MONTFORT SECONDARY SCHOOL

CHEMISTRY 6092
REVISION NOTES 2024

Name: _____ **Class:** _____

CONTENT PAGE

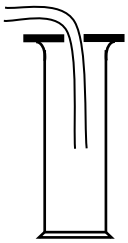
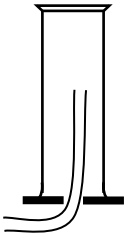
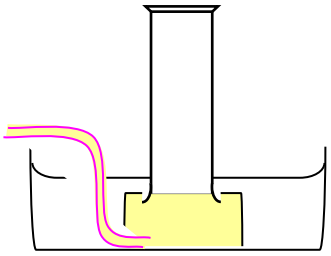
1.1 Experimental Design

Key Concept I

Quantity	S.I. Unit	Apparatus for measurement
time	second (s) [1h = 60 min] [1 min = 60 s]	Electronic stopwatch
temperature	Kelvin (K) more commonly used - degree Celsius (°C)	Laboratory thermometer
mass	Kilogram (kg) Other units – gram (g) [1 kg = 1000 g]	Electronic balance
volume	Cubic metre (m^3) Other units - litre (l), cubic centimetre (cm^3) [1 m^3 = 1000dm^3] [1 dm^3 = 1000cm^3]	<u>liquid</u> Pipette – fixed volumes (eg. 20.0 cm^3, 25.0 cm^3) Burette – accuracy to 2 d.p. (eg. 24.85 cm^3) Measuring cylinder – accuracy to 1 d.p (eg. 20.5 cm^3)
		<u>gas</u> gas syringe

Key Concept II

1. Collection of Gases

 **downward delivery of gas**	 **upward delivery of gas**	 **displacement of water**
For • gas **denser** than air (high density / heavier)	For • gas **less dense** than air (low density / lighter)	For • for gases that are **insoluble** / slightly soluble in water
Examples: Chlorine Hydrogen chloride	Examples: Ammonia	Examples: Hydrogen Oxygen Carbon dioxide

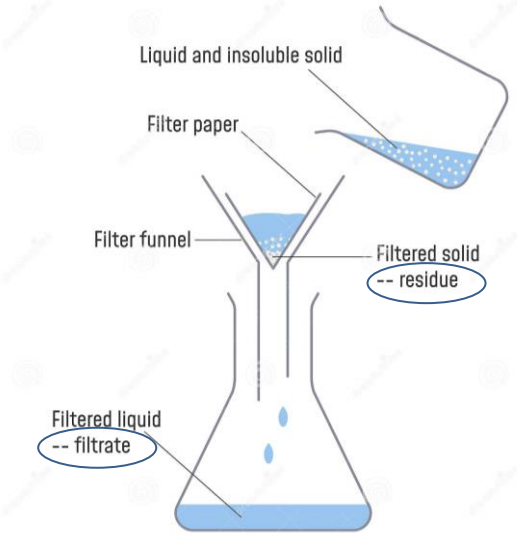
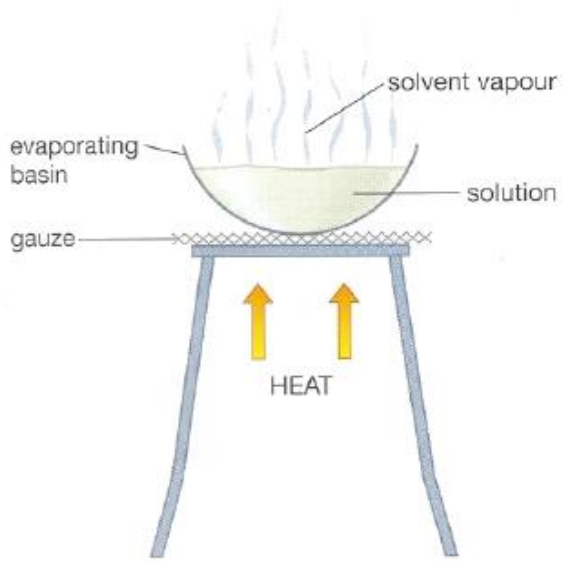
2. Collection of Gases <u>AND</u> Measuring the Volume of Gas

gas syringe

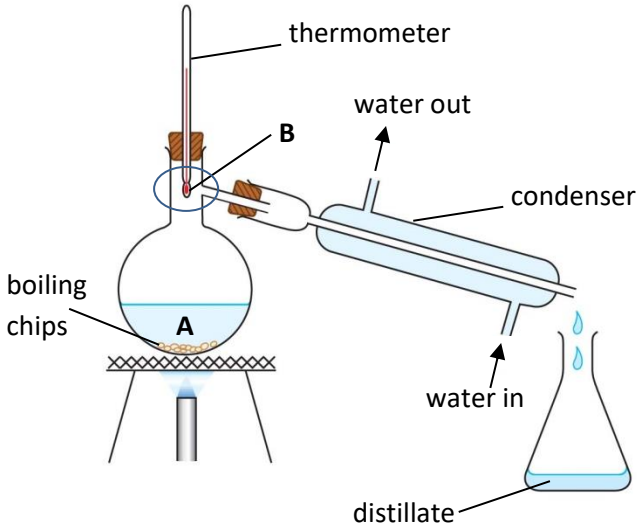
1.2 Methods of Purification and analysis

Key Concept I

Types of Separation Techniques

Separation technique	Commonly asked questions
Filtration  	1. Essay type question, where you are given an insoluble solid. 2. You will be expected to draw a labelled diagram. ** *Residue : insoluble solid* ** *Filtrate: liquid*
Evaporation / Crystallisation  	1. Essay type question, where you are required to obtain dry solid from a solid-liquid mixture. 2. You will be expected to draw a labelled diagram. 3. For evaporation, *heat to dryness* 4. For crystallisation, *heat to saturation point* / crystallisation point. ** Evaporation **cannot** be used for substances that decompose on heating.

Distillation (solid-liquid mixture, e.g. distillation of seawater)



1. At which position (**A** or **B**) would the temperature be 100^{o}C? *(Ans: B. A contains impurities, hence the boiling point is not at 100^{o}C)*

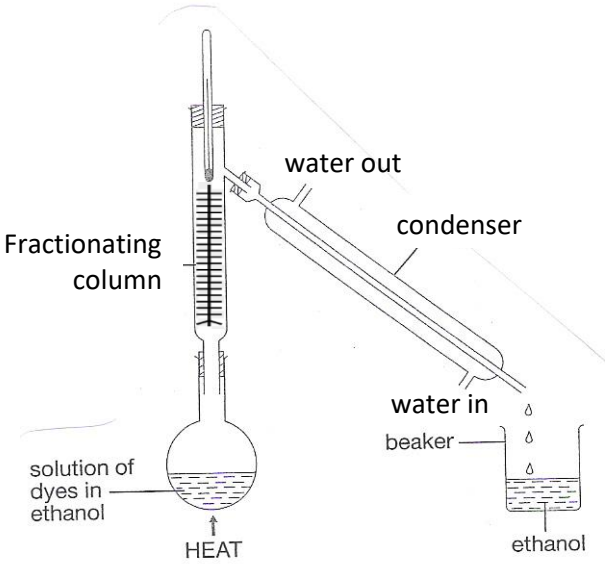
2. What is the purpose of the condenser? *(Ans: to <u>cool and condense vapours</u>.* ***REJECTED ANS: To condense the vapour**)*

3. Draw arrows to indicate where water enters and exit the condenser.
(BITO: **B**ottom **I**n **T**op **O**ut)

4. What is the purpose of boiling chips? *(Ans: To <u>ensure smooth boiling</u>)*

5. Identify the distillate in this distillation experiment
(Ans: water)

Fractional distillation (liquid-liquid mixture)

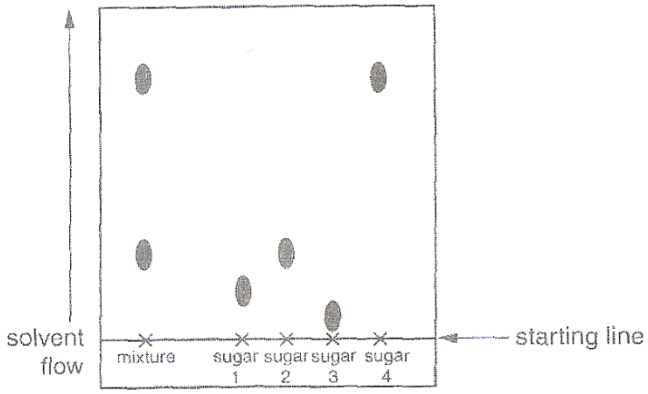


1. What is the purpose of fractional distillation?
*(Ans: to separate **liquids** with close boiling points)*

2. What are some industrial application of fractional distillation?
(Ans: Refining crude oil, separating air into its components)

****** liquid with **lower** boiling point will distil first.

Chromatography	1. Which substance has the highest solubility? *(Ans: sugar 4)* 2. Which substance has the lowest solubility? *(Ans: sugar 3)* 3. Which substances does the mixture contain? *(Ans: sugars 2 and 4)* 4. Why is the starting line drawn in pencil? *(Ans: Pencil lead is not soluble in most solvents)* 5. Name a suitable solvent. *(Ans: water, ethanol)* ** More soluble in solvent → travel faster and further up the chromatogram.



Key Concept II

Purity Check

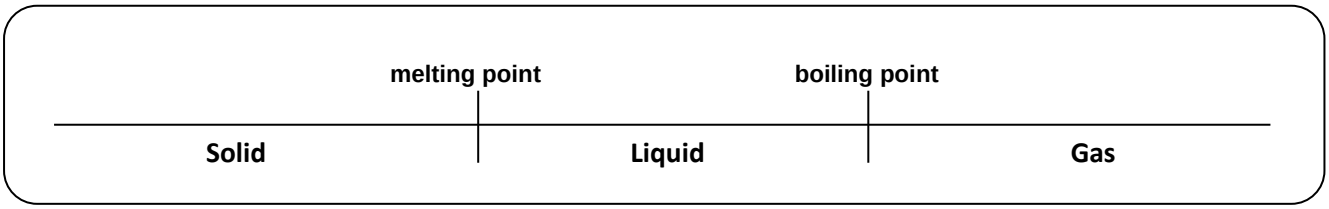
	Pure substance	**Mixture**
Composition	Fixed	Not fixed
Boiling point & melting point	Fixed	Varied/ Over a range of temperature
Chromatogram	One spot	More than one spot

Example: How does one find out of the substance is pure?
Determine the melting and boiling point of the substance. **If it melts and boils at a fixed temperature, it is a pure substance**.

2.1 Kinetic Particle Theory

Key Concept I

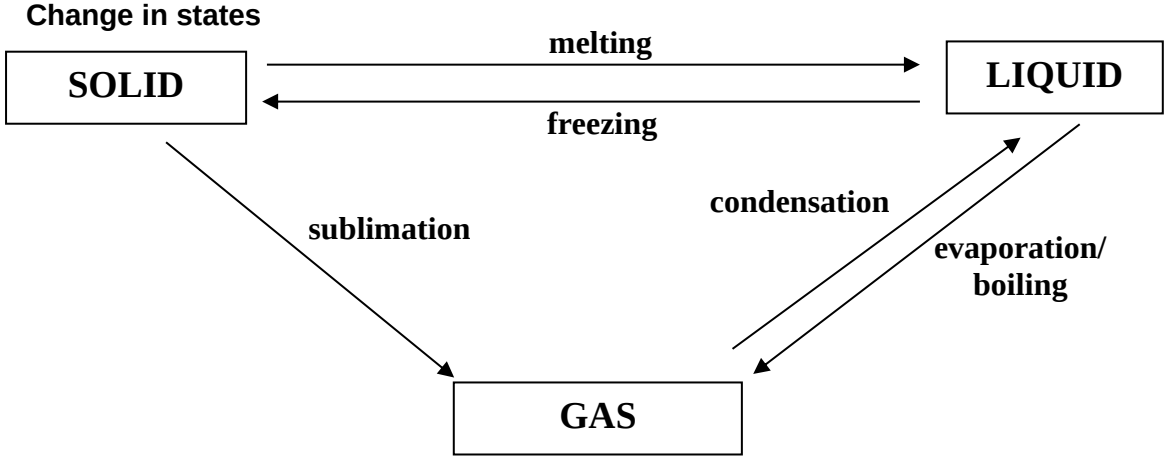
1. Physical state at room temperature (25°C) depends on the boiling and melting points.



Key Concept II
Properties of the 3 states of matter

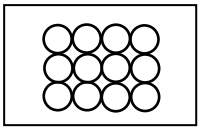
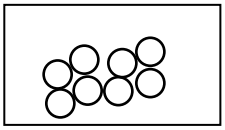
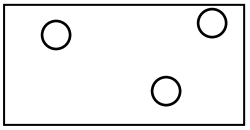
Property	Solid	Liquid	Gas
Shape	Fixed	Not fixed	Not fixed
Volume	Fixed	Fixed	Not fixed
Compressibility	Cannot be compressed	Cannot be compressed	Can be compressed

Key Concept III
Change in states



Key Concept IV

Kinetic Particle Theory***

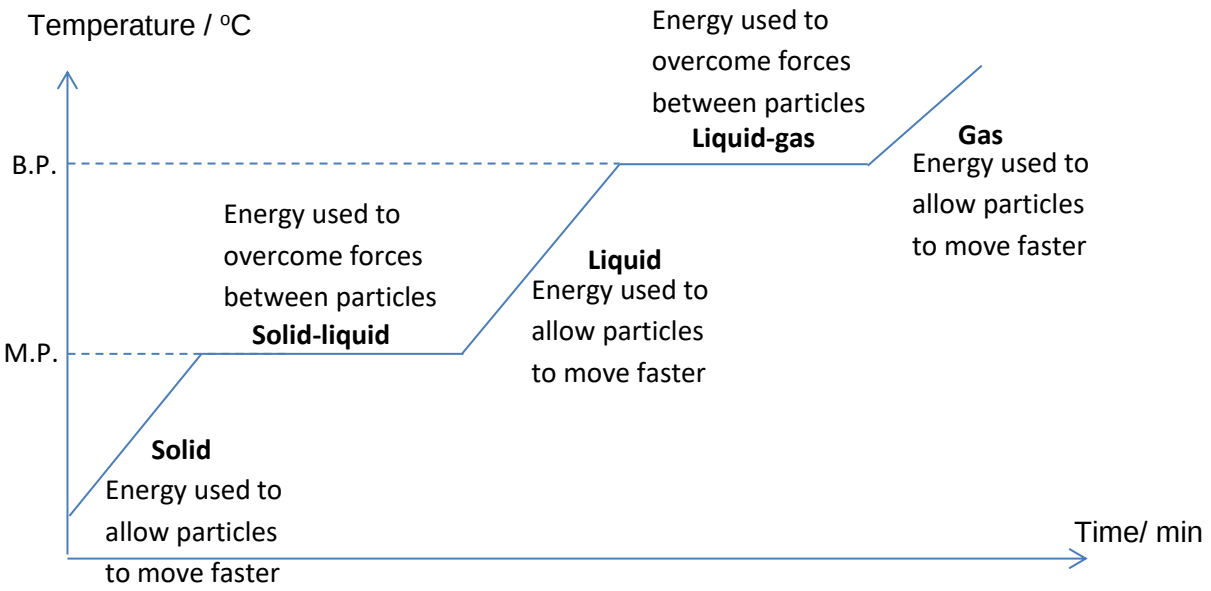
	Solid	Liquid	Gas
2-D representation of the arrangement of particles			
Arrangement of particles	• Closely packed • Regular/orderly	• Closely packed • Disorderly/ irregular	• Far apart • irregular
Attractive forces between particles	• Very strong	• Strong	• Weak
Kinetic energy of particles	• Very low	• High	• Very high
Particle motion	• <u>Vibrate about a fixed position</u>	• Slide over each another	• Rapidly in any direction

**** For changes from one state to another: emphasise on the changes in spacing/ movement (e.g. closer/ faster/ further)**

Example: What happens when a liquid boils? (Key focus: KPT, Energy changes)
Particles gain enough energy to overcome the force of attraction (between them). The particles move further apart from each other and move with greater speed.

Key Concept V

Heating curve



Key Concept VI
Differences between boiling and evaporation

Boiling	Evaporation
Occurs at the boiling point	Occurs at any temperature
Occurs throughout the liquid	Occurs at the surface of the liquid
Very fast	Very slow

Key Concept V
Diffusion

Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration until an equilibrium is achieved.

The rate of diffusion is affected by:

- Relative molecular mass (Mr) of the particles (higher Mr → slower rate of diffusion)
- Temperature of substance (higher temperature → faster rate of diffusion)

2.2 Atomic Structure

Key Concept I

1. Atoms are made up of 3 sub-atomic particles

Particle	Symbol	Relative mass	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	$\frac{1}{1840}$	–1

2. Protons and neutrons: in the nucleus
   Electrons: in electron shells outside the nucleus.

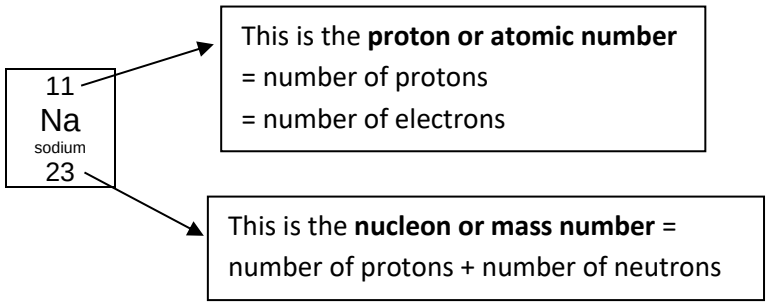
3. Proton /atomic number: number of protons in an atom. (unique for each element)

4. Nucleon / mass number: number of protons and neutrons in an atom.

   (Nucleon – nucleus, particles in the nucleus; mass number – sum of the relative masses, but mass of electron is negligible, hence not included)

   Atoms are **electrically neutral** because its number of protons is equal to its number of electrons (**p = e**). Hence the positive and negative charges cancel off each other and the atoms have no overall charge.

   For example, in the periodic table the element sodium is shown as follows:



   Therefore, an atom of sodium has **11 protons**, **11 electrons** and **12 neutrons**.
   How did we get 12 neutrons? Number of neutrons = nucleon number – proton number

> Sodium atom can also be represented as $^{23}_{11}Na$.
> Here, the position of the proton number and nucleon number have been swapped.
> But the counting of number of protons, electrons and neutrons is the same.

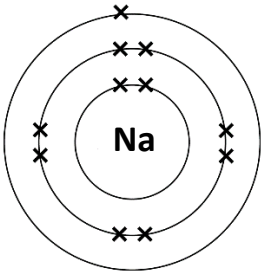
Key Concepts II

Arrangement of electrons

1. The 1st shell can hold up to 2 electrons, the 2nd and 3rd shells usually can hold up to 8 electrons.

2. The arrangement of electrons in an atom is represented using electronic configuration or electronic structure.

   For example, we know that an atom of sodium has 11 electrons. How are the 11 electrons arranged around the nucleus?

Electronic structure of sodium atom	Electronic configuration of sodium atom
	2,8,1

3. Electrons in the outer shell of an atom are called valence electrons.

4. The **chemical properties** of an element depend on the number of valence electrons.

5. Elements with the **same number of valence electrons** belong to the **same group** in the Periodic Table. Hence, they have **similar chemical properties**.

Key Concept III

Isotopes

1. Isotopes are atoms of the same element with the same number of protons and electrons, but different numbers of neutrons.

2. Isotopes have identical chemical properties but their physical properties are different.

3. Isotopes have different physical properties because they have different masses.

For example, chlorine has two isotopes – chlorine-35 and chlorine-37

$^{35}_{17}Cl$ $^{37}_{17}Cl$

Both isotopes of chlorine have **17 protons and 17 electrons**. Both have an electronic configuration of 2,8,7. The both have 7 valence electrons and so have the same chemical properties (they take part in the same chemical reactions)
However, **chlorine-35 has 18 neutrons**, while **chlorine-37 has 20 neutrons**. So chlorine-35 has a lower mass than chlorine-37, and so they have different physical properties (e.g. different density, different melting and boiling points)

3.1 & 3.2 Ionic and Covalent Bonding

Key Concept I

Type of Bonding	Elements	Electrons	Ions
Ionic	Metal and non-metal	Gain or loss	Ions formed
Covalent	Non-metals	Share	No ions formed

Why do atoms form ions or covalent bonds?
Atoms form ions or form covalent bonds so that they can obtain a stable, fully-filled valence electron shell similar to the noble gases.

