

3

Acids, Bases and Salts

SYLLABUS

Study of Acids, Bases and Salts

- (i) Simple definitions in terms of the molecules and their characteristic properties.
Self-explanatory.
- (ii) Ions present in mineral acids, alkalis and salts and their solutions: use of litmus and pH paper to test for acidity and alkalinity.
Examples with equation for the ionisation/dissociation of ions of acids, bases and salts: acids form hydronium ions (only positive ions) which turn blue litmus red, alkalis form hydroxyl ions (only negative ions) with water which turns red litmus blue. Salts are formed by partial or complete replacement of the hydrogen ion of an acid by a metal, should be explained with suitable examples. Introduction to pH scale to test for acidity, neutrality and alkalinity by using; pH paper or Universal indicator.
- (iii) Definition of salt; types of salts.
Types of salts: normal salts, acid salt, basic salt, definition and examples.
- (iv) General properties of salts:
 - *Deliquescence, Efflorescence, water of crystallization.*
Definition and example of each of the above.
 - *Decomposition of hydrogen carbonates, carbonates, chlorides and nitrates by appropriate acids with heating if necessary. (relevant laboratory work must be done).*
Action of dilute acids on carbonates, hydrogen carbonates and action of concentrated acid equation of formation of acid rain (Sulphuric acid) on chlorides and nitrates, to obtain carbon dioxide, hydrogen chloride and nitric acid respectively should be taught. This will assist the students in their practical work.
- (v) Preparation : laboratory preparation of salts (normal and acid salts) – relevant laboratory work is essential. (no apparatus details are required). *Laboratory preparation of salts (normal and acid salts) : Direct combination; decomposition; displacement; double decomposition; neutralization.*

3A. ACIDS AND BASES

3.1 INTRODUCTION

Lavoisier and Davy studied the terms acids, bases and salts on more scientific lines. Their findings say that

- all acids compulsorily have hydrogen as their constituent element in the form of proton (H^+).

Examples : Hydrochloric acid, sulphuric acid and nitric acid.

- all bases are either metallic oxide (O^{2-}) or metallic hydroxide (OH^-) or ammonia solution.

Examples : Copper oxide, sodium hydroxide and ammonium hydroxide.

- all salts are chemical compounds. They are formed by the interaction of acid and base replacing hydrogen ions present in an acid, by metallic ion [or non-metallic ammonium ions (NH_4^+)].

Examples : Sodium chloride, calcium carbonate and ammonium phosphate.

3.2 ACIDS

The word 'acid' comes from the Latin word *acidus*, meaning 'sour'. In fact, the sour taste of some foods is due to the acids they contain.

Some naturally occurring acids

Acetic acid	Vinegar
Citric acid	Citrus fruits (oranges and lemons)
Butyric acid	Rancid butter
Hydrochloric acid	Gastric juice
Formic acid	Sting of bees and ants
Lactic acid	Sour milk
Maleic acid	Apples
Oleic acid	Olive oil
Oxalic acid	Tomatoes
Stearic acid	Fats
Tartaric acid	Grapes, tamarind and apples
Uric acid	Urine

Acids are defined as **compounds which contain one or more hydrogen atoms and when dissolved in water, produce hydronium ions (H_3O^+) the only positively charged ions**. At first, the acid molecule furnishes hydrogen ion (or proton), i.e., H^+ ion in aqueous solution. But this H^+ ion cannot exist independently. Therefore, it combines with a water molecule to form hydronium ion (H_3O^+). This hydrated ion can now exist independently in solution.

Hydrogen is common in all acids.

For example : Hydrochloric acid, sulphuric acid, nitric acid and acetic acid when dissolved in water produce hydronium ions.



The hydrated hydrogen ion that exists in the solutions of acids is called a hydronium ion.



However, for simplicity, we shall use the simplified form, H^+ instead of H_3O^+ .

The acidic properties of an acid are actually the properties of hydronium ions present in it.

Note : In fact, the hydronium ion, H_3O^+ , can also be written as OH_3^+ , since the positive charge is due to the presence of extra hydrogen ion (H^+) with water, which is neutral.

3.3 CLASSIFICATION OF ACIDS

Acids can be classified in different ways.

1. Depending on their sources

(i) **Organic acids** : Acids which are obtained usually from plants are called organic acids. They contain carbon atom also along with hydrogen atom.

Examples : Oxalic acid (COOH)₂
Acetic acid (CH_3COOH).

They are weak acids and they do not ionise completely in solution. So they contain ions as well as molecules.

(ii) **Inorganic acids (mineral acids)** : Acids which are obtained usually from minerals are known as inorganic acids. They do not contain carbon (except carbonic acid H_2CO_3).

Examples : Hydrochloric acid (HCl),
Sulphuric acid (H_2SO_4),
Nitric acid (HNO_3).

They are strong acids. They ionise completely in solution producing a high concentration of hydronium ions [H_3O^+].

Carbonic acid (H_2CO_3) is a weak mineral acid. It turns blue litmus pink. It is non-corrosive, and so used in soft drinks.

Acids which contain oxygen along with hydrogen and some other element are **oxy-acids**.
Examples : Nitric acid (HNO_3), sulphuric acid (H_2SO_4).

Hydracids contain hydrogen and a non-metallic element, and no oxygen. Examples : Hydrochloric acid (HCl), hydrobromic acid (HBr).

2. Depending on their concentration

Concentration of an acid means the amount of acid present in a definite amount of its aqueous solution.

Concentrated acid : An acid which contains a very small amount of water or no water is called concentrated acid.

Dilute acid : An acid which contains far more

amount of water than its own mass is known as dilute acid.

In order to dilute an acid, pour acid into water in small amounts and stir constantly. Water is not added to acid as it is an exothermic process and in this process so much heat is produced that splashing of acidic solution may occur, also the container may break which can be fatal to the person.

Note : Concentration of an acid simply tells us the amount of water present in the acid and not at all the strength of the acid. Strength of an acid is the measure of concentration of hydronium ions it produces in its aqueous solution. Thus **dil. HCl is stronger acid than highly concentrated acetic acid.**

Note : The strength of an acid depends on the degree of ionisation $[\alpha]$ and concentration of hydronium ions $[\text{H}_3\text{O}^+]$ produced by that acid in aqueous solution.

1. Degree of ionisation (α)

$$= \frac{\text{No. of acid molecules ionised}}{\text{(Total no. of acid molecules present in aqueous solution)}} \times 100$$

2. If the degree of ionisation (α) for an acid, base or salt in aqueous solution is greater than 30%, it is strong and if it is less than 30%, it is weak.

3. Depending on their basicity

The basicity of an acid is defined as *the number of hydronium ions (H_3O^+) that can be produced by the ionisation of one molecule of that acid in aqueous solution.*

Therefore we have,

Examples : Hydrochloric acid (HCl),

(i) **Monobasic acids :** Acids which on ionisation in water produce one hydronium ion per molecule of the acid are known as monobasic acids.

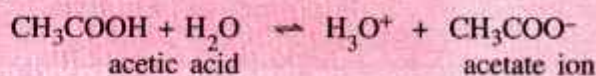
For example : Hydrochloric acid



Other similar examples are : HBr, HNO_3 , HI, CH_3COOH , etc.

It must be emphasized here that the basicity of an acid depends not on the number of hydrogen atoms in one molecule of that acid, but on the number of ionisable hydrogen atoms that it has per molecule.

For example, acetic acid (CH_3COOH or $\text{C}_2\text{H}_4\text{O}_2$) contains four hydrogen atoms in its molecule, but it is a monobasic acid because its molecule ionises by liberating only one hydronium ion.



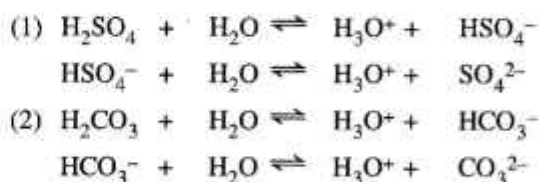
Monobasic acids ionise in one step and so form only one normal salt.



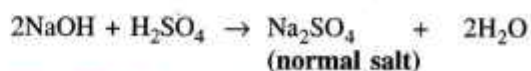
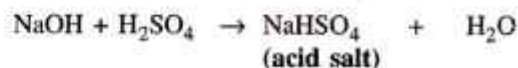
(ii) **Dibasic acids :** Acids which, on ionisation in water, produce two hydronium ions (H_3O^+) per molecule of the acid, are known as dibasic acids.

Examples : Sulphuric acid (H_2SO_4)
Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$)
Sulphurous acid (H_2SO_3)
Carbonic acid (H_2CO_3)

Dibasic acids ionise in *two steps*, as shown below:



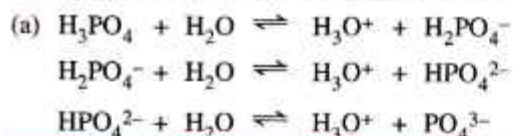
Dibasic acids have two replaceable hydrogen ions, therefore they form *one acid salt or one normal salt.*



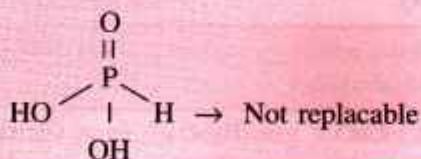
(iii) **Tribasic acids :** Acids which, on ionisation in water, produce three hydronium ions per molecule of the acid, are called tribasic acids.

Examples : Phosphoric acid (H_3PO_4)
Citric acid ($\text{C}_6\text{H}_8\text{O}_7$)

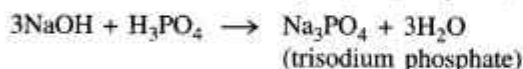
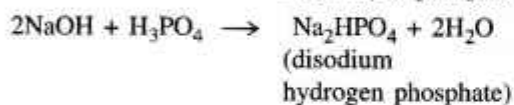
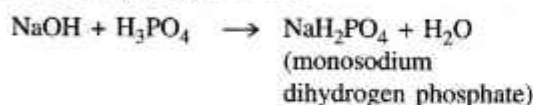
They ionise in **three steps**, as shown below :



H_3PO_3 is a **dibasic acid** because in oxyacids of phosphorus, hydrogen atoms which are attached to oxygen atoms are replaceable. Hydrogen atoms directly bonded to phosphorus atoms are not replaceable.



Tribasic acids can form **two different acid salts** and **one normal salt**, because they have three replaceable hydrogen ions.



Silicic acid H_4SiO_4 has basicity four.

3.4 PREPARATION OF ACIDS

(1) By synthesis

Binary acids (acids containing two elements) are prepared by this method.

Hydrogen		Non-metal	→	Acid
H_2	+	Cl_2	→	2HCl
H_2	+	Br_2	→	2HBr
H_2	+	S	→	H_2S

(2) By the action of water on non-metallic or acidic oxides

Acidic oxides (also called acidic anhydrides) dissolve in water to give acids.

Acidic oxide	+ Water	→	Acid
SO_3	+ H_2O	→	H_2SO_4 (Sulphuric acid)
SO_2	+ H_2O	→	H_2SO_3 (Sulphurous acid)
CO_2	+ H_2O	→	H_2CO_3 (Carbonic acid)
P_2O_5	+ $3\text{H}_2\text{O}$	→	$2\text{H}_3\text{PO}_4$ (Phosphoric acid)
N_2O_5	+ H_2O	→	2HNO_3 (Nitric acid)

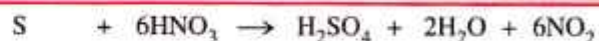
Note : (i) Only oxy-acids can be prepared by this method.

(ii) NO_2 is called mixed or double acid anhydride because two acids, nitrous acid and nitric acid, are formed when it reacts with water.



(3) By the oxidation of non-metals

A non-metal, e.g., sulphur or phosphorus is oxidised by conc. nitric acid to form sulphuric acid or phosphoric acid respectively.



(4) By displacement

Normal salts of more volatile acids are displaced by a less or non-volatile acid.

Examples : Both hydrochloric and nitric acids are formed by sulphuric acid.



3.5 PROPERTIES OF ACIDS

Physical properties

(1) **Taste :** Acids have a sour taste.

