Acids, Bases & Salts (Part 1)

Learning Objectives:

Acids and Bases

- describe the meanings of the terms acid and alkali in terms of the ions they produce in aqueous solution and their effects on Universal Indicator
- describe how to test hydrogen ion concentration and hence relative acidity using Universal Indicator and the pH scale
- describe qualitatively the difference between strong and weak acids in terms of the extent of ionisation
- describe the characteristic properties of acids as in reactions with metals, bases and carbonates
- state the uses of sulfuric acid in the manufacture of fertilisers and as a battery acid
- describe the reaction between hydrogen ions and hydroxide ions to produce water, $H^+ + OH^- \rightarrow H_2O$, as neutralisation
- describe the importance of controlling the pH in soils and how excess acidity can be treated using calcium hydroxide
- describe the characteristic properties of bases in reactions with acids and with ammonium salts
- classify oxides as acidic, basic, amphoteric or neutral based on metallic/non-metallic character

Salts

- describe the techniques used in the preparation, separation and purification of salts as examples of some of the separation techniques learnt previously (methods for preparation should include precipitation and titration together with reactions of acids with metals, insoluble bases and insoluble carbonates)
- describe the general rules of solubility for common salts to include nitrates, chlorides (including silver and lead), sulfates (including barium, calcium and lead), carbonates, hydroxides, group 1 cations and ammonium salts
- suggest a method of preparing a given salt from suitable starting materials, given appropriate information
- construct chemical equations, with state symbols, including ionic equations

1. Acids

1.1 Classes of Acids

Classes of Acids

Inorganic acids - prepared in the laboratory from mineral elements or inorganic matter

Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Carbonic acid	H_2CO_3

Organic Acids - obtained from plants and animals

$C_6H_8O_7$

Ethanoic acid, **CH_3COOH** – in vinegar
Citric acid – in citrus fruits and vegetables
Ascorbic acid – vitamin C
Tartaric acid – in grapes, baking powders
Lactic acid – in sour milk, yogurt and cottage cheese; produced in the muscles during vigorous activities
Tannic acid – in tea and oak leaves
Formic/methanoic acid – in bodies of red ants and stingers of bees, in textile dyeing

1.2 What are Acids?

- **Definition:**

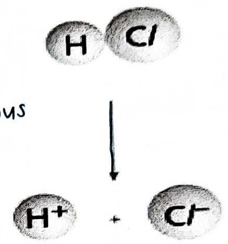
  An acid is a substance which dissociates (break up) **in** water / aqueous solution **to give** hydrogen ions. **All acids contain** H^+ ions.

- Pure acid (without water) exists as simple covalent molecules. These acid molecules react with water to produce ions. (i.e. acid molecules ionises / dissociates in presence of water)

- Equation to show the **dissociation** (break up) / **ionisation** of acid molecules: water ⇒ ~~popu~~ polar molecule

  - $\underset{\textit{acid molecules}}{HCl(g)} \xrightarrow{\text{add water}} HCl(aq) \rightarrow \underbrace{H^+(aq) + Cl^-(aq)}_{\textit{ions}}$

  (s) solid
  (l) liquid
  (g) gas
  (aq) aqueous



  Nitric acid
  - $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

  Sulfuric acid
  - $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

  Carbonic acid
  - $H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$

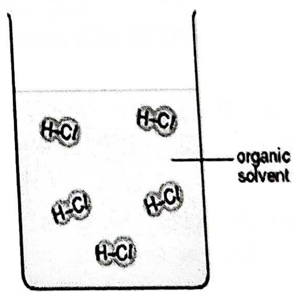
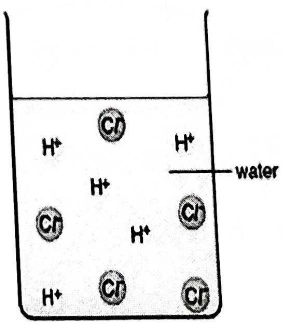
NaOH
↳ cannot be ionisation
only use dissociation

> Note[1]:
> → implies complete dissociation
> ⇌ implies partial dissociation

> Note[2]:
> ⇌ is known as reversible reaction arrow

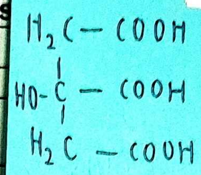
- An acid only exhibits acid properties when **dissolved in water.** The presence of hydrogen ions cause the acidity.

Compare hydrogen chloride when dissolved in (i) water and (ii) methylbenzene

Solutions	**Nature of solution**	**Reason**	
Hydrogen chloride dissolved in **methylbenzene** *[an organic solvent]*	Neutral solution	In absence of water, hydrogen chloride still exists as simple molecules. This means that no hydrogen ions are formed.	
Hydrogen chloride dissolved in **water**	Acidic solution	Hydrogen chloride dissociates in water to produce hydrogen ions and chloride ions. The hydrogen ions are responsible for the acidic property.	

1.3 Basicity of Acids

$C_6H_8O_7$
→ citric acid
→ weak acid
→ tribasic
Dissociation equation for citric acid
$C_6H_8O_7 \rightleftharpoons 3H^+_{(aq)} + 3C_6O_7^-{}_{(aq)} + 4C_6H_8O_7$
$3H^+(aq) + C_6H_5O_7^{3-}$
↓
molecules



- **Basicity** of an acid is the maximum number of [hidden] one molecule of the acid when the acid mole[cule] [hidden]
- Table showing the basicity of some common acids:

→ complete dissociation (strong acid)

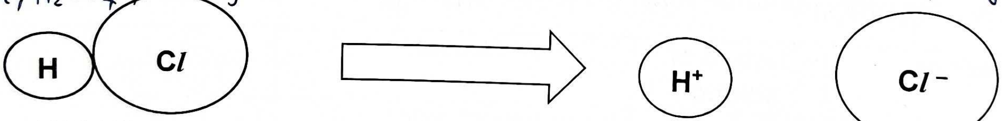
⇌ partial dissociation (weak acid)

Basicity of Acid	Examples	Equations
1 (monobasic)	hydrochloric acid strong	$HCl(aq)$
	nitric acid strong	$HNO_3(aq)$
	ethanoic acid weak	$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$
2 (dibasic)	sulfuric acid strong	$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
	carbonic acid weak	$H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$
3 (tribasic)	phosphoric acid weak	$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{3-}(aq)$

1.4 Strength of acids

- **Strength of an acid** depends on the extent of dissociation/ionisation of **the acid molecules in water / aqueous solution.**
- A **strong acid** is one that completely dissociates in water to give hydrogen ions.

e.g. HCl, H_2SO_4, HNO_3

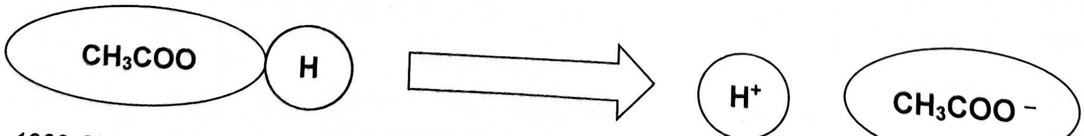


1000 HCl molecules *1000 H^+ ions and 1000 Cl^- ions*

- Equation: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

H_2CO_3, CH_3COOH

- A **weak acid** is one that is partially dissociates. in water to give hydrogen ions.

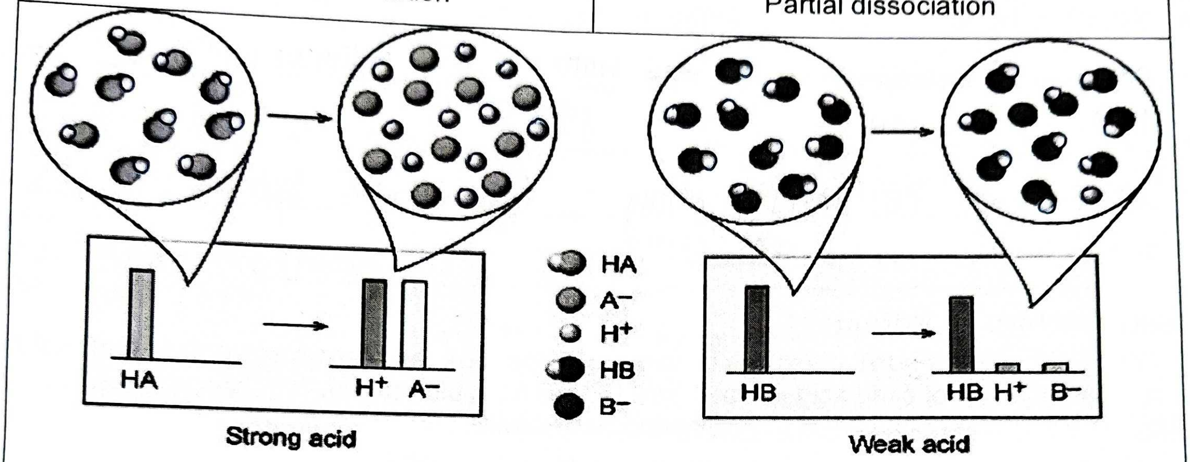


1000 CH_3COOH molecules *4 H^+ ions, 4 CH_3COO^- ions and 996 CH_3COOH molecules*

- Equation: $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

- Table comparing strong acids and weak acids:

Strong acids cannot be consumed	**Weak acids** can be consumed
E.g. Hydrochloric acid HCl Sulfuric acid H_2SO_4 Nitric acid HNO_3	E.g. Ethanoic acid CH_3COOH Carbonic acid H_2CO_3
All acid molecules form ions.	Most acid molecules remain unchanged; few ions are formed.
Complete dissociation	Partial dissociation



Comparing strong and weak acids with concentrated and dilute acid:

- **Note: "Strong" does not have the same meaning as "concentrated".**
- "**Strong**" and "**weak**" refer to the extent of dissociation of the acid in water.
- "**Concentrated**" and "**dilute**" refer to the amount of solute (acid molecule) in the solution.

