

## Discussion

This is a typical *calorimetry* problem—two bodies at different temperatures are brought in contact with each other and exchange heat until a common temperature is reached. Why is the final temperature so much closer to  $20.0^{\circ}\text{C}$  than  $150^{\circ}\text{C}$ ? The reason is that water has a greater specific heat than most common substances and thus undergoes a small temperature change for a given heat transfer. A large body of water, such as a lake, requires a large amount of heat to increase its temperature appreciably. This explains why the temperature of a lake stays relatively constant during a day even when the temperature change of the air is large. However, the water temperature does change over longer times (e.g., summer to winter).

### Take-Home Experiment: Temperature Change of Land and Water

What heats faster, land or water?

To study differences in heat capacity:

- Place equal masses of dry sand (or soil) and water at the same temperature into two small jars. (The average density of soil or sand is about 1.6 times that of water, so you can achieve approximately equal masses by using 50% more water by volume.)
- Heat both (using an oven or a heat lamp) for the same amount of time.
- Record the final temperature of the two masses.
- Now bring both jars to the same temperature by heating for a longer period of time.
- Remove the jars from the heat source and measure their temperature every 5 minutes for about 30 minutes.

Which sample cools off the fastest? This activity replicates the phenomena responsible for land breezes and sea breezes.

## CHECK YOUR UNDERSTANDING

If 25 kJ is necessary to raise the temperature of a block from  $25^{\circ}\text{C}$  to  $30^{\circ}\text{C}$ , how much heat is necessary to heat the block from  $45^{\circ}\text{C}$  to  $50^{\circ}\text{C}$ ?

### Solution

The heat transfer depends only on the temperature difference. Since the temperature differences are the same in both cases, the same 25 kJ is necessary in the second case.

## 14.3 Phase Change and Latent Heat

### LEARNING OBJECTIVES

By the end of this section, you will be able to:

- Examine heat transfer.
- Calculate final temperature from heat transfer.

So far we have discussed temperature change due to heat transfer. No temperature change occurs from heat transfer if ice melts and becomes liquid water (i.e., during a phase change). For example, consider water dripping from icicles melting on a roof warmed by the Sun. Conversely, water freezes in an ice tray cooled by lower-temperature surroundings.



**FIGURE 14.6** Heat from the air transfers to the ice causing it to melt. (credit: Mike Brand)

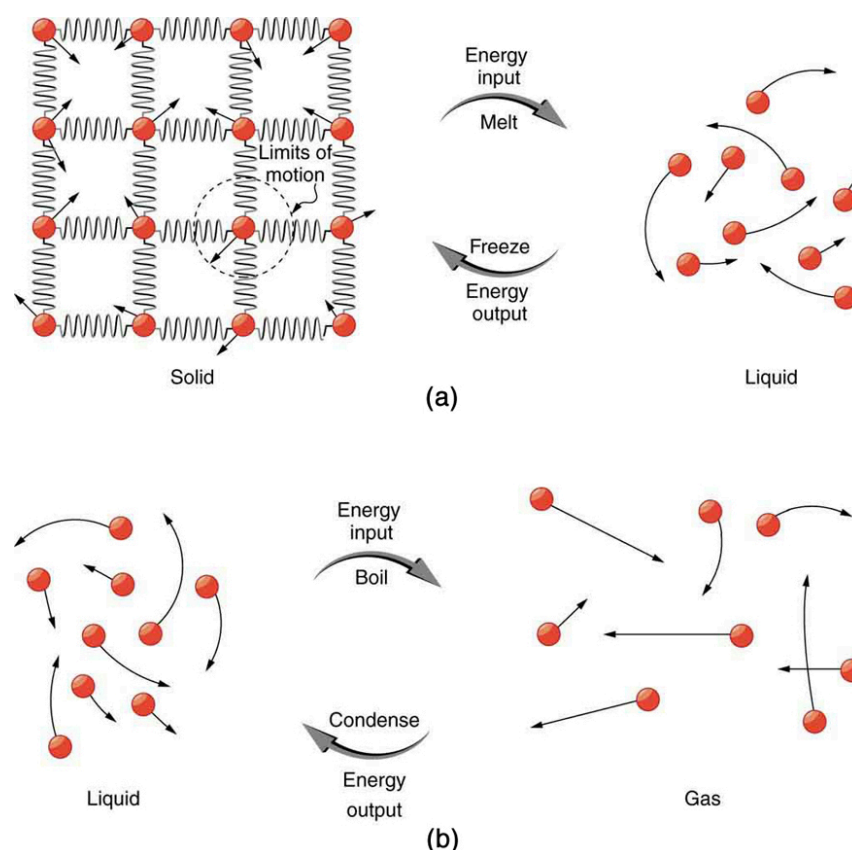
Energy is required to melt a solid because the cohesive bonds between the molecules in the solid must be broken apart such that, in the liquid, the molecules can move around at comparable kinetic energies; thus, there is no rise in temperature. Similarly, energy is needed to vaporize a liquid, because molecules in a liquid interact with each other via attractive forces. There is no temperature change until a phase change is complete. The temperature of a cup of soda initially at  $0^{\circ}\text{C}$  stays at  $0^{\circ}\text{C}$  until all the ice has melted. Conversely, energy is released during freezing and condensation, usually in the form of thermal energy. Work is done by cohesive forces when molecules are brought together. The corresponding energy must be given off (dissipated) to allow them to stay together [Figure 14.7](#).

