Average 70%

FIRST LETTER OF YOUR LAST NAME ☐

CHEMISTRY 1128 **EXAM II** **March 28, 2014**

Name ECE **Section** KEY

Signature ____________ **TA** ____________

ID # ____________

PLEASE READ THE FOLLOWING INSTRUCTIONS

Do NOT begin the exam until asked to do so.

There are 8 numbered pages including a table of equilibrium constants for weak acids and their conjugate bases, a periodic table and equations page in this exam. Check to see that they are all here before you begin the exam. Return all these papers when you are finished. Write your name on every page. Use a pen with blue or black ink for the entire exam.

Exams done in pencil, erasable ink, or where white-out, liquid paper, etc. have been used are *ineligible for regrades.*

Be sure to follow all directions. In working any numerical problem, you MUST SHOW ALL YOUR WORK. No credit will be given unless all work is clearly shown and the method of solution is logically correct. Pay attention to units and significant figures throughout.

Do not write below this line

Page	Total	Grader
1	______ / 26	
2	______ / 18	
3	______ / 10	
4	______ / 12	
5	______ / 23	
6	______ / 14	
7	______ / ~~31~~ 27	
8	______ / 20	

Tot al Grade ____________ **/150 Checked by** ______

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I. (54 points)

A. (10 points) Consider the following six beakers. All have 100 mL of aqueous 0.1 M solutions of the following compounds.

Beaker A has HI
Beaker B has HNO_2
Beaker C has NaOH
Beaker D has $Ba(OH)_2$
Beaker E NH_4Cl
Beaker F $C_2H_5NH_2$

Answer the questions below, using **LT** (for *is less than*), **GT** (for *is greater than*), **EQ** (for *is equal to*), or **MI** (for *more information required*).

2pts each

GT **1.** The pH of beaker B ________ the pH in beaker A.

LT **2.** The pH in beaker C ________ the pH in beaker D.

GT **3.** The pH in beaker C ________ the pH in beaker F.

LT **4.** The pH in beaker B ________ the pH in beaker E. (or more info, if the table is not provided)

EQ **5.** Percent ionization in beaker A ________ the % ionization in beaker C.

B. (8 points) What is the pH of a solution obtained by diluting 15.0 mL of 0.115 M solution of $Sr(OH)_2$ to 285 mL of solution?

8pts

$$\# \text{ of moles of } OH^- = 15.0\,mL\ Sr(OH)_2 \times \frac{0.115\ mol}{1000\,mL} \times \frac{2\,mol\ OH^-}{1\,mol\ Sr(OH)_2}$$

$$= 3.45 \times 10^{-3}\ mol \quad (2pts)$$

$$[OH^-] = \frac{3.45 \times 10^{-3}}{0.285} = 0.0121 \Rightarrow (2pts)$$

$$pOH = 1.92 \qquad pH = 12.08 \quad (2pts)$$

12.08 ⇐ (2pts)

C. (8 points) A 4.90 x 10^{-3} mol sample of a weak acid HY is dissolved in enough water. To form 0.095 L of solution. If the pH of the solution is 2.65, what is K_a for the weak acid?

8pts

$$[HY] = \frac{4.90 \times 10^{-3}}{0.095} = 0.052\,M \quad (2pts)$$

$$pH = 2.65 \qquad [H^+] = 2.24 \times 10^{-3} \quad (2pts)$$

$$K_a = \frac{(2.24 \times 10^{-3})^2}{0.052}$$

9.7×10^{-5}

2pts for squaring
2pts for set-up

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D. **(10 points)** Ascorbic acid ($HC_6H_7O_6$, MM = 176.124 g/mol) is also known as vitamin C, and it is a weak acid. A solution of ascorbic acid is prepared by dissolving 2.00 g in enough water to make 100.0 mL solution. If the K_a of ascorbic acid is 7.6×10^{-5} calculate the following

1. The concentration of $[H^+]$ in solution

# of mol of HA = 2.00g HA × $\frac{1 \text{ mol}}{176.124 \text{ g}}$ = 0.0114 mol (2pts)

$[HA] = \frac{0.0114}{0.1000} = 0.114\ M$ (2pts)

