CHEMISTRY
UNIT: 6 SALTS SOCHBADLOBYMAK
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Q: Define Salts and how are they formed?

Salts:

Definition Salt is an **ionic compound** formed through the **electrostatic attraction** between oppositely charged ions (cations and anions).

Formation of Salts

From Acid and Base reaction:

Acid + Base → Salt + Water

Example:

- $HCl + NaOH \rightarrow NaCl + H_2O$

Process of Formation:

- Base provides positive ions (cations).
- Acid provides negative ions (anions).
- Cation and anion develop electrostatic attraction.
- As a result, salt is formed.

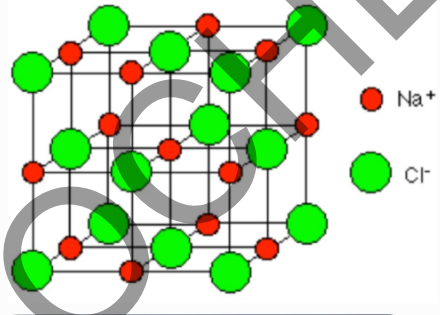
Q: What is the physical state add melting point at of solids at STP. And what is the Lattice structure of ionic compounds?

Physical State & Melting Point of Salts (at STP)

- STP means 0 °C temperature and 1 atm pressure.
- At standard temperature and pressure (STP), salts are solids.
- They are held together by strong electrostatic forces between oppositely charged ions.
- Because of these strong forces, salts have a high melting and boiling point.

Lattice Structure of Ionic Compounds

- Ionic compounds are crystalline solids with an ordered and repeating arrangement of ions.
- They form a 3-D crystal lattice structure.
- The lattice is held together by strong ionic (electrostatic) bonds between cations and anions.



Q: explain electrical conductivity, high melting and boiling points and hardness and bitterness of ironic compounds.?

1. Electrical Conductivity of Ionic Compounds

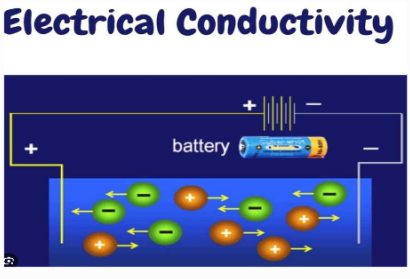
Electricity needs moving charged particles.

- In ionic compounds, the charges are the **ions**.
- When solid → ions are locked in the lattice → no movement → no conduction.
- When molten or dissolved → ions move freely → conduction happens.

So the movement of free ions is the only reason for their electrical conductivity.

- Example: Molten NaCl conducts, but solid NaCl does not.

- Ionic compounds conduct electricity because their ions are free to move in molten state or solution.
- Solid state: They do not conduct electricity because ions are fixed in the rigid lattice and cannot move.
- Example: Molten NaCl conducts, but solid NaCl does not.



2. High Melting and Boiling Points

- Ionic compounds have strong electrostatic forces between oppositely charged ions.
- A large amount of energy is required to break these bonds.
- Therefore, they have high melting and boiling points (e.g., NaCl melts at ~801°C).

3. Hardness and Brittleness

- Ionic compounds are hard due to their strong lattice structure.
- But they are also brittle: when force is applied, like charges (cation–cation or anion–anion) are forced close, causing repulsion → lattice breaks easily.
- Example salts: NaCl, KBr, $CaCO_3$.

Q: What is the preparation, separation, and purification of soluble salts?

It means making a salt, then removing extra substances, and finally getting pure salt crystals.

We use different methods depending on the reactants:

- Titration – using acid and alkali
- Excess metal – using acid and metal
- Excess insoluble base – using acid and base
- Excess insoluble carbonate – using acid and carbonate

Q: How Is a sodium chloride NACL prepared, using the titration of an acid and alkali.?

Preparation of Sodium Chloride (NaCl) by Titration

Sodium chloride is prepared by reacting hydrochloric acid (HCl) with sodium hydroxide (NaOH) through titration.Titration analyte is the solution of unknown concentration taken in the flask to be measured.