The energy involved in a phase change depends on two major factors: the number and strength of bonds or force pairs. The number of bonds is proportional to the number of molecules and thus to the mass of the sample. The strength of forces depends on the type of molecules. The heat  $Q$  required to change the phase of a sample of mass  $m$  is given by

$$Q = mL_f \text{ (melting/freezing),} \quad 14.17$$

$$Q = mL_v \text{ (vaporization/condensation),} \quad 14.18$$

where the latent heat of fusion,  $L_f$ , and latent heat of vaporization,  $L_v$ , are material constants that are determined experimentally. See ([Table 14.2](#)).



**FIGURE 14.7** (a) Energy is required to partially overcome the attractive forces between molecules in a solid to form a liquid. That same energy must be removed for freezing to take place. (b) Molecules are separated by large distances when going from liquid to vapor, requiring significant energy to overcome molecular attraction. The same energy must be removed for condensation to take place. There is no temperature change until a phase change is complete.

Latent heat is measured in units of J/kg. Both  $L_f$  and  $L_v$  depend on the substance, particularly on the strength of its molecular forces as noted earlier.  $L_f$  and  $L_v$  are collectively called **latent heat coefficients**. They are *latent*, or hidden, because in phase changes, energy enters or leaves a system without causing a temperature change in the system; so, in effect, the energy is hidden. [Table 14.2](#) lists representative values of  $L_f$  and  $L_v$ , together with melting and boiling points.

The table shows that significant amounts of energy are involved in phase changes. Let us look, for example, at how much energy is needed to melt a kilogram of ice at  $0^\circ\text{C}$  to produce a kilogram of water at  $0^\circ\text{C}$ . Using the equation for a change in temperature and the value for water from [Table 14.2](#), we find that

$Q = mL_f = (1.0\text{ kg})(334\text{ kJ/kg}) = 334\text{ kJ}$  is the energy to melt a kilogram of ice. This is a lot of energy as it represents the same amount of energy needed to raise the temperature of 1 kg of liquid water from  $0^\circ\text{C}$  to  $79.8^\circ\text{C}$ . Even more energy is required to vaporize water; it would take 2256 kJ to change 1 kg of liquid water at the normal boiling point ( $100^\circ\text{C}$  at atmospheric pressure) to steam (water vapor). This example shows that the energy for a phase change is enormous compared to energy associated with temperature changes without a phase change.

Substance	Melting point ( $^\circ\text{C}$ )	$L_f$		Boiling point ( $^\circ\text{C}$ )	$L_v$	
		kJ/kg	kcal/kg		kJ/kg	kcal/kg
Helium	-269.7	5.23	1.25	-268.9	20.9	4.99
Hydrogen	-259.3	58.6	14.0	-252.9	452	108

**TABLE 14.2** Heats of Fusion and Vaporization <sup>6</sup>

<sup>6</sup> Values quoted at the normal melting and boiling temperatures at standard atmospheric pressure (1 atm).

		$L_f$			$L_v$	
Nitrogen	-210.0	25.5	6.09	-195.8	201	48.0
Oxygen	-218.8	13.8	3.30	-183.0	213	50.9
Ethanol	-114	104	24.9	78.3	854	204
Ammonia	-75		108	-33.4	1370	327
Mercury	-38.9	11.8	2.82	357	272	65.0
Water	0.00	334	79.8	100.0	2430 <sup>4</sup>	580 <sup>5</sup>
Sulfur	119	38.1	9.10	444.6	326	77.9
Lead	327	24.5	5.85	1750	871	208
Antimony	631	165	39.4	1440	561	134
Aluminum	660	380	90	2450	11400	2720
Silver	961	88.3	21.1	2193	2336	558
Gold	1063	64.5	15.4	2660	1578	377
Copper	1083	134	32.0	2595	5069	1211
Uranium	1133	84	20	3900	1900	454
Tungsten	3410	184	44	5900	4810	1150

