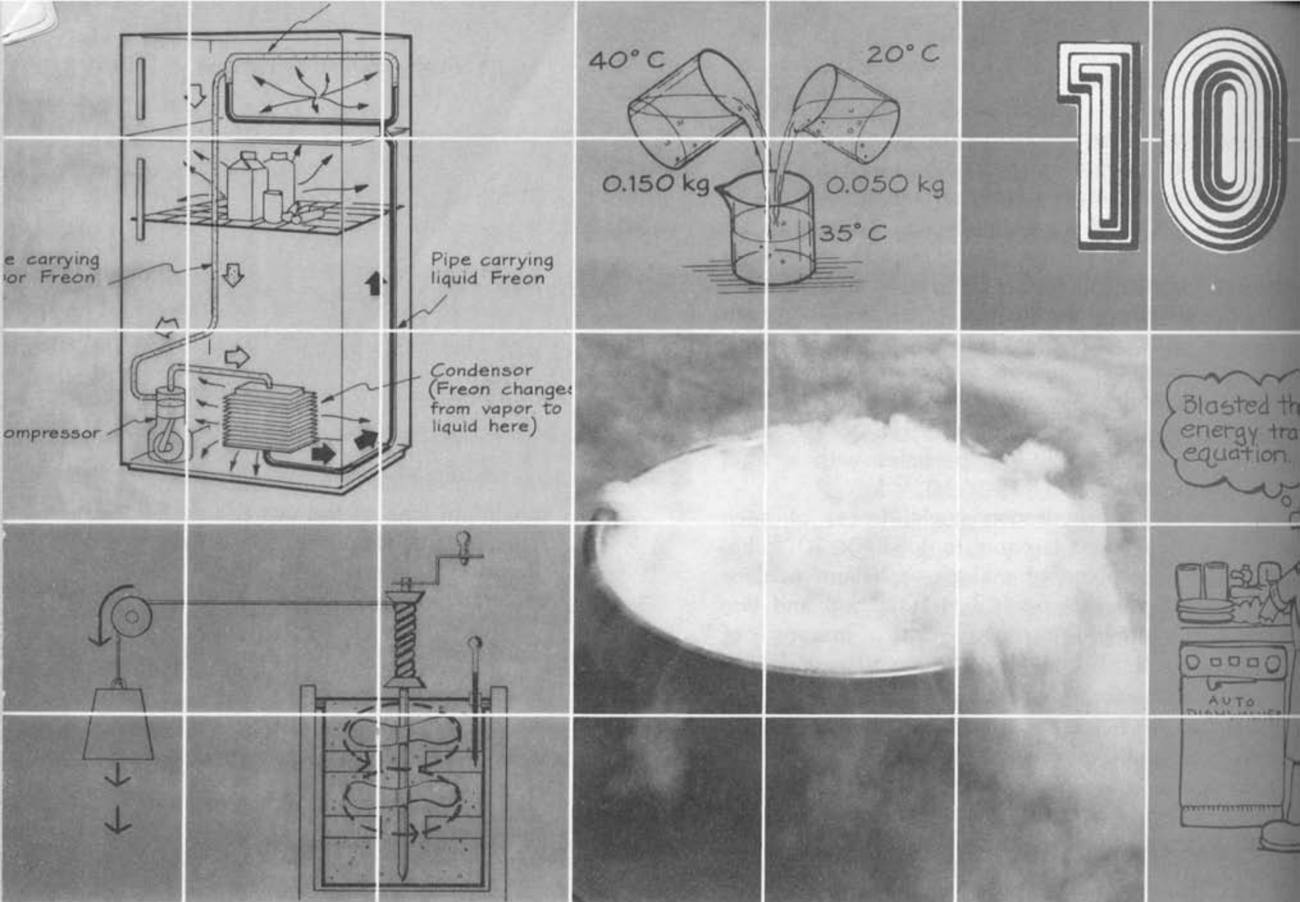


Thermal Energy in Matter

We rub our hands together; they become warmer. We rub two blocks of ice together; they melt. We rub two sticks together; they become hot—hot enough to ignite. For millenia people have known about the connection between motion and heat. Only within the last century, however, have we realized the reason for this connection. Heat and motion are both forms of energy. The energy of motion can be converted into the energy of heat.

Heat can cause a substance to change its temperature or to change its state from a solid to a liquid or a liquid to a gas—or even from a solid to a gas. In this chapter we look at interactions that involve changes in temperature and changes in state. Changes in temperature are described in terms of the different *specific heat capacities* of objects. Changes in state are described in terms of *latent heats of fusion* and *latent heats of vaporization*. The energy exchanged during these interactions comes from the energy stored internally in matter, called *thermal energy*.

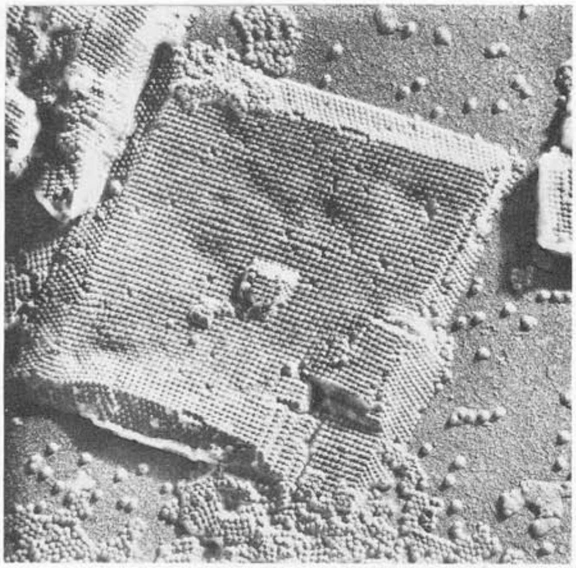


Figure 10-1
Magnified 80,000 times, the necrosis virus protein shows an orderly arrangement of molecules.

THE ENERGY STORED IN MATTER

A glass of hot water looks the same as a glass of cold water. Yet if you put your finger in each glass, you can feel a difference. A piece of ice looks different from a puddle of water. Yet both can exist at the same temperature—your finger would not notice any difference. Hot water, cold water, ice, and puddles all differ in the amount of energy stored in them internally. To understand how this energy is related to temperature and state, we need to examine matter microscopically.

Just looking at a piece of ice, it is hard to imagine that it contains billions of tiny particles. Yet an overwhelming amount of indirect evidence convinced scientists a century ago that matter is made up of molecules. Today, electron microscopes allow us to see directly what those before us merely imagined. Consisting of one to hundreds of atoms, molecules range in size from 10^{-10} meters (m) to 10^{-5} m. Billions of neatly stacked molecules are found in the virus protein shown in Figure 10-1. Individual molecules in the virus protein are about 10^{-8} m in diameter. The neat stacks arise from electrical forces that hold molecules together. The irregularities arise from the fact that individual molecules move about constantly, occasionally venturing out of position. Electrical potential energy is associated with the electrical forces, and kinetic energy is associated with the molecular motion. Changes in the state of matter—from water to ice, for example—are related to changes in electrical potential energy. Changes in temperature are related to changes in kinetic energy.

Differences in the strengths of the electrical forces among molecules are responsible for the three **states of matter**—solid, liquid, and gas (Figure 10-2). Very weak electrical forces lead to the lack of shape and volume characteristic of gases. The molecules of a gas move about independently of one

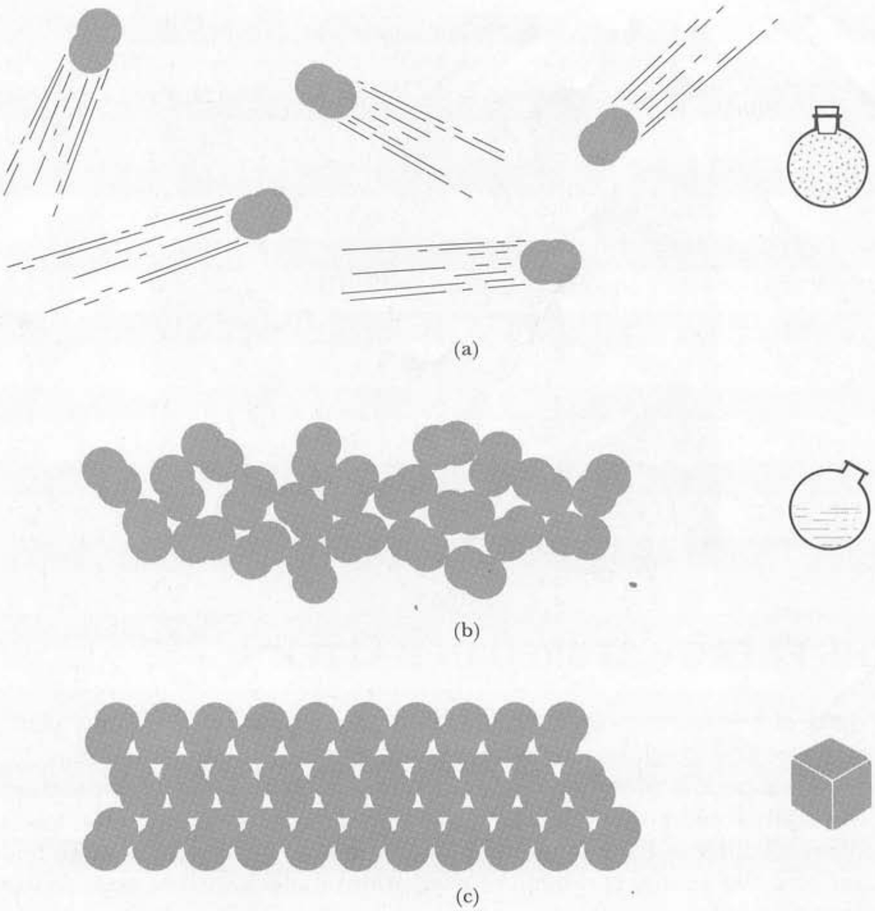


Figure 10-2
The three states of matter differ because the strength of the electrical forces among molecules varies.
(a) Very weak forces allow molecules in gases to move about almost at will.
(b) Slightly stronger forces allow the molecules in a liquid to move about some, but the liquid retains a definite volume.
(c) Stronger forces limit the molecular motions in solids to vibrations about a central location. A solid maintains a definite shape and volume.

