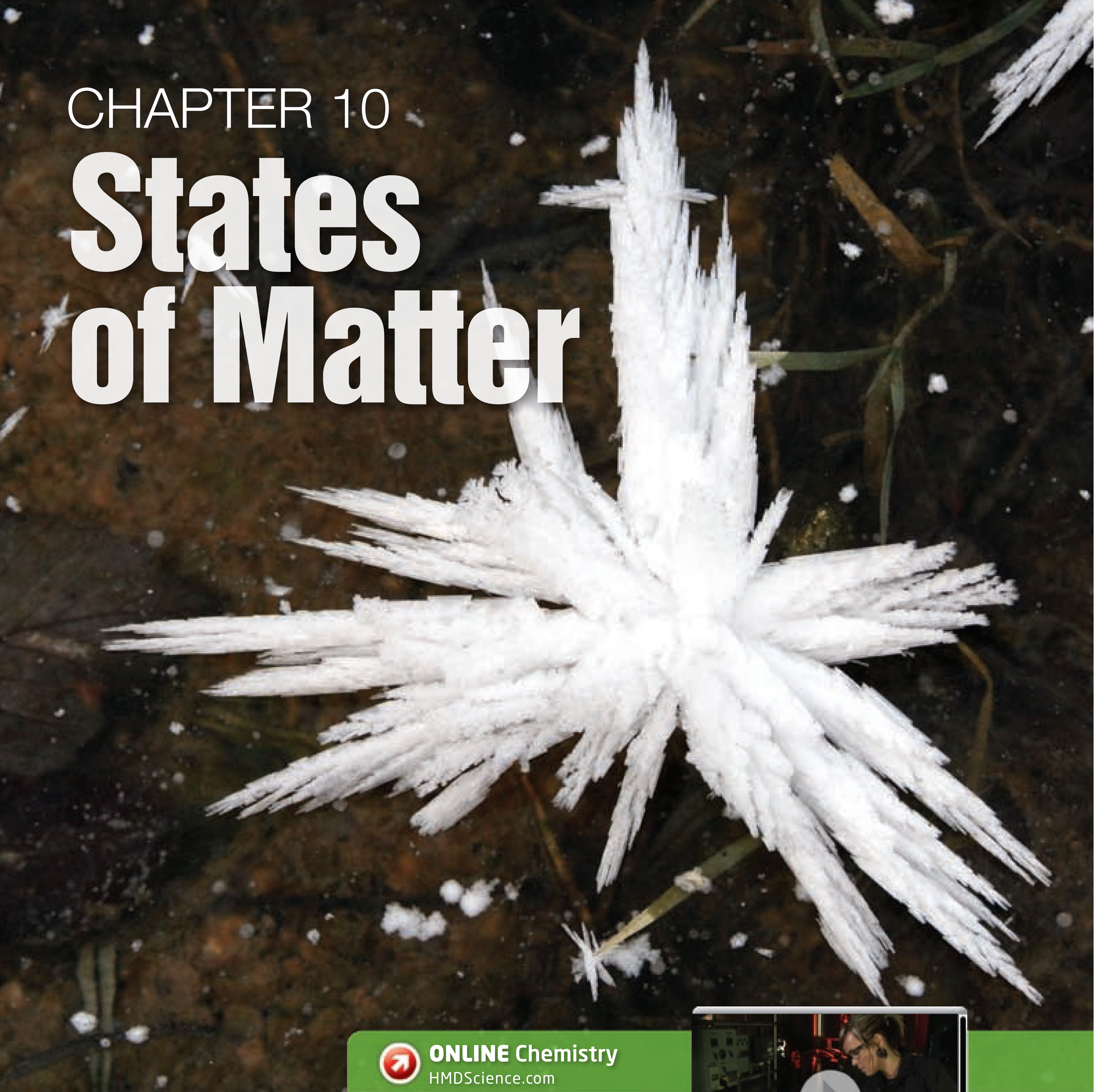
CHAPTER 10

States of Matter



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Changes of State

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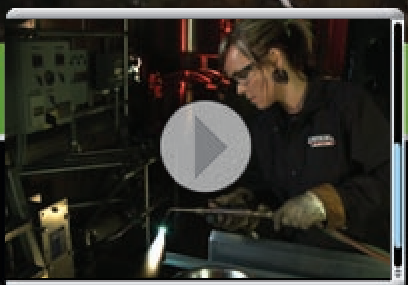

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States of Matter

The Kinetic-Molecular Theory of Matter

SECTION 1

Main Ideas

- The kinetic-molecular theory explains the constant motion of gas particles.
- The kinetic-molecular theory explains the physical properties of gases.
- Real gases do not behave according to the kinetic-molecular theory.

VIRGINIA STANDARDS

CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles.
CH.5.EKS-5

Key Terms

kinetic-molecular theory	elastic collision	effusion
ideal gas	diffusion	real gas

In the chapter "Matter and Change," you read that matter exists on Earth in the forms of solids, liquids, and gases. Although it is not usually possible to observe individual particles directly, scientists have studied large groups of these particles as they occur in solids, liquids, and gases.

In the late nineteenth century, scientists developed the kinetic-molecular theory of matter to account for the behavior of the atoms and molecules that make up matter. **The kinetic-molecular theory is based on the idea that particles of matter are always in motion.** The theory can be used to explain the properties of solids, liquids, and gases in terms of the energy of particles and the forces that act between them. In this section, you will study the theory as it applies to gas molecules.

MAIN IDEA

The kinetic-molecular theory explains the constant motion of gas particles.

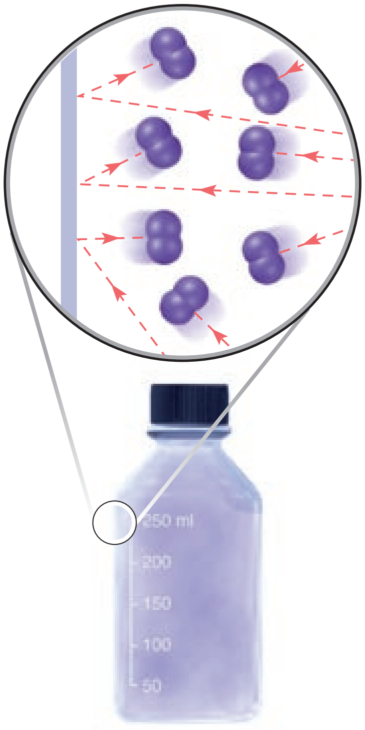
The kinetic-molecular theory can help you understand the behavior of gas molecules and the physical properties of gases. The theory provides a model of what is called an ideal gas. **An ideal gas is a hypothetical gas that perfectly fits all the assumptions of the kinetic-molecular theory.**

The kinetic-molecular theory of gases is based on the following five assumptions:

1. *Gases consist of large numbers of tiny particles that are far apart relative to their size.* These particles, usually molecules or atoms, typically occupy a volume that is about 1000 times greater than the volume occupied by an equal number of particles in the liquid or solid state. Thus, molecules of gases are much farther apart than molecules of liquids or solids. Most of the volume occupied by a gas is empty space, which is the reason that gases have a lower density than liquids and solids do. This also explains the fact that gases are easily compressed.
2. *Collisions between gas particles and between particles and container walls are elastic collisions.* **An elastic collision is one in which there is no net loss of total kinetic energy.** Kinetic energy is transferred between two particles during collisions. However, the total kinetic energy of the two particles remains the same as long as temperature is constant.

FIGURE 1.1

Elastic Collisions Gas particles travel in a straight-line motion until they collide with each other or the walls of their container.



3. *Gas particles are in continuous, rapid, random motion. They therefore possess kinetic energy, which is energy of motion.* Gas particles move in all directions, as shown in **Figure 1.1.** The kinetic energy of the particles overcomes the attractive forces between them, except near the temperature at which the gas condenses and becomes a liquid.
4. *There are no forces of attraction between gas particles.* You can think of ideal gas molecules as behaving like small billiard balls. When they collide, they do not stick together but immediately bounce apart.
5. *The temperature of a gas depends on the average kinetic energy of the particles of the gas.* The kinetic energy of any moving object, including a particle, is given by the following equation:

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

In the equation, m is the mass of the particle and v is its speed. Because all the particles of a specific gas have the same mass, their kinetic energies depend only on their speeds. The average speeds and kinetic energies of gas particles increase with an increase in temperature and decrease with a decrease in temperature.

All gases at the same temperature have the same average kinetic energy. Therefore, at the same temperature, lighter gas particles, such as hydrogen molecules, have higher average speeds than do heavier gas particles, such as oxygen molecules.

MAIN IDEA

The kinetic-molecular theory explains the physical properties of gases.

The kinetic-molecular theory applies only to ideal gases. Although ideal gases do not actually exist, many gases behave nearly ideally if pressure is not very high and temperature is not very low. As you will see, the kinetic-molecular theory accounts for the physical properties of gases.

Expansion

Gases do not have a definite shape or volume. They completely fill any container in which they are enclosed, and take its shape. A gas transferred from a one-liter vessel to a two-liter vessel will expand to fill the larger volume. The kinetic-molecular theory explains these facts. According to the theory, gas particles move rapidly in all directions (assumption 3), with no significant attraction between them (assumption 4).

Fluidity

Because the attractive forces between gas particles are insignificant (assumption 4), gas particles glide easily past one another. This ability to flow causes gases to behave as liquids do. Because liquids and gases flow, they are both referred to as *fluids*.

Low Density

The density of a gaseous substance at atmospheric pressure is about 1/1000 the density of the same substance in the liquid or solid state. The reason is that the particles are so much farther apart in the gaseous state (assumption 1).

Compressibility

During compression, the gas particles, which are initially very far apart (assumption 1), are crowded closer together. The volume of a given sample of a gas can be greatly decreased. Steel cylinders containing gases under pressure are widely used in industry. When they are full, such cylinders may contain more than 100 times as many particles of gas as nonpressurized containers of the same size could contain.

Diffusion and Effusion

Gases spread out and mix with one another, even without being stirred. If the stopper is removed from a container of ammonia in a room, ammonia gas will mix uniformly with the air and spread throughout the room. The random and continuous motion of the ammonia molecules (assumption 3) carries them throughout the available space. **Such spontaneous mixing of the particles of two substances caused by their random motion is called diffusion.**

Gases diffuse readily into one another and mix together due to the rapid motion of the molecules and the empty space between the molecules. The gas molecules in each of the two flasks in **Figure 1.2a** continuously move about in separate flasks because the stopcock is closed. When the stopcock is open, the gas molecules continuously diffuse back and forth from one flask to the other through the opening in the stopcock, as shown in **Figure 1.2b.**

Diffusion is a process by which particles of a gas spread out spontaneously and mix with other gases. **In contrast, effusion is a process by which gas particles pass through a tiny opening.**

FIGURE 1.2

Diffusion Gases diffuse readily into one another. The space between the molecules allows different gases to mix together easily.

