

β

Name Key  
(Please print family name last; e.g., Robert Boyle)

UMB Student Number \_\_\_\_\_

Chem 104 - Section 1  
Hour Examination III  
May 5, 2006

This test consists of seven (7) pages, including this cover page, a table of conjugate acid-base pairs with  $K_a$  values, and a periodic table. Be sure your copy is complete before beginning your work. If this test packet is defective, ask for another one. **Feel free to detach the acid-base table and/or periodic table to use for reference or scratch paper.**

**Give all numerical answers to the proper number of significant figures.**

$$K_w = 1.00 \times 10^{-14}$$

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DO NOT WRITE BELOW THIS LINE

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1.

2.

3 a - e

3 f & g  
+ bonus

TOTAL

β

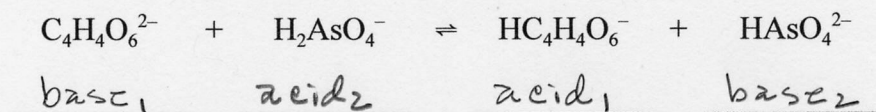
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1. (12 points; 3 points each part) Complete the following table by calculating the missing entries and indicating whether the solution is acidic or basic.

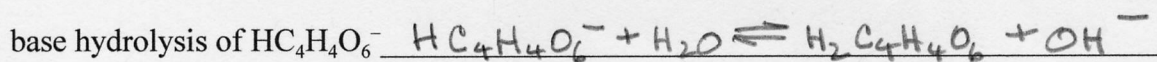
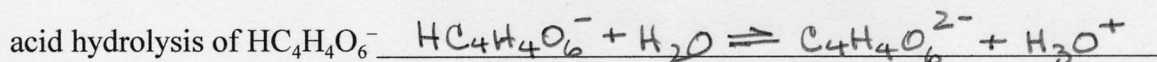
$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	pOH	acidic/basic?
$4.2 \times 10^{-6}$	$2.4 \times 10^{-9} \text{ M}$	5.38	8.62	acidic

2. (44 points; 4 points each part) Fill in the blanks.

- a. For the following reaction, label the conjugate acid-base pairs (i.e., acid<sub>1</sub>/base<sub>1</sub>; acid<sub>2</sub>/base<sub>2</sub>).



- b. Write a balanced chemical equation for each of the following equilibria:



- c. For the acid hydrolysis of tartaric acid,  $\text{H}_2\text{C}_4\text{H}_4\text{O}_6$ ,  $K_a = 1.0 \times 10^{-3}$ . What is the value of  $K_b$  for the base hydrolysis of the hydrogen tartrate ion,  $\text{HC}_4\text{H}_4\text{O}_6^-$ ?

$$K_b = \underline{1.0 \times 10^{-11}} \text{ for } \text{HC}_4\text{H}_4\text{O}_6^-$$

- d. For the acid hydrolysis of the hydrogen tartrate ion,  $\text{HC}_4\text{H}_4\text{O}_6^-$ ,  $K_a = 4.6 \times 10^{-5}$ . Judging from this and the  $K_b$  value you just calculated in question c, would a 0.10 M solution of  $\text{NaHC}_4\text{H}_4\text{O}_6(aq)$  be acidic or basic?

acidic (acidic/basic)

- e. Consider a 0.10 M solution of the diprotic acid  $\text{H}_2\text{C}_4\text{H}_4\text{O}_6(aq)$ , for which  $K_1 = 1.0 \times 10^{-3}$  and  $K_2 = 4.6 \times 10^{-5}$ . What is the concentration of  $\text{C}_4\text{H}_4\text{O}_6^{2-}$  ion in this solution?

$$[\text{C}_4\text{H}_4\text{O}_6^{2-}] = \underline{4.6 \times 10^{-5}} \text{ M}$$

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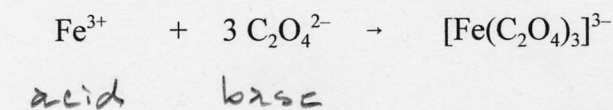
- f. Refer to the Table of Conjugate Acid-Base Pairs. Which one of the following solutions when added in excess to a solution containing 1.0 mmol  $\text{Na}_3\text{PO}_4(\text{aq})$  would produce 1.0 mmol  $\text{H}_2\text{PO}_4^-$  ion in solution:  $\text{HOI}(\text{aq})$ ,  $\text{Al}(\text{NO}_3)_3(\text{aq})$ ,  $\text{NH}_4\text{NO}_3(\text{aq})$ ?

Answer:  $\text{Al}(\text{NO}_3)_3$ 

- g. Assuming equal concentrations, which one of the following pairs would produce the more acidic solution?

