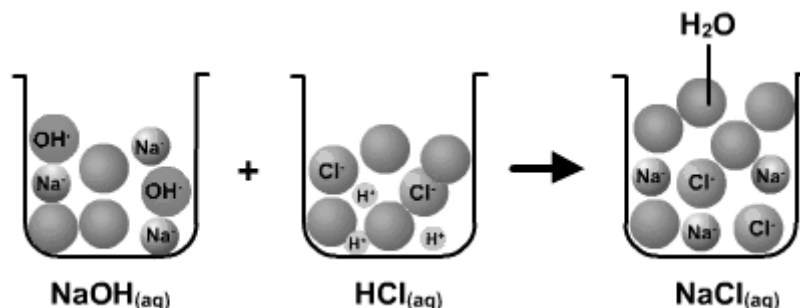


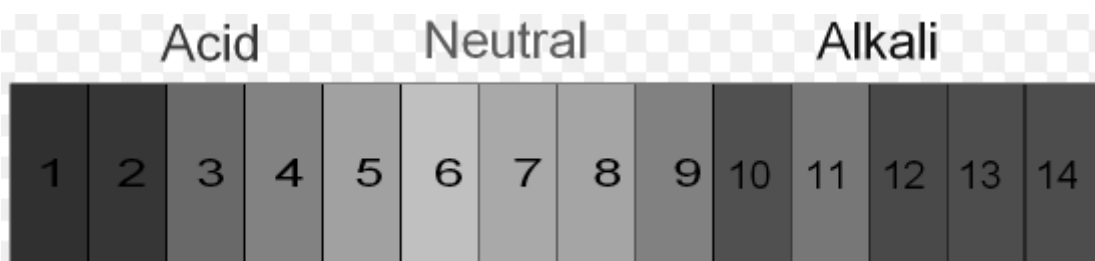
UNIT 5

ACIDS, BASES AND SALTS

Teacher Version



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Contents

- Acids, Bases, Salts and Neutralisation
- Physical Properties of Acids, Bases and Salts
- Strong and Weak Acids and Bases
- Further Reactions of Acids, Bases and Salts
- Laboratory Preparation of Salts by Neutralisation
- Qualitative Analysis 1a: Acid-Base Reactions
- Indicators and the pH scale
- Qualitative Analysis 1b: Indicators
- Quantitative Analysis: Acid-Base Titrations

Key words: Arrhenius, Bronsted-Lowry, acid, base, salt, neutralisation, hydrogen ion, hydroxide ion, proton donor, proton acceptor, alkali, amphoteric, hygroscopic, deliquescent, efflorescent, water of crystallisation, strong acid, weak Acid, strong base, weak Base, amphoteric, salt hydrolysis, acidic, alkaline, neutral, pH, indicator, litmus, phenolphthalein, methyl orange, universal indicator, titration, equivalence point, initial volume, final volume, titre volume, concordancy

Units which must be completed before this unit can be attempted:

Unit 1 – Atoms and the Periodic Table

Unit 2 – Particles, Bonding and Structures

Unit 3 – Amount of Substance and Measurement

Unit 4 – Introduction to Physical Chemistry

Estimated Teaching Time: 12 hours

Lesson 1 – What are acids, bases and salts?



Thinkabout Activity 1.1: What do you know about acids and alkalis?

- How many common acids can you name?
- Which foods and drinks contain acids?
- What do acids feel and taste like?
- How many common alkalis can you name?
- Which foods and drinks contain alkalis?
- What do alkalis feel and taste like?

Note: students should be encouraged to identify some common acids and alkalis and give some of their features

- acids: hydrochloric acid, sulphuric acid, nitric acid, lactic acid, citric acid
- vinegar (ethanoic acid), orange/lemon juice (citric acid); many fizzy sweet drinks (coke, sprite) also contain acids
- acids taste sour and bitter and they can sting if in contact with eyes or broken skin
- Alkalis: sodium hydroxide, calcium hydroxide, ammonia, soaps, detergents, bleach
- foods tend not to be strongly alkaline as they would be harmful
- alkalis are generally not good to eat; they feel and taste soapy

(a) Definition of Acids, Bases, Salts and Neutralisation

(i) Acids

- The **Arrhenius** definition of an acid is a substance which releases hydrogen (H^+) ions when dissolved in water; common examples are:
 - hydrogen chloride (HCl) dissolves in water to produce hydrogen (H^+) and chloride (Cl^-) ions:
 $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$; the resulting solution is known as **hydrochloric acid**
 - hydrogen nitrate (HNO_3) dissolves in water to produce hydrogen (H^+) and nitrate (NO_3^-) ions:
 $HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$; the resulting solution is known as **nitric acid**
 - hydrogen sulphate (H_2SO_4) dissolves in water to produce hydrogen (H^+) and nitrate (SO_4^{2-}) ions:
 $H_2SO_4(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$; the resulting solution is known as **sulphuric acid**
- Acids therefore mostly exist in aqueous solution (aq)
- the **Bronsted-Lowry** definition of an acid is a proton (H^+) donor; the acid reacts with other ions and molecules in solution, losing its H^+ in the process
 - HCl, HNO_3 and H_2SO_4 are all acids because they lose their H^+ ions (protons) when in solution
- the Bronsted-Lowry definition of an acid includes all substances which are classified as Arrhenius acids; it also includes some substances which are not classified as Arrhenius acids (usually because they do not dissolve in water)

Unit 5 – Acids, Bases and Salts

(ii) Bases and Alkalis

- The **Arrhenius** definition of a base is a substance which releases hydroxide (OH^-) ions when dissolved in water; common examples are:
 - some **metal oxides** such as CaO which react with water to produce hydroxide (OH^-) ions:
 $\text{CaO(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$
 - some **metal hydroxides** such as NaOH which dissolve in water to produce hydroxide (OH^-) ions: $\text{NaOH(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 - some **metal carbonates** such as Na_2CO_3 which react with water to produce hydroxide (OH^-) ions:
 $\text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O(l)} \rightarrow 2\text{Na}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
 - **ammonia** (NH_3), which reacts with water to produce hydroxide (OH^-) ions:
 $\text{NH}_3(\text{g}) + \text{H}_2\text{O(l)} \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
- The **Bronsted-Lowry** definition of a base is a proton (H^+) acceptor; the base accepts H^+ ions (protons) from other species in the solution (which are therefore behaving as acids); the four most common types of base are:
 - OH^- (hydroxide) ions, which accept H^+ ions to form water: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
 - CO_3^{2-} (carbonate) ions, which accept H^+ ions to form carbon dioxide and water: $\text{CO}_3^{2-} + 2\text{H}^+ \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
 - O^{2-} (oxide) ions, which accept H^+ ions to form water: $\text{H}^+ + \text{O}^{2-} \rightarrow \text{H}_2\text{O}$
 - NH_3 , which accepts H^+ ions to become NH_4^+ (ammonium): $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$
- The Bronsted-Lowry definition of a base includes all substances which are classified as Arrhenius bases; it also includes some substances which are not classified as Arrhenius bases, usually because they do not dissolve in water
- A Bronsted-Lowry base which dissolves in water to produce a solution containing OH^- ions is called an **alkali**; an alkali is therefore a soluble base; an alkali is therefore the same as an Arrhenius base
- According to Bronsted-Lowry, any reaction which involves the transfer of protons (ie from an acid to a base) can be classified as an acid-base reaction

(iii) Salts and Neutralisation

- When an acid reacts with a base, it loses its H^+ ion; the H^+ ion is replaced either with the metal ion from the base or with an ammonium ion; the resulting compound is called a **salt**; a salt is the product formed when the H^+ ion from an acid is replaced by a metal ion or an ammonium ion
- A reaction between an acid and a base to make a salt is called a **neutralisation** reaction
- There are four main types of base, so there are four main types of neutralisation reaction in which a salt can be made from an acid and a base:
 - from metal hydroxides: acid + metal hydroxide \rightarrow metal salt + water
Eg $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ (the salt produced is sodium chloride)
 - from metal oxides: acid + metal oxide \rightarrow metal salt + water
Eg $2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$ (the salt produced is calcium nitrate)
 - from metal carbonates: acid + metal carbonate \rightarrow metal salt + carbon dioxide + water
Eg $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$ (the salt produced is potassium sulphate)
 - from ammonia: acid + ammonia \rightarrow ammonium salt
Eg $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$ (the salt produced is ammonium chloride)



Test your knowledge 1.2: Acids, Bases and Salts

Deduce the formulae of the following substances and indicate whether they are acids, bases or salts:

- | | |
|------------------------|-----------------------|
| a) Sodium oxide | j) Calcium nitrate |
| b) Calcium hydroxide | k) Sulphuric acid |
| c) Ammonium nitrate | l) Ammonium chloride |
| d) Potassium carbonate | m) Nitric acid |
| e) Strontium sulphate | n) Potassium sulphate |
| f) Ammonium sulphate | o) Magnesium oxide |
| g) Hydrochloric acid | p) Caesium bromide |
| h) Rubidium hydroxide | q) Barium sulphate |
| i) Magnesium carbonate | r) Strontium nitrate |

- | | |
|---|---|
| a) Na ₂ O (base) | j) Ca(NO ₃) ₂ (salt) |
| b) Ca(OH) ₂ (base) | k) H ₂ SO ₄ (acid) |
| c) NH ₄ NO ₃ (salt) | l) NH ₄ Cl (salt) |
| d) K ₂ CO ₃ (base) | m) HNO ₃ (acid) |
| e) SrSO ₄ (salt) | n) K ₂ SO ₄ (salt) |
| f) (NH ₄) ₂ SO ₄ (salt) | o) MgO (base) |
| g) HCl (acid) | p) CsBr (salt) |
| h) RbOH (base) | q) BaSO ₄ (salt) |
| i) MgCO ₃ (base) | r) Sr(NO ₃) ₂ (salt) |



Test your knowledge 1.3: Neutralisation Reactions

Write balanced symbol equations for the following neutralisation reactions:

- | | |
|---|--|
| a) Nitric acid with potassium hydroxide | g) Nitric acid with calcium carbonate |
| b) Sulphuric acid with sodium hydroxide | h) Hydrochloric acid with barium carbonate |
| c) Hydrochloric acid with calcium hydroxide | i) Sulphuric acid with sodium carbonate |
| d) Nitric acid with calcium oxide | j) Nitric acid with ammonia |
| e) Hydrochloric acid with barium oxide | k) Sulphuric acid with ammonia |
| f) Sulphuric acid with magnesium oxide | l) Hydrochloric acid with ammonia |

- | | |
|---|---|
| a) $\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$ | g) $\text{HNO}_3 + \text{CaCO}_3 \rightarrow \text{Ca(NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}$ |
| b) $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$ | h) $2\text{HCl} + \text{BaCO}_3 \rightarrow \text{BaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$ |
| c) $2\text{HCl} + \text{Ca(OH)}_2 \rightarrow \text{CaCl}_2 + 2\text{H}_2\text{O}$ | i) $\text{H}_2\text{SO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$ |
| d) $2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca(NO}_3)_2 + \text{H}_2\text{O}$ | j) $\text{HNO}_3 + \text{NH}_3 \rightarrow \text{NH}_4\text{NO}_3$ |
| e) $2\text{HCl} + \text{BaO} \rightarrow \text{BaCl}_2 + \text{H}_2\text{O}$ | k) $\text{H}_2\text{SO}_4 + 2\text{NH}_3 \rightarrow (\text{NH}_4)_2\text{SO}_4$ |
| f) $\text{H}_2\text{SO}_4 + \text{MgO} \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$ | l) $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$ |

Lesson 2 – What are the physical properties of acids, bases and salts?

