11

Heat and Temperature

KEY IDEA

Heat is a form of energy that can be transferred from one object to another because of a difference in temperature.

PHYSICS AROUND US . . . A Hot Summer's Day

The midday sun blazes as you walk across the blistered black asphalt toward your car. You wipe sweat from your brow, grab your keys, open the door, and feel a blast of hot air. As you slide inside, the seat sears your skin. The metal buckle of the seat belt burns. The suffocating heat assaults you.

But relief is at hand. The air conditioner kicks in, and within seconds refreshing waves of cool air wash over you. It feels so good!

Every day of our lives, in ways both big and small, heat flows around us. Heat flows from hotter objects such as door handles and seats to our cooler skin. Also, heat flows from our hot skin to the cooler air pouring out of the air conditioner. Little wonder, then, that scientists and engineers are fascinated by the movement of heat.

TEMPERATURE

In Chapter 9 we explored different phases of matter—solids, liquids, and gases—as well as changes from one phase to another. In order to understand the causes of these dramatic transformations, we first have to examine the intertwined phenomena of heat and temperature. The differences between these two concepts were not clearly recognized until the early nineteenth century. However, understanding these differences led to important progress in our understanding of energy and its practical application in heat engines and machines.

Work, energy, inertia, momentum, force—over and over again in physics we encounter everyday words that have an exact scientific meaning. In this chapter we explore two more common terms, *temperature* and *heat,* that are often used interchangeably in day-to-day conversation, but which have very different scientific definitions. As you study this chapter, be sure to pay careful attention to the important distinctions between temperature and heat.

Temperature Defined

Temperature is one of the most familiar physical variables in our lives. You probably see or hear temperature measurements a dozen times a day. Weather reports document the current temperature and predict high and low temperatures for the coming week. You take your own body's temperature when you're feeling sick. And you set the temperature on your thermostat, your oven, and your refrigerator. As we'll see, temperature can be defined in different ways. But at the simplest level, **temperature** is a quantity that indicates how hot or cold an object is relative to a standard value. And, as we'll soon see, at a deeper level temperature is related to the speed at which the atoms and molecules in a substance are moving.

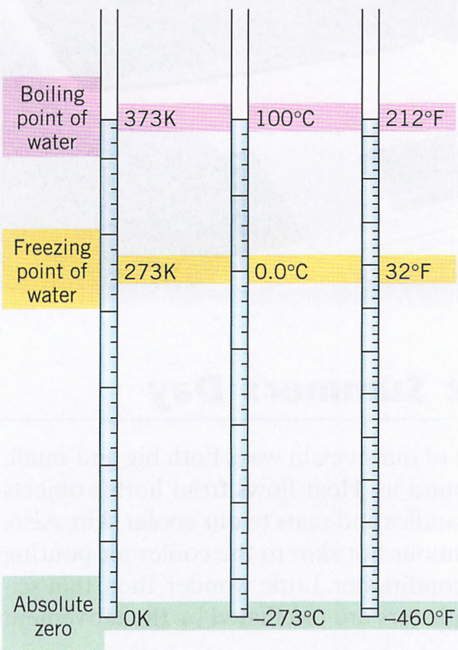


FIGURE 11-1. The Fahrenheit, Celsius, and Kelvin temperature scales compared.

Temperature scales provide a convenient way to compare the temperatures of two objects. Several temperature scales have been devised over the centuries. All of these numerical scales are somewhat arbitrary, but every scale requires two easily reproduced temperatures for calibrating the measuring instrument (the thermometer). Two especially convenient standards are the freezing and boiling points of pure water. These reference temperatures are used today in the Fahrenheit scale (32°F and 212°F for freezing and boiling, respectively; see Figure 11-1) and the Celsius scale (0°C and 100°C for freezing and boiling, respectively). The Kelvin temperature scale, most commonly used in scientific research, also uses 100-degree increments between the temperatures of freezing and boiling water, but it defines 0 kelvin as **absolute zero,** which is the coldest possible temperature. Note that temperatures in the Kelvin scale are reported as kelvins (K), not as degrees Kelvin, for historical reasons we won't go into here.

Absolute zero is the temperature at which it is impossible to extract any energy at all from the vibrations of atoms or molecules. The temperature of absolute zero is approximately −273°C or −460°F. It turns out, therefore, that freezing and boiling of water occur at about 273 and 373 kelvins, respectively (Figure 11-1). We will come back to the idea of absolute zero in a little while; for now, we note that it is impossible to reach a temperature of absolute zero. Physicists have come very close, to within one billionth of a degree. However, although an object can be at zero degrees or at a negative temperature on the Fahrenheit or Celsius scales, an object cannot achieve zero or a negative temperature on the Kelvin scale.

It's important to realize that while all of the conventional temperature scales involve the freezing and boiling points of water in some way, there is no reason that other fixed points could not be used. It's all just a matter of convenience. For example, until well into the twentieth century, brewers and distillers in parts of Europe used a temperature scale based on the freezing and boiling points of alcohol!

It's often necessary to convert from one temperature scale to another. American travelers, for example, often have to convert from degrees Celsius (used in most of the rest of the world) to degrees Fahrenheit. This conversion requires the following formula:

$$T \text{ (in °F)} = [1.8 \times T \text{ (in °C)}] + 32$$

The 1.8 in this formula reflects the fact that the Celsius degree is larger than the Fahrenheit degree by a factor of $\frac{9}{5}$ (= 1.8). There are 100 Celsius degrees between the freezing and boiling points of water, but 180 Fahrenheit degrees. Therefore the Fahrenheit degree is smaller. The 32 in the conversion formula reflects the fact that water freezes at 32°F but 0°C. To convert the opposite way, from Fahrenheit to Celsius, the formula is

$$T \text{ (in °C)} = \frac{T \text{ (in °F)} - 32}{1.8}$$

Scientists often convert between the Celsius and Kelvin scales, a simple process that just involves adding or subtracting 273:

$$T \text{ (in °C)} = T \text{ (in K)} - 273$$
$$T \text{ (in K)} = T \text{ (in °C)} + 273$$

Remember that all of these temperature scales measure the exact same phenomenon. They just use a different number scale to do so, much like measuring a distance in meters versus inches. Looking at Temperature on page 230 illustrates the range of temperatures physicists often deal with.

What To Wear?

EXAMPLE 11-1

You awake in Paris to a radio announcer forecasting a high temperature of 32°C. Should you wear gloves and an overcoat?

SOLUTION: Apply the equation to convert temperature from Celsius to Fahrenheit:

$$\begin{aligned} T \text{ (in °F)} &= [1.8 \times T \text{ (in °C)}] + 32 \\ &= (1.8 \times 32) + 32 \\ &= 89.6\text{°F} \end{aligned}$$

Looks like you won't need your overcoat today!

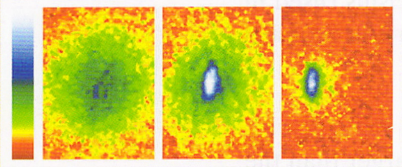
Thermometers

A **thermometer** is a device used to measure temperature. Most thermometers display temperature either digitally or on a numbered scale. Thermometers work by incorporating a material whose properties change with temperature. For example, many materials expand when heated and contract when cooled, and this

Looking at Temperature

Strange things happen to materials at very low and very high temperatures. We can't reach absolute zero, but we can come awfully close, to a billionth of a degree. Matter can exist in odd forms at these conditions, where atoms barely move at all. At 4 kelvins, helium becomes a liquid with unusual properties, such as spontaneously flowing up the walls of its container. At very high temperatures, solid rock melts and turns to lava; air conducts electricity, as in a lightning bolt; and hydrogen atoms fuse together, producing the energy in the Sun. Strange things indeed!

