
- amorphous solid (p. 424)
- crystalline solid (p. 420)
- surfactant (p. 419)
- surface tension (p. 418)
- unit cell (p. 421)
- viscosity (p. 417)

Key Concepts

- The kinetic-molecular theory explains the behavior of solids and liquids.
- Intermolecular forces in liquids affect viscosity, surface tension, cohesion, and adhesion.
- Crystalline solids can be classified by their shape and composition.

Section 12.4 Phase Changes

MAIN Idea Matter changes phase when energy is added or removed.

Vocabulary

- boiling point (p. 427)
- condensation (p. 428)
- deposition (p. 429)
- evaporation (p. 426)
- freezing point (p. 428)
- melting point (p. 426)
- phase diagram (p. 429)
- triple point (p. 429)
- vaporization (p. 426)
- vapor pressure (p. 427)

Key Concepts

- States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture.
- Energy changes occur during phase changes.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

Section 12.1

Mastering Concepts

34. What is an elastic collision?

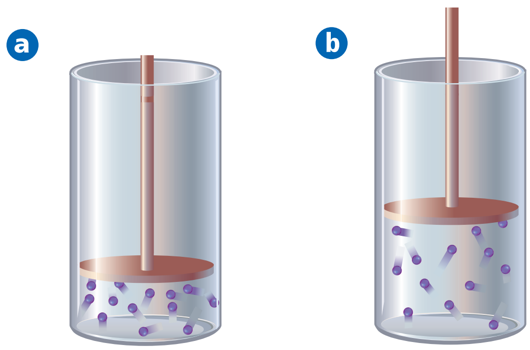
35. How does the kinetic energy of particles vary as a function of temperature?

36. Use the kinetic-molecular theory to explain the compression and expansion of gases.

37. List the three basic assumptions of the kinetic-molecular theory.

38. Describe the common properties of gases.

39. Compare diffusion and effusion. Explain the relationship between the rates of these processes and the molar mass of a gas.



■ **Figure 12.31**

40. In **Figure 12.31,** what happens to the density of gas particles in the cylinder as the piston moves from Position A to Position B?

41. Baking Explain why the baking instructions on a box of cake mix are different for high and low elevations. Would you expect to have a longer or a shorter cooking time at a high elevation?

Mastering Problems

42. What is the molar mass of a gas that takes three times longer to effuse than helium?

43. What is the ratio of effusion rates of krypton and neon at the same temperature and pressure?

44. Calculate the molar mass of a gas that diffuses three times faster than oxygen under similar conditions.

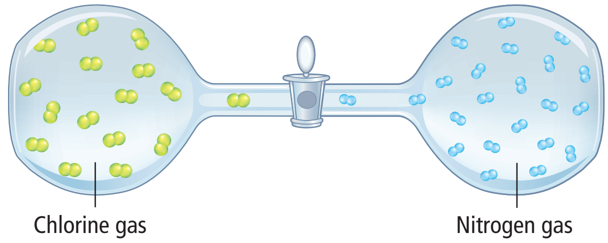
45. What is the partial pressure of water vapor in an air sample when the total pressure is 1.00 atm, the partial pressure of nitrogen is 0.79 atm, the partial pressure of oxygen is 0.20 atm, and the partial pressure of all other gases in air is 0.0044 atm?

46. What is the total gas pressure in a sealed flask that contains oxygen at a partial pressure of 0.41 atm and water vapor at a partial pressure of 0.58 atm?

47. Mountain Climbing The pressure atop the world's highest mountain, Mount Everest, is usually about 33.6 kPa. Convert the pressure to atmospheres. How does the pressure compare with the pressure at sea level?

48. High Altitude The atmospheric pressure in Denver, Colorado, is usually about 84.0 kPa. What is this pressure in atm and torr units?

49. At an ocean depth of 76.2 m, the pressure is about 8.4 atm. Convert the pressure to mm Hg and kPa units.



■ **Figure 12.32**

50. Figure 12.32 represents an experimental set-up in which the left bulb is filled with chlorine gas and the right bulb is filled with nitrogen gas. Describe what happens when the stopcock is opened. Assume that the temperature of the system is held constant during the experiment.

