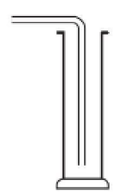
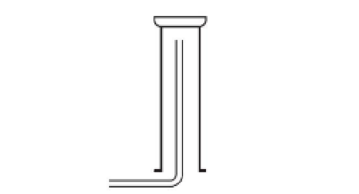
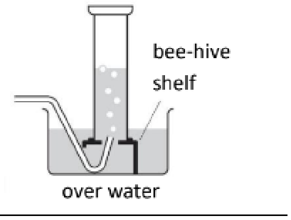
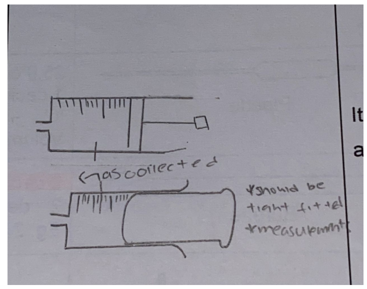
Units and measurements

Physical quantities	Apparatus	Accuracy
Temperature	Thermometer	1 decimal place Up to 0.5 degree Celsius (E.g.37.0 <to the accuracy of one d.p.>)
Mass	Electronic Balance	2 decimal place Up to 0.01g (e.g.2.56 to measure mass)
Time	Stopwatch	Nearest Second

- Volume (liquid)
  - ➔ Measuring cylinder (no decimal place, approximately nearest cm^3)
  - ➔ Pipette (25.0 cm^3, 1 decimal place always " 0 ", volume measured in multiples of 5)
  - ➔ Burette (2 decimal place, 2nd decimal is either " 0 " or " 5 ")
- Volume (gas)
  - ➔ Gas syringe (depends on the gas syringe)

- **Gas collection**

Method	**Experimental set up**	**Characteristics of gas collected**
Downward delivery		• Gas collected is more dense than air (hence the gas will sink soluble in water)
Upward delivery		• Gas collected is less dense than air (hence the gas will float soluble in water) ©wzkai

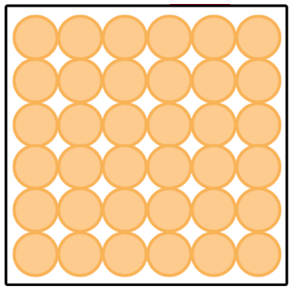
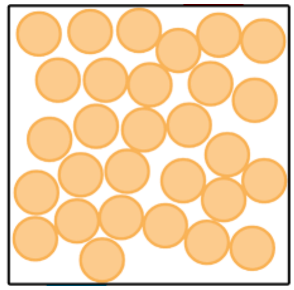
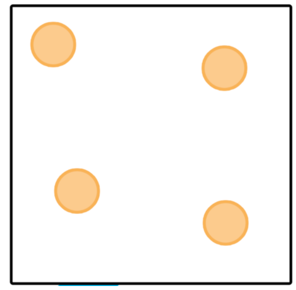
Displacement of water		• Gas collected is insoluble in water
Gas syringe		• It is suitable to collect and measure the volume of any type of gas

Kinetic Particle Theory

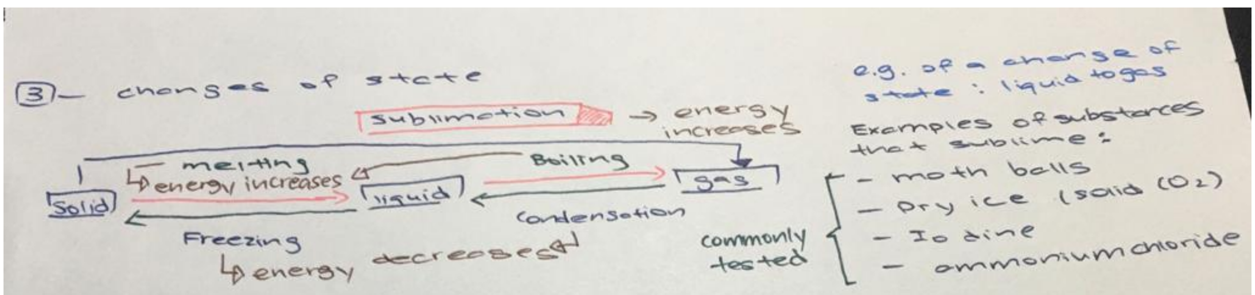
1. (a) All matter is made up of tiny discrete particles
   (b) The particles are always in constant random motion

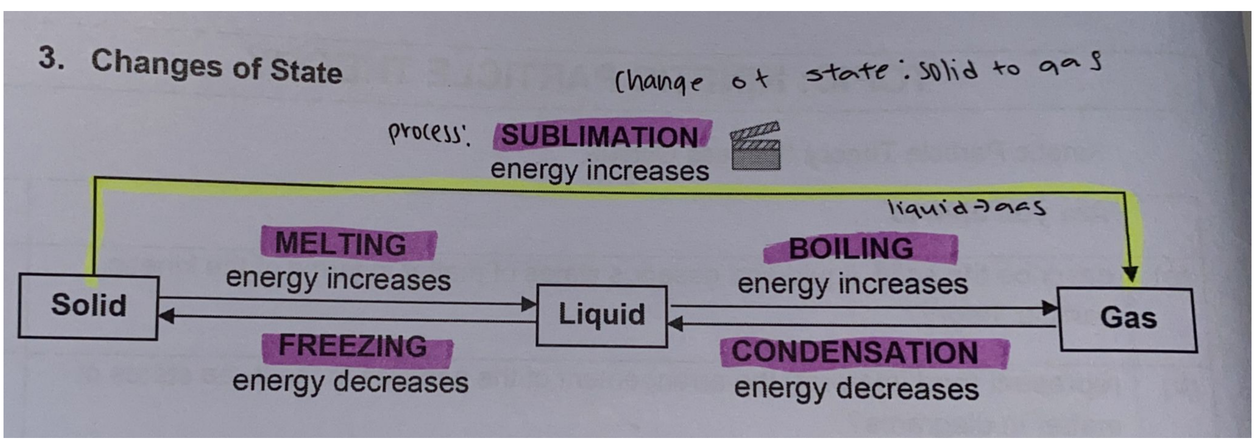
2. **State of matter**

	Solid	**Liquid**	**Gases**
Arrangement of particles	Very closely packed in an orderly manner	Closely packed but not in an orderly manner	Very far apart and in random arrangement
Movement of particles	Vibrate about in fixed position	Slide over each other	Move freely at high speed
Attractive forces between particles	Very strong	Strong	Very weak or almost negligible ©wzkai

Diagrams	 * 4 rows x 4 columns * similar sizes	 * should not have any pattern * should touch each other * should not be floating in the air (touching the ground)	 * very far apart * at least 6 particles