10^{-9} K

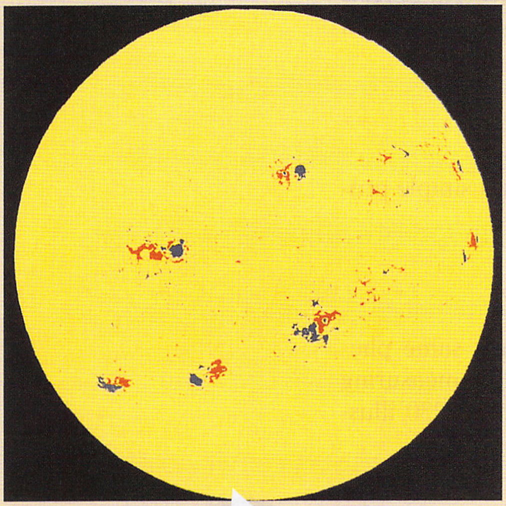


Bose-Einstein condensate, 2 billionths of a degree

10^{1} K

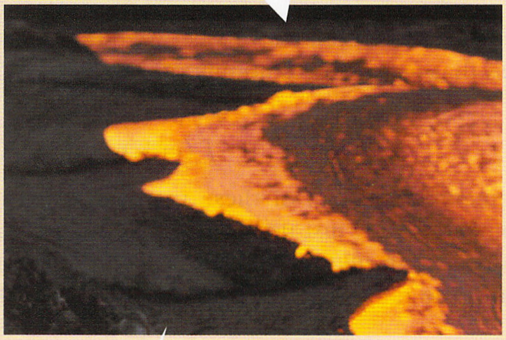
Liquid helium, 4 K

10^{7} K



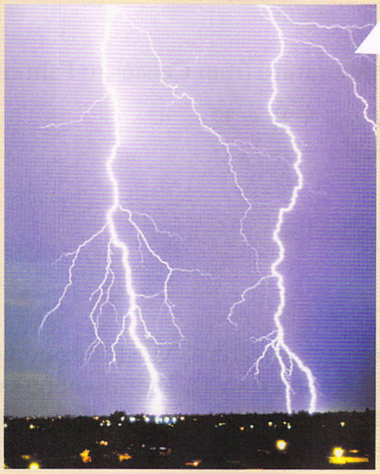
Center of the Sun, 15,000,000 K

10^{3} K



Lava, 2000 K

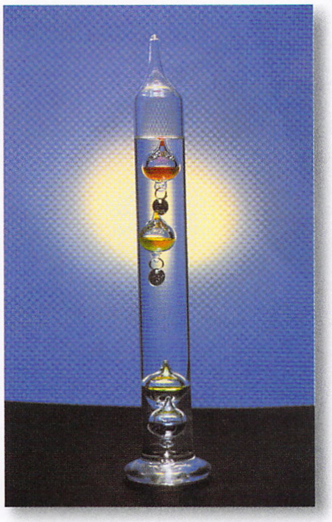
10^{5} K



Lightning, 30,000 K

can be used to gauge temperature. In the old-style mercury thermometer, a bead of mercury expands with increasing temperature into a thin glass column; you read the height of the mercury against a scale marked in degrees. Many other (much safer) thermometers rely on changes in the electric properties of a temperature sensor, called a *thermocouple*.

One of the most visually intriguing types of thermometers, invented by Galileo Galilei in the early 1600s, uses changes in liquid density as a function of temperature. The Galilean thermometer consists of a large sealed flask with a liquid that changes density as it is heated. Suspended in this liquid are dozens of small numbered weights, each of a slightly different density. At low temperature, most of the weights rise to the top of the flask. As the temperature increases, the denser weights sink one by one to the bottom. The temperature is read simply as the lowest number on the weights that remain floating at the top of the thermometer.



(*a*)

(*b*)

(*a*) A Galilean thermometer depends on the changes in density of a liquid with changes in temperature. (*b*) A modern digital fever thermometer depends on the changes in electrical properties of a crystal with changes in temperature.

The Atomic Basis of Temperature

We can use a thermometer to measure temperature, but what exactly is it that we are measuring? Temperature is actually a measure of the magnitude of the average speed of atoms and molecules, including their vibrations—that is, the kinetic energy of the particles that make matter. At lower temperatures a material's molecules have less kinetic energy than at higher temperature (Figure 11-2). On the other hand, if two materials are at the same temperature, then their atoms and molecules have the same kinetic energy. The important point here is that:

The higher an object's temperature, the faster its atoms or molecules move.

From this fact, it's clear that temperature is not a measure of the *total* kinetic energy of atoms and molecules in a substance. After all, if you double the amount of material, you double the total kinetic energy, even though the temperature remains the same.

This distinction between temperature and total molecular kinetic energy reveals an important point about using a thermometer. We use a thermometer to measure the temperature of a substance, such as a pot of water, the atmosphere,

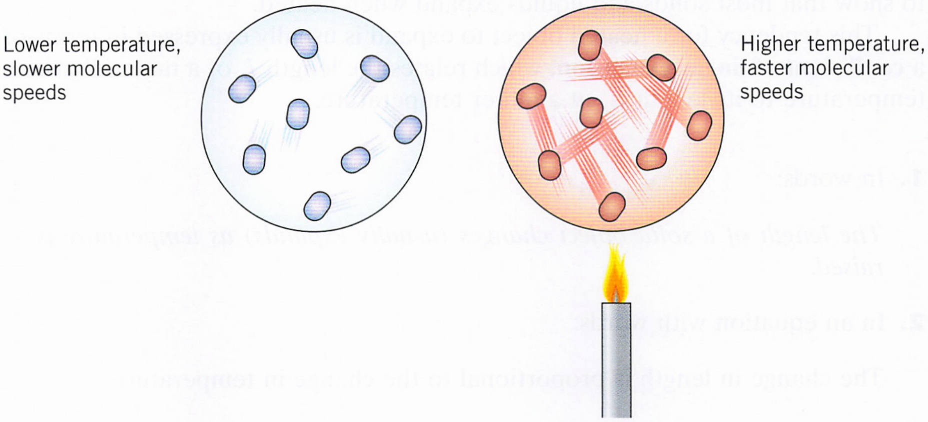


Figure 11-2. Atomic view of temperature (cold and hot). Temperature measures the average kinetic energy of the molecules that make up an object.

or your mouth. It works like this: when you put a thermometer in your mouth, high-energy molecules in your body collide with lower-energy molecules in the thermometer. In the process, the molecules in the thermometer pick up energy and start to move faster. This process goes on until the molecules in the thermometer have the same average kinetic energy as those in your mouth, at which point no further change in temperature occurs. In other words, the thermometer is actually measuring its own temperature! The thermometer works perfectly well if the surroundings have much more mass (and thus more molecular kinetic energy) than the thermometer itself, because the thermometer can drain away some kinetic energy from that object without changing the object's temperature significantly. However, a thermometer can't be used to measure the temperature of an object such as a single drop of liquid, which is much less massive than the thermometer itself. In that case, the thermometer changes the temperature of the drop it's supposed to be measuring.

In the nineteenth century, scientists argued that there was a lowest possible temperature, called *absolute zero,* which would be reached when atoms came to a standstill. An atom, after all, cannot have a speed less than zero. With the advent of quantum mechanics (see Chapter 22), we recognized that atoms can never actually stop moving but that they can move into a state from which it is impossible to extract more energy. This is the modern definition of absolute zero, which is denoted as 0 K. Note that zero degrees in the Fahrenheit and Celsius scales does not correspond to any such milestone at the atomic level.

Thermal Expansion

Some important evidence for the link between temperature and atomic structure comes from the phenomenon called "thermal expansion." For example, think about the structure of a crystalline solid, as we described in Chapter 9. In such a solid, the atoms are separated from one another by springlike chemical bonds that vibrate about their equilibrium positions. As the temperature of the solid is raised, these vibrations become more energetic so that departures from the equilibrium positions become more pronounced. These motions, in turn, mean that each atom requires more room for its motion. As a result, the size of the crystal increases to accommodate this requirement. A similar argument could be given to show that most solids and liquids expand when heated.

This tendency for a heated object to expand is usually expressed in terms of a **coefficient of linear expansion,** which relates the length L of a material at one temperature to its length L' at another temperature.

1. In words:

The length of a solid object changes (usually expands) as temperature is raised.

2. In an equation with words:

The change in length is proportional to the change in temperature.

3. In an equation with symbols:

$$L' = \mathrm{L}\,(1 + \alpha\,\Delta T)$$

where ΔT is the temperature change and α (the Greek letter alpha) is the coefficient of thermal expansion, a number that is different for different materials (Figure 11-3). This equation can also be written

$$\frac{\Delta L}{L} = \alpha \, \Delta T$$

where $\Delta L = L' - L$ is the change in length.

An important exception to this rule is water, which actually shrinks slightly when its temperature is raised from 0°C to 4°C, but expands for temperatures higher than 4°C. This fact is very important in nature because the shrinkage means that cold water is slightly more dense than warm water. As a result, when ponds of water begin to freeze over in the winter, the cold water at the surface sinks before turning to ice, allowing slightly warmer water to take its place. After all the water in the pond has reached 4°C, the water at the surface turns to ice first, since it floats on the denser 4°C water below it. The ice acts as further insulation for the water left in the pond. This makes it harder for ponds and lakes to freeze solid, which would be disastrous for the survival of fish and other aquatic organisms. Since biologists conjecture that life on Earth began in small pools and shallow lakes on the young planet, the thermal properties of water may have been an important factor in the early evolution of life.

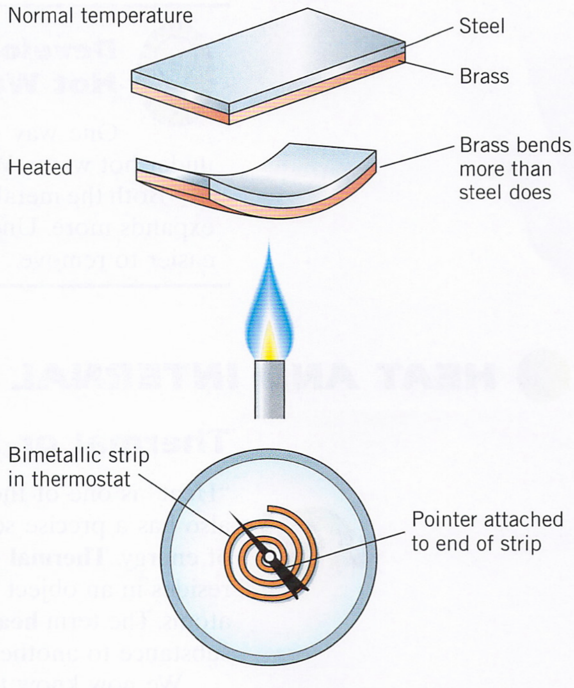


FIGURE 11-3. The thermostat in your home relies on the fact that different materials undergo different thermal expansions. At the heart of the thermostat is a bimetallic strip made of brass and steel. When heated, the brass expands more than the steel, causing the strip to bend. That bending, in turn, opens or closes an electrical switch that turns your furnace on or off.