 $\text{HClO}_3(\text{aq})$  or  $\text{HClO}_2(\text{aq})$   $\text{HClO}_3$  $\text{C}_6\text{H}_5\text{CO}_2\text{H}(\text{aq})$  or  $\text{ClC}_6\text{H}_4\text{CO}_2\text{H}(\text{aq})$   $\text{ClC}_6\text{H}_4\text{CO}_2\text{H}$ 

- h. Identify the Lewis acid and Lewis base in the following reaction



- i. Consider a 0.10 M solution of the weak acid HA, for which  $\text{p}K_a = 4.58$ . Would the expression  $[\text{H}_3\text{O}^+] = \sqrt{C_{\text{HA}}K_a}$  give a reasonably accurate estimate of the hydronium ion concentration (less than 5% error)?  $K_a = 2.6 \times 10^{-5}$

Answer: yes (yes/no)

- j. The base B has  $K_b = 1.0 \times 10^{-8}$ . What is the pH of a buffer solution prepared by mixing 1.0 mol B with 1.0 mol of the salt  $\text{HBCl}$  in enough water to make a liter of solution?

pH = 6.00

- k. Indicate whether 0.10 M aqueous solutions of each of the following solutions would have a pH > 7.0, pH < 7.0, or pH ≈ 7.0:

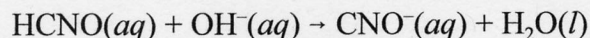
 $\text{Fe}(\text{NO}_3)_3$  < 7.0                   $\text{Ba}(\text{NO}_2)_2$  > 7.0



B

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3. (44 points)  $K_a = 3.5 \times 10^{-4}$  for cyanic acid, HCNO. Consider the titration of 25.0 mL of 0.160 M HCNO solution (the analyte), with 0.100 M NaOH(aq) solution (the titrant):



- a. (4 points) How many milliliters of 0.100 M NaOH(aq) solution must be added to reach the equivalence point?

$$V_b = \frac{V_a M_a}{M_b} = \frac{(25.0 \text{ mL})(0.160 \text{ M})}{0.100 \text{ M}} = 40.0 \text{ mL}$$

- b. (2 points) What is the total volume in the solution at the equivalence point? 65.0 mL

- c. (2 points) How many millimoles of HCNO are present in the analyte sample before any titrant has been added?

$$\text{millimoles HCNO} = \underline{4.00}$$

- d. (6 points) What is the initial pH of the HCNO solution, before adding any titrant?

$$[\text{H}_3\text{O}^+] = \sqrt{(0.160)(3.5 \times 10^{-4})} = \sqrt{5.6 \times 10^{-5}} = 7.48 \times 10^{-3}$$

$$\text{pH} = 2.1259 = 2.13$$

- e. (6 points) What is the pH of the resulting solution after adding 20.0 mL of 0.100 M NaOH(aq) solution? [Hint: How far along in the titration is this?]

Half-titration point  $\Rightarrow$  equimolar buffer solution

$$\text{pH} = \text{p}K_a = -\log(3.5 \times 10^{-4}) = 3.4559 = 3.46$$

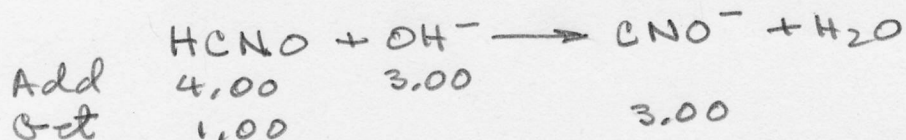
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- f. (10 points) What is the pH of the resulting solution after adding 30.0 mL of 0.100 M NaOH(aq) solution?

$$\text{mmol OH}^- \text{ added} = (30.0 \text{ mL})(0.100 \text{ M}) = 3.00 \text{ mmol}$$



$$K_2 = \frac{[\text{H}_3\text{O}^+][\text{CNO}^-]}{[\text{HCNO}]} = 3.5 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+](3.00)}{1.00}$$

$$[\text{H}_3\text{O}^+] = \frac{(1.00)(3.5 \times 10^{-4})}{3.00} = 1.167 \times 10^{-4} \Rightarrow \text{pH} = 3.93$$

- g. (14 points) What is the pH at the equivalence point?

All HCNO converted to 4.00 mmol CNO<sup>-</sup> in 65.0 mL.

$$[\text{CNO}^-] = \frac{4.00 \text{ mmol}}{65.0 \text{ mL}} = 6.1538 \times 10^{-2} \text{ M}$$

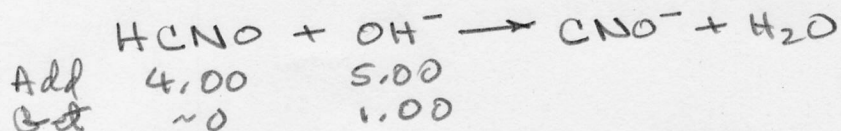
$$K_b^{\text{CNO}^-} = \frac{K_w}{K_a^{\text{HCNO}}} = \frac{1.0 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.857 \times 10^{-11} = 2.9 \times 10^{-11}$$

$$[\text{OH}^-] = \sqrt{(6.1538 \times 10^{-2} \text{ M})(2.857 \times 10^{-11})} = \sqrt{1.758 \times 10^{-12}} = 1.326 \times 10^{-6}$$

$$\text{pOH} = 5.88 \Rightarrow \text{pH} = 8.12$$

- BONUS (5 points) What is the pH of the solution after the addition of 50.0 mL of 0.100 M NaOH solution?

$$\text{mmol OH}^- \text{ added} = (50.0 \text{ mL})(0.100 \text{ M}) = 5.00 \text{ mmol}$$



$$V = 75.0 \text{ mL}$$

$$[\text{OH}^-] = \frac{1.00 \text{ mmol}}{75.0 \text{ mL}} = 1.33 \times 10^{-2} \Rightarrow \text{pOH} = 1.875 \Rightarrow \text{pH} = 12.125$$