(b) Physical Properties of Acids, Bases and Salts

(i) Solubility

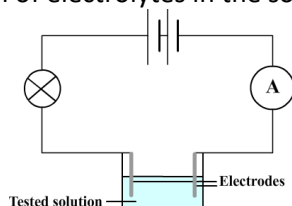
- Most acids are soluble in water; hydrochloric acid, sulphuric acid and nitric acid are all highly soluble in water; a small number of organic acids do not dissolve in water, but all acids will dissolve in alkalis such as NaOH
- The solubility of bases in water varies:
 - **metal hydroxides** are usually insoluble in water; they are therefore bases but not alkalis
the exceptions are the Group I hydroxides (LiOH, NaOH, KOH etc) and strontium and barium hydroxide ($\text{Sr}(\text{OH})_2$ and $\text{Ba}(\text{OH})_2$), which are highly soluble in water, and calcium hydroxide ($\text{Ca}(\text{OH})_2$), which is slightly soluble in water; these hydroxides are alkalis as well as bases
 - **metal oxides** are usually insoluble in water; they are therefore bases but not alkalis
the exceptions are the Group I oxides (Li_2O , Na_2O , K_2O etc), and strontium and barium oxide (SrO and BaO), which are highly soluble in water; and calcium oxide (CaO), which is slightly soluble in water; when these oxides dissolve in water, they react with water to form their respective hydroxides: Eg $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$
 - **metal carbonates** are usually insoluble in water; they are therefore bases but not alkalis
the exceptions are the Group I carbonates (Li_2CO_3 , Na_2CO_3 , K_2CO_3 etc) which are highly soluble in water; these hydroxides are alkalis as well as bases
 - ammonia is highly soluble in water; it is therefore an alkali as well as a base
 - all bases will dissolve in acids such as HCl, even if insoluble in water
- Most common salts are soluble, with some exceptions:
 - all nitrate salts are soluble
 - most common chloride salts are soluble (except AgCl and PbCl_2)
 - most common sulphate salts are soluble, except barium sulphate and strontium sulphate (which are completely insoluble) and calcium sulphate (which is only slightly soluble) and a few other less common exceptions
 - all Group I salts and all ammonium salts are solubleThe solubility of salts will be studied in more detail later in the course

(ii) Electrical conductivity

- Every ionic compound can be classified as an acid, base or salt, based on the cations and anions it contains:
 - If the cation is H^+ , the ionic compound is an **acid**
 - If the anion is O^{2-} , OH^- , CO_3^{2-} or HCO_3^- , the ionic compound is a **base**
 - Otherwise the ionic compound is a **salt**
- Most ionic compounds dissolve in water to produce free ions (note the exceptions above); as a result, aqueous solutions of ionic compounds (acids, bases or salts) behave as electrical conductors; molten ionic compounds also contain ions which can move; as a result, molten acids, bases and salts also behave as electrical conductors; any substance which can conduct electricity in the molten or aqueous states is called an **electrolyte**; molten and aqueous solutions of acids, bases and salts are **electrolytes**; any solution which shows electrical conductivity must therefore contain an acid, a base or a salt (or a mixture of them)

Unit 5 – Acids, Bases and Salts

- The electrical conductivity of a solution can easily be measured using a simple circuit; a cell, bulb and ammeter are all connected in series; the circuit is completed by two electrodes dipped into the solution but otherwise not in direct contact; if the bulb lights, the solution contains electrolytes; the greater the ammeter reading, the greater the concentration of electrolytes in the solution



(iii) Taste

- Acids have a sour taste, caused by the hydrogen ions; alkalis have a soapy taste, caused by the hydroxide ions; many salts have a salty taste, mainly caused by sodium ions but also caused by some other cations

(iv) Deliquescence, Efflorescence, Hygroscopy and Water of Crystallisation

- Many acids, bases and salts in the solid and aqueous states are hygroscopic, deliquescent, or efflorescent
- A **hygroscopic** substance is one which can absorb water from the atmosphere; this happens because the ions in the solid are strongly attracted to water; most concentrated solutions of acids and alkalis are hygroscopic and so are most solid salts if they are soluble in water; examples of hygroscopic substances are concentrated sulphuric acid, solid sodium chloride and solid sodium hydroxide; hygroscopic substances need to be stored in dry conditions to stop them from absorbing water
- Solids which are very hygroscopic are also deliquescent; a **deliquescent** solid is one which can absorb so much water from the atmosphere that it will eventually dissolve in the water it has absorbed; sodium hydroxide and calcium chloride are both deliquescent solids
- Hygroscopic and deliquescent solids are often used as **drying agents** because of their ability to absorb water from the atmosphere
- Many bases and salts, and some acids, in the solid state contain water molecules locked in between the ions in the crystal structure; these are known as **hydrated** solids; they often have a fixed chemical formula; the water is shown after the unit formula, separated by a dot:

Examples of hydrated solids	Unit formula
Hydrated copper sulphate	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
Hydrated sodium carbonate	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Hydrated ethanedioic acid	$\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

The water locked into the solid in these compounds is known as **water of crystallisation**; the water of crystallisation should be included in the molar mass of the solid:

- Eg the molar mass of hydrated sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, is $106 + 10 \times 18 = 286$

- Some of these compounds lose this water on exposure to air over a period of time; this is known as efflorescence; an **efflorescent** solid is one which contains water within its crystal structure but which will gradually release the water on exposure to the atmosphere; efflorescence is the opposite of hygroscopy

Unit 5 – Acids, Bases and Salts

- In many cases, there is an observable change in the solid as a result of efflorescence:
 - hydrated copper sulphate is blue; when exposed to the atmosphere, it gradually loses its water and becomes anhydrous copper sulphate, which is white
 - wet plaster contains hydrated calcium sulphate ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$); when plaster is left to dry, it loses its water and hardens



Test your knowledge 2.1: Some properties of acids, bases and salts

- Name two substances which are soluble bases (alkalis)
- Name two substances which are insoluble bases
- Name two soluble salts
- Name two insoluble salts
- Explain why soluble acids, alkalis and soluble salts are all electrolytes
- Describe the typical taste of acids, bases and salts and give the ion responsible for the taste
- Explain what is meant by the term hygroscopic and name one hygroscopic acid, one hygroscopic base and one hygroscopic salt
- Explain what is meant by the term deliquescent and name one deliquescent salt
- What is meant by the term “hydrated salt”?
- A sample of hydrated calcium nitrate is found to have a molar mass of 236.1. Deduce its dot formula.
- What is efflorescence? Give an example of an efflorescent salt.

- eg potassium hydroxide, sodium carbonate, ammonia etc
- eg magnesium oxide, calcium carbonate, aluminium hydroxide etc
- eg sodium chloride, magnesium nitrate, ammonium sulphate etc
- eg silver chloride, lead chloride, barium sulphate etc
- they contain ions and so conduct electricity when in solution
- acids are sour due to H^+ ; bases are soapy due to OH^- ; salts are salty, often due to Na^+
- absorbs water from the atmosphere; eg concentrated sulphuric acid, solid sodium hydroxide, solid calcium chloride
- absorbs water from the atmosphere and then dissolves in the water it has absorbed; eg calcium chloride
- a salt which has water locked into its crystal structure
- Mr of $\text{Ca}(\text{NO}_3)_2 = 164.1$; $x\text{H}_2\text{O} = 236.1 - 164.1 = 72$, so $x = 72/18 = 4$, so $\text{Ca}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$
- the loss of water from a crystal structure; eg $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$

Lesson 3 – What is the difference between strong and weak acids?

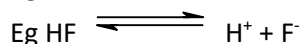
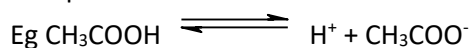
(c) Strong and Weak Acids and Bases

(i) strong and weak acids

- Acids react with water to form solutions containing H^+ ions:
Eg $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$ Eg $\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$
In hydrochloric acid and nitric acid, the molecule dissociates completely into the H^+ ion and the anion; there is no molecular HCl or HNO_3 remaining
An acid which completely dissociates in water to form hydrogen ions is called a **strong acid**; a **strong acid is a substance which completely dissociates in water to form H^+ ions**; HCl, HNO_3 and H_2SO_4 are strong acids

Unit 5 – Acids, Bases and Salts

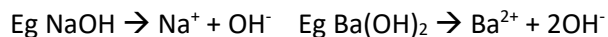
- Most acids, however, only dissociate slightly in water; the H^+ ions, anions and undissociated molecules form a dynamic equilibrium:



If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have dissociated), then the acid is called a **weak acid**; a **weak acid is a substance which only slightly dissociates in water to form H^+ ions**; all carboxylic acids, such as methanoic acid and ethanoic acid, are weak acids

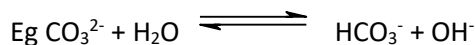
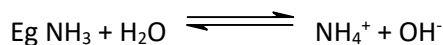
(ii) Strong and weak bases

- Bases react with water to form solutions containing OH^- ions:



In sodium hydroxide and barium hydroxide, the solid dissolves completely into the OH^- ion and the cation; there is no undissolved or undissociated NaOH or $\text{Ba}(\text{OH})_2$; a base which completely dissociates in water to form OH^- ions is called a **strong base**; a **strong base is a substance which completely dissociates in water to form OH^- ions**; NaOH , KOH and BaO are strong bases

- Some bases are highly soluble in water, but do not fully react with water to form hydroxide ions; the OH^- ions, cations and undissociated molecules form a dynamic equilibrium:



If the equilibrium position lies mainly on the left-hand side (ie only a small number of the molecules have reacted), then the bases is called a **weak base**; a **weak base is a substance which only slightly dissociates in water to form OH^- ions**; ammonia and sodium carbonate are weak bases.