A Bridge in the Summer

EXAMPLE 11-2

The coefficient of expansion for steel is 0.000056 per degree C. A particular bridge is 100 meters long when the temperature is 0°C. How long will that bridge be in the summer, when the temperature climbs to 40°C (over 100°F)?

SOLUTION: Putting the numbers into the thermal expansion equation, we find that the length will be

$$\begin{aligned} L' &= L\,(1 + \alpha\,\Delta T) \\ &= 100 \text{ m} \times [1 + (0.000056°\text{C}^{-1} \times 40°\text{C})] \\ &= 100.22 \text{ m} \end{aligned}$$

The length of a steel bridge, in other words, can change quite a bit from one season to the next—in this case by 22 cm. (almost 9 inches). Engineers take this fact into account by incorporating expansion joints into the bridge structure, like the one shown in the photo. As temperatures rise or fall these joints close together or open up. If the joints weren't there, the bridge might buckle on the first hot days of summer. ●

Expansion joints on a bridge allow the steel supports to expand in hot weather without buckling.

Develop Your Intuition: Hot Water and Bottle Caps

One way of getting a hard-to-remove cap off a bottle is to hold it under hot water. Why does this work?

Both the metal cap and the glass bottle expand when heated, but the metal expands more. Under the hot water, the cap expands, becomes looser, and is easier to remove.

HEAT AND INTERNAL ENERGY

Thermal or Internal Energy

"Heat" is one of those words like "energy" that has an everyday meaning, but also has a precise scientific definition. Physicists distinguish between two kinds of energy. **Thermal energy,** sometimes called **internal energy,** is the energy that resides in an object because of the motion and interactions of its molecules and atoms. The term **heat** is reserved for energy that is transferred from one body or substance to another due to a temperature difference.

We now know that atoms and molecules, the minute particles that make up all matter (see Chapter 9), constantly move and vibrate and, therefore, possess kinetic energy. If molecules in a material move more rapidly, they have more kinetic energy and they are capable of exerting greater forces on each other in collisions.

If you touch an object whose molecules are moving with greater kinetic energy than the molecules in your hand, the collisions that occur between molecules will transfer some of that energy to the molecules in your hand. As a result, you perceive the object to be hot. By the same token, if an object feels cold, then the molecules in your hand have greater average kinetic energy than the molecules in that object. What we call "heat" in everyday speech, therefore, such as the heat of boiling water, is actually a transfer of thermal (internal) energy, which is the kinetic energy of wiggling atoms and molecules. Since heat is energy in transit from a substance at a higher temperature to another substance at a lower temperature, matter itself does not contain heat.

By contrast, all objects, whether scalding hot or frozen solid, have internal energy—the total energy of all the atoms and molecules. That internal energy represents the sum of all the molecular motions (vibrations, rotations, and other movements), as well as energies associated with the interatomic forces that hold groups of atoms and molecules together. Any time a substance absorbs or emits heat, the internal energy of that substance changes. Usually, a change in internal energy causes a corresponding change in the material's temperature. Under special circumstances, however, as when ice melts at 0°C or water boils at 100°C, the absorbed heat helps to break interatomic bonds while the temperature remains constant—a phenomenon called a *change of phase* (see Chapter 9).

Heat

The most familiar everyday attribute of temperature is that it tends to even out. A hot cup of coffee gradually cools, while a glass of cold water slowly warms up to room temperature. In each case, and in hundreds of other occurrences in your everyday life, some of a material's molecular kinetic energy is transferred from one substance to another as temperature evens out, a process that involves what we have defined as heat.

The transfer of energy by atomic collisions goes on all the time because atoms never sit still. If you've ever tried to warm a house during a cold winter day you've experienced this process. The thermal energy in the house gradually moves (as heat) to the cooler outside surroundings. Then if you turn off the furnace, the house begins to get cold. The only way to stay warm is to keep the furnace on.

Like a furnace, our bodies constantly convert chemical energy to thermal energy to maintain our core body temperature close to 98.6°F (37°C). Both your furnace and your body produce thermal energy on the inside, and that energy inevitably flows to the cooler outside as heat.

Heat is often measured in a unit called the *calorie,* which is defined as the amount of heat required to raise the temperature of 1 gram of water by 1 degree Celsius. The more familiar "Calorie" (with a capital *C*) is the unit used to measure stored energy in food and is equal to 1000 calories (with a lowercase *c*), or 1 kilocalorie. Since heat is a form of energy, it can also be measured using the SI unit joule or kilojoule. In the English system, heat energy is measured using the British thermal unit (BTU), which is defined as the amount of energy needed to raise the temperature of 1 pound of water by 1 degree Fahrenheit (Figure 11-4). Be sure to use the proper conversion factors when dealing with all these different common units for heat.

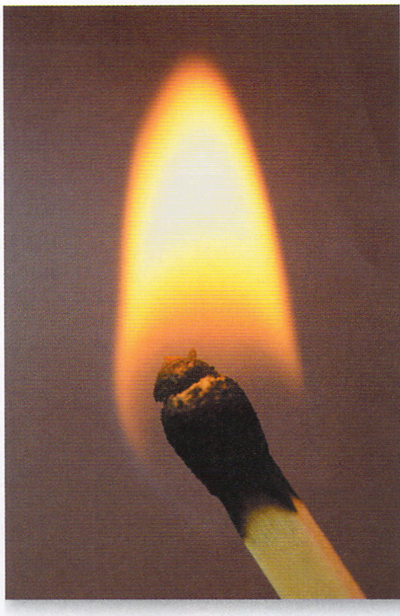


FIGURE 11-4. Comparison of heat units. One BTU is about the energy of a kitchen match.

Heat Capacity and Specific Heat

Heat capacity is a measure of how much heat an object can absorb. An object with a large mass has a higher heat capacity than an object of smaller mass. If we want to look at how much heat a particular kind of material can absorb, we need its specific heat. **Specific heat** is a measure of the ability of a material to absorb heat. It is defined as the quantity of heat required to raise the temperature of 1 gram of that material by 1°C. Water displays the largest specific heat of any familiar substance; by definition, 1 calorie is required to raise the temperature of 1 gram of water by 1°C. By contrast, you know that metals heat up quickly, so a small amount of heat can cause a significant increase in the metal's temperature.

Think about the last time you boiled water in a copper-bottom pot. It doesn't take long to raise the temperature of an empty copper pot to above the boiling point of water because copper, like most other metals, can't absorb much heat without having a large rise in temperature. In fact, 1 calorie of heat raises the temperature of a gram of copper by about 10°C. But water is a different matter; it must absorb 10 times more heat per gram than copper to raise its temperature by the same amount (Figure 11-5). Thus, even at the highest stove setting, heating a pot of water to boiling can take several minutes. This ability of water to absorb or release large amounts of heat plays a critical role in the Earth's climate, which is moderated by the relatively steady temperature of the oceans.

FIGURE 11-5. Heat capacities of water and copper. The pot gets hot quickly, so most of the heat goes to boil the water.

Connection
The Heat Capacities of the Land and Oceans

Earth's climate is affected by many factors, but one important ingredient is the thermal energy stored in the oceans. During warm weather, sunlight falling on the ocean raises the water's temperature. Since water has a high specific heat and the ocean has a huge amount of mass, it takes a long time for the ocean to warm up and a correspondingly long time for it to cool off. Materials on land, however, have a lower specific heat and only a relatively thin layer of land surface is heated by the Sun. Thus, the land warms up and cools off much more quickly than the ocean.

Weather patterns along an ocean coast are affected by the different heat capacities of land and water.

Several regular features of Earth's weather are a direct result of the relatively high heat capacity of Earth's oceans. For example, the lands surrounding the Indian Ocean experience a seasonal wind pattern known as the *monsoon.* During the summer, the land on the Indian subcontinent warms up quickly while the ocean warms more slowly. The warm air over the land rises, drawing cooler air in from the ocean (Figure 11-6*a*). These winds carry a great deal of moisture after traveling over the ocean. When they rise over the heated land, they produce clouds and regular rainfall (the "rainy season").

On the other hand, in winter the ocean cools more slowly than the land. Thus, in winter the air over the ocean is warmer than the air over the land, and the wind pattern reverses. Warm air rises over the ocean, drawing in cooler, drier air from over the land (Figure 11-6*b*). The rains stop and a 6-month dry season begins. In this way, the weather experienced over a large part of our planet can be traced directly to the high heat capacity of water. ●

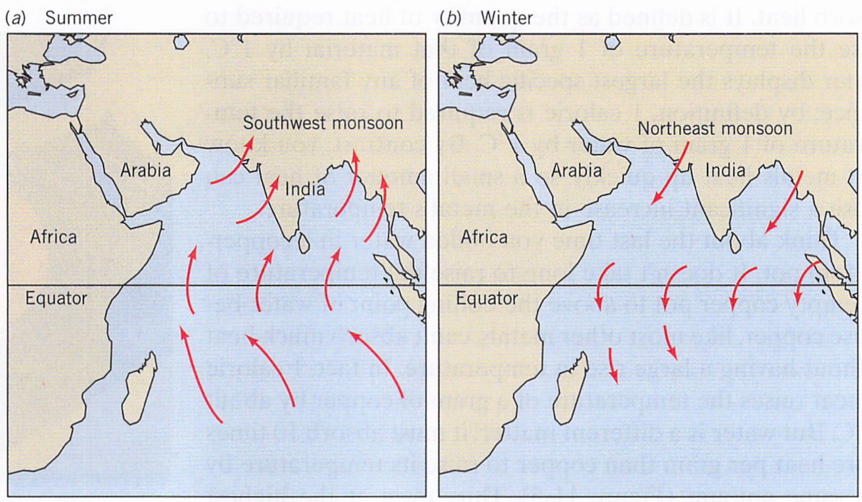


FIGURE 11-6. The Indian monsoon has a wet season and a dry season. (*a*) In summer, wet winds blow onshore toward the warmer land. (*b*) In winter, dry winds blow offshore toward the warmer ocean.

