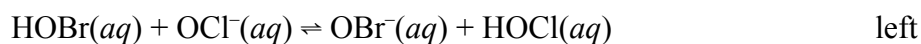
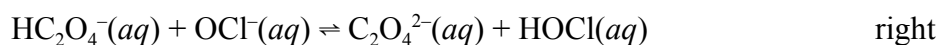
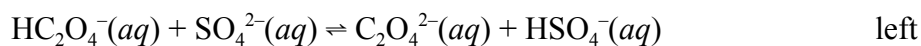
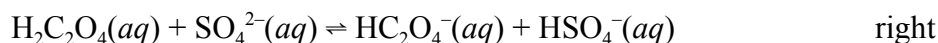


Chem 104
Test 3 Practice Problems
Solutions

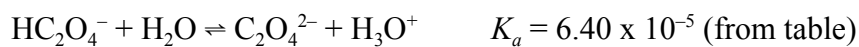
1. Complete the following table by calculating the missing entries and indicating whether the solution is acidic or basic.

$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	pOH	acidic or basic?
$5.0 \times 10^{-8} \text{ M}$	$2.0 \times 10^{-7} \text{ M}$	7.30	6.70	basic

2. Using the Table of Conjugate Acid-Base Pairs, decide whether each of the following equilibria lies to the left or right.



3. Using the Table of Conjugate Acid-Base Pairs, decide whether a solution of $\text{NaHC}_2\text{O}_4(\text{aq})$ is acidic or basic.

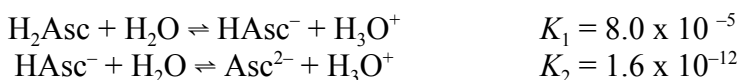


Calculate K_b for HC_2O_4^- from the K_a of its conjugate acid, $\text{H}_2\text{C}_2\text{O}_4$.

$$K_b = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{5.90 \times 10^{-2}} = 1.69 \times 10^{-13}$$

$K_a \gg K_b$, therefore the solution is acidic.

4. Vitamin C is ascorbic acid, a diprotic acid for which $K_1 = 8.0 \times 10^{-5}$ and $K_2 = 1.6 \times 10^{-12}$.
 (a) Using the abbreviation H_2Asc for ascorbic acid, write the hydrolysis equilibria that correspond to K_1 and K_2 .



- (b) Consider a 0.10 M solution of ascorbic acid. Calculate $[\text{H}_3\text{O}^+]$, pH, and the percent dissociation of the acid in this solution.

Assume the first dissociation is the only important source of hydronium ion in the solution; i.e., $[\text{H}_3\text{O}^+] = [\text{HAsc}^-]$ (Assumption I). The analytical concentration of the acid is much greater than K_1 . Therefore, ignore acid hydrolysis (Assumption II); i.e., $[\text{H}_2\text{Asc}] = 0.10 \text{ M}$.

$$[\text{H}_3\text{O}^+] = \sqrt{(0.10)(8.0 \times 10^{-5})} = 2.8_3 \times 10^{-3} \text{ M} = [\text{HAsc}^-]$$

$$\text{pH} = 2.55$$

$$\% \text{ dissociation} = \frac{2.8_3 \times 10^{-3}}{0.10} \times 100\% = 2.8\%$$

- (c) What is the concentration of ascorbate ion, $[\text{Asc}^{2-}]$, in a 0.10 M ascorbic acid solution?

From the initial calculation, $[\text{H}_3\text{O}^+] = [\text{HAsc}^-]$. Therefore, they cancel when substituted into the K_2 expression, and $[\text{Asc}^{2-}] = K_2 = 1.6 \times 10^{-12}$.

5. Consider the titration of 25.0 mL of 0.120 M acetic acid ($\text{CH}_3\text{CO}_2\text{H}$, $K_a = 1.76 \times 10^{-5}$) with 0.100 M $\text{NaOH}(aq)$.

- (a) How much 0.100 M $\text{NaOH}(aq)$ must be added to reach the equivalence point?

$$V_b = \frac{M_a V_a}{M_b} = \frac{(0.120 \text{ M})(25.00 \text{ mL})}{0.100 \text{ M}} = 30.0 \text{ mL}$$

- (b) How many millimoles of $\text{CH}_3\text{CO}_2\text{H}$ are present in the initial sample?

$$\text{mmol CH}_3\text{CO}_2\text{H} = (0.120 \text{ M})(25.0 \text{ mL}) = 3.00 \text{ mmol}$$

- (c) What is the initial pH of the sample solution?