3. **Changes of State**



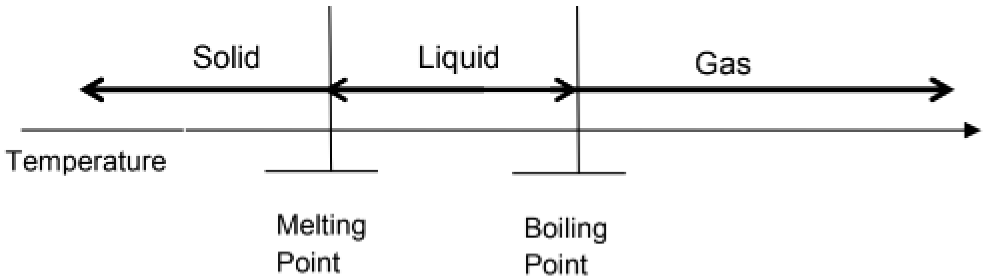


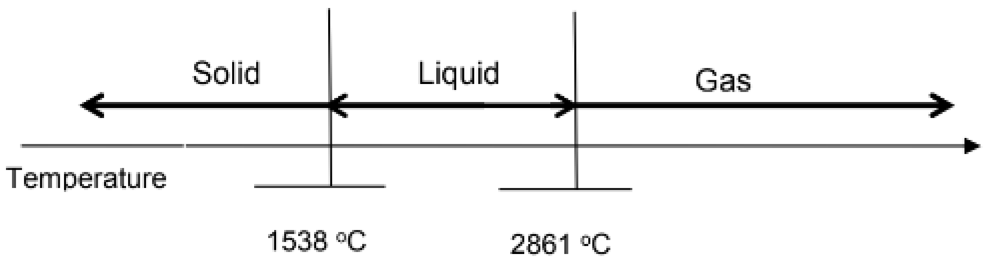
Examples of sublime are

- Moth balls
- Dry ice (solid CO2)
- Iodine

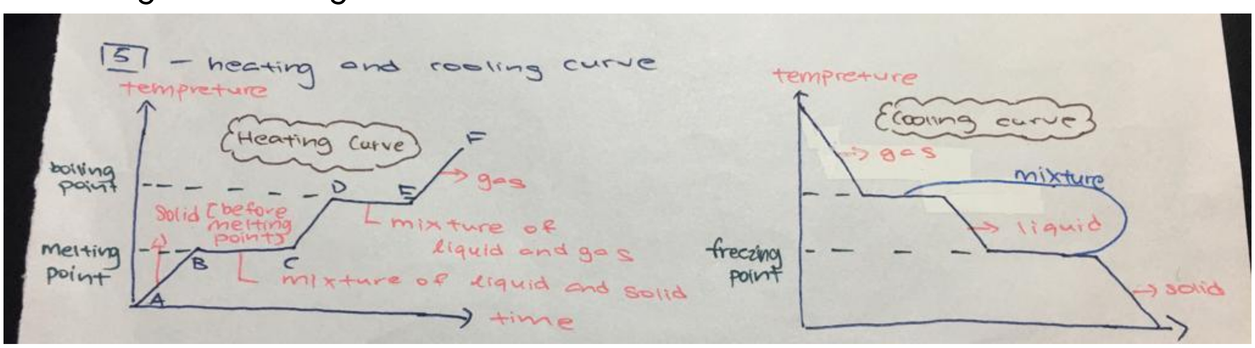
- Ammonium chloride (white solid)

4. **Determining the state of matter**





5. **Heating and cooling curve**



MELTING	CONDENSATION
A The particles are very closely packed in an orderly manner and vibrate at their fixed position.(Solid state)	A The particles are randomly arranged and move freely at high speed. (gaseous state)
A-B As the solid is heated, the particles gain kinetic energy and start to vibrate faster.	A-B As the gas is cooled, the particles lose kinetic energy and start to move slower.
B-C A mixture of solid and liquid exists here. During the melting process, the temperature of the substance remains constant. All the heat energy taken in by the particles is used to	B-C A mixture of gas and liquid exists here. The particles lose kinetic energy and they move slower and closer together. ©wzkai

overcome the forces of attraction. between particles.	
C-D The particles are randomly arranged and slide over one another. The substance becomes liquid.	C – D The particles become closer, randomly arranged and slide over one another. The substance becomes liquid.
BOILING	FREEZING
C The particles are closely packed but not in an orderly manner and slide over one another. (liquid state)	C The particles are randomly arranged and slide over one another. (liquid state)
C-D As the liquid is heated, the particles gain kinetic energy and start to vibrate/ move faster.	C – D As the liquid is cooled, the particles lose kinetic energy and start to slide over one another slower.
D-E A mixture of liquid and gas exists here. During the boiling process, the temperature of the substance remains constant. [All the heat energy taken in by the particles is used to overcome the forces of attraction between particles.]	D – E A mixture of liquid and solid exists here. The **particles lose energy**, the particles **start to move closer and slower.**
E – F They spread far apart and more rapidly in all directions. The substance is now a gas.	E – F The particles become closer and vibrate about fixed position one another. The substance becomes solid.

Separation techniques

1.Criteria for purity <Element, compound>

(a) A pure substance has

➔ A fixed melting and boiling point

➔ Only one spot on the chromatogram

(b) An impure substance has

➔ Melts and boils over a range of temperatures. It also has a lower melting point and higher boiling point than the pure substance

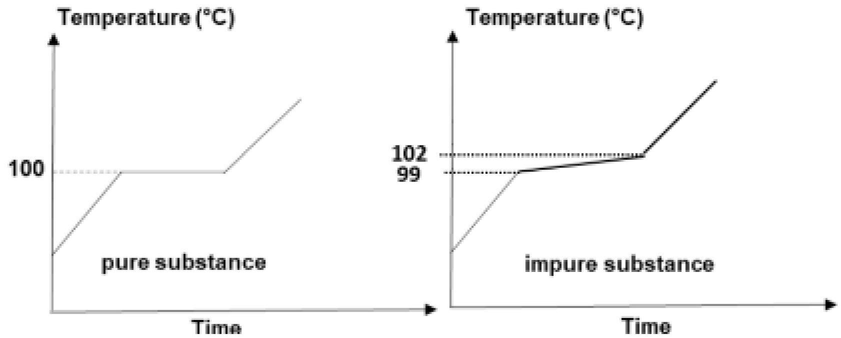
➔ More than one spot on the chromatogram.