Develop Your Intuition: At the Beach

If you spend time along the beach, you have probably noticed that there is usually a breeze blowing out to sea in the early morning, but that the breeze calms down during the day and then blows in toward the land by late afternoon. Why does this happen?

At night, the land cools off. In the morning, then, the air over the ocean is warmer and it rises, drawing the cooler air from the land to replace it (Figure 11-7*a*). As the day passes, the land warms up, eventually becoming warmer than the sea. Then warm air rises over the land, and the (now relatively cooler) air from the ocean moves in to replace it (Figure 11-7*b*). The midday calm occurs when the two temperatures are about the same. Note that the same pattern of onshore and offshore breezes occurs on all coasts, but may happen earlier or later in the day, depending on the season and the location on Earth.

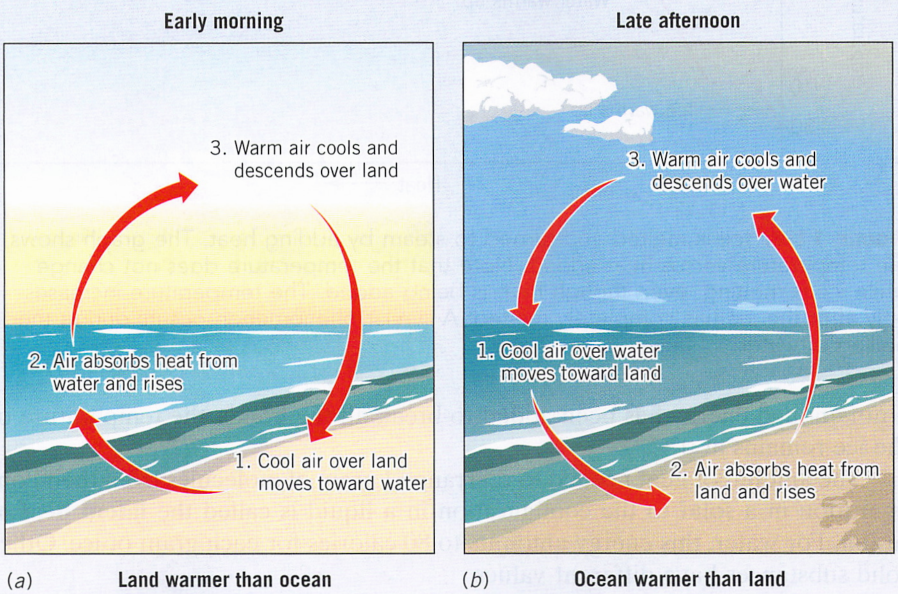


FIGURE 11-7. At the beach. (*a*) In the morning, the cool winds blow offshore; (*b*) by the afternoon and evening, the breeze has shifted to onshore.

Change of Phase

Specific heat gives us a good indication of how well any substance—gas, liquid, or solid—absorbs heat. However, if a change from one phase to another is involved, the situation is more complicated.

Suppose we take an ice cube from the freezer and transfer a small amount of heat energy to it each second. For a while, the added heat causes the temperature of the ice to rise, as shown in Figure 11-8. The molecules in the ice move faster and faster, but remain near their equilibrium positions. When the temperature reaches 0°C (32°F), however, another process comes into play. As heat energy flows into the ice, the bonds that hold the solid ice together begin to break.

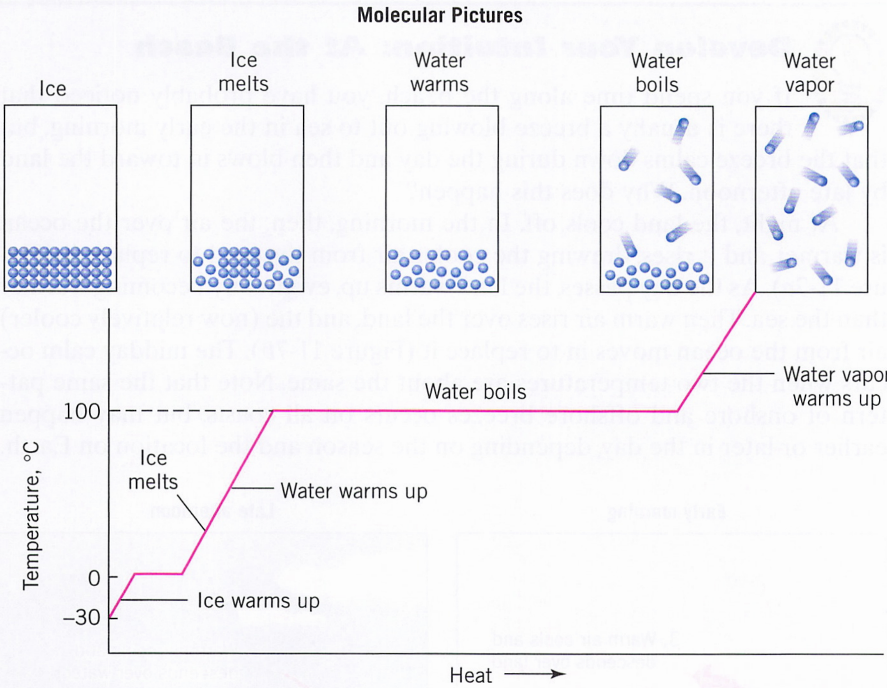


FIGURE 11-8. Ice is melted and turned to steam by adding heat. The graph shows the temperature versus heat added. Note that the temperature does not change while ice is melting, even though heat is being added. The temperature increases only after the ice has completely melted. A similar phenomenon occurs during the boiling of water.

Until enough energy has been added to break all the bonds, the temperature of the ice remains at 0°C.

The amount of heat needed to rearrange 1 gram of molecules from their configuration in a solid to the configuration in a liquid is called the **latent heat of fusion.** For water, this energy amounts to 80 calories for each gram of ice. Other solid substances have different values.

Once the conversion to a liquid is complete, the added heat begins to raise the temperature of the liquid water, as shown in the figure. (The fact that the slopes of the lines in Figure 11-8 are different for ice and liquid water is a reflection of the fact that solid ice and liquid water have different specific heats.) The molecules jitter faster and faster as the temperature rises. At 100°C (212°F), another phase transition takes place as the liquid water changes to water vapor. Once again, the liquid temperature remains constant while enough energy is added to the system to effect the change of phase. The amount of heat required to convert 1 gram of a liquid to a gas is called the **latent heat of vaporization** and is 540 calories per gram for water. Once all the water is in vapor form, there are no more changes of phase, and the normal increase of temperature with added heat goes on.

Note that the reverse processes—removing heat from steam to make liquid water and removing heat from liquid water to make ice—proceed in exactly the reverse order. The temperature remains fixed while enough energy is extracted from the system to allow the molecules to enter their new phase. Only when the change of phase has been completed does the temperature start to fall again.

Develop Your Intuition: Evaporation

People sweat, and the more active they are, the more they sweat. We've all seen pictures of athletes during a time-out with perspiration pouring off their bodies, soaking their uniforms. Why do people sweat? What purpose does it serve?

You can think about evaporation as a change of phase from liquid to gas. When water from your body's sweat glands comes through the pores of your skin to the surface, the water evaporates; that is, it turns from liquid water to water vapor. As it does so, the water absorbs heat from your body in an amount equal to the latent heat of vaporization. In other words, the evaporation of sweat from your skin acts to cool your body and keep your overall body temperature within normal bounds.

Many animals do not have sweat glands and cannot perspire from the skin. Instead, they must find other ways to cool their body temperature. For example, dogs pant heavily, allowing evaporation through the mouth and lungs.

In hospitals, patients running a high fever are sometimes given a rubdown with rubbing alcohol. Water has a higher heat of vaporization than alcohol, so it absorbs more heat per gram than alcohol does. However, alcohol evaporates very rapidly and so lowers the body temperature more quickly.

Physics in the Making

The Nature of Heat

What is heat? How would you apply the scientific method to determine its characteristics? That was the problem facing scientists 200 years ago.

They found that, in many respects, heat behaves like a fluid. It flows from place to place and seems to spread out evenly, like water that has been spilled on the floor. Some objects soak up heat faster than others, and many materials seem to swell up when heated, just like wood swells when it absorbs water. Thus, in the eighteenth century, after years of observations and experiments, many physicists mistakenly accepted the theory that heat is an invisible fluid, which they called "caloric." Supporters of the caloric theory of heat claimed that the best fuels, such as coal, are saturated with caloric, and they thought ice is virtually devoid of the substance.

