# Acid-Base Titrations

## *Standardization of Approximately 0.1M HCl Solution*

**Introduction**

HCl stock solution can be standardized using a primary or secondary standard solution. Two important methods (using primary substance) are usually used; against sodium carbonate, or sodium tetraborate decahydrate (borax). The second substance is preferred due to its relatively large molecular mass, superior purity, non-hygroscopic characters, and sharpness of the end point at room temperature.

**a. Against $Na_2CO_3$**

Sodium carbonate reacts with HCl according to the equation:

$$Na_2CO_3 + 2\,HCl = 2\,NaCl + CO_2 + H_2O$$

This reaction occurs if enough HCl is used which can be visualized using methyl orange indicator. Usually, an appropriate weight of $Na_2CO_3$ is dissolved in water and titrated against HCl to determine the normality of the acid.

**b. Against Sodium Tetraborate**

Sodium tetraborate (borax) can be used for neutralization of acids as mentioned previously. Usually, methyl red is the preferred indicator where boric acid is the result of the reaction as the following equation shows:

$$Na_2B_4O_7 + 2\,HCl + 5\,H_2O = 4\,H_3BO_3 + 2\,NaCl$$

In this experiment, we will use solid $Na_2CO_3$ (sodium carbonate), a primary standard, to standardize an HCl solution by titration. The reaction is an acid-base neutralization in which carbonic acid and sodium chloride are produced. The carbonic acid decomposes to give carbon dioxide and water. The equation for the reaction is:

$$2HCl_{(aq)} + Na_2CO_{3(aq)} \text{---> } 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

It is important to note that the stoichiometry between HCl and $Na_2CO_3$ is a 2 to 1 ratio.

Titration consists of the careful addition of one solution from a burette (HCl) to another substance in a conical flask ($Na_2CO_3$) until all of the substance in the flask has reacted. The solution added from the burette

is called the titrant. For every titration, there must be a way to determine when the titration reaction is complete. In acid-base titrations, this is accomplished by adding a small amount of an organic dye, called an indicator, to the solution to be titrated. If the indicator is chosen correctly, a color change, called the end point, occurs when the moles of acid equal the moles of base. Today, the indicator methyl orange will be used. It changes from yellow to pink. If you titrate to the pink color, you have gone to far. When $Na_2CO_3$ reacts with HCl, carbon dioxide ($CO_2$) is produced. This will give us an incorrect endpoint. To correct this, we must remove the $CO_2$ by boiling the solution (for half minute) therefore it turn to (original color) yellow. When the solution will turn back to yellow then can be titrated to the correct endpoint (pink color).

**Procedure**

**A. Preparation of 0.1 N HCl**

The normality of concentrated hydrochloric acid can be calculated from the information written on the bottle (percentage % , specific gravity , equivalent weight) according to the equation:-

$$N = \frac{\% \times Sp.gr \times 1000}{eq.wt}$$

*(Sp.gr=10188, eq.wt=36.5, %=37)*

Transfer (2.08) mL of concentrated hydrochloric acid into volumetric flask (250) mL using cylinder, complete to the mark with distilled water and mix well.

**CAUTION: HYDROCHLORIC ACID CAN CAUSE CHEMICAL BURNS IN ADDITION TO RUINING YOUR CLOTHING. IF YOU SPILL ANY OF THE ACID ON YOU, WASH THE CONTAMINATED AREA THOROUGHLY AND REPORT THE INCIDENT TO YOUR INSTRUCTOR.**

**B. Preparation of the Sodium Carbonate Solution**

1- Dissolve (2.64) gm of $Na_2CO_3$ into a small beaker with distilled water and transfer the solution after dissolution into the volumetric flask (250) mL , washing the beaker many times and adding the washing into the volumetric flask for quantitative transferring of the solution, complete to the mark and mix well.
2- Transfer (10) mL of $Na_2CO_3$ solution into a clean conical flask using pipette.
3- Add two or three drops of methyl orange indicator into the conical flask (the color is yellow).

## C. Doing the Titration

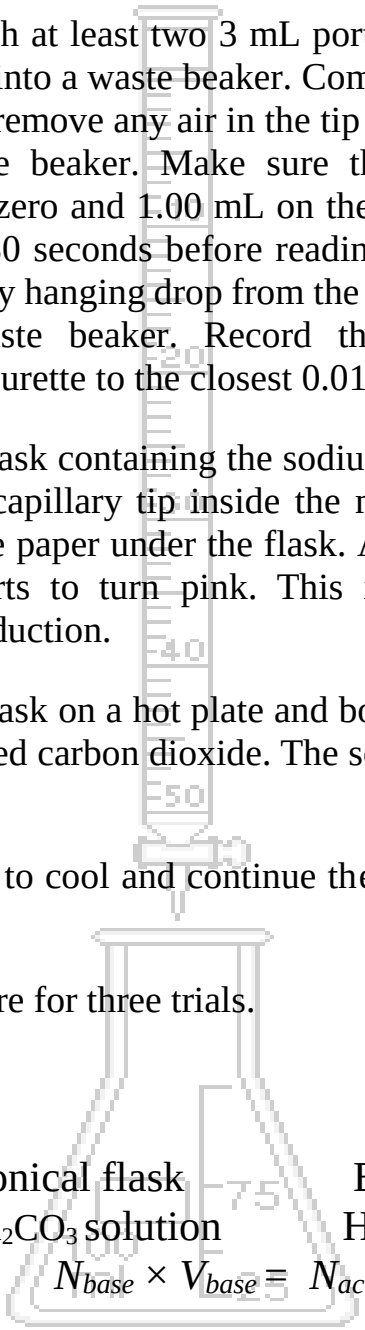
1. Rinse a burette with at least two 3 mL portions of your HCl solution. Discard each portion into a waste beaker. Completely fill the burette with the HCl solution and remove any air in the tip by running out some of the liquid into the waste beaker. Make sure that the lower part of the meniscus is between zero and 1.00 mL on the burette. Allow the burette to stand for at least 30 seconds before reading the exact position of the meniscus. Remove any hanging drop from the burette tip by touching it to the side of the waste beaker. Record the initial burette reading. Remember, read the burette to the closest 0.01 mL.

2. Place the conical flask containing the sodium carbonate solution under the burette with the capillary tip inside the mouth of the conical flask. Place a piece of white paper under the flask. Add the HCl carefully until the solution just starts to turn pink. This is the premature endpoint discussed in the introduction.

3. Place the conical flask on a hot plate and boil gently for a few minutes to remove the dissolved carbon dioxide. The solution should change back to the yellow color.

4. Allow the solution to cool and continue the titration until the solution turns pink.

5. Repeat the procedure for three trials.



## Calculations

Conical flask Burette

$Na_2CO_3$ solution HCl solution

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

$N_{base}$= 0.1 N

$V_{base}$= 10 mL

$N_{acid}$= unknown

$V_{acid}$= (x)mL from burette

So that

$$N_{acid} = \frac{N_{base} \times V_{base}}{V_{acid}}$$

***Calculate the mean normality of your three trials***.

$$Mean = \bar{X} = \frac{X_1 + X_2 + X_3 + .....}{N}$$

$N$ = number of trials

| **Volume of HCl** | **Normality of HCl** |
|---|---|
| | |
| | |
| | |
| X = | |