

**6 • Energy and Chemical Reactions****Station 1 – EXOTHERMIC AND ENDOTHERMIC**

Classify each statement as talking about an [EXO]thermic or [ENDO]thermic reaction:

EXO surroundings get hot

EXO  $\Delta H$  is negative

ENDO PE diagram is uphill

EXO PE diagram is downhill

EXO energy is a product

ENDO surroundings get cold

ENDO  $\Delta H$  is positive

ENDO products have more energy

EXO reactants have more energy

ENDO energy is a reactant

**6 • Energy and Chemical Reactions****Station 2 – HEAT CALCULATIONS**

**45.0 mL = 45.0 g (density of water is 1 g/mL)**

A 45.0 mL sample of water is heated from 15.0°C to 35.0°C. How many joules of energy have been absorbed by the water? (Show work)

$$q = mc\Delta T = (45.0 \text{ g})(4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{C}^{-1})(20.0^\circ\text{C}) = \boxed{3770 \text{ J}} = \boxed{3.77 \text{ kJ}}$$

If 5430 J of energy is used to heat 1.25 L of room temperature water (23.0°C), what is the final temperature of the water?

$$q = mc\Delta T$$

$$5430 \text{ J} = (1,250 \text{ g})(4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{C}^{-1})(x)$$

$$x = 1.0^\circ\text{C}$$

$$\boxed{T_{\text{final}} = 24.0^\circ\text{C}}$$

## 6 • Energy and Chemical Reactions

### Station 3 – HOT AND COLD OBJECTS

A 100. gram sample of aluminum (specific heat =  $0.900 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$ ) in boiling water is added to an insulated cup containing 50.0 grams of water at  $5.00^\circ\text{C}$ . What will the final temperature of the mixture be? The specific heat of water is  $4.184 \text{ J} \cdot \text{g}^{-1} \cdot ^\circ\text{C}^{-1}$ . *Let  $x = \text{final temp}$*

$$q_{\text{lost}} = -q_{\text{gained}}$$

$$(100. \text{ g Al}) \left( \frac{0.900 \text{ J}}{\text{g}^\circ\text{C}} \right) (x - 100^\circ\text{C}) = - (50.0 \text{ g H}_2\text{O}) \left( \frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \right) (x - 5.00^\circ\text{C})$$

$$90(x - 100) = -209.2(x - 5)$$

$$90x - 9000 = -209.2x + 1046$$

$$90x + 209.2x = 1046 + 9000$$

$$299.2x = 10046$$

$$x = \frac{10046}{299.2} = 33.57 = \boxed{33.6^\circ\text{C}}$$

## 6 • Energy and Chemical Reactions

### Station 4 – WORK, HEAT, AND ENERGY

Determine the change in energy,  $\Delta E$ , for each system:

A system gives off 25.0 kJ of heat and has 15.0 kJ of work done on it. **-10.0 kJ**

*lose 25*

*gain 15*

A system takes in 75.0 kJ of heat and has 25.0 kJ of work done on it. **+100.0 kJ**

*Gain 75*

*gain 25*

A system does 45.0 kJ of work and loses 80.0 kJ of heat. **-125 kJ**

*lose 45*

*lose 80*

## 6 • Energy and Chemical Reactions

### Station 5 – HEATS OF FUSION & VAPORIZATION

Knowing that the  $\Delta H_{\text{fus}}$  for water is  $6.02 \text{ kJ}\cdot\text{mol}^{-1}$ , calculate the following:

How much energy (in kJ) is absorbed by 45.0 g of ice as it melts?

$$45.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mole H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6.02 \text{ kJ}}{1 \text{ mole H}_2\text{O}} = \boxed{15.0 \text{ kJ}} \text{ absorbed}$$

What mass of ice can be melted with 75.0 kJ of energy?

$$75.0 \text{ kJ} \times \frac{1 \text{ mole H}_2\text{O}}{6.02 \text{ kJ}} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mole H}_2\text{O}} = \boxed{224.5 \text{ g}} = \boxed{225 \text{ g H}_2\text{O}}$$

## 6 • Energy and Chemical Reactions

### Station 6 – $\Delta H$ FROM DATA

When 10.0 grams of  $\text{C}_5\text{H}_{12}$  is burned, 453 kJ of energy is released.