Eventually, the caloric theory of heat was discarded as new observations failed to bear out the theory's predictions. In particular, the practical experience of machinists just did not support the idea of heat as a fluid. For instance, if heat is a fluid, then each object must contain a fixed quantity of that substance. However, Benjamin Thompson (later Count Rumford, 1753–1814), an eighteenth-century American who spent some time as a cannon maker, discovered that the amount of heat generated during cannon boring had nothing at all to do with the quantity of brass being drilled. Sharp tools, he found, cut brass quickly with minimum generation of heat, while dull tools made slow progress and produced prodigious amounts of heat. The amount of heat that could be generated seemed to be boundless. As long as the cannon borer was turned, it produced heat.

Benjamin Thompson (Count Rumford) helped disprove the caloric theory of heat by noting the transformation of mechanical energy into heat.

Thompson proposed an alternative hypothesis. He suggested that heat is nothing more than a consequence of the mechanical energy of friction, instead of an invisible fluid. He proved his point by immersing an entire cannon-boring machine in water, turning it on, and watching the heat that was generated turn the water to steam. British chemist and popular science lecturer Sir Humphry Davy (1778–1829) further dramatized Thompson's point when he generated heat by rubbing two pieces of ice together on a cold London day. These demonstrations could not be explained by the caloric theory and marked the beginning of the end for that theory about the nature of heat.

The Mechanical Equivalent of Heat

The work of Thompson, Davy, and others inspired the English researcher James Prescott Joule to devise an experiment to test the predictions of rival theories about the nature of heat. As shown in Figure 11-9, Joule's apparatus contained weights that were attached to ropes and lifted up. When the weight descended, the rope turned a paddle wheel that was immersed in a beaker of water. The weight had gravitational potential energy and, as it fell, that energy was converted into kinetic energy of the rotating paddle. The paddle wheel's kinetic energy, in turn, was transferred to internal energy of the agitated water molecules. As Joule suspected, the activity caused the water to heat up. Heat, he declared, is just another form of energy. In 1843, Joule even worked out how much heat was the equivalent of other kinds of energy. In modern units, Joule found that

$$1 \text{ calorie} = 4.186 \text{ joules}$$

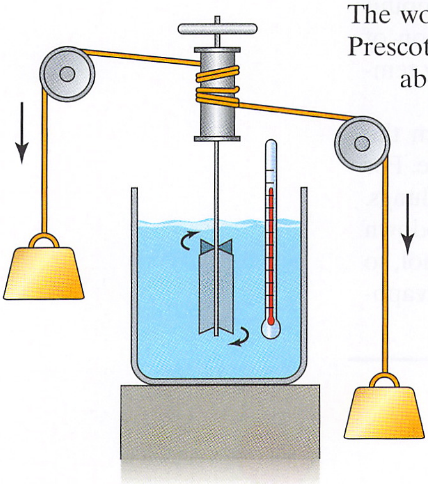


FIGURE 11-9. In his experiments, Joule demonstrated conclusively that heat is a form of energy. The kinetic energy of the turning paddle wheel heats the agitated water by a definite, measurable amount.

Joule's recognition of the equivalence of heat and mechanical energy was one of the great conceptual breakthroughs of physics. By disposing of the caloric theory once and for all, Joule showed that energy is the underlying concept governing all motions and interactions of objects. The study of energy—that is, *thermodynamics*—became one of the great fields of research for the rest of the nineteenth century. The practical benefits of this work included the improvement of the steam engine, the internal combustion engine, and electric generators—the sources of energy for modern industry and society. It is little wonder that the basic unit of energy was named for Joule.

HEAT TRANSFER

Whenever two objects are at different temperatures, heat moves from the hotter to the cooler object. This movement of energy in the form of heat occurs all around us, all the time. You can't prevent heat from moving; you can only slow down its movement. In fact, scientists and engineers have spent many decades attempting to improve thermal insulation by studying the phenomenon known as **heat transfer**—the processes by which heat moves from one place to another. Heat transfer occurs by three mechanisms: conduction, convection, and radiation.

Conduction

Have you ever reached for a pan on a hot stove, only to burn your fingers when you grasped the metal handle? If so, you have experienced **conduction,** which is the movement of heat by atomic collisions.

Conduction occurs because of the motion of individual atoms or molecules. If a piece of metal such as a pot is heated at one end, the atoms at that end begin to move faster. When they vibrate and collide with atoms farther away from the heat source, they are likely to transfer kinetic energy to those atoms, so that those molecules begin moving faster as well. A chain of collisions occurs, with atoms progressively farther and farther away from the hot end moving faster (Figure 11-10).

To an outside observer, it appears that heat somehow flows like a liquid from the hottest part of the metal pot into the handle. There's nothing particularly mysterious about this process. Heat conduction is a result of collisions between vibrating atoms or molecules. When a fast-moving object collides with a slow-moving one, the fast object usually slows down and the slow object usually speeds up.

Steelworkers test a molten sample from the furnace. Because heat transfers from the hot end of the rod to the cool end by the process of conduction, the worker at the left must wear thick insulated gloves to hold the cool end of the rod.

When we pay our home heating bills, we are in large measure paying for the conduction of heat. The process works like this: In the winter the air inside the house is kept warmer than the air outside, so the molecules of air inside are moving faster than those outside. When these molecules collide with materials in the wall (a windowpane, for example), they impart some of their energy to the molecules in the wall. At that point, conduction takes place in the wall itself and the heat is transferred to the outside of the wall. There, the heat energy is transferred outdoors mostly by convection and radiation, which are processes that we will describe in a moment. In essence, your house becomes a kind of conduit: heat flows from the interior to the outside.

One way of slowing down the flow of heat out of a house is to add insulation to the walls or use thermal glass for the windows. Both of these materials are effective because of their low **thermal conductivity,** which is their ability to transfer heat energy from one molecule to the next by conduction. Have you ever noticed that a piece of wood at room temperature feels "normal," while a piece of metal at the same temperature feels cold to the touch? The wood and metal are at exactly the same temperature, but the metal feels cold because it is a good heat conductor; it moves heat rapidly away from your skin, which is generally warmer than air temperature. The wood, on the other hand, is a good *heat insulator;* it impedes the flow of heat and so it feels normal. You wouldn't want to live in a house made entirely of metal—we usually use relatively good insulators such as wood or masonry as primary building materials. The insulation in your home, especially in newer homes, is designed to have especially low thermal conductivity, so that heat transfer is slowed down (but never completely stopped). Thus, when you use special insulated windowpanes or put certain kinds of insulation in your walls, you make it more difficult for heat to flow outside, and thereby you use heat more efficiently. The result is that you need less energy to heat your house.

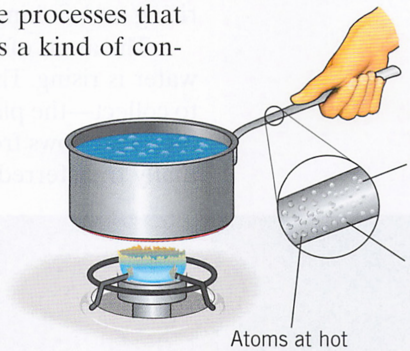


FIGURE 11-10. Conduction transfers heat by the collision of atoms in a heated object.

Convection

Let's look carefully again at a pot of boiling water on the stove, as in Chapter 9. On the surface of the water you see a rolling, churning motion as the water moves and mixes. If you put your hand above the water, you feel heat. Heat has moved from the water at the bottom of the pot to the top by **convection,** which is the transfer of heat by the bulk motion of a liquid or gas, as shown in Figure 11-11. Convection may be "forced"—for example, when a fan circulates cool air—or it may be driven by gravity, as in our pot of boiling water.

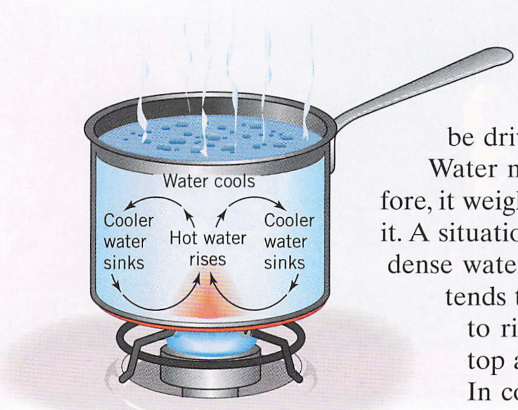


FIGURE 11-11. Convection in a pot of water transfers heat by the bulk motion of the water.

Water near the bottom of the pot expands as the flames heat it. Therefore, it weighs less per unit volume than the colder water immediately above it. A situation such as this, with colder, denser water above and warmer, less dense water below, is unstable. Under the force of gravity, the denser fluid tends to descend and displace the less dense fluid, which in turn begins to rise. Consequently, the warm water from the bottom rises to the top as shown, while the cool water from the top sinks to the bottom. In convection, masses of fluid move in bulk and carry the fast-moving molecules with them. Heat moves by the actual physical motion of these masses of fluid.

Convection is a continuous, cyclic process as long as heat flows into the fluid. As cool water from the top of the pot arrives at the bottom, it is heated by the burner. In the same way, when the hot water gets to the top, its heat flows off into the air. The water on the top cools and contracts, while the water on the bottom gets hotter and expands. The original situation repeats continuously, with the less dense fluid on the bottom always rising and the more dense fluid on the top always sinking. This transfer of fluids results in a kind of rolling motion, which you see when you look at the surface of boiling water. Each of these regions of rising and sinking water is called a **convection cell.**

The areas of clear water that seem to be bubbling are the places where warm water is rising. The places where old bubbles (and scum, if the pot is dirty) tend to collect—the places that look rather stagnant—are where the cool water is sinking. Heat flows from the burner through the convection of the water and is eventually transferred to the atmosphere.