1.5 Concentration of acids

- **Concentration of a solution is a measure of how much** solute **has** dissolved **in** 1dm³ **of the solution.**
- mol/dm³ **is the unit for concentration.**

Examples:

- A 10 mol/dm³ solution of hydrochloric acid is a concentrated solution of a strong acid.
- A 10 mol/dm³ solution of ethanoic acid is a concentrated solution of a weak acid.
- A 0.100 mol/dm³ solution of hydrochloric acid is a dilute solution of a strong acid.
- A 0.100 mol/dm³ solution of ethanoic acid is a dilute solution of a weak acid.

Free-moving charged particles (eg. ions, electrons) can conduct electricity.

Quick Check:

1. Explain why aqueous acids can conduct electricity.

   In the presence of water, acid molecules are able to dissociate to form ions. These ions can move freely in water to conduct electricity.

2. Why must blue litmus paper be moist to test for acidic gases such as sulfur dioxide?

- $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

  Sulfur dioxide molecules react with ~~react~~ water to form sulfurous acid.

- $H_2SO_3(aq) \rightleftharpoons 2H^+(aq) + SO_3^{2-}(aq)$

  Sulfurous acid dissociates partially in water to form hydrogen ions which turn blue litmus paper red.

No water → remains as molecule

3. Citric acid, a white solid, dissolves in water or propanone (an organic solvent) to form solutions.
   a) The solution of citric acid in water turns litmus from blue to red.
   b) The solution of citric acid in propanone has no effect on litmus paper.

   Explain each of the observation.
   a) In water, citric acid molecules partially dissociate to produce hydrogen ions which turn blue litmus to red.

   b) In propanone, citric acid still exists as simple molecules. The acid molecules do not dissociate to produce hydrogen ions. Since there is no hydrogen ions, there is no effect on litmus paper.

4. For the same concentration of acids, suggest why the reaction of magnesium with dilute hydrochloric acid is much faster than that between magnesium and dilute ethanoic acid.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Hydrochloric acid is a strong acid which completely dissociates in water to form hydrogen ions. Thus, there is a high concentration of hydrogen ions in solution to react with magnesium.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

However, aqueous ethanoic acid is a weak acid which partially dissociates in water to form hydrogen ions. Thus there is a lower concentration of hydrogen ions in solution to react with magnesium.

1.6 Properties of acids

Properties of acids

Physical -
- **sour**
- turn **blue** litmus paper **red**
- **pH < 7**
- can **conduct electricity** (are

Chemical -
Acids react with:
- **Metals** above hydrogen in the reactivity series
- **Metal carbonates / metal hydrogen carbonates**
- **Bases (metal oxides / metal hydroxides)**

Before looking at the chemical properties of acids, **how do we determine if the salt formed from the reaction is soluble or insoluble in water? (aq) or (s)?**

Solubility Table

Solubility of Salts in water

Salts	Solubility in water
All **sodium** (Na^+), **potassium** (K^+), **ammonium** (NH_4^+) ["S.P.A." salts] and salts of other Group 1 metals	are soluble (no exception)
All nitrates (NO_3^-) salts	are soluble (no exception)
All sulfate (SO_4^{2-}) salts	are soluble except barium sulfate, calcium sulfate, lead(II) sulfate, silver sulfate (sparingly soluble)
All chloride (Cl^-) salts	are soluble except silver chloride, lead(II) chloride (soluble in hot water)
All carbonates (CO_3^{2-}) salts	are insoluble except Sodium, Potassium, Ammonium, ["S.P.A." carbonates] and other carbonates of Group 1 metals

Solubility of Bases in water

Bases	Solubility in water
All oxides	are **insoluble** except Na_2O, K_2O, oxides of other Group 1 metals, BaO and CaO (Note: CaO is slightly soluble)
All hydroxides	are **insoluble** except NaOH, KOH, hydroxides of other Group 1 metals, $Ba(OH)_2$ and $Ca(OH)_2$. (Note: $Ca(OH)_2$ is slightly soluble)

Write the chemical formula of the compounds and state whether they are soluble or insoluble in water.

Compound	Chemical Formula	Solubility
copper(II) oxide	CuO	No
sodium carbonate	Na_2CO_3	Yes
iron(II) sulfate	$FeSO_4$	Yes
lead(II) sulfate	$PbSO_4$	No
potassium hydroxide	KOH	Yes
zinc nitrate	$Zn(NO_3)_2$	Yes

Name	Chemical Formula	Solubility
iron(III) chloride	$FeCl_3$	Yes
zinc carbonate	$ZnCO_3$	No
ammonium chloride	NH_4Cl	Yes
lithium sulfate	Li_2SO_4	Yes
iron(II) hydroxide	$Fe(OH)_2$	No
barium chloride	$BaCl_2$	Yes

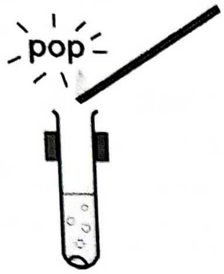
Chemical Properties of Acids

(A) Acid reacts with **metals above hydrogen in the reactivity series** to form **salt and hydrogen gas** only.

$$\boxed{\textbf{acid + metal} \rightarrow \textbf{salt + hydrogen}}$$

Eg. zinc + sulfuric acid → zinc sulfate + hydrogen

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$



What would you **observe** when the metal is added to the acid?

- Effervescence of colourless, odourless gas observed.
- Gas extinguishes a lighted/ burning splint with a 'pop' sound.
- The gas is **hydrogen**.

	Reactivity Series	**Reaction**
explosively	K Na	Acids (HCl & H_2SO_4) react with **some metals** (**above hydrogen** in the reactivity series) to form salt and liberate hydrogen gas.
vigorously	Ca Mg Al	
soluble (aq) insoluble (s) ~~mild~~ mildly	Zn Fe Sn Pb	
	H	
not reactive	***Cu*** ***Ag*** ***Au*** ***Pt***	*(No reaction with copper and metals below it.)*

Write balanced chemical equations, with state symbols, for the following reactions:

1) iron + hydrochloric acid

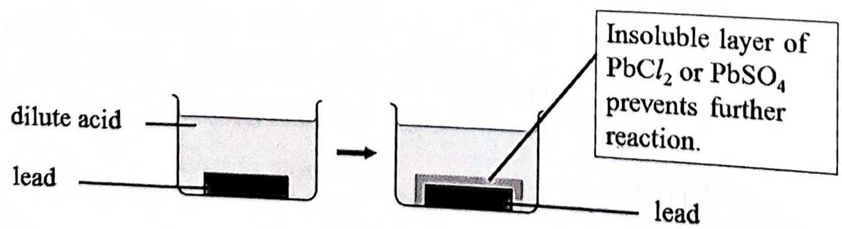
iron(II) is more common valency

$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

2) aluminium + sulfuric acid

$2Al(s) + 3H_2SO_4(aq) \rightarrow \cancel{Al_2(SO_4)_3(s)}\ Al_2(SO_4)_3(aq) + 3H_2(g)$

- Some reactive metals have no apparent reaction with acids.
  - Lead metal appears not to react with hydrochloric acid or sulfuric acid.
  - The initial reaction produces a layer of lead(II) chloride or lead(II) sulfate. This layer is insoluble in water and forms a layer around the metal, which prevents from further reaction between the metal and the acid.



$Pb(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2(g)$

(Enrichment)

- Nitric acid will not produce hydrogen gas with most metals. Instead, oxides of nitrogen will be formed.

  E.g. $Zn(s) + 4HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

- Some unreactive metals (like silver and copper) will undergo a different type of reaction with concentrated nitric acid to produce a salt, nitrogen dioxide and water.

  E.g. $Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

(B) Acid reacts with **metal carbonates** (and **hydrogen carbonates**) to form **salt, carbon dioxide and water**.

acid + metal carbonate → salt + carbon dioxide + water

acid + metal hydrogen carbonate → salt + carbon dioxide + water

E.g.

- hydrochloric acid + calcium carbonate → calcium chloride + carbon dioxide + water

  salt

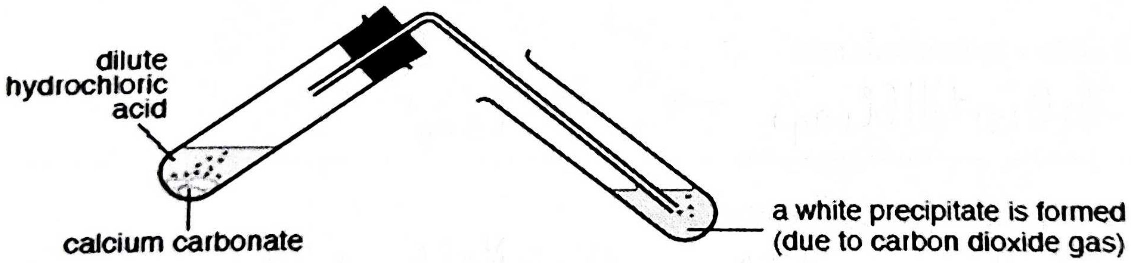
  $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

- nitric acid + sodium hydrogen carbonate → sodium nitrate + carbon dioxide + water

  salt

  $HNO_3(aq) + NaHCO_3(s) \rightarrow NaNO_3(aq) + CO_2(g) + H_2O(l)$

What would you **observe** when a metal carbonate is added to an acid?