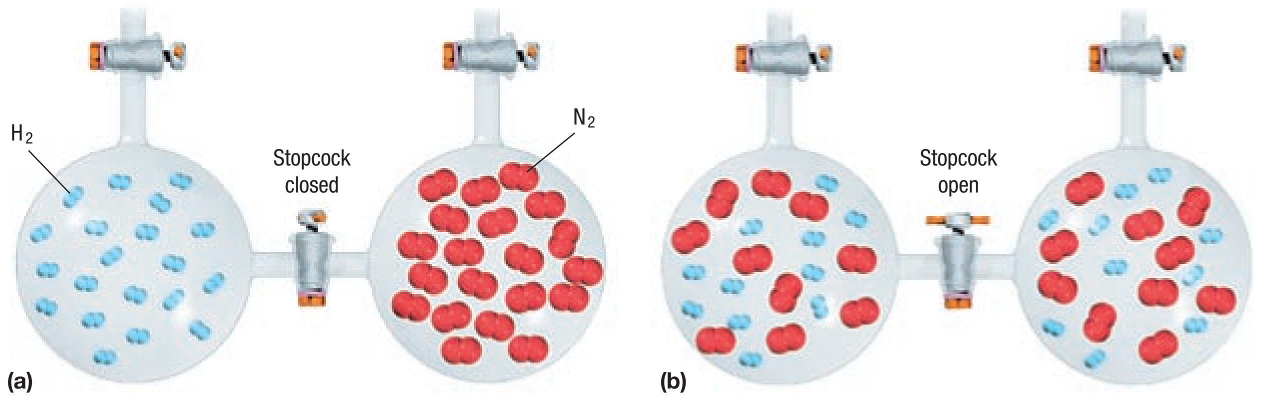
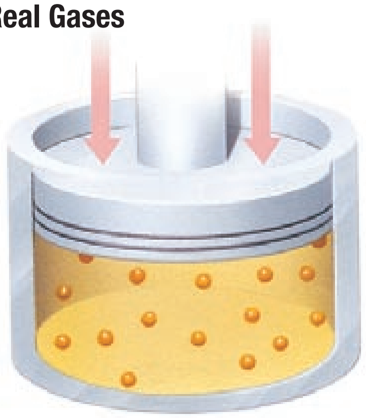
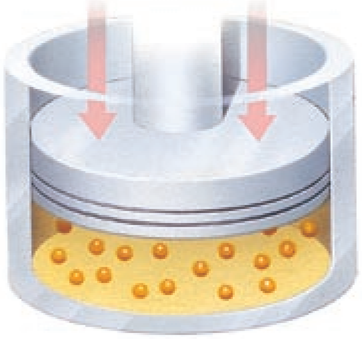


FIGURE 1.3

Real Gases



(a) Gas molecules in a car engine cylinder expand to fill the cylinder.



(b) As pressure is exerted on them, the gas molecules move closer together, reducing their volume.

CRITICAL THINKING

Describe What will eventually happen to a gas if it is compressed enough?

The rates of effusion of different gases are directly proportional to the velocities of their particles. Because of this proportionality, molecules of low mass effuse faster than molecules of high mass.

MAIN IDEA

Real gases do not behave according to the kinetic-molecular theory.

Because particles of gases occupy space and exert attractive forces on each other, all real gases deviate to some degree from ideal gas behavior. **A real gas is a gas that does not behave completely according to the assumptions of the kinetic-molecular theory.** At very high pressures and low temperatures, the gas particles will be closer together and their kinetic energy will be insufficient to completely overcome the attractive forces. At such conditions, the gas is most likely to behave like a non-ideal gas. These conditions are illustrated in **Figure 1.3.**

The kinetic-molecular theory is more likely to hold true for gases whose particles have little attraction for each other. The noble gases, such as helium, He, and neon, Ne, show essentially ideal gas behavior over a wide range of temperatures and pressures. The particles of these gases are monatomic and thus nonpolar. The particles of gases, such as nitrogen, N_2, and hydrogen, H_2, are nonpolar diatomic molecules. The behavior of these gases most closely approximates that of the ideal gas under certain conditions. The more polar the molecules of a gas are, the greater the attractive forces between them and the more the gas will deviate from ideal gas behavior. For example, highly polar gases, such as ammonia (NH_3) and water vapor, deviate from ideal behavior to a larger degree than nonpolar gases.

SECTION 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Use the kinetic-molecular theory to explain the following properties of gases: expansion, fluidity, low density, compressibility, and diffusion.
2. Describe the conditions under which a real gas is most likely to behave ideally.
3. Which of the following gases would you expect to deviate significantly from ideal behavior: He, O_2, H_2, H_2O, N_2, HCl, or NH_3?
4. How does the kinetic-molecular theory explain the pressure exerted by gases?
5. What happens to gas particles when a gas is compressed?
6. What happens to gas particles when a gas is heated?

Critical Thinking

7. **DRAWING CONCLUSIONS** Molecules of hydrogen escape from Earth, but molecules of oxygen and nitrogen are held to the surface and remain in the atmosphere. Explain.

SECTION 2

Main Idea

The intermolecular forces of liquids determine their properties.

Liquids

Key Terms

fluid
surface tension
capillary action
vaporization
evaporation
freezing

VIRGINIA STANDARDS

CH.5.d The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include: phase changes.

When you think of Earth's oceans, lakes, and rivers and the many liquids you use every day, it is hard to believe that liquids are the *least* common state of matter in the universe. Liquids are less common than solids and gases because a substance can exist in the liquid state only within a relatively narrow range of temperatures and pressures. In this section, you will examine the properties of the liquid state and compare them with those of the solid state and the gas state. These properties will be discussed in terms of the kinetic-molecular theory.

MAIN IDEA

The intermolecular forces of liquids determine their properties.

A liquid can be described as a form of matter that has a definite volume and takes the shape of its container. The properties of liquids can be understood by applying the kinetic-molecular theory.

As in a gas, particles in a liquid are in constant motion. However, the particles in a liquid are closer together than the particles in a gas are. Therefore, the attractive forces between particles in a liquid are more effective than those between particles in a gas. This attraction between liquid particles is caused by intermolecular forces, such as dipole-dipole forces, London dispersion forces, and hydrogen bonding.

Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of the liquid particles. According to the kinetic-molecular theory of liquids, the particles are not bound together in fixed positions, but move about constantly. This particle mobility explains why liquids and gases are referred to as fluids. **A fluid is a substance that can flow and therefore take the shape of its container.** Most liquids naturally flow downhill because of gravity. However, some liquids can flow in other directions as well. For example, liquid helium near absolute zero is able to flow uphill.

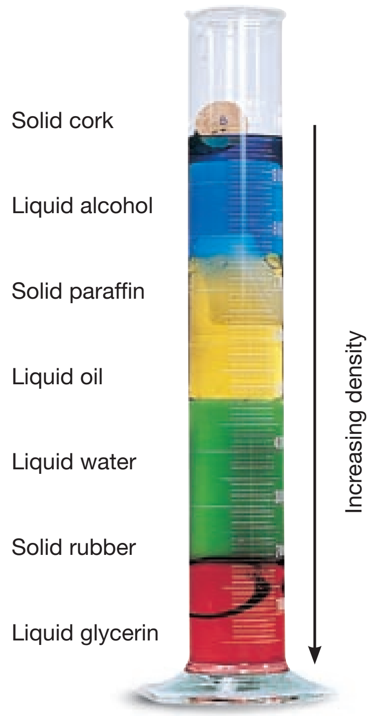
Relatively High Density

At normal atmospheric pressure, most substances are hundreds of times denser in a liquid state than in a gaseous state. This higher density is a result of the close arrangement of liquid particles. Most substances are only slightly less dense (about 10%) in a liquid state than in a solid state, however. Water is one of the few substances that becomes less dense when it solidifies, as will be discussed further in Section 5.

At the same temperature and pressure, different liquids can differ greatly in density. **Figure 2.1** shows some liquids and solids with different densities. The densities differ to such an extent that the liquids form layers.

FIGURE 2.1

Density Solids and liquids of different densities are shown. The densest materials are at the bottom. The least dense are at the top. (Dyes have been added to the liquids to make the layers more visible.)



Relative Incompressibility

When liquid water at 20°C is compressed by a pressure of 1000 atm, its volume decreases by only 4%. Such behavior is typical of all liquids and is similar to the behavior of solids. In contrast, a gas under a pressure of 1000 atm would have only about 1/1000 of its volume at normal atmospheric pressure. Liquids are much less compressible than gases because liquid particles are more closely packed together. Like gases, liquids can transmit pressure equally in all directions.

Ability to Diffuse

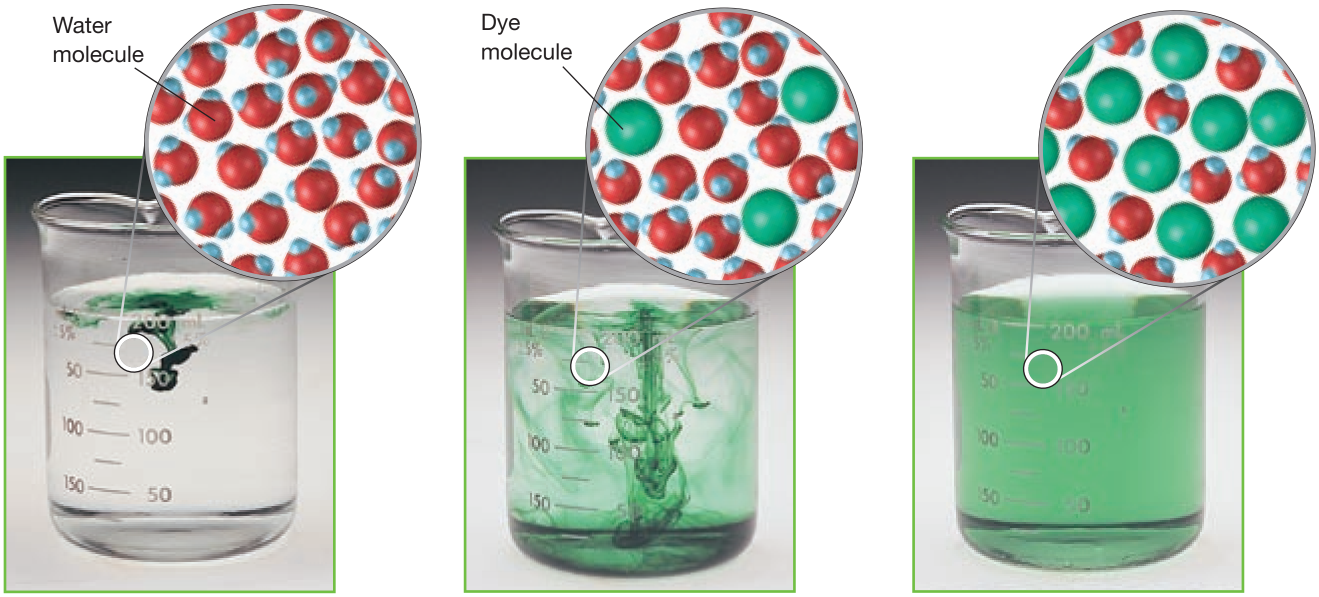
As described in Section 1, gases diffuse and mix with other gas particles. Liquids also diffuse and mix with other liquids, as shown in **Figure 2.2.** Any liquid gradually diffuses throughout any other liquid in which it can dissolve. The constant, random motion of particles causes diffusion in liquids, as it does in gases. Yet diffusion is much slower in liquids than in gases because liquid particles are closer together. Also, the attractive forces between the particles of a liquid slow their movement. As the temperature of a liquid is increased, diffusion occurs more rapidly. The reason is that the average kinetic energy, and therefore the average speed of the particles, is increased.

