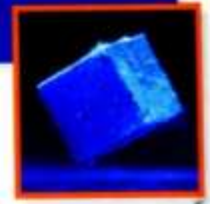
section 3.5

objective
- Convert between the Celsius and Kelvin temperature scales

key terms
- temperature
- Celsius scale
- Kelvin scale
- absolute zero

TEMPERATURE



In 1911, a Dutch physicist named Heike Kamerlingh Onnes discovered that when mercury is cooled to near absolute zero (−460 °F), it loses all resistance to electrical current. This state is known as superconductivity. If a material could be found that would superconduct at room temperature (around 64 °F), it would revolutionize the computer industry and other electronics industries. **Which temperature scale has its zero point at absolute zero?**

Measuring Temperature

When you hold a glass of hot water, the glass feels hot because heat transfers from the glass to your hand. When you hold an ice cube, it feels cold because heat transfers from your hand to the ice cube. The **temperature** of an object determines the direction of heat transfer. When two objects at different temperatures are in contact, heat moves from the object at the higher temperature to the object at the lower temperature. In Chapter 10, you will learn how the temperature of an object is related to the energy and motion of particles.

Almost all substances expand with an increase in temperature and contract as the temperature decreases (a very important exception is water). These properties are the basis for the common mercury-in-glass thermometer, shown in **Figure 3.19.** The liquid mercury in the thermometer expands and contracts more than the volume of the glass bulb that holds it, producing changes in the column height of the mercury.

Temperature Scales

Several temperature scales have been devised. The Celsius scale of the metric system is named after the Swedish astronomer Anders Celsius (1701–1744). It uses two readily determined temperatures as reference temperature values: the freezing point and the boiling point of water. The **Celsius scale** sets the freezing point of water at 0 °C and the boiling point of water at 100 °C. The distance between these two fixed points is divided into 100 equal intervals, or degrees Celsius (°C).

Another temperature scale often used in the physical sciences is the Kelvin scale, or absolute scale. This scale is named after Lord Kelvin (1824–1907), a Scottish physicist and mathematician. On the **Kelvin scale,** the freezing point of water is 273 kelvins (K), and the boiling point is 373 K. Notice that with the Kelvin scale, the degree sign is not used. The unit is simply called a kelvin. A change of 1° on the Celsius scale is the same as a change of 1 kelvin on the Kelvin scale. The zero point on the Kelvin scale, 0 K, or **absolute zero,** is equal to −273 °C. You can compare the Celsius and Kelvin temperature scales by looking at **Figure 3.19.** The relationship between a temperature on the Celsius scale and one on the Kelvin scale is given by the following equations.

$$K = °C + 273$$
$$°C = K - 273$$

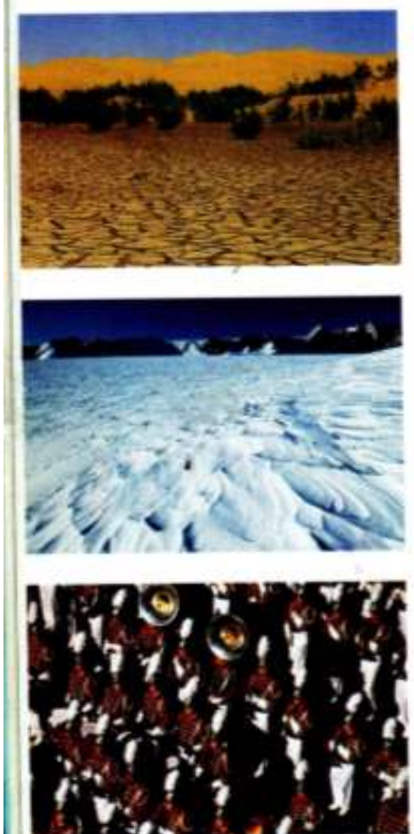


Figure 3.18
Temperatures on Earth range from the scorching heat of a desert, to the frigid cold of the Antarctic, to the normal human body temperature of 37 °C. What is the hottest temperature ever recorded? The coldest?

THE NATURE OF GASES

You are walking your dog in the woods, enjoying the great outdoors. Suddenly your dog begins to bark at what you believe is a black cat. But before you realize what it actually is, the damage is done. It's too late; the skunk has released its spray! Within seconds you smell that all-too recognizable smell. **How do the gaseous odor molecules travel from one place to another?**

objectives

- Describe the motion of gas particles according to the kinetic theory
- Interpret gas pressure in terms of kinetic theory

key terms

- kinetic energy
- kinetic theory
- gas pressure
- vacuum
- atmospheric pressure
- barometers
- pascal (Pa)
- standard atmosphere (atm)

Kinetic Theory

An ice cube is solid water; tap water is liquid water; and steam is gaseous water. Most substances commonly exist in only one of the three states of matter: solid, liquid, or gas. However, a substance in one state may change to another state with a change in temperature. You know that when cooled, water freezes to become ice, and steam condenses to form water. You also know that soon after you put on perfume, anyone nearby will be able to smell it. Obviously, the molecules of the perfume have moved through the air. How do you account for these behaviors? A model called the kinetic theory will help you find the answer.

