CHAPTER-9

ACIDS, BASES AND SALTS

You will learn in this chapter about:

* Acids and bases.
* Arrhenius theory.
* Bronsted lowry theory.
* Lewis concept.
* Properties of acids and bases.
* Neutralization.
* Basicity and acidity.
* Strength of acids and bases.
* Salts.
* Classification of salts and some important salts.
* Dissociation of water.
* The concept of pH and pH scale.
* The importance of pH.
* Acid-base titration.
* Standard solution.

9.1 ACIDS AND BASES:

Introduction:

By the 15th century, chemists recognized a group of substances which had sour taste called acids (In Latin acidus meaning 'sour'). They also recognized another group of substances which had bitter taste and were used as good cleaning agents, called Bases.

In the 16th century, it was recognized that bases react with acids to 'destroy' or "neutrailize" them, forming an ionic compound called salt.

Nearly all fruits and foods contain acids. Lemons, oranges, grape-fruits, contain citric acid. All citrous fruits contain large amount of ascorbic acid ($C_6H_8O_6$), or vitamin-C. Ascorbic acid also acts as an anti-oxidant. Apples contain maleic acid. The souring of milk produces lactic acid, butter on rancidity gives butyric acid. The extract of vinegar is acetic acid. Chemists prepare large quantities of important industrial acids. They are manufactured from minerals and are known as inorganic acids or more commonly mineral acids. These include hydrochloric acid (HCl), nitric acid (HNO_3), sulphuric acid (H_2SO_4) and phosphoric acid (H_3PO_4). The most important acid is H_2SO_4. The consumption of (H_2SO_4) is an index to the state of civilization and prosperity of a country. The important acid for making explosives and fertilizers is (HNO_3), and (HCl) is used as cleaning agent. It composes about 0.4% of gastric juice of our stomach and aids in digestion of food.

There are several substances found in almost every home, called bases. House hold ammonia (NH_3+H_2O solution) is a common cleaning agent. Lye is commercial (NaOH) used for cleaning, sink–drains. Lime water is a solution of $Ca(OH)_2$. Milk of magnesia ($Mg(OH)_2$) is used as an antacid, laxative, and an antidote, when strong acids are swallowed.

Salts have a positive ion other than (H^+) combined with a negative ion, other than (OH^-). However, one must keep in mind that all the salts are not neutral, some behave like acids and others like bases. Thus many substance were grouped into one of the three classes, acids, bases, and salts.

9.2 PROPERTIES OF ACIDS AND BASES

Acids:

1. Acids have a sour taste, vinegar gives the taste of acetic acid and lemons and other citrous fruits contain citric acid.
2. They change the colour of blue litmus to red.
3. Acids react with certain metals, such as (Zn, Mg and Fe) to produce (H_2) gas, for example, when diluted, (HCl) reacts with (Zn) metal, producing (H_2) gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(s) + H_2(g)$$

4. Acids react with carbonates and bicarbonates, such as Na_2CO_3, $CaCO_3$ and $NaHCO_3$ to produce CO_2 gas.

$$CaCO_3(s) + 2\,HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$
$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

5. Acids react with oxides and hydroxide of metals, forming salt and water.

$$Fe\,O(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2O(l)$$
$$Fe(OH)_3(s) + 3HCl(aq) \longrightarrow FeCl_3(aq) + 3H_2O(l)$$

6. Aqueous acid solutions, conduct electricity.
7. They react with bases to form salts and water.

Bases:

1. Bases have bitter taste.
2. Bases have slippery touch.
3. They change the colour of red litmus to blue
4. Aqueous solutions of bases, conduct electricity.
5. They react with acids to form salts. When acids and bases are mixed in right proportions, the acidic and basic properties disappear and new substances salt and water are formed and the reaction is neutralization reaction.

9.2.1 Acids and Bases (Arrhenius theory):

Svante-Arrhenius, a Swedish chemist in (1887) first gave the clue to chemical nature of acids in his theory of ionization. It states, that an acid can be defined as a substance, that yields hydrogen (H^+) ions when dissolved in water. A base can be defined as a substance that yields hydroxide (OH^-) ions, when dissolved in water. Thus (HCl) and (H_2SO_4) are acids and (NaOH) and (KOH) are bases.

$$\text{Arrhenius acid}\quad HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$$

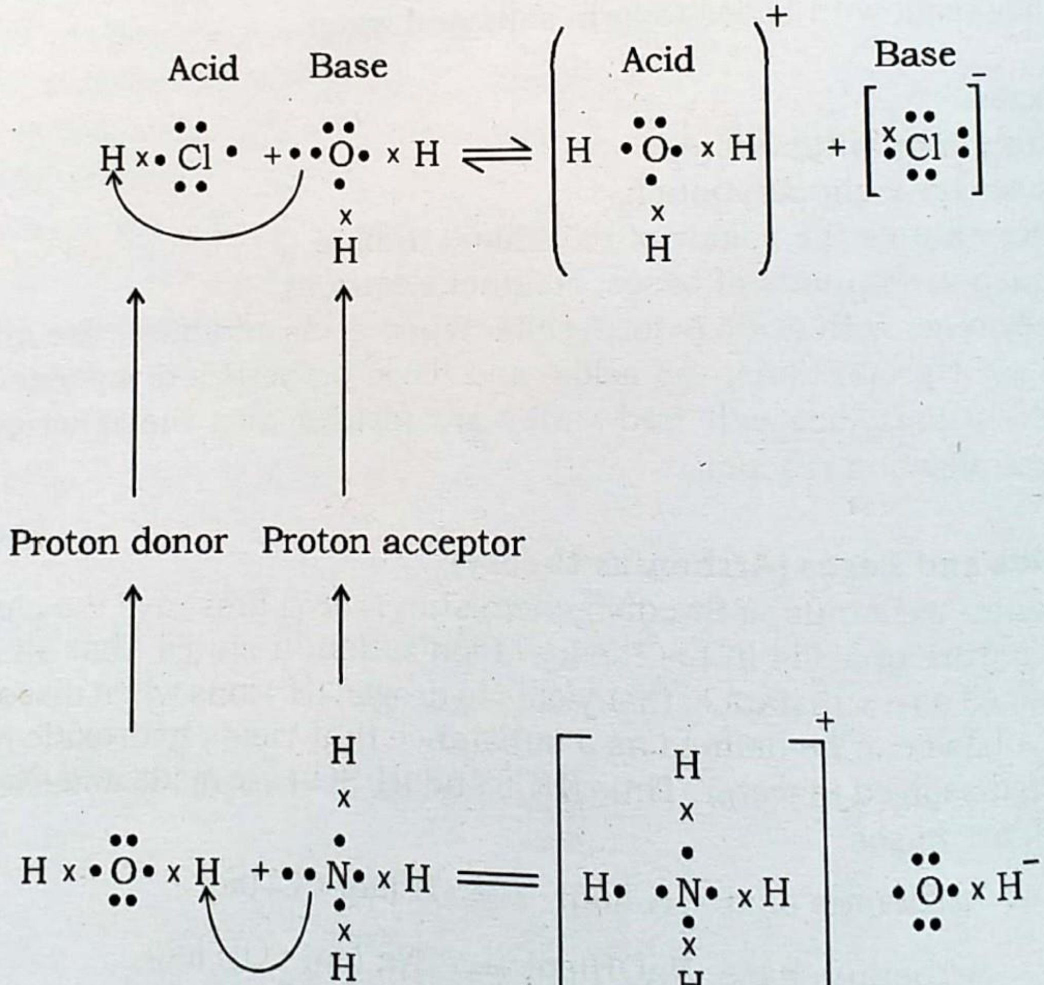
$$\text{Arrhenius base}\quad NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$$

Bronsted Lowry Theory (Proton-Donor and Acceptor Theory):

Arrhenius's definitions of acids and bases are limited in that, they apply only to water (aqueous) solutions and it also does not account for the basicity of ammonia (NH_3), that doesn't contain (OH) group. Broader definitions, which were proposed by the Danish Chemist Johannes Bronsted and English Chemist Thomas Lowry, in 1923, describe, that *An acid is a substance having a tendency to donate one or more protons and base is a substance, having a tendency to accept (add) protons.*

Bronsted-Lowry acid = A substance that can donate (H^+)
Bronsted-Lowry base = A substance that can accept (H^+)

In the above examples, the (HCl) and water (H_2O), are proton donors and act as Bronsted-Lowry acids, whereas (H_2O) and ammonia (NH_3) are proton acceptors and are known as Bronsted Lowry bases.



