CEAC 103
GENERAL CHEMISTRY

Experiment 5

Titration of Acids and Bases

Purpose: To become familiar with the techniques of titration, a volumetric method of analysis; to determine the molarity and pH of an acid solution.

Prelaboratory Work

Before the experiment in the laboratory, you should be able to answer these questions.

1. What is the definition of standardization?
2. What is the definition of titration?
3. What is the definition of molarity?
4. Write the difference between equivalence points and end points.
5. Find the molarity of a solution that contains 3.78 g of $H_2C_2O_4.2H_2O$ in 100 mL of solution?

Theory

According to Arrhenius, **acid** is a chemical substance which gives hydronium ion, H_3O^+, when dissolved in water. Also, Bronsted-Lowry defined that acid acts as a proton donor. On the other hand, Arrhenius defined **base** in a way that it gives hydroxide ion, OH^-, when dissolved in water. According to Bronsted-Lowry, base acts as a proton acceptor.

One of the most common and familiar reactions in chemistry is the reaction of an acid with a base. This reaction is named as ***neutralization*** reaction, and the essential feature of this process in aqueous solution is the combination of hydronium ions with hydroxide ions to form water.

$$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2H_2O_{(l)}$$

In this experiment you will perform this reaction to determine accurately the concentration of a sodium hydroxide solution that you have prepared. The process of determining the exact concentration (molarity) of a solution is called ***standardization***. Next you will measure the concentration of the unknown acid solution. For this purpose you are expected to measure the volume of your standard base that is required to exactly neutralize the unknown acid solution. The technique of accurately measuring the volume of a solution required to react with another reagent is termed ***titration***.

During titration, you will use an ***indicator*** solution to understand whether you could neutralize your acid with a base or vice versa. Indicators change colors at different pH values. For example, phenolphthalein changes color from colorless to pink at a pH of about 9; in slightly more acidic solutions it is colorless, whereas, in more alkaline solutions it is pink. The color change is termed the ***end point*** of the titration.

The point at which stoichiometrically equivalent quantities are brought together is known as the ***equivalence point*** of the titration. It should be noted that the equivalence point in a titration is a theoretical point.

Molarity (M) is used to define concentration of a solution more clearly ,and it is defined as the number of moles of solute per liter of solution, or the number of millimoles of solute per milliliter of solution:

$$M = \frac{moles.solute}{volume.of.solution} = \frac{10^{-3}\,mole}{10^{-3}\,liter} = \frac{mmol}{mL} \quad [1]$$

PH concept

pH is a concept that is used to measure the acidity of a solution. It is related to concentration of H_3O^+ ion in molarity.

$$HA + H_2O \leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$$

The concentration of $[H^+]$ may change over a wide range of values and these values are frequently expressed in terms of exponential numbers. For this reason, a simpler form of representation for $[H^+]$ is provided as follows: **pH = -Iog[H$^+$]**

There is another concept named aas **pOH**. It is used to measure the basicity of the solution and related to concentration of OH^- in molarity: **pOH = -log[OH$^-$]**

When water is self ionized, hydronium and hydroxide ion formed in equal amounts:

$$H_2O \leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

$K_w= [H^+][OH^-]$

$K_w= 1.0x10^{-14}$ at 25^0C

Therefore;

pH + pOH = 14

Example: What are the pH values for **a)** 0. 1 M HCl and **b)** 0. 1 M NaOH?

Solution:

a) HCI is a strong acid, it dissociates almost completely in aqueous solution. Therefore 0.1 M HCl gives 0.1 M H^+ and 0. 1 M Cl^-.

$$HCl \rightarrow H^+ + Cl^-$$

$[H^+] = 0.1$ M

pH = -Iog(0. 1) = 1.00

b) NaOH is a strong base; it dissociates almost completely in aqueous solution. 0.1 M NaOH gives 0.1 M Na^+ and 0. 1 M OH^-.

$$NaOH \rightarrow Na^+ + OH^-$$

$[OH^-] = 0.\ 1$ M

$[H^+] = Kw/ [OH^-] = 1 \times 10^{-14} /0.1 = 1.0 \times 10^{-13}$ M.

