

## Chapter 6 - Salts

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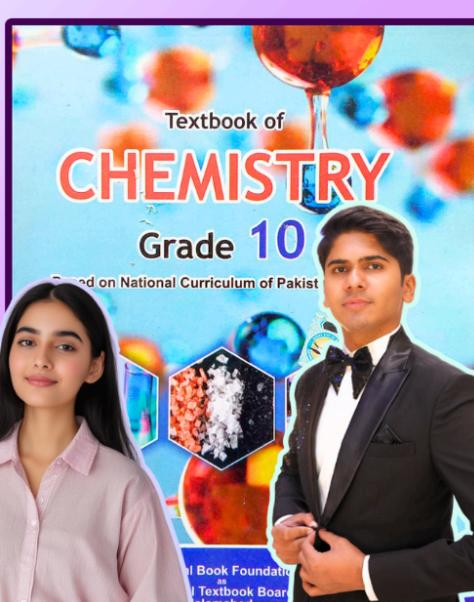
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### 6.1. Salts

Salts are ionic compounds.

**Formation:** They are formed by the electrostatic attraction between oppositely charged ions known as cations (positive ions) and anions (negative ions).

**Origin of Ions:** These ions originate from acids and bases. **Bases** provide the **cations (positive ions)**. **Acids** provide the **anions (negative ions)**. The combination of these positive and negative ions results in a neutral compound—the salt.