- There are also some bases, such as $\text{Ca}(\text{OH})_2$, which are only slightly soluble in water but which fully dissociate into OH^- ions once they are dissolves; these are also classified as weak bases because they do not produce solutions containing a lot of OH^- ions due to their low solubility

(iii) Properties of strong acids and bases compared to weak acids and bases

(α) Electrical conductivity:

- Strong acids and bases are completely dissociated into their ions:
 - a 1 mol dm^{-3} solution of HCl contains $1 \text{ mol dm}^{-3} \text{ H}^+$ and $1 \text{ mol dm}^{-3} \text{ Cl}^-$; it does not contain any undissociated HCl
 - a 1 mol dm^{-3} solution of NaOH contains $1 \text{ mol dm}^{-3} \text{ OH}^-$ and $1 \text{ mol dm}^{-3} \text{ Na}^+$ it does not contain any undissociated NaOH
 - these solutions both contain the same number of ions and so will both show the approximately the same electrical conductivity in aqueous solution
- Weak acids and bases are only slightly dissociated into their ions:
 - 1 mol dm^{-3} solution of CH_3COOH contains mostly undissociated CH_3COOH ; it contains a very small quantity of CH_3COO^- and H^+ ions (around 0.01 mol dm^{-3})
 - a 1 mol dm^{-3} solution of NH_3 contains mostly undissociated NH_3 . It contains a very small quantity of NH_4^+ and OH^- ions (around 0.01 mol dm^{-3})

Solutions of weak acids and bases contain far fewer ions than solutions of strong acids and alkalis of the same concentration; they therefore have a much lower electrical conductivity than **solutions of strong acids and alkalis of the same concentration**

Unit 5 – Acids, Bases and Salts

(β) Enthalpy of neutralisation:

- Solutions of strong acids contain fully dissociated H^+ ions; solutions of strong bases contain fully dissociated OH^- ions:
 - when a solution of a strong acid reacts with a solution of a strong base, the H^+ ions react with the OH^- ions to make H_2O : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
 - the remaining ions are not involved in the reaction; this is true of a neutralisation reaction between any strong acid and any strong base – in all cases, an identical chemical reaction is taking place
 - as a result, all reactions between strong acids and strong bases in solution have the same enthalpy of neutralisation (-58 kJmol^{-1})
 - eg $\text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$ ($\Delta H = -58 \text{ kJmol}^{-1}$); $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$ ($\Delta H = -58 \text{ kJmol}^{-1}$)
- Solutions of weak acids only contain small quantities of H^+ ions
 - when weak acids are neutralised by strong bases, the OH^- ions react with the H^+ ions which quickly run out
 - as a result, the weak acid dissociates more, producing more H^+ ions (according to Le Chatelier's principle); this process continues until the weak acid has fully dissociated; so during the neutralisation of weak acids, two reactions are taking place:
 - Step 1: the acid dissociates to form more H^+ : $\text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}^+$
 - Step 2: the resulting H^+ reacts with the OH^- : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
 - the first reaction, the dissociation, is endothermic by around 2 kJmol^{-1} ; the second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol^{-1} ; the overall molar enthalpy change of neutralisation is therefore -56 kJmol^{-1}
 - eg $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$ $\Delta H = -56 \text{ kJmol}^{-1}$
 - weak acids have a less exothermic enthalpy of neutralisation than strong acids, because the H^+ ion has to be removed from the acid first, which absorbs some of the energy
- Solutions of weak bases only contain small quantities of OH^- ions:
 - When weak bases are neutralised by strong acids, the H^+ ions react with the OH^- ions which quickly run out
 - as a result, the weak base dissociates more, producing more OH^- ions (according to Le Chatelier's principle); this process continues until the weak base has fully dissociated; so during the neutralisation of weak bases, two reactions are taking place:
 - Step 1: the base dissociates to form more OH^- : $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
 - Step 2: the resulting OH^- reacts with the H^+ : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
 - the first reaction, the dissociation, is endothermic by around 5 kJmol^{-1} ; the second reaction, the neutralisation, has a molar enthalpy change of -58 kJmol^{-1} ; the overall molar enthalpy change of neutralisation is therefore -53 kJmol^{-1}
 - eg $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$ $\Delta H = -53 \text{ kJmol}^{-1}$
 - Weak bases have a less exothermic enthalpy of neutralisation than strong bases, because the OH^- ion has to be produced from the base first, which absorbs some of the energy



Practical 3.1: Compare the enthalpy of neutralisation of a strong base and a weak base by the same

acid

- 1) Place a clean, dry polystyrene cup inside a glass beaker.
- 2) Pour 25 cm³ of 2 mol dm⁻³ HCl into the cup using a measuring cylinder. Record the initial temperature of the solution.
- 3) Measure out 25 cm³ of 2 mol dm⁻³ NaOH using a measuring cylinder and record the initial temperature of the solution. Hence deduce the average initial temperature of the solutions.
- 4) Add the NaOH solution to the HCl solution and stir continuously, until the temperature either remains constant for over a minute or starts to change in the opposite direction. Record the final temperature reached.
- 5) Deduce the maximum temperature change.
- 6) Repeat steps 1 – 5, but this time use 25 cm³ of 2 mol dm⁻³ NH₃ instead of NaOH.
- 7) Which reaction gives you a greater temperature change? Can you explain why?

Equipment needed per group: two measuring cylinders (25 cm³), polystyrene cup (250 cm³), beaker (250 cm³), thermometer, access to 2.0 mol dm⁻³ HCl (50 cm³ per group), 2 mol dm⁻³ NH₃ (25 cm³ each per group) and 2 mol dm⁻³ NaOH (25 cm³ each per group)

Caution: the alkalis are corrosive at this molarity and should be handled with great care

- **The temperature increase should be around 7 °C with NaOH and around 5 °C with NH₃**
- **The temperature increase is higher with NaOH because the OH⁻ ions are already in the solution; NH₃ dissociates during the neutralisation to give OH⁻ ions; this process is endothermic so the overall neutralisation process is less exothermic**



Test your knowledge 3.2: Strong and Weak Acids and Bases

- (a) What is the difference between a strong acid and a weak acid? Give one example of each.
- (b) Give one example of a strong base, one example of a base which is weak due to low solubility and one example of a base which is weak due to small degree of ionisation
- (c) Why do strong acids conduct electricity better than weak acids?
- (d) Why is the enthalpy of neutralisation of weak acids less exothermic than the enthalpy of neutralisation of weak acids?

- (a) Strong acid fully dissociates in water to give H⁺ ions (eg H₂SO₄, HNO₃, HCl); weak acid only partially dissociates in water to give H⁺ ions (eg ethanoic acid etc)
- (b) Strong base: NaOH; weak base due to low solubility: Ca(OH)₂, weak base due to limited dissociation: NH₃
- (c) Strong acids are fully dissociated into ions, so the concentration of ions is greater, and the conductivity is due to the presence of ions
- (d) Weak acids must dissociate into ions during neutralisation; this process is endothermic, so the overall reaction is less exothermic

Lesson 4 - What are the other important reactions of acids, bases and salts?

d) Further Reactions of Acids, Bases and Salts

(i) Amphoterism

- Proton donors (acids) will react with proton acceptors (bases), usually to form salts; as well as the common acids (HCl, HNO₃ and H₂SO₄), which contain H⁺ ions, all oxides and hydroxides of non-metals (ie all covalently bonded oxides and hydroxides) can behave as acids; some dissolve in water to form solutions containing H⁺ ions; some do not dissolve in water but react with alkalis to form salts; examples of acidic oxides are CO₂, NO₂, P₄O₆, P₄O₁₀, SO₂, SO₃, Cl₂O
 - Eg SO₃ + H₂O → H₂SO₄ and SO₃ + 2NaOH → Na₂SO₄ + H₂O (salt = sodium sulphate)
 - Eg SiO₂ + 2NaOH → Na₂SiO₃ + H₂O (salt = sodium silicate)
- Proton acceptors (bases) will react with proton donors (acids), usually to form salts; most oxides and hydroxides of metals (ie all ionically bonded oxides and hydroxides) can behave as bases; some dissolve in water to form solutions containing OH⁻ ions; some do not dissolve in water but react with acids to form salts:
 - Eg Na₂O + H₂O → 2NaOH and Na₂O + HCl → 2NaCl + H₂O
 - Eg Mg(OH)₂ + 2HCl → MgCl₂ + 2H₂O
- There are a small number of oxides and hydroxides of metals which can form salts by reacting with both acids and alkalis:
 - Eg Al(OH)₃ + 3HCl → AlCl₃ + 3H₂O and Al(OH)₃ + NaOH → NaAl(OH)₄
 - Eg ZnO + 2HCl → ZnCl₂ + H₂O and ZnO + 2NaOH → Na₂ZnO₂ + H₂O
 These substances can react with both acids and bases; they are proton donors and proton acceptors and are said to be **AMPHOTERIC**; the bonding in these oxides and hydroxides tends to be intermediate between ionic and covalent
(The acid-base properties of oxides and hydroxides will be studied in more detail later in the course)
- Other substances can show amphoteric properties by behaving as both proton donors and proton acceptors; one of the most common examples is the hydrogencarbonate ion (HCO₃⁻)

$$\text{NaHCO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{CO}_2 + \text{H}_2\text{O} \text{ and } \text{NaHCO}_3 + \text{NaOH} \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$$

(ii) Acid-Salt Reactions

- If a strong acid reacts with a salt of a weaker acid, the stronger acid will displace the weaker acid from its salt, by forcing the salt of the other acid to accept the hydrogen ion
- For example, sulphuric acid is a stronger acid than hydrochloric acid; as a result, if concentrated sulphuric acid is reacted with any salt of hydrochloric acid (eg potassium chloride), the sulphuric acid will give its hydrogen ion to the chloride ion, forming hydrochloric acid and a sulphate salt (eg potassium sulphate);

$$\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$$
- In some cases, a weak acid will even react with the salt of a stronger acid; for example, phosphoric acid can react with chloride salts to make phosphate salts, even though phosphoric acid is a weaker acid than hydrochloric acid; the reaction is possible because the HCl escapes as a gas, preventing the reverse reaction and pushing the position of equilibrium to the right according to Le Chatelier's principle:

$$\text{H}_3\text{PO}_4(\text{l}) + 3\text{KCl}(\text{s}) \rightarrow \text{K}_3\text{PO}_4(\text{s}) + 3\text{HCl}(\text{g})$$
- In acid-salt reactions, the anion in the salt is accepting a proton from the acid; it is therefore acting as a base; acid-salt reactions are just a special type of acid-base reaction

Unit 5 – Acids, Bases and Salts

(iii) Salt Hydrolysis

- Salts can be classified into three main types:
 - salts formed from strong acids and strong bases (eg NaNO_3 , K_2SO_4 and BaCl_2)
 - salts formed from weak acids and strong bases (eg CH_3COONa , KCN)
 - salts formed from strong acids and weak bases (eg NH_4Cl , NH_4NO_3)Salts can also be formed from weak acids and weak bases, but these tend not to be stable
- Anions produced from weak acids can react partially with the water in the solution to reform the original acid: eg $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$
 - the resulting solution, therefore, contains OH^- ions as well as the salt
 - this type of reaction occurs in salts made from weak acids; if the acid is reluctant to give up its H^+ in the first place, it is more likely that the anion will pick up an H^+ to reform the acid
 - salts containing anions which do this are known as basic salts; this is an example of salt hydrolysis
 - examples of basic salts are CH_3COONa and KCN ; they are made from weak acids and strong bases
- Cations produced from weak bases in the salt can react partially with the water in the solution to reform the original base: eg $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$
 - the resulting solution, therefore, contains H^+ ions as well as the salt
 - this type of reaction occurs in salts made from weak bases; if the base is reluctant to accept an H^+ ion in the first place, its cation is more likely release an H^+ and reform the original base
 - salts which do this are known as **acidic salts**; this is another example of salt hydrolysis
 - examples of acidic salts are NH_4Cl and NH_4NO_3 ; they are made from strong acids and weak bases
- Some highly charged cations also react with water to form H^+ ions, in order to reduce their charge; they are accepting OH^- ions from the water and partly reforming the original hydroxide:
 - eg $\text{Al}^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Al}(\text{OH})^{2+} + \text{H}^+$
 - the resulting solution contains H^+ ions as well as the salt
 - salts which do this are also known as acidic salts; this is another example of salt hydrolysis
 - an example of this type of acidic salt is AlCl_3 or CrCl_3
- Anions made from strong acids and cations made from strong bases do not undergo salt hydrolysis; salts made from strong acids and strong bases are therefore known as neutral salts
- The tendency of a salt to itself react with water and show acidic or basic properties is known as **salt hydrolysis**


