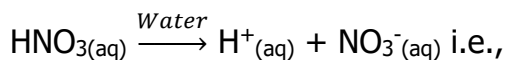


TOPIC 5: ACIDS, BASES AND SALTS

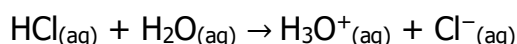
ACIDS

- Compounds that produce H^+ (protons) in solution as the only positively charged ions, e.g.,



- ✓ hydrochloric acid HCl gives H^+ and Cl^- ions in water (aqueous solution)
- ✓ sulphuric/sulphuric acid H_2SO_4 gives 2H^+ and SO_4^{2-} ions in water (aqueous solution)
- ✓ nitric acid HNO_3 gives H^+ and NO_3^- ions in water (aqueous solution)
- ✓ ethanoic acid CH_3COOH gives H^+ and CH_3COO^- ions (aqueous solution).

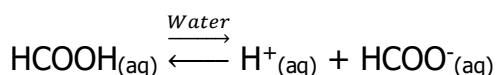
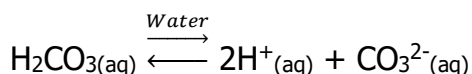
An acid can also be defined as *a proton donor*. For example, hydrochloric acid in aqueous solution donates a proton to the solvent, water, which results in the formation of the hydronium ion, H_3O^+ as shown in the equation below.



TYPES OF ACIDS

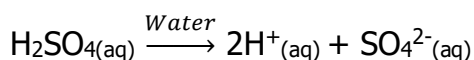
Depends on extent of ionisation.

1. **Weak acids:** is *an acid that ionise partially or partly when dissolved in water*, e.g.,



Generally, all **organic acids** such as (*Acetic acid, Propanoic acid, Butanoic acid, Citric acid, Lactic acid and Carbonic acid*) are weak acids.

2. **Strong acids:** is *an acid that ionises completely when dissolved in water*, e.g.,



Generally, all **mineral acids** such as (*Hydrochloric acid, Nitric acid and Sulphuric acid*) are strong acids.

Strength depends on extent of ionisation. If the given number of acid molecules are dissolved in water, then;

MORE ionises = Strong acid

FEW ionises = Weak acid

Dilute and concentrated acids

- **Concentration** - is *the amount of solute dissolved per unit volume of the solution*.

A concentrated acid – an acid containing greater number of dissolved H^+ per unit volume of the solution.

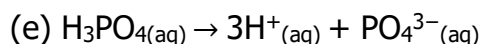
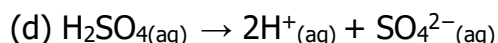
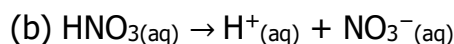
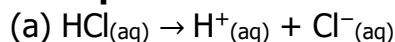
A dilute acid – an acid containing fewer number of dissolved H^+ per unit volume of the solution.

Ionization of acids

Ionization is *the process of forming ions*.

An ion is *a charged particle*. When acids dissolve in water, they **dissociate** to produce hydrogen ions, H^+ , which are responsible for the acidic properties.

Examples of ionization of acids



BASICITY OF AN ACID

Basicity is the number of H^+ produced per molecule of the acid when dissolved in water.

Monobasic acid

A monobasic acid *produces one hydrogen ion, H^+ when it dissolves in water*.

Dibasic acid

A dibasic acid *produces two hydrogen ions, $2H^+$ when it dissolves in water*.

Tribasic acid

A tribasic acid *produces three hydrogen ions, $3H^+$ when it dissolves in water*.

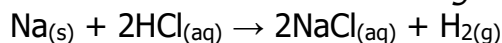
ACID	IONISATION	BASICITY
Nitric acid, HNO_3	$HNO_3 \rightarrow H^+ + NO_3^-$	1 – monobasic acid
Hydrochloric acid, HCl	$HCl \rightarrow H^+ + Cl^-$	1 – monobasic acid
Sulphuric acid, H_2SO_4	$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$	2 – dibasic acid
Carbonic acid, H_2CO_3	$H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$	2 – dibasic acid
Phosphoric acid, H_3PO_4	$H_3PO_4 \rightleftharpoons 3H^+ + PO_4^{3-}$	3 – tribasic acid

PHYSICAL PROPERTIES OF ACIDS

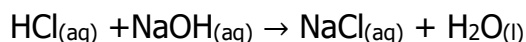
1. Acids have a sour taste.
2. Have a strong choking/unpleasant smell.
3. Some are corrosive and poisonous, i.e., they eat away surfaces of substances.
4. They are clear colourless solutions.
5. They have a pH value of less than 7
6. They are electrolytes, i.e., they conduct an electric current

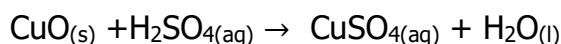
CHEMICAL PROPERTIES OF ACIDS

- Acids turn blue litmus paper red/pink.
- React with reactive metals to give SALT and HYDROGEN gas.



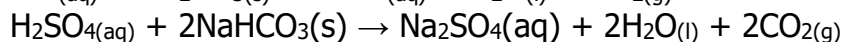
- React with Bases to form SALT and WATER only. This is called NEUTRALISATION.