A heat island or city creates weather patterns inland. Warm air rises up and cooler air moves in to take its place.

Convection is a very efficient way of transferring heat. If you carefully study water in a pot on the stove you will notice that for a while the temperature at the surface of the water doesn't change appreciably. During this period, the heat transfers to the surface by the rather slow process of conduction through the water molecules. Eventually, when the temperature difference between the top and bottom becomes large enough, the water starts to move. At this point convection takes over and transfers the heat through the water.

Convection is a very common process in nature. From the small-scale circulation of cold water in a glass of ice tea to air rising above a radiator or toaster to large-scale motions of Earth's atmosphere and ocean currents, convection is at work. You may even have seen convection cells in operation in large urban areas. When you're in the parking lot of a large shopping mall on a hot summer day, you can probably see the air shimmer. What you are seeing is air, heated by the hot asphalt, rising upward. Some place farther away, perhaps out in the countryside, cooler air is falling. The shopping center with all its concrete is called a "heat island" and is the source of upward-rising air. It is the hot part of a convection cell, while the rest of the convection cell is the downward-flowing air elsewhere.

You may have noticed that the temperature in big cities is usually a few degrees warmer than in the outlying suburbs. Cities influence their own weather through the creation of convection cells. Rainfall is typically higher in cities than in the surrounding areas precisely because the warmer cities result in convection cells that draw in cool moist air from the surrounding areas.

Connection

Home Insulation

Today's homebuilders take heat convection and conduction very seriously. An energy-efficient dwelling has to stay warm in the winter and remain cool in the summer. A variety of high-tech materials provide effective solutions to this insulation problem.

Fiberglass, the most widely used insulation, is made of loosely intertwined strands of glass (Figure 11-12). It works by taking advantage of the fact that motionless air is an outstanding insulator. By trapping air into pockets, fiberglass minimizes the opportunities for conduction and convection of heat out of your home. Solid glass, by itself, is a rather poor heat insulator, but it takes a long time for heat to move along a thin, twisted glass fiber and even longer for heat to transfer across the occasional contact points between pairs of crossed fibers. Furthermore, a clothlike mat of fiberglass disrupts airflow and prevents heat transfer by convection. A thick continuous layer of fiberglass in your walls and ceiling thus acts as an ideal barrier to the flow of heat.

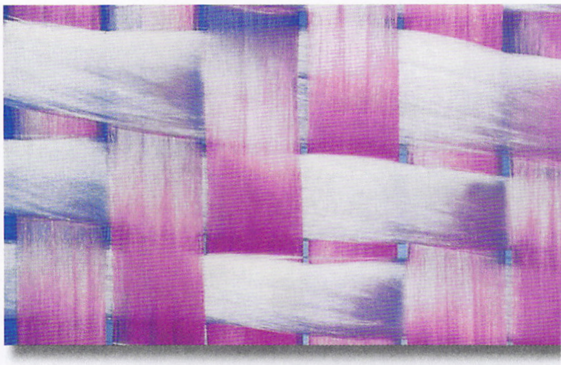


FIGURE 11-12. The interweaving of fibers in an ordinary piece of fiberglass, magnified in this photomicrograph, reduces heat transfer by convection and conduction.

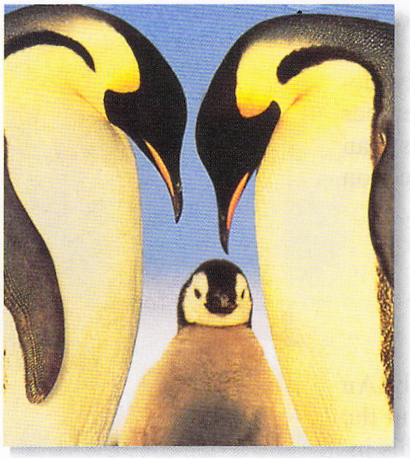
If our houses were constructed with solid walls, then fiberglass would serve all our needs. However, windows pose a special problem. Have you ever sat near an old window on a cold winter day? Old-style single-pane windows conduct heat rapidly, as you can tell by putting your hand on such a window in winter. But how do we let light in without letting heat out? One solution is dual-pane windows with sealed, airtight spaces between the panes that greatly restrict heat conduction. In addition, builders employ a variety of caulking and foam insulation to seal any possible leaks around windows and doors, reducing drafts and resultant heat loss by convection. As a result, modern homes can be almost completely airtight (although some air has to be let in and out to prevent the house from becoming too stuffy). ●

Connection

Animal Insulation: Fur and Feathers

Houses aren't the only places where heat insulation is important. Birds and mammals maintain constant body temperatures despite the temperature of their surroundings, and both have evolved methods to control the flow of heat into and out of their bodies. Part of these strategies involve their natural insulating materials, including fur, feathers, and fat. Since most of the time an animal's body is warmer than the environment, insulation usually works to keep heat in.

Whales, walruses, and seals are examples of animals that have thick layers of fat to insulate them from the cold arctic waters in which they swim. Fat is a



poor conductor of heat, and it plays much the same role in their bodies as the fiberglass insulation in your attic.

Feathers also provide insulation; in fact, most biologists believe that feathers evolved first to help birds maintain their body temperature and were only later adapted for flight. Feathers are made of light, hollow tubes connected to each other by an array of small interlocking spikes. They have some insulating properties themselves, but their main effect comes from the fact that they trap air next to the body, and stationary air is an excellent heat insulator.

Birds often react to extreme cold by contracting muscles in their skin so that the feathers fluff out. This action has the effect of increasing the thickness (and hence the insulating ability) of the layer of trapped air. (Incidentally, birds need insulation even more than we do because their normal body temperature is 41°C or 106°F.)

Animals use feathers and fur for insulation.

Hair (or fur) is actually made up of dead cells similar to those in the outer layer of the skin. Like feathers, hair serves as an insulator and traps a layer of air near the body. In some animals (such as polar bears) the insulating power of the hair is increased by the fact that each hair contains tiny bubbles of trapped air. The reflection of light from these bubbles makes polar bear fur appear white—the strands of hair are actually translucent.

Hair grows from follicles in the skin, and small muscles allow animals to make their hair stand up to increase its insulating power. Over time, human beings have lost much of their body hair, as well as the ability to make most of it stand up. However, our mammalian nature is revealed in the phenomenon of goose bumps, in which muscles in the skin attempt to make the nonexistent hair stand up.

The main purpose of human clothing is to trap air near the body for insulation, just as animals and birds use fur or feathers. (This function of clothing was first recognized by Benjamin Thompson, who helped determine the nature of heat; see Physics in the Making on page 239.) Indeed, the earliest forms of clothing consisted of furs taken from animals hunted for that purpose. ●

Radiation

Everyone has experienced coming in on a cold day and finding a fire in the fireplace or an electric heater glowing red-hot. The normal reaction is to walk up to the source of heat, hold out your hands, and feel the warmth on your skin. But how does the heat move from the fire to your hands? It can't do so by conduction—it's too hard to transfer heat through the air that way. It can't be convection either, because you don't feel a hot breeze. The air in the room is almost stationary.

What you experience is the third kind of heat transfer—**radiation,** or the transfer of heat by *electromagnetic radiation,* which is a kind of wave that we discuss in Chapter 19. A fire, an electric heater, and the Sun all transfer heat in this form. This radiation travels like light from a hot source to your hand, where it is absorbed and converted into internal energy. You feel the warmth because of the energy that the radiation carries to your hand (see Figure 11-13).

Every object in the universe radiates energy. Under normal circumstances, as an object gives off radiant energy to its surroundings, it also receives radiant energy from the surroundings. Thus, a kind of equilibrium is set up and there is no net loss or gain of energy because the object is at the same temperature as its surroundings. It receives the same amount of energy as it radiates. However, if the object is at a higher temperature than its surroundings, it radiates more energy than it receives. Your body, for example, constantly radiates heat into its cooler surroundings. This energy can be detected easily at night with infrared goggles. You continue to radiate this energy as long as your body processes the food that keeps you alive.

FIGURE 11-13. A fire transfers much of its energy by electromagnetic radiation.

Radiation is the only kind of energy that can travel through the emptiness of space. Conduction requires atoms or molecules that can vibrate and collide with each other. Convection requires atoms or molecules of bulk fluid, so that they can move. But radiation doesn't require anything in the environment to transport it; radiation can even travel through a vacuum. The energy that falls on Earth in the form of sunlight—almost all the energy that sustains life on Earth—travels through 93 million miles of intervening empty space in the form of radiation.

In the real world, all three types of heat transfer—conduction, convection, and radiation—occur constantly. Think about the heat generated by your body: heat conducts through your bones and teeth, heat convects as your blood circulates (an example of forced convection), and heat radiates from your skin into the cooler surroundings, where it is eventually absorbed. As your heat flows outward, it produces slightly higher temperatures in the surrounding air. The same three phenomena operate in a pot of boiling water. Heat moves through the metal bottom and sides of the pot by conduction, it moves through the boiling water by convection, and it moves from the sides and surface of the water by radiation. In fact, everywhere in the natural world, heat is constantly being transferred by these three mechanisms.

THINKING MORE ABOUT

Heat: Global Warming and the Greenhouse Effect

Heat constantly flows to Earth from the Sun, and it constantly flows outward from Earth into the cold blackness of space. Changes in the average temperature of Earth's surface arise from a complex combination of heat transfers that are not yet fully understood. But one possible source of change, called the greenhouse effect, is receiving a lot of attention.