Methods to test fo purity

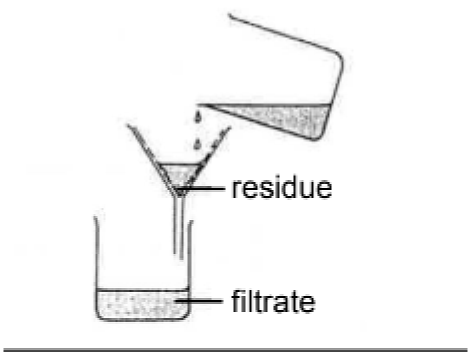
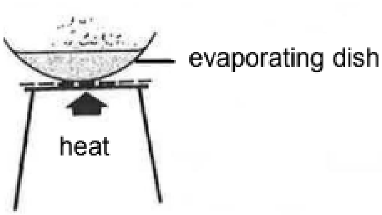
➔ Measure m.p / b,p

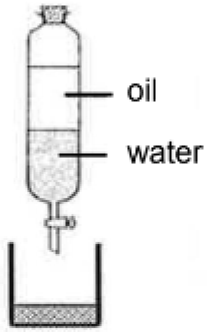
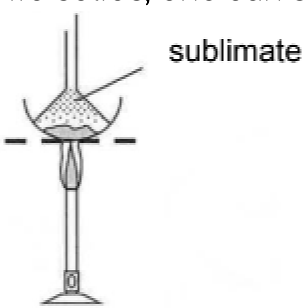
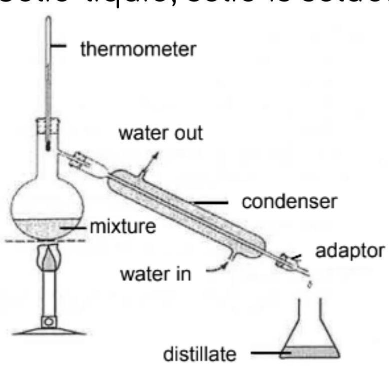
➔ Conduct a paper chromatography experiment.

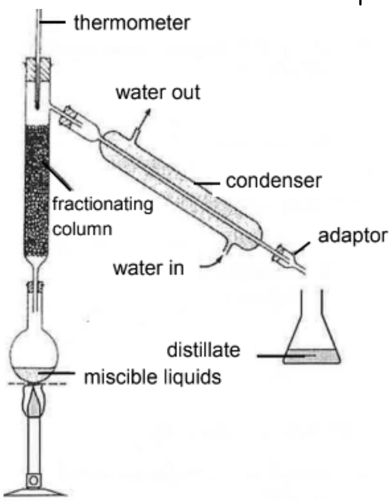
Boiling point of pure and impure substances



Chromatogram of pure and impure substances

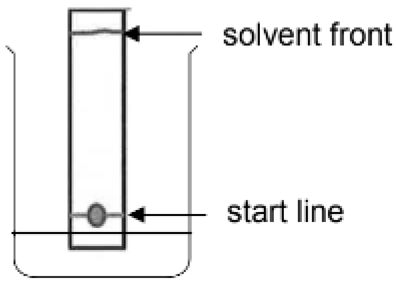
MIXTURE	Method of separation	
Solid-liquid mixture ; solid is insoluble 	**Filtration** Solid collected as the residue, liquid collected as filtrate.	
Solid-liquid mixture ; solid is soluble 	**Evaporation to dryness** 1. The solution is heated till dryness. (100% water removal) 2. Wash the crystals formed with cold distilled water and tap dry.	**Crystallisation** 1. The solution is heated till saturated. (most water being removed) 2. Cooled for the crystals to form. 3. Wash the crystals with cold distilled water and tap dry.
	For **heat-stable** solid (does not decompose on heating). Eg. sodium chloride salt.	For **heat-unstable** solid (decompose on heating). Eg. sugar.

Two immiscible liquids 	**Separating Funnel** Denser liquid is collected first from the tap below.	
Two solids, one can sublime 	**Sublimation** (From solid to gaseous state) Heat the mixture and collect the sublimate on a cold surface, like an inverted filter funnel. Examples of solid that sublime on heating: • Iodine • Ammonium chloride • Carbon dioxide	
Solid-liquid, solid is soluble 	**Simple distillation** 1. The solution boils in the flask. Boiling chips are added to ensure smooth boiling. 2. The thermometer measures the boiling point of the vapour before it enters the condenser. 3. In the condenser, the vapour condenses and changes from gaseous to liquid state. Water enters from the bottom and exits from the top of the condenser to ensure that it is fully filled. (to maximise efficiency of condensation) 4. Liquid collected is known as the	

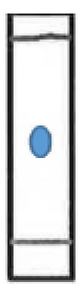
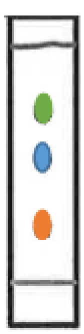
	distillate. 5. The solid is left at the bottom of the flask.	
Two or more miscible liquids 	**Fractional distillation** • A liquid of higher boiling point will condense at the fractionating column and fall back into the flask. A liquid with lower boiling point will be able to pass through the fractionating column. This ensures separation of liquids according to their boiling points. • Lower boiling point liquid is collected first distillate. • Example: alcohol and water, crude oil	

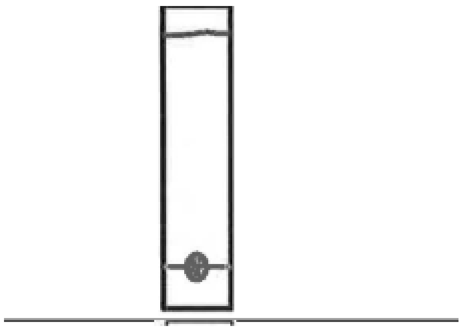
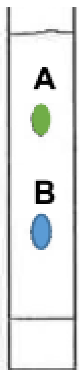
Condenser – To condense hot vapour from the delivery tube to liquid state.
Thermometer – Placed at the opening of the delivery tube to measure the temperature of the vapour entering the delivery tube.
Fractionating column – It is used to improve the efficiency of fractionating column.
Miscible Vs Immiscible –
Miscible liquids are ones that can mix together – like water and ethanol.
Immiscible liquids are ones that can't – like oil and water.

3.Paper chromatography

	• Start line must not be drawn in ink/pen (drawn with pencil) o The ink/pen is soluble in the solvent used and may interfere with the results. • Start line must be above the solvent • o To ensure that the sample does not dissolve directly in the solvent. • Ink samples must be small dots o To ensure clear and separate spots are obtained. • Solvent front must be as near to the top edge of the paper as possible o To ensure complete separation of dyes

Analysis of results

	1 single spot - The sample is pure. - It has only one component/spot/dot.
	2 or more spots - The sample is a mixture. - From the chromatogram, it has three components

	Remains at start line - The sample is insoluble in the solvent used. - To separate the mixture, change the solvent used. Universal solvent : water
	Different distance travelled - Component A is more soluble in the solvent than component B. - Thus component A travels further away from the start line.

