

Atomic Structure (chapter five)

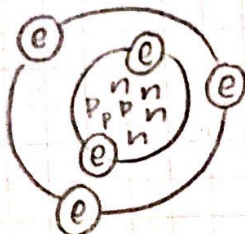
What is an atom?

The smallest particle of a substance.

What are atoms made up of?

3 types of sub-atomic particles:

- ① Proton
- ② Electron
- ③ Neutron

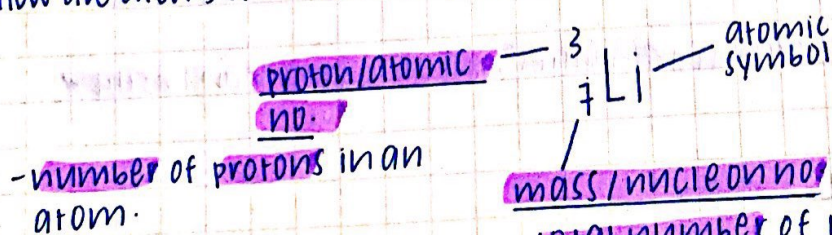


- Protons and neutrons are found in the nucleus of the atom.
- Electrons are found in electron shells outside the nucleus of the atom.

Particle	Relative Mass	Relative charge	Describe relative charge
Proton	1	+1	single positive charge
Electron	$\frac{1}{1840}$ / $\frac{1}{1836}$	-1	single negative charge
Neutron	1	0	no charge

- The number of protons and electrons is the same, hence an atom is electrically neutral ($P=E$)

How are atoms of elements represented in the periodic table?



- total number of protons and neutrons in an atom.
↳ also known as relative atomic mass.

* When writing symbol form of an element, the proton and nucleon number need to be flipped.

What are isotopes?

- Isotopes are atoms of the same element with the same number of protons but different number of neutrons.
- ↳ This means that isotopes have the same proton / atomic nos, but different mass / nucleon numbers.

Example 1: isotopes of Chlorine
Chlorine has 2 stable isotopes, chlorine-25, $^{35}_{17}\text{Cl}$ and chlorine-37, $^{37}_{17}\text{Cl}$.

Isotope	No. of Protons	No. of electrons	No. of neutrons
$^{35}_{17}\text{Cl}$	17	17	18
$^{37}_{17}\text{Cl}$	17	17	20

- 3) Boron have two isotopes, boron-11 and boron-12. Explain why ^{11}B and ^{12}B are isotopes.

They are atoms of the same element, boron. Both have 5 protons but boron-11 has 6 neutrons while boron-12 has 7 neutrons.

Properties of isotopes

- Isotopes of the same element have same chemical properties but different physical properties.
- Isotopes have same chemical properties because they have the same number of valence electrons.
- Isotopes have different physical properties due to different relative atomic masses (mass numbers).
- ↳ This results in the isotopes having different densities, melting points and boiling points.

Arrangement of Electrons in Atoms

- The electronic configuration / structure of an atom tells us how electrons are arranged among the various electron shells.

1st shell: 2 electrons

2nd shell: 8 electrons

3rd shell: 8 electrons

Valence Shell and Valence Electrons

- The outermost electron shell is called the valence shell. It is the shell that is the furthest away from the nucleus of the atom.
- The electrons that are found in the valence shell are known as the valence electrons.

Periodic Table

in the periodic table,

- elements are arranged in order of increasing proton number.
- horizontal rows are called periods. Elements in the same period ~~have the same period~~ have the same number of electron shells.
 - ↳ elements in period 2 all have 2 electron shells.
- vertical columns are called groups. Elements in the same group have the same number of valence electrons.
 - ↳ elements in group 1 all have 1 valence electron.

Chemical Bonding: (chapter six)

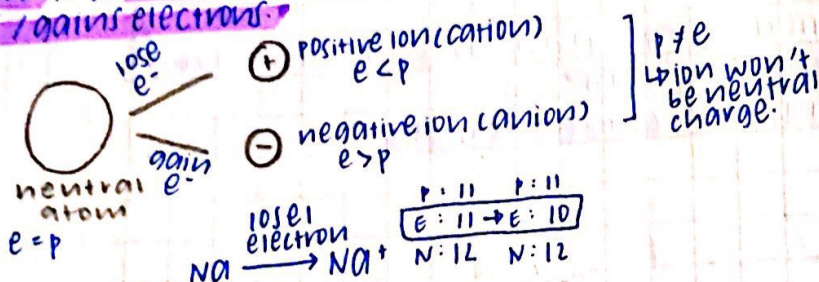
IONS & IONIC BONDING

IONS

In a neutral atom, the number of protons is the same as the number of electrons.

If an atom loses/gains electrons, then the atom is no longer electrically neutral. It becomes a charged particle, which is called an ion.

An ion is a charged particle formed when an atom/a group of atoms loses/gains electrons.



GROUP 0 ELEMENTS

Known as the Noble Gases.

↳ electronic structure & electronic configuration.

He: 2
Ne: 2.8
Ar: 2.8.8

fully filled valence shell

Noble Gases

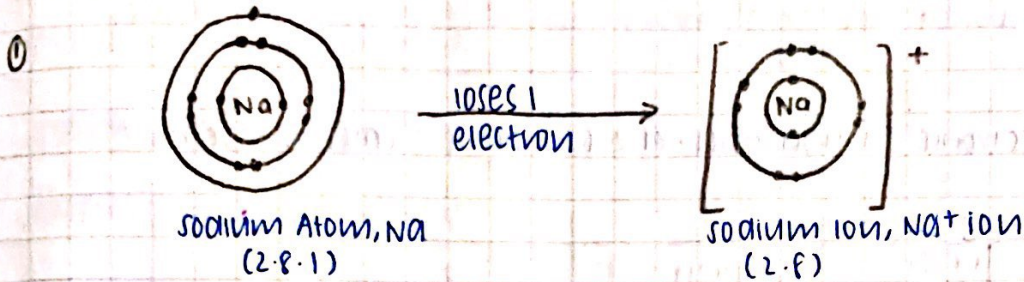
All the electron shells of the noble gases are fully-filled / complete. That is, each electron shell is occupied by the maximum allowable no. of electrons.

Atoms of noble gases do NOT need to gain, lose / share any electrons, hence they are chemically inert / stable / unreactive as compared to non-noble gases.

All non-noble gases want to achieve stable noble gas electronic configuration.

Formation of positive ions (cations)

- a positive ion is formed by losing valence electrons from an atom, such that the number of electrons is less than the number of protons.
- It has an overall positive charge.
- metals tend to form positive ions.
- metal atoms can easily lose valence electron(s) to achieve a sngec. This is because, it takes less energy for metals to lose electrons than to take in electrons.



② Explain why sodium is a metal, with reference to its e.c.

↳ Sodium has an electronic configuration of 2.8.1. It loses 1 valence electron to form a sodium ion with a single positive charge to achieve sngec.

Periodic Table

group = charge

I = +1
 II = +2
 III = +3

} lose electrons (cation)

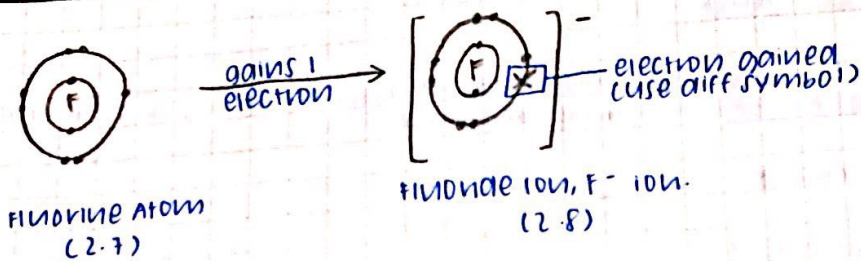
V = -3
 VI = -2
 VII = -1

} gain electrons (anion)

Group	NO. OF valence e ⁻	NO. OF e ⁻ lost	CATION
I	1	1	M ⁺
II	2	2	M ²⁺
III	3	3	M ³⁺

Formation of negative ions (anions)

- a negative ion is formed by gaining valence electrons from an atom, such that the number of electrons is more than the number of protons.
- It has an overall negative charge.
- non-metals form negative ions
- atoms of non-metals readily gain valence electrons to achieve stable noble gas electronic configuration. This is because it takes less energy to accept electrons than to give out electrons.



- for non-metal ions, change last 3 letters to -ide.

Explain the formation of the oxide ion, with reference to its electronic structure.

↳ Oxygen has an electronic configuration of 2.6. It gains 2 valence electrons to achieve 8/8/2.

Group	No. of valence e^-	No. of e^- gained	Anion
VII	7	1	X^-
VI	6	2	X^{2-}
V	5	3	X^{3-}

IONIC BONDS

• ionic bonds are formed when a metal atom transfers its valence electrons to non-metal atoms.

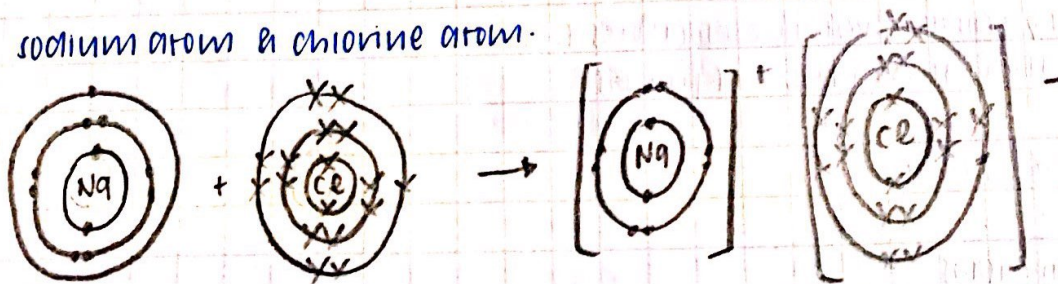
↳ when this happens, we call them ionic compounds.

→ sodium chloride (NaCl), table salt.

• It is the electrostatic forces of attraction that hold a positive & negative ion together.

[bot a cross diagram]

sodium atom & chlorine atom.



KEY

- — electron of Na
- × — electron of Cl

⇒ NaCl

ionic compound	cation	anion	chemical formula
sodium oxide	Na^+	O^{2-}	Na_2O
aluminium oxide	Al^{3+}	O^{2-}	Al_2O_3
copper (II) chloride	Cu^{2+}	Cl^-	CuCl_2

*common ions - cation

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

*silver $\rightarrow \text{Ag}^+$
zinc $\rightarrow \text{Zn}^{2+}$

ionic compounds: physical properties

- most common ionic compound $\rightarrow$ salt (sodium chloride)
 $\rightarrow \text{NaCl}$

ionic compounds:

- 3D giant lattice structures also called crystal lattice.
- The positive & negative ions are attracted very strongly.
- Ions are packed together in regular patterns.

Physical property 1: high melting & boiling points

property: ionic compounds have high melting & boiling point.

Explanation: large amount of energy is required to overcome strong ionic bonds between ions, hence ionic compounds have high melting and boiling points.

- Explain why magnesium oxide has a high melting point of 3600°C .
 $\rightarrow$ magnesium oxide is an ionic compound. large amount of energy is required to overcome strong ionic bonds between ions.

Property 2: Electrical conductivity

- Solid NaCl does not conduct electricity.
- molten & aqueous NaCl conducts electricity.

Why?

- presence of free moving / mobile ions.

Property 3: Solubility

- ionic compounds are soluble in water to form aqueous solution.

- ionic compounds are insoluble in organic solvents.

↓
alcohols, petrols
& kerosene.

Chemical Bonding: (chapter six)

covalent Bonding

- Covalent Bonds -

What is a covalent bond?

- the bond formed when non-metals share their valence electrons

Why share valence electrons?

- so that the stable noble gas electronic configuration can be obtained for each atom involved in the bond.

- when non-metal atoms share their electrons, molecules are formed.

