CHAPTER 1

Acids and Bases

Acids and Bases

First definition of acid was given by Svante Arrhenius (Sweden) in 1884, which states that an acid is a substance that can release a proton or hydrogen ion (H^+). Hydrogen chloride (HCl) in water solution ionizes and becomes hydrogen ions and chloride ions. The acids HCl, HNO_3, H_3PO_3, H_2SO_4, HBr, HI and HCN dissociate completely in water and thus are *strong acids*. A base or alkali, is a substance that can donate a hydroxide ion (OH^-). Sodium hydroxide (NaOH) in water solution becomes sodium ions and hydroxide ions. Bases like, NaOH, KOH, LiOH, $Sr(OH)_2$, $Ba(OH)_2$ and $Ca(OH)_2$ completely dissociates in water and are thus *strong bases*.

Three models of acids and bases

1. **Arrhenius model**

   Basis for the model-action in water

   - Acid definition: produces H^+ (hydrogen ion) in water solution
   - Base definition: produces OH^- (hydroxide ion) in water solution

2. **Bronsted-lowry model**

   Basis for the model-proton transfer

   - Acid definition: donates a proton
   - Base definition: accepts a proton
   - Conjugate acid definition: the acid becomes the conjugate base after it donates the proton because it can now accept it back.
   - Conjugate base definition: the base becomes the conjugate acid after it accepts the proton because it can now donate it back.

3. **Lewis model**

   Basis for model-electron pair transfer

   - Acid definition: accepts a pair of electrons
   - Base definition: donates a pair of electrons

Acid-Base Titration

Acid-base titration is a neutralization titration in which, titration of alkaline solution is carried out with a standard acid solution in order to determine the amount of acid which is exactly equivalent to the amount of base present. The point at which this is reached, is called **theoretical end point** and at this point whole acid reacts with base to form the corresponding salt. For any titration the correct end point will be characterized by a definite value of pH of the solution, which depends on the nature of acid, nature of the base and the concentration of the solution. A large number of substances, acid-base indicators, change the color of the solution according to the concentration of hydrogen ions present. Change in color from acid to alkaline or vice-versa is not sudden but takes place within a small interval of pH and is called **color change interval** of indicator. A list of indicators commonly used in acid-base titration is given in **Table 1.** For most acid-base titrations, an indicator that exhibits sharp color change at the pH close to the equivalent pH, is chosen.

Table 1 Common indicators used in acid-base titrations

Indicator	Color change pH interval	Acid color	Base Color
m-Cresol purple	0.5-2.5	Red	Yellow
Thymol Blue	1.2-2.8	Red	Yellow
Bromophenol blue	3.0-4.6	Yellow	Blue
Methyl orange	3.1-4.4	Red	Yellow
Methyl red	4.2-6.2	Red	Yellow
Phenol red	6.8-8.4	Yellow	Red
Thymol blue	8.0-9.6	Yellow	Blue
Phenolphthalein	8.2-10	Colorless	Red

EXPERIMENT 1.1

To Prepare 1M HCl and Standardize it and Perform Assay of Sodium Hydrogen Carbonate

Theory

Sodium hydrogen carbonate is a white crystalline or amorphous powder and saline in taste. It is freely soluble in water but practically insoluble in alcohol. It is used for its acid neutralizing properties. It is used as an antacid to treat heartburn, indigestion and other stomach disorders. Its aqueous solution is used as local applicant for burns, insect bites etc.

Sodium hydrogen carbonate is a base and it is titrated with dilute hydrochloric acid using methyl orange as indicator (pH: 3.1-4.4) and the equivalence point of the reaction is at 3.6 approximately.

Chemical reaction

$$Na_2CO_3 + HCl \longrightarrow NaHCO_3 + NaCl$$

$$NaHCO_3 + HCl \longrightarrow NaCl + CO_2 + H_2O$$

Chemicals required

1. 1.0 M HCl solution. Fill the volumetric flask up to $3/4^{th}$ with distilled water. Add 85 ml of concentrated hydrochloric acid. Stopper the volumetric flask and shake the solution. Fill the volumetric flask with distilled water up to 1000 ml mark.