Surface Tension

A property common to all liquids is surface tension, a force that tends to pull adjacent parts of a liquid's surface together, thereby decreasing surface area to the smallest possible size. Surface tension results from the attractive forces between particles of a liquid. The higher the force of attraction, the higher the surface tension.

FIGURE 2.2

Diffusion of Liquids Like gases, the two liquids in this beaker diffuse over time. The green liquid food coloring from the drop will eventually form a uniform solution with the water.



Water has a higher surface tension than most liquids. This is due in large part to the hydrogen bonds water molecules can form with each other. The molecules at the surface of the water are a special case. They can form hydrogen bonds with the other water molecules beneath them and beside them, but not with the molecules in the air above them. As a result, the surface water molecules are drawn together and toward the body of the liquid, creating a high surface tension. Surface tension causes liquid droplets to take on a spherical shape because a sphere has the smallest possible surface area for a given volume. An example of this phenomenon is shown in **Figure 2.3.**

Capillary action, the attraction of the surface of a liquid to the surface of a solid, is a property closely related to surface tension. A liquid will rise quite high in a very narrow tube and will wet the tube if a strong attraction exists between the liquid molecules and the molecules that make up the surface of the tube. This attraction tends to pull the liquid molecules upward along the surface and against the pull of gravity. This process continues until the attractive forces between the liquid molecules and the surface of the tube are balanced by the weight of the liquid. Capillary action can occur between water molecules and paper fibers, as shown in **Figure 2.4.** Capillary action is at least partly responsible for the transportation of water from the roots of a plant to its leaves. The same process is responsible for the concave liquid surface, called a *meniscus*, that forms in a test tube or graduated cylinder.

Evaporation and Boiling

The process by which a liquid or solid changes to a gas is vaporization. Evaporation is a form of vaporization. **Evaporation is the process by which particles escape from the surface of a nonboiling liquid and enter the gas state.**

FIGURE 2.3

Surface Tension As a result of surface tension, liquids form roughly spherical drops. The net attractive forces between the particles pull the molecules on the surface of the drop inward. The molecules are pulled very close together, which minimizes the surface area.

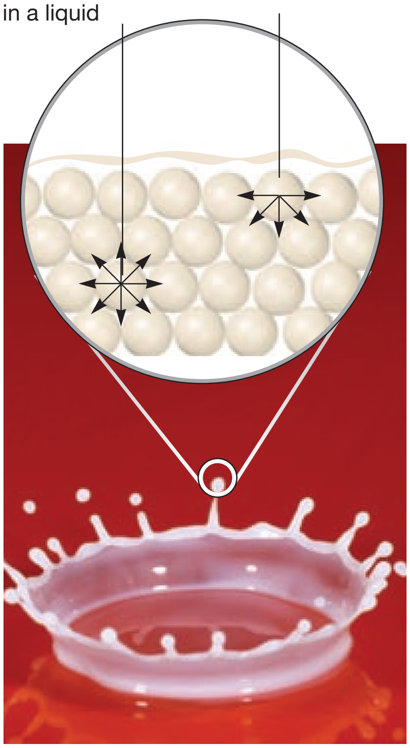


FIGURE 2.4

Polar Attraction

(a) Water-soluble ink is placed near the bottom of a piece of chromatography paper in shallow water.

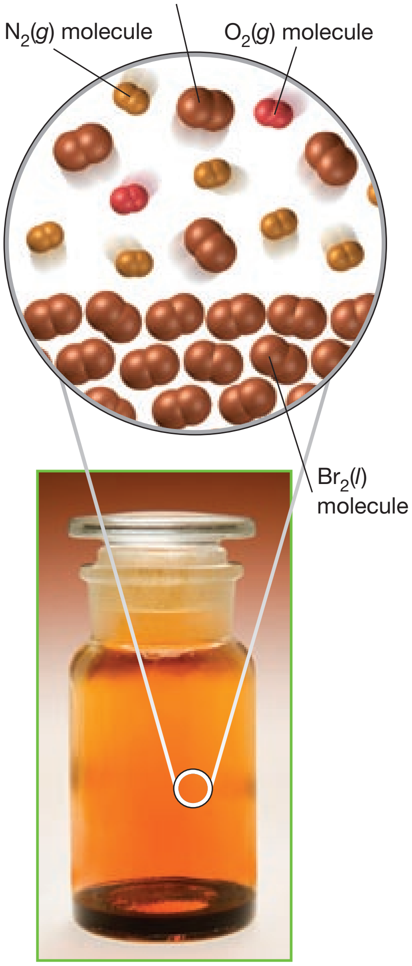
(b) The ink has separated and has risen up the paper. These actions are the result of specific properties of water molecules.

CRITICAL THINKING

Apply Using the principles of capillary action, describe why water-soluble ink rises up the chromatography paper, which is set into a small amount of water.

FIGURE 2.5

Evaporation Liquid bromine, Br_2, evaporates near room temperature. The resulting brownish-red gas diffuses into the air above the surface of the liquid.



A small amount of liquid bromine was added to the bottle shown in **Figure 2.5.** Within a few minutes, the air above the liquid bromine turned brownish-red because some bromine molecules escaped from the surface of the liquid. These molecules became gas molecules, or bromine vapor, which mixed with the air. A similar phenomenon occurs if you apply perfume to your wrist. Within seconds, you become aware of the perfume's fragrance. Scent molecules evaporate from your skin and diffuse through the air, where your nose detects them.

Evaporation occurs because the particles of a liquid have different kinetic energies. Particles with higher-than-average energies move faster. Some surface particles with higher-than-average energies can overcome the intermolecular forces that bind them to the liquid. They can then escape into the gas state.

Evaporation is a crucial process in nature. Evaporation removes fresh water from the surface of the ocean, leaving behind a higher concentration of salts. In tropical areas, evaporation occurs at a higher rate, causing the surface water to be saltier. All water that falls to Earth in the form of rain and snow previously evaporated from oceans, lakes, and rivers. Evaporation of perspiration plays an important role in keeping you cool. Perspiration, which is mostly water, cools you by absorbing body heat when it evaporates. Energy as heat is absorbed from the skin, causing the cooling effect.

Boiling is the change of a liquid to bubbles of vapor that appear throughout the liquid. Boiling differs from evaporation, as you will see in Section 4.

Formation of Solids

When a liquid is cooled, the average energy of its particles decreases. If the energy is low enough, attractive forces pull the particles into an even more orderly arrangement. The substance then becomes a solid. **The physical change of a liquid to a solid by removal of energy as heat is called freezing or solidification.** Perhaps the best-known example of freezing is the change of liquid water to solid water, or ice, at 0°C. Another familiar example is the solidification of paraffin at room temperature. All liquids freeze, although not necessarily at temperatures you normally encounter. Ethanol, for example, freezes near −114°C.

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Describe the liquid state according to the kinetic-molecular theory.
2. List the properties of liquids.
3. How does the kinetic-molecular theory explain the following properties of liquids: (a) relatively high density, (b) ability to diffuse, and (c) ability to evaporate?
4. Explain why liquids in a test tube form a meniscus.
5. Compare vaporization and evaporation.

Critical Thinking

6. **INTERPRETING CONCEPTS** The evaporation of liquid water from the surface of Earth is an important step in the water cycle. How do water molecules obtain enough kinetic energy to escape into the gas state?

SECTION 3

Solids

Main Ideas

- The particles in a solid hold relatively fixed positions.
- Crystal particles are arranged in a 3-dimensional lattice.
- The particles in amorphous solids are not arranged in a regular pattern.

Key Terms

crystalline solid
crystal
amorphous solid
melting
melting point
supercooled liquid
crystal structure
unit cell

The common expression "solid as a rock" suggests something that is hard or unyielding and has a definite shape and volume. In this section you will examine the properties of solids and compare them with those of liquids and gases. The properties of solids are explained in terms of the kinetic-molecular theory, as are the other states of matter.

VIRGINIA STANDARDS

CH.5.d The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include: phase changes.

MAIN IDEA

The particles in a solid hold relatively fixed positions.

The particles of a solid are more closely packed than those of a liquid or gas. Intermolecular forces between particles are thus much more effective in solids. All interparticle attractions, such as dipole-dipole attractions, London dispersion forces, and hydrogen bonding, exert stronger effects in solids than in the corresponding liquids or gases. Attractive forces tend to hold the particles of a solid in relatively fixed positions, with only vibrational movement around fixed points. Because the motions of the particles are restricted in this way, solids are more ordered than liquids and are much more ordered than gases. The importance of order and disorder in physical and chemical changes will be discussed in the chapter "Reaction Energy." Compare the physical appearance and molecular arrangement of the element in **Figure 3.1** in solid, liquid, and gas form.

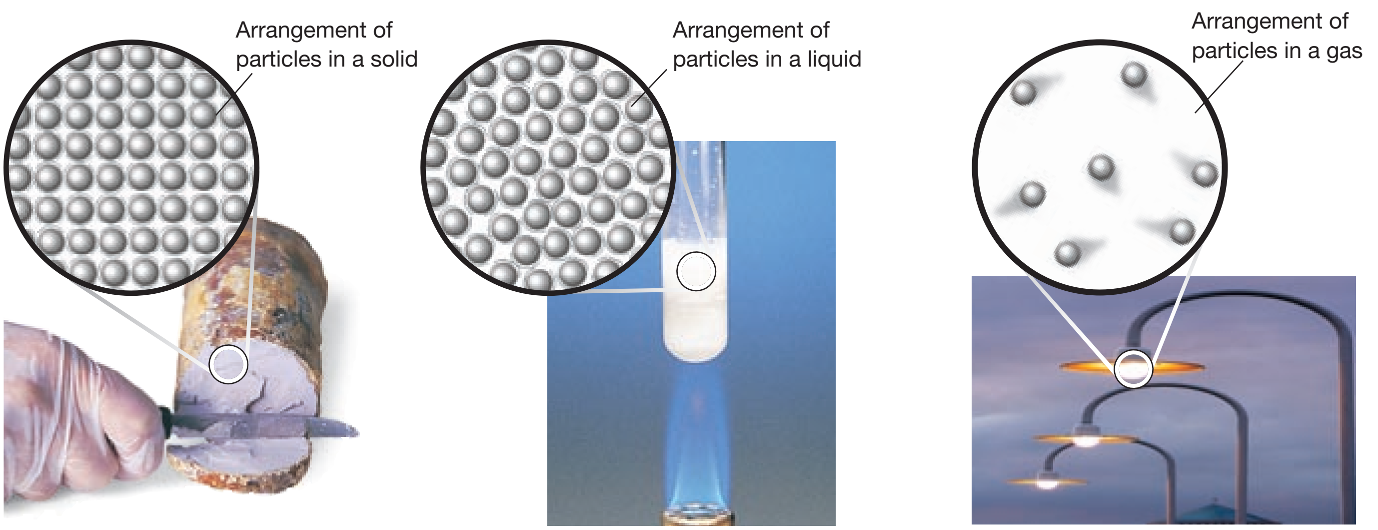
CRITICAL THINKING

Analyze Analyze the pictures below and describe in detail what each picture is showing.

