Practice Problems

5. How much heat does it take to raise the temperature of 10.0 kg of water by 1.0 °C ?
  a. 84 J
  b. 42 J
  c. 84 kJ
  d. 42 kJ

6. Calculate the change in temperature of 1.0 kg of water that is initially at room temperature if 3.0 kJ of heat is added.
  a. 358 °C
  b. 716 °C
  c. 0.36 °C
  d. 0.72 °C

Check Your Understanding

7. What causes heat transfer?
  a. The mass difference between two objects causes heat transfer.
  b. The density difference between two objects causes heat transfer.
  c. The temperature difference between two systems causes heat transfer.
  d. The pressure difference between two objects causes heat transfer.

8. When two bodies of different temperatures are in contact, what is the overall direction of heat transfer?
  a. The overall direction of heat transfer is from the higher-temperature object to the lower-temperature object.
  b. The overall direction of heat transfer is from the lower-temperature object to the higher-temperature object.
  c. The direction of heat transfer is first from the lower-temperature object to the higher-temperature object, then back again to the lower-temperature object, and so-forth, until the objects are in thermal equilibrium.
  d. The direction of heat transfer is first from the higher-temperature object to the lower-temperature object, then back again to the higher-temperature object, and so-forth, until the objects are in thermal equilibrium.

9. What are the different methods of heat transfer?
  a. conduction, radiation, and reflection
  b. conduction, reflection, and convection
  c. convection, radiation, and reflection
  d. conduction, radiation, and convection

10. True or false—Conduction and convection cannot happen simultaneously
  a. True
  b. False

11.3 Phase Change and Latent Heat

Section Learning Objectives

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

- Explain changes in heat during changes of state, and describe latent heats of fusion and vaporization
- Solve problems involving thermal energy changes when heating and cooling substances with phase changes

Section Key Terms

condensation	freezing	latent heat	sublimation
latent heat of fusion	latent heat of vaporization	melting	vaporization
phase change	phase diagram	plasma	

Phase Changes

So far, we have learned that adding thermal energy by heat increases the temperature of a substance. But surprisingly, there are situations where adding energy does not change the temperature of a substance at all! Instead, the additional thermal energy acts to loosen bonds between molecules or atoms and causes a **phase change**. Because this energy enters or leaves a system during a phase change without causing a temperature change in the system, it is known as **latent heat** (latent means *hidden*).

The three phases of matter that you frequently encounter are solid, liquid and gas (see Figure 11.8). Solid has the least energetic state; atoms in solids are in close contact, with forces between them that allow the particles to vibrate but not change position with neighboring particles. (These forces can be thought of as springs that can be stretched or compressed, but not easily broken.)

Liquid has a more energetic state, in which particles can slide smoothly past one another and change neighbors, although they are still held together by their mutual attraction.

Gas has a more energetic state than liquid, in which particles are broken free of their bonds. Particles in gases are separated by distances that are large compared with the size of the particles.

The most energetic state of all is **plasma**. Although you may not have heard much about plasma, it is actually the most common state of matter in the universe—stars are made up of plasma, as is lightning. The plasma state is reached by heating a gas to the point where particles are pulled apart, separating the electrons from the rest of the particle. This produces an ionized gas that is a combination of the negatively charged free electrons and positively charged ions, known as plasma.

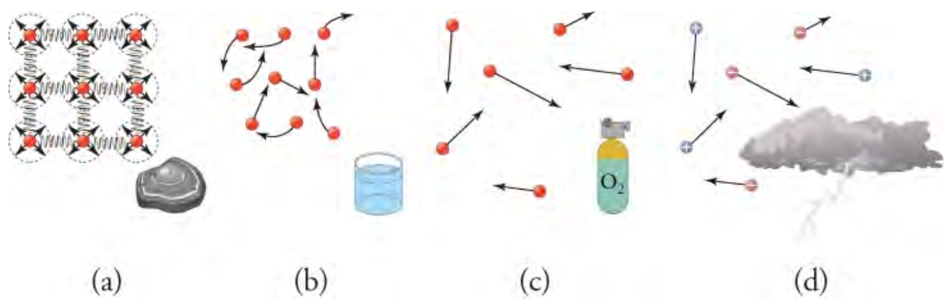

Figure 11.8 (a) Particles in a solid always have the same neighbors, held close by forces represented here by springs. These particles are essentially in contact with one another. A rock is an example of a solid. This rock retains its shape because of the forces holding its atoms or molecules together. (b) Particles in a liquid are also in close contact but can slide over one another. Forces between them strongly resist attempts to push them closer together and also hold them in close contact. Water is an example of a liquid. Water can flow, but it also remains in an open container because of the forces between its molecules. (c) Particles in a gas are separated by distances that are considerably larger than the size of the particles themselves, and they move about freely. A gas must be held in a closed container to prevent it from moving out into its surroundings. (d) The atmosphere is ionized in the extreme heat of a lightning strike.