Test your knowledge 4.1: Further Reactions of Acids, Bases and Salts

- (a) What is meant by the term “amphoteric”; give two examples of amphoteric substances
- (b) Write equations for the following reactions:
- sulphuric acid and potassium chloride
 - nitric acid and sodium fluoride
 - phosphoric acid (H_3PO_4) and potassium bromide
- (c) Predict whether the following salts will undergo salt hydrolysis; if they do, write an equation to show how one of the ions in the salt reacts with water to produce H^+ or OH^- ions
- sodium nitrate
 - ammonium nitrate
 - sodium ethanoate
 - potassium sulphate
 - aluminium sulphate
 - potassium cyanide

- (a) reacts with acids and bases/can behave as an acid or a base, eg NaHCO_3 or $\text{Al}(\text{OH})_3$ or ZnO etc
- (b) (i) $\text{H}_2\text{SO}_4 + 2\text{KCl} \rightarrow \text{K}_2\text{SO}_4 + 2\text{HCl}$; (ii) $\text{HNO}_3 + \text{NaF} \rightarrow \text{NaNO}_3 + \text{HF}$; (iii) $\text{H}_3\text{PO}_4 + 3\text{KBr} \rightarrow \text{K}_3\text{PO}_4 + 3\text{HBr}$
- (c) (i) no; (ii) yes $\text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}^+$; (iii) yes $\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{OH}^-$; (iv) no; (v) yes $\text{Al}^{3+} + \text{H}_2\text{O} \rightarrow \text{Al}(\text{OH})^{2+} + \text{H}^+$; (vi) yes $\text{CN}^- + \text{H}_2\text{O} \rightarrow \text{HCN} + \text{OH}^-$

Lesson 5 – How can we use acid-base reactions to prepare salts in the laboratory?

e) Laboratory Preparation of Salts by Neutralisation

- The most suitable method of preparing a salt by a neutralisation reaction depends on the solubility of the salt and the solubility of the acid and the base used to make it; insoluble salts are generally not prepared by neutralisation reactions, but soluble salts frequently are; there are two common types of preparation:
- Soluble salts can be prepared from **an acid and an insoluble base**:
 - if the salt being prepared is soluble, the easiest way to make it is from an acid and an insoluble base
 - an excess of the insoluble base must be added to ensure that the acid reacts completely
 - the salt solution can be separated from the excess base by simple filtration
 - the salt can be extracted from the salt solution by gradually evaporating off the water
 - copper sulphate is an example of a salt which can be prepared by this method



Practical 5.1: Prepare a salt by neutralising an acid with an insoluble base

- 1) Measure out 20 cm³ of 0.5 moldm⁻³ sulphuric acid into a 100 cm³ beaker.
- 2) Use a Bunsen burner to warm the beaker gently on a tripod until the temperature reaches 50 °C.
- 3) Add 1 g of copper oxide gradually to the beaker over a period of 2 minutes, stirring gently.
- 4) Heat gently for a few minutes until it is almost boiling.
- 5) Allow the mixture to cool.
- 6) Place a folded piece of filter paper inside a filter funnel, and then place the funnel into the neck of a 250 cm³ conical flask.
- 7) Pour the warm mixture into the filter funnel and allow the solution to filter through. A clear blue solution should collect in the conical flask.
- 8) Rinse the beaker and then pour the clear blue solution back into it. Label the beaker with your name and leave it for a week.
- 9) Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.
 - How many moles of sulphuric acid (H₂SO₄) are used in the reaction?
 - How many moles of copper oxide (CuO) are used in the reaction?
 - Why is it important that the CuO is in excess?

Equipment needed per group: measuring cylinder (25 cm³), beaker (250 cm³), Bunsen burner, tripod, gauze, thermometer, weighing boat, spatula, stirrer, filter paper, funnel, conical flask (250 cm³), access to a mass balance, access to 0.5 moldm⁻³ H₂SO₄ (20 cm³ per group), access to CuO (1 g per group)

- $0.02 \times 0.5 = 0.01$
- $1/79.5 = 0.0126$
- **To ensure the H₂SO₄ fully reacts; excess CuO can be removed by filtration; it would be much more difficult to remove excess H₂SO₄**

Cannot do this experiment? Watch this video: www.youtube.com/watch?v=qlOMlwBoe_4

- Group I and ammonium salts cannot be prepared by neutralisation of insoluble bases as ammonia and all Group I bases are soluble; these salts can only be prepared by reacting acids and alkalis
 - the acid and alkali must be added in the exact amounts required, as there is no way to separate the salt produced from any remaining acid or alkali; the final mixture should therefore only contain the salt and water
 - the salt can be extracted from the salt solution by gradually evaporating off the water
 - ammonium sulphate is an example of a salt which can only be prepared by this method



Practical 5.2: Prepare a salt by neutralising an acid with a soluble base

- 1) Measure out 25 cm³ of 2 moldm⁻³ ammonia solution into a 100 cm³ beaker.
- 2) Add 25 cm³ of 1 moldm⁻³ sulphuric acid solution to the same beaker. Stir the mixture gently.
- 3) Pour the resulting mixture into an evaporating dish and heat over a Bunsen burner until most (but not all) of the water has evaporated.
- 4) Leave the remaining solution for a week.
- 5) Use a spatula to remove any crystals in the beaker and dry them with a piece of filter paper.
 - Write an equation for the reaction between sulphuric acid and ammonia.
 - How many moles of sulphuric acid are used in the reaction?
 - How many moles of ammonia are used in the reaction?
 - Why is it important to use these exact quantities of the reactants?
 - Why is it easier to make a pure salt if the base is insoluble rather than soluble?

Equipment needed per group: two measuring cylinders (25 cm³), beaker (100 cm³), stirring rod, Bunsen burner, tripod, gauze, evaporating dish, spatula, filter paper access 1.0 moldm⁻³ H₂SO₄ (20 cm³ per group), and 2.0 moldm⁻³ NH₃ (25 cm³ each per group)



- $0.025 \times 1 = 0.025$

- $0.025 \times 2 = 0.05$

- to ensure that both reactants are fully used up and that there is no excess of either reactant in the solution at the end

- excess insoluble base can be removed easily by filtration; excess soluble base cannot be easily removed

Cannot do this experiment? Watch this video: <https://www.youtube.com/watch?v=Yh7z2ThG32o>

- Salts can also be prepared from acid-salt reactions; the salts most commonly prepared in this way are phosphate salts, by reacting concentrated H₃PO₄ with chloride salts:
Eg $3\text{CaCl}_2(\text{s}) + 2\text{H}_3\text{PO}_4(\text{l}) \rightarrow \text{Ca}_3(\text{PO}_4)_2(\text{s}) + 6\text{HCl}(\text{g})$

Lesson 6 – How are acid-base reactions useful in qualitative analysis?

f) Qualitative Analysis Part 1a: Acid-Base Reactions

- Qualitative analysis is the experimental identification of a substance or of a particular species present in a substance; this can include cations, anions, gases or organic molecules
- Some cations, anions and gases can be identified using simple acid-base reactions:
 - H⁺ ions with CO₃²⁻ and SO₃²⁻ ions**
 - H⁺ ions react with CO₃²⁻ and HCO₃⁻ ions to give carbon dioxide and water:
 $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$; $\text{H}^+ + \text{HCO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
 - H⁺ ions react with SO₃²⁻ and HSO₃⁻ ions to give sulphur dioxide and water:
 $2\text{H}^+ + \text{SO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$; $\text{H}^+ + \text{HSO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$
 - if a solid sample of a carbonate (eg CaCO₃) is added to a solution and bubbles/fizzing (also known as effervescence) is evolved, it shows that the solution contains H⁺ ions
 - if an acid such as HCl is added to a solid sample and effervescence is observed, then the gas should be tested; if the gas is shown to be CO₂ then CO₃²⁻ or HCO₃⁻ ions are present in the solid; if the gas is shown to be SO₂ then SO₃²⁻ or HSO₃⁻ ions are present in the solid
 - CO₂ and SO₂ with limewater**
 - CO₂ reacts with Ca(OH)₂ (limewater) to produce insoluble CaCO₃, which turns the limewater a milky colour: $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
 - SO₂ reacts with Ca(OH)₂ (limewater) to produce insoluble CaSO₃, which also turns the limewater a milky colour: $\text{SO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaSO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
 - if either gas is bubbled through the limewater for a longer period of time, a second reaction will happen and the milkiness will disappear:
 $\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$
 $\text{CaSO}_3(\text{s}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HSO}_3)_2(\text{aq})$
 - CO₂ and SO₂ are the only common gases which have this effect on limewater
 - Note: it is easy to distinguish between CO₂ and SO₂; CO₂ is odourless but SO₂ smells like a burning match

Unit 5 – Acids, Bases and Salts

(iii) NH_4^+ ions with OH^- ions

- NH_4^+ ions react with OH^- ions to produce NH_3 ; $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- when warmed gently, the NH_3 is given off as a pungent-smelling gas
- if NaOH is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains NH_4^+ ions
- if NH_4Cl is added to a solution and the mixture is warmed, and as a result a pungent smell is observed, then the solution contains OH^- ions

(iv) NH_3 with HCl

- NH_3 (an alkaline gas) and HCl (an acidic gas) react together to give a salt (NH_4Cl):
 $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$
- this salt is clearly visible as a white smoke which appears when the gases come into contact with each other
- If a piece of filter paper is soaked in concentrated HCl and placed close to a source of NH_3 , a white smoke can be seen, confirming the presence of NH_3
- If a piece of filter paper is soaked in concentrated NH_3 and placed close to a source of HCl, a white smoke can be seen, confirming the presence of HCl

(v) Concentrated H_2SO_4 and Cl^- ions

- If concentrated sulphuric acid is added to a solid sample containing Cl^- ions, an acid-salt reaction will take place and HCl (g) will be produced: $\text{H}_2\text{SO}_4 + \text{Cl}^- \rightarrow \text{HSO}_4^- + \text{HCl}$
- If a piece of filter paper is soaked in concentrated NH_3 and placed close to a source of HCl, a white smoke can be seen, confirming the presence of HCl


Practical 6.1: Qualitative Analysis Part 1a - use acid-base reactions to identify cations, anions and gases

You are given four solid samples to test, labelled A, B, C and D; each solid contains a different one of the following ions: NH_4^+ , OH^- , CO_3^{2-} , SO_3^{2-}