   Note: When diluting concentrated hydrochloric acid, remember to add concentrated acid (HCl) to water (and not water to HCl) in order to avoid spattering. Handle the containers carefully as dilution involves generation of heat.
2. Methyl orange as indicator
3. Anhydrous Sodium hydrogen carbonate

Procedure

1. **Standardization of HCl**
   - (i) Weigh accurately 1.5g of anhydrous sodium carbonate and transfer it to conical flask.
   - (ii) Dissolve it in 100mL of distilled water.
   - (iii) Now add 0.1mL of methyl red indicator.
   - (iv) Take 1.0 M HCl in a burette and fill the burette up to mark.
   - (v) Add the acid slowly from the burette to the conical flask containing sodium carbonate solution.
   - (vi) Keep on adding 1.0 M HCl with constant shaking until the appearance of faint pink color.
   - (vii) Heat the solution, cool and continue titration.
   - (viii) Repeat step vi and vii until the faint pink color is no longer affected by continuous heating.
2. **Determination of percent purity of Sodium Hydrogen Carbonate**
   - (i) Weigh accurately 1.5 g of sodium hydrogen carbonate and transfer it to conical flask.
   - (ii) Add 50 mL of distilled water to the conical flask to dissolve sodium hydrogen carbonate.
   - (iii) Fill the burette with standardized 1.0 M HCl up to the mark.
   - (iv) Add 0.2 mL methyl orange as an indicator.
   - (v) Titrate the solution in the conical flask with standardized 1.0 M HCl.
   - (vi) Stir the solution continuously until the appearance of end point i.e. color of the solution changes from orange to pink color.

Observation and Calculations

1. **Standardization of HCl**

   The molarity of HCl is obtained by using the following relation:

$$\text{Molarity of HCl} = \frac{\text{moles of HCl}}{\text{liter}} = \frac{\text{moles of } Na_2CO_3 \times 2}{(\text{ml HCl}/1000)}$$

$$\text{Molarity of HCl} = \frac{\text{wt. of } Na_2CO_3 \times 2}{\text{mol. wt. of } Na_2CO_3 \times \text{volume of HCl (liter)}}$$

2. Percentage purity of $NaHCO_3$

Percentage purity

$$= \frac{\text{Burette reading} \times 0.084 \times \text{Molarity (calculated)} \times 100}{\text{Weight of sample} \times 1.0 \text{ (molarity known)}}$$

EXPERIMENT 1.2

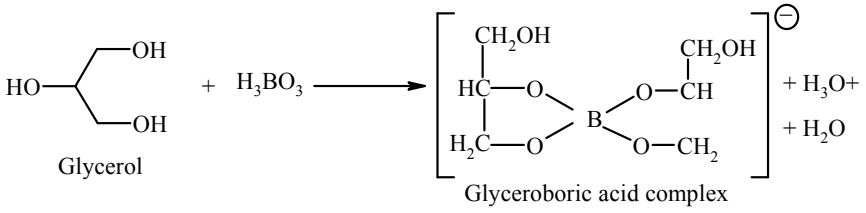
To Prepare and Standardize 0.1 M NaOH and Perform the Assay of Boric Acid

Theory

Boric acid is a solid which is available in three forms: colorless, odorless pearly scales; six-sided triclinic crystals and white odorless powder. Boric acid is a local anti-infective and is used in dusting powders, local antiseptic creams, ointments, lotions and is applied on skin, eyes and mucous membrane. Aqueous solution of boric acid have been used as mouth washes and eye lotions as it possess weak bactericide and fungicide properties.

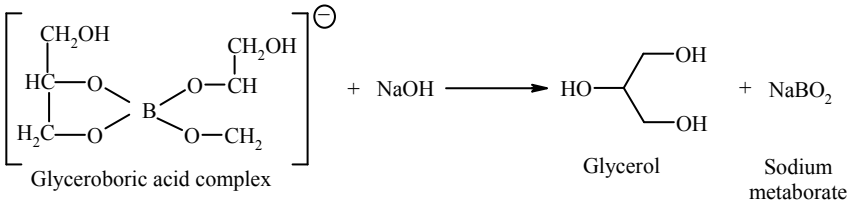
In this titration weak boric acid acts as a strong acid in the presence of glycerin due to the formation of Glyceroboric acid complex. This aids its titration with strong alkali, sodium hydroxide. Phenolphthalein is used as an indicator.