During a phase change, matter changes from one phase to another, either through the addition of energy by heat and the transition to a more energetic state, or from the removal of energy by heat and the transition to a less energetic state.

Phase changes to a more energetic state include the following:

- **Melting**—Solid to liquid
- **Vaporization**—Liquid to gas (included boiling and evaporation)
- **Sublimation**—Solid to gas

Phase changes to a less energetic state are as follows:

- **Condensation**—Gas to liquid
- **Freezing**—Liquid to solid

Energy is required to melt a solid because the bonds between the particles in the solid must be broken. Since the energy involved in a phase changes is used to break bonds, there is no increase in the kinetic energies of the particles, and therefore no rise in temperature. Similarly, energy is needed to vaporize a liquid to overcome the attractive forces between particles in the liquid. There is no temperature change until a phase change is completed. The temperature of a cup of soda and ice that is initially at 0 °C stays at 0 °C until all of the ice has melted. In the reverse of these processes—freezing and condensation—energy is released from the latent heat (see Figure 11.9).

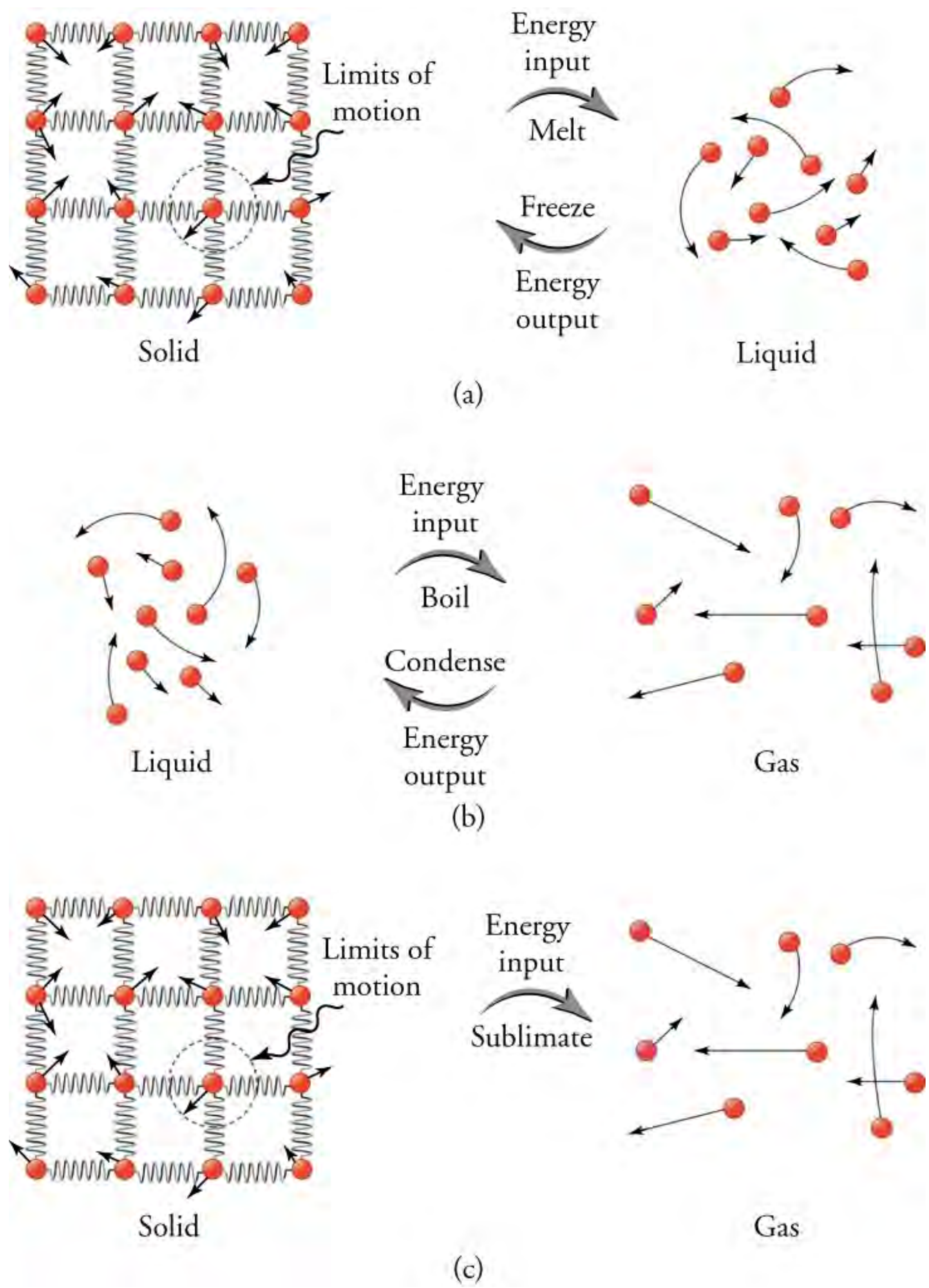

Figure 11.9 (a) Energy is required to partially overcome the attractive forces between particles in a solid to form a liquid. That same energy must be removed for freezing to take place. (b) Particles are separated by large distances when changing from liquid to vapor, requiring significant energy to overcome molecular attraction. The same energy must be removed for condensation to take place. There is no temperature change until a phase change is completed. (c) Enough energy is added that the liquid state is skipped over completely as a substance undergoes sublimation.

The heat, Q, required to change the phase of a sample of mass m is

$Q = mL_f$ (for melting/freezing),

$Q = mL_v$ (for vaporization/condensation),

