Solutions to the Extra Problems for Chapter 16

1. The equation is $\underline{CS_2 + 4H_2 \rightleftarrows 2H_2S + CH_4}$. To solve this problem, we just have to think backwards from Equation (16.1). The exponents are the coefficients in the chemical equation, and the products are in the numerator of the equation for K, while the reactants are in the denominator.

2. The concentration is 1.07 M. For this one, we don't have all the equilibrium concentrations, but we do have K. We can use that to calculate the one equilibrium concentration we are missing.

$$K = \frac{[CH_4]_{eq}[H_2O]_{eq}}{[CO]_{eq}[H_2]_{eq}^3}$$

Plugging in the numbers we know:

$$4.18\,\frac{1}{M^2} = \frac{[0.210\text{ M}][H_2O]_{eq}}{[0.212]\ [0.633\text{ M}]^3}$$

Solving for $[H_2O]_{eq}$:

$$[H_2O]_{eq} = \frac{4.18\,\frac{1}{\cancel{M^2}}[0.212\text{ M}][0.633\,\cancel{M}]^3}{[0.210\,\cancel{M}]} = 1.07\text{ M}$$

3. a. This K is really, really small. Thus, it's just the reverse reaction that occurs:

$$AgBr\ (s) \leftarrow Ag^+\ (aq) + Br^-\ (aq)$$

This isn't standard, since the products are supposed to be on the right. Thus, the one-way reaction is:

$$\underline{Ag^+\ (aq) + Br^-\ (aq) \rightarrow AgBr\ (s)}$$

b. The K is larger than 1, but it is not really large. Thus, this is an equilibrium that is weighted toward the products.

c. This K is huge. Thus, you can assume only the forward reaction happens:

$$\underline{CO\ (g) + Cl_2\ (g) \rightarrow COCl_2\ (g)}$$

d. The K is smaller than 1, but it is not really small. Thus, this is an equilibrium that is weighted toward the reactants.

4. a. We ignore solids and liquids in Equation (16.1), so the AgBr doesn't appear in the equation:

$$\underline{K = [Ag^+][Br^-]}$$

b. We ignore solids and liquids, so $Fe(OH)_3$, and H_2O don't appear in the equation:

$$K = \frac{[Fe_2(SO_4)_3]_{eq}}{[H_2SO_4]_{eq}^3}$$

5. a. <u>The reaction must shift towards the products</u>. We have to see what Equation (16.1) is with these concentrations and then compare it to K. For this equation, Equation (16.1) is:

$$K = \frac{[CH_4]_{eq}[H_2S]_{eq}^2}{[CS_2]_{eq}[H_2]_{eq}^4}$$

Plugging in the concentrations:

$$\frac{[0.050\ M][0.050\ M]^2}{[0.100\ M][0.100\ M]^4} = 13\ \frac{\cancel{M^3}}{M^{\cancel{5}\,2}} = 13\ \frac{1}{M^2}$$

This is smaller than the equilibrium constant, so the equilibrium must shift to make Equation (16.1) larger. This will happen if reactants are used up and products are made, so the reaction must shift towards the products.

b. <u>This reaction is at equilibrium</u>. Putting the numbers into Equation (16.1):

$$\frac{[2.2\ M][4.5\ M]^2}{[1.1\ M][1.1\ M]^4} = 28\ \frac{\cancel{M^3}}{M^{\cancel{5}\,2}} = 28\ \frac{1}{M^2}$$

Since this is the value of the equilibrium constant (to two significant figures), the reaction is at equilibrium.

c. <u>The reaction must shift towards the reactants</u>. Putting the numbers into Equation (16.1):

$$\frac{[7.1\ M][7.1\ M]^2}{[0.90\ M][0.90\ M]^4} = 610\ \frac{\cancel{M^3}}{M^{\cancel{5}\,2}} = 610\ \frac{1}{M^2}$$

This is larger than the equilibrium constant, so the equilibrium must shift to make Equation (16.1) smaller. This will happen if reactants are made and products are used up, so the reaction must shift towards the reactants.

6. a. <u>The concentration will decrease</u>. This is an exothermic reaction, so energy is a product:

$$2Fe(OH)_3\ (s)\ +\ 3H_2SO_4\ (aq)\ \rightleftarrows\ Fe_2(SO_4)_3\ (aq)\ +\ 6H_2O\ (l)\ +\ 115\ kJ$$

Increasing temperature is like adding a product, which increases the rate of the reverse reaction. Thus, the equilibrium will shift towards the reactants. That will use up $Fe_2(SO_4)_3$.

b. <u>The concentration will increase</u>. If $Fe_2(SO_4)_3$ is added, that will increase the rate of the reverse reaction. Thus, the equilibrium will shift towards reactants. That will make more H_2SO_4.

c. Nothing will happen, because we ignore solids in the equilibrium constant.

d. The concentration will decrease. If H_2SO_4 is removed, that will slow down the forward reaction. The reverse reaction will therefore be faster, and the equilibrium will shift towards the reactants. That will use up $Fe_2(SO_4)_3$.

7. a. The concentration will decrease. Increasing pressure causes the reaction to shift to the side that has fewer gas molecules. The reactants side has five gas molecules (one CS_2 and four H_2), while the products side has three (one CH_4 and two H_2S). The reaction, therefore, will shift towards products, which will use up H_2.

b. The concentration will decrease. If pressure is decreased, the reaction shifts to the side with more gas molecules, which is the reactants side. That will use up H_2S.

8. The acid ionization constant is the equilibrium constant for the acid's reaction with water. Acids donate an H^+, so the reaction is:

$$HCHO_2\ (aq) + H_2O\ (l) \rightleftarrows CHO_2^-\ (aq) + H_3O^+\ (aq)$$

Using Equation (16.1), the constant is:

$$K_a = \frac{[CHO_2^-]_{eq}[H_3O^+]_{eq}}{[HCHO_2]_{eq}}$$

9. The base ionization constant is the equilibrium constant for the base's reaction with water. Bases accept an H^+, so the reaction is:

$$C_2H_7N\ (aq) + H_2O\ (l) \rightleftarrows C_2H_8N^+\ (aq) + OH^-\ (aq)$$

Using Equation (16.1), the constant is:

$$K_b = \frac{[C_2H_8N^+]_{eq}[OH^-]_{eq}}{[C_2H_7N]_{eq}}$$

10. The higher pH will be made with nitrous acid. The stronger the acid, the lower the pH. Thus, the weaker acid will have the higher pH. Nitrous acid has a lower acid ionization constant and is therefore the weaker acid.