Elements, Compounds and Mixture

★ Diatomic (2 atoms) → Gp VII elements, hydrogen, oxygen, nitrogen.

★ Monoatomic (1 atoms) → Gp 0 elements (unreactive)

1. Definitions

a. An **element** is a substance that cannot be broken down into two or more simpler substances by physical or chemical processes. (An element contains only one type of atom.)

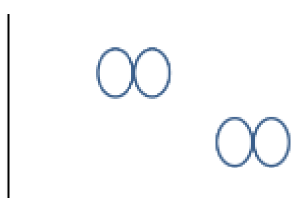
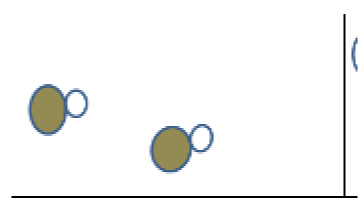
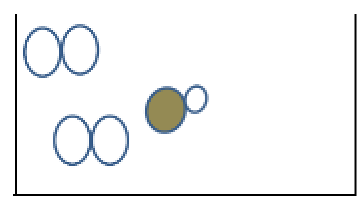
b. A **compound** is a substance made up of two or more elements combined together by chemical methods. (oxidation, combustion, thermal decomposition)

c. A **mixture** is a substance made up of two or more constituents which can be separated by physical methods. (filtration, paper chromatography)

i. An **alloy** is a mixture of a metal and other element(s).

E.g. brass, bronze, steel

2. Differences between elements, compound and mixtures

	Element	Compound	Mixture
Look out for the keyword or key phrases used	❖ Cannot decompose ❖ Cannot breakdown ❖ Simplest	❖ Constant Composition (fixed ratio) ❖ Decompose ❖ Breakdown	❖ Only some ❖ Speckle ❖ Filtration ❖ Chromatography ❖ Distillation/distilled
	❖ Test1 : Pure substance: fixed melting and boiling point		❖ Melts over a range of temperature
	❖ Test 2 : shows one spot (pure substance) for paper chromatography		❖ Shows more than one spot for paper chromatography
(reaction with O_2) On combustion/ burning with air	• Form only **ONE** Product E.g. $C+O_2 \rightarrow CO_2$	• Form two or more products E.g. Methane ($H_4+2O_2 \rightarrow C0_2+H_20$)	-
Diagram representation	• Only one type of particle	• 2 or more types of particles • **Combine in a fixed ratio**	• 2 or more types of particles • **Mix in variable ratio**
			

	Element	Compound	Mixture
Examples	Anything found in the Periodic Table	Ionic compound: sodium chloride potassium carbonate Covalent compound: water, carbon dioxide	Air Seawater Brass(alloy) Crude oil Steel (Fe and C) (alloy) ©wzkai

Atomic Structure

Subatomic Particles

subatomic particles	symbol	relative mass	relative charge	Where are they found? (in a nucleus or in an electron shell?)
Proton	p	1	+1	nucleus
Neutron	n	1	0	nucleus
Electron	e	1/1840	-1	electron shell

Important relationship!

❖ proton number (atomic number)
number of protons = number of electrons (in an atom)

❖ nucleon number (mass number/ relative atomic mass)
nucleon number = number of proton + number of neutrons

$^{A}_{Z}X$ or $^{Z}_{A}X$

A: nucleon number
Z: proton number (Always the smaller value than A)
X: Chemical symbol

Isotopes

Isotopes:

❖ Definition:
Isotopes are atoms of the same element with the **same number of proton** but **different number of neutrons.**
❖ They have **different physical properties** due to different number of neutrons
❖ They have **same chemical properties** due to the same number of valence electrons.

Atoms & Ions

Diagram of an atom:

❖ Protons and neutrons from the nucleus in the center.

❖ First electron shell can only hold up to two electrons

❖ Subsequent shells can hold up to an eight electrons

o **Period number** = total number of electron shells

❖ Valence shell = outermost shell

❖ Valence electrons = outermost shell electrons

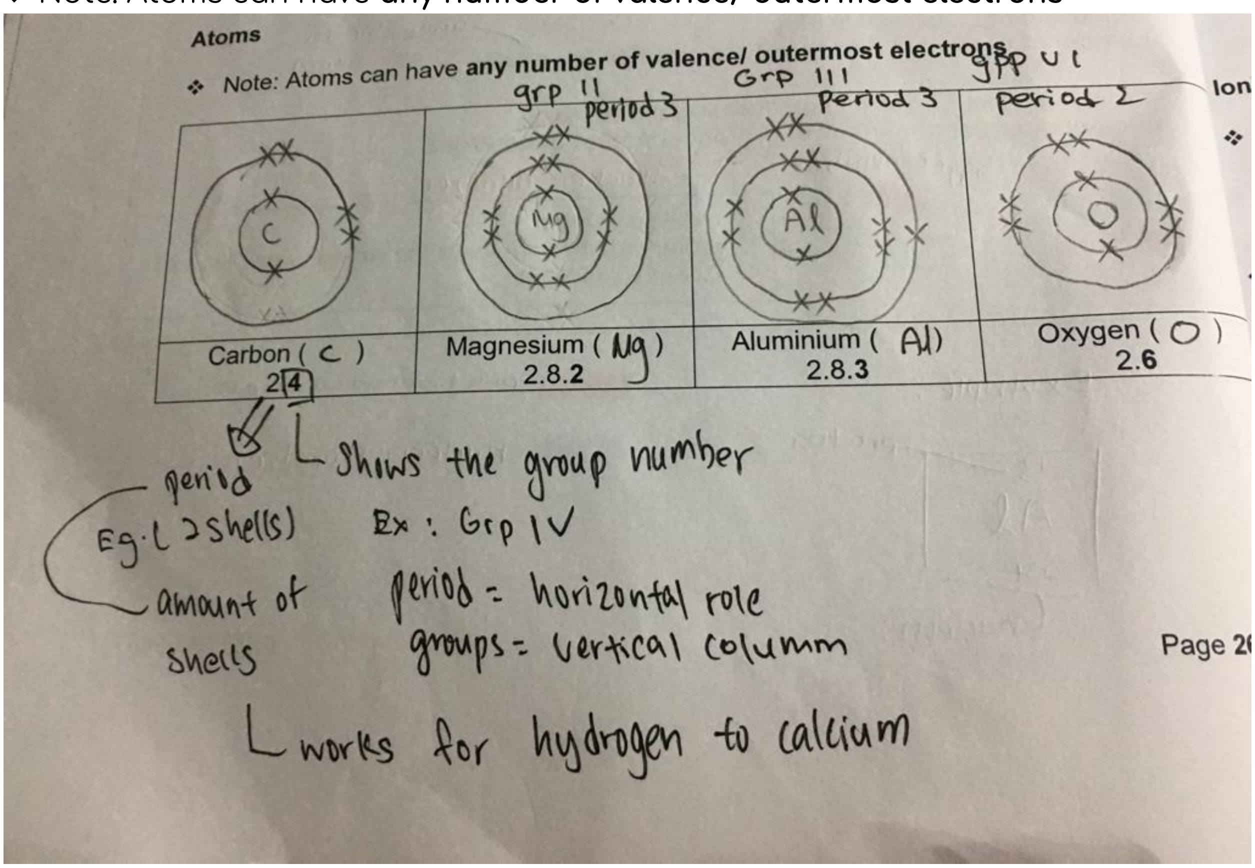
o **Group number** = number of valence electrons