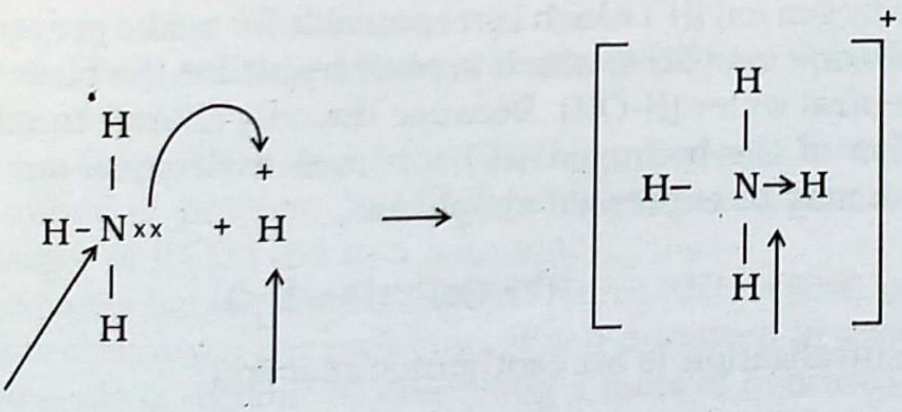
LEWIS CONCEPT OF ACIDS AND BASES
(Electronic explanation of Acids and Bases):

In 1923 G. N. Lewis proposed a more general concept of acids and bases. It explains the behaviour of acids in terms of electrons rather than protons, since electrons are responsible for chemical bonding. Lewis acids include not only (H^+) protons, but also other cations and neutral molecules, having vacant valence orbitals that can accept a pair of electrons donated by Lewis base.

According to the Lewis theory, an acid is any species (molecule or ion) which can accept a pair of electrons, and base is any species (molecule or ion) which can donate a pair of electrons. An acid-base reaction, in which electron pair donor is base and electron-pair acceptor is acid, they form a co-ordinate covalent bond between the two.

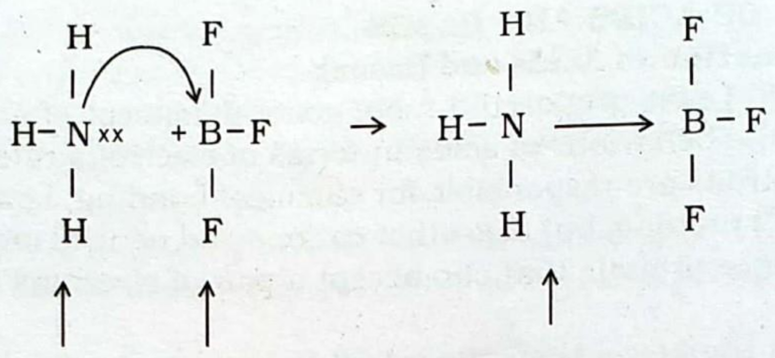
Lewis Acid = An electron pair acceptor.
Lewis Base = An electron pair donor.

For Example: When ammonia (NH_3) reacts with proton (H^+) to form ammonium ion (NH^+_4), in which the nitrogen of (NH_3) donates a pair of electrons whereas the (H^+) accepts that pair of electrons for bond formation, this is shown by curved arrow.



Electron pair donor (Lewis base) Electron pair acceptor (Lewis acid) Co-ordinate covalent bond

Another example is provided by the reaction of ammonia (NH_3) with boron tri-fluoride (BF_3), in which nitrogen of (NH_3) donate an electron pair and (B) of BF_3, which lacks a pair of electrons to complete its outer most shell (octet), accepts that pair of electrons and form a co-ordinate covalent bond.



(Lewis base) (Lewis acid) Acid base adduct

9.2.2 Neutralization

If we add an acid to a base drop by drop, the acidic character of the acid decreases gradually. A stage will come when the resultant solution will become neutral to litmus. This stage is called neutralization. In other words when equivalent quantities of an acid and a base are mixed, salt and water are formed. A common example is the reaction between HCl and NaOH.

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

The hydrogen ion (H^+) which is responsible for acidic properties, reacts with the hydroxide ion (OH^-), which is responsible for the basic properties, producing neutral water (H-OH). Because the only change that takes place is the reaction of the hydrogen (H^+) ion and hydroxide ion (OH^-), the neutralization may be expressed simply as:

$$H^+ + OH^- \longrightarrow H_2O$$

The neutralization is an exothermic reaction.

e.g. $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$ + 13700 cal/mol (57.24KJ/mol)

In case either acid or base is not completely ionized, i.e. (weak acid or weak base), the neutralization reaction may not go to completion, and the heat of neutralization may be less than 13700 cal. For example, when strong acid (HCl) reacts with weak base Ca $(OH)_2$, the amount of heat evolved is 24700 cal/mol.

$$2\,HCl + Ca\,(OH)_2 \longrightarrow CaCl_2 + 2H_2O + 24700 \text{ cals}$$

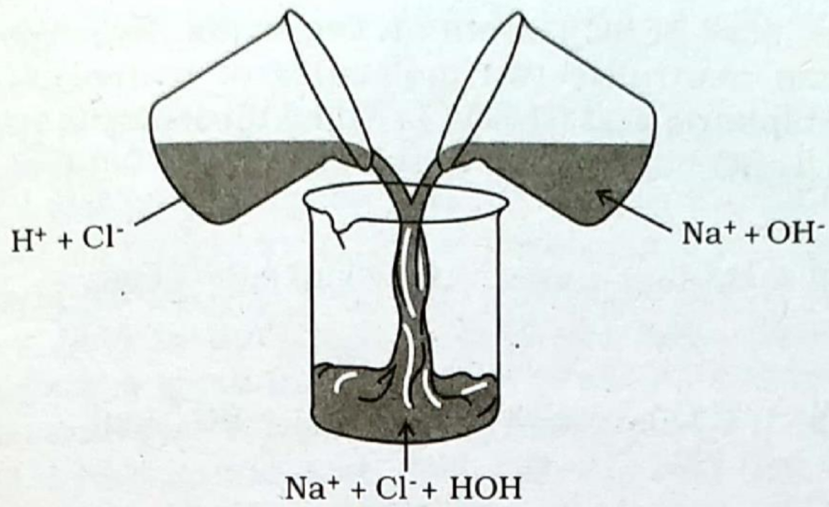


Fig. 9.1 Shows when solutions of hydrochloric acid and sodium hydroxide are mixed, only hydrogen ions and hydroxide ions react forming water. Therefore, we may write the equation: $H^+(aq) + OH^-(aq) \rightarrow HOH(l)$

REMEMBER

A reaction in which an acid and a base form an ionic compound (salt and water) is called neurtralization reaction. If water is formed, the reaction can be classified as a double-replacement reaction.