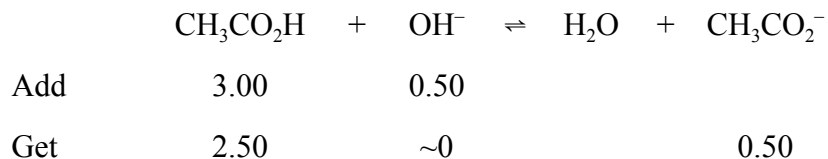
Use Assumptions I and II.

$$[\text{H}_3\text{O}^+] = \sqrt{(0.120)(1.76 \times 10^{-5})} = 1.45 \times 10^{-3}$$

$$\text{pH} = 2.838$$

(d) What is the pH of the solution after adding 5.00 mL of 0.100 M NaOH(aq)?

$$\text{mmol OH}^- \text{ added} = 5.00 \text{ mL} \times 0.10 \text{ M} = 0.50 \text{ mmol}$$



Use K_a to calculate the concentration of hydronium ion in equilibrium with these amounts of acid and conjugate base.

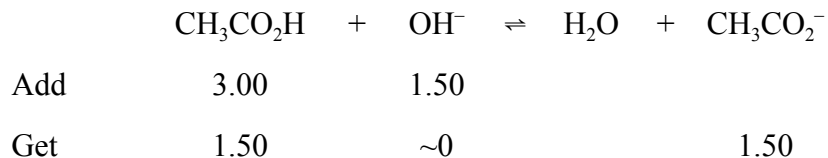
$$K_a = 1.76 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](0.50)}{2.50}$$

$$[\text{H}_3\text{O}^+] = 8.8 \times 10^{-5} \text{ M}$$

$$\text{pH} = 4.05_{56} = 4.06$$

(e) What is the pH of the solution after adding 15.0 mL of 0.100 M NaOH(aq)?

$$\text{mmol OH}^- \text{ added} = 15.0 \text{ mL} \times 0.10 \text{ M} = 1.50 \text{ mmol}$$



This is the half-titration point, where equal amounts of acid and conjugate base exist in the solution.

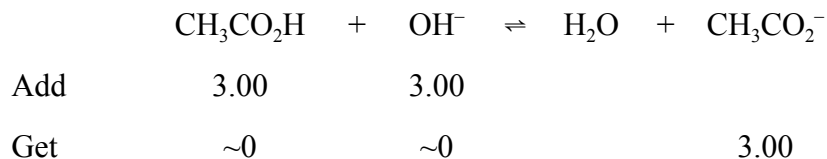
$$K_a = 1.76 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](1.50)}{1.50}$$

Therefore, $[\text{H}_3\text{O}^+] = K_a$ and $\text{pH} = \text{p}K_a$

$$\text{pH} = -\log(1.76 \times 10^{-5}) = 4.754$$

(f) What is the pH at the equivalence point?

All the acid has been converted to conjugate base. Therefore, calculate the K_b for acetate ion, calculate the analytical concentration of the acetate ion in the resulting solution, and use K_b to calculate the concentration of hydroxide ion. Then, calculate pOH, and by subtraction from 14.00 calculate pH.



$$K_b = \frac{1.00 \times 10^{-14}}{1.76 \times 10^{-5}} = 5.68 \times 10^{-10}$$

$$\text{volume} = (25.0 + 35.0) \text{ mL} = 55.0 \text{ mL}$$

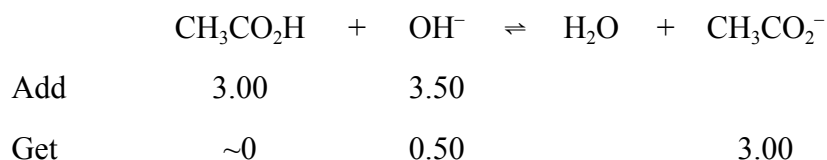
$$c = \frac{3.00 \text{ mmol}}{55.0 \text{ mL}} = 0.0545_{45} \text{ M}$$

$$[\text{OH}^-] = \sqrt{(0.0545)(5.68 \times 10^{-10})} = 5.57 \times 10^{-6} \text{ M}$$

$$\text{pOH} = 5.254 \Rightarrow \text{pH} = 8.746$$

(g) What is the pH when 5.00 mL of 0.100 M NaOH(aq) has been added beyond the equivalence point?

This is after adding 35.0 mL of 0.100 M NaOH, which supplies 3.50 mmol OH^- .



Only the excess sodium hydroxide is an important source of hydroxide ion.

$$\text{volume} = (25.0 + 35.0) \text{ mL} = 60.0 \text{ mL}$$

$$[\text{OH}^-] = \frac{0.50 \text{ mmol}}{60.0 \text{ mL}} = 8.3 \times 10^{-3} \text{ M}$$

$$\text{pOH} = 2.08 \Rightarrow \text{pH} = 11.92$$