

Please print:

Last name: Answer Key

First name: _____

Chem 1062 Exam 3

Spring 2005

Andy Aspaas, Instructor

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Equations:

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad x \text{ is sufficiently small if } \frac{c_a}{K_a} > 100$$
$$pH = -\log[H_3O^+] \quad K_b = \frac{[HB^+][OH^-]}{[B]} \quad ax^2 + bx + c = 0$$
$$pOH = -\log[OH^-] \quad pH + pOH = 14.00 \quad x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$
$$K_w = [H_3O^+][OH^-] \quad pH = pK_a + \log \frac{[A^-]}{[HA]}$$
$$K_w = K_a K_b = 1.0 \times 10^{-14}$$

Tables of K_a and K_b values are on page 2.

Contents:

I. Multiple choice , 5 questions, 3 points each.	15 points
II. Short answer , 10 questions, worth 25 points.	25 points
III. Problems , 3 questions, worth 20 points each.	<u>60 points</u>
Total: 100 points	

I, _____ have read and understand the directions given above, and pledge that I will follow all regulations with regard to Academic Dishonesty as outlined by this college when taking this exam.

Signature _____ Date and Time _____

Acid-Ionization Constants at 25 °C

Substance	Formula	K_a
Acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$	1.7×10^{-5}
Benzoic acid	$\text{HC}_7\text{H}_5\text{O}_2$	6.3×10^{-5}
Boric acid	H_3BO_3	5.9×10^{-10}
Carbonic acid	H_2CO_3	4.3×10^{-7}
	HCO_3^-	4.8×10^{-11}
Cyanic acid	HOCN	3.5×10^{-4}
Formic acid	HCHO_2	1.7×10^{-4}
Hydrocyanic acid	HCN	4.9×10^{-10}
Hydrofluoric acid	HF	6.8×10^{-4}
Hydrogen sulfate ion	HSO_4^-	1.1×10^{-2}
Hydrogen sulfide	H_2S	8.9×10^{-8}
	HS^-	1.2×10^{-13}
Hypochlorous acid	HClO	3.5×10^{-8}
Nitrous acid	HNO_2	4.5×10^{-4}
Oxalic acid	$\text{H}_2\text{C}_2\text{O}_4$	5.6×10^{-2}
	HC_2O_4^-	5.1×10^{-5}
Phosphoric acid	H_3PO_4	6.9×10^{-3}
	H_2PO_4^-	6.2×10^{-8}
	HPO_4^{2-}	4.8×10^{-13}
Phosphorous acid	H_2PHO_3	1.6×10^{-2}
	H_2PHO_3^-	7.0×10^{-7}
Propionic acid	$\text{HC}_3\text{H}_5\text{O}_2$	1.3×10^{-5}
Pyruvic acid	$\text{HC}_3\text{H}_3\text{O}_3$	1.4×10^{-4}
Sulfurous acid	H_2SO_3	1.3×10^{-2}
	HSO_3^-	6.3×10^{-8}

Base-Ionization Constants at 25 °C

Substance	Formula	K_b
Ammonia	NH_3	1.8×10^{-5}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2$	4.2×10^{-10}
Dimethylamine	$(\text{CH}_3)_2\text{NH}$	5.1×10^{-4}
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2$	4.7×10^{-4}
Hydrazine	N_2H_4	1.7×10^{-6}
Hydroxylamine	NH_2OH	1.1×10^{-8}
Methylamine	CH_3NH_2	4.4×10^{-4}
Pyridine	$\text{C}_5\text{H}_5\text{N}$	1.4×10^{-9}
Urea	NH_2CONH_2	1.5×10^{-14}

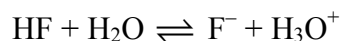
I. Multiple choice

Choose the best answer from the choices given, and clearly **circle** the letter of your choice. You may need to consult the tables on page 2 for some answers. (3 pts each)

1. Which of the following is the **strongest** acid in water?

- a. HBrO
- b. HBrO₂
- c. HBrO₃
- d. HBrO₄
- e. Not enough information to say.

2. For the following equilibrium, which of the following statements is true?



- a. The equilibrium lies predominately to the right.
- b. The equilibrium lies predominately to the left.
- c. This isn't an equilibrium at all, since HF is a strong acid.
- d. This reaction will not occur at all.
- e. None of these statements are true.