$7.6 \times 10^{-5} = \frac{[H^+]^2}{0.114}$ 2pts for set-up

(8pts) $[H^+] = 2.9 \times 10^{-3}\ M$

2.9×10^{-3} 2pts answer

2. The pH of the solution.

$pH = -\log(2.9 \times 10^{-3})$ based on answer in # 1

$= 2.53$

(2pts)

All or nothing

2.53

E. **(8 points)** State whether a 1 M solution of the following salts in water will be acidic, basic, or neutral.

(2pts each)

basic **1.** K_3PO_4

basic **2.** NH_4HCO_3 $K_a = 5.6 \times 10^{-10}$ $K_b = 2.3 \times 10^{-8}$

neutral **3.** $NaClO_4$

acidic **4.** LiH_2PO_4

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F. **(10 points)** Thiamine hydrochloride ($C_{12}H_{18}ON_4SCl_2$) is a water soluble form of thiamine. Its acid dissociation constant is 3.37 X 10^{-7}. A solution of thiamine hydrochloride has a pH of 3.50. How many grams of thiamine hydrochloride are needed in a 15.00 mL solution to give the same pH? (MM of $HC_{12}H_{17}ON_4SCl_2$ = 337.27g/mol)

$[H^+] = 10^{-3.50} = 3.16 \times 10^{-4}\,M$ 2pts

$$3.37 \times 10^{-7} = \frac{(3.16 \times 10^{-4})^2}{X - 3.16 \times 10^{-4}}$$ 2 pts for set-up

10 pts

$3.37 \times 10^{-7} x - 1.1 \times 10^{-10} = 1.0 \times 10^{-7}$

$3.37 \times 10^{-7} x = 1.0 \times 10^{-7}$

$x = 0.297\ M$ 2pts

mass of thiamine $= 15.00\,mL \times \frac{0.297\,mol}{1000\,mL} \times \frac{337.29\,g}{1\,mol}$

$= 1.50\,g$ 2pts 2pts

2pts for # of moles

2pts for mass

1.50 g

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II. (49 points)

A. (12 points) Answer questions 1 – 4 below, using **LT** (for *is less than*), **GT** (for *is greater than*), **EQ** (for *is equal to*), or **MI** (for *more information required*)

Beaker A has a weak acid ($K_a = 1 \times 10^{-5}$)
Beaker B has HCl

The volume and molarity of each acid in the beakers are the same. Both acids are to be titrated with a 0.1 M solutions of NaOH.

GT **1.** Before the titration starts, the pH of the solution is Beaker A is ________ the pH of the solution in beaker B.

MI **2.** At half neutralization (halfway to the equivalence point), the pH of the solution in Beaker A ________ the pH of the solution in beaker B.

GT **3.** When each solution has reached the equivalence point, the pH of the solution in Beaker A ________ the pH of the solution in Beaker B.

EQ **4.** At the equivalence point, the volume of NaOH used to titrate HCl in Beaker B ________ the volume of NaOH used to titrate the weak acid in beaker A.

Place your answer on the line provided.

C **5.** A buffer may be prepared from a weak acid. Which of the acids below is best for the preparation of a buffer with a pH of 4.00?

(A) sulfurous acid , H_2SO_3; $K_a = 1.7 \times 10^{-2}$
(B) hydrofluoric acid, HF; $K_a = 6.9 \times 10^{-4}$
(C) benzoic acid, $HC_7H_5O_2$; $K_a = 6.6 \times 10^{-5}$
(D) dihydrogen phosphate ion, $H_2PO_4^-$; $K_a = 6.2 \times 10^{-8}$
(E) ammonium ion, NH_4^+; $K_a = 5.6 \times 10^{-10}$

2 pts each

A **6.** All of the following statements are false for the titration of a strong base by a strong acid **EXCEPT**

(A) the equivalence point occurs at pH 7.
(B) there will be at least two equivalence points, one for the base and one for the acid.
(C) the equivalence point and the end point are identical, regardless of what indicator is used.
(D) equal masses of acid and base are required to neutralize the solution.
(E) it is impossible to neutralize a strong acid with a strong base.

