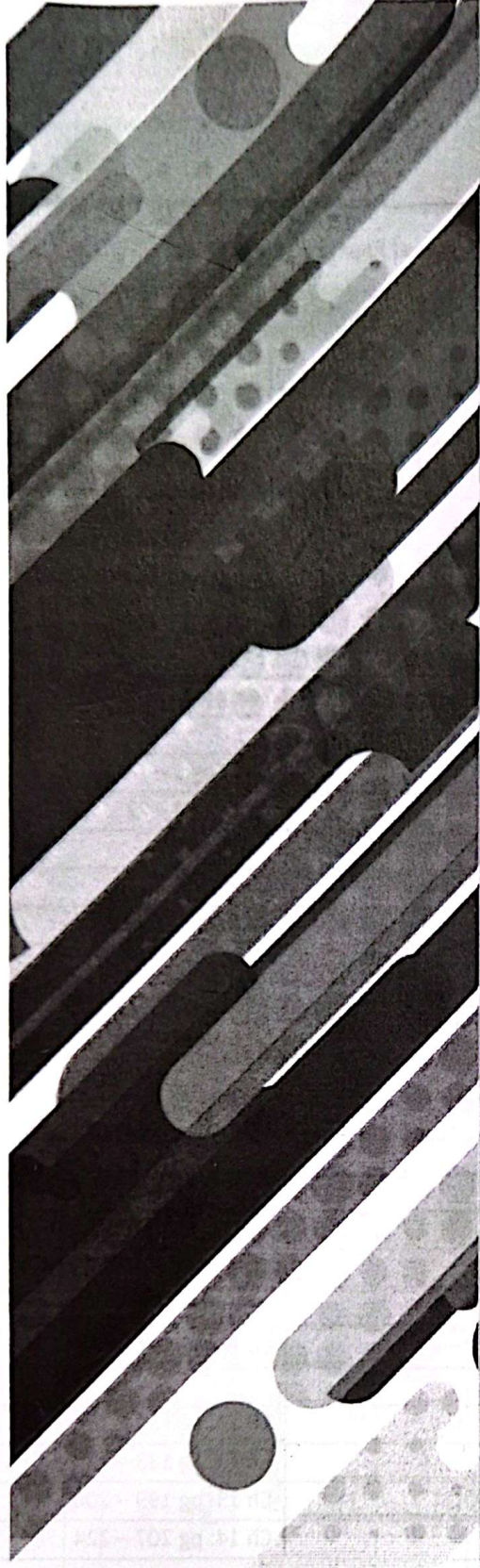




5086 Jasmine 4562

O SCIENCE (CHEMISTRY)

TOPICAL SUMMARY

The Periodic Table of Elements

Group																							
1	2	1 H hydrogen 1												13	14	15	16	17	18				
1+ 2+		Key proton (atomic) number atomic symbol name relative atomic mass																3- 2- 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 -	113 Nh nihonium -	114 Fl flerovium -	115 Mc moscovium -	116 Lv livermorium -	117 Ts tennessine -	118 Og oganeson -					
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.).
 The Avogadro constant, $L = 6.02 \times 10^{23} \text{ mol}^{-1}$.

NOTES FOR QUALITATIVE ANALYSIS

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.
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

EXPERIMENTAL CHEMISTRY

1.1 Experimental Design

- Measurements

To measure		Units	Apparatus		Remarks
mass		grams (g) kilograms (kg)	electronic balance mass balance		recorded to the nearest 0.01 g
volume	liquid	1 m ³ = 1 000 dm ³ = 1 000 000 cm ³	burette		measures up to 50.00 cm ³ , accurate to 0.05 cm ³
			pipette		very accurate set volumes, such as exactly 20.0 cm ³ or 25.0 cm ³
			measuring cylinder		more accurate than beaker, but accurate only to the nearest cm ³
			beaker		approximate volume of 100 cm ³ or 250 cm ³
	gas		gas syringe		
temperature		degrees Celcius (°C)	thermometer temperature sensor		recorded to the nearest 1 d.p.
time		seconds (s) minutes (min)	stopwatch		recorded to the nearest second

- Methods for collecting gases

collection method	characteristics	examples
displacement of water	suitable for collecting gases that are <u>insoluble</u> or <u>slightly soluble</u> in water	H ₂ , O ₂ , CO ₂
downward delivery (upward displacement of air)	used to collect gases that are <u>soluble</u> in water and <u>denser than air</u>	Cl ₂ (M _r = 71) HCl (M _r = 36.5)
upward delivery (downward displacement of air)	used to collect gases that are <u>soluble</u> in water and <u>less dense than air</u>	NH ₃ (M _r = 17)

- Drying a gas

drying agent	characteristics	examples
concentrated sulfuric acid	[acidic] used to dry most gases	Cl ₂ , HCl
quicklime (calcium oxide)	[basic] used to dry alkaline gas	NH ₃
fused calcium chloride	[neutral] used to dry most gases	H ₂

hello

1.2 Methods of Purification and Analysis

- Testing for purity:

- ❖ melting point:

Pure substances melt at a fixed (constant) temperature.

Impure substances melt below the melting point of the pure substance / over a range of temperatures.

- ❖ boiling point:

Pure substances boil at a fixed (constant) temperature.

Impure substances boil above the boiling point of the pure substance / distil over a range of temperatures.

- ❖ chromatography: **Pure** substances produce one spot on the chromatogram (i.e. contain only one substance).

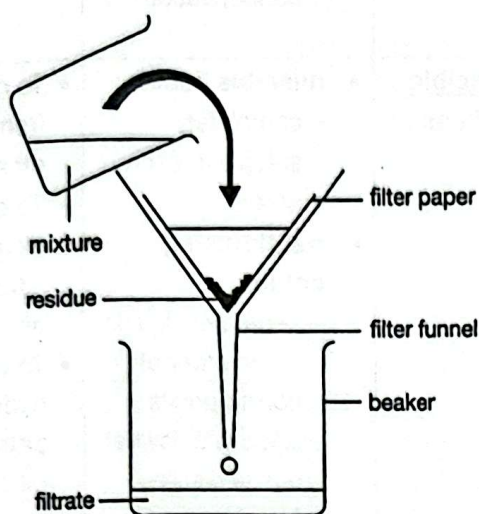
Impure substances produce more than one spot on the chromatogram (i.e. contain more than one substance).

- Methods of purification and analysis

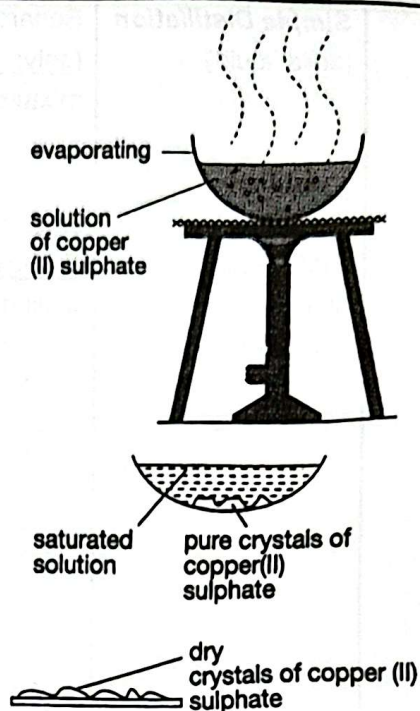
Technique	Application	Remarks	Examples
Use of a suitable solvent (solid-solid)	Separating two solids by using a solvent which dissolves only one of them.		• To obtain <u>sand</u> from a mixture of sand and copper(II) sulfate.
Filtration (solid-liquid)	Separating an insoluble solid from a mixture	• residue – <i>solid</i> collected in the filter paper • filtrate – <i>liquid</i> passing through the filter paper	• To obtain <u>sand</u> from a mixture of sand and water. • To remove solid impurities when making drinking water. • To obtain penicillin from yeast.
Crystallisation or Evaporation (solid-liquid)	Separating a soluble solid from a mixture	• solute – a <i>solid</i> that dissolves in a liquid • solvent – a <i>liquid</i> that dissolves the solid • solution – formed when a solid dissolves in a liquid	• To obtain pure <u>salt</u> from seawater. • To obtain pure sugar from sugar cane plant. • To obtain copper(II) sulfate crystals from its saturated solution.

Technique	Application	Remarks	Examples
Simple Distillation (solid-liquid)	Separating a <u>liquid</u> (solvent) from a mixture	• distillate – the pure liquid obtained from condensation	• To obtain pure <u>water</u> from seawater.
Fractional Distillation (liquid-liquid; miscible)	Separating <u>miscible liquids</u> with different boiling points	• miscible liquids – completely soluble in each other • fractionating column – separates liquids in the order of boiling points – liquid with lowest b.p. is distilled first; highest b.p. distilled last	• To obtain ethanol from a mixture of ethanol and water. • To obtain pure oxygen and pure nitrogen from liquid air. • To obtain useful hydrocarbons from petroleum (crude oil).
Paper Chromatography	• Separating <u>colours</u>, <u>dyes</u>, amino acids, solid particles, etc. • Separating complex mixtures.	• Various types: - Paper chromatography - Thin layer chromatography (TLC) - High performance liquid chromatography (HPLC) - Gas chromatography (GC)	• To separate the <u>dyes</u> present in black ink. • To identify the <u>sugars</u> present in fruit juice. • To identify artificial dyes in food. • To separate and identify complicated substances (e.g. dyes and drugs).
*Use of a Separating Funnel (liquid-liquid; immiscible)	Separation of <u>immiscible liquids</u>	• Immiscible liquids – insoluble in each other.	• To obtain oil from a mixture of oil and water (i.e. oil floats on top of water).
*Sublimation (solid-solid)	Separation of <u>solid</u> mixtures containing two substances of which one <u>sublimes</u>		• To obtain iodine, naphthalene or ammonium chloride.

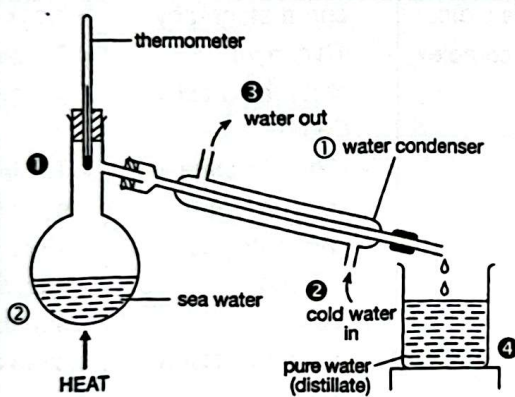
• Diagrams:



FILTRATION



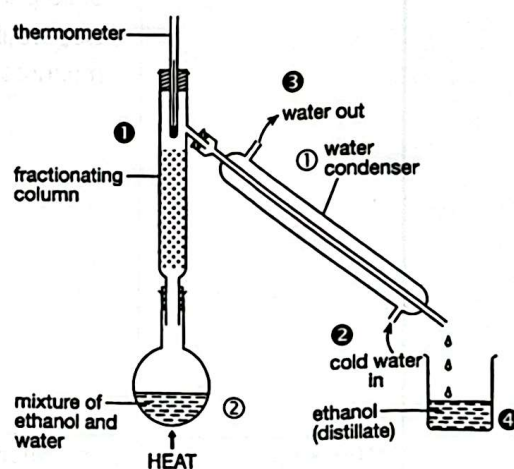
CRYSTALLISATION



SIMPLE DISTILLATION

Note:

- ① the bulb of thermometer should be placed at the mouth of the condenser
- ② cold water enters the condenser from the bottom
- ③ water leaves the condenser from the top
- ④ if the distillate is volatile (i.e. low b.p.), place the beaker in a larger container containing ice



FRACTIONAL DISTILLATION

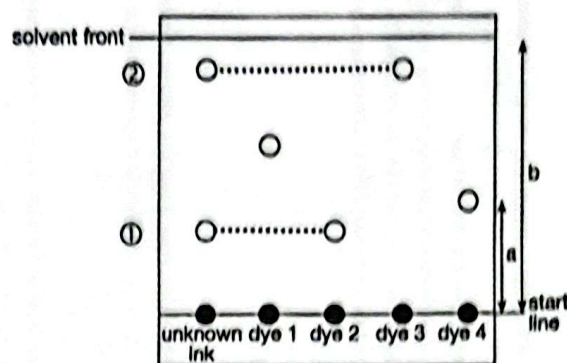
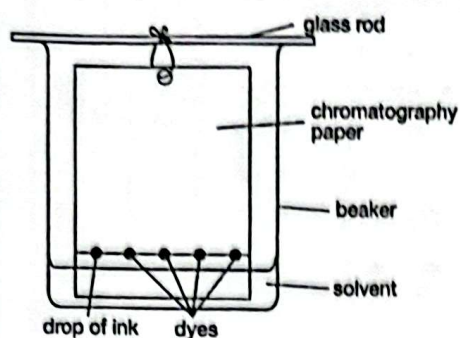
Note:

- ① Purpose of **condenser**.

To cool the hot vapour to form a pure liquid.

- ② Purpose of **boiling stones / chips**:

To ensure smooth boiling.



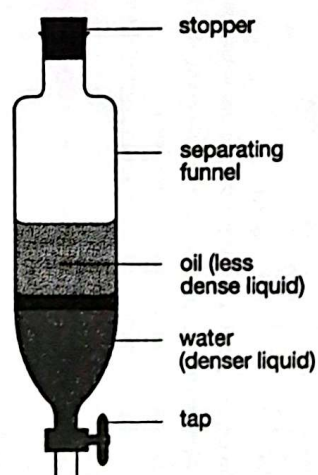
CHROMATOGRAPHY

Interpreting a chromatogram:

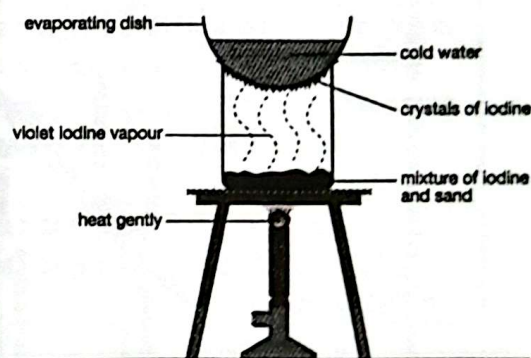
- The unknown ink is a **mixture**, as it produces two spots.
- Dyes 1, 2, 3 and 4 are **pure**, as each of the dyes produces one spot.
- The unknown ink consists of dyes 2 and 3 as the results matches those found in the unknown ink.
- In the unknown ink,
 - ① dye 2 is less soluble as the spot travelled a shorter distance
 - ② dye 3 is more soluble as the spot travelled a longer distance

Note:

- Draw **start line with pencil**, as pencil is not soluble in the solvent.
[Do not draw in ink, as the components in ink may be soluble in the solvent.]
- The **initial solvent level** must be **below the start line**.
[Otherwise, the sample would dissolve directly into the solvent.]
- Use a **locating agent** in the chromatography of colourless compounds.