another. Both the shape and the volume of a gas change easily. When a gas condenses into a liquid, the molecules move close enough to experience much larger electrical forces. The forces are large enough to keep molecules within a definite volume but not within a definite shape. A liquid, unlike a gas, stays confined in an open container; however, it does assume the shape of the container. As a liquid solidifies, its molecules move even closer together. The electrical force becomes large enough to keep the molecules in the orderly stacks that lead to a definite shape and volume. Molecular motions in solids are restricted to small vibrations about fixed centers.

The electrical potential energy found in matter depends on the strength of these intermolecular forces, much as the gravitational potential energy of an object depends upon the strength of the gravitational force acting on it. When we compared the gravitational potential energy of the pile driver on earth and on the moon, we found it to be greater on earth. Larger gravitational forces lead to larger amounts of energy stored in an object in a given position. By the same token, the amount of electrical potential energy stored in matter depends on the strength of the electrical force binding the molecules together. For any given substance, we will find more electrical potential energy in its solid state, less in its liquid state, and very little in its gaseous state. Changes in state involve changes in electrical potential energy.

Differences in the motion of molecules within matter are responsible for the temperature differences we feel with our hands or measure with a thermometer. Molecules move constantly. In gases like air, molecular motions are fairly obvious because they give rise to motions of visible particles. Dust particles, for example, do not simply fall to the ground. Air molecules collide with them constantly, keeping them suspended. Pollen grains suspended in a liquid also bounce around as the moving molecules of the liquid hit them. Molecules move within solids, too, although this is not so obvious. Kinetic energy is associated with these molecular motions. The faster the molecules move, the higher are their kinetic energies. A substance's **temperature** is proportional to the average kinetic energy of its molecules. Changes in temperature involve changes in kinetic energy.

The combination of the electrical potential energy and the kinetic energy associated with molecules is called the **thermal energy** of a substance. Interactions in which there is a change in temperature or a change in state are called thermal interactions. Many of the concepts used to describe thermal interactions were actually introduced before the molecular model of matter. In Chapters 10–12 we examine these concepts in connection with everyday situations. In Chapter 13 we discuss the molecular model of matter in more detail and see how these everyday events can be described at the molecular level.

SELF-CHECK 10A

Hot water and cold water differ in temperature. How do they differ in the thermal energy stored in their molecules? Ice at 32°F and water at 32°F differ in state. How do they differ in the thermal energy stored in their molecules?

TEMPERATURE

Suppose you are heating some chocolate milk on the stove. You wander off to watch a football game while thermal energy is being transmitted from the flame to the pot to the milk. You don't call it thermal energy, of course. You call it heat. Ordinarily we say that the flame heats up the milk. When you come back to the kitchen, you are not sure whether the milk is hot enough, so you test it with your finger. It is hot enough, all right! Thermal energy—heat—was transferred from the milk to your finger, raising your skin temperature noticeably.

Heat and Temperature

Heat and temperature—these two terms are so closely involved in describing interactions like making hot chocolate that we often confuse them. Heat and temperature, however, describe two very different concepts.

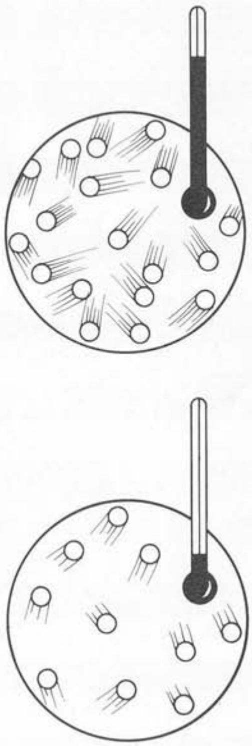


Figure 10-3
The temperature of a substance is a measure of the average kinetic energy of its molecules.

Heat is the name we give to thermal energy that is in the process of being transferred from one object to another. As happens in all energy transfer processes, one object acts as the energy source and the other as the energy receiver. The energy source loses thermal energy—molecular bonds tighten or molecules move around more slowly. The energy receiver gains thermal energy. Its bonds loosen or its molecules move about more rapidly. Heat simply describes the thermal energy as it is in transit from one to the other. Physicists now use the term *thermal energy* for heat as well as for the energy stored internally in the molecular bonds and motions found in matter.

Though everyone knows intuitively what temperature is, there is no simple definition. In everyday terms, **temperature** is the hotness or coldness of an object. Hot describes how we perceive high temperatures; cold describes how we perceive low temperatures. Scientifically, we define **temperature** as a measure of the average kinetic energy of an object's molecules. *Hot* describes objects in which the molecules are moving about rapidly. *Cold* describes objects in which the molecules are moving about much more slowly. The higher the temperature of a substance, the faster its molecules move about (Figure 10-3).

This rather general definition of temperature allows us to understand the distinction between heat and temperature. When two objects at different temperatures are brought into contact with one another, heat always moves from the hotter object to the colder one. Heat was transferred from the hot chocolate to your finger. The molecules in the hot chocolate begin to slow down, while the molecules in your skin begin to speed up. Heat is transferred until molecules in your finger and the hot chocolate are moving at about the same rate—until both are at the same temperature. Physicists say that thermal energy is transferred until molecules in your finger and the hot chocolate have about the same average kinetic energy.

Since thermal energy can be stored in molecular bonds as well as in the motion of an object's molecules, heat can be transferred even though an object does not change temperature. When you hold an ice cube, you hand gets colder and the ice cube melts. The molecules in your hand slow down, but the molecules in the ice move apart rather than speed up. The heat you supply goes into stretching the molecular bonds, thus decreasing the strength in the electrical forces between molecules. The ice cube stays at the same temperature, even though you constantly supply heat—thermal energy—to it.

Temperature Measurement

While your finger is a convenient temperature-measuring device, it is not very precise. The best it can give you is a qualitative assessment, such as warm. A variety of instruments have been invented to measure temperature more precisely, the most common of which is the thermometer. A thermometer consists of a small amount of fluid, usually mercury, inside a closed tube. A scale along the outside of the tube allows us to measure the expansion and contraction of the mercury as it gains or loses thermal energy.

When a substance is heated, its molecules start moving over a wider area (as well as faster), so the substance expands. When a substance is cooled, its

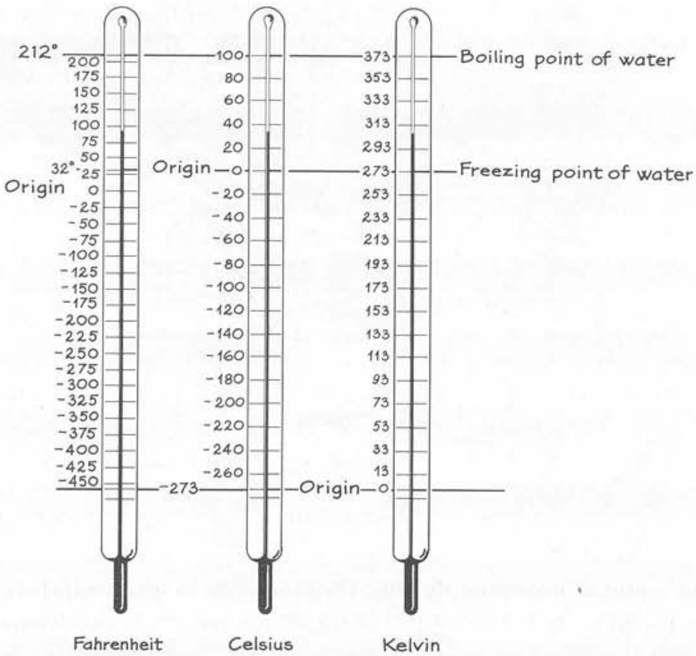


Figure 10-4

The Celsius, Kelvin, and Fahrenheit scales differ in the origin, degree size, and reference temperatures.

molecules move closer together again and the substance contracts slightly. This expansion and contraction allows us to measure temperature. When you place a thermometer in hot water, for example, thermal energy is transferred from the water to the mercury inside the tube. The mercury becomes hotter and expands, rising inside the glass tube. Thermal energy is transferred until the mercury reaches the same temperature as the hot water. We then associate the height of the mercury inside the column with the temperature of the water. If you then move the thermometer to a pan of cold water, thermal energy is transferred from the mercury to the water. The mercury cools, contracting, so its height in the glass tube drops. The mercury continues to lose thermal energy until it reaches the same temperature as the cold water. The scale along the outside of the glass tube allows you to associate a number with the height of the mercury column. You call this number the water's temperature.