Bases are metal oxides and metal hydroxides including ammonium hydroxide.

- React with metal carbonates and metal hydrogen carbonates to give SALT, WATER and CARBON DIOXIDE.



USES OF ACIDS

- Controlling pH of soil in agriculture if it is too alkaline
- Used in car batteries
- Used in pickling of metals
- Used in making of salts

Effects of Acids in the Environment

- Corrodes limestone buildings (acid rain)
- Increases soil acidity by lowering its pH
- Increases water acidity
- Destroys vegetation due to spillages caused by concentrated acids

BASES

- Compounds which neutralise acids
- Compounds which when reacted with acids produce SALT and WATER only.

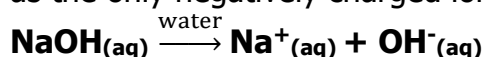
Note that all metal oxides and metal hydroxide are bases.

CLASSES OF BASES

i. Insoluble bases: Do not dissolve in water. All metal oxides and hydroxides are insoluble except oxides and hydroxides of Group I metals.

ii. Soluble bases: They dissolve in water hence are called ALKALIS (soluble base = alkali)

- An alkali is a soluble base that produces hydroxide (OH^-) ions in aqueous solution as the only negatively charged ions, e.g.,



A solution of a base in water is called an **alkaline solution**.

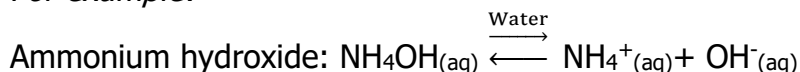
All alkalis are bases but not all bases are alkalis.

STRENGTH OF AN ALKALI (Types of Alkali)

The types of alkalis depend on the degree (extent) of ionisation.

1. **Weak alkalis:** ionise partially or partly in aqueous solution.

For example:



2. **Strong Alkalis:** Ionise completely in aqueous solution.

Examples include:



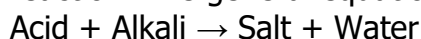
The strength of an acid or alkali can be shown by the **pH scale**.

Physical Properties of alkalis

1. Alkalis have a bitter taste
2. Alkalis are soapy and feel slippery
3. Alkalis turn red litmus paper blue
4. Have pH values greater than 7
5. Some are corrosive
6. they are electrolytes, i.e., they conduct an electric current

Chemical properties of Alkalis

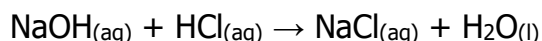
1. *Alkalis react with acids to produce salt and water only.* This is a neutralisation reaction. The general equation for this reaction is



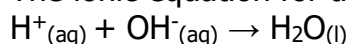
In neutralisation reaction, the hydrogen ions from an acid and the hydroxide ions from the alkali react to form water.

For example,

Sodium Hydroxide + hydrochloric acid \rightarrow sodium chloride + water



The ionic equation for this reaction is



This is the ionic equation for any neutralisation reaction.

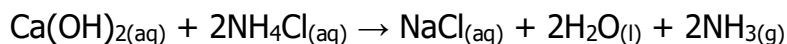
2. *Alkalis react with ammonium salts to produce ammonia gas, a salt and water.* The general equation for this reaction is

Alkali + Ammonia salt \rightarrow ammonia + salt + water

Ammonia can be recognised by its *characteristic pungent smell*. It also *turns moist/damp red litmus paper blue*.

For example, when calcium hydroxide is heated with ammonium chloride, ammonia gas is produced.

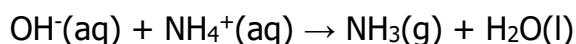
Calcium hydroxide + ammonium chloride \rightarrow calcium chloride + water + ammonia



When alkalis react with ammonia salts, hydroxide ions and ammonium ions react to produce ammonia gas.

The ionic equation for this reaction is

Hydroxide ions + ammonium ions \rightarrow ammonia + water

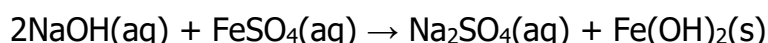
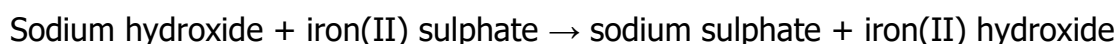


3. Precipitate insoluble metal hydroxides from their salt solution.

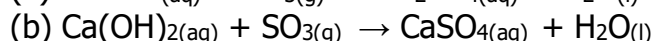
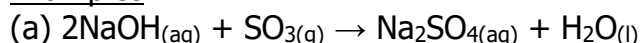
The general equation for this reaction is



For example, sodium hydroxide reacts with a solution of iron(II) sulphate to give a green precipitate of iron(II) hydroxide and sodium sulphate solution.



- Alkalis are therefore used to test for cations (Positive ions). They are also used as complexing reagent.

4. Alkalis react with acidic oxides to form a salt and water onlyExamples**USES OF ALKALIS**Agricultural Uses

- i. Slaked lime (Calcium Hydroxide) and quick lime (calcium oxide) are used to neutralise the acid in the soil.
- ii. Potash and lime are also used as fertilizers.