❖ 2.8.8... rule is only applicable to first twenty elements (H to Ca)

2.8.8 <maximum electrons per cells>

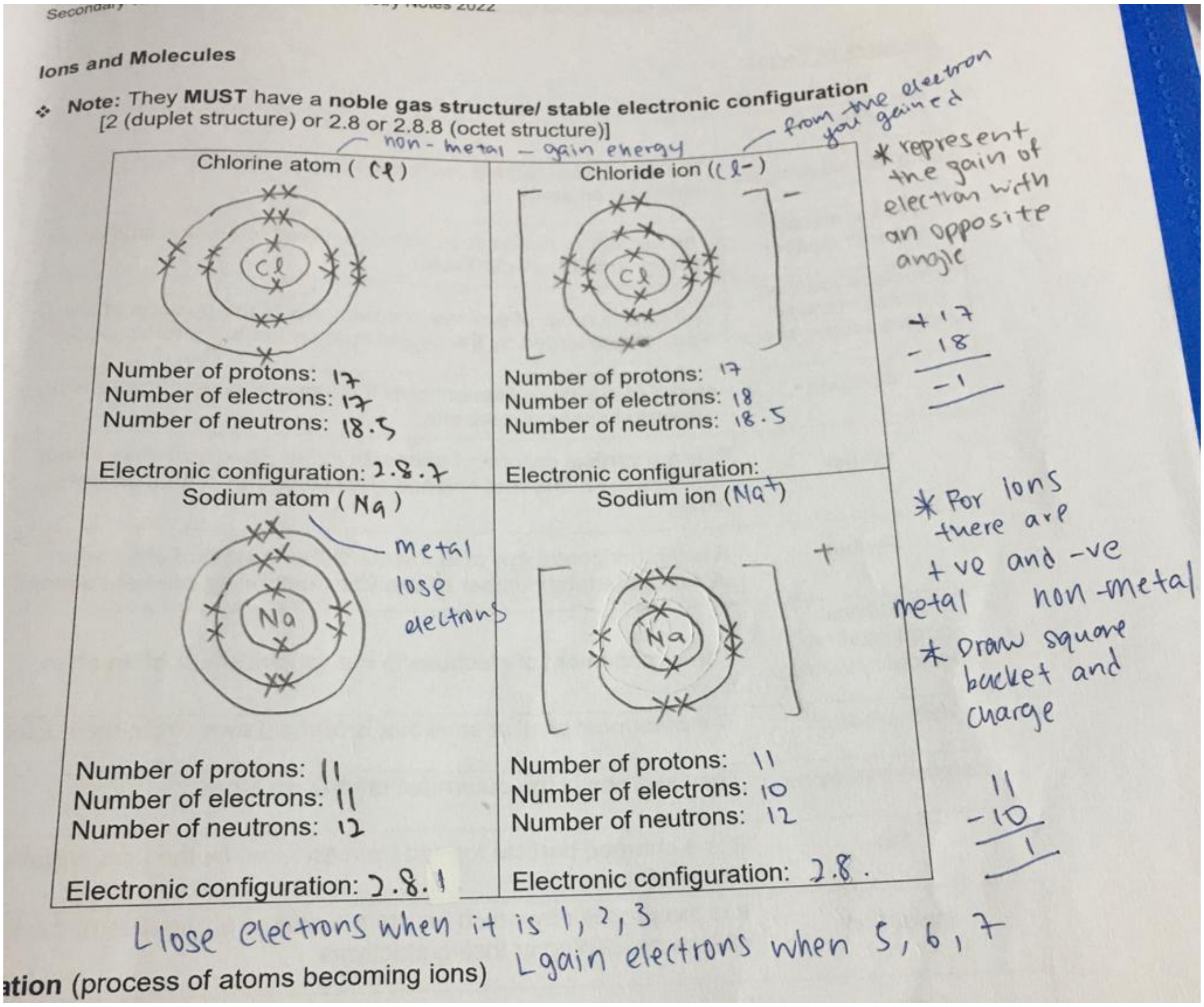
Atoms

❖ Note: Atoms can have **any number of valence/ outermost electrons**



Ions and Molecules

❖ **Note:** They **MUST** have a **noble gas structure/ stable electronic configuration** [2 (duplet structure) or 2.8 or 2.8.8 (octet structure)]



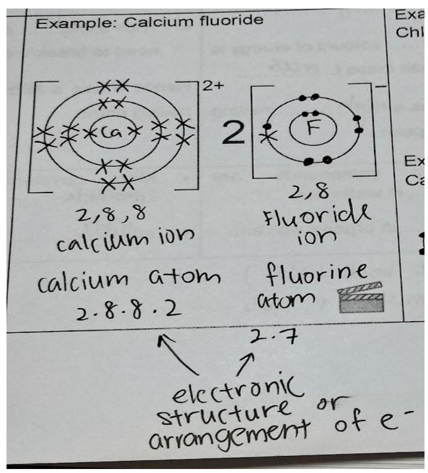
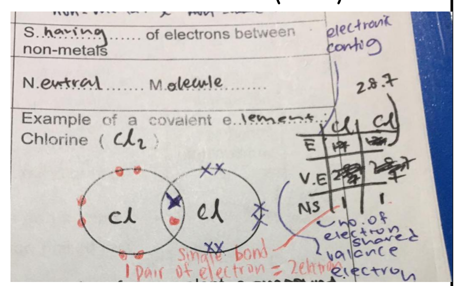
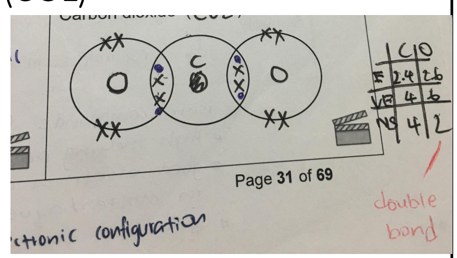
Ionisation (process of atoms becoming ions)

❖ Sodium loses one electron to become a positive ion in order to have a stable electronic configuration (This is a **positive** ion – more proton than electron.)