* **molecule**: a grp of atoms held together by covalent bonds

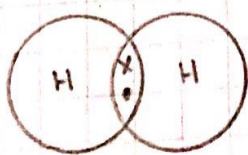
"dot & cross"

**
no. of e⁻ shared = no. of e⁻ needed

① Hydrogen (H_2) - 2 H atoms, covalent bond (non-metals)

- has 1 valence electron

↳ needs 1 more electron to achieve SNGEC (share 1)



key

x - e⁻ of H

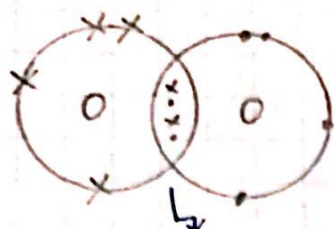
o - e⁻ of H

'dot and cross' diagram	structural formula	molecular formula	Model
	H-H	H ₂	

a single covalent bond
(1 pair of electron shared $\frac{x}{\div}$)

② Oxygen (O_2)

- has 6 valence electrons
- ↳ needs 2 more electrons (share 2)



key
 x - e^- of O
 • - e^- of O

* note: alternate the electrons shared.
 put the remaining electrons in "NESW"

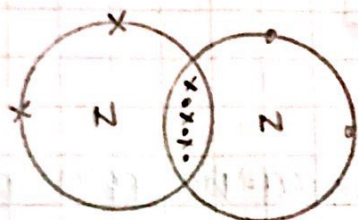
structural formula: $O=O$ → double covalent bond.

molecular formula: O_2



③ Nitrogen (N_2)

- has 5 valence electrons
- ↳ needs 3 more (share 3)



key
 x - e^- of N
 • - e^- of N

→ triple covalent bond

structural formula: $N \equiv N$

molecular formula: N_2



no. of valence e^- shared per atom	Type of bond formed	represented by: (for molecular diagram)	examples
one	single bond	—	F_2, Cl_2, H_2
two	double bond	=	O_2
three	triple bond	≡	N_2

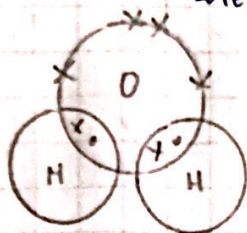
no. of electrons needed = no. of electrons shared

-Complex Covalent-

1. water $\rightarrow \text{H}_2\text{O}$
2. methane $\rightarrow \text{CH}_4$
3. carbon dioxide $\rightarrow \text{CO}_2$
4. Ammonia $\rightarrow \text{NH}_3$

memorise!!

① water (H_2O) $\begin{cases} \text{H: } 1 \rightarrow \text{need } 1e^- \rightarrow 1e^- \text{ shared} \\ \text{O: } 2.6 \rightarrow \text{need } 2e^- \rightarrow 2e^- \text{ shared} \\ \rightarrow \text{left 4 valence } e^- \end{cases}$

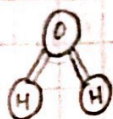


key
 $\times$ - e^- of O
 $\cdot$ - e^- of H

* note: must be in "mickey mouse" shape

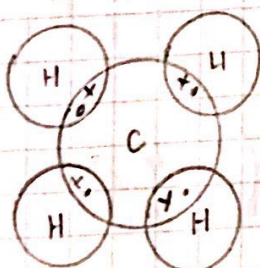
structural formula: $\begin{array}{c} \text{O}-\text{H} \\ | \\ \text{H} \end{array} \quad / \quad \begin{array}{c} \text{O} \\ / \quad \backslash \\ \text{H} \quad \text{H} \end{array}$

molecular formula: H_2O
 model



- two single covalent bond

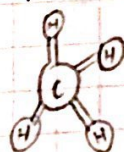
② methane (CH_4) $\begin{cases} \text{C: } 2.4 \rightarrow \text{need } 4e^- \rightarrow 4e^- \text{ shared} \\ \text{H: } 1 \rightarrow \text{need } 1e^- \rightarrow \text{share } 1e^- \end{cases}$



key
 $\times$ - e^- of C
 $\cdot$ - e^- of H

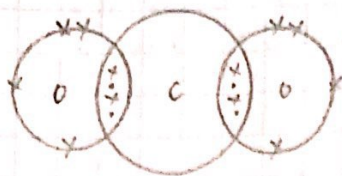
structural formula: $\begin{array}{c} \text{H} \\ | \\ \text{H}-\text{C}-\text{H} \\ | \\ \text{H} \end{array}$

molecular formula: CH_4
 model



- 4 single covalent bond.

③ carbon dioxide (CO_2) $\begin{matrix} 2.6 - \text{need } 2e^-, \text{ share } 2e^- \\ \text{left } 4e^- \\ 2.4 \\ - \text{need } 4e^-, \text{ share } 4e^- \end{matrix}$

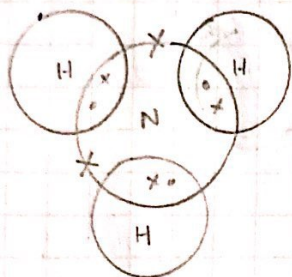


key
 $\cdot$ - e^- of O
 $\times$ - e^- of C

structural formula: $\text{O}=\text{C}=\text{O}$
 molecular formula: CO_2
 model

- two double covalent bond

④ ammonia (NH_3) $\begin{matrix} 2.5 - \text{need } 3e^-, \text{ share } 3e^- \\ \text{left } 2e^- \\ 1 \\ - \text{need } 1e^-, \text{ share } 1e^- \end{matrix}$



structural formula: $\text{H}-\text{N}-\text{H}$
 $\quad \quad \quad \text{H}$
 molecular formula: NH_3

- three single covalent bond

Physical Properties -

- ① melting & boiling points
- ② electrical conductivity
- ③ solubility

PROPERTY 1: low melting & boiling point

- many simple covalent substances that are made up of small molecules are liquids / gases at room temp.

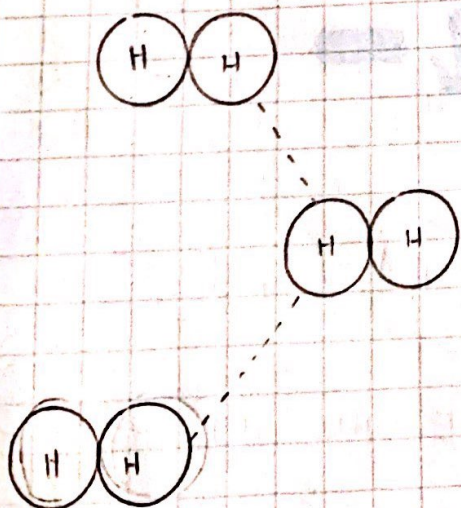
↳ they have low melting & boiling points.

↳ those that are liquids are volatile.

↓
they evaporate easily at rm. temp & pressure.

covalent substance	melting point (°C)	boiling point (°C)
CO ₂ carbon dioxide	-56	-79
Cl ₂ chlorine	-101	-35
H ₂ Hydrogen	-259	-253
CH ₄ methane	-183	-161
O ₂ Oxygen	-214	-183
H ₂ O water	0	100

Why do covalent compounds have low melting & boiling points?



- This occurs because the forces between the molecules are weak compared to the covalent bonds within molecules.

- small amount of energy is needed to overcome the weak intermolecular forces of attraction between molecules.

PROPERTY 2: electrical conductivity

- most covalent elements / compounds do not conduct electricity whether in the solid, liquid / gaseous state.

WHY?

There are no mobile / free moving ions / electrons present to carry electrical charges.

EXCEPTIONS:

Carbon, in the form of graphite, conducts electricity. Hydrogen chloride, sulphur dioxide and ammonia react with water to form solutions that conduct electricity.

PROPERTY 3: solubility

- most covalent molecules are insoluble in water & soluble in organic solvents.

EXCEPTIONS:

Chlorine, hydrogen chloride, alcohol and sugar are covalent substances that are soluble in water.

- Writing Chemical formula -

- The name of a covalent compound tells us
 - ↳ the elements it contains
 - ↳ the no. of each type of atom a molecule contains. (prefixes)

prefix	no. of atoms
mono	1
di	2
tri	3
tetra	4
penta	5

common covalent compounds:

1. ammonia (NH_3)
2. hydrogen peroxide (H_2O_2)
3. methane (CH_4)
4. ozone (O_3)
5. water (H_2O)

name the following compounds:

- NO → nitrogen oxide
- SO_2 → sulfur dioxide
- N_2O_4 → dinitrogen tetroxide
- H_2O_2 → hydrogen peroxide
- NH_3 → ammonia
- HCl → hydrogen chloride
- H_2 → hydrogen molecule
- Cl_2 → chlorine molecule
- O_2 → oxygen molecule

- elements -

- an element is a pure substance that cannot be broken down into two / more simpler substances by chemical processes / by electricity. (consist of the same type atoms of elements)

Physical properties of metals and non-metals.

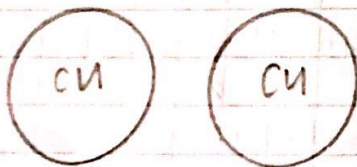
metals

1. Good conductors of electricity
2. Good heat conductors
3. Strong and hard
4. Malleable
5. Shiny appearance (lustrous)
6. High density
7. High melting and Boiling points (except sodium, potassium & mercury)

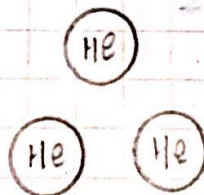
non-metals

1. Non-conductors of electricity (except carbon in the form of graphite)
2. Poor heat conductors
3. Weak and soft
4. Brittle if solid
5. Dull appearance
6. Low density
7. Low melting and Boiling points (except for carbon & silicon)

- Particles of elements can be atoms / molecules.



two copper atoms
(metal)

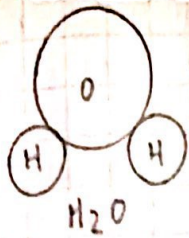


three helium atoms
(noble gases)

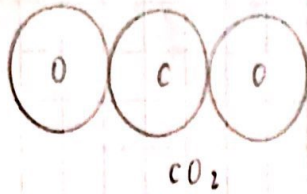
↓
usually monoatomic

- pure elements possess their own chemical properties.

Compounds



a water molecule



a carbon dioxide molecule

- a substance consisting of 2/more elements chemically combined together.
- can only be broken down into its constituent elements, by chemical reactions / electricity.
- physical properties diff from its constituent elements.
- possess unique chem properties to them

- Mixture -

a mixture consists of two/more substances which are not chemically combined together.

- does not have a fixed composition by mass; percentage of each element in a mixture is not always the same
- can be made up of different compounds and/or elements (can be element; a element)
- can be separated by physical methods, such as filtration, evaporation etc.
- consists of different types of atoms.
 - air: elements & compounds.
- the physical properties of each individual molecules, element & compound are retained. (no fixed physical properties)
- the chemical properties of each individual molecule, element & compound are retained.
- no chemical reaction present, little/no energy change involved.

add on: compound

- chemical reaction present.
- change in energy noted, either in terms of energy $\uparrow / \downarrow$.
- elements in a compound have a fixed proportion by mass / fixed percentage composition by mass.

-time-

SI unit: seconds (s)

other units: hours (h) / minutes (min)

type of instrument / apparatus:

1) analogue stopwatch

2) digital stopwatch

→ accuracy: milliseconds

* in chemistry, time is recorded to nearest seconds.

-temperature-

SI unit: kelvin (K)

other unit: degrees celsius ($^{\circ}\text{C}$)

converting from $^{\circ}\text{C}$ to K:

→ Add 273K to the value obtained in $^{\circ}\text{C}$.

type of instrument / apparatus:

1) Mercury thermometer

→ measures from -10°C to 110°C

→ each division measures 1°C

→ has accuracy of $\pm 0.5^{\circ}\text{C}$

↓ ↳ given 1 d.p

all temp readings
recorded to 1 d.p

2) Data Logger

→ temperature sensor connected to it

→ more accurate than mercury thermometer

→ able to record changing temperature over long periods of time.

→ e.g of use: measuring pH of different water sources, humidity of different environment

-mass-

SI Unit: Kilogram (kg)

Other Unit: grams (g), milligrams (mg)

In chemistry, smaller mass units are usually used.

Type of instrument/apparatus:

1) Beam Balance

2) Electronic Mass Balance

→ most accurate

→ use for smaller mass measurements

→ accurate of up to 0.01g (2 d.p)

→ used for precise and fast weighing.

