When it is cold, warm clothes act as insulators to reduce heat loss from the body to the environment by conduction and convection. Heat radiation from a campfire can warm you and your clothes. The fire can also transfer energy directly by heat convection and conduction to what you are cooking. Heat, like work, represents a transfer of energy. Heat is defined as a transfer of energy due to a difference of temperature. Internal energy U is the sum total of all the energy of all the molecules of the system.

CHAPTER 14

Heat

CONTENTS

CHAPTER-OPENING QUESTION—Guess now!

A 5-kg cube of warm iron (60°C) is put in thermal contact with a 10-kg cube of cold iron (15°C). Which statement is valid?

(a) Heat flows spontaneously from the warm cube to the cold cube until both cubes have the same heat content.
(b) Heat flows spontaneously from the warm cube to the cold cube until both cubes have the same temperature.
(c) Heat can flow spontaneously from the warm cube to the cold cube, but can also flow spontaneously from the cold cube to the warm cube.
(d) Heat flows from the larger cube to the smaller one because the larger one has more internal energy.

When a pot of cold water is placed on a hot burner of a stove, the temperature of the water increases. We say that heat "flows" from the hot burner to the cold water. When two objects at different temperatures are put in contact, heat spontaneously flows from the hotter one to the colder one. The spontaneous flow of heat is in the direction tending to equalize the temperature. If the two objects are kept in contact long enough for their temperatures to become equal, the objects are said to be in thermal equilibrium, and there is no further heat flow between them. For example, when a fever thermometer is first placed in your mouth, heat flows from your mouth to the thermometer. When the thermometer reaches the same temperature as the inside of your mouth, the thermometer and your mouth are then in equilibrium, and no more heat flows.

Heat and temperature are often confused. They are very different concepts, and in this Chapter we will make a clear distinction between them. We begin by defining and using the concept of heat. We also discuss how heat is used in calorimetry, how it is involved in changes of state of matter, and the processes of heat transfer—conduction, convection, and radiation.

14–1 Heat as Energy Transfer

We use the term "heat" in everyday life as if we knew what we meant. But the term is often used inconsistently, so it is important for us to define heat clearly, and to clarify the phenomena and concepts related to heat.

We commonly speak of the flow of heat—heat flows from a stove burner to a pot of soup, from the Sun to the Earth, from a person's mouth into a fever thermometer. Heat flows spontaneously from an object at higher temperature to one at lower temperature. Indeed, an eighteenth-century model of heat pictured heat flow as movement of a fluid substance called *caloric*. However, the caloric fluid could never be detected. In the nineteenth century, it was found that the various phenomena associated with heat could be described consistently using a new model that views heat as being akin to work, as we will discuss in a moment. First we note that a common unit for heat, still in use today, is named after caloric. It is called the **calorie** (cal) and is defined as *the amount of heat necessary to raise the temperature of 1 gram of water by 1 Celsius degree*. [To be precise, the particular temperature range from 14.5°C to 15.5°C is specified because the heat required is very slightly different at different temperatures. The difference is less than 1% over the range 0 to 100°C, and we will ignore it for most purposes.] More often used than the calorie is the **kilocalorie** (kcal), which is 1000 calories. Thus *1 kcal is the heat needed to raise the temperature of 1 kg of water by 1 C°*. Often a kilocalorie is called a **Calorie** (with a capital C), and this Calorie (or the kJ) is used to specify the energy value of food. In the British system of units, heat is measured in British thermal units (Btu). One **Btu** is defined as the heat needed to raise the temperature of 1 lb of water by 1 F°. It can be shown (Problem 5) that 1 Btu = 0.252 kcal = 1056 J. Also, one **therm** is 10^5 Btu.

CAUTION
Heat is not a fluid

The idea that heat is related to energy transfer was pursued by a number of scientists in the 1800s, particularly by an English brewer, James Prescott Joule (1818–1889). Joule and others performed a number of experiments that were crucial for establishing our present-day view that heat, like work, represents a transfer of energy. One of Joule's experiments is shown (simplified) in Fig. 14–1. The falling weight causes the paddle wheel to turn. The friction between the water and the paddle wheel causes the temperature of the water to rise slightly (barely measurable, in fact, by Joule). The same temperature rise could also be obtained by heating the water on a hot stove. In this and many other experiments (some involving electrical energy), Joule determined that a given amount of work done was always equivalent to a particular amount of heat input. Quantitatively, 4.186 joules (J) of work was found to be equivalent to 1 calorie (cal) of heat. This is known as the **mechanical equivalent of heat**:

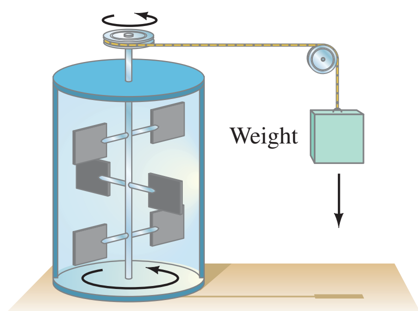


FIGURE 14–1 Joule's experiment on the mechanical equivalent of heat. [The work is transformed into internal energy (Section 14–2).]

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

$$4.186 \text{ kJ} = 1 \text{ kcal}.$$

As a result of these and other experiments, scientists came to interpret heat not as a substance, and not exactly as a form of energy. Rather, heat refers to a *transfer of energy*: when heat flows from a hot object to a cooler one, it is energy that is being transferred from the hot to the cold object. Thus, **heat** is *energy transferred from one object to another because of a difference in temperature*. In SI units, the unit for heat, as for any form of energy, is the joule. Nonetheless, calories and kcal are still sometimes used. Today the calorie is *defined* in terms of the joule (via the mechanical equivalent of heat, above), rather than in terms of the properties of water, as given previously. The latter is still handy to remember: 1 cal raises 1 g of water by 1 C°, or 1 kcal raises 1 kg of water by 1 C°.

CAUTION
Heat is energy transferred because of a ΔT

PHYSICS APPLIED
Working off Calories

EXAMPLE 14–1 ESTIMATE Working off the extra Calories. Suppose you throw caution to the wind and eat 500 Calories of ice cream and cake. To compensate, you want to do an equivalent amount of work climbing stairs or a mountain. How much total height must you climb?

APPROACH The work W you need to do in climbing stairs equals the change in gravitational potential energy: $W = \Delta \text{PE} = mgh$, where h is the vertical height climbed. For this estimate, let us approximate your mass as $m \approx 60$ kg.

SOLUTION 500 food Calories is 500 kcal, which in joules is

$$(500\,\text{kcal})(4.186 \times 10^3\,\text{J/kcal}) = 2.1 \times 10^6\,\text{J}.$$

The work done to climb a vertical height h is $W = mgh$. We solve for h:

$$h = \frac{W}{mg} = \frac{2.1 \times 10^6\,\text{J}}{(60\,\text{kg})(9.80\,\text{m/s}^2)} = 3600\,\text{m}.$$

This is a huge elevation change (over 11,000 ft).

NOTE The human body does not transform food energy with 100% efficiency—it is more like 20% efficient. As we shall discuss in the next Chapter, some energy is always "wasted," so you would actually have to climb only about $(0.2)(3600\,\text{m}) \approx 700$ m, which is more reasonable (about 2300 ft of elevation gain).

14–2 Internal Energy

The sum total of all the energy of all the molecules in an object is called its **internal energy**. (Sometimes **thermal energy** is used to mean the same thing.) We introduce the concept of internal energy now since it will help clarify ideas about heat.

Distinguishing Temperature, Heat, and Internal Energy

CAUTION
Distinguish heat from internal energy and from temperature

Using the kinetic theory, we can make a clear distinction between temperature, heat, and internal energy. Temperature (in kelvins) is a measure of the *average* kinetic energy of individual molecules (Eq. 13–8). Internal energy refers to the *total* energy of all the molecules within the object. (Thus two equal-mass hot ingots of iron may have the same temperature, but two of them have twice as much internal energy as one does.) Heat, finally, refers to a *transfer* of energy from one object to another because of a difference in temperature.

CAUTION
Direction of heat flow depends on temperatures (not on amount of internal energy)

Notice that the direction of heat flow between two objects depends on their temperatures, not on how much internal energy each has. Thus, if 50 g of water at 30°C is placed in contact (or mixed) with 200 g of water at 25°C, heat flows *from* the water at 30°C *to* the water at 25°C even though the internal energy of the 25°C water is much greater because there is so much more of it.

Internal Energy of an Ideal Gas

Let us calculate the internal energy of n moles of an ideal monatomic (one atom per molecule) gas. The internal energy, U, is the sum of the translational kinetic energies of all the atoms. This sum is equal to the average kinetic energy per molecule times the total number of molecules, N:

$$U = N\left(\tfrac{1}{2}m\overline{v^2}\right).$$

Using Eq. 13–8, $\overline{\text{KE}} = \frac{1}{2}m\overline{v^2} = \frac{3}{2}kT$, we can write this as

$$U = \tfrac{3}{2}NkT$$

or (recall Section 13–8)

$$\boxed{U = \tfrac{3}{2}nRT,} \qquad \left[\begin{array}{l}\text{internal energy of}\\ \text{ideal monatomic gas}\end{array}\right] \qquad \textbf{(14–1)}$$

where n is the number of moles. Thus, the internal energy of an ideal gas depends only on temperature and the number of moles (or molecules) of gas.

If the gas molecules contain more than one atom, then the rotational and vibrational energy of the molecules (Fig. 14–2) must also be taken into account. The internal energy will be greater at a given temperature than for a monatomic gas, but it will still be a function only of temperature for an ideal gas.

The internal energy of real gases also depends mainly on temperature, but where real gases deviate from ideal gas behavior, their internal energy depends also somewhat on pressure and volume (due to atomic potential energy).

The internal energy of liquids and solids is quite complicated, for it includes electrical potential energy associated with the forces (or "chemical" bonds) between atoms and molecules.

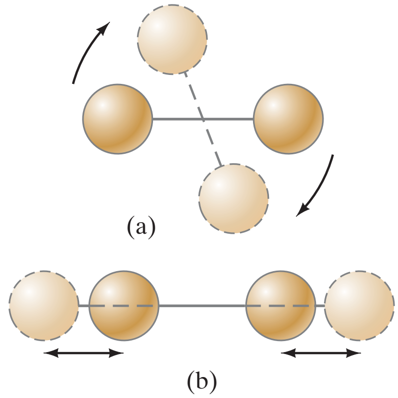


FIGURE 14–2 Besides translational kinetic energy, molecules can have (a) rotational kinetic energy, and (b) vibrational energy (both kinetic and potential).

14–3 Specific Heat

If heat flows into an object, the object's temperature rises (assuming no phase change). But how much does the temperature rise? That depends. As early as the eighteenth century, experimenters had recognized that the amount of heat Q required to change the temperature of a given material is proportional to the mass m of the material present and to the temperature change ΔT. This remarkable simplicity in nature can be expressed in the equation

$$Q = mc\,\Delta T, \tag{14–2}$$

where c is a quantity characteristic of the material called its **specific heat**. Because $c = Q/(m\,\Delta T)$, specific heat is specified in units† of $\mathrm{J/kg\cdot C^\circ}$ (the proper SI unit) or $\mathrm{kcal/kg\cdot C^\circ}$. For water at 15°C and a constant pressure of 1 atm, $c = 4.186 \times 10^3\,\mathrm{J/kg\cdot C^\circ}$ or $1.00\,\mathrm{kcal/kg\cdot C^\circ}$, since, by definition of the cal and the joule, it takes 1 kcal of heat to raise the temperature of 1 kg of water by 1 C°. Table 14–1 gives the values of specific heat for other substances at 20°C. The values of c depend to some extent on temperature (as well as slightly on pressure), but for temperature changes that are not too great, c can often be considered constant.

TABLE 14–1 Specific Heats
(at 1 atm constant pressure and 20°C unless otherwise stated)

Substance	**Specific Heat, c** J/kg · C°	kcal/kg · C° (= cal/g · C°)
Aluminum	900	0.22
Alcohol (ethyl)	2400	0.58
Copper	390	0.093
Glass	840	0.20
Iron or steel	450	0.11
Lead	130	0.031
Marble	860	0.21
Mercury	140	0.033
Silver	230	0.056
Wood	1700	0.4
Water		
Ice (−5°C)	2100	0.50
Liquid (15°C)	4186	1.00
Steam (110°C)	2010	0.48
Human body (average)	3470	0.83
Protein	1700	0.4

EXAMPLE 14–2 How heat transferred depends on specific heat. (*a*) How much heat input is needed to raise the temperature of an empty 20-kg vat made of iron from 10°C to 90°C? (*b*) What if the vat is filled with 20 kg of water?

APPROACH We apply Eq. 14–2 to the different materials involved.