Section 12.2

Mastering Concepts

51. Explain the difference between a temporary dipole and a permanent dipole.

52. Why are dispersion forces weaker than dipole-dipole forces?

53. Explain why hydrogen bonds are stronger than most dipole-dipole forces.

54. Compare intramolecular and intermolecular forces.

55. Hypothesize why long, nonpolar molecules would interact more strongly with one another than spherical nonpolar molecules of similar composition.

Mastering Problems

56. Polar Molecules Use relative differences in electronegativity to label the ends of the polar molecules listed as partially positive or partially negative.
a. HF **b.** HBr **c.** NO **d.** CO

57. Draw the structure of the dipole-dipole interaction between two molecules of carbon monoxide.

58. Decide which of the substances listed can form hydrogen bonds.
a. H_2O **b.** H_2O_2 **c.** HF **d.** NH_3

Chemistry Online Chapter Test glencoe.com

59. Decide which one of the molecules listed below can form intermolecular hydrogen bonds, and then draw it, showing several molecules attached together by hydrogen bonds.
a. NaCl **b.** $MgCl_2$ **c.** H_2O_2 **d.** CO_2

Section 12.3

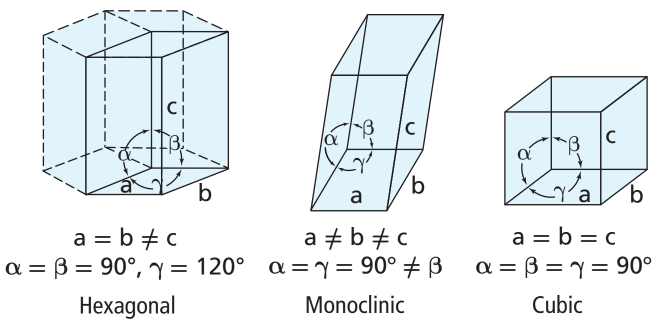
Mastering Concepts

60. What is surface tension, and what conditions must exist for it to occur?

61. Explain why the surface of water in a graduated cylinder is curved.

62. Which liquid is more viscous at room temperature, water or molasses? Explain.

63. Explain how two different forces play a role in capillary action.



■ **Figure 12.33**

64. Use the drawings in **Figure 12.33** to compare the cubic, monoclinic, and hexagonal crystal systems.

65. What is the difference between a network solid and an ionic solid?

66. Explain why most metals bend when struck but most ionic solids shatter.

67. List the types of crystalline solids that are usually good conductors of heat and electricity.

68. How does the strength of a liquid's intermolecular forces affect its viscosity?

69. Explain why water has a higher surface tension than benzene, whose molecules are nonpolar.

70. Compare the number of particles in one unit cell for each of the following types of unit cells.
a. simple cubic
b. body-centered cubic

71. Predict which solid is more likely to be amorphous—one formed by cooling a molten material over 4 h at room temperature or one formed by cooling a molten material quickly in an ice bath.

72. **Conductivity** Predict which solid will conduct electricity better—sugar or salt.

73. Explain why ice floats in water but solid benzene sinks in liquid benzene. Which behavior is more "normal"?

Mastering Problems

74. Given edge lengths and face angles, predict the shape of each of the following crystals.
a. $a = 3$ nm, $b = 3$ nm, $c = 3$ nm; $\alpha = 90°$, $\beta° = 90$, $\gamma = 90°$
b. $a = 4$ nm, $b = 3$ nm, $c = 5$ nm; $\alpha = 90°$, $\beta° = 100$, $\gamma = 90°$
c. $a = 3$ nm, $b = 3$ nm, $c =5$nm; $\alpha = 90°$, $\beta° = 90$, $\gamma = 90°$
d. $a = 3$ nm, $b = 3$ nm, $c = 5$ nm; $\alpha = 90°$, $\beta° = 90$, $\gamma = 120°$

