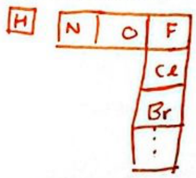
Upper Secondary Pure Chemistry

1. Formula of ions

- Group 1 ions : + (Na^+, K^+)
- Group 2 ions : 2+ (Mg^{2+}, Ca^{2+})
- Group 13 ions : 3+ (Al^{3+})
- Group 15 ions : 3- (N^{3-})
- Group 16 ions : 2- (O^{2-})
- Group 17 ions : - (F^-, Cl^-, I^-, Br^-)
- Roman numerals ~ only positive charge (iron (II) → Fe^{2+})

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

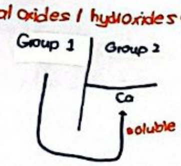
2. Diatomic formulas (X_2)



⇒ H_2, N_2, O_2, F_2 . . . etc

3. Solubility table.

- all nitrates are soluble
- all chlorides are soluble except $AgCl$ and $PbCl_2$ (iodide solubility same)
- all sulfates are soluble except $BaSO_4$, $CaSO_4$, $PbSO_4$
- all carbonates insoluble except group 1 and ammonium
- all metal oxides / hydroxides are insoluble except



4. Ionic Equations

(a) For all acid + alkali reactions

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

(b) For precipitate, all aq but 1 s product

Write formula of solid on RHS, work backwards to see what ions on LHS

5. Colours of compounds and QA cations, anions test.

- test for cation table (esp white precipitate)
- test for anion table
-

Substance & State	Colour
copper metal, Cu (s)	pink
copper(II) oxide, CuO (s)	black
copper(II) carbonate, $CuCO_3$ (s)	green
copper (II) nitrate, $Cu(NO_3)_2$ (aq)	blue
copper (II) sulfate, $CuSO_4$ (aq)	blue
copper (II) hydroxide, $Cu(OH)_2$ (s)	blue
copper (II) chloride, $CuCl_2$ (aq)	green
iron metal, Fe (s)	grey
iron (II) hydroxide, $Fe(OH)_2$ (s)	dirty green
iron (II) solutions	(pale) green
Iron (III) hydroxide, $Fe(OH)_3$	reddish brown
Iron (III) solutions	Yellow
Flourine gas, F_2 (g)	green
Chlorine gas, Cl_2 (g)	greenish-yellow
Bromine liquid, Br_2 (l)	reddish-brown
Iodine solid, I_2 (s)	(purplish-)black
Iodine aqueous, I_2 (aq)	brown
Iodine vapour, I_2 (g)	violet / purple