The word kinetic refers to motion. The energy an object has because of its motion is called **kinetic energy.** The **kinetic theory** states that the tiny particles in all forms of matter are in constant motion. The following are the basic assumptions of the kinetic theory as it applies to gases.

1. A gas is composed of particles, usually molecules or atoms. These particles are considered to be small, hard spheres that have insignificant volume and are relatively far apart from one another. Between the particles there is empty space. No attractive or repulsive forces exist between the particles.
2. The particles in a gas move rapidly in constant random motion. They travel in straight paths and move independently of each other. As a result, gases fill their containers regardless of the shape and volume of the containers; uncontained gases diffuse into space without limit. The gas particles change direction only when they rebound from collisions with one another or with other objects. Measurements indicate that the average speed of oxygen molecules in air at 20 °C is an amazing 1660 km/h! At these high speeds, the odor molecules from a hot cheese pizza in Washington, D.C., should reach Mexico City in about 90 minutes. That does not happen, however, because the odor molecules are constantly striking molecules of air and rebounding in other directions. Their path of uninterrupted travel in a straight line is very short. The aimless path the gas molecules take is called a random walk. **Figure 10.1** on the following page illustrates a typical random walk.
3. All collisions are perfectly elastic. This means that during collisions kinetic energy is transferred without loss from one particle to another, and the total kinetic energy remains constant.

As you will learn next, the kinetic theory of gases is very helpful in explaining gas pressure.

p. 271 AND p. 272

substance. The remaining absorbed energy speeds up the particles—that is, increases their average kinetic energy—which results in an increase in temperature. The particles in any collection of atoms or molecules at a given temperature have a wide range of kinetic energies, from very low to very high. Most of the particles have kinetic energies somewhere in the middle of this range. Therefore average kinetic energy is used when discussing the kinetic energy of a collection of particles in a substance. **Figure 10.4** on the following page shows the distribution of kinetic energies of gas particles at two different temperatures. Notice that at the higher temperature there is a wider range of kinetic energies.

An increase in the average kinetic energy of particles causes the temperature of a substance to rise. As a substance cools, the particles tend to move more slowly, and their average kinetic energy declines. You could reasonably expect the particles of all substances to stop moving at some very low temperature. The particles would have no kinetic energy at that temperature because they would have no motion. Absolute zero (0 K, or

−273 °C) is the temperature at which the motion of particles theoretically ceases. Absolute zero has never been produced in the laboratory, although temperatures of about 0.0001 K have been achieved. Would you expect to find negative temperatures on the Kelvin temperature scale?

The Kelvin temperature scale reflects the relationship between temperature and average kinetic energy. The Kelvin temperature of a substance is directly proportional to the average kinetic energy of the particles of the substance. For example, the particles in helium gas at 200 K have twice the average kinetic energy as the particles in helium gas at 100 K. As you will see later in this chapter, the effects of temperature on particle motion in liquids and solids are more complex than in gases. Nevertheless, at any given temperature the particles of all substances, regardless of physical state, have the same average kinetic energy.

section review 10.1

3. According to the assumptions of kinetic theory, how do the particles in a gas move?
4. Use kinetic theory to explain what causes gas pressure.
5. Express the pressure 545 mm Hg in kilopascals.
6. How can you raise the average kinetic energy of the water molecules in a glass of water?
7. A cylinder of oxygen gas is cooled from 300 K (27 °C) to 150 K (−123 °C). By what factor does the average kinetic energy of the oxygen molecules in the cylinder decrease?

Chem ASAP! **Assessment 10.1** Check your understanding of the important ideas and concepts in Section 10.1.

tion between temperature and kinetic energy.

