Experiment 2.

1. Preparing Standard Acid Solution

Hydrochloric acid and sodium hydroxide are the most common strong acids and bases used in the laboratory. Both reagents need to be standardized to learn their exact concentrations. Your textbook provides background information for the procedures described below. ***Please prepare an EXCEL file for your calculations in advance.***

Reagents

Methyl Red indicator

Concentrated (37 wt%) HCl

Primary standard: sodium tetraborate $Na_2B_4O_7 \times 10\ H_2O$ ("borax"). FW 381.422; Solid colorless prismatic crystals. Specific gravity is 1.72. Borax melts when heated loosing water. Non-hygroscopic. Reacts with acid:

$$Na_2B_4O_7 + 2HCl + 5H_2O \rightarrow 4H_3BO_3 + 2NaCl$$

2. Standardizing HCl

a. Use the table inside the cover of the textbook to calculate the volume of ~37 wt% HCl that should be added to 1 L of distilled water to produce 0.1 M HCl and prepare this solution.

b. Prepare primary standard grade sodium tetraborate.

c. Weigh four samples, each containing enough **$Na_2B_4O_7 \times 10\ H_2O$** to react with ~25 mL of 0.1 M HCl (around 0.5 g) and place each in a 125-mL flask. When you are ready to titrate each one, dissolve it in ~50mL of distilled water. Add 3 drops of methyl red indicator and titrate one to an orange-red color.

Perform one blank titration of 50 mL of approximately 0.05 M NaCl containing 3 drops of indicator.

Subtract the volume of HCl needed for the blank from that required to titrate $Na_2B_4O_7 \times 10\ H_2O$. Calculate the mean HCl molarity, standard deviation, and relative standard deviation.

$$c_{HCl} = \frac{m_{borax} \times 2 \times 1000}{V_{HCl} \times FW_{borax}}$$

Example: 27.65 mL of HCl were used to titrate 0.4916 g of Borax.

Molarity of HCl is:

$$C_{HCl} = \frac{0.4916 \times 2 \times 1000}{27.65 \times 381.42} = 0.0932_3$$

3. Analysis of a Sample of Impure Sodium Carbonate

Dry the sample at 150^{o}C and then cool it in a desiccator.

a. Weigh a sample of a proper size (0.3-0.5 g would be a good starting point) in to 200-250 mL Erlenmeyer flask. Dissolve it in 20-40 mL of deionized water and add 3 drops of bromcresole green indicator.

b. Titrate with standard HCl solution until indicator just begins to change to the green color. Boil the solution for 2-3 min and complete the titration.

$$Na_2CO_3 + 2HCl \rightarrow 2\ NaCl + H_2O + CO_2$$

c. For next titrations, recalculate the necessary amount of the sample weighed so the volume necessary for titration will be around 40 mL of **HCl**.

Repeat these titrations at least three more times, calculate the percentage of sodium carbonate in your sample with appropriate standard deviation.

$$m_{Na2CO3} = \frac{FW_{Na2CO3} \times C_{HCl} \times V_{HCl}}{2 \times 1000}$$

$$\%\%Na2CO3 = \frac{m_{Na2CO3}}{m_{sample}} \times 100\%$$

Experiment 2 - standartisation with TRIS instead of borax.

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Reagents

Methyl Red indicator

Concentrated (37 wt%) HCl

Primary standard: 2-Amino-2-(hydroxymethyl)propane-1,3-diol ("TRIS"). FW 121.136; Solid colorless crystals. Non-hygroscopic. Reacts with acid.

2. Standardizing HCl

a. Use the table inside the cover of the textbook to calculate the volume of ~37 wt% HCl that should be added to 1 L of distilled water to produce 0.1 M HCl and prepare this solution.

b. Prepare primary standard grade TRIS.

c. Weigh four samples, each containing enough **TRIS** to react with ~25 mL of 0.1 M HCl (around 0.3-0.35 g) and place each in a 125-mL flask. When you are ready to titrate each one, dissolve it in ~50mL of distilled water. Add 3 drops of methyl red indicator (more if necessary) and titrate one to an orange color.

Perform one blank titration of 50 mL of water containing 3 drops of indicator.

Subtract the volume of HCl needed for the blank from that required to titrate TRIS. Calculate the mean HCl molarity, standard deviation, and relative standard deviation.

$$C_{HCl} = \frac{m_{TRIS} \times 1000}{V_{HCl} \times FW_{TRIS}}$$