'Dot and cross' Diagram

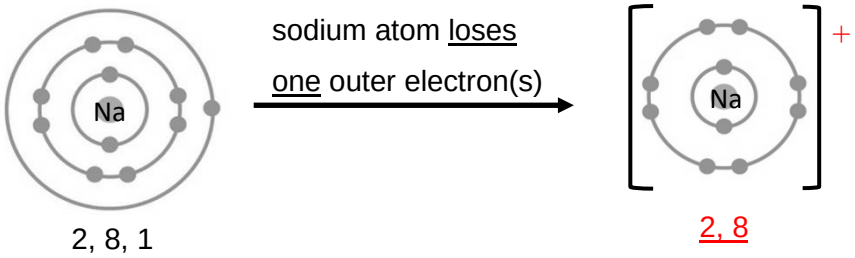
Step 1: Identify whether it is a covalent or ionic compound.
Step 2: Write the electronic configuration of each element.
Step 3:

Ionic compound	Covalent compound
Find the charge of each element	Find the number of electron(s) needed to obtain the noble gas configuration
Ensure that the positive and negative charges cancel off each other	Share 1 for 1 from each atom

Forming Ions

1. An atom loses or gains electrons to form ions. Ions are charged particles because they do not have equal numbers of protons and electrons.
2. Metallic atoms lose electrons to form positively-charged ions called cations.

Example: Formation of a sodium ion

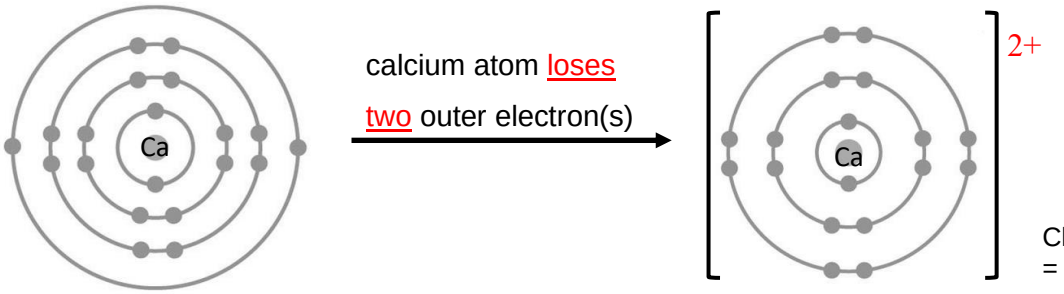


2, 8, 1

2, 8

Charge on sodium ion
= 11p + 10e
= (+11) + (–10)
= +1

Example: Formation of a calcium ion



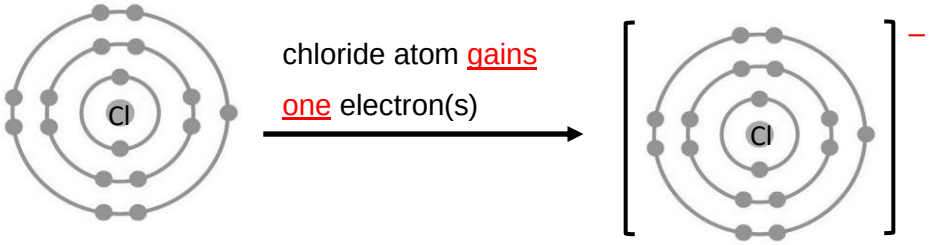
2, 8, 8, 2

2, 8, 8

Charge on calcium ion
= 20p + 18e
= (+20) + (–18)
= +2

3. Non-metallic atoms gain electrons to form negatively-charged ions called anions.

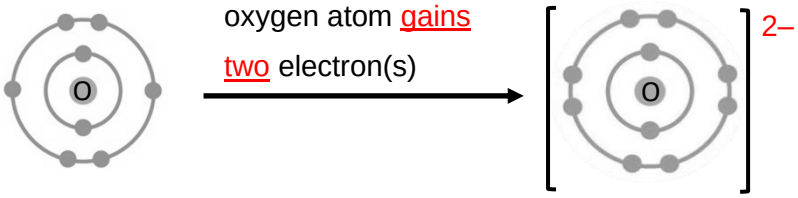
Example: Formation of a chloride anion



2, 8, 7 → 2, 8, 8

Charge on chloride ion = 17p + 18e
= (+17) + (–18)
= –1

Example: Formation of an oxide ion



2, 6 → 2, 8

Charge on oxide ion = 8p + 10e
= (+8) + (–10)
= –2

Ionic Bonds: Transferring Electrons

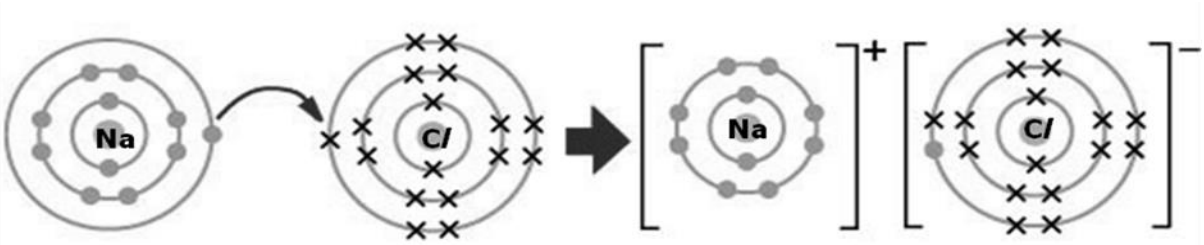
1. In the formation of an ionic bond, the metal atom transfers outer electrons to the non-metal atom. Positive ions (cations) and negative ions (anions) are formed.

2. The oppositely charged ions are held together by **strong** electrostatic forces of attraction called ionic bonds.

Example: Sodium chloride

(a) The sodium atom loses one valence electron to the chlorine atom.

ionic equations: $Na \rightarrow Na^+ + e^-$ $Cl + e^- \rightarrow Cl^-$



Na atom	Cl atom	Na^+ ion	Cl^- ion
(2, 8, 1)	(2, 8, 7)	(2, 8)	(2, 8, 8)

Chemical Formulae of Ionic Compounds

1. When writing the chemical formula of an ionic compound, the positive charges must balance the negative charges.
2. The chemical formula of an ionic compound can be deduced using the following steps:
3. **(a)**Write down the ions with the charges, e.g. $X^{m+}Y^{n-}$.
4. **(b)** Move the values m and n diagonally (without the charges).

The chemical formula is X_nM_m.

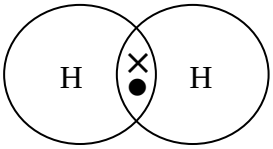
Covalent Bond: Sharing Electrons

1. Covalent bonding involves the sharing of electrons with no transfer or movement of electrons between two atoms. Molecules are formed, not ions. (MOLECULES ARE MADE UP OF 2 OR MORE ATOMS CHEMICALLY COMBINED)

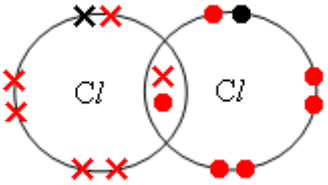
2. This type of bonding occurs when non-metal atoms join together so as to achieve the stable electronic configuration of a noble gas.

Molecules of Elements

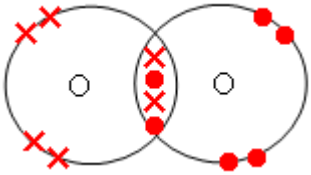
1. Hydrogen



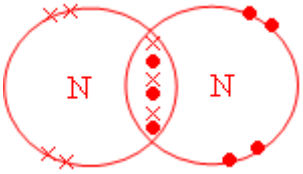
2. Chlorine



3. Oxygen

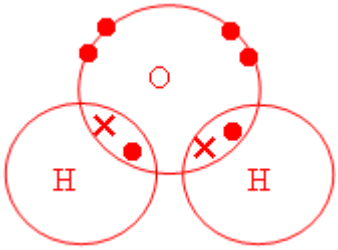


4. Nitrogen

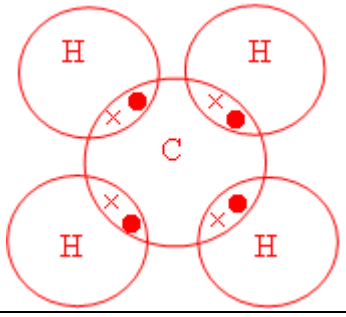


Molecules of Compounds

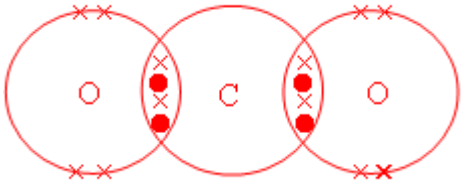
1. Water



2. Methane



3. Carbon dioxide



Key Concept II

Boiling and Melting point

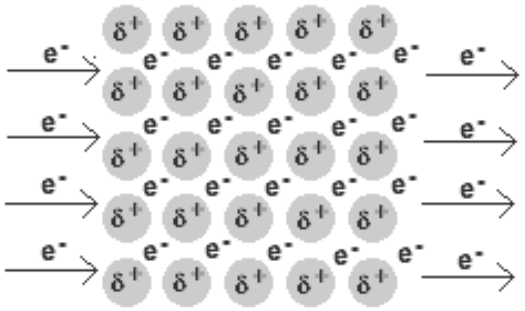
Type of Compound	Boiling and melting point	Force of attraction	Results
Ionic	High	Strong electrostatic force of attraction between ions	Large amount of energy is required to overcome this force.
Simple Covalent	Low	Weak intermolecular force of attraction between molecules	Only a small amount of energy is required to overcome this force.

Key Concept III

Electrical Conductivity

Type of Compound	Electrical conductivity	Explanation
Ionic	Solid (x)	Ions are held in a fixed position in a giant lattice structure No mobile ions to conduct electricity
	Molten/ liquid (√)	Ions are no longer held in a fixed position and are able to move around. Mobile ions available to conduct electricity
	Aqueous (√)	
Simple Covalent	Not in any states	No ions or electrons (charged particles) present to conduct electricity

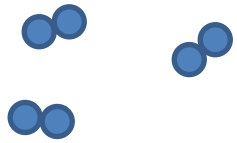
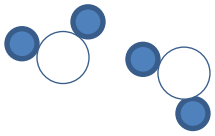
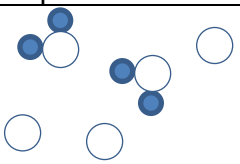
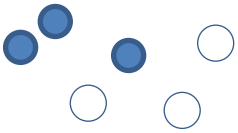
3.3 Metallic Bonding

	METAL
diagram	
structure	Metals consist of metal cations that are surrounded by a 'sea' of mobile valence electrons and are held together by strong electrostatic forces of attraction in a metal lattice structure.
high melting points / solid at room temperature	Metals consist of metal cations that are surrounded by a 'sea' of mobile valence electrons and are held together by strong electrostatic forces of attraction in a metal lattice structure. These strong electrostatic forces of attraction require large amount of energy to overcome them. Hence, metals have a high melting and boiling point.
conduct electricity	Metals consist of metal cations that are surrounded by a 'sea' of mobile valence electrons. The 'sea' of mobile valence electrons are able to move freely throughout the metal, allowing metals to conduct electricity.

3.4 Structure and Properties of materials

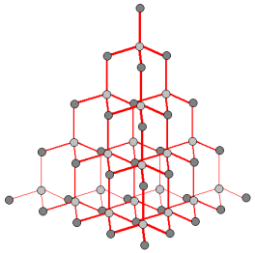
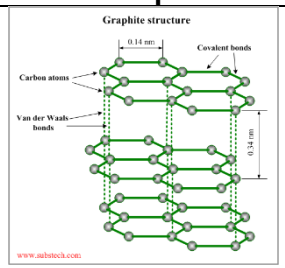
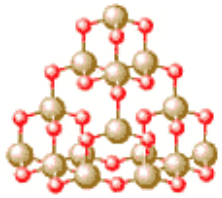
(Elements, Mixtures, Compounds)

Key Concept I

	Elements	Compounds	Mixtures
Diagrams	Example 1:	Example 1:	Example 1:
			
	Example 2:	Example 2:	Example 2:
			
Constituents	Made up of <u>one type</u> of element	Made up of <u>two or more elements chemically combined</u>	Made up of <u>two or more substances</u> **<u>not</u>** <u>chemically combined</u>
Composition by mass	-	<u>Fixed</u> composition by mass (i.e. ratio of sodium to chlorine in sodium chloride (NaCl) is 1:1)	Not fixed, can be mixed in any proportion
Properties	-	<u>Different</u> from its constituent elements	<u>Same</u> as its constituent elements
Separation	Cannot be broken down into any simpler substance	Can be broken down by <u>chemical methods</u> (i.e. electrolysis, decomposing by heating)	Can be separated by <u>physical methods</u> (i.e. filtration, evaporation, distillation and chromatography)
Energy changes	-	Formed in a chemical reaction involving heat or light	No chemical reaction involved in forming a mixture
Other notes:	Burns in oxygen to give one product (i.e. carbon + oxygen → carbon dioxide)		

3.4 Giant covalent (macromolecular) structures

A substance with a giant molecular structure is made up of atoms held together in an extensive network of covalent bonds. The large molecules are called macromolecules.

	Diamond	Graphite	Silicon(IV) oxide
diagram			
structure	In diamond, each carbon atom is bonded to four other carbon atoms by strong covalent bonds in a tetrahedral structure.	Graphite consists of layers of carbon atoms arranged in hexagonal arrangement. Within each layer, each carbon atom is bonded to three other carbon atoms by strong covalent bonds. Between layers, the layers are held together by weak intermolecular forces of attraction.	In silicon dioxide, each silicon atom is bonded to four other oxygen atoms and each oxygen atom is bonded to two other silicon atoms by strong covalent bonds in a tetrahedral structure.

melting point	High / solid state at room temperature In diamond, each carbon atom is bonded to other carbon atoms by strong covalent bonds in a tetrahedral structure. These strong covalent bonds require large amount of energy to overcome them.	High / solid state at room temperature Each carbon atom is bonded to other carbon atoms by strong covalent bonds in hexagonal layers. These strong covalent bonds require large amount of energy to overcome them.	High / solid state at room temperature In silicon dioxide, the silicon and oxygen atoms are held together by strong covalent bonds in a tetrahedral structure. These strong covalent bonds require large amount of energy to overcome them.
electrical conductivity	All the valence electrons of the carbon atoms are used for bonding. So, there are no mobile charged particles that move through the structure to conduct electricity.	Each carbon atom has a valence electron that is not used for bonding. These electrons are mobile and can move freely throughout to conduct electricity.	All valence electrons of silicon and oxygen atoms are used for bonding. There are no mobile charged particles that move through the structure to conduct electricity.
slippery		Graphite consists of layers of carbon atoms held together by weak intermolecular forces of attraction. When a force is applied, the weak forces are easily overcome and the layers of carbon atoms can slide over each other easily.	

4.1 Formulae & Equation Writing

Key Concept I

Chemical Formulae of Covalent Compounds

The molecular formula of a covalent compound will reflect the number of atoms of each element in the molecule.

Name of Compound	Formula	Name of Compound
Water	H_2O	2 hydrogen atoms and 1 oxygen atom
Methane	CH_4	1 carbon atom and 4 hydrogen atoms
Ammonia	NH_3	1 nitrogen atom and 3 hydrogen atoms

The name may contain prefixes to indicate the number of atoms of an element present in the compound.

Prefix	Number
mono	1
di	2
tri	3
tetra	4

Name of Compound	Formula
Carbon **mono**xide	CO
Carbon **di**oxide	CO_2
Sulfur **di**oxide	SO_2
Nitrogen **mono**xide	NO
Nitrogen **di**oxide	NO_2

Formula of Acids

name	formula
hydrochloric acid	HC*l*
sulfuric acid	H_2SO_4
nitric acid	HNO_3

Key Concept II

Chemical Formulae of Ionic Compounds

Charges of Ions:

- The charge of the ion can be deduced from its group number.

Group Number	1	2	13	14	15	16	17
Charge of ion	1+	2+	3+	NA	3-	2-	1-

1	**2**	H^+ hydrogen	**13**	**14**	**15**	**16**	**17**
Li^+ lithium	Be^{2+} beryllium				N^{3-} nitride	O^{2-} oxide	F^- fluoride
Na^+ sodium	Mg^{2+} magnesium		Al^{3+} aluminium			S^{2-} sulfide	Cl^- chloride
K^+ potassium	Ca^{2+} calcium						Br^- bromide
	Ba^{2+} barium						I^- iodide

Transition metal cations:

- The charge of some transition metal ions may be indicated by the roman numeral in the name of the ion.

Transition metals		
Fe^{2+} iron(II)	Cu^{2+} copper(II)	Zn^{2+} zinc
Fe^{3+} iron(III)	Ag^+ silver	Pb^{2+} lead(II)

Polyatomic ions:

- Polyatomic ion are ions which are made up of more than one atom <u>covalently bonded</u> together.

name	**formula**
hydroxide	OH^-
nitr<u>ate</u>	NO_3^-
ammonium	NH_4^+
peroxide	O_2^{2-}

name	**formula**
carbon<u>ate</u>	CO_3^{2-}
sulf<u>ate</u>	SO_4^{2-}
phosph<u>ate</u>	PO_4^{3-}
hydrogencarbonate	HCO_3^-

Note: For naming of non-metallic ions

- **-ide** is used when the ion is monoatomic
- **-ate** is used when the ion is polyatomic

General rules for writing chemical formulas of ionic compounds:

1. Compounds are electrically neutral. Hence, the overall positive and negative charges need to be balanced.
2. Symbol of metallic element is written first.
3. Number of atoms is written as subscript.
4. Subscripts are simplified to the lowest term.
5. If there is more than one unit of polyatomic ion, the formula of the polyatomic ion must be brackets.

Example:

Ionic compound	formula of cation	formula of anion	chemical formula of ionic compound
calcium chloride	Ca^{2+}	Cl^-	Ca^{2+} Cl^- → $CaCl_2$

Key Concept III

Writing Chemical Equations

- A chemical reaction summarises what happens during a chemical reaction. It describes:
  - the reactants (on the left-hand side) and the products (on the right-hand side)
  - the ratio of the reactants and products in a reaction
  - the physical states of the reactants and products of the reaction

- There are 3 types of chemical equations:

Word equation	sodium hydroxide + nitric acid → sodium nitrate + water
Chemical equation	$NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$
Ionic equation	$OH^- + H^+ \rightarrow H_2O$

- State symbols are written in the equation to indicate the **state** of each reactant and product. The state symbols are:
  - (s): **solid** eg. metal, precipitates, powder
  - (*l*): **liquid** eg. water
  - (g): **gas** eg. ammonia, carbon dioxide, water vapour
  - (aq): **aqueous, dissolved in water** eg. acids, alkali, solutions

Steps to writing balanced chemical equation:

<table>
<tr><td>Step 1</td><td>Write down the chemical formulae of the reactants and products to get the chemical equation.</td><td colspan="2">$Mg + HCl \rightarrow MgCl_2 + H_2$</td></tr>
<tr><td rowspan="2">Step 2</td><td rowspan="2">Check the number of atoms of each element in the formulae on both sides of the equation.
If the equation is not balanced, proceed to Step 3.</td><td>Left-hand side</td><td>Right-hand side</td></tr>
<tr><td>1 Mg
1 H
1 Cl</td><td>1 Mg
2 H
2 Cl</td></tr>
<tr><td>Step 3</td><td>Balance the equation.</td><td colspan="2">$Mg + 2HCl \rightarrow MgCl_2 + H_2$</td></tr>
<tr><td>Step 4</td><td>Add the state symbols.</td><td colspan="2">$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$</td></tr>
</table>

4.2 Mole Concept

Key Concept I
Relative Atomic Mass (A_r)

1. The relative atomic mass of an element is the mass of one atom of that element compared to $\frac{1}{12}$ of the mass of a carbon-12 atom.

$$\text{Relative atomic mass} = \frac{\text{mass of one atom of the element}}{\frac{1}{12}\text{ mass of an atom of carbon-12}}$$

2. The relative atomic mass of an element is represented by the symbol A_r and has no unit.

3. The relative atomic masses of some elements:

Element	Relative atomic mass (A_r)
hydrogen	1
carbon	12
oxygen	16
chlorine	35.5

Some A_r values are not whole numbers.

