**Some Background**

1. What is the only equation that we have been using that is reactants – products?

   **Bond Enthalpy**
   \[ \Delta G = -RT \ln K \]

   \[ \Delta G = -RT \ln K \]

2. List all of the equations that we have been utilizing throughout the unit on thermodynamics.

   \[ \Delta G = \Delta H - T \Delta S \]

3. Briefly describe the relationship between \( \Delta G \) and \( K \).

   \[ K > 1 \text{ if } \Delta G < 0 \]
   \[ K < 1 \text{ if } \Delta G > 0 \]

4. Under what conditions of \( \Delta H \) and \( \Delta S \) will a chemical reaction be thermodynamically favorable under all temperatures?

   - **Spontaneous**
     \[ \Delta H(-) \quad \Delta S(+) \]
   - **Nonspontaneous**
     \[ \Delta H(+) \quad \Delta S(-) \]

5. Under what conditions of \( \Delta H \) and \( \Delta S \) will a chemical reaction be thermodynamically unfavorable under all temperatures?

6. Under what conditions of \( \Delta H \) and \( \Delta S \) will a chemical reaction be thermodynamically favorable at high temperatures and thermodynamically unfavorable at low temperatures? Is this type of reaction enthalpy or entropy driven?

   \[ \Delta H(+) \quad \Delta S(+) \]

7. Under what conditions of \( \Delta H \) and \( \Delta S \) will a chemical reaction be thermodynamically favorable at high temperatures and thermodynamically unfavorable at low temperatures? Is this type of reaction enthalpy or entropy driven?

   \[ \Delta H(-) \quad \Delta S(-) \]

8. \( q = \Delta H \) under what conditions?

   \[ \text{Constant pressure} \]

**J9.42**

Predict the sign of \( \Delta S \) for each of the following processes:

- (a) molten gold solidifies
- (b) gaseous \( \text{Cl}_2 \) dissociates in the stratosphere to form gaseous \( \text{Cl} \) atoms
- (c) Gaseous CO reacts with gaseous \( \text{H}_2 \) to form liquid methanol
- (d) calcium phosphate precipitates upon mixing \( \text{Ca(NO}_3)_2 \) and \( \text{(NH}_4)_3\text{PO}_4 \)

**J9.56**

A certain reaction has \( \Delta H^\circ = 237 \text{ kJ} \) and \( \Delta S^\circ = 52.3 \text{ J/K} \)

- (a) Is the reaction exothermic or endothermic?
- (b) Does the reaction lead to an increase or decrease in the randomness of disorder of the system? Increase in disorder b/c \( \Delta S \) has a (+) number
- (c) Calculate \( \Delta G^\circ \) for the reaction at 298 K

   \[ \Delta G = \Delta H - T \Delta S \]

   \[ \Delta G = 8.1146 \text{ kJ} \]

- (d) Is the reaction spontaneous at 298 K under standard conditions? No b/c \( \Delta G \) is (+)

**J9.65**

A particular constant-pressure reaction is barely spontaneous at 890 K. The enthalpy change for the reaction is 23.7 kJ. Estimate \( \Delta S \) for the reaction.

\[ \Delta H \quad \Delta S \text{ should be (+)} \]

\[ \uparrow \text{ temp} \]
Calculate the value of the equilibrium constant $K$ for the following reaction utilizing the given information.

$$\Delta G^\circ = 0 \text{ kJ/mol} \quad \Delta G^\circ \text{ H}_2 \text{(g)} = 0 \text{ kJ/mole} \quad \Delta G^\circ \text{ H}_2 \text{O} \text{(g)} = 1.3 \text{ kJ/mol}$$

$$\text{H}_2 \text{(g)} + \text{I}_2 \text{(g)} \rightarrow 2 \text{ HI} \text{(g)}$$

$$\Delta G = \text{products} - \text{reactants} = \Delta G^\circ$$

$$\Delta G = (2 \times 2.6 \text{ kJ}) - (3 \times 0 \text{ kJ}) = 2600 = - \frac{8.3145 \times (288)(\text{ln}K)}{2600 \text{ J}}$$

$$\text{ln}K = 0.35$$

$K = 0.35$

For each of the following processes, indicate whether the signs of $\Delta H$ and $\Delta S$ are expected to be positive, negative, or about zero.

