**SALTS**

A salt is a compound formed when the hydrogen ions in an acid are replaced by metal ions or by ammonium ions. Compounds in which the $H^+$ ions in an acid have been replaced by ammonium ions ($NH_4^+$) are called ammonium salts.
NOTE:
• An anion is a negatively charged ion. Hence:-
• Hydrochloric acid gives chlorides. E.g. sodium chloride, ammonium chloride.
• Nitric acid gives nitrates. E.g. barium nitrate, copper nitrate.
• Sulphuric acid gives sulphates. E.g. silver sulphate, iron (ii) sulphate.
• Phosphoric acid gives phosphates. E.g. sodium phosphate, ammonium phosphate.
• Each acid gives rise to a series of salts named by the ANION which they contain.
Some acids can donate more than one replaceable hydrogen ion (proton)

E.g. sulphuric acid ($H_2SO_4$) and phosphoric acid ($H_3PO_4$).

Sulphuric acid can give rise to two series of salts and phosphoric acid to three series of salts depending on how many protons are replaced.

i.e. (a) $H_2SO_4 \rightarrow NaHSO_4 \rightarrow Na_2SO_4$

(b) $H_3PO_4 \rightarrow NaH_2PO_4 \rightarrow Na_2HPO_4 \rightarrow Na_3PO_4$
• **NORMAL SALTS** - These are salts which do not contain replaceable hydrogen. E.g. sodium chloride (NaCl).
• **ACIDIC SALTS** - These are salts which contain replaceable hydrogen. This is because when the hydrogen is replaced they are acting like an acid. E.g. sodium hydrogen sulphate (NaHSO₄).

• The method chosen to prepare a salt depends on its solubility.
• Solubility depends on the combination of positive and negative ions.
<table>
<thead>
<tr>
<th>SOLUBLE</th>
<th>INSOLUBLE</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 All nitrates</td>
<td>All carbonates except for sodium, potassium and ammonium carbonates</td>
</tr>
<tr>
<td>2 All chlorides except for silver chloride and lead (ii) chloride</td>
<td>All sulphides except for sodium, potassium and ammonium sulphides</td>
</tr>
<tr>
<td>3 All sulphates except for calcium, barium and lead (ii) sulphates</td>
<td>All oxides except for sodium potassium and ammonium oxides</td>
</tr>
<tr>
<td>4 All sodium, potassium (even other group I elements), ammonium and</td>
<td>All hydroxides except for sodium, potassium (group I), ammonium and</td>
</tr>
<tr>
<td>nitrate compounds are soluble</td>
<td>calcium hydroxides</td>
</tr>
</tbody>
</table>
NOTE:
• Insoluble salts are prepared by PRECIPITATION.
• Soluble salts are prepared by FILTRATION and CRYSTALLISATION method or by TITRATTION.

PREPARATION OF INSOLUBLE SALTS
• Insoluble salts are prepared by mixing solutions containing their positive and negative ions using the method of precipitation (Double decomposition).
• The reactants are chosen so that on exchanging ions the unwanted product is still soluble but the given insoluble salt will form as a precipitate.
• I.e. soluble salt + soluble salt → soluble salt + insoluble salt (precipitate).
NOTE:
• A precipitate is an insoluble solid formed when a chemical reaction occurs between two dissolved ionic substances.

CHOOSING REACTANT SOLUTIONS
• The other one must contain the negative ion of the insoluble salt required.

PROCEDURE OF PRECIPITATION
• Dissolve each reactant separately in water.
• Mix chemically equivalent quantities of the reactant solutions.
• Filter the solution and wash the precipitate in warm distilled water.
• One must contain the positive ion of the insoluble salt required.
• Dry the solid salt that was produced in an oven (105°C).

**EXAMPLES**

• PREPARATION OF BARIUM SULPHATE (BaSO₄)
• We can use barium nitrate and sodium sulphate.

**REACTANT IONS**

<table>
<thead>
<tr>
<th>Ion</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ba²⁺</td>
<td>NO₃⁻ → NaNO₃ (soluble)</td>
</tr>
<tr>
<td>Na⁺</td>
<td>SO₄²⁻ → BaSO₄ (Insoluble)</td>
</tr>
</tbody>
</table>
**PREPARATION OF SILVER CHLORIDE** ($\text{AgCl}$)

- We can use silver nitrate and sodium chloride.

### Reactant Ions

<table>
<thead>
<tr>
<th>Ion</th>
<th>Formula</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Ag}^+$</td>
<td>$\text{NO}_3^-$</td>
<td>$\text{NaNO}_3$ (soluble)</td>
</tr>
<tr>
<td>$\text{Na}^+$</td>
<td>$\text{Cl}^-$</td>
<td>$\text{AgCl}$ (insoluble)</td>
</tr>
</tbody>
</table>

**Note:**

- The easiest salt solution containing the positive ions is the metal nitrate since all nitrates are soluble.
- For a solution containing the negative ions, we can use the sodium salt as all sodium salts are soluble.
PREPARATION OF LEAD (II) IODIDE (PbI₂)

I.e. \( \text{Pb} (\text{NO}_3)_2 (aq) + 2\text{NaI} (aq) \rightarrow \text{PbI}_2 (s) + 2\text{NaNO}_3 (aq) \)

IONICALLY

I.e. \( \text{Pb}^{2+} (aq) + 2\text{I}^- (aq) \rightarrow \text{PbI}_2 (s) \)
LEAD NITRATE SOLUTION

SODIUM IODIDE SOLUTION

MIX

Na⁺ and NO₃⁻ are SPECTATOR IONS

A PRECIPITATE OF LEAD IODIDE
NOTE:

• The reactants involved in a precipitation reaction must be in solution form because the ions must be able to move and interact with one another when the reactants are mixed.

