

CHM1311 D: Principles of Chemistry (Prof. N. Goto)

Assignment #7

Due Nov 23rd, at the beginning of class. Late assignments will not be accepted.

Assignments can be submitted individually, or by groups of up to 4 students.

1) Last Name: _____ First Name: _____ Student ID: _____

2) Last Name: _____ First Name: _____ Student ID: _____

3) Last Name: _____ First Name: _____ Student ID: _____

4) Last Name: _____ First Name: _____ Student ID: _____

Solutions must be written legibly, in the space provided. Adequate detail to the calculation (including units, appropriate sig figs) must be provided to make it possible for other students to understand how you arrived at the final solution. If more space is needed, use the back of the page. Do not add extra pages, as they will not be marked. Assignment pages must be stapled together.

NOTE: For each question a hint, reference to an Office Hours video, or an Interactive LearningWare (ILW) problem in WileyPLUS is given in brackets.

You will need to use the data in Appendix E of the textbook for some of the questions.

Question 1. (ILW 16.13)

What mass of acetic acid ($\text{CH}_3\text{CO}_2\text{H}$) must be added to 1.45 L of 0.40 M sodium acetate to make a buffer solution whose $\text{pH}=5.25$? (3 marks)

$$\text{p}K_a = 4.75$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$\frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 10^{\text{pH} - \text{p}K_a} = 10^{5.25 - 4.75} = 3.16$$

$$[\text{CH}_3\text{COOH}] = \frac{[\text{CH}_3\text{COO}^-]}{3.16} = \frac{0.40 \text{ M}}{3.16} = 0.126 \text{ M}$$

$$M = nM = cVM = (0.126 \text{ M})(1.45 \text{ L})(60.05 \text{ g mol}^{-1})$$

$$= 11 \text{ g}$$

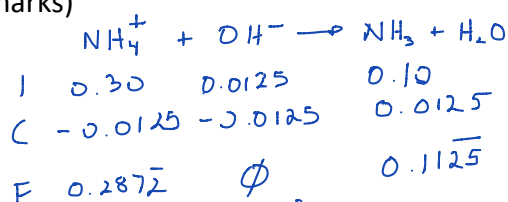
Question 2. (Similar to problem covered Nov 9th, 5th slide)Consider a buffer solution that contains 0.30 M NH_4Cl and 0.10 M NH_3 .

a) Calculate its pH. (2 marks)

$$\begin{aligned}
 \text{pH} &= \text{p}K_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \\
 &= 9.25 + \log \left(\frac{0.10 \text{ M}}{0.30 \text{ M}} \right) \\
 &= 8.77
 \end{aligned}$$

$\text{p}K_b = 4.75$
 $\text{p}K_a = 14 - 4.75 = 9.25$

b) Calculate the change in pH if 0.100 g of solid NaOH is added to 200.0 mL of this solution. (4 marks)



$$\begin{aligned}
 c_{\text{NaOH}} &= \frac{m}{M \cdot V} = \frac{0.100 \text{ g}}{(39.997 \text{ g mol}^{-1})(0.2000 \text{ L})} \\
 &= 0.0125 \text{ mol L}^{-1}
 \end{aligned}$$

$$\begin{aligned}
 \text{pH} &= \text{p}K_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \\
 &= 9.25 + \log \left(\frac{0.1125}{0.2872} \right) \\
 &= 8.84
 \end{aligned}$$

c) If the acceptable buffer range of the solution is ± 0.10 pH units, calculate how many moles of H_3O^+ can be neutralized by 250 mL of the initial buffer. (4 marks)

WE WANT TO KNOW HOW MUCH H_3O^+ CAN BE ADDED TO LOWER pH BY MORE THAN 0.10 - NEW pH WOULD BE $8.77 - 0.10 = 8.67$

$$\begin{aligned}
 \text{NEW } [\text{NH}_3] &= [\text{NH}_3]_0 - [\text{H}_3\text{O}^+]_{\text{ADDED}} \\
 [\text{NH}_4^+] &= [\text{NH}_4^+]_0 + [\text{H}_3\text{O}^+]_{\text{ADDED}}
 \end{aligned}$$

PUT INTO H-H EQUATION:

$$\begin{aligned}
 \text{pH} &= \text{p}K_a + \log \frac{[\text{NH}_3]_0 - [\text{H}_3\text{O}^+]_{\text{ADDED}}}{[\text{NH}_4^+]_0 + [\text{H}_3\text{O}^+]_{\text{ADDED}}} \\
 \frac{[\text{NH}_3]_0 - [\text{H}_3\text{O}^+]_{\text{ADDED}}}{[\text{NH}_4^+]_0 + [\text{H}_3\text{O}^+]_{\text{ADDED}}} &= 10^{\text{pH} - \text{p}K_a} = 10^{8.67 - 9.25} = 0.2648
 \end{aligned}$$

$$[\text{NH}_3]_0 - [\text{H}_3\text{O}^+]_{\text{ADDED}} = 0.2648 ([\text{NH}_4^+]_0 + [\text{H}_3\text{O}^+]_{\text{ADDED}})$$