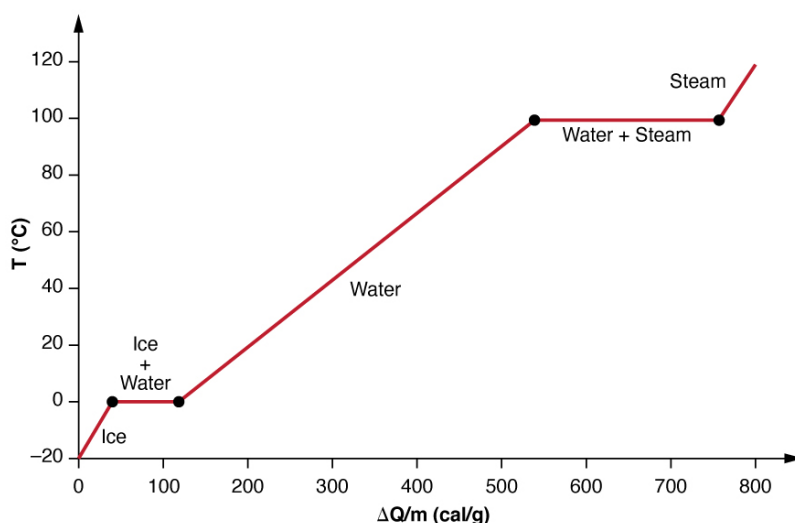
**TABLE 14.2** Heats of Fusion and Vaporization <sup>6</sup>

Phase changes can have a tremendous stabilizing effect even on temperatures that are not near the melting and boiling points, because evaporation and condensation (conversion of a gas into a liquid state) occur even at temperatures below the boiling point. Take, for example, the fact that air temperatures in humid climates rarely go above 35.0°C, which is because most heat transfer goes into evaporating water into the air. Similarly, temperatures in humid weather rarely fall below the dew point because enormous heat is released when water vapor condenses.

We examine the effects of phase change more precisely by considering adding heat into a sample of ice at -20°C (Figure 14.8). The temperature of the ice rises linearly, absorbing heat at a constant rate of 0.50 cal/g · °C until it reaches 0°C. Once at this temperature, the ice begins to melt until all the ice has melted, absorbing 79.8 cal/g of heat. The temperature remains constant at 0°C during this phase change. Once all the ice has melted, the temperature of the liquid water rises, absorbing heat at a new constant rate of 1.00 cal/g · °C. At 100°C, the water begins to boil and the temperature again remains constant while the water absorbs 539 cal/g of heat during this phase change. When all the liquid has become steam vapor, the temperature rises again, absorbing heat at a rate of 0.482 cal/g · °C.

<sup>4</sup> At 37.0°C (body temperature), the heat of vaporization  $L_v$  for water is 2430 kJ/kg or 580 kcal/kg

<sup>5</sup> At 37.0°C (body temperature), the heat of vaporization  $L_v$  for water is 2430 kJ/kg or 580 kcal/kg



**FIGURE 14.8** A graph of temperature versus energy added. The system is constructed so that no vapor evaporates while ice warms to become liquid water, and so that, when vaporization occurs, the vapor remains in of the system. The long stretches of constant temperature values at 0°C and 100°C reflect the large latent heat of melting and vaporization, respectively.

Water can evaporate at temperatures below the boiling point. More energy is required than at the boiling point, because the kinetic energy of water molecules at temperatures below 100°C is less than that at 100°C, hence less energy is available from random thermal motions. Take, for example, the fact that, at body temperature, perspiration from the skin requires a heat input of 2428 kJ/kg, which is about 10 percent higher than the latent heat of vaporization at 100°C. This heat comes from the skin, and thus provides an effective cooling mechanism in hot weather. High humidity inhibits evaporation, so that body temperature might rise, leaving unevaporated sweat on your brow.



### EXAMPLE 14.4

#### Calculate Final Temperature from Phase Change: Cooling Soda with Ice Cubes

Three ice cubes are used to chill a soda at 20°C with mass  $m_{\text{soda}} = 0.25$  kg. The ice is at 0°C and each ice cube has a mass of 6.0 g. Assume that the soda is kept in a foam container so that heat loss can be ignored. Assume the soda has the same heat capacity as water. Find the final temperature when all ice has melted.

