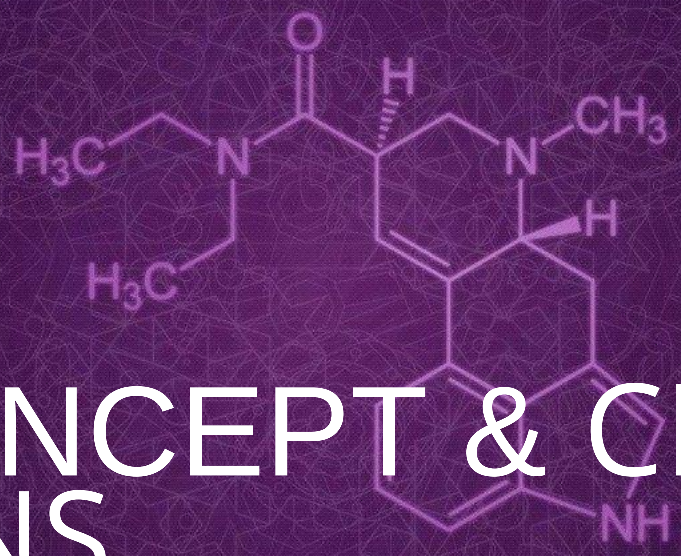


TOPIC 3:
MOLE CONCEPT & CHEMICAL EQUATIONS

THE ABOUT

CHAPTER ANALYSIS

TIME

- Need to practice a lot
- 5 **key** concepts

EXAM

- Heavily tested
- Tested as add-on to other chapters
  → Acid & Bases, Electrolysis etc…

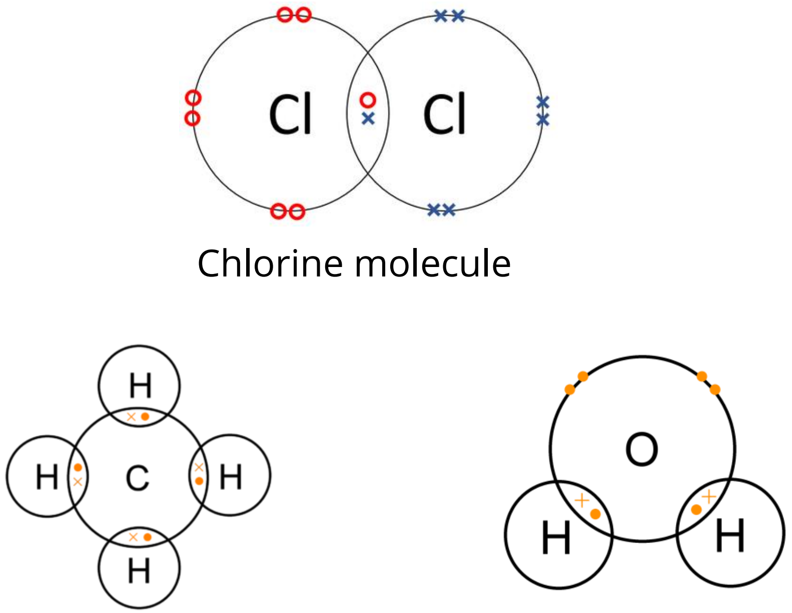
WEIGHTAGE

- Heavy overall weightage
- Constitute to **5.5%** of marks for past 5 year papers

KEY CONCEPT

CHEMICAL EQUATION
CHEMICAL FORMULA
BALANCING CHEMICAL EQUATION
IONIC EQUATION

CHEMICAL FORMULA



Chlorine molecule

Methane compound

Water compound

COVALENT COMPOUNDS

For simple covalent molecules, most elements exist as diatomic molecules as they undergo covalent bonding to **achieve the stable noble gas configuration.**

Prefixes are also commonly used to name compounds.

