

# AP Chemistry

## Unit 5- Homework Problems

### Thermodynamics

#### Specific Heat Problems

1. How many J would it take to raise the temperature of 200 grams of water from 5 °C to 85 °C?

$$Q = (200 \text{ g})(4.18)(80 \text{ }^\circ\text{C}) = 66880 \text{ J}$$

2. How many J would problem number 1 be if it was aluminum instead of water?

$$Q = (200 \text{ g})(0.897)(80 \text{ }^\circ\text{C}) = 14352 \text{ J}$$

3. How many grams of copper could be heated from 20 °C to 75 °C if 1200 J are applied to it?

$$1200 \text{ J} = (x)(0.385)(55 \text{ }^\circ\text{C}) = 56.7 \text{ g}$$

4. What is the specific heat capacity of a substance if 750 J caused 100 grams of it to go from 90 °C to 135 °C?

$$750 \text{ J} = (100 \text{ g})(x)(45 \text{ }^\circ\text{C}) = 0.167 \text{ J/g}^\circ\text{C}$$

5. What would the final temperature be if 500 J are applied to 150 grams of ice at -90 °C?

$$500 \text{ J} = (150 \text{ g})(2.06)(x - (-90 \text{ }^\circ\text{C})) = -88.4 \text{ }^\circ\text{C}$$

6. What would the temperature *change* by if a 90 gram piece of hot iron cooled by losing 200 J.

$$-200 \text{ J} = (90 \text{ g})(0.449)(x) = -4.95 \text{ }^\circ\text{C}$$

7. What was the initial temperature if 500 J were applied to 250 g of mercury and the final temperature was 50 °C?

$$500 \text{ J} = (250 \text{ g})(0.140)(50 \text{ }^\circ\text{C} - x) = 35.7 \text{ }^\circ\text{C}$$

#### Latent Heat Problems

1. Why can we not use the equation  $Q = mc\Delta T$  for phase changes of a substance; why won't it work?

There is no temperature change at a phase change

2. What is the Latent Heat of Fusion of a substance?

Energy needed to melt or freeze

3. What is the Latent Heat of Vaporization of a substance?

Energy needed to boil or condense

4. How many J are needed to melt 30 grams of copper?

$$Q = (30 \text{ g})(209 \text{ J/g}) = 6270 \text{ J}$$

5. How much heat is needed to boil 63 grams of ethanol?

$$Q = (63 \text{ g})(838 \text{ J/g}) = 52.8 \text{ kJ}$$

6. What is the Latent Heat of Fusion of a substance if 1200 J melts 40 grams of it?

$$1200 \text{ J} = (40 \text{ g})(x) = 30 \text{ J/g}$$

7. What is the Latent Heat of Vaporization of a substance if 50 J boiled 3 grams of it?

$$50 \text{ J} = (3 \text{ g})(x) = 16.7 \text{ J/g}$$

8. How many grams of copper could be melted if 700 J are applied to it?

$$700 \text{ J} = (x)(209 \text{ J/g}) = 3.35 \text{ g}$$

9. How many grams of water could be boiled if 8000 J are applied to it?

$$8000 \text{ J} = (x)(2260 \text{ J/g}) = 3.54 \text{ g}$$

## Combination Problems

1. How much energy is needed to heat 200 grams of diethyl ether,  $(\text{C}_2\text{H}_5)_2\text{O}$ , from  $5^\circ\text{C}$  to its boiling point and boil it?

A) Heat to boiling point  $Q = mc(T_f - T_i) = (200 \text{ g})(2.33)(34.6 - 5^\circ\text{C}) = 13.8 \text{ kJ}$

B) Boil  $Q = mL_v = (200 \text{ g})(357 \text{ J/g}) = 71.4 \text{ kJ}$

$$\text{Total} = 13.8 + 71.4 = 85.2 \text{ kJ}$$

2. How much energy is needed to turn 400 grams of liquid benzene,  $\text{C}_6\text{H}_6$ , at  $20^\circ\text{C}$  to gaseous benzene at  $150^\circ\text{C}$ ?

A) Heat to boiling point  $Q = mc(T_f - T_i) = (400 \text{ g})(1.74)(80 - 20^\circ\text{C}) = 41.8 \text{ kJ}$

B) Boil  $Q = mL_v = (400 \text{ g})(393 \text{ J/g}) = 157.2 \text{ kJ}$

C) Heat gas from bp to  $150^\circ\text{C}$   $Q = mc(T_f - T_i) = (400 \text{ g})(1.06)(150 - 80) = 29.7 \text{ kJ}$

$$\text{Total} = 41.8 + 157.2 + 29.7 = 228.7 \text{ kJ}$$

3. How much energy is needed to turn a 75 g block of ice at  $-40^\circ\text{C}$  to steam at  $250^\circ\text{C}$ ?

A) Heat to melting point  $Q = mc(T_f - T_i) = (75 \text{ g})(2.06)(0 - (-40^\circ\text{C})) = 6.2 \text{ kJ}$

B) Melt  $Q = mL_f = (75 \text{ g})(333 \text{ J/g}) = 25 \text{ kJ}$

C) Heat liq from mp to bp  $Q = mc(T_f - T_i) = (75 \text{ g})(4.18)(100 - 0^\circ\text{C}) = 31.4 \text{ kJ}$

B) Boil  $Q = mL_v = (75 \text{ g})(2260 \text{ J/g}) = 169.5 \text{ kJ}$

C) Heat gas from bp to  $250^\circ\text{C}$   $Q = mc(T_f - T_i) = (75 \text{ g})(1.86)(250 - 100^\circ\text{C}) = 20.9 \text{ kJ}$

$$\text{Total} = 6.2 + 25 + 31.4 + 169.5 + 20.9 \text{ kJ} = 253 \text{ kJ}$$

4) Which part of #3 contributes the most energy to the overall process?

