

Acids/Bases and Buffers

Worksheet 2

Acids Bases and Buffers Worksheet2_07answers.doc

Name _____

Date _____ Period _____

1. Define the following terms:

- a. Amphoteric: **Able to react with both bases and acids.**
- b. Amphiprotic: **Able to accept or donate a proton.**
- c. Endpoint: **pH at which an indicator changes color.**
- d. Equivalence point: **When equal mole ratios of H⁺ and OH⁻ have reacted**
- e. Anhydride: **Substance that when placed in water produces an acid or base**
- f. polyprotic: **Having more than 1 proton capable of dissociated to some degree.**

2. A 1M solution of a weak acid is found to dissociate by only 1.37%. Determine which of the following acids it is likely to be:

- Nitrous Acid (HNO₂) K_a = 6.0x10⁻⁴
- Formic Acid (HCHO₂) K_a = 1.9x10⁻⁴
- Acetic Acid (CH₃COOH) K_a = 1.8x10⁻⁵
- Hydrofluoric acid (HF) K_a = 6.9x10⁻⁴

$$K_a = \frac{[H_3O^+][B^-]}{[HB]}$$

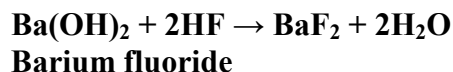
$$X = \frac{(.0137)(.0137)}{(1-.0137)}$$

$$1.9 \times 10^{-4} = \frac{(.0137)(.0137)}{(1-.0137)}$$

Formic acid

Answer: _____

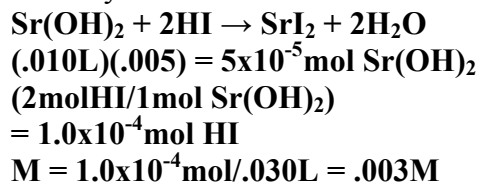
3. Write a balanced chemical equation for the neutralization reaction between barium hydroxide and hydrofluoric acid. Name the salt that is formed.



4. Would you expect the pH at the equivalence point of the solution in question 3 to be acidic, basic or neutral? Explain.

Basic. HF is a weak acid, therefore F⁻ in the salt BaF₂ will have some strength as a base and will hydrolyze to produce OH⁻ in solution and drive up the pH

5. A 30.0mL sample of a hydroiodic acid is titrated to the equivalence point with 10.0mL of .005M strontium hydroxide. What is the molarity of the acid solution?



Answer: _____ M

6. What was the pH of an original solution of hydrobromic acid if the pH of a 10.0mL sample of that solution that has been diluted with water to 85.0mL has a pH of 6.2?

$$[H_3O^+] = 10^{-6.2} = 6.31 \times 10^{-7} M$$

$$(6.31 \times 10^{-7} M)(.085L) = 5.363 \times 10^{-8} \text{ mol}$$

$$\text{Orig. Molarity} = (5.363 \times 10^{-8}) / (.010L)$$

$$= 5.36 \times 10^{-6}$$

$$pH = -\log(5.36 \times 10^{-6}) = 5.27$$

Answer: _____

7. How many grams of sodium fluoride should be added to 500.mL of a .02M hydrofluoric acid solution to create a buffer of pH = 3.16?

Assume the change in volume is negligible.

$$K_a = 6.9 \times 10^{-4}$$

$$pH = pK_a + \log[F^-]/[HF]$$

$$3.16 = 3.16 + \log[F^-]/[HF] \text{ (1:1 mole ratio)}$$

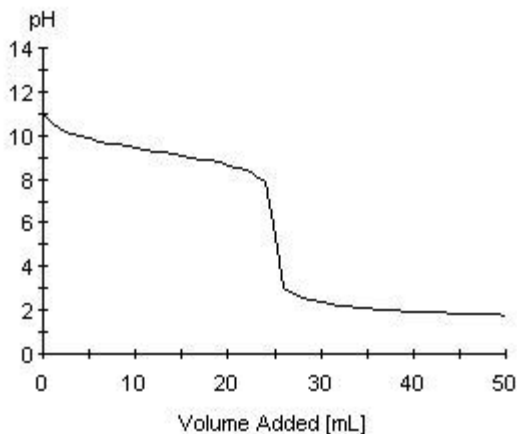
$$\text{mol NaF} = (.500L)(.02M) = .01 \text{ moles}$$

