12.2 Changes of State and the Laws of Thermodynamics

Eighteenth-century steam-engine builders used heat to turn liquid water into steam. The steam pushed a piston to turn the engine, and then the steam was cooled and condensed into a liquid again. Adding heat to the liquid water changed not only its temperature, but also its structure. You will learn that changing state means changing form as well as changing the way in which atoms store thermal energy.

▶ **Objectives**

- **Define** heats of fusion and vaporization.
- **State** the first and second laws of thermodynamics.
- **Distinguish** between heat and work.
- **Define** entropy.

▶ **Vocabulary**

heat of fusion
heat of vaporization
first law of thermodynamics
heat engine
entropy
second law of thermodynamics

Changes of State

The three most common states of matter are solids, liquids, and gases. As the temperature of a solid is raised, it usually changes to a liquid. At even higher temperatures, it becomes a gas. How can these changes be explained? Consider a material in a solid state. When the thermal energy of the solid is increased, the motion of the particles also increases, as does the temperature.

Figure 12-10 diagrams the changes of state as thermal energy is added to 1.0 g of water starting at 243 K (ice) and continuing until it reaches 473 K (steam). Between points A and B, the ice is warmed to 273 K. At some point, the added thermal energy causes the particles to move rapidly enough that their motion overcomes the forces holding the particles together in a fixed location. The particles are still touching each other, but they have more freedom of movement. Eventually, the particles become free enough to slide past each other.

Melting point At this point, the substance has changed from a solid to a liquid. The temperature at which this change occurs is the melting point of the substance. When a substance is melting, all of the added thermal energy goes to overcome the forces holding the particles together in the solid state. None of the added thermal energy increases the kinetic energy of the particles. This can be observed between points B and C in Figure 12-10, where the added thermal energy melts the ice at a constant 273 K. Because the kinetic energy of the particles does not increase, the temperature does not increase between points B and C.

Boiling point Once a solid is completely melted, there are no more forces holding the particles in the solid state. Adding more thermal energy again increases the motion of the particles, and the temperature of the liquid rises. In the diagram, this process occurs between points C and D. As the temperature increases further, some particles in the liquid acquire enough energy to break free from the other particles. At a specific temperature, known as the boiling point, further addition of energy causes the substance to undergo another change of state. All the added thermal energy converts the substance from the liquid state to the gaseous state.

Figure 12-10 A plot of temperature versus heat added when 1.0 g of ice is converted to steam. Note that the scale is broken between points D and E.

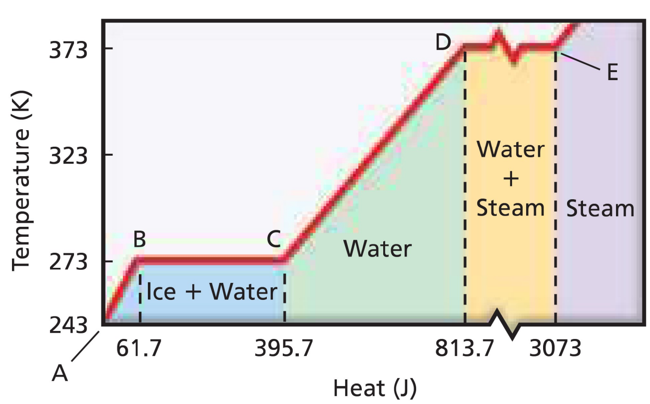


Table 12-2

Heats of Fusion and Vaporization of Common Substances

Material	Heat of Fusion H_f (J/kg)	Heat of Vaporization H_v (J/kg)
Copper	2.05×10^5	5.07×10^6
Mercury	1.15×10^4	2.72×10^5
Gold	6.30×10^4	1.64×10^6
Methanol	1.09×10^5	8.78×10^5
Iron	2.66×10^5	6.29×10^6
Silver	1.04×10^5	2.36×10^6
Lead	2.04×10^4	8.64×10^5
Water (ice)	3.34×10^5	2.26×10^6

As in melting, the temperature does not rise while a liquid boils. In Figure 12-10, this transition is represented between points D and E. After the material is entirely converted to gas, any added thermal energy again increases the motion of the particles, and the temperature rises. Above point E, steam is heated to temperatures greater than 373 K.

Heat of fusion The amount of energy needed to melt 1 kg of a substance is called the **heat of fusion** of that substance. For example, the heat of fusion of ice is 3.34×10^5 J/kg. If 1 kg of ice at its melting point, 273 K, absorbs 3.34×10^5 J, the ice becomes 1 kg of water at the same temperature, 273 K. The added energy causes a change in state but not in temperature. The horizontal distance in Figure 12-10 from point B to point C represents the heat of fusion.

Heat of vaporization At normal atmospheric pressure, water boils at 373 K. The thermal energy needed to vaporize 1 kg of a liquid is called the **heat of vaporization.** For water, the heat of vaporization is 2.26×10^6 J/kg. The distance from point D to point E in Figure 12-10 represents the heat of vaporization. Every material has a characteristic heat of vaporization.