Reaction (Neutralization):

Acid + Base → Salt + Water

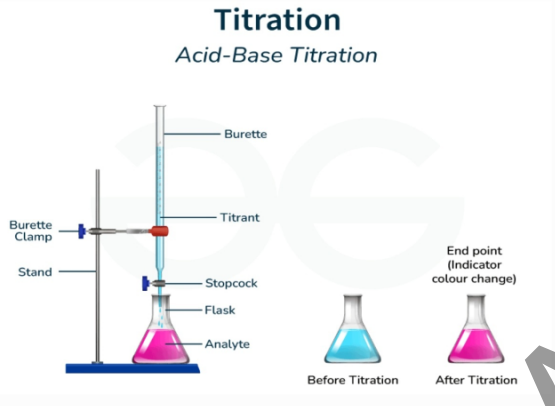
$$HCl + NaOH \rightarrow NaCl + H_2O$$

Method:

- Take HCl (acid) of unknown concentration in a conical flask.
- Add a few drops of phenolphthalein indicator (turns pink in base, colourless in acid).
- Fill burette with NaOH solution (base) of known concentration.
- Add NaOH slowly until the pink colour just disappears → neutralization point reached.
- (NaOH = titrant, HCl = analyte).
- (Titrant → the solution of known strength that we add from the burette).
- (Analyte → the solution of unknown strength that we measure in the flask).

Formation of Salt:

- The solution now contains NaCl (salt) + water.
- On evaporation/crystallization, pure NaCl crystals are obtained.



Q: How is zinc sulphate ($ZnSO_4$) prepared using excess metal?

Preparation of Zinc Sulphate by Excess Metal Method

Zinc sulphate can be prepared by reacting zinc metal (Zn) with dilute sulphuric acid (H_2SO_4).

Reaction:

Metal + Acid → Salt + Hydrogen gas

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$

Method:

- Take dilute sulphuric acid in a beaker.
- Add excess zinc metal pieces to the acid. Until (bubbles of hydrogen gas) will be observed.
- Continue until some zinc remains unreacted, showing that the acid has been fully used.
- Filter to remove extra zinc.
- Collect the filtrate
- Heat the solution, then allow to cool so that pure zinc sulphate crystals form.

Formation of Salt:

The final product is zinc sulphate ($ZnSO_4$) crystals along with hydrogen gas released during the reaction.

Q: How is copper(II) sulphate ($CuSO_4$) prepared using an excess insoluble base?

Preparation of Copper(II) Sulphate by Excess Insoluble Base Method

Copper(II) sulphate can be prepared by reacting copper(II) oxide (CuO) with dilute sulphuric acid (H_2SO_4).

Reaction:

Base (insoluble) + Acid → Salt + Water

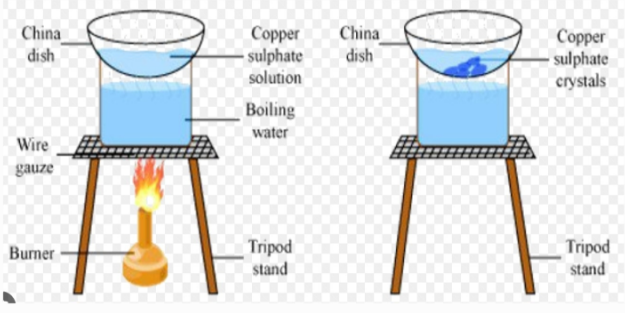
$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

Method (short & simple):

- Put dilute sulphuric acid in a beaker.
- Add excess copper(II) oxide; it reacts and forms copper sulphate + water.
- Keep adding until no more dissolves (acid is fully used).
- Filter to remove extra CuO.
- Heat the filtrate and let it cool → blue crystals of $CuSO_4$ form.

Formation of Salt:

The final product is copper(II) sulphate crystals and water.



Q: How is calcium chloride ($CaCl_2$) prepared using an excess insoluble carbonate?

Preparation of Calcium Chloride by Excess Insoluble Carbonate Method

Calcium chloride can be prepared by reacting calcium carbonate ($CaCO_3$) with dilute hydrochloric acid (HCl).

Reaction:

Carbonate (insoluble) + Acid → Salt + Carbon dioxide + Water

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 \uparrow + H_2O$

Method ;

- Put dilute hydrochloric acid in a beaker.
- Add excess calcium carbonate; bubbles of CO_2 gas appear.
- Keep adding until fizzing stops (acid fully used).
- Filter to remove extra $CaCO_3$.
- Heat the filtrate and let it cool → crystals of $CaCl_2$ form.

Formation of Salt:

The final product is calcium chloride crystals, with carbon dioxide gas released.

Q: What are soluble and insoluble salts? Give the main rules of solubility.

- **Soluble salts** dissolve in water at room temperature.
- **Insoluble salts** do not dissolve in water at room temperature.