***SEPARATING FUNNEL**



***SUBLIMATION**

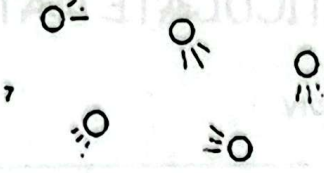
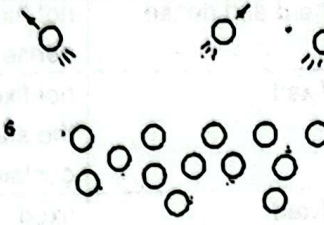
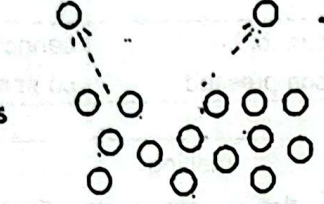
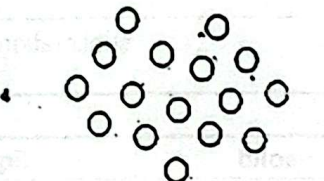
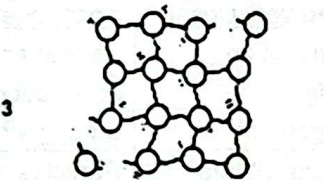
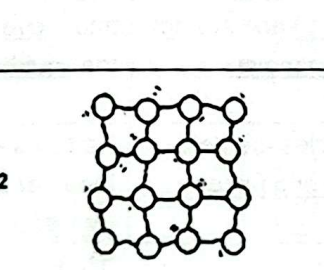
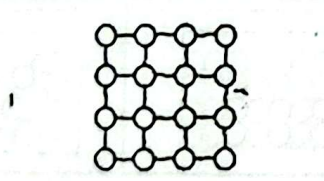
THE PARTICULATE NATURE OF MATTER

2.1 Kinetic Particle Theory

state	solid	liquid	gas
hardness and density	hard and dense	not hard, but dense	not hard, not dense
shape	fixed	not fixed (takes on the shape of the container)	not fixed (takes on the shape of the container)
volume	fixed	fixed	not fixed (takes on volume of the container)
compressibility	cannot be compressed	cannot be compressed	can be compressed
change of state	<pre> graph LR S[solid] -- melting --> L[liquid] L -- freezing --> S L -- boiling --> G[gas] G -- condensation --> L S -- sublimation --> G </pre>		

state	solid	liquid	gas
arrangement	Particles are packed <u>very closely</u> together in an <u>orderly</u> arrangement.	Particles are packed <u>close together</u> (farther apart than solid) but <u>not</u> in an <u>orderly</u> arrangement.	Particles are <u>far apart</u> in a <u>random</u> arrangement.
movement	Particles <u>vibrate</u> about a <u>fixed position</u> and are <u>not free to move</u> .	Particles are able to <u>slide</u> over each other and <u>move</u> throughout the liquid.	Particles are <u>free</u> to <u>move</u> throughout the container or the space that the gas occupies.
forces between particles	Particles are held by <u>strong forces</u> .	Particles are held together by <u>weaker forces</u> .	Force between particles is <u>very weak</u> or <u>negligible</u> .
K.E. of particles	<u>low</u>	<u>higher</u>	<u>highest</u>
diagrams			

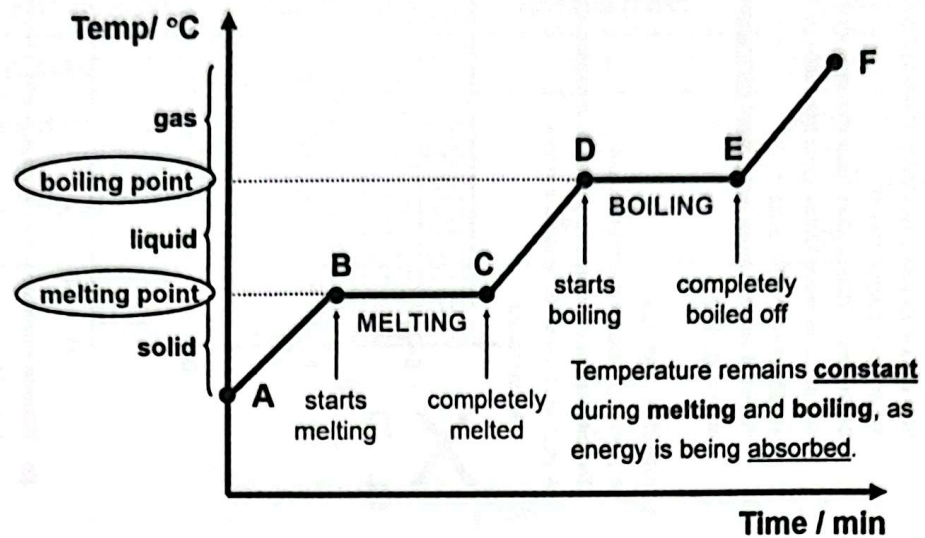
TEMPERATURE INCREASES

GAS		GAS
Liquid Boiling Bonds between particles are broken. Many particles escape.		Gas Condensing Gas particles move closer together. Bonds form between particles.
Liquid A few particles have enough energy to escape as a gas.		Liquid Most gas particles would have condensed.
LIQUID		LIQUID
Solid Melting Bonds holding particles rigidly together are broken.		Liquid Freezing Particles move closer together and bonds form between neighbouring particles.
Solid Particles vibrating more.		Solid Bonds formed between all neighbouring particles.
SOLID		SOLID

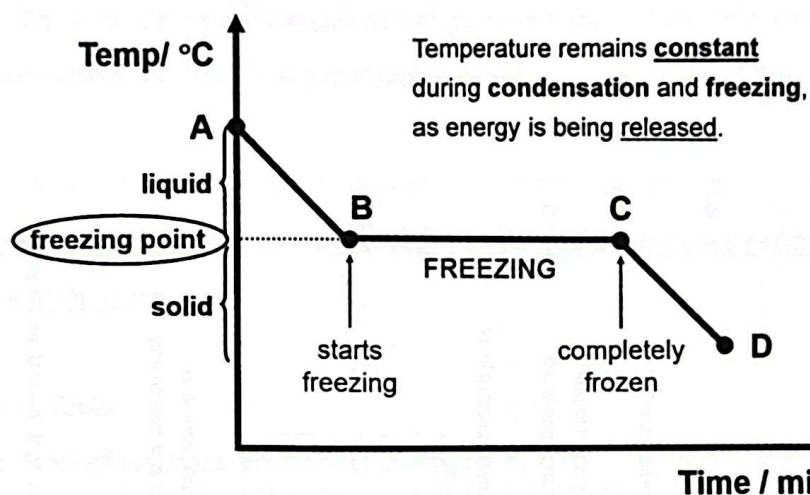
TEMPERATURE DECREASES

- **Diffusion** is the movement of particles from a region of higher concentration to a region of lower concentration.

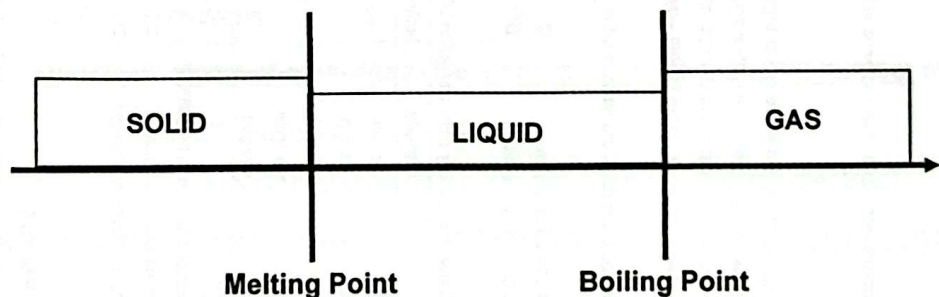
- Heating Curve



- Cooling Curve



- Physical State of Substances

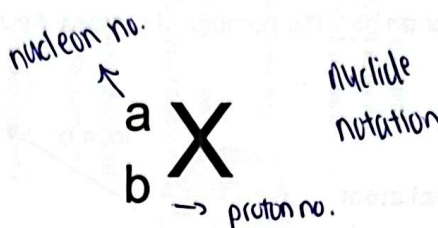


2.2 Atomic Structure

- Relative charge, relative mass and position of the sub-atomic particles:

sub-atomic particles	relative charge	relative mass	position
proton	+1	1	in the nucleus
neutron	0	1	in the nucleus
electron	-1	$\frac{1}{1840} \approx 0$	around nucleus (in electron shells)

- Definitions



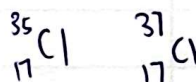
- ❖ **Proton (atomic) number** is the number of protons in the nucleus of an atom.

Note: For a neutral atom, the number of protons = the number of electrons

- ❖ **Nucleon (mass) number** is the number of protons and neutrons in the nucleus of an atom.

Note: No. of neutrons = nucleon (mass) no. – proton (atomic) no.

- ❖ **Isotopes** are atoms of the same element with the same number of protons, but different number of neutrons.



- The Periodic Table

- ❖ Arrangement of electrons for the first 20 elements:

	in shell 1	in shell 2	in shell 3
maximum number of electrons	2	8	8

- ❖ **Group trend (vertical):** atoms of elements in the same group have the same number of outer (valence) electrons

(e.g. atoms of elements in Group 1 have 1 valence electron)

- ❖ **Period trend (horizontal):** atoms of elements in the same period have the same number of electron shells

(e.g. atoms of elements in Period 2 have 2 electron shells)

- Ions

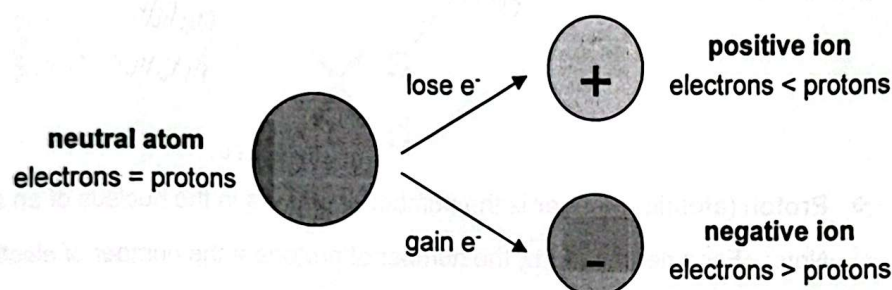
❖ Noble (inert) gases or Group 18 elements are chemically unreactive (do not lose or gain electrons), because their atoms have a stable electronic configuration.

- ${}^4_2\text{He}$ – duplet structure

- ${}^{20}_{10}\text{Ne}$ and ${}^{40}_{18}\text{Ar}$ – octet structure

❖ Atoms of elements in the other Groups (Groups 1–13 and 15–17) lose or gain electrons (i.e. chemically reactive) to achieve the stable electronic configuration.

Note: For an ion, the number of protons $\neq$ the number of electrons



❖ positive ion [cation]

Atoms of metals (Groups 1–13) lose electrons to form positively-charged ions, achieving a stable electronic configuration (i.e. number of protons > number of electrons).

Group	no. of valence e's	no. of e's lost	example of cation
1	1	1	Li^+ , Na^+ , K^+
2	2	2	Mg^{2+} , Ca^{2+}
13	3	3	Al^{3+}

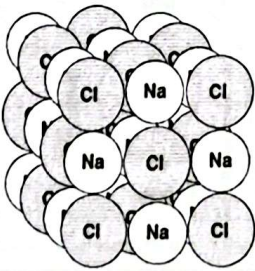
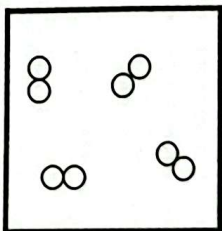
❖ negative ion [anion]

Atoms of non-metals (Groups 15–17) gain electrons to form negatively charged ions, achieving a stable electronic configuration (i.e. number of protons < number of electrons).

Group	no. of valence e's	no. of e's gained	example of anion
17	7	1	F^- , Cl^- , Br^-
16	6	2	O^{2-} , S^{2-}
15	5	3	N^{3-}

CHEMICAL BONDING AND STRUCTURE

3.1-3.2 Ionic and Covalent Bonding

type of bonding	IONIC	SIMPLE COVALENT
bonding between:	metal and non-metal [transfer of electrons]	non-metals <u>only</u> [sharing of electrons]
particles present:	ions	molecules
state at r.t.p:	solid	gas or volatile liquid
physical property #1: melting point & boiling point	high A <u>large</u> amount of heat energy is required to overcome the <u>strong electrostatic</u> forces of attraction between oppositely charged <u>ions</u> .	low A <u>small</u> amount of heat energy is required to overcome the <u>weak intermolecular</u> forces of attraction between <u>molecules</u> .
physical property #2: electrical conductivity	Conducts electricity in <u>molten</u> and <u>aqueous</u> states but not in solid state. In the <i>molten</i> and <i>aqueous</i> states, the presence of <u>mobile ions</u> allows the electric current to be carried. In the <i>solid</i> state, the <u>ions</u> are held rigidly in fixed positions. The <u>absence of mobile ions</u> prevents the electric current from being carried.	Does <u>not</u> conduct electricity in any state. Exists as <u>neutral molecules</u> . The <u>absence of mobile charged particles</u> prevents the electric current from being carried.
physical property #3: solubility	<u>soluble</u> in water	<u>insoluble</u> in water
examples:	NaCl, MgCl ₂	H ₂ , O ₂ , H ₂ O, CH ₄ , CO ₂
diagram:		

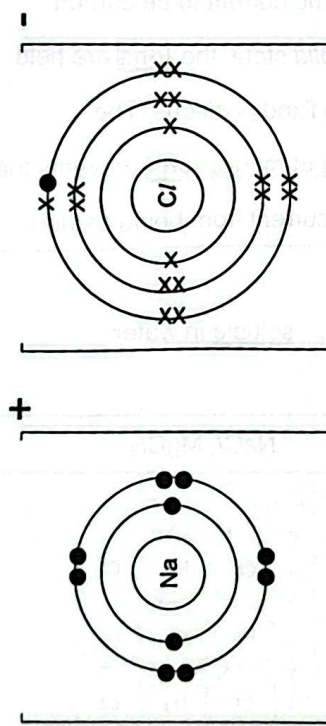
Dot-and-Cross Diagrams

- An ionic bond is formed as a result of the strong electrostatic force of attraction between oppositely charged ions. A metallic atom loses electrons to a non-metallic atom (which gains), and this transfer of electrons allow both atoms to achieve a stable electronic configuration.