The boiling part (going from liquid to gas or vice versa) requires the most energy

## Latent Heat Graphing

1. What do you call each of the following changes of state:

A. Solid to liquid **melt**

B. Liquid to gas **boil**

C. Gas to solid **deposition**

D. Solid to gas **sublimation**

E. Liquid to solid **freeze**

F. Gas to liquid **condense**

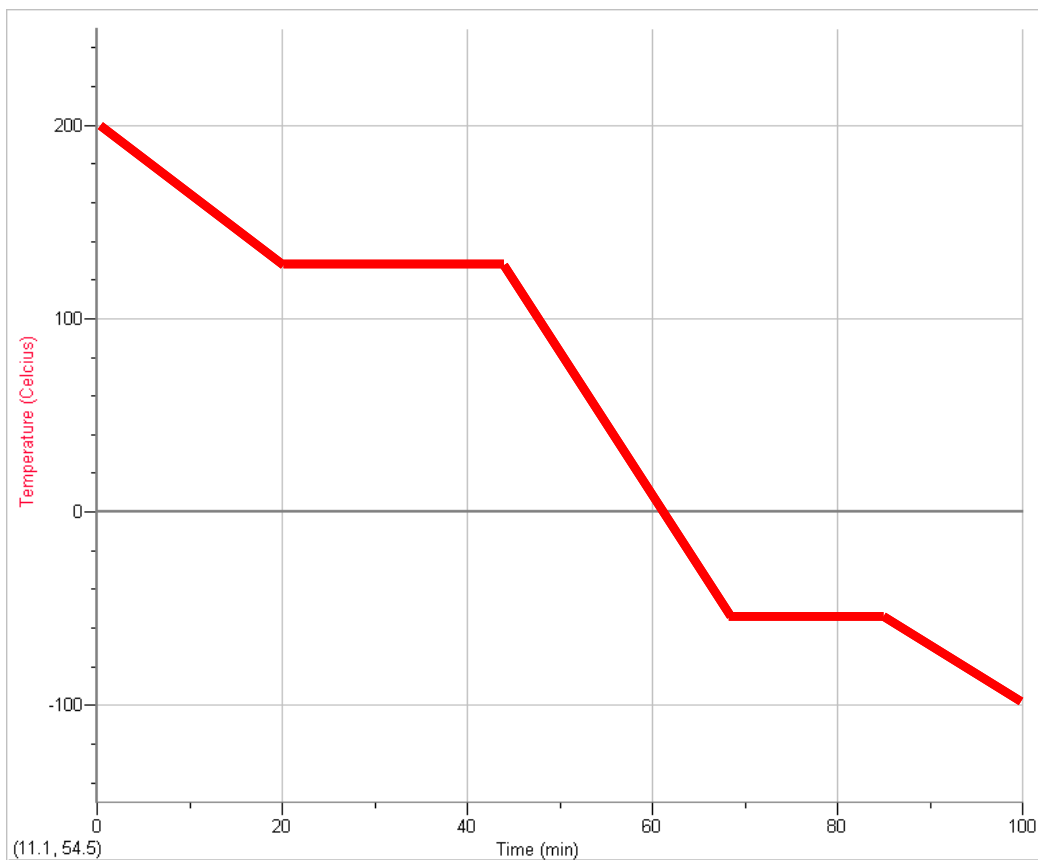
2. The melting point of a solid is the same as what other point?

**Freezing point**

3. The condensing point of a gas is the same as what other point?

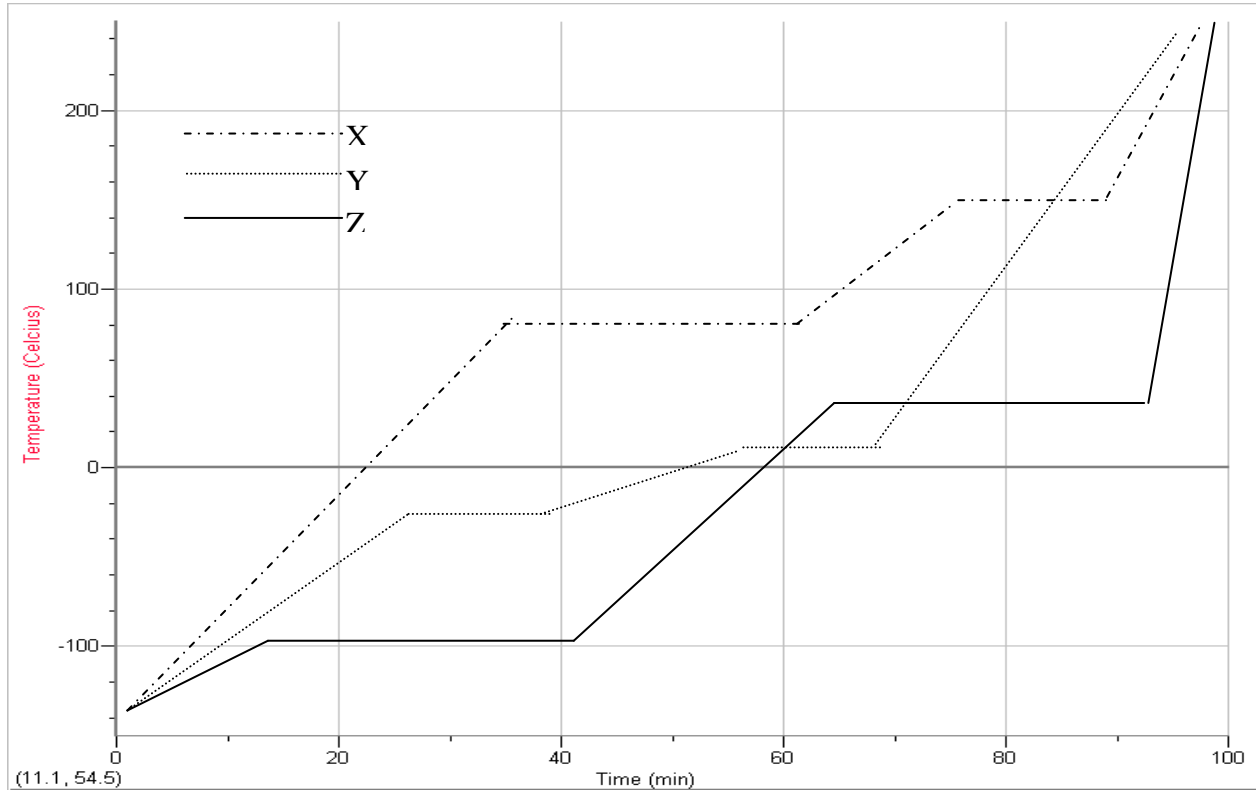
**Boiling point**

4. Octane, the major component in gasoline freezes at  $-57\text{ }^{\circ}\text{C}$  and boils at  $125\text{ }^{\circ}\text{C}$ . If gaseous octane was cooled from  $200\text{ }^{\circ}\text{C}$  to  $-100\text{ }^{\circ}\text{C}$ , draw what the graph would look like:



5. The following graph shows three different substances and their phase changes. Answer the following questions about them:

- A. What is the melting point of substance Y? @  $-25\text{ }^{\circ}\text{C}$
- B. What is the boiling point of substance Z? @  $40\text{ }^{\circ}\text{C}$
- C. Which substance is still a solid when the other two substances have turned to gases? X
- D. Which substance has the lowest boiling point? Y
- E. Which substance has the lowest freezing point? Z



6. Water boils at  $100\text{ }^{\circ}\text{C}$ . Why are steam burns potentially more dangerous than boiling water burns?

