Energy and Change in State

Recall the following:

Energy in the form of heat can be added or removed for a substance when it alters state.

EXAMPLE:

\[ \text{H}_2\text{O} \text{ (s)} \rightarrow \text{H}_2\text{O} \text{ (l)} \rightarrow \text{H}_2\text{O} \text{ (g)} \]

melting    vapor

In both phase transitions, energy is being absorbed making each process endothermic!!!
Sample Problem

Determine how much energy is needed to convert 20.0g of water from liquid to gas? Also, how much energy is needed to convert 20.0g of gaseous water to liquid water?

(From liquid to gas)
Given: $\Delta H_{\text{vap}} = 2.26 \text{ kJ/g}$ and $m = 20.0\text{g}$

Solution:
$$q = \Delta H_{\text{vap}} \times m$$
$$q = 2.26 \text{ kJ/g} \times 20.0\text{g} = 45.2 \text{ kJ}$$

Sample Problem

Determine how much energy is needed to convert 20.0g of water from liquid to gas? Also, how much energy is needed to convert 20.0g of gaseous water to liquid water?

(From gas to liquid)
Given: $\Delta H_{\text{vap}} = -2.26 \text{ kJ/g}$ and $m = 20.0\text{g}$

Solution:
$$q = \Delta H_{\text{vap}} \times m$$
$$q = -2.26 \text{ kJ/g} \times 20.0\text{g} = -45.2 \text{ kJ}$$
Heat of Fusion

It is quantity of energy needed to convert 1 gram of substance from solid to liquid or from liquid to solid.

It is represented by ($\Delta H_{\text{fus}}$) where the following is true:

$$q \propto m \rightarrow q = \Delta H_{\text{fus}} \times m$$

$$\Delta H = \frac{q}{m} \text{ where } m = \text{mass, } q = \text{heat}$$

Sample Problem

Determine how much energy is needed to convert 20.0g of water from solid to liquid? Also, how much energy is needed to convert 20.0g liquid water to solid ice?

(From liquid to gas)
Given: $\Delta H_{\text{fus}} = 333 \text{ kJ/g}$ and $m = 20.0g$

Solution:

$$q = \Delta H_{\text{fus}} \times m$$
$$q = 333 \text{ J/g} \times 20.0g = 6.66 \times 10^3 \text{ J}$$

Sample Problem

Determine how much energy is needed to convert 20.0g of water from solid to liquid? Also, how much energy is needed to convert 20.0g liquid water to solid ice?

(From gas to liquid)
Given: $\Delta H_{\text{fus}} = -333 \text{ J/g}$ and $m = 20.0g$

Solution:

$$q = \Delta H_{\text{fus}} \times m$$
$$q = -333 \text{ J/g} \times 20.0g = -6.66 \times 10^3 \text{ J}$$
Specific Heat

When one calculates heat, it is directly proportional to its amount of mass of the substance and \( \Delta T \).

Equation:

\[
q \propto m \text{ and } q \propto \Delta T
\]

Introducing a proportionality constant, \( c \):

\[
q = c m \times \Delta T
\]

The units of \( c \) may be found by solving the equation for that quantity:

\[
c = \frac{q}{m \times \Delta T}
\]

**Specific heat** is defined as the following:

The heat flow required to change the temperature of one gram of a substance by one degree Celsius.

A substance with a low specific heat gains little energy in warming through a temperature change, as compared with a substance with a higher specific heat.

### Table 15.5 Selected Specific Heats

<table>
<thead>
<tr>
<th>Elements</th>
<th>Specific heat (J/g °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al(Aluminum)</td>
<td>0.90</td>
</tr>
<tr>
<td>Cd(Cadmium)</td>
<td>0.33</td>
</tr>
<tr>
<td>Cu(Copper)</td>
<td>0.38</td>
</tr>
<tr>
<td>Ag(Argentum)</td>
<td>0.11</td>
</tr>
<tr>
<td>Zn(Zinc)</td>
<td>0.39</td>
</tr>
<tr>
<td>Sn(Tin)</td>
<td>0.31</td>
</tr>
</tbody>
</table>