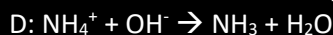
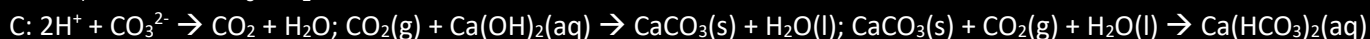
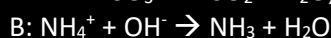
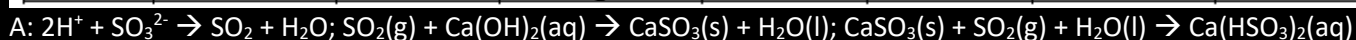
- Place a small spatula of sample A onto a watch glass; add a few drops of HCl and observe whether bubbles are formed; if bubbles are formed, observe whether or not the gas smells like a burning match; repeat for samples B, C, and D
- If bubbles are observed, half-fill a test-tube with limewater; add a spatula measure of the same sample into another test-tube; add 3 cm³ of HCl to the sample, attach a bung with a delivery tube to the test tube and insert the delivery tube into the limewater in the other test tube; wait for a few minutes until there is no further change in the limewater
- Place a small spatula of sample A onto a watch glass; add a few drops of NaOH; observe whether or not a pungent smell is given off; repeat for samples B, C and D
- Place a small spatula of sample A onto a watch glass; add a few drops of NH_4Cl ; observe whether or not a pungent smell is given off; repeat for samples B, C and D
- Place a very small spatula of sample A into a test tube and add 2 – 3 cm³ of water so that the solid is mostly dissolved; then add a small spatula measure of powdered CaCO_3 ; observe whether bubbles are formed; repeat for samples B, C and D
- Use your observations to identify which ion is present in which sample; present your answers in a table similar to the one below:

Sample	Bubbles with HCl? Y/N	Smell with HCl?	Effect of gas on limewater?	Smell with NaOH?	Smell with NH_4Cl ?	Bubbles with CaCO_3 ? Y/N	Ion present
A							
B							
C							
D							

- Write equations for the reaction occurring with each sample which results in a positive identification

Equipment needed per group: three watch glasses, four test tubes, test tube rack, bung to fit test tube with delivery tube connected; access to solid samples of: Na_2SO_3 (labelled A), KOH (labelled B), Na_2CO_3 (labelled C), NH_4Cl (labelled D), each with their own spatula; access to 1 moldm⁻³ HCl with dropping pipette (10 cm³ per group), limewater (10 cm³ per group), 1 moldm⁻³ NaOH (5 cm³ per group) with dropping pipette, 1 moldm⁻³ NH_4Cl (5 cm³ per group) with dropping pipette, powdered CaCO_3 (2 g per group) with its own spatula

Sample	Bubbles with HCl? Y/N	Smell with HCl?	Effect of gas on limewater?	Smell with NaOH?	Smell with NH_4Cl ?	Bubbles with CaCO_3 ? Y/N	Ion present
A	Y	Burning match	Milky then clear again	none	none	N	SO_3^{2-}
B	N (but it might get warm)	none	n/a	none	pungent	N	OH^-
C	N	none	n/a	none	none	Y	H^+
D	Y	none	Milky then clear again	none	none	N	CO_3^{2-}




Test your knowledge 6.2: Using acid-base reactions to identify certain cations and anions

Describe a simple test to show how you would confirm the presence of the following ions:

- CO_3^{2-} in a solid sample
- SO_3^{2-} in a solid sample
- Cl^- in a solid sample
- H^+ in solution
- OH^- in solution
- NH_4^+ in solution
- NH_3 as a gas
- HCl as a gas

In each case, give the equations for the reactions occurring

- add HCl (aq); odourless gas evolved which turns limewater milky and then clear again:
 $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$; $\text{CO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$; $\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$
- add HCl (aq); gas evolved with burning-match smell which turns limewater milky and then clear again: $2\text{H}^+ + \text{SO}_3^{2-} \rightarrow \text{SO}_2 + \text{H}_2\text{O}$; $\text{SO}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{CaSO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$; $\text{CaSO}_3(\text{s}) + \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HSO}_3)_2(\text{aq})$
- add concentrated H_2SO_4 ; if gas given off, test with filter paper soaked in concentrated NH_3 ; a white smoke should be seen: $\text{H}_2\text{SO}_4 + \text{Cl}^- \rightarrow \text{HSO}_4^- + \text{HCl}$; $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$
- add $\text{CaCO}_3(\text{s})$; gas evolved: $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- add $\text{NH}_4\text{Cl}(\text{aq})$ and warm; pungent smell given off: $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- add $\text{NaOH}(\text{aq})$ and warm; pungent smell given off: $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$
- test with filter paper soaked in concentrated HCl ; a white smoke should be seen: $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$
- test with filter paper soaked in concentrated NH_3 ; a white smoke should be seen: $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$

Lesson 7 - What is the pH scale?

g) Indicators and the pH Scale

(i) The pH Scale

- Water dissociates very slightly to produce H^+ and OH^- ions: $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$; as a result, all aqueous solutions contain both H^+ and OH^- ions
- Solutions which contain acids contain many more H^+ ions than pure water; as a result, these solutions also contain fewer OH^- ions than pure water, because the extra H^+ ions push the above equilibrium to the left; the higher the concentration of H^+ ions, the lower the concentration of OH^- ions; any solution in which the concentration of H^+ ions is greater than the concentration of OH^- ions is said to be an **acidic** solution
- Solutions which contain alkalis contain many more OH^- ions than pure water; as a result, these solutions also contain fewer H^+ ions than pure water, because the extra OH^- ions push the above equilibrium to the left; the higher the concentration of OH^- ions, the lower the concentration of H^+ ions; any solution in which the concentration of OH^- ions is greater than the concentration of H^+ ions is said to be an **alkaline** solution
- A solution in which the concentration of H^+ ions is equal to the concentration of OH^- ions, as is the case in pure water, is said to be a neutral solution

Unit 5 – Acids, Bases and Salts

- The level of acidity or alkalinity of a solution (ie the relative concentrations of H^+ and OH^- ions is measured on a scale called the **pH scale**:
 - the lower the pH, the more acidic the solution (the higher the concentration of H^+ ions and the lower the concentration of OH^- ions); the most acidic solutions possible have a pH of around -1
 - the higher the pH, the more alkaline the solution (the higher the concentration of OH^- ions and the lower the concentration of H^+ ions). The most alkaline solutions possible have a pH of around 15
 - Neutral solutions have a pH of 7
- The relationship between pH, acidity and alkalinity is summarised in the table below:

pH	-1	1	3	5	7	9	11	13	15
Acidity	strongly acidic		slightly acidic		Neutral	slightly alkaline		Strongly alkaline	
$[H^+]$	very high		quite high		Normal	quite low		very low	
$[OH^-]$	very low		quite low		Normal	quite high		very high	

- Examples of the pH of common solutions are:
 - 1 mol dm⁻³ HCl has a pH of 0
 - lemon juice and vinegar have a pH of approx. 2
 - orange juice has a pH of approx. 3
 - rainwater has a pH of approx. 6
 - pure water has a pH of 7
 - household bleach has a pH of approx. 12
 - 1 mol dm⁻³ NaOH has a pH of 14
- The pH of a solution can be measured directly using a **pH meter** (although these can be difficult to use)



Test your knowledge 7.1: Understanding the pH scale

Classify the following solutions as acidic, alkaline or neutral based on the information provided:

(a) it has a pH of 5 (b) it is a solution of ammonia (c) $[H^+] = [OH^-]$ (d) it is lemon juice (e) it has a pH of 11 (f) $[H^+] > [OH^-]$ (g) It is a solution of sodium hydroxide (h) It is a solution of sodium chloride	(i) it is a solution of sodium carbonate (j) it has a pH of 7 (k) it has a pH of -1 (l) $[OH^-] > [H^+]$ (m) it is rainwater (n) it is pure water (o) it is a solution of ammonium chloride (p) it is a solution of potassium ethanoate
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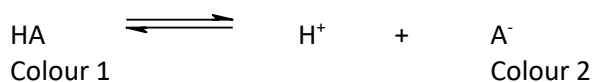
(a) acidic; (b) alkaline; (c) neutral; (d) acidic; (e) alkaline; (f) acidic; (g) alkaline; (h) neutral; (i) alkaline; (j) neutral; (k) (very) acidic; (l) alkaline; (m) acidic; (n) neutral; (o) acidic (due to salt hydrolysis); (p) alkaline (due to salt hydrolysis)

(i) Indicators

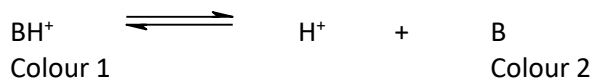
- An indicator is a substance which appears as one colour in acidic solutions and as a different colour in alkaline solutions; indicators can therefore be used to determine whether a solution is acidic or alkaline

Unit 5 – Acids, Bases and Salts

- Most indicators are weak acids in which the acid and its anion have different colours:



Some indicators are weak bases in which the base and its cation have different colours:



- at low pH, the H^+ ions shift the position of equilibrium to the left and the indicator will appear as colour 1
- at high pH, the $[\text{H}^+]$ is low so the equilibrium shifts to the right to produce more H^+ and the indicator will appear as colour 2
- **hence indicators therefore appear as one colour in low pH, and a different colour in high pH**

- The pH range over which an indicator changes colour varies from indicator to indicator:

Indicator	Colour 1 (low pH)	Colour 2 (high pH)	pH of colour change
Methyl orange	Pink	Yellow	3 – 5
Litmus	Red	Blue	5 – 8
Phenolphthalein	Colourless	Pink	8 – 10

- The effect of acids and bases on indicators depends on the exact pH of the solution, and the pH at which the indicator changes colour:
 - Strongly acidic solutions with a pH of 3 or below will cause all three indicators to appear as Colour 1
 - Weakly acidic solutions with a pH of 5 – 6 will cause methyl orange to appear as Colour 2 but will cause litmus and phenolphthalein to appear as colour 1
 - Neutral solutions (pH = 7) will cause methyl orange to appear as Colour 2, litmus to appear as an intermediate colour (or whatever colour it was before) but will cause phenolphthalein to appear as colour 1
 - Alkaline solutions with a pH of 10 or more will cause all three indicators to appear as Colour 2
- Single indicators, therefore, cannot give you the exact pH of a solution; they will simply tell you whether the pH is above or below a certain number



Practical 7.2: Investigate the effect of solutions of different pH values on different indicators

You have access to five different solutions, with pH values of 3, 5, 7, 9 and 11 and labelled accordingly.

You have access to three different indicators: litmus, methyl orange and phenolphthalein

- Pour 1 cm³ of the solution of pH 3 into each of three different test tubes
- Add two drops of methyl orange to the first test tube and note the colour; add two drops of litmus to the second test tube and note the colour; add two drops of phenolphthalein to the third test tube and note the colour
- Repeat steps (a) and (b) for the solutions with pH values of 5, 7, 9 and 11
- Record your results in a table similar to the one below:

pH of solution	methyl orange colour	litmus colour	Phenolphthalein colour
3			
5			
7			
9			
11			

- Deduce the approximate pH at which each of the three indicators changes colour

Unit 5 – Acids, Bases and Salts

Equipment needed per group: five test tubes and one test tube rack; access to: 0.1 mol dm⁻³ CH₃COOH (labelled pH 3); a solution containing 0.1 mol dm⁻³ CHCOOH and 0.5 mol dm⁻³ CHCOO⁻Na⁺ (labelled pH 5); distilled water (labelled pH = 7); a solution containing 0.1 mol dm⁻³ NH₃ and 0.5 mol dm⁻³ NH₄Cl (labelled pH 9); 0.1 mol dm⁻³ NH₃ (labelled pH = 11); 5 cm³ per group for each solution; each bottle should have its own dropping pipette; access to solutions of phenolphthalein, methyl orange and litmus, each with their own dropping pipette (1 cm³ per group)

pH of solution	methyl orange colour	litmus colour	Phenolphthalein colour
3	pink	red	colourless
5	yellow	red	colourless
7	yellow	red/blue	colourless
9	yellow	blue	pink/colourless
11	yellow	blue	pink

- MO changes colour at pH = 5; L changes colour at pH 7; PP changes colour at pH 9

Lesson 8 - What is universal indicator and why is it useful?