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B. (8 points) If added to 750.0 mL of 0.243 M HOCl, which of these would form a buffer? On the line provided write **YES** or **NO**.

2pts each

1. YES 0.10 mol NaOCl
2. NO 0.10 mol HNO_3
3. YES 0.10 mol KOH
4. NO 0.20 mol KOH

# of mol of HOCl = 750.0 mL × $\frac{0.243}{1000\text{mL}}$
= 0.182 mol

$[Ben^-] = \frac{0.104}{0.425}$

= 0.245 M

OR

C. (15 points) A buffer solution is prepared by dissolving 15.00 g of sodium benzoate ($NaC_7H_5O_2$, MM = 144.10) in 425 mL of 0.249 M $HC_7H_5O_2$. Assume not volume change after $NaC_7H_5O_2$ is dissolved. K_a of $HC_7H_5O_2 = 6.6 \times 10^{-5}$.

1. Calculate the pH of this buffer.

# of mol of $C_7H_5O_2^-$ = 15.00g $NaC_7H_7O_2$ × $\frac{1\text{mol}}{144.1\text{g}}$ × $\frac{Ben^-}{1\text{mol NaBen}}$

= 0.104 mol (2pts)

# of mol of $HC_7H_5O_2$ = 425 mL × $\frac{0.249\text{mol}}{1000\text{mL}}$ = 0.106 mol (2pts)

6pts

$6.6 \times 10^{-5} = \frac{[H^+]0.104}{0.106}$

$[H^+] = 6.7 \times 10^{-5}$ M

pH = $-\log 6.7 \times 10^{-5}$
= 4.17

4.17 (2pts)

2. Calculate the pH of the solution after 0.0255 moles HCl is added to this buffer. Assume no volume change.

9pts

	$Ben^-(aq)$ +	$H^+(aq)$ →	$HBen(aq)$
before	0.104	0.0255	0.106
	−0.0255	−0.0255	+0.0255
after	0.0785	0	0.1315
	3pts		3pts

$6.6 \times 10^{-5} = \frac{[H^+]0.0785}{0.1315}$

$[H^+] = 1.1 \times 10^{-4}$

pH = 3.96

3.96 (3pts)

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D. (14 points) A 25.0 mL solution of lactic acid, $HC_3H_5O_3$, is titrated with 15.00 mL of 0.186 M $Ba(OH)_2$. K_a of lactic acid is 1.38×10^{-4}.

1. Write a balanced net ionic equation

(2pts)

$HC_3H_5O_3(aq) + OH^-(aq) \rightarrow C_3H_5O_3^-(aq) + H_2O$

2. What are the species present at the equivalence point?

(2pts)

(1pt) (1pt)

$C_3H_5O_3^-$, Ba^{2+}

having H_2O is OK ✓

For extra species -1

3. What is the molarity of the initial lactic solution?

(4pts)

$$15.00 mL\ Ba(OH)_2 \times \frac{0.186 mol\ Ba(OH)_2}{1000 mL\ Ba(OH)_2} \times \frac{2 mol\ OH^-}{1 mol\ Ba(OH)_2} \times \frac{1 mol\ HLac}{1 mol\ OH^-} = 5.58 \times 10^{-3} mol$$

(2pts)

$$[HLac] = \frac{5.58 \times 10^{-3}}{0.0250}$$

0.223 M (2pts)

4. What is the pH of the solution at the equivalence point?

(6pts)

$$[Lac^-] = \frac{5.58 \times 10^{-3}}{0.0400} = 0.1395 M$$ (2pts)

$$\frac{1.0 \times 10^{-14}}{1.38 \times 10^{-4}} = 7.25 \times 10^{-11} = \frac{[OH^-]^2}{0.1395}$$

$$[OH^-] = 3.18 \times 10^{-6} M$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{3.18 \times 10^{-6}}$$

$$= 3.14 \times 10^{-9} M$$

(4pts)

1 mist. 2
2 mist. 0

pH = 8.50

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III. (47 points)

A. (8 points) Predict whether the **forward** or **reverse** reaction or **no change** will occur when the equilibrium is disturbed by

$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$ $\Delta H = 23.4$ kJ

2 pts each

reverse **1.** addition of 1 M $Pb(NO_3)_2$ solution
forward **2.** increase in temperature.
forward **3.** addition of Ag^+, forming AgCl.
no change **4.** Addition of 1 M HCl.