Mineral acids like H_2SO_4 and HNO_3 are highly corrosive in nature. Therefore, they should not be tasted.

(2) **Physical state :** Some acids are solids and some are liquids at room temperature.

For example :

Boric acid	H_3BO_3	Solid Acids
Oxalic acid	$(\text{COOH})_2$	
Tartaric acid	$\text{C}_4\text{H}_6\text{O}_6$	
Citric acid	$\text{C}_6\text{H}_8\text{O}_7$	
Phosphoric acid	H_3PO_4	
Acetic acid	CH_3COOH	Volatile acids vaporise easily, i.e., at room temperature or on heating at about 100°C .
Formic acid	HCOOH	
Carbonic acid	H_2CO_3	
Hydrochloric acid	HCl	
Nitric acid	HNO_3	
Sulphurous acid	H_2SO_3	
Sulphuric acid	H_2SO_4 (conc.)	Non-volatile acids

LIQUID ACIDS

(3) **Effect on skin** : All strong mineral acids have corrosive action on the skin and cause painful burns. Conc. H_2SO_4 stains the skin black, conc. HNO_3 makes it yellow and conc. HCl amber colour. Carbonic acid and organic acids are not corrosive.

(4) **They change the colours of indicators.** Indicators are complex substances that acquire separate colours in acidic and basic mediums. For example, litmus indicator is a natural colouring matter, obtained from certain lichens on dissolving in distilled water.

The red cabbage extract (natural indicator) remains red in acidic solutions but turns green in basic solutions.

Those substances whose smell (or odour) changes in acidic or basic solutions are called **OLFACTORY INDICATORS**. For example, onion, vanilla and clove oil. Onion has a characteristic smell. When a basic solution like sodium hydroxide solution is added to a cloth strip treated with onion (or onion extract), then the onion smell cannot be detected. An acidic solution does not destroy the smell of the onion.

Indicator	Colour change in acidic medium
Litmus	Blue to red
Methyl orange	Orange to pink
Phenolphthalein	Remains colourless

(5) They are electrolytes, *i.e.*, they conduct electricity in the aqueous state.

Chemical properties

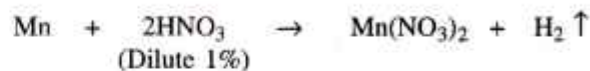
(1) Reaction with active metals

Both dilute HCl and dilute H_2SO_4 react explosively with active metals like potassium, sodium and calcium, and moderately with less active metals like magnesium, zinc, *etc.*, producing metallic salts and hydrogen.

Active metal	+ Acid	→	Salt	+ Hydrogen
Mg	+ 2HCl	→	MgCl_2	+ $\text{H}_2 \uparrow$
Zn	+ 2HCl	→	ZnCl_2	+ $\text{H}_2 \uparrow$
Fe	+ H_2SO_4	→	FeSO_4	+ $\text{H}_2 \uparrow$

Note : Nitric acid is a very strong oxidising agent so it is not used in the preparation of

hydrogen. It oxidises hydrogen and form water, thus defeats the purpose of the reaction. Only magnesium and manganese can produce hydrogen with very dilute nitric acid (1% acid).



(2) Reaction with bases – Neutralisation

Acids neutralise bases to form salt and water only.

Basic oxide/ hydroxide	+ Acid	→	Salt	+ Water
CuO	+ H_2SO_4	→	CuSO_4	+ H_2O
NaOH	+ HNO_3	→	NaNO_3	+ H_2O

(3) Reaction with carbonates and bicarbonates (hydrogen carbonates)

Acids liberate carbon dioxide on reaction with metallic carbonates and bicarbonates.

Carbonate/ bicarbonate	+ Acid	→	Salt	+ Water	+ Carbon dioxide
CaCO_3 (carbonate)	+ 2HCl	→	CaCl_2	+ H_2O	+ $\text{CO}_2 \uparrow$
$\text{Ca}(\text{HCO}_3)_2$ (bicarbonate)	+ 2HCl	→	CaCl_2	+ $2\text{H}_2\text{O}$	+ $2\text{CO}_2 \uparrow$

Note : (i) If the salt produced is insoluble, then the reaction does not proceed. So, we do not expect lead carbonate to react with hydrochloric acid or sulphuric acid. Nor does calcium carbonate react with sulphuric acid.

(ii) The hydrogen carbonates of sodium and potassium are the only common ones, which exist in the solid state.

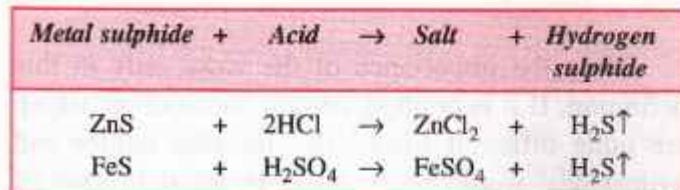
(4) Reaction with sulphites and bisulphites

Acids react with sulphites and bisulphites of metals to liberate sulphur dioxide.

Sulphite/ bisulphite	+ Acid	→	Salt	+ Water	+ Sulphur dioxide
CaSO_3 (sulphite)	+ 2HCl	→	CaCl_2	+ H_2O	+ $\text{SO}_2 \uparrow$
NaHSO_3 (bisulphite)	+ HCl	→	NaCl	+ H_2O	+ $\text{SO}_2 \uparrow$

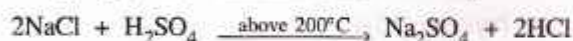
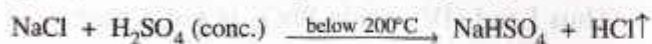
(5) Reaction with sulphides

Acids react with metal sulphides to liberate hydrogen sulphide.

**(6) Reaction with chlorides**

Chlorides do not react with any of the dilute acids. In general, chlorides react with concentrated sulphuric acid on warming to liberate hydrogen chloride.

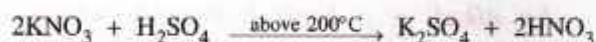
The reaction takes place in two steps since H₂SO₄ contains two replaceable hydrogen ions.

**(7) Reaction with nitrates**

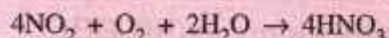
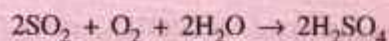
Nitrates do not react with dilute acids. However, lead nitrate solution reacts with both dilute HCl and dil. H₂SO₄ to form insoluble lead salts.



Nitrates when heated with *conc. sulphuric acid*, produce more volatile nitric acid.

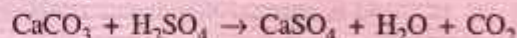
**ACID RAIN**

You have studied in class IX (Atmospheric Pollution) that acid rain is a by-product of a variety of human activities that releases oxides of sulphur and nitrogen in the atmosphere. Burning of fossil fuels, coal, oil, petrol and diesel produce sulphur dioxide and nitrogen oxide. They all pollute air. Polluted air also contain many oxidising agents which produces oxygen due to excessive heat. This oxygen further combines with the oxides of sulphur and nitrogen. When it rains, oxides of sulphur and nitrogen react with rain water and form acids.



Sulphuric acid and nitric acid formed are washed down to the earth with water. This is **Acid rain**. The highly acidic rain water has harmful effects.

- Acid rain damages the leaves of plants and trees.
- It removes the basic nutrients such as calcium from the soil.
- It causes respiratory ailments in human beings and animals.
- It affects plant and animal life in aquatic ecosystem.
- It corrodes water pipes resulting in leaching of heavy metals, such as iron, lead and copper, into the drinking water.
- Acid rain damages buildings as it reacts with stone or metals.

**3.6 GENERAL USES OF SOME ACIDS**

<i>Acid</i>	<i>Use</i>
1. Boric acid	Eye-wash/antiseptic.
2. Citric acid	Food preservation. Vitamin C preparation.
3. Oxalic acid	Ink stain remover.
4. Carbonic acid	Flavoured drinks.
5. Tartaric acid	Baking powder.
6. Acetic acid	Table vinegar and cooking.
7. Hydrochloric acid	Cleaning of metal items
8. Benzoic acid	Preservation of food and making of perfumes and medicines.
9. Nitric acid	Explosives.
10. Phosphoric acid	Fertilizers.

Intext Questions

- (a) What do you understand by the term, acid ?
(b) Name the positive ion formed when an acid is dissolved in water.
(c) Draw the structure of this ion.
- Write the ionisation of sulphuric acid showing the formation of hydronium ion.
- Water is never added to acid in order to dilute it. Why ?
- Define the term 'basicity' of an acid. Give the basicity of : nitric acid, sulphuric acid and phosphoric acid.
- Give two examples of each of the following :
(a) oxy-acid, (b) hydracid,
(c) monobasic acid, (d) dibasic acid,
(e) tribasic acid

6. Name the :
- (a) acidic anhydride of the following acids :
- (i) sulphurous acid, (ii) nitric acid,
(iii) phosphoric acid, (iv) carbonic acid.
- (b) acids present in vinegar, grapes and lemon.
7. What do you understand by the statement 'acetic acid is a monobasic acid'?
8. Give a balanced equation for reaction of nitrogen dioxide with water.
9. What do you understand by the strength of an acid ? On which factor does the strength of an acid depend ?
10. Explain the following :
- (a) Carbonic acid gives an acid salt but hydrochloric acid does not.
- (b) Dil. HCl acid is stronger than highly concentrated acetic acid.
- (c) H_3PO_3 is not a tribasic acid.
- (d) Lead carbonate does not react with dilute HCl.
- (e) Nitrogen dioxide is a double acid anhydride.
11. What is acid rain ? How is it formed ?
12. How is an acid prepared from a
(a) Non-metal (b) salt ? Give an equation for each.
13. Give equations to show how the following are made from their corresponding anhydrides.
(a) sulphurous acid, (b) phosphoric acid,
(c) carbonic acid, (d) sulphuric acid.
14. Name an acid used :
(a) to flavour and preserve food,
(b) in a drink,
(c) to remove ink spots,
(d) as an eyewash.

3.7 BASES

A base is either a metallic oxide or a metallic hydroxide or ammonium hydroxide which reacts with hydronium ions of an acid to form salt and water only.

For example, CuO (metallic oxide), $Mg(OH)_2$ (metallic hydroxide), etc.

Base	+	Acid	→	Salt	+	Water
CuO	+	2HCl	→	CuCl ₂	+	H ₂ O
Mg(OH) ₂	+	H ₂ SO ₄	→	MgSO ₄	+	2H ₂ O
Al(OH) ₃	+	3HNO ₃	→	Al(NO ₃) ₃	+	3H ₂ O
NH ₄ OH	+	HCl	→	NH ₄ Cl	+	H ₂ O

A basic oxide is a metallic oxide which contains the ion O^{2-} and reacts with an acid to form salt and water only.

Note the importance of the word *only* in this definition. If it is omitted, certain compounds, which are quite different from basic metallic oxides and hydroxides, would be included in the definition of the base.

For example : Lead (IV) oxide is a metallic oxide which reacts with hydrochloric acid to produce lead (II) chloride (a salt) and water, but the word *only* excludes it from the class of bases, because chlorine is also produced.



Thus Lead (IV) oxide PbO_2 is not a base.

Basic hydroxide

It is a metallic hydroxide which contains OH^- and will react with an acid to give salt and water only.

For example : Sodium hydroxide (NaOH) and Aluminium hydroxide $Al(OH)_3$.



3.8 ALKALIS

An alkali is a basic hydroxide which when dissolved in water produces hydroxyl (OH^-) ions as the only negatively charged ions.



An alkali is a base soluble in water.

All alkalis furnish OH^- ions in aqueous solution

All metallic oxides and hydroxide are insoluble except of sodium, potassium, calcium (slightly soluble). Thus

The most common soluble bases *i.e.* alkalis are :

*Sodium hydroxide (NaOH, caustic soda),
Potassium hydroxide (KOH, caustic potash),
Calcium hydroxide $[Ca(OH)_2]$, slaked lime] and
Ammonium hydroxide (NH_4OH).*

The first two are *strong alkalis* because they are completely ionised in their aqueous solutions and the last two are *weak alkalis*, because they are feebly ionised in their aqueous solutions.

