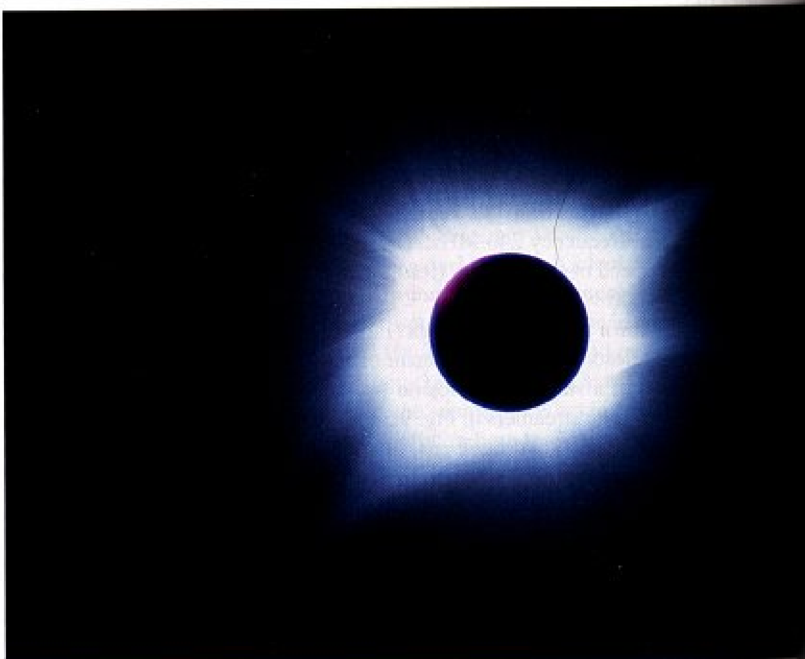
17 TEMPERATURE AND HEAT

Perhaps the highest-temperature material you will ever see is the sun's outer atmosphere, or corona. At a temperature of about 2,000,000°C (3,600,000°F), the corona glows with a light that is literally unearthly. But because the corona is also very thin, its light is rather faint. You can only see the corona during a total solar eclipse when the sun's disk is covered by the moon, as shown here.

? **Is it accurate to say that the corona contains heat?**



Whether it's a sweltering summer day or a frozen midwinter night, your body needs to be kept at a nearly constant temperature. It has effective temperature-control mechanisms, but sometimes it needs help. On a hot day you wear less clothing to improve heat transfer from your body to the air and for better cooling by evaporation of perspiration. You probably drink cold beverages, possibly with ice in them, and sit near a fan or in an air-conditioned room. On a cold day you wear more clothes or stay indoors where it's warm. When you're outside, you keep active and drink hot liquids to stay warm. The concepts in this chapter will help you understand the basic physics of keeping warm or cool.

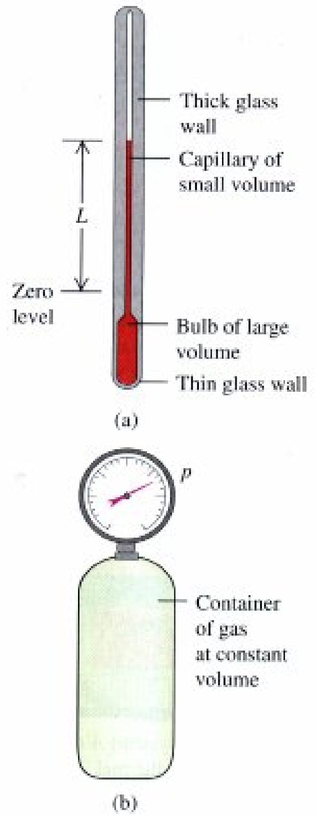
The terms *temperature* and *heat* are often used interchangeably in everyday language. In physics, however, these two terms have very different meanings. In this chapter we'll define temperature in terms of how it's measured and see how temperature changes affect the dimensions of objects. We'll see that heat refers to energy transfer caused by temperature differences and learn how to calculate and control such energy transfers.

Our emphasis in this chapter is on the concepts of temperature and heat as they relate to *macroscopic* objects such as cylinders of gas, ice cubes, and the human body. In Chapter 18 we'll look at these same concepts from a *microscopic* viewpoint in terms of the behavior of individual atoms and molecules. These two chapters lay the groundwork for the subject of **thermodynamics,** the study of energy transformations involving heat, mechanical work, and other aspects of energy and how these transformations relate to the properties of matter. Thermodynamics forms an indispensable part of the foundation of physics, chemistry, and the life sciences, and its applications turn up in such places as car engines, refrigerators, biochemical processes, and the structure of stars. We'll explore the key ideas of thermodynamics in Chapters 19 and 20.

17.1 | Temperature and Thermal Equilibrium

The concept of **temperature** is rooted in qualitative ideas of "hot" and "cold" based on our sense of touch. A body that feels hot usually has a higher temperature than a similar body that feels cold. That's pretty vague, and the senses can be deceived. But many properties of matter that we can *measure* depend on temperature. The length of a metal rod, steam pressure in a boiler, the ability of a wire to conduct an electric current, and the color of a very hot glowing object—all these depend on temperature.

Temperature is also related to the kinetic energies of the molecules of a material. In general this relationship is fairly complex, so it's not a good place to start in *defining* temperature. In Chapter 18 we will look at the relationship between temperature and the energy of molecular motion for an ideal gas. It is important to understand, however, that temperature and heat can be defined independently of any detailed molecular picture. In this section we'll develop a *macroscopic* definition of temperature.

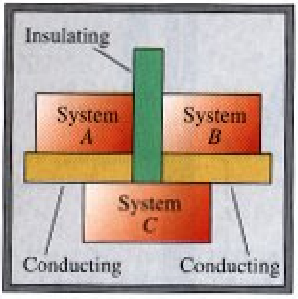


17.1 (a) A system whose temperature is specified by the value of length L. (b) A system whose temperature is given by the value of the pressure p.

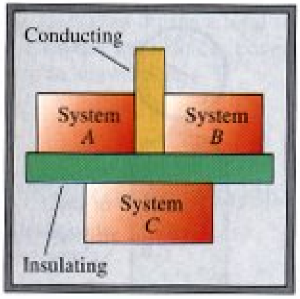
To use temperature as a measure of hotness or coldness, we need to construct a temperature scale. To do this, we can use any measurable property of a system that varies with its "hotness" or "coldness." Figure 17.1a shows a familiar system that is used to measure temperature. When the system becomes hotter, the colored liquid (usually mercury or ethanol) expands and rises in the tube, and the value of L increases. Another simple system is a quantity of gas in a constant-volume container (Fig. 17.1b). The pressure p, measured by the gauge, increases or decreases as the gas becomes hotter or colder. A third example is the electrical resistance R of a conducting wire, which also varies when the wire becomes hotter or colder. Each of these properties gives us a number (L, p, or R) that varies with hotness and coldness, so each property can be used to make a **thermometer.**

To measure the temperature of a body, you place the thermometer in contact with the body. If you want to know the temperature of a cup of hot coffee, you stick the thermometer in the coffee; as the two interact, the thermometer becomes hotter and the coffee cools off a little. After the thermometer settles down to a steady value, you read the temperature. The system has reached an *equilibrium* condition, in which the interaction between the thermometer and the coffee causes no further change in the system. We call this a state of **thermal equilibrium.**

If two systems are separated by an insulating material or **insulator** such as wood, plastic foam, or fiberglass, they influence each other more slowly. Camping coolers are made with insulating materials to delay the ice and cold food inside from warming up and attaining thermal equilibrium with the hot summer air outside. An *ideal insulator* is a material that permits no interaction at all between the two systems. It prevents the systems from attaining thermal equilibrium if they aren't in thermal equilibrium at the start. An ideal insulator is just that, an idealization; real insulators, like those in camping coolers, aren't ideal, so the contents of the cooler will warm up eventually.



(a) If systems A and B are each in thermal equilibrium with system C ...



(b) ... then systems A and B are in thermal equilibrium with each other

17.2 The zeroth law of thermodynamics. Blue slabs represent thermal insulators, and orange slabs represent thermal conductors.