### Table 15.5 Selected Specific Heats

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat (J/g °C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetone(l)</td>
<td>2.17</td>
</tr>
<tr>
<td>Benzene(l)</td>
<td>1.74</td>
</tr>
<tr>
<td>Water (solid)</td>
<td>2.06</td>
</tr>
<tr>
<td>Water (liquid)</td>
<td>4.18</td>
</tr>
<tr>
<td>Water (gas)</td>
<td>2.00</td>
</tr>
<tr>
<td>Common Substances</td>
<td></td>
</tr>
<tr>
<td>Concrete</td>
<td>0.88</td>
</tr>
<tr>
<td>Glass</td>
<td>0.84</td>
</tr>
<tr>
<td>Granite</td>
<td>0.79</td>
</tr>
<tr>
<td>Wood</td>
<td>1.76</td>
</tr>
</tbody>
</table>
Sample Question

How much energy is required to change the temperature of 25.0 g of water from 50.0°C to 90.0°C?

Given: Specific Heat (water) is 4.18 J/g°C

\[ T_{\text{initial}} = 50.0^\circ\text{C} \] & \[ T_{\text{final}} = 90.0^\circ\text{C} \]

\[ \text{mass} = 25.0 \text{g} \]

\[ q = m \times c \times \Delta t \]

\[ q = 25.0 \text{g} \times 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \times (90.0 - 50.0)^\circ\text{C} = 4.18 \times 10^3 \text{ J} \]

Calculation of Total Heat Flow for a Change in Temperature + Change in State

If you steadily apply heat to a pure substance in the solid state, five things will happen:

1. The solid will warm to its melting point temperature.
2. The solid will change to liquid at the melting point temperature.
3. The liquid will warm to its boiling point temperature.
4. The liquid will change to gas at the boiling point temperature.
5. The gas will become hotter.
Practice Question

Determine how much energy is needed to convert 20.0g of solid ice at -20.0°C to 110.0°C.

Given Information:
m = 20.0g
Initial T = -20.0°C, BP = 100.0°C, MP = 0.0°C
Final T = 110.0°C
C (ice) = 2.06 J/g °C, C (liquid water) = 4.18 J/g °C,
C (vapor water) = 2.00 J/g °C
ΔH_fus = 333 J/g & ΔH_vap = 2.26 kJ/g

Solution

Part 1: water in the solid state

\[ q_1 = m \times C \times \Delta T \]

\[ q_1 = 20.0 \times 2.06 \frac{J}{g \cdot ^\circ C} 	imes (0.0 - (-20.0))^\circ C \]

\[ q_1 = 824 \text{ J} \]

Solution

Part 2: convert ice to liquid water

\[ q_2 = m \times ΔH_{fus} \]

\[ q_2 = 20.0 \times 333 \frac{J}{g} \]

\[ q_2 = 6660 \text{ J} \]
Solution

Part 3: water in the liquid state

\[ q_1 = m \times C \times \Delta T \]
\[ q_1 = 20.0g \times 4.18 \times \frac{J}{g \cdot ^\circ C} \times (100.0 - 0.0)^\circ C \]
\[ q_1 = 8360 \text{ J} \]

Solution

Part 4: liquid water to vapor

\[ q_4 = m \times \Delta H_{\text{vap}} \]
\[ q_4 = 20.0g \times 2.26 \times \frac{kJ}{g} \times \frac{1000J}{1kJ} \]
\[ q_4 = 45200 \text{ J} \]

Solution

Part 5: water in the gas state

\[ q_5 = m \times C \times \Delta T \]
\[ q_5 = 20.0g \times 2.00 \times \frac{J}{g \cdot ^\circ C} \times (110.0 - 100.0)^\circ C \]
\[ q_5 = 400 \text{ J} \]
Solution

Part 6: Total energy \((q_1 \text{ to } q_5)\)

\[ q_{\text{total}} = q_1 + q_2 + q_3 + q_4 + q_5 \]

\[ q_{\text{total}} = 824J + 6660J + 8360J + 45200J + 400J \]

\[ q_{\text{total}} = 61444J \text{ or } 61.4 \text{ kJ} \]