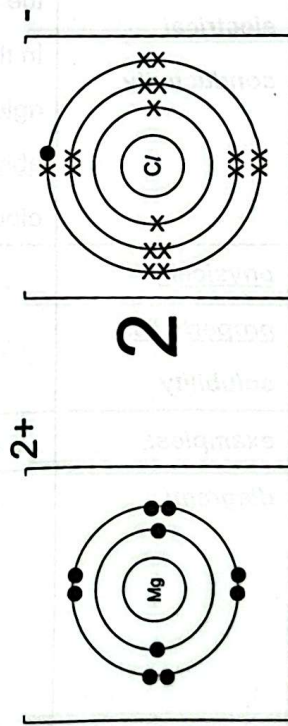
Steps:

- work out the formula of the compound (i.e. how many positive and negative ions)
- refer to the Periodic Table for the electronic configuration of the ions
- assign the correct charges to the respective ions

E.g. 1: sodium chloride (NaCl)



E.g. 2: magnesium chloride (MgCl_2)



The Mg atom loses 2 valence electrons to two Cl atoms, forming a Mg^{2+} ion, achieving a stable electronic configuration. On the other hand, the Cl atom gains 1 electron into its valence shell, forming a Cl^- ion, achieving a stable electronic configuration.

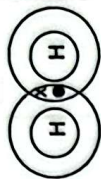
Summary

- A covalent bond is formed when atoms of non-metals share electrons to achieve the stable electronic configuration.

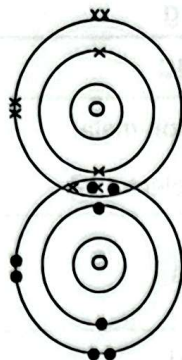
Steps:

- work out the molecular / structural formula of the molecule (i.e. how many atoms)
- draw the outline of the molecule
- refer to the Periodic Table for the electronic configuration of the atoms

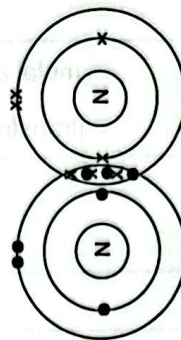
E.g. 1: hydrogen (H_2) [$\text{H}-\text{H}$]



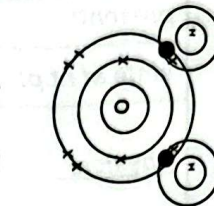
E.g. 2: oxygen (O_2) [$\text{O}=\text{O}$]



E.g. 3: nitrogen (N_2) [$\text{N}\equiv\text{N}$]



E.g. 4: water (H_2O) [$\text{H}-\text{O}-\text{H}$]



The O atom shares 2 valence electrons with two H atoms, each achieving a stable electronic configuration.

Key: ● = electrons from H atom
x = electrons from O atom

3.3 Structure and Properties of Materials

(A) Atoms and Molecules

- **Atoms** are the smallest particles of an element that have the chemical properties of that element.
- A **molecule** is a group of two or more atoms that are chemically combined (joined together).

(B) Elements

- An **element** is a pure substance that cannot be broken down into simpler substances by physical or chemical processes.

metals			non-metals				
1	2	13	14	15	16	17	18
sodium	magnesium	aluminium	carbon	nitrogen	oxygen	fluorine	neon
Na	Mg	Al	C	N ₂	O ₂	F ₂	Ne

- Metals and some non-metals are *monoatomic*.
e.g.: metals like sodium (Na), calcium (Ca), iron (Fe)
non-metals like helium (He), neon (Ne), argon (Ar), xenon (Xe)
- However, most non-metals exist as molecules.
 - ❖ *Diatomic* molecules are molecules formed by the combination of two atoms.
e.g.: hydrogen (H₂), nitrogen (N₂), oxygen (O₂), fluorine (F₂), chlorine (Cl₂),
bromine (Br₂) and iodine (I₂)

(C) Compounds and Mixtures

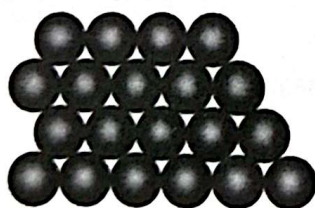
- A **compound** is a pure substance containing two or more different elements that are chemically combined in a fixed ratio.
e.g. ionic – sodium chloride, NaCl; magnesium chloride, MgCl₂
e.g. covalent – hydrogen chloride, HCl; water, H₂O; ammonia, NH₃; methane, CH₄,
carbon dioxide, CO₂
- A **mixture** is made up of two or more substances that are not chemically combined.
e.g. seawater; air; crude oil; ink; alloy

In summary,

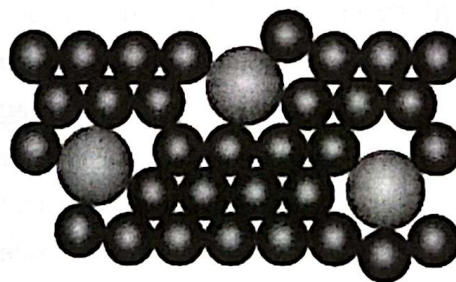
elements	compounds	mixtures
	formed by <u>chemical</u> change	<u>not</u> formed by <u>chemical</u> change
<u>fixed</u> composition	<u>fixed</u> composition	<u>variable</u> composition
	<u>different properties</u> from their constituent elements	<u>same properties</u> as their constituent elements
<u>cannot</u> be broken down further by chemical or physical methods	is separated by <u>chemical</u> methods	is separated by <u>physical</u> methods
	has <u>fixed</u> melting and boiling points	<u>variable</u> melting and boiling points

(D) Metals

- General physical properties of metals:
 - Metals have a high density. [except: Group 1, Al, Ti]
 - Metals have a high melting and boiling point. [except: Group 1, Hg]
 - Metals are malleable (i.e. can be bent and beaten into different shapes) and ductile (i.e. can be stretched to form wires).
 - Metals are good conductors of heat and electricity.
- An alloy is a mixture of a metal with another element.
 - The presence of atoms of a different size disrupts the orderly arrangement of atoms, causing the layers of atoms to be unable to slide over each other easily.
 - Hence, alloys are harder and stronger than pure metals.



metal
[e.g. iron]



alloy
[e.g. steel]

- Examples of alloys:
 - brass – copper and zinc
 - bronze – copper and tin
 - mild steel – carbon [nm] and iron
 - stainless steel – carbon [nm], iron, chromium and nickel

CHEMICAL CALCULATIONS

4.1 Formulae and Equation Writing

- In the Periodic Table, each element is represented by its own chemical **symbol**.

element	lithium	beryllium	boron	carbon	nitrogen	oxygen	fluorine	neon
symbol	Li	Be	B	C	N	O	F	Ne

- The chemical formula tells us the types of elements present and the number of atoms of each type of element.

formula	NaCl	H ₂ O	NH ₃	C ₂ H ₅ OH
information	1 sodium atom 1 chlorine atom	2 hydrogen atoms 1 oxygen atom	1 nitrogen atom 3 hydrogen atoms	2 carbon atoms 6 hydrogen atoms 1 oxygen atom

- Valency of metallic, non-metallic, polyatomic ions

	VALENCY							
	1		2		3		4	
metals	sodium potassium <i>silver</i>	Na ⁺ K ⁺ Ag ⁺	magnesium calcium barium <i>zinc</i> copper(II) iron(II) lead(II)	Mg ²⁺ Ca ²⁺ Ba ²⁺ Zn ²⁺ Cu ²⁺ Fe ²⁺ Pb ²⁺	aluminium iron(III) chromium(III)	Al ³⁺ Fe ³⁺ Cr ³⁺	lead(IV) manganese(IV)	Pb ⁴⁺ Mn ⁴⁺
non-metals	fluorine chlorine bromine iodine <i>hydrogen</i>	F ⁻ Cl ⁻ Br ⁻ I ⁻ H ⁺	oxygen sulfur	O ²⁻ S ²⁻	nitrogen phosphorous	N ³⁻ P ³⁻	carbon silicon	C ⁴⁻ Si ⁴⁻
radical groups	ammonium hydroxide nitrate	NH ₄ ⁺ OH ⁻ NO ₃ ⁻	carbonate sulfate IMPORTANT!!	CO ₃ ²⁻ SO ₄ ²⁻				

How to deduce the formulae of ionic compounds?

• Charge of ions

Group	1	2	13	14	15	16	17	18
	metals			non-metals				
charge	+	2+	3+	4-	3-	2-	-	
example	Na ⁺	Mg ²⁺	Al ³⁺	C ⁴⁻	N ³⁻	O ²⁻	Cl ⁻	

• Charge of polyatomic ions

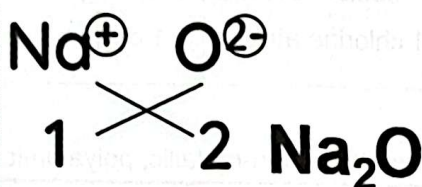
ammonium ion	hydroxide ion	nitrate ion	carbonate ion	sulfate ion
NH ₄ ⁺	OH ⁻	NO ₃ ⁻	CO ₃ ²⁻	SO ₄ ²⁻

• For example: sodium oxide

Steps:

- Write down the charge of the ions.
- Write down the valency of each element (i.e. no. of bonds it can form)
- Do a criss-cross (i.e. to balance the charges)
- Congratulations!

The formula is ready! ☺



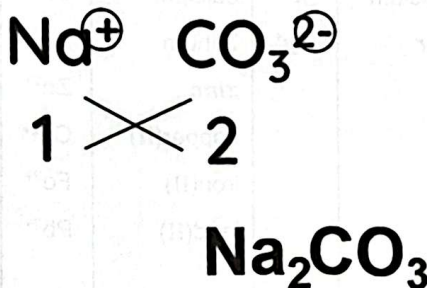
Na⁺ and O²⁻ ions are simple ions that are formed from single atoms. Such ions are also known as monoatomic ions.

• For example: sodium carbonate

Steps:

- Write down the charge of the ions.
- Write down the valency of each element / group of elements (i.e. no. of bonds it can form)
- Do a criss-cross (i.e. to balance the charges)
- Congratulations!

The formula is ready! ☺



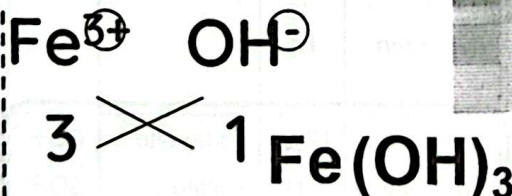
The CO₃²⁻ ion is an ion that contains more than one atom. Such ion is also known as polyatomic ion.

• For example: iron(III) hydroxide

Steps:

- Write down the charge of the ions.
- Write down the valency of each element / group of elements (i.e. no. of bonds it can form)
- Do a criss-cross (i.e. to balance the charges)
- Congratulations!

The formula is ready! ☺



The OH⁻ ion is an ion that contains more than one atom. Such ion is also known as polyatomic ion.

How to construct chemical equations?

Steps:

- 1 Write down the formulae of the reactants and products for the reaction.
- 2 Balance the equation by putting numbers IN FRONT of the formulae of the substances in the equation.
- 3 Ensure that the number of particles of each element remains the same before and after a reaction.

For e.g.:

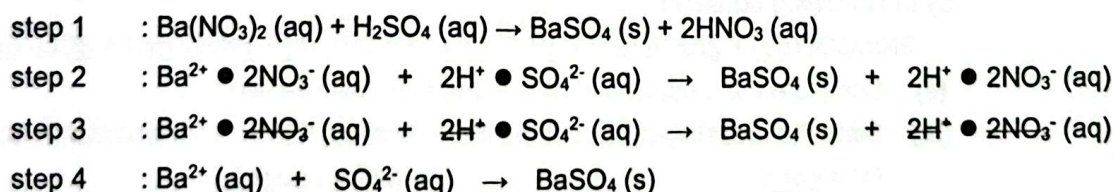
topic	name of reaction	equation
5	acid + metal	$\text{Mg (s)} + 2\text{HCl (aq)} \rightarrow \text{MgCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$
	acid + base	$\text{CuO (s)} + \text{H}_2\text{SO}_4 \text{ (aq)} \rightarrow \text{CuSO}_4 \text{ (aq)} + \text{H}_2\text{O (l)}$
	acid + carbonate	$\text{CaCO}_3 \text{ (s)} + 2\text{HNO}_3 \text{ (aq)} \rightarrow \text{Ca(NO}_3)_2 \text{ (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$
	alkali + acid	$\text{NaOH (aq)} + \text{HCl (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
	alkali + ammonium salt	$2\text{NaOH (aq)} + (\text{NH}_4)_2\text{SO}_4 \text{ (aq)} \rightarrow \text{Na}_2\text{SO}_4 \text{ (aq)} + 2\text{H}_2\text{O (l)} + 2\text{NH}_3 \text{ (g)}$
8.2	Group 1	$2\text{Na (s)} + 2\text{H}_2\text{O (l)} \rightarrow 2\text{NaOH (aq)} + \text{H}_2 \text{ (g)}$
	Group 17 displacement	$\text{Cl}_2 \text{ (g)} + 2\text{KI (aq)} \rightarrow 2\text{KCl (aq)} + \text{I}_2 \text{ (s)}$
11.2	combustion	$\text{C}_3\text{H}_8 \text{ (g)} + 5\text{O}_2 \text{ (g)} \rightarrow 3\text{CO}_2 \text{ (g)} + 4\text{H}_2\text{O (l)}$
	alkanes: substitution	$\text{CH}_4 \text{ (g)} + \text{Cl}_2 \text{ (g)} \rightarrow \text{CH}_3\text{Cl (g)} + \text{HCl (g)}$
	alkenes: addition of H_2	$\text{C}_2\text{H}_4 \text{ (g)} + \text{H}_2 \text{ (g)} \rightarrow \text{C}_2\text{H}_6 \text{ (g)}$
	alkenes: addition of Br_2	$\text{C}_2\text{H}_4 \text{ (g)} + \text{Br}_2 \text{ (aq)} \rightarrow \text{C}_2\text{H}_4\text{Br}_2 \text{ (l)}$

How to construct ionic equations?