All alkalis are bases but all bases are not alkalis.

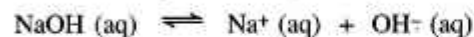
For example : Ferric hydroxide $[\text{Fe}(\text{OH})_3]$ and Cupric hydroxide $[\text{Cu}(\text{OH})_2]$ are bases, but not alkalis because they are **insoluble** in water.

3.9 CLASSIFICATION OF BASES

3.9.1 On the basis of their strength

(a) **A strong base** : It undergoes almost complete dissociation in aqueous solution to produce a high concentration of OH^- .

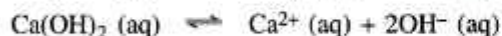
Examples : Sodium hydroxide (NaOH) and Potassium hydroxide (KOH).



(b) **A weak base** : It undergoes only partial dissociation or ionisation in aqueous solution to produce a low concentration of OH^- in solution.

Examples : Ammonium hydroxide (NH_4OH) and

Calcium hydroxide ($\text{Ca}(\text{OH})_2$).



2. On the basis of their acidity

The acidity of a base

The number of hydroxyl ions $[\text{OH}]^-$ which can be produced per molecule of the base in aqueous solution or the number of hydrogen ions (of an acid) with which a molecule of that base will react to produce salt and water only is known as acidity of the base.

(a) **Monoacidic base** : It is a base that dissociates in molten (fused) state or in aqueous solutions to produce one OH^- ion per molecule of that base.

For example : NaOH, KOH and NH_4OH .



(b) **Diacidic base** : It is a base that dissociates in molten state or in aqueous solutions to produce two OH^- ions per molecule of that base.

For example : $\text{Ca}(\text{OH})_2$ and $\text{Cu}(\text{OH})_2$.



For water-insoluble hydroxides, acidity is equal to the number of replaceable OH^- ions in one molecule of that base.

(c) **Triacidic base** : It is a base that dissociates in aqueous solutions to produce three OH^- ions per molecule of that base (or contains three replaceable hydroxyl ions per molecule of the base).

For example : $\text{Al}(\text{OH})_3$ and $\text{Fe}(\text{OH})_3$.

3.9.2 Preparation of bases

The common methods of preparation are :

(1) **From metals** : Metals when react with oxygen give bases.

Metal	+	Oxygen	→	Base
4Na	+	O_2	→	$2\text{Na}_2\text{O}$
2Mg	+	O_2	→	2MgO

(2) **By the action of water on reactive metals** like sodium, potassium and calcium.

Reactive metal	+	Water	→	Base/alkali	+	Hydrogen
2Na	+	$2\text{H}_2\text{O}$	→	2NaOH	+	$\text{H}_2\uparrow$
2K	+	$2\text{H}_2\text{O}$	→	2KOH	+	$\text{H}_2\uparrow$
Ca	+	$2\text{H}_2\text{O}$	→	$\text{Ca}(\text{OH})_2$	+	$\text{H}_2\uparrow$

(3) **By the action of water on soluble metallic oxides** like oxides of sodium, potassium and calcium .

Metal oxide	+	Water	→	Base (Alkali)
Na_2O	+	H_2O	→	2NaOH
K_2O	+	H_2O	→	2KOH
CaO	+	H_2O	→	$\text{Ca}(\text{OH})_2$

(4) **By double decomposition** : Aqueous solutions of salts with a base (alkali) precipitate the respective metallic hydroxide.

Salt solution	+ Base (Alkali)	→	Basic hydroxide	+ Normal salt
FeCl ₃	+ 3NaOH	→	Fe(OH) ₃ ↓	+ 3NaCl
CuSO ₄	+ 2NaOH	→	Cu(OH) ₂ ↓	+ Na ₂ SO ₄
AlCl ₃	+ 3NH ₄ OH	→	Al(OH) ₃ ↓	+ 3NH ₄ Cl

(5) **By the action of oxygen on metal sulphides :**

Metallic sulphide	+ Oxygen	$\xrightarrow{\Delta}$	Metallic oxide	+ Sulphur dioxide
2ZnS	+ 3O ₂	→	2ZnO	+ 2SO ₂
2PbS	+ 3O ₂	→	2PbO	+ 2SO ₂

(6) **By decomposition of salts :**

Metal carbonate	$\xrightarrow{\Delta}$	Basic oxide	+ Carbon dioxide
CaCO ₃	→	CaO	+ CO ₂ ↑
CuCO ₃	→	CuO	+ CO ₂ ↑

Note : Sodium and potassium carbonate do not decompose on heating.

Metal nitrate	$\xrightarrow{\Delta}$	Basic oxide	+ Nitrogen dioxide	+ Oxygen
2Ca(NO ₃) ₂	→	2CaO	+ 4NO ₂	+ O ₂
2Zn(NO ₃) ₂	→	2ZnO	+ 4NO ₂	+ O ₂

Note : Sodium and potassium nitrate do not give metal oxide when decomposed.

(7) **Ammonium hydroxide** can be prepared by simply dissolving ammonia gas in water. Ammonia is extremely soluble in water.



3.9.3 Properties of bases/alkalis

Physical properties

- (1) They have a *sharp and bitter taste*.
- (2) They *change the colour of indicators*.

Indicator	Colour change
Litmus	Red to blue
Methyl orange	Orange to yellow
Phenolphthalein	Colourless to pink
Insoluble base do not affect indicators.	

- (3) They are **soapy substances**, i.e., they are slippery to touch.

Caustic soda and caustic potash are strong alkalis. We know that alkalis react with oil to form soap. Since our skin contains oil (in the form of fat), hence when we touch caustic soda or caustic potash, a reaction takes place and soapy solutions are formed.

- (4) They are strong electrolytes.
- (5) They show a mild corrosive action (slight burn) on the skin.

Chemical properties

- (1) Strong alkalis absorb carbon dioxide from the air to form carbonates.

Strong alkali	+ Carbon dioxide	→	Carbonate	+ Water
2NaOH	+ CO ₂	→	Na ₂ CO ₃	+ H ₂ O
2KOH	+ CO ₂	→	K ₂ CO ₃	+ H ₂ O

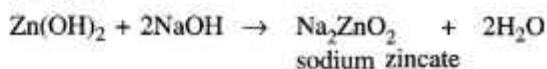
- (2) They neutralise acids to form salt and water.

Base/Alkali	+ Acid	→	Salt	+ Water
Ca(OH) ₂	+ 2HCl	→	CaCl ₂	+ 2H ₂ O
Fe(OH) ₂	+ 2HCl	→	FeCl ₂	+ 2H ₂ O

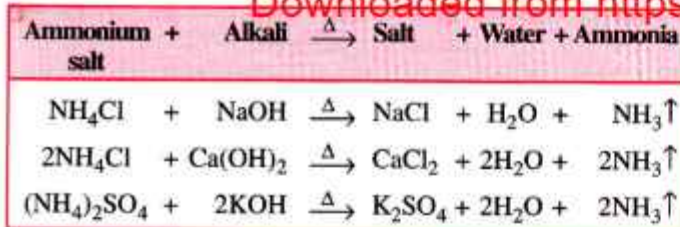
- (3) They precipitate as insoluble metallic hydroxides when added to the solutions of the salts of the heavy metals, viz. copper, iron, zinc, etc.

Metallic salt	+ Base/alkali	→	Salt	+ Insoluble hydroxide
CuSO ₄ (aq.)	+ 2NH ₄ OH (aq.)	→	(NH ₄) ₂ SO ₄ (aq.)	+ Cu(OH) ₂ ↓ (pale blue)
ZnSO ₄ (aq.)	+ 2 NaOH (aq.)	→	Na ₂ SO ₄ (aq.)	+ Zn(OH) ₂ ↓ (gelatinous white)
FeSO ₄ (aq.)	+ 2NaOH (aq.)	→	Na ₂ SO ₄ (aq.)	+ Fe(OH) ₂ ↓ (dirty green)

Note : The hydroxides of zinc, aluminium and lead, being amphoteric by nature, dissolve in excess of NaOH or KOH, but other hydroxides do not.



- (4) When alkalies are warmed with an ammonium salt, ammonia gas is given out.



3.9.4 General uses of some bases

Base	Use
1. Sodium hydroxide	Manufacture of soap.
2. Potassium hydroxide	Manufacture of salts and soaps and used in batteries.
3. Calcium hydroxide (slaked lime)	Manufacture of bleaching powder and softening of hard water.
4. Magnesium hydroxide	As an antacid [neutralizes acidity].
5. Aluminium hydroxide	As a foaming agent in fire extinguishers.
6. Ammonium hydroxide	To remove grease stains from clothes.
7. Magnesia	In making refractory bricks.
8. Calcium oxide	Drying agent, basic flux.

3.10 STRENGTH OF ACIDIC AND BASIC SOLUTIONS

The acidic, basic or neutral solution can be ascertained on the basis of ionisation. Water ionises to a very small extent. The concentration of $[\text{H}_3\text{O}^+]$ (aq.) and OH^- (aq.) ions in pure water is found to be 1×10^{-7} mol litre⁻¹. This means that all aqueous solutions contain both $[\text{H}_3\text{O}^+]$ (aq.) and OH^- (aq.) ions and the product of concentration of these ions is $(1 \times 10^{-7})(1 \times 10^{-7}) = 1 \times 10^{-14}$ mol² litre⁻² at 25°C. This is known as **ionic product of water (K_w)**.

On adding acid to water the concentration of H_3O^+ ion increases and that of OH^- decreases so that K_w remains the same and if an alkali is added the reverse is noticed. So the **nature of solution depends upon the concentration of H_3O^+ ion in solution.**

A Danish biochemist (1909) devised a scale known as pH scale which represents the $[\text{H}_3\text{O}^+]$ ion concentration of the given aqueous solution.

The 'p' in pH stands for 'Potenz' in German meaning *power*, so pH means the power of hydrogen.

The pH of a solution is the negative logarithm to the base 10 of the hydrogen ion concentration expressed in moles per litre.

$$\text{pH} = \log_{10} \frac{1}{[\text{H}^+]} = -\log_{10}[\text{H}^+]$$

A solution of H^+ ions of concentration equal to 10^{-12} moles/L will have $\text{pH} = 12$.

$$\text{i.e. } \text{pH} = \log_{10} \frac{1}{[\text{H}^+]} = \log_{10} \frac{1}{10^{-12}}$$

$$\log_{10} 10^{12} = 12 \log_{10} 10 = 12. \text{ (Since } \log_{10} 10 = 1)$$

The normal pH scale ranges from 0 to 14, as given in Fig. 3.1.

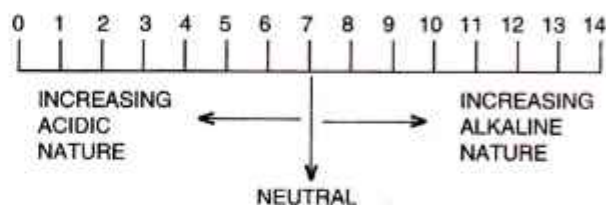


Fig. 3.1 The pH scale

The number 7 on the pH scale represents the neutrality of the solution. Numbers less than 7, i.e. pH 6, 5, 4, indicate that *acidic nature is increasing* as the numbers *decrease*. On the other hand, numbers greater than 7, i.e., pH 8, 9, 10, show that *alkalinity increases* as the numbers *increase*.

In a colourless liquid, a reasonably accurate pH value can be obtained by adding a *universal indicator*.

The **universal indicator** is a mixture of indicator dyes that gives a spectrum of colours depending on how acidic or alkaline a solution is. Universal indicators gives different colours at different concentrations of hydrogen ions in a solution.

For example, a universal indicator produces green colour in a neutral solution, i.e., when $\text{pH} = 7$. It changes in a basic solution progressively from blue to indigo to violet as pH increases progressively from 7 to 14. The colour change in acidic solution is from yellow to pink, and then to red, as pH progressively decreases from 7 to 1 (Fig. 3.2).