pH = -log(10 x 10^{13}) = 13.00

or, pH + pOH = 14   1 + pOH = 14 then, pOH = 13

Strong Acid-Base Titration

As an example for neutralization reaction between strong acid (e.g. HCl) and a strong base (e.g. NaOH);

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$$

As a result, for a monoprotonic acid and base at the end point;

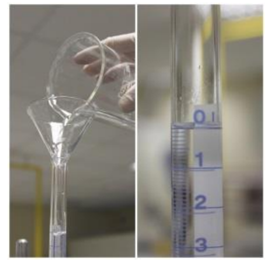
$M_{acid}V_{acid} = M_{base}V_{base}$

In this experiment, we use an acid-base indicator, phenolphthalein to determine the end point in the titration. We choose an indicator such that its color change occurs as closely as possible to the equivalence point.

Materials

Sodium hydroxide (NaOH)	600 mL beaker	Ring stand
Hydrochloric acid (HCl)	250 mL Erlenmeyer flasks	Buret clamp
Phenolphthalein solution	50 mL buret	Balance
Wash bottle	500 mL erlenmeyer flask	

Procedure

PART A: Standardization of Sodium Hydroxide Solution	
1. Fill the buret with the NaOH solution and remove the air from the tip by running out some of the liquid into an empty beaker. Make sure that the lower part of the meniscus is at the zero mark or slightly lower.	
2. Take 15.00 mL of standard HCl solution into a clean Erlenmeyer flask and add a few drops of phenolphthalein solution.	
3. Start to add the sodium hydroxide solution slowly to your flasks of HCl solution while gently swirling the contents of the flask. As the sodium hydroxide solution is added, a pink color appears where the drops of the base come in contact with the solution. This coloration disappears with swirling. As the end point is approached, the color disappears more slowly, at which time the sodium hydroxide should be added drop by drop. The end point is reached when the colour of the solution turns from colorless to pink.	
4. Read the buret and record it. Repeat this procedure for two times. From the data you obtain in the three titrations, calculate the molarity of the sodium hydroxide solution to four significant figures.	

PART B: Analysis of an Unknown Acid	
1. Obtain 15 mL of the unknown solutions. Add a few drops of phenolphthalein solution and titrate against the standard NaOH solution as in Part A.	

Questions

1. Write the balanced chemical equation for the reaction of HCl with NaOH.

2. A solution of malonic acid, $H_2C_3H_2O_4$, was standardized by titration with 0.100 M NaOH solution. If 21.82 mL of the NaOH solution were required to neutralize completely 12.12 mL of the malonic acid solution, what is the molarity of the malonic acid solution?

3. Sodium carbonate is a reagent that may be used to standardize acids in the same way. In such standardization it was found that a 0.432-g sample of sodium carbonate required 22.3 mL of a sulfuric acid solution to reach the end point for the reaction.

$Na_2CO_3(aq) + H_2SO_4(aq) \rightarrow H_2O(l) + CO_2(g) + Na_2SO_4(aq)$

What is the molarity of the H_2SO_4?

4. A solution contains 0.252 g of oxalic acid, $H_2C_2O_4.2H_2O$, in 500 mL. What is the molarity of this solution?

DATA SHEET

Titration of Acids and Bases

Student's Name : Date:

Laboratory Section/Group No :

Assistant's Name and Signature :

A. Standardization of Sodium Hydroxide Solution

1. Molarity of HCI solution : ……………
2. Volume of HCl solution: ……………
3. Volume of NaOH solution used for the titrations 1st: …….2nd: ……..3rd: ……..
4. Molarity of NaOH solution: …………….

B. Analysis of an Unknown Acid I

1. Volume of unknown acid solution: ……………
2. Volume of NaOH solution used for the titration : …………….
3. Molarity of unknown acid I:....................

C. Analysis of an Unknown Acid II

1. Volume of unknown acid solution: ……………
2. Volume of NaOH solution used for the titration : …………….
3. Molarity of unknown acid II:....................