9.2.3 Mono and Poly Acids and Bases (Mono Basic and Poly Basic Acids) (Basicity of Acids)

As we have seen that an acid yields the protons. Different acids have different number of protons (acidic-hydrogen) per molecule and yield different number of (H_3O^+) ion in a solution.

The common acids like (HCl) (HNO_3) and (CH_3COOH) contain only one acidic hydrogen atom per molecule. When dissolved in water, 1 mole of each of these acids is capable of producing 1 mole of hydrated H^+ = (H_3O^+) ions, and in order to neutralize this solution,1 mole of (OH^-)ions is required. Consequently these acids are called mono-basic acids, more commonly called mono-protic acids

e.g.

$$H^{+\delta} - Cl^{-\delta}(aq) + H^{+\delta} - OH^{-\delta}(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

The number of replaceable or ionizable hydrogen atoms present in a molecule of an acid is called 'Basicity" of the acid.

Sulphuric acid (H_2SO_4) contains two acidic hydrogen atoms per molecule. It can neutralize two molecules of hydroxide (OH^-) ions. Consequently sulphuric acid (H_2SO_4) is called dibasic acid (Diprotic acid). Sulphuric acid (H_2SO_4) dissociates in two steps :

$$H_2SO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{-2}(aq)$$

The acid like phosphoric acid (H_3PO_4), which contains three acidic hydrogen atoms per molecule can neutralize three molecules of (OH^-)ions, so phosphoric acid (H_3PO_4) is called tri basic acid or (Tri-protic acid).

$$H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + H_2PO_4^{-1}(aq)$$

$$H_2PO_4^{-1}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{-2}(aq)$$

$$H_2PO_4^{-2}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + PO_4^{-3}(aq)$$

Acids that contain two, three or more acidic hydrogen per molecule, are called poly-basic acids, or more commonly poly-protic acids.

9.2.4 Mono-Acid and Poly-Acid Bases (Acidity of Bases):

Similarly, bases that produce 1 mole of (OH^-) ions per mole of base (such as NaOH and KOH) are called mono-acid bases. Bases that produce 2 moles of (OH^-) ions per one mole of base (such as $Ca(OH)_2$ and $Ba(OH)_2$) are called di-acid bases, and bases that produce 3 moles of (OH^-) ions per 1 mole of base (such as $Al(OH)_3$ and $Cr(OH)_3$ are called tri-acid bases.

The number of ionizable or replaceable ($-OH^-$) ions, present in a molecule of base is called acidity of the base.

Bases that contain two three or more hydroxide (OH^-) ions per molecule are called poly-acid bases.

9.3 DISSOCIATION OF ACIDS AND BASES (Acid Strength and Base Strength)

Different acids differ in their ability to donate protons. A strong acid is one, that is almost completely dissociated (strong electrolytes) i.e. an acid that produces large number of (H^+) ions in aqueous solution is said to be a strong acid. Typical examples of strong acids are hydrochloric acid (HCl) nitric acid (HNO_3) and sulphuric acid (H_2SO_4).

A weak acid is one, that is only partially dissociated (weak-electrolytes). Only a small fraction of the weak acids transfer a proton to water. Typical examples of weak acids are nitrous acid (HNO_2), phosphoric acid (H_3PO_4), hydro fluoric acid (HF) carbonic acid (H_2CO_3) acetic acid (CH_3-COOH) and formic acid (HCOOH).

For example: In strong acid, greater is the extent of ionization in water.

$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$ (complete ionization in dilute solution)

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$ (8.5% in 0.1 M solution)

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$ (1.3% in 0.1 M solution)

Similarly, the strong base is one, that is almost dissociated completely (strong electrolytes), that is, a base which yields large number of (OH^-) ions in aqueous solution. Most metal hydroxides, such as NaOH, KOH, $Ba(OH)_2$ and $Ca(OH)_2$ are strong electrolytes and strong bases.

A weak base is one, that is only partially dissociated (weak electrolytes). Weak bases dissociate to a small extent in water to yield (OH^-) ions.

Typical examples of weak bases are, NH_4OH, $Mg(OH)_2$ and $Be(OH)_2$, etc. Hence, the relative strength of weak bases may be measured by the extent to which they dissociate in water to yield hydroxide ions (OH^-).

9.4 SALTS

A salt is ionic compound produced when an acid is neutralized by a base. For example, sodium hydroxide neutralizes hydrochloric acid to form sodium chloride (salt) and water.

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

Similarly potassium hydroxide neutralizes nitric acid to form potassium nitrate and water.

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

On the basis of their chemical nature, salts can be divided into three groups.

1. Normal Salts   2. Acidic salts.   3. Basic salts.

1. Normal Salts:

Salts which are formed by the complete neutralization of an acid by a base e.g. NaCI, $NaNO_3$, K_2SO_4 etc are normal salts. These salts do not have replaceable hydrogen atoms or hydroxyl groups.

2. Acidic Salts:

Salts which are formed by the partial neutralization of an acid by a base e.g. $NaHSO_4$, $KHCO_3$, etc are acidic salts.

$$NaOH(aq) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + H_2O(l)$$

$$KOH(aq) + H_2CO_3(aq) \longrightarrow KHCO_3(aq) + H_2O(l)$$

These salts contain replaceable hydrogen ion. They react further with bases to form normal salts.

3. Basic Salts:

Salts, which are formed by the partial neutralization of a base by an acid. e.g. Mg (OH) CI, Zn(OH)CI, etc are basic salts.

$$Mg(OH)_2 + HCI(aq) \longrightarrow Mg(OH)Cl + H_2O(l)$$

These salts have replaceable hydroxyl groups. They can further react with acids to form normal salts.

DOUBLE SALTS

The crystalline compounds which are obtained, when two specific salts are crystallized together are known as double salts. These salts have definite chemical composition. These compounds usually have definite number of water molecules with them. Typical examples of double salts are:

Potash Alum	$K_2SO_4 . Al_2(SO_4)_3 . 24H_2O$
Chrome Alum	$K_2SO_4 . Cr_2(SO_4)_3 . 24H_2O$
Carnalite	$KCl . MgCl_2 . 6H_2O$
Mohr's Salt	$FeSO_4 . (NH_4)_2SO_4 . 6H_2O$

9.4.1 Some important Commercial Preparation and uses of Salts.

1. Sodium carbonate: ($Na_2CO_3.10H_2O$)

Today sodium carbonate (Na_2CO_3) is commercially prepared by the **Solvay process** or ammonia soda process.

The raw materials are lime stone ($CaCO_3$), sodium chloride (NaCl), ammonia (NH_3) and water.