Three different temperature scales are in use today—the Celsius, Fahrenheit, and Kelvin scales. These scales differ in their origin, or zero point, and in the sizes of their divisions, or degrees. The origin and degree size are both determined by the choice of reference temperatures used to establish the scale. Since water is an abundant and readily available substance, the freezing and boiling points of water are convenient reference temperatures. Figure 10-4 compares the origin, degree size, and reference temperatures for the Celsius, Fahrenheit, and Kelvin scales.

The **Celsius** scale is by far the most widely used temperature scale. Its origin (0 °C) is the freezing point of water. The boiling point of water has been arbitrarily chosen to be 100 °C. Since the two reference points are separated by 100 units, each unit is 0.01 the temperature difference between freezing

and boiling water. Consequently, the Celsius scale is often referred to as the centigrade (centi = one-hundredth) scale. Since we often experience temperatures below the freezing point of water (0°C), both negative and positive temperatures are common.

The **Kelvin** scale assigns different values to the two reference temperatures. The freezing point of water is called 273 K and the boiling point of water is called 373 K. One hundred units still separate the freezing and boiling temperatures of water, so the size of the Kelvin degree is the same as the Celsius degree. But, as illustrated in Figure 10-4, the origin has been shifted. In the Kelvin scale, the origin (0 K) has been selected to be the temperature at which the molecules in a substance have the minimum possible kinetic energy. This means that the substance cannot get any colder. Consequently, negative Kelvin temperatures do not exist, and 0 K is an **absolute zero.** Additionally, the temperatures in the Kelvin scale represent actual proportions of kinetic energy. If the state of a substance has not changed, matter at 200 K has molecules with twice the average kinetic energy as matter at 100 K. Consequently, the Kelvin scale is frequently used in scientific work. The conversion between the Celsius scale and the Kelvin scale is given by:

$$K = {}^{\circ}C + 273$$

In the United States, the **Fahrenheit** scale is still used extensively. Fahrenheit also chose to divide his scale into 100 degree intervals, but he chose as the origin (0°F) the freezing temperature of a particular mixture of salt and water. Apparently this temperature (equal to about −18.9°C) was the coldest he could produce. He then chose human body temperature as the second reference point, designated by 100°F. We have kept this reference point, though we know now that human body temperature is about 98.6°F. (Fahrenheit had either a hot body or a bad thermometer!) Translating the reference temperatures of the Celsius and Kelvin scales to Fahrenheit units, we find that the freezing point of water is 32°F and the boiling point of water is 212°F. On the Fahrenheit scale, there are 180 degree units between the two reference temperatures, while both the Celsius and Kelvin scales have only 100 degree units. Consequently, the Fahrenheit degree is smaller than the Celsius or Kelvin degree. While we have grown accustomed to the Fahrenheit system, its reference temperatures are difficult to reproduce accurately. Consequently, scientific work uses only the Celsius and Kelvin scales.

THERMAL ENERGY TRANSFER AND CHANGES IN TEMPERATURE

Cold milk is poured into hot coffee. The coffee cools down and the milk warms up. Cold air from an air conditioner mixes with the warm summer air in the room. The mixture is neither warm nor cold, but somewhere in between. Cold vegetables are plunged into a pot of boiling water. The water stops boiling and the vegetables get warm. Objects at different temperatures transfer thermal energy and, in the process, change temperature. In this section we investigate the relationship between thermal energy transfer and temperature change.

Specific Heat Capacity

Your pizza has arrived straight from the oven. Carefully, you pick up a slice. The crust seems warm, but not too hot to eat. Of course, the hands are not outstanding temperature sensors, so you carefully touch the crust to your tongue. The pizza is quite warm, but it does not burn. So . . . you take a big bite. Ouch!!! The sauce burns the roof of your mouth. What happened? The sauce and crust were at the same temperature, yet the sauce burned your mouth while the crust did not. Both sauce and crust contained thermal energy, which was then transferred to your mouth, raising your mouth's temperature. Evidently, though, the sauce contained a lot more thermal energy than the crust.

Figure 10-5

Think! . . . About specific heats?

Our experience with pizza illustrates the fact that substances heated to the same temperature may store and later transfer very different amounts of thermal energy. If we heat equal amounts of pizza crust and sauce to exactly the same temperature, the sauce will have absorbed more thermal energy than the crust. On the other hand, if we transfer exactly the same amount of thermal energy to the same mass of pizza crust and sauce, the temperature of the sauce will be lower than the temperature of the crust. More thermal energy is required to raise the sauce to the same temperature as the crust. We use the concept of specific heat capacity to describe the different amounts of thermal energy stored in different substances.

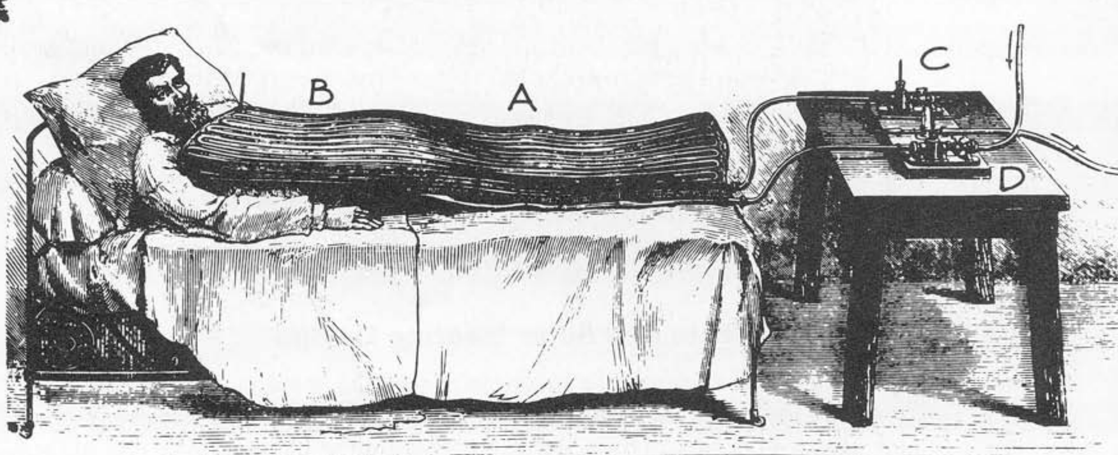
The **specific heat capacity** of a substance is defined as the amount of energy transferred to or from 1 kilogram (kg) of a substance for each 1° change in temperature. Specific heat capacities for a variety of substances are listed in Table 10-1. The specific heat capacity for tomatoes, for example, is 3.9 kilojoules per degree Celsius per kilogram (kJ/°C · kg). A kilojoule (kJ)

Table 10-1 Specific Heat Capacities of Common Substances

Substance	Specific Heat Capacity (kJ/°C · kg)
Water	4.2
Beer	4.0
Tomatoes	3.9
Cheese	3.4
Apples	3.3
Olive Oil	2.0
Wood (average)	1.7
Flour dough	1.7
Porcelain	1.1
Sugar	1.1
Air	1.0
Aluminum	0.9
Brick	0.8
Rock (average)	0.8
Glass	0.8
Iron	0.5
Copper	0.4
Silver	0.3
Tin	0.2

is 1000 joules. In other words, 1 kg of tomatoes must receive 3.9 kJ of thermal energy before its temperature will increase by 1°C. By contrast, copper has a specific heat capacity of 0.4 kJ/°C · kg. One kilogram of copper will increase its temperature by 1°C when only 0.4 kJ of thermal energy have been added. A material with a high specific heat capacity stores more thermal energy per kilogram than one with a lower specific heat capacity.

Specific heat capacity can help us understand the mouth-burning capabilities of pizza sauce. As shown in Table 10-1, tomatoes have a much higher specific heat capacity than dough. (This is primarily due to the fact that tomatoes have large quantities of water in them, and water has an extremely high heat capacity. By contrast, cooked pizza dough has a relatively small amount



BLANKETS TO KEEP YOU COOL

We have electric blankets to keep us warm in the winter, so why shouldn't we have a cooling blanket for the summer? Such a blanket would have to take thermal energy from us and deposit it somewhere else. In 1879 these thoughts led to the invention of the refrigerating blanket (*A*). Something was circulated through tubes (*B*) in this blanket by means of pumps (*C* and *D*). To be efficient, the material that circulated had to take away the maximum amount of thermal energy for each kilogram and each degree decrease in temperature. Thus the circulating material needed to have the highest specific heat capacity possible. The inventor of the blanket chose water for this reason. The water also allowed the blanket to be reversible. Its high specific heat capacity allowed the user to heat the water before it was circulated and transfer a lot of energy to the sleeper. It could be warm in the winter and cool in the summer. Try that with an electric blanket!

of water in it.) When the pizza is placed in the oven, the sauce and dough both heat to the same temperature as the oven. Because of its high specific heat capacity, however, the sauce absorbs more thermal energy per degree change in temperature than the dough. When you put a piece of the hot pizza in your mouth, it cools to the same temperature as your mouth. To reach this temperature, the sauce must transfer larger amounts of thermal energy to your mouth than the crust. That burning sensation is a consequence of the high specific heat capacity of tomatoes.