LINK TO MEDICINE

Cryogenics

Cryogenics is the science of producing very low temperatures and studying the behavior of matter at such temperatures. Biological reactions slow down or even stop at very low temperatures. Biologists use this fact to study reactions more easily. In medicine, some tissues can be removed from the body and preserved by rapid freezing. They can later be used when thawed. Blood and cartilage are examples of such materials. Doctors also perform cryosurgery by freezing tissue instead of cutting it. Cryosurgery minimizes bleeding during operations, and healing is rapid with minimal scarring.

p. 274

section 10.2

objectives

- Describe the nature of a liquid in terms of the attractive forces between the particles
- Differentiate between evaporation and boiling of a liquid, using kinetic theory

key terms

- vaporization
- evaporation
- vapor pressure
- boiling point
- normal boiling point

THE NATURE OF LIQUIDS

The Kilauea volcano in Hawaii is the most active volcano in the world—it has been erupting for centuries. The hot lava oozes and flows, scorching everything in its path, occasionally including nearby houses. When the lava cools, it solidifies into rock. **What makes a liquid different from a solid?**

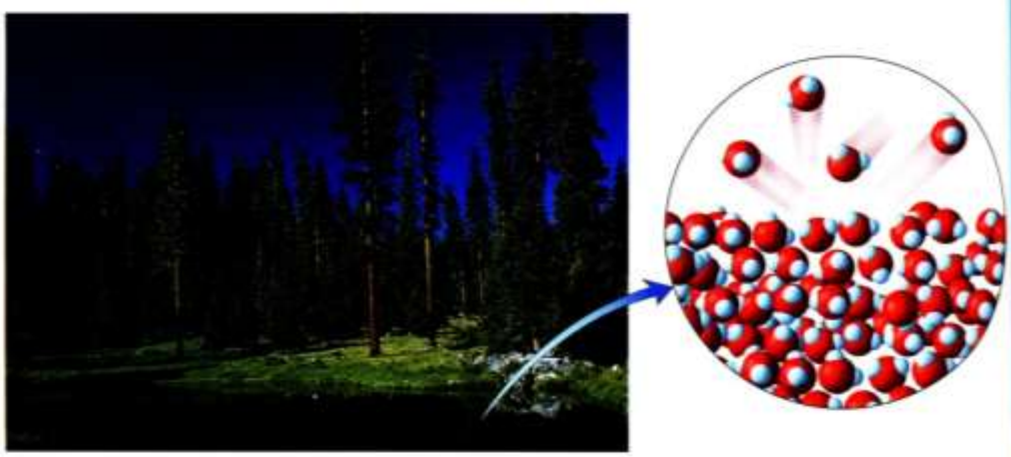
A Model for Liquids

You have just learned from the kinetic theory that gas pressure can be explained by assuming that gas particles have motion and that there is no attraction between the particles. The particles that make up liquids are also in motion, as **Figure 10.5** shows. Liquid particles are free to slide past one another. For that reason, both liquids and gases can flow, as you can see in **Figure 10.6.** However, the particles in a liquid are attracted to each other, while according to kinetic theory, those in a gas are not. The attractive forces between the molecules are called intermolecular forces.

The particles that make up liquids vibrate and spin while they move from place to place. All of these motions contribute to the average kinetic energy of the particles. Even so, most of the particles do not have enough kinetic energy to escape into the gaseous state. To do so, a particle must have sufficient kinetic energy to overcome the intermolecular forces that hold it together with the other particles. The intermolecular forces also reduce the amount of space between the particles in a liquid. Thus liquids are much more dense than gases. Increasing the pressure on a liquid has hardly any effect on its volume. The same is true of solids. For that reason, liquids and solids are known as condensed states of matter.

The interplay between the disruptive motions of particles and the attractive forces between them determines many of the physical properties of liquids. You will now explore two of these properties: vapor pressure and boiling point.

Figure 10.5
The particles in a liquid, such as the water in this lake, are close together but can move and slide around one another. The attractive forces between the particles generally prevent most of the particles from escaping the liquid and entering into the vapor state.



p. 280

section 10.3

objectives

- Describe how the degree of organization of particles distinguishes solids from gases and liquids
- Distinguish between a crystal lattice and a unit cell
- Explain how allotropes of an element differ

key terms

- melting point
- crystal
- unit cell
- allotropes
- amorphous solids
- glasses

THE NATURE OF SOLIDS

Imagine carbon atoms arranged in the shape of a soccer ball. Just such a structure, the buckminsterfullerene or buckyball, was discovered in 1985. Scientists are currently researching ways to make use of the buckyball's unique properties. Diamonds are also made of carbon atoms, and they too have unique properties that make them valuable for a variety of applications. **How many different molecular forms of carbon are there, and what are these forms called?**

A Model for Solids

The particles in liquids are relatively free to move. The particles in solids, however, are not, as **Figure 10.12** shows. Particles in a solid tend to vibrate about fixed points, rather than sliding from place to place. In most solids the particles are packed against one another in a highly organized pattern. Solids tend to be dense and incompressible. Because of the fixed positions about which their particles vibrate, solids do not flow to take the shape of their containers.