Industrial Uses

Name of base	Every day use
Ammonia	- Production of fertilizers (ammonium and nitrate salts). - Production of window cleaning agents
Aluminium hydroxide	- Manufacture other aluminium compound and to make gastric medicine (antacid)
Calcium hydroxide	- in making cement - in making limewater - in neutralization of the acidity of the soil - in bleaching powders - used in sewage treatment
Sodium hydroxide	- used in the manufacturing of soaps (saponification). - in making detergents - in making household cleaners
Magnesium hydroxide	- used as an antacid to relieve indigestion. - used as an antiperspirant armpit deodorant. - used in toothpaste to neutralise acid on teeth.

EFFECTS OF ALKALIS ON THE ENVIRONMENT

- Increases soil alkalinity by raising pH
- Affects certain crops that do not grow in too much alkaline soils
- Causes water alkalinity to increase.



pH Scale

pH stands for Hydrogen Potential.

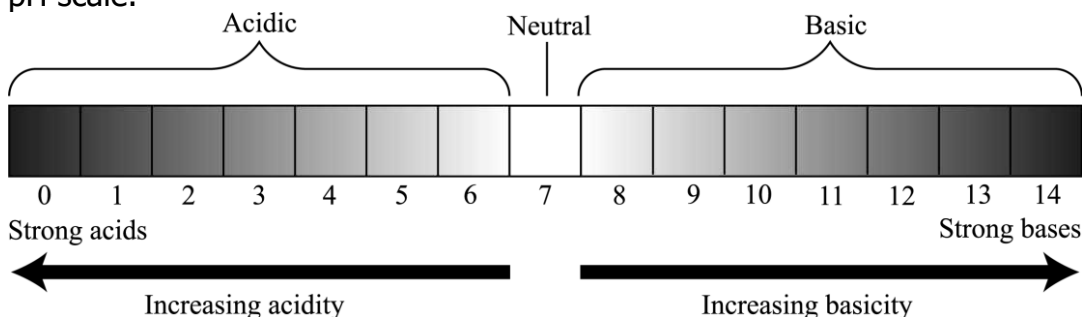
What is the pH scale?

The pH scale is the measure of the degree of acidity and alkalinity of a substance. The degree of acidity and alkalinity can be measured by the pH scale using the **universal indicator** and a **pH meter** for precise values.

Universal indicator is a mixture of pH indicators that show different colours at different pH values. It is available in *paper* and *solution form*. It undergoes various colour changes depending on the degree of acidity or alkalinity of a substance. The table below shows the colours of universal indicator at different pH values.

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Colour	Red			Orange		Yellow		Green	Blue			Violet			
Strength	Strong Acid							Neutral	Strong Base						
															

The pH scale consists of numbers ranging from 0 to 14. The figure below shows the pH scale.



- ✓ The smaller the pH value of an acid the stronger it is and vice versa, e.g., an acid with pH value of 3 is stronger than the one with pH value of 6.
- ✓ The smaller the pH value of an alkali the weaker it is and vice versa, e.g., an alkali with pH value of 8 is weaker than the one with pH value of 14.

How is the pH of a solution related to the concentration of hydrogen ions in the solution?

- ✓ The pH of a solution is calculated based on the number of hydrogen ions or hydroxide ions present in a solution.
- ✓ Acids with smaller pH value have higher concentration of hydrogen ions.
- ✓ When distilled water is added to an acid, the pH value of the acid increases towards seven. The solution becomes less acidic.

- ✓ Alkaline solutions with a larger pH value have a higher concentration of hydroxide ions.
- ✓ When distilled water is added to an alkaline solution, the pH value of the alkali decreases towards seven. The solution becomes less alkaline.

The Importance of pH Value in Agriculture

The pH value of the soil varies from 4 to 8 and it depends on the type of the soil. It is important to know the pH of the soil because it affects the growth and development of the plants. Most plants grow well when the soil pH is neutral or slightly acidic (around pH 6 or 7).

Plants will not grow in the soil that is too acidic. This can happen when too much fertilizer is added to the soil. Another reason why the soil becomes acidic is due to acid rain.

Chemicals are often added to the soil to adjust its pH. In areas where the soil is too acidic, it can be treated with bases such as slacked lime (calcium hydroxide) or quick lime (calcium oxide). This is known as 'liming' the soil. The bases will react with the acids in the soil and raise the pH so that the plants can grow healthy.

However, adding too much base will make the soil too alkaline and unsuitable for plant growth. In this case, a fertilizer such as ammonium sulphate is applied to improve the condition of the soil.

Indicators

Indicators are substances that show a particular colour in acidic or alkaline medium. The table below summarises the names of indicators and their colour changes in acidic and alkaline media.

Substance to be tested	Indicators			
	Blue litmus paper	Red litmus	Methyl orange	phenolphthalein
Acid	Blue to red	No effect	Orange to pink	Remains colourless
Alkalis/bases	No effect	Red to blue	Orange to yellow	Colourless to pink

SALTS

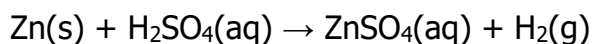
What is a SALT? *Is a compound formed when the hydrogen ions of an acid are fully or partially replaced by a metal or ammonium ions.* Or a compound made of positive metallic/ammonium ions and any negative ion of an acid.