- Effervescence of colourless, odourless gas observed.
- Gas forms white precipitate when passed into limewater (aqueous calcium hydroxide).
- The gas is **carbon dioxide**.

Why is a white precipitate seen when carbon dioxide gas is passed into limewater?

$CO_2(g)$ + $Ca(OH)_2$ (aq) → $CaCO_3$ (s) + H_2O (l)

Carbon dioxide gas (acidic gas) reacts with limewater (aqueous calcium hydroxide, an alkali) to give calcium carbonate, an insoluble salt which appears as white precipitate.

↗ same size as eqn

Write balanced chemical equations, with state symbols, for the following reactions:

1) sodium carbonate solution + nitric acid

$Na_2CO_{3\,(aq)} + 2HNO_{3\,(aq)} \rightarrow 2NaNO_{3\,(aq)} + CO_{2\,(g)} + H_2O_{(l)}$

2) magnesium hydrogen carbonate + hydrochloric acid

$Mg(HCO_3)_{2\,(s)} + 2HCl_{(aq)} \rightarrow 2CO_{2\,(g)} + 2H_2O_{(l)} + MgCl_{2\,(aq)}$

3) copper(II) carbonate + sulfuric acid

$CuCO_{3\,(s)} + H_2SO_{4\,(aq)} \rightarrow CO_{2\,(g)} + H_2O_{(l)} + CuSO_{4\,(aq)}$

(C) Acids react with **bases (metal oxides/hydroxides)** to form **salt and water only** – **(neutralisation)**

acid + bases → salt + water

E.g. hydrochloric acid + magnesium oxide → magnesium chloride + water

$2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$

Write balanced chemical equations, with state symbols, for the following reactions:

1) copper(II) oxide + sulfuric acid

$CuO_{(s)} + H_2SO_{4\,(aq)} \rightarrow CuSO_{4\,(aq)} + H_2O_{(l)}$

2) zinc oxide + hydrochloric acid

$ZnO_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\,(aq)} + H_2O_{(l)}$

3) aqueous sodium hydroxide + sulfuric acid

$2NaOH_{(aq)} + H_2SO_{4\,(aq)} \rightarrow Na_2SO_{4\,(aq)} + 2H_2O_{(l)}$

4) calcium hydroxide + nitric acid

$Cu(OH)_{2\,(s)} + 2HNO_{3\,(aq)} \rightarrow Ca(NO_3)_{2\,(aq)} + 2H_2O_{(l)}$

1.7 Uses of acids

- ***Match** appropriate uses to the acids below:*

Acids
carbonic acid
hydrochloric acid
ascorbic acid
nitric acid
sulfuric acid

Uses
vitamin C found in health supplements
making fertilizers and explosives.
in fizzy drinks
used in car batteries, for making fertilizers
produced by stomach glands, helps to kill bacteria and activate enzymes

- <u>**Uses of acids**</u>

  - **Manufacture of fertilizers** – Sulfuric acid and nitric acid are important starting materials for the production of ammonium sulfate and ammonium nitrate, which are active ingredients in fertilisers.

  - **Battery acid in cars** – Dilute sulfuric acid reacts with lead plates and lead(IV) oxide plates in the battery to release electrical energy.

  - **Pickling of metals** – Hydrochloric acid is used to remove rust which consists mainly of iron(III) oxide, Fe_2O_3 (a base). Hydrochloric acid will react with rust to form iron(III) chloride.

  - **Food production** – Phosphoric acid is added to food and beverages to give them a sour taste. Carbonic acid is used to make fizzy drinks. Ethanoic acid acts as a food preservative and flavour enhancer.

Acids, Bases & Salts (Part 2)

2. Bases / Alkalis

2.1 What is a base? What is an alkali?

insoluble soluble base

- A **base** is a substance that reacts with an acid to form salt and water only.
- Bases are usually metal oxides or metal hydroxides.
Exception: Aqueous ammonia, NH_3 (aq) is also a base.
- Examples of bases: magnesium oxide (s), copper(II) oxide (s), sodium oxide (s) and zinc hydroxide (s).
- Most metal oxides and metal hydroxides are **insoluble** in water but there are those that are **soluble** in water.
- Examples of metal oxides which are **soluble in water** include Group 1 metal oxides (e.g. sodium oxide and potassium oxide), barium oxide and calcium oxide.
- If a metal oxide dissolves and reacts with water, a solution of metal hydroxide (an alkali) is formed.

Equation: Sodium oxide
Na_2O (s) + H_2O (l) → $2NaOH$ (aq)
soluble base

calcium oxide
CaO (s) + H_2O (l) → $Ca(OH)_2$ (aq)
slightly soluble base

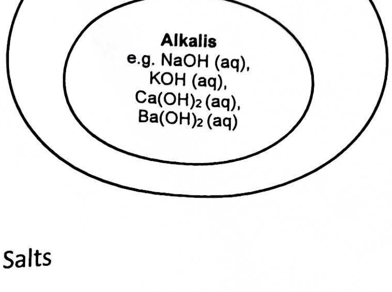
Note: CaO is only slightly soluble in water. Hence, only some CaO reacts with water.

- **Alkalis** are **bases** that dissociate **in water** to produce hydroxide **ions**, OH^- (aq)
- **Alkalis** are usually metal hydroxide solution with the **exception** of aqueous ammonia.

The table below shows some common alkalis:

Name of alkali	Chemical Formula	Ions produced in aqueous solution	
Aqueous sodium hydroxide	NaOH (aq)	Na^+	OH^-
Aqueous potassium hydroxide	KOH (aq)	K^+	OH^-
Aqueous calcium hydroxide	$Ca(OH)_2$ (aq)	Ca^{2+}	OH^-
Aqueous ammonia	NH_3 (aq)	NH_4^+	OH^-

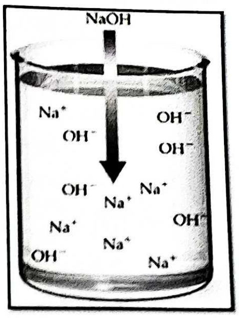
Note: All alkalis are bases but not all bases are alkalis.



$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

No ionisation takes place!

2.2 Strength of Alkalis



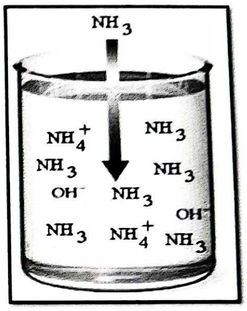
- A **strong alkali** completely dissociates in water **to give hydroxide ions.**

E.g. KOH (aq) → $K^+ (aq) + OH^- (aq)$

NaOH (aq) → $Na^+ (aq) + OH^- (aq)$

$Ca(OH)_2$ (aq) → $Ca^{2+} (aq) + 2OH^- (aq)$

Note: Calcium hydroxide (slaked lime) is not very soluble in water. However, those that dissolve will completely dissociate in water to give calcium ions, Ca^{2+} (aq), and hydroxide ions, OH^- (aq) ions. Hence, aqueous calcium hydroxide is still considered a strong alkali.



- A **weak alkali** partially dissociates in water **to give hydroxide ions.**

E.g. NH_3 (aq) + H_2O (*l*) ⇌ $NH_4^+ + OH^-$

Why does ammonia gas form a weak alkali when dissolved in water?

- When ammonia gas dissolves in water, only a **small fraction** of **ammonia** molecules react with water to produce ammonium, ~~NH_4^+~~ NH_4^+ (aq) and hydroxide ions, OH^- (aq).
- **Most ammonia** still remains as simple molecules in the solution.

2.3 Properties of Alkalis

Yes Yes!

Properties of alkalis

Physical -	*Chemical* -
Bitter **Soapy** Turns **red** litmus paper **blue** **pH > 7** Can **conduct electricity** (are electrolytes)	Alkalis react with: **Acids** **Ammonium salts** some **metal cations** to form precipitate

(A) **Bases** react with **acids** to form **salt and water only - the process is known as** neutralisation.

$$\boxed{\text{base + acid} \rightarrow \text{salt + water}}$$

CuO(s) + H_2SO_4(aq) → $H_2O (l) + CuSO_4$
insoluble base

Na_2O(s) + 2 HNO_3(aq) → $H_2O (l) + 2NaNO_3 (s)$
soluble base

NaOH(aq) + HC*l* (aq) → $H_2O (l) + NaCl (s)$
alkali

$Ba(OH)_2$(aq) + 2 HNO_3(aq) → $H_2O (l) + Ba(NO_3)_2 (s)$
alkali

On warming, alkalis react with **ammonium salt** to form salt, ammonia gas and water.

alkali + ammonium salt → salt + ammonia gas + water

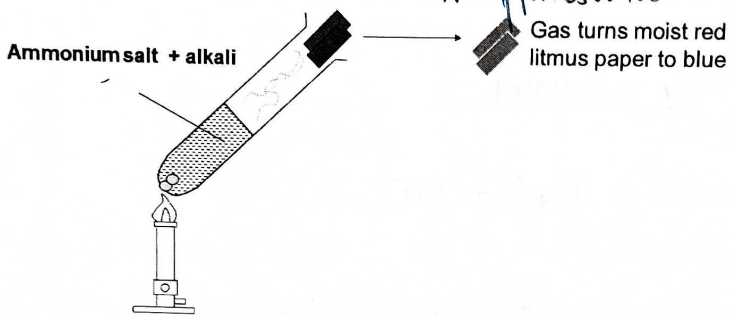
NH_4Cl (s) + NaOH(aq) → $NaCl + NH_3 + H_2O$

2 NH_4NO_3 (aq) + $Ca(OH)_2$(aq) → $Ca(NO_3)_2 + NH_3 + H_2O$

What would you **observe** when an ammonium salt is warmed with an alkali?