The A_r of an element takes into account the mass of all the isotopes of the element and their relative abundance.

Key Concept II
Relative Molecular Mass (M_r)

1. The relative molecular mass of a molecule is the mass of a molecule compared to $\frac{1}{12}$ of the mass of a carbon-12 atom.

$$\text{Relative molecular mass} = \frac{\text{mass of a molecule of the element/compound}}{\frac{1}{12}\text{ mass of an atom of carbon-12}}$$

2. The relative molecular mass of an element or compound is represented by the symbol M_r and has no unit.

3. Calculating the relative molecular masses of some compounds:

Compound	Relative molecular mass (M_r)
NH_3	**14 + (3 × 1) = 17**

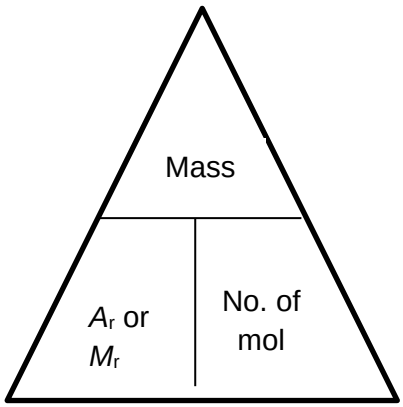
Key Concept III
Calculations Involving the Mole

1. The number of moles of an element, molecules or an ionic compound can be calculated using the following formula:

$$\text{Number of moles of a substance} = \frac{\text{mass of substance (in grams)}}{A\text{r or } M\text{r}}$$

2. By rearranging the formula, we get:

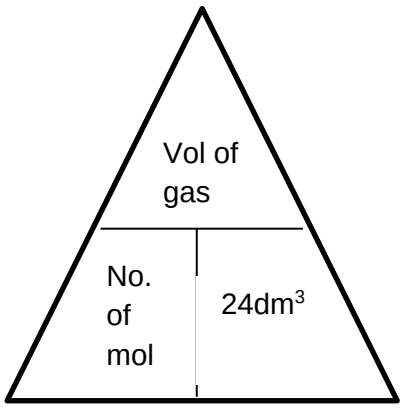
Mass of a substance = molar mass of the substance × number of moles of the substance



Key Concept IV
Calculations involving gases

1. 1 mole of gas occupies a volume of 24dm^3 at room temperature and pressure. This volume is also called molar volume.

2. 1 dm^3 = 1000 cm^3

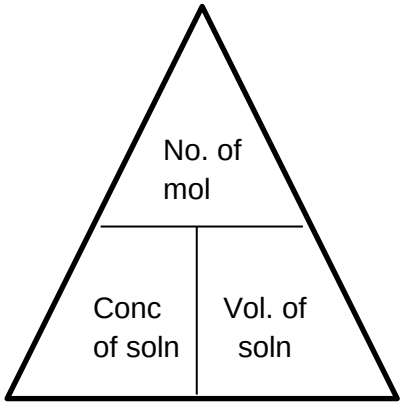


Key Concept V

Calculations involving aqueous solutions

1. Calculating the amount of reactant particles in a solution requires the concentration of solution.

2. concentration of solution in g/dm^3 = concentration of solution in mol/dm^3 x M_r



Key Concept VI

Percentage Yield and Percentage Purity

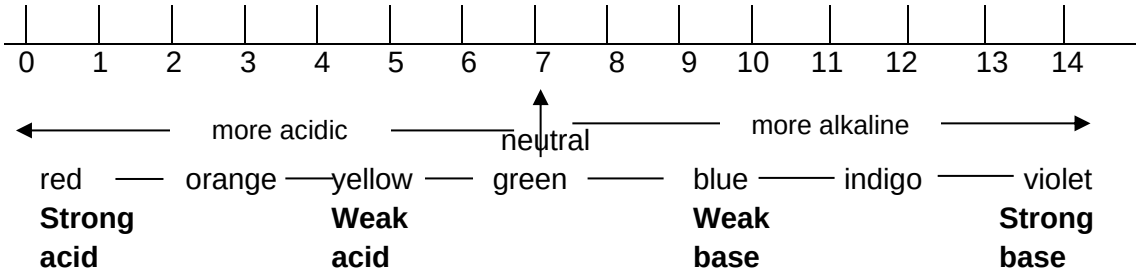
$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$\text{Percentage purity} = \frac{\text{Mass of pure substance}}{\text{Total mass of impure substance}} \times 100\%$$

5.1 Acids & Bases

Key Concept I

1. pH and colour of Universal indicator



Key Concept II

Acids

1. Definition: produces **hydrogen ions** (H^+) in aqueous solution (i.e. when dissolved in water).

2. Physical Properties:

- Conduct electricity (there are mobile ions in acids)

3. Chemical Properties:

- Turn moist blue litmus paper red
- Acid + metal → salt + hydrogen
- Acid + base → salt + water (neutralisation)
- Acid + carbonate → salt + water+ carbon dioxide

Key Concept III

Naming salt:

Name of salt is made of 2 parts.

For example:

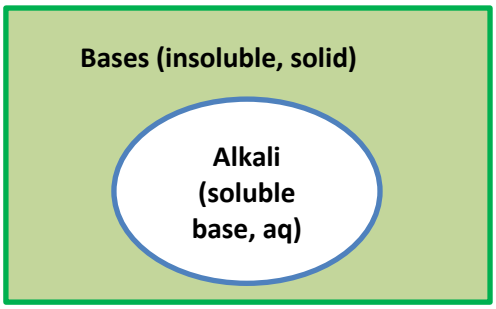
- Naming the **second** part:

Acid used	Type of salt
Hydrochloric acid (HCl: H^+, Cl^-)	Chloride salt
Nitric acid (HNO_3: H^+, NO_3^-)	Nitrate salt
Sulfuric acid (H_2SO_4: H^+, SO_4^{2-})	Sulfate salt

For instance: Zinc oxide + nitric acid → zinc nitrate + water

Key Concept IV

Bases

1. Bases are **oxides or hydroxides of metals.** (i.e. sodium oxide, sodium hydroxide)
2. **Alkali is a soluble base.**
3. 

4. **Alkali** produces **hydroxide ions** (OH^-) in aqueous solution (i.e. when dissolved in water).

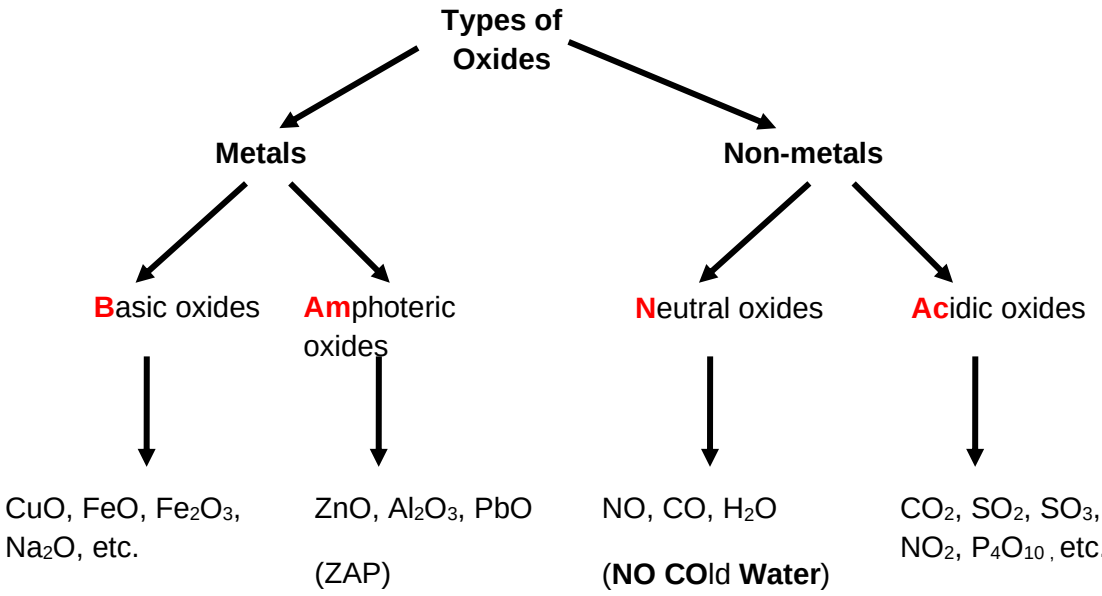
5. **Physical Properties**:
   - Conduct electricity (there are mobile ions in alkalis)

6. Chemical Properties:
   - Turn moist red litmus blue
   - Acid + base → salt + water (neutralisation)
   - Base + ammonium salt → salt + water + ammonia gas

7. Uses of bases:
Calcium hydroxide is used to neutralize the acidic soil as most plants do not grow well in acidic soils.

Key Concept V

Types of Oxides



1 **Basic oxides** (or bases) react with acids or acidic oxides to form salt and water.

2 **Amphoteric oxides** react with BOTH acids or acidic oxides to form salt and water.
(*Amphi* – both, i.e. amphibians live on both land and water)

3 **Acidic oxides** react with bases to form salt and water.

4 **Neutral oxides** are (neutral), that is, neither acidic nor basic
(i.e. does NOT react with acids or bases)

Key Concept VI

Neutralisation (ionic equation)

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

5.2 Salts

Key Concept I

1. A salt is formed when the hydrogen ions of an acid is replaced by a metal or ammonium ions.

2. Not all salts are soluble.

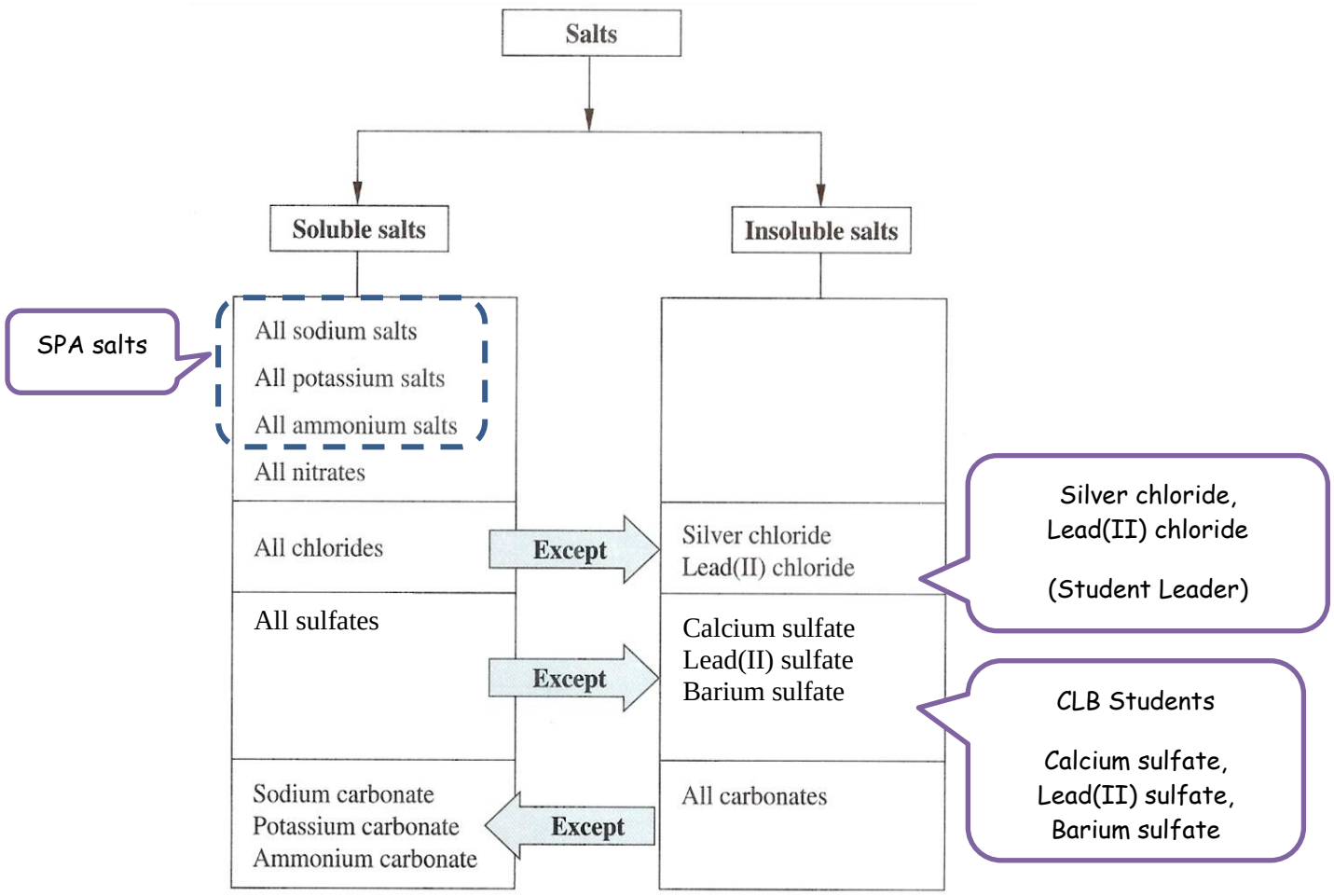
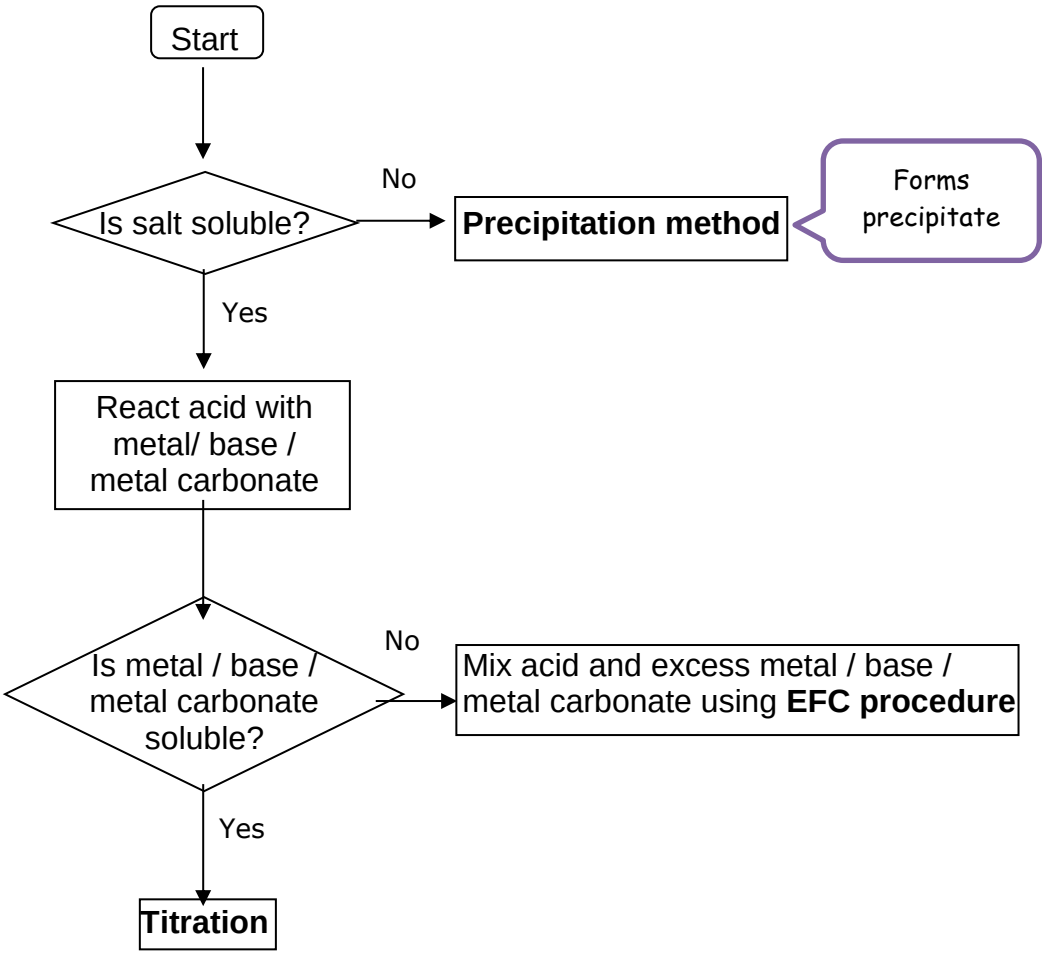


Table 9.2 The solubilities of common salts in water

Key Concept II
Preparation of Salts

Flow chart showing which method to use for preparation of salt



Preparation of **soluble salts**

Soluble salts are prepared from acids. There are four reactions possible:

(a) Action of acid on **reactive metals**, e.g. Mg, Zn, Fe, etc.
(b) Action of acid on **insoluble bases**, e.g. CuO, PbO, etc.
(c) Action of acid on **insoluble carbonates**, e.g. $CuCO_3$, $CaCO_3$, etc.

} EFC /MBC procedure

Note: MBC: Metal, Base, Carbonate
EFC: Excess, Filter, Crystallise

(d) Action of acid on **alkalis**, e.g. NaOH, KOH, Na_2CO_3, K_2CO_3, etc. } titration

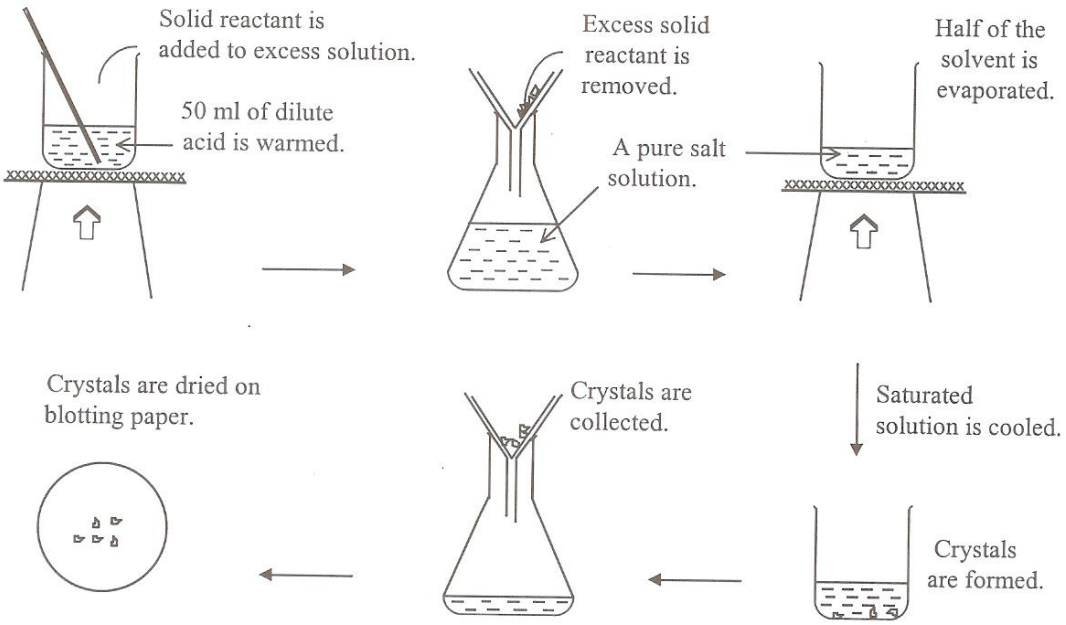
ANSWERING PREPARATION OF SALTS QUESTIONS

****If the question ask you to prepare a **soluble salt** using **solid reactant**.*

Method 1: Acid + MBC → soluble salt + water

For reactions (a), (b), (c), the second reactant is a solid insoluble in water. The procedure for preparation is illustrated below.

mbc procedure – dissolution – filtration – crystallization - filtration – drying



The steps for the preparation of a soluble salt are given below.

(a) Pour about 50 cm^3 of the acid in a beaker.

(b) Add the solid reactant to the acid. (Warm and stir the mixture to speed up the reaction.)

(c) The **solid reactant** is added till in **excess**, to ensure that the acid is completely neutralised.