# Formation of Common Salt



## Example Reaction:

**Acid:** Hydrochloric Acid (HCl) → provides Cl<sup>-</sup> anions.

**Base:** Sodium Hydroxide (NaOH) → provides Na<sup>+</sup> cations.

**Salt Formed:** Sodium Chloride (NaCl), common table salt.

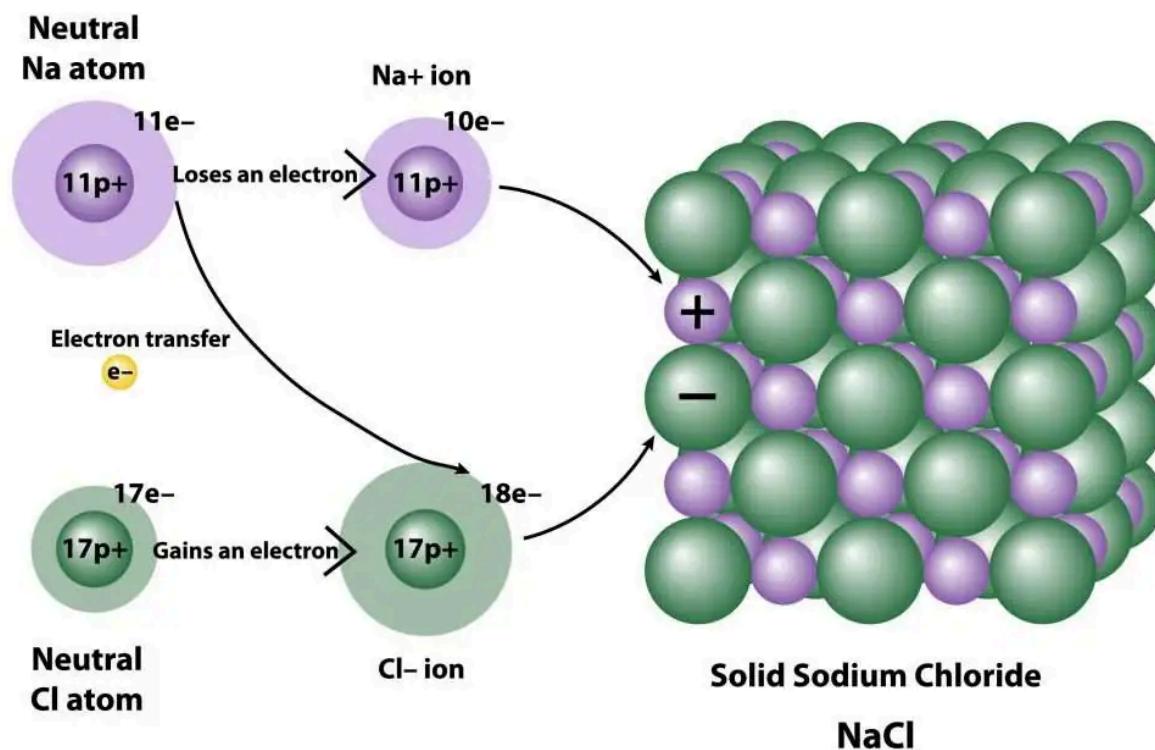


The driving force behind the formation is the **electrostatic attractive force** between the oppositely charged ions, leading to **ionic bonding**.

## 6.2. Physical State and Melting Points

### 6.2.1. Solids at STP

**Physical State:** At Standard Temperature and Pressure (STP), salts are usually **solids**. This solid state is due to the **strong electrostatic forces of attraction** between the oppositely charged ions.

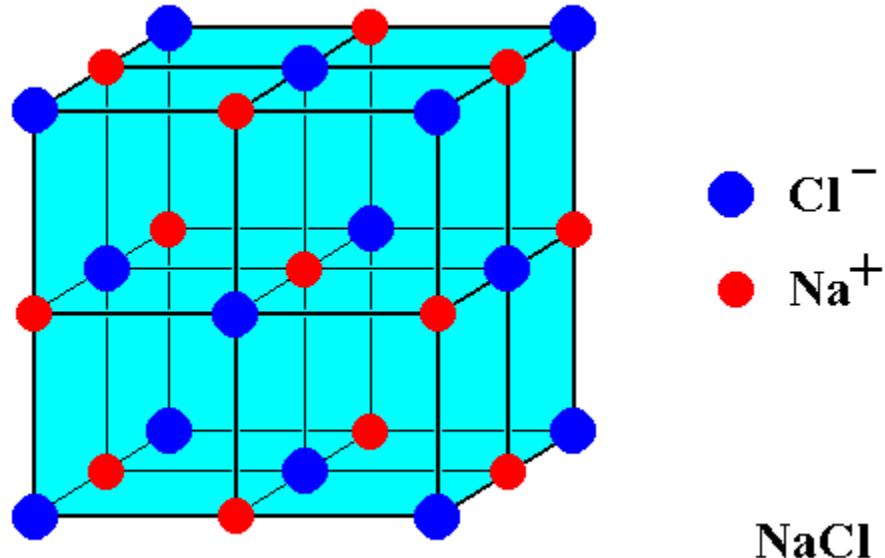


**Structure:** These forces create a **tightly bound lattice structure**.

**Melting Points:** A significant amount of energy is required to overcome these strong forces. This is why ionic compounds have **high melting points**.

### 6.2.2. Lattice Structure of Ionic Compounds

Under normal conditions, ionic compounds are solids with a well-defined structure. This is due to the nature of **ionic bonds**, which are the electrostatic attractions between cations and anions.



**Crystal Lattice:** The ions are arranged in a repeating, three-dimensional pattern called a **crystal lattice**.

**Purpose of the Lattice:** This specific arrangement is crucial because it **maximizes the electrostatic attractions** between oppositely charged ions while simultaneously **minimizing the repulsions** between like-charged ions. This leads to a very stable and strong structure.

