What quantity of heat is required to melt 500 g of ice at 0 °C and heat the water to steam at 100 °C?

1. To melt the ice:
   \[ q = (500 \text{ g})(333 \text{ J/g}) = 1.67 \times 10^5 \text{ J} \]
2. To raise the water \( T \) from 0 °C to 100 °C:
   \[ q = (500 \text{ g})(4.2 \text{ J/g•K})(100 - 0)\text{K} = 2.1 \times 10^5 \text{ J} \]
3. To evaporate the water at 100 °C:
   \[ q = (500 \text{ g})(2260 \text{ J/g}) = 1.13 \times 10^6 \text{ J} \]
4. Total heat energy = \( 1.51 \times 10^6 \text{ J} = 1510 \text{ kJ} \)
Energy transfer in a reaction also allows us to predict the degree of reactivity. In general, reactions that transfer energy to their surroundings are product-favored. We now begin to consider heat transfer in chemical processes.

Heat flows into the SYSTEM (solid CO$_2$) from the SURROUNDINGS in an ENDOOTHERMIC process. Gas molecules have higher kinetic energy.
Energy flowed into the system. This is an endothermic reaction.

Energy movement:
- Gas molecules have higher kinetic energy than molecules in the solid.
- Also, WORK is done by the system (expanding CO₂) in pushing aside the atmosphere.

First law of thermodynamics: energy is conserved. So the heat into the system equals the energy change for the system plus any energy out of the system:

\[ q = \Delta E + w_{by} \]

where \( q \) is heat energy transferred to the system, \( \Delta E \) is energy change of the system, and \( w_{by} \) is work done by the system.

\[ q = \Delta E - w_{on} \]

where \( q \) is heat energy transferred to the system, \( \Delta E \) is energy change of the system, and \( w_{on} \) is work done on the system.

**ENTHALPY**

Most chemical reactions occur at constant pressure \( P \). Heat transferred at constant \( P \) is called \( q_p \):

\[ q_p = \Delta H = \Delta E + w_{by} = \Delta E - w_{on} \]

where \( H \) is enthalpy.

- \( \Delta H \) is heat transferred at constant \( P \).
- \( \Delta H \) is change in heat content of the system.
- \( \Delta H = H_{\text{final}} - H_{\text{initial}} \)

enthalpy is defined for every state of the system, here initial and final.
\[ \Delta H = H_{\text{final}} - H_{\text{initial}} \]

If \( H_{\text{final}} > H_{\text{initial}} \) then \( \Delta H \) is positive and the process is **ENDOTHERMIC**

If \( H_{\text{final}} < H_{\text{initial}} \) then \( \Delta H \) is negative and the process is **EXOTHERMIC**

**Using Enthalpy**

Consider the decomposition of gaseous water

\[ H_2O(g) + 242 \text{ kJ} \rightarrow H_2(g) + \frac{1}{2} O_2(g) \]

For an endothermic reaction — heat is a “reactant”

\[ \Delta H = +242 \text{ kJ} \]
**USING ENTHALPY**

This is an example of Hess’s Law—

If a reaction is the sum of 2 or more other rxns, the net $\Delta H$ is the sum of the $\Delta H$’s of the other rxns.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$\Delta H$ (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$H_2O(l) \rightarrow H_2O(g)$</td>
<td>44</td>
</tr>
<tr>
<td>$H_2O(g) \rightarrow H_2(g) + 1/2 O_2(g)$</td>
<td>242</td>
</tr>
<tr>
<td>$H_2O(l) \rightarrow H_2(g) + 1/2 O_2(g)$</td>
<td>286</td>
</tr>
</tbody>
</table>

Calculate $\Delta H$ for $S(s) + 3/2 O_2(g) \rightarrow SO_3(g)$ knowing that

$S(s) + O_2(g) \rightarrow SO_2(g)$ \hspace{1cm} $\Delta H_1 = -320.5$ kJ

$SO_2(g) + 1/2 O_2(g) \rightarrow SO_3(g)$ \hspace{1cm} $\Delta H_2 = -75.2$ kJ

The two equations add up to give the desired net equation, so

$\Delta H_{net} = \Delta H_1 + \Delta H_2 = -395.7$ kJ

This is an exothermic reaction—the system releases heat.

**USING ENTHALPY**

$\Sigma \Delta H$ along one path = $\Sigma \Delta H$ along another path

- This equation is valid because $\Delta H$ is a state function.
- State functions depend only on the state of the system and not how it got there.
- $V$, $T$, $P$, $E$ and $H$ are state functions.
- Unlike $V$, $T$ and $P$, one cannot readily measure absolute $H$. It’s much easier to get $\Delta H$. 