We can discover an important property of thermal equilibrium by considering three systems, A, B, and C, that initially are not in thermal equilibrium (Fig. 17.2). We surround them with an ideal insulating box so that they cannot interact with anything except each other. We separate systems A and B with an ideal insulating wall (the blue slab in Fig. 17.2a), but we let system C interact with both systems A and B. This interaction is shown in the figure by an orange slab representing a thermal **conductor,** a material that *permits* thermal interactions through it. We wait until thermal equilibrium is attained; then A and B are each in thermal equilibrium with C. But are they in thermal equilibrium *with each other?*

To find out, we separate system C from systems A and B with an ideal insulating wall (Fig. 17.2b), and then we replace the insulating wall between A and B with a *conducting* wall that lets A and B interact. What happens? Experiment shows that *nothing* happens; there are no additional changes to A or B. We conclude that **if C is initially in thermal equilibrium with both A and B, then A and B are also in thermal equilibrium with each other.** This result is called the **zeroth law of thermodynamics.** (The importance of this law was recognized only after the first, second, and third laws of thermodynamics had been named. Since it is fundamental to all of them, the name "zeroth" seemed appropriate.)

Now suppose system C is a thermometer, such as the tube-and-liquid system of Fig. 17.1a. In Fig. 17.2a the thermometer C is in contact with both A and B. In thermal equilibrium, when the thermometer reading reaches a stable value, the thermometer measures the temperature of both A and B; hence A and B both have the *same* temperature. Experiment shows that thermal equilibrium isn't affected by adding or removing insulators, so the reading of thermometer C wouldn't change if it were in contact only with A or only with B. We conclude that **two systems are in thermal equilibrium if and only if they have the same temperature.** This is what makes a thermometer useful; a thermometer actually measures *its own* temperature, but when a thermometer is in thermal equilibrium with another body, the temperatures must be equal. When the temperatures of two systems are different, they *cannot* be in thermal equilibrium.

Test Your Understanding

Why does a nurse taking your temperature wait for the thermometer reading to stop changing? What object is it whose temperature the nurse is reading?

17.2 | Thermometers and Temperature Scales

To make the liquid-in-tube device shown in Fig. 17.1a into a useful thermometer, we need to mark a scale on the tube wall with numbers on it. These numbers are arbitrary, and historically many different schemes have been used. Suppose we label the thermometer's liquid level at the freezing temperature of pure water "zero" and the level at the boiling temperature "100," and divide the distance between these two points into 100 equal intervals called *degrees.* The result is the **Celsius temperature scale** (formerly called the *centigrade* scale in English-speaking countries). The Celsius temperature for a state colder than freezing water is a negative number. The Celsius scale is used, both in everyday life and in science and industry, almost everywhere in the world.

Another common type of thermometer uses a *bimetallic strip,* made by bonding strips of two different metals together (Fig. 17.3a). When the temperature of the composite strip increases, one metal expands more than the other and the strip bends. This strip is usually formed into a spiral, with the outer end anchored to the thermometer case and the inner end attached to a pointer (Fig. 17.3c). The pointer rotates in response to temperature changes.

In a *resistance thermometer* the changing electrical resistance of a coil of fine wire, a carbon cylinder, or a germanium crystal is measured. Because resistance can be measured very precisely, resistance thermometers are usually more precise than most other types.

Some thermometers don't require physical contact with the object they're measuring. An example is an ear thermometer (Fig. 17.4). Such a thermometer uses a device called a *thermopile* to measure the amount of infrared radiation emitted by the eardrum, which indicates the eardrum's temperature. (We'll see in Section 17.7 that *all* objects emit electromagnetic radiation as a consequence of their temperature.) The advantage of this technique is that it doesn't require touching the fragile and easily damaged eardrum.

In the **Fahrenheit temperature scale,** still used in everyday life in the United States, the freezing temperature of water is 32°F (thirty-two degrees Fahrenheit) and the boiling temperature is 212°F, both at standard atmospheric pressure. There are 180 degrees between freezing and boiling, compared to 100 on the Celsius scale, so one Fahrenheit degree represents only $\frac{100}{180}$, or $\frac{5}{9}$, as great a temperature change as one Celsius degree.

To convert temperatures from Celsius to Fahrenheit, note that a Celsius temperature T_C is the number of Celsius degrees above freezing; the number of Fahrenheit degrees above freezing is $\frac{9}{5}$ of this. But freezing on the Fahrenheit scale is at 32°F, so to obtain the actual Fahrenheit temperature T_F, multiply the Celsius value by $\frac{9}{5}$ and then add 32°. Symbolically,

$$T_F = \frac{9}{5}T_C + 32° \tag{17.1}$$

To convert Fahrenheit to Celsius, solve this equation for T_C:

$$T_C = \frac{5}{9}(T_F - 32°) \tag{17.2}$$

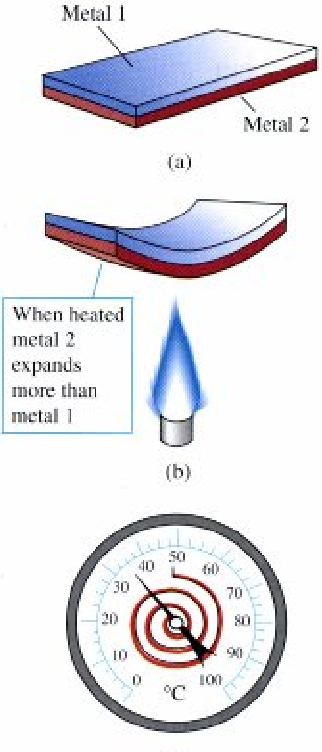
In words, subtract 32° to get the number of Fahrenheit degrees above freezing, and then multiply by $\frac{5}{9}$ to obtain the number of Celsius degrees above freezing, that is, the Celsius temperature.

We don't recommend memorizing Eqs. (17.1) and (17.2). Instead, try to understand the reasoning that led to them so that you can derive them on the spot when you need them, checking your reasoning with the relation 100°C = 212°F.

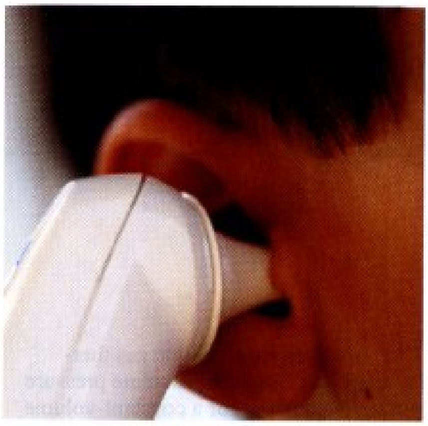
It is useful to distinguish between an actual temperature and a temperature *interval* (a difference or change in temperature). An actual temperature of 20° is stated as 20°C (twenty degrees Celsius), and a temperature *interval* of 10° is 10 C° (ten Celsius degrees). A beaker of water heated from 20°C to 30°C has a temperature change of 10 C°.

Test Your Understanding

Find the average Fahrenheit temperature on the planet Venus (average Celsius temperature 460°C), and find the temperature at which the Fahrenheit and Celsius scales agree.



17.3 (a) A bimetallic strip. (b) The strip bends when its temperature is raised. (c) A bimetallic strip used in a thermometer.



17.4 An ear thermometer measures infrared radiation from the eardrum, which is located far enough inside the head that it gives an excellent indication of the body's internal temperature.