Chemical reaction



Glycerol

Glyceroboric acid complex



Glyceroboric acid complex

Glycerol

Sodium metaborate

Chemicals required

1. 0.1 M sodium hydroxide solution
2. Boric acid
3. Potassium hydrogen phthalate
4. Glycerin
5. Phenolphthalein indicator

Procedure

1. Standardization of sodium hydroxide

(i) Weigh accurately 0.2 g of potassium hydrogen phthalate and transfer it to the conical flask.

(ii) Dissolve it in about 100 mL of water.

(iii) Add 0.1 ml of phenolphthalein as indicator.

(iv) Fill the burette up to mark with sodium hydroxide solution.

(v) Titrate potassium hydrogen phthalate by slowly adding sodium hydroxide solution from burette.

(vi) Continue the titration with continuous shaking until the appearance of light pink color persists for 30 secs.

(vii) Note the reading and again perform the experiment for two more times.

(viii) Take three concordant reading and calculate the morality of given sample of sodium hydroxide.

2. Assay of boric acid

(i) Take 25mL of glycerin in conical flask.

(ii) Add 2 drops of phenolphthalein in it.

(iii) Neutralize the solution by titrating with standardized sodium hydroxide solution.

(iv) Continue the titration till light pink color is obtained.

(v) Now add 50 mL of water in it.

(vi) Take 0.5 mg boric acid in another conical flask.

(vii) Add the above solution in it.

(viii) Shake the solution and add 1 drop of phenolphthalein in it.

(ix) Titrate it with sodium hydroxide by slowly adding the solution from burette.

(x) Continue the titration till permanent pink color is obtained.
(xi) Take the reading by performing the experiment three times.
(xii) Calculate the percentage purity.

Observation and Calculation

Determination of molarity of NaOH

Molarity of NaOH

$$= \frac{\text{Weight of potassium hydrogen phthalate}}{\text{Mol. wt. of pot. hydrogen phthalate} \times \text{volume of NaOH (liter)}}$$

EXPERIMENT 1.3

To Prepare and Standardize 0.1 M NaOH and Perform Assay of Ammonium Chloride

Theory

In this titration aqueous solution of ammonium chloride is first treated with formaldehyde which results in the liberation of hydrochloric acid equivalent to the amount of ammonium chloride present in the solution. This hydrochloric in turn reacts with sodium hydroxide.

Chemical equation

$$4NH_4Cl + 4H_2O \longrightarrow 4NH_4OH + 4HCl$$

Ammonium Chloride — Ammonium hydroxide

$$4NH_4OH + 6HCHO \longrightarrow (CH_2)_6N_4 + 10\ H_2O$$

Formaldehyde — Hexamine

$$4\ HCl + 4NaOH \longrightarrow 4NaCl + 4H_2O$$

Sodium hydroxide

Chemicals required

0.1 M sodium hydroxide solution

Potassium hydrogen phthalate

Phenolphthalein indicator

Procedure

Standardization of sodium hydroxide

(i) Weigh accurately 0.2 g of potassium hydrogen phthalate and transfer it to the conical flask.

(ii) Dissolve it in about 100 mL of water.

(iii) Add 0.1mL of phenolphthalein indicator.

(iv) Fill the burette up to mark with sodium hydroxide solution.

(v) Titrate the above solution by slowly adding sodium hydroxide solution from the burette.

(vi) Continue the titration with continuous shaking until the appearance light pink color persists for 30 secs.

(vii) Note the reading and again perform the experiment for two more times.

(viii) Take three concordant reading and calculate the morality of given sample of sodium hydroxide.

Determination of ammonium chloride

(i) Weigh accurately about 0.1 g ammonium chloride and transfer it to conical flask.

(ii) Dissolve it in 20 mL of distilled water.

(iii) Add a mixture of 5 mL formaldehyde, 0.1 mL phenolphthalein and 20 mL of distilled water.

(iv) Fill the burette with standardized 0.1 M NaOH solution.

(v) Titrate the solution with 0.1 M NaOH using further 0.2 mL phenolphthalein as indicator.