Colour	Example	Remarks
Colourless gas	Oxygen O_2 (g) Hydrogen H_2 (g) Nitrogen N_2 (g) Water vapour H_2O (g) Ammonia NH_3 (g) Sulfur dioxide SO_2 (g), etc	[except halogens: F_2 (g)—pale yellow, Cl_2 (g)—greenish-yellow, Br_2 (g) —reddish-brown, I_2 (g)—purple] [except NO_2 (g)—brown]
Colourless Solution (aq)	All Group I compounds ($NaCl$, Na_2CO_3, K_2SO_4, KNO_3, KI, etc) Group II compounds that are soluble in water ($MgSO_4$, $Mg(NO_3)_2$, $CaCl_2$, etc) Group III compounds that are soluble in water ($Al_2(SO_4)_3$, $Al(NO_3)_3$, etc)	Compounds of <u>all main group</u> metals (i.e. Group I, II, III)
	Silver compounds that are soluble Zinc compounds that are soluble	Compounds of <u>some transition metals</u> (others: coloured compounds)
	All ammonium compounds	
	All acids solutions	
	All alkalis solutions	
	Hydrogen peroxide H_2O_2	
	Mn^{2+} (aq)	Redox reaction (purple → colourless)
White (solid)	All Group I, II, III compounds	
	Most silver compounds ($AgCl$, Ag_2O, Ag_2CO_3, etc) Most zinc compounds ($ZnCl_2$, $ZnCO_3$, $ZnSO_4$, $Zn(OH)_2$, etc)	[except AgI (s)—pale yellow] [except ZnO(s)—yellow when hot, white when cold]
Black	Most transition metal oxides: Copper(II) oxide CuO (s) Iron(II) oxide FeO (s) Iron(III) oxide Fe_3O_4 (s) Manganese(IV) oxide MnO_2 (s)	Note: Fe_2O_3 – part of iron ore, haematite MnO_2 – catalyst in decomposition of H_2O_2
	Some non-metal elements: Iodine crystals I_2 (s) Astatine crystals At_2 (s) Carbon, graphite, coke, soot, C (s)	
Grey/ Silver	All metal elements: K, Na, Ca, Mg, Al, Zn, Fe, Sn, Pb, Ag, etc	[except Au (s)—gold] [except Cu (s)—pink (fresh), reddish-brown]
Blue	Copper(II) sulfate $CuSO_4$ (s, aq) Copper(II) nitrate $Cu(NO_3)_2$ (s, aq) Copper(II) chloride $CuCl_2$ (s, aq)—blue-green Copper(II) hydroxide $Cu(OH)_2$ (s)—cation test	Most copper(II) compounds
	Anhydrous cobalt(II) chloride $CoCl_2$ (s)	Test for water, turns pink when moist
Green	Iron(II) chloride $FeCl_2$ (s, aq) Iron(II) sulfate $FeSO_4$ (s, aq) Iron(II) hydroxide $Fe(OH)_2$ (s) dirty-green, cation test	Most iron(II) compounds
	Nickel(II) chloride $NiCl_2$ (s, aq) Nickel(II) sulfate $NiSO_4$ (s, aq), etc	Most nickel(II) compounds
	Copper(II) carbonate $CuCO_3$ (s)	
Greenish Yellow	Chlorine Cl_2 (g)	
Yellow	Lead(II) iodide PbI_2 (s) Silver iodide AgI (s) anion test	
Orange/Yellow	Iron(III) chloride $FeCl_3$ (s, aq) Iron(III) sulfate $Fe_2(SO_4)_3$ (s, aq)	Most iron(III) compound, Orange/yellow colour depends on concentration
Reddish Brown	Copper Cu (s)	Pink when fresh
	Iron (III) hydroxide $Fe(OH)_3$ (s) cation test Rust, hydrated iron (III) oxide $Fe_2O_3.xH_2O$ (s)	
	Bromine Br_2 (aq, *l*)	
Brown	Iodine solution I_2 (aq)	
	Nitrogen dioxide NO_2 (g)	
Purple	Potassium manganate(VII) $KMnO_4$ (s, aq)	Oxidizing agent

★Metals are usually solid @ r.t.p.★
★Monatomic elements = Left of dark black line★

The Periodic Table of Elements

Key: proton (atomic) number / atomic symbol / name / relative atomic mass

Charge: +1 Group 1	+2 2	3	4	5	6	7	8	9	10	11	12	+3 13	14	-3 15	-2 16	-1 17	0 18
							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 –	113 Nh Nihonium –	114 Fl flerovium –	115 Mc Moscovium –	116 Lv livermorium –	117 Ts Tennessine –	118 Og Oganesson –

	57–71														
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.).

- Group 1: Alkali metals; 1 electron in valence shell. Group 2: 2 electrons in valence shell. Group 13: 3 electrons in valence shell; Group 14: 4 electrons in valence shell; Group 15: 5 electrons in valence shell; Group 16: 6 electrons in valence shell; Group 17: 7 electrons in valence shell; Group 18: 8 electrons in valence shell – Noble gas configuration.
- Group 0: All single – Noble gases (unreactive).
- Group 17: Halogens – lowest m.p. and b.p., lightest colour; X_2.
- H (1) → H_2.
- Transition metals: high m.p. and high density; variable oxidation states; colour compounds; catalysts.
- Highlight: oxides and hydroxides that are soluble in water.
- Electron configurations: Li (2,1), Be (2,2), Na (2,8,1), Mg (2,8,2), K (2,8,8,1), Ca (2,8,8,2), N (2,5), F (2,7), P (2,8,5), Cl (2,8,7).
- Number of valence electron shells increases down the group; 1 electron shell, 2 electron shells, 3 electron shells.
- Top of group 1: highest m.p., lowest density, least reactive, lowest metallic properties, highest non-metallic properties. Bottom: lowest non-metallic properties, highest metallic properties, most reactive; down the group: lowest m.p., highest density.
- Group I properties: soft, can be cut easily; low m.p. and low b.p.; low density.
- metallic ← → ; basic oxide; amphoteric oxide; metalloid; non-metallic; acidic oxide.
- Group 18 properties: monoatomic, non-metallic, colourless gas at r.t.p., low m.p. and b.p., insoluble in water, unreactive.
- Group 17 properties: low m.p. and b.p., coloured; highest m.p. and b.p., darkest colour (bottom).
- Across the period: number of valence electrons increases; left: highest metallic properties, lowest non-metallic properties; right: lowest metallic properties, highest non-metallic properties.

