Version A

CHEM 1423 - Exam 3 – March 31, 2016

Name Solutions

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

1. 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) $2.5x10^{-11}$ (B) $1.6x10^{-5}$ (C) $4.0x10^{-11}$
   (D) None of the above

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

   (A) $1.3x10^{-16}$ (B) $7.4x10^{-6}$ (C) $3.2x10^{-10}$ (D) $1.6x10^{-11}$

3. 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) i & iv (B) ii & iv (C) i & iii (D) iv only

4. 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 $6.0x10^{-10}$.

5. 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

6. What is the approximate **percent protonation** in a 0.05 M solution of Quinoline?

   (A) $5.5x10^{-6}$ % (B) $1.1x10^{-2}$ % (C) $5.5x10^{-4}$ %
   (D) Cannot be determined without the hydroxide concentration, $[OH^-]$

7. 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

Version 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) 2.4×10^{-3} % (B) 3.5×10^{-1} % (C) 2.4×10^{-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, $K_a' = 3.0 \times 10^{-3}$ and $K_a'' = 2.0 \times 10^{-8}$

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 $[TeO_3^{2-}]/[HTeO_3^-]$ 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 pKa's of Histidine are: $pK_a' = 1.8$, $pK_a'' = 6.0$, and $pK_a''' = 9.2$.

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 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 H_3PO_4(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, $K_b = 1.8x10^{-9}$. 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.

$$K_a = \frac{10^{-14}}{K_b} = \frac{10^{-14}}{1.8\times10^{-9}} = 5.56\times10^{-6}$$

$$[H^+] = [Pyr] = 10^{-3.1} = 2.94\times10^{-4}$$

$$PyrH^+ \rightleftharpoons H^+ + Pyr$$

$$[PyrH^+]$$

$$K_a = \frac{[H^+][Pyr]}{[PyrH^+]} \rightarrow 5.56\times10^{-6} = \frac{(2.94\times10^{-4})^2}{[PyrH^+]}$$

$$\therefore [PyrH^+] = [PyrHBr] = \frac{(2.94\times10^{-4})^2}{5.56\times10^{-6}} = 0.113\ mol/L$$

Assume 1L = 1000 mL

$$n_{PyrHBr} = 0.113\ mol/L \times 1L = 0.113\ mol$$

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

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

$$m\% = \frac{18.1}{1000}\times 100 = 1.81\% = \boxed{1.8\%}$$

$H_3PO_4 \rightleftharpoons H_2PO_4^- \rightleftharpoons HPO_4^{2-} \rightleftharpoons PO_4^{3-}$ Version A

$pKa's \rightarrow 2.12 \quad 7.21 \quad 12.44$

(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.

$n_{H^+} = 0.60\,mol/L \times 0.35\,L = 0.21\,mol$

$n_{PO_4^{3-}} = 0.40\,mol/L \times 0.80\,L = 0.32\,mol$

	PO_4^{3-}	$+ H^+$	$\rightarrow HPO_4^{2-}$
Init	0.32	0.21	0
Chg.	−0.21	−0.21	+0.21
	0.11 mol	~0	0.21 mol

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

$= \boxed{12.16}$

(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.

Init

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

$n_{H_3PO_4} = 0.4 \times 0.7 = 0.28$

Step 1

	H_3PO_4	$+ OH^-$	$\rightarrow H_2PO_4^-$
Init	0.28	0.45	0
Chg.	−0.28	−0.28	+0.28
	0	0.17	0.28

Step 2

	$H_2PO_4^-$	$+ OH^-$	$\rightarrow HPO_4^{2-}$
Init	0.28	0.17	0
Chg.	−0.17	−0.17	+0.17
	0.11 mol	~0	0.17 mol

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

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

$= \boxed{7.40}$

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.

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

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

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

(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.

$$2H_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6H_2O$$

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

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

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

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

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

$$\approx \boxed{22\%}$$

Version B

CHEM 1423 - Exam 3 – March 31, 2016

Name Solutions

(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) 5.5×10^{-6} % (B) 5.5×10^{-4} % (C) 1.1×10^{-2} %

   (D) Cannot be determined without the hydroxide concentration, $[OH^-]$

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 $[TeO_3^{2-}]/[HTeO_3^-]$ 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) 2.4×10^{-3} % (B) 2.4×10^{-1} % (C) 3.5×10^{-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 pKa's of Histidine are: $pK_a' = 1.8$, $pK_a'' = 6.0$, and $pK_a''' = 9.2$.

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 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 $H_3PO_4(aq)$ is needed to completely neutralize 200 mL of an aqueous NaOH(aq) solution? What is the approximate Molarity of the NaOH(aq) solution?

    (A) 0.18 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.

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

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

$$PyrH^+ \rightleftarrows H^+ + Pyr.$$

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

$$\therefore [PyrH^+] = [PyrHBr] = \frac{(1.58 \times 10^{-3})^2}{5.56 \times 10^{-6}}$$

$$= 0.449 \text{ mol/L}$$

Assume 1L = 1000 mL

$$n_{PyrHBr} = 0.449 \text{ mol/L} \times 1L = 0.449 \text{ mol.}$$

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

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

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

$H_3PO_4 \rightleftharpoons H_2PO_4^- \rightleftharpoons HPO_4^{2-} \rightleftharpoons PO_4^{3-}$

pKa's 2.12 7.21 12.44

Version B

(20) 2. Phosphoric Acid (H_3PO_4) is a triprotic acid with acid dissociation constants, $K_a' = 7.5\times10^{-3}$, $K_a'' = 6.2\times10^{-8}$ and $K_a''' = 3.6\times10^{-13}$

(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 K_3PO_4.

Init

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

$n_{PO_4^{3-}} = 0.40 \times 0.85 = 0.34$ mol

	PO_4^{3-}	$+ H^+$	$\rightarrow HPO_4^{2-}$
init	0.34	0.21	0
chg	−0.21	−0.21	+0.21
	0.13 mol	~0	0.21

$$pH = 12.44 + \log\frac{[PO_4^{3-}]}{[HPO_4^{2-}]} = 12.44 + \log\frac{0.13}{0.21}$$

$$= \boxed{12.23}$$

(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 H_3PO_4.

Init

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

$n_{H_3PO_4} = 0.4 \times 0.65 = 0.26$

Step 1

	H_3PO_4	$+ OH^-$	$\rightarrow H_2PO_4^-$
init	0.26	0.45	0
chg	−0.26	−0.26	+0.26
	0	0.19	0.26

Step 2

	$H_2PO_4^-$	$+ OH^-$	$\rightarrow HPO_4^{2-}$
init	0.26	0.19	0
chg	−0.19	−0.19	+0.19
	0.07	~0	0.19

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

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

$$= \boxed{7.64}$$

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.

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

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

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

$= \boxed{0.29}$

(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.

$2H_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6H_2O$

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

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

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

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

$m\%(impur) = \frac{1.96}{7.80} \times 100 = 25.1\%$

$= \boxed{25\%}$