Between points A and B, there is a definite slope to the line as the temperature is raised. This slope represents the specific heat of the ice. The slope between points C and D represents the specific heat of water, and the slope above point E represents the specific heat of steam. Note that the slope for water is less than those of both ice and steam. This is because water has a greater specific heat than does ice or steam. The heat, Q, required to melt a solid of mass m is given by the following equation.

Heat Required to Melt a Solid $Q = mH_f$

The heat required to melt a solid is equal to the mass of the solid times the heat of fusion of the solid.

Similarly, the heat, Q, required to vaporize a mass, m, of liquid is given by the following equation.

Heat Required to Vaporize a Liquid $Q = mH_v$

The heat required to vaporize a liquid is equal to the mass of the liquid times the heat of vaporization of the liquid.

When a liquid freezes, an amount of heat, $Q = -mH_f$, must be removed from the liquid to turn it into a solid. The negative sign indicates that the heat is transferred from the sample to the external world. In the same way, when a vapor condenses to a liquid, an amount of heat, $Q = -mH_v$, must be removed from the vapor. The values of some heats of fusion, H_f, and heats of vaporization, H_v, are shown in **Table 12-2.**

MINI LAB

Melting

1. Label two foam cups *A* and *B*.

2. Measure and pour 75 mL of room-temperature water into each cup. Wipe up any spilled liquid.

3. Add an ice cube to cup A, and add ice water to cup B until the water levels are equal.

4. Measure the temperature of the water in each cup at 1-min intervals until the ice has melted.

5. Record the temperatures in a data table and plot a graph.

Analyze and Conclude

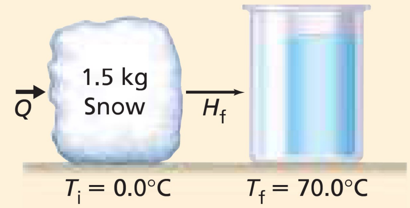
6. Do the samples reach the same final temperature? Why?

EXAMPLE Problem 3

Heat Suppose that you are camping in the mountains. You need to melt 1.50 kg of snow at 0.0°C and heat it to 70.0°C to make hot cocoa. How much heat will be needed?

1 Analyze and Sketch the Problem

- Sketch the relationship between heat and water in its solid and liquid states.
- Sketch the transfer of heat as the temperature of the water increases.



Known:		Unknown:
$m = 1.50$ kg	$H_f = 3.34\times10^5$ J/kg	$Q_{melt\ ice} = ?$
$T_i = 0.0°C$	$T_f = 70.0°C$	$Q_{heat\ liquid} = ?$
$C = 4180$ J/kg·°C		$Q_{total} = ?$

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2 Solve for the Unknown

Calculate the heat needed to melt ice.

$$Q_{melt\ ice} = mH_f$$
$$= (1.50\ kg)(3.34\times10^5\ J/kg) \quad \text{Substitute } m = 1.50\ kg,\ H_f = 3.34\times10^5\ J/kg$$
$$= 5.01\times10^5\ J$$
$$= 5.01\times10^2\ kJ$$

Calculate the temperature change.

$$\Delta T = T_f - T_i$$
$$= 70.0°C - 0.0°C \quad \text{Substitute } T_f = 70.0°C,\ T_i = 0.0°C$$
$$= 70.0°C$$

Calculate the heat needed to raise the water temperature.

$$Q_{heat\ liquid} = mC\Delta T$$
$$= (1.50\ kg)(4180\ J/kg\cdot°C)(70.0°C) \quad \text{Substitute } m = 1.50\ kg,\ C = 4180\ J/kg\cdot°C,\ \Delta T = 70.0°C$$
$$= 4.39\times10^5\ J$$
$$= 4.39\times10^2\ kJ$$

Calculate the total amount of heat needed.

$$Q_{total} = Q_{melt\ ice} + Q_{heat\ liquid}$$
$$= 5.01\times10^2\ kJ + 4.39\times10^2\ kJ \quad \text{Substitute } Q_{melt\ ice} = 5.01\times10^2\ kJ,\ Q_{heat\ liquid} = 4.39\times10^2\ kJ$$
$$= 9.40\times10^2\ kJ$$

3 Evaluate the Answer

- **Are the units correct?** Energy units are in joules.
- **Does the sign make sense?** Q is positive when heat is absorbed.
- **Is the magnitude realistic?** The amount of heat needed to melt the ice is greater than the amount of heat needed to increase the water temperature by 70.0°C. It takes more energy to overcome the forces holding the particles in the solid state than to raise the temperature of water.

PRACTICE Problems

Additional Problems, Appendix B

19. How much heat is absorbed by 1.00×10^2 g of ice at −20.0°C to become water at 0.0°C?

20. A 2.00×10^2-g sample of water at 60.0°C is heated to steam at 140.0°C. How much heat is absorbed?

21. How much heat is needed to change 3.00×10^2 g of ice at −30.0°C to steam at 130.0°C?