Section 12.4

Mastering Concepts

75. How does sublimation differ from deposition?

76. Compare boiling and evaporation.

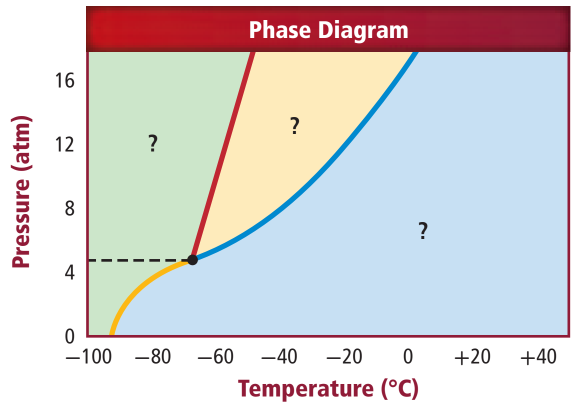
77. Define the term *melting point*.

78. Explain the relationships among vapor pressure, atmospheric pressure, and boiling point.

79. Explain why dew forms on cool mornings.

80. **Snow** Why does a pile of snow slowly shrink even on days when the temperature never rises above the freezing point of water?

Mastering Problems



■ **Figure 12.34**

81. Copy and label the solid, liquid, and gas phases, triple point, and critical point on **Figure 12.34.**

82. Why does it take more energy to boil 10 g of liquid water than to melt an equivalent mass of ice?

Mixed Review

83. Use the kinetic-molecular theory to explain why both gases and liquids are fluids.

84. Use intermolecular forces to explain why oxygen is a gas at room temperature and water is a liquid.

85. Use the kinetic-molecular theory to explain why gases are easier to compress than liquids or solids.

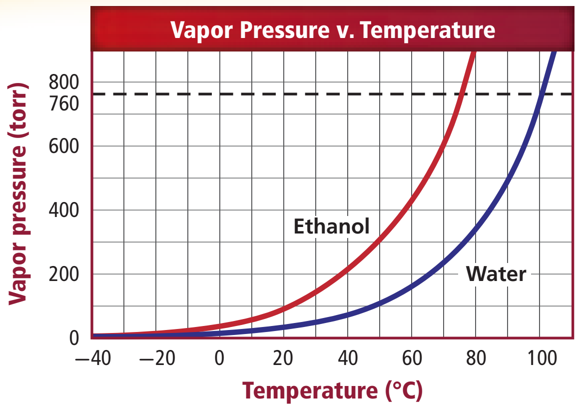
86. At 25°C and a pressure of 760 mm Hg, the density of mercury is 13.5 g/mL; water at the same temperature and pressure has a density of 1.00 g/mL. Explain this difference in terms of intermolecular forces and the kinetic-molecular theory.

87. If two identical containers each hold the same gas at the same temperature but the pressure inside one container is exactly twice that of the other container, what must be true about the amount of gas inside each container?

88. List three types of intermolecular forces.

89. When solid sugar crystals are dissolved in a glass of water, they form a clear homogeneous solution in which the crystals are not visible. If the beaker is left out at room temperature for a few days, the crystals reappear in the bottom and on the sides of the glass. Is this an example of freezing?

Think Critically



■ **Figure 12.35**

90. Interpret Graphs Examine **Figure 12.35,** which plots vapor pressure versus temperature for water and ethyl alcohol.

a. What is the boiling point of water at 1 atm?
b. What is the boiling point of ethyl alcohol at 1 atm?
c. Estimate the temperature at which water will boil when the atmospheric pressure is 0.80 atm.

91. Hypothesize What type of crystalline solid do you predict would best suit the following needs?

a. a material that can be melted and reformed at a low temperature
b. a material that can be drawn into long, thin wires
c. a material that conducts electricity when molten
d. an extremely hard material that is nonconductive

92. Compare and Contrast An air compressor uses energy to squeeze air particles together. When the air is released, it expands, allowing the energy to be used for purposes such as gently cleaning surfaces without using a more abrasive liquid or solid. Hydraulic systems essentially work the same way, but involve compression of liquid water rather than air. What do you think are some advantages and disadvantages of these two types of technology?