Of course, just giving a phenomenon a name such as *specific heat capacity* does not explain it. You may be wondering how objects of the same mass and temperature can contain different amounts of thermal energy, or how objects of the same mass can require different amounts of thermal energy to reach the same temperature. In a substance with a high specific heat capacity, where does the "extra" thermal energy go? We examine this in more detail in Chapter 13.

SELF-CHECK 10B

Twenty kilojoules of thermal energy are transferred to each of two spoons of the same mass—one made out of wood and the other made out of aluminum. Use the specific heat capacities in Table 10-1 to determine which spoon will reach a higher temperature.

Specific Heats and Solar Heating Design

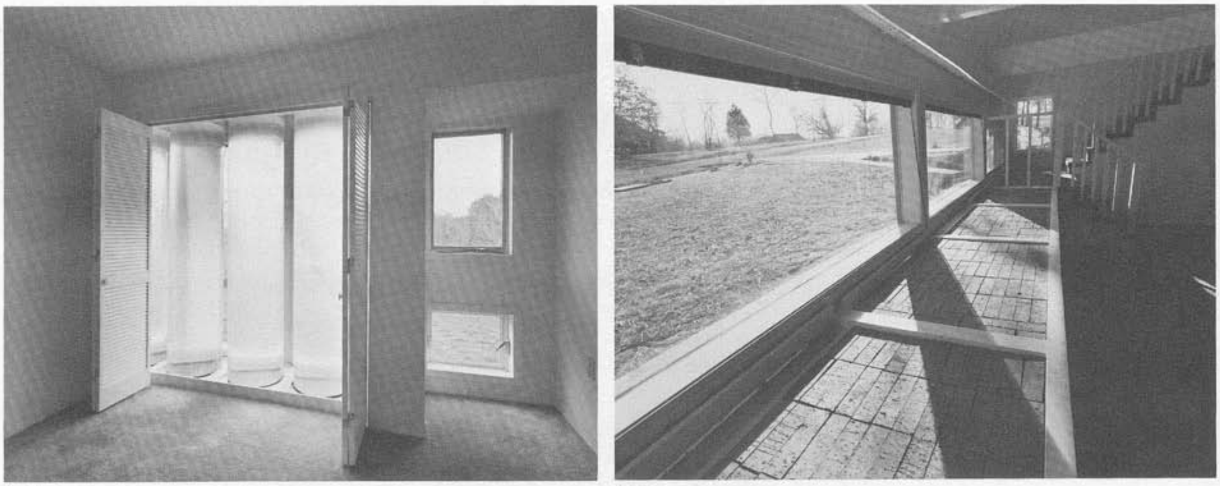
Specific heat capacity is a valuable concept to someone who is trying to store energy for future use. For example, a solar heating system stores energy, which is available on sunny days, for use at night or on cloudy days. This energy can be stored by converting the sun's energy into thermal energy stored in matter. Later, the thermal energy can be withdrawn and used to heat a home.

To be useful in storing solar energy, a substance must have several characteristics. Since the space devoted to home heating systems is rather small, the substance must have a small mass and volume. Additionally, since a material at a high temperature could boil away or become a fire hazard, the substance should remain at a fairly low temperature as its absorbs solar energy. A material with a high specific heat capacity stores more energy per kilogram and per degree increase in temperature than a material with a low specific heat capacity.

Water, with its extremely high specific heat capacity, is sometimes chosen for solar heating systems. Water will absorb a great deal of thermal energy before its temperature rises to the boiling point. However, since liquids can evaporate, cause rust in metal parts, or leak from their containers, water is not the most convenient material. Instead, many designers choose various types of rocks. Rocks are easily handled, inexpensive, and stable at everyday temperatures. These characteristics frequently outweigh the advantage of using water with its higher specific heat capacity (Figure 10-6).

How Much is Transferred?

Thermal energy is transferred from warmer objects to cooler ones. The flame transfers thermal energy to our chocolate milk. The oven transfers thermal energy to the pizza crust and sauce. Experience tells us something about the amount of thermal energy transferred in each of these processes. We can describe the amount of thermal energy transferred in terms of the energy receiver—the hot chocolate or the pizza. The hotter you want your hot chocolate, the more thermal energy you must supply. If you want to heat 2 cups of milk instead of just one to the same temperature, you have to supply more thermal energy. And, as illustrated by our pizza example, if you want different substances to reach the same temperature, you have to supply different amounts of thermal energy to them.



(a) (b)

Figure 10-6 Solar heat storage systems use materials that have relatively high specific heats. **(a)** Tubes that are filled with water are placed in the window. Energy stored in the water can enter the room when the louvered doors are open. **(b)** The brick surface stores solar energy.

Consistent with these observations, the *thermal energy transferred* in interactions involving a *change in temperature* is the product of the substance's mass, specific heat capacity, and change in temperature:

$$\begin{matrix}\text{Thermal energy} \\ \text{transferred}\end{matrix} = (\text{mass}) \times (\text{specific heat}) \times \begin{pmatrix}\text{change in} \\ \text{temperature}\end{pmatrix}$$

Thermal energy transferred

Specific heat capacity

$$TE = mc\,\Delta T$$

Mass

Change in temperature

To see how this definition can be applied, suppose you want to cool 0.35 kg of beer from room temperature (20°C) to 5°C. The amount of thermal energy you need to remove from the beer is the product of the beer's mass, its specific heat capacity, and the change in temperature—(0.35 kg)(4.0 kJ/°C · kg) (20°C − 5°C) = 21 kJ. We have to remove 21 kJ of energy from the beer in order to cool it from room temperature to 5°C. (Text continues on p. 220.)

SELF-CHECK 10C

In cooling the beer, we ignored the fact that we also have to cool the can. How much additional thermal energy must be removed to cool a 0.02 kg aluminum can from 20°C to 5°C? (See Table 10-1 for the specific heat capacity of aluminum.)

A STEP FURTHER—MATH

HOT, COLD, AND LUKEWARM

Have you ever wondered how people think up equations like the thermal energy transfer equation? While you can probably convince yourself that it makes sense, a few relatively simple experiments can show you how scientists control variables one at a time in order to build a complete description of a relatively complex interaction. Let's examine a series of experiments in which a hot substance is mixed with a cold substance.

First, consider the effect of temperature. Figure 1 shows three experiments in which 0.050 kg of hot water at different initial temperatures has been mixed with 0.050 kg of cold water at 20°C. The magnitudes of the change in temperature of the hot water always equal those of the change in temperature of the cold water. In Figure 1(a), for example, the hot water cools from 40°C to 30°C ($\Delta T = -10°C$), while the cold water warms up from 20°C to 30°C ($\Delta T = 10°C$). In equation form, we say that

$$-\Delta T_{hot} = \Delta T_{cold}$$

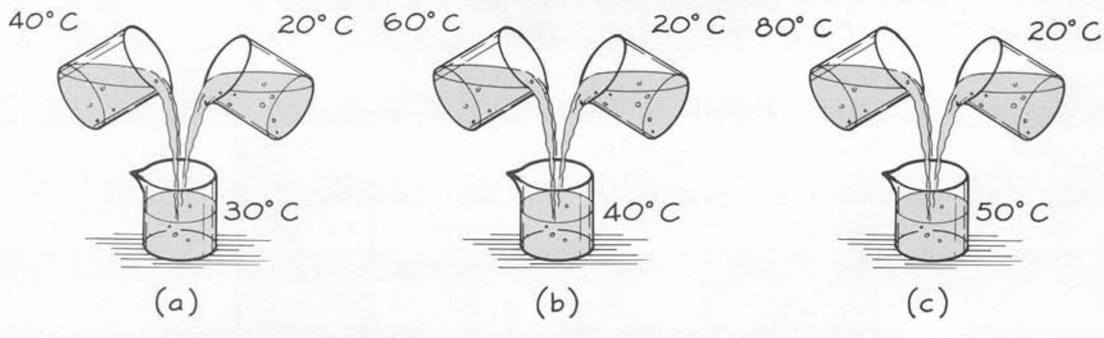
Next, add the effect of mass. Figure 2 shows three experiments in which different masses of hot water at the same initial temperature (40°C) have been mixed with 0.050 kg of cold water at 20°C. When we compare the changes in temperatures, our simple relationship no longer works. In Figure 2(c), for example, the hot water cools from 40°C to 35°C ($\Delta T = -5°C$), while the cold water warms up from 20°C to 35°C ($\Delta T = 15°C$)—three times the temperature decrease in the hot water. But, we added three times as much hot water as cold water. If we compare the products of mass and temperature change for the hot and cold water, we find them to be equal. Try it!

$$-m_{hot}\Delta T_{hot} = m_{cold}\Delta T_{cold}$$

Finally, add the effect of specific heat capacity. Figure 3 (page 219) shows three experiments in which equal masses (0.050 kg) of water, copper, and aluminum at the same initial temperature (40°C) have been mixed with 0.050 kg of cold water at 20°C. Hot water caused a temperature increase of 10°C in the cold water; aluminum, an increase of 3.5°C; and copper, an increase of only 1.6°C. If

Figure 1

Equal masses of hot and cold water are mixed.