- Colourless and pungent gas is produced.
- Gas turns moist red litmus paper blue.
- The gas is ammonia.

No effervescence



Note:
- Warming is necessary as ammonia is soluble in water.
- Solubility of gas decreases with increase in temperature.

Why did the moist red litmus paper turn blue when ammonia gas is in contact with it?

Equation: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Explanation: Ammonia reacts with water to produce hydroxide ions which turn the red litmus paper blue.

(C) **Test for cations**

The reaction produces **insoluble metal hydroxide**, which appears as precipitate. The reaction is called a precipitation reaction.

solution containing metal ion + alkali → insoluble metal hydroxide (precipitate) + salt

Note:
- **Precipitation is a chemical reaction** between which an **insoluble product (precipitate) is formed** when **two aqueous solutions** are mixed.

$CuSO_4$(aq) + 2 NaOH(aq) → $Na_2SO_4(aq) + Cu(OH)_2(s)$

$FeCl_3$(aq) + 3 KOH(aq) → $3KCl(aq) + Fe(OH)_3(s)$

$Ca(NO_3)_2$(aq) + 2 NaOH(aq) → $2NaNO_3(aq) + Ca(OH)_2(s)$

In the laboratory, alkalis such as aqueous sodium hydroxide and aqueous ammonia are often used to test for cations.

precipitation is a chemical reaction in which an insoluble product (precipitate) is formed when 2 aqueous solutions are mixed

2.4 Uses of bases

(i) **To neutralise acids**

- Toothpaste (mainly made of magnesium hydroxide) neutralises the acid on our teeth produced by the bacteria when it feeds on sugars in our food.
- Patients suffering from gastric problem due to the ***excess*** gastric juices (mainly due to hydrochloric acid) in the stomach are given milk of magnesia, which is magnesium hydroxide. This helps to neutralise the acidic gastric juice.
- Lime (calcium oxide) or slaked lime (solid calcium hydroxide) is used to neutralise acidic soil.

(ii) **To dissolve dirt and grease**

- Soap and detergents are mild alkalis
- Floor cleaners often contain sodium hydroxide.
- Ammonia is used in liquids for cleaning glass windows.

2.5 Neutralisation

What happens during neutralisation process between an acid and an alkali?

The word equation for **neutralisation** between an acid and an alkali:

$$\boxed{\text{acid} + \text{alkali} \rightarrow \text{salt} + \text{water}}$$

H^+ OH^-

Ionic equations show **only the reacting species that are taking part** in the reaction.
Ions that **do not undergo chemical changes** are known as **spectator ions.** These ions remain in the solution.

Example 1:

Step 1: Write a balanced chemical equation with state symbols.

NaOH (aq) + HCl (aq) → $NaCl_{(s)}$ + $H_2O_{(l)}$

$\rightarrow Na^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq)$

Step 2: Write the ions for substances in the aqueous state. $\rightarrow \underbrace{Na(aq) + Cl^-(aq)}$ + H_2O (l)

$Na^+ + OH^+ (aq) + H^+ + Cl^-_{(aq)} \rightarrow H_2O (l)$ spectator ions

[Na^+ (aq) and Cl^- (aq) are the **spectator ions**; OH^-(aq) + H^+(aq) are the **reacting particles**]

y don't undergo chemical changes.

Step 3: Remove spectator ions on both sides.

The ionic equation for the reaction is $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

Example 2:
Balanced chemical equation: 2 KOH (aq) + H_2SO_4 (aq) → K_2SO_4 (aq) + $2H_2O$ (l)

$2K^+(aq) + 2OH^-(aq) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow 2K^+(aq) + SO_4^{2-}(aq) + 2H_2O(l)$

The ionic equation for the reaction is $2OH^-(aq) + 2H^+(aq) \rightarrow 2H_2O$ (l)

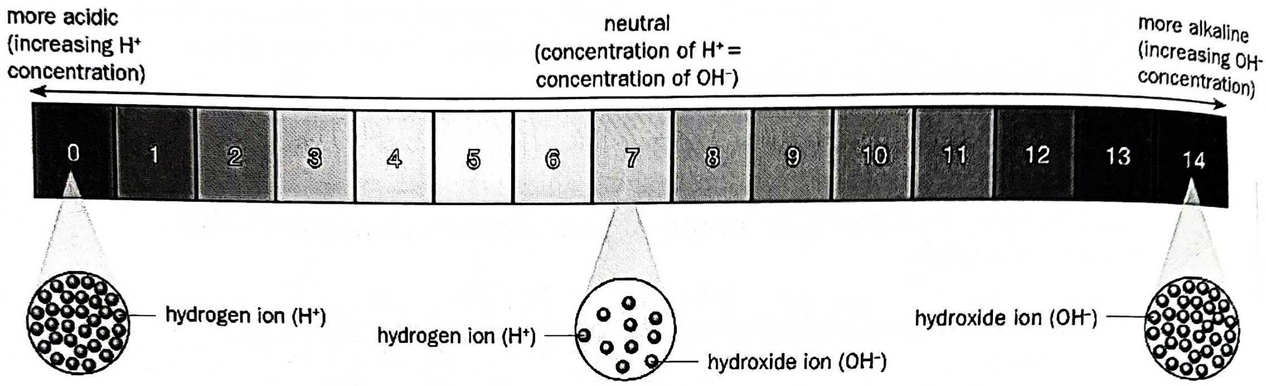
For all neutralisation reactions between an acid and an alkali, the ionic equation is:

$$\boxed{OH^-(aq) + H^+(aq) \rightarrow H_2O\ (l)}$$

3 Indicators and the pH Scale

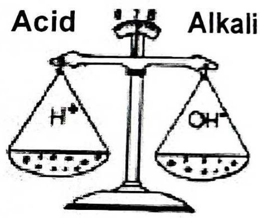
3.1 pH Scale

- pH is a measure of acidity or alkalinity in aqueous solution.
- pH measures the **concentration** of hydrogen **ions** in the aqueous solution.
- pH scale is numbered between 0 - 14.



- An aqueous solution with:
  pH < 7 is acidic; pH = 7 is neutral; pH >than 7 is alkaline.

- Note:
  (1) In a <u>neutral</u> solution, the concentration of $H^+(aq)$ is equal to the concentration of $OH^-(aq)$.
  (2) In an <u>acidic</u> solution, the concentration of $H^+(aq)$ is More than concentration of $OH^-(aq)$.
  (3) In an <u>alkaline</u> solution, the concentration of $H^+(aq)$ is less than concentration of $OH^-(aq)$.



- The smaller the pH, the more acidic the solution, the higher the concentration of hydrogen ions (H^+).
  E.g. A solution with pH 5 has a lower *concentration of* $H^+(aq)$ *than a solution with pH 1.*

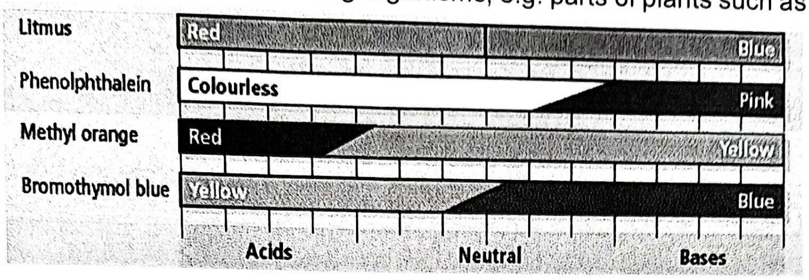
- The larger the pH, the more alkaline the solution, the higher the concentration of hydroxide ions (OH^-).
  E.g. A solution with pH 8 has a lower *concentration of* $OH^-(aq)$ *than a solution with pH 14.*

pH of some common substances:

pH range	Type	Examples
0 – 2	Strongly acidic	hydrochloric acid, nitric acid, sulfuric acid
4 – 6	Weakly acidic	carbonic acid, ethanoic acid, phosphoric acid
7	Neutral	water, aqueous sodium chloride
8 – 10	Weakly alkaline	aqueous sodium carbonate, aqueous sodium bicarbonate, aqueous ammonia
12 – 14	Strongly alkaline	aqueous sodium hydroxide, aqueous potassium hydroxide

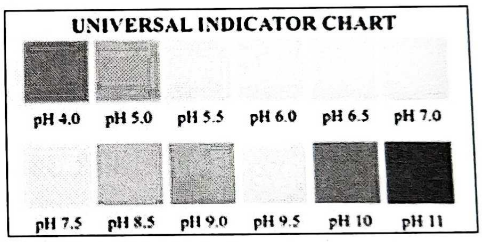
3.2 Indicators

- **An <u>indicator</u> is an organic compound which changes in colour in accordance with the pH of the solution.**
- Indicators:
  - o are substances that have **different colours in acidic and alkaline solution.**
  - o can be made from natural living organisms, e.g. parts of plants such as leaves, flowers.