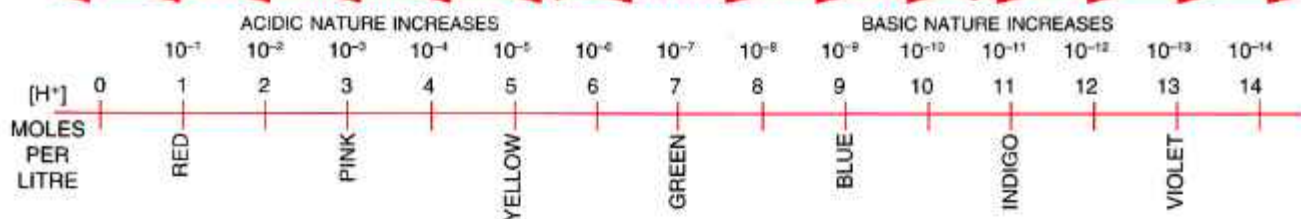


Fig. 3.2 Colour changes in the universal indicator (for different pH values).

Importance of pH in everyday life

1. Our body works within a narrow pH range of 7.0 to 7.8. Plants and animals also survive in a narrow range.
2. **In agriculture** : Every crop grows better in a particular pH range. *For example*, rice grows better in slightly acidic soil, sugarcane in neutral and citrus fruits in alkaline soil.
3. **In medicines** : Certain diseases are diagnosed only on the basis of the pH value of blood and urine.
4. **In digestive system** : Hydrochloric acid is produced in the stomach which helps in the digestion of food but if it becomes excess, the pH falls, and pain and irritation occurs. To get rid of this ANTACIDS like milk of magnesia ($Mg(OH)_2$) is generally used to adjust the pH.
5. In saving tooth decay : Substances like chocolates and sweets are degraded by bacteria present in our mouth. When the pH falls to 5.5 tooth decay starts. Tooth enamel (calcium phosphate) is the hardest substance in our body and it gets corroded. The saliva produced by salivary glands is slightly alkaline, it helps to increase the pH, to some extent, but tooth paste is used to neutralise excess acid in the mouth.
6. Bee sting leaves acid in the body. If baking soda, a base, is applied on the stung area, it gives relief.

Intext Questions

1. What do you understand by an alkali ?
Give two examples of :
(a) strong alkalis, (b) weak alkalis.
2. What is the difference between :
(a) an alkali and a base,
(b) the chemical nature of an aqueous solution of HCl and an aqueous solution of NH_3 .
3. Name the ions furnished by :
(a) bases in solution, (b) an acid.
4. Give one example in each case :
(a) A basic oxide which is soluble in water,
(b) A hydroxide which is highly soluble in water,
(c) A basic oxide which is insoluble in water,
(d) A hydroxide which is insoluble in water,
(e) A weak mineral acid,
- (f) A base which is not an alkali,
- (g) An oxide which is a base,
- (h) A hydrogen containing compound which is not an acid,
- (i) A base which does not contain a metal ion.
5. You have been provided with three test tubes. One of them contains distilled water and the other two have an acidic solution and a basic solution respectively. If you are given only red litmus paper, how will you identify the contents of each test tube?
6. HCl, HNO_3 , C_2H_5OH , $C_6H_{12}O_6$ all contain H atoms but only HCl and HNO_3 show acidic character. Why ?
7. (a) Dry HCl gas does not change the colour of dry litmus paper. Why ?
(b) Is PbO_2 a base or not ? Comment.
(c) Do basic solutions also have $H^+(aq)$? Explain why are they basic by taking an example?

8. How would you obtain :
- a base from other base,
 - an alkali from a base,
 - salt from another salt ?
9. Write balanced equations to satisfy each statement.
- Acid + Active metal \rightarrow Salt + Hydrogen
 - Acid + Base \rightarrow Salt + Water
 - Acid + Carbonate \rightarrow Salt + Water + Carbon dioxide
or bicarbonate
 - Acid + Sulphite \rightarrow Salt + Water + Sulphur dioxide
or bisulphite
 - Acid + Sulphide \rightarrow Salt + Hydrogen Sulphide
10. The skin *has and needs* natural oils. Why is it advisable to wear gloves while working with strong alkalis ?
11. Complete the table :

Indicator	Neutral	Acidic	Alkaline
Litmus	Purple
Phenolphthalein	Colourless

12. What do you understand by pH value ? Two solutions X and Y have pH values of 4 and 10 respectively. Which one of these two will give a pink colour with phenolphthalein indicator ?
13. You are supplied with five solutions : A, B, C, D and E with pH values as follows:
A = 1.8, B = 7, C = 8.5, D = 13, and E = 5
Classify these solutions as *neutral*, *slightly* or *strongly acidic* and *slightly* or *strongly alkaline*.

Which solution would be most likely to liberate hydrogen with :

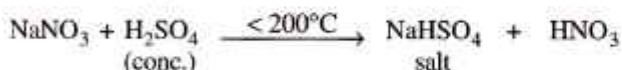
- magnesium powder,
 - powdered zinc metal. Give a word equation for each reaction.
14. Distinguish between :
- a common acid base indicator and a universal indicator,
 - acidity of bases and basicity of acids,
 - acid and alkali (other than indicators).
15. How does tooth enamel get damaged ? What should be done to prevent it ?
16. When you use universal indicator, you see that solutions of different acids produce different colours. Indeed, solutions of the same acid with different concentrations will also give different colours. Why ?
17. (a) A solution has a pH of 7. Explain how you would :
(i) increase its pH; (ii) decrease its pH;
(b) If a solution changes the colour of litmus from red to blue, what can you say about its pH ?
(c) What can you say about the pH of a solution, that liberates carbon dioxide from sodium carbonate ?
18. Solution P has a pH of 13, solution Q has a pH of 6 and solution R has a pH of 2.
Which solution :
(a) will liberate ammonia from ammonium sulphate on heating ?
(b) is a strong acid ?
(c) contains molecules as well as ions ?

3B. SALTS AND THEIR LABORATORY PREPARATIONS

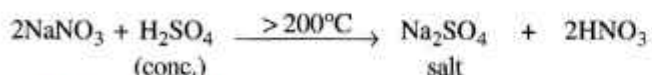
3.11 SALT

Salt is a compound formed by the partial or total replacement of the ionisable hydrogen atoms of an acid by a metallic ion or an ammonium ion.

Partial replacement



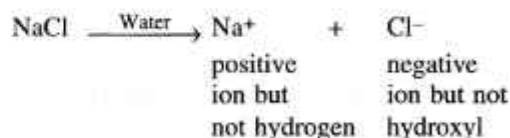
Complete replacement



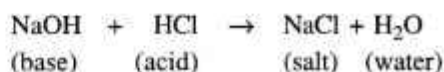
Ionic definition : Salt is an ionic compound, which dissociates in water to yield a positive ion

other than hydrogen ion (H^+) and a negative ion other than hydroxyl ion (OH^-).

For example, sodium chloride is a salt consisting of Na^+ and Cl^- .



It is prepared by neutralising sodium hydroxide with hydrochloric acid.



3.11.1 Classification of salts

There are six kinds of salts.

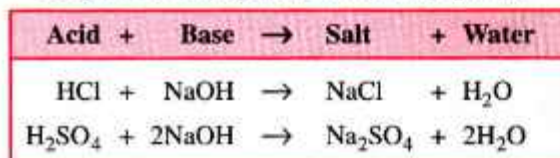
1. Normal salts
2. Acid salts
3. Basic salts
4. Double salts
5. Mixed salts
6. Complex salts

1. Normal salts : Normal salts are the salts formed by the complete replacement of the ionisable hydrogen atoms of an acid by a metallic or an ammonium ion.

For example :

Hydrochloric acid yields sodium chloride.

Sulphuric acid gives sodium sulphate.

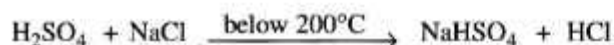


Normal salts have no ionisable hydrogen atoms.

2. Acid salts : Acid salts are formed by the partial replacement of the ionisable hydrogen atoms of a polybasic acid by a metal or an ammonium ion.

For example :

Dibasic sulphuric acid yields sodium hydrogen sulphate (NaHSO₄).



OR



Other examples of acid salts are :

NaHSO₃, Na₂HPO₄ and NaH₂PO₄.

Acid salts are usually formed when the amount of base taken is insufficient for the complete neutralisation of the acid.

Acid salts ionise in water solution to give hydronium ions, and therefore, they show all the properties of an acid.

3. Basic salts : Basic salts are formed by the partial replacement of the hydroxyl group of a di- or a tri- acidic base by an acid radical.

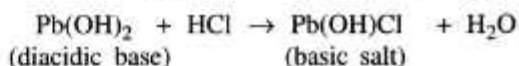
Examples :

Basic lead chloride [Pb(OH)Cl],

Basic Magnesium chloride [Mg(OH)Cl],

Basic copper chloride [Cu(OH)Cl], and

Basic copper nitrate [Cu(OH)NO₃]



The basic salt contains a metallic cation, a hydroxyl ion of the base and an anion of the acid.

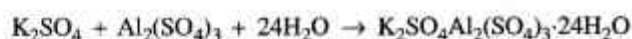
***4. Double salts :** Double salts are formed by the union of two simple salts that dissolve in water and crystallise.

Examples :

Potash alum : K₂SO₄ · Al₂(SO₄)₃ · 24H₂O

Mohr's salt : FeSO₄ · (NH₄)₂ SO₄ · 6H₂O

Dolomite : CaCO₃ · MgCO₃



***5. Mixed salts :** Mixed salts are those salts that contain more than one basic or acid radical.

Examples :

Sodium potassium carbonate : NaKCO₃;

Bleaching powder : CaOCl₂

***6. Complex salts :** Complex salts are those which on dissociation give one simple ion and one complex ion.

Complex salt ⇌ Simple ion + Complex ion



Sodium argentocyanide

Other examples are :

Silver amino chloride : [Ag(NH₃)₂]Cl

Tetrammine copper (II) sulphate : [Cu(NH₃)₄]SO₄

3.11.2 General properties of salts

- (i) Salts are electrovalent compounds. They conduct electricity in their molten state as well as in their aqueous solutions [due to their dissociation into ions].

* Not in syllabus

- (ii) Salts are non-volatile solids that form crystals.
- (iii) **Solubility of salts** : Most of the salts are soluble in water. Their degree of solubility in water varies with temperature (Fig. 3.3).

Note : All metallic oxides and hydroxides are insoluble except of sodium, potassium and ammonium. Calcium hydroxide is slightly soluble.

NaHCO_3 and KHCO_3 are sparingly soluble.

- (iv) **Hydrolysis of salts** :

The phenomenon due to which salt formed by a weak acid and a strong base, or by a strong acid and a weak base, reacts with water to give an acidic or an alkaline solution, is known as **hydrolysis**.

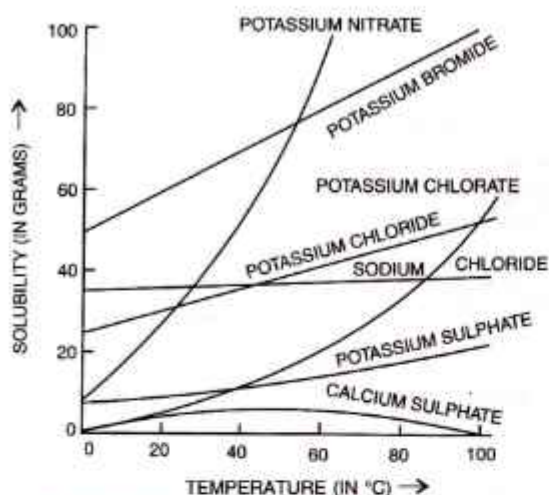


Fig. 3.3 Variation in solubility of some salts with change in temperature

If at least 1g of a substance can be dissolved in 100 mL of water at 298 K, it is called a **soluble salt**.

If 0.1 g to 1g of a substance can be dissolved in 100 mL water at 298 K, it is called a partially **soluble salt**.

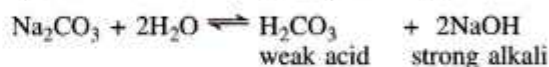
If less than 0.1 g of a substance can be dissolved in 100 mL water at 298 K, it is called an **insoluble salt**.