Solubility table

① All Nitrates, NO_3^- are soluble
② All Chlorides, Cl^- are soluble except Silver chloride ($AgCl$) and Lead (II) chloride ($PbCl_2$)
③ All Sulfates, SO_4^{2-} are soluble except calcium sulfate ($CaSO_4$), Lead (II) sulfate ($PbSO_4$) and Barium Sulfate ($BaSO_4$)
④ All Carbonates, CO_3^{2-} are insoluble except group 1 carbonates and Ammonium carbonate ($(NH_4)_2CO_3$)
⑤ All Metal oxides / hydroxides, O^{2-} / OH^- are insoluble except group 1 oxides / hydroxides and bottom half of group 2 oxides / hydroxides (Ca ↓)

Balancing chemical equation

① Balance METAL atoms
② Balance NON-METALS atoms, excluding H and O atoms
③ Balance H atoms
④ Balance O atoms → adjust O_2 if possible

} Double check LHS and RHS to ensure balance

Balancing ionic equation

① Identify compounds which are in 'aq' on both LHS and RHS of the chemical equation
② Cancel the same ions in (aq) that appear on both LHS and RHS → spectator ions
③ Form ionic equation → remaining ions in aq compound that is uncancelled (remember to put charge)

Ions to memorise

① Silver ion → Ag^+
② Zinc ion → Zn^{2+}
③ Ammonium ion → NH_4^+
④ Carbonate → CO_3^{2-}
⑤ Sulfate → SO_4^{2-}
⑥ Nitrate → NO_3^-
⑦ Hydroxide → OH^-

Acids/Alkalis

$HCl (aq)$, $HNO_3 (aq)$, $H_2SO_4 (aq)$, $NaOH (aq)$, $KOH (aq)$

↑ hydrochloric acid, ↑ nitric acid, ↑ sulfuric acid, ↑ sodium hydroxide, ↑ potassium hydroxide

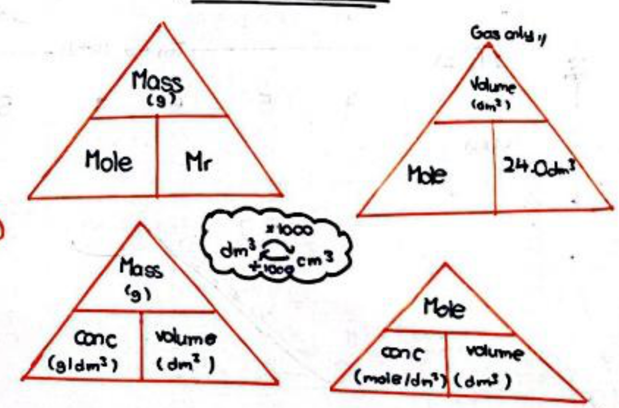
Formulas to memorise

① Water / Steam → H_2O
② Carbon Dioxide → CO_2
③ Carbon Monoxide → CO
④ Ammonia → NH_3
⑤ Methane → CH_4
⑥ Ethane → C_2H_6
⑦ Ethene → C_2H_4

Reactivity Series

Please	Potassium (K)
Stop	Sodium (Na)
Calling	Calcium (Ca)
Me	Magnesium (Mg)
A	(Aluminium) (Al)
Cute	Carbon (C)
Zebra	Zinc (Zn)
Instead	Iron (Fe)
Learn	Lead (Pb)
How	(Hydrogen) (H)
Copper	Copper (Cu)
Strengthen	Silver (Ag)
Gold	Gold (Au)
	Platinum (Pt)

Mole formulas



QA notes: will be given during practical

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.
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.