93. Graph Use **Table 12.6** to construct a phase diagram for ammonia.

Table 12.6 Phase Diagram for Ammonia

Selected Points	Pressure (atm)	Temperature (°C)
Triple point	0.060	−77.7
Critical point	112	132.2
Normal boiling point	1.0	−33.5
Normal freezing point	1.0	−77.7

94. Apply A solid being heated stays at a constant temperature until it is completely melted. What happens to the heat energy put into the system during that time?

95. Communicate Which process—effusion or diffusion—is responsible for your being able to smell perfume from an open bottle that is located across the room from you? Explain.

96. Infer A laboratory demonstration involves pouring bromine vapors, which are a deep red color, into a flask of air and then tightly sealing the top of the flask. The bromine is observed to first sink to the bottom of the beaker. After several hours have passed, the red color is distributed equally throughout the flask.

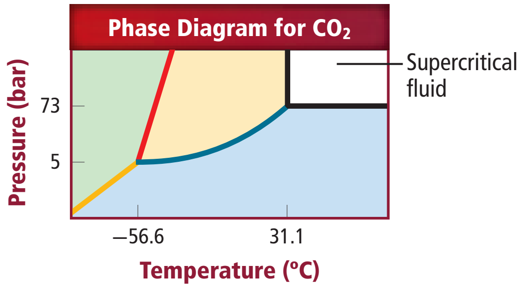
a. Is bromine gas more or less dense than air?
b. Would liquid bromine diffuse more or less quickly than gaseous bromine after you pour it into another liquid?

97. Analyze Use your knowledge of intermolecular forces to predict whether ammonia (NH_3) or methane (CH_4) will be more soluble in water.

98. Evaluate List three changes that require energy and three that release energy.

Chemistry Online Chapter Test glencoe.com

99. Evaluate Supercritical carbon dioxide is a liquid form of CO_2 used in the food industry to decaffeinate tea, coffee, and colas, as well as in the pharmaceutical industry to form polymer microparticles used in drug delivery systems. Use **Figure 12.36** to determine what conditions must be used to form supercritical carbon dioxide.



■ **Figure 12.36**

Challenge Problem

100. You have a solution containing 135.2 g of dissolved KBr in 2.3 L of water. What volume of this solution, in mL, would you use to make 1.5 L of a 0.1 mol/L KBr solution? What is the boiling point of this new solution?

Cumulative Review

101. Identify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture. *(Chapter 3)*

a. air
b. blood
c. antimony
d. ammonia
e. mustard
f. water

102. You are given two clear, colorless aqueous solutions. You are told that one solution contains an ionic compound, and one contains a covalent compound. How could you determine which is an ionic solution and which is a covalent solution? *(Chapter 8)*

103. Which branch of chemistry would most likely study matter and phase changes? *(Chapter 1)*

a. biochemistry
b. organic chemistry
c. physical chemistry
d. polymer chemistry

104. What type of reaction is the following? *(Chapter 9)*

$K_2CO_3(aq) + BaCl_2(aq) \longrightarrow 2KCl(aq) + BaCO_3(s)$

a. combustion
b. double-replacement
c. single-replacement
d. synthesis

105. Which chemist produced the first widely used and accepted periodic table? *(Chapter 6)*

a. Dmitri Mendeleev
b. Henry Moseley
c. John Newlands
d. Lothar Meyer

Additional Assessment

WRITING in Chemistry

106. Musk is the basic ingredient of many perfumes, soaps, shampoos, and even foods such as chocolates, licorice, and hard candies. Both synthetic and natural musk molecules have high molecular weights compared to other perfume ingredients, and as a result, have a slower rate of diffusion, assuring a slow, sustained release of fragrance. Write a report on the chemistry of perfume ingredients, emphasizing the importance of diffusion rate as a property of perfume.

107. Birthstones Find out what your birthstone is and write a brief report about the chemistry of that gem. Find out its chemical composition, which category its unit cell is in, how hard and durable it is, and what its approximate cost is at present.

108. Propane gas is a commonly used heating fuel for gas grills and homes. However, it is not packaged as a gas. It is liquefied and referred to as liquid propane or "LP gas." Make a poster explaining the advantages and disadvantages of storing and transporting propane as a liquid rather than a gas.