The process involves the following steps:

[i] Lime stone [$CaCO_3$] is heated to yield calcium oxide (quicklime CaO) and the CO_2 gas.

$$CaCO_3(s) \xrightarrow{\Delta H} CaO(s) + CO_2(g)$$

[ii] This (CO_2) is passed into aqueous solution of ammonia, and the ammonium bicarbonate is produced.

$$NH_3(g) + CO_2(g) + H_2O(l) \longrightarrow NH_4HCO_3(aq)$$

[iii] This (NH_4HCO_3) reacts with aqueous cold solution of (NaCl) at 15°C, called **Brine** to yield, sodium bicarbonate ($NaHCO_3$), which is not soluble at low temperature (15°C) and this precipitates out.

$$Na^+Cl^-(aq) + NH^+_4HCO^-_3(aq) \xrightarrow{\text{At } 15^\circ C \text{ brine}} NH^+_4Cl^-(aq) + NaHCO_3(s)$$

[iv] This ($NaHCO_3$) on heating yields sodium carbonate.

$$2NaHCO_3(s) \xrightarrow{\Delta H} Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

The ammonia (NH_3) which is used as a raw material in 2nd step, is recovered by reacting (CaO) with NH_4Cl.

$$2NH_4Cl(aq) + CaO(s) \longrightarrow 2NH_3(g) + CaCl_2(aq) + H_2O(l)$$

Anhydrous sodium carbonate (Na_2CO_3) is known as soda-ash and sodium carbonate decahydrate ($Na_2CO_3.10H_2O$) is commonly known as washing soda.

Uses of Sodium Carbonate ($Na_2\ CO_3$):

Sodium carbonate is soluble in water and has many important applications.

[a] It is used in the softening of water. Sodium carbonate furnishes carbonate ion (CO_3^{-2}) to precipitate calcium and magnesium ions,

$$CaCl_2 + Na_2\ CO_3 \longrightarrow CaCO_3\downarrow + 2NaCl$$
$$MgCl_2 + Na_2\ CO_3 \longrightarrow MgCO_3\downarrow + 2NaCl$$

[b] It is used as cleaning agent, and in making of soap, detergents and paper.

[c] It is used in making ordinary glass, which is used in bottles.

e.g. $$Na_2CO_3(s) + SiO_2(s) \longrightarrow \underset{\text{(water glass)}}{Na_2SiO_3(s)} + CO_2\uparrow$$

2. Sodium Hydrogen carbonate ($NaHCO_3$) (Baking Soda):

Preparation:

Sodium hydrogen carbonate ($NaHCO_3$) or baking soda is formed by *"Solvay process"*, but mostly it is prepared by passing the stream of CO_2 through concentrated aqueous (Na_2CO_3) solution.

$$Na_2CO_3(s) + CO_2(g) + H_2O(l) \longrightarrow 2NaHCO_3(s)$$

Uses:

1. Baking soda is used in the preparation of baking powder
2. In the preparation of effervescent drinks and fruit salts
3. In medicines to remove acidity of stomach (i.e. as Antacid).
4. In fire extinguishers.

3. Copper Sulphate ($CuSO_4.5H_2O$):

Copper sulphate or cupric sulphate which is also known as blue vitrol or blue stone may be prepared by reacting copper scraps with dilute sulphuric acid in the presence of air.

$$2Cu(s) + 2H_2SO_4(aq) + O_2(g) \longrightarrow 2CuSO_4(aq) + 2H_2O$$

It can also be prepared by the treatment of CuO or $CuCO_3$ with dilute sulphuric acid (H_2SO_4).

$$CuO + \text{dilute } H_2SO_4 \longrightarrow CuSO_4 + H_2O$$

$$CuCO_3 + \text{dilute } H_2SO_4 \longrightarrow CuSO_4 + CO_2 + H_2O$$

Uses of Copper Sulphate:

1. In textiles (mordant), tanning, electric batteries, hair dyes and in electroplating.
2. As germicide, insecticide, preservative for wood and paper pulp.
3. In calico printing, making synthetic rubber and copper salts e.g scheels, green paint,
4. In paint and varnish industry.
5. A mixture of copper sulphate and milk lime is used to kill fungus and molds.

4. Magnesium Sulphate ($MgSO_4.7H_2O$) (Epsom-Salt):

Preparation:

It is prepared by the action of H_2SO_4 and magnesite or dolomite, but nowadays it is prepared by heating kiesserite under pressure with water.

(i) $\underset{\text{Magnesite}}{MgCO_3} + H_2SO_4 \longrightarrow MgSO_4 + H_2O + CO_2(g)$

(ii) $\underset{\text{Dolomite}}{MgCO_3.CaCO_3} + 2H_2SO_4 \longrightarrow MgSO_4 + CaSO_4 + 2H_2O + 2CO_2(g)$

(iii) $\underset{\text{Kiesserite}}{MgSO_4.H_2O} + 6H_2O \xrightarrow{\text{pressure}} \underset{\text{(Epsom salt)}}{MgSO_4.7H_2O}$

Uses:

1. It is used as a mild purgative in medicines.
2. In dyeing and tanning processes.
3. In making fire proof fabrics.
4. As a filler in paper industry.
5. In manufacture of ceramics, glazed tiles and match boxes.

5. **Potash Alum** **($K_2SO_4.Al_2(SO_4)_3.24H_2O$):**

If equi-molecular quantities of potassium sulphate and aluminium sulphate are dissolved in water and the solution is allowed to evaporate. crystals of ($K_2SO_4.Al_2(SO_4)_3.24H_2O$) which is called ordinary alum or potash alum are seperated out.

Uses:

1. Alum is used in dyeing as mordant to fix insoluble dye to fibre.
2. It is also used in tanning leather.
3. In sizing paper.
4. In purifying water.
5. As an antiseptic and as a mouth wash.
6. It also used in medicines.

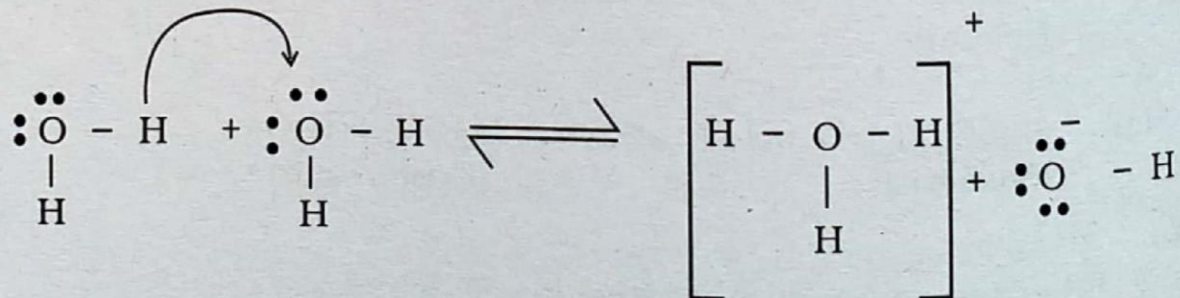
9.5 DISSOCIATION OF WATER (The acid-base properties of water):

One of the most important properties of water is its ability to act both as an acid and as a base. In presence of an acid, water acts as a base where as in the presence of base, water acts as an acid.

$$\text{Acid} + \text{Base} \rightleftharpoons \text{Acid} + \text{Base}$$
$$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$$

$$H_2O(l) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

A substance (such as water) that can behave as both an acid and a base is said to be an amphoteric substance. The amphoteric nature of water is best seen in its self ionization. A proton from one water molecule is transferred to another water molecule, leaving behind (OH^-) ion and forming (H_3O^+) ion.



9.5.1 The Ion Product of Water:

In the study of acid-base reaction in aqueous solution, the important quantity is the (H^+) ions concentration, expressing the proton as (H^+) rather than (H_3O^+). We can write the equilibrium constant (K_c) for the auto ionization water.

$$H_2O \rightleftharpoons H^+(aq) + OH^-(aq)$$

$$Kc = \frac{[H^+][OH^-]}{H_2O}$$

Since a very small fraction of water molecules are ionized, the concentration of water, that is $[H_2O]$ remains mostly unchanged, therefore,

$$K_c\,[H_2O] = K_W = [\overset{+}{H}]\,[O\overline{H}]$$

The equation constant (K_W) is called the ion-product constant, which is the product of molar concentration of (H^+) and (OH^-) ions at a particular temperature.