FIGURE 3.1

States of Matter

Particles of sodium metal in three different states are shown.



There are two types of solids: crystalline solids and amorphous solids. **Most solids are crystalline solids—they consist of crystals. A crystal is a substance in which the particles are arranged in an orderly, geometric, repeating pattern.** Noncrystalline solids, including glass and plastics, are called amorphous solids. **An amorphous solid is one in which the particles are arranged randomly.** The two types of solids will be discussed in more detail later in this section.

Definite Shape and Volume

Unlike liquids and gases, solids can maintain a definite shape without a container. In addition, crystalline solids are geometrically regular. Even the fragments of a shattered crystalline solid have distinct geometric shapes that reflect their internal structure. Amorphous solids maintain a definite shape, but they do not have the distinct geometric shapes of crystalline solids. For example, glass can be molded into any shape. If glass is shattered, the fragments can have a variety of irregular shapes.

The volume of a solid changes only slightly with a change in temperature or pressure. Solids have definite volume because their particles are packed closely together. There is very little empty space into which the particles can be compressed. Crystalline solids generally do not flow, because their particles are held in relatively fixed positions.

Definite Melting Point

Melting is the physical change of a solid to a liquid by the addition of energy as heat. The temperature at which a solid becomes a liquid is its melting point. At this temperature, the kinetic energies of the particles within the solid overcome the attractive forces holding them together. The particles can then break out of their positions in crystalline solids, which have definite melting points. In contrast, amorphous solids, such as glass and plastics, have no definite melting point. They have the ability to flow over a range of temperatures. Therefore, amorphous solids are sometimes classified as **supercooled liquids, which are substances that retain certain liquid properties even at temperatures at which they appear to be solid.** These properties exist because the particles in amorphous solids are arranged randomly, much like the particles in a liquid. Unlike the particles in a true liquid, however, the particles in amorphous solids are not constantly changing their positions.

High Density and Incompressibility

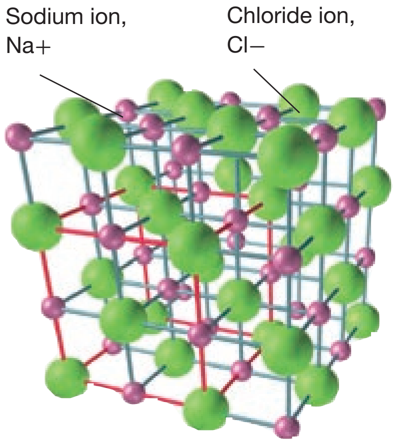
In general, substances are most dense in the solid state. Solids tend to be slightly denser than liquids and much denser than gases. The higher density results from the fact that the particles of a solid are more closely packed than those of a liquid or a gas. Solid hydrogen is the least dense solid; its density is about 1/320 that of the densest element, osmium, Os.

Solids are generally less compressible than liquids. For practical purposes, solids can be considered incompressible. Some solids, such as wood and cork, may *seem* compressible, but they are not. They contain pores that are filled with air. When subjected to intense pressure, the pores are compressed, not the solid matter in the wood or cork itself.

FIGURE 3.2

Crystal Lattice of Sodium Chloride

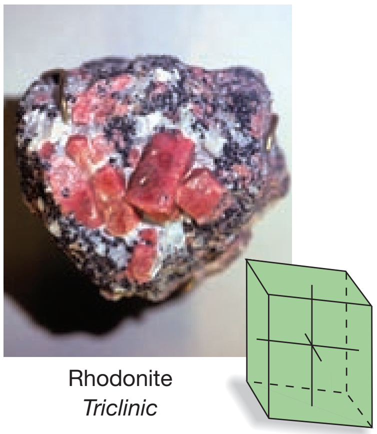
(a) This is a scanning electron micrograph (SEM) of sodium chloride crystals.



(b) The crystal structure of sodium chloride is made up of individual unit cells represented regularly in three dimensions. Here, one unit cell is outlined in red.

FIGURE 3.3

Crystal Systems in Minerals Shown are the seven basic crystal systems and representative crystals of each.



Low Rate of Diffusion

If a zinc plate and a copper plate are clamped together for a long time, a few atoms of each metal will diffuse into the other. This observation shows that diffusion does occur in solids. The rate of diffusion is millions of times faster in liquids than in solids, however.

MAIN IDEA

Crystal particles are arranged in a 3-dimensional lattice.

Crystalline solids exist either as single crystals or as groups of crystals fused together. **The total three-dimensional arrangement of particles of a crystal is called a crystal structure.** The arrangement of particles in the crystal can be represented by a coordinate system called a lattice. **The smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice is called a unit cell.** Each crystal lattice contains many unit cells packed together. **Figure 3.2** (on the previous page) shows the relationship between a crystal lattice and its unit cell. A crystal and its unit cells can have any one of seven types of symmetry. This fact enables scientists to classify crystals by their shape. Diagrams and examples of each type of crystal symmetry are shown in **Figure 3.3.**

FIGURE 3.4

MELTING AND BOILING POINTS OF REPRESENTATIVE CRYSTALLINE SOLIDS

Type of substance	Formula	Melting point (°C)	Boiling point at 1 atm (°C)
Ionic	NaCl MgF_2	801 1266	1413 2239
Covalent network	$(SiO_2)_x$ C_x (diamond)	1610 3500	2230 3930
Metallic	Hg Cu Fe W	−39 1083 1535 3410	357 2567 2750 5660
Covalent molecular (nonpolar)	H_2 O_2 CH_4 CCl_4 C_6H_6	−259 −218 −182 −23 −6	−253 −183 −164 77 80
Covalent molecular (polar)	NH_3 H_2O	−78 0	−33 100

Binding Forces in Crystals

Crystal structures can also be described in terms of the types of particles in the crystals and the types of chemical bonding between the particles. According to this method of classification, there are four types of crystals. These types are listed in **Figure 3.4.** Refer to this table as you read the following discussion.

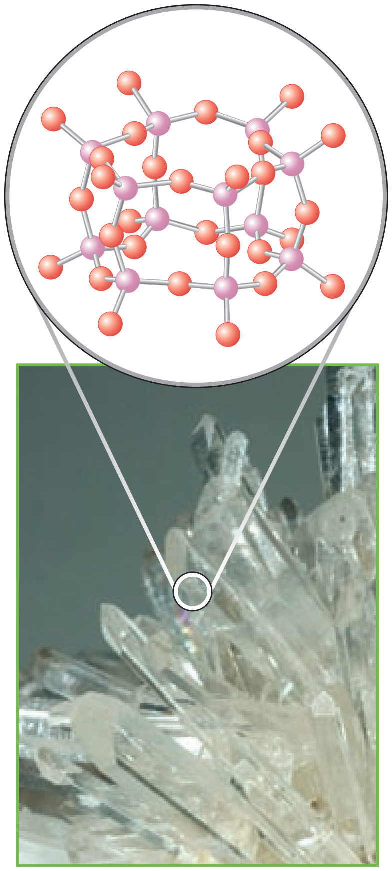
1. *Ionic crystals.* The ionic crystal structure consists of positive and negative ions arranged in a regular pattern. The ions can be monatomic or polyatomic. Generally, ionic crystals form when Group 1 or Group 2 metals combine with Group 16 or Group 17 nonmetals or nonmetallic polyatomic ions. The strong binding forces between the positive and negative ions in the crystal structure give the ionic crystals certain properties. For example, these crystals are hard and brittle, have high melting points, and are good insulators.
2. *Covalent network crystals.* In covalent network crystals, each atom is covalently bonded to its nearest neighboring atoms. The covalent bonding extends throughout a network that includes a very large number of atoms. Three-dimensional covalent network solids include diamond, C_x, quartz, $(SiO_2)_x$ (see **Figure 3.5** on the next page), silicon carbide, $(SiC)_x$, and many oxides of transition metals. Such solids are essentially giant molecules. The subscript x in these formulas indicates that the component within the parentheses extends indefinitely. The network solids are nearly always very hard and brittle. They have rather high melting points and are usually nonconductors or semiconductors.

3. *Metallic crystals.* The metallic crystal structure consists of metal cations surrounded by a sea of delocalized valence electrons. The electrons come from the metal atoms and belong to the crystal as a whole. This explains the high electrical conductivity of metals.
4. *Covalent molecular crystals.* The crystal structure of a covalent molecular substance consists of covalently bonded molecules held together by intermolecular forces. If the molecules are nonpolar—for example, hydrogen, H_2, methane, CH_4, and benzene, C_6H_6—then there are only weak London dispersion forces between molecules. In a polar covalent molecular crystal—for example, water, H_2O, and ammonia, NH_3—molecules are held together by dispersion forces, by somewhat stronger dipole-dipole forces, and sometimes by even stronger hydrogen bonding. The forces that hold polar or nonpolar molecules together in the structure are much weaker than the covalent chemical bonds between the atoms within each molecule. Covalent molecular crystals thus have low melting points. They are easily vaporized, are relatively soft, and are good insulators. Ice crystals, the most familiar molecular crystals, are discussed in Section 5.

FIGURE 3.5

Covalent Network Crystals

Covalent network crystals include three-dimensional network solids, such as this quartz, $(SiO_2)_x$, shown here with its three-dimensional atomic structure.



MAIN IDEA

The particles in amorphous solids are not arranged in a regular pattern.

The word *amorphous* comes from the Greek for "without shape." Unlike the atoms that form crystals, the atoms that make up amorphous solids, such as glasses and plastics, are not arranged in a regular pattern.

Glasses are made by cooling certain molten materials in a way that prevents them from crystallizing. The properties that result make glasses suitable for many uses, including windows, light bulbs, transformer cores, and optical fibers that carry telephone conversations.

Plastics, another type of amorphous solid, are easily molded at high temperatures and pressures. They are used in many structural materials.

Other, more recently created amorphous solids have been placed in many important applications. Amorphous semiconductors are used in electronic devices, including solar cells, copiers, laser printers, and flat-panel displays for computer monitors and television screens.

SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Describe the solid state according to the kinetic-molecular theory.
2. What is the difference between an amorphous solid and a crystalline solid?
3. Account for each of the following properties of solids: (a) the definite volume, (b) the relatively high density, (c) the extremely low rate of diffusion.
4. Compare and contrast the four types of crystals.
5. Why do crystalline solids shatter into regularly-shaped fragments when broken?

Critical Thinking

6. **RELATING IDEAS** Explain why ionic crystals melt at much higher temperatures than typical covalent molecular crystals.

SECTION 4

Main Ideas

- Substances in equilibrium change back and forth between states at equal speeds.
- A liquid boils when it has absorbed enough energy to evaporate.
- Freezing occurs when a substance loses enough heat energy to solidify.
- Under certain conditions, water can exist in all three phases at the same time.