-volume-

SI Unit: cubic meter (m^3)

Other Unit: cubic centimeter (cm^3), milliliter (mL)

conversion of units:

1 litre (L) = 1000 millilitres (mL)

1 litre (L) = 1 cubic decimeter (dm^3)

= 1000 cubic centimeter (cm^3)

1 cubic meter (m^3) = 1000 litres (L)

Type of instrument/apparatus:

1) Burette

→ accurately measures out the volume of a liquid to the nearest $\pm 0.10 \text{ cm}^3$, up to 2 d.p

→ scale marked (graduated) in 0.10 cm^3 divisions

→ used to deliver different volumes of liquids, e.g. 24.00 cm^3 / 38.90 cm^3

→ total volume of a burette = 50.00 cm^3

2) measuring cylinder

→ more accurate than a beaker

→ measures up to the nearest cm^3 , e.g. 99 cm^3 .

whole numbers

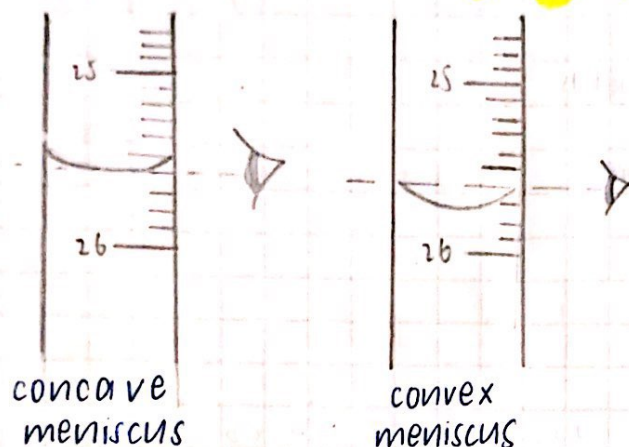
3) BEAKER

→ used to estimate the volume of a liquid, e.g. approximately 100cm^3 .

4) PIPETTE

→ accurately measures out fixed volumes of liquids, e.g. 5.0cm^3 , 10cm^3 , 20.0cm^3 or 25.0cm^3
↳ multiples of 5.

-Reading volume-



-collecting gases-

→ They tell us more about a reaction

→ They can be used in another reaction.

-Physical Properties-

[GASES]

1. Solubility of the gas

→ How soluble is the gas in water?

2. Density of the gas

→ How dense / heavy the gas is compared to the air?

* GAS

SOLUBILITY IN WATER

DENSITY TO AIR

1. Ammonia : extremely soluble : less dense
2. Carbon dioxide : slightly soluble : denser
3. Chlorine : soluble : denser
4. Hydrogen : not soluble : less dense
5. Hydrogen chloride : very soluble : denser
6. Oxygen : very slightly soluble : slightly denser
7. Sulfur dioxide : very soluble : denser

- Hydrogen, oxygen, carbon dioxide
↳ displacement of water

- chlorine, hydrogen chloride, sulfur dioxide
↳ downward delivery

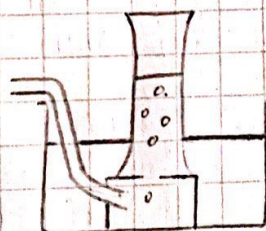
- ammonia
↳ upward delivery

collecting methods

[1]

displacement of water

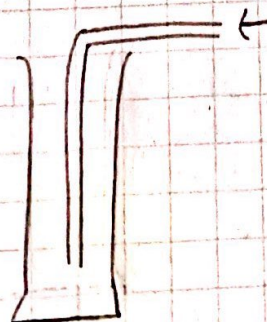
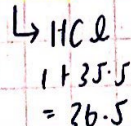
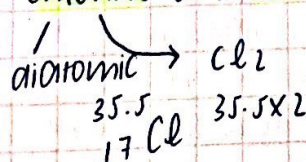
- used to collect gases that are insoluble / slightly soluble in water
- carbon dioxide, hydrogen & oxygen can be collected this way
- not for ammonia, chlorine & hydrogen chloride



[2]

downward delivery

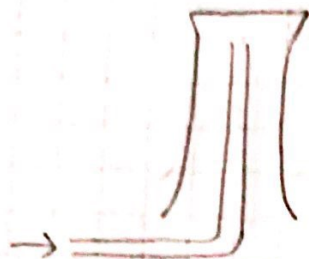
- used to collect gases that are soluble in water
- used to collect gases that are denser than air
- chlorine and hydrogen chloride can be collected this way.



[3]

upward delivery

- used to collect gases that are soluble in water
- used to collect gases that are less dense than air
- ammonia can be collected this way



dry gas -

dry gas → gas without the presence of water vapour.

How do we do it?

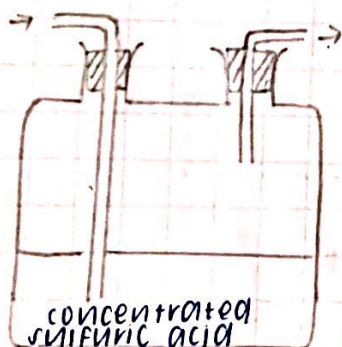
- by passing the gas through a drying agent
- the drying agent is to be used with our collection method.

3 commonly used drying agents:

- ① concentrated sulfuric acid → used to dry acidic gases
- ② quicklime (calcium oxide) → used to dry basic gases
- ③ fused calcium chloride
|
base

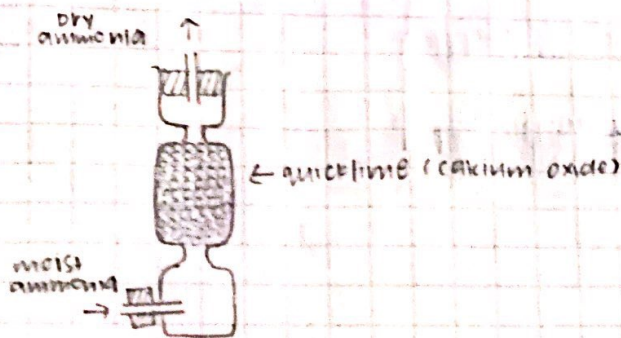
1.

concentrated sulfuric acid



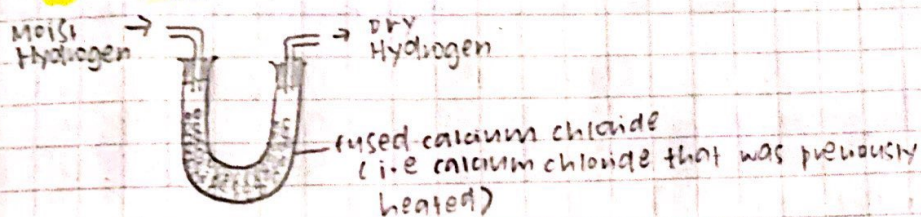
- used to dry many gases (except for ammonia)
- chlorine
- hydrogen chloride

2. Quicklime (calcium oxide)



→ we use quicklime to dry ammonia gas.

3. Fused calcium chloride

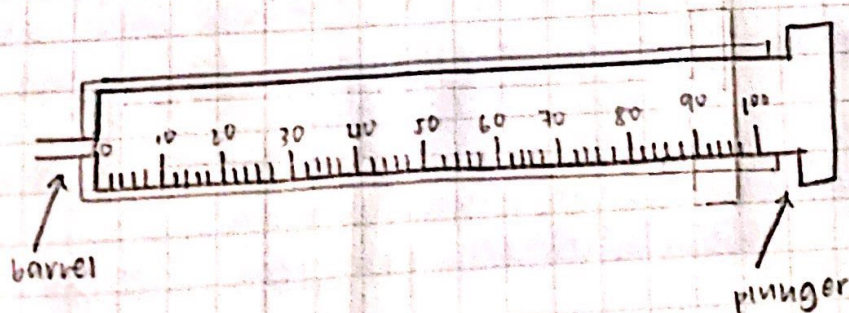


→ fused calcium chloride can be used to dry most gases.

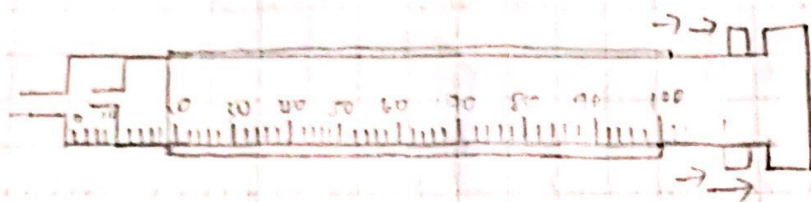
measuring volume of gas-

A gas syringe is used to measure the volume of a gas.
↳ measures a max volume of 100cm^3 .

(1) At the start, the plunger must be pushed down to expel all gases within the syringe.



② As gas is collected from an external source, the plunger is pushed outwards. When the reaction is completed, we can then read the vol of gas collected off the scale.



- determining purity -

What is a pure substance?

- made up of **only one substance** and is **not mixed** with **any substances**

e.g.: white diamonds

What is a mixture?

- **contains 2 / more other substances** **not chemically combined** together.

e.g.: Air (78% nitrogen, 21% oxygen, 1% others)

- identifying pure substances -

How can we identify?

- 1) look at **physical** properties
 - boiling point
 - melting point
- 2) look at **chemical** properties
 - chromatography

- Purity of Solids -

How do we tell if a solid substance is pure?

- a solid is **pure**
 - ↳ **exact & fixed melting point**
- ⇒ pure solid will melt completely at one temperature. And **not over a range of temperatures**

Example:

If we say pure solid water ice melts at 0°C . It will melt completely at 0°C .

↳ if it **does not** it means that **impurities are present**

impurities

- materials that must be removed in order to make a substance pure.

impurities affect the melting point in 2 ways:

- ① lower melting point of the substance.
 - ↳ ↑ amt of impurities, ↓ melting point of the substance
- ② cause melting to take place over a range of temperatures.

[PURITY OF LIQUIDS]

How do we tell if a liquid substance is pure?

- a liquid is pure if it has an exact & fixed boiling point
- ↳ a pure liquid will boil completely at one temperature, & not boil over a range of temperatures.

Example:

- if we say pure liquid water boils at 100°C . It will boil at 100°C
- ↳ if it does not it means that impurities are present

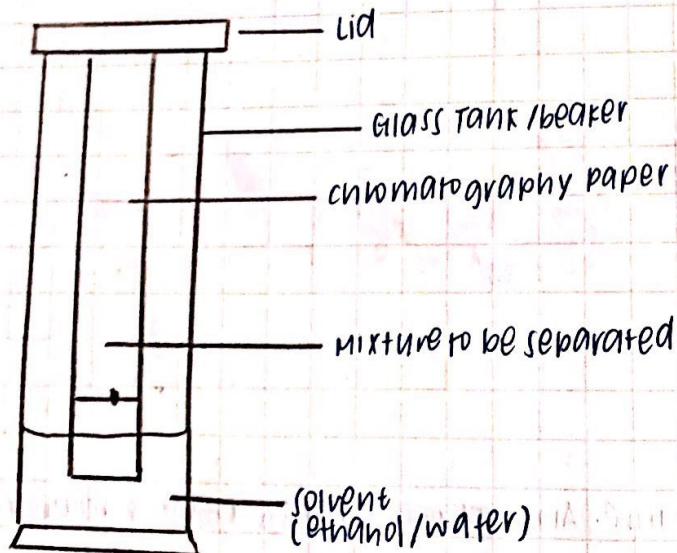
impurities affect the boiling point in 2 ways:

- ① increase the boiling point of the substance.
 - ↳ ↑ amt of impurities, ↑ the boiling point of the substance
- ② cause boiling to take place over a range of temperatures.

* pressure also has an effect on boiling point, the ↑ the surrounding pressure, ↑ the boiling point and vice versa.

Chromatography-

- technique of using a solvent to separate a mixture into its components.



uses of chromatography:

① separating components in a sample:

- separate dyes in ink
- separate pigments in plants
- separate amino acids from proteins

② identify components & no. of components present in sample:

- identify poisons / drugs
- detect traces of banned substances in food.

how does it work

[CHROMATOGRAPHY]

1. on the chromatography paper, draw a pencil line about 1.5cm from the end of the paper. ensure that the line drawn is high enough so that the solvent does not reach & cover it.

2. A spot of the sample is applied to the pencil line on the chromatography paper.

3. Allow the spot to dry, while doing so, prepare a beaker of ethanol / water the height of the liquid must be below the pencil line.