a) A solid sublimes $S \rightarrow l$ put energy in, more disorder $\Delta H(+) \Delta S(\pm)$

b) The temperature of a sample of Co (s) is lowered from 60 °C to 25 °C $\Delta H(-) \Delta S(-)$

c) Ethyl alcohol evaporates from a beaker $l \rightarrow g$ $\Delta H(+) \Delta S(\pm)$

d) A diatomic molecule dissociates into atoms $\text{Cl}_2 \rightarrow 2 \text{ Cl}$ $\Delta H(+) \Delta S(\pm)$

e) A piece of charcoal is combusted to form $\text{CO}_2 \text{(g)}$ and $\text{H}_2\text{O}_\text{(g)}$ $\Delta H(-) \Delta S(\pm)$

e) All combustions are exothermic

$$2 \text{ Mg (s)} + \text{O}_2 \text{(g)} \rightarrow 2 \text{ MgO (s)} \quad \Delta H = -1204 \text{ kJ}$$

a) Is this reaction exothermic or endothermic?

Exothermic $\Delta H(-)$

b) Calculate the amount of heat transferred when 3.55 grams of Mg (s) reacts at constant pressure.

$$3.55 \text{ g} \times \frac{1 \text{ mole}}{24.31 \text{ g}} \times \frac{-1204 \text{ kJ}}{1 \text{ mole}} = -149.03 \text{ kJ}$$

d) How many kilojoules of heat are absorbed when 40.3 g of MgO (s) is decomposed into Mg (s) and O$_2$ (g) at constant pressure?

$$2 \text{ mole} \times \frac{140.31 \text{ g}}{1 \text{ mole}} = \frac{280.62 \text{ kJ}}{1 \text{ mole}}$$

Consider the decomposition of liquid benzene, $\text{C}_6\text{H}_6$ to gaseous acetylene, $\text{C}_2\text{H}_2$.

$$\text{C}_6\text{H}_6 \rightarrow 3 \text{ C}_2\text{H}_2 \quad \Delta H = 630 \text{ kJ}$$

a) What is the enthalpy change for the reverse reaction?

$$\frac{-630 \text{ kJ}}{3} = -210 \text{ kJ}$$

b) What is the $\Delta H$ for the formation of 1 mole of acetylene?

c) Which is more likely to be thermodynamically favored, the forward reaction or the reverse reaction? Reverse is the way that provides a $(-) \Delta H$

d) If $\text{C}_6\text{H}_6 \text{(g)}$ were consumed instead of $\text{C}_6\text{H}_6 \text{(l)}$, would you expect the magnitude of $\Delta H$ to increase, decrease, or stay the same? Explain. Decrease, more energy is needed for a gas :: more heat
Without doing any calculations, predict the sign of $\Delta H$ for each of the following reactions:

a) $\text{NaCl}(s) \rightarrow \text{Na}^+(g) + \text{Cl}^-(g)$ $\Delta H(+) \quad$ breaking bonds
b) $2 \text{H}_2(g) \rightarrow \text{H}_2(g)$ $\text{H} \quad \text{H-H} \quad \Delta H(-) \quad$ making bonds
c) $\text{Na}(g) \rightarrow \text{Na}^+(g) + e^- \quad \Delta H(+) \quad$ put energy into system to pop off $e^-$
d) $\text{I}_2(g) \rightarrow \text{I}_2(l) \quad \Delta H(+) \quad$ put energy into system

A solid sample of $\text{Zn(OH)}_2$ is added to 0.350 L of 0.500 M aqueous $\text{HBr}$. The solution that remains is still acidic. It is then titrated with 0.500 M $\text{NaOH}$ solution, and it takes 88.5 mL of the $\text{NaOH}$ solution to reach the equivalence point. What mass of $\text{Zn(OH)}_2$ was added to the $\text{HBr}$ solution?

\[
\text{base} \quad \text{acid} \\
\text{Zn(OH)}_2 + \text{HBr} \quad \text{NaOH} \\
x \quad 0.350 \text{ L} \quad 0.500 \text{ m} \quad \text{mL} \\
0.0175 \text{ moles} \quad 0.0885 \text{ L} \quad 0.04425 \text{ moles} \\
\]

\[
0.0175 \text{ moles} - 0.04425 \text{ moles} = 0.013075 \text{ moles of } \text{HBr reacted with } \text{NaOH} \\
\]

\[
\begin{align*}
\text{Zn(OH)}_2 + 2\text{HBr} & \rightarrow \text{ZnBr}_2 + 2\text{H}_2\text{O} \\
0.065375 \text{ mole} & \quad 0.13075 \\
\text{mole} & \\
\text{mole} & \\
0.065375 \text{ mole} & \quad 99.4067 \text{ g} \\
1 \text{ mole} & \quad 6.49867 \text{ g} \\
\end{align*}
\]

\[
6.50 \text{ g} \\
\]