❖ Chlorine gains one electron to become a negative ion in order to have a stable electronic configuration . (This is a **negative** ion – more electrons than protons)

❖ Argon does not need to gain or lose electrons because it already has a stable electronic configuration

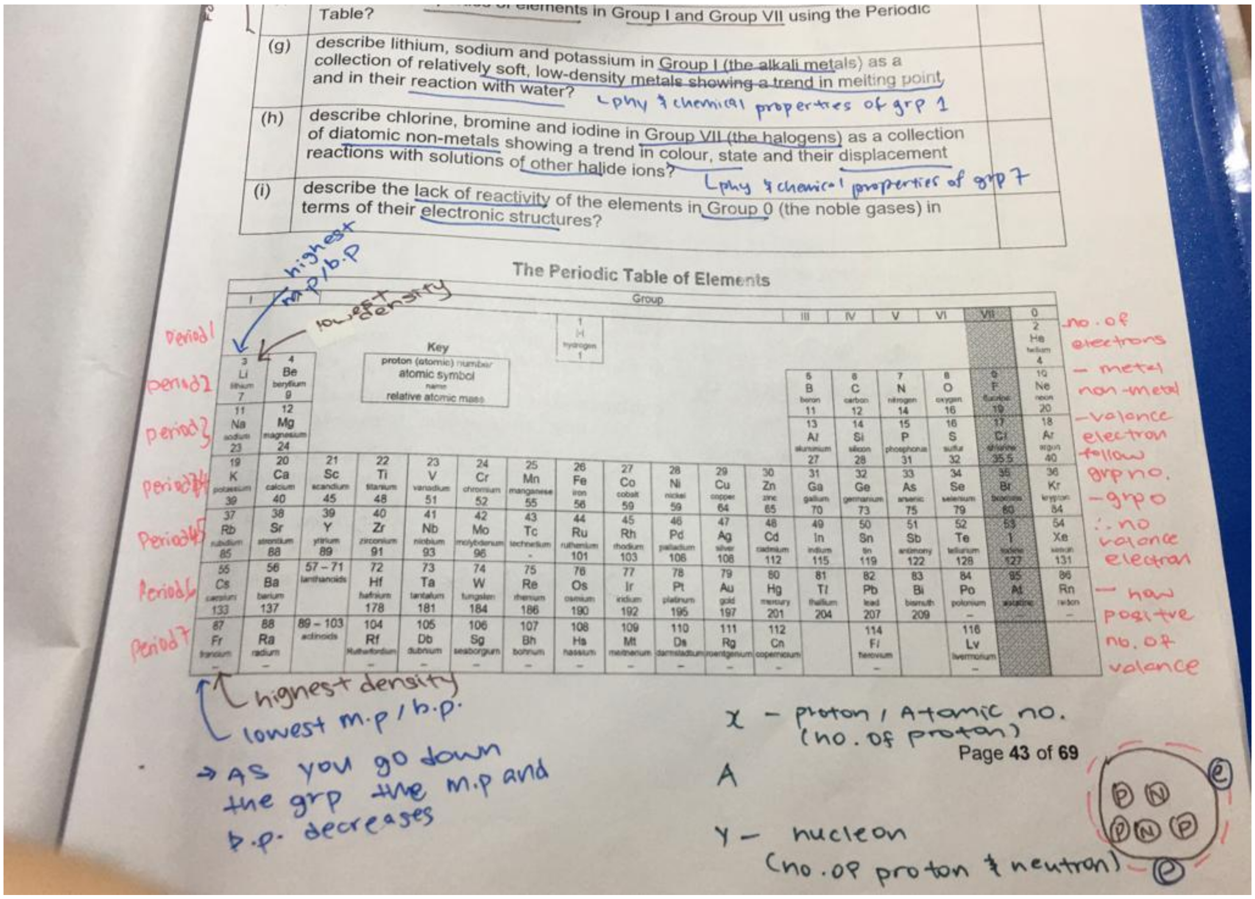
<u>Chemical Bonding</u>

	IONIC COMPOUNDS	SIMPLE COVALENT SUBSTANCE
Form between	Metal atoms (gain electrons) and non-metal atoms (lose electrons)	Non-metal atoms
Definition	<u>Transfer</u> of electrons from metal to non-metal	<u>Sharing</u> of electrons between non-metals
Product formed	Charged Ions	Neutral Molecules
Dot-and cross diagram (only valence electrons)	Example: Calcium fluoride  	Example of a covalent element Chlorine (Cl_2)  Example of a covalent compound Carbon dioxide (CO2) 

Electrical conductivity (structure and bonding)	In solid state, • It has no free moving ions • To carry electric charge from one end to the other. • Thus it is a poor conductor of electricity. In aqueous or molten state, • It has free moving ions • To carry electrical charges from one end to the other. • Thus it is a good conductor of electricity.	In every state, • It has no free moving ions or electrons • To carry electrical charges from one end to the other. • Thus it is a poor conductor of electricity.
Melting and boiling point (structure and bonding)	• It has a giant crystal lattice structure • With strong electrostatic attraction force between oppositely charged ions • Thus large amount of energy is need to break these forces • Hence it has a high melting and boiling point.	• It has a simple molecular structure • With weak attraction force between molecules • Thus small amount of energy is need to break these forces. Hence it has a low melting and boiling point.

Solubility	• Most ionic compounds are soluble in water but •insoluble in organic solvent.	• Most covalent substances are insoluble in water but • soluble in organic solvent.

Group Properties



	Alkalis Metals <High Ph level> (Group I)	Halogen (Group VII)	Noble Gases (Group 0)
Melting/ Boiling points	**General:** Low **Trend:** *Decreases* down the grp	**General:** Low **Trend:** *Increases* down the grp	**General:** Low **Trend:** *Increases* down the grp
Density	**General:** Low **Trend:** *Increases* down the grp	**General:** Low **Trend:** *Increases* down the grp	**General:** Low **Trend:** *Increases* down the grp ©wzkai

<table>
<tr>
<td></td>
<td>Note : The density and melting/boiling point is low in comparison to other metals</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Others <physical properties></td>
<td>Soft
Grey colour
Store in oil
(to prevent any reaction with oxygen or water)</td>
<td>fluorine g F_2 Yellow
chlorine g Cl_2 Green
bromine l Br_2 Brown
iodine s I_2 Black
astatine s At_2 Black
(Colour become darker, more intense)</td>
<td></td>
</tr>
<tr>
<td>Chemical reactivity <chemical properties></td>
<td>Reaction with water to form alkali (metal hydroxide) and Hydrogen gas.