Rules of Solubility :

- All sodium, potassium, and ammonium salts are soluble.
- All nitrates are soluble.
- Most chlorides are soluble, except $PbCl_2$ and AgCl.
- Most carbonates are insoluble, except sodium, potassium, ammonium carbonates.
- Most sulphates are soluble, except $PbSO_4$, $BaSO_4$, $CaSO_4$ (slightly soluble).

SLO Questions – Unit 6

1) Acid + Base → Salt + Water

- $HCl + NaOH \rightarrow NaCl + H_2O$

Hydrochloric acid + Sodium hydroxide → Sodium chloride + Water

- $H_2SO_4 + KOH \rightarrow K_2SO_4 + 2H_2O$

Sulfuric acid + Potassium hydroxide → Potassium sulfate + Water

- $H_2SO_4 + Mg(OH)_2 \rightarrow MgSO_4 + 2H_2O$

Sulfuric acid + Magnesium hydroxide → Magnesium sulfate + Water

- $HCl + KOH \rightarrow KCl + H_2O$

Hydrochloric acid + Potassium hydroxide → Potassium chloride + Water

2) Metal + Acid → Salt + H_2

- $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Zinc + Sulfuric acid → Zinc sulfate + Hydrogen

- $Cu + H_2SO_4 \rightarrow CuSO_4 + H_2$

Copper + Sulfuric acid → Copper sulfate + Hydrogen

- $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

Iron + Sulfuric acid → Iron sulfate + Hydrogen

- $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

Zinc + Hydrochloric acid → Zinc chloride + Hydrogen

- $Cu + 2HCl \rightarrow CuCl_2 + H_2$

Copper + Hydrochloric acid → Copper chloride + Hydrogen

3) Metal Oxide + Acid → Salt + Water

- $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

Copper(II) oxide + Sulfuric acid → Copper sulfate + Water

- $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

Magnesium oxide + Hydrochloric acid → Magnesium chloride + Water

- $CaO + H_2SO_4 \rightarrow CaSO_4 + H_2O$

Calcium oxide + Sulfuric acid → Calcium sulfate + Water

- $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$

Magnesium oxide + Sulfuric acid → Magnesium sulfate + Water

4) Metal Carbonate + Acid → Salt + CO_2 + H_2O

- $CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Calcium carbonate + Hydrochloric acid → Calcium chloride + Carbon dioxide + Water

- $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + CO_2 + H_2O$

Magnesium carbonate + Sulfuric acid → Magnesium sulfate + Carbon dioxide + Water

- $MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$

Magnesium carbonate + Hydrochloric acid → Magnesium chloride + Carbon dioxide + Water

- $BaCO_3 + 2H_3PO_4 \rightarrow Ba_3(PO_4)_2 + 3CO_2 + 3H_2O$

Barium carbonate + Phosphoric acid → Barium phosphate + Carbon dioxide + Water

- $3BaCO_3 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 3CO_2 + 3H_2O$

Calcium carbonate + Phosphoric acid → Calcium phosphate + Carbon dioxide + Water

Difference between Soluble and Insoluble Salts

Aspect	Soluble Salts	Insoluble Salts
Definition	Dissolve in water	Do not dissolve in water
Uses	Fertilizers (KNO_3, $(NH_4)_2SO_4$), table salt (NaCl)	Construction ($CaCO_3$), pigments ($PbSO_4$)
Conductivity	Form electrolytes → conduct electricity in solution	Poor conductivity in water
Examples	NaCl, KNO_3	AgCl, $BaSO_4$

CONCEPT ASSESSMENT: 6.1

Q: Which of the following salts are insoluble: Silver chloride, Barium sulphate, Sodium nitrate, Ammonium carbonate, and Lead nitrate?

Answer:

- Insoluble salts: Silver chloride (AgCl), Barium sulphate ($BaSO_4$)
- The others (Sodium nitrate, Ammonium carbonate, Lead nitrate) are soluble.

So, the correct insoluble salts are AgCl and $BaSO_4$.