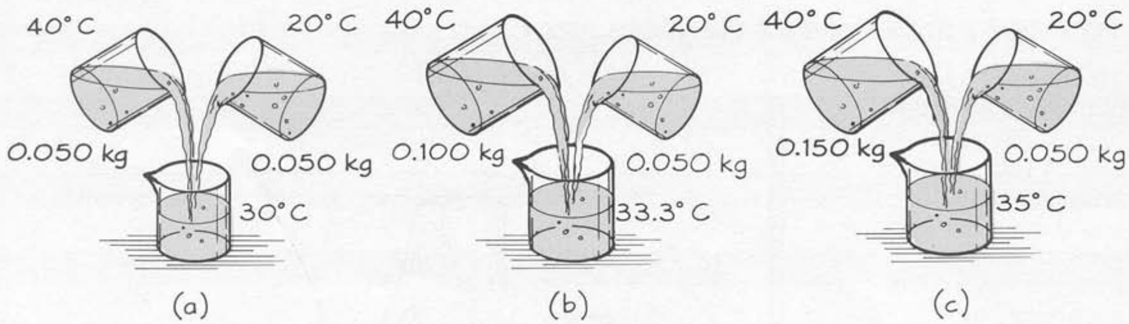


Figure 2
Different masses of hot water are added to 0.050 kg of cold water.

we compare the specific heat capacities of the three materials (Table 10-1), we find that water has the highest specific heat (4.2 kJ/°C · kg), aluminum has a significantly lower specific heat (0.9 kJ/°C · kg), and copper has the lowest specific heat (0.4 kJ/°C · kg). We modify our relationship one final time:

$$-m_{\text{hot}}c_{\text{hot}}\Delta T_{\text{hot}} = m_{\text{cold}}c_{\text{cold}}\Delta T_{\text{cold}}$$

Does it work? Try it and see!

Beneath this series of experiments lies a fundamental assumption—that energy is conserved. We assume that the hot object acts as the energy source and the cold object acts as the energy receiver. Whatever thermal energy is, we assume that the thermal energy lost by the hot substance equals the thermal energy gained by the cold substance. By manipulating the three variables one at a time, we could discover the role each plays in defining the amount of thermal energy transferred.

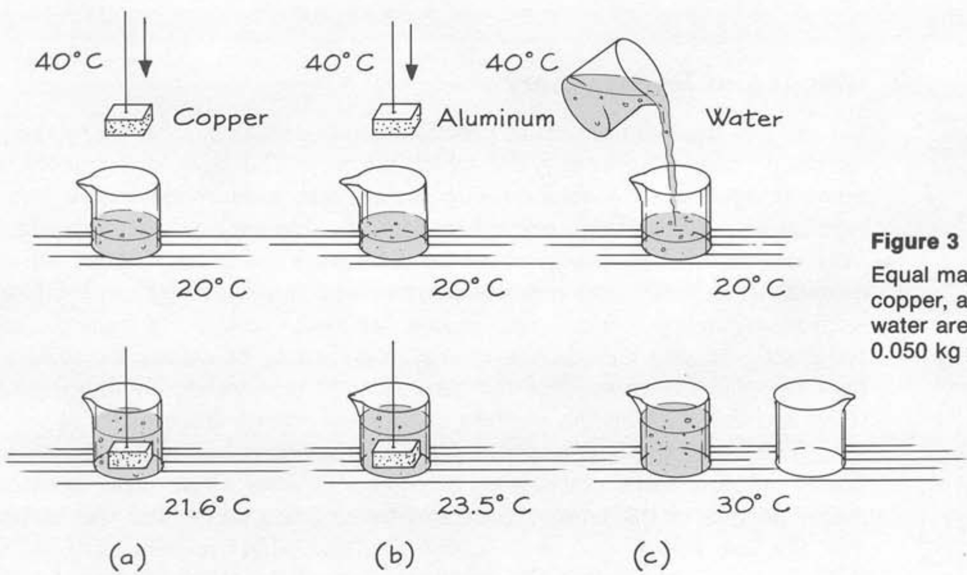
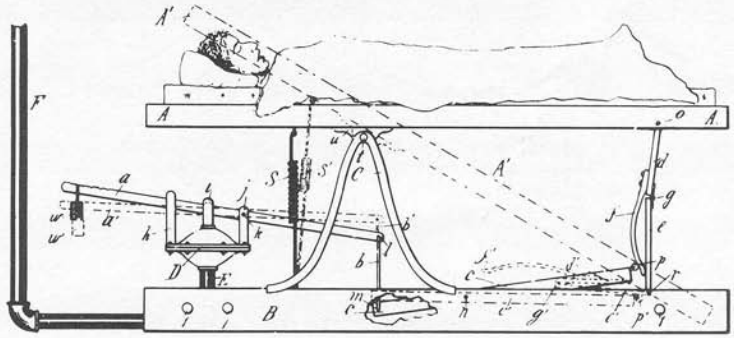


Figure 3
Equal masses of copper, aluminum, and water are added to 0.050 kg of cold water.

YOU CAN'T SLEEP IF THE FLOWERS ARE COLD



What would happen in the middle of the night if the steam heat in your greenhouse goes off? If you don't wake up, your roses might freeze. To avoid this possibility Ludwig Ederer invented, in 1900, the greenhouse-keeper's alarm bed. It relies on the volume changes that take place when a gas turns into a liquid. Under normal conditions, steam in the pipe (F) heats the greenhouse and applies a force to a series of levers, one of which (g) is holding up the bed (A). If the fire in the boiler goes out, the steam cools and turns into a liquid. Because the liquid takes up a smaller volume, it does not apply the same force to the levers. They collapse, and the bed moves to position A'. The force of gravity being what it is, the person is soon on the floor. The only way the bed can return to the horizontal position (A) is by changing the water back to steam. So, the greenhouse-keeper must get the fires burning. Then, both the keeper and the roses can sleep in comfort.

Dishpan or Dishwasher?

Figure 10-7

We can put this relationship to practical use in estimating the energy required for daily tasks—the energy for which we pay. The age-old argument about doing things the old-fashioned way versus with modern devices is bound to reoccur as energy becomes more expensive. One such argument could occur over washing dishes. Dishwasher manufacturers are quick to point out that a dishwasher actually uses less energy than washing dishes by hand. Of course, such an argument depends on just how far back you want to trace the energy required—do you include the energy required to construct the dishwasher, for example? But if we take the statement at face value, we can evaluate it using our definition of the amount of thermal energy transferred.

The basic argument is that a dishwasher uses less water than washing dishes by hand since you can do a full day's dishes all at once. Because the major portion of the energy goes into heating the water, the less water you use, the less energy you use. Automatic dishwashers require very hot water (60°C) to clean the dishes effectively and use about 40 kg of water for a complete cycle. If we assume that the water is initially at room temperature, then we must supply (40 kg)(4.2 kJ/°C · kg)(60°C − 20°C) = 6720 kJ. Washing dishes by hand usually requires water at about 38°C and uses 28 kg of water, two average-sized sinks half full (one for washing and one for rinsing). Again, if we assume that the water is initially at room temperature, then we must supply (28 kg)(4.2 kJ/°C · kg)(38°C − 20°C) = 2117 kJ. For a single washing, the difference is substantial. We did not even take into account the energy required if you use the dry cycle. Even if you wash dishes by hand three times a day, the dishwasher uses more energy—although not much more. But if you wash dishes by hand less often, the dishpan is clearly an energy saver!

How much energy do you save when you turn your thermostat down to 18°C? What kind of temperature rise will the waste heat from a power plant produce in a nearby lake? If you let the bath water cool to room temperature before letting it out, will it really heat your bathroom enough to matter? Like the dishwasher/dishpan issue, these questions often involve more factors than can be seen on the surface. But do not let the complexity keep you from making rough estimates like the ones we made for the dishwasher and dishpan. The definition of the amount of thermal energy transferred during changes in temperature allows us to check the validity of many statements made about energy.

THERMAL ENERGY TRANSFER AND CHANGE OF STATE

Place a 1 kg block of ice at −20°C on a stove. Turn on the burner so that a constant amount of thermal energy is transferred to the ice. Then monitor the temperature of the ice (Figure 10-8). As we expect, the thermal energy

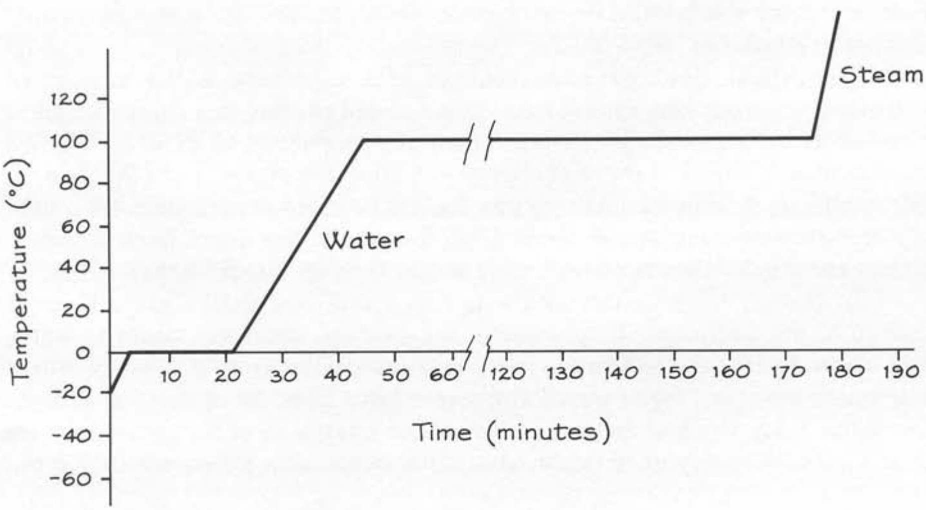


Figure 10-8
The temperature of a substance remains constant as it changes state, even though thermal energy is constantly being provided to it.

supplied by the burner causes an increase in the temperature of the ice. But, when the ice reaches 0°C, something surprising occurs. Its temperature stops increasing as it starts to melt, even though thermal energy is continually being supplied to it. Instead of seeing a change in temperature, you see a change of state. Once the ice has completely melted, the temperature of the water again begins to increase steadily. If you continue to heat the water, you see a similar result at 100°C, when it begins to boil. The temperature of the water remains constant until all the water has been converted to steam. At a substance's melting or boiling point, its temperature remains constant even though it is absorbing thermal energy. The thermal energy is not available for producing a temperature rise—it is needed to cause a change in state.