(d) Filter the mixture to remove the excess solid reactant. The filtrate collected is a pure salt solution.

(e) Evaporate the filtrate to about half its volume to get a hot saturated solution. Leave it to cool and allow crystals to form.

(f) Filter to collect the crystals. Dry them between sheets of filter paper.

***If the question ask you to prepare **soluble salt** using an **alkali**.*

Method 2 (Titration) : Acid + Alkali → soluble salt + water
(Soluble base)

This procedure is used for preparing salts of sodium, potassium and ammonia.

The soluble salt is obtained by reacting an acid with an alkali.

The steps for the preparation of a soluble salt are given below.

(a) Pipette 25 cm^3 of the alkali and put into a conical flask.

(b) Add a few drops of a suitable indicator into the alkali.

Don't have to name the indicator

(c) Fill a burette with the dilute acid. Note its initial reading (let's say V_1 cm^3).

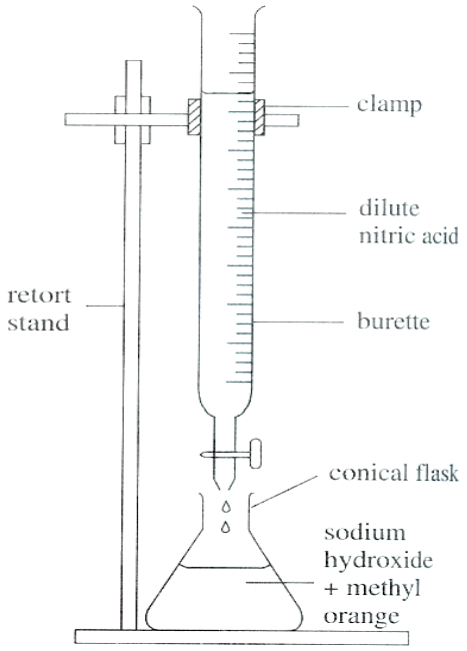
(d) Add the acid from the burette slowly and towards the end-point, drop by drop until the indicator suddenly changes its colour to show that the end-point is reached.
Note the final burette reading (let's say V_2 cm^3).
V_2-V_1 cm^3 of the acid is required to neutralize the alkali in the flask.

Don't have to name the colour change.

(e) REPEAT the experiment by pipetting a fresh volume of 25 cm^3 of the alkali and add to it, V_2-V_1 cm^3 of the acid from the burette. This mixture is the salt solution required.

(f) Evaporate the solution to get a saturated solution. Cool the saturated solution to obtain crystals.

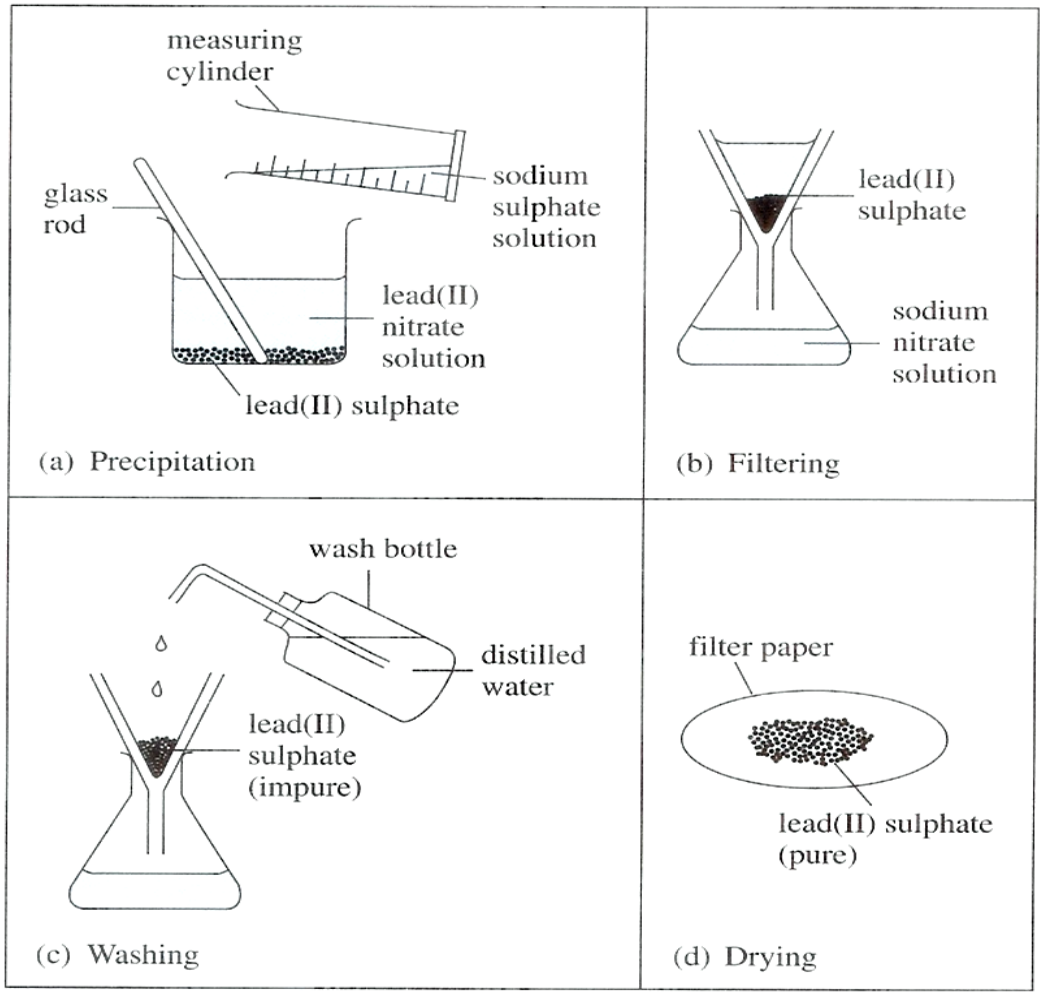
(g) Filter to collect the crystals. Wash crystals using distilled water, and dry using filter paper.



****If the question ask you to prepare an **<u>insoluble salt</u>** from soluble reactants.*

Method 3 (Precipitation)

This precipitation method involves mixing two aqueous solutions to obtain the **<u>insoluble salt</u>** as a precipitate, which is filtered and then washed with distilled water.



Tip: To get an insoluble salt MX which is made up of cation of M and anion of X, mix a nitrate solution of M with a sodium, potassium or ammonium solution of X.

e.g.

M-nitrate solution + sodium-X solution → MX precipitate + sodium nitrate solution

M-nitrate solution + potassium-X solution → MX precipitate + potassium nitrate solution

M-nitrate solution + ammonium-X solution → MX precipitate + ammonium nitrate solution

5.3 Ammonia

1. The Haber Process

Ammonia is an important chemical that is manufactured in large amounts through the Haber process. It is produced from nitrogen gas and hydrogen gas. Nitrogen gas is obtained directly from the air and hydrogen gas is obtained from the cracking of large hydrocarbons.

As the formation of ammonia from hydrogen and nitrogen is a reversible process, reaction conditions are controlled to maximise the yield of ammonia.

nitrogen + hydrogen ⇌ ammonia

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

The production of ammonia is favoured at low temperatures. These low temperatures however, are kinetically unfavourable as the reaction would proceed too slowly. Therefore a relatively high temperature of 450 °C is used.

Higher pressures result in a higher yield of ammonia. Despite this, it is expensive to generate and maintain high pressures and have equipment that can withstand extreme pressures. Considering these costs, the Haber process usually takes place at 250 atm.

An iron catalyst is also used to further increase the rate of reaction.

2. Displacement of Ammonia

Ammonia is displaced when ammonium salts are heated with alkalis.

ammonium chloride + sodium hydroxide → ammonia + sodium chloride + water

$NH_4Cl + NaOH \rightarrow NH_3 + NaCl + H_2O$

3. Uses of Ammonia

Ammonium nitrate, which is made from ammonia, is used as a fertiliser. Nitrogen is needed for the production of proteins for healthy plant growth. While nitrogen is abundant in the air, most plants cannot utilise atmospheric nitrogen. Nitrogen is supplied to plants in the form of ammonium salts and urea.

Ammonium fertilisers cannot be added alongside agricultural lime (calcium hydroxide and calcium oxide) as ammonia would be displaced from ammonium salts. This would cause wastage of the fertiliser as ammonia gas cannot be utilised by plants.

Ammonia is also used to manufacture nitric acid, explosives, plastics, textiles, pesticides and dyes.

6 Qualitative Analysis of Salt

Test for anions

anion	test	test result
carbonate (CO_3^{2-})	add dilute acid	effervescence, carbon dioxide produced
chloride (Cl^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
iodide (I^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate (NO_3^-) [in solution]	add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
sulfate (SO_4^{2-}) [in solution]	acidify with dilute nitric acid, then add aqueous barium nitrate	white ppt.

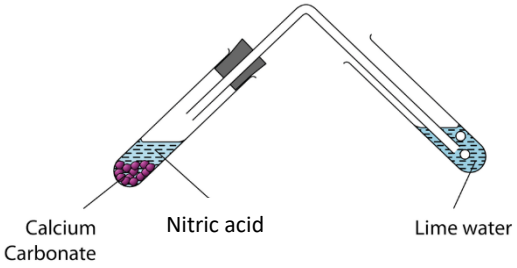
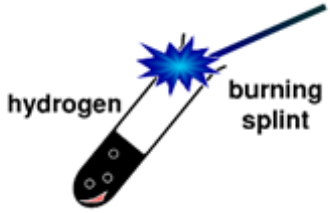
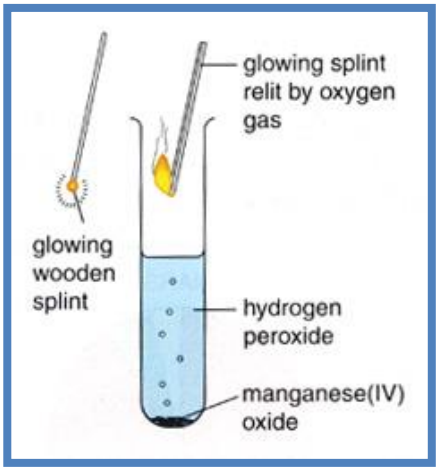
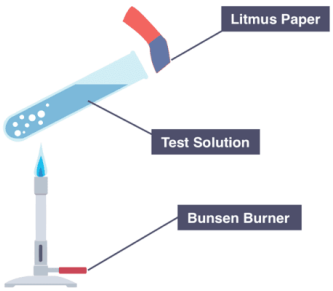
Test for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium (Al^{3+})	white ppt., soluble in excess giving a colourless solution	white ppt., insoluble in excess
ammonium (NH_4^+)	ammonia produced on warming	–
calcium (Ca^{2+})	white ppt., insoluble in excess	no ppt.
copper(II) (Cu^{2+})	light blue ppt., insoluble in excess	light blue ppt., soluble in excess giving a dark blue solution
iron(II) (Fe^{2+})	green ppt., insoluble in excess	green ppt., insoluble in excess
iron(III) (Fe^{3+})	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc (Zn^{2+})	white ppt., soluble in excess giving a colourless solution	white ppt., soluble in excess giving a colourless solution

Test for gases

gas	test and test result
ammonia (NH_3)	turns damp red litmus paper blue
carbon dioxide (CO_2)	gives white ppt. with limewater (ppt. dissolves with excess CO_2)
chlorine (Cl_2)	bleaches damp litmus paper
hydrogen (H_2)	'pops' with a lighted splint
oxygen (O_2)	relights a glowing splint
sulfur dioxide (SO_2)	turns aqueous acidified potassium manganate(VII) from purple to colourless

Identification of gases

Gas	Test	Result
Carbon dioxide	Bubble the gas through limewater 	White precipitate formed in the limewater
Hydrogen	Place a burning splint at the mouth of the test tube 	The burning splint extinguishes with a 'pop' sound
Oxygen	Insert a glowing splint 	Glowing splint relights
Ammonia	Place a damp red litmus paper at the mouth of test tube 	The damp red litmus paper turns blue.

7.1 Redox

Key Concept I: Identifying redox reaction

Redox refers to the processes of **red**uction and **ox**idation, which occurs **simultaneously**.

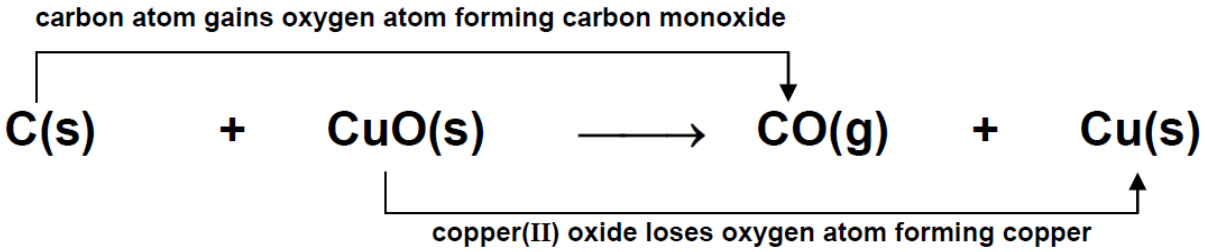
	oxidation	**reduction**
oxygen	gain	lost
hydrogen	lost	gain
electron	lost	gain
oxidation state	increase	decrease

- Oxidising agent: **acidified potassium manganate(VII)**
  Observation: Acidified potassium manganate(VII) solution changes from **purple** to **colourless** in the presence of a reducing agent

- Reducing agent: **aqueous potassium iodide**
  Observation: Aqueous potassium iodide changes from **colourless** to **brown** in the presence of an oxidising agent

Gain / Loss of Oxygen

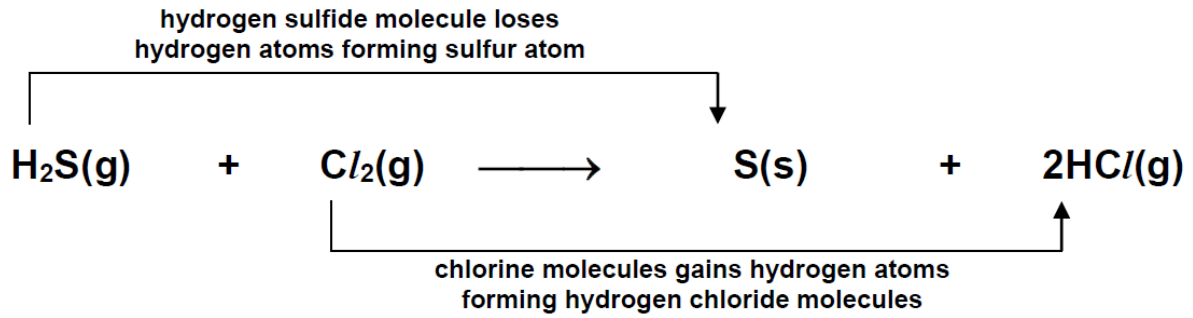
Example 1:



CuO is **reduced** because CuO **loses oxygen** to form **Cu**.
C is **oxidised** because C **gains oxygen** to form **CO**.

Gain / Loss of Hydrogen

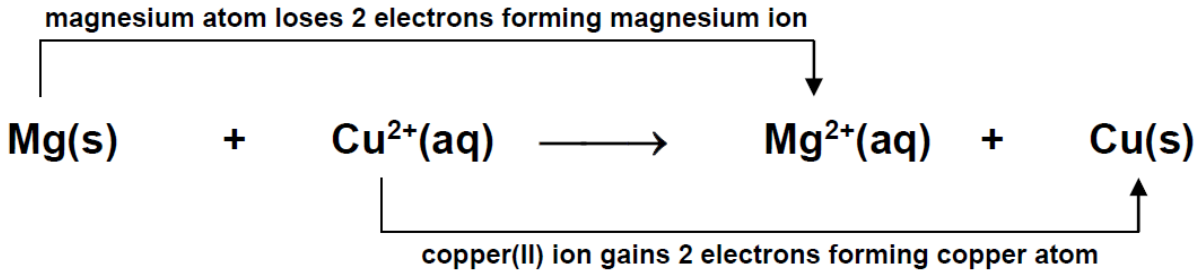
Example 2:



Cl_2 is **reduced** because Cl_2 **gains hydrogen** to form **HCl**.
H_2S is **oxidised** because H_2S **loses hydrogen** to form **S**.

Gain / Loss of Electrons

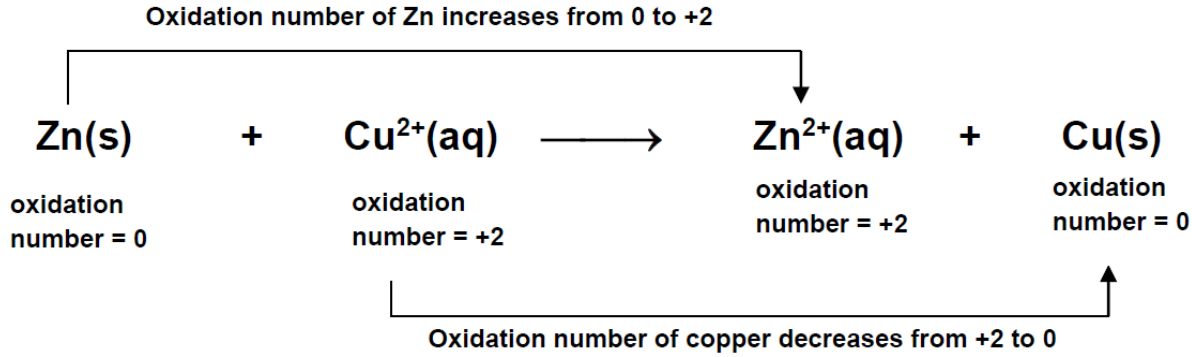
Example 3:



Cu^{2+} is **reduced** because Cu^{2+} **gains electrons** to form **Cu**.
Mg is **oxidised** because Mg **loses electrons** to form **Mg^{2+}**.

Increase / Decrease in Oxidation State

Example 4:



Cu^{2+} is **reduced** because the oxidation state of copper **decreased** from **+2** in **Cu^{2+}** to **0** in **Cu**.
Zn is **oxidised** because the oxidation state of zinc **increased** from **0** in **Zn** to **+2** in **Zn^{2+}**.

Oxidation State: It is the charge an atom of an element would have **if** it exists as a simple ion in a compound, even if it is covalently bonded.