In pure water at 25°C,the concentration of H^+ and OH^- ions is equal and found to be:

$$[H^+] = 1x10^{-7}\ M \text{ and } [O\overline{H}] = 1x10^{-7}\ M$$

Thus $$K_W = (1x10^{-7}) \times (1x10^{-7}) = 1x10^{-14}$$

Whenever $[\overset{+}{H}] = [O\overline{H}]$, the aqueous solution is said to be neutral, if the number of (H^+)ions increase, the aqueous solution is acidic. When the number of (OH^-) ions increase, the aqueous solution is basic.

For example, the (H^+) ion concentration of any solution is $1x10^{-4}$, then the (OH^-) ion concentration must change to,

$$K_W = [\overset{+}{H}]\,[O\overline{H}] = 1x10^{-14}$$

$$[OH^-] = \frac{K_W}{[H^+]} = \frac{1\times10^{-14}}{1\times10^{-4}}$$

$$= 1\times10^{-14}\times10^{4}$$

$$= 1\times10^{-10}M$$

$$\boxed{\therefore\quad OH^- \text{ ion concentration} = 1 \times 10^{-10}\ M}$$

9.5.2 The Concept of (pH):

It is well known fact that the strength of acidic solution is measured in moles per litre of a solution. A one molar (1M) solution of HCl contains 1mole of (H_3O^+) ions in each litre (dm^3) of solution. In pure water (or any neutral solution) there are $1x10^{-7}$M (moles per litre) of H_3O^+ ions and equal number i.e. ($1x10^{-7}$M) of OH^- ions. To avoid the use of complex numbers, such as $1x10^{-7}$M or $1x10^{-10}$ M, to express the concentration of H^+ and OH^- ions, it is convenient to express the acidity or basicity of a solution in terms of pH. In 1909 the Danish Chemist S.P.L. Sorensen proposed that only the number in the exponent be used to express the acidity, called (pH) from the French (Pouvoir hydrogen = hydrogen power). On this scale, a concentration of ($1x10^{-7}$) moles of H_3O^+ ions per litre of solution becomes a pH of 7. Similarly a concentration $1x10^{-10}$M, becomes a pH of 10, and so on. *Thus (pH) of a solution is defined as the negative logarithm of the hydrogen ion (H^+) concentration or (H_3O^+) ion concentration (in moles per litre).*

Mathematically, we can write pH = -log [H^+]

Thus, a pure water (or any neutral solution) in which (H^+) ion concentration is $1x10^{-7}$ M, has a pH of 7.

$$\text{Since } pH = -\log[H^+]$$
$$pH = -\log[10^{-7}]$$
$$pH = -[-7]$$
$$\therefore \quad pH = 7$$

Similarly pOH is the negative logarithm of hydroxide ion (OH^-) concentration.

Mathematically, we can write pOH = -log [OH^-]

The sum of pH and pOH of a solution is always equal to 14.

i.e $\boxed{pH + pOH = 14}$

Hydrogen and Hydroxide Ion Concentration in Various Acidic and Basic Water Solutions at 25°C

Degree of Acidity	$[H^+]$	$[OH^-]$	Preparation	pH	pOH
Very acidic	10^{+1}	10^{+15}		-1	15
	10^{0}	10^{-14}		0	14
	10^{-1}	10^{-13}		1	13
Medium Acidity	10^{-2}	10^{-12}	Add sufficient acid to water	2	12
	10^{-3}	10^{-11}		3	11
	10^{-4}	10^{-10}		4	10
Slightly Acidic	10^{-5}	10^{-9}		5	9
	10^{-6}	10^{-8}		6	8
Neutral	10^{-7}	10^{-7}	Pure water or solutions of Neutral substance	7	7
Slightly basic	10^{-8}	10^{-6}		8	6
	10^{-9}	10^{-5}		9	5
Medium alkalinity	10^{-10}	10^{-4}		10	4
	10^{-11}	10^{-3}		11	3
	10^{-12}	10^{-2}		12	2
Very basic	10^{-13}	10^{-1}	Add sufficient base to water	13	1
	10^{-14}	10^{0}		14	0
	10^{-15}	10^{+1}		15	-1

The following examples illustrate the calculation of pH, pOH and (H^+) ion concentrations.

Examples. (1) Calculate the pH of 0.01 M HCl solution?

Solution:

We know that, HCl is strong acid, so it is completely ionized in solution.

Then $pH = -\log [H^+]$

Express $[H^+]$ in exponential form

$0.01 \text{ M} \longrightarrow 1\times10^{-2}\text{ M}$

$pH = -\log [H^+]$

$pH = -\log [10^{-2}]$

$= -\log [-2]$

$pH = 2$ (Answer)

Example. (2) Calculate the pH of a solution whose (H^+)ion concentration is 5×10^{-4}M?

Solution:

pH = - log [H^+]

= - (log5 + log 10^{-4})

= - [.699 - 4]

= - .699+4

pH = 3.301 (Answer)

Example. (3) Calculate the pH and pOH of a solution whose (H^+) ion concentration is $3.0 x 10^{-2}$ moles/ litre?

Solution:

pH = - log [H^+]

= - log ($3x10^{-2}$)

= - (log 3 – log 10^{-2})

= - [0.477 - 2]

= [-0.477 +2]

pH = 1.523

As pH + pOH = 14

1.523+pOH= 14

pOH= 14 – 1.523

pOH= 12.477

pOH= 14 – 1.523

pOH= 12.477

$\boxed{\text{pH = 1.523 pOH = 12.477 (Answer)}}$

Example. (4) Calculate the (H^+) ion concentration of a solution whose pH 4.4?

Solution:

pH = - log [H^+]

pH = - log (H^+) = 4.4

pH = log (H^+) = -4.4

As we know that exponent should be whole number, so that –4.4 is equal to (-5+0.6), then we can write as:

log [H^+] = - 4.4

log [H^+] = (0.6-5)

H^+ = Antilog (0.6 –5)

H^+ = Antilog of 0.6 x Antilog of –5

= 3.98 x 10^{-5}

$\boxed{H^+ = 3.98 \times 10^{-5} \text{ M (Answer)}}$

9.5.3 The Measurement of pH with (pH Paper):

There are three methods to measure the pH of a solution. (i) By acid-base indicator's . (ii) By "pH-meters" and (iii) By "pH-paper". The method for the approximate determination of pH widely used is with the 'pH-paper'. In this method, paper-strips, that are treated with several different indicators can be used to estimate pH. These papers-strips are called "**pH-paper**". pH can be estimated by dipping the pH paper in a given solution, then by matching the colour appearing on the pH-paper with a colour corresponding to a known pH.

9.5.4 The Importance of pH:

The concept of pH plays essential role in the field of biology. Other areas in which pH information and control is necessary, include, water treatment, soil conditioning, swimming pool managements, corrosion, control, food processing and electroplating.

For example: The pH of human blood is normally maintained by the body between 7.35 and 7.45. If the blood pH drops to 7, as in some illness, the patient may go into coma, a pH below 6, death may occur. pH rises as high as 7.7 or 7.8 causes diabetes excess vomiting, diarrhea.

Hence the pH values of various body fluids are vey important for a doctor in diagnosing and treating many illnesses.

The pH values of several biological fluids.