Observations and Calculations

Standardization of sodium hydroxide

Molarity of NaOH

$$= \frac{\text{Weight of potassium hydrogen phthalate}}{\text{Mol. wt. of pot. hydrogen phthalate} \times \text{volume of NaOH (liter)}}$$

Assay of ammonium chloride

Molarity of NaOH

$$= \frac{\text{Volume of NaOH used (ml)} \times \text{molarity (cal)} \times 100}{\text{Weight of sample} \times \text{molarity (given)}}$$

EXPERIMENT 1.4

To Prepare and Standardize 0.5 M H_2SO_4 and Perform Assay of Sodium Carbonate

Theory

It is having not less than 99.0 % and not more than 105.0 % of sodium carbonate. It is used as an antacid and topically as lotion in dermatitis. Sodium carbonate being basic in nature is titrated with dilute sulphuric acid using bromophenol blue indicator. Due to the formation of carbonic acid during the titration, pH of the solution becomes acidic. Thus, bromophenol blue is an indicator of choice (pH range, 3.6-4.4).

Chemical equation

$$Na_2CO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O + CO_2$$

Chemicals required

0.5 M H_2SO_4 solution. Pour 15 mL of H_2SO_4 in 400 mL of distilled water in a 500 mL volumetric flask. Shake the solution and adjust the volume up to 500 mL mark with distilled water.

Bromophenol blue indicator

Anhydrous sodium carbonate

Methyl orange indicator

Procedure

Standardization of H_2SO_4

(i) Weigh accurately about 0.3 g of anhydrous sodium carbonate.

(ii) Transfer the weighed sodium carbonate to a conical flask and add 100 mL of distilled water and dissolve it completely.

(iii) Add 2-3 drops of methyl orange indicator.

(iv) Add 0.5M H_2SO_4 into the burette and fill it up to mark.

(v) Titrate the solution in conical flask with 0.5M H_2SO_4 until the color of solution changes from yellow to orange.

(vi) Take three concordant readings and calculate the molarity of H_2SO_4.

(vii) Each ml of 0.5M H_2SO_4 solution = 0.05299 g or 52.99 mg of sodium carbonate.

Assay of sodium carbonate

(i) Weigh accurately about 0.5 g of sodium carbonate and transfer it to the conical flask.

(ii) Dissolve it in 20 mL of water.

(iii) Titrate the solution with 0.5M H_2SO_4 slowly by adding the solution drop wise to the conical flask.

(iv) Bromophenol blue is used as an indicator.

(v) Continue titration until the color changes from blue to yellow at end point.

Observation and Calculations

Standardization of H_2SO_4

Molarity of H_2SO_4 =

$$\frac{\text{Weight of sod. carbonate (g)}}{\text{Mol. wt. of sod. carbonate} \times \text{vol. of } H_2SO_4 \text{ consumed (liter)}}$$

Assay of sodium carbonate

$$\text{Percentage purity} = \frac{\text{Volume of } H_2SO_4 \text{ used (mL)} \times \text{molarity (cal)} \times 100}{\text{Weight of sample} \times \text{molarity (given)}}$$

EXPERIMENT 1.5

To Determine Percentage Purity of Zinc Oxide

Theory

It does not have more than 99.0% of zinc oxide which has been calculated with reference of the substance which is ignited to constant weight. Zinc oxide occurs as a soft, white, fine powder, free from grittiness. it is used as a mild antiseptic and astringent. It is assistive in the treatment of eczema, ringworm, varicose ulcers and psoriasis.

Assay of zinc oxide is a type of back titration because its reaction with sulphuric acid is slow. Excess of sulphuric acid is added to the reaction mixture and excess of acid is back titrated with sodium hydroxide.

Chemical equation

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

$$H_2SO_4 + 2\ NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

Chemicals required

0.5M H_2SO_4 solution: Pour 15 mL of H_2SO_4 in 400 mL of distilled water in a 500 mL volumetric flask. Shake the solution and adjust the volume up to 500 mL mark with distilled water.

0.5 M NaOH solution.

Methyl red indicator.

Procedure

Standardization of H_2SO_4

(i) Weigh accurately about 0.3 g of anhydrous sodium carbonate.

(ii) Transfer the weighed sodium carbonate to a conical flask and add 100 ml of distilled water and dissolve it completely.

(iii) Add 2-3 drops of methyl orange indicator.

(iv) Titrate the solution with 0.5M H_2SO_4 by slowly adding the solution from burette.

(v) Continue titration until the color of solution changes from yellow to orange.

(vi) Take three concordant readings and calculate the molarity of H_2SO_4.

(vii) Each ml of 0.5M H_2SO_4 solution = 0.05299 g or 52.99 mg of sodium carbonate.

Determination of zinc oxide

(i) Weigh accurately about 1.5 g of zinc oxide and 2.5 g of ammonium chloride.