3. Which of the following is a diprotic acid?

- a. HC₂H₃O₂
- b. H₂CO₃
- c. HCHO₂
- d. C₆H₅NH₂
- e. HC₃H₃O₃

4. Which of the following solutions will have the highest pH?

- a. 0.1 M Na₂CO₃
- b. 0.1 M KCl
- c. 0.1 M NaNO₃
- d. 0.1 M NaCl
- e. 0.1 M NH₄Cl

5. Which of the following solution mixtures is a buffer?

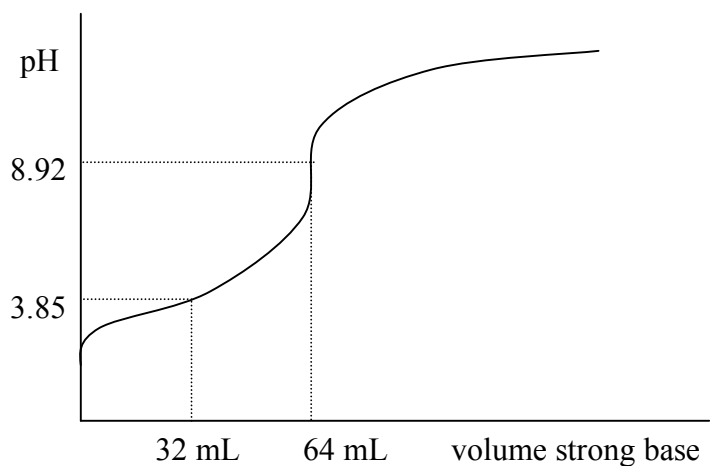
- a. HCl / NaCl
- b. KH₂PO₄ / KNO₃
- c. HC₂H₃O₂ / NaC₂H₃O₂
- d. HOCN / NaOH
- e. HOCN / HCN

II. Conceptual

Predict whether aqueous solutions of the following salts will be acidic, basic, or neutral.
(3 pts each)

- | | |
|--------------------------------------|----------------|
| 6. Na_2S | <u>basic</u> |
| 7. $\text{Cu}(\text{NO}_3)_2$ | <u>acidic</u> |
| 8. KClO_4 | <u>neutral</u> |
| 9. $\text{CH}_3\text{NH}_3\text{Cl}$ | <u>acidic</u> |
| 10. NH_4F | <u>acidic</u> |

An unidentified acid is titrated by a strong base, and its titration curve is shown below:



11. (3 pts) Is the acid strong or weak? weak
12. (3 pts) Is the acid monoprotic, diprotic, or triprotic? monoprotic
13. (4 pts) The identity of the acid can be determined from the values on this curve. Which acid is it?

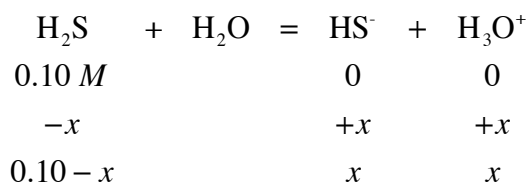
pH at halfway point = pKa of acid = 3.85; $K_a = 1.41 \times 10^{-4}$ (Pyruvic acid)

III. Problems

Choose 3 of the following problems to complete. If you answer more than three, only the first three will be graded. (20 points each)

I would like questions numbered _____, _____, and _____ to be graded.

14. a. Calculate $[\text{H}_3\text{O}^+]$, $[\text{HS}^-]$, and pH for a 0.10 M H_2S solution.



$$K_a = 8.9 \times 10^{-8} = \frac{x^2}{0.10 - x}$$

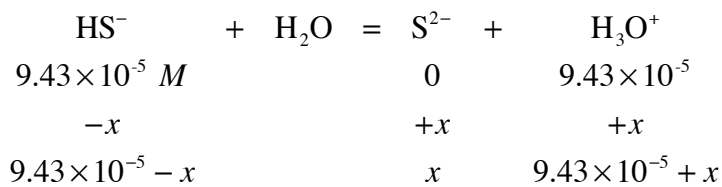
$$\frac{c_a}{K_a} = \frac{0.10}{8.9 \times 10^{-8}} \gg 100$$

$$8.9 \times 10^{-8} \cong \frac{x^2}{0.10}$$

$$x = \sqrt{(0.10)(8.9 \times 10^{-8})} = 9.43 \times 10^{-5} \text{ M} = [\text{H}_3\text{O}^+] = [\text{HS}^-]$$

$$\text{pH} = -\log(9.43 \times 10^{-5}) = 4.03$$

- b. Calculate $[\text{S}^{2-}]$ in this solution.



$$K_a = 1.2 \times 10^{-13} = \frac{x(9.43 \times 10^{-5} + x)}{9.43 \times 10^{-5} - x}$$

$$\frac{c_a}{K_a} = \frac{9.43 \times 10^{-5}}{1.2 \times 10^{-13}} \gg 100$$

$$1.2 \times 10^{-13} = \frac{x(9.43 \times 10^{-5})}{9.43 \times 10^{-5}} = x$$

$$[\text{S}^{2-}] = 1.2 \times 10^{-13} \text{ M}$$

15. a. Use the quadratic equation method to calculate the pH of a 0.0015 M solution of NaCN.

$$\begin{array}{rcccl}
 \text{CN}^- & + & \text{H}_2\text{O} & = & \text{HCN} & + & \text{OH}^- \\
 0.0015 \text{ M} & & & & 0 & & 0 \\
 -x & & & & +x & & +x \\
 0.0015 - x & & & & x & & x \\
 K_b = \frac{K_w}{K_{a, \text{HCN}}} = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.04 \times 10^{-5} = \frac{x^2}{0.0015 - x} \\
 \frac{c_b}{K_b} = \frac{0.0015}{2.04 \times 10^{-5}} = 73.5 < 100 \\
 (0.0015 - x)(2.04 \times 10^{-5}) = x^2 \\
 x^2 + 2.04 \times 10^{-5}x - 3.06 \times 10^{-8} = 0 \\
 x = \frac{-2.04 \times 10^{-5} \pm \sqrt{(2.04 \times 10^{-5})^2 + 4(3.06 \times 10^{-8})}}{2} \\
 x = 1.65 \times 10^{-4} = [\text{OH}^-] \\
 \text{pOH} = 3.78 \\
 \text{pH} = 10.22
 \end{array}$$