Have you ever visited a greenhouse on a sunny winter day? It may be freezing cold outside, yet the temperature inside is much warmer. This contrast occurs because the Sun transfers heat by the radiation of visible light, which passes through the glass and warms up the interior. When the interior radiates away its own energy, it must do so by sending out infrared radiation. However, while the glass in a greenhouse is transparent to visible light, it is opaque to infrared radiation. The Sun's energy is thus trapped, warming the greenhouse until its temperature rises to the point where as much energy leaks out through the glass as comes in from the Sun.

In exactly the same way, Earth's atmosphere transmits the Sun's incoming radiation but some molecules in the air (particularly carbon dioxide) trap much of the Earth's heat as it flows away from the surface (Figure 11-14). These molecules play the same role for Earth that the glass does for the greenhouse, so the phenomenon has come to be known as the "greenhouse effect." Its effect is to warm the planet, just as the glass warms the greenhouse.

The greenhouse effect is vital to life on Earth. Without a layer of greenhouse gases, our planet would be a lifeless frozen ball. Modern concerns arise from the relatively rapid increase in atmospheric carbon dioxide during the past century—a consequence of burning fossil fuels such as coal, petroleum, and natural gas. Will this change in atmospheric composition cause a rise in global temperatures? Has such a change already begun?

The greenhouse effect is an area of continuing research in many fields of science. Physicists have examined the mechanisms of heat transfer in the atmosphere; chemists have studied which gases persist in the atmosphere and how they absorb heat; geologists have looked at rock and fossil formations to determine the history of Earth's temperature changes; and astronomers have simulated atmospheric models on other planets, such as Venus, which may have once had an accelerated greenhouse effect. Most experts now conclude that global warming is well under way and that the big question is not whether but how much global temperatures will rise over the next century. The best estimates are that the average global temperature will rise by a few degrees Celsius over that period.

If greenhouse warming is indeed happening, what sorts of things do you want to know about its possible consequences? How much in the way of consequences are you willing to accept in order to keep the convenience of using fossil fuels?

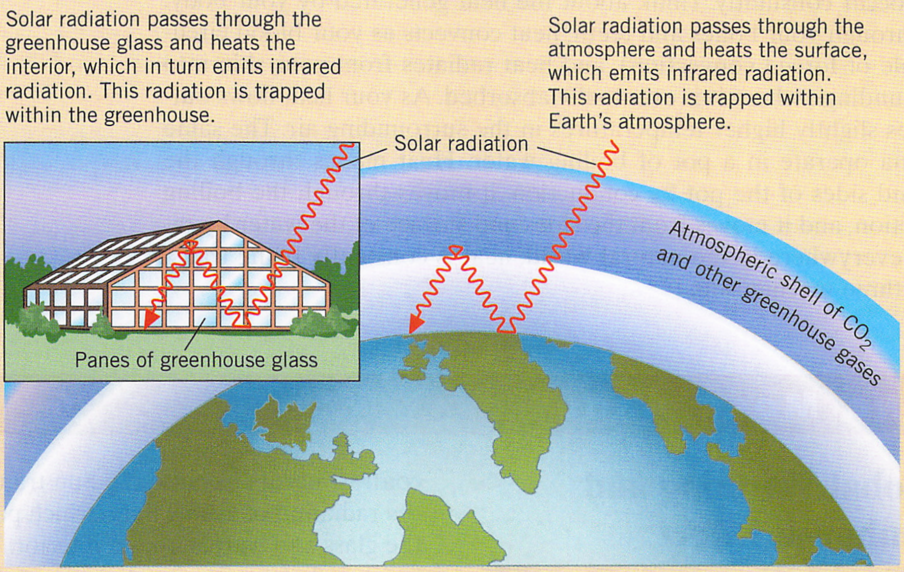


FIGURE 11-14. The greenhouse effect. Just as the Sun's energy passes through the glass of a greenhouse and becomes trapped inside as heat, so too does the atmosphere act as a greenhouse to warm up Earth.

Summary

In everyday life, **temperature** is a quantity that indicates how hot or cold an object is relative to a standard value. All objects in the universe are at a temperature above **absolute zero;** thus, they hold some internal energy—the kinetic energy of the moving atoms. A **thermometer** is a device that measures temperature, according to a **temperature scale.**

Heat is a transfer of energy that moves from an object at a higher temperature to an object at a lower temperature.

Objects themselves contain **thermal energy** (also called **internal energy**), which is the kinetic energy of wiggling atoms and molecules. Most solids and liquids expand when heated; the amount of expansion is proportional to the **coefficient of linear expansion.**

Specific heat is a measure of the ability of a material to absorb heat and is defined as the quantity of heat required to raise the temperature of 1 gram of that material by 1°C. An object's overall ability to absorb heat is measured by its **heat capacity.** When a change of phase is involved, the temperature of an object remains fixed while the bonds between its atoms and molecules are rearranged. The **latent heat of fusion** is the heat required to change 1 gram of a material from a solid to a liquid, and the **latent heat of vaporization** is the heat required to change 1 gram of a substance from a liquid to a gas.

There are three modes of **heat transfer** between two objects at different temperatures. **Conduction** involves the transfer of heat energy through the collision of individual atoms and molecules. **Thermal conductivity** measures the ability of substances to transfer heat. **Convection** involves the motion of a mass of fluid in a **convection cell,** in which atoms are physically transported from one place to another. Heat can also be transferred by **radiation,** which is electromagnetic energy that can travel across a room or across the vastness of space until it is absorbed.

Key Terms

absolute zero The lowest possible temperature, at which no energy can be extracted from atoms. (p. 228)

coefficient of linear expansion A quantity that relates the temperature change with the corresponding length change of a material. (p. 232)

conduction The transfer of heat due to atomic or molecular collisions. (p. 240)

convection Heat transfer due to the motion of a liquid or a gas. (p. 242)

convection cell A region of a fluid that is either rising or sinking due to the heat convection process. (p. 242)

heat The energy transferred from one body to another due to a difference in temperature between the two bodies. (p. 234)

heat capacity A measure of the change in temperature of an object on adding or removing heat; the amount of heat required to raise the temperature of the object by 1°C. (p. 235)

heat transfer The process by which thermal energy moves from one place to another. (p. 234)

latent heat of fusion The amount of heat required to change 1 gram of a solid material to a liquid when the solid is at its melting temperature; equivalently, the amount of heat that must be removed from 1 gram of liquid material to turn it into a solid when the liquid is at its freezing temperature. (p. 238)

latent heat of vaporization The amount of heat required to change 1 gram of a liquid material to a gas when the liquid is at its boiling temperature; equivalently, the amount of heat that must be removed from 1 gram of a gaseous material to turn it into a liquid when the gas is at its condensation temperature. (p. 238)

radiation Heat transfer due to the emission and absorption of electromagnetic waves between two bodies at different temperatures. (p. 244)

specific heat The quantity of heat required to raise the temperature of 1 gram of a material by 1°C. (p. 235)

temperature A quantity that reflects how vigorously atoms or molecules are moving and colliding in a material. (p. 228)

temperature scale A standard of measurement for estimating temperature; familiar examples are the Fahrenheit and the Celsius scales. (p. 228)

thermal conductivity The ability of a material to transfer heat. (p. 241)

thermal energy or **internal energy** The energy of an object that results from the vibrations of individual atoms and molecules. (p. 234)

thermometer A device used to measure temperature. (p. 229)

Key Equations

Temperature conversions:

$$T \text{ (in °F)} = [1.8 \times T \text{ (in °C)}] + 32$$

$$T \text{ (in °C)} = \frac{T \text{ (in °F)} - 32}{1.8}$$

$$T \text{ (in °C)} = T \text{ (in K)} - 273$$

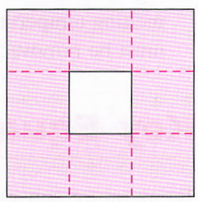
$$T \text{ (in K)} = T \text{ (in °C)} + 273$$

Review

1. What is temperature? Is it a relative or absolute quantity? Explain.
2. Describe three common temperature scales. What fixed points are used to calibrate them?
3. How do you convert degrees Fahrenheit to degrees Celsius? To kelvins?
4. What is meant by absolute zero? What happens to the movement of molecules at this temperature?
5. What is the numerical value of absolute zero in degrees Celsius? In degrees Fahrenheit?
6. In your own words, define temperature. How is absolute zero related to your definition of temperature?
7. What is the general principle behind the working of a thermometer? Give examples of different types of thermometers.
8. How did Galileo's thermometer work? How does the density of the liquid used affect the location of different weights suspended inside the liquid?
9. How does the kinetic energy of the atoms of a substance affect its temperature? If two objects have the same temperature, what can you say about the average kinetic energies of their molecules?
10. Can a thermometer be used to measure the temperature of something much smaller than itself? Why or why not?
11. What is the difference between temperature and heat?
12. What is a calorie? How many calories of heat are required to raise the temperature of 5 grams of water by 1°C?
13. How many calories make up 1 Calorie (capital C), the unit by which we measure the energy in our diets?
14. What is a BTU? How do you convert BTUs to calories?
15. What is the relation between the perception that something is hot and the motion of the molecules in it?
16. What is the difference, if any, between heat and internal energy?
17. How does the discovery of heat as a form of energy illustrate the scientific method?
18. What is specific heat? Is it the same for all materials?
19. What is the specific heat of water?
20. How does the relatively high specific heat of water affect Earth's climate? Specifically, what role does it play in the creation of monsoons?
21. How does the specific heat of water affect the winds blowing at different times of the day near large bodies of water?
22. Describe the historical process by which the notion of heat as a form of energy was developed.
23. What is a joule? What does it measure? How many calories make up 1 joule?
24. Identify the three ways that heat can be transferred and give examples of each.
25. What kind of heat transfer depends only on the collisions between individual atoms and molecules?
26. What kind of heat transfer depends on the bulk motion of large numbers of molecules?
27. By what process can heat be transferred across the vacuum of empty space?
28. Describe the kinds of heat transfer that occur when you cook a meal. Where does the heat energy come from? Where does it end up?
29. What is a fluid? Are all fluids liquid?