Steps:

- 1 Write a balanced chemical equation, including state symbols.
- 2 Re-write the equation in terms of ions for substances in the aqueous state.
[Leave substances in the solid / liquid / gaseous states alone.]
- 3 Cancel out the spectator ions.
- 4 Re-write the equation without the spectator ions.

For e.g.:



4.2 The Mole Concept and stoichiometry

Some useful definitions to remember:

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

Note: The A_r of an element can be found in the Periodic Table.
e.g.: A_r of Mg = 24

- The **relative molecular mass, M_r** , of a substance (i.e. element, compound) is the average mass of one molecule of the substance compared to $\frac{1}{12}$ the mass of a carbon-12 atom.

Note: The relative molecular mass (and relative formula mass), M_r , can be calculated as the sum of relative atomic masses, A_r .
e.g.: M_r of fayalite, $\text{Fe}_2\text{SiO}_4 = 2(56) + 28 + 4(16) = 204$

- One **mole** of any substance will always contain 6.02×10^{23} particles (i.e. atoms, molecules, ions and even subatomic particles such as electrons) of that substance.

Some useful formulae to remember:

- Calculate the percentage mass of an element in a compound

$$\text{Percentage mass of an element X in a compound} = \frac{(\text{No. of atoms of element X}) \times A_r}{M_r \text{ of compound}} \times 100\%$$

- Calculate the mass of an element in a compound

$$= \frac{\text{No. of atoms of the element in the formula} \times A_r \text{ of the element}}{M_r \text{ of the compound}} \times \text{Mass of the sample}$$

Example 1: The relative atomic mass of nitrogen is 14 and the relative molecular mass of ammonium nitrate is 80.

What is the percentage of nitrogen in ammonium nitrate?

$$\%N \text{ in } \text{NH}_4\text{NO}_3 = \frac{2(14)}{80} \times 100 = \underline{35\%}$$

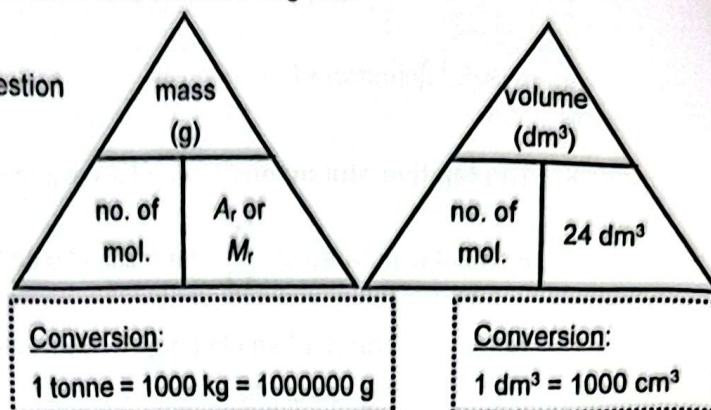
Example 2: What is the mass of iron that can be extracted from 116 g of Fe_3O_4 ?

$$\text{mass of iron} = \frac{3(56)}{232} \times 116 \text{ g} = \underline{84 \text{ g}}$$

- Calculate stoichiometric reacting masses and volumes of gases

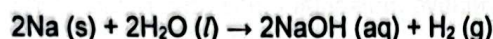
Steps:

- find the no. of mol. given in question
- compare the mole ratio
- answer question



Example 3: What is the volume of hydrogen produced at room temperature and pressure, when 4.6 g of sodium is reacted with an excess of water?

[1 mol of gas occupies 24 dm^3 at room temperature and pressure.]



Fr. Eqn.: 2

1

Given : 4.6 g

? dm^3

A_r / M_r : 23

① 0.2 mol

② 0.1 mol

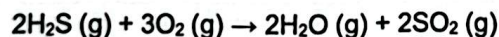
③ Volume of hydrogen produced = $0.1 \text{ mol} \times 24 \text{ dm}^3 = \underline{2.4 \text{ dm}^3}$

Example 4: Hydrogen sulfide burns in an excess of oxygen according to the following equation: $2\text{H}_2\text{S (g)} + 3\text{O}_2 \text{ (g)} \rightarrow 2\text{H}_2\text{O (g)} + 2\text{SO}_2 \text{ (g)}$

48 dm^3 of hydrogen sulfide is burned.

Which volume of sulfur dioxide will be formed at room temperature and pressure?

[All volumes are measured at the same temperature and pressure.]



Fr. Eqn.: 2

2

Given : 48 dm^3

? dm^3

Volume of sulfur dioxide formed = 48 dm^3

Example 5: 20 cm³ of ethene are reacted with 70 cm³ of oxygen.

The equation for the reaction is shown: $\text{C}_2\text{H}_4 (\text{g}) + 3\text{O}_2 (\text{g}) \rightarrow 2\text{CO}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{l})$

What is the total volume of gas remaining at the end of the reaction?

(all volumes are measured at r.t.p.)

	$\text{C}_2\text{H}_4 (\text{g})$	+	$3\text{O}_2 (\text{g})$	$\rightarrow$	$2\text{CO}_2 (\text{g})$	+	$2\text{H}_2\text{O} (\text{l})$
Fr. Eqn.:	1		3		2		
Given	: 20 cm ³		60 cm ³		? cm ³		
	0 cm ³ left		10 cm ³ left		40 cm ³ formed		
	(limiting)		(excess)				

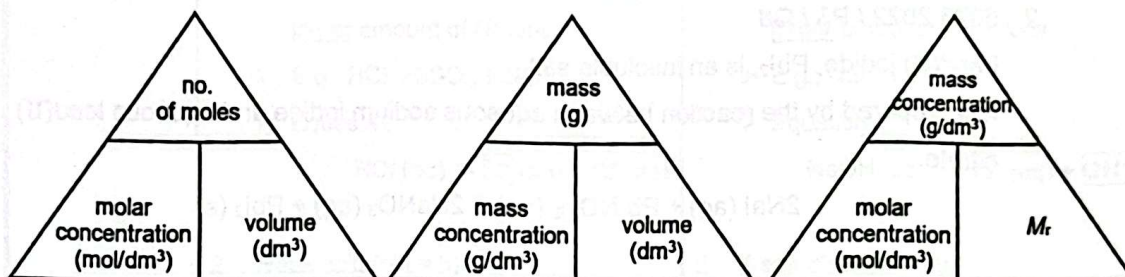
Total volume = 10 cm³ of O₂ left + 40 cm³ of CO₂ formed

- Apply the concept of solution concentration (in mol/dm³ or g/dm³) to process the results of volumetric experiments (e.g. titration) and to solve simple problems.

molar concentration

mass concentration

IMPORTANT!!



Example 6: 25 cm³ of aqueous 0.1 mol/dm³ hydrochloric acid exactly neutralises 20 cm³ of aqueous sodium hydroxide.

What is the concentration of the sodium hydroxide solution?

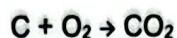
	$\text{NaOH} (\text{aq})$	+	$\text{HCl} (\text{aq})$	$\rightarrow$	$\text{NaCl} (\text{aq})$	+	$\text{H}_2\text{O} (\text{l})$
Fr. Eqn.:	1		1				
Given	: ? mol/dm ³		0.1 mol/dm ³				
	20 cm ³		25 cm ³				
	② 0.0025 mol		① 0.0025 mol				

③ Concentration of sodium hydroxide = 0.0025 mol ÷ 0.020 dm³ = 0.125 mol/dm³

Example

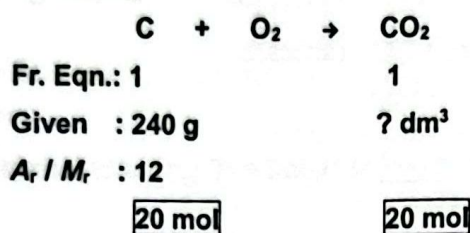
1 5076 2022 / P3 / Q9

(c) (iv) Carbon burns in air to form carbon dioxide.



A sample of 240 g of carbon is completely burned to form carbon dioxide. Calculate the volume of carbon dioxide formed at room temperature and pressure.

[This volume of one mole of any gas at room temperature and pressure is 24 dm³.]



[1]

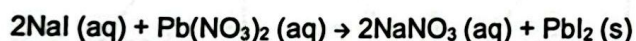
$$\text{volume of carbon dioxide} = 20 \text{ mol} \times 24 \text{ dm}^3 = \underline{480 \text{ dm}^3}$$

[1]

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Lead(II) iodide, PbI₂, is an insoluble salt.

It is prepared by the reaction between aqueous sodium iodide and aqueous lead(II) nitrate.



(a) Name the experimental technique used to separate the lead(II) iodide from the reaction mixture.

filtration

[1]

(b) A sample of 50.0 cm³ of 0.100 mol/dm³ sodium iodide is mixed with an excess of aqueous lead(II) nitrate.

(i) Calculate the number of moles of sodium iodide used.

$$\text{no. of mol.} = 0.100 \text{ mol/dm}^3 \times 0.05 \text{ dm}^3 = \underline{0.005 \text{ mol}}$$

[1]

(ii) Deduce the number of moles of lead(II) iodide made in the reaction.

$$\text{no. of mol.} = 0.005 \text{ mol} \div 2 = \underline{0.0025 \text{ mol}}$$

[1]

(iii) Calculate the mass, in grams, of lead(II) iodide made in the reaction.

$$\text{mass} = \underline{0.0025 \text{ mol}} \times 461$$

[1]

$$= 1.1525 \text{ g}$$

$$= \underline{1.15 \text{ g (to 3 s.f.)}}$$

[1]

ACID-BASE CHEMISTRY

Properties of Acids and Alkalis

	acid	alkali
ions	H^+ (aq)	OH^- (aq)
physical properties	1. Dilute acids have a <u>sour</u> taste. 2. Acids are soluble in water and give solutions with $pH < 7$. 3. Acids turn damp blue litmus paper <u>red</u>. 4. Acids turn Universal Indicator <u>red</u>. 5. Concentrated forms of strong acids like HCl and H_2SO_4 are <u>corrosive</u>. 6. Aqueous forms of acid are <u>good conductors of electricity</u>.	1. Alkalis have a <u>bitter</u> taste. 2. Alkalis are <u>soapy</u> to the touch. 3. Alkalis are soluble in water and give solutions with $pH > 7$. 4. Alkalis turn damp red litmus paper <u>blue</u>. 5. Alkalis turn Universal Indicator <u>blue</u>. 6. Concentrated forms of strong alkalis like KOH and $NaOH$ are <u>corrosive</u>.
strength	1. Strong acid ($pH \approx 1$) > An acid which undergoes <u>complete ionisation</u> in water to produce a <u>larger</u> amount of H^+ ions. > E.g.: HCl, H_2SO_4, HNO_3 > Equation: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$ 2. Weak acid ($pH \approx 5$) > An acid which undergoes <u>partial ionisation</u> in water to produce a <u>smaller</u> amount of H^+ ions. > E.g.: ethanoic acid, CH_3COOH > Equation: $CH_3COOH(aq) \leftrightarrow CH_3COO^-(aq) + H^+(aq)$	1. Strong alkali ($pH \approx 14$) > An alkali which undergoes <u>complete ionisation</u> in water to produce a <u>larger</u> amount of OH^- ions. > E.g.: $NaOH$, KOH > Equation: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$ 2. Weak alkali ($pH \approx 9$) > An alkali which undergoes <u>partial ionisation</u> in water to produce a <u>smaller</u> amount of OH^- ions. > E.g.: aqueous ammonia, NH_4OH > Equation: $NH_3(aq) + H_2O \leftrightarrow NH_4^+(aq) + OH^-(aq)$
chemical properties	1. acid + metal $\rightarrow$ salt + <u>hydrogen</u> 2. acid + base $\rightarrow$ salt + water 3. acid + carbonate $\rightarrow$ salt + water + <u>carbon dioxide</u>	1. alkali + acid $\rightarrow$ salt + water neutralisation $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ 2. alkali + ammonium salt $\rightarrow$ salt + water + <u>ammonia</u>
uses		Calcium hydroxide, $Ca(OH)_2$, is used to treat excess acidity in soil by increasing pH.