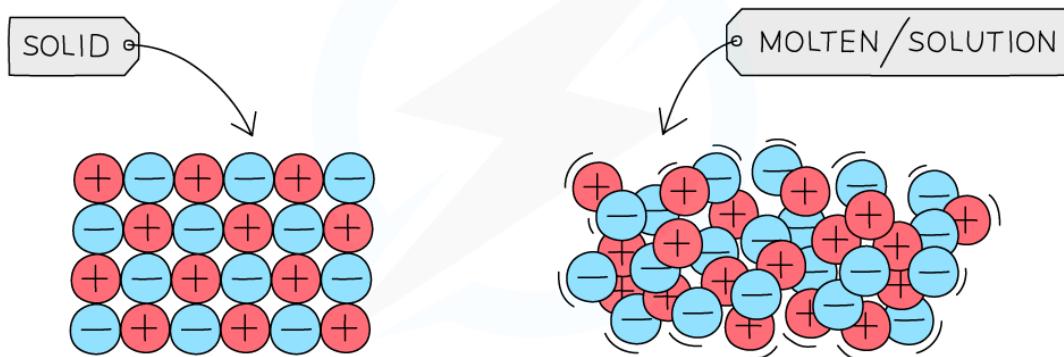
### 6.2.3. Electrical Conductivity

**Conducting States:** Molten (liquid) and aqueous (dissolved in water) solutions of salts are **good conductors of electricity**.

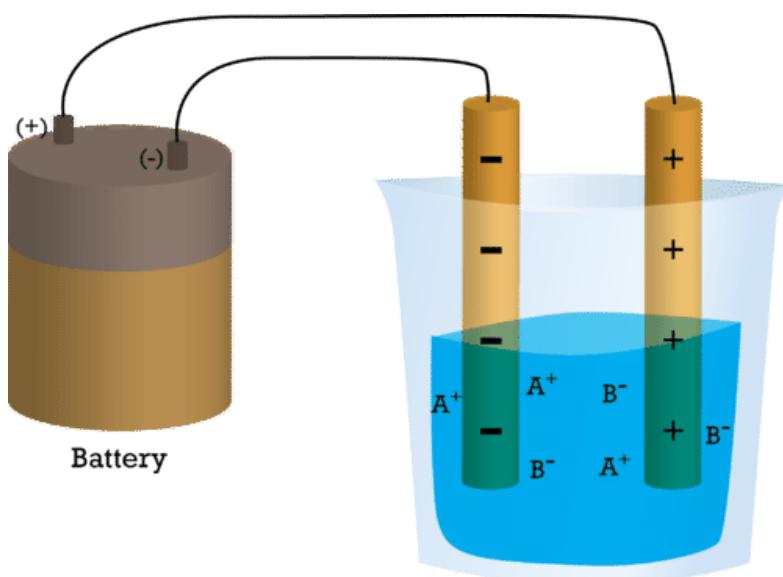
**Reason for Conductivity:** The presence of **free-moving ions** that can carry electrical charge.

**Solid State: Does it conduct? No.** In a solid salt, the ions are locked into a **rigid lattice structure** by strong ionic bonds. They are **not free to move**, and without mobile charge carriers, electricity cannot be conducted.

