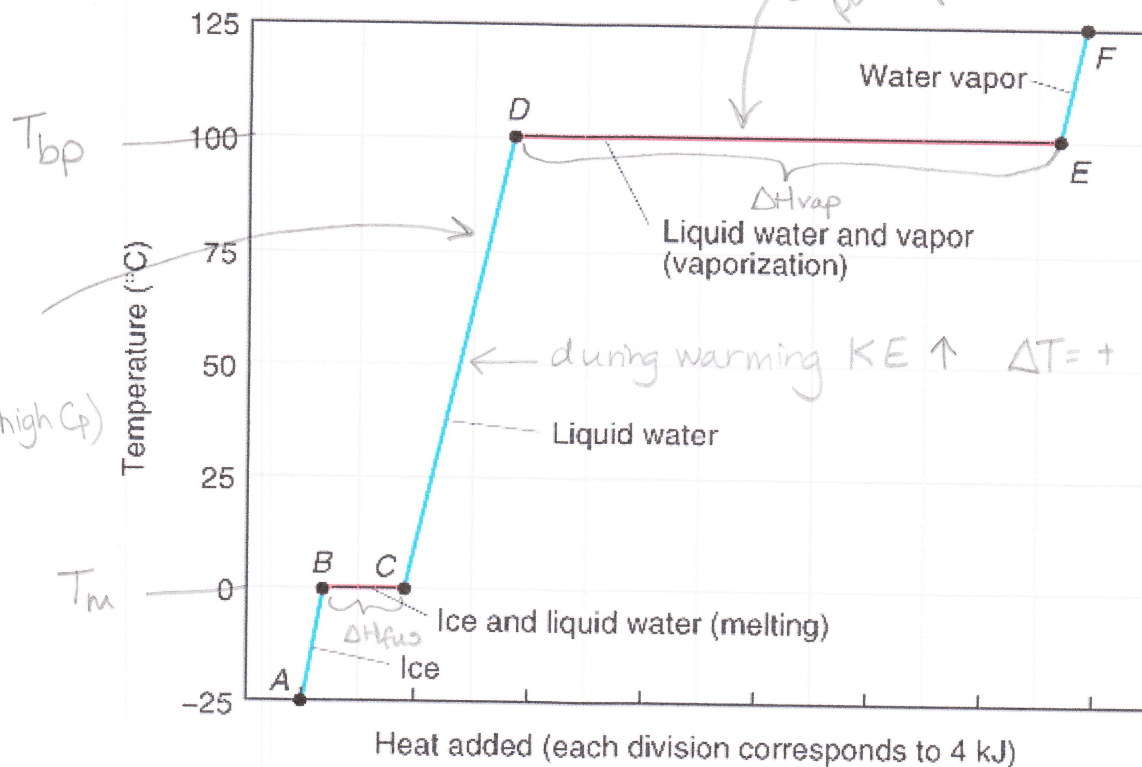


## Thermal Energy and Phase Changes



**EXAMPLE:** How much total energy does it take to convert 40.0 g of ice at  $-30^\circ\text{C}$  to steam at  $125^\circ\text{C}$ ?

Data you need to know:

$$\Delta H_{fus} = 333 \text{ J/g}$$

$$\Delta H_{vap} = 2256 \text{ J/g}$$

$$C_{ice} = 2.06 \text{ J/g}^\circ\text{C}$$

$$C_{water} = 4.18 \text{ J/g}^\circ\text{C}$$

$$C_{steam} = 1.92 \text{ J/g}^\circ\text{C}$$

Heat ice to melting point

$$q = m \cdot c \cdot \Delta T = 40\text{g} \cdot \frac{2.06\text{J}}{\text{g}^\circ\text{C}} \cdot (0^\circ\text{C} - (-30^\circ\text{C})) = 2472 \text{ J}$$

Melt ice

$$q = \Delta H_{fus} \cdot m = \frac{333\text{J}}{\text{g}} \cdot 40.0\text{g} = 13320 \text{ J}$$

Heat water to boiling point

$$q = m \cdot c \cdot \Delta T = 40\text{g} \cdot \frac{4.18\text{J}}{\text{g}^\circ\text{C}} \cdot (100^\circ\text{C} - 0^\circ\text{C}) = 16720 \text{ J}$$

Vaporize water

$$q = \Delta H_{vap} \cdot m = 2256 \text{ J/g} \cdot 40\text{g} = 90240 \text{ J}$$

Heat steam to  $125^\circ\text{C}$

$$q = m \cdot c \cdot \Delta T = 40.0\text{g} \cdot \frac{1.92\text{J}}{\text{g}^\circ\text{C}} \cdot (125^\circ\text{C} - 100^\circ\text{C}) = 1920 \text{ J}$$

$$\begin{array}{r} 2472.0 \\ 13320.0 \\ 16720.0 \\ 90240.0 \\ \hline 1920.0 \\ \hline 124672 \text{ J} \end{array}$$

**125 kJ**

**Clicker Question:**

50g of ethanol is at 20°C. I add 6385 J of heat. What is the final temperature and state of the ethanol? ( $T_{bp}=78.4^\circ\text{C}$ ,  $C_{\text{eth}}=1.43\text{J/g}\cdot^\circ\text{C}$ ,  $\Delta H_{\text{vap}}=837.0\text{J/g}$ )

Warm  $20^\circ\text{C} \rightarrow 78.4^\circ\text{C}$

$$q = m \cdot c \cdot \Delta T = 50\text{g} \cdot \frac{1.43\text{J}}{\text{g}^\circ\text{C}} \cdot (78.4^\circ\text{C} - 20^\circ\text{C}) = 4,175.60\text{J}$$

