

Vince A

CHEM 1423 - Exam 3 - March 31, 2016

Name Solu Kong

(60) PART I. MULTIPLE CHOICE (Circle the ONE correct answer)

- The pH of a 0.10 M solution of Hypoiodous acid, HIO, is 5.80. What is the approximate Acid Dissociation Constant of HIO?
(A) 2.5×10^{-11} (B) 1.6×10^{-5} (C) 4.0×10^{-11}
(D) None of the above
- The pH of a 0.05 M solution of a weak base, B, is 8.6. Therefore, the base equilibrium constant is approximately:
(A) 1.3×10^{-16} (B) 7.4×10^{-6} (C) 3.2×10^{-10} (D) 1.6×10^{-11}
- Which of the following aqueous solutions is/are basic (pH > 7)?
~~X~~ (i) Ammonium Chloride (NH_4Cl)
~~X~~ (ii) Potassium Nitrate (KNO_3)
~~X~~ (iii) Pyridinium Bromide (PyrHBr)
 (iv) Sodium Cyanide (NaCN)
(A) i & iv (B) ii & iv (C) i & iii (D) iv only
- If added to 2 L of 0.40 M HNO_3 , which one of the following would form a buffer?
(A) 1.20 mol of KOH (B) 1.30 mol of sodium acetate (NaAc)
(C) 1.30 mol of NH_4Cl (D) 0.60 mol of potassium lactate (KLac)

For #5 - #8: Consider the weak base, Quinoline (Quin). Its base equilibrium constant is 6.0×10^{-10} .

- What is the approximate pH of a 0.05 M solution of Quinoline?
(A) 11.0 (B) 5.3 (C) 8.7 (D) 9.4
- What is the approximate percent protonation in a 0.05 M solution of Quinoline?
(A) $5.5 \times 10^{-6} \%$ (B) $1.1 \times 10^{-2} \%$ (C) $5.5 \times 10^{-4} \%$
(D) Cannot be determined without the hydroxide concentration, $[\text{OH}^-]$
- What is the approximate pH of a 0.01 M solution of Quinolinium Chloride, (QuinHCl)?
(A) 3.4 (B) 5.6 (C) 8.4 (D) 6.3

Vincenzo A

8. What is the approximate pH of a solution containing 0.60 M Quinoline (Quin) and 0.20 M Quinolinium Chloride (QuinHCl)?

- (A) 4.3 (B) 5.3 (C) 8.7 (D) 9.7

For #9 - #10: Consider the weak acid, hypochlorous acid, HClO. Its acid dissociation constant is 3.0×10^{-8} .

9. What is the approximate percent dissociation of a 0.005 M solution of HClO?

- (A) $2.4 \times 10^{-3} \%$ (B) $3.5 \times 10^{-1} \%$ (C) $2.4 \times 10^{-1} \%$ (D) 3.5%

10. What is the approximate pH of a 0.10 M potassium hypochlorite, KClO, solution?

- (A) 3.7 (B) 4.3 (C) 9.7 (D) 10.3

For #11 - #15: Tellurous acid, H_2TeO_3 , is a diprotic acid with acid dissociation constants, $K_{a1} = 3.0 \times 10^{-3}$ and $K_{a2} = 2.0 \times 10^{-8}$.

11. What is the approximate pH of a solution containing 0.05 M Na_2TeO_3 ?

- (A) 10.2 (B) 9.5 (C) 3.8
(D) None of the above

12. What is the approximate pH of a solution containing pure KHTeO_3 ?

- (A) 7.7 (B) 5.1 (C) 2.5
(D) The pH depends upon the concentration of KHTeO_3

13. What is the approximate pH of a solution containing 0.50 M KHTeO_3 and 0.20 M Na_2TeO_3 ?

- (A) 7.3 (B) 2.9 (C) 2.1 (D) 8.1

14. What is the approximate pH of a solution prepared by adding 0.40 mol of HNO_3 to 2.0 L of 0.30 M KHTeO_3 ?

- (A) 2.0 (B) 2.8 (C) 2.2 (D) 3.4

15. Approximately what ratio of $[\text{TeO}_3^{2-}]/[\text{HTeO}_3^-]$ will give a pH of 7.30?

- (A) 0.20 (B) 2.5 (C) 1.5 (D) 0.40

Version A

For #16 - #18: Consider the amino acid, Histidine (His). The most positive form of Histidine is His^{2+} and the most negative form is His^{1-} . The three pK_a 's of Histidine are: $\text{pK}_a' = 1.8$, $\text{pK}_a'' = 6.0$, and $\text{pK}_a''' = 9.2$.