where L_f is the **latent heat of fusion**, and L_v is the **latent heat of vaporization**. The latent heat of fusion is the amount of heat needed to cause a phase change between solid and liquid. The latent heat of vaporization is the amount of heat needed to cause a phase change between liquid and gas. L_f and L_v are coefficients that vary from substance to substance, depending on the strength of intermolecular forces, and both have standard units of J/kg. See Table 11.3 for values of L_f and L_v of different substances.

Substance	Melting Point ($°C$)	*Lf* (kJ/kg)	Boiling Point ($°C$)	*Lv* (kJ/kg)
Helium	−269.7	5.23	−268.9	20.9
Hydrogen	−259.3	58.6	−252.9	452
Nitrogen	−210.0	25.5	−195.8	201
Oxygen	−218.8	13.8	−183.0	213
Ethanol	−114	104	78.3	854
Ammonia	−78	332	−33.4	1370
Mercury	−38.9	11.8	357	272
Water	0.00	334	100.0	2256
Sulfur	119	38.1	444.6	326
Lead	327	24.5	1750	871
Antimony	631	165	1440	561
Aluminum	660	380	2520	11400
Silver	961	88.3	2193	2336
Gold	1063	64.5	2660	1578
Copper	1083	134	2595	5069
Uranium	1133	84	3900	1900
Tungsten	3410	184	5900	4810

Table 11.3 Latent Heats of Fusion and Vaporization, along with Melting and Boiling Points

Let's consider the example of adding heat to ice to examine its transitions through all three phases—solid to liquid to gas. A phase diagram indicating the temperature changes of water as energy is added is shown in Figure 11.10. The ice starts out at −20 $°C$, and its temperature rises linearly, absorbing heat at a constant rate until it reaches 0 $°$. Once at this temperature, the ice gradually melts, absorbing 334 kJ/kg. 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. At 100 $°C$, the water begins to boil and the temperature again remains constant while the water absorbs 2256 kJ/kg during this phase change. When all the liquid has become steam, the temperature rises again at a constant rate.

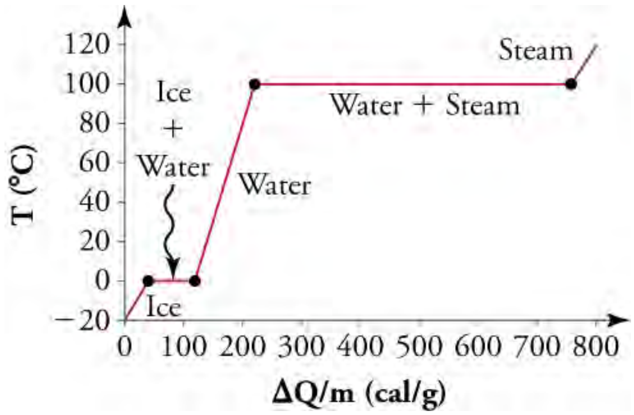

Figure 11.10 A graph of temperature versus added energy. The system is constructed so that no vapor forms while ice warms to become liquid water, and so when vaporization occurs, the vapor remains in the system. The long stretches of constant temperature values at 0 °C and 100 °C reflect the large latent heats of melting and vaporization, respectively.