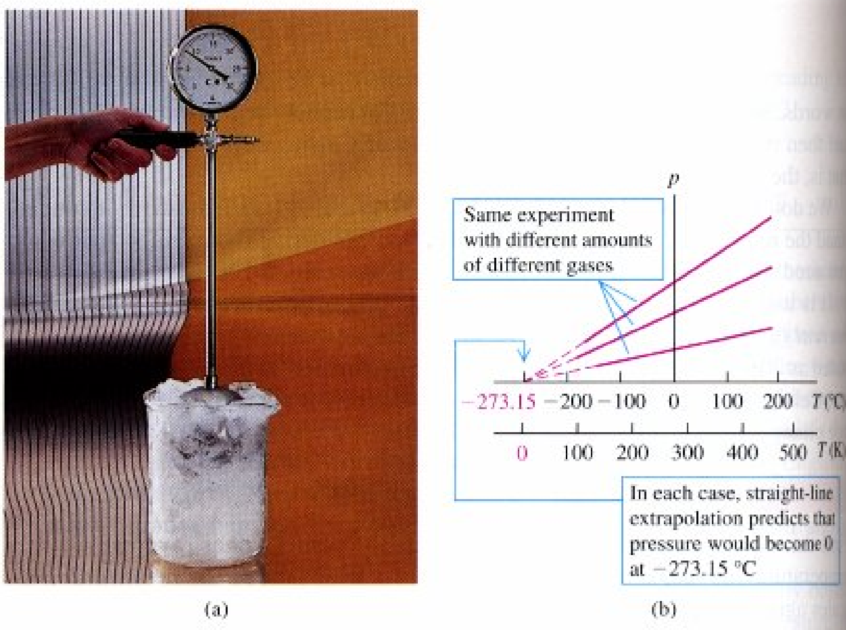
17.3 | Gas Thermometers and the Kelvin Scale

When we calibrate two thermometers, such as a liquid-in-tube system and a resistance thermometer, so that they agree at 0°C and 100°C, they may not agree exactly at intermediate temperatures. Any temperature scale defined in this way always depends somewhat on the specific properties of the material used. Ideally, we would like to define a temperature scale that *doesn't* depend on the properties of a particular material. To establish a truly material-independent scale, we first need to develop some principles of thermodynamics. We'll return to this fundamental problem in Chapter 20. Here we'll discuss a thermometer that comes close to the ideal, the *gas thermometer.*

The principle of a gas thermometer is that the pressure of a gas at constant volume increases with temperature. A quantity of gas is placed in a constant-volume container (Fig. 17.5a), and its pressure is measured by one of the devices described in Section 14.2. To calibrate a constant-volume gas thermometer, we measure the pressure at two temperatures, say 0°C and 100°C, plot these points on a graph, and draw a straight line between them. Then we can read from the graph the temperature corresponding to any other pressure. Figure 17.5b shows the results of three such experiments, each using a different type and quantity of gas.

By extrapolating this graph, we see that there is a hypothetical temperature, −273.15°C, at which the absolute pressure of the gas would become zero. We might expect that this temperature would be different for different gases, but it turns out to be the *same* for many different gases (at least in the limit of very low gas density). We can't actually observe this zero-pressure condition. Gases liquefy and solidify at very low temperatures, and the proportionality of pressure to temperature no longer holds.

We use this extrapolated zero-pressure temperature as the basis for a temperature scale with its zero at this temperature. This is the **Kelvin temperature scale,**



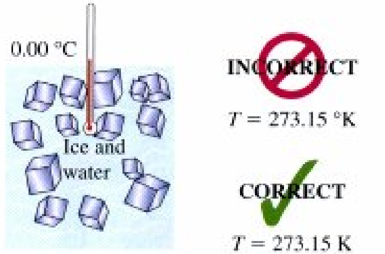
17.5 (a) A constant-volume gas thermometer. (b) Graphs of absolute pressure versus temperature for a constant-volume low-density gas thermometer. The three graphs are for experiments with different types and quantities of gas: the greater the amount of gas, the higher the graph. The dashed lines are extrapolations of the data to low temperature.

named for the British physicist Lord Kelvin (1824–1907). The units are the same size as those on the Celsius scale, but the zero is shifted so that 0 K = −273.15°C and 273.15 K = 0°C; that is,

$$T_K = T_C + 273.15 \quad (17.3)$$

This scale is shown in Fig. 17.5b. A common room temperature, 20°C (= 68°F), is 20 + 273.15, or about 293 K.

CAUTION In SI nomenclature, "degree" is not used with the Kelvin scale; the temperature mentioned above is read "293 kelvins," not "degrees Kelvin" (Fig. 17.6). We capitalize Kelvin when it refers to the temperature scale; however, the *unit* of temperature is the *kelvin*, which is not capitalized (but is nonetheless abbreviated as a capital K).



17.6 Kelvin temperatures are measured in kelvins (K), *not* "degrees Kelvin."

Example 17.1

Body temperature

You place a small piece of melting ice in your mouth. Eventually, the water all converts from ice at $T_1 = 32.00°\text{F}$ to body temperature, $T_2 = 98.60°\text{F}$. Express these temperatures as °C and K, and find $\Delta T = T_2 - T_1$ in both cases.

SOLUTION

IDENTIFY and SET UP: We convert Fahrenheit to Celsius temperatures using Eq. (17.2), and Celsius to Kelvin temperatures using Eq. (17.3).

EXECUTE: First we find the Celsius temperatures. We know that $T_1 = 32.00°\text{F} = 0.00°\text{C}$, and 98.60°F is 98.60 − 32.00 = 66.60 F° above freezing; multiply this by (5 C°/9 F°) to find 37.00 C° above freezing, or $T_2 = 37.00°\text{C}$.

To get the Kelvin temperatures, we just add 273.15 to each Celsius temperature: $T_1 = 273.15$ K and $T_2 = 310.15$ K. "Normal" body temperature is 37.0°C, but if your doctor says that your temperature is 310 K, don't be alarmed.

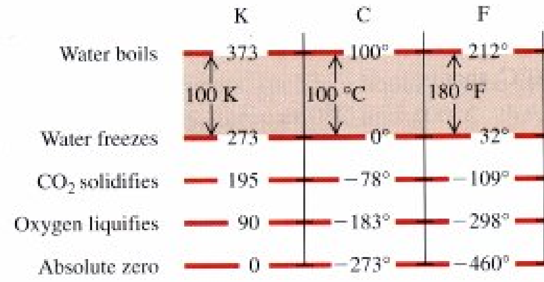
The temperature *difference* $\Delta T = T_2 - T_1$ is 37.00 C° = 37.00 K.

EVALUATE: The Celsius and Kelvin scales have different zero points but the same size degrees. Therefore any temperature difference is the *same* on the Celsius and Kelvin scales but not the same on the Fahrenheit scale.

The Celsius scale has two fixed points, the normal freezing and boiling temperatures of water. But we can define the Kelvin scale using a gas thermometer with only a single reference temperature. We define the ratio of any two temperatures T_1 and T_2 on the Kelvin scale as the ratio of the corresponding gas-thermometer pressures p_1 and p_2:

$$\frac{T_2}{T_1} = \frac{p_2}{p_1} \quad \text{(constant-volume gas thermometer, } T \text{ in kelvins)} \quad (17.4)$$

The pressure p is directly proportional to the Kelvin temperature, as shown in Fig. 17.5b. To complete the definition of T, we need only specify the Kelvin temperature of a single specific state. For reasons of precision and reproducibility the state



17.7 Relations among Kelvin, Celsius, and Fahrenheit temperature scales. Temperatures have been rounded off to the nearest degree.

chosen is the *triple point* of water. This is the unique combination of temperature and pressure at which solid water (ice), liquid water, and water vapor can all coexist. It occurs at a temperature of 0.01°C and a water-vapor pressure of 610 Pa (about 0.006 atm). (This is the pressure of the *water;* it has nothing to do directly with the gas pressure in the *thermometer.*) The triple-point temperature T_{triple} of water is *defined* to have the value $T_{triple} = 273.16\ K$, corresponding to 0.01°C. From Eq. (17.4), if p_{triple} is the pressure in a gas thermometer at temperature T_{triple} and p is the pressure at some other temperature T, then T is given on the Kelvin scale by

$$T = T_{triple}\frac{p}{p_{triple}} = (273.16\ K)\frac{p}{p_{triple}} \tag{17.5}$$

Low-pressure gas thermometers using various gases are found to agree very closely, but they are large, bulky, and very slow to come to thermal equilibrium. They are used principally to establish high-precision standards and to calibrate other thermometers.

The relationships among the three temperature scales we have discussed are shown graphically in Fig. 17.7. The Kelvin scale is called an **absolute temperature scale,** and its zero point ($T = 0\ K = -273.15°C$, the temperature at which $p = 0$ in Eq. (17.5)) is called **absolute zero.** At absolute zero a system of molecules (such as a quantity of a gas, a liquid, or a solid) has its *minimum* possible total energy (kinetic plus potential); because of quantum effects, however, it is *not* correct to say that all molecular motion ceases at absolute zero. To define more completely what we mean by absolute zero, we need to use the thermodynamic principles developed in the next several chapters. We will return to this concept in Chapter 20.