Note: For oxidation state, the **sign** is always placed **in front of** the **number**. E.g. oxidation state of chlorine in Cl^- = **–1**

Key Concept II: 8 Rules for Determining Oxidation State

rule	example	oxidation state
1. The oxidation state of an **element** is 0.	Cu	oxidation state of Cu = 0
	Cl_2	oxidation state of chlorine in Cl_2 = 0
2. The oxidation state of a **simple ion** is the same as the charge on the ion.	K^+	oxidation state of potassium in K^+ = +1
	Cu^{2+}	oxidation state of copper in Cu^{2+} = +2
	Fe^{3+}	oxidation state of iron in Fe^{3+} = +3
	N^{3-}	oxidation state of nitrogen in N^{3-} = –3
	O^{2-}	oxidation state of oxygen in O^{2-} = –2
	Cl^-	oxidation state of chlorine in Cl^- = –1
3. • Oxidation state of hydrogen with **non-metals** is **+1**. • Oxidation state of hydrogen with **metals** is **–1**. • Oxidation state of hydrogen in H_2 (**element**) = 0 (Rule 1)	NH_3	oxidation state of hydrogen in NH_3 = +1
	NaH sodium hydride	oxidation state of hydrogen in NaH = –1
	H_2	oxidation state of hydrogen in H_2 = 0
4. • Oxidation state of oxygen as an **oxide** (O^{2-}) = -2 (Rule 2) • Oxidation state of oxygen as a **peroxide** (O_2^{2-}) = –1 • Oxidation state of oxygen as a **superoxide** (O_2^-) = $-\frac{1}{2}$ • Oxidation state of oxygen in O_2 (**element**) = 0 (Rule 1)	Na_2O sodium oxide	oxidation state of oxygen in Na_2O = –2
	H_2O_2 hydrogen peroxide	oxidation state of oxygen in H_2O_2 = –1
	KO_2 potassium superoxide	oxidation state of oxygen in KO_2 = $-\frac{1}{2}$
	O_2	oxidation state of oxygen in O_2 = 0
5. Sum of the oxidation states of the atoms present in a formula of a **compound** = 0.	MgO +2 –2	+2 + (–2) = 0
	K_2O 2(+1) –2	2(+1) + (–2) = 0
6. Sum of oxidation states of the atoms in a **polyatomic ion** is equal to the charge on the ion.	OH^- –2 +1	–2 + 1 = –1
	NO_3^-	

rule	example	oxidation state
	CO_3^{2-}	
	SO_4^{2-}	
	NH_4^+	
7. Roman numerals are used for compounds containing elements with variable oxidation states.	Cu_2O copper(I) oxide	oxidation state of copper in Cu_2O = +1
	CuO copper(II) oxide	oxidation state of copper in CuO = +2
	K**Mn**O_4 potassium manganate(VII)	oxidation state of manganese in K**Mn**O_4 = +7
	K_2**Cr**$_2O_7$ potassium dichromate(VI)	oxidation state of chromium in K_2**Cr**$_2O_7$ = +6

8. Transition metals and some other elements have variable oxidation states.

oxidation state	–2	–1	0	+1	+2	+3	+4	+5	+6	+7
manganese			Mn		**Mn**Cl_2		**Mn**O_2			K**Mn**O_4 potassium manganate(VII)
chromium			Cr		**Cr**Cl_2	**Cr**Cl_3			K_2**Cr**$_2O_7$ potassium dichromate(VI)	
iron			Fe		**Fe**Cl_2	**Fe**Cl_3				
sulfur	Fe**S**		S				**S**O_2		H_2**S**O_4	
carbon			C		**C**O		Ca**C**O_3			

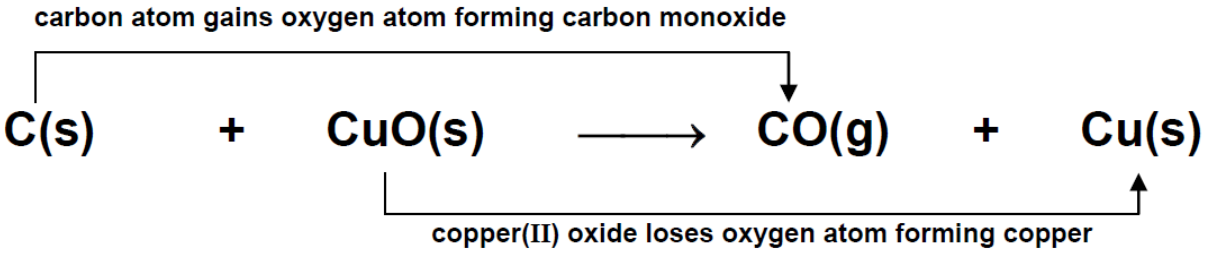
Key Concept III: Oxidising Agent and Reducing Agent

Oxidising Agent: a substance that causes another substance to be oxidised and is itself reduced.

Reducing Agent: a substance that causes another substance to be reduced and is itself oxidised.

Examples of Oxidising Agents and Reducing Agents

oxidising agent	reducing agent
potassium manganate(VII) Mn^{7+} → Mn^{2+} purple colourless	potassium iodide $2I^-$ → I_2 colourless brown
chlorine Cl_2 → $2Cl^-$ greenish-yellow colourless	
iron(III) chloride Fe^{3+} → Fe^{2+} pale yellow green	

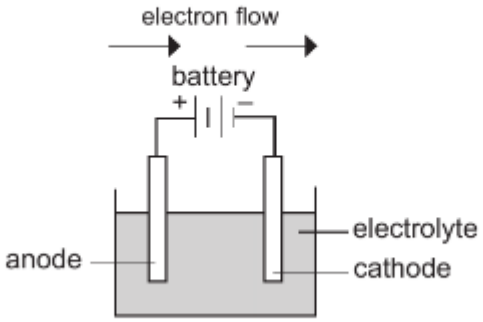


C is a **reducing agent** because C **gains oxygen** to form **CO**.
CuO is an **oxidising agent** because CuO **loses oxygen** to form **Cu**.

7.2 Electrolysis

1. Electrolytic Cell

Electrolysis is the use of electricity to break down a compound into its constituents. The process takes place in an electrolytic cell.



The battery provides a source of electricity for reactions to occur.

The electrodes used in electrolysis conduct electricity. Inert graphite or platinum electrodes are usually used.

The electrode connected to the positive terminal of the battery is the anode and the electrode connected to the negative terminal of the battery is the cathode. Reduction occurs at the cathode while oxidation occurs at the anode.

The electrolyte contains mobile ions to conduct electricity. It is usually an acid solution, or an ionic compound that is molten or dissolved in water. A solid ionic compound cannot be used as its ions are in fixed positions in the crystal lattice structure.

During electrolysis, electrons flow from the negative terminal of the battery to the cathode, and from the anode to the positive terminal of the battery.

2. Electrolysis of Molten Ionic Compounds

When an ionic compound is molten, it splits up into positive ions (cations) and negative ions (anions) which are free to move to the cathode and the anode respectively.

At the cathode, electrons are taken in by cations, while at the anode, electrons are lost by anions. To maintain a complete electrical circuit, the number of electrons taken in at the cathode must be the same as the number of electrons lost at the anode.

Since cations take in electrons, they are reduced. Anions are oxidised as they lose electrons.

3. Electrolysis of Solutions of Ionic Compounds

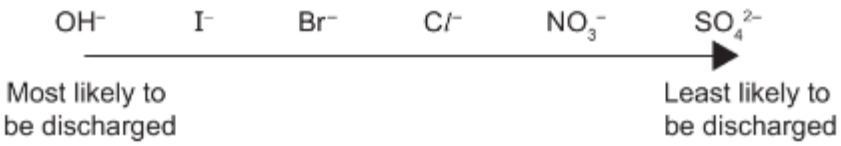
When a solution of an ionic compound is used instead, the autoionisation of water has to be taken into consideration as well.

Water partially dissociates to form hydrogen and hydroxide ions.

water ⇌ hydrogen ion + hydroxide ion
$H_2O \rightleftharpoons H^+ + OH^-$

These ions will compete with those of the ionic compound to be discharged at each of the electrodes.

The ease of discharge of cations can be predicted based on the reactivity series. As reactive metals tend to form ions, their ions are not easily discharged. Ions of less reactive metals have a higher tendency of getting discharged as they accept electrons more easily.



Hydroxide ions are most readily discharged in dilute solutions. Nitrates and sulfates are usually not discharged and tend to stay in the solution.

However, when the solution is concentrated, halide ions are preferentially discharged over hydroxide ions.

4. Purification of Copper

Copper can be purified by electrolysing copper(II) sulfate solution using copper electrodes. Impure copper is used as the anode while pure copper acts as the cathode.

At the anode, OH^- ions are not discharged since the electrode is not inert. Instead, copper atoms are oxidised to form Cu^{2+} ions.

Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

The impure copper anode gradually dissolves as the atoms are oxidised. The impurities are left behind to sink to the bottom of the cell as the anode dissolves.

At the cathode, Cu^{2+} ions in the electrolyte are discharged and deposited on the pure copper.

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

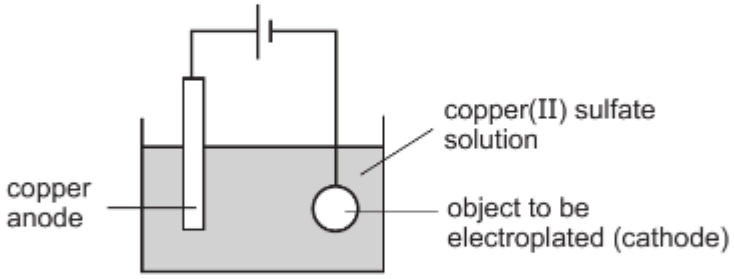
The pure copper cathode gains mass as a layer of pure copper is deposited.

5. Electroplating

Electroplating is done to coat a metal with another metal to improve its appearance or to improve its resistance to corrosion.

The metal used for plating is used as the anode and the object to be electroplated acts as the cathode. The electrolyte used is a salt solution of the metal used for plating.

The plating of an object with copper metal is shown below.



Copper metal acts as the anode as it is used to plate the object. The electrolyte used is a salt solution of the metal used for plating (copper(II) sulfate solution) and the object to be plated acts as the cathode.

At the anode, the copper atoms are oxidised to Cu^{2+} ions, which enter the electrolyte. At the cathode, Cu^{2+} ions are discharged and deposited on the object, plating it with copper metal.

6. Simple Cells

Simple cells convert chemical energy into electrical energy. The cell uses two different metals as electrodes and the voltage produced varies depending on the metals used.

In a simple cell, the more reactive metal acts as the anode while the less reactive metal acts as the cathode.

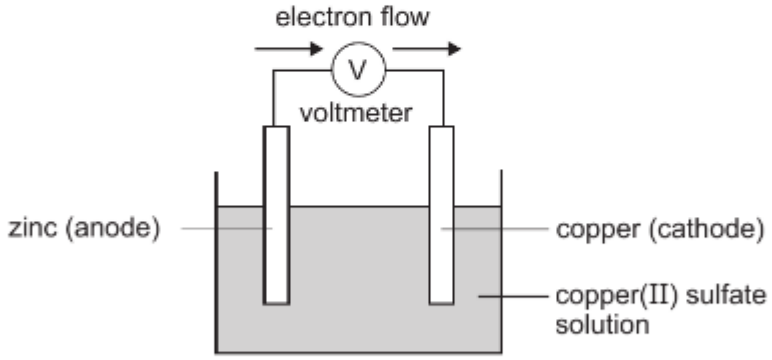
In a zinc-copper cell, zinc metal acts as the anode while copper acts as the cathode.

At the anode, zinc oxidises to form Zn^{2+} ions. In the process, electrons are released and they flow out of the anode through the wire.

Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

Electrons flow from the zinc anode to the copper cathode, where Cu^{2+} ions in the electrolyte are reduced and deposited as copper metal on the cathode.

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$



The zinc anode gradually loses mass as zinc atoms get oxidised while the copper cathode gains mass as Cu^{2+} ions are reduced and deposited. The overall equation of the reaction is obtained by adding the half-equations.

Overall equation: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

A greater voltage is produced when the two metals used are far apart in the reactivity series. A magnesium-copper cell generates a higher voltage than a zinc-copper cell since the difference in reactivity is greater in the magnesium-copper cell.

8.1 – 8.3 The Periodic Table

Key Concept I

1	2										13	14	15	16	17	18
							H									He
Li	Be										B	C	N	O	F	Ne
Na	Mg										Al	Si	P	S	Cl	Ar
K	Ca														Br	Kr
Rb	Sr						Transition metals								I	Xe
Cs	Ba														At	Rn
Fr	Ra															

1. Elements are arranged according to the **proton number**.
2. All members in the **same group** have the **same number of valence electrons**.
3. From left to right across the periodic table, it changes from **metallic to non-metallic**.

Key Concept II

1. All **Group 1** metals are called **alkali metals** (forms alkali with water)
2. **Physical** properties:
    a. Soft
    b. Shiny
    c. Low density (increases down the group)
    d. Low m.p. (decreases down the group)

I	Density (g/cm^3)	M.p. (^oC)
Li	0.53	180
Na	0.97	98
K	0.86	63
Rb	1.53	39

3. **Chemical** Properties:
    a. Similar chemical formula with chlorine (or any non-metals)
    b. Reactivity increases down the group
    c. Reaction with water
        i. Metal + water → alkali (metal hydroxide) + hydrogen

Key Concept III

1. All **Group 17** elements are called **halogens**
2. Exists as **diatomic** covalent molecules (i.e. F_2, Cl_2, Br_2, I_2, At_2)
3. **Physical** properties:
    a. Colour (darker down the group)
    b. Low m.p. (increases down the group)
4. **Chemical** Properties:
    a. decreases down the group
    b. A more reactive halogen will **displace** a less reactive halogen from its halide solution.

VII	Colour	M.p. (^oC)	State at r.t.p
F	Pale yellow	-219	Gas
Cl	Greenish-yellow	-101	Gas
Br	Reddish-brown	-7	Liquid
I	Purplish-black	114	Solid

Key Concept IV

1. All **Group 18** metals are called **noble gases or inert gases (chemically unreactive).**
2. **Chemical** Properties:
    a. Unreactive because they have full electronic structure.

Key Concept V

1. Transition elements are a block of metals found between Group 3 and 12
2. High melting points and high densities
3. Compounds of transition elements are usually coloured.
4. Transition elements have variable oxidation state, can form ions of different charges. Eg iron(II) Fe^{2+} and iron(III) Fe^{3+}
5. Transition elements and their compounds are good catalysts used industrially.

8.4 Reactivity Series (Metals)

Key Concept I

Properties of Metals

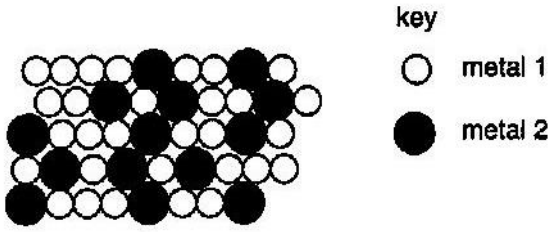
1. Physical properties:
    a. High m.p. and b.p. (therefore solid at room temperature)
    b. High density

    } Except Group I metals

    c. Good conductors of electricity
    d. Malleable (Can be beaten into different shapes)
    e. Ductile (Can bend without breaking)

2. Alloy is a mixture of a metal with another element.
   For example, bronze (copper and tin), brass (copper and zinc), stainless steel (iron, chromium, nickel and carbon)

3. Diagrammatic representation of alloy



Alloys are harder because it is made up of atoms of different sizes. The regular arrangement of atoms in pure metals is disrupted, hence the **layers of atoms** cannot slide over each other easily making alloy stronger and harder.

Key Concept II

Reactivity Series

Order of Reactivity	Mnemonics	Reaction with cold water	Reaction with steam	Reaction with dilute acids
Potassium (Highly reactive)	Please	Very violently	Explosively	Explosively
Sodium	Send	Violently		
Calcium	Camels	Readily		Violently
Magnesium	Monkeys	Slowly	Violently	Rapidly
Aluminium	And	Does not react with cold water	Readily	Moderately fast
(Carbon)	Cute			
Zinc	Zebras		Readily	Moderately fast
Iron	In		Slowly	Slowly
Lead*	Large		Slowly	Slowly
(Hydrogen)	Heavy			
Copper	Cages		Does not react with steam	Does not react with dilute acids
Silver (Least reactive)	Soon			
Gold				

*Reaction of lead with hydrochloric acid forms an <u>insoluble layer of lead(II) chloride</u> (Recall SL chloride), which stops further reaction from taking place.

1. Metals + water → metal hydroxide + hydrogen
2. Metals + steam → metal oxide + hydrogen
3. Metals + acid → salt + hydrogen

**Test for hydrogen: Extinguishes burning splint with a ‘pop’ sound

Key Concept III

Displacement of Metals (in relation to the reactivity series)

Displacement reaction takes place when a more reactive metal is placed in the salt solution of a less reactive metal. This is because a more reactive metal has a higher tendency to form ions.

For example, when copper metal is added in a solution of silver nitrate, copper displace silver from silver nitrate solution.

copper metal + silver nitrate → copper(II) nitrate + silver

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Observations:

1. A silver grey coating is formed on copper metal
2. The solutions changes from colourless to blue.

Key Concept IV

Thermal Stability of Metal Carbonates (in relation to the reactivity series)

Reactive metals form very stable carbonates which do not decompose easily upon heating. Less reactive metals form carbonates that are easily decomposed by heat.

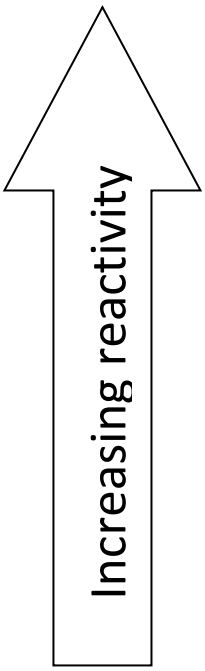
For example, copper(II) carbonate decompose upon heating to form copper(II) oxide and carbon dioxide gas.

$$\underset{\text{Green solid}}{CuCO_3} \rightarrow \underset{\text{Black solid}}{CuO} + CO_2$$

Observations:

1. The green solid turns black
2. Gas produced forms white precipitate in limewater

Key Concept V
Extraction of Metals

<table>
<tr><td rowspan="11">hard to extract
 

Easy to extract</td><td>Metals</td><td>Extraction method</td></tr>
<tr><td>POTASSIUM</td><td rowspan="5"><u>Very reactive metals</u>

Extracted by using electricity to decompose molten metal compounds

(Electrolysis)

This is because of the <u>strong bonds</u> in the <u>ores</u></td></tr>
<tr><td>SODIUM</td></tr>
<tr><td>CALCIUM</td></tr>
<tr><td>MAGNESIUM</td></tr>
<tr><td>ALUMINIUM</td></tr>
<tr><td>ZINC</td><td rowspan="5"><u>Less reactive metals</u>

Extracted by reducing the metal oxides using coke(carbon).

(reduction with carbon)</td></tr>
<tr><td>IRON</td></tr>
<tr><td>LEAD</td></tr>
<tr><td>COPPER</td></tr>
<tr><td>SILVER</td></tr>
<tr><td>GOLD</td><td><u>Unreactive metals</u>

Found naturally uncombined as metals, hence separated by physical methods.</td></tr>
</table>

Key Concepts VI
Rusting of Iron

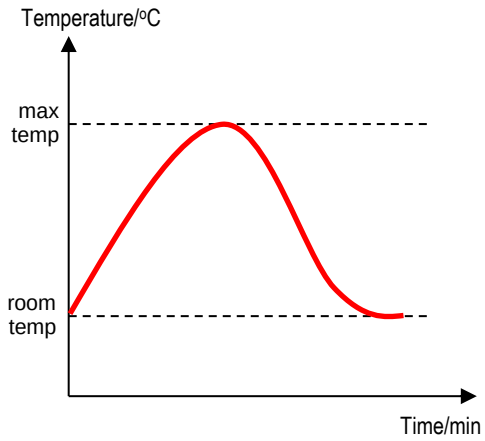
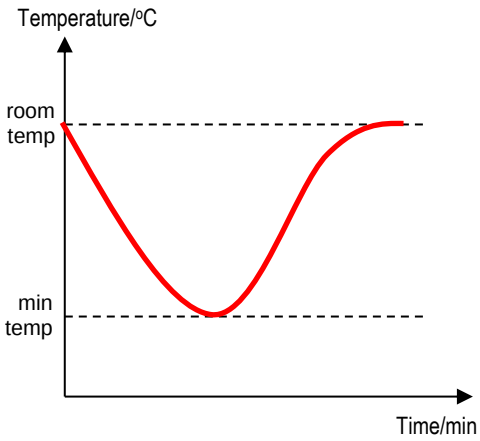
1. Rusting (Corrosion) of iron
2. Condition: <u>oxygen</u> and <u>water</u>
3. Prevention: Painting, greasing, plastic coating, sacrificial protection with a more reactive metal (where more reactive metal corrodes preferentially)

<u>NOTE:</u>
Boiled water does not contain oxygen as it has been driven off during boiling. Hence, rusting does not take place in boiled water.