#### Strategy

The ice cubes are at the melting temperature of 0°C. Heat is transferred from the soda to the ice for melting. Melting of ice occurs in two steps: first the phase change occurs and solid (ice) transforms into liquid water at the melting temperature, then the temperature of this water rises. Melting yields water at 0°C, so more heat is transferred from the soda to this water until the water plus soda system reaches thermal equilibrium,

$$Q_{\text{ice}} = -Q_{\text{soda}}. \quad 14.19$$

The heat transferred to the ice is  $Q_{\text{ice}} = m_{\text{ice}}L_f + m_{\text{ice}}c_W(T_f - 0^\circ\text{C})$ . The heat given off by the soda is  $Q_{\text{soda}} = m_{\text{soda}}c_W(T_f - 20^\circ\text{C})$ . Since no heat is lost,  $Q_{\text{ice}} = -Q_{\text{soda}}$ , so that

$$m_{\text{ice}}L_f + m_{\text{ice}}c_W(T_f - 0^\circ\text{C}) = -m_{\text{soda}}c_W(T_f - 20^\circ\text{C}). \quad 14.20$$

Bring all terms involving  $T_f$  on the left-hand-side and all other terms on the right-hand-side. Solve for the unknown quantity  $T_f$ :

$$T_f = \frac{m_{\text{soda}}c_W(20^\circ\text{C}) - m_{\text{ice}}L_f}{(m_{\text{soda}} + m_{\text{ice}})c_W}. \quad 14.21$$

#### Solution

1. Identify the known quantities. The mass of ice is  $m_{\text{ice}} = 3 \times 6.0$  g = 0.018 kg and the mass of soda is  $m_{\text{soda}} = 0.25$  kg.

2. Calculate the terms in the numerator:

$$m_{\text{soda}} c_W (20^\circ\text{C}) = (0.25 \text{ kg})(4186 \text{ J/(kg}\cdot^\circ\text{C)})(20^\circ\text{C}) = 20,930 \text{ J} \quad 14.22$$

and

$$m_{\text{ice}} L_f = (0.018 \text{ kg})(334,000 \text{ J/kg}) = 6012 \text{ J}. \quad 14.23$$

3. Calculate the denominator:

$$(m_{\text{soda}} + m_{\text{ice}}) c_W = (0.25 \text{ kg} + 0.018 \text{ kg})(4186 \text{ J/(kg}\cdot^\circ\text{C)}) = 1122 \text{ J/}^\circ\text{C}. \quad 14.24$$

4. Calculate the final temperature:

$$T_f = \frac{20,930 \text{ J} - 6012 \text{ J}}{1122 \text{ J/}^\circ\text{C}} = 13^\circ\text{C}. \quad 14.25$$

### Discussion

This example illustrates the enormous energies involved during a phase change. The mass of ice is about 7 percent the mass of water but leads to a noticeable change in the temperature of soda. Although we assumed that the ice was at the freezing temperature, this is incorrect: the typical temperature is  $-6^\circ\text{C}$ . However, this correction gives a final temperature that is essentially identical to the result we found. Can you explain why?

We have seen that vaporization requires heat transfer to a liquid from the surroundings, so that energy is released by the surroundings. Condensation is the reverse process, increasing the temperature of the surroundings. This increase may seem surprising, since we associate condensation with cold objects—the glass in the figure, for example. However, energy must be removed from the condensing molecules to make a vapor condense. The energy is exactly the same as that required to make the phase change in the other direction, from liquid to vapor, and so it can be calculated from  $Q = mL_V$ .



**FIGURE 14.9** Condensation forms on this glass of iced tea because the temperature of the nearby air is reduced to below the dew point. The rate at which water molecules join together exceeds the rate at which they separate, and so water condenses. Energy is released when the water condenses, speeding the melting of the ice in the glass. (credit: Jenny Downing)

### Real-World Application

Energy is also released when a liquid freezes. This phenomenon is used by fruit growers in Florida to protect oranges when the temperature is close to the freezing point ( $0^\circ\text{C}$ ). Growers spray water on the plants in orchards so that the water freezes and heat is released to the growing oranges on the trees. This prevents the temperature inside the orange from dropping below freezing, which would damage the fruit.





**FIGURE 14.10** The ice on these trees released large amounts of energy when it froze, helping to prevent the temperature of the trees from dropping below  $0^{\circ}\text{C}$ . Water is intentionally sprayed on orchards to help prevent hard frosts. (credit: Hermann Hammer)

**Sublimation** is the transition from solid to vapor phase. You may have noticed that snow can disappear into thin air without a trace of liquid water, or the disappearance of ice cubes in a freezer. The reverse is also true: Frost can form on very cold windows without going through the liquid stage. A popular effect is the making of “smoke” from dry ice, which is solid carbon dioxide. Sublimation occurs because the equilibrium vapor pressure of solids is not zero. Certain air fresheners use the sublimation of a solid to inject a perfume into the room. Moth balls are a slightly toxic example of a phenol (an organic compound) that sublimates, while some solids, such as osmium tetroxide, are so toxic that they must be kept in sealed containers to prevent human exposure to their sublimation-produced vapors.



