Titration Calculations

- **Titration Calculations:** When neutralized, moles H^+ = moles OH^-.

Primary Standards

- To perform a titration, the concentration of one of the solutions must be precisely and accurately known. These substances are called primary standards. A primary standard must meet the following criteria:
    1. It should be obtainable as a very pure solid at reasonable cost.
    2. The substance should be air stable. That is, it should not react with any component in the air, such as oxygen, carbon dioxide, or water vapour.
    3. The substance should be stable in solution for a reasonable length of time.
    4. A substance with a high molar mass is preferred. This minimizes weighing errors.

Example: In a titration, a few drops of bromothymol blue indicator are added to 16.80 mL of an aqueous sodium hydroxide solution of unknown concentration and is neutralized by 25.00 mL of a 0.190 M solution of sulfuric acid. What is the concentration (initial) of the sodium hydroxide solution? Remember: when neutralized, moles H^+ = moles OH^-

Example: A 20.0 mL aqueous solution of strontium hydroxide, has a drop of indicator added to it. The solution turns colour after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the strontium hydroxide?

Example: A bottle is labeled 2.00 M H2SO4 . You decide to titrate a 20.00 mL sample with a 1.85 M NaOH solution. What volume of NaOH solution would you expect to use if the label is correct?

See Titration Assignment

Ionization Constants for Acids and Bases

- Since weak acids and bases ________________ dissolve in water and reach equilibrium, equilibrium law expressions can be written.

$$HF_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + F^-_{(aq)}$$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \text{acid ionization constant}$$

- The larger the acid ionization constant, the ________________ the acid *(really strong acids almost completely dissociate so no equilibrium is reached)*.
- The base ionization constant is the equilibrium constant expression for a weak base.

- *Conjugate Acid-Base Pairs:*
- Ka x Kb = Kw here`s why:

$$HF_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + F^-_{(aq)} \qquad K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$F^-_{(aq)} + H_2O_{(l)} \longleftrightarrow HF_{(aq)} + OH^-_{(aq)} \qquad K_b = \frac{[HF][OH^-]}{[F^-]}$$

$$K_a \times K_b = \frac{[H_3O^+][F^-]}{[HF]} \times \frac{[HF][OH^-]}{[F^-]} = [H_3O^+][OH^-] = K_w$$

$$K_a \times K_b = K_w = 1.0 \times 10^{-14}$$

We can calculate the value of an acid ionization constant if we know the acid concentration and the pH of the solution. Conversely, if we know the concentration and acid ionization constant, we can calculate the expected pH of the solution.

Example: Acetylsalicylic acid (aspirin) is a weak monoprotic acid. A 0.100 M solution of the acid has a pH of 2.24. Calculate the acid ionization constant of the acid (K_a).

Example: Ascorbic acid (vitamin C) is a weak monoprotic acid. It has a $K_a = 8.0 \times 10^{-5}$. Calculate the pH of a 0.100 M solution.

Example: Caffeine is a weak base that is related to ammonia. It has a $K_b = 4.1 \times 10^{-4}$. Calculate the pH of a 0.70 M solution.

Buffers

- Buffers are mixtures of chemicals that make a solution resist a change in pH. Solutions that have a resistance to changes in their pH because of the presence of buffers are called ________________________________.
- In general, buffers are made up of (1) a weak acid and one of its soluble salts, or (2) a weak base and one of its soluble salts.
- Buffers of the first type buffer a solution in the acid range. Buffers of the second type buffer a solution in the base range. Both kinds of buffers make solutions in which an equilibrium exists that shifts in response to the addition of an acid or base so as to allow only small changes in the hydrogen-ion concentration.
- The body has a wide array of mechanisms to maintain **homeostasis** in the ________________and extracellular fluid. The most important way that the pH of the blood is kept relatively constant is by **buffers** dissolved in the blood.

Name: ______________________________

Titration

1. The following aqueous solutions were titrated with a 0.150 M aqueous sodium hydroxide solution. Write the overall reaction equations and calculate the concentrations of the acid solutions.
   a) 25.00 mL of hydrochloric acid requiring 16.50 mL of base solution

   b) 25.00 mL of sulfuric acid solution requiring 42.00 mL of the base solution

2. The following aqueous acid solutions were all titrated with a 0.180 M sodium hydroxide solution. In each case write the overall reaction equation for the neutralization reaction and calculate the concentration of the acid solution.
   a) 25.00 mL of formic acid, HCO_2H (with one acidic hydrogen per molecule) requiring 37.50 mL of the base solution

   b) 25.00 mL of a phosphoric acid solution requiring 62.00 mL of the base solution

3. The following aqueous base solutions were all titrated with a 0.126 M hydrochloric acid solution. Write the overall reaction equation for the neutralization reactions and calculate the concentration of the base solutions.

    a) 10.00 mL of an ammonia solution requiring 22.50 mL of the acid solution

    b) 10.00 mL of an diaminoethane, $H_2NCH_2CH_2NH_2$ requiring 36.00 mL of the acid solution

4. A solution of citric acid, a triprotic acid, is titrated with a sodium hydroxide solution. A 20.00 mL sample of the citric acid solution requires 17.03 mL of a 2.025 M solution of NaOH to reach the equivalence point. What is the molarity of the acid solution?

5. Calculate the pH of an aqueous solution of strong acid prepared by adding 50.00 mL of a 1.50 M hydrochloric acid to 100.0 mL of 0.500 M nitric acid.

6. A 12.00 mL sample of an ammonia solution is titrated with 1.499 M HNO3 solution. A total of 19.48 mL of acid is required to reach the equivalence point. What is the molarity of the ammonia solution?

7. A H_2SO_4 solution of unknown molarity is titrated with a 1.209 M NaOH solution. The titration requires 42.27 mL of the NaOH solution to reach the equivalent point with 25.00 mL of the H2SO4 solution. What is the molarity of the acid solution?

8. What volume of a 0.5200 M solution of H2SO4 would be needed to titrate 100.00 mL of a 0.1225 M solution of Sr(OH)2 ?

Name: ______________________________

Ionization Constants Acids and Bases

1. The reaction of oxalate ions, $C_2O_4^{2-}$, with water is an equilibrium reaction. Write the equation for the reaction and the equilibrium expression for the ionization constant. Given that the K_b of oxalate ions is 1.6×10^{-10}, calculate the value of K_a for the hydrogen oxalate ions.

2. An aqueous solution containing 1.00 *M* boric acid (H_3BO_4) is found to have a pH of 4.57. From this information, calculate the K_a value of boric acid. Consider only the first ionization step of boric acid.

3. Using the K_a value for nitrous acid, determine K_b for the basic anion, NO_2^-. Use this value to calculate the pH of a 0.15 *M* solution of sodium nitrite, $NaNO_2$.

4. A solution of a weak acid requires 25.40 mL of base for neutralization. After 12.70 mL of this base had been added during the titration, the pH of the solution was 4.82. Determine the K_a for the weak acid.

5. Calculate the ammonium ion concentration in 500.0 mL of a 0.050 M aqueous solution of ammonia to which 4.0 g of sodium hydroxide has been added. K_b for ammonia is 1.8×10^{-5}.

6. A solution of a weak acid requires 25.40 mL of base for neutralization. After 12.70 mL of this base had been added during the titration, the pH of the solution was 4.82. Determine the K_a for the weak acid.