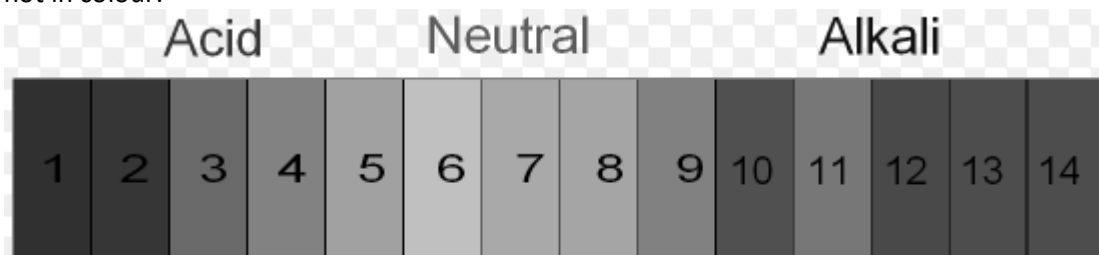
(ii) Universal Indicator

- Universal indicator is a mixture of several different indicators, each of which changes colour at a different pH; as a result, universal indicator appears at a different colour for very pH unit between 3 and 11:
 - at pH values of 3 and below, universal indicator appears red
 - between the pH values of 3 and 5, universal indicator appears orange
 - between the pH values of 5 and 6, universal indicator appears yellow
 - between the pH values of 6 and 8, universal indicator appears green
 - between the pH values of 8 and 9, universal indicator appears blue
 - between the pH values of 9 and 11, universal indicator appears indigo
 - at pH values of 11 and above, universal indicator appears violet
- Universal indicator is most useful for finding out the approximate pH of aqueous solutions; it can be used in paper form or in solution form; in most cases, a colour chart is also provided and the colour of the indicator can be compared to the chart to deduce the pH of the solution; it is not an accurate method of determining pH, but is very simple and easy to use
- Using universal indicator to measure pH is an example of **colorimetry** – the use of colour to make measurements
- pH meters measure pH more accurately, but they are not as simple or convenient to use



Activity 8.1: Universal Indicator and pH

The following chart shows the different colours shown by universal indicator in solutions with different pH values, but it is not in colour:



Copy the table above, but use coloured pencils, pens or crayons to make sure that each box is shaded in a colour which matches that shown by universal indicator at that pH.

Students will need access to the different colours shown by universal indicator.

Answer: 1 – 3 red; 4 - 5 orange; 6 – yellow; 7 – green; 8 – blue; 9 – 11 indigo; 12 – 14 violet



Extension 8.2: Universal Indicator and pH

Completed Activity 20.1? Underneath your chart, give an example of a common substance with a pH similar to that shown.

Eg pH 1 stomach acid (HCl); pH 2 lemon juice/vinegar; pH 3 orange juice; pH 4 tomato juice; pH 5 black coffee; pH 6 milk; pH 7 water; pH 8 sea water; pH 9 a solution of baking soda; pH 10 milk of magnesia; pH 11 ammonia solution; pH 12 soapy water; pH 13 bleach



Practical 8.3: Determine the pH value of various solutions by colorimetry

You have access to five different solutions, labelled A, B, C, D and E. Each solution has a different pH.

- 1) Using a dropping pipette, place around 1 cm³ of solution A into a test tube
- 2) Add three drops of universal indicator and note the colour
- 3) Repeat Steps 1 – 3 using solutions B, C, D and E
- 4) Record your results in the table below; state the pH of each solution and suggest what the solution might be.

Solution	Colour it turns UI	pH	Possible identity
A			
B			
C			
D			
E			

Equipment needed per group: five test tubes and one test tube rack; access to: rainwater (labelled A) (you can ensure a PH of 5 by bubbling CO₂ through it); distilled water (labelled B); a diluted solution of bleach (C); a solution of baking soda (D); diluted vinegar or lemon juice (E); the exact identity of the solution is not important but they should turn UI the following colours: A – orange/yellow ; B – green; C – violet; D – blue/indigo; E - red
5 cm³ per group for each solution; each bottle should have its own dropping pipette; access to universal indicator solution with dropping pipette (1 cm³ per group)

Solution	Colour it turns UI	pH	Possible identity
A	orange/yellow	5	rainwater (or urine)
B	green	7	water
C	violet	11	bleach
D	blue/indigo	9	soda or magnesia
E	red	3	lemon juice or vinegar



Practical 8.4: Determine the pH of different soil samples

You are given access to three different soil samples, taken from somewhere close to where you live.

- 1) Take 20 g of the soil to be tested and place it in a 100 cm³ beaker.
- 2) Add 40 cm³ of water to the beaker using a measuring cylinder.
- 3) Leave the beaker for around 30 minutes, stirring every few minutes.
- 4) Fold and insert a piece of filter paper into a funnel and pour the mixture through the funnel into a boiling tube. When around 10 cm³ of filtrate has been collected, add a few drops of UI solution and note the colour.
- 5) Repeat for any other soil samples you need to test.
- 6) Record your results in a suitable table.

Equipment needed per group: one beaker (100 cm³); one measuring cylinder (50 cm³); one stirring rod; one funnel; three pieces of filter paper; one boiling tube; access to three different soil samples (20 g per group), each with its own spoon; access to a mass balance; access to universal indicator with its own dropping pipette(1 cm³ per group)

Note: the soil samples should ideally cover a range of different types, ideally with a range of different acidities.

Lesson 9 – How can indicators be used in qualitative analysis?

h) Qualitative Analysis Part 1b – Using Indicators

- Indicators can be used in qualitative analysis to identify the presence of H⁺ ions or OH⁻ ions in a solution; the best indicator to use if you simply want to determine whether a solution is acidic, alkaline or neutral is litmus; this is because litmus changes colour at a pH of around 7; it is usually used soaked in paper, as **litmus paper**; litmus paper can be red or blue; red litmus paper contains litmus in Colour 1 form and blue litmus paper contains litmus in Colour 2 form
 - acidic solutions (containing H⁺) will turn blue litmus paper red but will have no effect on red litmus paper
 - alkaline solutions (containing CO₃²⁻ or HCO₃⁻) will turn red litmus paper blue but will have no effect on blue litmus paper
 - neutral solutions will have no effect on either type of litmus paper
- Litmus paper can also be used to test for certain gases which have acidic or alkaline properties:
 - HCl and SO₂ are acidic gases which turn blue litmus paper red
 - NH₃ is an alkaline gas which turns red litmus paper blue


Practical 9.1: Investigate the effect of acidic, alkaline and neutral solutions on litmus paper

You have access to four different solutions: HCl(aq); NaOH(aq), Na₂CO₃(aq) and NaCl(aq)

You have access to three different indicators: litmus, methyl orange and phenolphthalein

- Pour 5 cm³ of HCl(aq) into a beaker
- Dip a strip of red litmus paper into the solution; note any colour change
- Dip a strip of blue litmus paper into the solution; note any colour change
- Rinse out the beaker and measuring cylinder and repeat steps (a) – (c) using the other solutions (NaOH, Na₂CO₃ and NaCl)
- Pour 1 cm³ of each solution (HCl, NaOH, Na₂CO₃ and NaCl) separately into four test tubes
- Add two drops of methyl orange indicator to each test tube; note any colour change
- Rinse out the test tubes and repeat steps (e) and (f), but this time add two drops of phenolphthalein; note any colour changes
- Record your results in a table similar to the one below:

Solution	Effect on red litmus paper	Effect on blue litmus paper	Effect on methyl orange	Effect on phenolphthalein
HCl				
NaOH				
Na ₂ CO ₃				
NaCl				

Equipment needed per group: one small beaker (<50 cm³) and one small measuring cylinder (10 cm³); four strips of red litmus paper, four strips of blue litmus paper, access to solutions of 0.1 mol dm⁻³ HCl, NaOH, Na₂CO₃ and NaCl, access to phenolphthalein and methyl orange indicator solutions, 8 test tubes, 1 test tube rack

Solution	Effect on red litmus paper	Effect on blue litmus paper	Effect on methyl orange	Effect on phenolphthalein
HCl	stays red	turns red	pink	colourless
NaOH	turns blue	stays blue	yellow	pink
Na ₂ CO ₃	turns blue	stays blue	yellow	pink
NaCl	stays red	stays blue	Yellow	colourless


Test your knowledge 9.2: Summary of Qualitative Analysis of Cations, Anions and Gases

- Describe a test for H⁺ ions which uses a simple chemical reaction
- Describe a test for H⁺ ions which uses an indicator
- Describe a test for OH⁻ which uses a simple chemical reaction and a smell
- Describe a test for OH⁻ which uses an indicator
- Name three gases which turn blue litmus paper red
- Name one gas which turns red litmus paper blue
- Name two gases which turn limewater milky and then clear again and suggest how the gases could be distinguished
- Describe a test for HCl gas which does not use an indicator
- Describe a test for NH₃ gas which does not use an indicator
- Describe a test for Cl⁻ ions in a solid sample
- Describe a test for NH₄⁺ ions
- Describe a test for CO₃²⁻ ions in a solid sample
- Describe a test for SO₃²⁻ ions in a solid sample
- Suggest how you could demonstrate that a solution contains a neutral salt

(a) Add $\text{CaCO}_3(\text{s})$ - effervescence; (b) add blue litmus paper – it turns red; (c) add NH_4Cl and warm – a pungent smell; (d) add red litmus paper – it turns blue; (e) hydrogen chloride, sulphur dioxide and nitrogen dioxide; (f) ammonia; (g) carbon dioxide and sulphur dioxide, SO_2 smells of burning matches and also turns blue litmus red; (h) hang filter paper soaked in concentrated NH_3 close to the source – white smoke formed; (i) hang filter paper soaked in concentrated HCl close to the source – white smoke formed; (j) add concentrated H_2SO_4 and then test the gas either with filter paper soaked in concentrated NH_3 – white smoke formed, or with blue litmus paper – it turns red; (k) add NaOH and warm – pungent smell; (l) add HCl and bubble gas through limewater – odourless gas which turns limewater milky and then clear again; (m) add HCl and bubble gas through limewater – burning-match-smelling gas which turns limewater milky and then clear again and also turns blue litmus paper red; (n) conducts electricity which proves that ions are present; turns universal indicator green which shows that H^+ , OH^- and CO_3^{2-} are not present

Lesson 10 – How I determine how much of an acid or a base is present in a sample (practical)?

(i) Quantitative Analysis - Acid-Base Titrations

- **Quantitative analysis** is the experimental determination of how much of a substance is present in a sample; it is possible to carry out the quantitative analysis of acids and bases using a technique called **titration**
- A **titration** is an experiment used to find out what exact volume of one solution is required to react completely with a fixed volume of another solution in a flask; it is the most common example of **volumetric analysis** (a type of quantitative analysis which measures volumes); from the titration result, you can deduce the concentration of the alkali if the concentration of the acid is known, or vice versa



Summary Activity 10.1: What can you remember about the different instruments used to measure the volume of a solution?

- What are the names of the different instruments used to measure the volume of solution?
- What are the advantages and disadvantages of each instrument?
- Which instruments would be most useful when carrying out a titration?