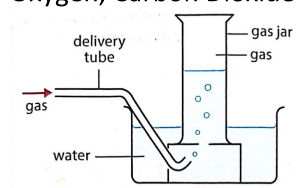
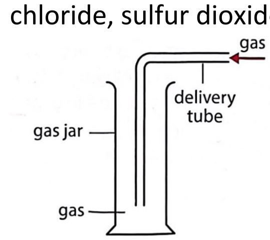
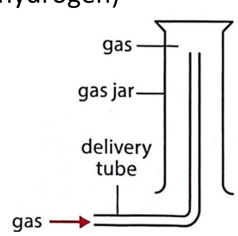
Test for aqueous cations

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

Chapter 1: Experimental Chemistry

Apparatus	Pipette	Measures accurate fixed volumes (e.g. 10.0cm^3 or 25.0cm^3) – 1 d.p.
	Volumetric Flask	Measures accurate fixed volumes that are larger (e.g. 100 cm^3 or 250 cm^3)
	Measuring Cylinder	Measures a range of volumes to the nearest 0.5 cm^3 (e.g. 31.5 cm^3 or 23.0 cm^3)
	Burette	Measures a range of volumes to the nearest 0.05 cm^3 (e.g. 31.55 cm^3 or 23.00 cm^3) – most accurate (when 2 dp value, usually burette)
Collection Method	Water Displacement	Gas must be insoluble/slightly soluble in water (e.g. Hydrogen, Oxygen, Carbon Dioxide) 
	Downward Delivery	Gas must be denser than air – Mr/Ar > 28 (e.g. chlorine, hydrogen chloride, sulfur dioxide) 
	Upward Delivery	Gas must be less dense than air – Mr/Ar < 28 (e.g. Ammonia, helium, hydrogen) 
Drying Agent	Concentrated Sulfuric Acid	Most gases except basic gas (e.g. ammonia)
	Quicklime (Calcium Oxide)	Gases (ammonia can be dried here) except acidic gases (e.g. Sulfur Dioxide, Carbon Dioxide, Nitrogen Dioxide, Chlorine, Hydrogen Chlorine)
	Fused Calcium Chloride	Gases except those who react with Calcium Chloride (e.g. Ammonia) → must be freshly heated before use

Separation of Mixtures	Magnetic Attraction	A magnet can be used to separate magnetic solids from non-magnetic solids (Solid – Solid)
	Sieving	A sieve can be used to separate solids of different particle sizes (Solid – Solid)
	Suitable Solvents	A suitable solvent can be used to separate solid-solid mixtures in which only one of the solids is soluble in the solvent (Solid – Solid)
	Sublimation	Sublimation can be used to separate a substance that changes from the solid to gaseous state directly (Solid – Solid) (e.g. pure iodine, dry ice…)
	Filtration	Filtration can be used to separate insoluble solids from liquids (Solid – Liquid)
	Evaporation to Dryness	Evaporation to dryness is used to separate a dissolved solid from its solvent by heating the mixture until all the solvent has vapourised (Solid – Liquid) (O levels, only have NaCl)
	Crystallisation	Crystallisation is used to obtain a pure solid from its saturated solution. A saturated solution is one in which no more solute can be dissolved (Solid – Liquid) This method is used instead of evaporation to dryness to prevent the crystals from losing its water of crystallization and becoming anhydrous
	Simple distillation	Simple distillation is used to separate a pure solvent from a solution (Solid – Liquid) Why bulb of thermometer at entrance of condenser? → To record the boiling point of vapour collected Why water in this direction (condenser)? → To ensure condenser is fully filled with water and allows efficient cooling of vapour
	Separating Funnel	A separating funnel is used to separate immiscible liquids (Liquid – Liquid) (immiscible = different density)

<table>
<tr><td rowspan="2"></td><td>Chromatography</td><td>Chromatography is used to separate a mixture of substances which have different solubilities in a given solvent

Why start line in pencil? → Graphite is insoluble in all solvents and will not be separated together with the samples

Why start line not in ink? → Ink is a mixture and dyes will be separated along with the sample

R_f value = $\frac{distance\ travelled\ by\ the\ substance}{distance\ travelled\ by\ the\ solvent} \leq 1$ (no units, 2d.p)</td></tr>
<tr><td>Fractional distillation</td><td>Fractional distillation is used to separate miscible liquids with different boiling points</td></tr>
<tr><td colspan="2">Purity of a substance</td><td>A pure substance has a specific melting and/or boiling point under fixed conditions
(impure substances melt/boil at a range of temperatures)</td></tr>
</table>