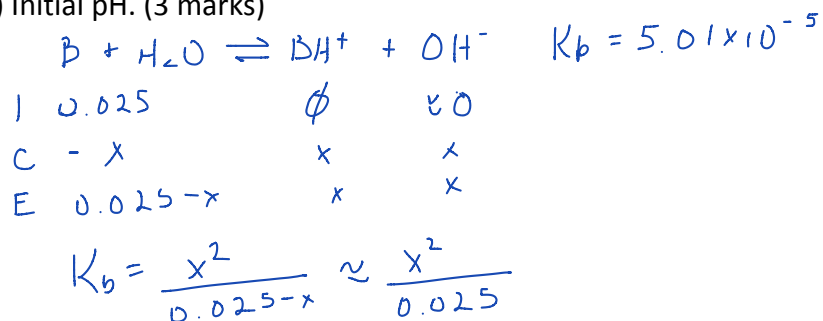
$$\frac{[\text{NH}_3]_0 - 0.2648 [\text{NH}_4^+]_0}{1.2648} = [\text{H}_3\text{O}^+]_{\text{ADDED}}$$

$$\begin{aligned}
 [\text{H}_3\text{O}^+]_{\text{ADDED}} &= 0.01626 \text{ M} \quad n = cV \\
 &= (0.01626 \text{ M})(0.250 \text{ L}) \\
 &= 4.07 \times 10^{-3} \text{ mol}
 \end{aligned}$$

Question 3. (Video 16.23)

Metolazone is a weakly basic ($pK_b = 4.30$) diuretic used to treat congestive heart failure. For analysis of the metolazone content of a generic brand pill, a titration is carried out with strong acid. It takes 50.00 mL of 0.100 M acid to reach the stoichiometric point. The initial metolazone concentration is 0.025 M and the solution volume is 200. mL. Calculate the pH of this solution at the following points:

a) Initial pH. (3 marks)



$[B]_{eq} \approx [B]_0$ VALID?

$$\% \text{ ERROR} = \frac{x}{0.025-x} \times 100\%$$

$$= 4.7\%$$

\therefore VALID

$$x = \sqrt{0.025 K_b} = 1.12 \times 10^{-3} \text{ M} = [OH^-]$$

$$pOH = 2.95$$

$$pH = 11.05$$

b) After addition of 19.50 mL of acid (4 marks)

$$n_{H_3O^+} = C_{H_3O^+} V_{H_3O^+} = (0.100 \text{ M})(19.50 \times 10^{-2} \text{ L})$$

$$= 1.950 \times 10^{-3} \text{ mol}$$

$$n_B = C_B V_B = (0.025 \text{ M})(0.200 \text{ L}) = 0.005 \text{ mol}$$

$$n_{B, \text{EXCESS}} = n_B - n_{H_3O^+} = 0.005 \text{ mol} - 0.00195 \text{ mol}$$

$$= 0.00305 \text{ mol}$$

$$[B] = \frac{n_{B, \text{EXCESS}}}{V_B + V_{H_3O^+}} = \frac{0.00305 \text{ mol}}{0.200 \text{ L} + 0.01950 \text{ L}} = 0.0139 \text{ M}$$

$$[BH^+] = \frac{n_{H_3O^+}}{V_{\text{TOTAL}}} = \frac{0.001950 \text{ mol}}{0.21950 \text{ L}} = 0.00888 \text{ M}$$

$$pH = pK_a + \log \frac{[B]}{[BH^+]} = 9.70 + \log \left(\frac{0.0139}{0.00888} \right) = 9.89$$

c) After addition of 25.00 mL of base solution. (1 mark)

\therefore 25.00 mL IS HALF-WAY TO STOICHIOMETRIC POINT

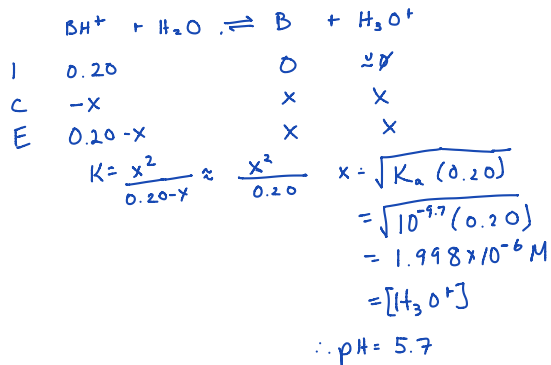
$$\therefore n_B = n_{BH^+}$$

$$\therefore \frac{[B]}{[BH^+]} = 1 \quad \therefore pH = pK_a$$

$$= 9.70$$

d) At the stoichiometric point. (5 marks) ALL B CONVERTED TO BH^+

$$[BH^+] = \frac{n_B}{V_{H_3O^+} + V_B} = \frac{0.005 \text{ mol}}{0.200 \text{ L} + 0.050 \text{ L}} = 0.02 \text{ M}$$