16. What is the isoelectric point (pI) of Histidine?
(A) 3.9 (B) 6.0 (C) 7.6 (D) 1.8
17. At what pH does one have 50% His^{1+} and 50% His^0 ?
(A) 6.0 (B) 3.9 (C) 1.8 (D) 7.6
18. What is the average charge on the Histidine molecule at $\text{pH} = 9.2$?
(A) +1.5 (B) +1.0 (C) +0.5 (D) -0.5
19. If one mixes 99. mL of 0.10 M HCl to 100. mL of 0.10 M NaOH, the pH of the resultant solution is approximately:
(A) 9.0 (B) 10.7 (C) 5.0 (D) 3.3
20. 180 mL of 0.20 M $\text{H}_3\text{PO}_4(\text{aq})$ is needed to completely neutralize 200 mL of an aqueous NaOH(aq) solution? What is the approximate Molarity of the NaOH(aq) solution?
(A) 0.54 M (B) 0.06 M (C) 0.18 M
(D) None of the above

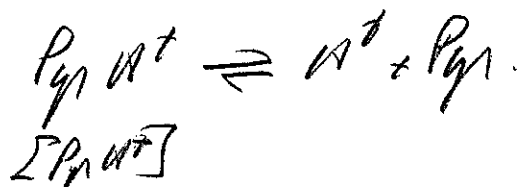
PART II. THREE (3) PROBLEMS BELOW: REMEMBER TO SHOW WORK FOR CREDIT

Version A

- (10) 1. Pyridine [$C_5H_5N = Pyr$] is a weak base with a base equilibrium constant, $K_b = 1.8 \times 10^{-9}$. The pH of an aqueous solution containing Pyridinium Bromide [$C_5H_5NHBr = PyrHBr$, $M = 160$.] is $pH = 3.1$. Calculate the mass percent of $PyrHBr$ in the aqueous solution.

Note: Assume that the density of the aqueous solution is 1.0 g/mL .

$$K_a = \frac{10^{-14}}{K_b} = \frac{10^{-14}}{1.8 \times 10^{-9}} = 5.56 \times 10^{-6}$$



$$\begin{aligned} [H^+] &= [Pyr] \\ &= 10^{-3.1} = 2.94 \times 10^{-4} \end{aligned}$$

$$K_a = \frac{[H^+][Pyr]}{[PyrH^+]} \rightarrow 5.56 \times 10^{-6} = \frac{(2.94 \times 10^{-4})^2}{[PyrH^+]}$$

$$\therefore [PyrH^+] = [PyrHBr] = \frac{(2.94 \times 10^{-4})^2}{5.56 \times 10^{-6}} = 0.113 \text{ mL}$$