B. (8 points) The solubility of an ionic compound M_2Y_3, (MM of M_2Y_3 = 245.0 g/mol) is 3.6×10^{-7} g/L. What is the solubility product constant, K_{sp}, for this compound?

8 pts

$$s \text{ (in mol/L)} = 3.6 \times 10^{-7} \frac{g}{L} \times \frac{1 \text{ mol } M_2Y_3}{245.0 g} = 1.5 \times 10^{-9} \text{ mol}$$

(2 pts)

$$M_2Y_3(s) \rightleftharpoons 2M^{3+}(aq) + 3Y^{2-}(aq)$$

$$Ksp = [M^{3+}]^2[Y^{2-}]^3 = \underbrace{(2s)^2}_{\text{2 pts}}\underbrace{(3s)^3}_{\text{2 pts}} = 108 \times (1.5 \times 10^{-9})^5$$

$$= 7.4 \times 10^{-43}$$

7.4×10^{-43}

C. (6 points) What is the solubility of solid $Ca_3(PO_4)_2$ ($K_{sp} = 1.3 \times 10^{-32}$) in a 0.20 M Na_3PO_4 solution.

6 pts

$$1.3 \times 10^{-32} = [Ca^{2+}]^3[PO_4^{3-}]^2$$

1 mist. 3
2 mist. 0

$$= [Ca^{2+}]^3(0.20)^2$$

$$[Ca^{2+}] = 6.9 \times 10^{-11} \text{ M} = 3s$$

$$s = 2.3 \times 10^{-11} \text{ mol/L}$$

2.3×10^{-11} mol/L

D. (5 points) A solution containing potassium bromide is mixed with a solution containing lead acetate to form a solution that is 0.013 M in KBr and 0.0035 M in $Pb(C_2H_3O_2)_2$. Will a precipitate form? K_{sp} for $PbBr_2$ is 4.0×10^{-5}. **(SHOW YOUR CALCULATION)**

5 pts

$$Q = (0.013)^2(0.0035)$$

$$= 5.9 \times 10^{-7} < 4.0 \times 10^{-5}$$

no ppt

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E. (8 points) Consider the following equation:

$Cu(OH)_2(s) + 4NH_3\,(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq) + 2OH^-(aq)$

K_{sp} for $Cu(OH)_2 = 2 \times 10^{-19}$ and K_f for $Cu(NH_3)_4^{2+} = 2 \times 10^{12}$

1. What is the equilibrium constant, K, for the reaction given above?

2 pts All or nothing

$2 \times 10^{-19} \times 2 \times 10^{12} = 4 \times 10^{-7}$

2. What is $[NH_3]$ at equilibrium when 1.68 g of copper(II) hydroxide (97.561 g/mol) are dissolved in 500.0 mL of solution?

\# of mol of $Cu(OH)_2 = 1.68\,g\ Cu(OH)_2 \times \frac{1\,mol\ Cu(OH)_2}{97.561\,g}$

$= 0.0172\ mol$

6 pts

$S = \frac{0.0172}{0.5000} = 0.0344$ ---→ 3 pts

$4 \times 10^{-7} = \frac{4 \times (0.0344)^3}{[NH_3]^4}$ 3 pts All or nothing

4.5 M

F. (12 points) Solid Na_2CrO_4 is slowly added to a solution that is 0.0100 M in Pb^{2+} and 0.0062 M Ag^+. K_{sp} of $PbCrO_4 = 2.0 \times 10^{-14}$ K_{sp} of $Ag_2CrO_4 = 1.0 \times 10^{-12}$