What is the  $\Delta H_{\text{combustion}}$  for  $\text{C}_5\text{H}_{12}$ ?

$$\frac{-453 \text{ kJ}}{10.0 \text{ g C}_5\text{H}_{12}} = \frac{72.17 \text{ g C}_5\text{H}_{12}}{1 \text{ mole C}_5\text{H}_{12}} = -3269.3 = \boxed{-3270 \text{ kJ}\cdot\text{mol}^{-1}}$$

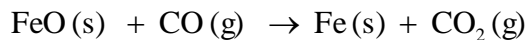
When 10.0 grams of aluminum melts, 3.929 kJ of energy is required. What is the  $\Delta H_{\text{fus}}$  of Al?

$$\frac{3.929 \text{ kJ}}{10.0 \text{ g Al}} = \frac{26.98 \text{ g Al}}{1 \text{ mole Al}} = \boxed{10.6 \text{ kJ}\cdot\text{mol}^{-1}}$$

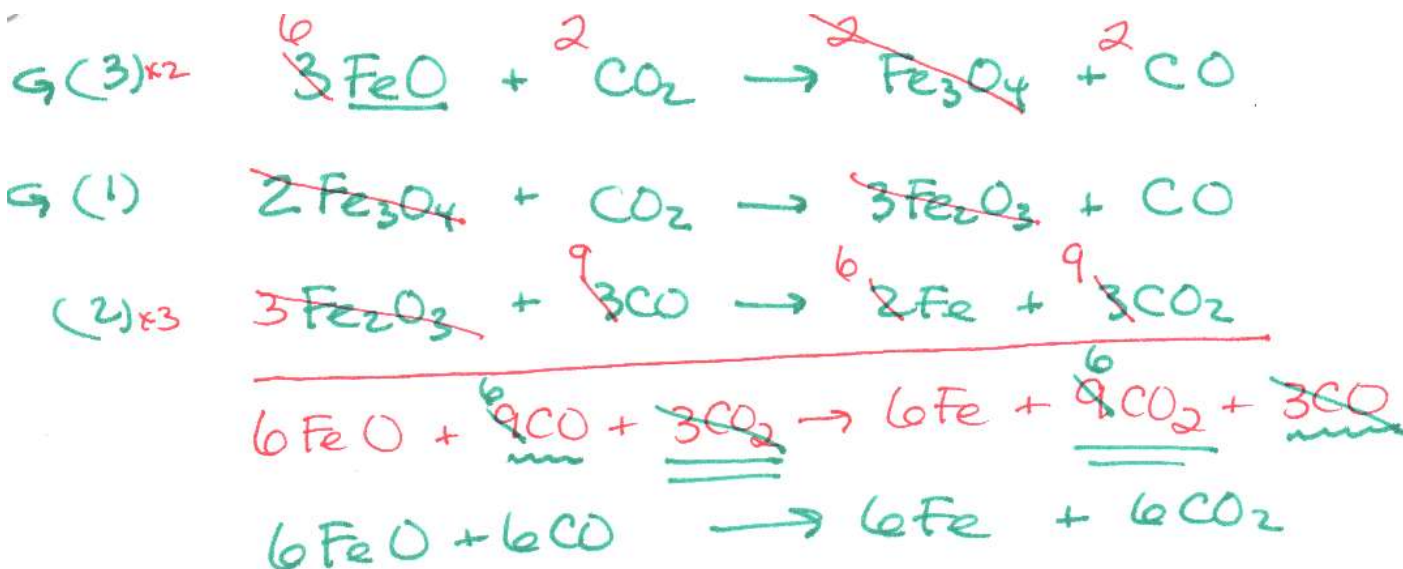
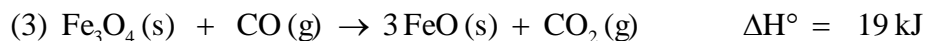
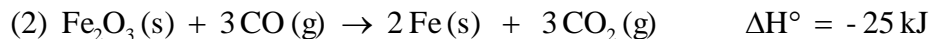
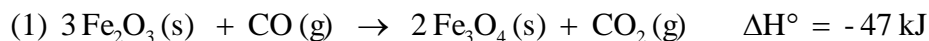
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### Station 7 – HESS'S LAW—LONG VERSION