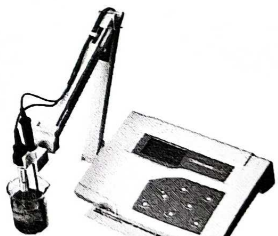
Indicator	Colour in acidic solution	Colour in alkaline solution	pH at which colour changes
Litmus	red	blue	7
Phenolphthalein	colourless	pink	9
Methyl orange	red	yellow	4 (turns orange)
Screened methyl orange	purple	green	4 (turns grey)
Bromothymol blue	yellow	blue	7 (turns green)

- Universal Indicator
  - o The Universal Indicator contains a mixture of dyes and can come in the form of a solution or pH paper.
  - o It is able to give a spectrum of different colours across pH values.
  - o The pH of the substance can be ascertained by comparing the colour obtained against a colour chart.

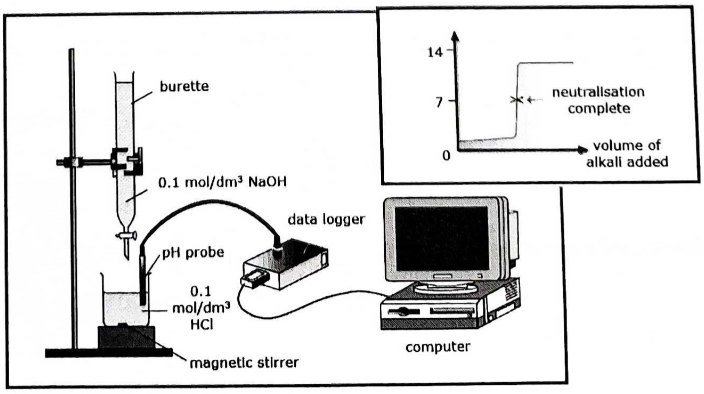


3.3 pH meter

- An electrical method of measuring the pH of a solution. It consists of a pH electrode connected to a meter. The pH electrode is dipped into solution and pH value is then shown on the meter.
- Advantages of using a pH meter over indicators:
  - o More reliable and accurate.
  - o Can be used in data logging to record rapid changes in pH.



- Variation of pH during a neutralisation reaction can be studied using a data logger and a computer.
  - o E.g. A sensor measures the pH of the mixture as an alkali is added to an acid in the beaker.



3.4 pH and agriculture

- Most plants grow best at a specific soil pH.
- Problem: Soils tend to become acidic from acid rain and extensive use of chemical fertilisers (e.g. ammonium sulfate).
- Solution: Adding of slaked lime (calcium hydroxide) to the soil to neutralise the acids, which is also known as 'liming the soil'.

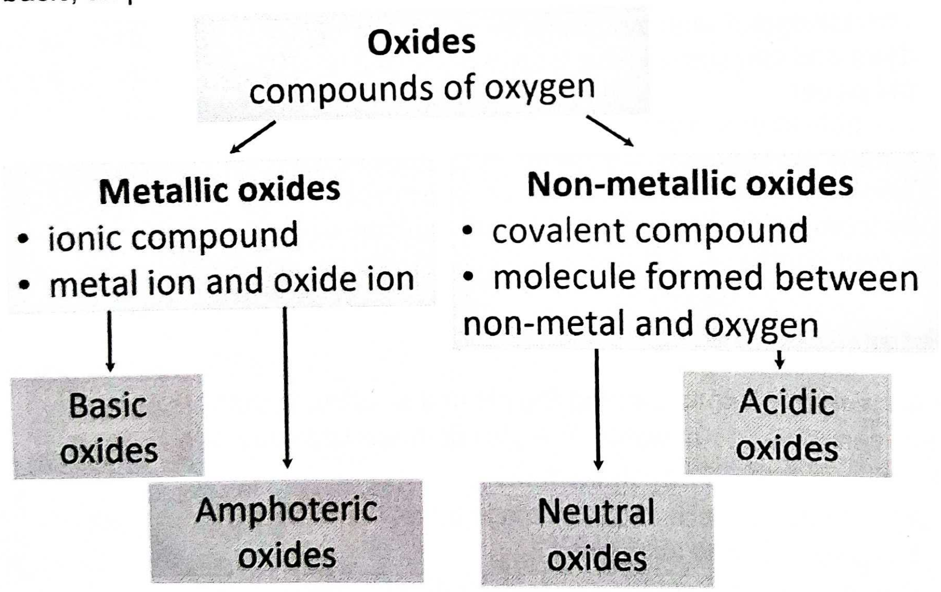
Question:
A farmer added slaked lime (calcium hydroxide) and chemical fertiliser (e.g. ammonium nitrate) to the soil at the same time.
Suggest why it is not advisable for the farmer to add these two compounds at the same time.

Chemical equation: $Ca(OH)_2 + 2NH_4NO_3 \rightarrow Ca(NO_3)_2 + \cancel{2H} \; 2NH_3 + 2H_2O$

Explanation: Other than neutralising the soil, the slaked lime will react with NH_4NO_3 to produce $NH_3(g)$ which will escape into the air. This will result in the loss of nitrogen content from the fertiliser added to the soil.

4 Types of Oxides

- An oxide is a compound of oxygen with one other element.
- Depending on their ability to react with acids and/or alkalis, there are four categories of oxides: acidic, basic, amphoteric and neutral.



→ compounds of oxygen

4.1 Acidic Oxides

- are **oxides of non-metals.** E.g. CO_2, SO_2, SO_3, NO_2, P_4O_{10}.
- **usually dissolve** and **react water** to **produce** acids (pH < 7)

E.g $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ carbonic acid

$SO_2(g) + H_2O(l) \rightarrow$ ~~H2C~~ $H_2SO_3(aq)$ sulfurous acid

$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ sulfuric acid.

Acidic oxide	Formula	Acid produced with water
sulfur trioxide	SO_3	sulfuric acid, H_2SO_4
sulfur dioxide	SO_2	sulfurous acid, H_2SO_3
carbon dioxide	CO_2	carbonic acid, H_2CO_3
phosphorus(V) oxide	P_4O_{10}	phosphoric acid, H_3PO_4

- **react with alkali** to produce salt and water only

acidic oxide + alkali → salt + water

E.g. $CO_2(g) + Ca(OH)_2(aq) \rightarrow H_2O(l) + CaCO_3(s)$

$SO_2(g) + 2\,NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$

4.2 Basic Oxides

- are **oxides of metals**. E.g. Na_2O, MgO, CuO, Fe_2O_3.
- **if soluble, dissolve** and **react with water** to **produce** alkalis ~~acids~~ (pH > 7).

pH > 7

E.g $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Note: most basic oxides are insoluble in water.

- **react with acid** to produce a salt and water only

basic oxide + acid → salt + water

E.g. $MgO(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$

$CuO(s) + 2\,HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l)$

4.3 Amphoteric Oxides

- are **oxides of metals.** E.g. ZnO, Al_2O_3, PbO.
- do not dissolve in water.
- can behave as acidic oxides or as basic oxides, i.e. it can react with acid or alkali to produce a salt and water only.

amphoteric oxide + acid → salt + water
(role as a base)

E.g. $ZnO(s)$ + 2 $HCl(aq)$ → $ZnCl_2(aq) + H_2O(l)$

$Al_2O_3(s)$ + 3 $H_2SO_4(aq)$ → $Al_2(SO_4)_3(aq) + 3H_2O(l)$

$PbO(s)$ + 2 $HNO_3(aq)$ → $Pb(NO_3)_2(aq) + H_2O(l)$

amphoteric oxide + alkali → salt + water
(role as an acid.)

E.g. $ZnO(s)$ + 2 $NaOH(aq)$ → $Na_2ZnO_2(aq) + H_2O(l)$

sodium zincate

$Al_2O_3(s)$ + 2 $NaOH(aq)$ → 2 $NaAlO_2(aq)$ + $H_2O(l)$

sodium aluminate

$PbO(s)$ + 2 $NaOH(aq)$ → $Na_2PbO_2(aq)$ + $H_2O(l)$

sodium plumbate

4.4 Neutral Oxides

- are **oxides of non-metals.** E.g. H_2O, CO, NO, N_2O.
- do **not** dissolve in water
- do **not react** with acid or alkali.

Summary table to compare the characteristics of the four types of oxides:

Types of oxides	Characteristics	Examples
Acidic	• Usually oxides of **non-metals** • **React with water** to form acids • **React with alkalis** to produce salts and water only	sulfur dioxide, SO_2 (g) carbon dioxide, CO_2 (g) nitrogen dioxide, NO_2 (g) phosphorus(V) oxide, P_4O_{10} (s)
Basic	• Usually oxide of **metals** • If soluble, will **react with water** to form alkalis • **React with acids** to produce salts and water only	sodium oxide, $Na_2O(s)$ calcium oxide, $CaO(s)$ copper(II) oxide, $CuO(s)$
Amphoteric	• Usually oxide of **metals** • Do **not** dissolve in water • **React with acids** to produce salts • **React with alkalis** to produce salts	zinc oxide, $ZnO(s)$ aluminium oxide, $Al_2O_3(s)$ lead(II) oxide, $PbO(s)$
Neutral	• Oxides of **non-metals** • Do **not** react with acids or alkalis • Do **not** dissolve in water	Water, $H_2O(l)$ carbon monoxide, $CO(g)$ dinitrogen monoxide, $N_2O(g)$ nitrogen monoxide, $NO(g)$

Quick Check:

1. Classify the following oxides:

Oxide	Type of oxide
potassium oxide	Basic
water	Neutral
nitrogen dioxide NO_2	~~Neutral~~ Acidic
zinc oxide	Amphoteric
magnesium oxide	Basic
iron(II) oxide	Basic
sulfur dioxide	~~Basic~~ Acidic
aluminium oxide	Amphoteric
carbon monoxide	Neutral
lead(II) oxide	Amphoteric

2. Element **L** burns in oxygen giving a product that dissolves in water producing an alkaline solution. What is element **L**?