9 Energy Changes

Key Concept I:

COMPARISON OF EXOTHERMIC AND ENDOTHERMIC REACTIONS

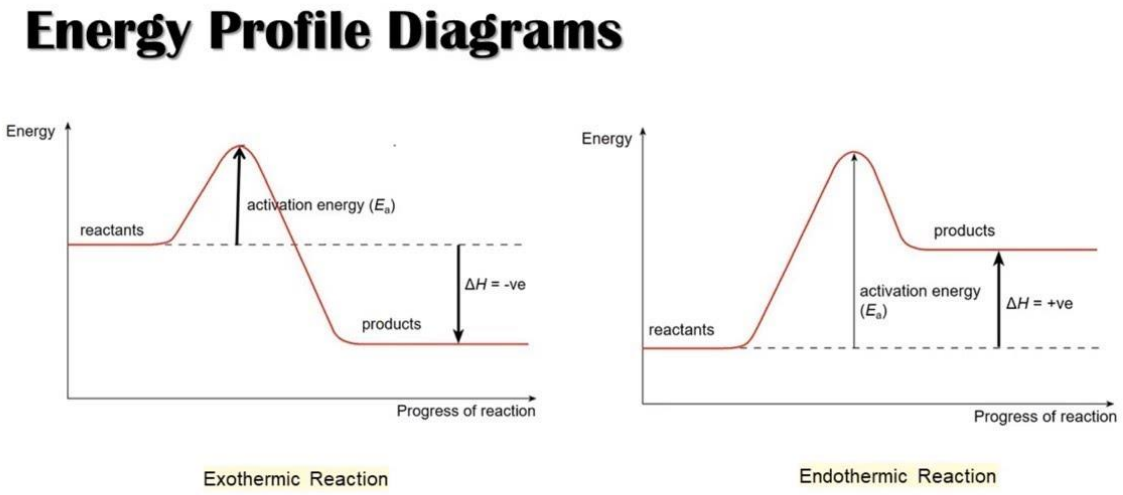
<table>
<tr><th>Differences in characteristics in terms of..</th><th colspan="2">EXOTHERMIC REACTION</th><th colspan="2">ENDOTHERMIC REACTION</th></tr>
<tr><td>net energy transfer
(energy absorbed or given out)</td><td colspan="2">energy given out</td><td colspan="2">energy absorbed</td></tr>
<tr><td>temperature change of surrounding</td><td colspan="2">increases</td><td colspan="2">decreases</td></tr>
<tr><td>temp vs time graph</td><td colspan="2"> 
</td><td colspan="2"> 
</td></tr>
<tr><td>examples</td><td>physical
freezing
condensation
dissolving acid in water</td><td>chemical
combustion of fuels
respiration
neutralisation reactions</td><td>physical
evaporation
melting
dissolving of ammonium chloride in water</td><td>chemical
photosynthesis
thermal decomposition</td></tr>
</table>

Key Concept II: Bond Making and Bond Breaking

Bond Forming of products – energy released (exothermic process)
Bond Breaking of reactants – energy absorbed (endothermic process)

$$\Delta H = \text{Total energy absorbed for bond breaking} - \text{Total energy released for bond making}$$

Key Concept III: Activation Energy and Energy Profile Diagram



Key concept VI: Fuels

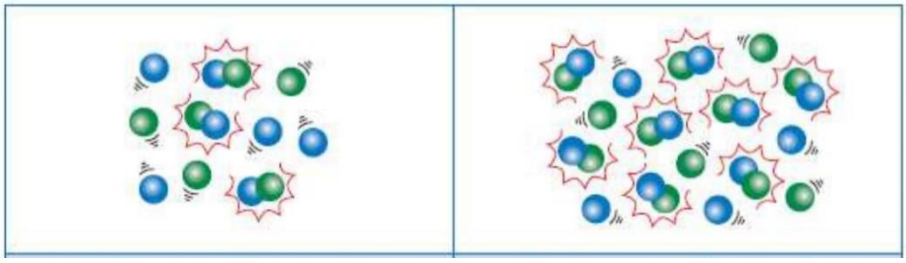
Hydrogen oxygen fuel cell can be used to release energy.

Anode: $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$
Cathode: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
Overall: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

10 Speed of Reaction

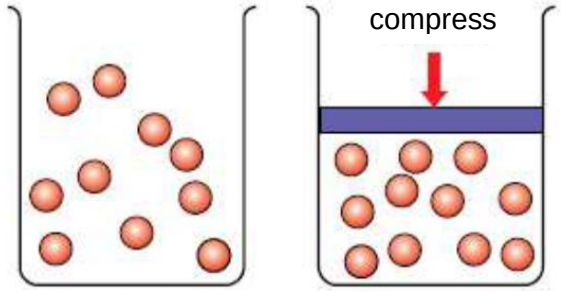
Key Concept I: Factors Affecting Speed of Reaction

1. Concentration (for liquids)



low concentration	high concentration
When the concentration is **decreased**, there will be **fewer** particles per unit volume. So, there will be **fewer** frequent effective collisions per unit time. Hence, the speed of reaction **decreases**.	When the concentration is **increased**, there will be **more** particles per unit volume. So, there will be **more** frequent effective collisions per unit time. Hence, the speed of reaction **increases**.

2. Pressure (for gases)

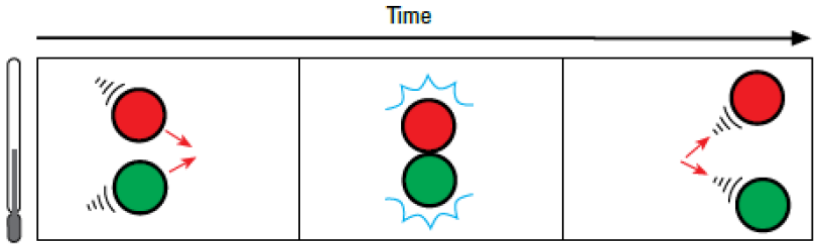


Container **A** (normal pressure)

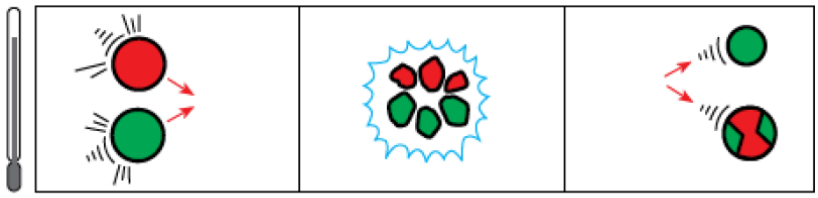
Container **B** (increased pressure)

low pressure	high pressure
When the pressure is **decreased**, there will be **fewer** particles per unit volume. So, there will be **fewer** frequent effective collisions per unit time. Hence, the speed of reaction **decreases**.	When the pressure is **increased**, there will be **more** particles per unit volume. So, there will be **more** frequent effective collisions per unit time. Hence, the speed of reaction **increases**.

3. Temperature (for liquids)



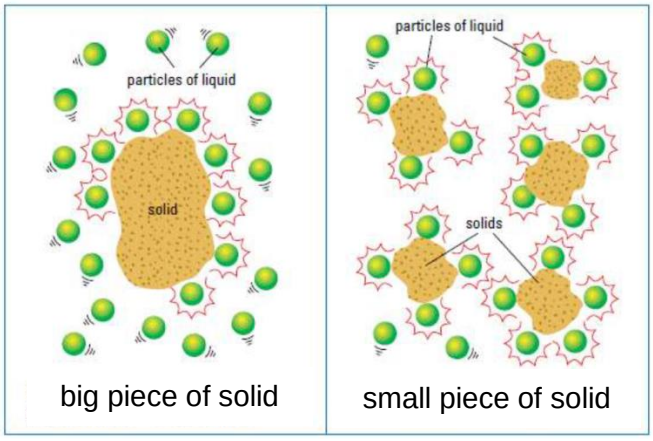
Particles bounce apart at low temperatures without reacting



High-speed collisions at higher temperatures more often produce a reaction.

low temperature	high temperature
When the temperature is **decreased**, the particles **lose** energy and move **slower**. So, there will be **fewer** frequent effective collisions per unit time. Hence, the speed of reaction **decreases**.	When the temperature is **increased**, the particles **gain** energy and move **faster**. So, there will be **more** frequent effective collisions per unit time. Hence, the speed of reaction **increases**.

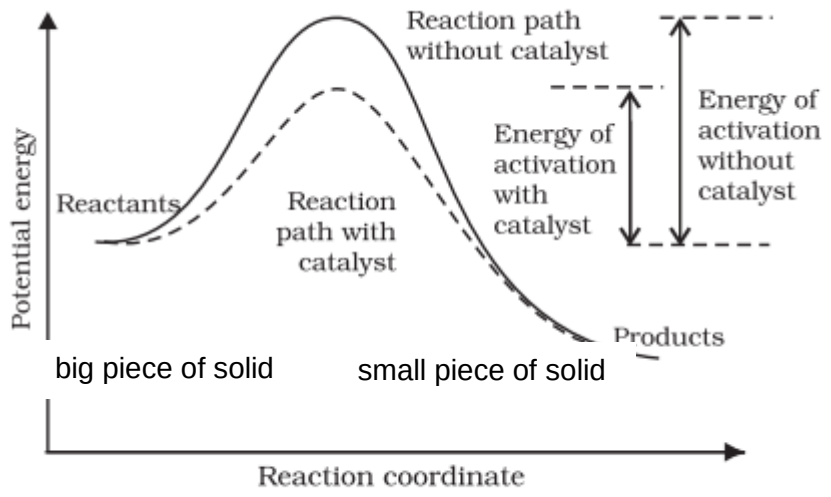
4. Particle size (for solids)



big piece of solid | small piece of solid

large particle size	small particle size
When the particle size is **increased**, the surface area **decreases**. So, there will be **fewer** frequent collisions effective per unit time. Hence, the speed of reaction **decreases**.	When the particle size is **decreased**, the surface area **increases**. So, there will be **more** frequent collisions effective per unit time. Hence, the speed of reaction **increases**.

5. Catalyst



Effect of catalyst on activation energy

Catalyst provides an alternative pathway with lower activation energy.

This results **more** frequent collisions effective per unit time. Hence, the speed of reaction **increases**.

Catalyst	Application
Iron	Haber process (manufacture of ammonia)
Platinum	Catalytic converters
Aluminium oxide (alumina) or silicon dioxide (silica)	Cracking of large hydrocarbons
Nickel	Hydrogenation of alkenes

Key Concept II: Measuring Speed of Reaction

Speed of reaction tells us how fast or slow a reaction takes place.

reactant + reactant → product + product

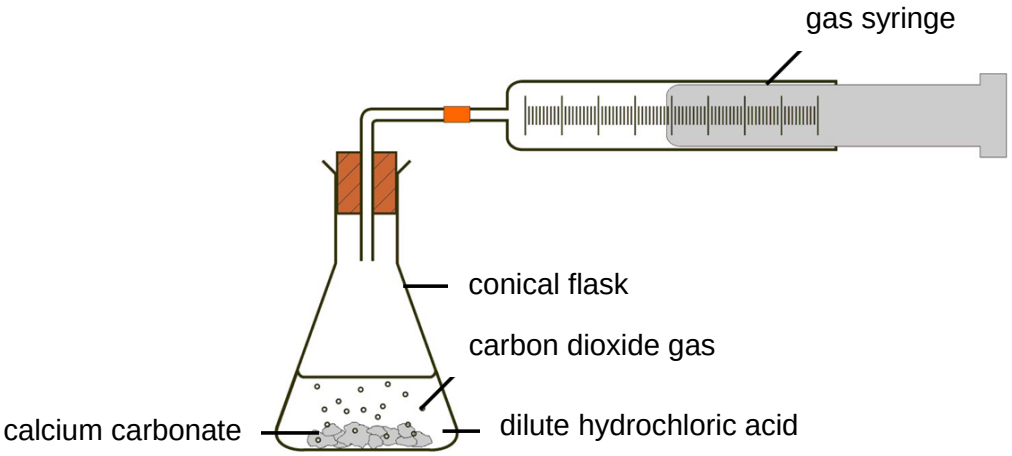
We can find the speed of reaction by:

- measuring the **amount of product** formed in a certain time;
- measuring the **amount of reactant(s)** used up in a certain time.

Method 1: *Measuring the volume of gas produced (amount of product)*

Consider the following reaction between marble (calcium carbonate) and dilute hydrochloric acid.

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

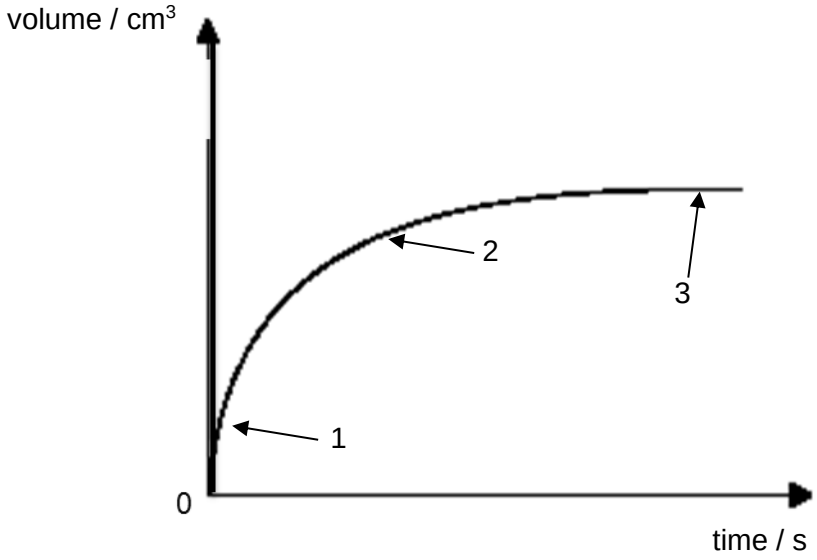


(i) Set up the experiment as shown above.

(ii) Collect and measure the volume of gas produced every 30 seconds.

time/ s	volume/ cm^3
30	13
60	18
90	22
120	25
150	28
180	31
210	33
240	34
270	35
300	35
330	35
360	35

(iii) From the results, plot a graph of volume of gas against time.



Interpretation of Graph

1: The gradient is the steepest at the start of the reaction, so the speed of reaction is the **fastest**.

2: The gradient decreases. The speed of reaction **decreases** with time.

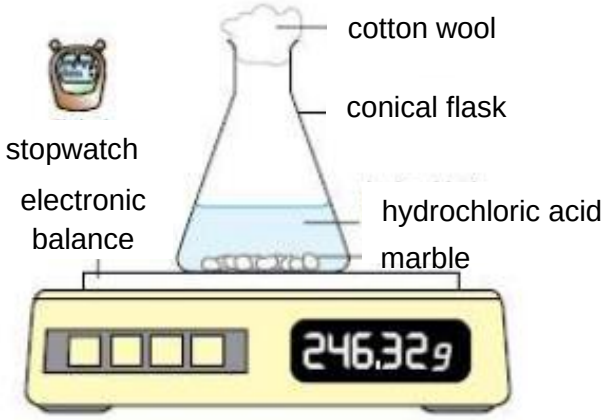
3: The gradient is 0 as the reaction has stopped. The limiting reactant is used up. No more gas is produced. The speed of reaction is **0**.

Method 2: Measuring Change in Mass (amount of reactants left)

Consider the following reaction between marble (calcium carbonate) and dilute hydrochloric acid.

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

There will be a decrease in mass as **carbon dioxide escapes into the surrounding**.



(i) Set up the experiment as shown above. The **cotton wool** in the mouth of the conical flask is to **prevent acid spray** (prevent acid from splashing out as reaction takes place)

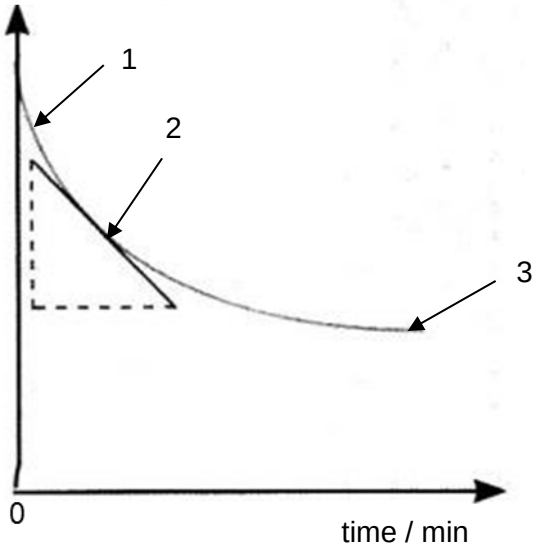
(ii) Record the mass of the conical flask and its contents.

(iii) Record the mass of the conical flask and its contents every 1 minute.

time/ min	mass/ g
1	58.8
2	58.3
3	57.9
4	57.5
5	57.4
6	57.2
7	57.1
8	57.1
9	57.1
10	57.1

(iv) From the results, plot a graph of mass of the conical flask and its contents against time.

mass of conical flask and its contents / g



Interpretation of Graph

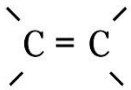
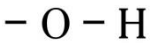
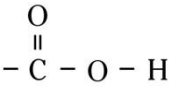
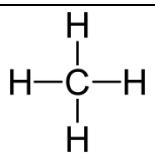
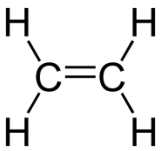
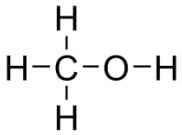
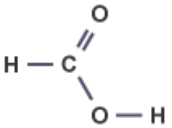
1: The gradient is the steepest at the start of the reaction, so the speed of reaction is the **fastest**.

2: The gradient decreases. The speed of reaction **decreases** with time.

3: The gradient is 0 as the reaction has stopped. The limiting reactant is used up. No more change in mass. The speed of reaction is **0**.

11 Organic Chemistry

Homologous Series

homologous Series	alkanes	alkenes	alcohols	carboxylic acids
functional group	no functional group	C=C  carbon-carbon double bond	–OH  hydroxyl group	–COOH  carboxyl group
general formula (*n* represents number of carbon atoms)	C_nH_{2n+2} $n \geq 1$	C_nH_{2n} $n \geq 2$	$C_nH_{2n+1}OH$ $n \geq 1$	$C_nH_{2n+1}COOH$ $n \geq 0$
type of reaction	undergo combustion and substitution short-chain alkanes are produced by cracking	undergo combustion and addition short-chain alkenes are produced by cracking	undergo combustion and oxidation produced by fermentation	undergo reactions with reactive metals, bases and carbonates
full structural formula, chemical name and molecular formula of first member	 methane CH_4	 ethene C_2H_4	 methanol CH_3OH	 methanoic acid HCOOH
remarks	saturated hydrocarbon	unsaturated hydrocarbon	-	-

Fuels and Crude Oil

Key Concept I

1. Crude oil (petroleum) is a very valuable natural resource.

2. Petroleum, coal and natural gas (mainly methane) are **fossil** fuels. They are the residue left behind from the decomposition of ancient plant life over millions of years ago.

3. **Hydrocarbons** are compounds that are made up of carbon and hydrogen atoms only.

4. Petroleum or crude oil is a thick, black liquid. It is a complex mixture consisting mainly of alkanes ranging from methane to large molecules with more than 70 carbon atoms.

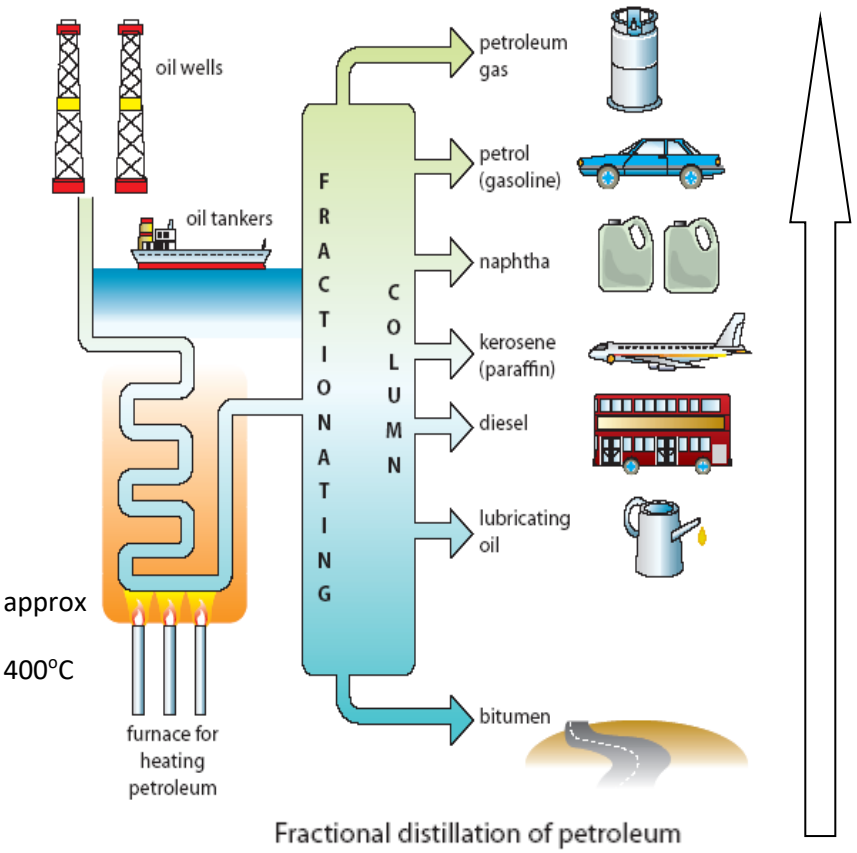
5. A process called **fractional distillation** is used to separate petroleum into useful fractions.