We have seen that vaporization requires heat transfer to a substance from its surroundings. Condensation is the reverse process, where heat in transferred *away from* a substance *to* its surroundings. This release of latent heat increases the temperature of the surroundings. Energy must be removed from the condensing particles to make a vapor condense. This is why condensation occurs on cold surfaces: the heat transfers energy away from the warm vapor to the cold surface. The energy is exactly the same as that required to cause the phase change in the other direction, from liquid to vapor, and so it can be calculated from $Q = mL_v$. Latent heat is also released into the environment when a liquid freezes, and can be calculated from $Q = mL_f$.

FUN IN PHYSICS

Making Ice Cream

Figure 11.11 With the proper ingredients, some ice and a couple of plastic bags, you could make your own ice cream in five minutes. (ElinorD, Wikimedia Commons)

Ice cream is certainly easy enough to buy at the supermarket, but for the hardcore ice cream enthusiast, that may not be satisfying enough. Going through the process of making your own ice cream lets you invent your own flavors and marvel at the physics firsthand (Figure 11.11).

The first step to making homemade ice cream is to mix heavy cream, whole milk, sugar, and your flavor of choice; it could be as simple as cocoa powder or vanilla extract, or as fancy as pomegranates or pistachios.

The next step is to pour the mixture into a container that is deep enough that you will be able to churn the mixture without it spilling over, and that is also freezer-safe. After placing it in the freezer, the ice cream has to be stirred vigorously every 45 minutes for four to five hours. This slows the freezing process and prevents the ice cream from turning into a solid block of ice. Most people prefer a soft creamy texture instead of one giant popsicle.

As it freezes, the cream undergoes a phase change from liquid to solid. By now, we're experienced enough to know that this means that the cream must experience a loss of heat. Where does that heat go? Due to the temperature difference between the freezer and the ice cream mixture, heat transfers thermal energy from the ice cream to the air in the freezer. Once the temperature in the freezer rises enough, the freezer is cooled by pumping excess heat outside into the kitchen.

A faster way to make ice cream is to chill it by placing the mixture in a plastic bag, surrounded by another plastic bag half full of ice. (You can also add a teaspoon of salt to the outer bag to lower the temperature of the ice/salt mixture.) Shaking the bag for five minutes churns the ice cream while cooling it evenly. In this case, the heat transfers energy out of the ice cream mixture and into the ice during the phase change.

This video (http://www.openstax.org/l/28icecream) gives a demonstration of how to make home-made ice cream using ice and plastic bags.

GRASP CHECK

Why does the ice bag method work so much faster than the freezer method for making ice cream?

a. Ice has a smaller specific heat than the surrounding air in a freezer. Hence, it absorbs more energy from the ice-cream mixture.
b. Ice has a smaller specific heat than the surrounding air in a freezer. Hence, it absorbs less energy from the ice-cream mixture.
c. Ice has a greater specific heat than the surrounding air in a freezer. Hence, it absorbs more energy from the ice-cream mixture.
d. Ice has a greater specific heat than the surrounding air in a freezer. Hence, it absorbs less energy from the ice-cream mixture.

Solving Thermal Energy Problems with Phase Changes

WORKED EXAMPLE

Calculating Heat Required for a Phase Change

Calculate a) how much energy is needed to melt 1.000 kg of ice at 0 $^\circ\mathrm{C}$ (freezing point), and b) how much energy is required to vaporize 1.000 kg of water at 100 $^\circ\mathrm{C}$ (boiling point).

STRATEGY FOR (A)

Using the equation for the heat required for melting, and the value of the latent heat of fusion of water from the previous table, we can solve for part (a).

Solution to (a)

The energy to melt 1.000 kg of ice is

$$Q = mL_f = (1.000\text{ kg})(334\text{ kJ/kg}) = 334\text{ kJ}. \tag{11.18}$$

STRATEGY FOR (B)

To solve part (b), we use the equation for heat required for vaporization, along with the latent heat of vaporization of water from the previous table.