Under normal conditions, liquid water can only get up to  $100\text{ }^{\circ}\text{C}$ . Steam, because it is a gas, can get hotter than  $100\text{ }^{\circ}\text{C}$  and is potentially more dangerous because of it.

# Calorimetry

1. What is the specific heat capacity of water? **4.18 J/g°C**

2. A piece of ice is placed on the sidewalk on a hot summer day. Describe what happens to energy content of the piece of ice, does it gain or lose energy? What happens to the air around the ice, does it gain or lose energy?

**The ice gains energy from the surrounding air and increases its energy (endothermic).**

**The air loses energy to the ice and decreases its energy (exothermic)**

3. 100 grams of hot water at 80 °C is combined with 100 grams of cool water at 20 °C. What is the final temperature of the combined water?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mc(T_f - T_i) \quad (100 \text{ g})(4.18)(x - 80 \text{ }^\circ\text{C}) = - (100 \text{ g})(4.18)(x - 20 \text{ }^\circ\text{C}) \quad T_f = 50 \text{ }^\circ\text{C}$$

4. 100 grams of hot water at 80 °C is combined with 50 grams of cool water at 20 °C. What is the final temperature of the combined water?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mc(T_f - T_i) \quad (100 \text{ g})(4.18)(x - 80 \text{ }^\circ\text{C}) = - (50 \text{ g})(4.18)(x - 20 \text{ }^\circ\text{C}) \quad T_f = 60 \text{ }^\circ\text{C}$$

5. A 25 gram piece of hot metal at 97 °C is plunged into a 35 gram cup of cool water at 19 °C. The metal gives up its heat to the water until they are both at 22 °C. What is the specific heat capacity of the metal?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mc(T_f - T_i) \quad (25 \text{ g})(x)(22 - 97 \text{ }^\circ\text{C}) = - (35 \text{ g})(4.18)(22 - 19 \text{ }^\circ\text{C}) \quad C = 0.234 \text{ J/g}^\circ\text{C}$$

6. A 75 gram piece of hot metal at 100 °C is dropped into a 50 gram cup of cool water at 22 °C. The final temperature of the system is 30 °C. What is the specific heat capacity of the metal?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mc(T_f - T_i) \quad (75 \text{ g})(x)(30 - 100 \text{ }^\circ\text{C}) = - (50 \text{ g})(4.18)(30 - 22 \text{ }^\circ\text{C}) \quad C = 0.318 \text{ J/g}^\circ\text{C}$$

7. A 120 gram piece of copper at 75 °C is put into a 20 gram sample of water at 10 °C. What is the final temperature of the system after the copper releases all its extra heat?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mc(T_f - T_i) \quad (120 \text{ g})(0.385)(x - 75 \text{ }^\circ\text{C}) = - (20 \text{ g})(4.18)(x - 10 \text{ }^\circ\text{C}) \quad T_f = 33.1 \text{ }^\circ\text{C}$$

8. 100 g of water at 22 °C was combined with 20 grams of ice. The final temperature of the system was 6 °C. Calculate the heat of fusion of ice based upon this data.

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mL_f \quad (100 \text{ g})(4.18)(6 - 22 \text{ }^\circ\text{C}) = - (20 \text{ g})(x) \quad L_f = 334.4 \text{ J/g}$$

9. 200 g of water at 22 °C was combined with 15 grams of ice. Knowing that the heat of fusion of ice is 333 J/g, calculate the final temperature of the system.

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mL_f \quad (200 \text{ g})(4.18)(x - 22 \text{ }^\circ\text{C}) = - (15 \text{ g})(333 \text{ J/g}) \quad T_f = 16 \text{ }^\circ\text{C}$$

10. Suppose you heat a 50 g piece of silver to 99.8 °C and then drop it onto ice. When the metal's temperature has dropped to 0 °C, it is found that 3.54 g of ice has melted. What is the specific heat capacity of silver?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mL_f \quad (50 \text{ g})(x)(0 - 99.8 \text{ }^\circ\text{C}) = - (3.54 \text{ g})(333 \text{ J/g}) \quad c = 0.236 \text{ J/g}^\circ\text{C}$$

11. A 9.36 g piece of Pt is heated to 98.6 °C and then dropped onto a block of ice. When the temperature of the metal has dropped to 0 °C, it is found that 0.37 g of ice melted. What is the specific heat capacity of Pt?

$$Q_{\text{lost}} = - Q_{\text{gain}} \quad mc(T_f - T_i) = - mL_f \quad (9.36 \text{ g})(x)(0 - 98.6 \text{ }^\circ\text{C}) = - (0.37 \text{ g})(333 \text{ J/g}) \quad T_f = 0.134 \text{ J/g}^\circ\text{C}$$

12. Describe each of the reactions below as either endothermic or exothermic. Some are tricky, be careful. Decide whether heat is entering or leaving the system in question:

A. Exploding fireworks	endothermic	<b>exothermic</b>
B. Melting snow	<b>endothermic</b>	exothermic
C. Lava cooling	endothermic	<b>exothermic</b>
D. Paper burning	endothermic	<b>exothermic</b>
E. An ice cube freezing	endothermic	<b>exothermic</b>
F. Water evaporating	<b>endothermic</b>	exothermic

# Enthalpy Problems

1. For the equation:  $\text{Mg} + \frac{1}{2} \text{CO}_2 \rightarrow \text{MgO} + \frac{1}{2} \text{C}$

If 16.7 kJ of heat is given off per 1.0 gram of magnesium reacted, calculate the  $\Delta H_{\text{rxn}}$

$$\Delta H_{\text{rxn}}^{\circ} = Q_{\text{rxn}}/\text{moles} \quad \Delta H_{\text{rxn}}^{\circ} = -16.7 \text{ kJ}/0.0411 \text{ moles} = -406 \text{ kJ/mole}$$