VIRGINIA STANDARDS

CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include:
CH.5.c vapor pressure.
CH.5.d phase changes.
CH.5.e molar heats of fusion and vaporization.
CH.5.EKS-4; **CH.5.EKS-5**; **CH.5.EKS-6**; **CH.5.EKS-7**

Changes of State

Key Terms

phase
condensation
equilibrium
equilibrium vapor pressure
volatile liquid
boiling
boiling point
molar enthalpy of vaporization
freezing
freezing point
molar enthalpy of fusion
sublimation
deposition
phase diagram
triple point
critical point
critical temperature
critical pressure

Matter on Earth can exist in any of three states—gas, liquid, or solid—and can change from one state to another. **Figure 4.1** lists the possible changes of state. In this section, you will examine these changes of state and the factors that cause them.

MAIN IDEA

Substances in equilibrium change back and forth between states at equal speeds.

Some liquid chemical substances, such as rubbing alcohol, have an odor that is very easily detected. This is because some molecules at the upper surface of the liquid have enough energy to overcome the attraction of neighboring molecules, leave the liquid phase, and evaporate. **A phase is any part of a system that has uniform composition and properties.** In a closed bottle of rubbing alcohol, gas molecules under the cap strike the liquid surface and reenter the liquid phase through *condensation*. **Condensation is the process by which a gas changes to a liquid.** A gas in contact with its liquid or solid phase is often called a *vapor*.

If the temperature of the liquid remains constant and the cap remains closed, the rate at which molecules move from the liquid to vapor remains constant. Near the beginning of the evaporation process, very few molecules are in the gas phase, so the rate of condensation is very low.

FIGURE 4.1

POSSIBLE CHANGES OF STATE

Change of state	Process	Example
Solid $\longrightarrow$ liquid	melting	ice $\longrightarrow$ water
Solid $\longrightarrow$ gas	sublimation	dry ice $\longrightarrow$ CO_2 gas
Liquid $\longrightarrow$ solid	freezing	water $\longrightarrow$ ice
Liquid $\longrightarrow$ gas	vaporization	liquid bromine $\longrightarrow$ bromine vapor
Gas $\longrightarrow$ liquid	condensation	water vapor $\longrightarrow$ water
Gas $\longrightarrow$ solid	deposition	water vapor $\longrightarrow$ ice

FIGURE 4.2

Liquid-Vapor Equilibrium A liquid-vapor equilibrium develops in a closed system.

(a) At first there is only liquid present, but molecules are beginning to evaporate.

(b) Evaporation continues at a constant rate. Some vapor molecules are beginning to condense to liquid.

(c) Equilibrium has been reached between the rate of condensation and the rate of evaporation.

As more liquid evaporates, the increasing number of gas molecules causes the rate of condensation to increase, until the rate of condensation equals the rate of evaporation, and a state of equilibrium is established (see **Figure 4.2**). **Equilibrium is a dynamic condition in which two opposing changes occur at equal rates in a closed system.** Even though molecules are constantly moving between liquid and gas phases, there is no net change in the amount of substance in either phase.

Equilibrium Vapor Pressure of a Liquid

Vapor molecules in equilibrium with a liquid in a closed system exert a pressure proportional to the concentration of molecules in the vapor phase. **The pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature is called the equilibrium vapor pressure of the liquid.**

The increase in equilibrium vapor pressure with increasing temperature can be explained in terms of the kinetic-molecular theory for the liquid and gaseous states. Increasing the temperature of a liquid increases the average kinetic energy of the liquid's molecules. This increases the number of molecules that have enough energy to escape from the liquid phase into the vapor phase. The resulting increased evaporation rate increases the number of molecules in the vapor phase, which in turn increases the equilibrium vapor pressure.

Every liquid has a specific equilibrium vapor pressure at a given temperature. The stronger these attractive forces are, the smaller is the percentage of liquid particles that can evaporate at any given temperature. A low percentage of evaporation results in a low equilibrium vapor pressure. **Volatile liquids,** **which are liquids that evaporate readily,** have **relatively weak forces of attraction between their particles.**

Ether is a typical volatile liquid. Nonvolatile liquids, such as molten ionic compounds, do not evaporate readily and have relatively strong attractive forces between their particles.

MAIN IDEA

A liquid boils when it has absorbed enough energy to evaporate.

Equilibrium vapor pressures can be used to explain and define the concept of boiling, which is **the conversion of a liquid to a vapor within the liquid as well as at its surface.**

If the temperature of the liquid is increased, the equilibrium vapor pressure also increases. **The boiling point of a liquid is the temperature at which the equilibrium vapor pressure of the liquid equals the atmospheric pressure.** The lower the atmospheric pressure is, the lower the boiling point is.

At boiling point, all of the energy absorbed is used to evaporate the liquid, and the temperature remains constant as long as the pressure does not change. If the pressure above the liquid being heated is increased, the temperature of the liquid will rise until the vapor pressure equals the new pressure and the liquid boils once again. A pressure cooker is sealed so that steam pressure builds up over the surface of the boiling water inside. The boiling temperature of the water increases, resulting in shorter cooking times. A device called a *vacuum evaporator* causes boiling at lower-than-normal temperatures. Vacuum evaporators used to prepare evaporated and sweetened condensed milk remove water from milk and sugar solutions. Under reduced pressure, the water boils away at a temperature low enough to avoid scorching the milk or sugar.

At normal atmospheric pressure (1 atm, 760 torr, or 101.3 kPa), water boils at exactly 100°C. This is the *normal* boiling point of water.

FIGURE 4.3

Boiling Point The vapor pressure of any liquid increases as its temperature increases. A liquid boils when its vapor pressure equals the pressure of the atmosphere.

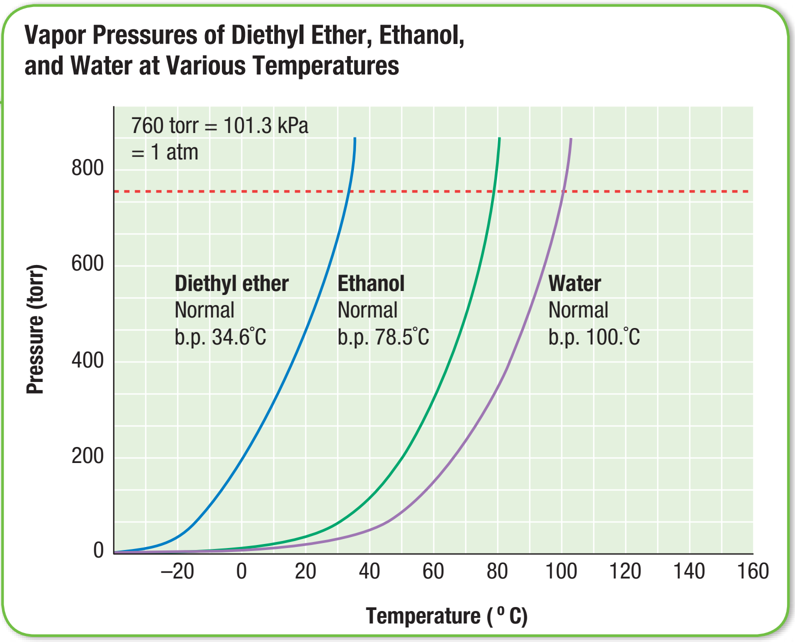


FIGURE 4.4

Energy Distribution The number of molecules in a liquid with various kinetic energies is represented at two different temperatures.

CRITICAL THINKING

Analyze Look at the graph, and describe what is happening in the shaded area at the lower right.

Energy Distribution of Molecules in a Liquid at Different Temperatures

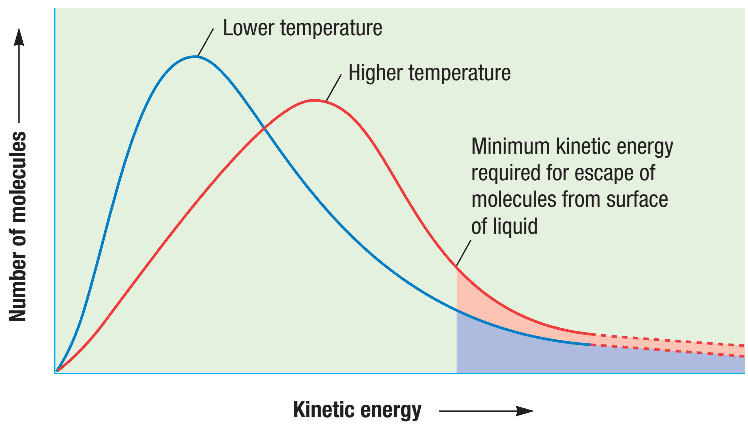


Figure 4.3 (see previous page) shows that the normal boiling point of each liquid occurs when its equilibrium vapor pressure equals 760 torr. Energy must be added continuously in order to keep a liquid boiling. The temperature of a liquid and its vapor at the boiling point remains constant despite the continuous addition of energy. The added energy is used to overcome the attractive forces between molecules of the liquid during the liquid-to-gas change and is stored in the vapor as potential energy.

Molar Enthalpy of Vaporization

The amount of energy as heat that is needed to vaporize one mole of liquid at the liquid's boiling point at constant pressure is called the liquid's molar enthalpy of vaporization, ΔH_v. Molar enthalpy of vaporization measures the attraction between particles of the liquid. The stronger the attraction is, the more energy is required to overcome it, resulting in a higher molar enthalpy of vaporization. Each liquid has a characteristic molar enthalpy of vaporization. Liquid water's molar enthalpy of vaporization is very high, due to its extensive hydrogen bonding, This makes water an effective cooling agent. When water evaporates, the escaping molecules carry away with them a great deal of energy as heat. **Figure 4.4** shows the distribution of the kinetic energies of molecules in a liquid at two different temperatures. At higher temperatures, a greater portion of the surface molecules have the kinetic energy required to escape and become vapor.

MAIN IDEA

Freezing occurs when a substance loses enough heat energy to solidify.

The physical change of a liquid to a solid is called freezing. Freezing involves a loss of energy in the form of heat by the liquid.

$$\text{liquid} \longrightarrow \text{solid} + \text{energy}$$

In the case of a pure crystalline substance, this change occurs at constant temperature. **The normal freezing point is the temperature at which the solid and liquid are in equilibrium at 1 atm (760 torr, or 101.3 kPa) pressure.** At the freezing point, particles of the liquid and the solid have the same average kinetic energy, and the energy lost during freezing is the potential energy that was present in the liquid. At the same time energy decreases, there is a significant increase in particle order, because the solid state of a substance is much more ordered than the liquid state, even at the same temperature.

Melting, the reverse of freezing, also occurs at constant temperature. As a solid melts, it continuously absorbs energy as heat, as represented by the following equation.

$$\text{solid} + \text{energy} \longrightarrow \text{liquid}$$

For pure crystalline solids, the melting point and freezing point are the same. At equilibrium, melting and freezing proceed at equal rates. The following general equilibrium equation can be used to represent these states.

$$\text{solid} + \text{energy} \rightleftarrows \text{liquid}$$

At normal atmospheric pressure, the temperature of a system containing ice and liquid water will remain at 0.°C as long as both ice and water are present, no matter what the surrounding temperature. Adding energy in the form of heat to such a system shifts the equilibrium to the right. That shift increases the proportion of liquid water and decreases that of ice. Only after all the ice has melted will the addition of energy increase the temperature of the system.

WHY IT MATTERS

Surface Melting S.T.E.M.