Iron ore can be converted to iron metal with CO gas.



Calculate the standard enthalpy change for this reaction from these reactions of iron oxides with CO:



$$\begin{aligned}
 6 \Delta H &= 2(-47) + 1(-25) + 3(19) \\
 &= 2(-47) + -(-25) + 3(19)
 \end{aligned}$$

$$\Delta H = \frac{-66}{6} = \boxed{-11 \text{ kJ}}$$

## 6 • Energy and Chemical Reactions

### Station 8 – HESS'S LAW--SHORTCUT

<b>chemical</b>	CO <sub>2</sub> (g)	H <sub>2</sub> O(g)	C <sub>5</sub> H <sub>12</sub> (l)	C <sub>2</sub> H <sub>5</sub> OH(l)
<b>ΔH<sub>f</sub></b>	-393.5 kJ·mol <sup>-1</sup>	-241.8 kJ·mol <sup>-1</sup>	-173.1 kJ·mol <sup>-1</sup>	-277.6 kJ·mol <sup>-1</sup>

Given the above ΔH<sub>f</sub><sup>o</sup>'s, calculate the ΔH<sub>combustion</sub> of pentane, C<sub>5</sub>H<sub>12</sub>.



$$\begin{aligned}\Delta H_{\text{combustion}} &= 5(\Delta H_f \text{CO}_2) + 6(\Delta H_f \text{H}_2\text{O}) - \Delta H_f \text{C}_5\text{H}_{12} \\ &= 5(-393.5) + 6(-241.8) - (-173.1) \\ &= \boxed{-3245.2 \text{ kJ}\cdot\text{mol}^{-1}}\end{aligned}$$

Calculate the ΔH<sub>combustion</sub> of ethyl alcohol, C<sub>2</sub>H<sub>5</sub>OH(l)



$$\begin{aligned}\Delta H_{\text{combustion}} &= 2(\Delta H_f \text{CO}_2) + 3(\Delta H_f \text{H}_2\text{O}) - \Delta H_f \text{C}_2\text{H}_5\text{OH} \\ &= 2(-393.5) + 3(-241.8) - (-277.6) \\ &= \boxed{-1234.8 \text{ kJ}\cdot\text{mol}^{-1}}\end{aligned}$$

## 6 • Energy and Chemical Reactions

### Station 9 – MORE HESS'S LAW

<b>chemical</b>	CO <sub>2</sub> (g)	H <sub>2</sub> O(g)	C <sub>8</sub> H <sub>18</sub> (l)
<b>ΔH<sub>f</sub></b>	-393.5 kJ·mol <sup>-1</sup>	-241.8 kJ·mol <sup>-1</sup>	??? kJ·mol <sup>-1</sup>

Knowing that the ΔH<sub>combustion</sub> of octane, C<sub>8</sub>H<sub>18</sub>, is -5508.9 kJ·mol<sup>-1</sup> calculate the ΔH<sub>f</sub> of octane.



$$\begin{aligned}\Delta H_{\text{combustion}} &= 8(\Delta H_f \text{CO}_2) + 9(\Delta H_f \text{H}_2\text{O}) - \Delta H_f \text{C}_8\text{H}_{18} \\ -5508.9 &= 8(-393.5) + 9(-241.8) - (x) \\ x &= \boxed{+184.7 \text{ kJ}\cdot\text{mol}^{-1}}\end{aligned}$$