The Nature of Oxides

Group	1	2	13	14	15	16	17
element	Na	Mg	Al	Si	P	S	Cl
formula of oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀ P ₄ O ₆	SO ₃ SO ₂	Cl ₂ O ₇ Cl ₂ O
nature of oxide	basic		amphoteric	acidic			

Basic oxide

- **Metallic** oxides that react with acids to form a salt and water.
- Most basic oxides are insoluble in water. Those that are soluble are known as alkalis.
- e.g.: sodium oxide [Na₂O], magnesium oxide [MgO], copper(II) oxide [CuO]

Amphoteric oxide

- **Metallic** oxides that react with both acids and bases to form a salt and water.
- e.g.: LAZY – lead(II) oxide [PbO], aluminium oxide [Al₂O₃], zinc oxide [ZnO]

Acidic oxide

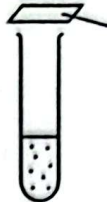
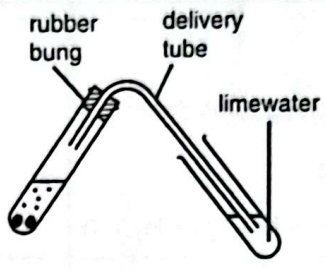
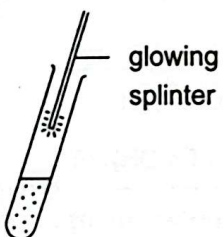
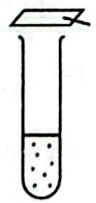
- **Non-metallic** oxides that react with bases to form a salt and water.
- They dissolve in water to form acidic solutions with pH < 7.
- e.g.: carbon dioxide [CO₂], nitrogen dioxide [NO₂], sulfur dioxide [SO₂], sulfur trioxide [SO₃]

Neutral oxide

- **Non-metallic** oxides that DO NOT react with both acids and alkalis.
- They dissolve in water to form neutral solutions with pH = 7.
- e.g.: water [H₂O], carbon monoxide [CO], nitrogen monoxide [NO]

QUALITATIVE ANALYSIS

Test for Gases

gas	test	positive test result
1. ammonia [NH ₃]	 damp red litmus paper	Gas evolved turns damp red litmus paper blue . alkali + ammonium salt → salt + water + ammonia
2. carbon dioxide [CO ₂]	 rubber bung delivery tube limewater	Gas evolved gives white precipitate (ppt.) with limewater . (ppt. dissolves with excess CO ₂) acid + carbonate → salt + water + carbon dioxide [CO₂ (g) + Ca(OH)₂ (aq) → CaCO₃ (s) + H₂O (l)]
3. chlorine [Cl ₂]	 damp blue litmus paper	Gas evolved bleaches damp litmus paper .
4. hydrogen [H ₂]	 lighted splinter	Gas evolved extinguishes a lighted splint with a ' pop ' sound. metal + water → metal hydroxide + hydrogen metal + steam → metal oxide + hydrogen metal + acid → salt + hydrogen
5. oxygen [O ₂]	 glowing splinter	Gas evolved relights a glowing splint .
6. sulfur dioxide [SO ₂]	 acidified aqueous potassium manganate(VII)	Gas evolved turns aqueous acidified potassium manganate(VII) from purple to colourless .

Test for Cations

effect of aqueous sodium hydroxide		effect of aqueous ammonia
cation	[NaOH]	[NH ₄ OH]
no ppt. formed	Na ⁺ No ppt. No gas evolved.	No ppt. No gas evolved.
	K ⁺ No ppt. No gas evolved.	No ppt. No gas evolved.
	NH ₄ ⁺ No ppt. On warming, gas evolved turns damp red litmus paper blue. $\text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{NH}_3 (\text{g}) + \text{H}_2\text{O} (\text{l})$	-
white ppt. formed	Ca ²⁺ <u>White</u> ppt., <i>insoluble</i> in excess. $\text{Ca}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Ca}(\text{OH})_2 (\text{s})$	No ppt. or very slight white ppt.
	Al ³⁺ <u>White</u> ppt., <i>soluble</i> in excess giving a colourless solution. $\text{Al}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Al}(\text{OH})_3 (\text{s})$	<u>White</u> ppt., <i>insoluble</i> in excess. $\text{Al}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Al}(\text{OH})_3 (\text{s})$
	Zn ²⁺ <u>White</u> ppt., <i>soluble</i> in excess giving a colourless solution. $\text{Zn}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Zn}(\text{OH})_2 (\text{s})$	<u>White</u> ppt., <i>soluble</i> in excess giving a colourless solution. $\text{Zn}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Zn}(\text{OH})_2 (\text{s})$
coloured ppt. formed	Cu ²⁺ <u>Light blue</u> ppt., <i>insoluble</i> in excess. $\text{Cu}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Cu}(\text{OH})_2 (\text{s})$	<u>Light blue</u> ppt., <i>soluble</i> in excess giving a dark blue solution. $\text{Cu}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Cu}(\text{OH})_2 (\text{s})$
	Fe ²⁺ <u>Green</u> ppt., <i>insoluble</i> in excess. $\text{Fe}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_2 (\text{s})$	<u>Green</u> ppt., <i>insoluble</i> in excess. $\text{Fe}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_2 (\text{s})$
	Fe ³⁺ <u>Red-brown</u> ppt., <i>insoluble</i> in excess. $\text{Fe}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_3 (\text{s})$	<u>Red-brown</u> ppt., <i>insoluble</i> in excess. $\text{Fe}^{3+} (\text{aq}) + 3\text{OH}^- (\text{aq}) \rightarrow \text{Fe}(\text{OH})_3 (\text{s})$

Tip: to help recall the three cations that give a white ppt.

ions	aq. NaOH		aq. NH ₄ OH	
	a few drops	in excess	a few drops	in excess
C [Ca ²⁺]			O	O
A [Al ³⁺]		O		
Z [Zn ²⁺]		O		O

key:

| = ppt. formed

O = ppt. dissolves

Summary

Test for Anions

anion	test	observation	conclusion
carbonate [CO ₃ ²⁻]	(a) Heat the solid salt	(a) Gas evolved gives a <u>white ppt.</u> with <u>limewater</u> .	<u>CO₂</u> evolved, therefore CO ₃ ²⁻ is present. [Sodium and potassium carbonate are stable to heat.]
	(b) Add dilute acid to the solid salt [HCl, HNO ₃ , H ₂ SO ₄]	(b) Effervescence observed. Gas evolved gives a <u>white ppt.</u> with <u>limewater</u> .	<u>CO₂</u> evolved, therefore CO ₃ ²⁻ is present.
chloride [Cl ⁻]	To a solution of the salt, acidify with dilute nitric acid [HNO ₃], then add aqueous silver nitrate [AgNO ₃].	<u>White ppt.</u> formed.	Cl ⁻ present. $\text{Ag}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) \rightarrow \text{AgCl} (\text{s})$ white ppt.
nitrate [NO ₃ ⁻]	To the solid salt or solution of the salt, add aqueous sodium hydroxide [NaOH], then aluminium foil [Al]; warm carefully.	Gas evolved turns <u>damp red litmus</u> <u>paper</u> <u>blue</u> .	<u>NH₃</u> evolved, therefore NO ₃ ⁻ is present.
sulfate [SO ₄ ²⁻]	To a solution of the salt, acidify with dilute nitric acid [HNO ₃], then add aqueous barium nitrate [Ba(NO ₃) ₂].	<u>White ppt.</u> formed.	SO ₄ ²⁻ present. $\text{Ba}^{2+} (\text{aq}) + \text{SO}_4^{2-} (\text{aq}) \rightarrow \text{BaSO}_4 (\text{s})$ white ppt.

REDOX CHEMISTRY

	Oxidation	Reduction
definition	- A <u>gain</u> of oxygen - A <u>loss</u> of hydrogen - A <u>loss</u> of electrons - An <u>increase</u> in oxidation states	- A <u>loss</u> of oxygen - A <u>gain</u> of hydrogen - A <u>gain</u> of electrons - A <u>decrease</u> in oxidation states

	Redox Reactions	Non-Redox Reactions
examples	1. Displacement reactions $\text{Mg (s)} + \text{FeSO}_4 \text{ (aq)} \rightarrow \text{MgSO}_4 \text{ (aq)} + \text{Fe (s)}$ $\text{Ca (s)} + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$ $2\text{Na (s)} + 2\text{H}_2\text{O (l)} \rightarrow 2\text{NaOH (aq)} + \text{H}_2 \text{ (g)}$ $\text{Cl}_2 \text{ (g)} + 2\text{KI (aq)} \rightarrow 2\text{KCl (aq)} + \text{I}_2 \text{ (s)}$ 2. Reactions involving addition and removal of oxygen or hydrogen $2\text{SO}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{SO}_3 \text{ (aq)}$ $\text{N}_2 \text{ (g)} + 3\text{H}_2 \text{ (g)} \rightarrow 2\text{NH}_3 \text{ (g)}$ 3. Electrolytic reactions	1. Neutralisation reactions $2\text{KOH (aq)} + \text{H}_2\text{SO}_4 \text{ (aq)} \rightarrow \text{K}_2\text{SO}_4 \text{ (aq)} + 2\text{H}_2\text{O (l)}$ $2\text{HCl (aq)} + \text{CuO (s)} \rightarrow \text{CuCl}_2 \text{ (aq)} + \text{H}_2\text{O (l)}$ $\text{CaCO}_3 \text{ (s)} + 2\text{HNO}_3 \text{ (aq)} \rightarrow \text{Ca(NO}_3)_2 \text{ (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$ 2. Precipitation reactions $2\text{NaOH (aq)} + \text{CuSO}_4 \text{ (aq)} \rightarrow \text{Cu(OH)}_2 \text{ (s)} + \text{Na}_2\text{SO}_4 \text{ (aq)}$ $\text{BaCl}_2 \text{ (aq)} + \text{K}_2\text{SO}_4 \text{ (aq)} \rightarrow \text{BaSO}_4 \text{ (s)} + 2\text{KCl (aq)}$ $\text{AgNO}_3 \text{ (aq)} + \text{NaCl (aq)} \rightarrow \text{AgCl (s)} + \text{NaNO}_3 \text{ (aq)}$

PATTERNS IN THE PERIODIC TABLE

8.1 Periodic Trends

- The elements are arranged in the Periodic Table in order of **increasing proton (atomic) number**.
 - A **Period** is the horizontal row of elements (i.e. Period 1, 2, 3, etc).
 - A **Group** is the vertical column of elements (i.e. Group 1, 2, 13, etc).
- The position of an element in the Periodic Table is related to proton number and electronic structure.
 - The atoms of elements in the **same period** have the same number of electron shells.

e.g.	Na	Mg	S	Cl	
	2.8.1	2.8.2	2.8.6	2.8.7	The atoms of elements in <u>Period 3</u> have <u>three electron shells</u> .

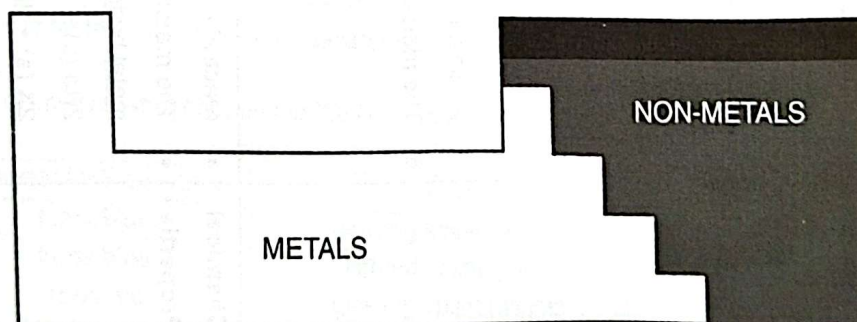
- The atoms of elements in the **same group** have the same number of outer (valence) electrons.

e.g.	Be	Mg	Ca	
	2.2	2.8.2	2.8.8.2	The atoms of elements in <u>Group 2</u> have <u>two outer (valence) electrons</u> .

- The atoms of elements in the **same group** lose/gain the same number of outer (valence) electrons to form ions with the same ionic charge, achieving a stable electronic configuration.

e.g.	1	2	13	15	16	17
	metal (loses 1e ⁻)	metal (loses 2e ⁻)	metal (loses 3e ⁻)	non-metal (gains 3e ⁻)	non-metal (gains 2e ⁻)	non-metal (gains 1e ⁻)
	Na ⁺	Mg ²⁺	Al ³⁺	N ³⁻	O ²⁻	F ⁻
	K ⁺	Ca ²⁺		P ³⁻	S ²⁻	Cl ⁻

- The elements change from **metals (left)** to **non-metals (right)** across a period of the Periodic Table.
 - e.g.: metals – Groups 1, 2, 13
 - non-metals – Groups 14, 15, 16, 17, 18



8.2 Group Properties

	Group 1 (the alkali metals)	Group 17 (the halogens)																								
Physical Properties	- The metals are <u>silvery, shiny</u> solids. - The metals are <u>soft</u> and <u>can be easily cut with a knife</u>. - The metals are <u>good conductors</u> of <u>heat</u> and <u>electricity</u>. - The metals have a <u>low density</u>. - The metals have a <u>low melting point</u>, which <u>decreases</u> down the group. ➤ solid – Li, Na, K, Rb, Cs ➤ liquid – Fr - The metals are stored under oil as they tarnish easily in air.	- The elements exist as <u>diatomic molecules</u>. [simple covalent] - The halogens have <u>low melting and boiling points</u>, which <u>increases</u> down the group. - The <u>colour intensity</u> of the halogens <u>increases</u> down the group. (i.e. the elements become darker down the group) <table><tr><th>element</th><th>molecular formula</th><th>physical state</th><th>colour</th></tr><tr><td>fluorine</td><td>F₂</td><td>gas</td><td>pale yellow</td></tr><tr><td>chlorine</td><td>Cl₂</td><td>gas</td><td>greenish-yellow</td></tr><tr><td>bromine</td><td>Br₂</td><td>liquid</td><td>red-brown</td></tr><tr><td>iodine</td><td>I₂</td><td>solid</td><td>shiny black</td></tr><tr><td>astatine</td><td>At₂</td><td>solid</td><td>dull black</td></tr></table>	element	molecular formula	physical state	colour	fluorine	F ₂	gas	pale yellow	chlorine	Cl ₂	gas	greenish-yellow	bromine	Br ₂	liquid	red-brown	iodine	I ₂	solid	shiny black	astatine	At ₂	solid	dull black
element	molecular formula	physical state	colour																							
fluorine	F ₂	gas	pale yellow																							
chlorine	Cl ₂	gas	greenish-yellow																							
bromine	Br ₂	liquid	red-brown																							
iodine	I ₂	solid	shiny black																							
astatine	At ₂	solid	dull black																							
Chemical Properties	- reactivity: Fr (most) > Cs > Rb > K > Na > Li (least) - The metals react vigorously with <u>cold water</u> to form a <u>soluble metal hydroxide</u> (alkali) and an <u>insoluble hydrogen</u> gas. 2Na (s) + 2H₂O (l) → 2NaOH (aq) + H₂ (g) 2K (s) + 2H₂O (l) → 2KOH (aq) + H₂ (g) - The metals react with <u>acid</u> to form a <u>salt</u> and <u>hydrogen</u> gas. - The metals react with <u>chlorine</u> to form <u>ionic chlorides</u>.	- reactivity: F (most) > Cl > Br > I > At (least) - The halogens react with <u>metals</u> to form <u>ionic salts</u>. - The halogens undergo <u>displacement</u> reactions, where a <u>more reactive</u> halogen displaces a <u>less reactive</u> halogen from its halide solution. Cl₂ + 2KBr → 2KCl + Br₂ Br₂ + 2KI → 2KBr + I₂ Cl₂ + 2NaI → 2NaCl + I₂ Br₂ + NaCl → * no reaction																								

Group 18 (the noble gases)

- ❖ The noble gases exist as monoatomic gases which have a stable electronic configuration.
- ❖ Noble gases are chemically unreactive and hence are important in providing an inert atmosphere.
- ❖ Except for helium, all members have eight electrons in their valence shell.
e.g.: He (2) Ne (2,8) Ar (2,8,8)
- ❖ Uses:
 - (a) helium : to fill balloons
 - (b) neon : to fill light bulbs
 - (c) argon : to fill light bulbs
 - : used in the manufacture of steel

Examples:

1 How are the elements arranged in the Periodic Table?

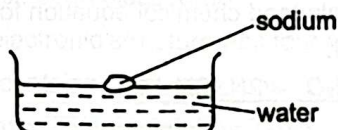
- A in order of chemical reactivity C in order of proton number
B in order of electron shells D in order of relative atomic mass (C)

2 Three elements each have two valency electrons.

Which statement about the elements is correct?

