<u>**EXPERIMENT 30A7: VINEGAR TITRATION**</u>

Learning Outcomes

Upon completion of this lab, the student will be able to:

    1) Measure the amount of acetic acid in a solution of vinegar

Introduction

The molar concentration of an acid or a base can be determined by the method of titration. In a titration, a solution of known concentration is slowly added to a known volume of a solution of unknown concentration until the two have completely reacted with each other. There are many different kinds of titration; the type used in this experiment is called an acid-base titration.

Consider for instance, the reaction between aqueous solutions of sulfuric acid and potassium hydroxide. The balanced chemical equation for the reaction is shown below:

$H_2SO_{4(aq)}$ + $2KOH_{(aq)}$ ➜ $K_2SO_{4(aq)}$ + $2H_2O_{(l)}$
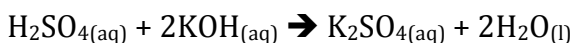
Assume that the molar concentration of KOH is 0.100 M and that of H_2SO_4 is unknown. In order to determine the molar concentration of H_2SO_4, one would need to add the KOH solution to a known volume of H_2SO_4.

In a certain experiment, the KOH was added to 10.00 mL of H_2SO_4. The volume of KOH needed to completely react with the H_2SO_4 will enable the determination of the molarity of H_2SO_4. In the same experiment, assume that 13.75 mL of KOH was needed to completely react with the H_2SO_4. The molarity of the acid is calculated as follows:

Molarity of H_2SO_4 =

$$0.100 \frac{mol}{L} KOH \times 13.75 ml \times \frac{1L}{1000mL} \times \frac{1 H_2SO_4}{2KOH} \times \frac{1}{10.00mL} \times \frac{1000mL}{1L} = 0.0688 \frac{mol}{L}$$

As seen from the above calculation, the stoichiometric ratio between the two reactants is the key to the determination of the molarity of the unknown solution.

In order to conduct the above experiment, typically the H_2SO_4 is in an Erlenmeyer flask, and the KOH is in a burette. The KOH is added one drop at a time from the burette into the acid solution with constant stirring to ensure that the reagents combine and react.

<u>**Vinegar Titration**</u>

In this experiment, a solution of vinegar has been provided for analysis. The active ingredient in vinegar is acetic acid, CH_3COOH. In order to determine the amount of acetic acid in the vinegar, the acetic acid will be titrated with a solution of known concentration of sodium hydroxide.

The chemical reaction between acetic acid and sodium hydroxide is given below:

$CH_3COOH_{(aq)} + NaOH_{(aq)}$ ➜ $CH_3COONa_{(aq)} + H_2O_{(l)}$
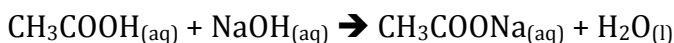
The balanced chemical equation shows that one mole of CH_3COOH reacts with exactly one mole of NaOH. The experiment is performed by adding NaOH of known molarity to a known volume of vinegar until the reaction is complete.

Determining the completion of an acid/base reaction

In order to obtain the molarity or moles of the unknown reactant the solution whose concentration is known (in this experiment that would be NaOH), must be added until the reaction is complete. This means, exactly one mole of NaOH must be added to one mole of acetic acid. This point in the titration is called the **Equivalence Point**.

The equivalence point is defined as that point in the titration when stoichiometrically equal amounts of acid and base are present. In the CH_3COOH/NaOH titration, that would be when one mole of NaOH has been added to one mole of CH_3COOH. In the H_2SO_4/KOH example shown previously, that would be when two moles of KOH have been added to one mole of H_2SO_4. Therefore the equivalence point depends on the reaction stoichiometry.

At the beginning of the titration, the solution in the Erlenmeyer flask is acidic. As the base is added, it completely reacts with the acid and the solution in the Erlenmeyer flask continues to be acidic. But, at the equivalence point, the acid has completely reacted with the base. If even one tiny drop of base is added beyond that needed to arrive at the equivalence point, the solution in the Erlenmeyer flask is basic. This difference in the acid/base property of the solution in the Erlenmeyer flask is used to visually determine the end of the titration.

An **indicator** is a chemical substance whose color depends on the acid/base property of the medium it is present in. Phenolphthalein is an indicator, which is colorless in an acidic medium and has a pink color in a basic medium.