SOLUTION (*a*) Our system is the iron vat alone. From Table 14–1, the specific heat of iron is $450\,\mathrm{J/kg\cdot C^\circ}$. The change in temperature is $(90°\mathrm{C} - 10°\mathrm{C}) = 80\,\mathrm{C^\circ}$. Thus,

$$Q_{\mathrm{vat}} = mc\,\Delta T = (20\,\mathrm{kg})(450\,\mathrm{J/kg\cdot C^\circ})(80\,\mathrm{C^\circ}) = 7.2 \times 10^5\,\mathrm{J} = 720\,\mathrm{kJ}.$$

(*b*) Our system is the vat plus the water. The water alone would require

$$Q_{\mathrm{w}} = mc\,\Delta T = (20\,\mathrm{kg})(4186\,\mathrm{J/kg\cdot C^\circ})(80\,\mathrm{C^\circ}) = 6.7 \times 10^6\,\mathrm{J} = 6700\,\mathrm{kJ},$$

or almost 10 times what an equal mass of iron requires. The total, for the vat plus the water, is $720\,\mathrm{kJ} + 6700\,\mathrm{kJ} = 7400\,\mathrm{kJ}$.

NOTE In (*b*), the iron vat and the water underwent the same temperature change, $\Delta T = 80\,\mathrm{C^\circ}$, but their specific heats are different.

If the iron vat in part (*a*) of Example 14–2 had been *cooled* from 90°C to 10°C, 720 kJ of heat would have flowed *out* of the iron. In other words, Eq. 14–2 is valid for heat flow either in or out, with a corresponding increase or decrease in temperature.

We saw in part (*b*) of Example 14–2 that water requires almost 10 times as much heat as an equal mass of iron to make the same temperature change. Water has one of the highest specific heats of all substances, which makes it an ideal substance for hot-water space-heating systems and other uses that require a minimal drop in temperature for a given amount of heat transfer. It is the water content, too, that causes the apples rather than the crust in hot apple pie to burn our tongues, through heat transfer.

PHYSICS APPLIED
Practical effects of water's high specific heat

†Note that $\mathrm{J/kg\cdot C^\circ}$ means $\dfrac{\mathrm{J}}{\mathrm{kg\cdot C^\circ}}$ and *not* $(\mathrm{J/kg})\cdot\mathrm{C^\circ} = \mathrm{J\cdot C^\circ/kg}$ (otherwise we would have written it that way).

CONCEPTUAL EXAMPLE 14–3 **A very hot frying pan.** You accidentally let an empty iron frying pan get very hot on the stove (200°C or even more). What happens when you dunk it into a few inches of cool water in the bottom of the sink? Will the final temperature be midway between the initial temperatures of the water and pan? Will the water start boiling? Assume the mass of water is roughly the same as the mass of the frying pan.

RESPONSE Experience may tell you that the water warms up—perhaps by as much as 10 or 20 degrees. The water doesn't come close to boiling. The water's temperature increase is a lot less than the frying pan's temperature decrease. Why? Because the mass of water is roughly equal to that of the pan, and iron has a specific heat nearly 10 times smaller than that of water (Table 14–1). As heat leaves the frying pan and enters the water, the iron pan's temperature change will be about 10 times greater than that of the water. If, instead, you let a few drops of water fall onto the hot pan, that very small mass of water will sizzle and boil away (the pan's mass may be hundreds of times larger than that of the water drops).

EXERCISE A Return to the Chapter-Opening Question, page 390, and answer it again now. Try to explain why you may have answered differently the first time.

*Specific Heats for Gases

Specific heats for gases are more complicated than for solids and liquids, which change in volume only slightly with a change in temperature. Gases change strongly in volume with a change in temperature at constant pressure, as we saw in Chapter 13 with the gas laws; or, if kept at constant volume, the pressure in a gas changes strongly with temperature. The specific heat of a gas depends very much on how the process of changing its temperature is carried out. Most commonly, we deal with the specific heats of gases kept (a) at constant pressure (c_P) or (b) at constant volume (c_V). Some values are given in Table 14–2, where we see that c_P is always greater than c_V. For liquids and solids, this distinction is usually negligible. More details are given in Appendix D on molecular specific heats and the equipartition of energy.

TABLE 14–2
Specific Heats of Gases
(kcal/kg · C°)

Gas	c_P (constant pressure)	c_V (constant volume)
Steam (100°C)	0.482	0.350
Oxygen	0.218	0.155
Helium	1.15	0.75
Carbon dioxide	0.199	0.153
Nitrogen	0.248	0.177

14–4 Calorimetry—Solving Problems

In discussing heat and thermodynamics, we often consider a particular **system**, which is any object or set of objects we choose to consider. Everything else in the universe is its "environment" (or the "surroundings"). There are several categories of systems. A **closed system** is one for which no mass enters or leaves (but energy may be exchanged with the environment). In an **open system**, mass may enter or leave (as may energy). Many (idealized) systems we study in physics are closed systems. But many systems, including plants and animals, are open systems since they exchange materials (food, oxygen, waste products) with the environment. A closed system is said to be **isolated** if no energy in any form (as well as no mass) passes across its boundaries; otherwise it is not isolated.

A perfectly isolated system is an ideal, but we often try to set up a system that can be closely approximated as an isolated system (preferably one we can deal with fairly easily). When different parts of an isolated system are at different temperatures, heat will flow (energy is transferred) from the part at higher temperature to the part at lower temperature—that is, within the system—until **thermal equilibrium** is reached, meaning the entire system is at the same temperature. For an isolated system, no energy is transferred into or out of it. So we can apply *conservation of energy* to such an isolated system. A simple intuitive way to set up a conservation of energy equation is to write that the heat lost by one part of the system is equal to the heat gained by the other part:

$$\text{heat lost} = \text{heat gained} \qquad \text{[isolated system]}$$

or

$$\text{energy } \textit{out} \text{ of one part} = \text{energy } \textit{into} \text{ another part.} \qquad \text{[isolated system]}$$

EXAMPLE 14–4 **The cup cools the tea.** If 200 cm^3 of tea at 95°C is poured into a 150-g glass cup initially at 25°C (Fig. 14–3), what will be the common final temperature T of the tea and cup when equilibrium is reached, assuming no heat flows to the surroundings?

APPROACH We apply conservation of energy to our system of tea plus cup, which we are assuming is isolated: all of the heat that leaves the tea flows into the cup. We can use the specific heat equation, Eq. 14–2, to determine how the heat flow is related to the temperature changes.

SOLUTION Because tea is mainly water, we can take its specific heat as $4186\ \mathrm{J/kg\cdot C^\circ}$ (Table 14–1), and its mass m is its density times its volume $(V = 200\ \mathrm{cm^3} = 200 \times 10^{-6}\ \mathrm{m^3})$: $m = \rho V = (1.0 \times 10^3\ \mathrm{kg/m^3})(200 \times 10^{-6}\ \mathrm{m^3}) =$ 0.20 kg. We use Eq. 14–2, apply conservation of energy, and let T be the as yet unknown final temperature:

$$\begin{aligned} \text{heat lost by tea} &= \text{heat gained by cup} \\ m_{\text{tea}}\, c_{\text{tea}}(95^\circ\mathrm{C} - T) &= m_{\text{cup}}\, c_{\text{cup}}(T - 25^\circ\mathrm{C}). \end{aligned}$$

Putting in numbers and using Table 14–1 ($c_{\text{cup}} = 840\ \mathrm{J/kg\cdot C^\circ}$ for glass), we solve for T, and find

$$\begin{aligned} (0.20\ \mathrm{kg})(4186\ \mathrm{J/kg\cdot C^\circ})(95^\circ\mathrm{C} - T) &= (0.15\ \mathrm{kg})(840\ \mathrm{J/kg\cdot C^\circ})(T - 25^\circ\mathrm{C}) \\ 79{,}500\ \mathrm{J} - (837\ \mathrm{J/C^\circ})T &= (126\ \mathrm{J/C^\circ})T - 3150\ \mathrm{J} \\ T &= 86^\circ\mathrm{C}. \end{aligned}$$

The tea drops in temperature by 9 C° by coming into equilibrium with the cup.

NOTE The cup increases in temperature by $86^\circ\mathrm{C} - 25^\circ\mathrm{C} = 61\ \mathrm{C^\circ}$. Its much greater change in temperature (compared with that of the tea water) is due to its much smaller specific heat compared to that of water.

NOTE In this calculation, the ΔT (of Eq. 14–2, $Q = mc\,\Delta T$) is a positive quantity on both sides of our conservation of energy equation. On the left is "heat lost" and ΔT is the initial minus the final temperature $(95^\circ\mathrm{C} - T)$, whereas on the right is "heat gained" and ΔT is the final minus the initial temperature.

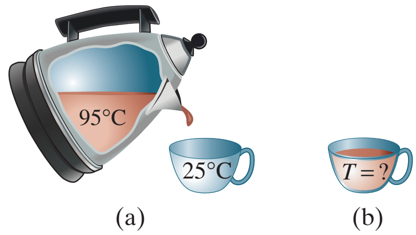


FIGURE 14–3 Example 14–4.

CAUTION

When using heat lost = heat gained, ΔT is positive on both sides

Another, perhaps more general, way to set up the energy conservation equation for heat transfer within an isolated system is to write that the sum of all internal heat transfers within the system adds up to zero:

$$\Sigma Q = 0. \qquad \text{[isolated system]} \qquad \textbf{(14–3)}$$

Each Q represents the heat entering or leaving one part of the system. Each term is written as $Q = mc(T_\mathrm{f} - T_\mathrm{i})$, and $\Delta T = T_\mathrm{f} - T_\mathrm{i}$ is always the final minus the initial temperature. ΔT can be either positive or negative, depending on whether heat flows into or out of that part. Let us redo Example 14–4 using $\Sigma Q = 0$.

PROBLEM SOLVING

Alternate approach: $\Sigma Q = 0$

EXAMPLE 14–4′ **Alternate Solution, $\Sigma Q = 0$.**

APPROACH We use Eq. 14–3, $\Sigma Q = 0$.

SOLUTION For each term, $\Delta T = T_\mathrm{f} - T_\mathrm{i}$. If $T_\mathrm{f} < T_\mathrm{i}$, then $\Delta T < 0$. Equation 14–3, $\Sigma Q = 0$, becomes

$$m_{\text{cup}}\, c_{\text{cup}}(T - 25^\circ\mathrm{C}) + m_{\text{tea}}\, c_{\text{tea}}(T - 95^\circ\mathrm{C}) = 0.$$

The second term is negative because T will be less than 95°C. Solving the algebra gives the same result, $T = 86^\circ\mathrm{C}$.

You are free to use either approach. They are entirely equivalent, algebraically. For example, if you move the first term in the displayed equation of the alternate Example 14–4 over to the other side of the equals sign, you obtain the "heat lost = heat gained" equation in the first version of Example 14–4.

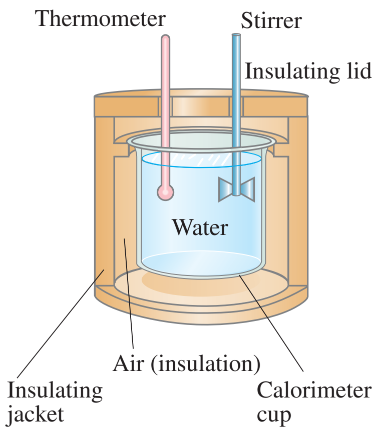


FIGURE 14–4 Simple water calorimeter.

The exchange of energy, as exemplified in Example 14–4, is the basis for a technique known as **calorimetry**, which is the quantitative measurement of heat exchange. To make such measurements, a **calorimeter** is used; a simple water calorimeter is shown in Fig. 14–4. It is very important that the calorimeter be well insulated so that almost no heat is exchanged with the surroundings. One important use of the calorimeter is in the determination of specific heats of substances. In the technique known as the "method of mixtures," a sample of a substance is heated to a high temperature, which is accurately measured, and then quickly placed in the cool water of the calorimeter. The heat lost by the sample will be gained by the water and the calorimeter cup. By measuring the final temperature of the mixture, the specific heat can be calculated, as illustrated in the following Example.

EXAMPLE 14–5 **Unknown specific heat determined by calorimetry.** An engineer wishes to determine the specific heat of a new metal alloy. A 0.150-kg sample of the alloy is heated to 540°C. It is then quickly placed in 0.400 kg of water at 10.0°C, which is contained in a 0.200-kg aluminum calorimeter cup. The final temperature of the system is 30.5°C. Calculate the specific heat of the alloy.

APPROACH We apply conservation of energy to our system, which we take to be the alloy sample, the water, and the calorimeter cup. We assume this system is isolated, and apply Eq. 14–3, $\Sigma Q = 0$.