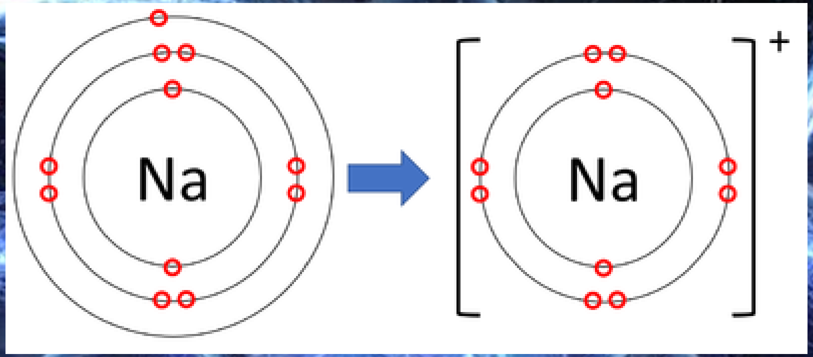
Prefix:

Mono – 1
Di – 2
Tri – 3
Tetra – 4
Pent – 5

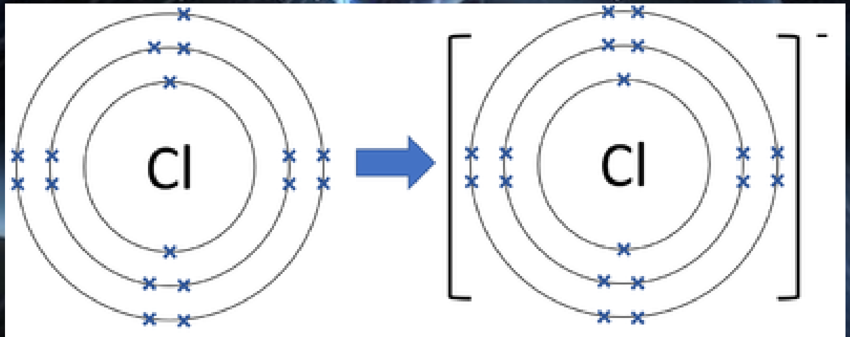
For example,
Carbon monoxide – CO
Carbon dioxide – CO_2

CHEMICAL FORMULA

Cation:



Anion:



IONIC COMPOUNDS

Ionic compounds are **electrically neutral**, where total positive charge will be equal to the total negative charge.

As such the positive charge(s) on the cation(s) and the negative charge(s) on the anion(s) in the compound must be **balanced out**.

Some common anions:

Carbonate CO_3^{2-}
Nitrate NO_3^-
Sulfate SO_4^{2-}
Phosphate PO_4^{3-}
Chloride Cl^-

Forming of ionic compounds:

For example,

Cation: Mg^{2+}
Anion: NO_3^-

To balance out charges,

1 x Mg^{2+} & 2 x NO_3^-

Compound:

$Mg(NO_3)_2$

CHEMICAL EQUATION

State of substance	State symbol	Usage
Solid	**(s)**	For substances with a high melting point. Most metals, ionic compounds and elements/compounds with giant molecular structure fall under this category. E.g. Al (s) or Si (s) or NaCl (s)
Liquid	**(l)**	For substances that are solvents. E.g. H_2O (l)
Gaseous	**(g)**	For substances with a low melting and boiling points. E.g. H_2 (g), Br_2 (g)
Aqueous	**(aq)**	For substances that dissolve in water to form ions. Most ionic compounds fall under this category. Acids and alkalis are of aqueous state. E.g. KOH (aq), H_2SO_4 (aq)

STATE SYMBOLS

Solid (s)
Liquid (l)
Gaseous (g)
Aqueous (aq) – solution form, water was added.

BALANCING EQUATIONS

When balancing equations, ensure that the number of atoms for each element is equal on both side.

Add a **coefficient** in front of the compound when balancing the equation.

For example,

$$CaCl_2\ (aq) + Na_2CO_3\ (aq) \rightarrow CaCO_3\ (s) + \mathbf{2}\ NaCl\ (aq)$$

$$CaCO_3\ (s) + \mathbf{2}\ HCl\ (aq) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

$$Fe_2O_3\ (s) + \mathbf{3}\ CO\ (g) \rightarrow \mathbf{2}\ Fe\ (s) + \mathbf{3}\ CO_2\ (g)$$

Practice makes perfect!