Types of salts

a) Normal salt

This is *a salt produced when all the replaceable hydrogen ions of an acid are replaced by metal or ammonium ions.* Examples include; sodium chloride, NaCl; ammonium nitrate, NH_4NO_3 ; Magnesium sulphate, MgSO_4 ; lead (II) bromide, PbBr_2 and sodium phosphate, Na_3PO_4 .

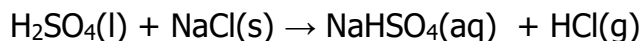
The reaction of Zinc granules reacting with dilute sulphuric acid forms a normal salt (Zinc sulphate).



b) Acid salts

Is a salt formed when the hydrogen ions in an acid are only partially replaced by metal or ammonium ions. These salts contain ionisable hydrogen and examples include: sodium hydrogen sulphate, NaHSO_4 ; calcium hydrogen carbonate, $\text{Ca}(\text{HCO}_3)_2$; and potassium hydrogen carbonate, KHCO_3 .

When Sodium chloride reacts with concentrated sulphuric acid an acid salt is formed as shown below.



c) Basic salts

Basic salt is formed when insufficient acid is present to neutralize the available base. E.g., basic zinc chloride, $(\text{Zn}(\text{OH})\text{Cl})$ and basic magnesium chloride $(\text{Mg}(\text{OH})\text{Cl})$.

Note that, all monobasic acids form normal salts while dibasic and tribasic acids form both normal and acid salts.

Note that a salt contains positive metal ion (or ammonium ion) and a negative non-metal ion.

The following are some examples of salts made from acid-base reactions.

Salt	Chemical formula	Possible reactants	
		Base	Acid
Calcium chloride	CaCl_2	CaO	HCl
Ammonium sulphate	$(\text{NH}_4)_2\text{SO}_4$	Aqueous NH_3	H_2SO_4
Copper(II) nitrate	$\text{Cu}(\text{NO}_3)_2$	CuCO_3	HNO_3

Preparation of salts

There are several methods of preparing salts. The method chosen for preparing a salt entirely depends on the solubility of that particular salt.

Soluble and insoluble salts

A salt is described as soluble if it can dissolve in a given solvent and insoluble if it cannot dissolve in the solvent. Salts have varying degree of solubility in water.

Solubility is the ability of a substance to dissolve in water.

Solubility rules of salts

The following are solubility rules that are put into consideration when preparing salts.

- ✓ All sodium, potassium, ammonium and nitrate salts are soluble in water.
- ✓ All chloride salts are soluble in water except silver chloride, lead (II) chloride (sparingly soluble) and mercury (I) chloride.
- ✓ All sulphate salts are soluble in water except lead (II) sulphate and barium sulphate. Calcium sulphate is sparingly soluble in water.
- ✓ All carbonate salts are insoluble in water except sodium, potassium and ammonium carbonates.

Soluble salts are prepared either by **crystallization method** (with an insoluble base) or **titration method** or simply; neutralisation (with a soluble base).

All the insoluble salts are prepared by **precipitation method (double decomposition)** and filtered by filtration.

Preparation of soluble salts

There are four ways of preparing soluble salts. These are the same as the typical reactions of acids:

- (a) *acid + metal \rightarrow salt + hydrogen*
- (b) *acid + insoluble base \rightarrow salt + water*
- (c) *acid + soluble base \rightarrow salt + water - (titration method)*
- (d) *acid + carbonate \rightarrow salt + carbon dioxide + water*

General procedure

1. Place some dilute acid in a beaker
2. Warm the acid and add the metal, base or metal carbonate bit by bit until excess solid remains and effervescence stops (if a metal or a carbonate is used) to ensure that the acid has all reacted. **Warm** acid is used so as to speed up the reaction when using a metal or a base.
3. Filter the excess metal, metal oxide, metal hydroxide or metal carbonate and collect the filtrate.
4. Dip a piece of blue litmus paper into the solution to ensure all the acid has reacted. The litmus should remain blue.
5. Saturate the filtrate by evaporating and allow the solution to cool as it cools to form the salt crystals.

Precaution: do not heat the solution to dryness or else there would be no water for crystallisation.

6. Filter to collect the crystals. Then wash the crystals with cold distilled water so as to remove surface solution (impurities) from the crystals.
7. Dry the crystals in an oven, or under sun shine or between few sheets of filter papers.

1. Preparation of salts from metals and dilute acids

Note that this method is suitable for moderately reactive metals like **magnesium, zinc, and iron**: However, it is not suitable for

- ❖ Very more reactive metals such as potassium, sodium and calcium because these metals react violently with acid, so the reaction is very dangerous.
- ❖ Unreactive metals such as copper and silver. This is because these metals cannot displace hydrogen from an acid, hence no reaction will occur.