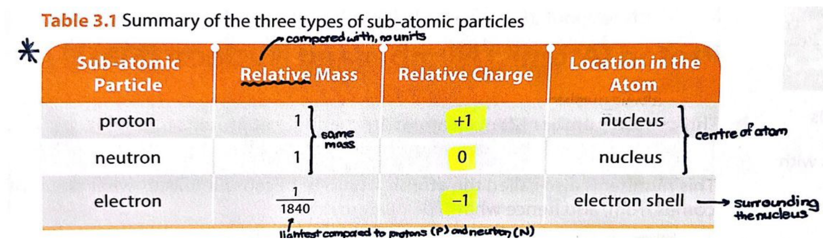
Chapter 2: Kinetic Particle Theory

<table>
<tr><td colspan="2">Kinetic Particle Theory</td><td>The kinetic particle theory states that all matter is made up of tiny particles and these particles are in constant random motion</td></tr>
<tr><td colspan="2">Solid, Liquid and Gas</td><td>Table 2.1 Summary of the differences between particles in solids, liquids and gases.
<table>
<tr><th>State of Matter</th><th>Solid</th><th>Liquid</th><th>Gas</th></tr>
<tr><td>Particle Arrangement</td><td>very closely packed in an orderly manner</td><td>closely packed in a disorderly manner</td><td>very far apart in a disorderly manner</td></tr>
<tr><td>Attractive Forces Between Particles</td><td>very strong</td><td>less strong</td><td>very weak</td></tr>
<tr><td>Kinetic Energy of Particles</td><td>very low</td><td>low</td><td>high</td></tr>
<tr><td>Particle Movement</td><td>vibrate and rotate about fixed positions</td><td>slide past one another freely throughout the liquid</td><td>move quickly and randomly in any direction</td></tr>
<tr><td>Shape</td><td>definite</td><td>indefinite</td><td>indefinite</td></tr>
<tr><td>Volume</td><td>definite</td><td>definite</td><td>indefinite</td></tr>
<tr><td>Compressibility</td><td>no</td><td>no</td><td>yes</td></tr>
</table>
Why temperature remains constants in change of state? → The energy provided is equal to the energy needed to overcome the forces of attraction between particles

What happens when impurities are added? → The boiling point increases and melting point decreases (both occurs at a range of temperatures) (substance at liquid state at wider temp range)</td></tr>
<tr><td rowspan="2">Diffusion</td><td>Definition</td><td>Diffusion is the <u>net</u> movement of particles from a region of higher concentration to a region of lower concentration</td></tr>
<tr><td>Conditions that affect R.O.D</td><td>Temperature → higher temp, higher KE, movement increase, higher ROD
Mr → higher Mr = heavier, slower ROD</td></tr>
</table>

Chapter 3: Atomic Structure

<table>
<tr><td rowspan="3">Atoms</td><td>Definition</td><td>An atom is the smallest particle that can still have the chemical characteristics of an element</td></tr>
<tr><td>Charges</td><td>Atoms are electrically neutral (charge = 0)

Number of protons in an atom = number of electrons in that atom</td></tr>
<tr><td>Sub-atomic particles</td><td>
Table 3.1 Summary of the three types of sub-atomic particles
<table>
<tr><th>Sub-atomic Particle</th><th>Relative Mass</th><th>Relative Charge</th><th>Location in the Atom</th></tr>
<tr><td>proton</td><td>1</td><td>+1</td><td>nucleus</td></tr>
<tr><td>neutron</td><td>1</td><td>0</td><td>nucleus</td></tr>
<tr><td>electron</td><td>$\frac{1}{1840}$</td><td>−1</td><td>electron shell</td></tr>
</table>
</td></tr>
<tr><td rowspan="2">Sub-atomic particles</td><td>Proton number</td><td>The Proton Number of an atom is the number of protons in its nucleus of an atom.</td></tr>
<tr><td>Nucleon number</td><td>The nucleon number is the total number of protons and neutrons in the nucleus of an atom</td></tr>
<tr><td colspan="2">Ions</td><td>An ion is the particle formed when an atom or a group of atoms gains or loses electron(s), but the number of protons and neutrons remains the same</td></tr>
<tr><td colspan="2">Isotopes</td><td>Isotopes are <u>atoms</u> of the same element that have the same proton number but different nucleon number. This means they have different number of neutrons.</td></tr>
<tr><td>Electrons</td><td>Electronic configuration</td><td>2 , 8 , 8 , 18 …

outermost electron shell = valence shell
electrons in outermost electron = valence electrons</td></tr>
</table>

