Take-Home Assignment Four

Please work out and turn in a solution to the following problem. If you make approximations in solving equilibrium problems, be sure to state them and to verify that they are reasonable; if any approximations you make are not reasonable, then continue working the problem until you find a valid solution. You may assume errors of ±5% are acceptable.

You may choose to work with one or two partners (no more than two) and turn in one set of solutions. This assignment is due before you leave for spring break.

1. Suppose I hand you two beakers, one containing a solution of 0.1 M HCl and one containing a solution of 0.1 M CH_3COOH.

   (a) Describe how can determine which solution is which by measuring the pH of each solution.

   The two acids are equal in concentration, so we need consider only their relative strengths. As HCl is a strong acid and CH_3COOH is a weak acid, the solution of HCl will have the a lower pH.

   (b) Describe how you can determine which solution is which by observing the reaction of each solution with solid Zn, which reacts with H_3O^+ to form Zn^{2+} and $H_2(g)$.

   Because the two acids are equal in concentration, each will dissolve the same amount of zinc; however, as HCl is the stronger acid, its solution has a higher concentration of H_3O^+ and will react with zinc more quickly.

2. Quinine, $C_{20}H_{24}N_2O_2$, is a natural product that has potent anti-malarial properties. Although it is now synthesized using simple organic starting materials, it once was extracted from the bark of cinchona trees.

   (a) Quinine is a weak base with a pK_b of 5.1. It is not very soluble and a saturated solution of quinine is 1.6×10^{-3} M. What is the pH of this solution?

   The equilibrium reaction of interest is $Q(aq) + H_2O(l) \leftrightarrows OH^-(aq) + HQ^+(aq)$, where Q is quinine, and where the equilibrium constant is 7.94×10^{-6}. Using an ICE table, we find equilibrium concentrations of x for OH^- and for HQ^+, and of $0.0016 - x$ for Q. Substituting into the K_b expression gives

   $$K_b = [OH^-][HQ^+]/[Q] = (x)(x)/(0.0016 - x) = 7.94 \times 10^{-6}$$

   If we assume $0.0016 - x \approx 0.0016$ and solve for x we find that $x = 8.91 \times 10^{-5}$ with an error of

   $$\% \text{ error} = 100 \times \{0.0016 - (0.0016 - x)\}/0.0016 = 100 \times (8.91 \times 10^{-5})/0.0016 = 5.57\%$$

   This error is too large, so we make a new assumption that $0.0016 - x \approx 0.0016 - 8.91 \times 10^{-5} = 0.001511$. Solving for x we find that $x = 1.095 \times 10^{-4}$. The error in this assumption is

   $$\% \text{ error} = 100 \times \{0.001511 - (0.0016 - 1.095 \times 10^{-4})\}/0.001511 = 1.36\%$$

   This is an acceptable error; thus, the $[OH^-]$ is 1.095×10^{-4}, the pOH is 3.96, and the pH is 10.04.

   (b) Because of its low solubility, quinine is dispensed as a weak acid using its hydrochloride salt, $C_{20}H_{24}N_2O_2$•HCl, which is called quinine hydrochloride. What is the pK_a for quinine's weak acid form, and what is the pH of a 1.5% w/v aqueous solution of quinine hydrochloride? You may assume the solution has a density equivalent to water.

The equilibrium reaction of interest is $HQ^+(aq) + H_2O(l) \leftrightharpoons H_3O^+(aq) + Q(aq)$, where HQ^+ is weak acid form of quinine. A solution that contains 1.5 g quinine hydrochloride per 100 mL has a molarity of HQ^+ of 0.0416 M. Using an ICE table, we find equilibrium concentrations of x for H_3O^+ and for Q, and of $0.0416 - x$ for HQ^+. The pK_a for HQ^+ is $14.0 - 5.1 = 8.9$, which gives K_a for HQ^+ as 1.26×10^{-9}. Substituting into the K_a expression gives

$$K_a = [H_3O^+][Q]/[HQ^+] = (x)(x)/(0.0416 - x) = 1.26\times10^{-9}$$

If we assume $0.0416 - x \approx 0.0416$ and solve for x we find that $x = 7.23\times10^{-6}$ with an error of

$$\%\ \text{error} = 100\times\{0.0416 - (0.0416 - x)\}/0.0416 = 100\times(7.23\times10^{-6})/0.0416 = 0.017\%$$

This is an acceptable error; thus, the $[H_3O^+]$ is 7.23×10^{-6} and the pH is 5.14.