Table 3.1 : Solubility of salts

Salts	Exceptions
1. All the compounds of ammonium, sodium and potassium are <i>soluble</i> in water.	_____
2. All nitrates and nitrites are <i>soluble</i> .	_____
3. All chlorides, bromides and iodides are <i>soluble</i> in water.	Hg_2Cl_2 ; AgCl ; PbCl_2 (PbCl_2 is soluble in hot water, not in cold water).
4. All sulphates are <i>soluble</i> .	CaSO_4 ; PbSO_4 ; BaSO_4 .
5. All carbonates sulphides, sulphites and phosphates are insoluble.	Except of ammonium, sodium and potassium.

- (a) Salts formed from strong bases like NaOH and KOH and weak acids like H_2CO_3 , $\text{CH}_3\text{CO}_2\text{H}$, etc., **hydrolyse in water to give alkaline solutions. They have pH more than 7.**

Examples : Sodium carbonate (Na_2CO_3), potassium acetate (CH_3COOK),



- (b) The salts of strong acids and weak bases hydrolyse in water to give an **acidic solution** (pH less than 7).

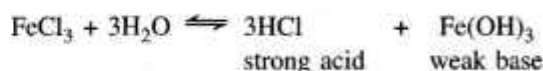
Examples :

Iron (III) chloride (FeCl_3),

Copper sulphate (CuSO_4),

Aluminium chloride (AlCl_3), and

Ammonium sulphate [$(\text{NH}_4)_2\text{SO}_4$].



- (c) Salts derived from strong acids and strong bases, such as NaCl , Na_2SO_4 , K_2SO_4 and KNO_3 gives a **neutral solution** in water.
- (d) Salts formed from weak bases like ammonium hydroxide and weak acids like acetic acid may be acidic, alkaline or neutral depending upon the dissociated ions and undissociated molecules.

Examples :

Ammonium acetate [$\text{CH}_3\text{COONH}_4$] and

Ammonium carbonate [$(\text{NH}_4)_2\text{CO}_3$].

3.11.3 Preparation of soluble salts

METHOD	REACTION
1. DIRECT COMBINATION (i.e., by heating two elements together)	$\begin{array}{l} \text{Metal} + \text{Non metal} \rightarrow \text{Salt} \\ 2\text{Na}(\text{molten}) + \text{Cl}_2 \rightarrow 2\text{NaCl} \\ 2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3 \end{array}$
2. SIMPLE DISPLACEMENT* (Action of dilute acids on active metals)	$\begin{array}{l} \text{Active metal} + \text{Acid} \rightarrow \text{Salt} + \text{Hydrogen} \\ \text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2\uparrow \\ \text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2\uparrow \end{array}$
3. DECOMPOSITION**	
(a) Decomposition of bicarbonates by acids.	$\begin{array}{l} \text{(i) NaHCO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2\uparrow \\ \text{sodium salt} \\ \text{bicarbonate} \\ \text{(ii) Ca(HCO}_3)_2 + 2\text{HNO}_3 \rightarrow \text{Ca(NO}_3)_2 + 2\text{H}_2\text{O} + 2\text{CO}_2 \\ \text{calcium salt} \\ \text{bicarbonate} \end{array}$
(b) Decomposition of carbonates by acids.	$\begin{array}{l} \text{(i) CuCO}_3 + 2\text{HCl} \rightarrow \text{CuCl}_2 + \text{H}_2\text{O} + \text{CO}_2\uparrow \\ \text{copper salt} \\ \text{carbonate} \\ \text{(ii) MgCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2\text{O} + \text{CO}_2\uparrow \\ \text{magnesium salt} \\ \text{carbonate} \end{array}$
(c) Decomposition of chlorides by conc. H_2SO_4 .	$\begin{array}{l} \text{(i) NaCl(s)} + \text{H}_2\text{SO}_4(\text{l}) \xrightarrow{\text{below } 200^\circ\text{C}} \text{NaHSO}_4(\text{aq}) + \text{HCl}(\text{g}) \\ \text{sodium (conc.) sodium} \\ \text{chloride bisulphate (acid salt)} \\ \text{(ii) 2NaCl(s)} + \text{H}_2\text{SO}_4(\text{l}) \xrightarrow{\text{above } 200^\circ\text{C}} \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{HCl}(\text{g}) \\ \text{sodium sulphuric sodium hydrogen} \\ \text{chloride acid (conc.) sulphate chloride gas} \\ \text{(iii) ZnCl}_2(\text{s)} + \text{H}_2\text{SO}_4(\text{l}) \rightarrow \text{ZnSO}_4(\text{aq}) + 2\text{HCl}(\text{g}) \\ \text{zinc chloride zinc sulphate} \end{array}$
(d) Decomposition of nitrates by conc. H_2SO_4 .	$\text{KNO}_3 + \text{H}_2\text{SO}_4 \xrightarrow{\text{below } 200^\circ\text{C}} \text{KHSO}_4 + \text{HNO}_3$ <p style="text-align: center;">acid salt</p>
NEUTRALISATION	
4. (i) Neutralising an acid with a soluble base is also known as titration . (ii) Neutralising an acid with an insoluble base.	$\begin{array}{l} \text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{H}_2\text{O} \\ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \\ 2\text{HCl} + \text{MgO} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O} \end{array}$
5. By the action of Alkalis (NaOH or KOH) on certain metals (Al , Zn) their oxides and their hydroxides.	$\begin{array}{l} 2\text{Al} + 2\text{NaOH} + 2\text{H}_2\text{O} \rightarrow 2\text{NaAlO}_2 + 3\text{H}_2 \\ \text{Zn} + 2\text{NaOH} \rightarrow \text{Na}_2\text{ZnO}_2 + \text{H}_2 \\ \text{Al}_2\text{O}_3 + 2\text{NaOH} \rightarrow 2\text{NaAlO}_2 + \text{H}_2\text{O} \\ \text{Al(OH)}_3 + \text{NaOH} \rightarrow \text{NaAlO}_2 + 2\text{H}_2\text{O} \\ \text{ZnO} + 2\text{KOH} \rightarrow \text{K}_2\text{ZnO}_2 + 2\text{H}_2\text{O} \\ \text{Zn(OH)}_2 + 2\text{KOH} \rightarrow \text{K}_2\text{ZnO}_2 + 2\text{H}_2\text{O} \end{array}$

Note : Soluble salts are prepared in solution only.

Soluble salts are obtained by the evaporation of water, followed by crystallisation.

Some soluble salts, like zincates and aluminates, are found only in solution. These salts cannot be obtained in their pure forms by crystallisation.

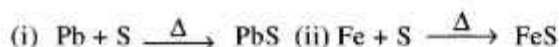
* Displacement is a chemical change in which a more active element displaces a less active element from its salt solution

** Decomposition is a chemical change in which a compound breaks up into its elements or simpler compounds

3.11.4 Preparation of insoluble salts

1. By direct combination

Reaction :



2. By the combination of an acidic oxide with a basic oxide.

Reaction :

	Acidic oxide	+	Basic oxide	→	Salt
(i)	SO ₂	+	CaO	$\xrightarrow{\Delta}$	CaSO ₃
(ii)	CO ₂	+	CaO	$\xrightarrow{\Delta}$	CaCO ₃

3. By the precipitation reaction (DOUBLE DECOMPOSITION)*

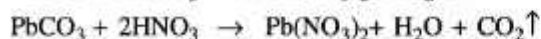
Reaction :

	Soluble salt	+	Acid or soluble salt	→	Insoluble salt	+	Acid or salt
(i)	BaCl ₂	+	H ₂ SO ₄	→	BaSO ₄ ↓	+	2HCl
(ii)	AgNO ₃	+	HCl	→	AgCl↓	+	HNO ₃
(iii)	Pb(NO ₃) ₂	+	Na ₂ CO ₃	→	PbCO ₃ ↓	+	2NaNO ₃
(iv)	Na ₂ SO ₄	+	BaCl ₂	→	BaSO ₄ ↓	+	2NaCl
(v)	CuSO ₄	+	H ₂ S	→	CuS↓	+	H ₂ SO ₄
(vi)	Pb(NO ₃) ₂	+	2HCl	→	PbCl ₂ ↓	+	2HNO ₃

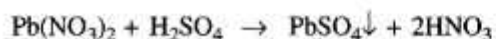
Note : An insoluble salt can also be obtained from another insoluble salt, by double decomposition. Insoluble salt is first to be converted into a soluble salt, which can then be used for preparing the desired salt.

For example :

- (i) Insoluble **lead sulphate** is prepared from insoluble lead carbonate or insoluble lead oxide, by first converting either of them into soluble lead nitrate (with the help of dilute nitric acid).



The resulting solution is then treated with sulphuric acid or sodium sulphate solution and a **white precipitate** of lead sulphate is obtained.



*A chemical change, in which two compounds in solution react to form two other compounds by the mutual exchange of radicals. A solid precipitate is formed as a result of the reaction.

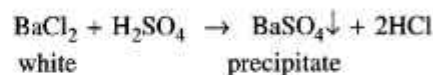
Adding dilute H₂SO₄ to PbCO₃ in order to prepare PbSO₄ is an **impractical method** because an insoluble crust of lead sulphate is formed on lead carbonate which prevents further reaction.

PbCl₂ is prepared from insoluble lead oxide or lead carbonate by the same method.

- (ii) **Silver chloride and barium sulphate** silver chloride is usually prepared by reacting of aqueous silver nitrate with dilute hydrochloric acid and precipitate of silver chloride is obtained.



Barium sulphate is prepared by reacting aqueous barium chloride with dilute sulphuric acid.



3.12 LABORATORY PREPARATION OF SOME SALTS (NORMAL AND ACID SALTS)

1. Iron (III) chloride [FeCl₃]

Method of preparation : Synthesis, i.e., direct combination of elements.

Procedure (Fig. 3.4) :

Powdered iron is taken in a combustion tube. Dry chlorine gas (obtained by passing the gas through concentrated sulphuric acid) is passed through the combustion tube. The air inside the apparatus is expelled.

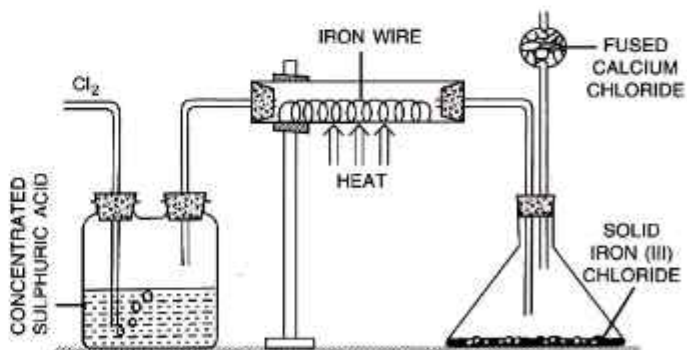
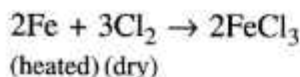


Fig. 3.4. Preparation of Iron (III) chloride.

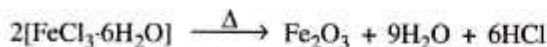
The iron turns red hot since the reaction is exothermic. Now, external heating is suspended.

Iron (III) chloride is formed, which volatilizes and condenses in the receiver as **brown scale**.

Reaction :

Iron (III) chloride is highly deliquescent, so it is kept dry with the help of fused calcium chloride (drying agent).

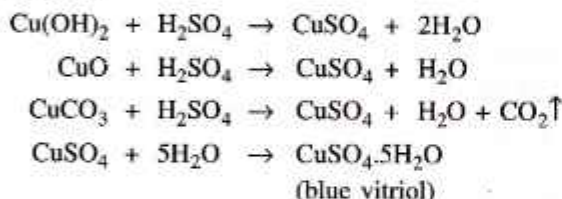
Note : Anhydrous FeCl_3 cannot be prepared by simply heating the hydrated ferric chloride $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$, because on heating, $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$ produces $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$ and HCl .



Anhydrous Iron (III) chloride is used as a catalyst in certain reactions.

2. Copper (II) sulphate (or blue vitriol)

Method of preparation : By the action of dilute acid (dil H_2SO_4) on an insoluble base copper hydroxide, copper oxide or copper carbonate.