Chapter 4: Chemical Bonding

<table>
<tr><td rowspan="2">Ions</td><td>Positive ions (cations)</td><td>Positive ions (cations) have a net positive charge and usually have a noble gas electronic configuration</td></tr>
<tr><td>Negative ions (anions)</td><td>Negative ions (anions) have a net negative charge and have a noble gas electronic configuration</td></tr>
<tr><td rowspan="2">Ionic bonding</td><td>Ionic bond</td><td>An ionic bond is the mutual electrostatic attraction between ions of opposite charges (cation and anion)</td></tr>
<tr><td>Ionic structures</td><td>A giant ionic crystal lattice is a three-dimensional structure of alternating positive and negative ions along the x, y and z axis</td></tr>
<tr><td>Covalent bonding</td><td>Valency (Covalent bond)</td><td>Valency refers to the number of electron(s) that must be lost, gained or shared in order for the atom to attain a noble gas electronic configuration</td></tr>
<tr><td>Metallic bonding</td><td>Metallic bond</td><td>The metallic bond is the mutual electrostatic attraction between positively charged ions in a metal and the “sea of delocalised electrons) (cation and electron)</td></tr>
</table>

Chapter 5: Structure and Properties of material

<table>
<tr><td colspan="2">EMC</td><td>
Table 5.1 Comparison between elements, compounds and mixtures
<table>
<tr><th></th><th>Element</th><th>Compound</th><th>Mixture</th></tr>
<tr><td>What Is It Made Of?</td><td>only one element</td><td>two or more elements that are chemically combined</td><td>two or more elements and/ or compounds that are not chemically combined</td></tr>
<tr><td>How Is It Formed?</td><td>mostly naturally occurring</td><td>from a chemical reaction</td><td>usually from physical mixing</td></tr>
<tr><td>What Is the Ratio of Its Constituents?</td><td>–</td><td>fixed ratio</td><td>no fixed ratio</td></tr>
<tr><td>What Are Its Properties Like?</td><td>–</td><td>has different properties from its constituent elements</td><td>usually has similar properties to its constituent substances</td></tr>
<tr><td>Melting and Boiling Points</td><td>fixed</td><td>fixed</td><td>melt and boil over a range of temperatures</td></tr>
</table>
</td></tr>
<tr><td>Properties</td><td>Ionic</td><td>- Giant ionic crystal lattice structure → strong electrostatic forces of attraction
- High Melting point, High Boiling Point → overcome strong electrostatic forces of attraction between <u>(cation)</u> and <u>(anion)</u>
- Hard
- Soluble in water, insoluble in organic solvents
- Conducts electricity in molten/liquid and aqueous states (mobile ions). Does not conduct electricity in solid (held in fixed position)</td></tr>
</table>

	Simple Covalent / Simple molecular	- Covalent bonds → weak intermolecular forces of attraction - Low melting point, Low boiling point → **overcome** weak intermolecular forces of attraction between ____ molecules - Insoluble in water, Soluble in organic solvents - Does not conduct electricity in any state
	Giant covalent	- Tetrahedral arrangement (1 carbon atom is covalently bonded to four other carbon atoms) – Diamond, layered arrangement (1 carbon atom is covalently bonded to three other carbon atoms) – Graphite → **break** strong covalent bonds between ____ atoms - Diamond is hard, Graphite is soft and slippery - Both have high boiling and melting point - Insoluble in water and insoluble in organic solvents
	Macromolecules	- Many covalent molecules joined into chains to form a larger molecule → **overcome** weak intermolecular forces of attraction - High melting and High boiling point (melts over a range of temperatures) - Insoluble in water and soluble in organic solvents - Not able to conduct electricity in any states
	Metals and Alloys	- Giant metallic lattice → lattice of strong metallic bonds - Pure metals are malleable and ductile - Alloys are harder and stronger than pure metals - High melting and boiling point - Both are good electrical conductors in any state → "sea of delocalized electrons"
Allotropes		Different forms of the same element with different structural arrangement of atoms
Alloys		A mixture of a metal with one or more other elements (other elements can be non-metal)