Note that, *Nitrate salts cannot be prepared using this method because dilute nitric acid being an oxidizing agent, does not react with metals to liberate hydrogen gas except magnesium metal.*

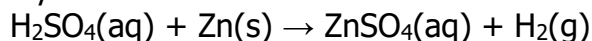
Example

Laboratory preparation of zinc sulphate crystals from zinc metal/powder

- Put dilute sulphuric acid in a beaker and heat it gently until when it is warm.
- Add zinc powder to the warm acid bit by bit while stirring until when the zinc powder is in excess.
- Filter off the excess zinc powder to obtain zinc sulphate solution as the filtrate.
- Saturate the filtrate by evaporating.
- Allow it to cool and form crystals of the salt.
- Filter the crystals and then wash them with cold distilled water.

- Dry the crystals between few sheets of filter papers.

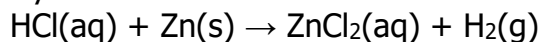
Equation



Other salts formed in similar ways are:

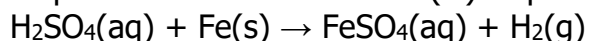
Zinc chloride

Hydrochloric acid + zinc \rightarrow Zinc chloride + Hydrogen gas



Iron (II) sulphate

Sulphuric acid + Iron \rightarrow Iron (II) sulphate + Hydrogen gas



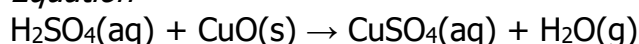
2. Preparation of salts from metal oxides and dilute acids

Example

Preparation of copper(II) sulphate from copper(II) oxide in the laboratory

- Put dilute sulphuric acid in a beaker and heat it gently until when it is warm.
- Add copper(II) oxide to the hot acid bit by bit while stirring until when the copper(II) oxide is in excess.
- Filter off the excess copper(II) oxide to obtain copper sulphate solution as the filtrate.
- Saturate the filtrate by evaporating using a water bath.
- Leave it to cool and form crystals of the salt.
- Filter the crystals and wash them with cold distilled water.
- Dry the crystals either in an oven or under sunshine or between filter papers.

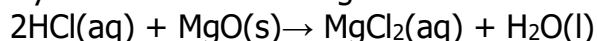
Equation



Other examples of salts formed from metal oxides are:

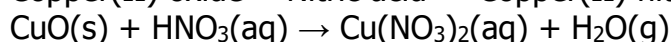
Magnesium chloride

Hydrochloric acid + Magnesium oxide \rightarrow Magnesium chloride + Water



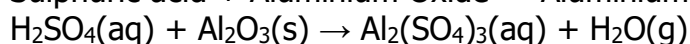
Copper (II) nitrate

Copper(II) oxide + Nitric acid \rightarrow Copper(II) nitrate + Water



Aluminium sulphate

Sulphuric acid + Aluminium Oxide \rightarrow Aluminium Sulphate + Water



3. Preparation of salts from insoluble metal carbonates

Example

Preparation of lead(II) nitrate from lead(II) carbonate

- Pour dilute nitric acid in a beaker and warm it gently.
- Add lead (II) carbonate a little at a time. Effervescence occurs as carbon dioxide is evolved.
- Continue adding the carbonate until when it is in excess and no more effervescence occurs.
- Filter off the excess carbonate to get a colourless filtrate.
- Evaporate the filtrate by heating gently to obtain a saturated solution.
- Cool the saturated solution to obtain white crystals of lead(II) nitrate salts.
- Wash the crystals with cold distilled water and dry either on sun shine, in an oven or between sheets of filter papers.

Equation

Lead (II) carbonate + Nitric acid → Lead (II) nitrate + Water + carbon dioxide
 $\text{HNO}_3(\text{aq}) + \text{PbCO}_3(\text{s}) \rightarrow \text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g})$

Other salts prepared in similar way are:

Barium chloride

Hydrochloric acid + Barium carbonate → Barium chloride + Water + carbon dioxide
 $2\text{HCl}(\text{aq}) + \text{BaCO}_3(\text{s}) \rightarrow \text{BaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

Copper (II) sulphate

Sulphuric acid + copper (II) carbonate → Copper (II) sulphate + water + carbon dioxide

**Laboratory preparations of salts by Titration method (neutralization).**

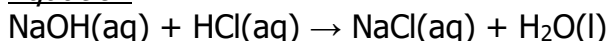
Neutralization is a reaction between an acid and a base to produce a salt and water only.