e) After addition of 56.8 mL base solution. (3 marks)

$$C_{H_3O^+, \text{ EXCESS}} = \frac{n_{H_3O^+} - n_B}{V_{H_3O^+} + V_B} = \frac{C_{H_3O^+} V_{H_3O^+} - n_B}{V_T}$$

$$= \frac{(0.100 \text{ M})(0.0568 \text{ L}) - 0.005 \text{ mol}}{0.2568 \text{ L}}$$

$$= 2.65 \times 10^{-3} \text{ M}$$

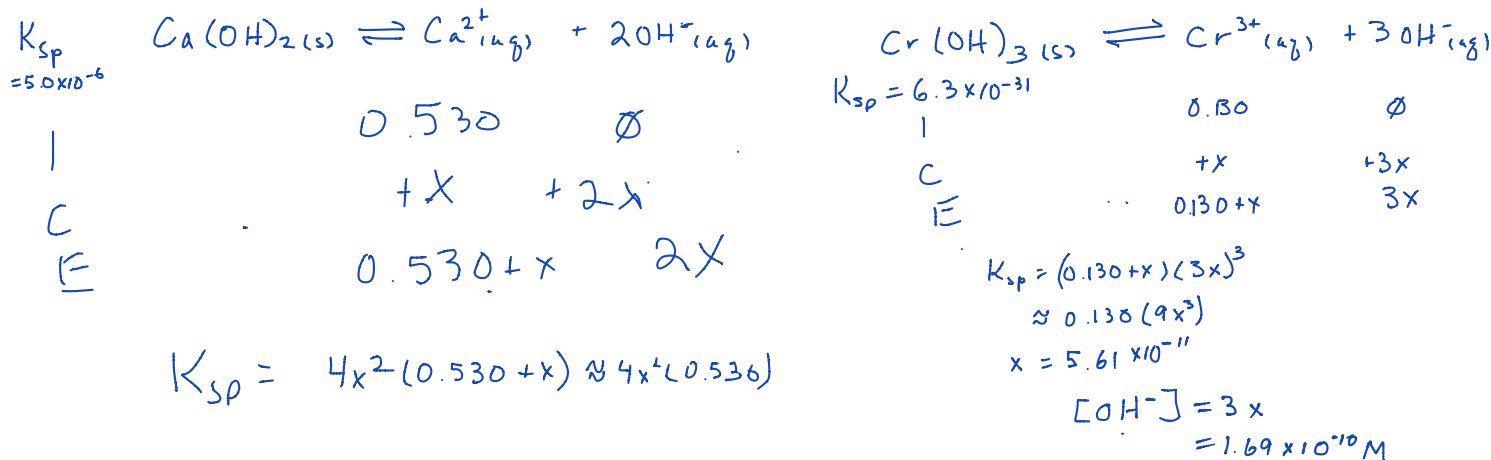
$$pH = 2.58$$

f) Choose an indicator from Table 6.2 of the textbook (Chapter 16, slide 30) that would be most appropriate for this titration. (1 mark)

\therefore pH @ STOICHIOMETRIC POINT IS 5.7
 \therefore USE BROMOCRESOL PURPLE (ITS pK_a IS CLOSEST @ 6.3)

Question 6. (Hint: Calculate the maximum concentration of hydroxide ions that can be present in the solution before each ion precipitates)

a) Can Cr^{3+} be separated from Ca^{2+} by the addition of an NaOH solution to an acidic solution that contains 0.530 M Ca^{2+} and 0.130 M Cr^{3+} ? Assume that the initial concentration of OH^- is zero for these calculations. (5 marks)



$$x = 0.00154 \text{ M} \quad [\text{OH}^-] = 2(0.00154 \text{ M}) = 0.00307 \text{ M}$$

IT IS POSSIBLE TO PRECIPITATE THE CHROMIUM BY RAISING THE $[\text{OH}^-]$ UP TO THE CONCENTRATION FOUND @ THE SOLUBILITY LIMIT OF Ca(OH)_2 .

b) At what hydroxide ion concentration will the second cation precipitate? (i.e. the cation of the more soluble salt.) This was calculated in part (a). (1 mark)

WE WOULD START TO SEE PRECIPITATION OF Ca(OH)_2 WHEN $[\text{OH}^-] = 0.00307 \text{ M}$

c) What is the concentration of the less soluble cation at the hydroxide ion concentration in (b)? (2 marks)

$$\begin{aligned}
 K_{sp} &= [\text{Cr}^{3+}][\text{OH}^-]^3 \\
 [\text{Cr}^{3+}] &= \frac{K_{sp}}{[\text{OH}^-]^3} \\
 &= \frac{6.3 \times 10^{-31}}{(0.00307)^3} \\
 &= 2.17 \times 10^{-23} \text{ M}
 \end{aligned}$$