A carbon **B** lead **C** sodium **D** sulfur (C) ✓

3. Element **R** reacts with air to form a gas, **T**. **T** changes the colour of damp litmus paper from blue to red. What is **R**? Acidic gas

A potassium **B** lead **C** hydrogen **D** sulfur (D) ✓

Acidic oxide

4. Excess nitrogen dioxide gas is bubbled through two different beakers, **A** and **B**, both containing pure water. Then, Universal Indicator is added into beaker **A** and solid zinc oxide is added into beaker **B**. What observation will be made?

	Beaker A	Beaker B
A	Red solution Acid	Effervescence, solid decreases in size
B	Red solution Acid	No effervescence, solid decreases in size
C	Purple solution Alkali	Effervescence, solid decreases in size
D	Purple solution Alkali	No effervescence, solid decreases in size

(B)

(D) +

5. Construct chemical equations and ionic equations (with state symbols) for the following reactions:

   a) Reaction of zinc with dilute sulfuric acid
   b) Reaction of ethanoic acid with aqueous potassium hydroxide
   c) Reaction of aqueous sodium carbonate with dilute hydrochloric acid
   d) Warming of a mixture of aqueous ammonium chloride and aqueous sodium hydroxide
   e) Mixing of aqueous copper(II) sulfate with aqueous potassium hydroxide
   f) Mixing of silver nitrate solution with sodium chloride solution

Acids, Bases & Salts (Part 3)

5 Salts

5.1 What is a salt?

- A salt is formed when the hydrogen ions, H^+, of the acid is partially or completely replaced by a metal ion or an ammonium ion.
- Some examples of salts formed from different acids:

Acids	Examples of salts that can be formed
Hydrochloric acid (HCl)	sodium chloride, $NaCl$ ammonium chloride, NH_4Cl copper(II) chloride, $CuCl_2$
Sulfuric acid (H_2SO_4)	sodium sulfate, Na_2SO_4 sodium hydrogen sulfate, $NaHSO_4$ copper(II) sulfate, $CuSO_4$ ammonium sulfate, $(NH_4)_2SO_4$
Nitric acid (HNO_3)	sodium nitrate, $NaNO_3$ zinc nitrate, $Zn(NO_3)_2$ ammonium nitrate, NH_4NO_3
Carbonic acid (H_2CO_3)	sodium carbonate, Na_2CO_3 sodium hydrogen carbonate, $NaHCO_3$ copper(II) carbonate, $CuCO_3$ ammonium carbonate, $(NH_4)_2CO_3$

(Enrichment)

- An **acid salt** is formed due to **incomplete replacement** of the hydrogen ion of an acid.
- Examples of **acid** salts:
  - sodium hydrogen sulfate, $NaHSO_4$
  - sodium hydrogen carbonate, $NaHCO_3$
  - sodium dihydrogen phosphate, NaH_2PO_4 and sodium hydrogen phosphate, Na_2HPO_4

5.2 Water of crystallisation

- Many salts have water molecules loosely associated with them. These water molecules are known as **water of crystallisation**.
  - Salts that do not contain water of crystallisation are called **anhydrous salts**.
  - Salts that contain water of crystallisation are called **hydrated salts**.

Name of salt	Anhydrous salt	Hydrated salt
Copper(II) sulfate	$CuSO_4$	$CuSO_4.5H_2O$
Cobalt(II) chloride	$CoCl_2$	$CoCl_2.6H_2O$

- Water of crystallisation can be lost by heating. As such, evaporation to dryness should not be used to obtain hydrated crystals during the separation of a soluble salt from water. For example, a white powder of anhydrous copper(II) sulfate will be obtained, instead of blue copper(II) sulfate crystals.
- On the other hand, water of crystallisation can also be easily taken in. For example, cobalt(II) chloride paper used to test for the presence of water. Blue anhydrous cobalt(II) chloride turns to pink in the presence of water.

cobalt (II) chloride paper is reusable.

5.3 Recap: Solubility of salts

Salts	Solubility in water	
All salts of group 1 metals (e.g. Na^+, K^+) and ammonium (NH_4^+) salts	are soluble	(no exception)
All nitrates (NO_3^-) salts	are soluble	(no exception)
All sulfate (SO_4^{2-}) salts	are soluble	except $BaSO_4$, $CaSO_4$, $PbSO_4$, Ag_2SO_4 (sparingly soluble)
All chloride (Cl^-) salts	are soluble	except $AgCl$, $PbCl_2$ (soluble in hot water)
All carbonates (CO_3^{2-}) salts	are **insoluble**	except carbonates of group 1 metals (e.g. Na_2CO_3, K_2CO_3) and $(NH_4)_2CO_3$

6 Preparation of salts

- When deciding on how to prepare salt, you must consider two factors:
  1. **Is the salt soluble in water?**
  2. **Are the reactants soluble in water?**

- There are three methods used in the preparation of salts:

Method	1 **Precipitation**	2 **Reaction of acids with <u>excess</u> insoluble carbonates, bases or metals**	3 **Titration**
Solubility of salt to be prepared	Insoluble	Soluble (Except group 1 salts or ammonium salts)	Soluble (Group 1 salts or ammonium salts)
Examples	Barium sulfate, $BaSO_4$ (s) Silver chloride, $AgCl$ (s) Lead(II) iodide, PbI_2 (s)	Copper(II) sulfate, $CuSO_4$ (aq) Zinc nitrate, $Zn(NO_3)_2$ (aq) Magnesium chloride, $MgCl_2$ (aq)	Sodium chloride, $NaCl$ (aq) Potassium nitrate, KNO_3 (aq) Ammonium sulfate, $(NH_4)_2SO_4$ (aq)
Reaction	Precipitation	Acid + Insoluble carbonate Acid + Insoluble base Acid + Metal	Acid + Alkali Acid + Soluble carbonate
State of reactants	Solution + Solution	Solution + Solid	Solution + Solution
Salt obtained as / from ...	Precipitate	Solution	Solution

6.1 Preparation of insoluble salts – Precipitation

- This method involves mixing **two aqueous solutions** to obtain an insoluble salt.
- **Precipitation** is a chemical reaction where an **insoluble product (precipitate)** is formed when two aqueous solutions are mixed.

$$AB\,(aq) + XY\,(aq) \rightarrow AY\,(s) + XB\,(aq)$$

solution containing cation +	**solution** containing anion →	**precipitate** + **solution**
(soluble salt)	*(soluble salt / acid)*	

Example:
Preparation of barium sulfate

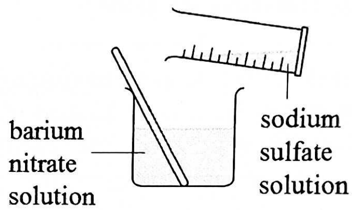
Reagent 1: aqueous barium nitrate

Reagent 2: aqueous sodium sulfate

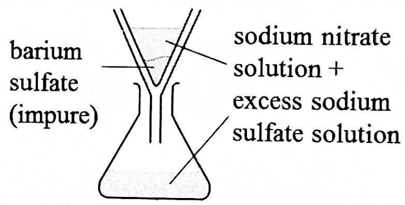
∵ all sodium salts are soluble

Chemical equation with state symbols: $Ba(NO_3)_2\,(aq) + Na_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2NaNO_3\,(aq)$

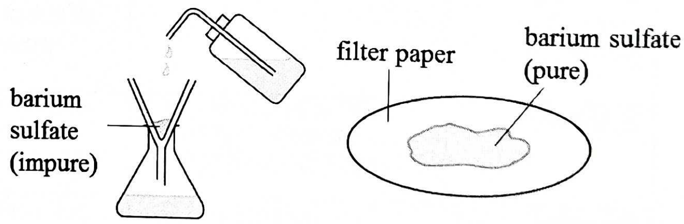
1. Mix the two aqueous solutions in a beaker.



2. Stir well. A white precipitate (insoluble salt) would be formed.
3. Filter the mixture to obtain barium sulfate as the residue.



4. Wash the residue with deionised water.
5. Dry the residue between sheets of filter paper.



Quick Check:
Name two reagents you would use to prepare the following insoluble salts.
Write chemical equation and ionic equation (including state symbols) for each of the preparation.

(i) Salt to be prepared: silver chloride

Reagent 1: aqueous silver nitrate   Reagent 2: aqueous sodium chloride

Chemical equation: $AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(s) + NaNO_3\,(aq)$

Ionic equation: $Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$

(ii) Salt to be prepared: calcium sulfate

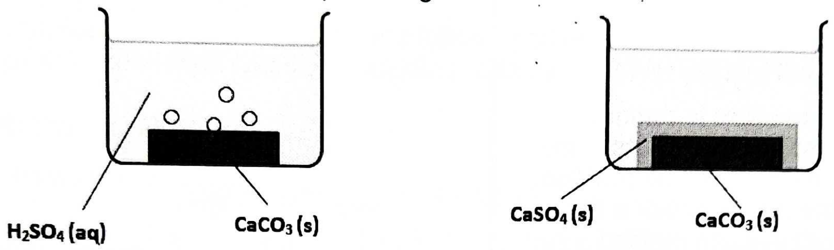
Reagent 1: aqueous calcium nitrate Reagent 2: aqueous potassium sulfate

Chemical equation: $Ca(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow CaSO_4(s) + 2KNO_3(aq)$

Ionic equation: $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$

Question: Why is insoluble reactant unsuitable for preparing an insoluble salt?