- Pipette: very accurate but can only deliver one volume
- Volumetric flask: very accurate but can only store one volume
- Burette: slightly less accurate than a pipette but can deliver any volume up to 50 cm^3
- Measuring cylinder: not accurate
- Pipettes and burettes are most useful for carrying out titrations

- In most titrations, a fixed volume (usually 25 cm^3) of alkali is prepared in a conical flask using a pipette and an indicator added to the alkali; the acid is then added gradually from the burette into the conical flask until the indicator changes colour (it is better to place the acid in the burette as alkalis can leave residue in the burette which makes it harder to use)
- Burettes are calibrated to measure the volume of solution delivered – the readings on the burette do not tell you the volume of solution actually in the burette; you must first record the burette reading before any solution has been delivered, and then record it again after the reaction is complete; the difference between these two volumes is the volume of solution delivered; this is also known as the titre volume: titre volume = final burette reading – initial burette reading

Unit 5 – Acids, Bases and Salts

- It is important to choose the indicator correctly:
 - the point at which the alkali has been completely neutralised by the acid is called the **equivalence point** of the titration, and the pH of the solution changes sharply at this point; the precise pH change at the end point depends on the type of titration being carried out
 - different indicators change colour at different pH values; for example methyl orange changes colour at pH 3 -5 and phenolphthalein changes colour at pH 8 – 10; the pH range over which the indicator changes colour is called the **end-point** of the indicator

Indicator	Colour 1 (low pH)	Colour 2 (high pH)	End point pH range
Methyl orange	Red	Yellow	3 - 5
Phenolphthalein	Colourless	Pink	8 - 10

- an indicator will only change colour at the equivalence point of a titration if the end-point pH range of the indicator lies within the pH change at the equivalence point of the titration

Type of titration	pH change at equivalence point	methyl orange suitable?	phenolphthalein suitable?
strong acid - strong base	11 to 3	Yes	yes
weak acid - strong base	11 to 7	No	Yes
strong acid - weak base	7 to 3	Yes	No
weak acid-weak base	No sharp change	No	No

- either methyl orange or phenolphthalein can be used for strong acid-strong base titrations
- only phenolphthalein can be used for weak acid-strong base titrations
- only methyl orange can be used for strong acid-weak base titrations
- there is no suitable indicator for weak acid-weak base titrations



Practical 10.2: Determine the concentration of a solution of NaOH by titration against 1.0 mol dm⁻³ HCl

1) Add 0.1 mol dm⁻³ hydrochloric acid to a burette and record the initial volume

Set up the burette for the first time (you only have to do this before the first titration)

- Ensure that the tap is closed
- Rinse the burette with acid to remove any impurities: pour around 10 cm³ of acid into the burette using a funnel; turn the burette as far horizontally as you can without spilling the acid and then rotate the burette a few times so that the acid covers all parts of the inside of the glass; then empty the burette by holding the tip of the burette downwards over a sink and opening the tap until the acid has come out; then close the tap
- Use a boss to attach a clamp to a stand and clamp a burette vertically into position; the tip of the burette should be pointing downwards and should be around 20 cm above the bench surface; make sure the tap is closed
- Using a funnel, add the acid solution gradually to the burette until it is around half-full; open the tap of the burette and allow acid to escape until the tip of the burette is completely full (there should be no air bubbles); you may need to flick the tip a few times to remove air bubbles; once the tip is completely full, close the tap again

Prepare the burette for a titration (you need to do this before every titration)

- Add more acid into the burette, using the funnel, until the level of solution in the burette is in between the 1 cm³ and the 5 cm³ mark; record this level accurately, to the nearest 0.05 cm³, as the "initial volume"
- Remove the funnel from the burette
- Once you have reached this stage for the first time, stop and ask the teacher to check your progress

2) Pipette 25.0 cm³ of NaOH(aq) into a conical flask and add a suitable indicator

- Pour around 100 cm³ of the NaOH into a beaker
- Use a pipette with a pipette filler to suck the alkali into the pipette until it is just above the graduation mark; remove the pipette filler and quickly place your thumb over the top of the pipette; release the pressure gradually until the base of the meniscus lies on the mark; then increase the thumb pressure to fix the solution at that level (it takes practice to get this right)
- Still applying pressure with your thumb, move the pipette out of the alkali solution and over a clean, empty 250 cm³ conical flask; remove your thumb and allow the pipette to empty itself under gravity
- Once the solution has stopped draining out of the pipette, dip the tip of the pipette into the solution in the conical flask for a few seconds and then remove the pipette completely from the conical flask
- Place the empty pipette to one side
- Add two drops of phenolphthalein indicator to the alkali; the indicator should turn pink
- Once you have reached this stage for the first time, stop and ask the teacher to check your progress

3) Add the HCl to the NaOH until the indicator changes colour

- Place the conical flask containing 25.0 cm³ of the alkali and the indicator and place it below the burette; the rim of the conical flask should be just below the tip of the burette
- open the tap of the burette and add the solution gradually into the conical flask, swirling the conical flask continuously until around 10 cm³ has been added
- close the tap gradually until the solution is dripping slowly into the conical flask at the rate of one drop per second; keep swirling the mixture and continue until the indicator changes colour (into Colour 1); then close the tap
- Record the new volume of solution accurately (to the nearest 0.05 cm³) as the “final volume”
- Once you have reached this stage for the first time, stop and ask the teacher to check your progress

4) Repeat the titration until concordancy is achieved

- After rinsing out the conical flask, steps 1 to 3 (except the burette set-up phase) should be repeated; both titre volumes for both titrations should be calculated and compared
- If the two titre volumes are within 0.10 cm³ of each other, the results can be described as concordant and you have completed the experiment
- If the values differ by more than 0.10 cm³, the results are not concordant and you must repeat the experiment until you have obtained two concordant titre volumes
- Record all your results in a table similar to the one below:

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Final volume (cm ³)					
Initial volume (cm ³)					
Titre volume (cm ³)					
Concordant? (Y or N)					

5) Hence determine the concentration of the sodium hydroxide solution

- Calculate the average of your concordant values
- Hence calculate the number of moles of HCl used in the titration ($n = C \times V$)
- Hence calculate the number of moles of NaOH in the conical flask (same as the moles of HCl)
- Hence calculate the molarity of the NaOH solution ($C = n/V$)

Need more assistance with this practical? Try these videos: www.youtube.com/watch?v=Y-5QJlr7Xm4 and www.youtube.com/watch?v=VADhf7M8bJ8

Unit 5 – Acids, Bases and Salts

Equipment needed per group: 1 burette, 1 25 cm³ pipette, 1 pipette filler, one conical flask, one funnel, two beakers and the means to label them; access to 0.1 moldm⁻³ HCl, a solution of NaOH of approximately 0.08 moldm⁻³ but with the concentration not labelled), phenolphthalein indicator and suitable dropping pipette
Note: this practical is the most important practical in SS Chemistry; it is also the most challenging in terms of the techniques required; it is recommended that when carrying out this practical for the first time, each group performs each step together so that the teacher can check that the student is performing the practical correctly
Particular things to check in each group are: no air bubbles in burette tip, initial reading is recorded to 0.05 cm³ and is correct, funnel is removed

- if the NaOH molarity is 0.08 moldm⁻³ then a typical titre volume should be around 15 cm³
- a correctly completed table would look something like this: it may be an idea to share this with the students beforehand:

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Final volume (cm ³)	23.30	24.85	23.15		
Initial volume (cm ³)	1.25	23.30	1.70		
Titre volume (cm ³)	22.05	21.55	21.45		
Concordant? (Y or N)	NO	YES	YES		

- the average titre volume (using concordant results only) should be $(21.55 + 21.45)/2 = 21.40$ cm³ (using the above results)
- moles of HCl = $0.1 \times \text{titre volume} / 1000$
- moles of NaOH = moles of HCl
- molarity of NaOH = moles of NaOH / 0.025

- Here is a simple worked example of how to use a titre volume to deduce the molarity of the unknown solution:

Question: In a titration 28.3 cm³ of a 0.10 moldm⁻³ solution of NaOH was required to react with 25 cm³ of a solution of H₂SO₄. What was the concentration of the H₂SO₄ solution?

Worked answer:

- Equation: $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
- Moles of NaOH used in titration = $28.3/1000 \times 0.1 = 2.8 \times 10^{-3}$
- 2:1 ratio so moles of H₂SO₄ present in conical flask = $2.8 \times 10^{-3}/2 = 1.4 \times 10^{-3}$
- so concentration of H₂SO₄ = $1.4 \times 10^{-3}/25 \times 1000 = 0.056$ moldm⁻³

Lesson 11 – How I determine how much of an acid or a base is present in a sample II?



Summary Activity 11.1: How do you prepare a standard solution?

- In SS1 you prepared some standard solutions, either by dissolving a solid or by diluting a concentrated solution
- How would you prepare 250 cm³ of a 0.1 moldm⁻³ solution of NaOH, starting with solid NaOH?
- How would you prepare 250 cm³ of a 0.2 moldm⁻³ solution of H₂O₂, starting with 2.0 moldm⁻³ H₂O₂?

Unit 5 – Acids, Bases and Salts

- Moles of NaOH required = $0.25 \times 0.1 = 0.025$ so mass of NaOH required = $0.025 \times 40 = 1.0$ g
Weigh out exactly 1.0 g of NaOH on a mass balance using a weighing boat
Pour the NaOH into a beaker
Rinse the weighing boat with distilled water and ensure all the rinsings run into the beaker
Dissolve the NaOH in a small quantity of distilled water
Pour the solution into a 250 cm³ volumetric flask using a funnel
Rinse the beaker with distilled water and pour the rinsings into the volumetric flask
Add distilled water until the meniscus rests exactly on the graduation mark, shaking continuously
- Dilution factor = $2/0.2 = 10$ so volume needed = $250/10 = 25$ cm³
Pipette 25 cm³ of 2.0 moldm⁻³ H₂O₂ into a 250 cm³ volumetric flask
Add distilled water until the meniscus rests exactly on the graduation mark, shaking continuously

- Once the titre volume has been determined, other desired quantities can be determined using techniques from the amount of substance topic; the simplest quantity to determine is the concentration of the unknown solution (acid or base); other quantities which can be determined from titration results include:
 - the molar masses of acids and bases and water of crystallization
 - the solubility of acids and bases
 - the percentage purity of acids and bases
- In many cases, the solution in the pipette needs to be prepared by dissolving a solid sample or by diluting a more concentrated solution; this requires a standard solution of the substance to be prepared first



Practical 11.2: Determine the relative formula mass, and hence water of crystallisation, of hydrated sodium carbonate, Na₂CO₃.xH₂O, by titration (Na₂CO₃ + 2HCl → 2NaCl + CO₂ + H₂O)