Exercise: Multiple Choice Questions (MCQs)

i. Which of the following salts is soluble in water?

a) Silver chloride (AgCl)
b) Lead carbonate ($PbCO_3$)
c) Sodium nitrate ($NaNO_3$) ✅
d) Lead chloride ($PbCl_2$)

ii. What type of ions do bases provide for salt formation?

a) Negative ions
b) Positive ions ✅
c) Neutral molecules
d) Complex ions

iii. What is the primary reason salts have high melting points?
a) Weak van der Waals forces
b) Strong electrostatic forces ✅
c) Hydrogen bonding
d) Covalent bonding

iv. In what state are salts generally good conductors of electricity?
a) Solid
b) Gas
c) Molten ✅
d) Powdered

v. Which of the following is a general solubility rule for chlorides?
a) All chlorides are insoluble
b) All chlorides are soluble
c) Chlorides are soluble except lead and silver chlorides ✅
d) Chlorides are soluble except sodium and potassium chlorides

vi. What happens to the ions in a salt when it is dissolved in water?
a) They form a gas
b) They become fixed in place
c) They become mobile ✅
d) They form a solid

vii. Which of the following methods is used to prepare a salt by titration?
a) Acid + Metal
b) Acid + Insoluble Base
c) Acid + Insoluble Carbonate
d) Acid + Alkali ✅

viii. Which of these salts is insoluble in water?
a) Potassium carbonate (K_2CO_3)
b) Ammonium nitrate (NH_4NO_3)
c) Calcium chloride ($CaCl_2$)
d) Lead chloride ($PbCl_2$) ✅

ix. What is formed when an acid reacts with an excess of an insoluble base?
a) Salt and hydrogen
b) Salt and water ✅
c) Salt and carbon dioxide
d) Salt and oxygen

x. Which ion is commonly found in soluble nitrates?
a) NH_4^+
b) NO_3^- ✅
c) Cl^-
d) CO_3^{2-}

Exercise Short Questions Answers

i) What is a salt (in terms of ions)?

Definition:

- An ionic compound made of positive ions (cations) and negative ions (anions).
- Formed when H^+ of an acid is replaced by a metal or NH_4^+ (neutralization).
- Overall electrically neutral (total + = total −).

Example: NaCl (Sodium chloride), KNO_3 (Potassium nitrate).

ii) Why do salts have high melting points?

Reason:

- Ions are held by strong electrostatic forces (ionic bonds).
- Arranged in a tight giant lattice → hard to separate.
- Needs a lot of heat energy to overcome these forces.

Example: NaCl melts at ~801 °C.

iii) Describe the lattice of an ionic compound.

Definition:

- Regular 3-D repeating arrangement of ions.
- Alternating cations and anions (each surrounded by opposite charges).
- Forms a rigid, brittle, giant ionic structure.

Example: NaCl lattice structure (cube arrangement).

iv) How do salts conduct electricity when molten or in solution?

Reason:

- In solid state → ions are fixed → no conduction.
- In molten/aqueous state → ions separate and move freely.
- Moving ions carry charge, so the liquid conducts.

Example: Molten NaCl conducts, but solid NaCl does not.

v) List three salts generally soluble in water.

Examples:

- Sodium chloride (NaCl).
- Potassium nitrate (KNO_3).
- Ammonium sulphate ($(NH_4)_2SO_4$).

vi) What happens to ions when an ionic compound dissolves?

Process:

- The crystal dissociates into free ions.
- Water molecules surround (hydrate) the ions.
- Ions become mobile, so the solution can conduct.

Example: $NaCl \rightarrow Na^+ (aq) + Cl^- (aq)$.

vii) Are most carbonates soluble in water?

Definition:

- No, most carbonates are insoluble.

Exceptions:

- Sodium carbonate (Na_2CO_3).
- Potassium carbonate (K_2CO_3).
- Ammonium carbonate ($(NH_4)_2CO_3$).

Reason:

- Most carbonates have very strong lattice forces that water cannot break.

viii) How to prepare a soluble salt using excess insoluble base?

Method (e.g., $CuO + H_2SO_4$):

- Add insoluble base to dilute acid until some base remains (acid all used up).
- Filter off unreacted base → get salt solution.
- Evaporate and cool → salt crystals form.

Example: $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ (Copper sulfate).

ix) Which salts are typically soluble (general rules)?

Rules:

- All nitrates (NO_3^-) are soluble.
- All sodium, potassium, and ammonium salts are soluble.
- Most chlorides are soluble except AgCl and $PbCl_2$.
- Most sulfates are soluble except $PbSO_4$ and $BaSO_4$ ($CaSO_4$ only slightly soluble).

Examples:

- Soluble: $NaNO_3$, KCl, $(NH_4)_2SO_4$.
- Insoluble: AgCl, $PbSO_4$, $BaSO_4$.