↳ because the sample spot will dissolve in the solvent and affect the accuracy of the results if submerged in the solvent.

4. place the prepared chromatography paper into the beaker of ethanol / water & place an evaporating dish on the top of the beaker.

↳ to prevent the solvent from evaporating, thus ensuring accuracy of results.

5. The chromatography paper soaks up the ethanol / water and as ethanol / water soaked up, it dissolves the sample and continues to travel up the paper carrying the sample with it.

→ if the sample is not very soluble in ethanol / water, it will not be carried up far along the paper.

→ if the sample is very soluble in ethanol, it will be carried far along the paper.

→ coloured spots will be left at different places on the paper at the end of the experiment.

why need use pencil?

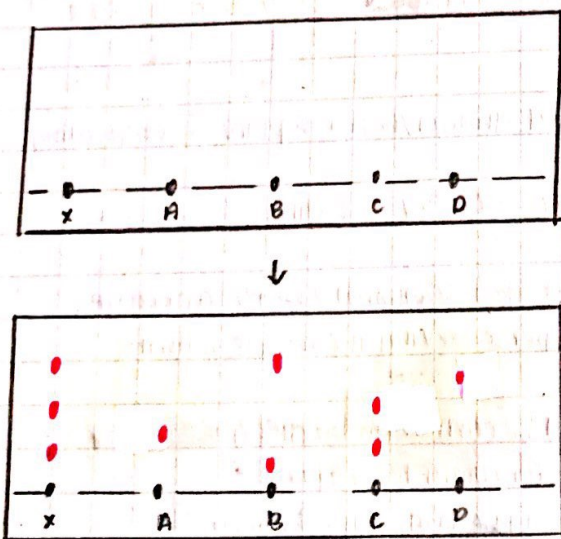
- Graphite is insoluble in water, therefore it won't dissolve in the solvent, affecting the accuracy of results.

interpreting results

- the chromatography paper with the separated components $\rightarrow$ chromatogram
- the chromatogram shows that the food colouring is a mixture of 2 dyes.
 $\rightarrow$ not pure
- A pure substance will only reveal one spot ^{above the start line} on the chromatogram.

analysing chromatograms

- the basis for the use of chromatography in analyzing samples is that identical dyes produce spots at the same height & in the same colour on the paper when the same solvent & temp is used.



$\therefore$ from the chromatogram, we can tell that both A and B are pure substances because there is only one spot for them. X has 3 components and that one of the component matches part of dye B. X however does contain dye C.

$\rightarrow$ If after performing the necessary steps for chromatography, you don't see any spots, it means the substance is insoluble in the solvent.

$\rightarrow$ change solvent.

- separating solid from liquid -

[a.] decanting

- used to separate a **dense, insoluble, solid** from a liquid.

example: **separating water from pebbles**



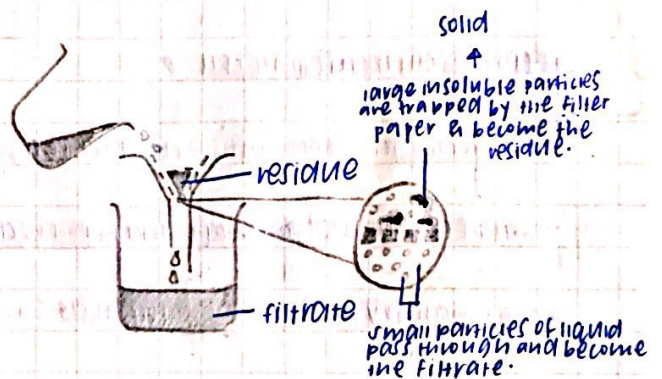
[b.] Filtration

- used to **separate (small solid particles)** from a liquid.

↓
sand, clay, dust particles
and **precipitates**

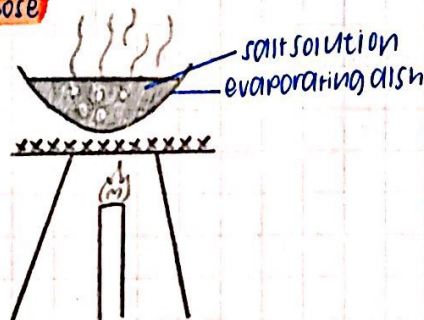
↓
small solid particles produced in a liquid by chemical reactions.

- a solid can be separated from a liquid by filtration because the **filter paper acts as a sieve**, where the **liquid is small enough to pass through the pores** (small holes) of the filter paper but the **solid cannot**.



[c.] Evaporation to dryness

- used to **recover salt from salt solution** (NaCl → sodium chloride & LiCl → lithium chloride)
- ↳ **can only be used if the substance can withstand high heat / temperature.**
- ↳ **otherwise, it will decompose**



- when a substance decomposes on heating / loses its water of crystallisation
→ do not evaporate it to dryness.

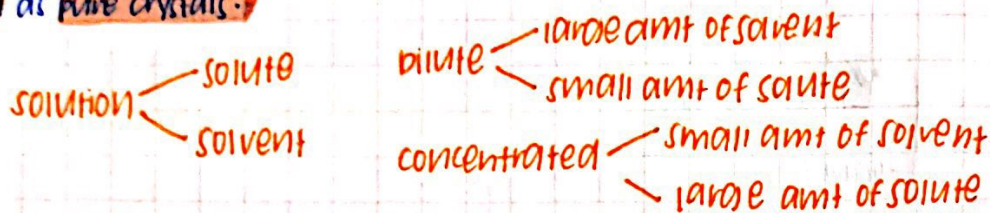
↳ sugar decomposes on heating.
⇒ crystallize it!

[d.]

crystallisation

mass
↑
max amt of solute in a fixed amt of solvent.
↳ volume

- water is removed by heating the solution.
- heating is stopped at the stage when a hot saturated solution is formed.
- if the remaining solution is allowed to cool to room temperature, the dissolved solid will be formed as pure crystals.



TEST FOR SATURATION POINT:

- a clear glass rod can be used to test whether a solution is saturated.
 - dipped into the solution and removed. There will be a small amt of solution on the rod.
 - if small crystals form on the rod as the solution cools, the solution is saturated.
- = solution is at its saturation point / crystallisation point

- separating solids -

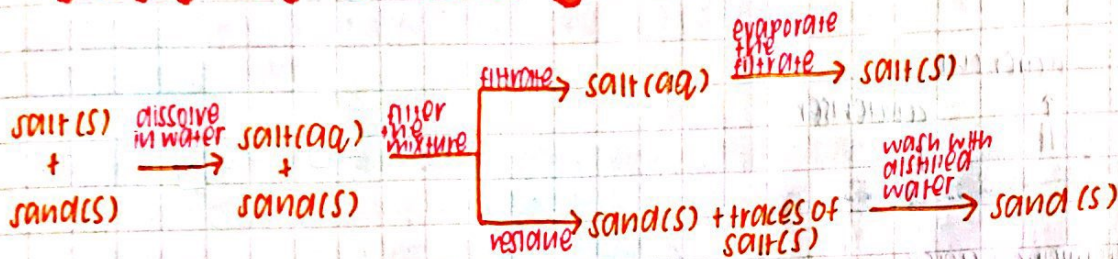
to separate 2 solids, it is important to be able to make use of what we know about the differences in their properties.

for e.g.: 2 solids can be separated by filtration if one of them is soluble in a solvent while the other is not.

procedure

1. pour some distilled water into the mixture of common salt & sand. stir & warm the mixture.
2. filter the mixture.
3. wash the residue with a little distilled water to remove all the salt solution from it. the residue is sand.
4. evaporate the filtrate to dryness.
5. salt is recovered.

summary - separating solids [separating salt from sand]

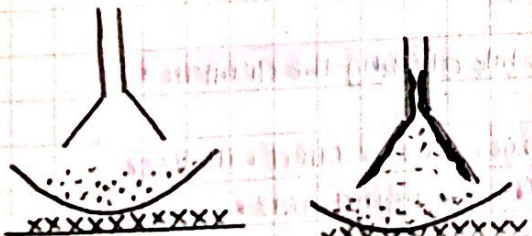


magnets

- some metals have magnetic properties
iron, nickel, steel?
- ↳ use these properties to separate metals from mixtures.

sublimation

- certain substances sublime when heated.
(ammonium chloride, iodine, dry ice [solid CO_2])
- ↳ use this property to separate a substance
- make sure that the other substance has a high melting point, lest it decomposes during this process.



- separating liquids -

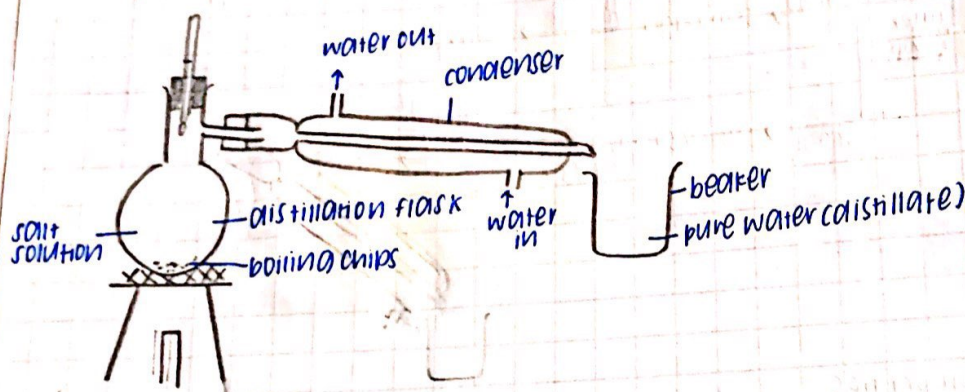
What is solution?

- when a solid or liquid dissolves in a solvent, a solution is formed.
- ↳ imp't that the solvent collected be pure
- ↳ simple distillation

Simple Distillation

What is distillation?

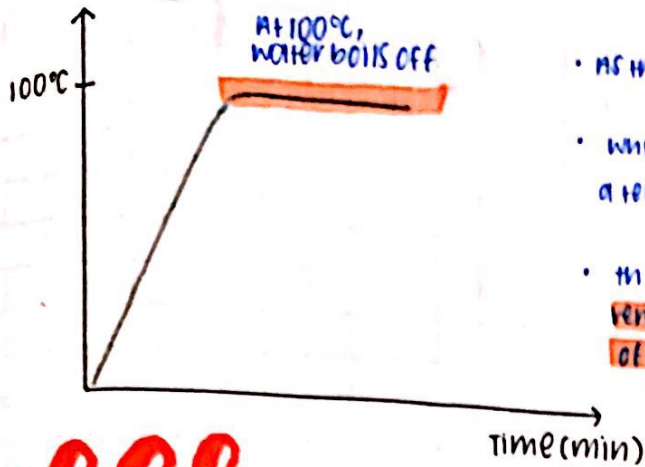
- the process of boiling a liquid and condensing its vapour.



Steps:

- boiling chips are placed in the round-bottom flask, with the solution.
- ↳ ensure that smooth boiling takes place.
- thermometer should be placed beside the side arm of the distillation flask.
- should not be dipped into the solution.
- ensures that the thermometer measures the boiling point of the substance that is being distilled.
- cold running water is allowed to enter the water jacket from the bottom of the condenser and leave from the top.
- ensure that maximum cooling of the vapour takes place while collecting the distillate.
- if the distillate is volatile, the receiver can be put in a large container filled with ice.
- keeps the temp of the distillate low so that it remains in the liquid state.

graph - simple distillation



- as the **sol. solution** is heated, its **temperature increases**
- when the **sol. finally boils**, the **thermometer records** a temp of **100°C**
- this is the temp of the **water vapour** & the temp **remains unchanged** until all the **water has boiled off**.

uses : - simple distillation

- **recover a solvent** from a **non-volatile solute** (sodium chloride / any substance that has a **high boiling point**.)
- **separate** the **mixtures of liquids** with **different boiling points**
 - ↳ **difficult to separate mixtures of liquids** whose **boiling points differ by less than 25°C**
 - ↳ **overcome** by using **fractional distillation**.