109. Other States of Matter Research and prepare an oral report about one of the following topics: plasma, superfluids, fermionic condensate, or Bose-Einstein condensate. Share your report with your classmates and prepare a visual aid that can be used to explain your topic.

DBQ Document-Based Questions

Iodine *Solid iodine that is left at room temperature sublimates from a solid to a gas. But when heated quickly, a different process takes place, as described here.*

"About 1 g of iodine crystals is placed in a sealed glass ampoule and gently heated on a hot plate. A layer of purple gas is formed at the bottom, and the iodine liquefies. If one tilts the tube, this liquid flows along the wall as a narrow stream and solidifies very quickly."

Data obtained from: Leenson, 2005. Sublimination of Iodine at Various Pressures: Multipurpose Experiments in Inorganic and Physical Chemistry. *Journal of Chemical Education* 82(2):241–245.

110. Why does solid iodine sublime readily? Use your knowledge of intermolecular forces to explain.

111. Why is liquid iodine not usually visible if crystals are heated in the open air?

112. Why is it necessary to use a sealed ampoule in this investigation?

113. Infer why the iodine solidifies when the tube is tilted.

Cumulative
Standardized Test Practice

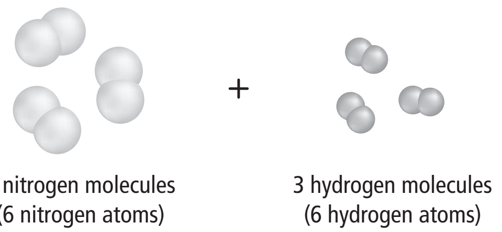
Multiple Choice

1. What is the ratio of diffusion rates for nitric oxide (NO) and nitrogen tetroxide (N_2O_4)?
   A. 0.326
   B. 0.571
   C. 1.751
   D. 3.066

2. Which is NOT an assumption of the kinetic-molecular theory?
   A. Collisions between gas particles are elastic.
   B. All the gas particles in a sample have the same velocity.
   C. A gas particle is not significantly attracted or repelled by other gas particles.
   D. All gases at a given temperature have the same average kinetic energy.

3. A sealed flask contains neon, argon, and krypton gas. If the total pressure in the flask is 3.782 atm, the partial pressure of Ne is 0.435 atm, and the partial pressure of Kr is 1.613 atm, what is the partial pressure of Ar?
   A. 2.048 atm
   B. 1.734 atm
   C. 1556 atm
   D. 1318 atm

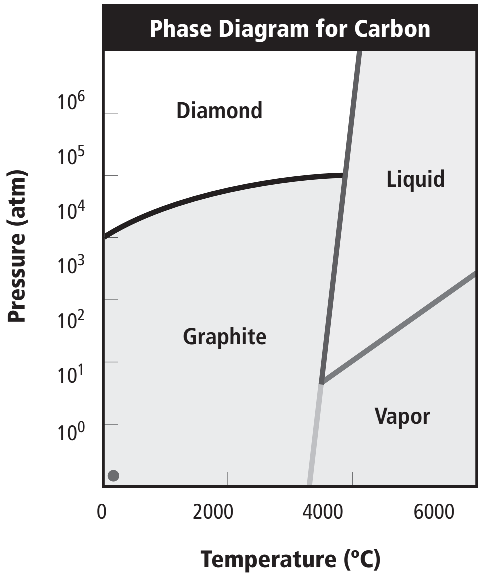
Use the figure below to answer Question 4.



4. Hydrogen and nitrogen react as shown to form ammonia (NH_3). What is true of this reaction?
   A. Three ammonia molecules are formed, with zero molecules remaining.
   B. Two ammonia molecules are formed, with two hydrogen molecules remaining.
   C. Six ammonia molecules are formed, with zero molecules remaining.
   D. Two ammonia molecules are formed, with two nitrogen molecules remaining.

5. Which does not affect the viscosity of a liquid?
   A. intermolecular attractive forces
   B. size and shape of molecules
   C. temperature of the liquid
   D. capillary action

Use the graph below to answer Questions 6 to 8.