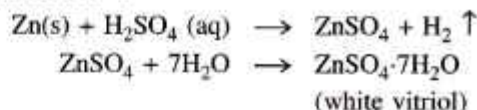
Reaction :**Procedure :**

- Step 1.** Take dilute sulphuric acid in a beaker and heat it on a wire gauze.
- Step 2.** Add copper hydroxide or black cupric oxide or copper carbonate, in small quantities at a time and keep on stirring the solution till no more of it dissolves and the excess compound settles to the bottom.
- Step 3.** Filter it hot, and collect the filtrate in a china dish. Evaporate the filtrate by heating to the point of crystallisation, and then allow it to cool.
- Step 4.** Collect the bright blue crystals of copper (II) sulphate penta-hydrate $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (blue vitriol), and dry the crystals.

Note : Other salts which can be prepared by the same method are : zinc sulphate, lead nitrate, calcium chloride, calcium nitrate and magnesium sulphate.

3. Zinc sulphate and Iron (II) sulphate

Method of preparation : By the action of dilute acid (dil H_2SO_4) on an active metal (Zinc, Iron).

(a) Zinc sulphate (or white vitriol)**Reaction :****Procedure :**

- Step 1.** Take dilute sulphuric acid (1 volume of acid : 5 volumes of water) in a beaker and heat it on a wire gauze. Add some granulated zinc pieces, a little at a time, with constant stirring.
- Step 2.** Effervescence takes place because of the liberation of hydrogen gas. Add zinc till the metal settles at the base of the beaker. When effervescence stops, it indicates that all the acid has been used up.
- Step 3.** The excess of zinc is filtered off.
- Step 4.** Collect the solution in a china dish. Evaporate the solution to get crystals. Filter, wash them with water and dry them between the folds of a filter paper.

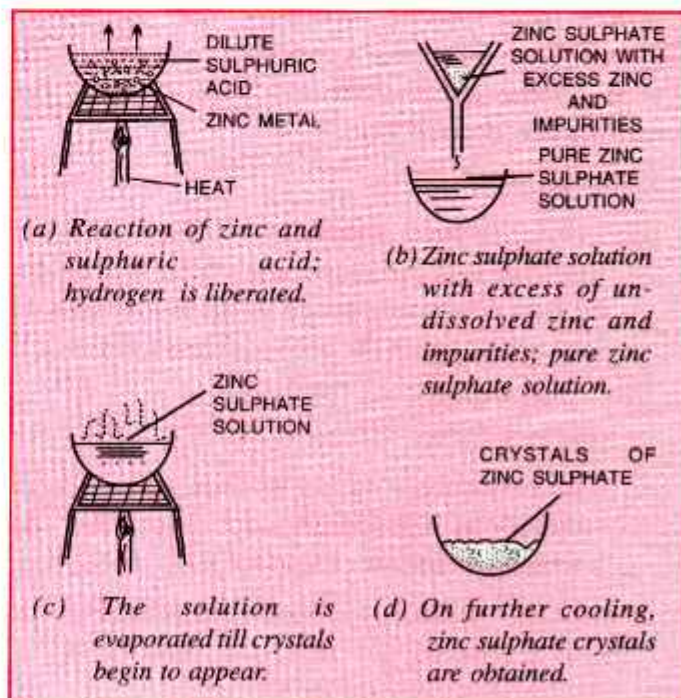


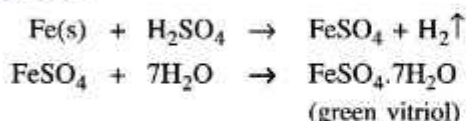
Fig. 3.5 Preparation of zinc sulphate crystals from zinc and sulphuric acid.

The white, needle-shaped crystals are of hydrated zinc sulphate, commonly known as **white vitriol**.

(b) Iron (II) sulphate (or green vitriol)

Iron (II) sulphate is prepared by the same method as white vitriol.

Reaction :



Filter the hot solution formed and allow the filtrate to evaporate at room temperature. Pale green crystals of hydrated iron (II) sulphate are formed.

4. Lead Chloride (PbCl_2) and Calcium Carbonate (CaCO_3)

Method of preparation : Precipitation or double decomposition

An insoluble salt can be prepared by mixing the solutions of two appropriate salts, each containing one of its ions. This mixing results in the formation of an insoluble solid precipitate, which can be filtered, washed and dried.

(a) Lead chloride (PbCl_2)

It is prepared by adding either dilute hydrochloric acid or sodium chloride solution to a solution of lead nitrate (Fig 3.6).

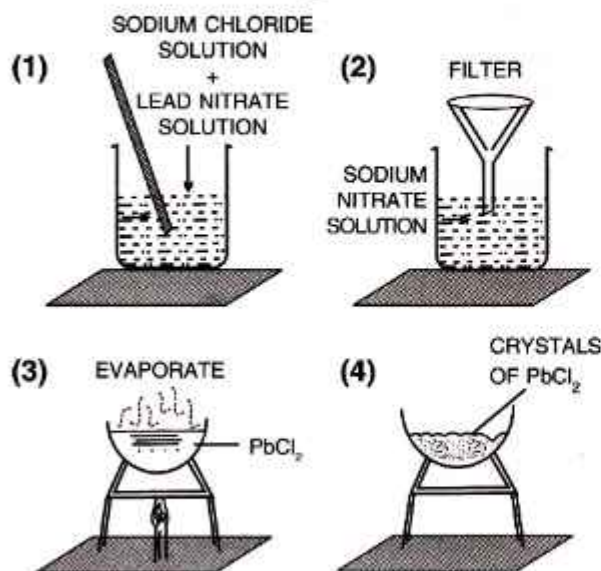
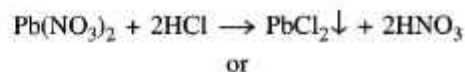


Fig. 3.6 Preparation of lead chloride.

Reaction :



The solution of lead nitrate and sodium chloride is then filtered. The heavy, white precipitate of lead chloride accumulate as residue on the filter paper. It is then washed repeatedly with cold water. The precipitate is taken in a china dish with some water and is heated till the precipitate of lead chloride dissolves. The solution is then cooled.

Pure, needle shaped crystals of lead chloride are thus obtained.

Note : lead chloride dissolves in hot water.

The chlorides of Pb, Ag and Hg (ous), and the sulphates of Ba, Pb and Ca are also prepared using this method.

(b) Calcium Carbonate (CaCO_3)

It is prepared by adding sodium carbonate solution to a hot solution of calcium chloride in a beaker, till the former is in excess.

Reaction :



The white precipitate formed by the interchange of radicals is filtered and washed repeatedly with cold water and collected. The washed precipitate is dried to obtain an amorphous powder, *i.e.*, calcium carbonate.

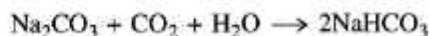
The carbonates of all the metals (except Na, K and NH_4^+) are prepared by the above method.

5. Preparation of an acid salt sodium bicarbonate (NaHCO_3)

Principle :

By passing carbon dioxide gas into a cold solution of sodium carbonate.

Reaction :



Procedure :

- Step 1.** Dissolve nearly 5 grams of anhydrous sodium carbonate in about 25 cm³ of distilled water in a flask.
- Step 2.** Cool the solution by keeping the flask in a freezing mixture.
- Step 3.** Pass carbon dioxide gas in the solution.
- Step 4.** Crystals of sodium bicarbonate will precipitate out after sometime.
- Step 5.** Filter the crystals and dry it in the folds of filter paper.

The procedure is shown in Fig. 3.7.

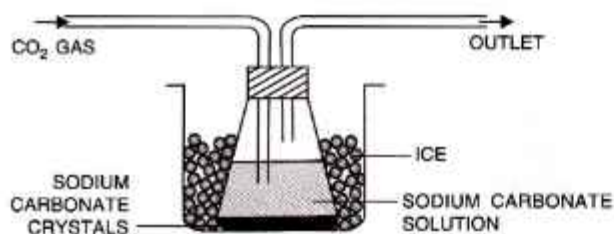


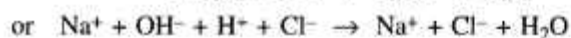
Fig. 3.7 Preparation of sodium bicarbonate NaHCO_3 .

Note : Sodium hydrogen carbonate is used in **health salts**. Health salts are the mixture of dry citric acid or tartaric acid and sodium hydrogen carbonate. On adding water, acid reacts with NaHCO_3 to liberate CO_2 which forms an effervescence drink and releases gas pressure in the stomach.

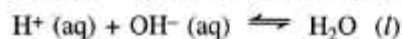
6. Neutralisation

Neutralisation is the process by which H^+ ions of an acid react completely with the $[\text{OH}^-]$ ions of a base to give salt and water only.

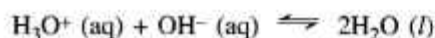
For example, in aqueous solutions :



Na^+ and Cl^- are present in the reactants as well as in the principal product. Thus, by cancelling out the ions present in the reactants as well as in one of the products, we get the following equation :



OR



Similarly,



Heat of neutralisation : It is the amount of heat liberated when one gram equivalent of an acid or a base is completely neutralised.

7. Laboratory preparation of sodium sulphate crystals ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$)

Principle :

Neutralisation of caustic soda with dilute sulphuric acid.

Reaction :



This method is useful for preparing the soluble salts of only sodium, potassium and ammonium.

Since the reactants as well as the products are soluble, a *titration* is conducted to determine the completion of the neutralisation reaction (Fig. 3.8).

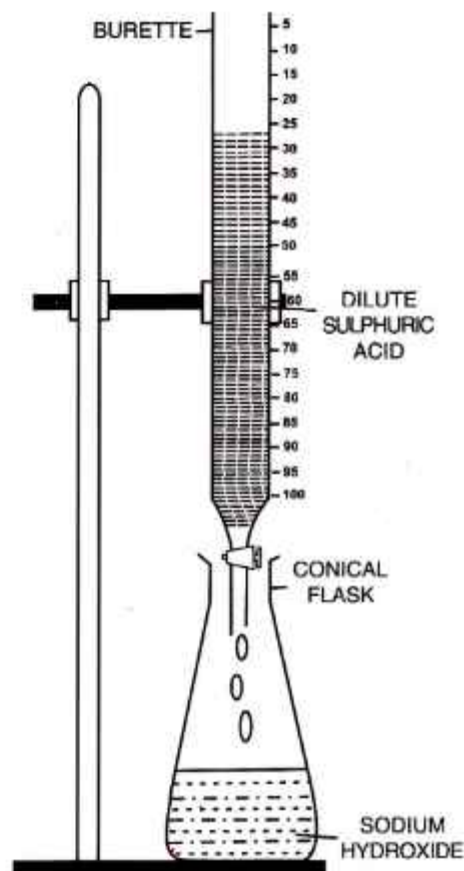


Fig. 3.8 Neutralization of an acid with a base.

Procedure :

Step 1. 25 cm³ of sodium hydroxide solution is measured and transferred to a conical flask (see Fig. 3.8). A drop of indicator **phenolphthalein** is added. The solution will turn pink.

Step 2. A burette is filled with dilute sulphuric acid, and the reading on the burette is noted (Zero cc). The acid is run in drops into the conical flask (alkali) till the last drop of the acid turns the solution colourless. This is known as the end point.

Step 3. Note the burette reading at the end point. The difference between the two readings is the volume of the acid required.

The acid and the alkali (in the same ratio as found above) are now mixed in an evaporating basin. The resulting solution contains sodium sulphate and water.

On evaporating this solution, crystals of Na₂SO₄.10H₂O are obtained.

Note :

- (i) In place of sodium hydroxide, sodium carbonate can be used in this reaction. However, methyl orange is then used as the indicator.
- (ii) Since sodium hydroxide and sulphuric acid are both soluble, an excess of either of them cannot be removed by filtration. Therefore, it is necessary to find out on a small scale, the ratio of the solutions of the two reactants required for complete neutralisation before preparation.