6. Crude oil is heated up to about 400^{o}C and fed into the fractionating column near the bottom.

7. The different components of the petroleum distill over (boil and then condense) depending on the **difference in their boiling points.**

8. The one with the **lowest boiling point** (petroleum gases) is collected first and near the **top** as a gas.

9. Bitumen with the **highest boiling point** is collected as a solid at the **bottom** of the column.



Fractional distillation of petroleum

REMEMBER:

Peter Pan N Kingkong Don't Like Batman

Going up the fractionating column,

the fractions have

- lower boiling point
- lower viscosity (less viscous if liquid)
- less carbon atoms in their molecules
- burn more easily

Name of fraction	Number of carbon atoms	Boiling point /°C	Uses
petroleum gases (LPG)	1 to 4	-160 to 20	liquefied petroleum gas for cooking; bottled gas for cooking (butane); higher pressure cylinders (propane)
petrol (gasoline)	5 to 11	20 to 60	car fuel
naphtha	7 to 13	60 to 180	feedstock for chemical industry
kerosene (paraffin)	10 to 16	120 to 240	aircraft fuel, fuel for oil stoves and heating
diesel	15 to 25	220 to 250	fuel for diesel engines in buses, trains and ships
lubricating oil	20 to 70	250 to 350	lubricants, source of waxes and polishes
bitumen	over 70	over 350	surfacing roads

Alkanes

Key Concept I

Homologous Series

1. A homologous series is a series of compounds with the same **functional group and general formula.**
2. Characteristics of members in the same series:

- Each member in the same homologous series differs from the next by the unit $–CH_2$.
- All members have **similar chemical properties**.
- There is a **gradation of physical properties**. For e.g. the boiling point increases as the number of carbon atoms increase in one molecule.

Key Concept II

Alkanes

1. general formula C_nH_{2n+2} where n is an integer.
2. Alkanes are **saturated** hydrocarbons, i.e. contain **only single** carbon to carbon covalent bonds.

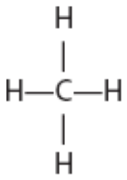
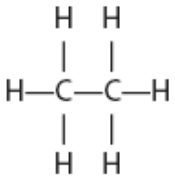
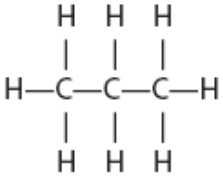
No. of C atoms	Name	Molecular formula	Structural formula	Condensed structural formula
1	methane	CH_4		CH_4
2	ethane	C_2H_6		CH_3CH_3
3	propane	C_3H_8		$CH_3CH_2CH_3$

Table 16.1 The first three alkanes of the alkane homologous series

3. Trend of alkanes

Name	Molecular formula	b.p. / ^{o}C	Physical state at rtp	Solubility in water	Viscosity	Flammability
methane	CH_4	-164	gas	Insoluble	↓ increases ↓	↑ increases ↑
ethane	C_2H_6	-88	gas			
propane	C_3H_8	-42	gas			
butane	C_4H_{10}	-0.5	gas			

Key Concepts III

Physical Properties of alkanes

1. low boiling points (increase as the molecules become larger)

Chemical Properties of alkanes

Generally, alkanes are **unreactive** due to their strong single covalent bonds.

Combustion

Alkanes burn to form carbon dioxide, water and a lot of heat. Their main use is in combustion to produce energy.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Substitution

Alkanes undergo substitution **with halogens (i.e. chlorine)** to form chloroalkanes in the **presence of UV light**.

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$
methane chlorine chloromethane hydrogen chloride

$CH_3Cl(g) + Cl_2(g) \rightarrow CH_2Cl_2(g) + HCl(g)$
dichloromethane

$CH_2Cl_2(g) + Cl_2(g) \rightarrow CHCl_3(g) + HCl(g)$
trichloromethane

$CHCl_3(g) + Cl_2(g) \rightarrow CCl_4(g) + HCl(g)$
tetrachloromethane

A hydrogen atom in methane is replaced by a chlorine atom.

Alkenes

Key Concept I

1. general formula C_nH_{2n} where n is an integer.

2. Alkenes are **unsaturated** hydrocarbons, i.e. contain carbon to carbon **double** covalent bonds.

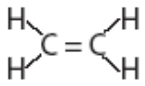
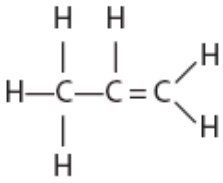
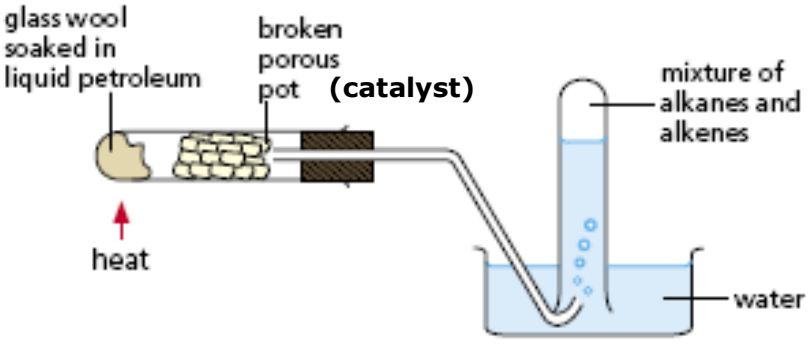
No. of C atoms	Name	Molecular formula	Structural formula	Condensed structural formula
2	ethene	C_2H_4		$CH_2=CH_2$
3	propene	C_3H_6		$CH_3CH=CH_2$

Table 16.3 The first two alkenes

Key Concept II

Preparation of alkenes – cracking

1. **Cracking** is a process where large long-chain alkane molecules are **broken into smaller molecules** using heat (high temperature) and / or catalysts.
2. Cracking of hydrocarbons produces smaller **alkanes, alkenes and hydrogen**.
3. The process usually requires a **catalyst** such as **aluminium oxide** (alumina) or silicon(IV) oxide (silica) and a temperature of **400°C to 500°C**.
4. Cracking of petroleum is important to meet the demands for <u>smaller sized molecules that are more useful.</u>



<u>Set-up for cracking of petroleum</u>

Key Concepts III
Properties of alkenes

1. Alkenes are more reactive than alkanes.

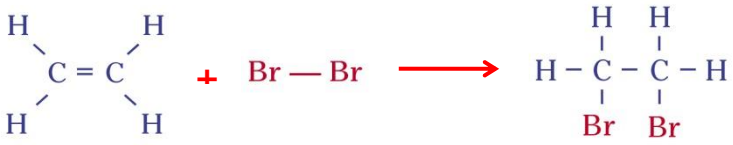
2. **Combustion** of alkenes
   Alkenes burn in air like alkanes to produce carbon dioxide, water and heat. More soot is produced than alkanes.

   $C_2H_4 + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

3. **Addition Reactions**
   Addition reactions involve the breaking of the double bond with the subsequent addition of a molecule to form a new saturated molecule as a single product.

(a) Addition of bromine

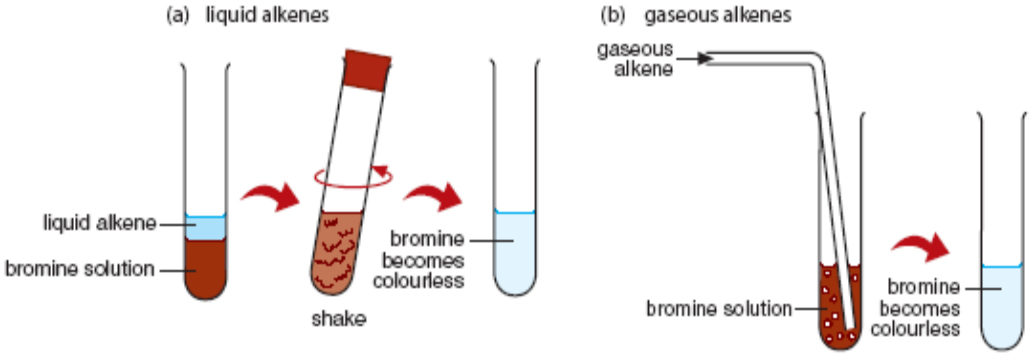
C_2H_4	+	$Br_2(aq)$	→	$C_2H_4Br_2(l)$
ethene				dibromoethane



Alkenes decolourize bromine in the **dark** at room temperature.
When ethene is bubbled through aqueous bromine, the reddish-brown bromine solution becomes colourless.

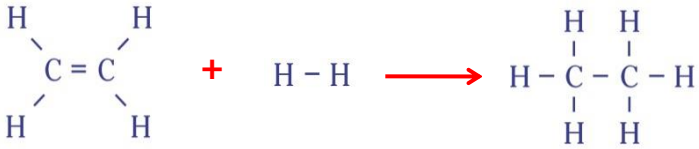
Decolourization of bromine is used as a test for unsaturation.

Alkenes decolourize aqueous bromine in the dark at room temperature. Alkanes do not.



(b) Addition of hydrogen

Ethene reacts with hydrogen in the presence of a **nickel catalyst** and a temperature of **200ºC** to form ethane.



$C_2H_4 + H_2(g) \rightarrow C_2H_6(g)$

ethene ethane

Unsaturated vegetable oils contain a number of carbon-carbon double bonds. It is said to be **polyunsaturated**.

It can be partially saturated by hydrogenation to produce a soft solid known as margarine.

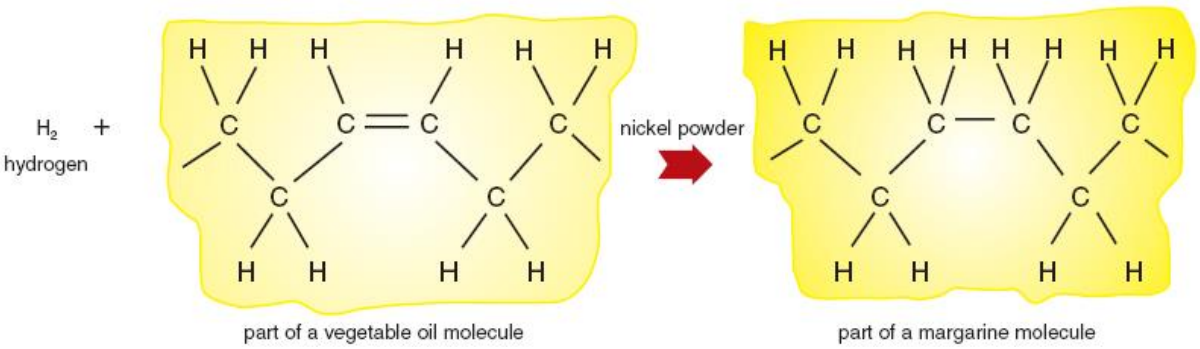
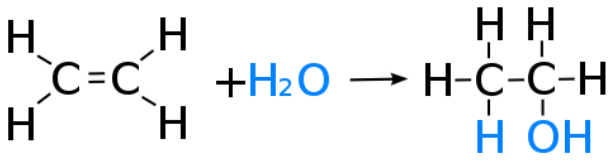


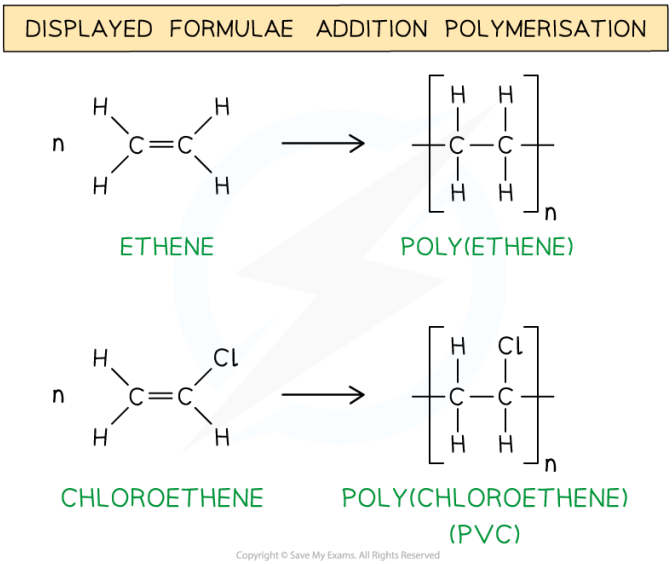
Figure 16.15 Formation of margarine from vegetable oil

(c) Addition of water

Ethene reacts with steam in the presence of a **phosphoric(V) acid catalyst** and a temperature of **300ºC, 60 atm** to form ethanol.



(d) Addition Polyermisation

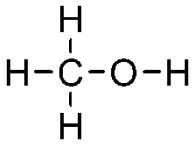
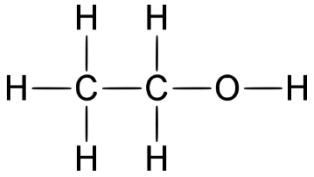
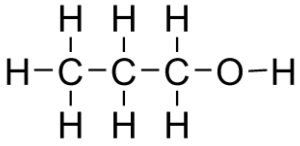


Alcohols

Key Concept I

1. Alcohols are organic compounds which have the **hydroxyl** (**–OH)** functional group and the general formula is **$C_nH_{2n+1}OH$**.

2. First 3 members of alcohols:

chemical name	molecular formula	full structural formula
methanol	CH_3OH	
ethanol	C_2H_5OH	
propanol	C_3H_7OH	

Key Concept II: Reactions of alcohols

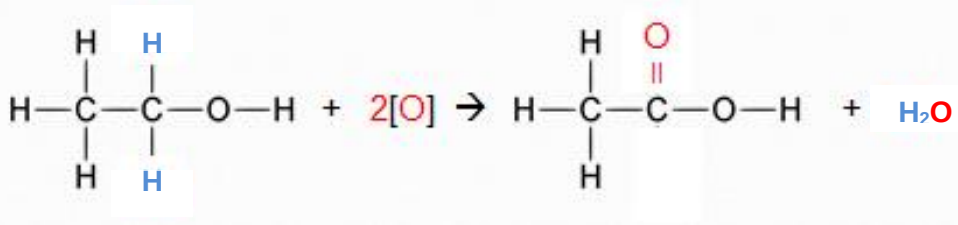
3. Alcohols undergo **combustion** and **oxidation**.

(a) Combustion

- **alcohol + oxygen → carbon dioxide + water**
- E.g. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
- Applications: Alcohols can be used as fuels.

(b) Oxidation

- An alcohol can be oxidised by warming it with acidified potassium manganate(VII) or oxygen in the air.
- Note: Sulfuric acid is used to acidify potassium manganate(VII).
- **alcohol + oxidising agent [O] → carboxylic acid + water**
- E.g. $C_2H_5OH + 2[O]$ or $O_2 \rightarrow CH_3COOH + H_2O$



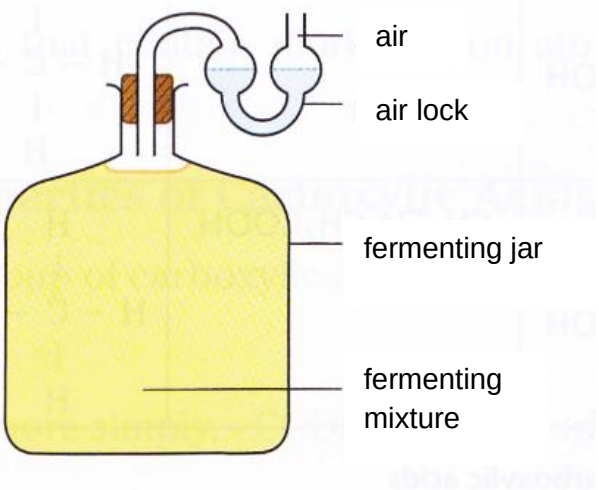
- Observation: **Purple** acidified potassium manganate(VII) turns **colourless**.

Key Concept III: Formation of alcohol (ethanol)

4. Fermentation of glucose

- Fermentation is a chemical process in which yeast acts on carbohydrates to produce ethanol and carbon dioxide.
- **glucose solution → ethanol + carbon dioxide**

  $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$



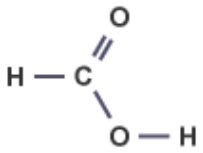
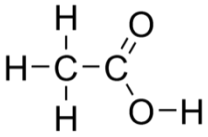
- Process:
  - A glucose solution is mixed with **yeast** and the mixture is kept at a temperature of about **37°C** in the **absence of oxygen**.
  - Yeast contains enzymes which break down glucose to form ethanol and carbon dioxide by fermentation.
  - Ethanol is obtained from the mixture by **fractional distillation**.
- Note:
  - When the alcohol content exceeds 15%, the yeast will die. Hence, fermentation can only produce a low concentration of ethanol.
  - 37°C is the optimum temperature at which enzymes in yeast will work best.
    - If the temperature is raised beyond 37°C, the enzymes will be **denatured** and fermentation will stop.
    - If the temperature is too low, the enzymes will be **dormant**.
  - Alcohol will turn sour if it is exposed to air for a few days as oxygen oxidises ethanol into ethanoic acid. Hence, the apparatus needs to be **air-tight** to **prevent oxygen from oxidising ethanol to ethanoic acid**.
- Applications:
  - To make alcoholic beverages
  - To be used as a fuel
  - To be used as a solvent for making paints or perfumes

Carboxylic Acids

Key Concept I

1. Carboxylic acids are organic acids which have the **carboxyl** (**–COOH)** functional group and the general formula is **$C_nH_{2n+1}COOH$**.

2. First 2 members of carboxylic acids

chemical name	molecular formula	full structural formula
methanoic acid	HCOOH	
ethanoic acid	CH_3COOH	

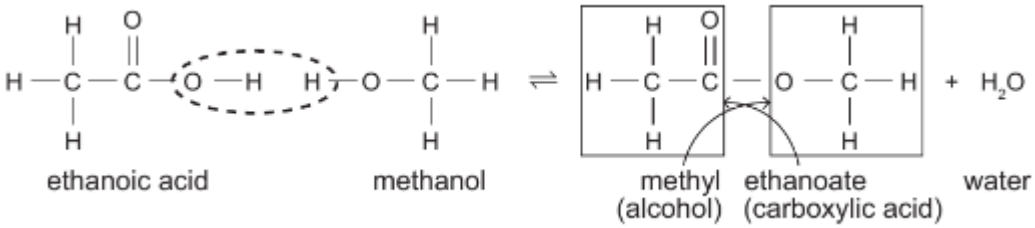
Key Concept II (Formation of ethanoic acid)

3. Ethanoic acid is produced from the oxidation of ethanol by atmospheric oxygen or acidified potassium manganate(VII). (Refer to 10.4 Alcohols)

Key Concept III (Reactions of carboxylic acids)

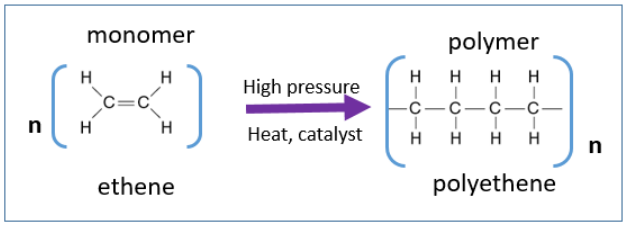
4. Carboxylic acids undergo reactions with reactive metals, bases and carbonates.