Assume $1L = 1000 \text{ mL}$

$$n_{PyrHBr} = 0.113 \text{ mL} \times 1L = 0.113 \text{ mL}$$

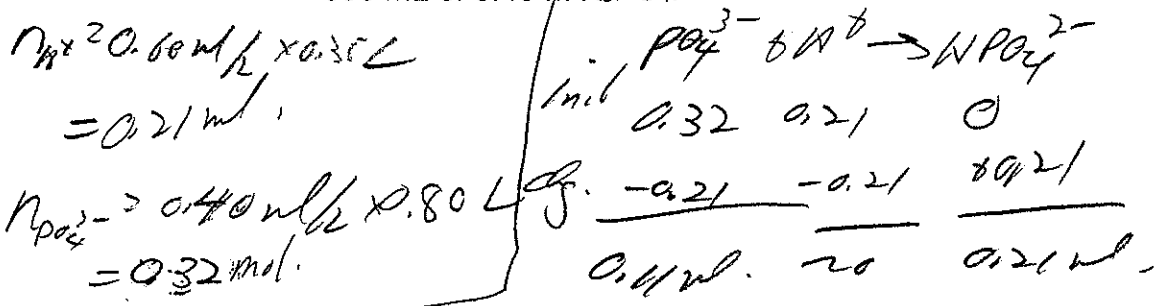
$$m_{PyrHBr} = 0.113 \text{ mL} \times 160 \text{ g/mL} = 18.1 \text{ g}$$

$$m_{tot} = 1000 \text{ mL} \times \frac{1 \text{ g}}{\text{mL}} = 1000 \text{ g}$$

$$m\% = \frac{18.1}{1000} \times 100 = 1.81\% = \boxed{1.81\%}$$

$H_3PO_4 \rightleftharpoons H_2PO_4^- \rightleftharpoons HPO_4^{2-} \rightleftharpoons PO_4^{3-}$ Version A
 $pK_a' \rightarrow 2.12 \quad 7.21 \quad 12.44$
 (20) 2. Phosphoric Acid (H_3PO_4) is a triprotic acid with acid dissociation constants, $K_a' = 7.5 \times 10^{-3}$, $K_a'' = 6.2 \times 10^{-8}$ and $K_a''' = 3.6 \times 10^{-13}$

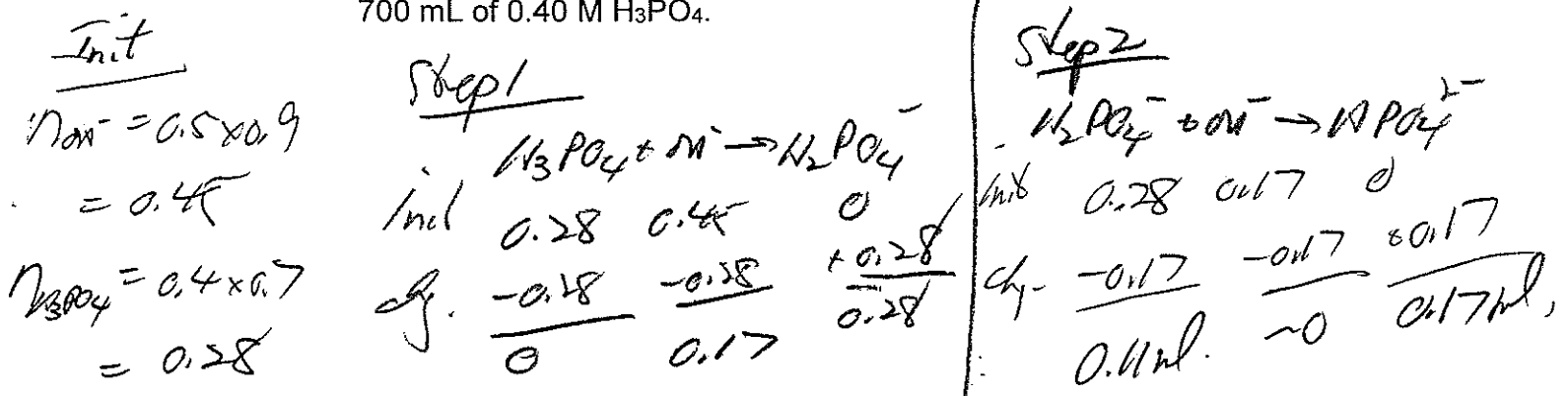
Init (7) (a) Calculate the pH of a solution prepared by mixing 350 mL of 0.60 M HCl with 800 mL of 0.40 M K_3PO_4 .



$$pH = 12.44 + \log \frac{[PO_4^{3-}]}{[HPO_4^{2-}]} = 12.44 + \log \left(\frac{0.11}{0.21} \right)$$

$$= \boxed{12.16}$$

(7) (b) Calculate the pH of a solution prepared by mixing 900 mL of 0.50 M KOH with 700 mL of 0.40 M H_3PO_4 .



