Solutions for Acid-Base Titrations Questions

Check for Understanding 9.1

1. If a weak acid or base is used as a titrant, the titration reaction may not even favor the products, making a quantitative determination of the analyte difficult. Furthermore, the solution will tend to be highly buffered so the pH change around the equivalence point may not be very large. This makes it more difficult to find a suitable end point.

Check for Understanding 9.2

1. Estimate the equivalence point pH by averaging the "plateau" pH values before and after the equivalence point. The NH_4^+/NH_3 buffer pH ~ pK_a = 9.2. The pH of the 0.10 M HCl titrant is about 1. The average of these values is (9.2 + 1)/2 = 5.1, which is very close to the actual pH found at the equivalence point (pH = 5.4).

Check for Understanding 9.3

1. a) The equivalence point pH ~ 8.7 for the $HC_2H_3O_2$/NaOH titration. Use phenolphthalein.

    b) The equivalence point pH ~ 5.4 for the NH_3/HCl titration. Use methyl red.

Check for Understanding 9.4

1. At the second equivalence point, the pH ~ (6.35 + 1.5)/2 = 3.9.

2. The first equivalence point pH ~ 8.3. Use thymol blue or phenolphthalein.

    The second equivalence point pH ~ 3.9. Use methyl orange.

Solutions for Acid-Base Titrations Questions

Check for Understanding 9.5

1.　　titration reaction:　$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2CO_3$

You have an aqueous solution of a weak acid and NaCl. Since $K_{a2} << K_{a1}$ for this diprotic weak acid, you can ignore the effect of the second acid dissociation on pH.

initial mol Na_2CO_3 = 0.200 g Na_2CO_3 (1 mol/106 g Na_2CO_3) = 1.89 x 10^{-3} mol

volume HCl to reach equivalence point:

$$1.89 \text{ } x \text{ } 10^{-3} \text{ } mol \text{ } Na_2CO_3 \text{ } x \text{ } \frac{2 \text{ } mol \text{ } HCl}{1 \text{ } mol \text{ } Na_2CO_3} \text{ } x \text{ } \frac{1000 \text{ } mL}{0.100 \text{ } mol \text{ } HCl} = 37.7 \text{ } mL \text{ } HCl$$

at the second equivalence point,

$[H_2CO_3]_{init}$ = 1.89 x 10^{-3} mol/0.0977 L = 1.93 x 10^{-2} M

the only important source of ions is the strong electrolyte NaCl

$[NaCl] = 2 \text{ } x \text{ } [H_2CO_3]$ = 0.039 M and μ = 0.039 M

Thus,　γ_{HCO3-} = 0.902 -0.059 = 0.843
　　　　　γ_{H3O+} = 0.914 - 0.039 = 0.875
　　　　　γ_{H2CO3} ~1

$$
\begin{array}{lllll}
 & H_2CO_3 & + \text{ } H_2O & \leftrightharpoons \text{ } H_3O^+ & + \text{ } HCO_3^- \\
\text{init} & 0.0193 \text{ M} & & \sim 10^{-7} \text{ M} & 0 \\
\text{equil} & 0.0193\text{-x} & & \sim\text{x} & \text{x}
\end{array}
$$
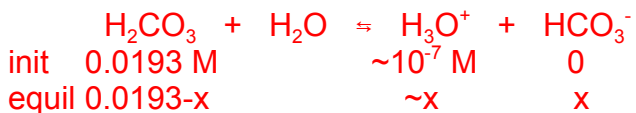
$$K_a = 4.45 \text{ } x \text{ } 10^{-7} = \frac{[H_3O^+]\gamma_{H_3O^+}[HCO_3^-]\gamma_{HCO_3^-}}{[H_2CO_3]\gamma_{H_2CO_3}} = \frac{(x)(0.875)(x)(0.843)}{(0.0193-x)(1)} = \frac{(x)^2(0.875)(0.843)}{(0.0193)}$$

x = 1.0_8 x 10^{-4} M = $[H_3O^+]$ and pH = -log(1.0_8 x 10^{-4} M)(0.875) = 4.02