(a)



(b)

**FIGURE 14.11** Direct transitions between solid and vapor are common, sometimes useful, and even beautiful. (a) Dry ice sublimates directly to carbon dioxide gas. The visible vapor is made of water droplets. (credit: Windell Oskay) (b) Frost forms patterns on a very cold window, an example of a solid formed directly from a vapor. (credit: Liz West)

All phase transitions involve heat. In the case of direct solid-vapor transitions, the energy required is given by the equation  $Q = mL_s$ , where  $L_s$  is the **heat of sublimation**, which is the energy required to change 1.00 kg of a substance from the solid phase to the vapor phase.  $L_s$  is analogous to  $L_f$  and  $L_v$ , and its value depends on the substance. Sublimation requires energy input, so that dry ice is an effective coolant, whereas the reverse process (i.e., frosting) releases energy. The amount of energy required for sublimation is of the same order of magnitude as

that for other phase transitions.

The material presented in this section and the preceding section allows us to calculate any number of effects related to temperature and phase change. In each case, it is necessary to identify which temperature and phase changes are taking place and then to apply the appropriate equation. Keep in mind that heat transfer and work can cause both temperature and phase changes.

### Problem-Solving Strategies for the Effects of Heat Transfer

1. *Examine the situation to determine that there is a change in the temperature or phase. Is there heat transfer into or out of the system?* When the presence or absence of a phase change is not obvious, you may wish to first solve the problem as if there were no phase changes, and examine the temperature change obtained. If it is sufficient to take you past a boiling or melting point, you should then go back and do the problem in steps—temperature change, phase change, subsequent temperature change, and so on.
2. *Identify and list all objects that change temperature and phase.*
3. *Identify exactly what needs to be determined in the problem (identify the unknowns).* A written list is useful.
4. *Make a list of what is given or what can be inferred from the problem as stated (identify the knowns).*
5. *Solve the appropriate equation for the quantity to be determined (the unknown).* If there is a temperature change, the transferred heat depends on the specific heat (see [Table 14.1](#)) whereas, for a phase change, the transferred heat depends on the latent heat. See [Table 14.2](#).
6. *Substitute the knowns along with their units into the appropriate equation and obtain numerical solutions complete with units.* You will need to do this in steps if there is more than one stage to the process (such as a temperature change followed by a phase change).
7. *Check the answer to see if it is reasonable: Does it make sense?* As an example, be certain that the temperature change does not also cause a phase change that you have not taken into account.

### CHECK YOUR UNDERSTANDING

Why does snow remain on mountain slopes even when daytime temperatures are higher than the freezing temperature?

#### **Solution**

Snow is formed from ice crystals and thus is the solid phase of water. Because enormous heat is necessary for phase changes, it takes a certain amount of time for this heat to be accumulated from the air, even if the air is above 0°C. The warmer the air is, the faster this heat exchange occurs and the faster the snow melts.

## 14.4 Heat Transfer Methods

### LEARNING OBJECTIVES

By the end of this section, you will be able to:

- Discuss the different methods of heat transfer.

Equally as interesting as the effects of heat transfer on a system are the methods by which this occurs. Whenever there is a temperature difference, heat transfer occurs. Heat transfer may occur rapidly, such as through a cooking pan, or slowly, such as through the walls of a picnic ice chest. We can control rates of heat transfer by choosing materials (such as thick wool clothing for the winter), controlling air movement (such as the use of weather stripping around doors), or by choice of color (such as a white roof to reflect summer sunlight). So many processes involve heat transfer, so that it is hard to imagine a situation where no heat transfer occurs. Yet every process involving heat transfer takes place by only three methods:

1. **Conduction** is heat transfer through stationary matter by physical contact. (The matter is stationary on a macroscopic scale—we know there is thermal motion of the atoms and molecules at any temperature above absolute zero.) Heat transferred between the electric burner of a stove and the bottom of a pan is transferred by conduction.
2. **Convection** is the heat transfer by the macroscopic movement of a fluid. This type of transfer takes place in a forced-air furnace and in weather systems, for example.
3. Heat transfer by **radiation** occurs when microwaves, infrared radiation, visible light, or another form of electromagnetic radiation is emitted or absorbed. An obvious example is the warming of the Earth by the Sun. A