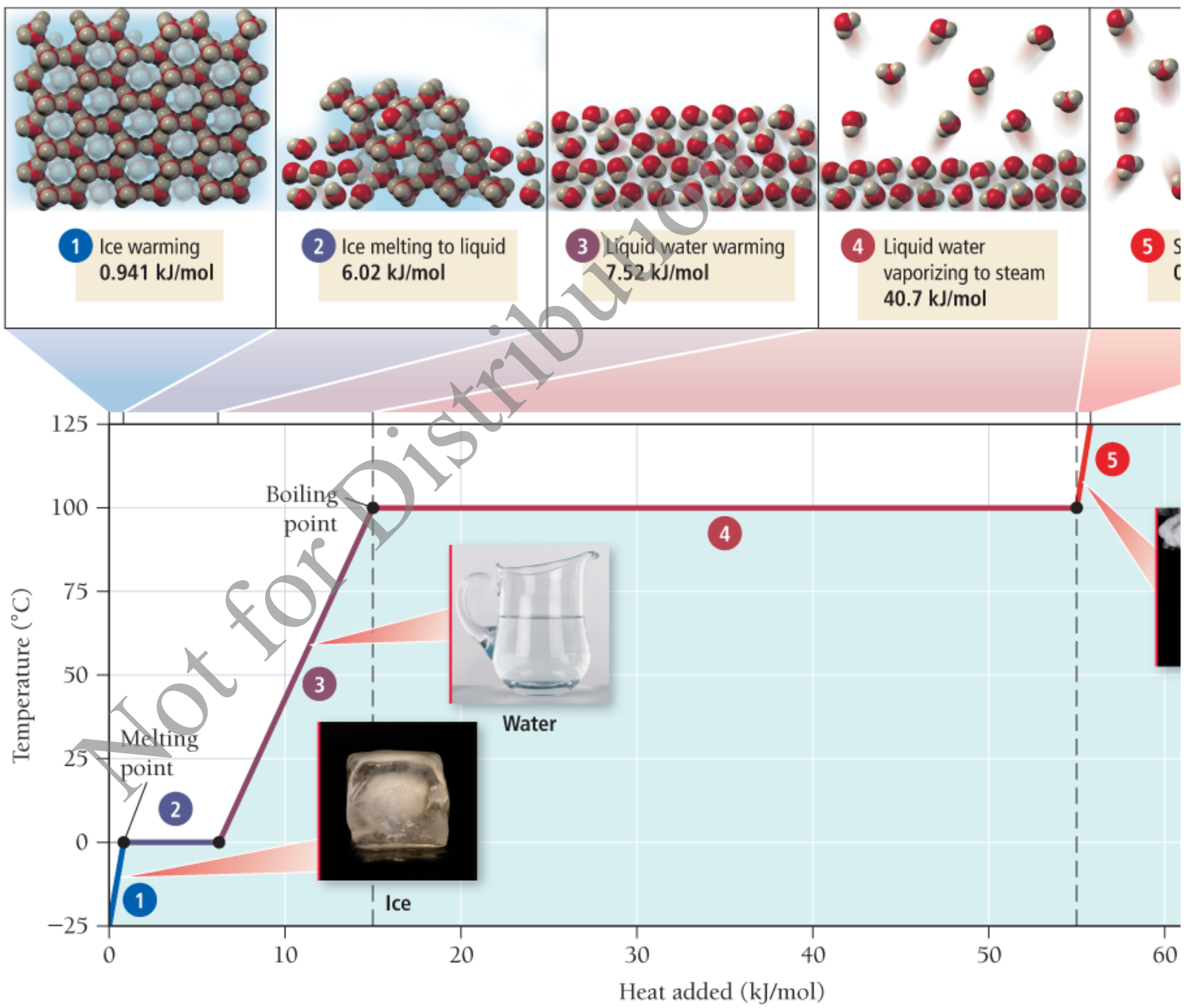
Pearson

11.7: Heating Curve for Water

Key Concept Video Heating Curve for Water

We can combine and build on the concepts from Sections 11.5 and 11.6 by examining the *heating curve* for 1.00 mol of water at 1.00 atm pressure shown in Figure 11.33 on the next page. The *y*-axis of the heating curve represents the temperature of the water sample. The *x*-axis represents the amount of heat added (in kilojoules) during heating. In the diagram, we divide the process into five segments: (1) ice warming; (2) ice melting into liquid water; (3) liquid water warming; (4) liquid water vaporizing into steam; and (5) steam warming.

Figure 11.33 Heating Curve for Water



In two of the segments in Figure 11.33 (Segments 2 and 4), the temperature is constant as heat is added because the added heat goes into producing the transition between states, not into increasing the temperature. The two states are in equilibrium during the transition, and the temperature remains constant. The amount of heat required to achieve the state change is given by $q = n\Delta H$.