   (a) carboxylic acid + reactive metal → salt + hydrogen
   (b) carboxylic acid + base → salt + water
   (c) carboxylic acid + carbonate → salt + carbon dioxide + water
   (d) carboxylic acid + alcohol → esters + water

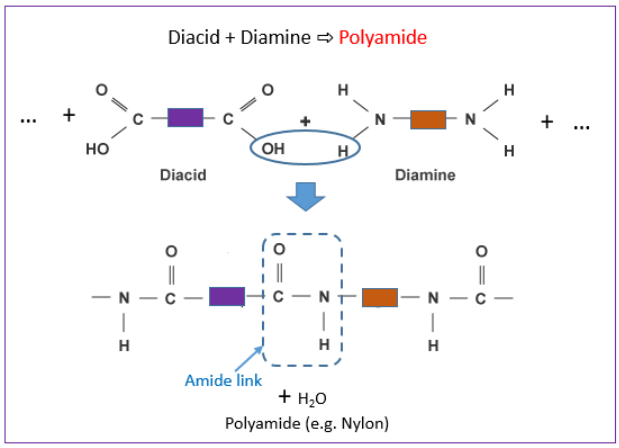


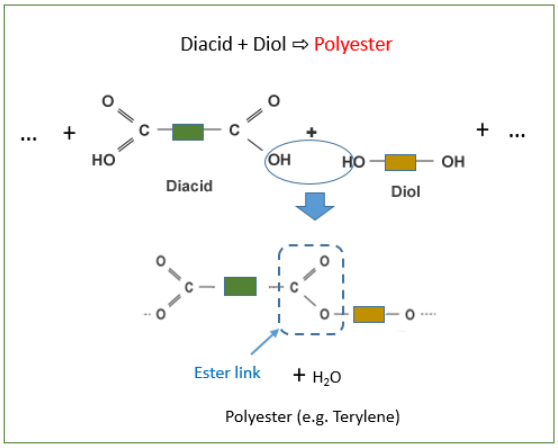
Macromolecules

Addition Polymerisation



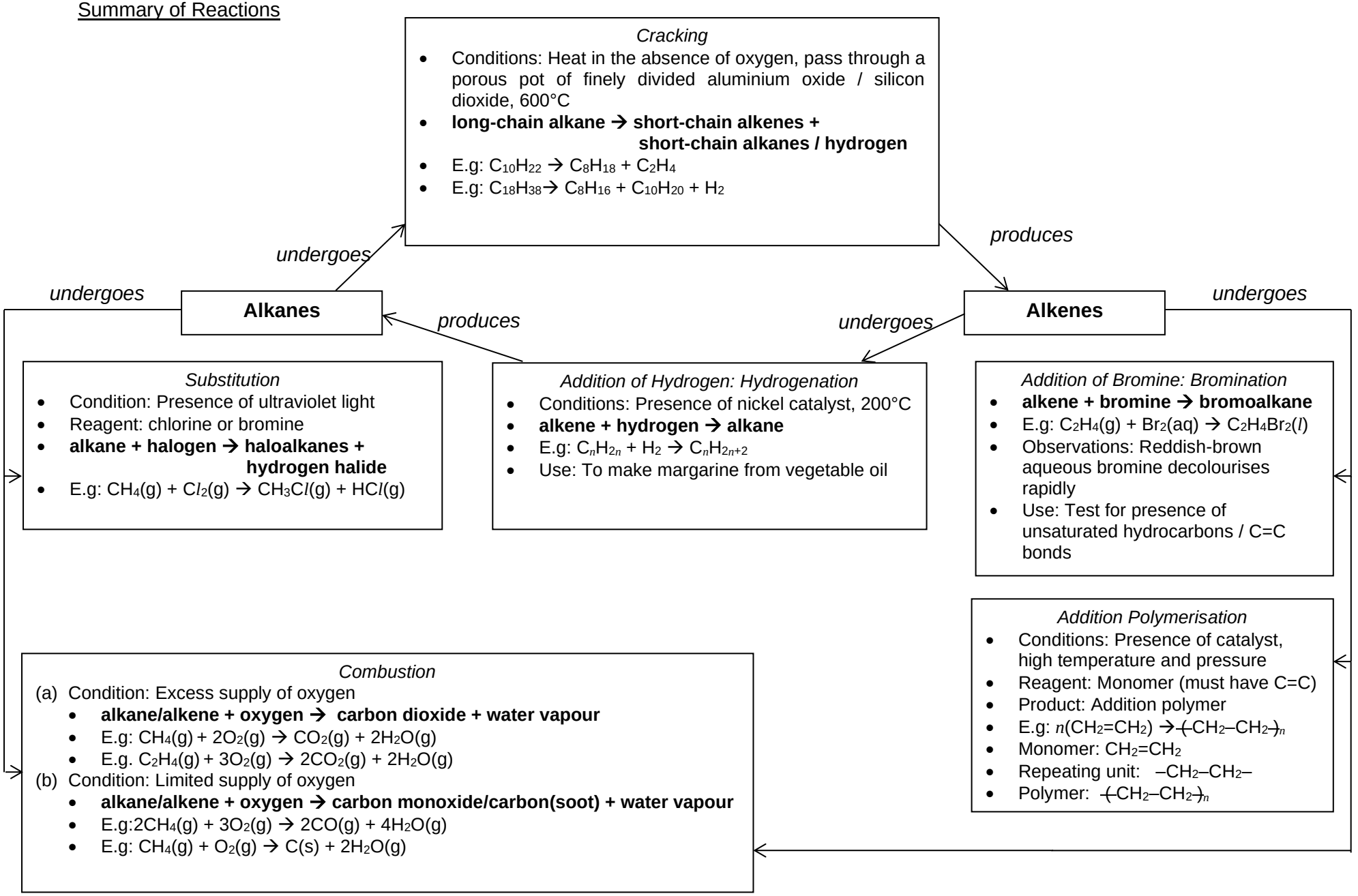
Condensation Polymerisation

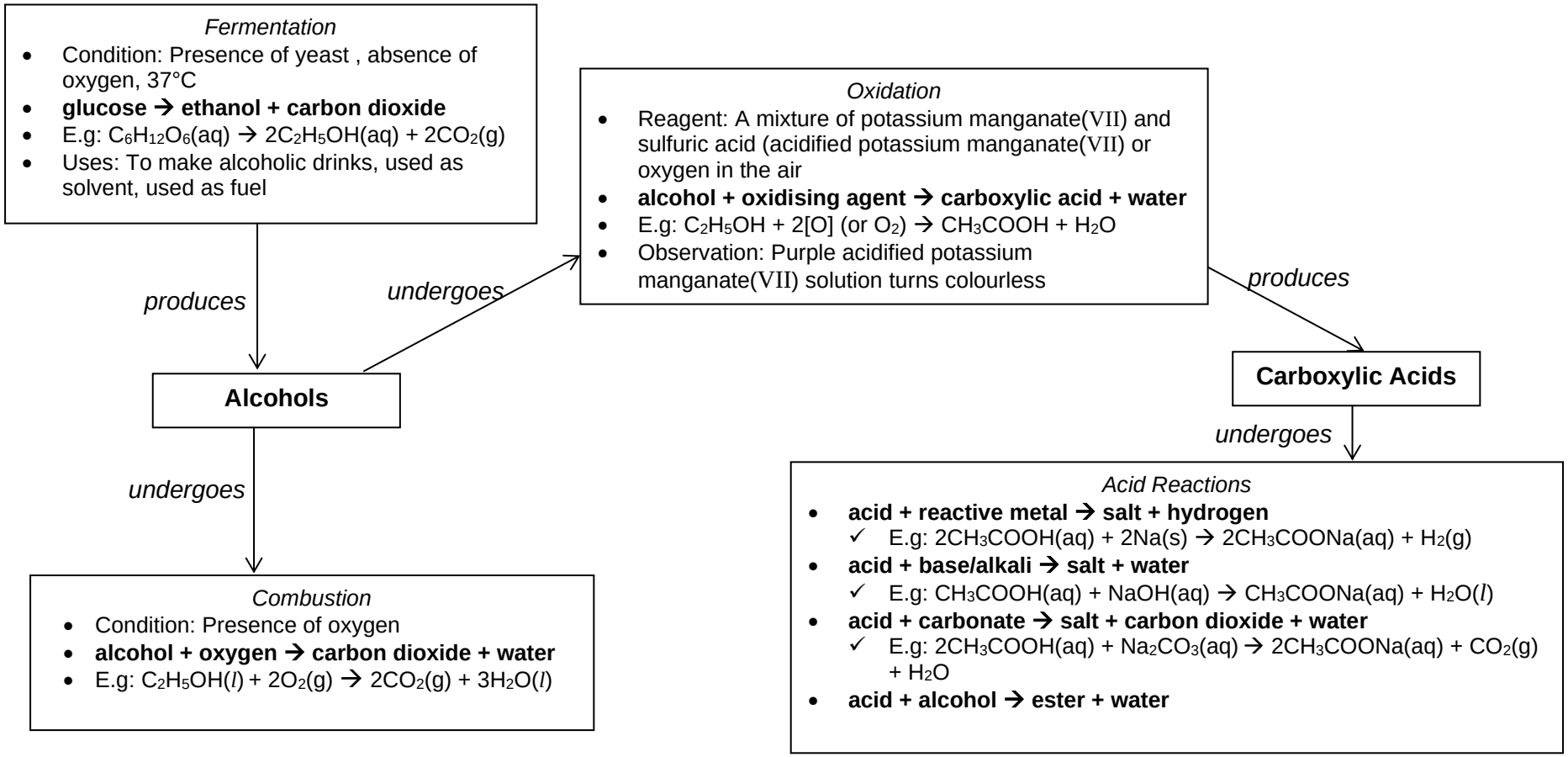




Addition polymer	Condensation polymer
Formed by addition reaction with no elimination of small molecules	Formed by condensation reaction with elimination of small molecules eg H2O
Molecular mass is a whole number multiple of the monomer	Molecular mass is not a whole number multiple of the monomer
Generally, involve one monomer	Generally, involve more than one monomer
Monomers are unsaturated molecules	Monomers must have two active functional groups

Summary of Reactions




Fermentation
• Condition: Presence of yeast , absence of oxygen, 37°C
• **glucose → ethanol + carbon dioxide**
• E.g: $C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$
• Uses: To make alcoholic drinks, used as solvent, used as fuel
produces
undergoes
Oxidation
• Reagent: A mixture of potassium manganate(VII) and sulfuric acid (acidified potassium manganate(VII) or oxygen in the air
• **alcohol + oxidising agent → carboxylic acid + water**
• E.g: $C_2H_5OH + 2[O]$ (or O_2) $\rightarrow CH_3COOH + H_2O$
• Observation: Purple acidified potassium manganate(VII) solution turns colourless
produces
Alcohols
Carboxylic Acids
undergoes
undergoes
Combustion
• Condition: Presence of oxygen
• **alcohol + oxygen → carbon dioxide + water**
• E.g: $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$
Acid Reactions
• **acid + reactive metal → salt + hydrogen**
✓ E.g: $2CH_3COOH(aq) + 2Na(s) \rightarrow 2CH_3COONa(aq) + H_2(g)$
• **acid + base/alkali → salt + water**
✓ E.g: $CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$
• **acid + carbonate → salt + carbon dioxide + water**
✓ E.g: $2CH_3COOH(aq) + Na_2CO_3(aq) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O$
• **acid + alcohol → ester + water**

12 Maintaining Air Quality

Key Concept I

1. Air is a mixture of gases which can be separated by fractional distillation. Its composition is as shown in the table below.

Component	Percentage by volume
nitrogen	78
oxygen	21
• noble gases (mainly argon • carbon dioxide • water vapour	1

Popular application question in MCQ. Will be tasked to calculate air left behind after reaction uses up O_2

2. Name, source and harmful effects of pollutants

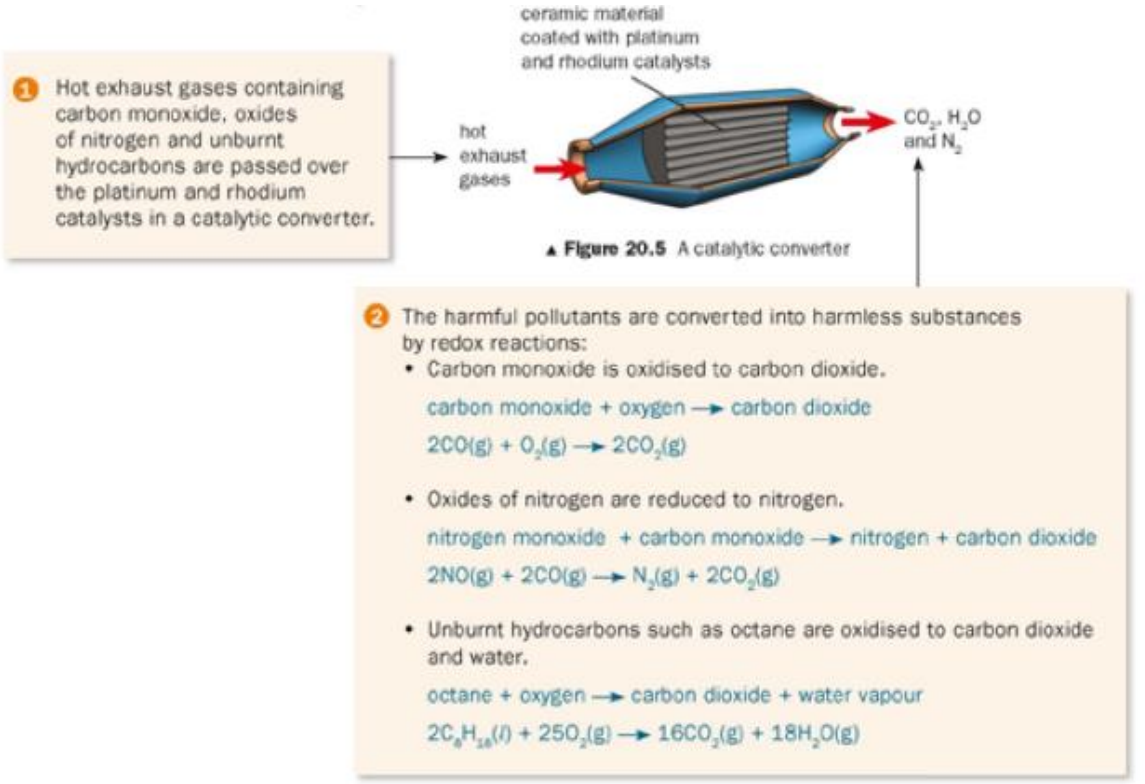
Pollutant	Main sources	Harmful effects
Carbon monoxide, CO	**Incomplete** combustion of carbon-containing fuel like petrol in motor vehicles, coal and natural gas in industries	**Poisonous** - Carbon monoxide is absorbed by haemoglobin in the blood and prevents the blood from absorbing oxygen, causing headaches and damaging the heart, and eventually leading to death.
Nitrogen oxides, (nitrogen monoxide and nitrogen dioxide, NO and NO_2)	Combustion of fuels in internal combustion engines of motor vehicles, power stations and industries; During lightning activity; Nitrogen reacts with oxygen in the air under **high temperature** to form oxides of nitrogen.	Irritate our lungs and eyes ; Forms **'acid rain'** when dissolved in water vapour. Acid rain destroys buildings, plants and aquatic life.
Sulfur dioxide, SO_2	Combustion of fossil fuels containing **sulfur** in power stations and industries; extraction of metals in industries; during volcanic eruptions	It irritates the eyes and makes it difficult for people to breathe. Forms **'acid rain'** when dissolved in water vapour. Acid rain destroys buildings, plants and aquatic life.
Ozone	Produced at ground level due to the reaction of nitrogen oxides and unburnt hydrocarbons in the presence of sunlight	Forms photochemical smog that causes respiratory problems and eye irritation

Unburned hydrocarbons	Released due to incomplete combustion of fossil fuels	Long term exposure can cause cancers
Methane	Naturally produced during the decay of organic matter by microorganisms or decay of rubbish in landfills.	A greenhouse gas that traps heat from sunlight and causes global warming.

3. Catalytic Converters

Installed in exhaust systems of cars to reduce the emissions of carbon monoxide, nitrogen oxides and unburnt hydrocarbons.

These substances undergo **redox reactions** in the presence of platinum and rhodium catalyst to form substances that are less harmful.



4. Flue Gas Desulfurization

Sulfur dioxide gas can be removed by passing waste gases through calcium carbonate or calcium oxide. Sulfur dioxide is removed as it reacts with the basic substances.

sulfur dioxide + calcium carbonate → calcium sulfite + carbon dioxide
$SO_2(g) + CaCO_3(s) \rightarrow CaSO_3(s) + CO_2(g)$

sulfur dioxide + calcium oxide → calcium sulfite
$SO_2(g) + CaO(s) \rightarrow CaSO_3(s)$

Calcium sulfite undergoes further oxidation with atmospheric oxygen to form calcium sulfate.

calcium sulfite + oxygen → calcium sulfate
$2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$

5. Depletion of ozone layer

Ozone layer in Earth's atmosphere has an important role of absorbing excess UV radiation from the Sun.

Substances like chlorofluocarbons (CFCs) deplete the ozone later.
CFCs: compounds containing chlorine, fluorine and carbon.

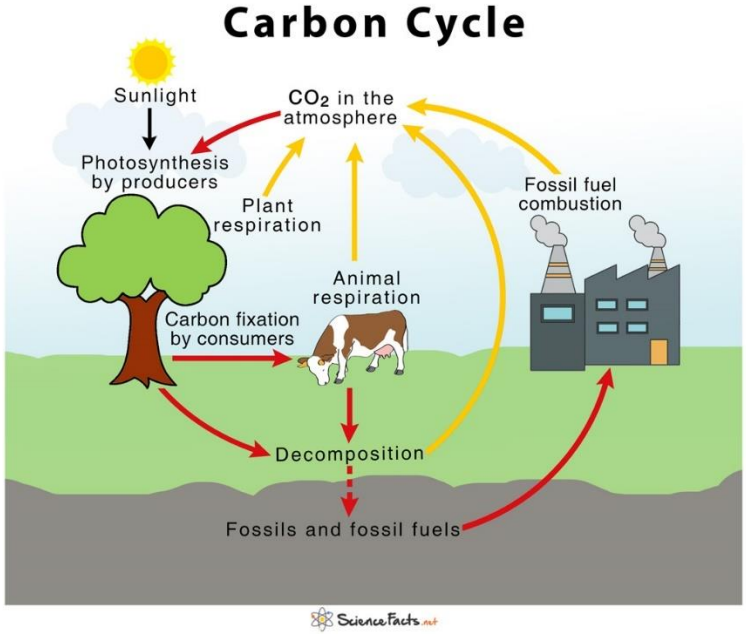
Chlorine radicals are produced when CFCs are exposed to UV radiation. These free radicals destroy the ozone layer by reacting with ozone molecules.

Chlorine atoms cause the breakdown of ozone in a two-step reaction.

$$Cl + O_3 \rightarrow ClO + O_2$$
$$ClO + O_3 \rightarrow 2O_2 + Cl$$

6. Carbon Cycle

The level of carbon dioxide in the atmosphere is maintained by the carbon cycle.



Respiration occurs in living things. It involves breaking down glucose to obtain energy, carbon dioxide and water.

glucose + oxygen → carbon dioxide + water + energy

$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$ + energy

At the same time, carbon dioxide is removed from the atmosphere by photosynthesis. This occurs in plants and involves converting carbon dioxide and water into glucose and oxygen in the presence of sunlight.

carbon dioxide + water → glucose + oxygen

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

A large quantity of carbon dioxide also dissolves in ocean water and is converted into calcium carbonate to form shells and skeletons of marine life.

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The Periodic Table of Elements

+1	+2											+3		-3	-2	-1	
Group																	
I	II											III	IV	V	VI	VII	0
							1 H hydrogen 1										2 He helium 4
3 Li lithium 7	4 Be beryllium 9											5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24											13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 Co cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium –	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57 – 71 lanthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au gold 197	80 Hg mercury 201	81 Tl thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89 – 103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –		114 Fl flerovium –		116 Lv livermorium –		

Key

proton (atomic) number
atomic symbol
name
relative atomic mass

list of important polyatomic ions:

NH_4^+ : ammonium
SO_4^{2-}: sulfate
NO_3^-: nitrate
CO_3^{2-}: carbonate
OH^-: hydroxide

lanthanoids	57 La lanthanum 139	58 Ce cerium 140	59 Pr praseodymium 141	60 Nd neodymium 144	61 Pm promethium –	62 Sm samarium 150	63 Eu europium 152	64 Gd gadolinium 157	65 Tb terbium 159	66 Dy dysprosium 163	67 Ho holmium 165	68 Er erbium 167	69 Tm thulium 169	70 Yb ytterbium 173	71 Lu lutetium 175
actinoids	89 Ac actinium –	90 Th thorium 232	91 Pa protactinium 231	92 U uranium 238	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.)