SOLUTION Each term is of the form $Q = mc(T_\mathrm{f} - T_\mathrm{i})$. Thus $\Sigma Q = 0$ gives

$$m_\mathrm{a} c_\mathrm{a}\, \Delta T_\mathrm{a} + m_\mathrm{w} c_\mathrm{w}\, \Delta T_\mathrm{w} + m_\mathrm{cal} c_\mathrm{cal}\, \Delta T_\mathrm{cal} = 0$$

where the subscripts a, w, and cal refer to the alloy, water, and calorimeter, respectively, and each ΔT is the final temperature (30.5°C) minus the initial temperature for each object. When we put in values and use Table 14–1, this equation becomes

$$(0.150\,\mathrm{kg})(c_\mathrm{a})(30.5°\mathrm{C} - 540°\mathrm{C}) + (0.400\,\mathrm{kg})(4186\,\mathrm{J/kg\cdot C°})(30.5°\mathrm{C} - 10.0°\mathrm{C})$$
$$+ (0.200\,\mathrm{kg})(900\,\mathrm{J/kg\cdot C°})(30.5°\mathrm{C} - 10.0°\mathrm{C}) = 0,$$

or

$$-(76.4\,\mathrm{kg\cdot C°})c_\mathrm{a} + 34{,}300\,\mathrm{J} + 3690\,\mathrm{J} = 0.$$

Solving for c_a we obtain:

$$c_\mathrm{a} = 497\,\mathrm{J/kg\cdot C°} \approx 500\,\mathrm{J/kg\cdot C°}.$$

NOTE We rounded off because we have ignored any heat transferred to the thermometer and the stirrer (which is used to quicken the heat transfer process and thus reduce heat loss to the outside). To take them into account we would have to add (small) additional terms to the equation.

PROBLEM SOLVING
Be sure to consider all possible sources of energy transfer

In all Examples and Problems of this sort, be sure to include *all* objects that gain or lose heat (within reason). For simplicity, we have ignored very small masses, such as the thermometer and the stirrer, which will affect the energy balance only very slightly.

Bomb Calorimeter

PHYSICS APPLIED
Measuring Calorie content

A **bomb calorimeter** is used to measure the thermal energy released when a substance burns (including foods) to determine their Calorie content. A carefully weighed sample of the substance, with an excess amount of oxygen, is placed in a sealed container (the "bomb"). The bomb is placed in the water of the calorimeter and a fine wire passing into the bomb is then heated to ignite the mixture. The Calorie content of foods determined in this way can be unreliable because our bodies may not metabolize all the available energy (which would be excreted). Careful measurements and calculations need to take this into account.

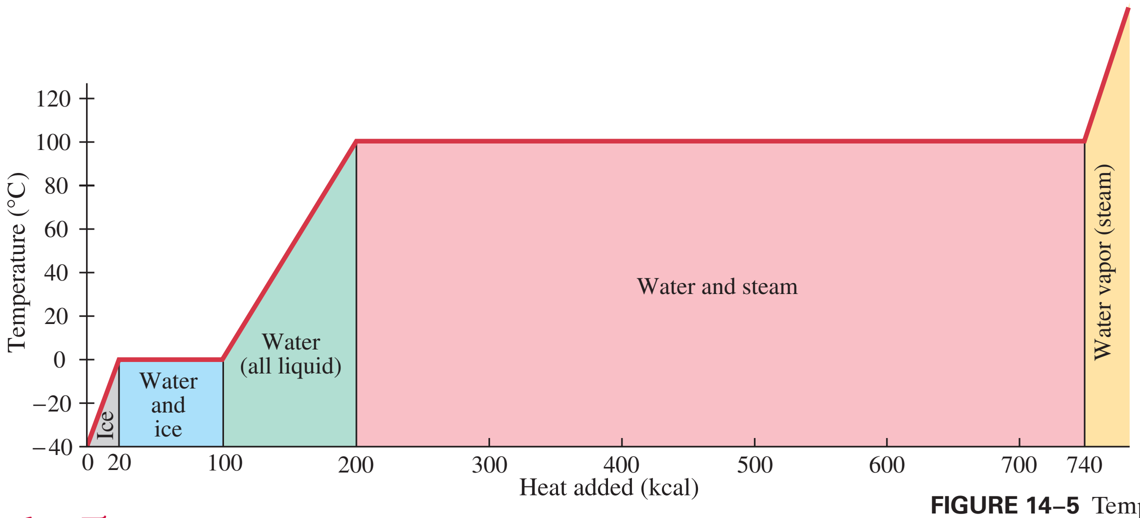


FIGURE 14–5 Temperature as a function of the heat added to bring 1.0 kg of ice at −40°C to steam above 100°C.

14–5 Latent Heat

When a material changes phase from solid to liquid, or from liquid to gas (see also Section 13–11), a certain amount of energy is involved in this **change of phase**. For example, let us trace what happens when a 1.0-kg block of ice at −40°C is heated at a slow steady rate until all the ice has changed to water, then the (liquid) water is heated to 100°C and changed to steam, and heated further above 100°C, all at 1 atm pressure. As shown in the graph of Fig. 14–5, as the ice is heated starting at −40°C, its temperature rises at a rate of about 2 C°/kcal of heat added (since for ice, $c \approx 0.50\,\mathrm{kcal/kg \cdot C^\circ}$). However, when 0°C is reached, the temperature stops increasing even though heat is still being added. The ice gradually changes to water in the liquid state, with no change in temperature. After about 40 kcal has been added at 0°C, half the ice remains and half has changed to water. After about 80 kcal, or 330 kJ, has been added, all the ice has changed to water, still at 0°C. Continued addition of heat causes the water's temperature to again increase, now at a rate of 1 C°/kcal. When 100°C is reached, the temperature again remains constant as the heat added changes the liquid water to vapor (steam). About 540 kcal (2260 kJ) is required to change the 1.0 kg of water completely to steam, after which the graph rises again, indicating that the temperature of the steam rises as heat is added.

The heat required to change 1.0 kg of a substance from the solid to the liquid state is called the **heat of fusion**; it is denoted by L_F. The heat of fusion of water is 79.7 kcal/kg or, in proper SI units, 333 kJ/kg $(= 3.33 \times 10^5\,\mathrm{J/kg})$. The heat required to change a substance from the liquid to the vapor phase is called the **heat of vaporization**, L_V. For water it is 539 kcal/kg or 2260 kJ/kg. Other substances follow graphs similar to Fig. 14–5, although the melting-point and boiling-point temperatures are different, as are the specific heats and heats of fusion and vaporization. Values for the heats of fusion and vaporization, which are also called the **latent heats**, are given in Table 14–3 for a number of substances.

TABLE 14–3 Latent Heats (at 1 atm)

Substance	Melting Point (°C)	Heat of Fusion		Boiling Point (°C)	Heat of Vaporization	
		kJ/kg	kcal/kg†		kJ/kg	kcal/kg†
Oxygen	−218.8	14	3.3	−183	210	51
Nitrogen	−210.0	26	6.1	−195.8	200	48
Ethyl alcohol	−114	104	25	78	850	204
Ammonia	−77.8	33	8.0	−33.4	137	33
Water	0	333	79.7	100	2260	539
Lead	327	25	5.9	1750	870	208
Silver	961	88	21	2193	2300	558
Iron	1538	289	69.1	3023	6340	1520
Tungsten	3410	184	44	5900	4800	1150

†Numerical values in kcal/kg are the same in cal/g.

The heats of vaporization and fusion also refer to the amount of heat *released* by a substance when it changes from a gas to a liquid, or from a liquid to a solid. Thus, steam releases 2260 kJ/kg when it changes to water, and water releases 333 kJ/kg when it becomes ice. [In these cases of heat release, $Q < 0$ when using the $\Sigma Q = 0$ approach, Eq. 14–3.]

The heat involved in a change of phase depends not only on the latent heat but also on the total mass of the substance. That is,

$$Q = mL, \tag{14–4}$$

where L is the latent heat of the particular process and substance, m is the mass of the substance, and Q is the heat added or released during the phase change. For example, when 5.00 kg of water freezes at 0°C, $(5.00\,\text{kg})(3.33 \times 10^5\,\text{J/kg}) = 1.67 \times 10^6\,\text{J}$ of energy is released.

EXERCISE B A pot of water is boiling on a gas stove, and then you turn up the heat. What happens? (*a*) The temperature of the water starts increasing. (*b*) There is a tiny decrease in the rate of water loss by evaporation. (*c*) The rate of water loss by evaporation increases. (*d*) There is an appreciable increase in both the rate of boiling and the temperature of the water. (*e*) None of these.

Calorimetry sometimes involves a change of state, as the following Examples show. Indeed, latent heats are often measured using calorimetry.

EXAMPLE 14–6 **Making ice.** How much energy does a freezer have to remove from 1.5 kg of water at 20°C to make ice at −12°C?

APPROACH We need to calculate the total energy removed by adding the heat outflow (1) to reduce the water temperature from 20°C to 0°C, (2) to change the liquid water to solid ice at 0°C, and (3) to lower the ice temperature from 0°C to −12°C.

SOLUTION The heat Q that needs to be removed from the 1.5 kg of water is

$$\begin{aligned} Q &= mc_{\text{w}}(20°\text{C} - 0°\text{C}) + mL_{\text{F}} + mc_{\text{ice}}[0° - (-12°\text{C})] \\ &= (1.5\,\text{kg})(4186\,\text{J/kg}\cdot\text{C}°)(20\,\text{C}°) + (1.5\,\text{kg})(3.33 \times 10^5\,\text{J/kg}) \\ &\quad + (1.5\,\text{kg})(2100\,\text{J/kg}\cdot\text{C}°)(12\,\text{C}°) \\ &= 6.6 \times 10^5\,\text{J} = 660\,\text{kJ}. \end{aligned}$$

EXERCISE C Which process in Example 14–6 required the greatest heat loss?

EXAMPLE 14–7 **ESTIMATE** **Will all the ice melt?** At a reception, a 0.50-kg chunk of ice at −10°C is placed in 3.0 kg of "iced" tea at 20°C. At what temperature and in what phase will the final mixture be? The tea can be considered as water. Ignore any heat flow to the surroundings, including the container.

PROBLEM SOLVING
First determine (or estimate) the final state

APPROACH Before we can write down an equation applying conservation of energy, we must first check to see if the final state will be all ice, a mixture of ice and water at 0°C, or all water. To bring the 3.0 kg of water at 20°C down to 0°C would require an energy release of

$$m_{\text{w}}c_{\text{w}}(20°\text{C} - 0°\text{C}) = (3.0\,\text{kg})(4186\,\text{J/kg}\cdot\text{C}°)(20\,\text{C}°) = 250{,}000\,\text{J}.$$

On the other hand, to raise the ice from −10°C to 0°C would require

$$\begin{aligned} m_{\text{ice}}c_{\text{ice}}[0°\text{C} - (-10°\text{C})] &= (0.50\,\text{kg})(2100\,\text{J/kg}\cdot\text{C}°)(10\,\text{C}°) \\ &= 10{,}500\,\text{J}, \end{aligned}$$

and to change the ice to water at 0°C would require

$$m_{\text{ice}}L_{\text{F}} = (0.50\,\text{kg})(333\,\text{kJ/kg}) = 167{,}000\,\text{J}.$$

The sum of the last two quantities is $10.5\,\text{kJ} + 167\,\text{kJ} = 177\,\text{kJ}$. This is not enough energy to bring the 3.0 kg of water at 20°C down to 0°C, so we see that the mixture must end up all water, somewhere between 0°C and 20°C.

SOLUTION To determine the final temperature T, we apply conservation of energy. We present both of the techniques discussed in Section 14–4.

Method 1: "$\Sigma Q = 0$" gives

$$\begin{pmatrix}\text{heat to raise}\\ \text{0.50 kg of ice}\\ \text{from } -10^\circ\text{C}\\ \text{to } 0^\circ\text{C}\end{pmatrix} + \begin{pmatrix}\text{heat to change}\\ \text{0.50 kg}\\ \text{of ice}\\ \text{to water}\end{pmatrix} + \begin{pmatrix}\text{heat to raise}\\ \text{0.50 kg of water}\\ \text{from } 0^\circ\text{C}\\ \text{to } T\end{pmatrix} + \begin{pmatrix}\text{heat lost by}\\ \text{3.0 kg of}\\ \text{water cooling}\\ \text{from } 20^\circ\text{C to } T\end{pmatrix} = 0.$$

Using some of the results from the "Approach" above, we obtain

$$10{,}500\text{ J} + 167{,}000\text{ J} + (0.50\text{ kg})(4186\text{ J/kg}\cdot\text{C}^\circ)(T - 0^\circ\text{C}) + (3.0\text{ kg})(4186\text{ J/kg}\cdot\text{C}^\circ)(T - 20^\circ\text{C}) = 0.$$

Solving for T we obtain

$$T = 5.0^\circ\text{C}.$$

Method 2: "heat gained = heat lost" produces a word equation like the one above (the last plus sign becomes an equals sign and we lose the "= 0"). That is,

$$10{,}500\text{ J} + 167{,}000\text{ J} + (0.50\text{ kg})(4186\text{ J/kg}\cdot\text{C}^\circ)(T - 0^\circ\text{C}) = (3.0\text{ kg})(4186\text{ J/kg}\cdot\text{C}^\circ)(20^\circ\text{C} - T).$$

The term on the right is heat lost by 3.0 kg of water cooling from 20°C to T; here $\Delta T = T_\text{i} - T_\text{f} = 20^\circ\text{C} - T$ for this approach. Algebraically, this equation is identical to the one above in the first method.†

EXERCISE D How much more ice at −10°C would be needed in Example 14–7 to bring the tea down to 0°C, while just melting all the ice?