$$pH = 7.21 + \log \frac{[HPO_4^{2-}]}{[H_2PO_4^-]}$$

$$= 7.21 + \log \frac{0.17}{0.11}$$

$$= \boxed{7.40}$$

Version A

Prob. 2 (Cont'd)

- (6) (c) Calculate the ratio, $[H_3PO_4]/[H_2PO_4^-]$ required to prepare a buffer solution with pH = 2.62.

$$pH = pK_a + \log \frac{[H_2PO_4^-]}{[H_3PO_4]} \rightarrow 2.62 = 2.12 + \log \frac{[H_2PO_4^-]}{[H_3PO_4]}$$

$$\log \frac{[H_2PO_4^-]}{[H_3PO_4]} = 2.62 - 2.12 = 0.50$$

$$\frac{[H_2PO_4^-]}{[H_3PO_4]} = 10^{0.50} = 3.16 \rightarrow \frac{[H_3PO_4]}{[H_2PO_4^-]} = \frac{1}{3.16} = \boxed{0.32}$$

- (10) 3. When 7.50 grams of a sample of impure Calcium Hydroxide $[Ca(OH)_2]$, $M = 74.1$ is titrated with 0.35 M H_3PO_4 , it takes 150. mL of H_3PO_4 to completely titrate the base. Calculate the **mass percent of impurity** in the Calcium Hydroxide sample.



$$n_{H_3PO_4} = 0.35 \text{ mol/L} \times 0.15 \text{ L} = 0.0525 \text{ mol}$$

$$n_{Ca(OH)_2} = 0.0525 \text{ mol } H_3PO_4 \times \frac{3 \text{ mol } Ca(OH)_2}{2 \text{ mol } H_3PO_4} = 0.0788 \text{ mol}$$

$$m_{Ca(OH)_2} = 0.0788 \text{ mol} \times 74.1 \text{ g/mol} = 5.84 \text{ g } Ca(OH)_2$$

$$m_{\text{impur}} = 7.50 - 5.84 = 1.66 \text{ g}$$

$$m\%(\text{impur}) = \frac{1.66}{7.50} \times 100 = 22.1\% \\ \approx \boxed{22\%}$$

Version B

CHEM 1423 - Exam 3 – March 31, 2016

Name Solution

(60) PART I. MULTIPLE CHOICE (Circle the ONE correct answer)

1. Which of the following aqueous solutions is/are basic (pH > 7)?

- ~~(i)~~ Ammonium Chloride (NH_4Cl)
~~(ii)~~ Potassium Nitrate (KNO_3)
~~(iii)~~ Pyridinium Bromide (PyrHBr)
~~(iv)~~ Sodium Cyanide (NaCN)

(A) iv only (B) ii & iv (C) i & iii (D) i & iv

2. If added to 2 L of 0.40 M HNO_3 , which one of the following would form a buffer?

- (A) 1.20 mol of KOH (B) 0.60 mol of potassium lactate (KLac)
(C) 1.30 mol of NH_4Cl (D) 1.30 mol of sodium acetate (NaAc)

3. The pH of a 0.05 M solution of a weak base, B, is 8.6. Therefore, the base equilibrium constant is approximately:

- (A) 1.3×10^{-16} (B) 3.2×10^{-10} (C) 7.5×10^{-6} (D) 1.6×10^{-11}

4. The pH of a 0.10 M solution of Hypoiodous acid, HIO, is 5.80. What is the approximate Acid Dissociation Constant of HIO?

- (A) 4.0×10^{-11} (B) 1.6×10^{-5} (C) 2.5×10^{-11}
(D) None of the above

For #5 - #8: Consider the weak base, Quinoline (Quin). Its base equilibrium constant is 6.0×10^{-10} .

5. What is the approximate pH of a 0.01 M solution of Quinolinium Chloride, (QuinHCl)?

- (A) 5.6 (B) 3.4 (C) 8.4 (D) 6.3

6. What is the approximate pH of a solution containing 0.60 M Quinoline (Quin) and 0.20 M Quinolinium Chloride (QuinHCl)?