### ELECTRICAL CONDUCTIVITY OF IONIC COMPOUNDS



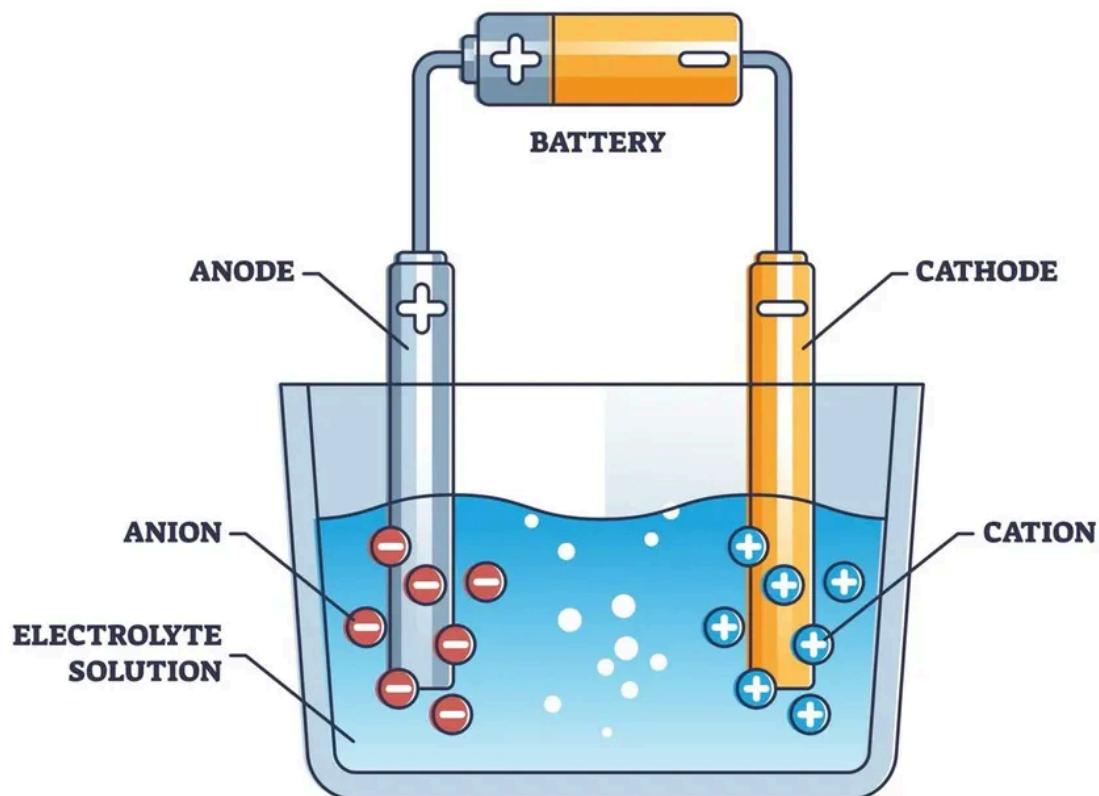
**Molten (Liquid) State: Does it conduct? Yes.** High temperature breaks down the rigid lattice. The ions are now **free to move**. The free-moving **cations (positive ions)** and **anions (negative ions)** can carry charge. When an electric potential is applied, they move, allowing conduction.



**Aqueous Solution (Dissolved in Water):** Does it conduct? Yes The polar water molecules cause the salt to **dissociate** into its constituent ions (e.g.,  $\text{NaCl} \rightarrow \text{Na}^+$  and  $\text{Cl}^-$ ). These ions are surrounded by water molecules but are **free to move** within the solution. When an electric potential is applied:

- Cations ( $\text{Na}^+$ ) migrate towards the **cathode (negative electrode)**.
- Anions ( $\text{Cl}^-$ ) migrate towards the **anode (positive electrode)**.

This movement of ions completes the electrical circuit, allowing conduction.



In both molten and aqueous states, the key condition is met: the **ionic lattice is disrupted, liberating the ions**. These free-moving ions, under the influence of an electric field, enable electrical conduction.

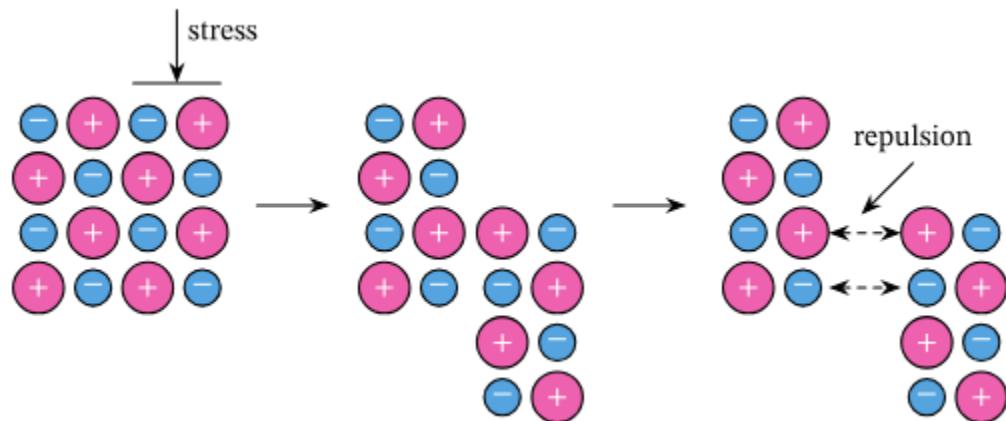
#### 6.2.4. High Melting Points and Boiling Points

The **strong electrostatic forces** holding the ionic lattice together require a **significant amount of energy** to break. This high energy requirement translates directly into **high melting points and high boiling points** for ionic compounds.

