Chapter 9 - Lecture Worksheet 3

1. For the following reaction ΔH^0 = -197.8 kJ/mol and ΔS^0 = -187.9 J/mol K. Assuming that ΔH^0 and ΔS^0 are independent of temperature, calculate the equilibrium constant for this reaction at 262 K.

   $2SO_2(g) + O_2(g)$ <------> $2SO_3(g)$

ΔH^0 = –197.8 kJ/mol
ΔS^0 = –187.9 J/mol K
T = 262 K
K = ?

$$\Delta G^0 = \Delta H^0 - T\Delta S^0 = -RT\ln K$$
$$= -197.8 \text{ kJ/mol} - (262\text{ K})(-187.9 \text{ J/mol K})(1 \text{ kJ}/1000 \text{ J}) = -148.57 \text{ kJ/mol}$$

$$K = e^{\frac{-\Delta G^0}{RT}} = e^{\frac{-(-148.57 kJ/mol)}{(8.314 J/molK)(262 K)}\left(\frac{1000 J}{kJ}\right)} = e^{+68.21} = 4.2 \times 10^{29}$$

2A. Carbonyl bromide, $COBr_2$, decomposes to CO and Br_2 at 73^0C. If you begin with 0.10 moles of $COBr_2$ in a 2.5 Liter flask and find that there are 0.015 moles of $COBr_2$ at equilibrium, what are the concentrations of CO and Br_2 at equilibrium ? **You must use the ICE method to solve this problem.**

	$COBr_2(g)$	<--->	CO(g) +	$Br_2(g)$
Initial	(0.10 mol/2.5 L) = 0.040 M		0	0
Change	–x		+x	+x
Equilibrium	0.040 – x = (0.015 mol / 2.5 L)		x	x

$$(0.015 \text{ mol} / 2.5 \text{ L}) = 0.0060 \text{ M} = 0.040 - x$$
$$= 0.0060 \text{ M} = [COBr_2(g)]_{eq}$$
$$x = 0.040 - 0.0060 \text{ M} = 0.034 \text{ M} = [CO(g)]_{eq} = [Br_2(g)]_{eq}$$

5. What is the value of the equilibrium constant K_c at 73^0C ?

$$Kc = \frac{[CO(g)][Br_2(g)]}{[COBr_2]} = \frac{(0.034)(0.034)}{(0.0060)} = 0.19$$

Since Kc is between 10^{-3} and 10^3, will get significant concs of both reactants and products at equilibrium.

6. This reaction is **1.** REACTANT FAVORED **2.** PRODUCT FAVORED at equilibrium.

Question 4	Question 5
1. 0.0030 M	**1.** 0.18
2. 0.0060 M	**2. 0.19**
3. 0.040 M	**3.** 5.2
4. 0.085 M	**4.** 5.7
5. 0.034 M	
6. 0.068 M	

D. You would expect ΔG^0 to be

1. GREATER THAN ZERO **2.** LESS THAN ZERO **3.** EQUAL TO ZERO

E. This is because K < 1

F. Calculate ΔG^0 for this reaction.

$$\Delta G^0 = -RT \ln K = -(8.314\ \text{J/molK})(262\text{K}) \ln(0.19) = +3{,}618\ \text{J/mol} = 3.6\ \text{kJ/mol}$$

G. Do you think ΔG^0 will change sign at high temperature vs low temperature ? **1. Yes** **2. No**
Explain.

$$\Delta G^0 = \Delta H^0 - T\Delta S^0$$

For this reaction $\Delta S^0 > 0$ because there are more moles of product gas than reactant ($\Delta n_{gas} = 2 - 1 = 1$)
Therefore as T increases expect that ΔG^0 will become negative.

H. Sketch a Gibbs Free Energy diagram *v.s.* Extent of reaction diagram for this reaction. *(Label everything!)* Show the region where Q < K, Q = K and Q > K.