PROBLEM SOLVING

Then determine the final temperature

PROBLEM SOLVING

Calorimetry

1. Be sure you have sufficient information to apply energy conservation. Ask yourself: **is the system isolated** (or nearly so, enough to get a good estimate)? Do we know or can we calculate all significant sources of energy transfer?
2. Apply **conservation of energy**. Either write
$$\text{heat gained} = \text{heat lost},$$
or use
$$\Sigma Q = 0.$$
3. If **no phase changes** occur, each term in the energy conservation equation will have the form
$$Q = mc\,\Delta T.$$
4. If **phase changes** do or might occur, there may be terms in the energy conservation equation of the form $Q = mL$, where L is the latent heat. But *before* applying energy conservation, determine (or estimate) in which phase the final state will be, as we did in Example 14–7 by calculating the different contributing values for heat Q.
5. Note that when the system reaches thermal **equilibrium**, the final **temperature** of each substance will have the *same* value. There is only one T_f.
6. **Solve** your energy equation for the unknown.

Evaporation

The latent heat to change a liquid to a gas is needed not only at the boiling point. Water can change from the liquid to the gas phase even at room temperature. This process is called **evaporation** (see previous Chapter, Section 13–12). The value of the heat of vaporization of water increases slightly with a decrease in temperature: at 20°C, for example, it is 2450 kJ/kg (585 kcal/kg) compared to 2260 kJ/kg (= 539 kcal/kg) at 100°C. When water evaporates, the remaining liquid cools, because the energy required (the latent heat of vaporization) comes from the water itself; so its internal energy, and therefore its temperature, must drop.‡

†For any algebraic equation $A = B$, if you subtract B from both sides you get $A - B = B - B = 0$, or $A - B = 0$. Thus, if you move a term from one side of an equals sign to the other, the sign (+ or −) changes.

‡Also from the point of view of kinetic theory, evaporation is a cooling process because it is the fastest-moving molecules that escape from the surface (Section 13–12). Hence the average speed of the remaining molecules is less, so by Eq. 13–8 the temperature is less.

PHYSICS APPLIED
Body temperature

Evaporation of water from skin is one of the most important ways the body uses to control its temperature. When the temperature of the blood rises slightly above normal, the hypothalamus region of the brain detects this temperature increase and sends a signal to the sweat glands to increase their production. The energy (latent heat) needed to vaporize this water comes from the body, so the body cools.

Kinetic Theory of Latent Heats

We can make use of kinetic theory to see why energy is needed to melt or vaporize a substance. At the melting point, the latent heat of fusion does not act to increase the average kinetic energy (and the temperature) of the molecules in the solid, but instead is used to overcome the potential energy associated with the forces between the molecules. That is, work must be done against these attractive forces to break the molecules loose from their relatively fixed positions in the solid so they can freely roll over one another in the liquid phase. Similarly, energy is required for molecules held close together in the liquid phase to escape into the gaseous phase where they are far apart. This process is a more energetic reorganization of the molecules than is melting (the average distance between the molecules is greatly increased), and hence the heat of vaporization is generally much greater than the heat of fusion for a given substance.

14–6 Heat Transfer: Conduction

Heat is transferred from one place or object to another in three different ways: by *conduction,* by *convection*, and by *radiation*. We now discuss each of these in turn; but in practical situations, any two or all three may be operating at the same time. This Section deals with conduction.

When a metal poker is put in a hot fire, or a silver spoon is placed in a hot bowl of soup, the exposed end of the poker or spoon soon becomes hot as well, even though it is not directly in contact with the source of heat. We say that heat has been *conducted* from the hot end to the cold end.

Heat **conduction** in many materials can be visualized as being carried out via molecular collisions. As one end of an object is heated, the molecules there move faster and faster (= higher temperature). As these faster molecules collide with slower-moving neighbors, they transfer some of their kinetic energy to them, which in turn transfer some energy by collision with molecules still farther along the object. Thus the kinetic energy of thermal motion is transferred by molecular collision along the object. In metals, collisions of free electrons are mainly responsible for conduction. Conduction between objects in physical contact occurs similarly.

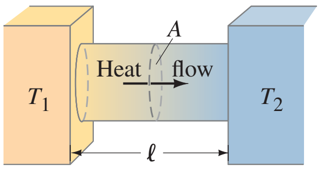


FIGURE 14–6 Heat conduction between regions at temperatures T_1 and T_2. If T_1 is greater than T_2, the heat flows to the right; the rate is given by Eq. 14–5.

Heat conduction from one point to another takes place only if there is a difference in temperature between the two points. Indeed, it is found experimentally that the rate of heat flow through a substance is proportional to the difference in temperature between its ends. The rate of heat flow also depends on the size and shape of the object. To investigate this quantitatively, let us consider the heat flow through a uniform cylinder, as illustrated in Fig. 14–6. It is found experimentally that the heat flow Q over a time interval t is given by the relation

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{\ell} \tag{14–5}$$

where A is the cross-sectional area of the object, ℓ is the distance between the two ends, which are at temperatures T_1 and T_2, and k is a proportionality constant called the **thermal conductivity** which is characteristic of the material. From Eq. 14–5, we see that the rate of heat flow (units of J/s) is directly proportional to the cross-sectional area and to the temperature gradient† $(T_1 - T_2)/\ell$.

†Equation 14–5 is quite similar to the relations describing diffusion (Section 13–13) and the flow of fluids through a pipe (Section 10–12). In those cases, the flow of matter was found to be proportional to the concentration gradient $(C_1 - C_2)/\ell$, or to the pressure gradient $(P_1 - P_2)/\ell$. This close similarity is one reason we speak of the "flow" of heat. Yet we must keep in mind that no substance is flowing in the case of heat—it is energy that is being transferred.

The thermal conductivities, k, for a variety of substances are given in Table 14–4. Substances for which k is large conduct heat rapidly and are said to be good **thermal conductors**. Most metals fall in this category, although there is a wide range even among them, as you may observe by holding the ends of a silver spoon and a stainless-steel spoon immersed in the same hot cup of soup. Substances for which k is small, such as wool, fiberglass, polyurethane, and goose down, are poor conductors of heat and are therefore good **thermal insulators**. The relative magnitudes of k can explain simple phenomena such as why a tile floor is much colder on the feet than a rug-covered floor at the same temperature. Tile is a better conductor of heat than the rug. Heat that flows from your foot to the rug is not conducted away rapidly, so the rug's surface quickly warms up to the temperature of your foot and feels good. But the tile conducts the heat away rapidly and thus can take more heat from your foot quickly, so your foot's surface temperature drops.

TABLE 14–4
Thermal Conductivities

Substance	Thermal Conductivity, k: $\frac{\text{J}}{(\text{s}\cdot\text{m}\cdot\text{C}^\circ)}$	Thermal Conductivity, k: $\frac{\text{kcal}}{(\text{s}\cdot\text{m}\cdot\text{C}^\circ)}$
Silver	420	10×10^{-2}
Copper	380	9.2×10^{-2}
Aluminum	200	5.0×10^{-2}
Steel	40	1.1×10^{-2}
Ice	2	5×10^{-4}
Glass	0.84	2.0×10^{-4}
Brick	0.84	2.0×10^{-4}
Concrete	0.84	2.0×10^{-4}
Water	0.56	1.4×10^{-4}
Human tissue	0.2	0.5×10^{-4}
Wood	0.1	0.3×10^{-4}
Fiberglass	0.048	0.12×10^{-4}
Cork	0.042	0.10×10^{-4}
Wool	0.040	0.10×10^{-4}
Goose down	0.025	0.060×10^{-4}
Polyurethane	0.024	0.057×10^{-4}
Air	0.023	0.055×10^{-4}

EXAMPLE 14–8 **Heat loss through windows.** A major source of heat loss from a house in cold weather is through the windows. Calculate the rate of heat flow through a glass window 2.0 m $\times$ 1.5 m in area and 3.2 mm thick, if the temperatures at the inner and outer surfaces are 15.0°C and 14.0°C, respectively (Fig. 14–7).

APPROACH Heat flows by conduction through the 3.2-mm thickness of glass from the higher inside temperature to the lower outside temperature. We use the heat conduction equation, Eq. 14–5.

SOLUTION Here $A = (2.0\,\text{m})(1.5\,\text{m}) = 3.0\,\text{m}^2$ and $\ell = 3.2 \times 10^{-3}\,\text{m}$. Using Table 14–4 to get k, we have

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{\ell}$$
$$= \frac{(0.84\,\text{J/s}\cdot\text{m}\cdot\text{C}^\circ)(3.0\,\text{m}^2)(15.0^\circ\text{C} - 14.0^\circ\text{C})}{(3.2 \times 10^{-3}\,\text{m})}$$
$$= 790\,\text{J/s}.$$

NOTE This rate of heat flow is equivalent to $(790\,\text{J/s})/(4.19 \times 10^3\,\text{J/kcal}) = 0.19\,\text{kcal/s}$, or $(0.19\,\text{kcal/s}) \times (3600\,\text{s/h}) = 680\,\text{kcal/h}$.

PHYSICS APPLIED
Heat loss through windows

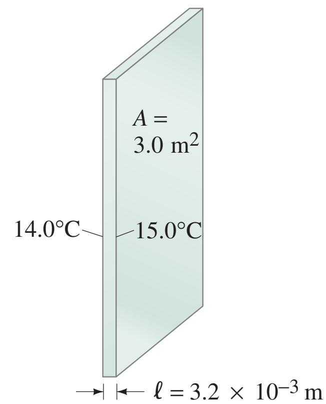


FIGURE 14–7 Example 14–8.

You might notice in Example 14–8 that 15°C is not very warm for the living room of a house. The room itself may indeed be much warmer, and the outside might be colder than 14°C. But the temperatures of 15°C and 14°C were specified as those at the window surfaces, and there is usually a considerable drop in temperature of the air in the vicinity of the window both on the inside and the outside. That is, the layer of air on either side of the window acts as an insulator, and normally the major part of the temperature drop between the inside and outside of the house takes place across the air layer. If there is a heavy wind, the air outside a window will constantly be replaced with cold air; the temperature gradient across the glass will be greater and there will be a much greater rate of heat loss. Increasing the width of the air layer, such as using two panes of glass separated by an air gap, will reduce the heat loss more than simply increasing the glass thickness, since the thermal conductivity of air is much less than that for glass. Such "double-pane windows" are often called *thermal windows*.

PHYSICS APPLIED
Thermal windows

The insulating properties of clothing come from the insulating properties of air. Without clothes, our bodies in still air would heat the air in contact with the skin and would soon become reasonably comfortable because air is a very good insulator. But since air moves—there are breezes and drafts, and people move about—the warm air would be replaced by cold air, thus increasing the temperature difference and the heat loss from the body. Clothes keep us warm by trapping air so it cannot move readily. It is not the cloth that insulates us, but the air that the cloth traps. Goose down is a very good insulator because even a small amount of it fluffs up and traps a great amount of air.

PHYSICS APPLIED
Clothes insulate by trapping an air layer

EXERCISE E Explain why drapes in front of a window reduce heat loss from a house.

PHYSICS APPLIED
R-values of thermal insulation

TABLE 14–5 *R*-values

Material	Thickness	*R*-value ($ft^2 \cdot h \cdot F°/Btu$)
Glass	$\frac{1}{8}$ inch	1
Brick	$3\frac{1}{2}$ inches	0.6–1
Plywood	$\frac{1}{2}$ inch	0.6
Fiberglass insulation	4 inches	12

* *R*-values for Building Materials

The insulating properties of building materials are often specified by *R*-values (or "thermal resistance"), defined for a given thickness ℓ of material as:

$$R = \frac{\ell}{k}.$$

The *R*-value of a given piece of material combines the thickness ℓ and the thermal conductivity k in one number. Larger R means better insulation from heat or cold. In the United States, *R*-values are given in British units as $ft^2 \cdot h \cdot F°/Btu$ (for example, *R*-19 means $R = 19\, ft^2 \cdot h \cdot F°/Btu$). Table 14–5 gives *R*-values for some common building materials. *R*-values increase directly with material thickness: for example, 2 inches of fiberglass is *R*-6, whereas 4 inches is *R*-12.

