Equilibrium Practice Problems

1. Many small, highly charged cations are slightly acidic when dissolved in water. For example, a 0.10 M solution of $AlCl_3$ has a pH of 2.9. The reaction responsible for the acidity is

$$Al(H_2O)_6^{3+}(aq) + H_2O(l) \leftrightharpoons H_3O^+(aq) + Al(OH)(H_2O)_5^{2+}(aq)$$

   where $Al(H_2O)_6^{3+}$ is a metal-ligand complex of Al^{3+} and H_2O. Determine the K_a value for $Al(H_2O)_6^{3+}$.

2. Quinine, $C_{20}H_{24}N_2O_2$, is an alkaloid derived from a tree that grows in tropical rain forests. In addition to being an ingredient in tonic water, quinine also is used in the treatment of malaria. Like all alkaloids, quinine is a sparingly soluble weak base; its pK_b is 5.1 and its solubility is 1.00 g in 190 L of water. What is the pH of a saturated solution of quinine?

3. Magnesium hydroxide, $Mg(OH)_2$, is a slightly soluble ionic compound with a K_{sp} of 1.8×10^{-11}. It is often used in antacid products, such as milk of magnesia, to reduce stomach acidity. It is also an important industrial source of magnesium. (a) What is the molar solubility of $Mg(OH)_2$ in pure water? (b) Sea water is an important industrial source of magnesium with typical concentrations of 5.4×10^{-2} M. What is the minimum pH at which $Mg(OH)_2$ begins to precipitate from sea water?

4. An important industrial reaction is the Haber process in which N_2 and H_2 react to make ammonia. What is frustrating about this reaction is that at low temperatures, where the reaction's equilibrium constant is very large, the reaction's rate is too slow to be of practical use. At higher temperatures the reaction is faster, but the equilibrium constant is less favorable. (a) Based on the above statement, do you expect that the Haber process is exothermic or endothermic? Briefly explain and then check your answer by calculating ΔH. (b) An inventor claims to have discovered a new catalyst that increases the percentage of nitrogen and hydrogen that is converted to ammonia by the Haber process without needing to change the temperature. A member of your family is eager to invest in the inventor's discovery. Explain to your relative why this is a bad idea. (c) To maximize the yield of ammonia, would it be better to run this reaction in a small volume container or a large volume container? Explain.

5. Virtually all investigations in cell biology and biochemistry are carried out in buffered aqueous solutions. Imagine you are studying an enzyme that is active only between a pH of 7.1 and 7.4 and that you need to prepare 1.50 L of a phosphate buffer at a pH of 7.25, limiting the total concentration of phosphate species to 0.085 M. The following compounds are available to you: 14.75 M H_3PO_4, KH_2PO_4 and K_2HPO_4. Which two phosphate species will you use? What quantities of each will you use? How many moles of H_3O^+ and of OH^- can this buffer absorb without the pH falling below 7.1 or above 7.4?

6. The solubility of the mineral Ag_2CrO_4 depends upon pH because chromate, CrO_4^{2-} is a weak base. For this problem limit yourself to the following equilibrium reactions:

$$Ag_2CrO_4(s) \leftrightharpoons 2Ag^+(aq) + CrO_4^{2-}(aq) \qquad K_{sp} = 2.57\times10^{-12}$$

$$H_2CrO_4(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HCrO_4^-(aq) \qquad K_{a1} = 1.6$$

$$HCrO_4^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CrO_4^{2-}(aq) \qquad K_{a2} = 3.1\times10^{-7}$$

Calculate the molar solubility for Ag_2CrO_4 at pH levels buffered to 4.0, 5.0, 8.0, and 9.0. Be sure to clearly indicate the reaction responsible for the molar solubility of Ag_2CrO_4 at each pH level and to clearly show your work.