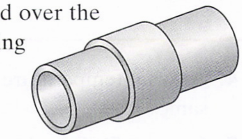
Questions

1. A glass of water sits on a table. The temperature of the water is the same as that of the glass. What can you say about the average kinetic energy of the water (H_2O) molecules compared to the silicon dioxide (SiO_2) molecules that make up the glass? Which is moving faster, the silicon dioxide molecules or the water molecules? (*Hint:* Look at the periodic table to determine the mass of a SiO_2 molecule versus a water molecule.)
2. A golf ball is dropped onto hard ground and, after a few bounces, comes to rest. Use the atomic basis of temperature to explain why the golf ball's temperature is slightly higher after it comes to rest.
3. A square hole is cut out of a piece of sheet metal, as shown in the figure. When the temperature of the metal is raised, the metal expands. What happens to the size of the square hole? (*Hint:* Break up the piece of metal into eight smaller square pieces of sheet metal, then raise the temperature, then put them back together.)

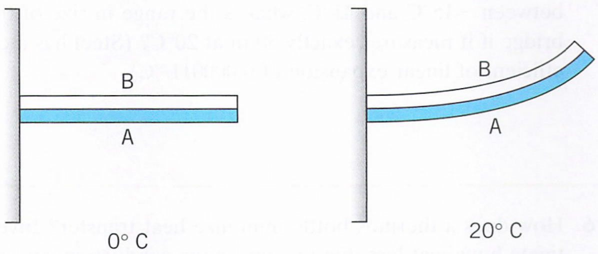


4. Suppose your gold wedding ring became stuck on your finger. Some home remedies suggest soaking your finger in ice water and then trying to remove the ring. If this remedy relies on thermal expansion effects, what does this tell you about the relative expansion coefficients of gold and your finger?
5. Two pieces of copper pipe are stuck together, as shown. One way to separate them is to run water at different tem-

peratures inside the inner pipe and over the outer pipe. Should the water running over the outer pipe be hotter or colder than the water running through the inner pipe? Explain.



6. Two thin strips of metal (A and B) are glued together at 0°C as shown in the figure. At 20°C they bend upward because the metals expand differently. Which metal, A or B, has a higher thermal expansion coefficient?



7. A mercury thermometer consists of a mercury-filled glass bulb that is connected to a narrow glass tube. Mercury thermometers are based on the thermal expansion of mercury: As the mercury expands, it rises up the tube. What can you say about the thermal expansion coefficient of mercury relative to the thermal expansion coefficient of glass? What would happen if their thermal expansion coefficients were the same?

8. A certain amount of heat is added to some water, and its temperature rises. The same amount of heat is added to a piece of aluminum with the same mass as the water. Compare the temperature change of aluminum to the temperature change of water.

9. If water had a lower specific heat, would your chances of enjoying a long, hot bath be greater or less? Explain?

10. Suppose a new liquid were discovered that is identical to water in every way except that it has a lower latent heat of vaporization. If you had to cook your pasta using either ordinary water or this new liquid, which would you choose and why?

11. Suppose a new liquid were discovered that is identical to water in every way except that it has a lower latent heat of fusion. Would it take a longer or shorter time to make ice out of this liquid in your freezer? Would this necessarily be a more desirable situation?

12. Suppose a new liquid were discovered that is identical to water in every way except that it has a lower specific heat. Consider taking a hot shower with this liquid. Would insulating the pipes from the hot water heater to the shower head be more or less important with this new liquid?

13. Why do some animals roll up into a ball when they are cold?

14. Why are feather beds warm, and why is goose down considered the best filling for a parka?

15. Imagine lying on a hot beach on a sunny summer day. In what different ways is heat transferred to your body? In each case, what was the original source of the heat energy?

16. Why do human beings wear clothes? Compare our behavior in this regard to other warm-blooded animals.

17. What is the difference between heat capacity and heat transfer?

18. Outline the three major modes of heat transfer. For each case, state if a medium is necessary, compare the motion of molecules in this medium, and relate the heat transfer to the presence or absence of a temperature difference in the medium.

19. The average specific heat of the human body is 83% of the specific heat of water. This value is higher than for most other solids, liquids, or gases. Why do you think the specific heat of the human body is closest to water?

20. Human beings must lose heat so that their internal temperatures do not increase substantially above 37°C. The main mechanism for losing heat is sweating. Explain why this is an efficient mechanism to lose energy from the body.

21. Absolute zero, 0 K, is the lowest possible temperature. Temperatures below absolute zero do not exist. In terms of the molecular motion of a substance (a gas, for example), explain why there is a lowest temperature (absolute zero) but not a highest temperature.

22. Temperatures below absolute zero do not exist. However, suppose you are given an object that is right at absolute zero. Would it be possible to use this object to cool another object down to absolute zero?

23. Three identical potatoes are taken out of a hot oven to cool. The first is placed on the countertop. The second is wrapped in aluminum foil and placed inside a jar, and then the air is removed from the jar. The third potato is wrapped in aluminum foil and then placed on the countertop alongside the first. Can you place the potatoes in order of which will cool fastest?

24. One hundred grams of liquid A is at a temperature of 100°C. One hundred grams of liquid B is at a temperature of 0°C. When the two liquids are mixed, the final temperature is 50°C. What can you say about the specific heats of the two liquids? Explain your reasoning.

25. Two hundred grams of liquid A is at a temperature of 100°C. One hundred grams of liquid B is at a temperature of 0°C. When the two liquids are mixed, the final temperature is 50°C. Which material has a higher specific heat? Explain your reasoning.

26. On a very cold day you find that your key does not fit into your car door lock. Assuming this has happened because of thermal expansion effects, what can you say about the thermal expansion coefficient of your key relative to the thermal expansion coefficient of the lock?

Problems

1. Convert the following Fahrenheit temperatures to kelvins.
   a. 120°F c. 11,500°F
   b. −40°F d. −456°F
2. Convert the following Celsius temperatures to Fahrenheit.
   a. 300°C c. 6,000°C
   b. −180°C d. 40°C
3. Convert the following kelvin temperatures to Celsius.
   a. 80 K c. 6000 K
   b. 300 K d. 545 K
4. At what temperature is the Celsius and Fahrenheit value the same?
5. Convert 70°F to degrees Celsius and to kelvins.
6. The coefficient of linear expansion for a silver strip is 19×10^{-6}/°C. What is its length on a hot day when the temperature is 37°C if the strip is 0.20000 m long when it is −10°C?
7. If a 50-m steel footbridge experiences extreme temperatures between −15°C and 45°C, what is the range in size of this bridge if it measures exactly 50 m at 20°C? (Steel has a coefficient of linear expansion of 0.000011/°C)

Investigations

1. Research the daily high and low temperatures for the past week in a nearby big city and one of its surrounding smaller towns. On average, what is the difference in temperature? What causes this difference between the city and town?
2. Repeat Investigation 1, but compare the daily high and low temperatures for the past week in a coastal town with a nearby inland town.
3. Investigate the kind of insulation that is installed in your home. What could you do to improve your home's insulation?
4. Get an aluminum cup, a ceramic coffee mug, and a plastic drinking cup. Simply by feeling these objects, can you guess their relative thermal conductivities? What experiment could you perform to test your guess?
5. Visit a building supply store and look at the doors and windows they sell. What steps have manufacturers taken to reduce heat flow from homes?
6. How does a thermos bottle minimize heat transfer? Investigate how heat loss due to convection, conduction, and radiation is prevented by looking at the design of a thermos. Compare cheaper versions to the design of more expensive versions and explain the differences in terms of the concepts used in this chapter.
7. Boil a pan of water on your stove. Can you identify the convection cells in the pan? How hot does the water have to be before convection starts? (*Hint:* A small amount of food coloring can reveal the formation of convection cells.)
8. Investigate the history of temperature scales. Describe an obsolete temperature scale and its fixed points. Why was it abandoned? When was the Fahrenheit scale introduced, and what were its original fixed points?

WWW Resources

See the *Physics Matters* home page at **www.wiley.com/college/trefil** for valuable web links.

1. **http://unidata.ucar.edu/staff/blynds/tmp.html** A tutorial on temperature from the National Center for Atmospheric Research.
2. **http://microgravity.grc.nasa.gov/combustion/index.htm** The NASA microgravity combustion science page describes the effects of microgravity on a candle flame, illustrating the effects of convection on combustion.
3. **http://www.psrc-online.org/classrooms/papers/coleman.html** A sample class on the physics of the greenhouse effect.
4. **http://jersey.uoregon.edu/vlab/Thermodynamics/** A simulation experiment on thermodynamic equilibrium of two ideal gases.