AP Chemistry Problem Set #13

1. (**1998 - #7**)

$$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g) \qquad \Delta H° = +131 \text{ kJ}$$

A rigid container holds a mixture of graphite pellets (C(s)), $H_2O(g)$, CO(g), and $H_2(g)$ at equilibrium. State whether the number of moles of CO(g) in the container will increase, decrease, or remain the same after each of the following disturbances is applied to the original mixture. For each case, assume that all other variables remain constant except for the given disturbance. Explain each answer with a short statement.

a. Additional $H_2(g)$ is added to the equilibrium mixture at constant volume. **Decrease. Additional H_2 causes the reaction to shift left.**

b. The temperature of the equilibrium mixture is increased at constant volume. **Increase. The reaction is endothermic, additional heat shifts the reaction to the right.**

c. The volume of the container is decreased at constant temperature. **Decrease. Decreasing the volume increases the pressure and shifts the reaction left to where there are fewer moles of gas.**

d. The graphite pellets are pulverized. **Remain the same. Pulverizing carbon does not affect the equilibrium position.**

e. Hydrogen gas (H_2) is leaked out of the container. **Increase. The reaction will shift right to replace the lost H_2.**

2. Refer to the system $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$. To an empty 15.0 L cylinder, 0.500 moles of gaseous PCl_5 are added and allowed to reach equilibrium. The concentration of PCl_3 is found to be 0.0220M. Assume a temperature of 375 K.

a. How many moles of PCl_5 remain at equilibrium? **0.165 moles**

b. Write the equilibrium constant expression for the above reaction. $\dfrac{[PCl_5]}{[PCl_3][Cl_2]}$

c. Determine the value of K_c. **K_c = 23**

d. Determine the value of K_p for this same system at the same temperature. **K_p = 0.75**

e. How would the value of K_p be effected by increasing the temperature of the system at equilibrium for this exothermic reaction? **K_p would decrease. Increasing the temperature of an exothermic reaction would shift the reaction to the left, increasing the denominator and decreasing the numerator.**

3. Consider the reaction $2HI(g) \rightleftharpoons H_2(g) + I_2(s)$ which is at equilibrium.

a. How will adding more hydrogen gas at constant volume affect the equilibrium? **Additional H_2 would cause the reaction to shift left to achieve equilibrium.**

b. Some of the iodine is removed. How will this affect the equilibrium? Explain. **No effect. Iodine is a solid and does not affect equilibrium.**

c. The pressure is increased by pumping in pure neon gas. How will this affect the equilibrium? Explain. **Neon is inert and does not react. The addition of an inert gas increases the total pressure but has no effect on the concentrations or partial pressures of the reactants.**

d. The pressure is increased by decreasing the volume. How will this affect the equilibrium? Explain. **A decrease in volume increases the pressure and shifts the equilibrium to where there are fewer moles of a gas. This reaction would shift right.**

e. A suitable catalyst is added. How will this affect the equilibrium? **No effect. Catalysts affect the rate of the reaction (Kinetics) but not the equilibrium position.**

4. (**2004B - #1)**

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

For the reaction represented above, the value of the equilibrium constant, K_p, is 3.1 x 10^{-4} at 700. K.

a. Write the expression for the equilibrium constant, K_p, for the reaction. $\dfrac{(NH_3)^2}{(N_2)(H_2)^3}$

b. Assume that the initial partial pressures of the gases are as follows:
P_{N2} = 0.411 atm, P_{H2} = 0.903 atm, and P_{NH3} = 0.224 atm.

(i) Calculate the value of the reaction quotient, Q, at these initial conditions. **Q = 0.166**

(ii) Predict the direction in which the reaction will proceed at 700. K if the initial partial pressures are those given above. Justify your answer. **The reaction will shift left to reach equilibrium because Q > K.**

c. Calculate the value of the equilibrium constant, K_c, given that the value of K_p for the reaction at 700. K is 3.1 x 10^{-4}. **K_c = 1.0**

d. The value of K_p for the reaction represented below is 8.3 x 10^{-3} at 700. K.

$$NH_3(g) + H_2S(g) \rightleftharpoons NH_4HS(g)$$

Calculate the value of K_p at 700. K for each of the reactions represented below.

(i) $NH_4HS(g) \rightleftharpoons NH_3(g) + H_2S(g)$ **K_p = 120**

(ii) $2\ H_2S(g) + N_2(g) + 3H_2(g) \rightleftharpoons 2NH_4HS(g)$ **K_p = 2.1 x 10^{-8}**

5. Hydrogen bromide decomposes according to the equation: $2HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$

a. Write the equilibrium law in terms of concentrations, K_c, and partial pressures, K_p.

$\dfrac{[H_2][Br_2]}{[HBr]^2}$ & $\dfrac{(H_2)(Br_2)}{(HBr)^2}$

b. A 55.5 gram sample of HBr (molar mass = 80.9) is transferred to an evacuated 22.0 L flask at 142°C.

(i) What is the initial concentration in moles per liter of HBr? **0.0312 M**

(ii) What is the initial pressure of HBr in mm Hg? **806 mm Hg (1.06 atm)**

c. When the system comes to equilibrium, 22.4 g of Br_2 is found to be in the flask. Calculate the value of K_c and K_p. **K_c = K_p = 0.119**

d. At another temperature, K_c = 17.7. 1.37 mol HBr(g), 2.31 mol $H_2(g)$ and 0.551 mol Br_2 are introduced into an evacuated 15.0 L flask. Determine whether the reaction proceeds in forward or reverse direction. **Q < K so the reaction proceeds in the forward direction (right).**

e. Calculate the K_p and K_c for the following reaction: $\frac{1}{2} H_2(g) + \frac{1}{2} Br_2(g) \rightleftharpoons HBr(g)$ **K_c = K_p = 2.90**