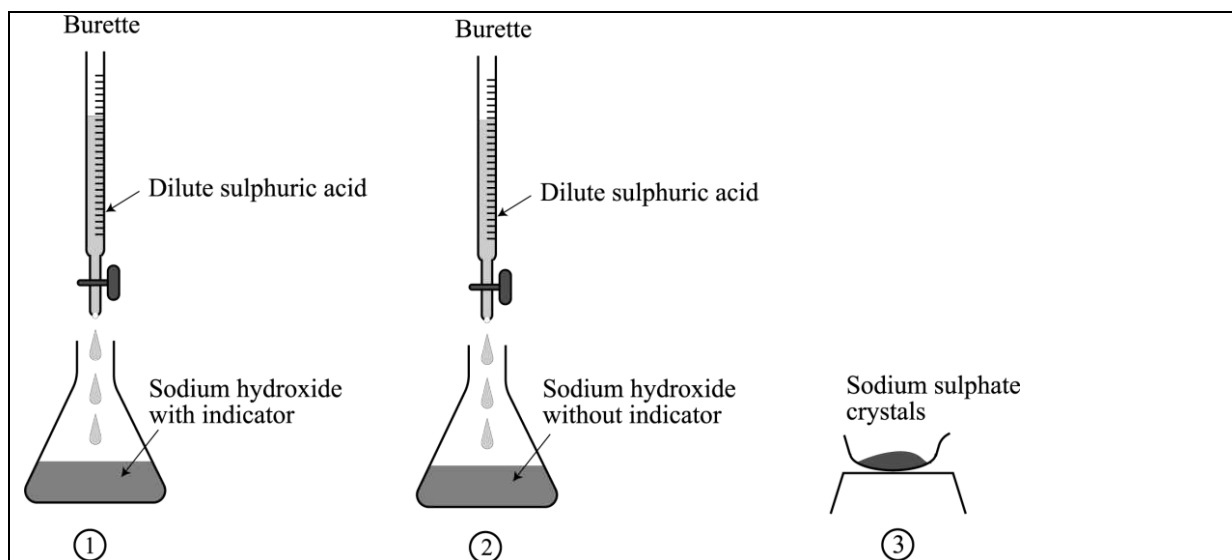
General procedure

- Put a known volume of hydroxide of a metal in a conical flask.
- Add 2 or 3 drops of an indicator solution.
- Run a suitable acid from the burette until when the colour of the mixture just changes. Note and record the volume of acid used.
- Measure accurately the same volume of hydroxide as before and titrate with exactly the same volume of acid as recorded above but without an indicator.
- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between sheets of filter papers.

Preparation of sodium chloride (NaCl) crystals in the laboratory*Procedure*

- Put a known volume of sodium hydroxide in a conical flask.
- Add 2 or 3 drops of an indicator.
- Titrate the sodium hydroxide with hydrochloric acid from the burette until when the end point is reached (when the indicator changes colour). Note and record the volume of acid used.
- Measure accurately the same volume of sodium hydroxide as before and titrate with exactly the same volume of hydrochloric acid as recorded above without using an indicator.
- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between filter papers.

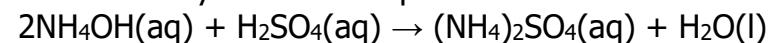
Equation



Other salts prepared in a similar way are:

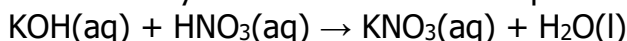
Ammonium sulphate ((NH₄)₂SO₄)

Ammonium hydroxide + sulphuric acid → ammonium sulphate



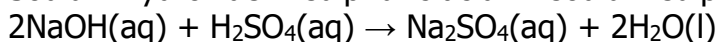
Potassium nitrate, KNO₃

Potassium hydroxide + nitric acid → potassium nitrate + water



Sodium sulphate, Na₂SO₄

Sodium hydroxide + sulphuric acid → sodium sulphate + water



Preparation of insoluble salts

Insoluble salts are prepared by **double decomposition** or **precipitation method**.

Insoluble salts, e.g., BaSO₄, CaSO₄, PbSO₄, PbCl, AgCl and most carbonates, can be prepared by reacting compound containing the wanted **cation** with another compound containing the wanted **anion**. This is **precipitation reaction**.

In this method, two soluble salts are mixed to form two new salts by exchange of radicals. One of the new salts formed is a soluble salt and one is an insoluble salt that appear as precipitates. The precipitate is filtered off and washed then dried.

Precipitation is the formation of solids when solutions are mixed.

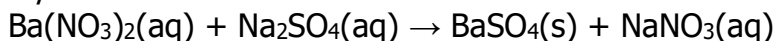
A precipitate is an insoluble solid formed when a chemical reaction occurs between two dissolved ionic substances.

Example

Preparation of barium sulphate (by reacting barium nitrate and sodium sulphate)

Procedure

- Put a solution of barium nitrate in a beaker and add a solution of sodium sulphate to it. A white precipitate of barium sulphate immediately appears,
- Filter off the precipitate and wash with distilled water.
- Dry the precipitate (salt formed) under sun shine or in an oven or between filter papers.

Equation

This method can be used to prepare salts such as *lead sulphate*, *aluminium chloride*, *silver chloride*, *silver carbonate* and *barium sulphate*.

Preparation of lead (II) sulphate (by reacting lead (II) nitrate and sulphuric acid)

Procedure

- Add dilute nitric acid to lead (II) nitrate solution in a beaker and stir the mixture. White precipitates of lead (II) sulphate are formed.
- Filter off the precipitates and wash with cold distilled water to remove traces of the acid.
- Dry the precipitates in a steam oven or on pieces of filter papers.

Water of Crystallisation

Many salts combine with water molecules to form **crystals**. These water molecules are called **water of crystallisation**.