- (A) 4.3 (B) 9.7 (C) 8.7 (D) 5.3

7. What is the approximate pH of a 0.05 M solution of Quinoline?

- (A) 8.7 (B) 5.3 (C) 11.0 (D) 9.4

Version B

8. What is the approximate **percent protonation** in a 0.05 M solution of Quinoline?
(A) $5.5 \times 10^{-6} \%$ (B) $5.5 \times 10^{-4} \%$ (C) $1.1 \times 10^{-2} \%$
(D) Cannot be determined without the hydroxide concentration, $[\text{OH}^-]$

For #9 - #13: Tellurous acid, H_2TeO_3 , is a diprotic acid with acid dissociation constants, $K_a' = 3.0 \times 10^{-3}$ and $K_a'' = 2.0 \times 10^{-8}$

9. What is the approximate pH of a solution containing pure KHTeO_3 ?
(A) 5.1 (B) 7.7 (C) 2.5
(D) The pH depends upon the concentration of KHTeO_3
10. What is the approximate pH of a solution containing 0.05 M Na_2TeO_3 ?
(A) 9.5 (B) 10.2 (C) 3.8
(D) None of the above
11. What is the approximate pH of a solution containing 0.50 M KHTeO_3 and 0.20 M Na_2TeO_3 ?
(A) 8.1 (B) 2.9 (C) 2.1 (D) 7.3
12. What is the approximate pH of a solution prepared by adding 0.40 mol of HNO_3 to 2.0 L of 0.30 M KHTeO_3 ?
(A) 2.0 (B) 2.8 (C) 2.2 (D) 3.4
13. Approximately what ratio of $[\text{TeO}_3^{2-}]/[\text{HTeO}_3^-]$ will give a pH of 7.30 ?
(A) 0.20 (B) 2.5 (C) 1.5 (D) 0.40

For #14 - #15: Consider the weak acid, hypochlorous acid, HClO . Its acid dissociation constant is 3.0×10^{-8} .

14. What is the approximate pH of a 0.10 M potassium hypochlorite, KClO , solution?
(A) 10.3 (B) 4.3 (C) 9.7 (D) 3.7
15. What is the approximate percent dissociation of a 0.005 M solution of HClO ?
(A) $2.4 \times 10^{-3} \%$ (B) $2.4 \times 10^{-1} \%$ (C) $3.5 \times 10^{-1} \%$ (D) 3.5%

Version B

For #16 - #18: Consider the amino acid, Histidine (His). The most positive form of Histidine is His^{2+} and the most negative form is His^{1-} . The three pK_a 's of Histidine are: $\text{pK}_a' = 1.8$, $\text{pK}_a'' = 6.0$, and $\text{pK}_a''' = 9.2$.

16. At what pH does one have 50% His^{1+} and 50% His^0 ?
(A) 1.8 (B) 3.9 (C) 6.0 (D) 7.6
17. What is the average charge on the Histidine molecule at $\text{pH} = 9.2$?
(A) +1.5 (B) +1.0 (C) +0.5 (D) -0.5
18. What is the isoelectric point (pI) of Histidine?
(A) 7.6 (B) 6.0 (C) 3.9 (D) 1.8
19. 180 mL of 0.20 M $\text{H}_3\text{PO}_4(\text{aq})$ is needed to completely neutralize 200 mL of an aqueous $\text{NaOH}(\text{aq})$ solution? What is the approximate Molarity of the $\text{NaOH}(\text{aq})$ solution?
(A) 0.18 M (B) 0.06 M (C) 0.54 M
(D) None of the above
20. If one mixes 99. mL of 0.10 M HCl to 100. mL of 0.10 M NaOH , the pH of the resultant solution is approximately:
(A) 9.0 (B) 10.7 (C) 5.0 (D) 3.3

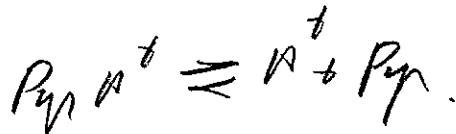
PART II. THREE (3) PROBLEMS BELOW: REMEMBER TO SHOW WORK FOR CREDIT

Version B

- (10) 1. Pyridine [$C_5H_5N = Pyr$] is a weak base with a base equilibrium constant, $K_b = 1.8 \times 10^{-9}$. The pH of an aqueous solution containing Pyridinium Bromide [$C_5H_5NHBr = PyrHBr$, $M = 160$.] is $pH = 2.8$. Calculate the mass percent of $PyrHBr$ in the aqueous solution.