Latent Heat

Water changes from solid to liquid at 0°C and from liquid to gas at 100°C. Other substances undergo the same changes in state, though at different temperatures. If we measure the amount of thermal energy needed to melt 1 kg of ice with that needed to vaporize 1 kg of water, we find them to be considerably different. Only 320 kJ of energy will melt 1 kg of ice, while 2160 kJ are needed to vaporize 1 kg of water. If we compare the amount of energy needed to melt 1 kg of ice with that needed to melt 1 kg of iron, we see substantial differences again. It takes 320 kJ of energy to melt 1 kg of ice, while only 33 kJ melt 1 kg of iron. The amount of thermal energy needed to cause a change in state varies with whether the change is from a solid to a liquid or from a liquid to a gas and with the particular substance involved. The concept of latent heat is used to describe the thermal energy needed to produce a change of state in a specific substance while the temperature of the substance remains constant.

The **latent heat of fusion** of a substance is the amount of thermal energy per kilogram needed to change it from a solid to a liquid at its melting point. Ice has a latent heat of fusion of 320 kJ/kg. That means that we must supply 320 kJ of thermal energy to each kilogram of ice at 0°C in order to change it from a solid to a liquid. Latent heats of fusion for a variety of substances are listed in Table 10-2.

The **latent heat of vaporization** of a substance is the amount of thermal energy per kilogram of substance needed to change it from a liquid to a gas at its boiling point. Water has a heat of vaporization of 2160 kJ/kg. We must supply 2160 kJ of thermal energy to a kilogram of water at 100°C in order to change it from a liquid to a gas. Latent heats of vaporization for a variety of materials are given in Table 10-2. Generally, the latent heat of vaporization for a substance is considerably larger than its latent heat of fusion.

To change ice to water and water to steam, we must supply thermal energy to the substance. If we reverse the process, changing steam to water and water to ice, the substance must lose thermal energy. If 1 kg of steam condenses to form 1 kg of water, the steam loses 2160 kJ of thermal energy. Similarly, 1 kg of water must lose 320 kJ of energy in order to become ice at 0°C. The latent heat of fusion and vaporization of a substance define the

Table 10-2 Latent Heat of Fusion/Latent Heat of Vaporization for Common Substances

Substance	Latent Heat of Fusion (kJ/kg)	Latent Heat of Vaporization (kJ/kg)
Water	320	2160
Freon	—	156
Alcohol	104	853
Oxygen	14	213
Hydrogen	59	452
Nitrogen	26	201
Mercury	12	213
Iron	33	—
Lead	25	871
Silver	88	2336
Gold	64	1578
Copper	134	5069

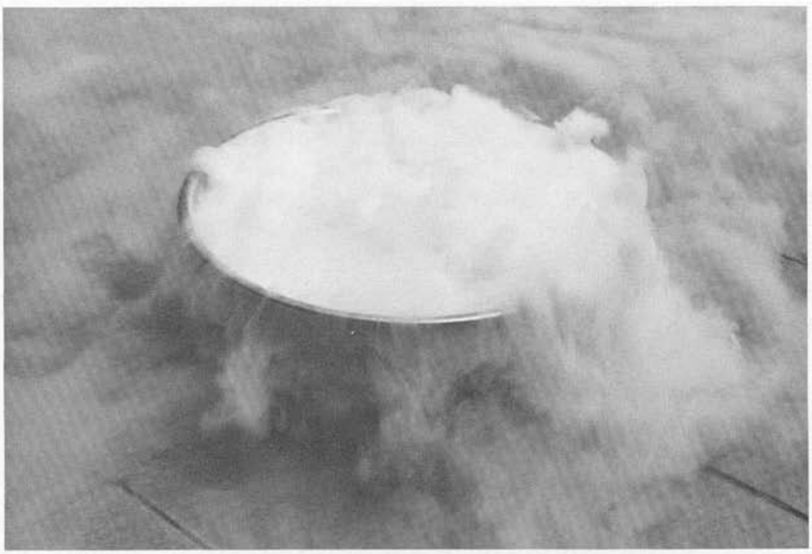


Figure 10-9
Some substances, like dry ice, change directly from a solid to a gas, absorbing energy along the way!

quantity of thermal energy involved in a change of state regardless of the direction in which the change is occurring. In one direction, the substance increases its thermal energy, acting as an energy receiver. In the other direction, the substance decreases its thermal energy, acting as an energy source.

Once again, giving a name like *latent heat* to the thermal energy transferred during a change of state does not really explain it. You may be wondering why the energy is needed and where it goes, since the temperature of the substance remains constant. In essence, the thermal energy is needed to change the average separation between molecules and, hence, the strength of their electrical bonds. As we see in Chapter 13, these bonds are stronger between some molecules and weaker between others, leading to the differences in latent heats found in Table 10-2.

SELF-CHECK 10D

Use the values given in Table 10-2 to determine the amount of thermal energy transferred when 0.5 kg of Freon (a substance used in refrigerators and air conditioners) changes from a liquid to a gas. Does Freon act as an energy source or energy receiver?

Change of State in Ice Boxes and Refrigerators

The thermal energy associated with a change of state has long been applied to the problem of preserving foods. Most foods spoil rapidly at room temperature. To retard spoilage, we place foods in cooler storage locations. Originally, families used underground cellars. Then came the icebox and, more recently, the refrigerator.

As its name implies, the icebox is simply a storage container that holds ice. The melting point of ice (0°C) is well below the temperature at which we like to store foods. As it melts, the ice acts as an energy receiver, absorbing 320 kJ of energy from the surrounding air for each 1 kg of ice melted. The air inside the icebox becomes cooler and the ice gradually melts. As long as fresh ice is added every few days to replace the ice that has melted, the inside of the box stays cool. The system is not very convenient, however, and the icebox has given way to the refrigerator.

Modern refrigerators also take advantage of the thermal energy associated with change of state. Instead of using ice as the energy receiver, refrigerators use a liquid called Freon, which changes to a gas as it absorbs energy from the surrounding air. The latent heat of vaporization of Freon is 156 kJ/kg, so each kilogram of Freon absorbs 156 kJ of energy.

The use of a liquid instead of a solid allows designers to circulate the Freon throughout the refrigerator. The complete cycling system for a typical

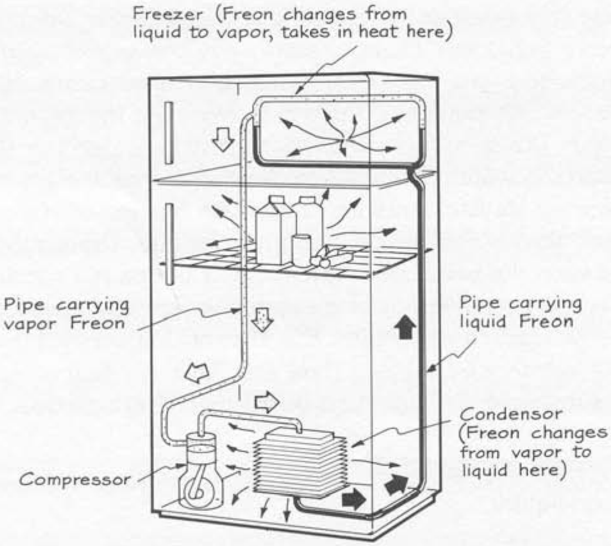


Figure 10-10 Liquid Freon is circulated through the refrigerator and freezer. It absorbs thermal energy, changing into a gas. A compressor at the bottom of the refrigerator converts the gaseous Freon back into a liquid, releasing energy. The thermal energy released at the condenser exceeds the thermal energy absorbed by the Freon. Freon is then recycled through the refrigerator.

refrigerator is shown in Figure 10-10. As the liquid circulates through the interior pipes, it absorbs thermal energy from the air inside the refrigerator. This cools the air and vaporizes the Freon. The gas is then pumped down to the condenser outside the refrigerator, where it is condensed back into a liquid. The thermal energy released by the gas as it changes back into a liquid is exhausted outside the refrigerator. (Most condensers are located at the bottom and toward the back of the refrigerator. The warm air you feel there is due to the energy released as the gas is converted back into a liquid.) The liquid Freon is then pumped back through the interior of the refrigerator to absorb more thermal energy.

How Much is Transferred?

It is a hot day and you want some cold tea. You have already made the tea and it is at room temperature. Now, how should you chill it? If you add a cooler substance, it will act as an energy receiver, absorbing thermal energy from the tea. From experience you know that ice is an effective coolant. But why not simply use cold water instead of ice? After all, the ice usually melts and dilutes the tea just as much as the cold water would. Why go to the trouble of making ice if cold water could serve the same function?