58.4

Change State @ 78.4°C

$$q = \Delta H_{\text{vap}} \cdot m = \frac{837.0\text{J}}{\text{g}} \cdot 50\text{g} = 41,850\text{J}$$

Started with

$$\begin{array}{r} 6385\text{J} \\ - 4175\text{J} \text{ to warm} \\ \hline 2210\text{J} \text{ left} \end{array}$$

not enough to vaporize  
so I'll have liquid @  $78.4^\circ\text{C}$

zero in constant P situation

**Calorimetry**

$$H = E + PV$$

**1. Constant Pressure**

Gives  $\Delta H$

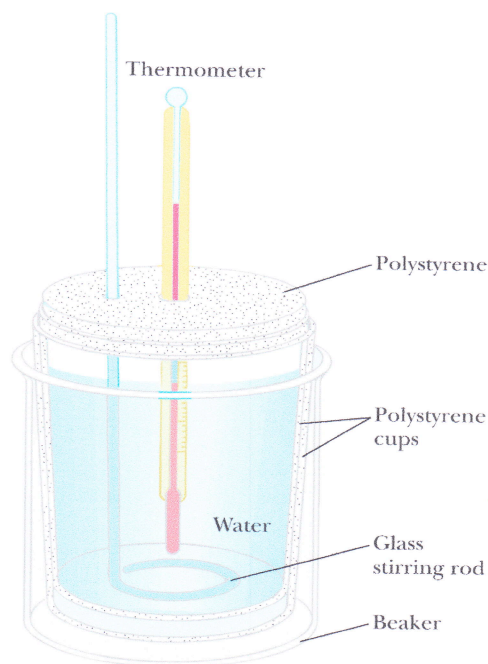
This week's lab:

$$\Delta H = \frac{\text{Joules transferred}}{\text{moles reacting}} = \frac{m \cdot c \cdot \Delta T}{g \cdot \frac{\text{mol}}{\text{g}}}$$

$$J_1 = \left[ \begin{array}{l} \text{Mg mass} \\ + \\ \text{water mass} \end{array} \right] \cdot \frac{4.18\text{J}}{\text{g}^\circ\text{C}} \cdot \Delta T$$

~100g

$$\Delta H_1 = \left( \quad \right) \left( \frac{\text{mol}}{24.305\text{g}} \right)$$



$$\begin{array}{l} \text{Mg} = 24.305\text{g/mol} \\ \text{MgO} = 40.30\text{g/mol} \end{array}$$

$$H = E + PV$$

2. Constant Volume  $\rightarrow$  Combustion Reactions Include Calorimeter!

Gives  $\Delta E$



Hydrazine  
(air bags, polymers)

$$E = E_{\text{water, abs.}} + E_{\text{calorimeter, abs.}}$$

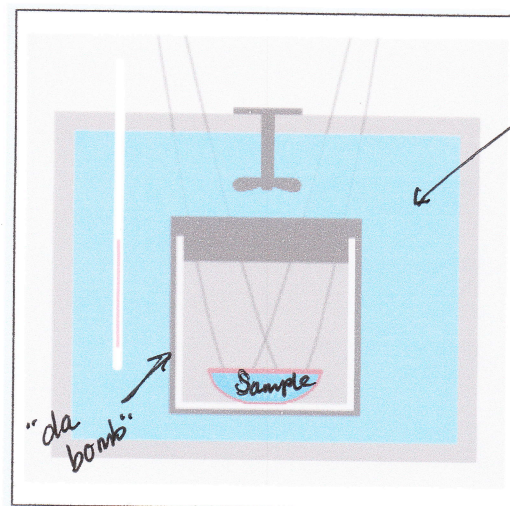
$$E_w = m \cdot c \cdot \Delta T = 600\text{g} \cdot 4.18\frac{\text{J}}{\text{g}^\circ\text{C}} \cdot 2.8^\circ\text{C}$$

$$E_w = 7022.4\text{J}$$

$$E_{\text{cal}} = \frac{420\text{J}}{^\circ\text{C}} \cdot 2.8^\circ\text{C} = 1176\text{J}$$

$$E = 1176\text{J} + 7022.4\text{J} = 8198.4\text{J}$$

$$\Delta H = \frac{E}{\text{moles}} = \frac{8198.4\text{J}}{\left(\frac{0.5\text{g}}{32.5\text{g/mol}}\right)} = \boxed{532.9\text{kJ}}$$



water

$\text{N}_2\text{H}_4$   
molar mass  
= 32.5 g/mol

0.5 g  $\text{N}_2\text{H}_4$

600g  $\text{H}_2\text{O}$

420 J/ $^\circ\text{C}$

The heat capacity  
of the calorimeter  
vessel is 420 J/ $^\circ\text{C}$

20.0  $^\circ\text{C}$   $\rightarrow$  22.8  $^\circ\text{C}$

$$\Delta T = 22.8^\circ\text{C} - 20.0^\circ\text{C} = 2.8^\circ\text{C}$$

Constant  
P

	exothermic	endothermic
$\Delta T$	+	-
$\Delta H_{\text{reaction}}$	-	+
$q_{\text{system}}$	-	+
$q_{\text{surroundings}}$	+	-

} absorbs heat  
from surr.

### Hess' Law

- Enthalpy is a state function.
- If a reaction can be written as the sum of two or more reactions, the  $\Delta H$  for the net reaction is the sum of the  $\Delta H$ 's for the individual steps.
- If a reaction is reversed, the sign of  $\Delta H$  is reversed.
- If the coefficients of the reactants/products are multiplied by a number, then the  $\Delta H$  must be multiplied by that number, as well.