Freezing of water and melting of ice are phase changes that are familiar to all of us. Yet physicists and chemists have only recently begun to understand the basic aspects of these phase changes, through experimental and theoretical studies of a phenomenon known as surface melting. Experimental studies in the mid-1980s confirmed that the rigid surface arrangements of metals can become increasingly disordered several degrees below the melting point of the metal, forming a "quasi-liquid layer." Many different techniques have now shown that ice also has such a fluid surface layer just a few molecules thick. This surface melting of ice might explain observations as diverse as the origin of lightning, the unique shapes of snowflakes, and ice skating.

Molar Enthalpy of Fusion

The amount of energy as heat required to melt one mole of solid at the solid's melting point is the solid's molar enthalpy of fusion, ΔH_f. The energy absorbed increases the solid's potential energy as its particles are pulled apart, overcoming the attractive forces holding them together. At the same time, there is a significant decrease in particle order as the substance makes the transformation from solid to liquid. Similar to the molar enthalpy of vaporization, the magnitude of the molar enthalpy of fusion depends on the attraction between the solid particles.

Sublimation and Deposition

At sufficiently low temperature and pressure conditions, a liquid cannot exist. Under such conditions, a solid substance exists in equilibrium with its vapor instead of its liquid, as represented by the following equation.

$$\text{solid} + \text{energy} \rightleftarrows \text{vapor}$$

The change of state from a solid directly to a gas is known as sublimation. The reverse process is called **deposition, the change of state from a gas directly to a solid.** Among the common substances that sublime at ordinary temperatures are dry ice (solid CO_2) and iodine.

Ordinary ice sublimes slowly at temperatures lower than its melting point (0.°C). This explains how a thin layer of snow can eventually disappear, even if the temperature remains below 0.°C. Sublimation occurs in frost-free refrigerators when the temperature in the freezer compartment is periodically raised to cause any ice that has formed to sublime. A blower then removes the water vapor that has formed. The formation of frost on a cold surface is a familiar example of deposition.

MAIN IDEA

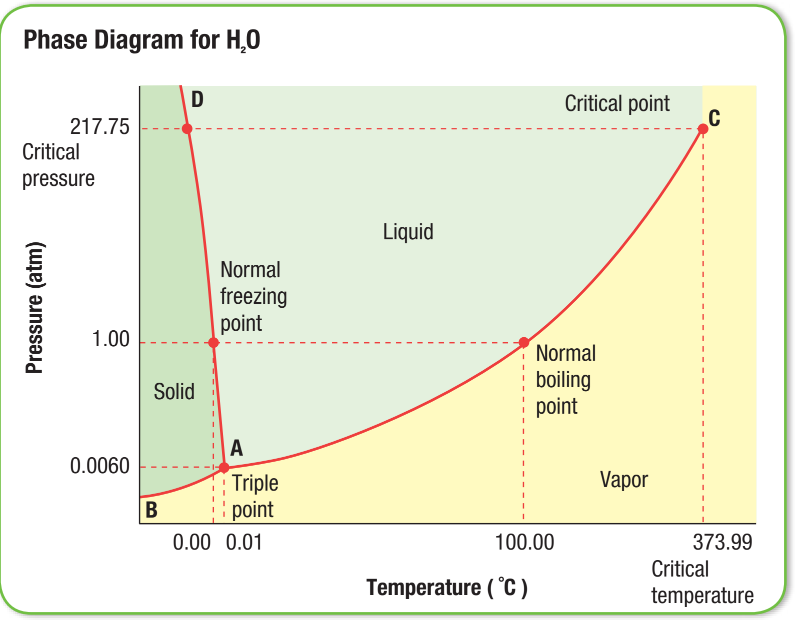
Under certain conditions, water can exist in all three phases at the same time.

A phase diagram is a graph of pressure versus temperature that shows the conditions under which the phases of a substance exist. A phase diagram also reveals how the states of a system change with changing temperature or pressure.

Figure 4.5 shows the phase diagram for water over a range of temperatures and pressures. Note the three curves, AB, AC, and AD. Curve AB indicates the temperature and pressure conditions at which ice and water vapor can coexist at equilibrium. Curve AC indicates the temperature and pressure conditions at which liquid water and water vapor coexist at equilibrium. Similarly, curve AD indicates the temperature and pressure conditions at which ice and liquid water coexist at equilibrium. Because ice is less dense than liquid water, an increase in pressure lowers the melting point. (Most substances have a positive slope for this curve.) Point A is the triple point of water. **The triple point of a substance indicates the temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium.** Point C is the critical point of water. **The critical point of a substance indicates the critical temperature and critical pressure. The critical temperature (t_c) is the temperature above which the substance cannot exist in the liquid state.**

FIGURE 4.5

Phase Diagram This phase diagram shows the relationships between the physical states of water and its pressure and temperature.



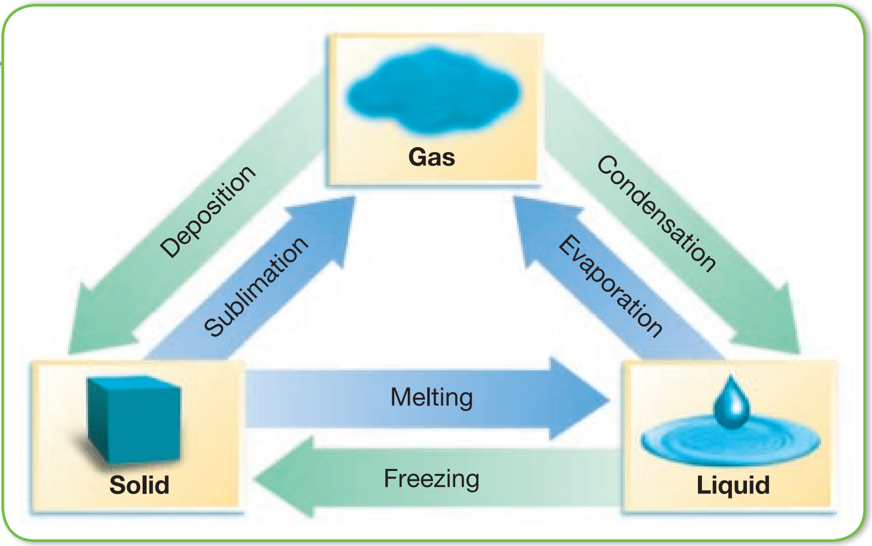


FIGURE 4.6

Changes of State Solids, liquids, and gases can undergo various changes of state. The changes shown in green are exothermic, and those shown in blue are endothermic.

The critical temperature of water is 373.99°C. Above this temperature, water cannot be liquefied, no matter how much pressure is applied. **The critical pressure (P_c) is the lowest pressure at which the substance can exist as a liquid at the critical temperature.** The critical pressure of water is 217.75 atm.

The phase diagram in **Figure 4.5** on the previous page indicates the normal boiling point and the normal freezing point of water. It also shows how boiling point and freezing point change with pressure. As shown by the slope of line AD, ice melts at a higher temperature with decreasing pressure. Below the triple point, the temperature of sublimation decreases with decreasing pressure. **Figure 4.6** summarizes the changes of state of solids, liquids, and gases.

SECTION 4 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. What is equilibrium?

2. What happens when a liquid-vapor system at equilibrium experiences an increase in temperature? What happens when it experiences a decrease in temperature?

3. What would be an example of deposition?

4. What is the equilibrium vapor pressure of a liquid? How is it measured?

5. What is the boiling point of a liquid?

6. In the phase diagram for water, what is meant by the triple point and the critical point?

Critical Thinking

7. INTERPRETING GRAPHICS Refer to the phase diagram for water (**Figure 4.5**) to answer the following questions.

a. Describe all the changes a sample of solid water would undergo when heated from −10°C to its critical temperature at a pressure of 1.00 atm.

b. Describe all the changes a sample of water vapor would undergo when cooled from 110°C to 5°C at a pressure of 1.00 atm.

c. At approximately what pressure will water be a vapor at 0°C?

d. Within what range of pressures will water be a liquid at temperatures above its normal boiling point?

Water

SECTION 5

Main Ideas

- The properties of water in all phases are determined by its structure.
- The molar enthalpy of water determines many of its physical characteristics.

VIRGINIA STANDARDS

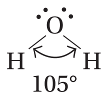
CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles.
CH.5.EKS-3; **CH.5.EKS-7**

Water is a familiar substance in all three physical states: solid, liquid, and gas. On Earth, water is by far the most abundant liquid. Oceans, rivers, and lakes cover about 75% of Earth's surface. Significant quantities of water are also frozen in glaciers. Water is an essential component of all organisms; 70% to 90% of the mass of living things is water. The chemical reactions of most life processes take place in water, and water is frequently a reactant or product in such reactions. In order to better understand the importance of water, let us take a closer look at its structure and its properties.

MAIN IDEA

The properties of water in all phases are determined by its structure.

Water molecules consist of two atoms of hydrogen and one atom of oxygen united by polar-covalent bonds. Research shows that a water molecule is bent. The structure can be represented as follows.



The angle between the two hydrogen-oxygen bonds is about 105°. This is close to the angle expected for sp^3 hybridization of the oxygen-atom orbitals.

The molecules in solid or liquid water are linked by hydrogen bonding. The number of linked molecules decreases with increasing temperature, because increases in kinetic energy make hydrogen bond formation difficult. Nevertheless, there are usually from four to eight molecules per group in liquid water, as shown in **Figure 5.1.** If it were not for these molecular groups, water would be a gas at room temperature. Nonpolar molecules, such as methane, CH_4, that are similar in size and mass to water molecules do not undergo hydrogen bonding. Such substances are gases at room temperature.

Ice consists of water molecules in the hexagonal arrangement shown in **Figure 5.2** on the next page. The empty spaces between molecules in this pattern account for the relatively low density of ice. As ice is heated, the increased energy of the molecules causes them to move and vibrate more vigorously. When the melting point is reached, the energy of the molecules is so great that the rigid open structure of the ice crystals breaks down, and ice turns into liquid water.

Figures 5.1 and **Figure 5.2** also show that the hydrogen bonds between molecules of liquid water at 0°C are fewer and more disordered than those between molecules of ice at the same temperature. Because the rigid open structure of ice has broken down, water molecules can crowd closer together. Thus, liquid water is denser than ice.

FIGURE 5.1

Liquid Water Within the water molecule, oxygen and hydrogen are covalently bonded to each other. Hydrogen bonds hold the molecules together in groups.