For example, when water is added to anhydrous copper (II) sulphate heat is produced and a blue solution is formed:

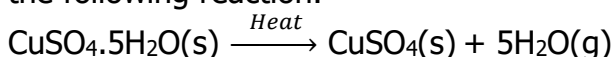


This process is called **hydration**.

Salts that contain water of crystallisation are called **hydrated salts**. *Salts that do not contain water of crystallisation* are called **anhydrous salts**. Anhydrous salts are usually in powder form. The table below shows the formulae of some anhydrous salts and some hydrated salts.

Name of salt	Formula of anhydrous salt	Formula of hydrated salt
Copper(II) sulphate	CuSO_4	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Magnesium sulphate	MgSO_4	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Sodium carbonate	Na_2CO_3	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Zinc sulphate	ZnSO_4	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$

When a hydrated salt is heated, water of crystallisation is given off as illustrated in the following reaction.

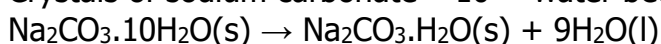


Efflorescence

This is the loss of water of crystallization to the atmosphere.

Example

Crystals of sodium carbonate – 10 – water become Powderly when exposed to air.



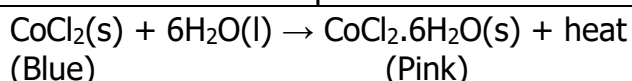
Deliquescence

This is the absorption of water from the atmosphere to form a solution. Calcium chloride is a deliquescent salt. It is used as a drying agent in desiccators. A desiccator is a piece of equipment used to dry substances.

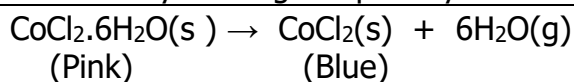
Hygroscopic

A hygroscopic substance absorbs water from the air but does not change its state.

Anhydrous cobalt chloride is a hygroscopic salt. Water changes anhydrous cobalt chloride from blue to pink.



This reaction is often used as a test for the presence of water. The process can be reversed by heating the pink hydrated salt:



Concentrated sulphuric acid (H_2SO_4) is also hygroscopic. It can be used to dehydrate blue crystals of hydrated copper(II) sulphate forming the pale blue anhydrous salt.

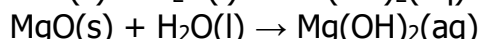
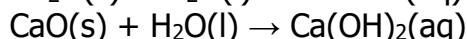
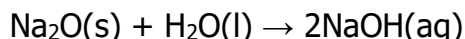
OXIDES

An oxide is a compound of oxygen with another element. The following are major classes of oxides (types of oxides).

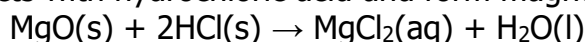
1. Basic oxides

Basic oxides are oxides with basic properties. These are oxides of metals meaning that they are formed when a metal reacts with oxygen. Examples of basic oxides include; magnesium oxide (MgO), calcium oxide (CaO), sodium oxide (Na_2O) and copper (II) oxide (CuO).

- Soluble basic oxides dissolve in water to form an alkaline solution.



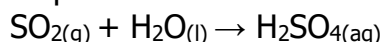
- Basic oxides react with acid to form salt and water. For example, magnesium oxide reacts with hydrochloric acid and form magnesium chloride and water.



2. Acidic oxides

Acidic oxides are oxides with acidic properties. Most of the non-metals form **acidic oxides**. Most acidic oxides dissolve in water to form an acid. One example is sulphur dioxide, which dissolves readily in water to form sulphurous acid.

Sulphur dioxide + water \rightarrow sulphurous acid



Other examples of acidic oxides are *carbon dioxide* (CO_2), *sulphur trioxide* (SO_3) and *phosphorus(V) oxide* (P_4O_{10}), *nitrogen dioxide* (NO_2) and *silicon(IV) oxide* (SiO_2).

3. Neutral oxides

Neutral oxides are oxides with neither acidic nor basic properties.

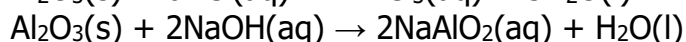
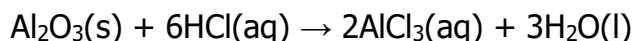
- They do not react with either acids or bases.
- Examples of neutral oxides include; water (H_2O), carbon monoxide (CO) and nitric oxide (NO)

4. Amphoteric oxides

Amphoteric oxides are oxides with both acidic and basic properties.

- Examples of amphoteric oxides include; aluminium oxide (Al_2O_3), zinc oxide (ZnO) and lead(II) oxide (PbO).
- The reaction between amphoteric oxides and acids or bases produce salt and water.

- For example, aluminium oxide reacts with hydrochloric acid to form aluminium chloride and water. Likewise, aluminium oxide reacts with concentrated sodium hydroxide to form sodium aluminate and water.



QUALITATIVE ANALYSIS TESTS

All salts are ionic compounds. This means that, they contain cations (positive ions) and anions (negative ions). In order to identify the cations and anions in the salt, certain reagents (chemicals) are added to separate portions of the salt solution. The process of identifying cations and anions present in an ionic compound is called **qualitative analysis**.