6. Under what conditions is diamond most likely to form?
   A. temperatures > 5000 K and pressures < 100 atm
   B. temperatures > 6000 K and pressures < 25 atm
   C. temperatures < 3500 K and pressures > 10^5 atm
   D. temperatures < 4500 K and pressures < 10 atm

7. Find the point on the graph at which carbon exists in three phases: solid graphite, solid diamond, and liquid carbon. What are the approximate temperature and pressure at that point?
   A. 4700 K and 10^6 atm
   B. 3000 K and 10^3 atm
   C. 5100 K and 10^5 atm
   D. 3500 K and 80 atm

8. In what form or forms does carbon exist at 6000 K and 10^5 atm?
   A. diamond only
   B. liquid carbon only
   C. diamond and liquid carbon
   D. liquid carbon and graphite

Short Answer

Use the table below to answer Questions 9 and 10.

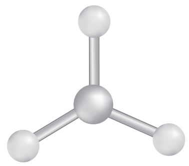
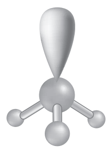
Properties of Single Bonds		
Bond	**Strength (kJ/mol)**	**Length (pm)**
H – H	435	74
Br – Br	192	228
C – C	347	154
C – H	393	104
C – N	305	147
C – O	356	143
Cl – Cl	243	199
I – I	151	267
S – S	259	208

9. Create a graph to show how bond length varies with bond strength. Place bond strength on the *x*-axis.

10. Summarize the relationship between bond strength and bond length.

Extended Response

Use the table below to answer Question 11.

Geometry of $AlCl_3$ and PCl_3		
Compound	$AlCl_3$	PCl_3
Molecular Shape		

11. What are the names of the shapes of the molecules for each compound? Explain how the atomic arrangements in each compound result in their different shapes despite their similar formulas.

SAT Subject Test: Chemistry

12. Potassium chromate and lead(II) acetate are both dissolved in a beaker of water, where they react to form solid lead(II) chromate. What is the balanced net ionic equation describing this reaction?
   - **A.** $Pb^{2+}(aq) + C_2H_3O_2^{-}(aq) \longrightarrow Pb(C_2H_3O_2)_2(s)$
   - **B.** $Pb^{2+}(aq) + 2CrO_4^{-}(aq) \longrightarrow Pb(CrO_4)_2(s)$
   - **C.** $Pb^{2+}(aq) + CrO_4^{2-}(aq) \longrightarrow PbCrO_4(s)$
   - **D.** $Pb^{+}(aq) + C_2H_3O_2^{-}(aq) \longrightarrow PbC_2H_3O_2(s)$
   - **E.** $Pb^{2+}(aq) + CrO_4^{-}(aq) \longrightarrow PbCrO_5(s)$

13. The solid phase of a compound has a definite shape and volume because its particles
   - **A.** are not in constant motion.
   - **B.** are always more tightly packed in the liquid phase.
   - **C.** can vibrate only around fixed points.
   - **D.** are held together by strong intramolecular forces.
   - **E.** have no intermolecular forces.

Use the table below to answer Questions 14 and 15.

Properties of Sulfuric Acid	
Formula	H_2SO_4
Molar mass	98.08 g/mol
Density	1.834 g/mL

14. What is the mass of 75.0 mL of sulfuric acid?
   - **A.** 40.9 g
   - **B.** 138 g
   - **C.** 98.08 g
   - **D.** 180 g
   - **E.** 198.4 g

15. How many atoms of oxygen are present in 235 g of sulfuric acid?
   - **A.** 9.42×10^{22} atoms
   - **B.** 2.35×10^{26} atoms
   - **C.** 1.44×10^{24} atoms
   - **D.** 5.78×10^{24} atoms
   - **E.** 6.02×10^{23} atoms

NEED EXTRA HELP?															
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Review Section . . .	12.1	12.1	12.1	11.1	12.3	12.4	12.4	12.4	8.1	8.1	8.4	9.3	12.3	2.1	10.3