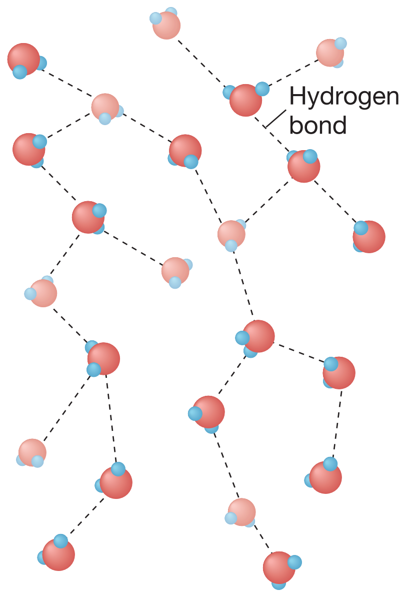
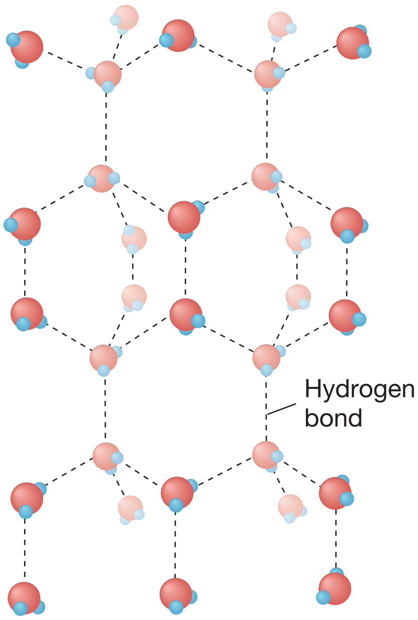


FIGURE 5.2

Solid Water Ice contains the same types of bonding as liquid water. However, the structure of the hydrogen bonding is much more rigid and open than it is in liquid water.



As the liquid water is warmed from 0°C, the water molecules crowd still closer together. Water molecules are as tightly packed as possible at 3.98°C. At temperatures above 3.98°C, the increasing kinetic energy of the water molecules causes them to overcome molecular attractions. The molecules move farther apart as the temperature continues to rise. As the temperature approaches the boiling point, groups of liquid water molecules absorb enough energy to break up into separate molecules. Because of hydrogen bonding between water molecules, a high kinetic energy is needed, causing water's boiling point to be relatively high (100°C) compared to other liquids that have similar molar masses.

MAIN IDEA

The molar enthalpy of water determines many of its physical characteristics.

At room temperature, pure liquid water is transparent, odorless, tasteless, and almost colorless. Any observable odor or taste is caused by impurities, such as dissolved minerals, liquids, or gases.

As shown by its phase diagram in **Figure 4.5** (in Section 4), water freezes and ice melts at 0°C at a pressure of 1 atm (101.3 kPa). The molar enthalpy of fusion of ice is 6.009 kJ/mol. That value is relatively large compared with the molar enthalpy of fusion of other solids. As you have read, water has the unusual property of expanding in volume as it freezes, because its molecules form an open rigid structure. As a result, ice at 0°C has a density of only about 0.917 g/cm^3, but liquid water at 0°C has a density of 0.999 84 g/cm^3.

This lower density explains why ice floats in liquid water. The insulating effect of floating ice is particularly important in the case of large bodies of water. If ice were more dense than liquid water, it would sink to the bottom of lakes and ponds, where it would be less likely to melt completely. The water of such bodies of water in temperate climates would eventually freeze solid, killing nearly all the living things in it.

Under a pressure of 1 atm (101.3 kPa), water boils at 100°C. At this temperature, water's molar enthalpy of vaporization is 40.79 kJ/mol. Both the boiling point and the molar enthalpy of vaporization of water are quite high compared with those of nonpolar substances of comparable molecular mass, such as methane. The values are high because of the strong hydrogen bonding that must be overcome for boiling to occur. The high molar enthalpy of vaporization makes water useful for household steam-heating systems. The steam (vaporized water) stores a great deal of energy as heat. When the steam condenses in radiators, great quantities of energy are released. In living organisms, the high molar enthalpy of vaporization helps them to resist dehydration when sources of water intake are restricted and provides a large degree of evaporative cooling to carry excess heat away from their bodies. This property especially allows warm-blooded organisms to better regulate their internal temperatures, reduce cell death from water loss, and maintain homeostasis.

PREMIUM CONTENT

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Using Molar Enthalpy of Vaporization

Sample Problem A **How much energy is absorbed when 47.0 g of ice melts at STP? How much energy is absorbed when this same mass of liquid water boils?**

1 ANALYZE

Given: mass of $H_2O(s)$ = 47.0 g
mass of $H_2O(l)$ = 47.0 g
molar enthalpy of fusion of ice = 6.009 kJ/mol
molar enthalpy of vaporization = 40.79 kJ/mol

Unknown: energy absorbed when ice melts; energy absorbed when liquid water boils

2 PLAN

First, convert the mass of water from grams to moles.

$$47.0\ \cancel{g}\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g}\ H_2O} = 2.61\ mol\ H_2O$$

Then, use the molar enthalpy of fusion of a solid to calculate the amount of energy absorbed when the solid melts. Multiply the number of moles by the amount of energy needed to melt one mole of ice at its melting point (the molar enthalpy of fusion of ice). Using the same method, calculate the amount of energy absorbed when water boils by using the molar enthalpy of vaporization.

amount of substance (mol) × molar enthalpy of fusion or vaporization (kJ/mol) = energy (kJ)

3 SOLVE

$$2.61\ \cancel{mol} \times 6.009\ kJ/\cancel{mol} = 15.7\ kJ\ \text{(on melting)}$$
$$2.61\ \cancel{mol} \times 40.79\ kJ/\cancel{mol} = 106\ kJ\ \text{(on vaporizing or boiling)}$$

4 CHECK YOUR WORK

Units have canceled correctly. The answers have the proper number of significant digits and are reasonably close to estimated values of 18 (3 × 6) and 120 (3 × 40), respectively.

Practice

Answers in Appendix E

1. What quantity of energy is released when 506 g of liquid water freezes?
2. What mass of steam is required to release 4.97×10^5 kJ of energy on condensation?

SECTION 5 FORMATIVE ASSESSMENT

Reviewing Main Ideas

1. Why is a water molecule polar?
2. How is the structure of water responsible for some of water's unique characteristics?
3. Describe the arrangement of molecules in liquid water and in ice.
4. Why does ice float? Why is this phenomenon important?
5. Why is ice less dense than liquid water?
6. Is more energy required to melt one gram of ice at 0°C or to boil one gram of water at 100°C? How do you know?

Critical Thinking

7. **RELATING IDEAS** Why is exposure to steam dangerous?

Math Tutor

Calculations Using Enthalpies of Fusion

When one mole of a liquid freezes to a solid, energy is released as attractive forces between particles pull the disordered particles of the liquid into a more orderly crystalline solid. When the solid melts to a liquid, the solid must absorb the same quantity of energy in order to separate the particles of the crystal and overcome the attractive forces opposing separation. This quantity of energy used to melt or freeze one mole of a substance at its melting point is called its molar enthalpy of fusion, ΔH_f.

Problem-Solving TIPS

- The enthalpy of fusion of a substance can be given as either joules per gram or kilojoules per mole.
- *Molar* enthalpy of fusion is most commonly used in calculations.
- The enthalpy of fusion is the energy absorbed or given off as heat when a substance melts or freezes at the melting point of the substance.
- There is no net change in temperature as the change of state occurs.

Samples

7.30 kJ of energy is required to melt 0.650 mol of ethylene glycol ($C_2H_6O_2$) at its melting point. Calculate the molar enthalpy of fusion, ΔH_f, of ethylene glycol and the energy absorbed.

$$\text{molar enthalpy of fusion} = \Delta H_f = \frac{\text{energy absorbed}}{\text{moles of substance}}$$

$$\Delta H_{f,\ ethylene\ glycol} = \frac{7.30\ \text{kJ}}{0.065\ \text{mol}} = 11.2\,\frac{\text{kJ}}{\text{mol}}$$

Determine the quantity of energy that will be needed to melt 2.50×10^5 kg of iron at its melting point, 1536°C. The ΔH_f of iron is 13.807 kJ/mol.

To calculate the number of moles of iron, use the equation below.

$$\text{moles of substance} = \frac{\text{mass of substance}}{\text{molar mass of substance}}$$

Next, use the following equation for energy as heat absorbed.

$$\text{energy absorbed} = \Delta H_f \times \text{moles of substance}$$

Now, substitute the calculation for moles of substance, and solve.

$$\text{energy absorbed} = \Delta H_f \times \frac{\text{grams of substance}}{\text{molar mass of substance}}$$

$$= 13.807\,\frac{\text{kJ}}{\text{mol}} \times \frac{2.50 \times 10^8\ \text{g Fe}}{55.847\ \text{g Fe/mol Fe}}$$

$$\text{energy absorbed} = 6.18 \times 10^7\ \text{kJ}$$

Practice

Answers in Appendix E

1. Calculate the molar enthalpy of fusion of silver if 1.940 mol of silver requires 22.60 kJ of energy to change from a solid to a liquid at its melting point, 961°C.
2. What quantity of energy in kJ must be absorbed by 6.47 mol of solid acetic acid, $C_2H_4O_2$, to melt it at its melting point, 16.7°C? The ΔH_f for acetic acid is 11.54 kJ/mol.

CHAPTER 10 Summary

PREMIUM CONTENT
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Concept Maps

SECTION 1 The Kinetic-Molecular Theory of Matter

- The kinetic-molecular theory of matter can be used to explain the properties of gases, liquids, and solids.
- The kinetic-molecular theory of gases describes a model of an ideal gas.
- Gases consist of large numbers of tiny, fast-moving particles that are far apart relative to their size.

KEY TERMS

kinetic-molecular theory
ideal gas
elastic collision
diffusion
effusion
real gas

SECTION 2 Liquids

- The particles of a liquid are closer together and more ordered than those of a gas and are less ordered than those of a solid.
- Liquids have a definite volume and a fairly high density, and they are relatively incompressible. Like gases, liquids can flow and thus are considered to be fluids.

KEY TERMS

fluid
surface tension
capillary action
vaporization
evaporation
freezing

SECTION 3 Solids

- The particles of a solid are not nearly as free to move about as those of a liquid or a gas are.
- Solids have a definite shape and may be crystalline or amorphous. They have a definite volume and are generally nonfluid.
- A crystal structure is the total three-dimensional array of points that describes the arrangement of the particles of a crystal.
- Unlike crystalline solids, amorphous solids do not have a highly ordered structure or a regular shape.

KEY TERMS

crystalline solid
crystal
amorphous solid
melting
melting point
supercooled liquid
crystal structure
unit cell

SECTION 4 Changes of State

- A liquid in a closed system will gradually reach a liquid-vapor equilibrium as the rate at which molecules condense equals the rate at which they evaporate.
- When two opposing changes occur at equal rates in the same closed system, the system is said to be in dynamic equilibrium.

KEY TERMS

phase
condensation
equilibrium
equilibrium vapor pressure
volatile liquid
boiling
boiling point
molar enthalpy of vaporization
freezing
freezing point
molar enthalpy of fusion
sublimation
deposition
phase diagram
triple point
critical point
critical temperature
critical pressure

SECTION 5 Water

- Water is a polar covalent compound.
- The structure and the hydrogen bonding in water are responsible for its relatively high melting point, molar enthalpy of fusion, boiling point, and molar enthalpy of vaporization.

CHAPTER 10 Review

SECTION 1

The Kinetic-Molecular Theory of Matter

REVIEWING MAIN IDEAS

1. What idea is the kinetic-molecular theory based on?
2. What is an ideal gas?
3. State the five basic assumptions of the kinetic-molecular theory.
4. How do gases compare with liquids and solids in terms of the distance between their molecules?
5. What is the relationship between the temperature, speed, and kinetic energy of gas molecules?
6. **a.** What is diffusion?
   b. What factors affect the rate of diffusion of one gas through another?

