REVIEW QUESTIONS
Chapter 16

1. Write equilibrium constant expressions for each reaction below:

A) $N_2\ (g) + 2\ H_2O\ (g) \rightleftharpoons NO\ (g) + 2\ H_2\ (g)$

$$K_{eq} = \frac{[NO][H_2]^2}{[N_2][H_2O]^2}$$

B) $2\ N_2O_5\ (g) \rightleftharpoons 2\ N_2O_4\ (g) + O_2\ (g)$

$$K_{eq} = \frac{[N_2O_4]^2[O_2]}{[N_2O_5]^2}$$

2. At 1500 K the following equilibrium is established:

$$2\ BrF_5\ (g) \rightleftharpoons Br_2\ (g) + 5\ F_2\ (g)$$

At equilibrium,
$[BrF_5] = 0.0064$ M
$[Br_2] = 0.0018$ M
$[F_2] = 0.0090$ M.

Calculate the equilibrium constant for this reaction.

$$K_{eq} = \frac{[Br_2][F_2]^5}{[BrF_5]^2} = \frac{(0.0018)(0.0090)^5}{(0.0064)^2} = 2.6 \times 10^{-9}$$

3. For each reaction below, determine whether reactants or products are favored at equilibrium:

A) $N_2\ (g) + O_2\ (g) \rightleftharpoons 2\ NO\ (g)$ $K_{eq} = 4.5 \times 10^{-31}$

Equilibrium constant is small, therefore reactants are favored at equilibrium

B) $N_2\ (g) + 3\ H_2\ (g) \rightleftharpoons 2\ NH_3\ (g)$ $K_{eq} = 3.6 \times 10^{8}$

Equilibrium constant is large, therefore products are favored at equilibrium

C) $CH_4\ (g) + Cl_2\ (g) \rightleftharpoons CH_3Cl\ (g) + HCl\ (g)$ $K_{eq} = 1.2 \times 10^{18}$

Equilibrium constant is large, therefore products are favored at equilibrium

4. In the equilibrium shown below, identify four changes that would increase the concentration of ammonia equilibrium:

$$\text{heat} + N_2\,(g) + 3\,H_2\,(g) \rightleftharpoons 2\,NH_3\,(g)$$

   1. **Add N_2 (equilibrium would shift forward and produce more NH_3)**

   2. **Add H_2 (equilibrium would shift forward and produce more NH_3)**

   3. **Increase temperature (equilibrium would shift forward and produce more NH_3)**

   4. **Decrease volume (pressure would increase and equilibrium would shift forward and produce m**

5. For the equilibrium shown below, describe the changes that occur after each action listed:

$$A\,(g) + 3\,B\,(g) \rightleftharpoons 2\,C\,(g) + 3\,D\,(g) + \text{heat}$$

   a) Temperature is increased

   Equilibrium would shift reverse; A and B would increase; C and D would decrease.

   b) More A is added

   Equilibrium would shift forward; B would decrease; C and D would increase.

   c) Some D is removed

   Equilibrium would shift forward; A and B would decrease; C would increase

   d) Volume is decreased

   Equilibrium would shift reverse; A and B would increase; C and D would decrease

6. 0.100 moles of HI was placed in a 1.0-L flask and allowed to decompose as shown below:

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

At equilibrium, the 0.011 moles of H_2 was present. Calculate the equilibrium constant for this reaction. (Hint: set up a reaction table)

	$2\,\text{HI (g)} \rightleftharpoons$	**$\text{H}_2\text{ (g)}$ +**	**$\text{I}_2\text{ (g)}$**
Initial	**0.100**	**0**	**0**
Change (Δ)	**–0.022**	**+0.011**	**+0.011**
Equilibrium	**0.078**	**0.011**	**0.011**

$$K_{eq} = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.011\text{ M})(0.011\text{ M})}{(0.078\text{ M})^2} = 2.0\text{x}10^{-2}\text{ M}$$

7. A 0.15 M solution of butanoic acid ($HC_4H_7O_2$) has $[H_3O^+] = 1.51 \text{ x } 10^{-3}$ M. Calculate K_a for this acid.