Fluid	pH
Lemon juice	2.3
Vinegar	2.8
Tomato juice	4.2
Human urine	5.0-7.0
Cow's milk	6.5
Saliva	7.0
Human blood	7.35-7.45
Egg white	7.8

9.6 ACID-BASE TITRATION

There are two ways to make a solution of known molarity. The first and most convenient way to make the solution, by dissolving exactly 1 mole of solute in a litre (dm^3) of solution. The second way is to make up a solution quickly, using an estimated amount of solute and an estimated amount of solution, and then determine the solution's exact molarity by titration.

Titration is the chemical process by which we can determine the concentration of unknown solution, that reacts with a standard solution, whose concentration is known.

Titrations play a major role in determining, amounts of solutes present in a solutions. Titration process is an important tool of the analytical chemist.

9.6.1 Molarity (M) and Molar Solution:

The most generally useful means of expressing a concentration of solution is molarity.

It is defined as the number of moles of solute dissolved per 1 litre ($1dm^3$) of a solution, it is denoted by (M).

Thus 1 mole of H_2SO_4 (i.e. its gram formula mass) 98g dissolved in one litre ($1dm^3$) of solution is said to be 1 molar (1M) solution, if only one-half the mole i.e. 49g of H_2SO_4 are dissolved in one litre of solution, the solution is said to be one-half molar (i.e. 0.5 M)

The molarity of any solution is found by dividing the number of moles of solute by the number of moles/litres of the solution.

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Litres of solution}}$$

We known that,

(i) $$\text{Number of moles} = \frac{\text{Mass of solute}}{\text{Gram formula mass}}$$

(ii) $$\text{Litres of solution} = \frac{\text{Volume of solution in } cm^3\text{(ml)}}{1000}$$

$$\text{Molarity} = \text{(M)} = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{Volume of solution in } cm^3\text{(ml)}}$$

For example:

(i) Calculate the molarity of a solution, containing 1.5g of NaOH in 250 cm^3 of solution?

Solution:

$$M = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{Volume of solution in cm}^3}$$

Data:

(i)	Molarity	=	M = ?
(ii)	Mass of solute (NaOH)	=	1.5g
(iii)	Gram formula mass of solute (23+16+1)	=	40g
(iv)	Volume of solution	=	250 cm^3

$$M = \frac{1.5g \times 1000}{40 \times 250} = 0.15M$$

$\boxed{\text{Result: Molarity of NaOH solution is = M = 0.15M}}$

Example. 2

(i) What mass of(NaOH) must be dissolved in 500cm^3 of solution to make 1.5 M-solution?

Solution:

$$M = \frac{\text{Mass of solute}}{\text{Gram formula mass of solute}} \times \frac{1000}{\text{cm}^3 \text{ of solution}}$$

By cross-multiplication, we get,

$$\text{Mass of solute} = \frac{\text{M x G.F.mass x cm}^3 \text{of solution}}{1000}$$

Data:

(i)	Mass of solute (NaOH)	=	?
(ii)	Molarity = M	=	1.5 M
(iii)	G.F mass of solute (23+16+1)	=	40 g
(vi)	Cm^3 of solution	=	500 cm^3

$$\text{Mass of solute} = \frac{1.5g \times 40 \times 500cm^3}{1000\ cm^3} = 30g$$

$\boxed{\text{Result: Mass of solute i.e. NaOH = 30g}}$

9.6.2 Preparation of Solution of Known Molarity:

For example, how to prepare 1 M-solution of NaOH.

To prepare 1M solution of NaOH you first weigh out one mole of NaOH i.e. 23 + 16 + 1 = 40g of NaOH, place 40g of NaOH in a 1 litre (1 dm^3) volumetric flask, and add some water to bring the level of solution to calibration mark as shown in figure given below, finally the solution is shaken untill uniform.

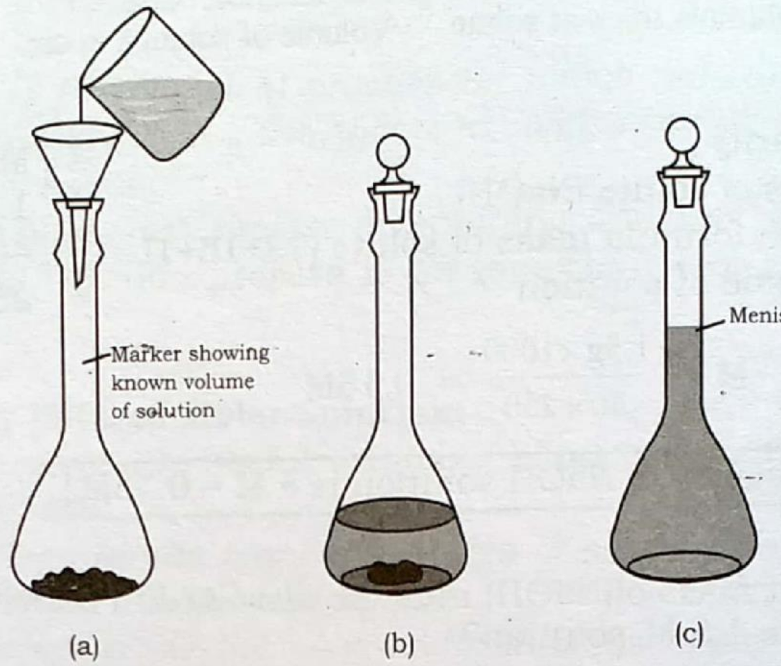


Fig. 9.2 Preparation of solution of known molarity

9.6.3 Standard Solution:

A solution whose molarity or strength is known, is called standard solution. For example 1M.KOH solution, contains 1mole of (KOH) i.e. formula mass expressed in grams =39+16+1 = 56g of KOH are dissolved in 1 litre of solution is said to be 1 molar (1M). If we use only one-half mole i.e. 28g of KOH in 1-litre of solution our solution would be one-half molar i.e. 0.5 M.

9.6.4 Acid-Base Titration:

In acid-base titration, a solution of known concentration (say base) is added gradually to a solution of unknown concentration (say an acid) so as to determine the concentration of unknown solution. The point at which the reaction is completed is called the end point.

Steps for Carrying out Titration:

1. Fill the burette with the given solution of NaOH; read and record the initial reading of the burette (Remember to read the lower meniscus) if the burette is not full refill it with NaOH solution. Then let the burette drain into beaker down to the zero mark before using it. The solution in the burette is called **"Titre"**.

2. Pipette out 10 cm^3 of HCl in a conical flask, (titration flask) and add one or two drops phenolphthalein indicator. The solution in titration flask is called **"Titrant"**. A suitable acid-base indicator, such a methyl orange, litmus, or phenolphthalein is used to detect the end point.

S.No.	Indicator	Colour in acids	pH-range	Colour in bases
1.	Methyl orange	Red	3 - 5	Yellow
2.	Litmus	Red	6 - 8	Blue
3.	Phenolphthalein	Colourless	8 -10	Pink(red)

3. Add slowly the NaOH solution, from the burette into the conical flask (titration flask) with constant shaking. Stop adding the NaOH solution, when the mixture in the titration flask becomes light pink. Record the final burette reading. If you have been adding the reactant rapidly and think you have over run the end point, repeat the entire titration slowing down when you think you are near the end point. Rinse out the flask with water before the next sample is titrated.