14–7 Heat Transfer: Convection

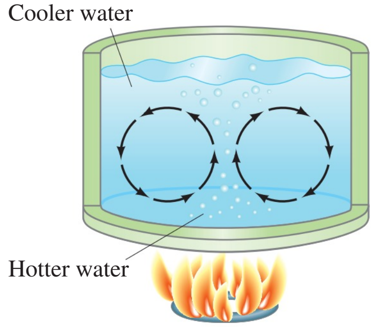


FIGURE 14–8 Convection currents in a pot of water being heated on a stove.

Although liquids and gases are generally not very good conductors of heat, they can transfer heat rapidly by convection. **Convection** is the process whereby heat flows by the bulk movement of molecules from one place to another. Whereas conduction involves molecules (and/or electrons) moving only over small distances and colliding, convection involves the movement of large numbers of molecules over large distances.

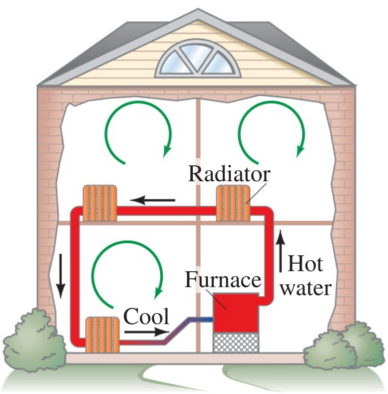
A forced-air furnace, in which air is heated and then blown by a fan into a room, is an example of *forced convection. Natural convection* occurs as well, and one familiar example is that hot air rises. For instance, the air above a radiator (or other type of heater) expands as it is heated (Chapter 13), and hence its density decreases. Because its density is less than that of the surrounding cooler air, it rises via buoyancy, just as a log submerged in water floats upward because its density is less than that of water. Warm or cold ocean currents, such as the balmy Gulf Stream, represent natural convection on a global scale. Wind is another example of convection, and weather in general is strongly influenced by convective air currents.

PHYSICS APPLIED
Ocean currents and wind

When a pot of water is heated (Fig. 14–8), convection currents are set up as the heated water at the bottom of the pot rises because of its reduced density. That heated water is replaced by cooler water from above. This principle is used in many heating systems, such as the hot-water radiator system shown in Fig. 14–9. Water is heated in the furnace, and as its temperature increases, it expands and rises as shown. This causes the water to circulate in the heating system. Hot water then enters the radiators, heat is transferred by conduction to the air, and the cooled water returns to the furnace. Thus, the water circulates because of convection; pumps are sometimes used to improve circulation. The air throughout the room also becomes heated as a result of convection. The air heated by the radiators rises and is replaced by cooler air, resulting in convective air currents, as shown by the green arrows in Fig. 14–9.

PHYSICS APPLIED
Convective home heating

FIGURE 14–9 Convection plays a role in heating a house. The circular arrows show convective air currents in the rooms.



Other types of furnaces also depend on convection. Hot-air furnaces with registers (openings) near the floor often do not have fans but depend on natural convection, which can be appreciable. In other systems, a fan is used. In either case, it is important that cold air can return to the furnace so that convective currents circulate throughout the room if the room is to be uniformly heated.

PHYSICS APPLIED
Body heat: convection by blood

The human body produces a great deal of thermal energy. Of the food energy transformed within the body, at best 20% is used to do work, so over 80% appears as thermal energy. During light activity, for example, if this thermal energy were not dissipated, the body temperature would rise about 3 C° per hour. Clearly, the heat generated by the body must be transferred to the outside. Is the heat transferred by conduction? The temperature of the skin in a comfortable environment is 33 to 35°C, whereas the interior of the body is at 37°C. Calculation shows (Problem 52) that, because of this small temperature difference, plus the low thermal conductivity of tissue, direct conduction is responsible for very little of the heat that must leave the body. Instead, the heat is carried to the skin by the blood. In addition to all its other important responsibilities, blood acts as a convective fluid to transfer heat to just beneath the surface of the skin. It is then conducted (over a very short distance) to the surface. Once at the surface, the heat is transferred to the environment by convection, evaporation, and radiation (Section 14–8).

14–8 Heat Transfer: Radiation

Convection and conduction require the presence of matter as a medium to carry the heat from the hotter to the colder region. But a third type of heat transfer occurs without any medium at all. All life on Earth depends on the transfer of energy from the Sun, and this energy is transferred to the Earth over empty (or nearly empty) space. This form of energy transfer is heat—since the Sun's surface temperature (6000 K) is much higher than Earth's ($\approx$ 300 K)—and is referred to as **radiation** (Fig. 14–10). The warmth we receive from a fire is mainly radiant energy.

FIGURE 14–10 The Sun's surface radiates at 6000 K—much higher than the Earth's surface.

As we shall see in later Chapters, radiation consists essentially of electromagnetic waves. Suffice it to say for now that radiation from the Sun consists of visible light plus many other wavelengths that the eye is not sensitive to, including infrared (IR) radiation, which is mainly responsible for heating the Earth.

The rate at which an object radiates energy has been found to be proportional to the fourth power of the Kelvin temperature, T. That is, an object at 2000 K, as compared to one at 1000 K, radiates energy at a rate $2^4 = 16$ times as much. The rate of radiation is also proportional to the area A of the emitting object, so the rate at which energy leaves the object, Q/t, is

$$\frac{Q}{t} = \epsilon\sigma AT^4. \tag{14–6}$$

This is called the **Stefan-Boltzmann equation**, and σ is a universal constant called the **Stefan-Boltzmann constant** which has the value

$$\sigma = 5.67 \times 10^{-8}\,\mathrm{W/m^2 \cdot K^4}.$$

The factor ϵ, called the **emissivity**, is a number† between 0 and 1 that is characteristic of the surface of the radiating material. Very black surfaces, such as charcoal, have emissivity close to 1, whereas shiny metal surfaces have ϵ close to zero and thus emit correspondingly less radiation. The value of ϵ depends somewhat on the temperature of the material.

Not only do shiny surfaces emit less radiation, but they absorb little of the radiation that falls upon them (most is reflected). Black and very dark objects are good emitters ($\epsilon \approx 1$); they also absorb nearly all the radiation that falls on them—which is why light-colored clothing is usually preferable to dark clothing on a hot day. Thus, **a good absorber is also a good emitter**.

PHYSICS APPLIED
Dark vs. light clothing

Any object not only emits energy by radiation but also absorbs energy radiated by other objects. If an object of emissivity ϵ and area A is at a temperature T_1, it radiates energy at a rate $\epsilon\sigma AT_1^4$. If the object is surrounded by an environment at temperature T_2, the rate at which the surroundings radiate energy is proportional to T_2^4; then the rate that energy is *absorbed* by the object is proportional to T_2^4. The *net* rate of radiant heat flow from the object is given by the equation

$$\frac{Q}{t} = \epsilon\sigma A\left(T_1^4 - T_2^4\right), \tag{14–7}$$

where A is the surface area of the object, T_1 its temperature and ϵ its emissivity (at temperature T_1), and T_2 is the temperature of the surroundings. This equation is consistent with the experimental fact that equilibrium between the object and its surroundings is reached when they come to the same temperature. That is, Q/t must equal zero when $T_1 = T_2$, so ϵ must be the same for emission and absorption. This confirms the idea that a good emitter is a good absorber. Because both the object and its surroundings radiate energy, there is a net transfer of energy from one to the other unless everything is at the same temperature. From Eq. 14–7 it is clear that if $T_1 > T_2$, the net flow of heat is from the object to the surroundings, so the object cools. But if $T_1 < T_2$, the net heat flow is from the surroundings into the object, and the object's temperature rises. If different parts of the surroundings are at different temperatures, Eq. 14–7 becomes more complicated.

†ϵ is the Greek letter epsilon.

PHYSICS APPLIED
The body's radiative heat loss

EXAMPLE 14–9 ESTIMATE **Cooling by radiation.** An athlete is sitting unclothed in a locker room whose dark walls are at a temperature of 15°C. Estimate the body's rate of heat loss by radiation, assuming a skin temperature of 34°C and $\epsilon = 0.70$. Take the surface area of the body not in contact with the chair to be $1.5\ \mathrm{m}^2$.

APPROACH We use Eq. 14–7, which requires Kelvin temperatures.

PROBLEM SOLVING
Must use the Kelvin temperature

SOLUTION We have

$$\begin{aligned}\frac{Q}{t} &= \epsilon\sigma A\left(T_1^4 - T_2^4\right) \\ &= (0.70)\left(5.67 \times 10^{-8}\ \mathrm{W/m^2 \cdot K^4}\right)\left(1.5\ \mathrm{m}^2\right)\left[(307\ \mathrm{K})^4 - (288\ \mathrm{K})^4\right] = 120\ \mathrm{W}.\end{aligned}$$

PROBLEM SOLVING
$T_1^4 - T_2^4 \neq (T_1 - T_2)^4$

NOTE This person's "output" is a bit more than what a 100-W bulb uses.

NOTE Avoid a common error: $\left(T_1^4 - T_2^4\right) \neq \left(T_1 - T_2\right)^4$.

PHYSICS APPLIED
Room comfort

A resting person naturally produces heat internally at a rate of about 100 W, as we will see in Section 15–3, less than the heat loss by radiation as calculated in Example 14–9. Hence, our person's temperature would drop, causing considerable discomfort. The body responds to excessive heat loss by increasing its metabolic rate, and shivering is one method by which the body increases its metabolism. Naturally, clothes help a lot. Example 14–9 illustrates that a person may be uncomfortable even if the temperature of the air is, say, 25°C, which is quite a warm room. If the walls or floor are cold, radiation to them occurs no matter how warm the air is. Indeed, it is estimated that radiation accounts for about 50% of the heat loss from a sedentary person in a normal room. Rooms are most comfortable when the walls and floor are warm and the air is not so warm. Floors and walls can be heated by means of hot-water conduits or electric heating elements. Such first-rate heating systems are becoming more common today, and it is interesting to note that 2000 years ago the Romans, even in houses in the remote province of Great Britain, made use of hot-water and steam conduits in the floor to heat their houses.

PHYSICS APPLIED
Prefer warm walls, cool air

EXAMPLE 14–10 ESTIMATE **Two teapots.** A ceramic teapot ($\epsilon = 0.70$) and a shiny one ($\epsilon = 0.10$) each hold 0.75 L of tea at 95°C. (*a*) Estimate the rate of heat loss from each, and (*b*) estimate the temperature drop after 30 min for each. Consider only radiation, and assume the surroundings are at 20°C.

APPROACH We are given all the information necessary to calculate the heat loss due to radiation, except for the area. The teapot holds 0.75 L, and we can approximate it as a cube 10 cm on a side (volume = 1.0 L), with five sides exposed. To estimate the temperature drop in (*b*), we use the concept of specific heat and ignore the contribution of the pots compared to that of the water.

SOLUTION (*a*) The teapot, approximated by a cube 10 cm on a side with five sides exposed, has a surface area of about $5 \times (0.1\ \mathrm{m})^2 = 5 \times 10^{-2}\ \mathrm{m}^2$. The rate of heat loss would be about

$$\begin{aligned}\frac{Q}{t} &= \epsilon\sigma A\left(T_1^4 - T_2^4\right) \\ &= \epsilon\left(5.67 \times 10^{-8}\ \mathrm{W/m^2 \cdot K^4}\right)\left(5 \times 10^{-2}\ \mathrm{m}^2\right)\left[(368\ \mathrm{K})^4 - (293\ \mathrm{K})^4\right] \approx \epsilon(30)\ \mathrm{W},\end{aligned}$$

or about 20 W for the ceramic pot ($\epsilon = 0.70$) and 3 W for the shiny one ($\epsilon = 0.10$).

(*b*) To estimate the temperature drop, we use the specific heat of water and ignore the contribution of the pots. The mass of 0.75 L of water is 0.75 kg. (Recall that $1.0\ \mathrm{L} = 1000\ \mathrm{cm}^3 = 1 \times 10^{-3}\ \mathrm{m}^3$ and $\rho = 1000\ \mathrm{kg/m^3}$.) Using Eq. 14–2 and Table 14–1, we get

$$\frac{Q}{t} = mc\,\frac{\Delta T}{t}.$$

Then

$$\frac{\Delta T}{t} = \frac{Q/t}{mc} \approx \frac{\epsilon(30)\ \mathrm{J/s}}{(0.75\ \mathrm{kg})(4186\ \mathrm{J/kg \cdot C^\circ})} \approx \epsilon(0.01)\ \mathrm{C^\circ/s}.$$

After 30 min (1800 s), $\Delta T = \epsilon(0.01\ \mathrm{C^\circ/s})t = \epsilon(0.01\ \mathrm{C^\circ/s})(1800\ \mathrm{s}) \approx 18\epsilon\ \mathrm{C^\circ}$, or about 12 C° for the ceramic pot ($\epsilon = 0.70$) and about 2 C° for the shiny one ($\epsilon = 0.10$). The shiny one clearly has an advantage, at least in terms of radiation.