• When the ions in the insoluble salt encounter each other, they get attracted and form a solid that will sink to the bottom of the container and be collected as a PRECIPITATE.
PREPARATION OF SOLUBLE SALTS.

• Soluble salts are prepared using two methods:
  • Filtration and crystallisation method. i.e. neutralising an ACID with EXCESS INSOLUBLE REACTANT.
  • Titration. i.e. neutralising an ACID with the EXACT amount of ALKALI.
• **FILTRATION AND CRYSTALLISATION METHOD**

• This method is used when a suitable insoluble starting material has been found.

• The acid reacts with an excess of insoluble reactant that can be:
  - Metal
  - Base (insoluble)
  - Carbonate

• Therefore, to prepare a given salt, we need to choose the correct acid and a suitable insoluble reactant. (metal, oxide, hydroxide or carbonate).
**PROCEDURE**

- Neutralise an acid with an excess of the insoluble reactant.
- Filter off any unreacted reagent.
- Evaporate the solution to the crystallisation point.
- Cool to produce crystals of the salt.
- Filter, wash and dry the crystals before collection.
EXAMPLE

PREPARATION OF COPPER (II) SULPHATE

• Starting materials include copper (ii) oxide and dilute sulphuric acid.

• i.e. \( \text{CuO}_\text{(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}_\text{(l)} \)
STEP 1.
• An excess reactant ensures that all the acid has been used up.
• Don't boil. The is is just waemed to increase the rate of reaction between the reactants.
STEP 2.

MIXTURE: COPPER (II) SULPHATE SOLUTION AND EXCESS COPPER (II) OXIDE

RESIDUE: EXCESS COPPER (II) OXIDE

FILTRATE: COPPER (II) SULPHATE AND WATER
Dont evaporate all the water. The filtrate is heated until a thin crust of crystals form on the surface of the liquid.
NOTE:
• If a metal carbonate is used to prepare a salt using this method, there will be bubbles of carbon dioxide gas as the metal carbonate is added to the acid in step 1.
• When there is no more bubbles, it means that all the acid has been used up and we can proceed to the next step.
**TITRATION**

Soluble salts of ammonium and group I metals (sodium, potassium and lithium) are prepared using this method. This is because all their compounds are soluble (including metals themselves) and very reactive.

Group I metals are very reactive resulting in too violent reactions that we cannot use excess reactant.
NOTE:

• Titration means using the EXACT quantities of reactants for the reaction.
• This method is used when it is not possible to find a suitable insoluble starting material like a metal, metal oxide or a carbonate that can be easily filtered off at the end of the reaction.
**INDICATOR**

- In titration, an indicator is needed to show the endpoint of one reactant needed to exactly neutralise a given volume of the other reactant.
- A common indicator used in the laboratory is the screened methyl orange.
<table>
<thead>
<tr>
<th>ACIDIC SOLUTION</th>
<th>END POINT (NEUTRAL)</th>
<th>ALKALINE SOLUTION</th>
</tr>
</thead>
<tbody>
<tr>
<td>RED</td>
<td>GREY/COLOURLESS</td>
<td>GREEN</td>
</tr>
</tbody>
</table>

**PROCEDURE OF TITRATION**

- The most common procedure is to react the alkali solution with the dilute acid using a burette.
- An indicator is used to determine when the exact amount of reactant has been added.
EXAMPLE: PREPARATION OF SODIUM NITRATE.

• Starting materials include sodium hydroxide and dilute nitric acid.
• i.e. \( \text{NaOH}_{(aq)} + \text{HNO}_3(aq) \rightarrow \text{NaNO}_3(aq) + \text{H}_2\text{O}(l) \)

STEP 1.
• Place the soluble dilute acid in a conical flask.
• Add a few drops of indicator. E.g. methyl orange.
**STEP 2.**

- From a burette, slowly add the alkali solution. Ensure that the solution is mixed well.
- When the indicator begins to change colour, the reaction should be slowed to a drip.
- At this point, just enough acid is added to neutralise the alkali, all the alkali has reacted.
i.e.
STEP 3.

• Once the colour change is complete, the reaction is complete (END POINT).
• The burette should be turned off.
• The resulting mixture contains only sodium nitrate and water.

NOTE:

• From the titration result, we can know the exact volume of nitric acid needed to react with 25.0 cm$^3$ of sodium hydroxide.
• From the burette, volume of nitric acid needed will be: $V_n = V_2 - V_1$. 
**STEP 4.**

- Evaporate the sodium to crystallisation point.
- Cool to produce crystals of the salt.
- Filter, wash and dry the crystals.

**NOTE:**

- In a strict titration, a second titration must be carried out.
- The salt solution obtained in the first titration is thrown away because it is affected by the indicator.
- The second titration is done without the indicator. The exact volume of acid to be added is obtained from the first titration.