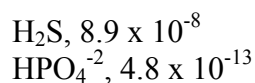


**Chemistry 1046**  
**Exam 3**  
**Fall 1999**

**KEY**

MULTIPLE CHOICE - CIRCLE THE BEST ANSWER (4 points each)

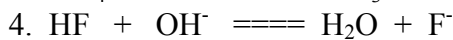
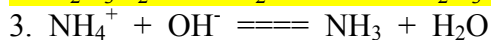
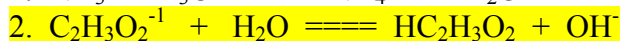
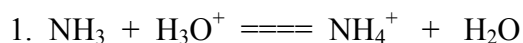
- Which is the best definition of a Lewis base?
  - An electron pair donor
  - A substance that releases hydrogen ions
  - A proton acceptor
  - An electron pair acceptor
  - A proton donor
- For a given weak acid, HA, the value of  $K_a$ 
  - $= [HA]/[A]$ .
  - $= [H_3O^+][A] / [H_2O][HA]$ .
  - will NOT change with temperature.
  - cannot be greater than  $10^{-7}$ .
  - will be a number less than 1 and is often much less than 1.
- A solution has a pH of 11. What is the pOH of the solution?
  - 3
  - 14
  - 11
  - 7
  - 13
- What is its hydrogen ion concentration of a solution of pH = 8.6?
  - $4.0 \times 10^{-8} M$
  - $8.6 \times 10^0 M$
  - $2.5 \times 10^{-9} M$
  - $1.3 \times 10^{-6} M$
  - $5.4 M$
- A solution in which the pH is 8.12 would be described as
  - very acidic.
  - very basic.
  - slightly acidic.
  - slightly basic.
  - neutral.
- A solution of the salt (conjugate base  $A^-$ ) of a weak acid (HA) will normally have a pH greater than 7. This is because
  - $A^-$  reacts with water to produce HA and  $OH^-$ .
  - no HA is formed.
  - the  $[H^+]$  is less than  $10^{-7}$ .
  - $A^-$  is a weak base.
  - the salt of a weak acid makes a strong acid.
- Consider the  $K_a$  values for the following acids:
  - HF,  $6.8 \times 10^{-4}$
  - HNO<sub>2</sub>,  $4.5 \times 10^{-4}$
  - HOBr,  $2.5 \times 10^{-9}$



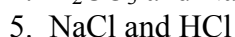
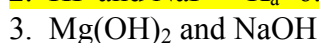
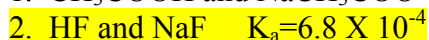
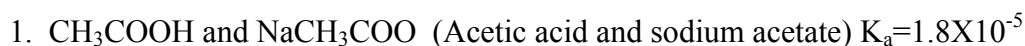
Which of the following is the weakest base? (Careful!)



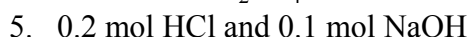
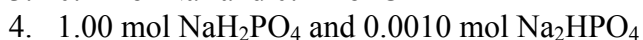
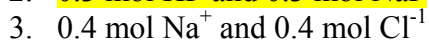
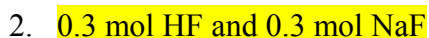
8. For which of the following reactions can we describe an equilibrium constant called a basicity constant,  $K_b$ ?



9. Which of the following mixtures is best for making a buffer with a pH of about 3.6? (Careful!)



10. Which of the following will be a buffer when dissolved in a liter of water?



11. Which of the following when added to water will decrease the pH of the solution?



5. None of the above will change the pH.

12. Which of the following statements is true of the titration of a weak acid by  $\text{NaOH}$ ?

1. At the equivalence point the pH is 7.

2. The moles of acid (HA) are greater than the moles of base (A) at the equivalence point.

3. The moles of the weak base (A) are equal to the moles of the weak acid (HA) at the equivalence point.

4. The pH is greater than 7 at the equivalence point.

5. The solution is acidic at the equivalence point.

13. In an aqueous solution of equal amounts HCl, HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> and NaC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> each whose concentration (except the water) is 2.00 M, which of the following is the greatest contributor to the [H<sup>+</sup>]?  
 1. HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub>                      4. HCl  
 2. C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup>                      5. Na<sup>+</sup>  
 3. H<sub>2</sub>O
14. When one equivalent of hydroxide ion is added to a sample of a weak monoprotic acid, the solution will then  
 1. contain one equivalent of the salt of the weak acid .  
 2. be neutral.  
 3. be slightly acidic.  
 4. contain equal amounts of the HA and A forms of the acid.  
 5. have a pH equal to the pK of the acid.
15. Regarding the equilibrium associated with the self-ionization of pure water at 25°C, it can be said that  
 1. the concentration of [H<sup>+</sup>] > [OH<sup>-</sup>].  
 2. the product of the concentrations of hydrogen and hydroxide ions is equal to 10<sup>-14</sup>.  
 3. the K<sub>w</sub> is equal to the K<sub>b</sub>/K<sub>a</sub> for water.  
 4. The pH – pOH will equal 7.  
 5. The K<sub>b</sub> must be divided by the concentration of water to get the K<sub>a</sub>.
15. What is the pH of a 2.0 x 10<sup>-2</sup> M HCl solution?  
 1. -2.00  
 2. 5.00  
 3. 1.70  
 4. -3.00  
 5. 2.30
16. If one wishes to use the approximation [HA – x] = [HA] in a calculation of the pH of a solution of a weak acid, the value of [HA]/K<sub>a</sub> should be  
 1. >1  
 2. >100  
 3. <1  
 4. <10  
 5. =1
17. What is the pH of a 1.20 M nitrous acid solution? (K<sub>a</sub> = 4.50 x 10<sup>-4</sup>)  
 1. 4.20  
 2. 4.50  
 3. 1.20  
 4. 4.00  
 5. 1.63
18. As the concentration of weak acids decrease, the % dissociation increases.  
 1. True  
 2. False