E.g. When calcium carbonate is reacted with sulfuric acid to prepare calcium sulfate, the effervescence seen initially will stop forming.



- Calcium carbonate reacts with sulfuric acid to form an insoluble salt, calcium sulfate, which will coat onto the surface of the calcium carbonate, (**preventing further reaction** between the calcium carbonate and sulfuric acid.)
- This means that the calcium sulfate salt formed will be contaminated with unreacted calcium carbonate. The calcium sulfate salt prepared is not pure.

6.2 Preparation of soluble salts – Reaction of acid with excess insoluble carbonate, insoluble base or metal

- This method involves the reaction of an acid with excess insoluble carbonate, insoluble base or moderately reactive metal.

(1)	**Acid and insoluble carbonate** acid + insoluble carbonate ***(in excess)*** → **soluble salt** + water + carbon dioxide
(2)	**Acid and insoluble base (insoluble metal oxide / insoluble metal hydroxide)** acid + insoluble base ***(in excess)*** → **soluble salt** + water
(3)	**Acid and moderately reactive metal** acid + metal ***(in excess)*** → **soluble salt** + hydrogen Note: • Only moderately reactive metals, such as Mg, Zn, Fe, should be used. • Reactive metals, such as Na, K, will react explosively with acids. • Metals below hydrogen in the reactivity series, such as Cu, Ag, Au, will not react with acid.

Example:
Preparation of zinc sulfate

Reagent 1: ~~zinc~~ zinc metal

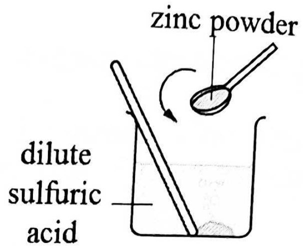
Reagent 2: Dilute sulfuric acid

Chemical equation with state symbols:

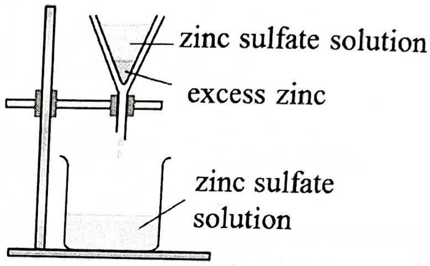
$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

excess

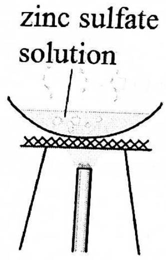
1. Heat the acid in a beaker.
2. Add **excess** insoluble base / insoluble carbonate / moderately reactive metal to the acid.
3. Stir the mixture well.



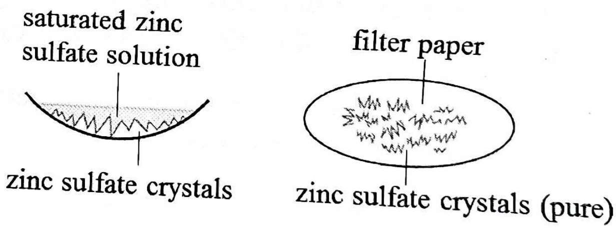
4. **Filter the mixture** to remove excess insoluble base / insoluble carbonate / metal as the **residue**.
5. Collect the **filtrate** which is the soluble salt solution needed.



6. **Heat the filtrate** to obtain a hot, saturated solution.
7. **Cool** the hot, saturated solution to obtain the crystals.



8. **Filter the mixture** to collect crystals as the residue.
9. **Dry** the crystals **between sheets of filter paper.**



Questions about the thinking behind the doing:

(Step 1) Why is dilute acid heated?

To speed up the reaction.

(Step 2) Explain why excess solid reactant is used.

To ensure that all the acid has been completely reacted with the metal / insoluble base / insoluble carbonate. This will ensure that the resulting salt solution is free from acid.

(Step 3) Why is stirring necessary?

To mix the reactants well for complete reaction.

(Step 7) Why must the hot solution be cooled to form crystals?

As temperature decreases, solubility of salt decreases. Hence, salt will crystallise from the solution.

(Step 9) Explain why the crystals must be dried between sheets of filter paper, instead of using evaporation to dryness to obtain the crystals.

Upon heating, hydrated salt may lose water of crystallisation and becomes an anhydrous salt. Some salts may also decompose.

Quick Check:
Name two reagents you would use to prepare the following soluble salts.
Write chemical equation and ionic equation (including state symbols) for each of the preparation.

(i) Salt to be prepared: zinc chloride

Reagent 1: Hydrochloric acid

Reagent 2: Zinc oxide

Chemical equation: $ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Ionic equation: $Zn(s) + 2H^+(aq) \rightarrow Zn^{2-}(aq) + H_2(g)$

(ii) Salt to be prepared: copper(II) sulfate

Reagent 1: copper(II) carbonate

Reagent 2: Sulfuric acid

Chemical equation: $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$

Ionic equation: $CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$

green solid ⇒

What would you observe when solid reactant is added?

- Effervescence of colourless, odourless gas seen
- The amount of green solid decreases
- Blue solution will be formed.

How do you tell when the reaction has stopped?

- Effervescence of colourless, odourless gas stops.
- There is green solid left behind

6.3 Preparation of soluble salts – Titration

Question: Why do we **not** prepare sodium sulfate (Na_2SO_4) the same way as zinc sulfate ($ZnSO_4$)?

i.e. Acid + metal: $H_2SO_4 (aq) + 2Na (s) \rightarrow Na_2SO_4 (aq) + H_2 (g)$

The reaction between group 1 metals and acids is highly explosive and not experimentally feasible to carry out safely. (Na is explosive)

i.e. Acid + excess carbonate: $H_2SO_4 (aq) + Na_2CO_3 (s) \rightarrow Na_2SO_4 (aq) + H_2O (l) + CO_2 (g)$
Acid + excess base: $H_2SO_4 (aq) + 2NaOH (s) \rightarrow Na_2SO_4 (aq) + H_2O (l)$

- Using the previous method, we are able to tell that the reaction is complete and that all the acid is used up because the excess unreacted carbonate, base or metal remains insoluble in water.
- However, when excess soluble base or soluble carbonate (e.g. sodium hydroxide, sodium carbonate) is added to acid, no residue will be observed. This excess soluble reactant cannot be removed by filtration. Hence, it is important to determine the exact amount of reactant needed to react with the acid.

- To do so, we will need to use an indicator to tell us exactly how much alkali or soluble carbonate is required to neutralise the acid.
- The method to prepare group 1 salts or ammonium salts is called titration.

(1)	**Acid and alkali** acid + alkali → **soluble salt** + water
(2)	**Acid and soluble carbonate** acid + soluble carbonate → **soluble salt** + water + carbon dioxide

Example:
Preparation of sodium sulfate

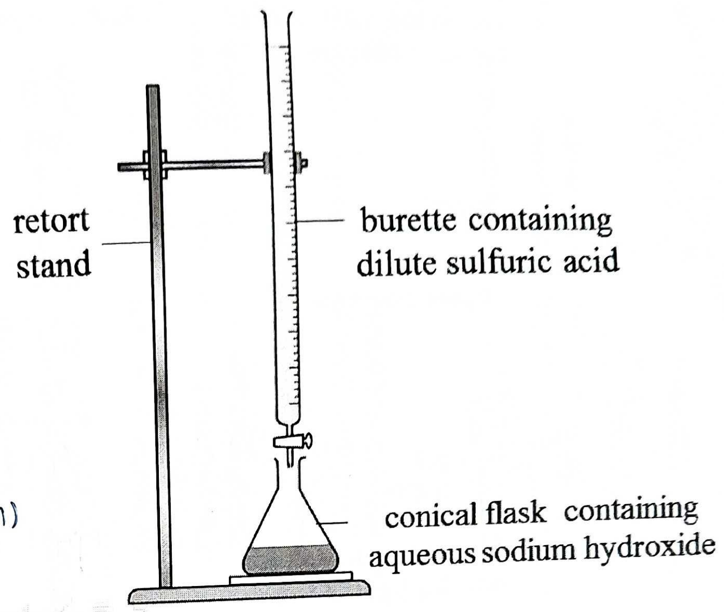
Reagent 1: Aqueous sodium hydroxide

Reagent 2: dilute sulfuric acid

Chemical equation with state symbols:

$2NaOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)$

1. Pipette a known volume of alkali / soluble carbonate into a conical flask.
2. Add a few drops of indicator.
3. Add acid from burette to the alkali / soluble carbonate until the end-point. End-point is reached when the indicator changes colour.
4. Record the volume of acid needed to neutralise the alkali / soluble carbonate.
5. Repeat titration until consistent results are obtained.
6. Repeat the experiment without adding indicator (no contamination) by adding the recorded volume of acid into the conical flask containing the known volume of alkali / soluble carbonate.
7. Heat the salt solution to obtain a hot, saturated solution.
8. Allow hot, saturated salt solution to cool for the salt crystals to form.
9. Filter the mixture to collect the crystals as residue.
10. Dry the crystals between sheets of filter paper.