Test Your Understanding

The temperature of the solar corona (see the photograph that opens this chapter) is 2.0×10^7 °C, and the temperature at which helium becomes a liquid at standard pressure is −268.93°C. Express these temperatures in kelvins, and explain why the Kelvin scale is usually employed for very high and very low temperatures.

17.4 | Thermal Expansion

Most materials expand when their temperatures increase. Rising temperatures make the liquid expand in a liquid-in-tube thermometer (Fig. 17.1a) and bend

Solving for the tensile stress F/A required to keep the rod's length constant, we find

$$\frac{F}{A} = -Y\alpha\,\Delta T \qquad \text{(thermal stress)} \qquad (17.12)$$

For a decrease in temperature, ΔT is negative, so F and F/A are positive; this means that a *tensile* force and stress are needed to maintain the length. If ΔT is positive, F and F/A are negative, and the required force and stress are *compressive*.

If there are temperature differences within a body, non-uniform expansion or contraction will result and thermal stresses can be induced. You can break a glass bowl by pouring very hot water into it; the thermal stress between the hot and cold parts of the bowl exceeds the breaking stress of the glass, causing cracks. The same phenomenon makes ice cubes crack when dropped into warm water. Heat-resistant glasses such as Pyrex™ have exceptionally low expansion coefficients and high strength.

Example 17.5 Thermal stress

An aluminum cylinder 10 cm long, with a cross-sectional area of 20 cm^2, is to be used as a spacer between two steel walls. At 17.2°C it just slips in between the walls. When it warms to 22.3°C, calculate the stress in the cylinder and the total force it exerts on each wall, assuming that the walls are perfectly rigid and a constant distance apart.

SOLUTION

IDENTIFY and SET UP: We use Eq. (17.12) to relate the stress (our target variable) to the temperature change. The relevant values of Young's modulus Y and the coefficient of linear expansion α are those for aluminum, the material of which the cylinder is made; we find these values from Tables 11.1 and 17.1, respectively.

EXECUTE: For aluminum, $Y = 7.0 \times 10^{10}$ Pa and $\alpha = 2.4 \times 10^{-5}\ \text{K}^{-1}$. The temperature change is $\Delta T = 22.3°\text{C} - 17.2°\text{C} = 5.1\ \text{C}° = 5.1$ K. The stress is F/A; from Eq. (17.12),

$$\frac{F}{A} = -Y\alpha\,\Delta T = -(0.70 \times 10^{11}\ \text{Pa})(2.4 \times 10^{-5}\ \text{K}^{-1})(5.1\ \text{K})$$
$$= -8.6 \times 10^{6}\ \text{Pa}\ (\text{or } -1200\ \text{lb/in.}^2)$$

The negative sign indicates that compressive rather than tensile stress is needed to keep the cylinder's length constant. This stress is independent of the length and cross-sectional area of the cylinder. The total force F is the cross-sectional area times the stress:

$$F = A\left(\frac{F}{A}\right) = (20 \times 10^{-4}\ \text{m}^2)(-8.6 \times 10^{6}\ \text{Pa})$$
$$= -1.7 \times 10^{4}\ \text{N}$$

or nearly two tons. The negative sign indicates compression.

EVALUATE: The stress on the cylinder and the force it exerts on each wall are immense. This points out the importance of accounting for such thermal stresses in engineering.

Test Your Understanding

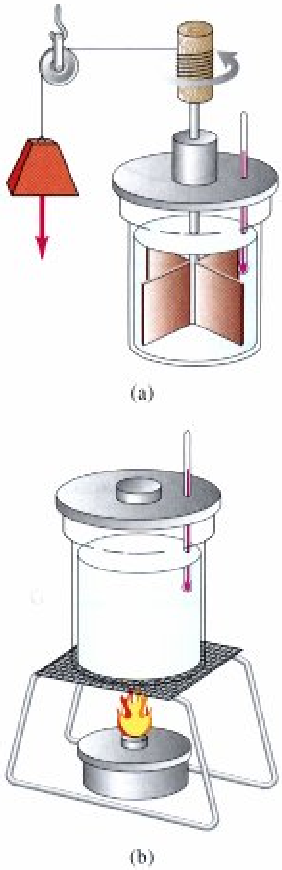
In the bimetallic strip shown in Fig. 17.3, metal 1 is copper. What are two different materials that could be used for metal 2?

17.5 | Quantity of Heat

When you put a cold spoon into a cup of hot coffee, the spoon warms up and the coffee cools down as they approach thermal equilibrium. The interaction that causes these temperature changes is fundamentally a transfer of *energy* from one substance to another. Energy transfer that takes place solely because of a tempe ature difference is called *heat flow* or *heat transfer*, and energy transferred in this way is called **heat.**

An understanding of the relation between heat and other forms of energy emerged gradually during the 18th and 19th centuries. Sir James Joule (1818–1889) studied how water can be warmed by vigorous stirring with a paddle wheel (Fig. 17.13a). The paddle wheel adds energy to the water by doing *work* on it, and Joule found that *the temperature rise is directly proportional to the amount of work done*. The same temperature change can also be caused by putting the water in contact with some hotter body (Fig. 17.13b); hence this interaction must also involve an energy exchange. We will explore the relation between heat and mechanical energy in greater detail in Chapters 19 and 20.

CAUTION It is absolutely essential for you to keep clearly in mind the distinction between *temperature* and *heat*. Temperature depends on the physical state of a material and is a quantitative description of its hotness or coldness. In physics the term "heat" always refers to energy in transit from one body or system to another because of a temperature difference, never to the amount of energy contained within a particular system. We can change the temperature of a body by adding heat to it or taking heat away, or by adding or subtracting energy in other ways, such as mechanical work (Fig. 17.13a). If we cut a body in half, each half has the same temperature as the whole; but to raise the temperature of each half by a given interval, we add *half* as much heat as for the whole.



17.13 The same temperature change of the same system may be accomplished by (a) doing work on it or (b) adding heat to it.

We can define a *unit* of quantity of heat based on temperature changes of some specific material. The **calorie** (abbreviated cal) is defined as *the amount of heat required to raise the temperature of one gram of water from* 14.5°C *to* 15.5°C. The kilocalorie (kcal), equal to 1000 cal, is also used; a food-value calorie is actually a kilocalorie. A corresponding unit of heat using Fahrenheit degrees and British units is the **British thermal unit,** or Btu. One Btu is the quantity of heat required to raise the temperature of one pound (weight) of water 1 F° from 63°F to 64°F.

Because heat is energy in transit, there must be a definite relation between these units and the familiar mechanical energy units such as the joule. Experiments similar in concept to Joule's have shown that

$$1 \text{ cal} = 4.186 \text{ J}$$

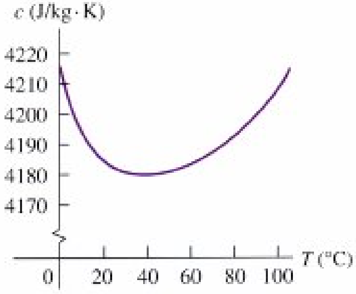
$$1 \text{ kcal} = 1000 \text{ cal} = 4186 \text{ J}$$

$$1 \text{ Btu} = 778 \text{ ft} \cdot \text{lb} = 252 \text{ cal} = 1055 \text{ J}$$

The calorie is not a fundamental SI unit. The International Committee on Weights and Measures recommends using the joule as the basic unit of energy in all forms, including heat. We will follow that recommendation in this book.

Specific Heat

We use the symbol Q for quantity of heat. When it is associated with an infinitesimal temperature change dT, we call it dQ. The quantity of heat Q required to increase the temperature of a mass m of a certain material from T_1 to T_2 is found to be approximately proportional to the temperature change $\Delta T = T_2 - T_1$. It is also proportional to the mass m of material. When you're heating water to make



17.14 Specific heat of water as a function of temperature. The value of c varies by less than 1% between 0°C and 100°C.

tea, you need twice as much heat for two cups as for one if the temperature interval is the same. The quantity of heat needed also depends on the nature of the material; raising the temperature of one kilogram of water by 1 C° requires 4190 J of heat, but only 910 J is needed to raise the temperature of one kilogram of aluminum by 1 C°.