<table>
<tr><th></th><th><u>Observation</u></th></tr>
<tr><td>Li</td><td>Fizzes</td></tr>
<tr><td>Na</td><td>Vigorously
Yellow flame</td></tr>
<tr><td>K</td><td>Violently
Lilac flame</td></tr>
<tr><td>Rb</td><td>Explosively</td></tr>
<tr><td>Cs</td><td>Explosively</td></tr>
<tr><td>Fr</td><td>Explosively</td></tr>
</table>

Note: chemical reactivity <u>increases</u> down grp I.

Eg

$2Na + 2H_2O \rightarrow$ <u>2NaOH</u>+ <u>H2</u>

Qn:
What is the colour of the solution when a few drops of Universal Indicator is added after the reaction?</td>
<td>Halogen Displacement

Idea:
A more reactive halogen displace a less reactive halogen from its salt solution.

<table>
<tr><th></th></tr>
<tr><td>fluorine</td></tr>
<tr><td>chlorine</td></tr>
<tr><td>bromine</td></tr>
<tr><td>iodine</td></tr>
<tr><td>astatine</td></tr>
<tr><td>Chemical reactivity <u>decreases</u> down grp VII</td></tr>
</table>

Observation:
Colourless solution turns brown. (Note: the brown colouration is due to either aq Br_2 or I_2 formed)</td>
<td>Explanation:
Unreactive as it has a <u>stable electronic configuration</u>.

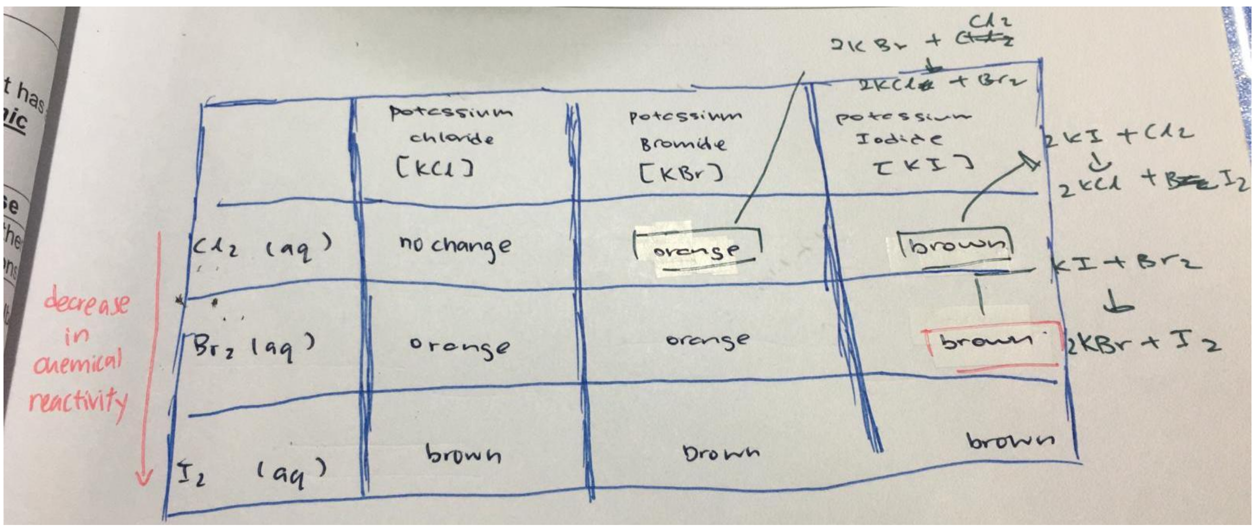
<table>
<tr><th></th><th><u>Use</u></th></tr>
<tr><td>He</td><td>Weather balloons</td></tr>
<tr><td>Ar</td><td>Light bulbs</td></tr>
<tr><td>Ne</td><td>Advertising light</td></tr>
</table>

Atoms are trying to achieve stability (8e- configuration)

- Noble gases already have 8e- (or 2e- → He) ∴ no reaction

</td>
</tr>
</table>

	Purple.	**Explanation:** • X_2 is more reactive than Y_2. • Thus X_2 displaces Y_2 from its salt solution. **Eg** Cl_2 + KBr → 2KCL + Br2	



Acid and Bases

Chemical and Physical Properties of Acids & Bases

	Acids	Bases (soluble in alkali and insoluble in water)
Definition	A substance that produces Hydrogen ions (H+) in water. Particle responsible for acidic properties: Hydrogen ions (H+)	A metal oxide hydroxide that reacts with acids to form salt and water only. Soluble bases: Alkalis Particle responsible for alkaline properties: hydrogen ions (OH-) ©wzkai

Physical properties	• Taste sour • pH less. than 7	• Taste bitter • pH more than 7 • Soapy feeling
Chemical properties	1. Reaction with metal 2. Reaction with carbonate 3. Reaction with alkalis (Neutralisation) Ionic equation: H+(aq) +OH-(aq)-->H2O(l). e.g.Hcl+NaOH →Nacl + H2O	1. Reaction with acid 2. Reaction with ammonium salt
Examples	In daily life: vinegar, lemons. Chemicals: hydrochloric acid (HCL), sulfuric acid (H2So4) nitric acid (HNO3)	In daily life: soap, detergent, most cleaning agent Chemicals: sodium hydroxide (NaOH) calcium hydroxide (Ca(OH)2)

Chemical Properties:

	Chemical Reaction	Observation
1	Acid + Metal → Salt + H_2	Lighted splint will extinguish with a pop sound
2	Acid + Metal Carbonate → Salt + CO_2 + H_2O	Bubbles formed When gas produced is bubbled into limewater a white precipitate is formed.
3	Acid + Alkali → Salt + H_2O (Neutralisation)	No visible change
4	Alkalis + Ammonium salt → Salt + NH_3 + H_2O	Pungent gas is produced Gas produced turns moist red litmus paper will turn blue

Test for Gases

	Hydrogen gas	Carbon dioxide gas	Ammonia gas
Test	Place a lighted splint at the mouth of the test tube.	Bubble the gas produced through limewater	Place a moist red litmus paper at the mouth of the test tube.
Observation	If hydrogen gas is produced, the lighted splint will extinguish with a pop sound	If carbon dioxide gas is produced, a white precipitate will be formed.	If ammonia gas is produced, the moist . red litmus paper will turn blue