**State at STP:** Reinforces that they are solid at STP due to these strong forces.

#### 6.2.5. Hardness and Brittleness

**Hardness:** The stiffness and strength of the lattice structure make ionic compounds **hard**.



**Brittleness:** However, they are also **brittle** and tend to shatter. When a force is applied, it can shift the layers of the lattice. This movement can bring ions of the **same charge (like-charged ions) close together**. When this happens, they **repel each other** strongly. This repulsive force causes the lattice to split and break apart.

#### Examples of Ionic Compounds:

- o Sodium Chloride (NaCl)
- o Potassium Bromide (KBr)
- o Calcium Carbonate (CaCO<sub>3</sub>)
- o Magnesium Sulphate (MgSO<sub>4</sub>)

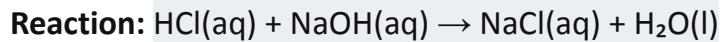
Each of these compounds forms a distinct **crystal lattice** that determines their structure and many of their physical properties, such as hardness and brittleness.

### 6.3. Preparation, Separation, and Purification of Soluble Salts

Soluble salts can be prepared through various methods, depending on the nature of the reactants involved. The general process often involves a reaction to form the salt, followed by separation and purification via crystallization.

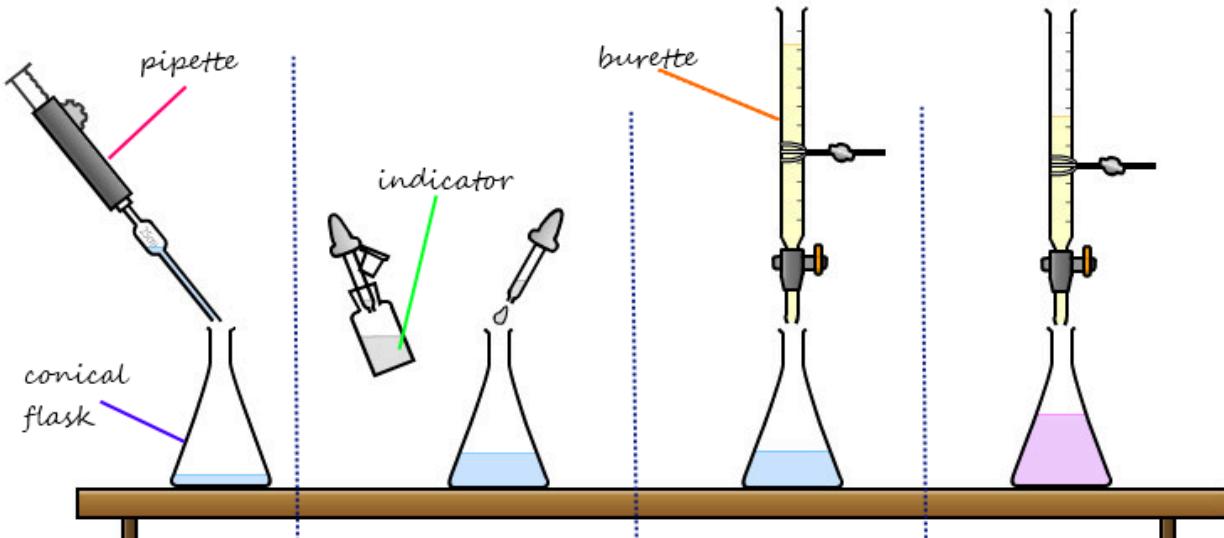
#### Preparation by Titration (Acid and Alkali)

Preparation of Sodium Chloride (NaCl).



**Principle:** This method is used when both the acid and the base are **soluble**. Titration ensures that the reactants are mixed in exactly the correct proportions to achieve complete neutralization without any excess.