Indicator	Colour in alkali	Colour at end point
Methyl orange	Yellow	Orange
Phenolphthalein	Pink	Colourless

Quick Check:
Name two reagents you would use to prepare the following soluble salts.
Write chemical equation and ionic equation (including state symbols) for each of the preparation.

acid
(i) Salt to be prepared: potassium nitrate

Reagent 1: Aqueous potassium hydroxide Reagent 2: Dilute Nitric acid

Chemical equation: $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$

Ionic equation: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

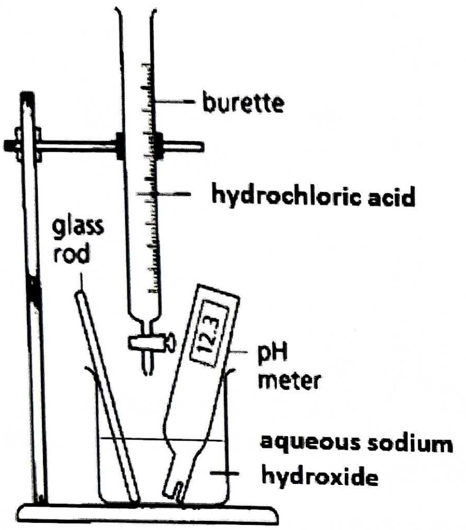
carbonate
(ii) Salt to be prepared: sodium chloride

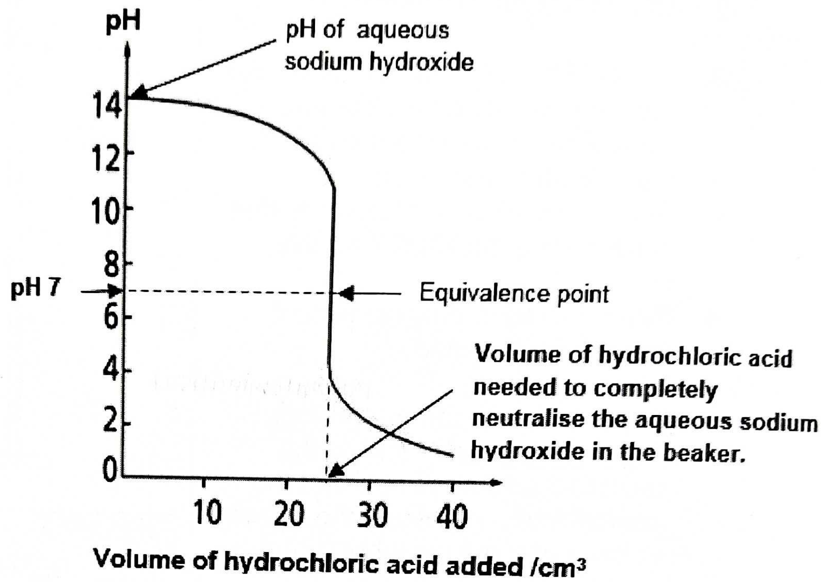
Reagent 1: aqueous Sodium carbonate Reagent 2: Dilute Hydrochloric acid

Chemical equation: $Na_2CO_3(aq) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$

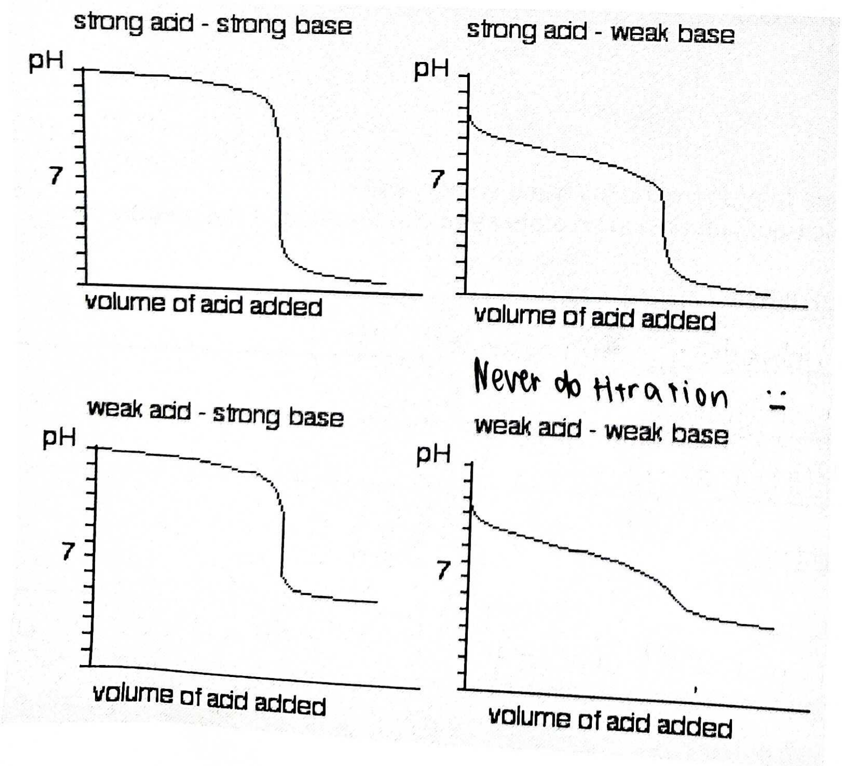
Ionic equation: ____

- Titration can also be carried out using a **pH meter** (instead of an indicator) to detect the end-point.
    - In the titration of a 1.00 mol/dm^3 of aqueous sodium hydroxide with a 1.00 mol/dm^3 of hydrochloric acid, the pH decreases from pH 14 (pH of aqueous sodium hydroxide) as more hydrochloric acid is added.
    - It finally ends at a pH near 0 (pH of hydrochloric acid).
    - When the values of pH are plotted against the volume of hydrochloric acid used, a pH curve is obtained.





- The **titration curve** above is for a titration involving a strong acid and a strong alkali. The equivalence point or point of neutralisation is at pH 7.
- However, the titration curves for a strong acid–weak alkali and a weak acid–strong alkali titration would look different.



SUMMARY - HOW TO MAKE SALTS

IS THE SALT SOLUBLE

YES – For cases of salts where the base or carbonate of the metal ion of this salt is ***insoluble*** or if the metal is moderately reactive.

1. Acid + excess insoluble carbonate
2. Acid + excess insoluble metal oxide
3. Acid + excess insoluble metal hydroxide
4. Acid + moderately reactive metal

Note-

- ***Add excess of insoluble reactant will help to ensure that all acid has completely reacted so as not to contaminate the salt solution.***
- ***Filtration needed to remove unreacted solid reactant and to collect the filtrate which contains the soluble salt solution.***

YES - For cases of Na^+, K^+ or any Group 1, or NH_4^+ salts where the base or carbonate is soluble

Method Name - TITRATION

1. Acid + Alkali / Soluble carbonate
   (Indicator is used to detect point of neutralisation)

Note: Method needs to be done twice, once with indicator to determine volume of acid/alkali needed for neutralisation and the second time without indicator as the presence of indicator will contaminate the resulting salt solution.

NO

Method Name – PRECIPITATION (Use of two soluble reactants)

Example:

X nitrate solution(aq) + Sodium Y solution(aq) → XY insoluble salt (s) + sodium nitrate solution(aq)

(precipitate)
Collect as residue after filtration

Description of the 3 methods to prepare salts

Precipitation For **insoluble salts** such as $AgCl$; $CaCO_3$ etc	**Acids + *excess* insoluble base /insoluble carbonate / moderately reactive metal** For **soluble salts** such as $CuSO_4$; $MgCl_2$; $Zn(NO_3)_2$	**Titration** For **soluble Na^+; K^+ and NH_4^+ salts** such as $NaCl$; $(NH_4)_2SO_4$; KNO_3
1. Place a known volume of **one salt solution** in a beaker. 2. Add **solution of the other salt**. 3. **Stir well** to mix. 4. **Filter mixture**. 5. **Collect residue** which is the **insoluble salt** needed. 6. **Wash residue** with **deionised water**. 7. **Dry residue** between sheets of filter papers.	1. **Heat some acid** in a beaker. 2. Add **excess of insoluble base / insoluble carbonate / moderately reactive metal** and **stir well** to mix. 3. **Filter mixture** to remove excess insoluble base / insoluble carbonate / unreacted moderately reactive metal. 4. **Collect filtrate** which is the **soluble salt solution needed**. 5. **Heat filtrate** until **saturation point**. 6. Allow **hot saturated filtrate to cool** for **salt crystals** to form. 7. **Filter mixture to collect crystals** as residue. 8. **Dry crystals** between sheets of filter papers.	1. **Pipette** a **known volume of alkali** in a conical flask. Add a **few drops of indicator**. (e.g. methyl orange – yellow in alkali ; phenolphthalein – pink in alkali) 2. **Add acid from burette** to the alkali **until indicator changes colour**. (e.g. methyl orange – turns orange at endpoint; phenolphthalein – turns colorless at endpoint) 3. **Record the volume of acid** needed to neutralize the alkali. 4. **Repeat experiment** by placing same volume of alkali in conical flask but **do not add indicator**. 5. **Add the recorded volume of acid** from burette as before to produce the soluble salt solution needed. 6. **Heat salt solution** until **saturation point**. 7. Allow **hot saturated salt solution to cool** for salt crystals to form. 8. **Filter mixture to collect crystals** as residue. 9. **Dry crystals** between sheets of filter papers.