Note: Assume that the density of the aqueous solution is 1.0 g/mL.

$$K_a = \frac{10^{-14}}{K_b} = 5.56 \times 10^{-6}$$



$$\begin{aligned} [H^+] &= [Pyr] \\ &= 10^{-2.8} = 1.58 \times 10^{-3} \end{aligned}$$

$$K_a = \frac{[H^+][Pyr]}{[PyrH^+]} \rightarrow 5.56 \times 10^{-6} = \frac{(1.58 \times 10^{-3})^2}{[PyrH^+]}$$

$$\begin{aligned} \therefore [PyrH^+] &= [PyrHBr] = \frac{(1.58 \times 10^{-3})^2}{5.56 \times 10^{-6}} \\ &= 0.449 \text{ mol/L} \end{aligned}$$

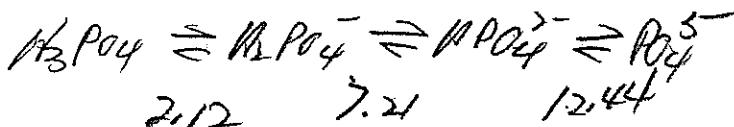
Assume $V = 1000 \text{ mL}$

$$n_{PyrHBr} = 0.449 \text{ mol/L} \times 1000 \text{ mL} = 449 \text{ mol}$$

$$m_{PyrHBr} = 0.449 \text{ mol} \times 160 \text{ g/mol} = 71.8 \text{ g}$$

$$m_{H_2O} = 1000 \text{ mL} \times \frac{1 \text{ g}}{\text{mL}} = 1000 \text{ g}$$

$$\% = \frac{71.8}{1000} \times 100 = 7.18\% \approx \boxed{7.2\%}$$



Version B

pk_a's

2.12

7.21

12.44

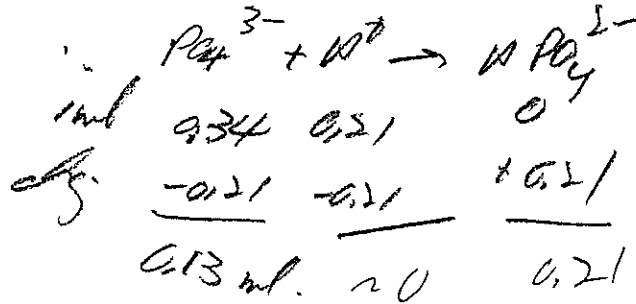
(20) 2. Phosphoric Acid (H₃PO₄) is a triprotic acid with acid dissociation constants, K_a' = 7.5x10⁻³, K_a'' = 6.2x10⁻⁸ and K_a''' = 3.6x10⁻¹³

inl

(7) (a) Calculate the pH of a solution prepared by mixing 350 mL of 0.60 M HCl with 850 mL of 0.40 M K₃PO₄.

$$n_{\text{H}^+} = 0.60 \times 0.35 = 0.21 \text{ mol}$$

$$n_{\text{PO}_4^{3-}} = 0.40 \times 0.85 = 0.34 \text{ mol}$$



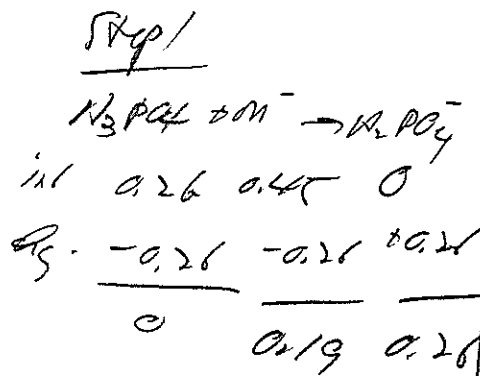
$$\text{pH} = 12.44 + \log \frac{[\text{PO}_4^{3-}]}{[\text{HPO}_4^{2-}]} = 12.44 + \log \frac{0.13}{0.21} = \boxed{12.23}$$

(7) (b) Calculate the pH of a solution prepared by mixing 900 mL of 0.50 M KOH with 650 mL of 0.40 M H₃PO₄.