- 1) Weigh out approximately 3.5 g of hydrated sodium carbonate into a weighing boat. Record the exact mass of the solid.
- 2) Dissolve the solid in approximately 100 cm³ of distilled water.
- 3) Transfer this solution into a 250 cm³ volumetric flask, and make the volume in the flask up to 250 cm³ by using washings from the beaker; shake continuously to ensure that the contents are thoroughly mixed
- 4) Pipette a 25.0 cm³ sample of this solution into a conical flask and add a few drops of methyl orange indicator (the solution should turn yellow)
- 5) Fill a burette with 0.1 moldm⁻³ HCl and record the initial volume
- 6) Add the HCl solution to the sodium carbonate solution until the indicator changes colour (from yellow to pink); record the final volume and deduce the titre volume
- 7) Repeat the titration until concordancy is achieved
- 8) Hence determine the moles of sodium carbonate present in the conical flask, and in the original volumetric flask
- 9) Hence determine the molar mass of the hydrated sodium carbonate sample
- 10) Hence calculate the value of x, given that the formula of sodium carbonate is Na₂CO₃.10H₂O

Equipment needed per group: 1 burette, 1 25 cm³ pipette, 1 pipette filler, one conical flask, two funnels, two 100 cm³ beakers with the means to label them, one 250 cm³ volumetric flask; access to 0.1 moldm⁻³ HCl, a sample of hydrated sodium carbonate with the formula not labelled, methyl orange indicator and suitable dropping pipette

- If the student uses 3.5 g of solid the titre volume should be around 24 cm³
- Using 24 cm³, moles of HCl = $0.1 \times 0.0224 = 0.00224$ so moles of Na₂CO₃ in conical flask is $0.00224/2 = 0.00112$
- Moles of Na₂CO₃ in volumetric flask = $0.00112 \times 10 = 0.0112$
- Molar mass of Na₂CO₃ = $3.5/0.0112 = 292$
- $106 + 18x = 292$ so $x = 10$



Practical 11.3: Determine the percentage purity of a sample of vinegar ($\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$)

- 1) A solution of vinegar is labelled as 62.3 gdm^{-3} . Pipette 25 cm^3 of this vinegar sample into a 250 cm^3 volumetric flask
- 2) Make the volume in the flask up to the mark, shaking continuously to ensure that the contents are thoroughly mixed
- 3) Transfer some of this diluted acid into a burette and record the initial volume
- 4) Pipette a 25.0 cm^3 sample of the 0.05 moldm^{-3} sodium hydroxide solution into a conical flask and add a few drops of phenolphthalein indicator (the solution should turn pink)
- 5) Add the diluted vinegar solution to the NaOH solution until the indicator changes colour (from pink to colourless); record the final volume and deduce the titre volume
- 6) Repeat the titration until concordancy is achieved
- 7) Hence determine the concentration of the diluted vinegar solution
- 8) Hence determine the concentration of the original vinegar solution
- 9) Hence determine the percentage purity of the original vinegar solution
- 10) Explain why phenolphthalein was used as the indicator for this titration

Equipment needed per group: 1 burette, 1 25 cm^3 pipette, 1 pipette filler, one conical flask, two funnels, two 100 cm^3 beakers with the means to label them, one 250 cm^3 volumetric flask; access to 0.1 moldm^{-3} NaOH, a sample of ethanoic acid (approx 1 moldm^{-3}) labelled "vinegar 62.3 gdm^{-3} ", phenolphthalein indicator and suitable dropping pipette

- If the ethanoic acid is around 1 moldm^{-3} the titre volume should be around 13 cm^3
- moles of NaOH = $0.05 \times 25/1000 = 0.00125$ so moles of CH_3COOH in titration = 0.00125
- molarity of = $\text{CH}_3\text{COOH} / (\text{titre volume}/1000)$; using 13 cm^3 , $C = 0.00125/0.013 = 0.096 \text{ moldm}^{-3}$
- so molarity before dilution = $0.096 \times 250/25 = 0.96 \text{ moldm}^{-3}$
- mass concentration = $0.96 \times 60 = 57.7 \text{ gdm}^{-3}$
- percentage purity = $57.7/62.3 \times 100 = 92.6 \%$
- phenolphthalein necessary because acid is weak so methyl orange will not work


Test your knowledge 11.4: Volumetric Analysis – Titrations

- (a) Succinic acid has the formula $(\text{CH}_2)_n(\text{COOH})_2$ and reacts with dilute sodium hydroxide as follows:
 $(\text{CH}_2)_n(\text{COOH})_2 + 2\text{NaOH} \rightarrow (\text{CH}_2)_n(\text{COONa})_2 + 2\text{H}_2\text{O}$
 2.0 g of succinic acid were dissolved in water and the solution made up to 250 cm^3 . This solution was placed in a burette and 18.4 cm^3 was required to neutralise 25 cm^3 of 0.1 mol dm^{-3} NaOH. Deduce the molecular formula of the acid and hence the value of n .
- (b) Sodium carbonate exists in hydrated form, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, in the solid state. 3.5 g of a sodium carbonate sample was dissolved in water and the volume made up to 250 cm^3 . 25.0 cm^3 of this solution was titrated against 0.1 mol dm^{-3} HCl and 24.5 cm^3 of the acid were required. Calculate the value of x given the equation: $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$
- (c) 25 cm^3 of a sample of vinegar (CH_3COOH) was pipetted into a volumetric flask and the volume was made up to 250 cm^3 . This solution was placed in a burette and 13.9 cm^3 were required to neutralise 25 cm^3 of 0.1 mol dm^{-3} NaOH. Calculate the molarity of the original vinegar solution and its concentration in g dm^{-3} , given that it reacts with NaOH in a 1:1 ratio.
- (d) 2.5 g of a sample of impure ethanedioic acid, $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$, was dissolved in water and the solution made up to 250 cm^3 . This solution was placed in a burette and 21.3 cm^3 were required to neutralise 25 cm^3 of 0.1 mol dm^{-3} NaOH. Given that ethanedioic acid reacts with NaOH in a 1:2 ratio, calculate the percentage purity of the sample.
- (e) A toilet cleaner containing sodium hydrogensulphate, NaHSO_4 is believed to have been contaminated. 5.678 g of the sample were dissolved in water and the solution was made up to 250 cm^3 . This solution was placed in a burette and 23.1 cm^3 of it were required to neutralise 25 cm^3 of 0.1 mol dm^{-3} sodium hydroxide.
 Calculate the percentage purity of the sample given that the species react in a 1:1 ratio.

- (a) Moles of NaOH = 0.0025 ; moles of SA used = $0.0025/2 = 0.00125$; moles of SA in volumetric flask = $0.00125 \times 250/18.4 = 0.016984$, mr of SA = $2/0.016984 = 118$; $(\text{CH}_2)_n = 118 - 90 = 24$ so $n = 24/14 = 1.71$
- (b) Moles of HCl = $0.0245 \times 0.1 = 0.00245$; moles of Na_2CO_3 used = $0.00245/2 = 0.001225$; moles of in volumetric flask = $0.001225 \times 10 = 0.01225$, molar mass of $\text{Na}_2\text{CO}_3 = 3.5/0.001225 = 286$; $x\text{H}_2\text{O} = 286 - 106 = 180$; $x = 180/18 = 10$
- (c) Moles of NaOH = $0.025 \times 0.1 = 0.0025$; moles of CH_3COOH used = 0.0025 ; molarity of CH_3COOH used = $0.0025/(13.9/1000) = 0.180$; molarity before dilution = $0.180 \times 250/25 = 1.80 \text{ mol dm}^{-3}$; mass concentration = $1.8 \times 60 = 108 \text{ g dm}^{-3}$
- (d) Moles of NaOH = 0.0025 ; moles of acid used = $0.0025/2 = 0.00125$; moles of acid in volumetric flask = $0.00125 \times 250/21.3 = 0.0147$; mass of pure acid = $0.0147 \times 126 = 1.85 \text{ g}$; % purity = $1.85/2.5 \times 100 = 73.9 \%$
- (e) Moles of NaOH = 0.0025 ; moles of acid used = 0.0025 ; moles of acid in volumetric flask = $0.0025 \times 250/21.3 = 0.0271$; mass of acid = $0.0271 \times 120.1 = 3.25$; % purity = $3.25/5.678 \times 100 = 57.2 \%$

Lesson 12 – What have you understood about Acids, Bases and Salts?

12.1 END-OF-TOPIC QUIZ

UNIT 5 – ACIDS, BASES AND SALTS



- Write symbol equations for the following neutralisation reactions:
 - hydrochloric acid and sodium hydroxide
 - sulphuric acid and magnesium oxide
 - nitric acid and calcium carbonate
- Explain the meaning of the term deliquescent, hygroscopic and efflorescent and give one example of an acid, base or salt with each property.
- State the difference between a strong acid and a weak acid. Explain how 1.0 mol dm^{-3} HCl and 1.0 mol dm^{-3} CH_3COOH would differ in terms of their:
 - pH
 - enthalpy of neutralisation by NaOH
 - rate of reaction with calcium carbonate
 - electrical conductivity
- Suggest whether the following salts undergo salt hydrolysis, and if so, whether they are acidic or basic salts:
 - sodium chloride
 - ammonium chloride
 - sodium ethanoate
- Describe a chemical test which would identify the following:
 - H^+ ions in solution
 - OH^- ions in solution
 - NH_4^+ ions in solution
 - Cl^- ions from a solid sample
 - SO_3^{2-} ions from a solid sample
 - NH_3 gas
- Describe how:
 - universal indicator could distinguish between samples of water, lemon juice and bleach
 - litmus paper could distinguish between HCl and NH_3
- Deduce the identity of the Group 1 metal M based on the following experiment:

2.5 g of the metal carbonate M_2CO_3 was dissolved in water and the volume made up to 250 cm^3 . A 25 cm^3 sample of this solution was titrated against 0.2 mol dm^{-3} HCl and 10.8 cm^3 of the acid was required to neutralise the sample.

Unit 5 – Acids, Bases and Salts

- $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
 - $\text{H}_2\text{SO}_4 + \text{MgO} \rightarrow \text{MgSO}_4 + \text{H}_2\text{O}$
 - $\text{CaCO}_3 + 2\text{HNO}_3 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}$
- Deliquescent – absorbs water from atmosphere and dissolves in it (eg NaOH or CaCl₂)
Hygroscopic – absorbs water from atmosphere (eg conc. H₂SO₄ or any deliquescent substance)
Efflorescent – contains water which it releases (eg CaSO₄·2H₂O)
- strong acid – lower pH
 - strong acid – more exothermic enthalpy of neutralisation
 - strong acid – faster reaction with calcium carbonate
 - strong acid – greater electrical conductivity
- no - neutral (salt of strong acid and strong base)
 - yes - acidic (salt of weak base)
 - yes – basic (salt of weak acid)
- Add CaCO₃; observe fizzing
 - Add NH₄Cl and warm – pungent smell
 - Add NaOH and warm – pungent smell
 - Add concentrated H₂SO₄ – white fumes which turn blue litmus red and give white smoke in presence of filter paper soaked in NH₃
 - Add acid; gas given off which smells like burning matches
 - Gives white smoke in presence of filter paper soaked in concentrated HCl
- water – green; lemon juice – red; bleach – violet
 - HCl turns blue litmus red; NH₃ turns red litmus blue
- Moles of HCl = $0.2 \times 10.8/1000 = 0.00216$; $\text{M}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{MCl} + \text{CO}_2 + \text{H}_2\text{O}$ so moles of M₂CO₃ used = $0.00216/2 = 0.00108$; moles of M₂CO₃ in volumetric flask = $0.00108 \times 10 = 0.0108$; molar mass = $2.5/0.0108 = 231$; $2M = 231 - 60 = 171$; $M = 171/2 = 85.5$; $M = \text{Rb}$