2 types of liquids

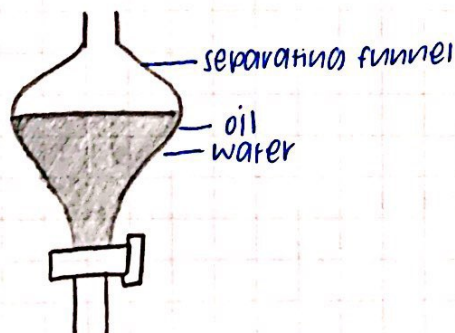
- ① immiscible liquids
- ② miscible liquids

immiscible: refers to **liquids that do not dissolve in each other**

miscible: **liquids that are soluble in / dissolve in each other**

immiscible liquids

to **separate immiscible liquids**, we need the **separating funnel**.



miscible liquids

- unlike oil & water, miscible liquids are liquids that mix together completely to form a sol.
 - ↳ water & ethanol.
 - ↳ fractional distillation

boiling VS evaporation

BOILING

- occurs at a fixed temp
- occurs throughout the liquid

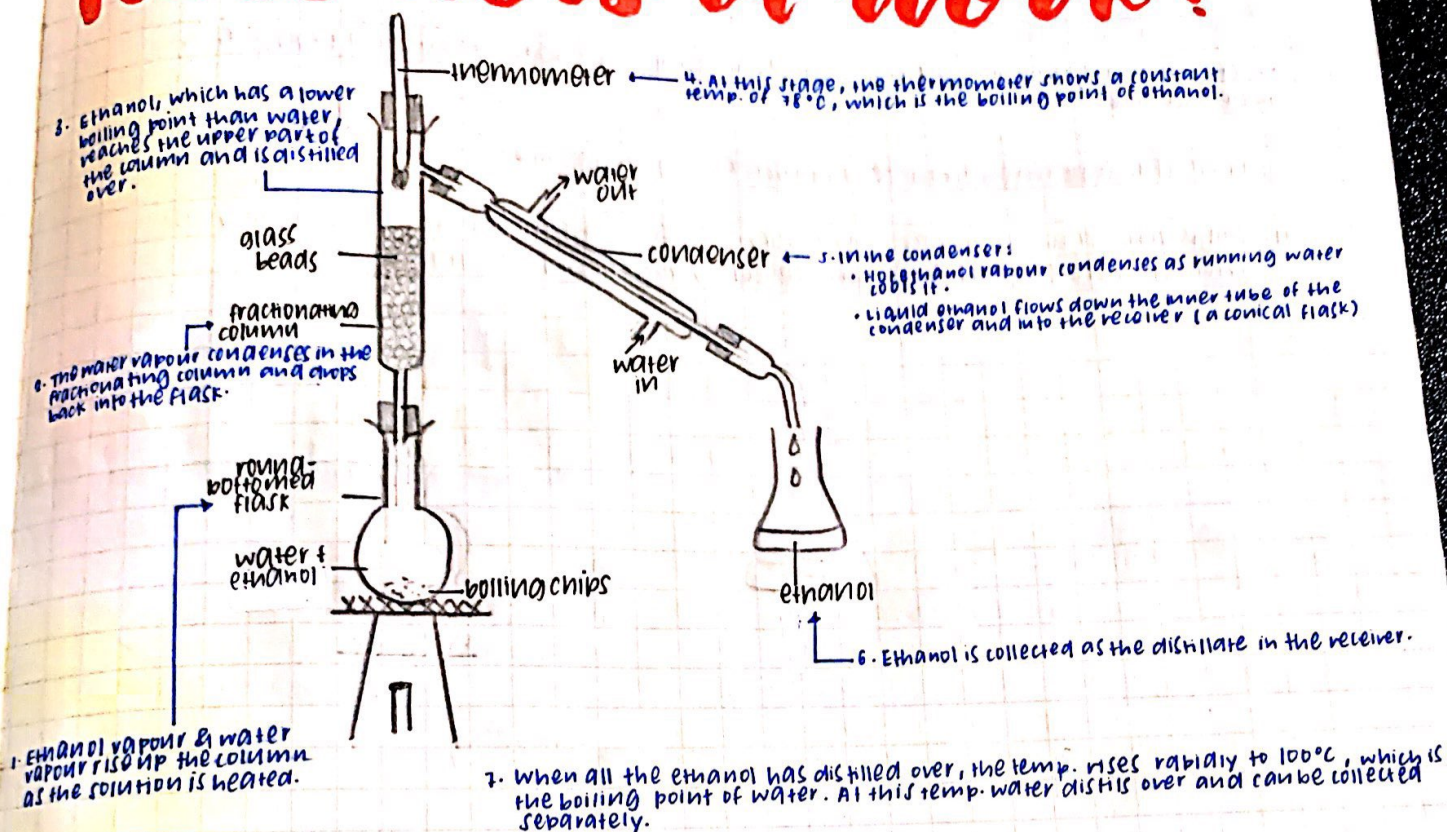
EVAPORATION

- takes place at any temperature
- occurs only at the surface of the liquid.

Fractional distillation

- A fractionating column, is attached to the round-bottomed flask and the condenser for fractional distillation.
- many glass beads in the fractionating column provide a large surface area for vapour to condense on.
- other than glass beads, a fractionating column may be filled with plates or a spiral.

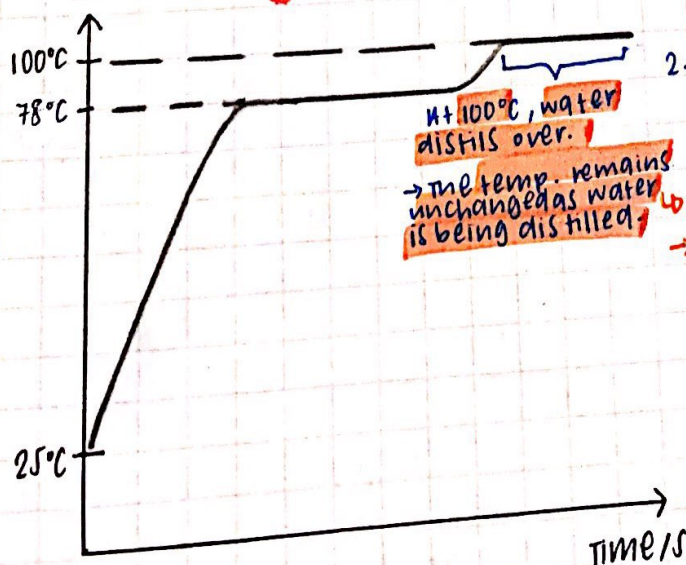
how does it work?



- the boiling point of ethanol is 78°C while the boiling point of water is 100°C .
- in fractional distillation, the liquid with lower boiling point (ethanol) will be collected as distillate first.

graph

- fractional distillation



1. The temp. of the mixture increases as it is heated.

2. At 78°C , ethanol distils over. The temp. remains constant until all the ethanol has distilled out of the round-bottomed flask.

→ recall: (KPT)
→ temp. remains constant during boiling

∴ heat energy is used to overcome F_{DA} between particles.

3. the temp. then increases until 100°C .

applications

- fractional distillation is used in industries to obtain nitrogen, argon and oxygen from liquefied air.
- can be used to separate mixtures of liquids such as crude oil
- ethanol is formed when glucose sol. undergoes fermentation in the presence of yeast. we can separate ethanol from glucose sol. by fractional distillation.

Kinetic Particle Theory

(chapter one)

CLASSIFICATION OF MATTER

- matter exists in 3 different physical states.
 - ↳ solid, liquid, gas.
- matter is made up of particles such as ions, atoms & molecules.
- the particles have kinetic energy and are in constant & random motion

CHARACTERISTICS OF MATTER

① Solid

- very closely packed in an orderly manner.
- vibrate & rotate at fixed positions.
- attractive forces: very strong
- fixed shape & volume

② Liquid

- disorderly manner & closely packed
- slide past each other & move randomly.
- attractive forces: strong
- no fixed shape & fixed volume

- ① gas
 - spread far apart from one another in disorderly manner.
 - moves about rapidly in any direction.
 - attractive forces: very weak / weak
 - no fixed shape / fixed volume.

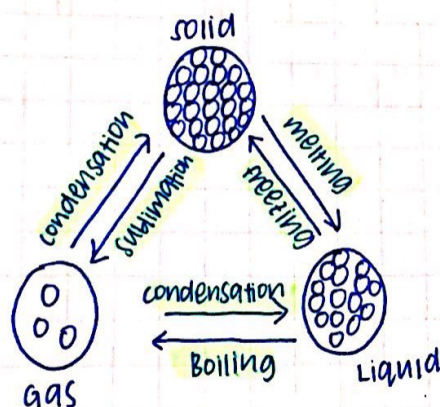
1001msd

USING KPT TO EXPLAIN COMMON PROPERTIES

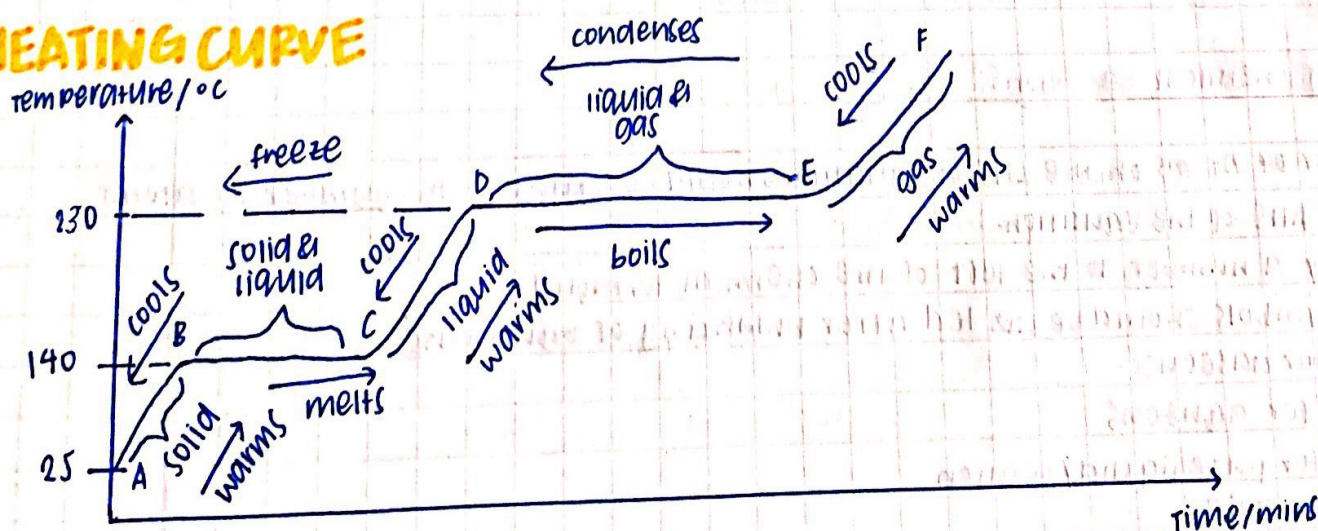
- ① volume & shape
 - arrangement of particles
 - attractive forces between particles.
- ② compressed
 - spaces between particles
- ③ Able to flow? / Take the shape & vol of container?
 - movement / attractive forces of particles.

CHANGE IN STATE

- as temp of the matter ↑, particles gained kinetic energy, distance between particles ↑ & the attractive forces between particles ↓.
- as temp ↓, kinetic energy ↓, distance between particles ↑, attractive forces between particles ↑.



HEATING CURVE



- mp : 140°C
- BP : 230°C

Why is the temp constant during melting & boiling?
 ↳ The heat energy gained is used to overcome the forces of attraction between the particles.

Chemical formulae and equations.