IONIC EQUATION

An ionic equation is a **simplified chemical equation that shows only the ion**s that take part in a chemical reaction in an aqueous solution.

Only ionic compounds that are dissolved to form an aqueous solution can be written as ions.

Ions that remain in aqueous state on both the left hand side and the right hand side of the equation are known as spectator ions.

Step 1

Write the balanced chemical equation for the reaction. Include the state symbols.

$$Ba(NO_3)_2\ (aq) + CuSO_4\ (aq) \rightarrow BaSO_4\ (s) + Cu(NO_3)_2\ (aq)$$

Step 2

Breakdown the chemical equation in terms of ions for substances in the aqueous state. Balance the chemical equation.

$$Ba^{2+}\ (aq) + 2NO_3^-\ (aq) + Cu^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s) + Cu^{2+}\ (aq) + 2NO_3^-\ (aq)$$

Step 3

Cancel the spectator ions.

Ba^{2+} (aq) + ~~$2NO_3^-$ (aq)~~ + ~~Cu^{2+} (aq)~~ + SO_4^{2-} (aq) → $BaSO_4$ (s) + ~~Cu^{2+} (aq)~~ + ~~$2NO_3^-$ (aq)~~

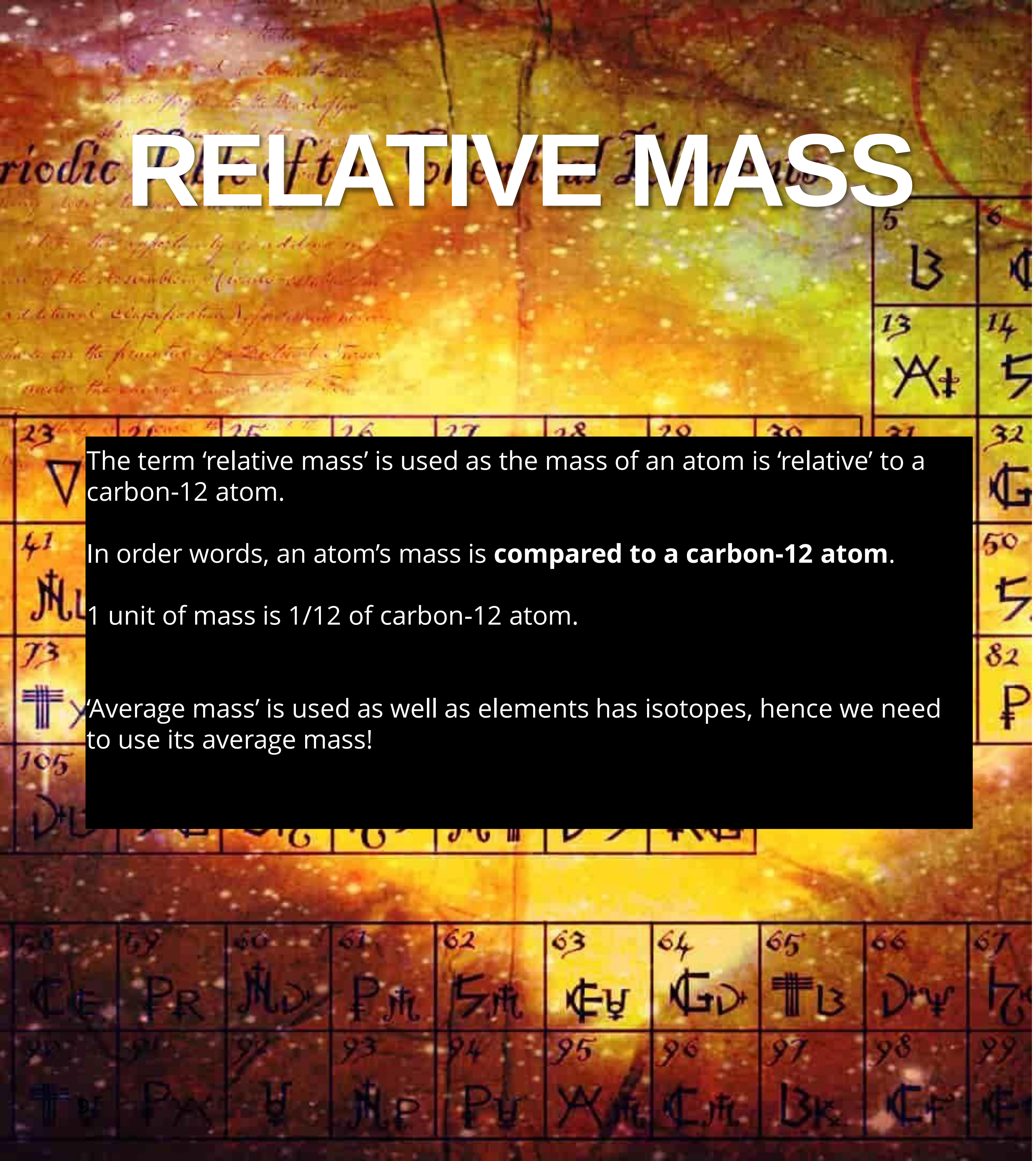
Step 4

Rewrite the equation without the spectator ions.

$$Ba^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s)$$

KEY CONCEPT

MOLE CONCEPT
Ar, Mr
MOLE
CONCENTRATION



RELATIVE MASS

The term ‘relative mass’ is used as the mass of an atom is ‘relative’ to a carbon-12 atom.

In order words, an atom’s mass is **compared to a carbon-12 atom**.

1 unit of mass is 1/12 of carbon-12 atom.

‘Average mass’ is used as well as elements has isotopes, hence we need to use its average mass!

Relative atomic mass (Ar)

It is the **average mass** of one atom of that element **compared to 1/12 of the mass of one carbon-12 atom.**

Relative molecular mass (Mr)