NOTE Convection and conduction could play a greater role than radiation.

Heating of an object by radiation from the Sun cannot be calculated using Eq. 14–7 since this equation assumes a uniform temperature, T_2, of the environment surrounding the object, whereas the Sun is essentially a point source. Hence the Sun must be treated as a separate source of energy. About 1350 J of energy from the Sun strikes Earth's atmosphere per second per square meter of area at right angles to the Sun's rays. This number, $1350\,\mathrm{W/m^2}$, is called the **solar constant**. The atmosphere may absorb as much as 70% of this energy before it reaches the ground, depending on the cloud cover. On a clear day, about $1000\,\mathrm{W/m^2}$ reaches the Earth's surface. An object of emissivity ϵ with area A facing the Sun absorbs energy from the Sun at a rate, in watts, of about

$$\frac{Q}{t} = (1000\,\mathrm{W/m^2})\epsilon A \cos\theta, \qquad \textbf{(14–8)}$$

where θ is the angle between the Sun's rays and a line perpendicular to the area A (Fig. 14–11). That is, $A\cos\theta$ is the "effective" area, at right angles to the Sun's rays.

The explanation for the **seasons** and the polar ice caps (see Fig. 14–12) depends on this $\cos\theta$ factor in Eq. 14–8. The seasons are *not* a result of how close the Earth is to the Sun—in fact, in the Northern Hemisphere, summer occurs when the Earth is farthest from the Sun. It is the angle (i.e., $\cos\theta$) that really matters. Furthermore, the reason the Sun heats the Earth more at midday than at sunrise or sunset is also related to this $\cos\theta$ factor.

PHYSICS APPLIED
Radiation from the Sun

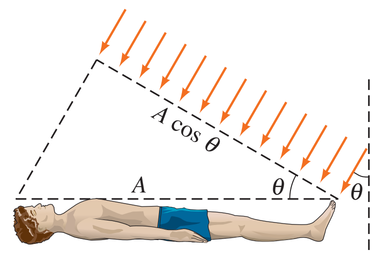
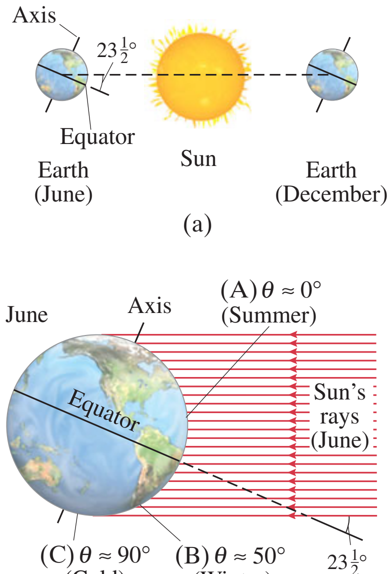


FIGURE 14–11 Radiant energy striking a body at an angle θ.

FIGURE 14–12 (a) Earth's seasons arise from the tilt of Earth's axis relative to its orbit around the Sun. (b) June sunlight makes an angle of about $23\frac{1}{2}°$ with the equator. Thus θ in the southern U.S. (label A) is near 0° (direct summer sunlight), but in the Southern Hemisphere (B), θ is 50° or 60°, and less heat can be absorbed—hence it is winter. Near the poles (C), there is never strong direct sunlight: $\cos\theta$ varies from about $\frac{1}{2}$ in summer to 0 in winter; so with little heating, ice can form.



EXAMPLE 14–11 ESTIMATE Getting a tan—energy absorption. What is the rate of energy absorption from the Sun by a person lying flat on the beach on a clear day if the Sun makes a 30° angle with the vertical? Assume that $\epsilon = 0.70$ and that $1000\,\mathrm{W/m^2}$ reaches the Earth's surface.

APPROACH We use Eq. 14–8 and estimate a typical human to be roughly 2 m tall by 0.4 m wide, so $A \approx (2\,\mathrm{m})(0.4\,\mathrm{m}) = 0.8\,\mathrm{m^2}$. See Fig. 14–11.

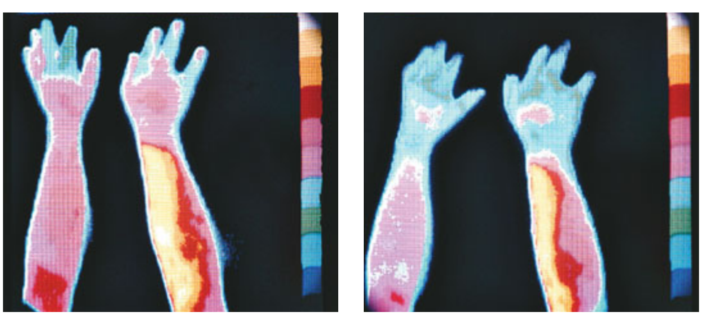
SOLUTION Since $\cos 30° = 0.866$, we have

$$\begin{aligned}\frac{Q}{t} &= (1000\,\mathrm{W/m^2})\epsilon A\cos\theta \\ &= (1000\,\mathrm{W/m^2})(0.70)(0.8\,\mathrm{m^2})(0.866) = 500\,\mathrm{W}.\end{aligned}$$

NOTE If a person wears light-colored clothing, ϵ is much smaller, so the energy absorbed is less.

An interesting application of thermal radiation to diagnostic medicine is **thermography**. A special instrument, the thermograph, scans the body, measuring the intensity of infrared† radiation from many points and forming a picture that resembles an X-ray (Fig. 14–13). Areas where metabolic activity is high, such as in tumors, can often be detected on a thermogram as a result of their higher temperature and consequent increased radiation.

†Infrared radiation is light whose wavelengths are longer than visible light (see Fig. 22–8).



(a) (b)

PHYSICS APPLIED
Thermography

FIGURE 14–13 Thermograms of a healthy person's arms and hands (a) before and (b) after smoking a cigarette, showing a temperature decrease due to impaired blood circulation associated with smoking. The thermograms have been color-coded according to temperature; the scale on the right goes from blue (cold) to white (hot).

PHYSICS APPLIED
Astronomy—size of a star

EXAMPLE 14–12 ESTIMATE Star radius. The giant star Betelgeuse emits radiant energy at a rate 10^4 times greater than our Sun, whereas its surface temperature is only half (2900 K) that of our Sun. Estimate the radius of Betelgeuse, assuming $\epsilon = 1$. The Sun's radius is $r_S = 7 \times 10^8$ m.

APPROACH We assume both stars are spherical, with surface area $4\pi r^2$.

SOLUTION We solve Eq. 14–6 for A:

$$4\pi r^2 = A = \frac{(Q/t)}{\epsilon\sigma T^4}.$$

Then

$$\frac{r_B^2}{r_S^2} = \frac{(Q/t)_B}{(Q/t)_S}\cdot\frac{T_S^4}{T_B^4} = \left(10^4\right)\left(2^4\right) = 16 \times 10^4.$$

Hence $r_B = \sqrt{16 \times 10^4}\, r_S = (400)\left(7 \times 10^8\,\text{m}\right) \approx 3 \times 10^{11}$ m. If Betelgeuse were our Sun, it would envelop us (Earth is 1.5×10^{11} m from the Sun).

EXERCISE F Fanning yourself on a hot day cools you by (*a*) increasing the radiation rate of the skin; (*b*) increasing conductivity; (*c*) decreasing the mean free path of air; (*d*) increasing the evaporation of perspiration; (*e*) none of these.

Summary

Internal energy, U, refers to the total energy of all the molecules in an object. For an ideal monatomic gas,

$$U = \tfrac{3}{2}NkT = \tfrac{3}{2}nRT \qquad \textbf{(14–1)}$$

where N is the number of molecules or n is the number of moles.

Heat refers to the transfer of energy from one object to another because of a difference of temperature. Heat is thus measured in energy units, such as joules.

Heat and internal energy are also sometimes specified in calories or kilocalories (kcal), where

$$1\text{ kcal} = 4.186\text{ kJ}$$

is the amount of heat needed to raise the temperature of 1 kg of water by 1 C°.

The **specific heat**, c, of a substance is defined as the energy (or heat) required to change the temperature of unit mass of the substance by 1 degree; as an equation,

$$Q = mc\,\Delta T, \qquad \textbf{(14–2)}$$

where Q is the heat absorbed or given off, ΔT is the temperature increase or decrease, and m is the mass of the substance.

When heat flows between parts of an isolated system, conservation of energy tells us that the heat gained by one part of the system is equal to the heat lost by the other part of the system. This is the basis of **calorimetry**, which is the quantitative measurement of heat exchange.

Exchange of energy occurs, without a change in temperature, whenever a substance changes phase. The **heat of fusion** is the heat required to melt 1 kg of a solid into the liquid phase; it is also equal to the heat given off when the substance changes from liquid to solid. The **heat of vaporization** is the energy required to change 1 kg of a substance from the liquid to the vapor phase; it is also the energy given off when the substance changes from vapor to liquid.

Heat is transferred from one place (or object) to another in three different ways: conduction, convection, and radiation.

In **conduction**, energy is transferred through a substance by means of collisions between hotter (faster) molecules or electrons with their slower moving neighbors.

Convection is the transfer of energy by the mass movement of molecules even over considerable distances.

Radiation, which does not require the presence of matter, is energy transfer by electromagnetic waves, such as from the Sun. All objects radiate energy in an amount that is proportional to the fourth power of their Kelvin temperature $\left(T^4\right)$ and to their surface area. The energy radiated (or absorbed) also depends on the nature of the surface (dark surfaces absorb and radiate more than do bright shiny ones), which is characterized by the emissivity, ϵ.

Radiation from the Sun arrives at the surface of the Earth on a clear day at a rate of about 1000 W/m^2.

Questions

1. What happens to the work done on a jar of orange juice when it is vigorously shaken?
2. When a hot object warms a cooler object, does temperature flow between them? Are the temperature changes of the two objects equal? Explain.
3. (*a*) If two objects of different temperatures are placed in contact, will heat naturally flow from the object with higher internal energy to the object with lower internal energy? (*b*) Is it possible for heat to flow even if the internal energies of the two objects are the same? Explain.
4. In warm regions where tropical plants grow but the temperature may drop below freezing a few times in the winter, the destruction of sensitive plants due to freezing can be reduced by watering them in the evening. Explain.
5. The specific heat of water is quite large. Explain why this fact makes water particularly good for heating systems (that is, hot-water radiators).
6. Why does water in a metal canteen stay cooler if the cloth jacket surrounding the canteen is kept moist?
7. Explain why burns caused by steam at 100°C on the skin are often more severe than burns caused by water at 100°C.

8. Explain why water cools (its temperature drops) when it evaporates, using the concepts of latent heat and internal energy.
9. Will pasta cook faster if the water boils more vigorously? Explain.
10. Very high in the Earth's atmosphere, the temperature can be 700°C. Yet an animal there would freeze to death rather than roast. Explain.
11. Explorers on failed Arctic expeditions have survived by covering themselves with snow. Why would they do that?
12. Why is wet sand at a beach cooler to walk on than dry sand?
13. If you hear that an object has "high heat content," does that mean that its temperature is high? Explain.
14. When hot-air furnaces are used to heat a house, why is it important that there be a vent for air to return to the furnace? What happens if this vent is blocked by a bookcase?
15. Ceiling fans are sometimes reversible, so that they drive the air down in one season and pull it up in another season. Explain which way you should set the fan (*a*) for summer, (*b*) for winter.
16. Goose down sleeping bags and parkas are often specified as so many inches or centimeters of *loft*, the actual thickness of the garment when it is fluffed up. Explain.
17. Microprocessor chips have a "heat sink" glued on top that looks like a series of fins. Why are they shaped like that?
18. Sea breezes are often encountered on sunny days at the shore of a large body of water. Explain, noting that the temperature of the land rises more rapidly than that of the nearby water.
19. The floor of a house on a foundation under which the air can flow is often cooler than a floor that rests directly on the ground (such as a concrete slab foundation). Explain.
20. A 22°C day is warm, while a swimming pool at 22°C feels cool. Why?
21. Explain why air temperature readings are always taken with the thermometer in the shade.
22. A premature baby in an incubator can be dangerously cooled even when the air temperature in the incubator is warm. Explain.
23. Does an ordinary electric fan cool the air? Why or why not? If not, why use it?
24. Heat loss occurs through windows by the following processes: (1) through the glass panes; (2) through the frame, particularly if it is metal; (3) ventilation around edges; and (4) radiation. (*a*) For the first three, what is (are) the mechanism(s): conduction, convection, or radiation? (*b*) Heavy curtains reduce which of these heat losses? Explain in detail.
25. A piece of wood lying in the Sun absorbs more heat than a piece of shiny metal. Yet the metal feels hotter than the wood when you pick it up. Explain.
26. The Earth cools off at night much more quickly when the weather is clear than when cloudy. Why?
27. An "emergency blanket" is a thin shiny (metal-coated) plastic foil. Explain how it can help to keep an immobile person warm.
28. Explain why cities situated by the ocean tend to have less extreme temperatures than inland cities at the same latitude.
29. A paper cup placed among hot coals will burn if empty (note burn spots at top of cup in Fig. 14–14), but won't burn if filled with water. Explain. Forget the marshmallows.