(Chapter seven)

Chemical Symbols of Ions

Metals	Valency	Symbols of Cations
Li, Na, K, H	1	Li^+ , Na^+ , K^+ , H^+
Be, Mg, Ca	2	Be^{2+} , Mg^{2+} , Ca^{2+}
Al	3	Al^{3+}
Iron(II), Iron(III) Fe^{2+} , Fe^{3+}	2, 3	Fe^{2+} , Fe^{3+}
Lead(II), Pb^{2+} , Zn	2	Pb^{2+} , Zn^{2+}
Copper(I), Copper(II) Cu^+ , Cu^{2+}	1, 2	Cu^+ , Cu^{2+}
Ag	1	Ag^+

Non-metals	Valency	Symbols of Anions	Name
F, Cl, Br, I	1	F^- , Cl^- , Br^- , I^-	Fluoride, chloride, bromide, iodide
H	1	H^-	Hydride
O, S	2	O^{2-} , S^{2-}	Oxide, sulfide

CASH'N & more

1. Carbonate $\rightarrow CO_3^{2-}$
2. Hydrogen carbonate $\rightarrow HCO_3^-$
3. Ammonium $\rightarrow NH_4^+$
4. Sulfate $\rightarrow SO_4^{2-}$
5. Sulfite $\rightarrow SO_3^{2-}$
6. Hydroxide $\rightarrow OH^-$
7. Nitrate $\rightarrow NO_3^-$
8. Nitrite $\rightarrow NO_2^-$

Balancing Chemical Equations

- (1) Number of Atoms on the LHS of equation should be equal to the number of atoms on the RHS of the equation.
- (2) Multiply a number to the left of the chemical formula.
- (3) State symbols should be added after balancing of equations:
 - (g) for gaseous
 - (aq) for aqueous
 - (l) for pure liquid/molten
 - (s) for solid

Acids, Bases, Salts (Chapter ten)

ACID

What is an acid?

1. An acid is a substance that produces hydrogen ions (H^+) when dissolved in water.
2. It is the hydrogen ions that are responsible for the properties of acid.

list of common acids and the ions they produce when dissolved in water:

Acid	Formula	Ions in solution
Hydrochloric Acid	HCl	H^+ Cl^-
Nitric Acid	HNO_3	H^+ NO_3^-
Sulfuric Acid	H_2SO_4	H^+ SO_4^{2-}
Ethanoic Acid	CH_3COOH	H^+ CH_3COO^-

*note:

- acids have the properties of an ionic compound when dissolved in water.
- behave like a covalent compound in other states of matter.

Physical properties of acids

1. Acids have a sour taste.
2. Acids conduct electricity in aqueous state.
3. Acids turn damp / moist blue litmus paper red.

Chemical properties - reaction of acids

1. acid + reactive metal $\rightarrow$ salt + hydrogen

Test for hydrogen gas:

Test $\rightarrow$ place a lighted splint at the mouth of the test tube.

observation $\rightarrow$ Hydrogen gas extinguishes the lighted splint with a 'pop' sound.

* unreactive metals: copper, silver, gold & lead.

2. acid + carbonate $\rightarrow$ salt + water + carbon dioxide

Test for carbon dioxide gas:

Test $\rightarrow$ bubble the gas into (limewater) $\rightarrow$ aqueous calcium hydroxide.

3. acid + base $\rightarrow$ salt + water (neutralisation)

metal oxide

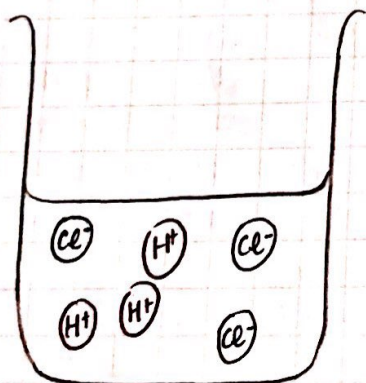
metal hydroxide

deriving the salts formed

Acid	Salt formed	conversion
Hydrochloric acid, (HCl)	Metal / Ammonium chloride	Chloric $\rightarrow$ chloride
Nitric acid, (HNO ₃)	Metal / Ammonium nitrate	Nitric $\rightarrow$ nitrate
Sulfuric acid, H ₂ SO ₄	Metal / Ammonium Sulfate	Sulfuric $\rightarrow$ sulfate

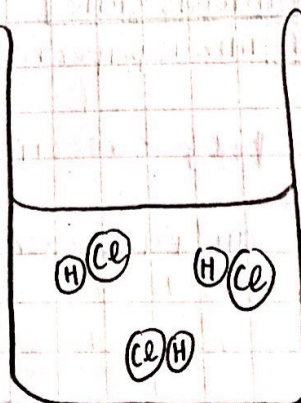
role of water in acidity

- Acids only show their acidic nature when dissolved in water. This is because they dissociate in water to produce the hydrogen ions.
- The hydrogen ions are responsible for the acid reactions.



in water:

Hydrogen chloride dissociates to give hydrogen ions.



in an organic solvent:

Hydrogen chloride does not dissociate.

i.e. acid properties are NOT displayed when an organic solvent is used.

Bases & Alkalines

what is a base?

- A base is a substance that reacts with an acid to form salt and water only.
→ A base come in the form of metal oxides / ammonium oxide or metal hydroxides / ammonium hydroxide.

relationship between bases & alkalis

- i. An alkali is a base that is soluble in water.
- ii. An alkali produces hydroxide (OH^-) ions when dissolved in water.

physical properties of alkalis

- i. They taste bitter.
- ii. They have a soapy feel.
- iii. Alkalis turn damp / moist red litmus paper blue.

Chemical properties of alkalis / bases

- i. acid + base $\rightarrow$ salt + water (neutralization)
ii. base + ammonium salt $\rightarrow$ salt + ammonia gas + water

test for ammonia gas (pungent smell)

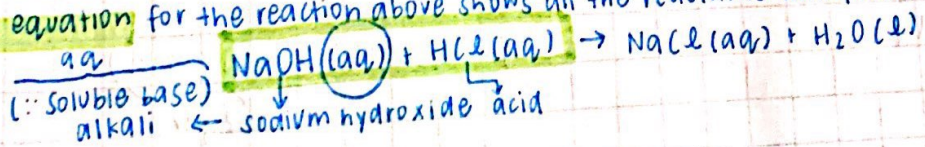
Test: Place damp red litmus paper at the mouth of the test tube.

Observation: damp/moist red litmus paper turns blue.

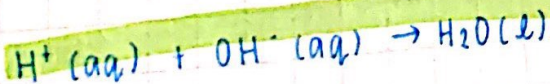
ionic equation for neutralisation reaction

- An ionic equation only shows the reaction between ions that produce a certain product.
eg: sodium hydroxide reacts with hydrochloric acid to form sodium chloride and water.

1. Chemical equation for the reaction above shows all the reactants and products:

$$2\text{HCl(aq)} + \text{Na}_2\text{CO}_3\text{(s)} \rightarrow 2\text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$$


2. The ionic equation for the reaction only shows the reaction between hydrogen ions & hydroxide ions to form water:
- $$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$$



This is why a reaction between an alkali and acid is called 'neutralisation': a neutral (pH 7) solution is formed.

pH scale and indicators

1. The pH scale is a scale that measures the acidity and alkalinity of a solution.
[one interpretation of pH is the power of hydrogen]

2. The pH scale

0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Acidic						Neutral		Alkaline						

The pH ranges from 0 to 14.

- i) $\text{pH} < 7$, the solution is acidic
- ii) $\text{pH} = 7$, the solution is neutral
- iii) $\text{pH} > 7$, the solution is alkaline

3. The lower the pH, the more acidic a substance is. The higher the pH, the more alkaline a substance is.

4. Strong acids VS weak acids

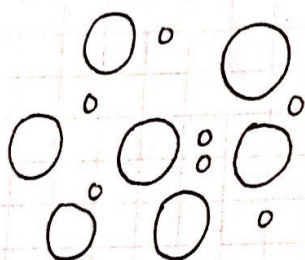
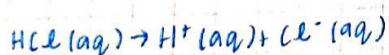
Strong acid

VS

Weak acid

- completely dissociates in water to form ions

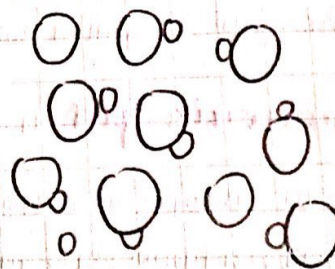
- only dissociates partially in water to form ions



All the hydrochloric acids dissociates to form H^+ and Cl^- ions.

pH of HCl is approximately 1.

E.g: HCl, HNO_3 , H_2SO_4



Most of the ethanoic acid remains in its CH_3COOH form.
Small amount of H^+ and CH_3COO^- ions are formed.

pH of ethanoic acid is around 5.

E.g: CH_3COOH (ethanoic acid)
vinegar

Indicators

To test for how acidic / alkaline a substance is, we use an indicator.

a) Universal Indicator

(i) The Universal Indicator is used to measure pH of solutions.

(ii) The colour of the universal indicator shows the pH of the solution.

0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
red			orange		yellow		green	blue			violet			

b) Examples of other indicators:

Indicator	Colour in acidic sol.	Colour in alkaline sol.	pH for colour change
Methyl Orange	Red	Yellow	3-5
Phenolphthalein	Colourless	Pink	8-10

c) Controlling acidity in soil

(i) The pH of soil affects the growth and development of plants.

(ii) Frequent adjustment of pH of soil is done by addition of chemicals.

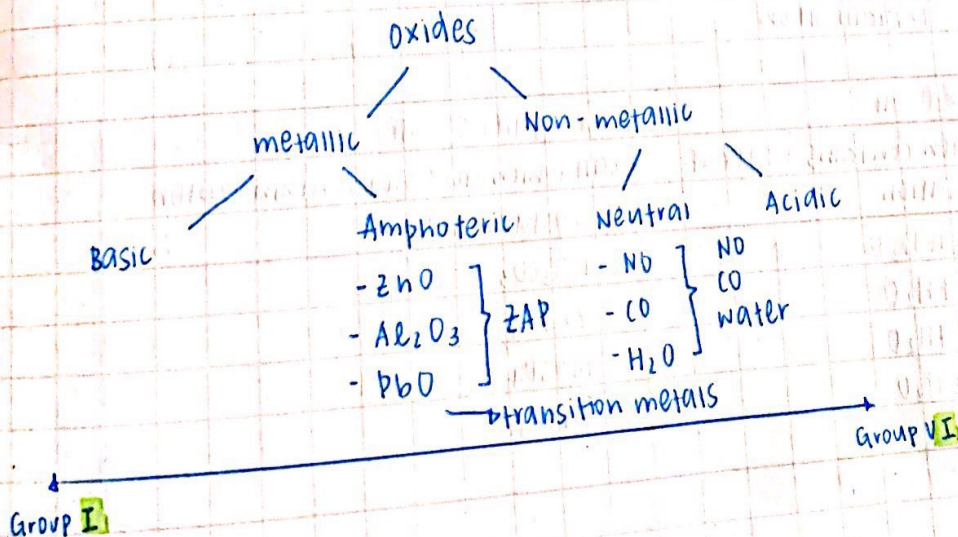
(iii) For soil that is too acidic, solid calcium hydroxide / slaked lime ($\text{Ca}(\text{OH})_2$) is added to increase its pH.

↳ 'liming' the soil.

Types of oxides

1. An oxide is a compound containing oxygen with another element.

2. There are 4 types of oxides: Basic, Amphoteric, Neutral, Acidic. (BANA)



3. **Basic oxides**
 - metal oxides
 - dissolve in water to form **alkaline solutions** ($\text{pH} > 7$)
 - react with acids to form **salt and water**.
4. **Amphoteric oxides** (ZAP)
 - a special group of **metal oxides** that can react with **both acids and bases** to form **salt and water**.
 - e.g. zinc oxide (ZnO), Aluminium oxide (Al_2O_3), Lead(II) oxide (PbO).
5. **Neutral oxides** (NO CO water)
 - non-metal oxides
 - do not react with acids/bases or display **acidic/basic properties**.
6. **Acidic oxides**
 - non-metal oxides
 - dissolve in water to form an **acidic solution** ($\text{pH} < 7$)
 - reacts with **alkalis** to form **salt and water**.

Salts (chapter eleven)

What are salts?