**Materials Required:** Dilute HCl, Dilute NaOH, Phenolphthalein indicator, Burette, Pipette, Conical flask, Funnel, Evaporating dish, Burette stand, Tripod stand, Heating source.



### **Procedure:**

1. **Measurement:** Precisely measure 25 cm<sup>3</sup> of HCl solution using a pipette and place it in a conical flask.
2. **Indicator:** Add a few drops of **phenolphthalein** indicator to the acid. It will be colourless.
3. **Filling Burette:** Fill a burette with the NaOH solution.
4. **Titration:** Slowly add NaOH from the burette to the HCl in the flask while swirling continuously. The endpoint is reached when a **light pink colour** persists, indicating the solution is neutral.

### **Separation & Purification:**

1. **Crystallization:** The resulting solution contains dissolved NaCl.
  - Gently **evaporate** some of the water to form a hot, saturated solution. ▪ **Allow the solution to cool.** As it cools, the solubility of NaCl decreases, and crystals form.
2. **Collection:** Filter the mixture to collect the pure NaCl crystals.

### **Preparation with Excess Metal**

Preparation of Zinc Sulphate (ZnSO<sub>4</sub>).



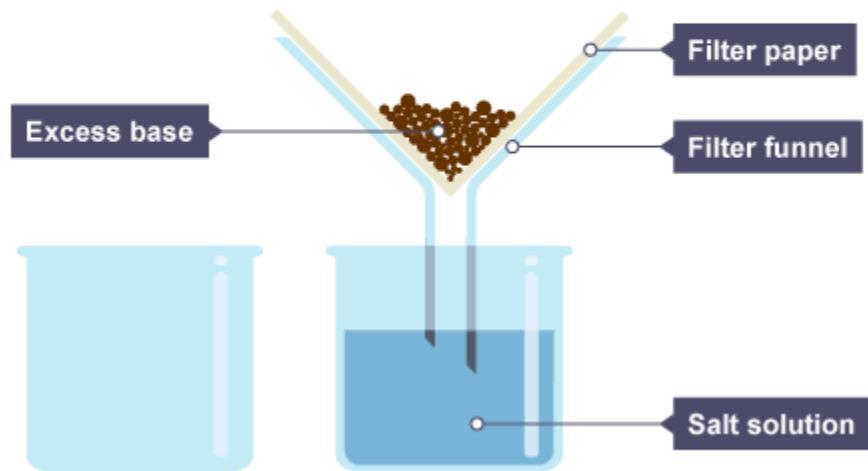
**Principle:** This method is used to prepare a salt from a reaction between an acid and a metal. The **metal is used in excess** to ensure **all of the acid is completely reacted**.

**Observation:** Bubbles of **hydrogen gas** are produced.

**Materials Required:** Zinc metal, Dilute sulphuric acid, Beaker, Glass rod, Filter funnel and paper, Evaporating dish, Tripod stand, Wire gauze, Heating source.

### Procedure:

- 1. Measurement:** Measure 25 cm<sup>3</sup> of dilute sulphuric acid into a beaker.
- 2. Reaction:** Add **excess zinc metal** to the acid. Bubbles of hydrogen gas will appear.
- 3. Confirm Completion:** Continue adding zinc until some solid remains unreacted, confirming that no more acid is left to react.



### Separation & Purification:

- 1. Filtration:** Filter the mixture to **remove the unreacted excess zinc metal**.

#### Crystallization:

- Evaporate the **filtrate** (the solution that passes through the filter, containing ZnSO<sub>4</sub>) to obtain a hot, saturated solution.
- Allow it to cool and crystallize.

