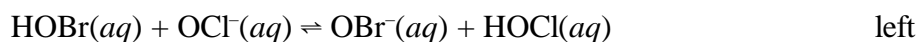
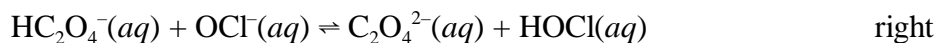
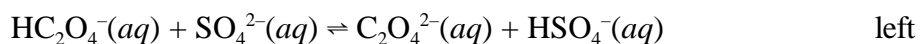
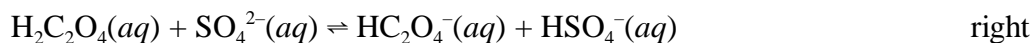


Chem 116
Test 3 Practice Problems
Solutions

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. The K_a of HPO_4^{2-} is 3.6×10^{-13} .

(a) What is the value of K_b for the phosphate ion, PO_4^{3-} ?

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{3.6 \times 10^{-13}} = 2.777 \times 10^{-2} = 2.8 \times 10^{-2}$$

(b) Calculate the concentration of hydroxide ion, $[\text{OH}^-]$, in a 0.10 M solution of Na_3PO_4 .

The analytical concentration of the base, PO_4^{3-} , is too close to the value of K_b to ignore hydrolysis. The more exact mass balance expression must be substituted into the denominator of the K_b expression.

$$K_b = \frac{[\text{OH}^-]^2}{0.10 - [\text{OH}^-]} = 2.8 \times 10^{-2}$$

$$[\text{OH}^-]^2 + 2.8 \times 10^{-2}[\text{OH}^-] - 2.8 \times 10^{-3} = 0$$

Solving the quadratic equation and taking the positive root:

$$[\text{OH}^-] = 0.0407 \text{ M} = 0.041 \text{ M}$$

(c) What is the percent hydrolysis of phosphate ion in a 0.10 M solution of Na_3PO_4 ?

$$\% \text{ hydrolysis} = \frac{0.0407}{0.10} \times 100\% = 41\%$$

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?

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

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

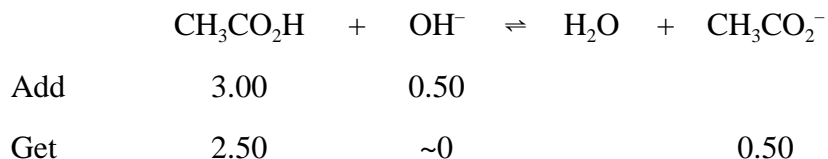
Use Assumptions I and II.

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

$$\text{pH} = 2.838$$

(d) What is the pH of the solution after adding 5.0 mL of 0.100 M NaOH(aq)?

$$\text{mmol OH}^- \text{ added} = 5.0 \text{ mL} \times 0.100 \text{ M} = 0.50 \text{ mmol}$$



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

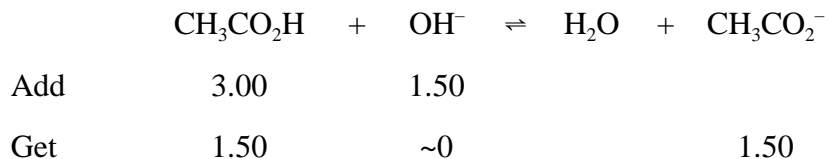
$$K_a = 1.76 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](0.50)}{2.50}$$

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

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

(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.

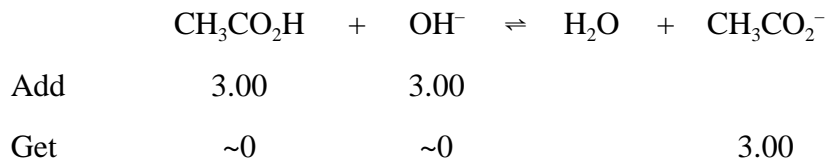
$$K_a = 1.76 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+](1.50)}{1.50}$$

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.



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

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

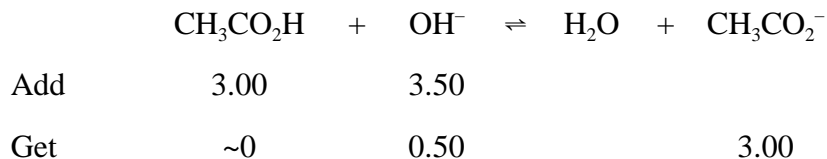
$$C = \frac{3.00 \text{ mmol}}{55.0 \text{ mL}} = 0.0545_{45} \text{ M}$$

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

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

(g) What is the pH when 5.0 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.

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

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

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