2. For the equation:  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 \rightarrow \text{N}_2 + 4 \text{H}_2\text{O} + \text{Cr}_2\text{O}_3$

If 3 grams of  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$  gives off 3.57 kJ of energy, calculate the  $\Delta H_{\text{rxn}}$

$$\Delta H_{\text{rxn}}^{\circ} = Q_{\text{rxn}}/\text{moles} \quad \Delta H_{\text{rxn}}^{\circ} = -3.57 \text{ kJ}/0.0119 \text{ moles} = -300 \text{ kJ/mole}$$

3. For the equation:  $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \frac{1}{2} \text{H}_2$

If 5 grams of sodium is placed into 100 grams of water at 20 °C and the final temperature of the system reaches 27 °C, calculate the  $\Delta H_{\text{rxn}}$

$$\Delta H_{\text{rxn}}^{\circ} = Q_{\text{rxn}}/\text{moles} \quad Q_{\text{rxn}} = -Q_{\text{water}} = -(100 \text{ g})(4.18 \text{ J/g}^{\circ}\text{C})(7^{\circ}\text{C}) = -2926 \text{ J}$$

$$\Delta H_{\text{rxn}}^{\circ} = -2926 \text{ J}/0.217 \text{ moles} = -13.5 \text{ kJ/mole}$$

4. For the equation:  $\text{Ca} + 2 \text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2$

If 2 grams of calcium is placed into 75 mL of 1 M HCl at 18 °C and the final temperature of the system reaches 23 °C, calculate the  $\Delta H_{\text{rxn}}$ . (Assume the density of the solution is 1 g/mL and the specific heat capacity of the HCl = 4.18 J/g°C)

$$\Delta H_{\text{rxn}}^{\circ} = Q_{\text{rxn}}/\text{moles} \quad Q_{\text{rxn}} = -Q_{\text{water}} = -(75 \text{ g})(4.18 \text{ J/g}^{\circ}\text{C})(5^{\circ}\text{C}) = -1567.5 \text{ J}$$

$$\Delta H_{\text{rxn}}^{\circ} = -1567.5 \text{ J}/0.0499 \text{ moles} = -31.4 \text{ kJ/mole}$$

5. For the equation:  $\text{NaNO}_3 (\text{s}) \rightarrow \text{NaNO}_3 (\text{aq})$

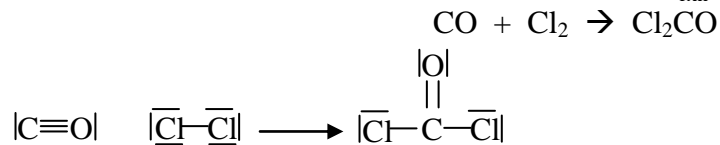
If 20 grams of  $\text{NaNO}_3$  were placed into 200 grams of water at 22 °C, and the temperature dropped to 12 °C, what is the  $\Delta H_{\text{rxn}}$ ?

$$\Delta H_{\text{rxn}}^{\circ} = Q_{\text{rxn}}/\text{moles} \quad Q_{\text{rxn}} = -Q_{\text{water}} = -(200 \text{ g})(4.18 \text{ J/g}^{\circ}\text{C})(-10^{\circ}\text{C}) = +8360 \text{ J}$$

$$\Delta H_{\text{rxn}}^{\circ} = +8360 \text{ J}/0.235 \text{ moles} = +35.6 \text{ kJ/mole}$$

# Bond Energies

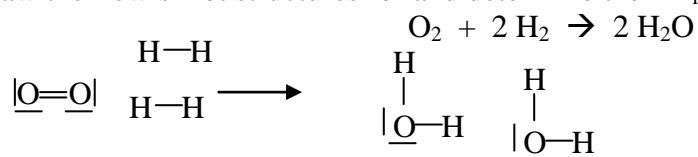
1. Draw the Lewis Dot structures and determine the  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = \Sigma \text{Bonds}_{\text{react}} - \Sigma \text{Bonds}_{\text{prod}}$$

$$\Delta H_{\text{rxn}} = (1046 + 242) - (2*339 + 745) = -135 \text{ kJ}$$

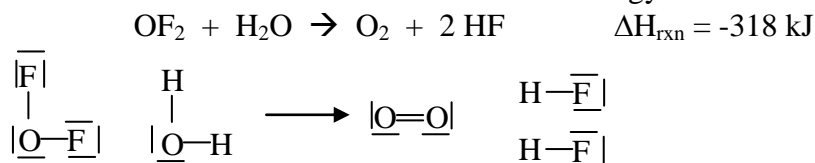
2. Draw the Lewis Dot structures for and determine the  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = \Sigma \text{Bonds}_{\text{react}} - \Sigma \text{Bonds}_{\text{prod}}$$

$$\Delta H_{\text{rxn}} = (498 + 2*436) - (4*463) = -482 \text{ kJ}$$

3. Draw the Lewis Dot structures for and determine the energy for the O-F bond for:

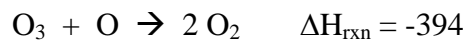


$$\Delta H_{\text{rxn}} = \Sigma \text{Bonds}_{\text{react}} - \Sigma \text{Bonds}_{\text{prod}}$$

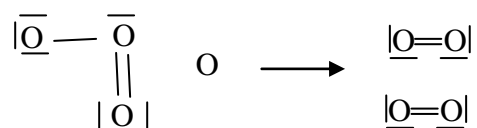
$$-318 = (2X + 2*463) - (498 + 2*565)$$

$$X = 192 \text{ kJ}$$

4. For the reaction:



a. Draw the Lewis dot structures for all species



b. Calculate the O-O bond energy in  $\text{O}_3$

$$\Delta H_{\text{rxn}} = \sum \text{Bonds}_{\text{react}} - \sum \text{Bonds}_{\text{prod}}$$

$$-394 = (2X) - (2 \cdot 498)$$

$$X = 301 \text{ kJ}$$

c. What is the bond order of the O-O bond in  $\text{O}_3$  (remember it is a resonance structure)?