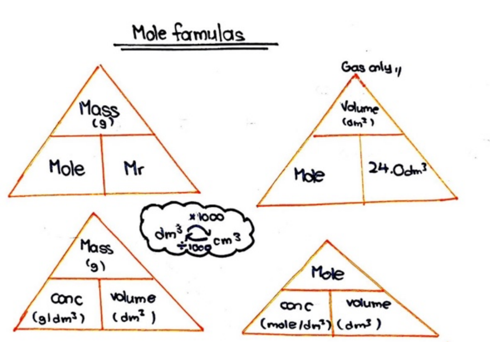
Chapter 6: Chemical Formulae and Equations

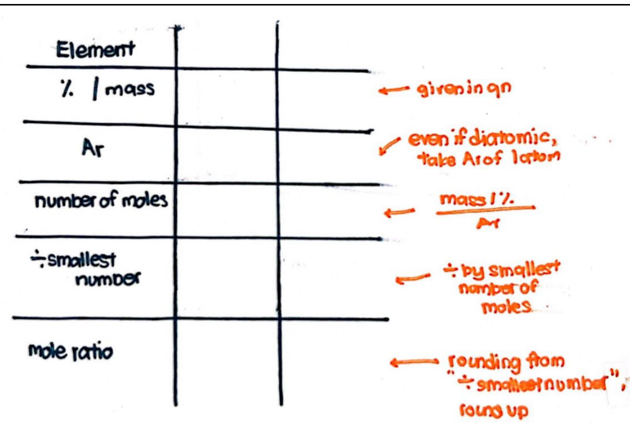
Balancing chemical equations	1. Balance Metal atoms 2. Balance Non-Metal atoms, excluding H and O atoms 3. Balance H atoms 4. Balance O atoms → adjust O_2 if possible
Forming Ionic equations	1. Identify compounds which are in "aq" on both LHS and RHS of the chemical equation 2. Cancel the same ions in (aq) that appear on both LHS and RHS → spectator ions 3. From ionic equation → remaining ions in aq compound that is uncancelled

Chapter 7: Mole concept and Stoichiometry

<table>
<tr><td rowspan="3">Mr and Ar

(definition not important)</td><td>Relative Atomic Mass</td><td>The relative atomic mass (A_r) of an element is the average mass of one atom of that element relative to $\frac{1}{12}$ the mass of an atom of carbon – 12.</td></tr>
<tr><td>Relative Molecular Mass</td><td>The relative molecular mass (M_r) of a molecular substance is the average mass of one molecule of that substance to $\frac{1}{12}$ the mass of an atom of carbon – 12.</td></tr>
<tr><td>Relative Formula Mass</td><td>The relative formula mass (M_r) of an ionic compound is the average mass of one unit of that compound relative to $\frac{1}{12}$ the mass of an atom of carbon – 12.</td></tr>
<tr><td rowspan="4">Mole calculations</td><td>Mole</td><td>One mole of any substance will always contain 6.02 x 10^{23} particles of that substance. The particles can be atoms, molecules, ions and even sub-atomic particles such as electrons.</td></tr>
<tr><td>Molar mass</td><td>One mole of any substance has a mass equal to its relative atomic mass (A_r), relative molecular mass or relative formula mass (M_r) in grams.</td></tr>
<tr><td>Concentration</td><td>The concentration of a solution is the amount of a solute dissolved in a unit volume of the solvent</td></tr>
<tr><td>Formulas