(ii) Transfer the weighed substance to the conical flask.

(iii) Dissolve them in 50 mL of 0.5 mL sulphuric acid by gentle heating.

(iv) Add few drops of methyl red indicator

(v) Titrate the solution with 0.5M NaOH solution until yellow color appears at the end point.

Observation and Calculations

Standardization of H_2SO_4

Molarity of H_2SO_4 =

$$\frac{\text{Weight of sod. carbonate (g)}}{\text{Mol. wt. of sod. carbonate} \times \text{vol. of } H_2SO_4 \text{ consumed (liter)}}$$

Assay of zinc oxide

Molecular weight of zinc oxide = 81.41

Weight of sample = W g

Volume of 0.5M NaOH consumed by unreacted sulphuric acid = M mL

If the molarity of sodium hydroxide and sulphric acid is same

Volume of 0.5M sulphuric acid reacted with zinc oxide = 50 – M = V mL

From equation

1 mole of ZnO = 1 mole of H_2SO_4 = 1000 mL of 1M H_2SO_4 solution

Each mL of 0.5M H_2SO_4 = 0.04068 g of ZnO

$$\text{Percentage purity} = \frac{0.04068 \times V \times \text{molarity (cal)} \times 100}{W \times \text{molarity (known)}}$$

EXPERIMENT 1.6

To Determine Percentage Purity of Sodium Phosphate

Theory

Sodium phosphate is dodecahydrate of disodium hydrogen orthophosphate. It contains not less than 98.5 % and not more than 101.0 % sodium phosphate. It occurs in the form of colorless transparent crystals, with a saline taste. It is used as saline laxative, cathartic and buffering agent.

In this titration basic sodium phosphate reacts with hydrochloric acid to form sodium dihydrogen phosphate in the presence of bromocresol green indicator.

Chemical equation

$$\underset{\text{Sodium phosphate}}{Na_2HPO_4} + HCl \longrightarrow \underset{\text{Sodium dihydrogen phosphate}}{NaH_2PO_4} + NaCl$$

Chemicals required

1. 0.5M HCl solution. To prepare this solution add 45 mL of concentrated HCl solution into water and make up the volume to 1 liter.
2. Bromocresol green.
3. Sodium phosphate

Procedure

1. **Standardization of 0.5M HCl solution**
   (i) Weigh accurately 1.2 g of anhydrous sodium carbonate and transfer it to the volumetric flask containing 100 mL of distilled water.
   (ii) Dissolve sodium carbonate completely and make up the volume of solution up to 250 mL with distilled water.

(iii) Take 25 mL of this solution in a conical flask and add 2-3 drops of methyl orange indicator.

(iv) Fill the burette with hydrochloric acid solution and make it up to mark.

(v) Titrate the solution with given hydrochloric acid solution until the color changes from yellow to orange at the end point.

(vi) Take three concordant readings and calculate the molarity of hydrochloric acid.

2. Assay of sodium phosphate

(i) Weigh accurately 1.5 g of sodium phosphate

(ii) Transfer the weighed sodium phosphate to the conical flask and dissolve it in 100 mL of distilled water.

(iii) Add a few drops of bromocresol green indicator to the above solution.

(iv) Fill the burette up to mark with standardized hydrochloric acid.

(v) Titrate the solution in the conical flask with standardized hydrochloric acid by adding the solution drop wise from the burette.

(vi) Continue the titration until the appearance of green color at the end point.

Observation and Calculations

1 mole of sodium phosphate = 1 mole of HCl = 1000 ml of 1M HCl solution

Each ml of 0.5M HCl solution = 0.01791 g of sodium phosphate

% purity of sodium phosphate =

$$\frac{0.01791 \times \text{vol.of HCl used (ml)} \times \text{molarity (cal)} \times 100}{\text{Weight of the sample} \times \text{molarity (known)}}$$

EXPERIMENT 1.7

To Determine Percentage Purity of Sodium Benzoate

Theory

Sodium benzoate is sodium salt of benzoic acid which is derived from a reaction of benzoic acid with sodium hydroxide. It is used as a preservative in carbonated drinks, jams, fruit juices, pickles and condiments. Sodium benzoate liberates sodium hydroxide which in turn reacts with hydrochloric acid. Reaction is carried out using bromophenol blue as indicator and ether is added to dissolve benzoic acid formed in the reaction.