$$HC_4H_7O_2 \rightleftharpoons H^+ + C_4H_7O_2^-$$

$$[H^+] = [C_4H_7O_2^-] = 1.51\text{x}10^{-3}\text{ M}$$

$$K_a = \frac{[H^+][C_4H_7O_2^-]}{[HC_4H_7O_2]} = \frac{(1.51\text{x}10^{-3})^2}{0.15} = 1.5\text{x}10^{-5}$$

8. A 0.035 M solution of a weak acid (HA) has a pH of 4.88. Calculate the K_a for this acid.

$$HA \rightleftharpoons H^+ + A^-$$

$$[H^+] = [A^-] = 10^{-4.88} = 1.\underline{3}2\text{x}10^{-5}\text{ M}$$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(1.\underline{3}2\text{x}10^{-5})^2}{0.035} = 5.0\text{x}10^{-9}$$

9. In a reaction, 0.15 moles of NH_3 is placed in a 1.5-L flask and allowed to decompose as shown below:

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

At equilibrium, 0.035 moles of N_2 were present. Determine the percent yield of this reaction.

$$[NH_3] = \frac{0.15\text{ mol}}{1.5\text{ L}} = 0.10\text{ M} \qquad [N_2] = \frac{0.035\text{ mol}}{1.5\text{ L}} =$$

	$2\,NH_3\,(g) \rightleftharpoons$	$N_2\,(g)$ +	$3\,H_2\,(g)$
Initial	0.10	0	0
Change (Δ)	-------	+0.023	--------
Equilibrium	--------	0.023	--------
No Equil.	0	0.050	--------

$$\%\text{ Yield} = \frac{0.023}{0.050} \times 100 = 46\%$$

10. Determine the pH and percent ionization of 1.0 M acetic acid (K_a=1.8 x 10^{-5}).

$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$

$$[H^+] = [C_2H_3O_2^-] = x$$

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{x^2}{1.0} = 1.8x10^{-5}$$

$$x = \sqrt{1.8x10^{-5}} = 4.\underline{2}4x10^{-3}\text{ M}$$

$$pH = -\log(4.\underline{2}4x10^{-3}) = 2.37$$

$$\%\text{ Ionization} = \frac{[H^+]}{[HA]} x\,100 = \frac{4.\underline{2}4x10^{-3}}{1.0} x\,100 = 0.42\%$$

11. What is the pH of a buffer solution consisting of 0.55 M acetic acid and 0.68 M sodium acetate? (K_a for acetic acid = 1.8×10^{-5})

$$\mathbf{HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-}$$

0.55 M **0.68 M**

$$\mathbf{[H^+] = K_a \text{ x } \frac{[HA]}{[A^-]} = 1.8x10^{-5} \text{ x } \frac{0.55}{0.68} = 1.\underline{4}6x10^{-5}}$$

$$\mathbf{pH = -log\,(1.\underline{4}6x10^{-5}) = 4.84}$$

12. The K_{sp} value for AgI is 8.5×10^{-17}. What is the molar solubility of AgI?

$$\mathbf{AgI\,(s) \rightleftharpoons Ag^+ + I^-}$$

$$\mathbf{K_{sp} = [Ag^+][I^-] = 8.5x10^{-17}}$$

$$\mathbf{[Ag^+] = \sqrt{8.5x10^{-17}} = 9.2x10^{-9}\ M}$$

13. The solubility of CaF_2 is 2.14×10^{-4} M. Calculate the K_{sp} for CaF_2.

$$\mathbf{CaF_2\,(s) \rightleftharpoons Ca^{2+} + 2\,F^-}$$

$$\mathbf{[Ca^{2+}] = 2.14x10^{-4}}$$

$$\mathbf{[F^-] = 2 \text{ x } [Ca^{2+}] = 4.28x10^{-4}\ M}$$

$$\mathbf{K_{sp} = [Ca^{2+}][F^-]^2 = (2.14x10^{-4})(4.28x10^{-4})^2 = 3.92x10^{-11}}$$