Problem Set 5 (For February 10)

Recommended Text Problems: 8.46, 8.53, 8.69, 8.71, 8.76, 8.103

Note: This problem set focuses on Chapter 8 material (acid-base and solubility equilibria). Problem Set 6 will be available Thursday, Feb. 10 (along with the solutions to these problems). There will be an x-hour lecture on Feb. 10, but there will be no class or office hour on Friday, Carnival Holiday. We have an exam Tuesday, Feb. 22, and a practice exam will be available soon.

1 . A saturated solution of carbon dioxide in water produces carbonic acid, H_2CO_3, at a concentration of 0.034 M. Calculate the pH of this solution as well as the concentrations of all the species in the carbonic acid dissociation chain, H_2CO_3, HCO_3^-, and CO_3^{2-}. The relevant acid dissociation constants are in Table 7.4 in the text. Suppose the pH of this solution is raised to 10.0. What are the concentrations of H_2CO_3, HCO_3^-, and CO_3^{2-} now?

2 . 20.0 mL of 0.10 M $Hg(NO_3)_2$ solution and 30.0 mL of 0.20 M NaI solution are mixed. Assuming that no reaction takes place, calculate the concentration of each of the ions present in the solution. Actually, a yellow-orange precipitate of mercury(II) iodide ($K_{sp} = 3 \times 10^{-26}$) forms. Calculate the concentrations of Hg^{2+} and I^- once equilibrium has been established. (There is actually more to this story: see Problem 8.89, but ignore the chemistry described in that problem as you work this one!)

3 . Here's a challenging problem on simultaneous solubility equilibria. The oxalate ion, $C_2O_4^{-2}$, forms many insoluble compounds with metal ions. For example, copper oxalate, CuC_2O_4, has $K_{sp} = 3 \times 10^{-8}$, and barium oxalate, BaC_2O_4, has $K_{sp} = 2 \times 10^{-8}$. In a solution saturated with both compounds at the same time, what are $[Cu^{2+}]$, $[Ba^{2+}]$, and $[C_2O_4^{-2}]$? Take care to note that $C_2O_4^{-2}$ comes from two sources, and check your answers by verifying that $[Cu^{2+}][C_2O_4^{-2}] = 3 \times 10^{-8}$ and $[Ba^{2+}][C_2O_4^{-2}] = 2 \times 10^{-8}$.

4 . The solubility product for $Zn(OH)_2$ is 4.5×10^{-17}. Calculate the solubility (in g L^{-1} units) of $Zn(OH)_2$ in a solution buffered at pH = 6.00. Assume the buffer capacity is large enough to maintain the pH at 6.00 as the $Zn(OH)_2$ dissolves.

5 . Solutions of the very soluble salt $Fe(NO_3)_3$, iron(III) nitrate, in pure water are acidic. If one dissolves 0.100 mol of this salt in 1.00 L of water, the solution has a pH of 1.6. This surprising result is explained by the reaction

$$Fe^{3+}(aq) + H_2O(l) \rightleftarrows H^+(aq) + FeOH^{2+}(aq)$$

Calculate the equilibrium constant for this reaction. *Note*: you will find that our usual simplifying assumption that the amount of material that has reacted is a small fraction of the initial amount is not very good here. An exact treatment is called for. What is the ratio $[FeOH^{2+}]/[Fe^{3+}]$?