Identification of cations

The cations can be identified by adding sodium hydroxide solution and aqueous ammonia to their portions of the salt solution. Note that all cations (except all metal ions of Group I and NH_4^+) give a precipitate (ppt) with these alkalis.

A cation can be identified by noting:

- the colour of the ppt produced,*
- whether the precipitate is soluble or insoluble in an excess of the reagent,*
- whether ammonia gas is liberated on addition of sodium hydroxide solution.*

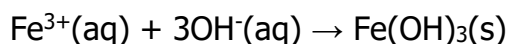
The table shown below summarises what we would observe when sodium hydroxide solution and ammonia solution are added separately to solutions containing different cations.

	Sodium hydroxide solution, NaOH(aq)	Aqueous ammonia, NH₃(aq)
	Observations on adding: a) a few drops of sodium hydroxide b) excess sodium hydroxide	Observations on adding: a) a few drops of aqueous ammonia b) excess aqueous ammonia
Al^{3+}	a) A white ppt is formed. b) The ppt dissolved in excess to give a colourless solution.	a) A white ppt is formed. b) The ppt is insoluble in excess.
NH_4^+	a) No ppt is formed. b) On warming, ammonia gas is given off.	-
Ca^{2+}	a) A white ppt is formed. b) The ppt is insoluble in excess.	a) No ppt (no change). b) No ppt (no change).
Cu^{2+}	a) A light blue ppt is formed. b) The ppt is insoluble in excess.	a) A light blue ppt is formed. b) The ppt dissolves in excess, giving a dark blue solution.
Fe^{2+}	a) A green ppt is formed. b) The ppt is insoluble in excess.	a) A green ppt is formed. b) The ppt is insoluble in excess
Fe^{3+}	a) A reddish-brown ppt is formed. b) The ppt is insoluble in excess.	a) A reddish-brown ppt is formed. b) The ppt is insoluble in excess.
Zn^{2+}	a) A white ppt is formed. b) The ppt dissolves/soluble in excess forming a colourless solution.	a) A white ppt is formed. b) The ppt dissolves/soluble in excess forming a colourless solution.
Pb^{2+}	a) A white ppt is formed. b) The ppt dissolves/soluble in excess forming a colourless solution.	a) A white ppt is formed. b) The ppt is insoluble in excess.

Why do we get the above observations?

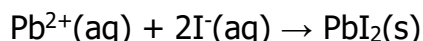
The ppt in each of the reactions above is the *hydroxide of the metal ion*. For example, the reddish-brown ppt seen on adding sodium hydroxide solution or aqueous ammonia to a solution containing iron(III) ions is iron(III) hydroxide, $\text{Fe}(\text{OH})_3$.

Iron(III) ion + hydroxide ion \rightarrow iron(III) hydroxide



Distinguishing between the Al^{3+} ion and the Pb^{2+} ion

Combining the results of the tests using sodium hydroxide solution and aqueous ammonia, the only cations which cannot be differentiated are the Al^{3+} ion and Pb^{2+} ion. Both ions form a *white* ppt with sodium hydroxide solution that is *soluble* in excess and a *white* ppt with aqueous ammonia that is *insoluble* in excess. Therefore, it is necessary to perform a further test to differentiate the two ions. The test can be done by adding a few drops of *potassium iodide* solution to a solution containing the metal cation. The Al^{3+} ion does not produce a ppt, however, the Pb^{2+} ion reacts with the iodide ion (I^{-}) to produce a *bright yellow* ppt of lead(II) iodide:



Identification of Anions

The reactions shown in the table below can be used to test/identify some common anions.

Anion	Test	Observations & Inference
CO_3^{2-}	Add dilute HCl acid. Pass the gas given off into limewater.	Effervescence is observed. Gas given off forms a white ppt with limewater. CO_2 gas is given off.
Cl^{-}	Add dilute nitric acid, then add silver nitrate solution.	A white ppt of silver chloride is formed.
I^{-}	Add dilute nitric acid, then add silver nitrate solution.	A yellow ppt of silver iodide is formed.
NO_3^{-}	Add dilute sodium hydroxide. Then add a piece of aluminium foil. Warm the mixture. Test the gas given off with a piece of damp red litmus paper.	The damp red litmus paper turns blue. Ammonia gas is given off.
SO_4^{2-}	Add dilute nitric acid, then add barium nitrate solution.	A white ppt of barium sulphate is formed.

Identification of Gases

The reactions shown in the table below can be used to test/identify some common gases.

Gas	Test	Observations & Inference
Ammonia ⁻	Introduce damp red litmus paper to the gas.	Turns damp red litmus paper blue.
Carbon dioxide (CO ₂)	Bubble the gas through limewater.	White precipitate formed.
Chlorine (Cl ₂)	Introduce damp blue litmus paper to the gas.	Turns litmus paper red then bleaches it.
Hydrogen (H ₂)	Introduce a lighted splint into the gas.	Puts out the lighted splint with a 'pop' sound.
Oxygen (O ₂)	Introduce a glowing splint into the gas.	Glowing splint relighted.
Sulphur dioxide (SO ₂)	Bubble the gas through acidified potassium dichromate (VI).	Turns orange potassium dichromate green.