To answer this question, we need to determine the thermal energy lost or gained when a substance changes state. You already know that if you add cold water to the tea, you can determine the amount of thermal energy the cold water absorbs. All you need to do is to calculate the product of the mass of the cold water, the specific heat capacity of water, and the change in temperature of the cold water. We need to develop an equivalent expression for the thermal energy transferred when there is a change of state.

Thermal energy transferred (solid-liquid)
Latent heat of fusion
$TE_{(s\text{-}L)} = mH_f$
Mass

The latent heat of fusion and the latent heat of vaporization provide us with the expression we need. Both latent heats define the amount of thermal energy transferred per kilogram of a substance that changes state. Thus, the only variables needed to determine the **thermal energy transferred** are the mass of the substance changing state and its latent heat of fusion or vaporization. For a substance undergoing a **solid-liquid transition:**

$$\text{Thermal energy transferred (solid-liquid)} = \text{mass} \times (\text{heat of fusion})$$

For a substance undergoing a **liquid-gas transition:**

Thermal energy transferred (liquid-gas)
Latent heat of vaporization
$TE_{(L\text{-}G)} = mH_v$
Mass

$$\text{Thermal energy transferred (liquid-gas)} = \text{mass} \times (\text{heat of vaporization})$$

For our glass of tea, the amount of thermal energy that is absorbed by 0.050 kg of ice that completely melts is the product of the mass of the ice and the latent heat of fusion of water—(0.050 kg)(320 kJ/kg) = 16 kJ. The ice absorbs 16 kJ of energy as it melts.

SELF-CHECK 10E

How much thermal energy is released to the environment when 10 kg of water at 0°C freezes and becomes 10 kg of ice at 0°C?

Evaporation

Although a change in state occurs most rapidly when the substance is at its melting or boiling point, changes can occur more gradually. Puddles in the road gradually dry up. The water level in a flower vase gradually drops. **Evaporation** is the gradual change of a liquid into a gas that occurs at temperatures below the boiling point.

Evaporation is an important process in cooling the human body (Figure 10-11). When we become hot, either from exertion or from our surroundings, one of the body's mechanisms for reducing body temperature is perspiration. Water on the surface of the skin absorbs thermal energy from the body and evaporates. Removing this energy helps us keep cool.

A STEP FURTHER—MATH

REALLY COOL TEA!

Remember the problem of cooling the drink? Suppose we have two glasses of tea at room temperature (20°C), and we want to cool them. In one glass we place 0.050 kg of water at 0°C; in the second glass we place 0.050 kg of ice at 0°C. A short time later we measure the temperature of the two mixtures. The water-tea mixture is at 16°C, while the ice-tea mixture, in which the ice has melted completely, is at 2°C. As we know from experience, ice is a better coolant than cold water. To see why, we calculate the thermal energy absorbed from the tea in each glass. (For simplicity, we assume that no energy is transferred to the air in the room.)

When we added cold water, the temperatures of the water and the tea changed. Using the definition for the thermal energy transferred during a change in temperature, we find that the cold water absorbed 3.4 kJ of energy from the tea:

$$\begin{aligned} TE &= (m) \times (c) \times (\Delta T) \\ &= (0.050\ \text{kg})(4.2\ \text{kJ}/°\text{C} \cdot \text{kg})(16°\text{C} - 0°\text{C}) \\ &= 3.4\ \text{kJ} \end{aligned}$$

When we added ice to the second glass, two things happened. First, the ice melted. Secondly, the melted ice increased its temperature from 0°C to 2°C. Because both a change of state and a change in temperature occurred, we have to calculate the total thermal energy absorbed by the ice in two steps.

$$\begin{aligned} \text{Change in state:}\quad TE &= (m) \times (H_f) \\ &= (0.050\ \text{kg})(320\ \text{kJ}/\text{kg}) \\ &= 16\ \text{kJ} \end{aligned}$$

$$\begin{aligned} \text{Change in temperature:}\quad TE &= (m) \times (c) \times (\Delta T) \\ &= (0.050\ \text{kg})(4.2\ \text{kJ}/°\text{C} \cdot \text{kg})(2°\text{C} - 0°\text{C}) \\ &= 0.42\ \text{kJ} \end{aligned}$$

The total thermal energy absorbed by the ice as it melts and warms up to 2°C is 16.42 kJ, more than four times the thermal energy absorbed when cold water is added instead. Most of that energy went into changing the state of the ice. We use ice to cool not just because it supplies cold water, but because the melting process absorbs a tremendous amount of energy.

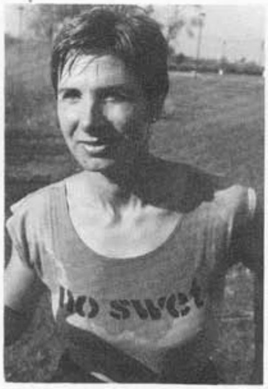


Figure 10-11
Evaporation is an important process in cooling the human body.

The thermal energy associated with evaporation is approximately the same as the latent heat of vaporization. Our body must supply 2160 kJ (2.16×10^3 kJ) of thermal energy per kilogram of water evaporated from the skin. Measurements show that the human body dissipates about 10.5×10^3 kJ of thermal energy per day. Using the latent heat of vaporization of water, the amount of perspiration that would have to be evaporated in order to remove this much thermal energy is $(10.5 \times 10^3 \text{ kJ})/(2.16 \times 10^3 \text{ kJ/kg}) = 4.86$ kg. This is equivalent to about 5 quarts of water! If evaporation were our only mechanism for cooling ourselves, we would need to drink 5 quarts of water per day just to replenish what is lost to perspiration. While evaporation is an important cooling mechanism, it is clearly not the only mechanism we use. We consider some of the other mechanisms in the next chapter.

In some respects we can view matter as an energy-storing device. The waters of the ocean store energy from the sun and maintain a mild temperature range within which life has evolved. Freon stores energy long enough to transfer it from inside to outside the refrigerator. Perspiration stores energy, removing it from our bodies and dispersing it throughout the atmosphere. The electrical potential energy and kinetic energy of billions of tiny molecules provide a convenient and efficient method for storing and transporting energy.

CHAPTER SUMMARY

Thermal energy is the energy stored internally in matter. It includes the electrical potential energy of the bonds holding the molecules together and the kinetic energy of molecular motion. The *state* of a substance—solid, liquid, or gas—is related to the amount of electrical potential energy stored in each molecular bond. The temperature of a substance is related to the average kinetic energy of its molecules.

Thermal energy that is being transferred from one place to another is called *heat*. The direction in which heat flows is always from the object at a higher temperature to the object at a lower temperature. *Temperature* is the perceived hotness or coldness of a substance and is measured with a thermometer. Temperature scales in use today include the *Celsius* scale, the *Kelvin* scale, and the *Fahrenheit* scale. These scales differ in the reference temperatures used to establish the scales' origins and in their degree sizes.

When a substance gains or loses thermal energy, it may change temperature. Two objects that undergo the same change in temperature may gain or lose different amounts of thermal energy. *Specific heat capacity* is defined as the amount of thermal energy required to change the temperature of 1 kg of a substance by 1°C. The greater the specific heat capacity of a substance, the more useful it is to us as a material with which to store energy. The *thermal energy* gained or lost when a substance *changes temperature* is given by:

$$\begin{matrix}\text{Thermal}\\ \text{energy}\end{matrix} = (\text{mass}) \times \begin{pmatrix}\text{specific}\\ \text{heat}\end{pmatrix} \times \begin{pmatrix}\text{change in}\\ \text{temperature}\end{pmatrix}$$

When a substance gains or loses thermal energy, it may change state. No temperature change occurs during a change in state. The *latent heat of* *fusion* of a substance is the amount of thermal energy required to change 1 kg of a substance from a solid to a liquid at its melting point. The *latent heat of vaporization* of a substance is the amount of thermal energy required to change 1 kg of the substance from a liquid to a gas at its boiling point. The *thermal energy* (TE) gained or lost when a substance *changes state* is given by:

(Solid-liquid) TE = mass × (heat of fusion)

(Liquid-gas) TE = mass × (heat of vaporization)

ANSWERS TO SELF-CHECKS

10A. The temperature of the hot water is higher than the temperature of the cold water. Since temperature is a measure of the average kinetic energy of the molecules, hot water has molecules with a greater average kinetic energy than cold water.

Ice is a solid and water is a liquid. Solids and liquids differ in the amounts of energy stored in the electrical bonds holding the molecules together. Water has more electrical potential energy stored in its molecules.

10B. Wood: specific heat = 1.7 kJ/°C · kg

$$\Delta T = 20 \text{ kJ}/1.7 \text{ kJ}/°\text{C} \cdot \text{kg}$$
$$\Delta T = 11.8°\text{C per kilogram of wood}$$

Aluminum: specific heat = 0.9 kJ/°C · kg

$$\Delta T = 20 \text{ kJ}/0.9 \text{ kJ}/°\text{C} \cdot \text{kg}$$
$$\Delta T = 22°\text{C per kilogram of aluminum}$$

The change in temperature of the aluminum is nearly twice that of wood. The aluminum spoon will reach a higher temperature.