Chemical equation

$$C_6H_5COONa + HCl \longrightarrow C_6H_5COOH + NaCl$$

Chemicals required

1. 0.5M *HCl solution*: To prepare this solution add 45 mL of concentrated HCl solution into water and make up the volume to 1 liter.
2. Bromophenol blue indicator.
3. Ether.

Procedure

1. **Standardization of 0.5M HCl solution**
   (i) Weigh accurately 1.2 g of anhydrous sodium carbonate and transfer it to the volumetric flask containing 100 mL of distilled water.
   (ii) Dissolve sodium carbonate completely and make up the volume of the solution up to 250 mL with distilled water.
   (iii) Take 25 mL of this solution in a conical flask and add 2-3 drops of methyl orange indicator.
   (iv) Fill the burette with given hydrochloric acid solution.

(v) Titrate the solution in the conical flask with hydrochloric acid by adding the solution drop wise from the burette.

(vi) Continue titration until the color changes from yellow to orange at the end point.

(vii) Take three concordant readings and calculate molarity of hydrochloric acid.

2. Assay of sodium phosphate

(i) Weigh accurately 1.5 g of sodium benzoate and transfer it to a 250 mL stoppered-conical flask.

(ii) Dissolve it in 25 mL of distilled water and add 75 mL of ether.

(iii) Add 2-3 drops of bromophenol blue indicator.

(iv) Fill the burette with standardized 0.5M hydrochloric acid solution.

(v) Titrate the solution with 0.5M hydrochloric acid with continuous shaking in order to mix water and ether layer.

(vi) Continue titration until light green color persists in the water layer.

Observation and Calculations

Assay of sodium benzoate

1 mole of sodium benzoate = 1 mole of HCl = 1000 mL of 1M HCl solution

Each ml of 0.5M HCl solution = 0.07205 g of sodium benzoate

% purity

$$= \frac{0.07205 \times \text{vol.of 0.5M HCl used (mL)} \times \text{molarity (cal)} \times 100}{\text{Weight of sodium benzoate used} \times \text{molarity (known)}}$$

EXPERIMENT 1.8

Determination of Percent Purity of Sodium Potassium Tartrate

Theory

Sodium potassium tartrate is a double salt and is popularly called Rochelle salt. It has been used medicinally as a laxative. It is also used in silvering of mirrors. It is an ingredient of Fehling solution. In this experiment, sodium potassium tartrate is ignited and salt gets converted to respective carbonates (sodium carbonate and potassium carbonate) equivalent to the amount of salt. If the salt is heated strongly, alkali carbonates may be converted to respective oxides. These carbonates formed are basic in nature and reacts with sulphuric acid during the titration. It is a type of back titration and the excess of acid is titrated with sodium hydroxide.

Chemical equation

$$2\ C_4H_4O_6\ NaK\ .\ 4H_2O \longrightarrow Na_2CO_3 + K_2CO_3$$

Chemicals required

1. Sodium potassium tartrate.
2. Methyl red-methylene blue TS
3. 0.5N NaOH

Procedure

1. **Standardization of 0.5N sodium hydroxide**
   (i) Weigh accurately 0.2 g of potassium hydrogen phthalate.
   (ii) Transfer the weighed potassium hydrogen phthalate to the conical flask and dissolve it in about 100 mL of water.
   (iii) Add 0.1mL of phenolphthalein indicator.
   (iv) Fill the burette up to mark with given sodium hydroxide solution.

(v) Titrate the solution in conical flask with given sodium hydroxide solution by adding the solution drop wise.

(vi) Continue the titration with continuous shaking until the light pink color persists for 30 sec.

(vii) Note the reading and again perform the experiment twice.

(viii) Take three concordant reading and calculate the morality of given sample of sodium hydroxide.

2. Assay of sodium potassium tartrate

(i) Weigh accurately 2 g of sodium potassium tartrate and transfer it to a dry crucible.

(ii) Gently ignite the crucible in order to carbonize the salt thoroughly.

(iii) Cool the crucible and place it in a glass beaker.

(iv) Transfer the carbonized mass to the glass beaker with glass rod.

(v) Add 50 mL of water and 50 mL 0.5N sulphuric acid solution.

(vi) Boil the solution for 30 min.

(vii) Filter the solution and wash with water until it is neutral to litmus.