SECTION 2

Liquids

REVIEWING MAIN IDEAS

7. What is a fluid?
8. What is surface tension?
9. Give two reasons why evaporation is a crucial process in nature.

SECTION 3

Solids

REVIEWING MAIN IDEAS

10. List six properties of solids, and explain each in terms of the kinetic-molecular theory of solids.
11. List four common examples of amorphous solids.
12. List and describe the four types of crystals in terms of the nature of their component particles and the type of bonding between them.

SECTION 4

Changes of State

REVIEWING MAIN IDEAS

13. Using **Figure 4.3,** estimate the approximate equilibrium vapor pressure of each of the following at the specified temperature.
    a. water at 80°C
    b. diethyl ether at 20°C
    c. ethanol at 60°C
14. **a.** What is sublimation?
    b. Give two examples of common substances that sublime at ordinary temperatures.
15. What is meant by the normal freezing point of a substance?
16. Explain how the attractive forces between the particles in a liquid are related to the equilibrium vapor pressure of that liquid.
17. Explain the relationship between atmospheric pressure and the actual boiling point of a liquid.
18. Explain the relationship between the molar enthalpy of fusion of a solid and the strength of attraction between that solid's particles.

PRACTICE PROBLEMS

19. **a.** The molar enthalpy of vaporization for water is 40.79 kJ/mol. Express this enthalpy of vaporization in joules per gram.
    b. The molar enthalpy of fusion for water is 6.009 kJ/mol. Express this enthalpy of fusion in joules per gram.
20. Calculate the molar enthalpy of vaporization of a substance, given that 0.433 mol of the substance absorbs 36.5 kJ of energy when it is vaporized.
21. Given that a substance has a molar mass of 259.0 g/mol and a 71.8 g sample of the substance absorbs 4.307 kJ when it melts,
    a. calculate the number of moles in the sample.
    b. calculate the molar enthalpy of fusion.

22. **a.** Calculate the number of moles in a liquid sample of a substance that has a molar enthalpy of fusion of 3.811 kJ/mol, given that the sample releases 83.2 kJ when it freezes.
b. Calculate the molar mass of this substance if the mass of the sample is 5519 g.

SECTION 5

Water

REVIEWING MAIN IDEAS

23. Describe the structure of a water molecule.

24. List at least eight physical properties of water.

PRACTICE PROBLEMS

25. Which contains more molecules of water: 5.00 cm^3 of ice at 0°C or 5.00 cm^3 of liquid water at 0.°C? How many more? What is the ratio of the numbers of molecules in these two samples?

26. **a.** What volume and mass of steam at 100.°C and 1.00 atm would release the same amount of energy during condensation as 100. cm^3 of liquid water would release during freezing?
b. What do you note, qualitatively, about the relative volumes and masses of steam and liquid water required to release the same amount of heat? (Hint: See Sample Problem A.)

Mixed Review

REVIEWING MAIN IDEAS

27. Find the molar enthalpy of vaporization for a substance, given that 3.21 mol of the substance absorbs 28.4 kJ of energy as heat when the substance changes from a liquid to a gas.

28. Water's molar enthalpy of fusion is 6.009 kJ/mol. Calculate the amount of energy as heat required to melt 7.95×10^5 g of ice.

29. A certain substance has a molar enthalpy of vaporization of 31.6 kJ/mol. How much of the substance is in a sample that requires 57.0 kJ to vaporize?

30. Given that water has a molar enthalpy of vaporization of 40.79 kJ/mol, how many grams of water could be vaporized by 0.545 kJ?

31. Calculate the amount of energy released as heat by the freezing of 13.3 g of a liquid substance, given that the substance has a molar mass of 82.9 g/mol and a molar enthalpy of fusion of 4.60 kJ/mol.

32. What volume and mass of steam at 100.°C and 760. torr would release the same amount of energy as heat during condensation as 65.5 cm^3 of liquid water would release during freezing?

33. The following liquid-vapor system is at equilibrium at a given temperature in a closed system.

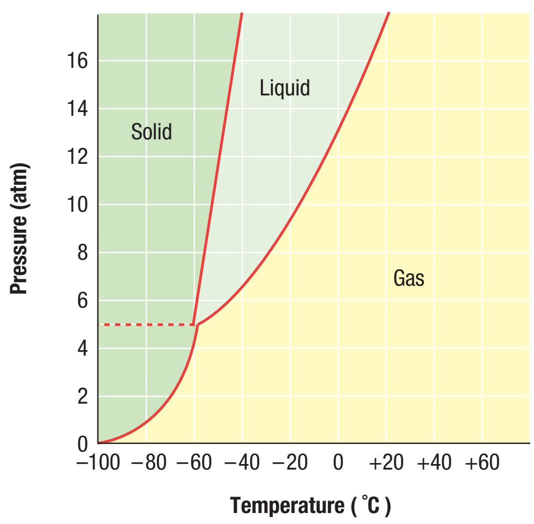
$$\text{liquid} + \text{energy} \rightleftharpoons \text{vapor}$$

Suppose the temperature is increased, and equilibrium is established at the higher temperature. How does the final value of each of the following compare with its initial value? (In each case, answer either higher, lower, or the same.)
a. the rate of evaporation
b. the rate of condensation
c. the final concentration of vapor molecules
d. the final number of liquid molecules

34. Given a sample of water at any point on curve AB in **Figure 4.5,** what effect would each of the following changes have on that sample?
a. adding energy at constant pressure
b. decreasing the volume at constant temperature
c. removing energy at constant pressure
d. increasing the volume at constant temperature

35. Using the phase diagram for CO_2, describe all the phase changes that would occur when CO_2 is heated from −100°C to −10°C at a constant pressure of 6 atm.

CO_2 Phase Diagram



CRITICAL THINKING

36. Interpreting Concepts During the freezing of a substance, energy is being removed from that substance. Yet the temperature of the liquid-solid system remains constant. Explain this phenomenon.

37. Applying Models At normal atmospheric pressure, the temperature of an ice-water system remains at 0°C as long as both ice and liquid water are present, regardless of the surrounding temperature. Explain why this occurs.

38. Predicting Outcomes Given a sample of water at any point on curve AD in **Figure 4.5**, how could more of the liquid water in that sample be converted into a solid without changing the temperature? Explain your reasoning.

39. Interpreting Diagrams Refer to the phase diagram in question 35.

a. Explain what happens when solid CO_2 ("dry ice") warms up to room temperature at normal atmospheric pressure.

b. Is there a pressure below which liquid CO_2 cannot exist? Estimate that pressure from the graph.

USING THE HANDBOOK

40. The *Elements Handbook* (Appendix A) contains a table of properties for each group that includes information on the crystal structures of the elements. Most metals crystallize in one of three lattice arrangements: body-centered cubic, face-centered cubic, or hexagonal close-packed. **Figure 3.2** shows a model of the face-centered cubic lattice for sodium chloride. Use this figure and the information in the *Elements Handbook* (Appendix A) to answer the following.

a. What elements in Group 2 have the same lattice structure as sodium chloride?

b. How would the model of an element in a face-centered cubic lattice differ from the compound shown in **Figure 3.2**?

c. The body-centered cubic lattice is the least-efficient packing structure of the metals. What elements in Groups 1 and 2 show this arrangement?

RESEARCH AND WRITING

41. Ceramics are formed from silicates found in the soil. Artists use them to create pottery, but engineers and scientists have created ceramics with superconductive properties. Investigate the growing field of superconductive ceramics.

42. Liquid crystals are substances that possess the combined properties of both liquids and crystals. Write a report on these substances and the various uses we are finding for them.

ALTERNATIVE ASSESSMENT

43. Compile separate lists of crystalline and amorphous solids found in your home. Compare your lists with those of your classmates.

44. Design an experiment to grow crystals of various safe, common household materials. Record the conditions under which each type of crystal is best grown.

Standards-Based Assessment

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Surface tension is
A. skin on the surface of a liquid.
B. the tendency of the surface of liquids to decrease the area.
C. the spontaneous mixing of two substances.
D. the same as vapor pressure.

2. Pure liquids boil at higher temperatures under high pressures than they do under low pressures, because
A. the molecules of liquid are closer together under higher pressures.
B. it takes a higher temperature for the vapor pressure to equal the higher external pressure.
C. the molecules of vapor are farther apart under higher pressures.
D. the vapor diffuses more rapidly at higher pressures.

3. The formation of frost is an example of
A. condensation.
B. evaporation.
C. deposition.
D. melting point.

4. A graph that shows the pressure and temperature conditions under which the phases of a substance exist is called
A. a phase diagram.
B. a vapor pressure curve.
C. a unit cell.
D. the kinetic-molecular theory of matter.

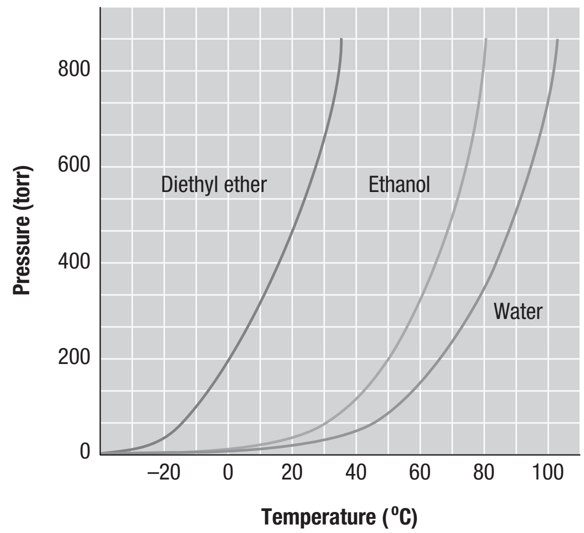
5. Water boils at 100°C. Ethanol boils at 78.5°C. Which of the following statements is true?
A. Water has the higher vapor pressure at 78.5°C.
B. Ethanol has the higher vapor pressure at 78.5°C.
C. Both have the same vapor pressure at 78.5°C.
D. Vapor pressure is not related to boiling point.

6. Which of the following is not a property of typical solids?
A. definite melting point
B. high density
C. easily compressible
D. low rate of diffusion

7. The kinetic-molecular theory states that ideal gas molecules
A. are in constant, rapid, random motion.
B. have mass and take up space.
C. exert forces of attraction and repulsion on each other.
D. have high densities compared with liquids and solids.

SHORT ANSWER

8. Using this graph of vapor pressures of substances at various temperatures, estimate the boiling point of ethanol at an applied (external) pressure of 300 torr.

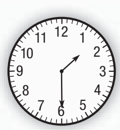


9. It is found that 60.0 J of energy are required to melt 15 g of a substance. The molar mass of the substance is 120 g/mol. Calculate the enthalpy of fusion of the substance in kilojoules per mole.

EXTENDED RESPONSE

10. Describe how a pressure cooker works.

11. What is meant by the statement that a liquid and its vapor in a closed container are in a state of dynamic equilibrium?



Test Tip
Test questions are not necessarily arranged in order of increasing difficulty. If you are unable to answer a question, mark it and move on to other questions.