4. Record the initial and final burette readings for each experiment.

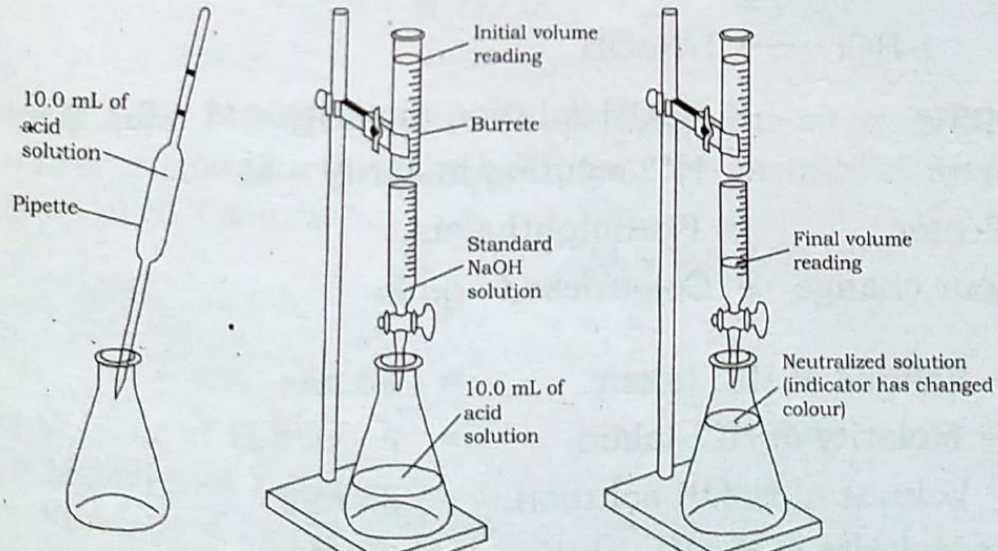


Fig. 9.3 Shows procedure for titrating an acid against a standardized solution of NaOH (base)

To see how titration works, let us consider that we have (HCl) solution (an acid) whose concentration we want to find by allowing it to react with a standard base such as NaOH.

We begin the titration by measuring out a known volume of HCl in a conical flask (titration flask)and adding few drops of an indicator, such as phenolphthalein (colourless liquid in acidic solutions but turns pink colour in basic solutions).

Next, we fill a burette with the(NaOH) standard solution (of known concentration) and we slowly add the NaOH to the HCl, until the phenolphthalein just begins to turn pink, indicating that all the HCl has been reacted and that the solution is starting to become basic. Then by recording the reading from the burette, to find the volume of the NaOH (standard solution) that has reacted with a known volume of HCl, we can calculate the concentration (molarity) of the HCl.

Let us assume for example that we take 10 mls of HCl-solution and find that we have to add 30 mls of 0.1 M NaOH from the burette to obtain complete reaction. Using the following formula we can calculate the molarity of unknown solution of HCl?

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

The balanced equation shows that:

1 mole of HCl reacts with 1 mole of NaOH

$$1\text{-}HCl \longrightarrow 1\text{-}NaOH$$

Burette = NaOH solution molarity = M = 0.1 M.
Pipette = HCl solution molarity = M = ?
Indicator = Phenolphthalein.
Colour change = Colourless to pink.

V_1 = Volume of HCl taken = 10 mls
M_1 = Molarity of HCl taken = ?
V_2 = Volume of NaOH solution = 30 mls
M_2 = Molarity of NaOH solution = 0.1M
n_1 = Number of moles of HCl = 1
n_2 = Number of moles of NaOH = 1

$$\frac{V_1 \times M_1}{n_1} = \frac{V_2 \times M_2}{n_2}$$

$$\frac{10 \times ?}{1} = \frac{30 \times 0.1}{1} = \frac{30 \times 0.1}{10} = \frac{3}{10}$$

M_1 = .3M

$\boxed{\text{Molarity of HCl} = 0.3\ \text{M}}$

We can also find out the amount of HCl dissolved per 1-litre = ($1dm^3$) of a solution by the following formula.

Amount of HCl per dm^3 = molarity x F. mass in grams

Amount of HCl per dm^3 = 0.3 x 36.5 = 10.95g

(As F. mass of HCl = 1+35.5 = 36.5g)

Result:	Molarity of HCl = 0.3M Amount of HCl/dm^3 = 10.95g

Example.

A flask contain 30 ml of NaOH solution, it require 50ml of 0.15M-H_2SO_4, to complete the reaction. Calculate the molarity of NaOH and how many grams of NaOH were in the flask?

Solution.

The balance equation for the reaction is:

$$\underset{2\ \text{mole}}{2NaOH(aq)} + \underset{1\text{mole}}{H_2SO_4(aq)} \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Since 2 mole of NaOH $\equiv$ 1mole of H_2SO_4 we need twice the amount of NaOH to react completely with H_2SO_4 solution as needed for the same concentration of HCl solution.

$$\frac{M_1 \times V_1}{n_1} = \frac{M_2 \times V_2}{n_2}$$

Where:

V_1 = Volume of H_2SO_4	=	50 ml
M_1 = Molarity of H_2SO_4	=	0.15M
n_1 = Number of moles of H_2SO_4	=	1
V_2 = Volume of NaOH	=	30 ml
M_2 = Molarity of NaOH	=	?
n_2 = Number of moles of NaOH	=	2

$$\therefore \quad \frac{M_1 \times V_1}{n_1} = \frac{M_2 \times V_2}{n_2} \qquad \frac{0.15 \times 50}{1} = \frac{M_2 \times 30}{2}$$

$$M_2 = \frac{0.15 \times 50 \times 2}{30} = 0.5\text{ M}$$

$$\text{Amount of NaOH per 30 ml} = \frac{M \times \text{G.F.mass} \times \text{volume of solution}}{1000}$$

$$= \frac{0.5 \times 40 \times 30}{1000} = 0.6\text{g}$$

Result:	Molarity of NaOH = 0.5M Amount of NaOH in the flask = 0.6g

SUMMARY

(1) Acids are substances that have the following properties in aqueous solution.
   - (i) They have a sour taste.
   - (ii) Change the colour of litmus from blue to red.
   - (iii) React with active metals (such as iron, tin and zinc) to liberate hydrogen gas.
   - (iv) React with bases to form salt (ionic compounds) and water.
   - (v) Their aqueous solutions conduct an electric current because they contain ions, they are electrolytes.

(2) Similarly bases are defined as substances that have the following properties, when dissolved in water.
   - (i) They have bitter taste.
   - (ii) Feel soapy and slippery on the skin.
   - (iii) Turn colour of litmus paper from red to blue.
   - (iv) React with acids to form salt and water.
   - (v) They give aqueous solutions which conduct an electricity, they are electrolytes.

(3) Arrhenius proposed that an acid is a substance that yields hydrogen (H^+) ions in water solution, while base produces hydroxide (OH^-) ions.

(4) The Lewis theory defines a base, as any substance that has one or more pair of electrons for bond formation. A Lewis base is an electron-pair donor, while lewis acid is electron-pair acceptor.

(5) Strong acid is one that ionizes in dilute solution almost completely and weak acid is one that ionizes only slightly. Thus a solution of strong acid has relatively high concentration of (H^+) ions and a solution of weak acid has a relatively low concentration of (H^+)ions. Similarly strong base is one that ionizes in dilute solution almost completely and weak base is one which ionizes only slightly. Thus a solution of strong base has a relatively high concentration of (OH^-) ions and a solution of weak base has a relatively low concentration of (OH^-) ions.

(6) Acids may be classified in terms of number of proton per molecule of acid that can given up in a reaction. Acids $HClO_4$, HI, HBr, HCl, HNO_3 and HCN that contain one ionizable hydrogen atom in a molecule of acid are called monoprotic acids. Di-protic acids contain two ionizable hydrogen atoms in one molecule of the acid, ionization of such acids occurs in two steps. Tri-protic acids, such as phosphoric acid, ionizes in three steps. Mono protic, Di-protic and tri-protic acids are commonly called mono basic, dibasic and tribasic. Similarly, bases that produces one (OH^-) ion per molecule are called mono-acidic bases, that produces two (OH^-) ions per molecule are called di-acidic bases and that produce three (OH^-) ions per molecule are called tri-acidic bases.