In order to prepare acid salt NaHSO₄ add the same volume of sulphuric acid again to the product formed in preparing the normal salt Na₂SO₄. Crystals of sodium hydrogen sulphate are obtained from the resulting solution. *i.e.* for a given amount of alkali, the addition of twice the normal amount of acid will result in the formation of the **acid salt**.



or

**Intext Questions**

1. Define the following and give two examples in each case : (a) a normal salt, (b) an acid salt, (c) a mixed salt.
2. Answer the following questions related to salts and their preparations :
 - (a) What is a 'salt' ?
 - (b) What kind of salt is prepared by precipitation?
 - (c) Name a salt prepared by direct combination. Write an equation for the reaction that takes place in preparing the salt you have named.
 - (d) Name the procedure used to prepare a sodium salt such as sodium sulphate.
3. Describe giving all practical details, how would you prepare :
 - (a) copper sulphate crystals from a mixture of charcoal and black copper oxide,
 - (b) zinc sulphate crystals from zinc dust (powdered zinc and zinc oxide),
 - (c) sodium hydrogen carbonate crystals.
 - (d) Calcium sulphate from calcium carbonate.
4. The following is a list of methods for the preparation of salts.
 - A – direct combination of two elements.
 - B – reaction of a dilute acid with a metal.
 - C – reaction of a dilute acid with an insoluble base.
 - D – titration of a dilute acid with a solution of soluble base.
 - E – reaction of two solutions of salts to form a precipitate.

Choose from the above list A to E, the best method of preparing the following salts by giving a suitable equation in each case :

 1. Anhydrous ferric chloride, 2. Lead chloride,
 3. Sodium sulphate, 4. Copper sulphate.
5. Name :
 - (a) a chloride which is insoluble in cold water but dissolves in hot water,
 - (b) a chloride which is insoluble,
 - (c) two sulphates which are insoluble,
 - (d) a basic salt, (e) an acidic salt,
 - (f) a mixed salt, (g) a complex salt,
 - (h) a double salt,
 - (i) salt whose solubility increases with temperature,
 - (j) a salt whose solubility decreases with temperature.

6. Fill in the blanks with suitable words :

An acid is a compound which when dissolved in water forms hydronium ions as the only ions. A base is a compound which is soluble in water contains ions. A base reacts with an acid to form a and water only. This type of reaction is known as

7. What would you observe when :

- blue litmus is introduced into a solution of ferric chloride,
- red litmus paper is introduced into a solution of sodium sulphate,
- red litmus paper is introduced in sodium carbonate solution ?

8. Explain why :

- it is necessary to find out the ratio of reactants required in the preparation of sodium sulphate,
- fused calcium chloride is used in the preparation of FeCl_3 ?
- Anhydrous FeCl_3 cannot be prepared by heating hydrated Iron (III) chloride.

9. Give the preparation of the salt shown in the left column by matching with the methods given in the right column. Write a balanced equation for each preparation.

Salt	Method of preparation
Zinc sulphate	Precipitation
Ferrous sulphide	Oxidation
Barium sulphate	Displacement
Ferric sulphate	Neutralisation
Sodium sulphate	Synthesis.

10. (a) Give the pH value of pure water. Does it change if common salt is added to it ?

(b) Classify the following solutions as acids, bases or salts.

Ammonium hydroxide, barium chloride, sodium chloride, sodium hydroxide, H_2SO_4 and HNO_3 .

11. Define the term neutralisation :

- Give a reaction, mentioning clearly acid and base used in the reaction,
- If one mole of a strong acid reacts with one mole of a strong base, the heat produced is always the same. Why ?

12. Write the balanced equations for the preparation of the following salts in the laboratory :

- A soluble sulphate by the action of an acid on an insoluble base,

- An insoluble salt by the action of an acid on another salt,

- An insoluble base by the action of a soluble base on a soluble salt,

- A soluble sulphate by the action of an acid on a metal.

13. You are provided with the following chemicals :

NaOH , Na_2CO_3 , H_2O , $\text{Zn}(\text{OH})_2$, CO_2 , HCl , Fe , H_2SO_4 , Cl_2 , Zn .

Using the suitable chemicals from the given list only, state briefly how you would prepare :

- iron (III) chloride,
- sodium sulphate,
- sodium zincate,
- iron (II) sulphate,
- sodium chloride ?

14. For each of the salt : A, B, C and D, suggest a suitable method of its preparation.

- A is a sodium salt.
- B is an insoluble salt.
- C is a soluble salt of copper.
- D is a soluble salt of zinc.

15. Choosing only substances from the list given in the box below, write equations for the reactions which you would use in the laboratory to obtain :

- Sodium sulphate,
- Copper sulphate,
- Iron(II) sulphate,
- Zinc carbonate,

Dilute sulphuric acid	Copper	Copper carbonate
	Iron	Sodium carbonate
	Sodium	
	Zinc	

16. From the formula listed below, choose one, in each case, corresponding to the salt having the given description :- AgCl , CuCO_3 , $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, KNO_3 , NaCl , NaHSO_4 , $\text{Pb}(\text{NO}_3)_2$, ZnCO_3 , $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$.

- an acid salt.
- an insoluble chloride.
- on treating with concentrated sulphuric acid, this salt changes from blue to white.
- on heating, this salt changes from green to black.
- this salt gives nitrogen dioxide on heating.

17. (a) $\text{Ca}(\text{H}_2\text{PO}_4)_2$ is an example of a compound called _____ (acid salt / basic salt / normal salt).

- Write the balanced equation for the reaction of : A named acid and a named alkali.

16. State the terms defined by the following sentences :
- A soluble base.
 - The insoluble solid formed when two solutions are mixed together.
 - An acidic solution in which there is only partial ionization of the solute molecules.
17. Which of the following methods, A, B, C, D or E is generally used for preparing the chlorides listed below from (i) to (v). Answer by writing down the chloride and the letter pertaining to the corresponding method. Each letter is to be used only once.
- Action of an acid on a metal,
 - Action of an acid on an oxide or carbonate,
 - Direct combination,
 - Neutralization of an alkali by an acid
 - Precipitation (double decomposition)
 - Copper (II) chloride

- Iron (II) chloride
- Iron (III) chloride
- Lead (II) chloride
- Sodium chloride

18. Complete the following table :

Reactants	Products	Method
Soluble base + Acid (dil)	Salt + water	Neutralisation Titration
Metal + Non-metal	Salt (soluble/insoluble)
Insoluble base +	Salt (soluble) + water
Active metal + Acid (dil) +
Soluble salt solution (A) + Soluble salt solution (B)	Precipitated salt + Soluble salt
Carbonate/bicarbonate + Acid (dil)	Salt + +	Decomposition of carbonate
Chlorides/nitrates + Acid (conc.) +	Decomposition of chlorides and nitrates

3C. WATER OF CRYSTALLISATION; EFFLORESCENCE; DELIQUESCENT

3.13 WATER OF CRYSTALLISATION

Some salts, while crystallising out from their solutions, unite with a definite quantity of water, which is known as the **water of crystallisation**.

This water of crystallisation is in loose chemical combination with the salts, and it can be driven out by heating the powdered crystals of these salts above 100°C.

Experiment : To show that copper sulphate crystals contain water of crystallisation.

Procedure : Take some powdered copper sulphate crystals in a clean, dry test tube, and clamp it in a tilted position (Fig. 3.9) so that the condensed water may not slip back and crack the test tube at its hot end. Heat the crystals. Drops of a colourless liquid will condense on the cooler parts of the test tube, and finally a little of it can be collected in a dish.

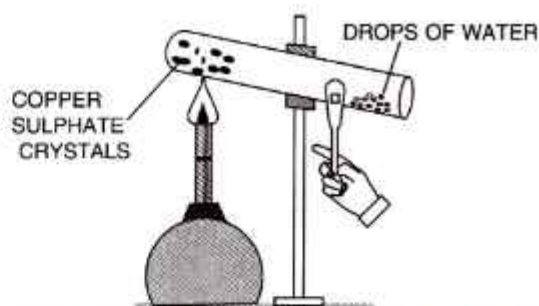
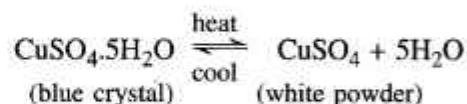


Fig. 3.9 Removal of water of crystallisation by heating.

Reaction :

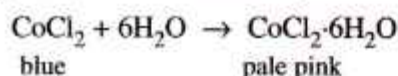


Test for condensed liquid :

- It is neutral to litmus.
- Spread some anhydrous copper sulphate (white) on a glass slide in a thin layer and add a few drops of the liquid in the dish. A **blue patch will be formed, thereby showing that the liquid is water.**

Common name	Chemical name	Formula
Washing soda crystals	Sodium carbonate decahydrate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Epsom salt	Magnesium sulphate heptahydrate	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Potash alum	Hydrated potassium aluminium sulphate	$\text{K}_2\text{SO}_4 \cdot \text{Al}_2(\text{SO}_4)_3 \cdot 24\text{H}_2\text{O}$
Glauber's salt	Sodium sulphate decahydrate	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
Blue vitriol	Copper (II) sulphate pentahydrate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Plaster of Paris	Calcium sulphate semihydrate	$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$
White vitriol	Zinc sulphate heptahydrate	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
Gypsum	Hydrated calcium sulphate	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

(iii) The liquid (water) turns blue CoCl_2 paper pale pink (almost colourless).



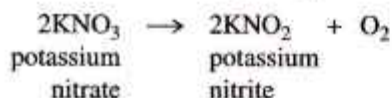
Crystalline solids not containing water of crystallisation.

Common salt (NaCl),
Nitrate (KNO_3),
Sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$),
Potassium permanganate (KMnO_4), and
Ammonium chloride (NH_4Cl), etc.

Experiment : Heat gently a few crystals of pure sodium chloride and potassium nitrate separately, in different test tubes.

Observation for sodium chloride : On heating, crystals of sodium chloride produce a crackling sound, called **decrepitation**. This is due to the bigger crystals breaking up into smaller ones.

Observation for potassium nitrate : The crystals of potassium nitrate melt into a colourless liquid and, on further heating, produce oxygen which rekindles a glowing splinter. Potassium nitrate, a pale yellow solid residue, is left behind.



Note : These crystalline salts do not produce water vapour on heating.

Conclusion : All crystalline salts do not contain water of crystallisation.

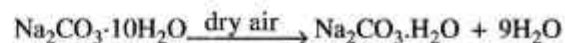
3.14 EFFLORESCENCE

Efflorescence is the property of some substances to lose wholly, or partly their water of crystallisation when their crystals are exposed to dry air even for a short time. They become powdery. Such substances are called **efflorescent substances**.

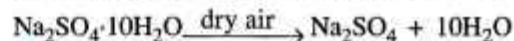
Efflorescent substances lose their water of crystallisation and thus their crystalline shape becomes powdery, when exposed to air.

Examples of efflorescent substances :

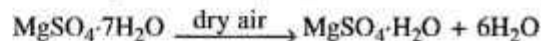
(a) **Washing soda** [hydrated sodium carbonate], when exposed to dry air, becomes a monohydrate.



(b) **Glauber's salt** [hydrated sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$] becomes a powdery anhydrous sodium sulphate when exposed to air.



(c) **Epsom salt** [magnesium sulphate heptahydrate], when exposed to dry air, becomes a monohydrate.



Note : The higher the temperature of the air, the higher the efflorescence. This is because the air absorbs more water with rising temperature and decreasing moisture.

3.15 HYGROSCOPIC SUBSTANCES

Certain substances absorb moisture (water vapour) from the atmosphere without dissolving in it. Such substances are called

hygroscopic substances and the phenomenon is called **hygroscopy**.

Some examples of hygroscopic substances are :

- (i) Conc. sulphuric acid (H_2SO_4)
- (ii) Phosphorus pentoxide (P_2O_5)
- (iii) Quicklime (CaO)
- (iv) Silica gel

In the laboratory, these substances are used generally for drying gases, *i.e.*, for removal of moisture from gases.

3.16 DELIQUESCENT SUBSTANCES

Certain water-soluble substances, when exposed to the atmosphere at ordinary temperature, absorb moisture from the atmospheric air to become moist, and ultimately dissolve in the absorbed water, forming a saturated solution. Such a substance is called a **deliquescent** substance, and the phenomenon is called **deliquescence**.