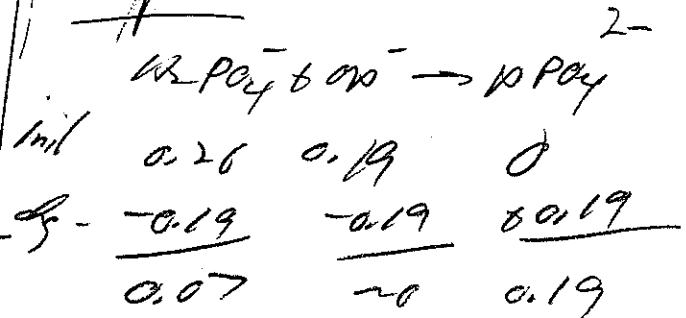
inl

$$n_{\text{OH}^-} = 0.5 \times 0.9 = 0.45$$

$$n_{\text{H}_3\text{PO}_4} = 0.4 \times 0.65 = 0.26$$



Step 2



$$\text{pH} = 7.21 + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

$$= 7.21 + \log \left(\frac{0.19}{0.07} \right)$$

$$= \boxed{7.64}$$

Version B

Prob. 2 (Cont'd)

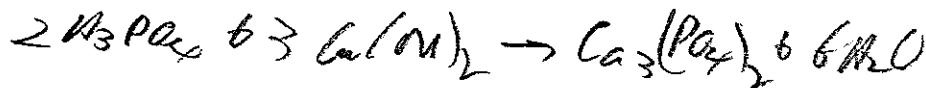
- (6) (c) Calculate the ratio, $[H_3PO_4]/[H_2PO_4^-]$ required to prepare a buffer solution with $pH = 2.66$.

$$pH = pK_a' + \log \frac{[H_2PO_4^-]}{[H_3PO_4]} \rightarrow 2.66 = 2.12 + \log \frac{[H_2PO_4^-]}{[H_3PO_4]}$$

$$\log \frac{[H_2PO_4^-]}{[H_3PO_4]} = 2.66 - 2.12 = +0.54$$

$$\frac{[H_2PO_4^-]}{[H_3PO_4]} = 10^{+0.54} = 3.47 \rightarrow \frac{[H_3PO_4]}{[H_2PO_4^-]} = \frac{1}{3.47} = \boxed{0.29}$$

- (10) 3. When 7.80 grams of a sample of impure Calcium Hydroxide $[Ca(OH)_2, M = 74.1]$ is titrated with 0.35 M H_3PO_4 , it takes 150. mL of H_3PO_4 to completely titrate the base. Calculate the **mass percent of impurity** in the Calcium Hydroxide sample.



$$n_{H_3PO_4} = 0.35 \text{ mol/L} \times 0.15 \text{ L} = 0.0525 \text{ mol}$$

$$n_{Ca(OH)_2} = 0.0525 \text{ mol } H_3PO_4 \times \frac{3 \text{ mol } Ca(OH)_2}{2 \text{ mol } H_3PO_4} = 0.788 \text{ mol } Ca(OH)_2$$

$$m_{Ca(OH)_2} = 0.788 \text{ mol} \times 74.1 \text{ g/mol} = 5.84 \text{ g } Ca(OH)_2$$

$$m_{\text{impur}} = 7.80 - 5.84 = 1.96 \text{ g}$$

$$M\%(\text{impur}) = \frac{1.96}{7.80} \times 100 = 25.1\% = \boxed{25\%}$$