1. Salts are **ionic compounds** which consists of 2 parts, a **positive ion** and a **negative ion**.
 2. A salt is formed when the hydrogen ion of an acid is replaced by a **metal ion / ammonium ion**.
- cation → formed by metals
 can also be NH_4^+
- anion → formed by non-metals
- positive ions

water of crystallisation

1. Salts can combine with **water molecules** to form crystals.
- water of crystallisation

Type of salt	Hydrated salt	Anhydrous salt
Description	Salt which contains water of crystallisation	Salt which does not contain water of crystallization.
Examples	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	Na_2CO_3 ZnSO_4 MgSO_4 CuSO_4

Solubility Table

Salts	Soluble	Insoluble
Nitrates (NO_3^-)	All	None
Chlorides (Cl^-)	All except	- Lead(II) chloride - Silver chloride
Sulfates (SO_4^{2-})	All except	- Barium sulfate - Calcium sulfate - Lead(II) sulfate
Carbonates (CO_3^{2-})	- Group I carbonates - Ammonium carbonates	All others
Hydroxides (OH^-) - sparingly soluble (i.e. need a lot of solvent to dissolve it)	- Group I hydroxides - Ammonium hydroxide - Barium hydroxide - Calcium hydroxide ↳ Group II	All others

↳ i.e. barium hydroxide & calcium hydroxide can be solid / aqueous.

All Group I metal salts and ammonium salts are soluble and nitrates.

* note: metal oxides are insoluble EXCEPT SPAG oxides

turning the salts that are formed

A. For reactions with acids with:

- 1) Reactive metals to form salt and hydrogen
- 2) Ammonium hydroxides / metal hydroxides to form salt and water.
- 3) Ammonium oxides / metal oxides to form salt and water.
- 4) carbonates to form salt, carbon dioxide and water.

Acid used	Salt formed
- sulfuric acid, H_2SO_4	→ Metal / Ammonium sulfate
- hydrochloric acid, HCl	→ Metal / Ammonium chloride
- nitric acid, HNO_3	→ Metal / Ammonium nitrate

B. For reactions with alkalis:

- ALKALIS react with ammonium salts to produce salt, water and ammonia gas.

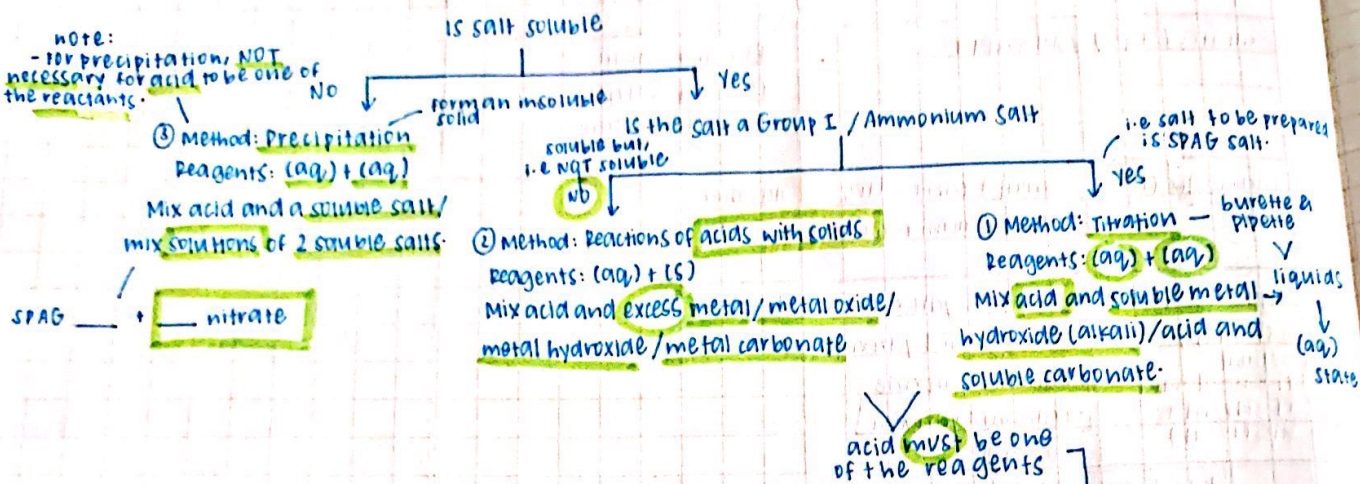
Ammonium salt	Chemical formula
- Ammonium chloride	→ NH_4Cl
- Ammonium nitrate	→ NH_4NO_3
- Ammonium sulfate	→ $(\text{NH}_4)_2\text{SO}_4$

methods of salt preparation

2 Questions:

- 1) Is the salt soluble?
- 2) If the salt is soluble, which category does it fall under?

note:
- for precipitation, NOT necessary for acid to be one of the reactants.



example 1:

describe the steps taken to obtain a sample of pure, dry potassium chloride crystals.

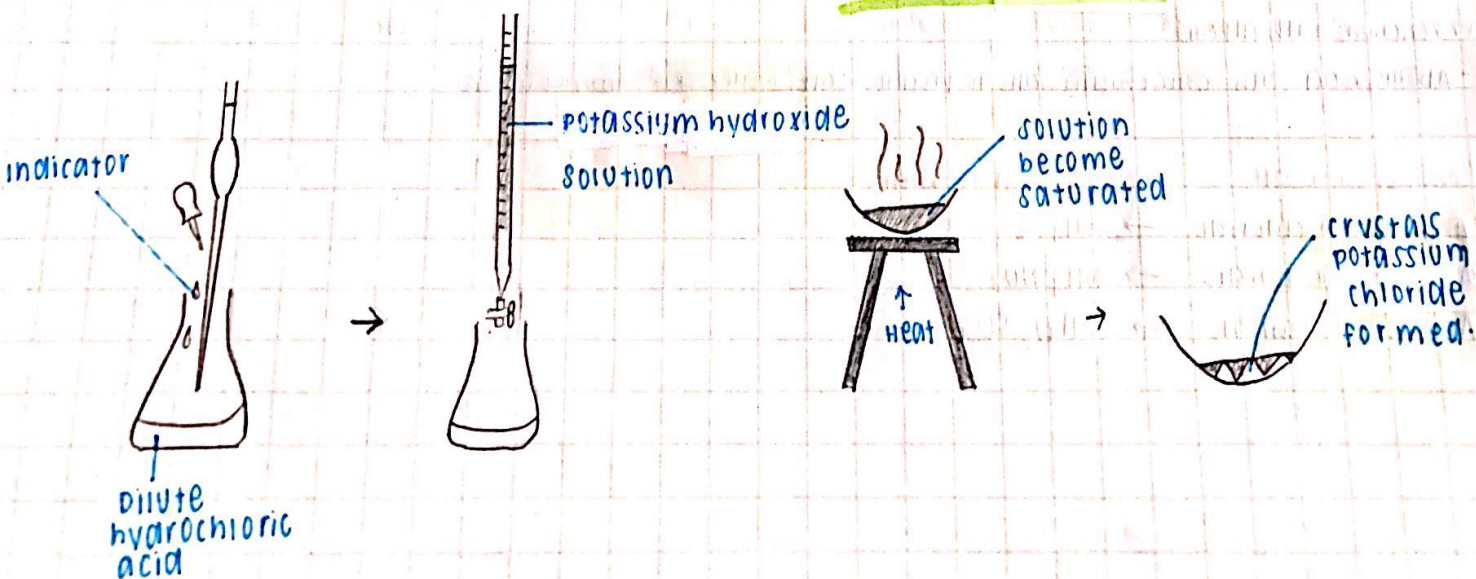
Here must be one of the reactants.

Salt Solubility : soluble

Method: Titration

Reagents: Acid: dilute hydrochloric acid

Alkali: Aqueous potassium hydroxide / Aqueous potassium carbonate.



- Step 1: Pipette 25.0 cm³ of dilute hydrochloric acid into a conical flask.
 Step 2: Add a suitable indicator. (used to indicate end-point of reaction) → via colour change.
 Step 3: Fill burette with aqueous potassium hydroxide / potassium carbonate.
 Step 4: Add aqueous potassium hydroxide / potassium carbonate to the hydrochloric acid until there is a permanent colour change. (i.e. rxn complete)
 Step 5: Record the volume of aqueous potassium hydroxide / potassium carbonate used.
 Step 6: Repeat steps 1-5 without the indicator. (to prevent contamination of the salt that is being prepared.)
 Step 7: Heat the potassium chloride solution until it is saturated.
 Step 8: Cool the solution and crystals form.
 Step 9: Filter the mixture to obtain crystals.
 Step 10: Dry crystals between filter papers.

→ Steps 1-6 = titration

→ Steps 7-10 = crystallisation

example 2:

Describe the steps taken to obtain a sample of pure, dry copper(II) sulfate.

Salt solubility: soluble

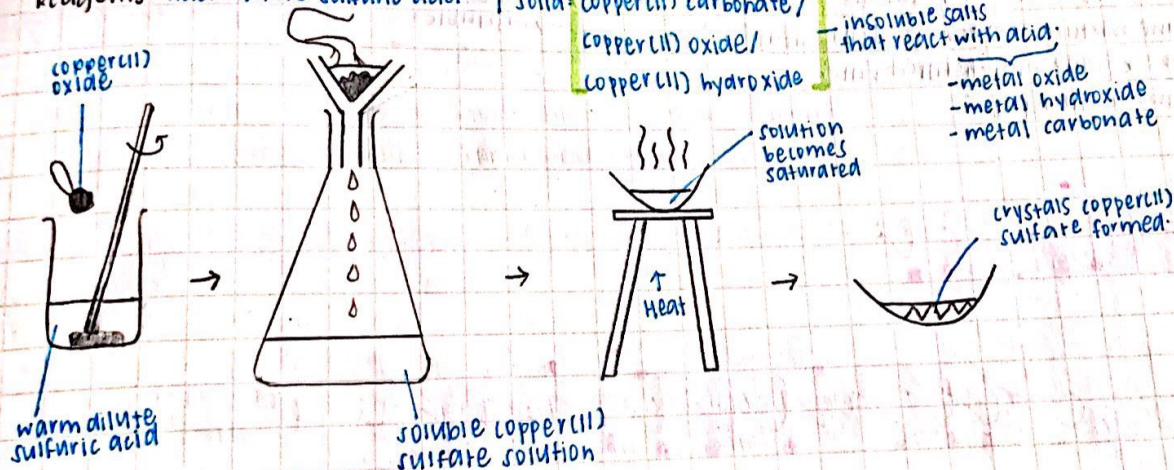
Method: reaction of acids with solids

Reagents: Acid: dilute sulfuric acid

Solid: copper(II) carbonate / copper(II) oxide / copper(II) hydroxide

insoluble salts that react with acid:

- metal oxide
- metal hydroxide
- metal carbonate



Step 1: Add * excess solid copper(II) oxide to dilute sulfuric acid.

Note: Excess solid is added to ensure that all the acid is reacted.

Step 2: Filter the mixture. (To remove the unreacted copper(II) oxide.)

Step 3: Collect the filtrate, which is copper(II) sulfate solution.

Step 4: Heat the filtrate until it is saturated.

Step 5: Cool the solution and crystals form.

Step 6: Filter the mixture to obtain crystals.

Step 7: Dry crystals between filter papers.

→ steps 4-7 = crystallisation

example 3:

Describe the steps taken to obtain a sample of pure, dry lead(II) sulfate.

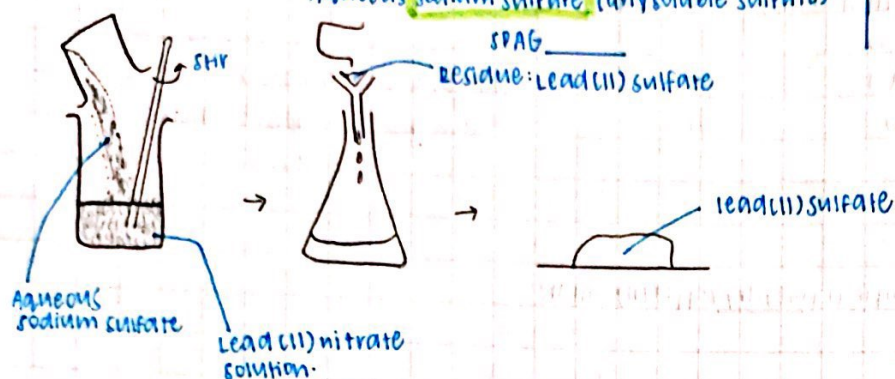
solubility: insoluble

Method: precipitation

Reagents: Solution A: dilute sulfuric acid /
Aqueous sodium sulfate (any soluble sulfate)

Solution B: Aqueous lead(II) nitrate.