Deliquescent substances

Examples : Caustic soda ($NaOH$),
Caustic potash (KOH),
Magnesium chloride ($MgCl_2$),
Zinc chloride ($ZnCl_2$),
Calcium chloride ($CaCl_2$),
Ferric chloride ($FeCl_3$),
Zinc nitrate [$Zn(NO_3)_2$], and
Copper nitrate [$Cu(NO_3)_2$].

Deliquescence occurs when the *vapour pressure of the salt* is much *lower* compared to **atmospheric vapour pressure**. Thus, *deliquescence is minimized*

during dry conditions, whereas *efflorescence is maximized in dry conditions*.

Note : Efflorescence substances loses their weight while hygroscopic and deliquescence substances gains weight when exposed to atmosphere.

Note : Table salt [sodium chloride] turns moist and ultimately forms a solution, on exposure to air [especially during the rainy season]. Though pure sodium chloride is not deliquescent, the commercial version of the salt contains **impurities**, like **magnesium chloride** and **calcium chloride**, which are deliquescent substances.

This kind of impurity can be removed by passing a current of dry hydrogen chloride gas through a saturated solution of the affected salt. Pure sodium chloride is produced as a precipitate, which can be recovered by filtering and washing first with a little water and finally with alcohol.

3.17 DRYING AND DEHYDRATING AGENTS

Drying or desiccating agents

Certain substances remove moisture from other substances and are, therefore, called **desiccants** or **desiccating agents** or **drying agents**. Almost all hygroscopic substances are desiccating agents.

Examples : Conc. sulphuric acid, phosphorus pentoxide, silica gel and quicklime.

Dehydrating agents are substances that can remove water molecules even from compounds. Concentrated sulphuric acid can remove water molecules from blue vitriol ($CuSO_4 \cdot 5H_2O$), so it is a dehydrating agent as well.

Differences between drying agent and dehydrating agent

Drying agent	Dehydrating agent
(i) They remove moisture from other substances.	(i) They remove chemically combined elements of water in the ratio of 2 : 1 (hydrogen : oxygen) from a compound.
(ii) They are used to dry gases like chlorine, sulphur dioxide, hydrogen chloride, etc. They are also used in desiccators to keep substances dry.	(ii) They prepare substances like carbon monoxide, sugar charcoal, etc.
(iii) They represent physical change.	(iii) They represent chemical change.
Examples : Phosphorus pentoxide P_2O_5 , fused calcium chloride $CaCl_2$, calcium oxide CaO , conc. sulphuric acid H_2SO_4 .	For example : Conc. Sulphuric acid H_2SO_4 .

Intext Questions

- What do you understand by water of crystallisation?
Give four substances which contain water of crystallisation and write their common names.
- (a) Define efflorescence. Give examples.
(b) Define deliquescence. Give examples.
- Distinguish between drying and dehydrating agent.
- Explain clearly how conc. H_2SO_4 is used as dehydrating as well as drying agent.
- M is an element in the form of a powder. M burns in oxygen and the product obtained is soluble in water. The solution is tested with litmus. Write down *only the word* which will correctly complete each of the following sentences.
 - If M is a metal, then the litmus will turn _____.
 - If M is a non-metal, then the litmus will turn _____.
 - If M is a reactive metal, then _____ will be evolved when M reacts with dilute sulphuric acid.
 - If M is a metal, it will form _____ oxide, which will form _____ solution with water.
 - If M is a non-metal, it will not conduct electricity in the form of _____.
- Give reasons for the following :
 - Sodium hydrogen sulphate is not an acid but it dissolves in water to give hydrogen ions, according to the equation
$$NaHSO_4 \rightleftharpoons H^+ + Na^+ + SO_4^{2-}$$
 - Anhydrous calcium chloride is used in a desiccator.
- State whether a sample of each of the following would increase or decrease in mass if exposed to air.

(a) Solid NaOH,	(b) Solid $CaCl_2$,	(c) Solid $Na_2CO_3 \cdot 10H_2O$,
(d) Conc. Sulphuric acid,	(e) Iron (III) Chloride	
- (a) Why does common salt get wet during the rainy season?
(b) How can this impurity be removed ?
(c) Name a substance which changes the blue colour of copper sulphate crystals to white.
(d) Name two crystalline substances which do not contain water of crystallisation.
- Answer the questions below, relating your answers only to salts in the following list : Sodium chloride, anhydrous calcium chloride, copper sulphate-5-water ?
 - What name is given to the water in the compound copper sulphate-5-water ?
 - If copper sulphate-5-water is heated, anhydrous copper sulphate is formed. What is its colour ?
 - By what means, other than heating, could you dehydrate copper sulphate-5-water and obtain anhydrous copper sulphate ?
 - Which one of the salts in the given list is deliquescent ?

EXERCISE

MISCELLANEOUS QUESTIONS BASED ON ICSE EXAMINATIONS

- Write the balanced equations for the preparation of the following compounds (as the major product) starting from iron and using only one other substance :

(a) Iron (II) chloride,	(b) Iron (III) chloride,
(c) Iron (II) sulphate,	(d) Iron (II) sulphide.
- Write balanced reactions for the following conversions (A, B, C, D). (2015)

$$Fe \xrightarrow{A} FeCl_2 \xrightarrow{B} FeCO_3 \xrightarrow{C} Fe(NO_3)_2 \xrightarrow{D} Fe(OH)_2$$
- The preparation of lead sulphate from lead carbonate is a two-step process. (lead sulphate cannot be prepared by adding dilute sulphuric acid to lead carbonate.)
 - What is the first step that is required to prepare lead sulphate from lead carbonate ?
 - Write the equation for the reaction that will take place when this first step is carried out.
 - Why is the direct addition of dilute sulphuric acid to lead carbonate an impractical method of preparing lead sulphate ?

4. Match the following :

Column A	Column B
(a) Acid salt	A. Sodium potassium carbonate
(b) Mixed salt	B. Alum
(c) Complex salt	C. Sodium carbonate
(d) Double salt	D. Sodium zincate
(e) Normal salt	E. Sodium hydrogen carbonate

5. Write balanced equations for the following reactions :

- Lead sulphate from lead nitrate solution and dilute sulphuric acid,
- Copper sulphate from copper and concentrated sulphuric acid,
- Lead chloride from lead nitrate solution and sodium chloride solution,
- Ammonium sulphate from ammonia and dilute sulphuric acid,
- Sodium chloride from sodium carbonate solution and dilute hydrochloric acid.

2008

- What are the terms defined by the following ?
 - A salt containing a metal ion surrounded by other ions or molecules,
 - A base which is soluble in water.
- Making use only of substances chosen from those given below :

Dilute sulphuric acid	sodium carbonate
Zinc	sodium sulphite
Lead	calcium carbonate

 give the equations for the reactions by which you could obtain :
 - hydrogen,
 - sulphur dioxide,
 - carbon dioxide,
 - zinc carbonate (two steps required).

2009

- The acid which contains four hydrogen atoms
 - Formic
 - Sulphuric
 - Nitric
 - Acetic-acid
- A black coloured solid which on reaction with dilute sulphuric acid forms a blue coloured solution is
 - Carbon
 - Manganese [IV] oxide
 - Lead [II] oxide
 - Copper [II] oxide
- Solution A is strong acid B weak acid C strong alkali
 - Which solution contains solute molecules in addition to water molecules

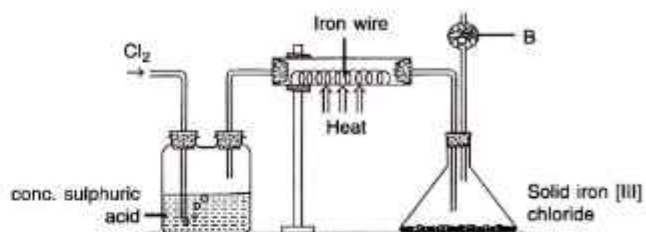
(ii) Which solution gives gelatinous white ppt with zinc sulphate, ppt disappears in excess

(iii) Give example of weak alkali.

(d) Write the equations[s] for the reaction[s] to prepare lead sulphate from lead carbonate.

(e) Define the following term — Neutralization.

(f) The diagram given below is to prepare Iron [III] chloride in the laboratory :

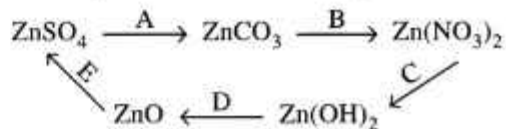


- What is substance B.
- What is the purpose of B.
- Why is iron [III] chloride to be stored in a closed container.
- Write the equation for the reaction between iron and chlorine.

2010

- Select the correct answer from A, B, C, D and E —
 - Nitroso iron [II] sulphate
 - Iron [III] chloride
 - Chromium sulphate
 - Lead chloride
 - Sodium chloride
 - A deliquescent compound
 - A compound soluble in hot water but insoluble in cold water
 - A compound which in the aqueous solution state, is neutral in nature.
- Select the correct answer from A, B, C and D —
 - A weak organic acid is :

A. Formic acid	B. Sulphuric acid
C. Nitric acid	D. Hydrochloric acid
 - A complex salt is
 - Zinc sulphate
 - Sodium hydrogen sulphate
 - Iron [ammonium sulphate]
 - Tetrammine copper [II] sulphate
- Give equations for the following conversions A to E —



- (d) For the preparation of the following salts — give a balanced equation in each case.
- Copper [II] sulphate from copper [II] oxide
 - Iron [III] chloride from the metal iron
 - Potassium sulphate from KOH solution
 - Lead [II] chloride from lead carbonate [give two equations].

2011

- (a) Write the balanced chemical equation — Lead nitrate solution is added to sodium chloride solution.
- (b) State what happens to crystals or washing soda when exposed to air. Name the phenomenon exhibited.
- (c) Name the method used for the preparation of the following salts from the list given below :
- Sodium nitrate
 - Iron (III) chloride
 - Lead chloride
 - Zinc sulphate
 - Sodium hydrogen sulphate

List :

- Simple displacement
- Neutralisation
- Decomposition by acid
- Double decomposition
- Direct synthesis

2012

- (a) Match the following :

Column A	Column B
(i) Acid salt	(A) Ferrous ammonium sulphate
(ii) Double salt	(B) Contains only ions
(iii) Ammonium hydroxide solution	(C) Sodium hydrogen sulphate
(iv) Dilute hydrochloric acid	(D) Contains only molecules
(v) Carbon tetrachloride	(E) Contains ions and molecules

2013

- (a) Select the word/s given which are required to correctly complete the blanks —
[ammonia, ammonium carbonate, carbon dioxide, hydrogen, hydronium, hydroxide, precipitate, salt, water]

- A solution M turns blue litmus red, so it must contain (1) ions; another solution turns red litmus blue and hence, must contain (2) ions.
- When solution M and O are mixed together, the products will be (3) and (4)
- If a piece of magnesium was put into a solution M, (5) gas would be evolved.

- (b) Give a suitable chemical term for :

- A salt formed by incomplete neutralisation of an acid by a base.
- A definite number of water molecules bound to some salts.
- The process in which a substance absorbs moisture from the atmospheric air to become moist, and ultimately dissolves in the absorbed water.

- (c) Choosing the substances from the list given : dil. sulphuric acid, copper, iron, sodium copper [II] carbonate, sodium carbonate, sodium chloride, zinc nitrate.

Write balanced equations for the reactions which would be used in the laboratory to obtain the following salts

- sodium sulphate
- zinc carbonate
- copper [II] sulphate
- iron [II] sulphate

2014

- Fill in the blank from the choices given : The basicity of acetic acid is [3, 1, 4].
- Draw the structure of the stable positive ion formed when an acid dissolves in water.
- State the inference drawn from the observation :
Salt S is prepared by reacting dilute sulphuric acid with copper oxide. Identify S.
- Give balanced chemical equations for the preparation of the following salts :
 - Lead sulphate – from lead carbonate.
 - Sodium sulphate – using dilute sulphuric acid.
 - Copper chloride – using copper carbonate.