1. Which compound will begin to precipitate first? Determine the concentration of CrO_4^{2-} to precipitate each cation?

7 pts

$1.0 \times 10^{-12} = (0.0062)^2[CrO_4^{2-}] \Rightarrow [CrO_4^{2-}] = 2.6 \times 10^{-8}$

$2.0 \times 10^{-14} = (0.0100)[CrO_4^{2-}] \Rightarrow [CrO_4^{2-}] = 2.0 \times 10^{-12}$

$Pb^{2+} \Leftarrow$ 2 pts

3 pts

$[CrO_4^{2-}]$ to ppt Ag_2CrO_4 = 2.6×10^{-8} $[CrO_4^{2-}]$ to ppt $PbCrO_4$ = 2.0×10^{-12} 2 pts

2. What percentage of first cation precipitated when the second cation starts to precipitate?

$1.0 \times 10^{-12} = [Pb^{2+}](2.6 \times 10^{-8})$

$[Pb^{2+}] = 3.8 \times 10^{-5}\ M$ 2 pts

5 pts

$\frac{3.8 \times 10^{-5}}{0.0100} \times 100 = 0.38\%$

99.62 % 3 pts

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Equilibrium Constants for Weak Acids and Their Conjugate Bases

	Acid	**K_a**	**Base**	**K_b**
Sulfurous acid	H_2SO_3	1.7×10^{-2}	HSO_3^-	5.9×10^{-13}
Hydrogen sulfate ion	HSO_4^-	1.0×10^{-2}	SO_4^{2-}	1.0×10^{-12}
Phosphoric acid	H_3PO_4	7.1×10^{-3}	$H_2PO_4^-$	1.4×10^{-12}
Hexaaquairon(III) ion	$Fe(H_2O)_6^{3+}$	6.7×10^{-3}	$[Fe(H_2O)_5OH]^{2+}$	1.5×10^{-12}
Hydrofluoric acid	HF	6.9×10^{-4}	F^-	1.4×10^{-11}
Nitrous acid	HNO_2	6.0×10^{-4}	NO_2^-	1.7×10^{-11}
Formic acid	$HCHO_2$	1.9×10^{-4}	CHO_2^-	5.3×10^{-11}
Lactic acid	$HC_3H_5O_3$	1.4×10^{-4}	$C_3H_5O_3^-$	7.1×10^{-11}
Benzoic acid	$HC_7H_5O_2$	6.6×10^{-5}	$C_7H_5O_2^-$	1.5×10^{-10}
Acetic acid	$HC_2H_3O_2$	1.8×10^{-5}	$C_2H_3O_2^-$	5.6×10^{-10}
Hexaaquaaluminum (III) ion	$A\ell(H_2O)_6^{3+}$	1.2×10^{-5}	$[Al(H_2O)_5OH]^{2+}$	$8.\ 3 \times 10^{-10}$
Carbonic acid	H_2CO_3	4.4×10^{-7}	HCO_3^-	2.3×10^{-8}
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}	HPO_4^{2-}	1.6×10^{-7}
Hydrogen sulfite ion	HSO_3^-	6.0×10^{-8}	SO_3^{2-}	1.7×10^{-7}
Hypochlorous acid	$HC\ell O$	2.8×10^{-8}	$C\ell O^-$	3.6×10^{-7}
Hydrocyanic acid	HCN	5.8×10^{-10}	CN^-	1.7×10^{-5}
Ammonium ion	NH_4^+	5.6×10^{-10}	NH_3	1.8×10^{-5}
Tetraaquazinc (II) ion	$Zn(H_2O)_4^{2+}$	3.3×10^{-10}	$[Zn(H_2O)_3OH]^+$	3.0×10^{-5}
Hydrogen carbonate ion	HCO_3^-	4.7×10^{-11}	CO_3^{2-}	2.1×10^{-4}
Hydrogen phosphate ion	HPO_4^{2-}	$4.5\ 10^{-13}$	PO_4^{3-}	2.2×10^{-2}