- A They are all metallic and belong to the same group of the Periodic Table.
B They are all metallic but belong to different groups of the Periodic Table.
C They are all non-metallic and belong to the same group of the Periodic Table.
D They are all non-metallic but belong to different groups of the Periodic Table. (A)

3 When sodium reacts with water, a solution and a gas are produced.



The solution is tested with Universal Indicator paper and the gas is tested with a splint.

Which row describes what happens to the Universal Indicator paper and to the splint?

	Universal Indicator paper	splint
A	turns blue	glowing splint relights
B	turns blue	lighted splint 'pops'
C	turns red	glowing splint relights
D	turns red	lighted splint 'pops'

(B)

8.3 Reactivity Series

- Reaction of metals with water, steam and dilute hydrochloric acid

	metal	reactions			extraction
please	potassium	cold water $\text{H}_2\text{O} (l)$	steam $\text{H}_2\text{O} (g)$	acid $\text{HCl} (aq)$ $\text{H}_2\text{SO}_4 (aq)$ $\text{HNO}_3 (aq)$	electrolysis
send	sodium				
cats	calcium				
monkeys	magnesium				
and	aluminium	do not react	steam $\text{H}_2\text{O} (g)$	acid $\text{HCl} (aq)$ $\text{H}_2\text{SO}_4 (aq)$ $\text{HNO}_3 (aq)$	heating with carbon / hydrogen
zebras	zinc				
in	iron				
large	lead				
hired	hydrogen [nm]				
cages	copper	do not react	do not react	do not react	physical method
make	mercury				
sure	silver				
good	gold				

- Ease of obtaining metals from their ores

K	Na	Ca	Mg	Al	Zn	Fe	Pb	(H)	Cu	Hg	Ag	Au
electrolysis					heating with carbon / hydrogen						physical method	
- very reactive metals - form very stable compounds					- less reactive metals - form less stable compounds						unreactive metal	

- Corrosion (rusting) of iron
 - ❖ Essential conditions: (a) oxygen (b) water
 - ❖ Prevention of rusting:

	method	what is it used on?	remarks
surface barrier protection (no reaction involved!)	1. painting	❖ large iron and steel objects ❖ e.g.: motorcars, ships, bridges, steel pipes and street poles	
	2. greasing	❖ machinery	❖ helps in lubrication
	3. plastic coating	❖ small iron and steel objects ❖ e.g.: wires, metal clips	
sacrificial protection (reaction involved!)	4. sacrificial protection	❖ underwater pipes	❖ Mg , being <u>more reactive</u> than Fe, corrodes in place of Fe
	5. galvanising	❖ hulls of ships / ocean liners	❖ Zn , being <u>more reactive</u> than Fe, corrodes in place of Fe

1. Chemical properties:

1. Reaction with water or steam



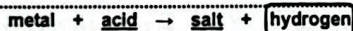
Test for hydrogen (H₂):

test : Insert a **lighted splinter** into the mouth of test-tube.

observation : Gas evolved **extinguishes** the lighted splint with a 'pop' sound.

metal	reaction with		observations / equations
	water	steam	
K	explosive	extremely explosive	Potassium reacts very vigorously with cold water. Potassium hydroxide solution and hydrogen gas are produced. A lot of heat is also produced which causes the hydrogen to catch fire and explode. $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
Na	vigorously	explosive	Sodium reacts vigorously with cold water. Sodium hydroxide solution and hydrogen gas are produced. A lot of heat is also produced which causes the hydrogen to catch fire and explode. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
Ca	moderate	explosive	Calcium reacts readily with cold water. Calcium hydroxide solution and hydrogen gas are produced. $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
Mg	v. slowly	vigorous	Magnesium reacts very slowly with cold water. Magnesium hydroxide solution and hydrogen gas are produced. $Mg(s) + 2H_2O(l) \rightarrow 2Mg(OH)_2(aq) + H_2(g)$ Magnesium reacts vigorously with steam, to produce magnesium oxide and hydrogen gas. $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$
Al	no reaction	no reaction	
Zn	no reaction	moderate	Zinc does not react with cold water. Hot zinc reacts with steam, to produce zinc oxide and hydrogen gas. $Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$
Fe	no reaction	slow	Iron does not react with cold water. Hot iron reacts slowly with steam, to produce iron(II, III) oxide and hydrogen gas. $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
Pb	no reaction	slow	
Cu	no reaction	no reaction	
Ag	no reaction	no reaction	

2. Reaction with dilute acid

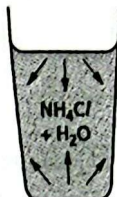
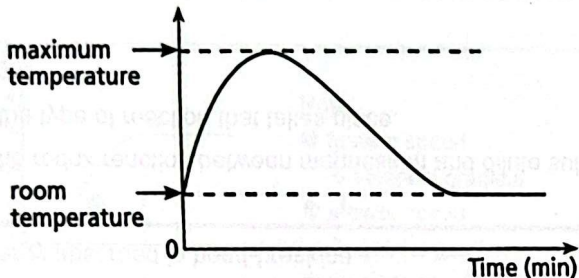
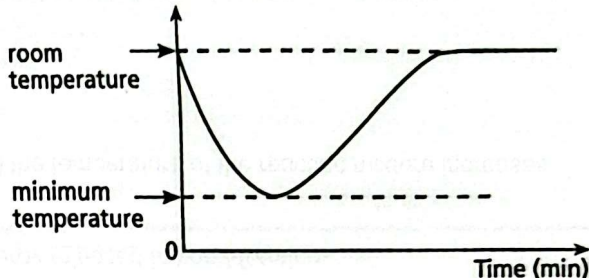


metal	reaction with dilute HCl	observations / equations
K	explosive	Potassium chloride solution and hydrogen gas are produced. $2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$
Na		Sodium chloride solution and hydrogen gas are produced. $2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
Ca	v. fast	Calcium chloride solution and hydrogen gas are produced. $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
Mg	fast	Magnesium chloride solution and hydrogen gas are produced. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Al	no reaction	
Zn	moderately fast	Zinc chloride solution and hydrogen gas are produced. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
Fe	slowly	A pale green solution of iron(II) chloride and hydrogen gas are produced. $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
Pb	v. slowly	
Cu	no reaction	
Ag		

❖ Note: Although Al is high up in the reactivity series, it does not react with cold water, steam or dilute acid.

This is due to its **protective oxide layer** (Al₂O₃), which prevents any reaction.

CHEMICAL ENERGETICS

	Exothermic (ΔH negative)	Endothermic (ΔH positive)
definition	A process or chemical reaction which <u>releases</u> heat energy to the surroundings and may be detected by an <u>increase</u> in temperature, e.g. the reaction between sodium hydroxide and hydrochloric acid. 	A process or chemical reaction which <u>absorbs</u> heat energy from the surroundings and may be detected by a <u>decrease</u> in temperature, e.g. the dissolving of ammonium nitrate in water. 
graph	Temperature ($^{\circ}\text{C}$)  Time (min)	Temperature ($^{\circ}\text{C}$)  Time (min)
examples	1. condensation and freezing 2. combustion 3. neutralisation 4. bond making / forming	1. melting, boiling and sublimation 2. dissolving of ammonium salts in water 3. thermal decomposition of compounds 4. bond breaking

examples	$2\text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ The reaction is exothermic . The energy <u>released</u> in <u>bond-forming</u> is greater than the energy <u>absorbed</u> in <u>bond-breaking</u> .	$\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$ The reaction is endothermic . The energy <u>absorbed</u> in <u>bond-breaking</u> is greater than the energy <u>released</u> in <u>bond-forming</u> .
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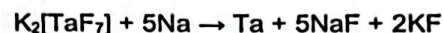
Example 1: During the redox reaction between magnesium and dilute sulfuric acid the temperature of the reaction mixture increases. Identify the type of reaction that takes place.

Exothermic

Example 2: Tantalum capacitors are found in nearly every piece of electronic equipment, including cell phones and laptops.

The metal is extracted by first converting its ore to potassium heptafluorotantalate, $\text{K}_2[\text{TaF}_7]$.

Tantalum can then be extracted by treating this compound with sodium under extreme conditions.



This reaction is exothermic. Define the term *exothermic*.

An exothermic reaction releases heat energy to the surroundings and may be detected by an increase in temperature.

Example 3: Briefly describe the method you would use and the results you would expect in distinguishing between an endothermic and an exothermic reaction.

method : Insert a thermometer into each of the solutions.

result for an endothermic reaction : The temperature of solution decreases.

result for an exothermic reaction : The temperature of solution increases.

RATE OF REACTIONS

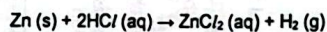
	Method 1: By measuring the vol. of gas evolved in a given time	Method 2: By measuring mass loss in a given time
example	$\text{Mg (s)} + 2\text{HCl (aq)} \rightarrow \text{MgCl}_2 \text{ (aq)} + \text{H}_2 \text{ (g)}$	$\text{CaCO}_3 \text{ (s)} + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 \text{ (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2 \text{ (g)}$
expt. set-up	Labels: conical flask, small test tube, dilute hydrochloric acid, magnesium ribbon, gas syringe, thread, stopwatch, retort stand.	Labels: thread, cotton wool, small test tube, marble chip, conical flask, dilute hydrochloric acid, electronic balance (60.00g), stopwatch.
graph	Note: ① fastest speed ↳ steepest gradient ② slower speed ↳ gentler gradient ③ no more reaction ↳ zero gradient	Question: How does the speed of reaction change with time? Answer: 1m – The speed of reaction <u>decreases</u> with time. 2m – The speed of reaction was initially <u>fast</u>. As the reaction proceeds, the speed <u>slows down</u> and eventually <u>stops</u>.

FACTORS	EFFECT	EXPLANATION (in terms of collisions between reacting particles)
concentration [solution]	• faster speed – higher concentration • slower speed – lower concentration	• Concentration (g/dm^3) = $\frac{\text{Mass of solute (g)}}{\text{Volume of solution (dm}^3\text{)}}$ • An increase in the concentration of reactant ➢ <u>increases the number of reactant particles</u> in a given volume, ➢ leading to a <u>higher frequency of effective collisions</u> between the particles, ➢ resulting in a <u>higher rate of reaction</u>.
pressure [gas]	• faster speed – higher pressure • slower speed – lower pressure	• Pressure $\propto \frac{1}{\text{Volume}}$ • An increase in pressure ➢ causes the reactant particles to be <u>closer</u> to each other, ➢ leading to a <u>higher frequency of effective collisions</u> between the particles, ➢ resulting in a <u>higher rate of reaction</u>.
particle size [solid]	• faster speed – finely divided powder • slower speed – big lumps / chips / chunks / granules	• Surface area $\propto \frac{1}{\text{Particle size}}$ • A decrease in particle size ➢ <u>increases the available surface area</u> of contact between reactant particles, ➢ leading to a <u>higher frequency of effective collisions</u> between the particles, ➢ resulting in a <u>higher rate of reaction</u>.

temperature [solution]	• faster speed – higher temperature • slower speed – lower temperature	• Temperature $\propto$ Rate • Temperature $\propto$ Kinetic Energy (K.E.) • At higher temperatures, ➢ the reactant particles move <u>faster</u> / <u>possess more K.E.</u>, ➢ leading to a <u>higher frequency of effective collisions</u> between the particles, ➢ resulting in a <u>higher rate of reaction</u>.
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Example 1: When hydrochloric acid reacts with lumps of zinc, hydrogen gas is given off.

State and explain **three** ways of increasing the speed of this reaction. Use your knowledge of reacting particles in your explanations.



Method 1: Increase the concentration of hydrochloric acid.

An increase in the concentration of hydrochloric acid increases the number of reactant particles in a given volume, leading to a higher frequency of effective collisions between the particles.

Method 2: Use finely divided zinc.

Using finely divided zinc increases the available area of contact between reactants, leading to a higher frequency of effective collisions between the particles.

Method 3: Increase the temperature of the reaction system.

An increase in the temperature causes the reactant particles to move faster as they possess more K.E., leading to a higher frequency of effective collisions between the particles.

ORGANIC CHEMISTRY

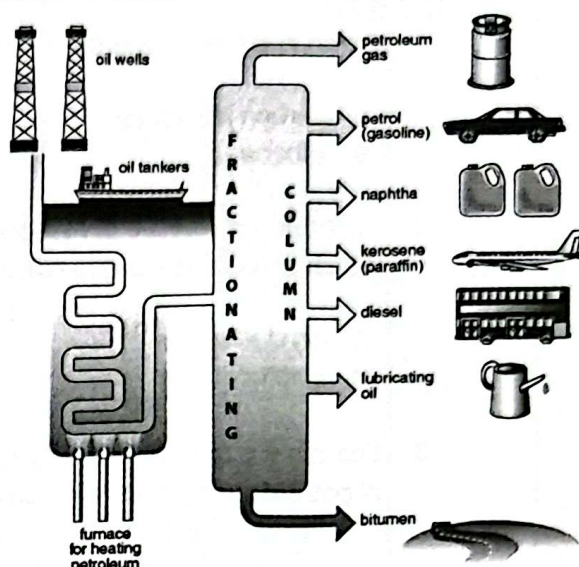
11.1 Fuels and Crude Oil

- Non-renewable sources of energy – natural gas (mainly methane) and crude oil
- Petroleum is a mixture of hydrocarbons.