(Most important)</td><td>
</td></tr>
<tr><td colspan="2">Empirical formula

(Most important)</td><td>
</td></tr>
</table>

Chapter 8: Acid and Bases

Acids	Definition	An acid is a substance that produces hydrogen ions, H^+, in aqueous solutions.
	Properties	1. Acids have a sour taste *2. Acids produce ions (H^+ ion and anion) when dissolved in water. These ions are mobile and can act as mobile charge carries to allow the resulting aqueous solution to conduct electricity. *3. Acids turn blue litmus paper red (litmus paper MOIST)
	Acid reactions	Metal + acid → salt + hydrogen gas (not all metals are reactive – copper, platinum, LEAD APPEARS UNREACTIVE) Base + Acid → salt + water (neutralisation) Carbonate + Acid → salt + water + carbon dioxide
	Strong vs weak acid	Strength refers to the extent of ionisation of an acid, when dissolved in water A strong acid is an acid that is completely ionized in an aqueous solution (e.g. HCl, H_2SO_4, HNO_3 – only three in syllabus) A weak acid is an acid that is only partially ionized in an aqueous solution (e.g. CH_3COOH, carboxylic acids)

Base	Definition	A base is any metal oxide or hydroxide. They contain either the oxide ion (O^{2-}) or the hydroxide ion (OH^-)
	Base reactions	Base + Acid → salt + water (neutralisation)
Alkalis	Definition	Alkalis are bases that are soluble in water that produces hydroxide ions, OH^-, in aqueous solutions.
	Strong vs weak alkali	Strength refers to the extent of ionisation of an alkali, when dissolved in water A strong alkali is an alkali that is completely ionised in an aqueous solution (e.g. NaOH) A weak alkali is an alkali that is only partially ionised in an aqueous solution (e.g. NH_3)
	Properties	1. Alkalis taste bitter 2. Alkalis feel slippery and soapy *3. Alkalis dissolve in water to form solutions that contain mobile ions which conduct electricity (have OH^- ions + cations) *4. Alkalis turn red litmus paper blue (litmus paper MOIST)
	Alkali reactions	Alkali + Acid → Salt + Water (14neutralization) Alkali + Ammonium salt → salt + water + ammonia gas Ammonia gas: If no water present in litmus paper, litmus paper will not change colour. Ammonia remains a molecule and no OH^- ions produced
	"Fertiliser question"	Qn: Fertilisers are usually ammonium salts. Farmers have been advised not to add fertilisers together with calcium oxide/hydroxide. Why? Ans: Calcium oxide/hydroxide reacts with ammonium salt to produce ammonia gas that escapes from the soil. Thus less nitrogen for plants to absorb through their roots.

Oxides	Basic Oxides (Metal + O)	Most metal oxides are basic oxides. They are insoluble in water and exist as solids in room temperature. They react with acids to form a salt and water.
	Amphoteric Oxides	Metallic oxides that react with both acid and bases to form salts and water. Zinc Oxide Aluminum Oxide Lead (II) Oxide } Only 3 in syllabus
	Acidic Oxides (Non-metal + O)	Most non-metal oxides are acidic oxides. They are able to dissolve in water to form acids. They do not react with acids but react with alkalis to form a salt and water.
	Neutral Oxides	Some non-metal oxides form oxides that show neither basic nor acidic properties. Water Carbon Monoxide Nitric oxide } Only 3 in syllabus

Chapter 9: Salts

Salt		A salt is an ionic compound that consists of a cation and an anion.
Salt preparation methods	Reaction of acid with an insoluble substance	This is when the salt is soluble in water and it is not a group 1/ammonium salt.
		Steps 1. Add excess (named solid reactant) to a fixed volume of (name acid). Stir. 2. Filter the mixture and collect (named salt) as the filtrate 3. Heat the filtrate until saturated 4. Let the saturated solution cool and crystallise 5. Filter out the crystals, was with a little cold distilled water and press dry with filter paper
	Titration	This is when the salt is soluble in water and it is a group 1/ammonium salt.
		Steps 1. Pipette 25.0cm^3 (can be any value, but reasonable) of (named alkali) into a conical flask and add in a few drops of suitable indicator 2. Add the (named acid) from the burette into the conical flask until the indicator changes colour. Note the volume of acid used. 3. Repeat the titration without the indicator and add in the pre-determined volume from step 2. 4. Heat the filtrate until saturated 5. Let the saturated solution cool and crystallise 6. Filter out the crystals, was with a little cold distilled water and press dry with filter paper
	Ionic Precipitation	This is when the salt is insoluble in water (follow solubility table, once insoluble confirm ionic precipitation)
		Steps 1. Add aqueous (named reactant 1) to aqueous (named reactant 2) 2. , stir 3. Filter the mixture and collect (named salt) as a residue 4. Wash the residue with cold distilled water and press dry with filter paper.