(7) When acids react with bases, salt and water are formed, and is known as neutralization. Many of these reactions can be denoted by a single net-ionic equation.

$$H^+ + OH^- \longrightarrow H_2O.$$

Salts are formed by any positive ion except hydrogen, combined with any negative ion, except hydroxide ion. Salts are ionic substances which are completely dissociated in water solution and are known as electrolytes.

(8) The acidity of an aqueous solution is expressed by its pH, which is defined as the negative logarthim of the hydrogen ions concentration. (in moles per litre). The neutral solution has pH=7, acidic solution has a pH is less then 7 and basic solution has pH is more then 7. Nowadays pH of a solution is determined by pH-meter.

(9) Most salts are strong electrolytes and dissociate completely into ions in aqueous solutions. The reaction of these ions, with water lead to the acidic or basic solution, called hydrolysis, the solution of strong acids and weak bases are acidic and the solutions of weak acids and strong bases are basic in nature, but the solutions of strong acids and strong bases are neutral because of the formation of equal concentration of strong acid and strong base.

(10) Water is amphoteric because it is both an acid and a base. It undergoes self ionization to give small quantities of hydrogen ion (H^+) and hydroxide ion (OH^-).

(11) Standard solution is that whose concentration is known.

(12) The concentration of solution is usually expressed as molarity (M), which is defined as the number of moles of solute dissolved per litre of solution. A solution of exact concentration can often be determined by titration. Titration is a chemical process by which we can determine the concentration of unknown (acid) solution by reacting it with a solution of known concentration of (base) solution.

EXERCISE

1. Fill in the blank.

i. The formula of baking soda is ....

ii. The formula of epsom salt is ....

iii. $K_2SO_4Al_2(SO_4)_3 . 24\,H_2O$ is the formula of ....

iv. .... is the most convenient way of expressing concentration.

v. The molarity of solution is denoted by ....

vi. A solution whose strength is known is called ....

vii. If H^+ ion concentration of a solution is 1×10^{-14} M, the solution is ....

viii. If the OH^- ion concentration of a solution is 1×10^{-10} M, solution is...........

ix. .... is the process by which we can determine the concentration of un-known solution with the help of standard solution.

x. The solution whose H^+ ion concentration is 1×10^{-4} M, then its pH is ....

xi. The solution whose pH is 6, then its H^+ ion concentration is ....

xii. The volume of a pipette is generally .... ml or cm^3.

xiii. Molarity is defined as number of moles per ....

2. Tick the Correct Answer:

(i) The substances whose aqueous solution change the blue litmus to red:

(a)	Acids.	(b)	Bases.
(c)	Neutral.	(d)	Salts.

(ii) The substances having a tendency to lose one or more protons are called:

(a)	Acids.	(b)	Bases.
(c)	Neutral.	(d)	Salts.

(iii) The substances which donate the pair of electrons for bond formation, are known as:

(a)	Acids.	(b)	Bases.
(c)	Neutral.	(d)	Salts.

(iv) When equivalent quantities of acid and base are mixed, salt and water are formed, the reaction is termed as:

(a)	Hydration.	(b)	Hydrolysis.
(c)	Neutralization.	(d)	None of these.

(v) The acids which contain one acidic hydrogen are called:

(a)	Mono-protic.	(b)	Di-protic.
(c)	Tri-protic.	(d)	Poly protic.

(vi) The number of acidic hydrogen atoms present in a molecule of an acid is called:

(a)	Acidity.	(b)	Basicity.
(c)	Neutral.	(d)	Hydrolysis.

(vii) The number of replaceable $[OH]^-$ ions present in a molecule of base, is called:

(a)	Acidity.	(b)	Basicity.
(c)	Neutral.	(d)	Hydrolysis.

(viii) An acid that produces large number of (H^+) ions in aqueous solution is called:

(a) Strong base. (b) Weak base.
(c) Strong acid. (d) Weak acid.

(ix) An ionic compound, that is formed when an acid neutralizes a base, is called:

(a) Acids. (b) Bases.
(c) Neutral. (d) Salts.

(x) Salts that formed by the reaction of strong acid with weak base are:

(a) Neutral. (b) Acidic.
(c) Basic. (d) Normal.

(xi) Salts that formed by the reaction of weak acid with strong base are:

(a) Acidic. b) Basic.
(c) Neutral. d) Normal.

(xii) Alums are:

(a) Single salts. b) Double salts
(c) Triple salts. d) Normal salts

xiii. The formula of Washing Soda is:

(a) Na_2CO_3. (b) $Na_2CO_3 . 6H_2O$.
(c) $Na_2CO_3 . 10H_2O$. (d) $NaHCO_3$.

3. **Write answer of the following questions:**

(i) What is Arrhenius theory of acids and bases? Why is the Arrhenius theory not satisfactory for acids and bases?

(ii) What is Lewis theory of acids and bases?

(iii) List the main general properties of acids and bases.

(iv) Write the formulas of four strong acids and four weak acids.

(v) Sulphuric acid (H_2SO_4) is strong acid, (HSO_4^-) is a weak acid? Account for the difference in strength?

(vi) Explain with illustrations what are strong acids and bases and weak acid and bases?

(vii) (a) Define acidic, basic and neutral solutions in terms of (H^+) ion concentration. Indicate whether each of the following solution will be acidic, basic or neutral.

(b) Strong acid and strong base.
(c) Strong acid and weak base.
(d) Weak acid and strong base.

(viii) What is salt? Give four examples of Salt?

(ix) Give an example each of mono-protic acid, di-protic acid and a tri-protic acid?

(x) Identify the following as a weak or strong acids or bases?
(a) NH_3 (b) H_3PO_4 (c) LiOH
(d) HCOOH (formic acid) (e) H_2SO_4
(f) H_2CO_3 (g) $Ba(OH)_2$

(xi) Define molarity? What is molar solution?

(xii). What is molarity of H_2SO_4 solution containing 9.8g of H_2SO_4 per 500 ml?

(xiii) Define pH? Explain why a neutral solution is said to have a pH of seven.

(xiv) Give an equation to show the dissociation of water.

(xv) Define the term "Amphoteric". Give an example?

(xvi) Calculate the pH of the following solutions.
(a) .001 M-HCl (b) 5.2×10^{-4} M-HNO_3

(xvii) Calculate (H^+) ion concentration of solutions, whose
(a) pH = 5.2 (b) pH = 9.63

(xviii) Fill in the word "acidic" "basic" or "neutral" for the following solutions.
(a) pH = 7 ....
(b) pH = > 7 ....
(c) pH = < 7 ....

(xix) The pOH of a solution is 9.40? Calculate the (H^+) ion concentration.

(xx) Define acid-base neutralization. Does the definition imply that the resulting solution is always neutral?

(xxi) Describe clearly how a solution of HCl could be titrated with a solution of NaOH?

(xxii) Define acid-base titrations, standard solution and equivalence point?

(xxiii) What volume of 0.5M KOH solution is needed to neutralize completely in each of the following.
(a) 10.0 ml of 0.3 M HCl-Solution.
(b) 10.0 ml of 0.2 M H_2SO_4-Solution.
(c) 10.0 ml of 0.25M H_3PO_4-Solution.