It can be separated into useful fractions by fractional distillation in the oil refinery, as its fractions are miscible liquids with different boiling points.

- Fractional distillation of crude oil

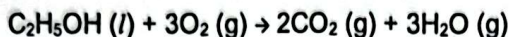
- ❖ Crude oil (petroleum) is heated in a furnace. The oil vapourises and passes up the fractionating column.
- ❖ The fractions condense and come out of the column at different heights depending on their boiling points.
 - The petroleum gas fraction (smaller hydrocarbons) is collected at the top of the column as its molecules have the lowest boiling points.
 - The bitumen fraction (bigger hydrocarbons) is collected at the bottom of the column as the residue as its molecules have the highest boiling points.



- Fractions and their uses:

	fraction	uses
P	<u>p</u> etroleum gas	as a <u>f</u> uel for <u>c</u> ooking
P	<u>p</u> etrol (gasoline)	as a <u>f</u> uel in <u>c</u> ars
N	<u>n</u> aphtha	as the feedstock and main source of hydrocarbons used for the production of a wide range of organic compounds in the petrochemical industry
K	<u>k</u> erosene (paraffin)	as a <u>f</u> uel for heating and cooking and for <u>a</u> ircraft engines
D	<u>d</u> iesel	as a <u>f</u> uel for <u>d</u> iesel engines
L	<u>l</u> ubricating oil	as <u>l</u> ubricants and as a source of <u>p</u> olishes and <u>w</u> axes
B	<u>b</u> itumen	for <u>m</u> aking road surfaces

- Renewable source of energy – biofuel (e.g. bioethanol from sugarcane)
 - ❖ Burning bioethanol is considered more environmentally sustainable compared to burning fossil fuels like petrol.
 - As sugarcane grows, it takes in carbon dioxide during photosynthesis.
 - This offsets the carbon dioxide produced when the bioethanol is burnt.



Examples:

1 Which is the main source of energy in natural gas?

- | | | |
|--------------------------|-------------------|-------|
| A carbon monoxide | C hydrogen | |
| B ethene | D methane | (D) |

2 Naphtha is a mixture of hydrocarbons obtained from petroleum.
Which process is used to obtain naphtha?

- | | | |
|--------------------------|----------------------------------|-------|
| A crystallisation | C filtration | |
| B displacement | D fractional distillation | (D) |

3 The table shows some fractions obtained from the fractional distillation of petroleum together with some of their uses.

fraction	uses
petroleum gas	fuel for cooking
gasoline	fuel for cars
1	making chemicals
2	fuel for aircraft
3	making polishes and waxes
bitumen	making roads

Which row identifies fractions 1, 2 and 3?

	1	2	3	
A	lubricating oil	diesel oil	paraffin	
B	naphtha	paraffin	diesel oil	
C	naphtha	paraffin	lubricating oil	
D	paraffin	naphtha	lubricating oil	(C)

11.2-11.3 Hydrocarbons, Alcohols and Carboxylic Acids

- A **homologous series** is a group of compounds with a general formula, similar chemical properties and showing a gradation in physical properties as a result of increase in the size and mass of the molecules, e.g. melting and boiling points ↑; viscosity ↑; flammability ↓.

homologous series	hydrocarbons		non-hydrocarbons	
	alkanes	alkenes	alcohols	carboxylic acids
general formula	C_nH_{2n+2}	C_nH_{2n}	$C_nH_{2n+1}OH$	$C_nH_{2n+1}COOH$
nomenclature	~ ane	~ ene	~ ol	~ oic acid
functional group	nil [saturated]	$C = C$ [unsaturated]	$-O-H$	$\begin{array}{c} O \\ \\ -C-O-H \end{array}$
chemical properties	1. combustion 2. substitution by chlorine	1. combustion 2. addition (a) $+ H_2 \rightarrow$ alkane (b) $+ Cl_2 \rightarrow$ dichloroalkane (c) polymerisation	1. combustion 2. oxidation $\rightarrow$ carboxylic acid	1. combustion 2. acidic reactions (a) $+ metal \rightarrow salt + H_2$ (b) $+ base \rightarrow salt + H_2O$ (c) $+ carbonate \rightarrow salt + H_2O + CO_2$
formation	fractional distillation of crude oil	cracking of long-chain hydrocarbons	<u>ethanol</u> fermentation of glucose	<u>ethanoic acid</u> oxidation of ethanol by: (i) atmospheric oxygen (ii) acidified $KMnO_4$

Structural Formulae

alkanes	alkenes	alcohols	carboxylic acids
C_nH_{2n+2}	C_nH_{2n}	$C_nH_{2n+1}OH$	$C_nH_{2n+1}COOH$
$\begin{array}{c} H \\ \\ H-C-H \\ \\ H \end{array}$ methane, CH_4	-	$\begin{array}{c} H \\ \\ H-C-O-H \\ \\ H \end{array}$ methanol, CH_3OH	$\begin{array}{c} O \\ \\ H-C-O-H \end{array}$ methanoic acid, $HCOOH$
$\begin{array}{c} H & H \\ & \\ H-C-C-H \\ & \\ H & H \end{array}$ ethane, C_2H_6	$\begin{array}{c} H & H \\ & \\ H-C=C-H \end{array}$ ethene, C_2H_4	$\begin{array}{c} H & H \\ & \\ H-C-C-O-H \\ & \\ H & H \end{array}$ ethanol, C_2H_5OH	$\begin{array}{c} H & O \\ & \\ H-C-C-O-H \\ \\ H \end{array}$ ethanoic acid, CH_3COOH
$\begin{array}{c} H & H & H \\ & & \\ H-C-C-C-H \\ & & \\ H & H & H \end{array}$ propane, C_3H_8	$\begin{array}{c} H & H & H \\ & & \\ H-C=C-C-H \\ \\ H \end{array}$ propene, C_3H_6	$\begin{array}{c} H & H & H \\ & & \\ H-C-C-C-O-H \\ & & \\ H & H & H \end{array}$ propanol, C_3H_7OH	$\begin{array}{c} H & H & O \\ & & \\ H-C-C-C-O-H \\ & \\ H & H \end{array}$ propanoic acid, C_2H_5COOH

Example 1: Aldehydes are a homologous series of organic compounds.

The table shows the names, formulae and boiling points of some aldehydes.

name	formula	boiling point / °C
methanal	HCHO	-19
ethanal	CH ₃ CHO	20
propanal	C ₂ H ₅ CHO	49
pentanal	C ₄ H ₉ CHO	103

Use the information in the table to give **two** pieces of evidence that suggest that the aldehydes are a homologous series.

All the aldehydes have the same general formula (i.e. $C_nH_{2n+1}CHO$) and the same functional group (i.e. CHO).

Example 2: Organic compounds are placed in an homologous series.

(a) Give **two** general properties of an homologous series.

Members of the same homologous series have the same general formula / same functional group and undergo similar chemical reactions.

(b) Write the general formula for the homologous series of alkanes. C_nH_{2n+2}

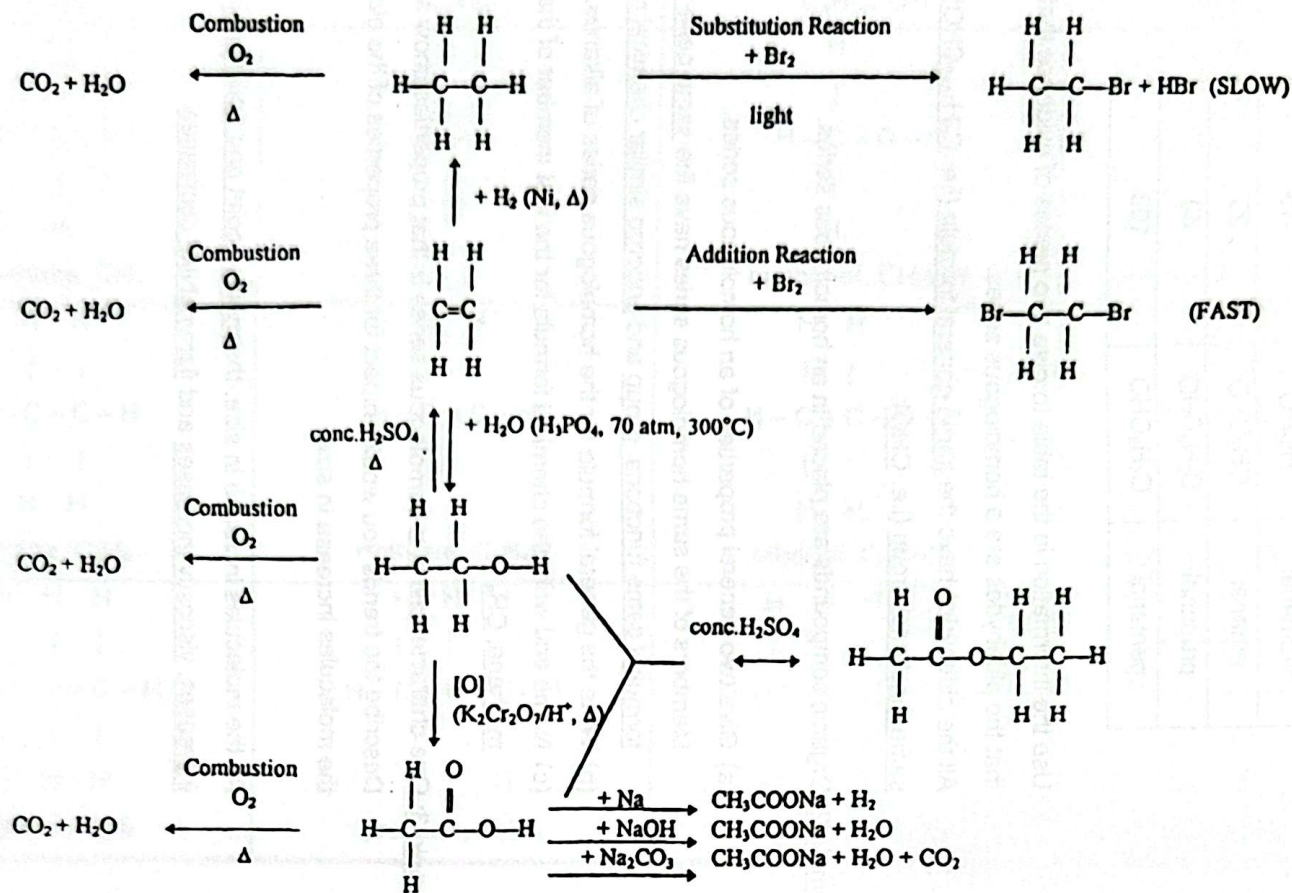
(c) Name and write the chemical formula for the first member of the alkanes.
methane, CH₄

Example 3: One characteristic of a homologous series is that properties show a trend.

Describe the trends you would expect for three properties of the alcohols as the molecules increase in size.

As the molecules increase in size, the melting point and boiling point increases, viscosity increases and flammability decreases.

ORGANIC CHEMISTRY – Overall View of Chemical Reactions



(A) ALKANES

Type of Reaction	Equation	Conditions
combustion	complete: $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ incomplete: $2\text{CH}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}(\text{g}) + 4\text{H}_2\text{O}(\text{l})$	
substitution	substitution by chlorine $\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array} + \text{Cl}_2 \xrightarrow{\text{sunlight}} \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{Cl} \\ \\ \text{H} \end{array} + \text{H}-\text{Cl}$ methane chloromethane	• sunlight

(B) ALKENES

Type of Reaction	Equation	Conditions
combustion	complete: $\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ incomplete: $\text{C}_2\text{H}_4(\text{g}) + 4\text{O}_2(\text{g}) \rightarrow 2\text{CO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$	
addition	(a) addition of hydrogen [to form alkanes] $\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}=\text{C}-\text{H} \end{array} + \text{H}_2 \rightarrow \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$ ethene ethane	• 150-200°C • nickel catalyst
	(b) addition of halogens [to form halogenoalkanes] $\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}=\text{C}-\text{H} \end{array} + \text{Cl}_2 \rightarrow \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{Cl} \quad \text{Cl} \end{array}$ ethene dichloroethane	
Uses as a typical plastic: • plastic bags • clingfilm	(c) addition polymerisation $\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{C}=\text{C} + \text{C}=\text{C} + \text{C}=\text{C} + \dots \rightarrow \left(\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ -\text{C}-\text{C}- \end{array} \right)_n$ ethene poly(ethene) [monomer] [polymer]	• high temp • high pressure • catalyst

- Pollution problems caused by the disposal of non-biodegradable plastics
 - ❖ **land pollution** – as plastics are non-biodegradable (i.e. cannot be decomposed by bacteria easily), they stay on land for a long period of time and cause landfill problems
 - ❖ **air pollution** – as plastics undergo incomplete combustion, toxic gases like carbon monoxide are released into the atmosphere

(C) ALCOHOLS

Type of Reaction	Equation
combustion	complete: $\text{C}_2\text{H}_5\text{OH} (l) + 3\text{O}_2 (g) \rightarrow 2\text{CO}_2 (g) + 3\text{H}_2\text{O} (l)$
oxidation	(a) atmospheric oxygen $\text{C}_2\text{H}_5\text{OH} + \text{O}_2 \rightarrow \text{CH}_3\text{COOH} + \text{H}_2\text{O}$ (b) acidified potassium manganate(VII) $[\text{H}^+/\text{KMnO}_4]$ [The <u>purple</u> solution turns <u>colourless</u>.] $ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} + \text{H}^+/\text{KMnO}_4 \rightarrow \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{O} - \text{H} \\ \\ \text{H} \end{array} + \text{Mn}^{2+} + \text{pdt} $ ethanol ethanoic acid

(D) CARBOXYLIC ACIDS

Type of Reaction	Equation	Conditions
combustion	complete: $\text{CH}_3\text{COOH} (l) + 2\text{O}_2 (g) \rightarrow 2\text{CO}_2 (g) + 2\text{H}_2\text{O} (l)$	
acidic reactions Carboxylic acids are weak acids , which undergo partial ionisation in water to produce a smaller amount of H^+ ions.	(a) acid + metal $\rightarrow$ salt + H_2 $2\text{CH}_3\text{COOH} + 2\text{Na} \rightarrow 2\text{CH}_3\text{COONa} + \text{H}_2$ $2\text{CH}_3\text{COOH} + \text{Mg} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg} + \text{H}_2$	
	(b) acid + base $\rightarrow$ salt + H_2O $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O}$ $2\text{CH}_3\text{COOH} + \text{MgO} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg} + \text{H}_2\text{O}$	
	(c) acid + carbonate $\rightarrow$ salt + $\text{H}_2\text{O} + \text{CO}_2$ $2\text{CH}_3\text{COOH} + \text{Na}_2\text{CO}_3 \rightarrow 2\text{CH}_3\text{COONa} + \text{H}_2\text{O} + \text{CO}_2$	

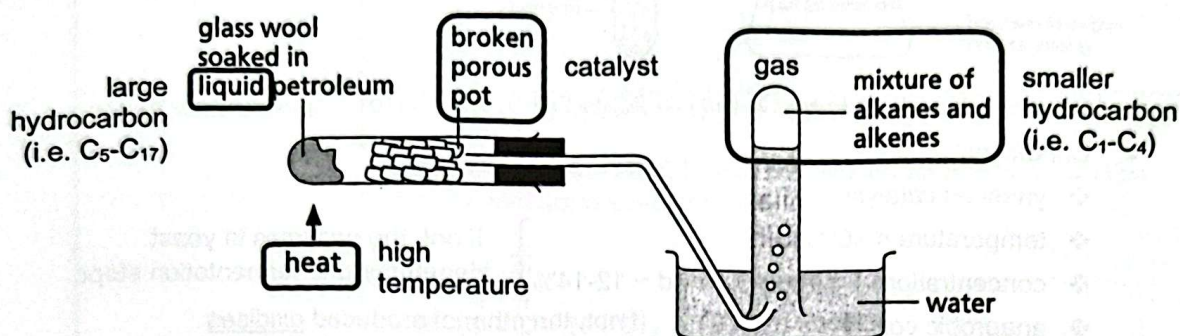
Cracking

- Describe the cracking process.