(ive) nitrate



Step 1: React / Mix aqueous sodium sulfate with aqueous lead(II) nitrate.

Step 2: Filter the mixture to obtain lead(II) sulfate as the residue → insoluble.

Step 3: Rinse the residue with distilled water.

Step 4: Dry between filter papers.

Chemical

Calculations

(chapter
nine)

Relative Atomic Mass (A_r)

- Atoms also have very small masses, so it is not practical to use the actual masses of atoms in calculations.
- Chemists often compare masses of different atoms with the carbon-12 atom. (an isotope of carbon)
- The masses of all other atoms are compared with one-twelfth the mass of one carbon-12 atom.

definition: The relative atomic mass of any atom is the number of times the mass of 1 atom of an element is greater than $\frac{1}{12}$ of the mass of one carbon-12 atom.

$$\text{relative Atomic mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

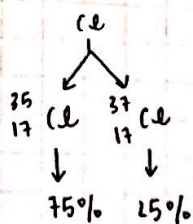
- symbol for relative mass is A_r .
- relative atomic mass is a ratio and has no units
- The relative atomic mass of each element is given in the periodic table.

Why are some Ar values not whole numbers?

- This is because such elements occur as mixtures of isotopes.

For example: Chlorine exists in two isotopic forms

→ Chlorine-35
→ Chlorine-37



$$\begin{aligned} \therefore \text{Relative Atomic Mass of Cl} &= (0.75 \times 35) + (0.25 \times 37) \\ &= 26.25 + 9.25 \\ &= 35.5 \end{aligned}$$

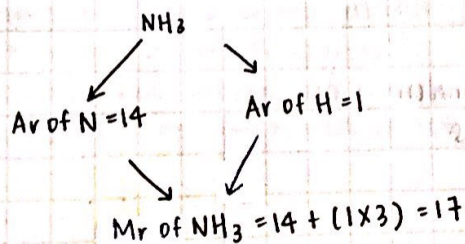
Relative Molecular Mass

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

$$\text{Relative Molecular Mass} = \frac{\text{Mass of 1 molecule}}{\text{Mass of } \frac{1}{12} \text{ of a carbon-12 atom}}$$

- Like A_r , it is a ratio and therefore has no unit.

How do we calculate the relative molecular mass of ammonia?



Relative Formula Mass

- Many substances such as water are covalent and exist as molecules.

- However, substances like sodium chloride are ionic and do not exist as molecules.

- The 'relative molecular mass' of an ionic compound is more accurately known as relative formula mass.

- Like relative molecular mass,

- relative formula mass has the symbol M_r and
- relative formula mass has no units.

The mole (chapter eight)

- Atoms are too small and numerous to be counted one at a time. Instead, the quantity of atoms is measured by mass.

- unit: moles, symbol: mol

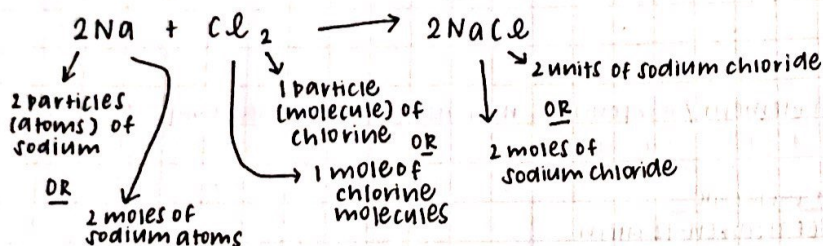
How many particles are there in a mole?

- There are approximately 6×10^{23} particles in one mole of substance.
- 6×10^{23} is called the Avogadro's constant / Avogadro's number.
- one mole of particles means 6×10^{23} (particles.)
 - atoms
 - molecules
 - ions
 - electrons

How do we convert between number of moles and number of particles?

$$\text{No. of moles} = \frac{\text{no. of particles}}{6 \times 10^{23}}$$

How is the mole related to chemical equations?



What is the mass of one mole of atoms of an element?

- The molar mass of an element is the mass of one mole of atoms of the element.

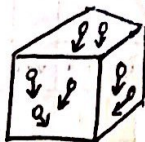
Element	Ar	Molar Mass
Aluminium	27	27g/mol
Carbon	12	12g/mol
Neon	20	20g/mol
Oxygen	16	16g/mol

A relationship between Ar & molar mass.

What is the relationship between mole and molar mass?

$$\text{No. of moles} = \frac{\text{mass of element (g)}}{\text{molar mass of element (g/mol)}}$$

Molar Gas Volume



one mole of any gas occupies 24 dm^3 (24000 cm^3) at r.t.p (25°C) and pressure (1 atm)

$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

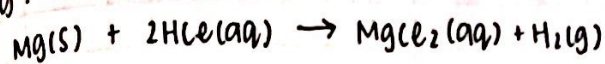
How do we calculate the number of moles of a gas?

$$\text{No. of moles} = \frac{\text{volume of gas in } \text{cm}^3 / \text{dm}^3 \text{ at r.t.p}}{24000 \text{ cm}^3 / 24 \text{ dm}^3}$$

Stoichiometry

- stoichiometry refers to the relationship between the amounts (in moles) of substance and products involved in a chemical reaction.

E.g.



based on stoichiometry,

1 mole of Mg reacts with 2 mole of HCl to form 1 mole of MgCl_2 and 1 mole of H_2

$$\text{No. of mole} = \frac{\text{mass}}{\text{molar mass}}$$

and molar mass of a substance has the same value as its M_r , we can say that:

24g of Mg reacts with 73g of HCl to form 95g of MgCl_2 and 2g of H_2

mole ratio of: $\text{Mg} : \text{HCl} : \text{MgCl}_2 : \text{H}_2$
1 : 2 : 1 : 1

- we can use a balanced equation to calculate mass of:

↳ reactant used up

↳ product formed in reaction

Steps for Chemical calculations:

- (1) calculate the number of moles for the reactant/product with given information.
- (2) state the mole ratio of all reactants and products from the balanced chemical equation.
- (3) calculate the number of moles for required reactant/product using mole ratio.
- (4) calculate the mass of the required reactant/product.

Volumes of Reacting Gases

- molar vol. of any gas at r.t.p and pressure = 24 dm^3
↳ vol. of a gas $\propto$ no. of moles

- we can use a balanced equation to calculate the volume of:

↳ gas used up.

↳ gas formed in a reaction.

Steps for chemical calculations:

- (1) calculate the no. of moles for the reactant/product with given info.
- (2) state the mole ratio of all reactants and products from the balanced chemical equation.
- (3) calculate the number of moles for required reactant/product using mole ratio.
- (4) calculate the volume of the required reactant/product.

Concentration of a solution

- concentration of a sol. is defined as the amount of a solute (g or mole) dissolved in a given vol. ($\text{cm}^3 / \text{dm}^3$) of a sol.

$$\textcircled{1} \text{ concentration (g/dm}^3\text{)} = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3\text{)}}$$

$$\textcircled{2} \text{ concentration (mol/dm}^3\text{)} = \frac{\text{no. of mol of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

$$\textcircled{3} \text{ concentration (mol/dm}^3\text{)} = \frac{\text{concentration (g/dm}^3\text{)}}{\text{molar mass of solute (g/mol)}}$$

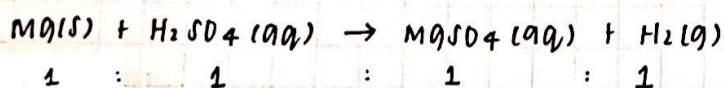
$\therefore$ conc. of a solution depends on

- ① the amt. of a solute (g or mole)
- ② vol. of a solution (dm^3)

Limiting Reactant

- limiting reactant refers to the reactant that is completely used up in the rxn.
 $\rightarrow$ determines (limits) the amt. of products formed.

e.g.



if 2 moles of Mg is made to react with 1 mole of sulfuric acid, there is extra Mg present, and Mg is said to be in excess.

$\rightarrow$ i.e. the reactant which will be completely used up in the reaction and determines when the rxn stops completely is H_2SO_4 . (limiting reactant)

where to find limiting reactant?

info. on both reactants given, & this info. allows you to calculate no. of mol for the reactants.

steps for calculations involved LR:

- ① calculate no. of moles present for both reactants
- ② identify LR (assume one reactant as LR, and compare against the other reactant using mole ratio)
- ③ Base subsequent calculations on no. of moles of LR present.

Volumetric Analysis

- volumetric analysis is used to find out the unknown concentration of a solution required to completely react with a solution of known concentration.

$\rightarrow$ involves the measurements of volume for both solutions.

$\rightarrow$ method used for volumetric analysis is titration.

How Titration is performed:

1. The titrant (sol. of known conc.) is placed in the burette. A fixed vol. of the sol. of the unknown conc. is placed into a conical flask using a pipette.
2. 1 or 2 drops of indicator is then added into the conical flask containing the sol. of unknown conc.
3. The titrant is then allowed to react completely with the sol. of unknown conc., i.e. when the end-point is reached.
 - a. End-point is reached when the indicator permanently changes colour.
4. Volume of the titrant used is then noted. Conc. of the unknown sol. can now be calculated.

Energy changes (chapter fifteen)

HANDS-ON EXPERIMENT

Reaction	Fill half a boiling tube with water & add 1 spatula of sodium carbonate	Fill half a boiling tube with water & add 1 spatula of ammonium chloride	Fill a boiling tube with 2cm ³ of HCl, followed by an equal amt. of sodium hydroxide
Boiling tube feels warm/cold aft. rxn	warm	cold	warm
Temp. of rxn mixture increase/decrease	increase	decrease	increase
Endothermic / Exothermic rxn.	exothermic	endothermic	exothermic

IMPORTANT

- A thermometer measures the temperature of the surroundings, not the reactant particles.
* unless an specifies & asks for reactant / reacting particles.
- Surroundings include: container (containing chemicals), water and air.

EXOTHERMIC REACTIONS

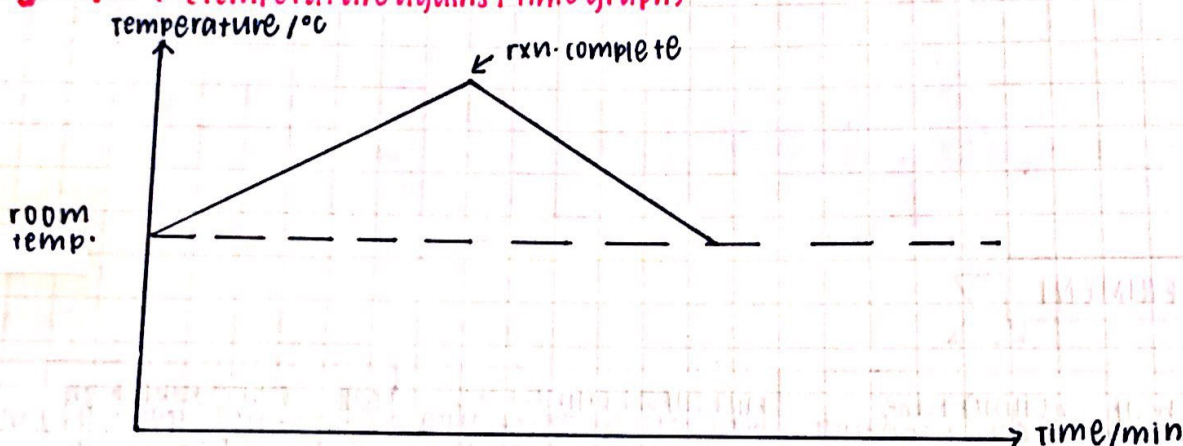
- temperature increases (of surroundings)
- Heat energy is lost or given out / released to the surroundings.

examples:

- burning

1. combustions reactions
2. change of state: freezing of liquids
3. Respiration
4. Neutralisations reactions between alkalis and acids
5. change of state: condensation of gases

graph: (temperature against time graph)



ENDOTHERMIC REACTIONS

- temperature decreases (of surroundings)
- Heat energy is taken in or absorbed from the surroundings

examples:

1. decomposition reactions
2. change of state: melting of solids
3. photosynthesis
4. Dissolving / dissolution of solid ammonium chloride
5. change of state: Boiling of liquids

graph : (temperature against time graph)