$$.01 \text{ moles (42.0g/mol)} = .42g$$

Answer: _____ g

8. Sketch a titration curve for a weak base being titrated with a strong acid in the space below.

Titration Curve



9. Which of the following pH indicators would be best for the titration curve you sketched in question 8? Explain.

Methyl Red (pH at endpoint = 5)

Bromothymol Blue (pH at endpoint = 7)

Phenolphthalein (pH at endpoint = 9)

Methyl Red. Its endpoint corresponds most closely to the equivalence point.

10. Determine the final pH of a buffer system in which 50.0mL of 1.2M HCl is added to 500.mL of a buffer solution that is 1.0M acetic acid and 1.0M sodium acetate. What would the final pH have been if the HCl had been added to 500.mL of pure water instead? (You will need to look up the K_a for acetic acid)

$$K_a = 1.8 \times 10^{-5} \quad pK_a = 4.74$$

$$pH = pK_a + \log \frac{[CH_3COO^-]}{[CH_3COOH]}$$

$$(.050L)(1.2MHCl) = .06mol \text{ HCl}$$

$$.50mol \text{ acetate} - .06mol = .44mol \text{ acetate}$$

$$.50mol \text{ acetic acid} + .06mol = .56mol \text{ acetic acid}$$

$$pH = 4.74 + \log (.44/.56) = 4.74 - .1047$$

$$= 4.635 = \underline{4.64}$$

Final pH of buffer system: _____

$$.06mol \text{ HCl} / .550L = .109M$$

$$pH = -\log(.109) = \underline{.963}$$

Final pH of non-buffered system: _____

11. Calculate the pH of a buffer solution prepared by adding 20.5g of CH_3COOH and 17.8g of CH_3COONa to enough water to make 5.00×10^2 mL of solution.

($CH_3COOH = 60.052g/mol$; $CH_3COONa = 82.034g/mol$)

$$mol \text{ } CH_3COOH = 20.5g / 60.052g/mol = .341mol$$

$$mol \text{ } NaCH_3COO = 17.8g / 82.034g/mol = .217mol$$

$$pH = 4.74 + \log (.217/.341) = \underline{4.54}$$

12. The pH of a sodium acetate/acetic acid buffer is 4.50. Calculate the ratio of $[CH_3COO^-] / [CH_3COOH]$ (K_a acetic acid = 1.8×10^{-5})

$$4.50 = 4.74 + \log [CH_3COO^-] / [CH_3COOH]$$

$$-.24 = \log [CH_3COO^-] / [CH_3COOH]$$

$$10^{-.24} = [CH_3COO^-] / [CH_3COOH]$$

$$= \underline{.58:1}$$

13. Calculate the pH of a buffer system composed of .05M hydrofluoric acid and .07M potassium fluoride solution. Is this system better at buffering acids or bases?

K_a for HF is 7.1×10^{-4}

$$pK_a = 3.15$$

$$pH = 3.15 + \log (.07/.05) = \underline{3.30}$$

Better at buffering acids because the concentration of the conjugate base salt is greater and can neutralize more acid.

14. Calculate the ratio of lithium nitrite to nitrous acid to create a buffer system with a pH of 5.0. If the solution is .025M HNO_2 , how many grams of lithium nitrite would you need to add to the solution to create 1.0L of buffer (assume constant volume)? $K_a = 4.5 \times 10^{-4}$

$$5.0 = 3.35 + \log [LiNO_2] / [HNO_2]$$

$$10^{1.65} = [LiNO_2] / [HNO_2] = 44.7:1$$

$$X / .025 = 44.7$$

$$X = 1.12M(1.0L) = 1.12mol$$

$$1.12mol \text{ } LiNO_2(52.95g/mol) = \underline{59.3g}$$