**3. Collection:** Filter the mixture to collect the pure  $\text{ZnSO}_4$  crystals.

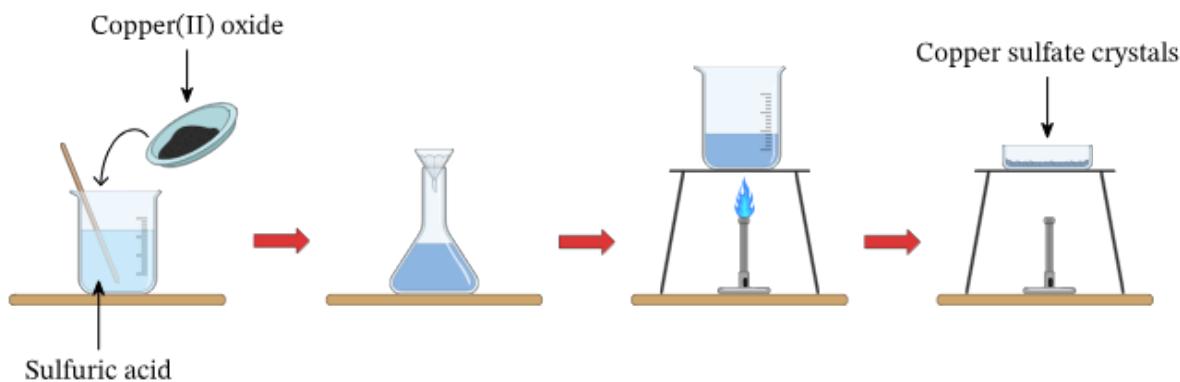
### Preparation with Excess Insoluble Base

Preparation of Copper(II) Sulphate ( $\text{CuSO}_4$ ).



**Principle:** This method is used when the base is **insoluble** in water. The **insoluble base (Copper(II) oxide)** is used in excess to ensure all the acid is used up.

**Materials Required:** Copper(II) oxide, Dilute sulphuric acid, Beaker, Glass rod, Filter funnel and paper, Evaporating dish, Tripod stand, Wire gauze, Heating source.



### Procedure:

**1. Measurement:** Measure  $25 \text{ cm}^3$  of dilute sulphuric acid into a beaker.

**Reaction:** Add **excess copper(II) oxide** to the acid.

**3. Confirm Completion:** Heat the mixture and continue adding  $\text{CuO}$  until no more solid dissolves. The presence of undissolved solid indicates that all the acid has reacted.

**4. Filtration:** Filter the mixture to remove the unreacted excess copper(II) oxide ( $\text{CuO}$ ).

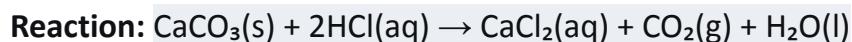
**5. Crystallization:**

- Evaporate the **filtrate** (the solution that passes through the filter, containing  $\text{CuSO}_4$ ) to obtain a hot, saturated solution.
- Allow the solution to cool and crystallize.

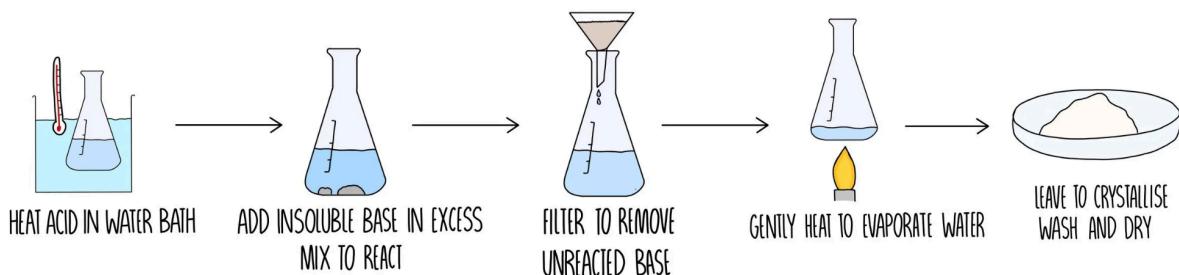
**6. Collection:** Filter the mixture to collect the pure  $\text{CuSO}_4$  crystals.

### Preparation with Excess Insoluble Carbonate

Preparation of Calcium Chloride ( $\text{CaCl}_2$ ).