FIGURE 14–14 Question 29.

30. On a cold windy day, a window will feel colder than on an equally cold day with no wind. This is true even if no air leaks in near the window. Why?

MisConceptual Questions

1. When you put an ice cube in a glass of warm tea, which of the following happens?
   (*a*) Cold flows from the ice cube into the tea.
   (*b*) Cold flows from the ice cube into the tea and heat flows from the tea into the ice cube.
   (*c*) Heat flows from the tea into the ice cube.
   (*d*) Neither heat nor cold flows. Only temperature flows between the ice and the tea.
2. Both beakers A and B in Fig. 14–15 contain a mixture of ice and water at equilibrium. Which beaker is the coldest, or are they equal in temperature?
   (*a*) Beaker A.
   (*b*) Beaker B.
   (*c*) Equal.

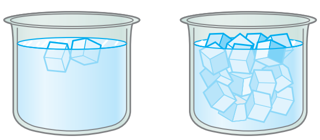


FIGURE 14–15 MisConceptual Question 2.

3. For objects at thermal equilibrium, which of the following is true?
   (*a*) Each is at the same temperature.
   (*b*) Each has the same internal energy.
   (*c*) Each has the same heat.
   (*d*) All of the above.
   (*e*) None of the above.
4. Which of the following happens when a material undergoes a phase change?
   (*a*) The temperature changes.
   (*b*) The chemical composition changes.
   (*c*) Heat flows into or out of the material.
   (*d*) The molecules break apart into atoms.
5. As heat is added to water, is it possible for the temperature measured by a thermometer in the water to remain constant?
   (*a*) Yes, the water could be changing phase.
   (*b*) No, adding heat will always change the temperature.
   (*c*) Maybe; it depends on the rate at which the heat is added.
   (*d*) Maybe; it depends on the initial water temperature.

6. A typical thermos bottle has a thin vacuum space between the shiny inner flask (which holds a liquid) and the shiny protective outer flask, often stainless steel. The vacuum space is excellent at preventing
(*a*) conduction.
(*b*) convection.
(*c*) radiation.
(*d*) conduction and convection.
(*e*) conduction, convection, and radiation.

7. Heat is
(*a*) a fluid called caloric.
(*b*) a measure of the average kinetic energy of atoms.
(*c*) the amount of energy transferred between objects as a result of a difference in temperature.
(*d*) an invisible, odorless, weightless substance.
(*e*) the total kinetic energy of an ideal gas.

8. Radiation is emitted
(*a*) only by glowing objects such as the Sun.
(*b*) only by objects whose temperature is greater than the temperature of the surroundings.
(*c*) only by objects with more caloric than their surroundings.
(*d*) by any object not at 0 K.
(*e*) only by objects that have a large specific heat.

9. Ten grams of water is added to ten grams of ice in an insulated container. Will all of the ice melt?
(*a*) Yes. (*b*) No. (*c*) More information is needed.

10. Two objects are made of the same material, but they have different masses and temperatures. If the objects are brought into thermal contact, which one will have the greater temperature change?
(*a*) The one with the higher initial temperature.
(*b*) The one with the lower initial temperature.
(*c*) The one with the greater mass.
(*d*) The one with the lesser mass.
(*e*) The one with the higher specific heat.
(*f*) Not enough information.

11. It has been a hot summer, so when you arrive at a lake, you decide to go for a swim even though it is nighttime. The water is cold! The next day, you go swimming again during the hottest part of the day, and even though the air is warmer the water is still almost as cold. Why?
(*a*) Water is fairly dense compared with many other liquids.
(*b*) Water remains in a liquid state for a wide range of temperatures.
(*c*) Water has a high bulk modulus.
(*d*) Water has a high specific heat.

12. Two equal-mass liquids, initially at the same temperature, are heated for the same time over the same stove. You measure the temperatures and find that one liquid has a higher temperature than the other. Which liquid has the higher specific heat?
(*a*) The cooler one.
(*b*) The hotter one.
(*c*) Both are the same.

For assigned homework and other learning materials, go to the MasteringPhysics website.

Problems

14–1 Heat as Energy Transfer

1. (I) To what temperature will 8200 J of heat raise 3.0 kg of water that is initially at 10.0°C?

2. (I) How much heat (in joules) is required to raise the temperature of 34.0 kg of water from 15°C to 95°C?

3. (II) When a diver jumps into the ocean, water leaks into the gap region between the diver's skin and her wetsuit, forming a water layer about 0.5 mm thick. Assuming the total surface area of the wetsuit covering the diver is about 1.0 m^2, and that ocean water enters the suit at 10°C and is warmed by the diver to skin temperature of 35°C, estimate how much energy (in units of candy bars = 300 kcal) is required by this heating process.

4. (II) An average active person consumes about 2500 Cal a day. (*a*) What is this in joules? (*b*) What is this in kilowatt-hours? (*c*) If your power company charges about 10 ¢ per kilowatt-hour, how much would your energy cost per day if you bought it from the power company? Could you feed yourself on this much money per day?

5. (II) A British thermal unit (Btu) is a unit of heat in the British system of units. One Btu is defined as the heat needed to raise 1 lb of water by 1 F°. Show that

$$1\text{ Btu} = 0.252\text{ kcal} = 1056\text{ J}.$$

6. (II) How many joules and kilocalories are generated when the brakes are used to bring a 1300-kg car to rest from a speed of 95 km/h?

7. (II) A water heater can generate 32,000 kJ/h. How much water can it heat from 12°C to 42°C per hour?

8. (II) A small immersion heater is rated at 375 W. Estimate how long it will take to heat a cup of soup (assume this is 250 mL of water) from 15°C to 75°C.

14–3 and 14–4 Specific Heat; Calorimetry

9. (I) An automobile cooling system holds 18 L of water. How much heat does it absorb if its temperature rises from 15°C to 95°C?

10. (I) What is the specific heat of a metal substance if 135 kJ of heat is needed to raise 4.1 kg of the metal from 18.0°C to 37.2°C?

11. (II) (*a*) How much energy is required to bring a 1.0-L pot of water at 20°C to 100°C? (*b*) For how long could this amount of energy run a 60-W lightbulb?

12. (II) Samples of copper, aluminum, and water experience the same temperature rise when they absorb the same amount of heat. What is the ratio of their masses? [*Hint*: See Table 14–1.]

13. (II) How long does it take a 750-W coffeepot to bring to a boil 0.75 L of water initially at 11°C? Assume that the part of the pot which is heated with the water is made of 280 g of aluminum, and that no water boils away.

14. (II) What will be the equilibrium temperature when a 265-g block of copper at 245°C is placed in a 145-g aluminum calorimeter cup containing 825 g of water at 12.0°C?

15. (II) A 31.5-g glass thermometer reads 23.6°C before it is placed in 135 mL of water. When the water and thermometer come to equilibrium, the thermometer reads 41.8°C. What was the original temperature of the water? Ignore the mass of fluid inside the glass thermometer.

16. (II) A 0.40-kg iron horseshoe, just forged and very hot (Fig. 14–16), is dropped into 1.25 L of water in a 0.30-kg iron pot initially at 20.0°C. If the final equilibrium temperature is 25.0°C, estimate the initial temperature of the hot horseshoe.

FIGURE 14–16 Problem 16.

17. (II) When a 290-g piece of iron at 180°C is placed in a 95-g aluminum calorimeter cup containing 250 g of glycerin at 10°C, the final temperature is observed to be 38°C. Estimate the specific heat of glycerin.

18. (II) The *heat capacity, C*, of an object is defined as the amount of heat needed to raise its temperature by 1 C°. Thus, to raise the temperature by ΔT requires heat Q given by

$$Q = C\,\Delta T.$$

(a) Write the heat capacity C in terms of the specific heat, c, of the material. (b) What is the heat capacity of 1.0 kg of water? (c) Of 45 kg of water?

19. (II) The 1.20-kg head of a hammer has a speed of 7.5 m/s just before it strikes a nail (Fig. 14–17) and is brought to rest. Estimate the temperature rise of a 14-g iron nail generated by eight such hammer blows done in quick succession. Assume the nail absorbs all the energy.

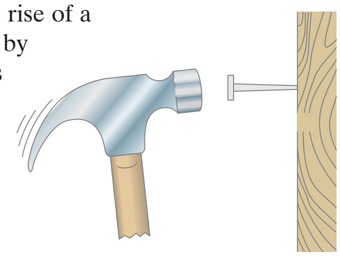


FIGURE 14–17 Problem 19.

20. (II) A 215-g sample of a substance is heated to 330°C and then plunged into a 105-g aluminum calorimeter cup containing 185 g of water and a 17-g glass thermometer at 10.5°C. The final temperature is 35.0°C. What is the specific heat of the substance? (Assume no water boils away.)

21. (II) A 0.095-kg aluminium sphere is dropped from the roof of a 55-m-high building. If 65% of the thermal energy produced when it hits the ground is absorbed by the sphere, what is its temperature increase?

22. (II) Estimate the Calorie content of 65 g of candy from the following measurements. A 15-g sample of the candy is placed in a small aluminum container of mass 0.325 kg filled with oxygen. This container is placed in 1.75 kg of water in an aluminum calorimeter cup of mass 0.624 kg at an initial temperature of 15.0°C. The oxygen–candy mixture in the small container (a "bomb calorimeter") is ignited, and the final temperature of the whole system is 53.5°C.

23. (II) Determine the energy content of 100 g of Karen's fudge cookies from the following measurements. A 10-g sample of a cookie is allowed to dry before putting it in a *bomb calorimeter* (page 396). The aluminum bomb has a mass of 0.615 kg and is placed in 2.00 kg of water contained in an aluminum calorimeter cup of mass 0.524 kg. The initial temperature of the system is 15.0°C, and its temperature after ignition is 36.0°C.

14–5 Latent Heat

24. (I) If 3.40×10^5 J of energy is supplied to a container of liquid oxygen at −183°C, how much oxygen can evaporate?

25. (II) How much heat is needed to melt 23.50 kg of silver that is initially at 25°C?

26. (II) During exercise, a person may give off 185 kcal of heat in 25 min by evaporation of water (at 20°C) from the skin. How much water has been lost? [*Hint*: See page 399.]

27. (II) What mass of steam at 100°C must be added to 1.00 kg of ice at 0°C to yield liquid water at 30°C?

28. (II) A 28-g ice cube at its melting point is dropped into an insulated container of liquid nitrogen. How much nitrogen evaporates if it is at its boiling point of 77 K and has a latent heat of vaporization of 200 kJ/kg? Assume for simplicity that the specific heat of ice is a constant and is equal to its value near its melting point.

29. (II) High-altitude mountain climbers do not eat snow, but always melt it first with a stove. To see why, calculate the energy absorbed from your body if you: (a) eat 1.0 kg of −15°C snow which your body warms to body temperature of 37°C; (b) melt 1.0 kg of −15°C snow using a stove and drink the resulting 1.0 kg of water at 2°C, which your body has to warm to 37°C.

30. (II) An iron boiler of mass 180 kg contains 730 kg of water at 18°C. A heater supplies energy at the rate of 58,000 kJ/h. How long does it take for the water (a) to reach the boiling point, and (b) to all have changed to steam?

31. (II) Determine the latent heat of fusion of mercury using the following calorimeter data: 1.00 kg of solid Hg at its melting point of −39.0°C is placed in a 0.620-kg aluminum calorimeter with 0.400 kg of water at 12.80°C; the resulting equilibrium temperature is 5.06°C.

32. (II) At a crime scene, the forensic investigator notes that the 6.2-g lead bullet that was stopped in a doorframe apparently melted completely on impact. Assuming the bullet was shot at room temperature (20°C), what does the investigator calculate as the minimum muzzle velocity of the gun?