Putting all these relationships together, we have

$$Q = mc\,\Delta T \quad \text{(heat required for temperature change of mass } m\text{)} \qquad (17.13)$$

where c is a quantity, different for different materials, called the **specific heat** of the material. For an infinitesimal temperature change dT and corresponding quantity of heat dQ,

$$dQ = mc\,dT \qquad (17.14)$$

$$c = \frac{1}{m}\frac{dQ}{dT} \quad \text{(specific heat)} \qquad (17.15)$$

In Eqs. (17.13), (17.14), and (17.15), Q (or dQ) and ΔT (or dT) can be either positive or negative. When they are positive, heat enters the body and its temperature increases; when they are negative, heat leaves the body and its temperature decreases.

CAUTION Remember that dQ does not represent a change in the amount of heat *contained* in a body; this is a meaningless concept. Heat is always energy *in transit* as a result of a temperature difference. There is no such thing as "the amount of heat in a body."

The specific heat of water is approximately

$$4190 \text{ J/kg}\cdot\text{K} \qquad 1 \text{ cal/g}\cdot\text{C}° \qquad \text{or } 1 \text{ Btu/lb}\cdot\text{F}°$$

The specific heat of a material always depends somewhat on the initial temperature and the temperature interval. Figure 17.14 shows this variation for water. In the problems and examples in this chapter we will usually ignore this small variation.

Example 17.6 Feed a cold, starve a fever

During a bout with the flu an 80-kg man ran a fever of 39.0°C (102.2°F) instead of the normal body temperature of 37.0°C (98.6°F). Assuming that the human body is mostly water, how much heat is required to raise his temperature by that amount?

SOLUTION

IDENTIFY: This problem uses the relationship between heat (the target variable), mass, specific heat, and temperature change.

SET UP: We are given the values of $m = 80$ kg, $c = 4190$ J/kg·K (for water), and $\Delta T = 39.0°\text{C} - 37.0°\text{C} = 2.0\ \text{C}° = 2.0$ K. We use Eq. (17.13) to determine the required heat.

EXECUTE: From Eq. (17.13),

$$Q = mc\,\Delta T = (80 \text{ kg})(4190 \text{ J/kg}\cdot\text{K})(2.0 \text{ K}) = 6.7 \times 10^5 \text{ J}$$

EVALUATE: This corresponds to 160 kcal, or 160 food-value calories. In fact, the specific heat of the human body is more nearly equal to 3480 J/kg·K, about 83% that of water. The difference is due to the presence of protein, fat, and minerals, which have lower specific heats. With this value of c, the required heat is 5.6×10^5 J = 133 kcal. Either result shows us that were it not for the body's temperature-regulating systems, taking in energy in the form of food would produce measurable changes in body temperature. (In the case of a person with the flu, the elevated temperature results from the body's extra activity in fighting infection.)

Example 17.11 Combustion, temperature change, and phase change

In a particular gasoline camp stove, 30% of the energy released in burning the fuel actually goes to heating the water in the pot on the stove. If we heat 1.00 L (1.00 kg) of water from 20°C to 100°C and boil 0.25 kg of it away, how much gasoline do we burn in the process?

SOLUTION

IDENTIFY and SET UP: In this problem, Eqs. (17.13) and (17.20) apply to the water, all of which undergoes a temperature change and part of which also undergoes a phase change from liquid to gas. This requires a certain amount of heat, which we use to determine the amount of gasoline that must be burned (the target variable).

EXECUTE: The heat required to raise the temperature of the water from 20°C to 100°C is

$$Q_1 = mc\,\Delta T = (1.00\ \text{kg})(4190\ \text{J/kg}\cdot\text{K})(80\ \text{K}) = 3.35 \times 10^5\ \text{J}$$

To boil 0.25 kg of water at 100°C requires

$$Q_2 = mL_v = (0.25\ \text{kg})(2.256 \times 10^6\ \text{J/kg}) = 5.64 \times 10^5\ \text{J}$$

The total energy needed is the sum of these, or 8.99×10^5 J. This is only 0.30 of the total heat of combustion, so that energy is $(8.99 \times 10^5\ \text{J})/0.30 = 3.00 \times 10^6$ J. As we mentioned earlier, one gram of gasoline releases 46,000 J, so the mass of gasoline required is

$$\frac{3.00 \times 10^6\ \text{J}}{46{,}000\ \text{J/g}} = 65\ \text{g}$$

or a volume of about 0.09 L of gasoline.

EVALUATE: This result is a testament to the tremendous amount of energy that can be released by burning even a small quantity of gasoline. Note that most of the heat delivered was used to boil away 0.25 L of water. Can you show that another 123 g of gasoline would be required to boil away the remaining water?

Test Your Understanding

You take a block of ice at 0°C and add heat to it at a steady rate. After a time t, the block of ice has been converted completely to steam at 100°C. What is the temperature of the (ice? water? steam?) at time $t/2$, and what is its phase?

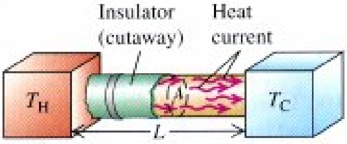
17.7 | Mechanisms of Heat Transfer

We have talked about *conductors* and *insulators,* materials that permit or prevent heat transfer between bodies. Now let's look in more detail at *rates* of energy transfer. In the kitchen you use an aluminum pot for good heat transfer from the stove to whatever you're cooking, but your refrigerator is insulated with a material that *prevents* heat from flowing into the food inside the refrigerator. How do we describe the difference between these two materials?

The three mechanisms of heat transfer are conduction, convection, and radiation. *Conduction* occurs within a body or between two bodies in contact. *Convection* depends on motion of mass from one region of space to another. *Radiation* is heat transfer by electromagnetic radiation, such as sunshine, with no need for matter to be present in the space between bodies.

Conduction

If you hold one end of a copper rod and place the other end in a flame, the end you are holding gets hotter and hotter, even though it is not in direct contact with the flame. Heat reaches the cooler end by **conduction** through the material. On the atomic level, the atoms in the hotter regions have more kinetic energy, on the average, than



17.20 Steady-state heat flow due to conduction in a uniform rod.

their cooler neighbors. They jostle their neighbors, giving them some of their energy. The neighbors jostle *their* neighbors, and so on through the material. The atoms themselves do not move from one region of material to another, but their energy does.

Most metals also use another, more effective mechanism to conduct heat. Within the metal, some electrons can leave their parent atoms and wander through the crystal lattice. These "free" electrons can rapidly carry energy from the hotter to the cooler regions of the metal, so metals are generally good conductors of heat. A metal rod at 20°C feels colder than a piece of wood at 20°C because heat can flow more easily from your hand into the metal. The presence of "free" electrons also causes most metals to be good electrical conductors.

Heat transfer occurs only between regions that are at different temperatures, and the direction of heat flow is always from higher to lower temperature. Figure 17.20 shows a rod of conducting material with cross-sectional area A and length L. The left end of the rod is kept at a temperature T_H and the right end at a lower temperature T_C, so heat flows from left to right. The sides of the rod are covered by an ideal insulator, so no heat transfer occurs at the sides.

When a quantity of heat dQ is transferred through the rod in a time dt, the rate of heat flow is dQ/dt. We call this rate the **heat current,** denoted by H. That is, $H = dQ/dt$. Experiments show that the heat current is proportional to the cross-sectional area A of the rod and to the temperature difference $(T_H - T_C)$ and is inversely proportional to the rod length L. Introducing a proportionality constant k called the **thermal conductivity** of the material, we have

$$H = \frac{dQ}{dt} = kA\frac{T_H - T_C}{L} \quad \text{(heat current in conduction)} \quad (17.21)$$

The quantity $(T_H - T_C)/L$ is the temperature difference *per unit length;* it is called the magnitude of the **temperature gradient.** The numerical value of k depends on the material of the rod. Materials with large k are good conductors of heat; materials with small k are poor conductors or insulators. Equation (17.21) also gives the heat current through a slab or through *any* homogeneous body with uniform cross section A perpendicular to the direction of flow; L is the length of the heat-flow path.