(viii) Cool the combined filtrate, add methyl red-methylene blue TS as indicator.

(ix) Titrate the above solution by slowly adding 0.5M sodium hydroxide solution from the burette.

Observation and Calculations

$$\% \text{ purity of sod. Pot. Tartrate} = \frac{a - b \times \text{eq. wt.} \times 100}{\text{Sample weight}}$$

a = amount of sulphuric acid used

b = amount of NaOH used in back titration

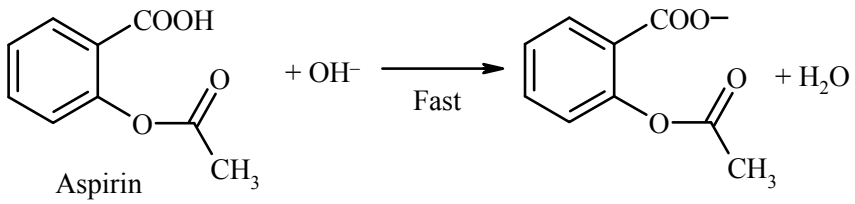
EXPERIMENT 1.9

To Determine Aspirin Content in Tablet Formulation

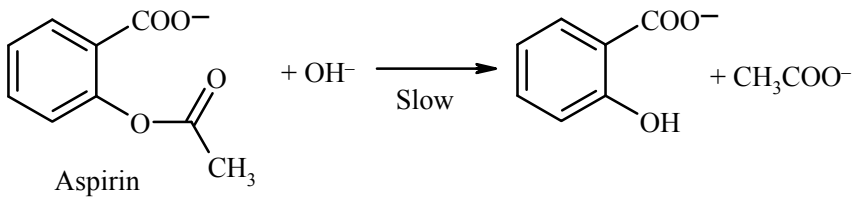
Theory

Aspirin belongs to a group of drugs called salicylates. It inhibits the synthesis of substances (prostaglandins) in the body which are responsible for pain, fever, and inflammation. Aspirin is used to treat mild to moderate pain, and also to reduce fever or inflammation. It is sometimes used to treat or prevent heart attacks, strokes and angina. Aspirin is a weak acid that undergoes slow hydrolysis; i.e., each aspirin molecules react with two hydroxide ions. To overcome this problem, a known excess amount of base is added to the sample solution and HCl titration is carried out to determine the amount of unreacted base. This is subtracted from the initial amount of base to find the amount of base that is actually reacted with the aspirin and hence the quantity of aspirin in the analyte.

Chemical equation



Aspirin



Aspirin

Chemicals required

1. 0.1M HCl solution. Add 8.5 mL of HCl solution to 100 mL of water. Make up the volume to 1 liter with water.
2. 0.1M NaOH solution. Dissolve 4.2 g of NaOH in 1000 mL of water.
3. Aspirin.
4. Phenolphthalein indicator.
5. Phenol red indicator.

Procedure

1. Weigh accurately 20 tablets and calculate the average weight of tablets.
2. Powder the tablets.
3. Accur weigh powder equivalent to 0.5 g of aspirin.
4. Dissolve in 10 to 15 mL alcohol and add 4 drops of phenolphthalein indicator.
5. Titrate each sample quickly to the first persistent faint pink color with standard 0.1M NaOH solution.
6. Record this volume.
7. Then add, same volume again +5 mL excess from the burrette.
8. Place the flasks on the steam bath for 45 minutes to allow reaction to proceed to completion.
9. Then back-titrate the excess base with your standard 0.1M HCl solution using phenol red as indicator.

Observation and Calculations

1. **Determination of aspirin content in tablet**

   Volume of 0.1M NaOH added = 50 mL.

   Moles of NaOH = $0.050 \times 0.10 = 5.0 \times 10^{-3}$.

   Suppose, volume of 0.1M HCl required to react with excess of unreacted 0.1M NaOH solution is 30 mL.

   Moles of HCl = Moles of excess of NaOH = $0.030 \times 0.10 = 3.0 \times 10^{-3}$

   Moles of NaOH which reacted with aspirin:

   $$5.0\times10^{-3} - 3.0\times10^{-3} = 2.0\times10^{-3} \text{ or } \frac{2.0\times10^{-2}}{2} \text{ moles of aspirin}$$

   Mass of aspirin = $1.0 \times 10^{-3} \times 180 = 0.180$ g of aspirin