33. (II) A 64-kg ice-skater moving at 7.5 m/s glides to a stop. Assuming the ice is at 0°C and that 50% of the heat generated by friction is absorbed by the ice, how much ice melts?

34. (II) A cube of ice is taken from the freezer at −8.5°C and placed in an 85-g aluminum calorimeter filled with 310 g of water at room temperature of 20.0°C. The final situation is all water at 17.0°C. What was the mass of the ice cube?

35. (II) A 55-g bullet traveling at 250 m/s penetrates a block of ice at 0°C and comes to rest within the ice. Assuming that the temperature of the bullet doesn't change appreciably, how much ice is melted as a result of the collision?

14–6 to 14–8 Conduction, Convection, Radiation

36. (I) Calculate the rate of heat flow by conduction through the windows of Example 14–8, assuming that there are strong gusty winds and the external temperature is −5°C.

37. (I) One end of a 56-cm-long copper rod with a diameter of 2.0 cm is kept at 460°C, and the other is immersed in water at 22°C. Calculate the heat conduction rate along the rod.

38. (II) (a) How much power is radiated by a tungsten sphere (emissivity $\epsilon = 0.35$) of radius 19 cm at a temperature of 25°C? (b) If the sphere is enclosed in a room whose walls are kept at −5°C, what is the *net* flow rate of energy out of the sphere?

39. (II) How long does it take the Sun to melt a block of ice at 0°C with a flat horizontal area $1.0\ \mathrm{m}^2$ and thickness 1.0 cm? Assume that the Sun's rays make an angle of 35° with the vertical and that the emissivity of ice is 0.050.

40. (II) *Heat conduction to skin*. Suppose 150 W of heat flows by conduction from the blood capillaries beneath the skin to the body's surface area of $1.5\ \mathrm{m}^2$. If the temperature difference is 0.50 C°, estimate the average distance of capillaries below the skin surface.

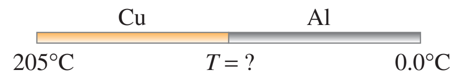
41. (II) Two rooms, each a cube 4.0 m per side, share a 14-cm-thick brick wall. Because of a number of 100-W lightbulbs in one room, the air is at 30°C, while in the other room it is at 10°C. How many of the 100-W bulbs are needed to maintain the temperature difference across the wall?

42. (II) A 100-W lightbulb generates 95 W of heat, which is dissipated through a glass bulb that has a radius of 3.0 cm and is 0.50 mm thick. What is the difference in temperature between the inner and outer surfaces of the glass?

43. (III) Approximately how long should it take 8.2 kg of ice at 0°C to melt when it is placed in a carefully sealed Styrofoam ice chest of dimensions 25 cm × 35 cm × 55 cm whose walls are 1.5 cm thick? Assume that the conductivity of Styrofoam is double that of air and that the outside temperature is 34°C.

44. (III) A copper rod and an aluminum rod of the same length and cross-sectional area are attached end to end (Fig. 14–18). The copper end is placed in a furnace maintained at a constant temperature of 205°C. The aluminum end is placed in an ice bath held at a constant temperature of 0.0°C. Calculate the temperature at the point where the two rods are joined.

FIGURE 14–18 Problem 44.



*45. (III) Suppose the insulating qualities of the wall of a house come mainly from a 4.0-in. layer of brick and an R-19 layer of insulation, as shown in Fig. 14–19. What is the total rate of heat loss through such a wall, if its total area is $195\ \mathrm{ft}^2$ and the temperature difference across it is 35 F°?

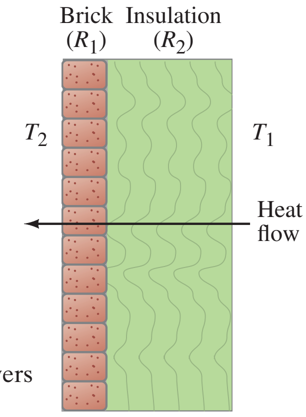


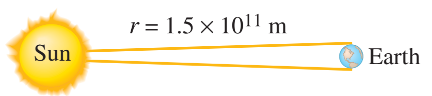
FIGURE 14–19 Problem 45. Two layers insulating a wall.

General Problems

46. A soft-drink can contains about 0.35 kg of liquid at 5°C. Drinking this liquid can actually consume some of the fat in the body, since energy is needed to warm the liquid to body temperature (37°C). How many food Calories should the drink have so that it is in perfect balance with the heat needed to warm the liquid (essentially water)?

47. (*a*) Estimate the total power radiated into space by the Sun, assuming it to be a perfect emitter at $T = 5500\ \mathrm{K}$. The Sun's radius is $7.0 \times 10^8\ \mathrm{m}$. (*b*) From this, determine the power per unit area arriving at the Earth, $1.5 \times 10^{11}\ \mathrm{m}$ away (Fig. 14–20).

FIGURE 14–20 Problem 47.



48. To get an idea of how much thermal energy is contained in the world's oceans, estimate the heat liberated when a cube of ocean water, 1 km on each side, is cooled by 1 K. (Approximate the ocean water as pure water for this estimate.)

49. What will be the final result when equal masses of ice at 0°C and steam at 100°C are mixed together?

50. A mountain climber wears a goose-down jacket 3.5 cm thick with total surface area $0.95\ \mathrm{m}^2$. The temperature at the surface of the clothing is −18°C and at the skin is 34°C. Determine the rate of heat flow by conduction through the jacket assuming (*a*) it is dry and the thermal conductivity k is that of goose down, and (*b*) the jacket is wet, so k is that of water and the jacket has matted to 0.50 cm thickness.

51. During light activity, a 70-kg person may generate 200 kcal/h. Assuming that 20% of this goes into useful work and the other 80% is converted to heat, estimate the temperature rise of the body after 45 min if none of this heat is transferred to the environment.

52. Estimate the rate at which heat can be conducted from the interior of the body to the surface. As a model, assume that the thickness of tissue is 4.0 cm, that the skin is at 34°C and the interior at 37°C, and that the surface area is $1.5\ \mathrm{m}^2$. Compare this to the measured value of about 230 W that must be dissipated by a person working lightly. This clearly shows the necessity of convective cooling by the blood.

53. A bicyclist consumes 9.0 L of water over the span of 3.5 hours during a race. Making the approximation that 80% of the cyclist's energy goes into evaporating this water (at 20°C) as sweat, how much energy in kcal did the rider use during the ride? [*Hint*: See page 399.]

54. If coal gives off 30 MJ/kg when burned, how much coal is needed to heat a house requiring 2.0×10^5 MJ for the whole winter? Assume that 30% of the heat is lost up the chimney.

55. A 15-g lead bullet is tested by firing it into a fixed block of wood with a mass of 35 kg. The block and imbedded bullet together absorb all the heat generated. After thermal equilibrium has been reached, the system has a temperature rise measured as 0.020 C°. Estimate the bullet's entering speed.

56. A 310-kg marble boulder rolls off the top of a cliff and falls a vertical height of 120 m before striking the ground. Estimate the temperature rise of the rock if 50% of the heat generated remains in the rock.

57. A 2.3-kg lead ball is placed in a 2.5-L insulated pail of water initially at 20.0°C. If the final temperature of the water–lead combination is 32.0°C, what was the initial temperature of the lead ball?

58. A microwave oven is used to heat 250 g of water. On its maximum setting, the oven can raise the temperature of the liquid water from 20°C to 100°C in 1 min 45 s (= 105 s). (*a*) At what rate does the oven put energy into the liquid water? (*b*) If the power input from the oven to the water remains constant, determine how many grams of water will boil away if the oven is operated for 2 min (rather than just 1 min 45 s).

59. In a typical squash game (Fig. 14–21), two people hit a soft rubber ball at a wall. Assume that the ball hits the wall at a velocity of 22 m/s and bounces back at a velocity of 12 m/s, and that the kinetic energy lost in the process heats the ball. What will be the temperature increase of the ball after one bounce? (The specific heat of rubber is about 1200 J/kg·C°.)

FIGURE 14–21 Problem 59.

60. The temperature within the Earth's crust increases about 1.0 C° for each 30 m of depth. The thermal conductivity of the crust is 0.80 J/s·C°·m. (*a*) Determine the heat transferred from the interior to the surface for the entire Earth in 1.0 h. (*b*) Compare this heat to the 1000 W/m^2 that reaches the Earth's surface in 1.0 h from the Sun.

61. An iron meteorite melts when it enters the Earth's atmosphere. If its initial temperature was −105°C outside of Earth's atmosphere, calculate the minimum velocity the meteorite must have had before it entered Earth's atmosphere.

62. The temperature of the glass surface of a 75-W lightbulb is 75°C when the room temperature is 18°C. Estimate the temperature of a 150-W lightbulb with a glass bulb the same size. Consider only radiation, and assume that 90% of the energy is emitted as heat.

63. In a cold environment, a person can lose heat by conduction and radiation at a rate of about 200 W. Estimate how long it would take for the body temperature to drop from 36.6°C to 35.6°C if metabolism were nearly to stop. Assume a mass of 65 kg. (See Table 14–1.)

64. A 12-g lead bullet traveling at 220 m/s passes through a thin wall and emerges at a speed of 160 m/s. If the bullet absorbs 50% of the heat generated, (*a*) what will be the temperature rise of the bullet? (*b*) If the bullet's initial temperature was 20°C, will any of the bullet melt, and if so, how much?

65. A leaf of area 40 cm^2 and mass 4.5×10^{-4} kg directly faces the Sun on a clear day. The leaf has an emissivity of 0.85 and a specific heat of 0.80 kcal/kg·K. (*a*) Estimate the energy absorbed per second by the leaf from the Sun, and then (*b*) estimate the rate of rise of the leaf's temperature. (*c*) Will the temperature rise continue for hours? Why or why not? (*d*) Calculate the temperature the leaf would reach if it lost all its heat by radiation to the surroundings at 24°C. (*e*) In what other ways can the heat be dissipated by the leaf?

66. Using the result of part (*a*) in Problem 65, take into account radiation from the leaf to calculate how much water must be transpired (evaporated) by the leaf per hour to maintain a temperature of 35°C.

67. After a hot shower and dishwashing, there seems to be no hot water left in the 65-gal (245-L) water heater. This suggests that the tank has emptied and refilled with water at roughly 10°C. (*a*) How much energy does it take to reheat the water to 45°C? (*b*) How long would it take if the heater output is 9500 W?

68. A house thermostat is normally set to 22°C, but at night it is turned down to 16°C for 9.0 h. Estimate how much more heat would be needed (state as a percentage of daily usage) if the thermostat were not turned down at night. Assume that the outside temperature averages 0°C for the 9.0 h at night and 8°C for the remainder of the day, and that the heat loss from the house is proportional to the temperature difference inside and out. To obtain an estimate from the data, you must make other simplifying assumptions; state what these are.

Search and Learn

1. Create graphs similar to Fig. 14–5, but for lead and ethyl alcohol. Compare and contrast them with each other and with the graph for water. Are there any temperature ranges for which all three substances are liquids? All vapors? All solids? For convenience, use the specific heats given in Table 14–1 for all states of lead and ethyl alcohol.

2. (*a*) Using the solar constant, estimate the rate at which the whole Earth receives energy from the Sun. (*b*) Assume the Earth radiates an equal amount back into space (that is, the Earth is in equilibrium). Then, assuming the Earth is a perfect emitter ($\epsilon = 1.0$), estimate its average surface temperature. [*Hint*: Discuss why you use area $A = \pi r_E^2$ or $A = 4\pi r_E^2$ in each part.]

3. Calculate what will happen when 1000 J of heat is added to 100 grams of (*a*) ice at −20°C, (*b*) ice at 0°C, (*c*) water at 10°C, (*d*) water at 100°C, and (*e*) steam at 110°C.

4. A house has well-insulated walls 19.5 cm thick (assume conductivity of air) and area 410 m^2, a roof of wood 5.5 cm thick and area 250 m^2, and uncovered windows 0.65 cm thick and total area 33 m^2. (*a*) Assuming that heat is lost only by conduction, calculate the rate at which heat must be supplied to this house to maintain its inside temperature at 23°C if the outside temperature is −15°C. (*b*) If the house is initially at 15°C, estimate how much heat must be supplied to raise the temperature to 23°C within 30 min. Assume that only the air needs to be heated and that its volume is 750 m^3. (*c*) If natural gas costs \$0.080/kg and its heat of combustion is 5.4×10^7 J/kg, what is the monthly cost to maintain the house as in part (*a*) for 24 h each day, assuming 90% of the heat produced is used to heat the house? Take the specific heat of air to be 0.24 kcal/kg·C°.

ANSWERS TO EXERCISES

A: (*b*).
B: (*c*).
C: The phase change, liquid to ice (second process).
D: 0.21 kg.
E: The drapes trap a layer of air between the inside of the window and the room, which acts as an excellent insulator.
F: (*d*).