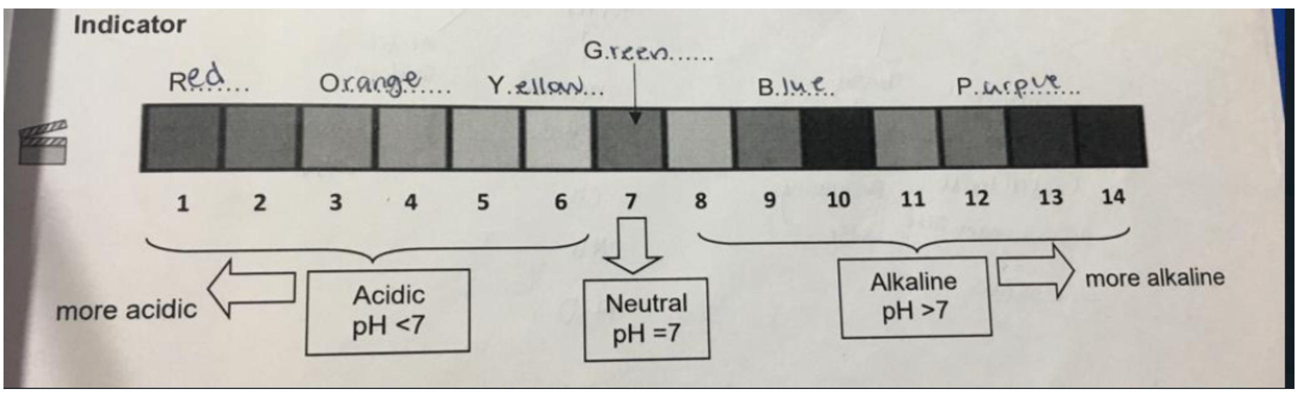
ρH Scale and Universal Indicators

• The ρH of a solution is a measure of the acidity , alkalinity or neutrality of a solution.

• ρH of a substance can be measured using Universal Indicator paper or

solution • A ρH scale ranges from 1 to 14 .

Indicator



Reducing Acidity of Soil

• Most plants grow well in neutral soil of ρH around 6.5.
• Excess acidity in soil is neutralised by bases such as quicklime (calcium oxide) / slaked lime (calcium hydroxide). Sometimes limestone (calcium carbonate) is used.
• This process is known as 'liming the soil'.

Types of Oxides

Acidic Oxides	Basic Oxides	Amphoteric Oxides	Neutral Oxides
Non-metal oxides Eg CO_2, SO_2 Dissolves in water to form an acid. Reacts with alkalis/bases to form salt and water only.	Metal oxides Eg CaO, Na_2O, MgO Reacts with acid to form salt and water only.	Three specific examples: Lead (ll) oxide [PbO] Aluminium oxide [Al2O3] Zinc oxide [ZnO] Reacts with BOTH acid and bases to form salt and water.	➔ Three specific examples: CO (carbon monoxide) , NO (nitrogen monoxide) , H_2O (dihydrogen monoxide) Reacts with **NEITHER** acid **NOR** bases

Salt preparation

Memorize the salt solubility table

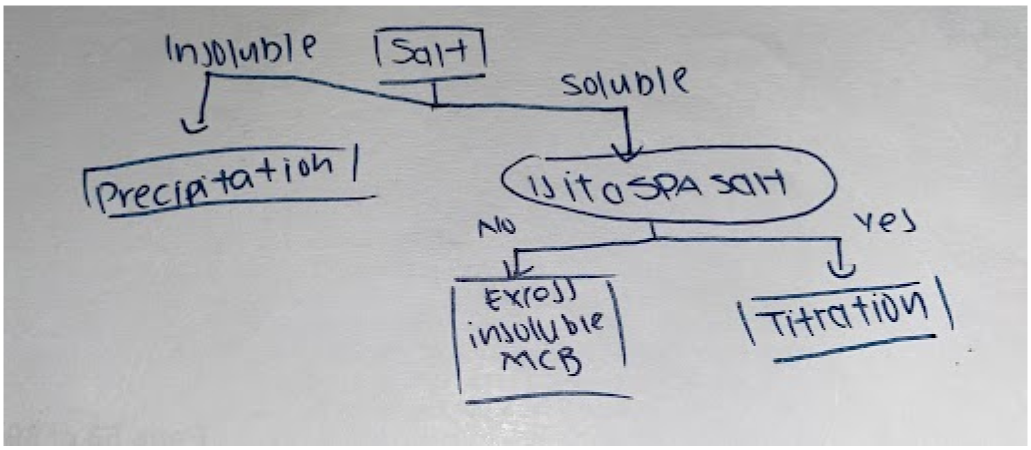
salt	soluble	Insoluble
chloride	All except	Lead(II) chloride PbCl2 Silver chloride AgCl
sulfate	All except	Calcium sulfate CaSO4 Lead(II) sulfate PbSO4 Barium sulfate BaSO4
carbonate	Sodium carbonate Potassium carbonate Ammonium carbonate	All except
nitrate	All	

Note:
All Group I metal salts, Ammonium salts and Nitrate salts are soluble in water.

Steps to determine the method to prepare the given salt.
Question1: Is the salt soluble or insoluble in water?
Answer 1:Insoluble → Precipitation
Answer 2:Soluble → Titration or Excess MCB with acid

Question 2 : If the salt is soluble, is it a SPA salt?
Answer 1:No → Excess insoluble MCB with acid
Answer 2:Yes → Titration

Summary



Can i use the insoluble MCB + Acid method to make Nacl from NaOH and Hcl?
-Because the substance will dissolve like NaOH and Hcl

- ☐ If any starting material is in excess, it cannot be removed from the solutions by filtration
  - ➔ Salt produced will be impure
  - ➔ Therefore volumes of each material used must be exact

SALT PRECIPITATION

Excess insoluble MCB

1. Add excess insoluble MCB into a beaker containing a fixed volume of acid.
2. Filter the mixture and collect the filtrate.
3. Heat the filtrate till saturated and allow it to cool for crystals to form.
4. Wash the crystals formed with cold distilled water and tap dry.

Precipitation

1. Add two salt solution in a beaker and stir.
2. Filter the mixture and collect the residue.
3. Wash the residue with distilled water and tap dry.

Titration

1. Pipette 25.0cm3 of acid into a conical flask.
2. Add a few drops of suitable indicator into the acid.
3. Fill a burette with alkali and record the initial reading (V1cm3).
4. Introduce the alkali into the acid until a permanent colour change is observed and record the final reading (V2cm3)
5. Repeat the experiment with V2-V1 cm3 of alkali and without indicator.
6. Heat the filtrate till saturated and allow it to cool for crystals to form.
7. Wash the crystals formed with cold distilled water and tap dry.