Green University of Bangladesh

Dept. of Computer Science and Engineering

Lab report-3

Course Code: CHE-102

Course Title: Chemistry Laboratory

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Remarks

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Experiment number: 03

Experiment name: Standardization of Hydrochloric Acid with Standard Sodium Carbonate Solution.

Objectives:

1. To study the strength of hydrochloric acid with standard sodium carbonate.

2. To observe the molar ratio of acid or base needed for neutralization.

3. To study effect of different indicators and end point..

Learning Outcome:

After completing this experiment the students will be able to:

1. Determine the strength of an acid with the help of a primary standard base.

2. Observe the end point of reaction by color change.

Theory:

In this experiment we shall determine the strength of commercial Hydrochloric Acid solution by a secondary standard solution of Na_2CO_3. This is done by means of 'Titration'. The important matters that are related with the experiment are stated below:

Titration: In presence of a suitable indicator, the volumetric analysis in which a standard solution is added in another solution (whose strength is not known) to reach its end point to determine the strength of that solution is called 'titration'.

Standard Solution: A solution of known concentration is called a 'standard solution'.

Indicator: In our acid-base titration there is an important use of indicator. An 'indicator' is a chemical substance that detects the equivalent point.

(i.e. the end point) of reaction by changing its color. Indicators have different structures in acidic and in basic solution.

Equivalent Point: The 'equivalent point' is the point in a titration when a stoichiometric amount of reactant has been added.

Normality: The number of gram equivalent weight of a solute per litre of solution is called normality.

Normality (N) = gm equivalent of solute /liters solution. It is known to us that both alkalimetry and acidimetry are based on neutralization reaction. If an acid-base reaction is such like that, a ACID + b BASE = PRODUCT .

Then we know that,

$V_{base} \times S_{base} = V_{acid} \times S_{acid}$

So, $S_{acid} = (V_{base} \times S_{base}) / V_{acid}$

Here ‘V’ represents the volume and ‘S’ represents the strength of the substance. Reaction: Neutralization reaction between Na_2CO_3 and HCl acid takes place into two steps-

$Na_2CO_3+2HCl \Rightarrow NaHCO_3+2NaCl$

$NaHCO_3+HCl \Rightarrow NaCl + H_2CO_3$

The ultimate reaction, $Na_2CO_3 + 2HCl \Rightarrow 2NaCl + H_2CO_3$

In the first step,- the solution is basic due to the formation of a salt where the basic part is stronger than the acidic part ($NaHCO_3$).So, in order to determine the equivalent point of this reaction Phenolphthalein is used. As the salt that forms due to the neutralization reaction, produces more OH-, so the solution becomes a basic one and thus it have a pH range above 7.We know that the working environment needed for phenolphthalein is basic ; thus Phenolphthalein becomes the perfect indicator for determining the end point of the first step of the reaction. In the second reaction, NaCl and Carbonic Acid is formed. Because of the presence of Carbonic Acid in the solution, it becomes acidic.

So, ‘Methyl Orange’ (pH range2.9 - 4.6) is used as indicator to determine the equivalent point.

Name of the Indicator	pH Range	Color in Alkaline solution	Color in Acid solution
Phenolphthalein	8.3 - 10.0	Pink	Colorless
Methyl Orange	2.9 - 4.6	Yellow	Pink

Apparatus:

1. Conical flask,

2. Burette,

3. Pipette,

4. Volumetric flask,

5. Stand,

6. Funnel .

Indicator:

1.Phenolphthalein

2. Methyl Orange

Chemical Reagents:

1) Standardized Na_2CO_3 solution

2) HCl solution

3) Distilled water

Procedure:

1) Take 10 ml of standard Na_2CO_3 solution in a conical flask and dilute it to about 50 ml. Add one drop of phenolphthalein and titrate against HCl (prepared as expt no. 2) Contained in a burette. Now note the burette reading when just one drop of HCl discharges the pink color of the solution. This is the first end point. Then add 3-4 drops of methyl orange inside the same conical flask and continue titration against the same HCl acid. The end point reached when the yellow color of the solution just changes to faint pink (orange). Note the burette reading. This is the second end point. The difference of the burette reading from initial to second end point will be the volume of the acid required in the titration.

Repeat the whole experiment twice and these should agree within ± 0.1 ml initial to second end point. Calculate the strength of dil. HCl and then find out the normality of commercial conc.
HCl acid Data and calculation:
Standardization of HCl acid with standard Na_2CO_3 solution:

Data and calculation:

Number of Observation	Volume of Na2CO3 (ml)	Burette Reading (ml)			Volume of Acid in ml	Average Reading in ml
		Initial reading	First end point	Second end point		
1	10	0	6	9	9	
2	10	9	15.2	18.6	9.6	9.05
3	10	9.6	24	27.7	9.1	

Weight of Na_2CO_3 in gram=0.53gm.

Calculation:

Strength of Na_2CO_3 = .5M.

We know that,

$$2V_{acid} \times S_{acid} = V_{base} \times S_{base}$$

Determination of the normality

Here,

$V_{base} = 10$ ml
$S_{base} = .5$M
$V_{acid} = 9.05$ml
S_{acid}=?

So,

S acid= $(10\times2)/(.5\times9.05)$ M =4.52M

Result:

The strength of the supplied HCl is : 4.52 M

Discussion:

The following causes can be assumed for the possible cause of error: While the solution of Na_2CO_3 was prepared, a little amount of extra water might have been

added into the volumetric flask, this can be one of the reasons. If these causes could be avoided we could have get a perfect result of the concentration of Commercial Hydrochloric Acid.