Problems (Show all work and pay attention to units and significant figures)

1. What are the concentrations of the major species in a 0.200M Nitrous acid buffer, pH = 3.00?  $pK_a = 3.35$  (Hint HH eq) (6 points)

$$pH = pK_a + \log(A/HA) \quad A = NO_2^{-1} \quad HA = HNO_2$$

$$3.00 = 3.35 + \log(A/HA)$$

$$A/HA = .447 / 1 \quad \text{or} = 1 / 2.24$$

$$0.200M / 3.24 = 0.0617M = A \quad HA = .200M - 0.0617M = 0.138M$$

$$[HNO_2] = 0.138M$$

$$[NO_2^{-1}] = 0.0617M$$

$$[H_3O^+] = 1.00 \times 10^{-3}M$$

2. Sulfuric acid ( $H_2SO_4$ ) is a diprotic acid that is characterized by a strong acid dissociation of the first hydrogen ion and a weak acid dissociation of the 2<sup>nd</sup> proton. ( $K_{a1}$ =strong acid,  $K_{a2}=1.2 \times 10^{-2}$ ,  $pK=1.92$ )

If you placed 0.100L of 1.00M  $H_2SO_4$  in a beaker and measured the pH, what would it be? (Ignore the dissociation of  $HSO_4^{-1}$  for this calculation only.)

$$1.00M \text{ strong acid} = [H_3O^+] = 1.00M \quad pH = 0.00$$

The actual pH would be slightly lower due to the dissociation of the  $HSO_4^{-1}$  formed.

If you added 0.200L of 0.500M KOH to the beaker, what would be the major species present in the beaker?

Before the addition the solution would contain 100mmols each of  $H_3O^+$  and  $HSO_4^{-1}$ . 100.mmols of  $OH^{-1}$  would react with 100.mmols of  $H_3O^+$  leaving  $HSO_4^{-1}$  as the major species.

If you placed another 0.100L of 0.500M KOH in the beaker, what major reaction would occur?

The 50.mmols of  $OH^{-1}$  would convert 50.mmols of the  $HSO_4^{-1}$  into 50.mmols of  $SO_4^{-2}$ .  
 $HSO_4^{-1} + OH^{-1} \rightarrow H_2O + SO_4^{-2}$

What would be the major species in the solution, what ratio would they be in, and what would the pH of the solution be?

The solution would now contain equal mmols of both  $HSO_4^{-1}$  and  $SO_4^{-2}$ . The HH equation tells us that the pH would be equal to the  $pK_a$  of  $HSO_4^{-1}$ , which is 1.92.

If you added an additional 10.0 mL of 0.500M KOH, what would the pH of the solution be?

The 5.0mmols of hydroxide would convert 5.0mmols of  $HSO_4^{-1}$  into 5.0mmols of  $SO_4^{-2}$ . This would decrease the  $[HSO_4^{-1}]$  to 45mmols and increase the  $SO_4^{-2}$  to 55mmols. Using the HH equation:

$$pH = 1.92 + \log(55\text{mmols} / 45 \text{mmols}) = 2.01$$

Extra Credit: What would be the final CONCENTRATIONS of all major species in the solution, including hydrogen and hydroxide ions? (Work on back of page)

45mmols of  $HSO_4^{-1}$  in 410mL (100mL + 200mL + 100mL + 10mL)

$$[\text{HSO}_4^-] = 45\text{mmols} / 410\text{mL} = 0.109\text{M}$$

55mmols of  $\text{SO}_4^{2-}$  in 360mL (100mL + 200mL + 50mL + 10mL)

$$[\text{SO}_4^{2-}] = 55\text{mmols} / 360\text{mL} = 0.134\text{M}$$

$$[\text{H}_3\text{O}^+] = 10^{-2.01} = 9.77 \times 10^{-3}\text{M}$$

$$[\text{OH}^-] = 10^{-14}\text{M} / 9.77 \times 10^{-3}\text{M} = 1.02 \times 10^{-12}\text{M}$$