In the other three segments (Segments 1, 3, and 5), temperature increases linearly. These segments represent the heating of single states in which the deposited heat raises the temperature in accordance with the substance's heat capacity ($q = mC_s\Delta T$). Let's examine each segment individually.

Segment 1

In Segment 1, solid ice is warmed from −25 °C to 0 °C. Since no transition between states occurs here, the amount of heat required to heat the solid ice is given by $q = mC_s\Delta T$ (see Section 9.4), where C_s is the specific heat capacity of ice ($C_{s,\text{ ice}} = 2.09\ \text{J/g}\cdot{}^\circ\text{C}$). For 1.00 mol of water (18.0 g), we calculate the amount of heat as follows:

$$\begin{aligned} q &= mC_{s,\text{ice}}\Delta T \\ &= 18.0\ \cancel{\text{g}}\left(2.09\frac{\text{J}}{\cancel{\text{g}}\cdot{}^\circ\cancel{\text{C}}}\right)\left[0.0^\circ\ \cancel{\text{C}} - \left(-25.0^\circ\ \cancel{\text{C}}\right)\right] \\ &= 941\ \text{J} = 0.941\ \text{kJ} \end{aligned}$$

So in Segment 1, 0.941 kJ of heat is added to the ice, warming it from −25 °C to 0 °C.

Segment 2

In Segment 2, the added heat does not change the temperature of the ice and water mixture because the heat is absorbed by the transition from solid to liquid. The amount of heat required to convert the ice to liquid water is given by $q = n\Delta H_{\text{fus}}$, where n is the number of moles of water and ΔH_{fus} is the heat of fusion (see Section 11.6):

$$\begin{aligned} q &= n\Delta H_{\text{fus}} \\ &= 1.00\ \cancel{\text{mol}}\left(\frac{6.02\ \text{kJ}}{\cancel{\text{mol}}}\right) \\ &= 6.02\ \text{kJ} \end{aligned}$$

In Segment 2, 6.02 kJ is added to the ice, melting it into liquid water. Notice that the temperature does not change during melting. The liquid and solid coexist at 0 °C as the melting occurs.

Segment 3

In Segment 3, the liquid water is warmed from 0 °C to 100 °C. Because no transition between states occurs here, the amount of heat required to heat the liquid water is given by $q = mC_s\Delta T$ as in Segment 1. However, now we use the heat capacity of liquid water (not ice) for the calculation. For 1.00 mol of water (18.0 g), we calculate the amount of heat as follows:

$$\begin{aligned} q &= mC_{s,\text{liq}}\Delta T \\ &= 18.0\ \cancel{\text{g}}\left(4.18\frac{\text{J}}{\cancel{\text{g}}\cdot{}^\circ\cancel{\text{C}}}\right)\left(100.0^\circ\ \cancel{\text{C}} - 0.0^\circ\ \cancel{\text{C}}\right) \\ &= 7.52\times 10^3\text{J} = 7.52\ \text{kJ} \end{aligned}$$

In Segment 3, 7.52 kJ of heat is added to the liquid water, warming it from 0 °C to 100 °C.

Segment 4

In Segment 4, the water undergoes a second transition between states, this time from liquid to gas. The amount of heat required to convert the liquid to gas is given by $q = n\Delta H_{\text{vap}}$, where n is the number of moles and ΔH_{vap} is the heat of vaporization (see Section 11.5):

$$\begin{aligned} q &= n\Delta H_{\text{vap}} \\ &= 1.00\ \cancel{\text{mol}}\left(\frac{40.7\ \text{kJ}}{\cancel{\text{mol}}}\right) \\ &= 40.7\ \text{kJ} \end{aligned}$$

Thus, in Segment 4, 40.7 kJ is added to the water, vaporizing it into steam. Notice that the temperature does not change during boiling. The liquid and gas coexist at 100 °C as the boiling occurs.

Segment 5

In Segment 5, the steam is warmed from 100 °C to 125 °C. Because no transition between states occurs here, the amount of heat required to heat the steam is given by $q = mC_s\Delta T$ (as in Segments 1 and 3) except that we use the heat capacity of steam (not water or ice) (2.01 J/g · °C):

$$\begin{aligned} q &= mC_{\mathrm{s,steam}}\Delta T \\ &= 18.0\ \cancel{\mathrm{g}}\left(2.01\frac{\mathrm{J}}{\cancel{\mathrm{g}}\cdot{}^\circ\ \cancel{\mathrm{C}}}\right)\left(125.0^\circ\ \cancel{\mathrm{C}} - 100.0^\circ\ \cancel{\mathrm{C}}\right) \\ &= 904 = 0.904\ \mathrm{kJ} \end{aligned}$$

In Segment 5, 0.904 kJ of heat is added to the steam, warming it from 100 °C to 125 °C.

Conceptual Connection 11.6 **Cooling of Water with Ice**

Interactive