$$\text{Bond order} = 1.5$$

d. Compare the answer you got in part b to the O-O bond energy and the O=O bond energy you can look up. Does your answer make sense? Why or why not?

$$\text{O-O bond} = 146 \quad \text{bond order} = 1$$

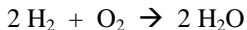
$$\text{O=O bond} = 498 \quad \text{bond order} = 2$$

O---O bond in  $\text{O}_3$       bond order 1.5 has energy of 301 kJ so that makes sense as it is between the 146 and 498 or the bond orders 1 and 2

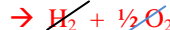


# Hess' Law Problems

1. Given the following equations:



$$\Delta H_{\text{rxn}} = -572 \text{ kJ}$$



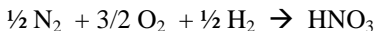
$$\Delta H_{\text{rxn}} = +286 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -74 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = +74 \text{ kJ}$$

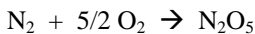


$$\Delta H_{\text{rxn}} = -174 \text{ kJ}$$



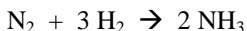
$$\Delta H_{\text{rxn}} = 2(-174 \text{ kJ})$$

Calculate  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = +12 \text{ kJ}$$

2. Given the following equations:



$$\Delta H_{\text{rxn}} = -92 \text{ kJ}$$



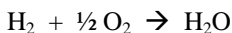
$$\Delta H_{\text{rxn}} = -46 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -906 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -226 \text{ kJ}$$

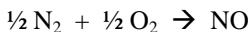


$$\Delta H_{\text{rxn}} = -242 \text{ kJ}$$



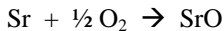
$$\Delta H_{\text{rxn}} = +363 \text{ kJ}$$

Calculate the  $\Delta H_{\text{rxn}}$  for:

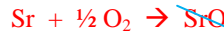


$$\Delta H_{\text{rxn}} = +91 \text{ kJ}$$

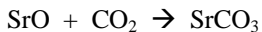
3. Given the following equations:



$$\Delta H_{\text{rxn}} = -592 \text{ kJ}$$



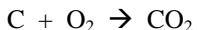
$$\Delta H_{\text{rxn}} = -592 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -234 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -234 \text{ kJ}$$

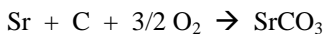


$$\Delta H_{\text{rxn}} = -394$$



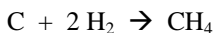
$$\Delta H_{\text{rxn}} = -394$$

Calculate the  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = -1220 \text{ kJ}$$

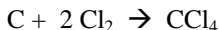
4. Given the following equations:



$$\Delta H_{\text{rxn}} = -75 \text{ kJ}$$



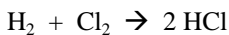
$$\Delta H_{\text{rxn}} = +75 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -96 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -96 \text{ kJ}$$

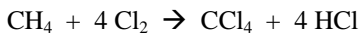


$$\Delta H_{\text{rxn}} = -92 \text{ kJ}$$



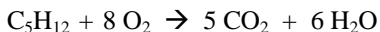
$$\Delta H_{\text{rxn}} = 2(-92) \text{ kJ}$$

Calculate the  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = -205 \text{ kJ}$$

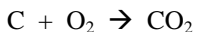
5. Given the following equations:



$$\Delta H_{\text{rxn}} = -3506 \text{ kJ}$$



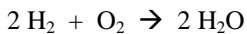
$$\Delta H_{\text{rxn}} = +3506 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -394 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = 5(-394) \text{ kJ}$$

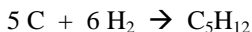


$$\Delta H_{\text{rxn}} = -484 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = 3(-484) \text{ kJ}$$

Calculate the  $\Delta H_{\text{rxn}}$  for:



$$\Delta H_{\text{rxn}} = +84 \text{ kJ}$$

# $\Delta H_f^\circ$

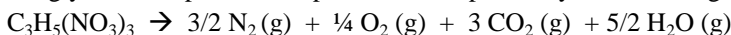
1. For each of the substances below, write a balanced equation showing the formation of 1 mole of the compound from its elements in their standard states. Look up the value for  $\Delta H_f^\circ$  in the Appendix in the back of a book or online.

- a)  $\text{Al}_2\text{O}_3$                      $2 \text{Al (s)} + 3/2 \text{O}_2 \text{(g)} \rightarrow \text{Al}_2\text{O}_3 \text{(s)}$                      $\Delta H_f^\circ = -1676 \text{ kJ/mole}$   
b)  $\text{Mg(OH)}_2$                      $\text{Mg (s)} + \text{H}_2 \text{(g)} + \text{O}_2 \text{(g)} \rightarrow \text{Mg(OH)}_2 \text{(s)}$                      $\Delta H_f^\circ = -925 \text{ kJ/mole}$   
c)  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$                      $12 \text{C (s)} + 11 \text{H}_2 \text{(g)} + 11/2 \text{O}_2 \text{(g)} \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11} \text{(s)}$                      $\Delta H_f^\circ = -2226 \text{ kJ/mole}$   
d)  $\text{NaHCO}_3$                      $\text{Na (s)} + 1/2 \text{H}_2 \text{(g)} + \text{C (s)} + 3/2 \text{O}_2 \text{(g)} \rightarrow \text{NaHCO}_3 \text{(s)}$                      $\Delta H_f^\circ = -951 \text{ kJ/mole}$

2. For each of the reactions below, calculate  $\Delta H_{\text{rxn}}^\circ$ . Look up the values for  $\Delta H_f^\circ$  in the Appendix in the back of a book or online.