In this titration, a few drops of phenolphthalein should be added to the acid in the Erlenmeyer flask. The solution will remain colorless until the equivalence point. When the equivalence point has been crossed and the solution becomes basic, the phenolphthalein will take on a pink color. This is the reason to add the base drop by drop, so that even though the equivalence point will be crossed, the titration can be stopped at the appearance of the first permanent pale pink color.

This **End Point** is the point in the titration when the indicator changes color. It is important to note that the End Point of a titration is slightly beyond the Equivalence Point. In the case of the phenolphthalein, the intensity of the pink color increases as the solution becomes more and more basic. Therefore, in order to stay as close to the Equivalence Point as possible, it is important to stop the titration at the appearance of a permanent pale pink color. In order to easily observe the color changes in the solution, it is a good idea to place a sheet of plain white paper beneath the flask.

Experimental Design

A solution of known concentration of sodium hydroxide will be provided for this experiment. This solution will be used to titrate the acetic acid in the vinegar. Phenolphthalein is used as an indicator in both of these titrations.

Reagents and Supplies

Aqueous sodium hydroxide, commercial vinegar solution, 0.1% phenolphthalein solution

(See posted Material Safety Data Sheets)

Procedure

<u>TITRATION OF THE ACETIC ACID IN VINEGAR WITH THE STANDARDIZED SODIUM HYDROXIDE SOLUTION</u>

1. Record the exact concentration of the NaOH solution provided for this experiment and obtain about 20 ml of this solution.

2. Obtain approximately 10-mL of commercial vinegar solution.

3. Clamp two microburettes to a burette stand and label one burette as "NaOH" and the other burette as "vinegar".

4. Condition each burette with the respective reagent.

5. Record the initial burette readings of both the burettes.

6. Obtain a clean, dry, small Erlenmeyer flask.

7. Record the mass of the empty Erlenmeyer flask.

8. Dispense approximately 0.1-mL of vinegar into the Erlenmeyer flask.

9. Record the mass of the Erlenmeyer flask with the vinegar.

10. Add one drop of **0.1%-phenolphthalein** solution into the Erlenmeyer flask containing the vinegar.

11. Rinse the sides of the Erlenmeyer flask with deionized water.

12. Titrate the contents of the Erlenmeyer flask with the NaOH solution until a permanent pale pink color is obtained. After the addition of each drop of NaOH, be sure to swirl the Erlenmeyer flask thoroughly to ensure mixing of the reagents. In case any NaOH solution falls on the side of the Erlenmeyer flask, rinse the sides of the flask with deionized water.

13. Record the final burette reading of the NaOH burette.

14. Repeat steps 5-13 two or three more times.

15. Dispose all waste into appropriate waste disposal containers as instructed by your instructor.

Data Table

TITRATION OF THE ACETIC ACID IN VINEGAR WITH THE STANDARDIZED SODIUM HYDROXIDE SOLUTION

Vinegar

	Trial 1	Trial 2	Trial 3	Trial 4
Mass of empty Erlenmeyer flask (grams)				
Mass of Erlenmeyer flask + vinegar (grams)				

NaOH solution

	Trial 1	Trial 2	Trial 3	Trial 4
Initial burette reading (mL)				
Final burette reading (mL)				

Calculations

<u>TITRATION OF THE ACETIC ACID IN VINEGAR WITH THE STANDARDIZED SODIUM HYDROXIDE SOLUTION</u>

$CH_3COOH_{(aq)} + NaOH_{(aq)}$ ➜ $CH_3COONa_{(aq)} + H_2O_{(l)}$
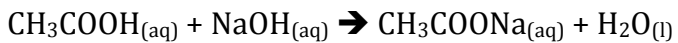
At equivalence point: moles of NaOH = moles of CH_3COOH

Molarity of NaOH (from the label on the bottle) =

	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**
Volume of NaOH (liters)				
Moles of NaOH (M × V)				
Moles of CH_3COOH				
Molar Mass of CH_3COOH (g/mol)				
Mass of CH_3COOH (grams)				
Mass of vinegar (grams)				
Mass percent of acetic acid in vinegar (%)				

Average mass percent of acetic acid in vinegar = _____