10C. $\text{TE} = (m) \times (c) \times (\Delta T)$

$$= (0.02 \text{ kg})(0.9 \text{ kJ}/°\text{C} \cdot \text{kg})(20°\text{C} - 5°\text{C})$$
$$= 0.27 \text{ kJ}$$

10D. The latent heat of vaporization of Freon is 156 kJ/kg. Since we have just $\frac{1}{2}$ kg of Freon, we need $\frac{1}{2}$ of 156 kJ, or 78 kJ.

The Freon is changing from a liquid to a gas, so thermal energy must be supplied to the Freon. It acts as an energy receiver.

10E. TE = (mass) × (heat of fusion)

$$= (10 \text{ kg})(320 \text{ kJ/kg})$$
$$= 3200 \text{ kJ}$$

PROBLEMS AND QUESTIONS

A. Review of Chapter Material

A1. Define the terms listed below:
Thermal energy
Celsius scale
Kelvin scale
Fahrenheit scale
Absolute zero
Temperature
Specific heat capacity
Latent heat of fusion
Latent heat of vaporization
Evaporation

A2. With what form of energy is a change of state associated?

A3. With what form of energy is a change in temperature associated?

A4. How do temperature scales differ?

A5. Use the concept of specific heat capacity to explain why pizza sauce burns your mouth, while the crust, at the same temperature, does not.

A6. When substances at different temperatures are mixed, they exchange thermal energy until both substances reach a common temperature. In which direction is the thermal energy transferred: from the warmer substance to the cooler substance or from the cooler substance to the warmer substance?

A7. List the three variables that determine the amount of thermal energy transferred when a substance changes temperature but does not change state.

A8. Thermal energy is required to melt or vaporize a substance even though the substance remains at a constant temperature. Why is this energy required to change the state of a substance?

A9. What two variables affect the amount of thermal energy transferred when a substance changes state?

A10. What two changes provide evidence that thermal energy has been exchanged between substances?

B. Using the Chapter Material

B1. Use the molecular model to describe how the thermal energy stored in ice at $-50°C$ is different from the thermal energy stored in ice at $-20°C$.

B2. Use the molecular model to describe how the thermal energy stored in boiling water at 100°C is different from the thermal energy stored in steam at 100°C.

B3. Suppose you establish your own temperature scale. You choose the freezing point of water to be 25°X and the boiling point of water to be 125°X. Compare the size of the degree on your scale with the Celsius scale. How is your scale different from the Celsius scale?

B4. You have 0.1 kg of cooked apples sitting in a 0.1 kg aluminum pan. Both the apples and the pan are at the same temperature. In terms of the specific heat capacity of apples and aluminum, which is more likely to burn your hand?

B5. In the good old days, bedrooms were unheated. You could take hot objects to bed with you. Which would provide you with the most thermal energy: 1 kg of wood at 70°C or 1 kg of bricks at 70°C?

B6. The specific heat capacity of the human body is about 3.3 kJ/°C · kg. When you have a fever, how much thermal energy is needed to raise your temperature by 2°C? The average person has a mass of about 70 kg.

B7. After sitting in the sun all day, the inside of a car reaches a temperature of 50°C. The mass of air in the car is 1 kg. The car's air conditioner removes the air at 50°C and replaces it with air at 10°C. How much air must be removed and replaced to obtain a final temperature of 30°C?

B8. If you place your finger in water at 100°C, it will be burned. However, the burn will be much more severe if you place your finger in steam at 100°C. Why?

B9. How much energy is removed from a refrigerator when 0.3 kg of Freon change from a liquid to a gas?

B10. As water freezes, does the temperature of the air inside the freezer increase or decrease? What about when the ice melts?

B11. The human body can generate approximately 6000 kJ per hour during heavy physical activity. If the mass of the average person is 70 kg and the specific heat capacity of the human body is about 3.3 kJ/°C · kg, how much would your body temperature rise if no thermal energy was released?

C. Extensions to New Situations

C1. A major use of chemical potential energy is the heating of water for baths, dishes, and cooking. With some care we can decrease the energy needed for this purpose.

a. Water for baths or showers has a temperature of 40°C and is heated from 20°C. An average shower requires 20 kg less water than a bath. How much energy is saved by taking showers instead of baths?

b. Usually the hot water used for a bath or shower is allowed to go down the drain, thus keeping the sewers warm. A bathtub will hold 165 kg of water. If this water were allowed to cool from 40°C to 20°C before leaving the tub, how much energy would be transferred to the air in the bathroom? How could this procedure decrease your utility bill?

C2. The amount of energy needed to heat an oven and its contents can be determined from the information in this chapter. Typically, an oven contains 0.4 kg of air, which is raised from 20°C to 170°C. The oven is enclosed in 50 kg of metal (specific heat capacity = 0.5 kJ/°C · kg). An average dinner includes 2 kg of food (average specific heat capacity = 3.4 kJ/°C · kg).

a. How much energy is needed to heat the air in the oven?

b. How much energy is needed to heat the metal in the oven from 20° to 170°C?

c. How much energy is needed to heat the food?

d. What is the total energy required to get the contents and oven from 20°C to 170°C? (As we will see in Chapter 11, more energy is actually needed for cooking.)

e. A microwave oven heats only the food, not the oven or air. How much less energy is required with a microwave oven?

C3. Many items can be cooked in small, countertop ovens rather than the larger stove ovens. A countertop oven holds 0.01 kg of air, while a regular oven holds about 0.4 kg.

a. How much energy is required to get the air in each oven from 20°C to a typical cooking temperature of 170°C?

b. Which oven is more practical for small meals?

C4. For lack of something better to do, you wish to turn 2 kg of ice at 0°C into 2 kg of steam at 100°C.

a. How much energy is required to change the ice at 0°C into water at 0°C?

b. How much energy must you supply to increase the water's temperature from 0°C to 100°C?

c. What is the energy required to change the water at 100°C into steam at 100°C?

d. Which of the above processes took the most energy? The least energy?

e. What is the total energy required?

C5. In recent years aluminum beverage containers have replaced steel and glass containers. One of the reasons for this change is related to the masses and specific heat capacities of the containers. Typical values are: aluminum, mass = 0.02 kg, specific heat capacity = 0.9 kJ/°C · kg; glass, mass = 0.20 kg, specific heat capacity = 0.8 kJ/°C · kg; steel, mass = 0.05 kg, specific heat capacity = 0.5 kJ/°C · kg. Beverages stored in these cans are usually cooled from 20°C to 5°C.

a. How much energy must be removed from the beverage? Assume the mass of the beverage is 0.35 kg.

b. How much energy must be removed from each of the containers?

c. What is the total energy (container plus beverage) removed from each type of container?

d. If you were the proprietor of a store and had to pay electric bills to cool drinks, which container would you prefer?

C6. If you drop a snowball from a high enough location, it will melt completely upon impact with the ground.

a. Neglecting the frictional interaction with the air, explain why the snowball can melt.

b. How much energy is needed to melt a 2 kg snowball completely?

c. Which of the following heights will be great enough to melt the 2 kg snowball: 4 m, 400 m, or 40,000 m? (Remember the answer to (b) is in kilojoules!)

d. What would happen to the water after it melted if the snowball were dropped from a height greater than that needed to melt it?

C7. The coolant used in refrigerators, Freon, comes in many different varieties. Each type has a slightly different chemical form and has different thermal properties. For example, Freon-11 has a boiling point of 23.8°C, a specific heat capacity of 1.29 kJ/°C · kg, and a latent heat of vaporization of 181 kJ/kg; Freon-14 boils at 3.8°C and has a specific heat capacity and latent heat of vaporization of 1.02 kJ/°C · kg and 136 kJ/kg, respectively.

a. Suppose the inside of your refrigerator is kept at a constant 4°C. In what state would each of these two Freons be when inside the refrigerator?

b. How much energy would 1 kg of each absorb at this temperature if each entered the refrigerator as a liquid?

c. Which would be a better coolant for a refrigerator?

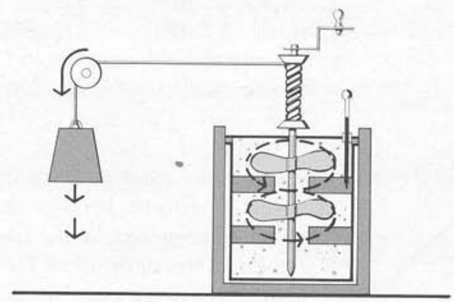
C8. In a famous thermal energy experiment, Humphry Davy kept all items in a laboratory at 0°C. Then, handling two blocks of ice by long sticks so that he transferred no body heat to them, he rubbed the ice blocks together. The ice melted. Why should the ice melt?

C9. James Joule performed experiments to show that gravitational potential energy could be converted into thermal energy. In one such experiment, he connected a 26-kg mass to a mixer, which was submerged in water (Figure 10-C9). The mass was dropped 1.5 m. Each drop resulted in a 0.015°C increase in the temperature of the water.

a. Why should we expect the water temperature to increase?

b. If all the gravitational potential energy of the mass was transferred to the water, how much thermal energy was gained by the water?

c. Joule stated that this experiment showed that heat was a form of energy. Why could he make that statement?



D. Activities

D1. Use specific heat capacities, latent heats of vaporization, and latent heats of fusion to estimate the total amount of energy needed to prepare your food for a day.

D2. Design a system that would store energy coming through a window on a sunny day. Explain how you would choose the material for the system.