Solution to (b)

The energy to vaporize 1.000 kg of liquid water is

$$Q = mL_v = (1.000\text{ kg})(2256\text{ kJ/kg}) = 2256\text{ kJ}. \tag{11.19}$$

Discussion

The amount of energy need to melt a kilogram of ice (334 kJ) is the same amount of energy needed to raise the temperature of 1.000 kg of liquid water from 0 °C to 79.8 °C . This example shows that the energy for a phase change is enormous compared to energy associated with temperature changes. It also demonstrates that the amount of energy needed for vaporization is even greater.

WORKED EXAMPLE

Calculating Final Temperature from Phase Change: Cooling Soda with Ice Cubes

Ice cubes are used to chill a soda at 20 °C and with a mass of $m_{soda} = 0.25 \text{ kg}$. The ice is at 0 °C and the total mass of the ice cubes is 0.018 kg. Assume that the soda is kept in a foam container so that heat loss can be ignored, and that the soda has the same specific heat as water. Find the final temperature when all of the 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 they are the same temperature. Since the amount of heat leaving the soda is the same as the amount of heat transferred to the ice.

$$Q_{ice} = -Q_{soda} \tag{11.20}$$

The heat transferred to the ice goes partly toward the phase change (melting), and partly toward raising the temperature after melting. Recall from the last section that the relationship between heat and temperature change is $Q = mc\Delta T$. For the ice, the temperature change is $T_f - 0 \text{ °C}$. The total heat transferred to the ice is therefore

$$Q_{ice} = m_{ice} L_f + m_{ice} c_w (T_f - 0 \text{ °C}). \tag{11.21}$$

Since the soda doesn't change phase, but only temperature, the heat given off by the soda is

$$Q_{soda} = m_{soda} c_w (T_f - 20 \text{ °C}). \tag{11.22}$$

Since $Q_{ice} = -Q_{soda}$,

$$m_{ice} L_f + m_{ice} c_w (T_f - 0 \text{ °C}) = -m_{soda} c_w (T_f - 20 \text{ °C}). \tag{11.23}$$

Bringing all terms involving T_f to the left-hand-side of the equation, and all other terms to the right-hand-side, we can solve for T_f .

$$T_f = \frac{m_{soda} c_w (20 \text{ °C}) - m_{ice} L_f}{(m_{soda} + m_{ice}) c_w} \tag{11.24}$$

Substituting the known quantities

$$T_f = \frac{(0.25 \text{ kg})(4186 \text{ J/kg} \cdot \text{°C})(20 \text{ °C}) - (0.018 \text{ kg})(334{,}000 \text{ J/kg})}{(0.25 \text{ kg} + 0.018 \text{ kg})(4186 \text{ K/kg} \cdot \text{°C})} = 13 \text{ °C} \tag{11.25}$$

Discussion

This example shows the enormous energies involved during a phase change. The mass of the ice is about 7 percent the mass of the soda, yet it causes a noticeable change in the soda's temperature.

TIPS FOR SUCCESS

If the ice were not already at the freezing point, we would also have to factor in how much energy would go into raising its temperature up to 0 °C, before the phase change occurs. This would be a realistic scenario, because the temperature of ice is often below 0 °C .

Practice Problems

11. How much energy is needed to melt 2.00 kg of ice at 0 °C ?
    a. 334 kJ
    b. 336 kJ
    c. 167 kJ
    d. 668 kJ

12. If 2500 kJ of energy is just enough to melt 3.0 kg of a substance, what is the substance's latent heat of fusion?
    a. $7500 \text{ kJ} \cdot \text{kg}$
    b. 7500 kJ/kg
    c. $830 \text{ kJ} \cdot \text{kg}$
    d. 830 kJ/kg

Check Your Understanding

13. What is latent heat?
    a. It is the heat that must transfer energy to or from a system in order to cause a mass change with a slight change in the temperature of the system.
    b. It is the heat that must transfer energy to or from a system in order to cause a mass change without a temperature change in the system.
    c. It is the heat that must transfer energy to or from a system in order to cause a phase change with a slight change in the temperature of the system.
    d. It is the heat that must transfer energy to or from a system in order to cause a phase change without a temperature change in the system.

14. In which phases of matter are molecules capable of changing their positions?
    a. gas, liquid, solid
    b. liquid, plasma, solid
    c. liquid, gas, plasma
    d. plasma, gas, solid