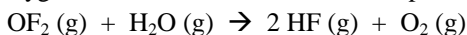
- a)  $\text{SO}_2 \text{(g)} + 1/2 \text{O}_2 \text{(g)} \rightarrow \text{SO}_3 \text{(g)}$                      $\Delta H_{\text{rxn}}^\circ = (-396) - (-297) = -99 \text{ kJ}$   
b)  $4 \text{NH}_3 \text{(g)} + 5 \text{O}_2 \text{(g)} \rightarrow 4 \text{NO (g)} + 6 \text{H}_2\text{O (g)}$                      $\Delta H_{\text{rxn}}^\circ = [4(91) + 6(-242)] - 4(-46) = -904 \text{ kJ}$   
c)  $\text{NH}_4\text{NO}_3 \text{(s)} \rightarrow \text{N}_2\text{O (g)} + 2 \text{H}_2\text{O (g)}$                      $\Delta H_{\text{rxn}}^\circ = [(82) + 2(-242)] - (-366) = -36 \text{ kJ}$

3. Nitroglycerin is a powerful explosive that explodes by the following equation while giving off 4200 kJ of heat.



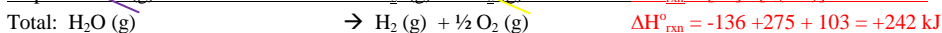
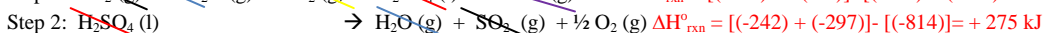
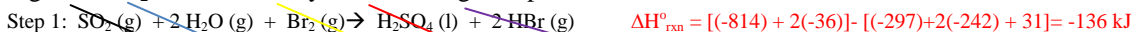
Calculate the  $\Delta H_f^\circ$  of nitroglycerin.  $-4200 \text{ kJ} = [3(-394) + 5/2(-242)] - x$                      $x = 2413 \text{ kJ}$

4. Oxygen difluoride reacts with water vapor to produce 318 kJ of heat by the following equation:



Calculate the  $\Delta H_f^\circ$  of  $\text{OF}_2$                      $-318 \text{ kJ} = [2(-273)] - [x + (-242)]$                      $x = 14 \text{ kJ}$

5. Large scale  $\text{H}_2$  can be made by the following 3 steps:



Calculate the  $\Delta H_{\text{rxn}}^\circ$  of each step. Calculate the  $\Delta H_{\text{rxn}}^\circ$  of the overall reaction. Is the overall reaction endo or exothermic?

Endothermic

## Bomb Calorimetry

1. You burn 0.300 g of C in an excess of O<sub>2</sub> in a bomb calorimeter. The temperature of the calorimeter which contains 775 grams of water increases from 25 to 27.38 °C. The heat capacity of the bomb is 893 J/K. What is ΔH<sub>rxn</sub> per mole of carbon?

$$Q_{\text{rxn}} = - [q_{\text{water}} + q_{\text{bomb}}] = - [mc\Delta T_{\text{water}} + k\Delta T_{\text{bomb}}]$$

$$Q_{\text{rxn}} = -[(775 \text{ g})(4.18 \text{ J/gC})(2.38 \text{ }^\circ\text{C}) + (893 \text{ J/K})(2.38 \text{ K})] = -9.835 \text{ kJ}$$

$$0.30 \text{ g C} * (1 \text{ mole}/12 \text{ g}) = 0.025 \text{ moles C}$$

$$\Delta H_{\text{rxn}} = q/\text{mole} = -9.835 \text{ kJ}/0.025 \text{ mole} = -393 \text{ kJ/mole}$$

2. You burn 1.50 g of benzoic acid (C<sub>6</sub>H<sub>5</sub>CO<sub>2</sub>H) in a bomb calorimeter which has a constant of 789 J/K. The calorimeter is filled with 775 g of water and the temperature increases from 22.50 to 31.69 °C. What is ΔH<sub>rxn</sub> per mole of benzoic acid?

$$Q_{\text{rxn}} = - [q_{\text{water}} + q_{\text{bomb}}] = - [mc\Delta T_{\text{water}} + k\Delta T_{\text{bomb}}]$$

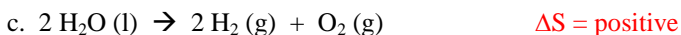
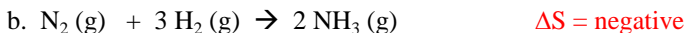
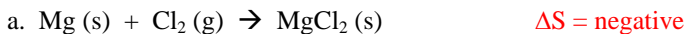
$$Q_{\text{rxn}} = -[(775 \text{ g})(4.18 \text{ J/gC})(9.19 \text{ }^\circ\text{C}) + (789 \text{ J/K})(9.19 \text{ K})] = -37.0 \text{ kJ}$$

$$1.50 \text{ g C}_6\text{H}_5\text{CO}_2\text{H} * (1 \text{ mole}/122 \text{ g}) = 0.0123 \text{ moles C}_6\text{H}_5\text{CO}_2\text{H}$$

$$\Delta H_{\text{rxn}} = q/\text{mole} = -37 \text{ kJ}/0.0123 \text{ mole} = -3008 \text{ kJ/mole}$$

## Entropy Problems

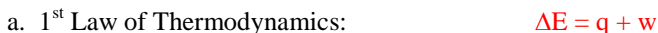
1. Without doing any calculations, determine the sign of each of the entropy changes below:



2. Calculate the ΔS° of each of the above equations and compare the value to your predictions from #1.



3. Summarize each of the laws of Thermodynamics:



4. List three things the entropy value of a substance depends upon:

a. Temperature

b. State of matter

c. Molecular complexity

5. Define entropy. A measure of the energy randomization or energy dispersal in a system (disorder)

# Spontaneity and Gibb's Free Energy

1. Describe how the sign of  $\Delta G$  determines the spontaneity of a reaction:

- $\Delta G > 0$  Non-spontaneous (positive)
- $\Delta G < 0$  Spontaneous (negative)
- $\Delta G = 0$  At equilibrium

2. Given the following thermodynamic values, determine how temperature effects the spontaneity of the reaction. Is the reaction enthalpy driven, entropy driven, both, or neither.