The units of heat current H are units of energy per time, or power; the SI unit of heat current is the watt $(1\ \text{W} = 1\ \text{J/s})$. We can find the units of k by solving Eq. (17.21) for k; you can show that the SI units are W/m · K. Some numerical values of k are given in Table 17.5.

The thermal conductivity of "dead" (that is, nonmoving) air is very small. A wool sweater keeps you warm because it traps air between the fibers. In fact, many insulating materials such as Styrofoam and fiberglass are mostly dead air. Figure 17.21 shows a ceramic material with very unusual thermal properties, including very small conductivity.

If the temperature varies in a non-uniform way along the length of the conducting rod, we introduce a coordinate x along the length and generalize the temperature gradient to be dT/dx. The corresponding generalization of Eq. (17.21) is

$$H = \frac{dQ}{dt} = -kA\frac{dT}{dx} \quad (17.22)$$

Table 17.5 Thermal Conductivities

Substance	k (W/m · K)
Metals	
Aluminum	205.0
Brass	109.0
Copper	385.0
Lead	34.7
Mercury	8.3
Silver	406.0
Steel	50.2
Various solids (representative values)	
Brick, insulating	0.15
Brick, red	0.6
Concrete	0.8
Cork	0.04
Felt	0.04
Fiberglass	0.04
Glass	0.8
Ice	1.6
Rock wool	0.04
Styrofoam	0.01
Wood	0.12–0.04
Gases	
Air	0.024
Argon	0.016
Helium	0.14
Hydrogen	0.14
Oxygen	0.023
Water	0.6

The negative sign shows that heat always flows in the direction of *decreasing* temperature.

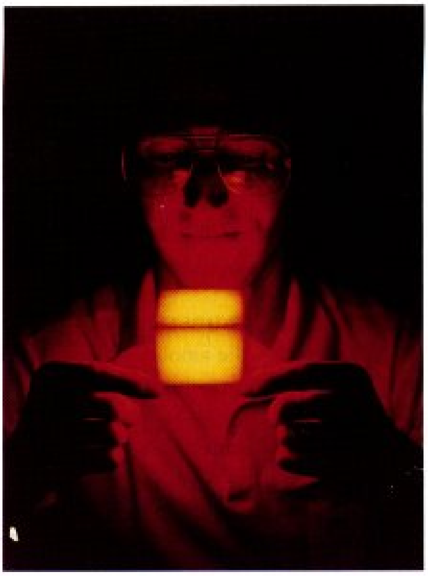
For thermal insulation in buildings, engineers use the concept of **thermal resistance,** denoted by R. The thermal resistance R of a slab of material with area A is defined so that the heat current H through the slab is

$$H = \frac{A(T_H - T_C)}{R} \tag{17.23}$$

where T_H and T_C are the temperatures on the two sides of the slab. Comparing this with Eq. (17.21), we see that R is given by

$$R = \frac{L}{k} \tag{17.24}$$

where L is the thickness of the slab. The SI unit of R is $1\ \text{m}^2 \cdot \text{K/W}$. In the units used for commercial insulating materials in the United States, H is expressed in Btu/h, A is in ft^2, and $T_H - T_C$ in F°. $(1\ \text{Btu/h} = 0.293\ \text{W}.)$ The units of R are then $\text{ft}^2 \cdot \text{F}° \cdot \text{h/Btu}$, though values of R are usually quoted without units; a 6-inch-thick layer of fiberglass has an R value of 19 (that is, $R = 19\ \text{ft}^2 \cdot \text{F}° \cdot \text{h/Btu}$), a two-inch-thick slab of polyurethane foam has a value of 12, and so on. Doubling the thickness doubles the R-value. Common practice in new construction in severe northern climates is to specify R values of around 30 for exterior walls and ceilings. When the insulating material is in layers, such as a plastered wall, fiberglass insulation, and wood exterior siding, the R values are additive. Do you see why? (See Problem 17.110.)



17.21 This protective tile, developed for use in the space shuttle, has extraordinary thermal properties. The extremely small thermal conductivity and small heat capacity of the material make it possible to hold the tile by its edges, even though its temperature is high enough to emit the light for this photograph.

Problem-Solving Strategy

Heat Conduction

IDENTIFY *the relevant concepts:* The concept of heat conduction comes into play whenever two objects at different temperature are placed in contact.

SET UP *the problem using the following steps:*

1. Identify the direction of heat flow in the problem (from hot to cold). In Eq. (17.21), L is always measured along this direction, and A is always an area perpendicular to this direction. Often when a box or other container has an irregular shape but uniform wall thickness, you can approximate it as a flat slab with the same thickness and total wall area.
2. Identify the target variable.

EXECUTE *the solution as follows:*

1. If heat flows through a single object, use Eq. (17.21) to solve for the target variable.
2. In some problems the heat flows through two different materials in succession. The temperature at the interface between the two materials is then intermediate between T_H and T_C; represent it by a symbol such as T. The temperature differences for the two materials are then $(T_H - T)$ and $(T - T_C)$. In steady-state heat flow, the same heat has to pass through both materials in succession, so the heat current H must be *the same* in both materials.
3. If there are two *parallel* heat-flow paths, so that some heat flows through each, then the total H is the sum of the quantities H_1 and H_2 for the separate paths. An example is heat flow from inside to outside a house, both through the glass in a window and through the surrounding frame. In this case the temperature difference is the same for the two paths, but L, A, and k may be different for the two paths.
4. As always, it is essential to use a consistent set of units. If you use a value of k expressed in W/m·K, don't use distances in centimeters, heat in calories, or T in degrees Fahrenheit!

EVALUATE *your answer:* As always, ask yourself whether the results are physically reasonable.

Example 17.12 Conduction through a picnic cooler

A Styrofoam box used to keep drinks cold at a picnic has total wall area (including the lid) of $0.80\ \text{m}^2$ and wall thickness 2.0 cm. It is filled with ice, water, and cans of Omni-Cola at 0°C. What is the rate of heat flow into the box if the temperature of the outside wall is 30°C? How much ice melts in one day?

SOLUTION

IDENTIFY and SET UP: The first target variable is the heat current H. The second is the amount of ice melted, which depends on the heat current (heat per unit time), the elapsed time, and the heat of fusion.

EXECUTE: We assume that the total heat flow is approximately the same as it would be through a flat slab of area $0.80\ \text{m}^2$ and thickness 2.0 cm = 0.020 m (Fig. 17.22). We find k from Table 17.5. From Eq. (17.21) the heat current (rate of heat flow) is

$$H = kA\frac{T_H - T_C}{L} = (0.010\ \text{W/m}\cdot\text{K})(0.80\ \text{m}^2)\frac{30^\circ\text{C} - 0^\circ\text{C}}{0.020\ \text{m}}$$
$$= 12\ \text{W} = 12\ \text{J/s}$$

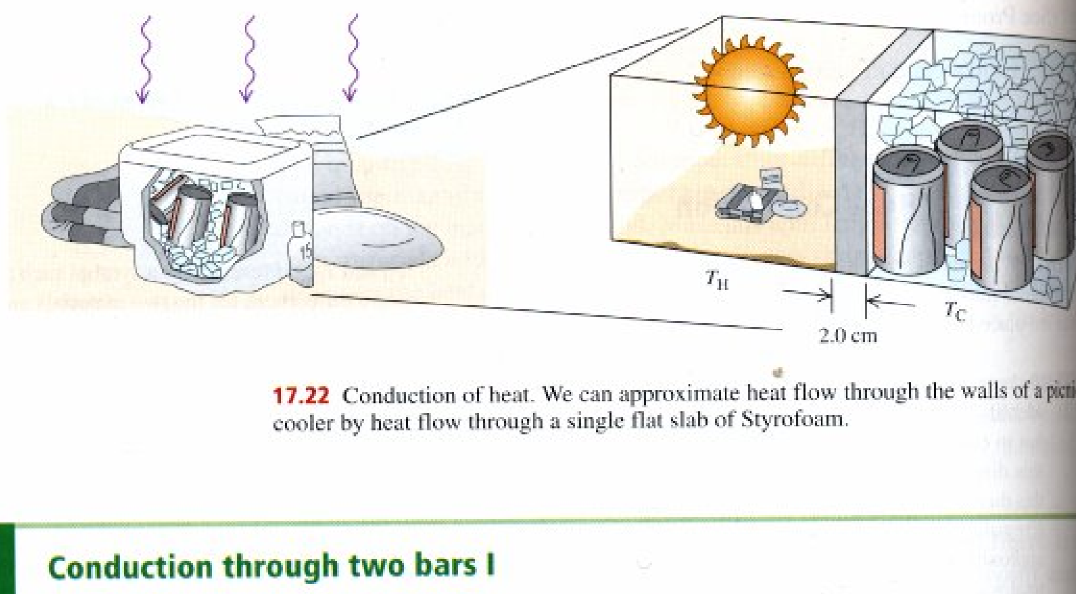
The total heat flow Q in one day (86,400 s) is