Cracking is the break-down of **large** (long-chain) hydrocarbons into **smaller** (short-chain) molecules, in the presence of **high temperature**, **high pressure**, usually in the presence of a **finely divided catalyst** (e.g. aluminium oxide, Al_2O_3 , silicon dioxide, broken porous pot, broken glass).

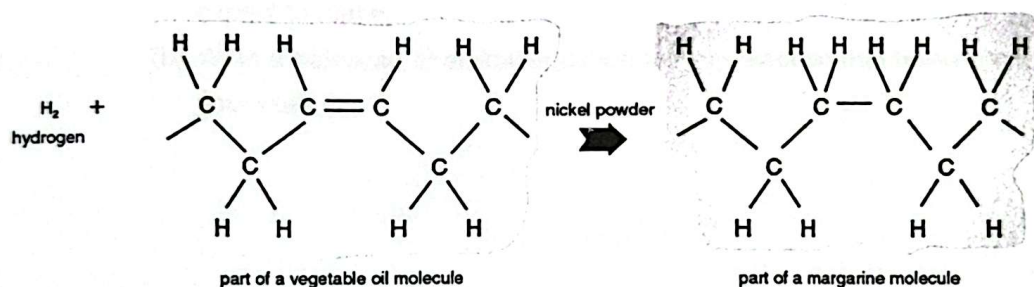
- Explain why cracking is an important process in an oil refinery.

Cracking is essential to **match the demand** for fractions containing smaller molecules from the refinery process.



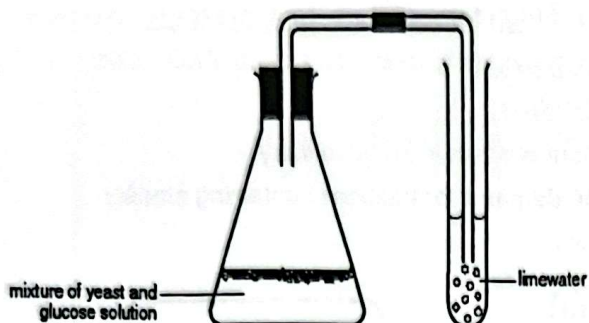
Manufacture of Margarine

- Vegetable oils are **polyunsaturated** (i.e. contains many $C=C$ double bonds).
- Margarine is made by reacting vegetable oil with **hydrogen** in the presence of a **nickel** catalyst.



Fermentation

- Fermentation of glucose is the process of making ethanol in a laboratory at room temperature.



- Chemical equation: $\text{C}_6\text{H}_{12}\text{O}_6 (\text{aq}) \rightarrow 2\text{C}_2\text{H}_5\text{OH} (\text{l}) + 2\text{CO}_2 (\text{g})$
- Conditions:
 - ❖ yeast as catalyst
 - ❖ temperature $< 40^\circ\text{C}$
 - ❖ concentration of ethanol formed $\sim 12\text{-}14\%$
 - ❖ anaerobic conditions (no O_2)

If not, the enzymes in yeast denatures and fermentation stops.

If not, the ethanol produced oxidises to form ethanoic acid.

To differentiate between alkanes and alkenes

chemical test	: Deliver both alkane and alkene into <u>aqueous bromine</u> .
result with alkane	: <u>No visible reaction</u>. [Alkane is <i>saturated</i>, hence undergoes a <i>substitution</i> reaction slowly with aqueous bromine, in the presence of sunlight.] $ \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{H}-\text{C}-\text{C}-\text{H} \\ & \\ \text{H} & \text{H} \end{array} + \text{Br}_2 \rightarrow \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{H}-\text{C}-\text{C}-\text{Br} \\ & \\ \text{H} & \text{H} \end{array} + \text{H}-\text{Br} $
result with alkene	: The <u>red-brown</u> aqueous bromine turns <u>colourless</u> immediately. [Alkene is <i>unsaturated</i>, hence undergoes an <i>addition</i> reaction rapidly with aqueous bromine.] $ \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{H}-\text{C}=\text{C}-\text{H} \end{array} + \text{Br}_2 \rightarrow \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{H}-\text{C}-\text{C}-\text{H} \\ & \\ \text{Br} & \text{Br} \end{array} $

Example 4: You are to distinguish between two gaseous compounds.

One is propane, an alkane, and the other is propene, an alkene.

- Describe the method you could use and the observations you would expect to make.
- Write a balanced chemical equation for any reaction that takes place in your method.

To differentiate between alcohols and carboxylic acids

Method 1

chemical test	: Add both alcohol and carboxylic acid to solid <u>sodium carbonate</u> .
result with alcohols	: <u>No visible reaction</u> .
result with carboxylic acids	: <u>Effervescence</u> observed. Gas evolved gives <u>white precipitate</u> with limewater. $ \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ 2 \text{H}-\text{C}-\text{C}-\text{O}-\text{H} + \text{Na}_2\text{CO}_3 \\ \\ \text{H} \end{array} $ $ \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \rightarrow 2 \text{H}-\text{C}-\text{C}-\text{O}-\text{Na} + \text{H}_2\text{O} + \text{CO}_2 \\ \\ \text{H} \end{array} $

Method 2

chemical test	: Add both alcohol and carboxylic acid to aqueous <u>acidified potassium manganate(VII)</u> .
result with alcohols	: The <u>purple</u> solution turns <u>colourless</u> . $ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} + \text{H}^+/\text{KMnO}_4 \\ \quad \\ \text{H} \quad \text{H} \end{array} $ $ \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \rightarrow \text{H}-\text{C}-\text{C}-\text{O}-\text{H} + \text{Mn}^{2+} + \text{product} \\ \\ \text{H} \end{array} $
result with carboxylic acids	: <u>No visible reaction</u> .

Example 5: Ethanoic acid and ethanol can be distinguished by chemical tests.

Describe **two** of these chemical tests and the observations that allow you to make the distinctions.

11.4 Polymers

- Methods of Recycling Plastics

- ❖ Method 1: Physical Method (Mechanical Recycling)

- Using this method, the chemical composition of the recycled plastic remains unchanged, i.e. if waste poly(ethene) bottles are collected, the recycled plastic products will still be made of poly(ethene).
 - After pre-treatment, small pieces of plastics such as poly(ethene) are melted, cooled, pulled into long, thin strands, and cut into pellets. These pellets can then be made into new products.

- ❖ Method 2: Chemical Method

- The chemical method of recycling plastics converts plastic waste into different raw materials through chemical reactions such as cracking.
 - Plastic waste, such as poly(ethene) waste, can undergo cracking to form short chains of alkanes and alkenes. The short-chain alkanes can be used as fuel, while the short-chain alkenes can be used to make other useful chemicals.

- Social, Economic and Environmental Issues of Recycling Plastics

- ❖ Different ways to minimise plastic waste: **reduce – reuse – recycle**

- ❖ Environmental issues of recycling plastics

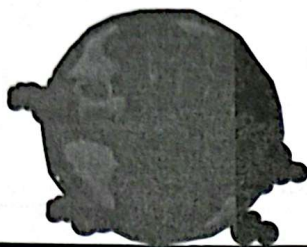
- The recycling of plastics may cause environmental issues.
For e.g.: If the wastewater containing harmful substances is not treated before it enters the oceans, lakes or rivers, it can cause water pollution.

- ❖ Economic issues of recycling plastics

- Recycling plastic waste can be expensive, costing much more than disposal in landfills or incineration.
For e.g.: There is the cost of transporting the waste to the processing plant, cost of sorting and cleaning the waste, and carrying out either physical or chemical recycling processes, cost of manpower, machines and energy.
 - It may not be worthwhile or economically viable to recycle plastics if the cost of recycling plastic waste is higher than the value of the recycled plastics.

- ❖ Social issues of recycling plastics

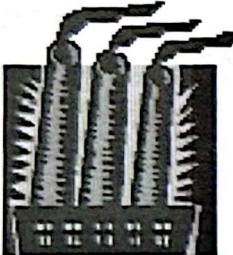
- In Singapore, many people find it more convenient to throw all their waste away. As a result, recyclable plastics are disposed of and are not recycled.
 - Many people do not know the proper way to use the recycling bins. This will affect the recycling process of plastic waste.
 - It is hard to get people to form new habits. Hence, time and effort are needed for communities to adopt recycling as a lifestyle.



Maintaining Air Quality

DRY AIR

- ❖ Air is a mixture of elements (78% N₂, 21% O₂, Ar) and compounds (CO₂, H₂O).
- ❖ The components of air can be obtained by the fractional distillation of liquefied air, as they are miscible liquids with different boiling points.

Air Pollutants	Source	Occurrence	Effects	Prevention
1. Carbon monoxide [neutral] $C + O_2 \rightarrow CO_2$ $2C + O_2 \rightarrow 2CO$ 	❖ <u>Incomplete combustion</u> of <u>fuel</u> in <u>cars</u> . ❖ Incomplete combustion of <u>carbon-containing</u> substances.	❖ On streets crowded with vehicles ❖ Around fires with poor ventilation (e.g. kitchens, bathrooms and gas heaters)	Effects on health: TOXIC!! ❖ In <u>low</u> concentrations: causes dizziness, headaches, fatigue and breathing difficulties ❖ In <u>high</u> concentrations: causes death <u>CO combines with haemoglobin, causing the body to be starved of oxygen, leading to death.</u>	❖ Fitting <u>catalytic converters</u> onto cars (CO is <u>oxidised</u> to CO ₂)
2. Sulfur dioxide [acidic] $S + O_2 \rightarrow SO_2$ 	❖ <u>Combustion</u> of <u>fossil fuels</u> (i.e. coal, crude oil and natural gas) in industries and <u>power stations</u> . ❖ Volcanoes		Effects on health: ❖ Causes irritation to eyes and breathing difficulties ❖ When inhaled, SO ₂ reacts with water to form H ₂ SO ₃ , which corrodes the body's internal organs [SO ₂ + H ₂ O → H ₂ SO ₃] Effect on the environment: ❖ When released in air, SO ₂ reacts with oxygen to form SO ₃ , which dissolves in water to form acid rain $2SO_2 + O_2 \rightarrow 2SO_3$ $SO_3 + H_2O \rightarrow H_2SO_4$	❖ Removal of sulfur from fossil fuels before they are burnt (expensive) ❖ Flue Gas <u>Desulfurisation</u> (a) Calcium carbonate $[CaCO_3 + SO_2 \rightarrow CaSO_3 + CO_2]$ $2CaSO_3 + O_2 \rightarrow 2CaSO_4$ (b) Calcium oxide $[CaO + SO_2 \rightarrow CaSO_3]$

3. Nitrogen oxides $\text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}$ [neutral] $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$ [acidic] 	➤ <u>Combustion of fuels</u> in <u>motor vehicles</u> (internal combustion engines) at <u>high temperature</u>. ➤ Lightning activity	➤ Occurs only around places of high temperature, which causes the N_2 in the air to be oxidised	<u>Effect on health:</u> ➤ Causes irritation to eyes, breathing difficulties, and lung damage <u>Effects on the environment:</u> ➤ Reacts with other pollutants damaging the ozone layer ➤ Readily soluble in water to form acid rain	➤ Fitting <u>catalytic converters</u> onto cars (NO is reduced to N_2) Both SO_2 and NO_2 cause the formation of <u>acid rain</u>, which results in the <u>corrosion of limestone buildings / metal structures</u> (e.g. steel bridges).
4. Chlorofluorocarbons [CFCs] 	➤ Aerosol propellants, refrigerator coolants, foam plastic		➤ Destroys ozone layer - The ozone layer is a protective layer that filters away UV rays - Exposure to UV radiation causes skin cancer, genetic mutations and eye damage	➤ Ban on the use of CFCs
5. Lead compounds 	➤ Paints; dried paint chips ➤ Burning of <i>leaded</i> gasoline ➤ Dust and fumes released by industries using lead		➤ <u>Low level:</u> mild diarrhea, irritability, weakness, abdominal pains, vomiting, constipation, headaches ➤ <u>Higher level:</u> coma, convulsions, paralysis, brain damage, mental retardation ➤ <u>High level:</u> death	➤ Use unleaded petrol

❖ some other common atmospheric pollutants: methane, ozone, unburned hydrocarbons