- |                              |                           |                 |                |      |         |
|------------------------------|---------------------------|-----------------|----------------|------|---------|
| a. $\Delta H = + 327$ kJ/mol | $\Delta S = + 127$ J/molK | enthalpy driven | entropy driven | both | neither |
| b. $\Delta H = - 187$ kJ/mol | $\Delta S = - 298$ J/molK | enthalpy driven | entropy driven | both | neither |
| c. $\Delta H = + 194$ kJ/mol | $\Delta S = - 4$ J/molK   | enthalpy driven | entropy driven | both | neither |
| a. $\Delta H = - 397$ kJ/mol | $\Delta S = + 653$ J/molK | enthalpy driven | entropy driven | both | neither |

3. For each reaction below, determine the temperature at which the reaction will be spontaneous.

- |                              |                           |                                                                 |
|------------------------------|---------------------------|-----------------------------------------------------------------|
| a. $\Delta H = + 245$ kJ/mol | $\Delta S = + 48$ J/molK  | $T = \Delta H/\Delta S = 245/0.048$ $T > 5104$ K spontaneous    |
| b. $\Delta H = + 187$ kJ/mol | $\Delta S = + 365$ J/molK | $T = \Delta H/\Delta S = 187/0.365$ $T > 512$ K spontaneous     |
| c. $\Delta H = - 456$ kJ/mol | $\Delta S = - 38$ J/molK  | $T = \Delta H/\Delta S = -456/-0.038$ $T < 12000$ K spontaneous |
| d. $\Delta H = - 547$ kJ/mol | $\Delta S = - 97$ J/molK  | $T = \Delta H/\Delta S = -547/-0.097$ $T < 5639$ K spontaneous  |

4. Define Gibb's Free Energy **The energy available to do work in a thermodynamic system**

5. Using the values of  $\Delta G_f^\circ$ , calculate the  $\Delta G_{rxn}^\circ$  of each of the equations below. Is each reaction spontaneous at standard conditions?

- |                                                            |                                                           |
|------------------------------------------------------------|-----------------------------------------------------------|
| a. $N_2O_4(g) \rightarrow 2 NO_2(g)$                       | $\Delta G = [2(51)] - [100] = 2$ kJ Non-spontaneous       |
| b. $NH_4Cl(s) \rightarrow HCl(g) + NH_3(g)$                | $\Delta G = [(-95) + (-16)] - [-203] = 92$ kJ Non-spontan |
| c. $3 H_2(g) + Fe_2O_3(s) \rightarrow 2 Fe(s) + 3 H_2O(g)$ | $\Delta G = [3(-229)] - [-742] = 55$ kJ Non-spontaneous   |
| d. $N_2(g) + 3 H_2(g) \rightarrow 2 NH_3(g)$               | $\Delta G = [2(-16)] - [0] = -32$ kJ Spontaneous          |

## Energy & Voltage

1. Write the equation that relates  $\Delta G^\circ$  and  $E^\circ$ ?  $\Delta G^\circ = -nFE^\circ$
2. What is n? **the number of moles of electrons transferred**
3. What is F? **Faraday's constant, 96500 J/voltmole**
4. For each of the following equations, what must n be?
  - a.  $\text{Cu}^{+2} + \text{Mg} \rightarrow \text{Mg}^{+2} + \text{Cu}$  **n = 2**
  - b.  $2 \text{Ag}^{+1} + \text{Sn} \rightarrow \text{Sn}^{+2} + 2 \text{Ag}$  **n = 2**
  - c.  $2 \text{Al} + 3 \text{Pb}^{+2} \rightarrow 3 \text{Pb} + 2 \text{Al}^{+3}$  **n = 6**
5. What is the  $\Delta G^\circ$  of the following batteries?
  - a.  $\text{Cu}^{+2} + \text{Mg} \rightarrow \text{Mg}^{+2} + \text{Cu}$   
 $E^\circ = 2.71 \text{ V}$   $\Delta G^\circ = -(2 \text{ moles})(96500 \text{ J/Vmol})(2.71 \text{ V}) = -523 \text{ kJ}$
  - b.  $2 \text{Ag}^{+1} + \text{Sn} \rightarrow \text{Sn}^{+2} + 2 \text{Ag}$   
 $E^\circ = 0.94 \text{ V}$   $\Delta G^\circ = -(2 \text{ moles})(96500 \text{ J/Vmol})(0.94 \text{ V}) = -181 \text{ kJ}$
  - c.  $2 \text{Al} + 3 \text{Pb}^{+2} \rightarrow 3 \text{Pb} + 2 \text{Al}^{+3}$   
 $E^\circ = 1.53 \text{ V}$   $\Delta G^\circ = -(6 \text{ moles})(96500 \text{ J/Vmol})(1.53 \text{ V}) = -886 \text{ kJ}$
6. What sign must each of the following be for a battery?
  - a.  $E^\circ$  **positive**
  - b.  $\Delta G^\circ$  **negative**

## Combination Thermo Problems

1. For the reaction:  $2 \text{NO} (\text{g}) + \text{O}_2 (\text{g}) \rightarrow 2 \text{NO}_2 (\text{g})$   $\Delta H_{\text{rxn}} = -114.1 \text{ kJ}$   $\Delta S_{\text{rxn}} = -146.5 \text{ J/K}$ 
  - a. How much heat is released when 73.1 g of NO is converted to  $\text{NO}_2$ ?  
 $73.1 \text{ g NO} * (1 \text{ mole}/30 \text{ g}) = 2.44 \text{ moles NO} * (-114.1 \text{ kJ}/2 \text{ mole NO}) = -139 \text{ kJ}$
  - b. Indicate whether the value of  $\Delta G_{\text{rxn}}$  would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.  
 $\Delta G = \Delta H - T\Delta S$  since  $\Delta S$  is negative, as T increased, it will make  $\Delta G$  more positive as you subtract a negative value making it more positive
  - c. The value of  $S^\circ = 210.8 \text{ J/Kmol}$  for NO. The value of  $S^\circ = 240.1 \text{ J/Kmol}$  for  $\text{NO}_2$ . Calculate the value of  $S^\circ$  for  $\text{O}_2$   
 $\Delta S = (2 * S_{\text{NO}_2}) - (2 * S_{\text{NO}} + S_{\text{O}_2})$   
 $-146.5 = (2 * 240.1) - (2 * 210.8 + X)$   
 $X = 205 \text{ J/K}$
  - d. The bond energy of the N-O bond in NO is 607 kJ/mol. The bond energy of the O-O bond in  $\text{O}_2$  is 495 kJ/mol. Calculate the bond energy of the N-O bond in  $\text{NO}_2$ .  
 $\Delta H_{\text{rxn}} = \sum \text{Bonds}_{\text{react}} - \sum \text{Bonds}_{\text{prod}}$   
 $-114.1 = (2 * 607 + 495) - 2(2 * X)$   
 $X = 456 \text{ kJ/mole}$

2. For the substance propane, C<sub>3</sub>H<sub>8</sub>:

a. Write a balanced equation for the complete combustion of propane with oxygen to yield carbon dioxide and water.