$$Q = Ht = (12\ \text{J/s})(86{,}400\ \text{s}) = 1.04 \times 10^6\ \text{J}$$

The heat of fusion of ice is 3.34×10^5 J/kg, so the quantity of ice melted by this quantity of heat is

$$m = \frac{Q}{L_f} = \frac{1.04 \times 10^6\ \text{J}}{3.34 \times 10^5\ \text{J/kg}} = 3.1\ \text{kg}$$

EVALUATE: The low heat current is a result of the low thermal conductivity of Styrofoam. A substantial amount of heat flows in 24 hours, but a relatively small amount of ice melts because the heat of fusion is high.



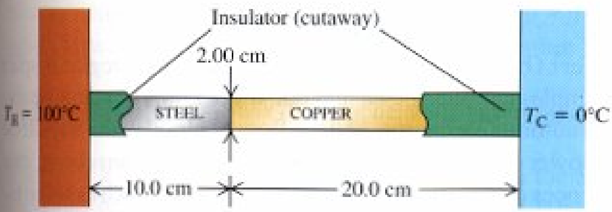
17.22 Conduction of heat. We can approximate heat flow through the walls of a picnic cooler by heat flow through a single flat slab of Styrofoam.

Example 17.13 Conduction through two bars I

A steel bar 10.0 cm long is welded end-to-end to a copper bar 20.0 cm long (Fig. 17.23). Both bars are insulated perfectly on their sides. Each bar has a square cross section, 2.00 cm on a side. The free end of the steel bar is maintained at 100°C by placing it in contact with steam, and the free end of the copper bar is maintained at 0°C by placing it in contact with ice. Find the temperature at the junction of the two bars and the total rate of heat flow.

SOLUTION

IDENTIFY and SET UP: As we discussed in the Problem-Solving Strategy, the heat currents in the two bars must be equal; this idea is the key to the solution. We write Eq. (17.21) twice, once for each bar, and set the heat currents H_{steel} and H_{copper} equal to each other. Both expressions for the heat current involve the temperature T at the junction, which is one of our target variables.



17.23 Heat flow along two metal bars, one of steel and one of copper, connected end-to-end.

EXECUTE: Setting the two heat currents equal,

$$H_{steel} = \frac{k_{steel}A(100°C - T)}{L_{steel}} = H_{copper} = \frac{k_{copper}A(T - 0°C)}{L_{copper}}$$

The areas A are equal and may be divided out. Substituting $L_{steel} = 0.100$ m, $L_{copper} = 0.200$ m, and numerical values of k from Table 17.5, we find

$$\frac{(50.2\ \text{W/m}\cdot\text{K})(100°C - T)}{0.100\ \text{m}} = \frac{(385\ \text{W/m}\cdot\text{K})(T - 0°C)}{0.200\ \text{m}}$$

Rearranging and solving for T, we find

$$T = 20.7°C$$

We can find the total heat current by substituting this value for T back into either of the above expressions:

$$H_{steel} = \frac{(50.2\ \text{W/m}\cdot\text{K})(0.0200\ \text{m})^2(100°C - 20.7°C)}{0.100\ \text{m}}$$
$$= 15.9\ \text{W}$$

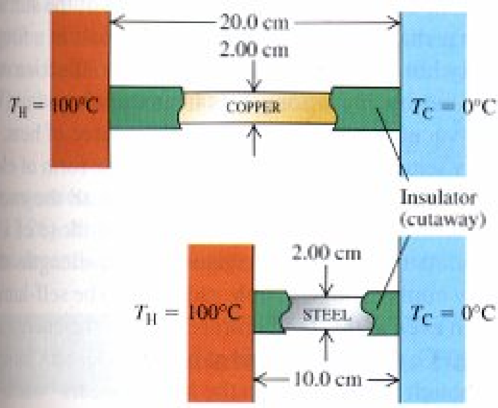
or

$$H_{copper} = \frac{(385\ \text{W/m}\cdot\text{K})(0.0200\ \text{m})^2(20.7°C)}{0.200\ \text{m}} = 15.9\ \text{W}$$

EVALUATE: Even though the steel bar is shorter, the temperature drop across it is much greater than across the copper bar (from 100°C to 20.7°C in the steel versus from 20.7°C to 0°C in the copper). This difference arises because steel is a much poorer conductor than copper.

Example 17.14 Conduction through two bars II

In Example 17.13, suppose the two bars are separated. One end of each bar is maintained at 100°C and the other end of each bar is maintained at 0°C (Fig. 17.24). What is the *total* rate of heat flow in the two bars?



17.24 Heat flow along two metal bars, one of steel and one of copper, parallel to and separated from each other.

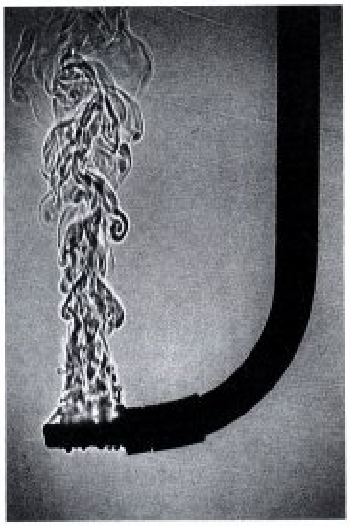
SOLUTION

IDENTIFY and SET UP: In this case the bars are in parallel rather than in series. The total heat current is now the *sum* of the currents in the two bars, and for each bar, $T_H - T_C = 100°C - 0°C = 100$ K.

EXECUTE: We write the heat currents for the two rods individually, then add them to get the total heat current:

$$H = H_{steel} + H_{copper} = \frac{k_{steel}A(T_H - T_C)}{L_{steel}} + \frac{k_{copper}A(T_H - T_C)}{L_{copper}}$$
$$= \frac{(50.2\ \text{W/m}\cdot\text{K})(0.0200\ \text{m})^2(100\ \text{K})}{0.100\ \text{m}}$$
$$+ \frac{(385\ \text{W/m}\cdot\text{K})(0.0200\ \text{m})^2(100\ \text{K})}{0.200\ \text{m}}$$
$$= 20.1\ \text{W} + 77.0\ \text{W} = 97.1\ \text{W}$$

EVALUATE: The heat flow in the copper bar is much greater than that in the steel bar, even though it is longer, because the thermal conductivity of copper is much larger. The total heat flow is much greater than in Example 17.13, partly because the total cross section for heat flow is greater and partly because the full 100-K temperature difference appears across each bar.



17.25 A heating element in the tip of this submerged tube warms the surrounding water, producing a complex pattern of free convection.

Convection

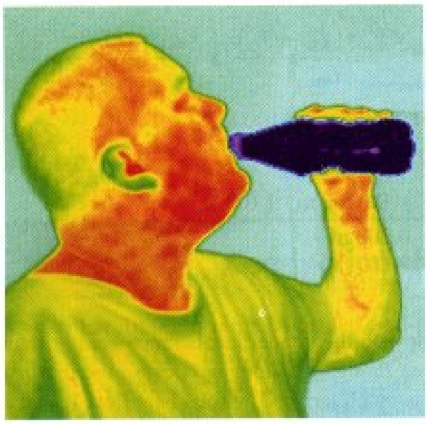
Convection is the transfer of heat by mass motion of a fluid from one region of space to another. Familiar examples include hot-air and hot-water home heating systems, the cooling system of an automobile engine, and the flow of blood in the body. If the fluid is circulated by a blower or pump, the process is called *forced convection*; if the flow is caused by differences in density due to thermal expansion, such as hot air rising, the process is called *natural convection* or *free convection* (Fig. 17.25).