- b. Perform the same calculation, this time using the approximation method.

$$\begin{array}{l}
 2.04 \times 10^{-5} = \frac{x^2}{0.0015} \\
 x = \sqrt{(0.0015)(2.04 \times 10^{-5})} = 1.749 \times 10^{-4} \\
 \text{pOH} = 3.76 \\
 \text{pH} = 10.24
 \end{array}$$

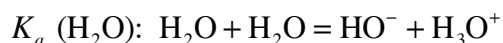
- c. Do the answers differ? Explain why or why not, in terms of the ionization and equilibrium that is occurring here.

Yes, the answers are slightly different. By assuming the equilibrium CN^- concentration is equal to the initial CN^- concentration, we're dividing by a larger number and x becomes slightly larger. Since $c_b / K_b < 100$, this assumption cannot be made.

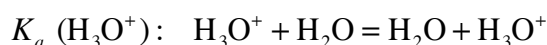
16. a. Calculate the molar concentration of pure water, assuming its density is 1.00 g/mL.

$$\frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ mol}}{18.016 \text{ g}} = 55.51 \text{ mol/L}$$

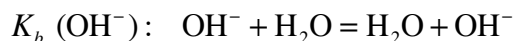
- b. Use this value along with the value of K_w to calculate the K_a of H_2O , the K_a of H_3O^+ , and the K_b of OH^- . (Hint: no ICE tables are required!)



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = \frac{K_w}{[\text{H}_2\text{O}]} = \frac{1.0 \times 10^{-14}}{55.51} = 1.80 \times 10^{-16}$$



$$K_a = \frac{[\text{H}_2\text{O}][\text{H}_3\text{O}^+]}{[\text{H}_3\text{O}^+]} = [\text{H}_2\text{O}] = 55.51 = 5.551 \times 10^1$$



$$K_b = \frac{[\text{H}_2\text{O}][\text{OH}^-]}{[\text{OH}^-]} = [\text{H}_2\text{O}] = 5.551 \times 10^1$$

- c. How do the K_a values you calculated compare in size to the K_a values on the table on page 2? What does this mean about the relative strengths of H_3O^+ and H_2O as acids?

K_a of H_2O is much smaller than any other K_a on the table, therefore water is a weak acid.

K_a of H_3O^+ is much larger than any other K_a on the table, therefore hydronium is a relatively strong acid.

K_b of OH^- is much larger than any other K_b on the table, therefore hydroxide is a relatively strong base.

17. A buffer solution with a total volume of 100.0 mL is 0.100 *M* in acetic acid and 0.100 *M* in sodium acetate.

- a. Calculate the pH of this solution.

$$\frac{c_a}{K_a} = \frac{0.100}{1.7 \times 10^{-5}} \gg 100$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$\text{p}K_a = -\log(1.7 \times 10^{-5}) = 4.77$$

$$\text{pH} = 4.77 + \log \frac{.100}{.100} = 4.77$$

(note $\text{pH} = \text{p}K_a$ when acid and base are at equal concentration)

- b. Calculate the pH of this same solution after 20.0 mL of 0.100 *M* HNO₃ is added.

Initial moles of all components

$$n_{\text{HA}} = 0.100 \text{ L} \times \frac{0.100 \text{ mol}}{\text{L}} = 0.0100 \text{ mol HA}$$

$$n_{\text{A}^-} = 0.100 \text{ L} \times \frac{0.100 \text{ mol}}{\text{L}} = 0.0100 \text{ mol A}^-$$

$$n_{\text{HNO}_3} = 0.0200 \text{ L} \times \frac{0.100 \text{ mol}}{\text{L}} = 0.00200 \text{ mol}$$

$$\text{New total volume} = 0.100 \text{ L} + 0.0200 \text{ L} = 0.120 \text{ L}$$

Moles of acid and base after HNO₃ is added:

$$n_{\text{A}^-} = 0.0100 \text{ mol} - 0.00200 \text{ mol} = 0.008 \text{ mol}$$

$$n_{\text{HA}} = 0.0100 \text{ mol} + 0.00200 \text{ mol} = 0.012 \text{ mol}$$

$$[\text{A}^-] = \frac{0.008 \text{ mol}}{0.120 \text{ L}} = 0.0667 \text{ M}$$

$$[\text{HA}] = \frac{0.012 \text{ mol}}{0.120 \text{ L}} = 0.100 \text{ M}$$

$$\text{pH} = 4.77 + \log \frac{0.0667}{0.100} = 4.59$$

-END OF EXAM-