When you heat a solid, its particles vibrate more rapidly as their kinetic energy increases. The organization of particles within the solid breaks down and eventually the solid melts. The **melting point** (mp) is the temperature at which a solid turns into a liquid. At this temperature, the disruptive vibrations of the particles are strong enough to overcome the interactions that hold them in fixed positions. The melting process can be reversed by cooling the liquid so it freezes. The freezing of a liquid is thus the reverse of melting the corresponding solid. The melting and freezing points of a substance are at the same temperature. At that temperature, the liquid and solid substance are in equilibrium with each other.

$$\text{Solid} \underset{\text{freezing}}{\overset{\text{melting}}{\rightleftharpoons}} \text{Liquid}$$

In general, ionic solids have high melting points. This is because relatively strong forces hold them together. Sodium chloride, an ionic compound, has a rather high melting point of 801 °C. By contrast, molecular solids have relatively low melting points. For example, hydrogen chloride, a molecular compound, melts at −112 °C. What is the freezing point of hydrogen chloride? Not all solids melt, however. Wood and cane sugar, for example, decompose when heated.

Figure 10.12
The particles in a solid vibrate about fixed points. A solid melts when its temperature is raised to a level at which the vibrations of the particles become so intense that they disrupt the ordered structure.

Crystal Structure and Unit Cells

Most solid substances are crystalline. In a **crystal**, such as the one shown in **Figure 10.13**, the atoms, ions, or molecules that make up the solid substance are arranged in an orderly, repeating, three-dimensional pattern called the crystal lattice. All crystals have a regular shape. The shape of a crystal reflects the arrangement of the particles within the solid. The type of bonding that exists between the atoms determines the melting points of crystals.

p. 284

section 10.4

CHANGES OF STATE

objectives

- Interpret the phase diagram of water at any given temperature and pressure
- Describe the behavior of solids that change directly to the vapor state and recondense to solids without passing through the liquid state

key terms

- phase diagram
- triple point
- sublimation

All life on Earth requires water, which cycles through our ecosystem in three forms—solid ice, liquid water, and water vapor. You are familiar with these states in a variety of ways: raindrops in a spring shower, snowflakes from a winter blizzard, and water vapor making a summer day humid. **How are the solid, liquid, and vapor states related?**

Phase Diagrams

The relationships among the solid, liquid, and vapor states (or phases) of a substance in a sealed container are best represented in a single graph called a phase diagram. A **phase diagram** gives the conditions of temperature and pressure at which a substance exists as solid, liquid, and gas (vapor). **Figure 10.18** shows the phase diagram for water. Each of the three regions represents a pure phase of water. A line that separates any two regions gives the conditions at which those two phases exist in equilibrium. The curving line that separates water's vapor phase from its liquid phase reveals the equilibrium conditions for liquid and vapor; it also illustrates how the vapor pressure of water varies with temperature. Similarly, the other two curving lines give the conditions for equilibrium between liquid water and ice and between water vapor and ice. A unique feature of the diagram is the point at which all three curves meet. This meeting point, called the **triple point,** describes the only set of conditions at which all three phases can exist in equilibrium with one another. For water, the triple

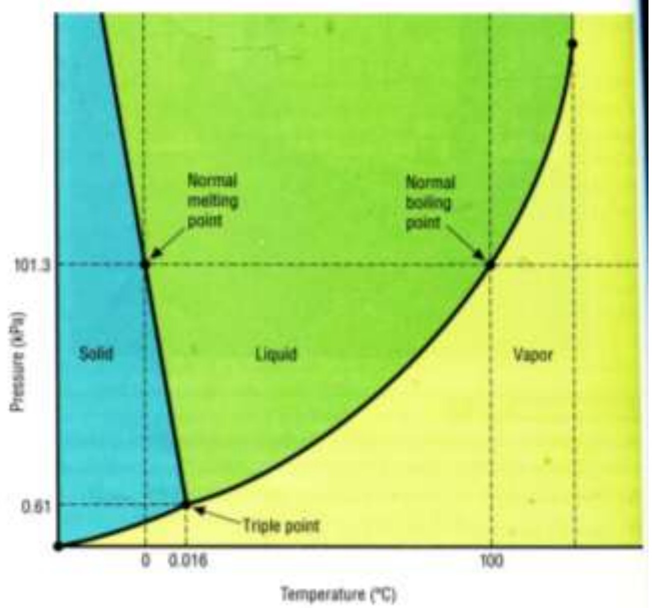


Figure 10.18
The phase diagram for water shows the relationship among pressure, temperature, and physical state. At the triple point, ice, liquid water, and water vapor can exist at equilibrium. Freezing, melting, boiling, and condensation can all occur at the same time, as shown in the photograph.