Free convection in the atmosphere plays a dominant role in determining the daily weather, and convection in the oceans is an important global heat-transfer mechanism. On a smaller scale, soaring hawks and glider pilots make use of thermal updrafts from the warm earth. The most important mechanism for heat transfer within the human body (needed to maintain nearly constant temperature in various environments) is *forced* convection of blood, with the heart serving as the pump.

Convective heat transfer is a very complex process, and there is no simple equation to describe it. Here are a few experimental facts:

1. The heat current due to convection is directly proportional to the surface area. This is the reason for the large surface areas of radiators and cooling fins.
2. The viscosity of fluids slows natural convection near a stationary surface, giving a surface film that on a vertical surface typically has about the same insulating value as 1.3 cm of plywood (R value $= 0.7$). Forced convection decreases the thickness of this film, increasing the rate of heat transfer. This is the reason for the "wind-chill factor"; you get cold faster in a cold wind than in still air with the same temperature.
3. The heat current due to convection is found to be approximately proportional to the $\frac{5}{4}$ power of the temperature difference between the surface and the main body of fluid.

Radiation



17.26 This false-color infrared photograph reveals radiation emitted by various parts of the man's body. The strongest emission (colored red) comes from the warmest areas, while there is almost no emission from the bottle of cold beverage.

Radiation is the transfer of heat by electromagnetic waves such as visible light, infrared, and ultraviolet radiation. Everyone has felt the warmth of the sun's radiation and the intense heat from a charcoal grill or the glowing coals in a fireplace. Most of the heat from these very hot bodies reaches you not by conduction or convection in the intervening air but by *radiation*. This heat transfer would occur even if there were nothing but vacuum between you and the source of heat.

Every body, even at ordinary temperatures, emits energy in the form of electromagnetic radiation. At ordinary temperatures, say 20°C, nearly all the energy is carried by infrared waves with wavelengths much longer than those of visible light (see Figs. 17.4 and 17.26). As the temperature rises, the wavelengths shift to shorter values. At 800°C a body emits enough visible radiation to be self-luminous and appears "red-hot," although even at this temperature most of the energy is carried by infrared waves. At 3000°C, the temperature of an incandescent lamp filament, the radiation contains enough visible light so the body appears "white-hot."

The rate of energy radiation from a surface is proportional to the surface area A. The rate increases very rapidly with temperature, depending on the fourth power of the absolute (Kelvin) temperature. The rate also depends on the nature of the surface; this dependence is described by a quantity e called the **emissivity**. A dimensionless number between 0 and 1, it represents txhe ratio of the rate of radiation from a particular surface to the rate of radiation from an equal area of an ideal radiating surface at the same temperature. Emissivity also depends somewhat on temperature. Thus the heat current $H = dQ/dt$ due to radiation from a surface area A with emissivity e at absolute temperature T can be expressed as

$$H = Ae\sigma T^4 \qquad \text{(heat current in radiation)} \qquad (17.25)$$

where σ is a fundamental physical constant called the **Stefan-Boltzmann constant.** This relation is called the **Stefan-Boltzmann law** in honor of its late-19th-century discoverers. The current best numerical value of σ is

$$\sigma = 5.670400(40) \times 10^{-8}\ \mathrm{W/m^2 \cdot K^4}$$

We invite you to check unit consistency in Eq. (17.25). Emissivity (e) is often larger for dark surfaces than for light ones. The emissivity of a smooth copper surface is about 0.3, but e for a dull black surface can be close to unity.

Example 17.15

Heat transfer by radiation

A thin square steel plate, 10 cm on a side, is heated in a blacksmith's forge to a temperature of 800°C. If the emissivity is 0.60, what is the total rate of radiation of energy?

SOLUTION

IDENTIFY and SET UP: We use Eq. (17.25). The target variable is H, the rate of emission of energy. All of the other quantities are given.

EXECUTE: The total surface area, including both sides, is $2(0.10\ \mathrm{m})^2 = 0.020\ \mathrm{m^2}$. We must convert the temperature to the Kelvin scale; 800°C = 1073 K. Then Eq. (17.25) gives

$$\begin{aligned} H &= Ae\sigma T^4 \\ &= (0.020\ \mathrm{m^2})(0.60)(5.67 \times 10^{-8}\ \mathrm{W/m^2 \cdot K^4})(1073\ \mathrm{K})^4 \\ &= 900\ \mathrm{W} \end{aligned}$$

EVALUATE: A blacksmith standing nearby will easily feel heat being radiated from this plate.

While a body at absolute temperature T is radiating, its surroundings at temperature T_s are also radiating, and the body *absorbs* some of this radiation. If it is in thermal equilibrium with its surroundings, $T = T_s$ and the rates of radiation and absorption must be equal. For this to be true, the rate of absorption must be given in general by $H = Ae\sigma T_s^4$. Then the *net* rate of radiation from a body at temperature T with surroundings at temperature T_s is

$$H_{net} = Ae\sigma T^4 - Ae\sigma T_s^4 = Ae\sigma(T^4 - T_s^4) \qquad (17.26)$$

In this equation a positive value of H means a net heat flow *out of* the body. Equation (17.26) shows that for radiation, as for conduction and convection, the heat current depends on the temperature *difference* between two bodies.

Example 17.16

Radiation from the human body

If the total surface area of the human body is 1.20 m² and the surface temperature is 30°C = 303 K, find the total rate of radiation of energy from the body. If the surroundings are at a temperature of 20°C, what is the *net* rate of heat loss from the body by radiation? The emissivity of the body is very close to unity, irrespective of skin pigmentation.

SOLUTION

IDENTIFY and SET UP: The rate of radiation of energy from the body is given by Eq. (17.25), and the net rate of heat loss is given by Eq. (17.26).

EXECUTE: Taking $e = 1$ in Eq. (17.25), we find that the body radiates at a rate

$$\begin{aligned} H &= Ae\sigma T^4 \\ &= (1.20\ \text{m}^2)(1)(5.67 \times 10^{-8}\ \text{W/m}^2\cdot\text{K}^4)(303\ \text{K})^4 \\ &= 574\ \text{W} \end{aligned}$$

This loss is partly offset by *absorption* of radiation, which depends on the temperature of the surroundings. The *net* rate of radiative energy transfer is given by Eq. (17.26):

$$\begin{aligned} H_{net} &= Ae\sigma(T^4 - T_s^4) \\ &= (1.20\ \text{m}^2)(1)(5.67 \times 10^{-8}\ \text{W/m}^2\cdot\text{K}^4) \\ &\quad \times [(303\ \text{K})^4 - (293\ \text{K})^4] = 72\ \text{W} \end{aligned}$$

EVALUATE: The value of H_{net} is positive because the body is losing heat to its colder surroundings.

Heat transfer by radiation is important in some surprising places. A premature baby in an incubator can be cooled dangerously by radiation if the walls of the incubator happen to be cold, even when the *air* in the incubator is warm. Some incubators regulate the air temperature by measuring the baby's skin temperature.

A body that is a good absorber must also be a good emitter. An ideal radiator, with an emissivity of unity, is also an ideal absorber, absorbing *all* of the radiation that strikes it. Such an ideal surface is called an ideal black body or simply a **blackbody.** Conversely, an ideal *reflector,* which absorbs *no* radiation at all, is also a very ineffective radiator.

This is the reason for the silver coatings on vacuum ("Thermos") bottles, invented by Sir James Dewar (1842–1923). A vacuum bottle has double glass walls. The air is pumped out of the spaces between the walls; this eliminates nearly all heat transfer by conduction and convection. The silver coating on the walls reflects most of the radiation from the contents back into the container, and the wall itself is a very poor emitter. Thus a vacuum bottle can keep coffee or soup hot for several hours. The Dewar flask, used to store very cold liquefied gases, is exactly the same in principle.

Test Your Understanding

The ear thermometer in Fig. 17.4 measures the radiation emitted by the eardrum. By what percentage does the radiation rate increase if the eardrum's temperature increases from 37.00°C to 37.10°C?