b. Calculate the volume of air at 30 °C and 1.0 atm that is needed to burn completely 10.0 grams of propane. Assume that air is 21% O<sub>2</sub>.

$$10 \text{ g C}_3\text{H}_8 * (1 \text{ mole}/44 \text{ g}) = 0.227 \text{ moles C}_3\text{H}_8 * (5 \text{ mole O}_2/1 \text{ mole C}_3\text{H}_8) = 1.136 \text{ mole O}_2$$

$$V = nRT/P = (1.136 \text{ mole})(0.0821)(303 \text{ K})/(1 \text{ atm}) = 28.3 \text{ L of O}_2 \text{ needed}$$

Since air is only 21% O<sub>2</sub> you need 134.6 L air

c. The heat of combustion of propane is -2220.1 kJ/mole. Knowing that ΔH<sub>f</sub> = -285.3 kJ/mole for H<sub>2</sub>O and ΔH<sub>f</sub> = -393.5 kJ/mole for CO<sub>2</sub>, calculate the ΔH<sub>f</sub> for propane.

$$\Delta H_{\text{rxn}} = \Sigma \Delta H_{\text{f}}^{\circ} \text{ prod} - \Sigma \Delta H_{\text{f}}^{\circ} \text{ react}$$

$$\Delta H_{\text{rxn}} = (3 * \text{CO}_2 + 4 * \text{H}_2\text{O}) - X$$

$$-2220.1 = (3 * (-393.5) + 4 * (-285.3)) - X$$

$$-2220.1 = -1180.5 - 1141.2 - X$$

$$X = -101.6 \text{ kJ/mole}$$

d. Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kg of water (C<sub>water</sub> = 4.18 J/gK), calculate the increase in temperature of the water.

$$0.682 \text{ moles C}_3\text{H}_8 * (-2220.1 \text{ kJ/mole}) = -1514 \text{ kJ or } 1514000 \text{ J given to water}$$

$$q_{\text{water}} = mc\Delta T \quad 1514000 \text{ J} = (8000 \text{ g})(4.18 \text{ J/gK})\Delta T \quad \Delta T = 45.3 \text{ }^{\circ}\text{C}$$

3. For the reaction:  $2 \text{Fe} (s) + 3/2 \text{O}_2 (g) \rightarrow \text{Fe}_2\text{O}_3 (s)$   $\Delta H_f^\circ = -824 \text{ kJ/mol}$

A 75.0 g sample of Fe is mixed with 11.5 L of  $\text{O}_2$  at 2.66 atm and 298 K.

a. Calculate the number of moles of:

i) Fe **1.34 moles**

ii)  $\text{O}_2$  **1.25 moles**

b. Identify the limiting reagent. Justify your answer.

**Try to use all 1.34 moles of Fe:**

**1.34 moles Fe  $\times$  (3/2 moles  $\text{O}_2$ /2 moles Fe) = 1.005 moles  $\text{O}_2$  needed which we have enough of so Fe is limiting reagent**

c. Calculate the number of moles of  $\text{Fe}_2\text{O}_3$  produced if the reaction goes to completion.

**1.34 moles Fe  $\times$  (1 mole  $\text{Fe}_2\text{O}_3$ /2 mole Fe) = 0.67 moles  $\text{Fe}_2\text{O}_3$**

d. If  $\Delta G_f^\circ = -740 \text{ kJ/mole}$  for  $\text{Fe}_2\text{O}_3$ :

i) Calculate the  $\Delta S_f^\circ$  for  $\text{Fe}_2\text{O}_3$

$$\Delta G = \Delta H - T\Delta S$$

$$-740,000 \text{ J} = -824,000 \text{ J} - (298\text{K})\Delta S$$

$$\Delta S = -282 \text{ J/K}$$

ii) Which is more responsible for the spontaneity of the reaction, the enthalpy or the entropy? Justify your answer.

**Enthalpy ( $\Delta H$ ) tends to be negative for spontaneous reactions**

**Entropy ( $\Delta S$ ) tends to be positive for spontaneous reactions**

**Since the enthalpy is negative and the entropy is negative, the reaction is enthalpy-driven only**

e. Knowing that:  $2 \text{FeO} (s) + 1/2 \text{O}_2 (g) \rightarrow \text{Fe}_2\text{O}_3 (s)$   $\Delta H_{\text{rxn}} = -280 \text{ kJ}$

Calculate  $\Delta H_f^\circ$  of  $\text{FeO} (s)$

$$\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ \text{ prod} - \sum \Delta H_f^\circ \text{ react}$$

$$\Delta H_{\text{rxn}} = (\text{Fe}_2\text{O}_3) - 2*(\text{FeO})$$

$$-280 = (-824) - 2X$$

$$2X = -544$$

$$X = -272 \text{ kJ/mole}$$

4. Consider pentane ( $C_5H_{12}$  MW = 72.15 g/mole)

a. Write the balanced equation for the combustion of pentane to yield  $CO_2$  and  $H_2O$



b. What volume of  $CO_2$  at 25 °C and 785 mm Hg will result from combustion of 2.5 g of pentane?

$$2.5 \text{ g} * (1 \text{ mole} / 72.15 \text{ g}) = 0.03465 \text{ moles } C_5H_{12} * (5 \text{ mole } CO_2 / 1 \text{ mole } C_5H_{12}) = 0.173 \text{ mole } CO_2$$

$$V = nRT/P = (0.173 \text{ mole } CO_2)(0.0821)(298 \text{ K}) / (1.033 \text{ atm}) = 4.103 \text{ L}$$

c. Combustion of 5 g of pentane releases 243 kJ of heat. Calculate  $\Delta H_{\text{combustion}}$  pentane.

$$\Delta H = q/\text{mole}$$

$$5 \text{ g } C_5H_{12} * (1 \text{ mole} / 72.15 \text{ g}) = 0.0693 \text{ moles}$$

$$\Delta H = (-243 \text{ kJ} / 0.0693 \text{ mole}) = -3506 \text{ kJ/mole}$$

d. Under identical conditions, an unknown gas effuses twice as fast as pentane. Calculate the molar mass of this gas.

$$\text{Rate X} / \text{Rate } C_5H_{12} = (72.15/X)^{1/2}$$

$$2/1 = (72.15/X)^{1/2}$$

$$X = 18 \text{ g/mole}$$

e. Draw three structural isomers of pentane.