**Principle:** This method is used to prepare a salt from a reaction between an acid and a carbonate. The **insoluble carbonate is used in excess** to ensure **all of the acid is completely reacted**.



**Observation:** Effervescence (bubbles) are observed as carbon dioxide gas is evolved.

**Materials Required:** Calcium carbonate (limestone), Dilute hydrochloric acid, Beaker, Glass rod, Filter funnel and paper, Evaporating dish, Tripod stand, Wire gauze, Heating source.

### **Procedure:**

1. **Measurement:** Measure 25 cm<sup>3</sup> of dilute hydrochloric acid into a beaker.
2. **Reaction:** Add **excess calcium carbonate (CaCO<sub>3</sub>)** to the acid. Effervescence will be observed due to CO<sub>2</sub> production.
3. **Confirm Completion:** Continue adding CaCO<sub>3</sub> until **no more effervescence is observed**. This indicates that all the acid has been used up.

### **Separation & Purification:**

1. **Filtration:** Filter the mixture to **remove the unreacted excess calcium carbonate (CaCO<sub>3</sub>)**.
2. **Crystallization:**
  - Evaporate the **filtrate** (the solution that passes through the filter, containing CaCl<sub>2</sub>) to obtain a hot, saturated solution.
  - Allow it to cool and crystallize.
3. **Collection:** Filter the mixture to collect the pure CaCl<sub>2</sub> crystals.

## **6.4. Soluble and Insoluble Salts**

**Soluble Salts:** Salts that **dissolve** in water at room temperature. **Insoluble Salts:** Salts that **do not dissolve** in water at room temperature.

**General Solubility Rules:** These rules help predict whether a given ionic compound will dissolve in water. They are essential for understanding reactions in aqueous solutions.



**All Sodium ( $\text{Na}^+$ ), Potassium ( $\text{K}^+$ ), and Ammonium ( $\text{NH}_4^+$ ) salts are soluble.**

**Examples:**

- Sodium Chloride ( $\text{NaCl}$ )
- Sodium Sulphate ( $\text{Na}_2\text{SO}_4$ )
- Potassium Chloride ( $\text{KCl}$ )
- Potassium Bromide ( $\text{KBr}$ )
- Ammonium Chloride ( $\text{NH}_4\text{Cl}$ )
- Ammonium Sulphate ( $(\text{NH}_4)_2\text{SO}_4$ )

**All Nitrate ( $\text{NO}_3^-$ ) salts are soluble.**

**Examples:**

- Potassium Nitrate ( $\text{KNO}_3$ )
- Calcium Nitrate ( $\text{Ca}(\text{NO}_3)_2$ )

**Most Chloride ( $\text{Cl}^-$ ) salts are soluble.**

**Exceptions: Lead(II) Chloride ( $\text{PbCl}_2$ ) and Silver Chloride ( $\text{AgCl}$ ) are insoluble.**

**Most Carbonate ( $\text{CO}_3^{2-}$ ) salts are insoluble.**

**Exceptions: Sodium, Potassium, and Ammonium carbonates are soluble.**

**Examples of Insoluble Carbonates:**

- Calcium Carbonate ( $\text{CaCO}_3$ )
- Magnesium Carbonate ( $\text{MgCO}_3$ )

**Most Sulphate ( $\text{SO}_4^{2-}$ ) salts are soluble.**

**Exceptions: Lead(II) Sulphate ( $\text{PbSO}_4$ ), Barium Sulphate ( $\text{BaSO}_4$ ), and Calcium Sulphate ( $\text{CaSO}_4$ ) are insoluble.**

To summarize:

Soluble salts	Insoluble salts
All sodium salts	
All potassium salts	
All ammonium salts	
All nitrates	
All chlorides	Silver chloride Lead(II) chloride
Except	
All sulphates	Barium sulphate Lead(II) sulphate Calcium sulphate (sparingly soluble)
Except	
Sodium carbonate Potassium carbonate Ammonium carbonate	All carbonates
Except	



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