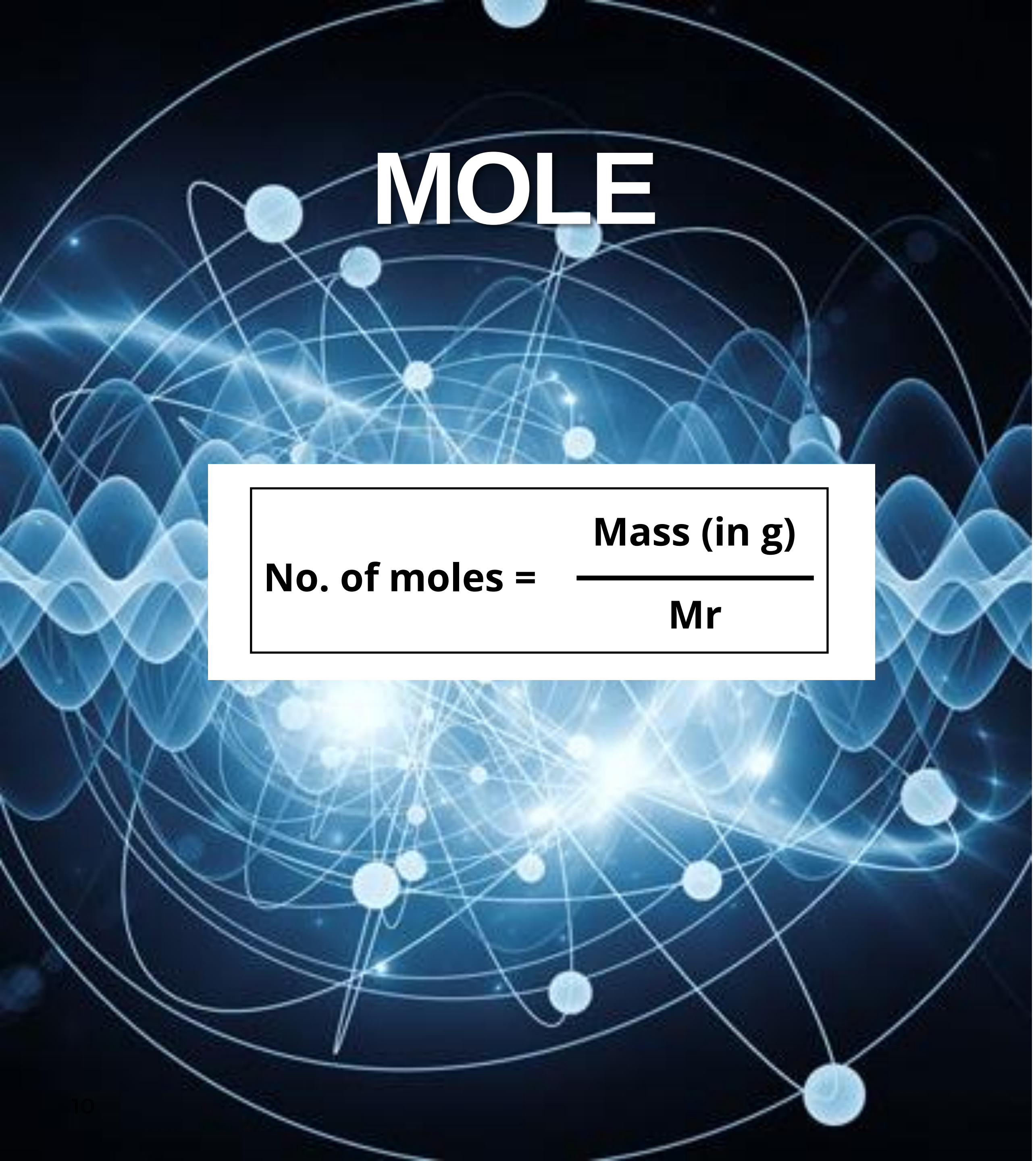
It is the **average mass** of one molecule of the substance **compared to 1/12 of the mass of one carbon-12 atom.**

*Carbon–12 is used as the point of reference as it is the most commonly available element on Earth.

*There are **no units** for Ar & Mr as it is relative mass to Carbon-12, effectively a ratio of comparison.

Percentage by mass of an element in a compound:

$$\frac{\text{Ar x (no. of atoms)}}{\text{Mr of compound}} \text{ X } 100\%$$



MOLE

$$\text{No. of moles} = \frac{\text{Mass (in g)}}{\text{Mr}}$$

<u>WHAT IS MOLE?</u>

One mole of any substance contains 6.02×10^{23} particles.

The value $\mathbf{6.02 \times 10^{23}}$ is called Avogadro's constant.

$$\textbf{No. of particles = mole x } \mathbf{6.02 \times 10^{23}}$$

<u>MOLAR VOLUME OF GASES</u>

At room conditions (25 °C and 1 atmosphere), one mole of gas has a volume of **$24\ dm^3$ or $24\ 000\ cm^3$.**

All gases have the same molar volume, regardless of their chemical formula & Mr.

$$\textbf{1 mole of gas} = \mathbf{24dm^3}$$

Concentration

$$\text{Concentration} = \frac{\text{Mole / mass}}{\text{volume}}$$

No. of moles = Concentration x volume

CONCENTRATION

Concentration of a solution refers to the amount of solute in a set amount of solution.

Concentration is typically measured in two ways:

1) The mass (in grams) of solute in 1 dm^3 of a solution (**gdm^{-3}**).
2) The number of moles of solute in 1 dm^3 of solution (**$moldm^{-3}$**).

Example:
Calculate the mass of solute in 300 cm^3 of 0.5 $moldm^{-3}$ copper(II) sulfate solution.

Volume of solution = 300 cm^3 = 0.30 dm^3

Number of moles of $CuSO_4$
= Concentration ($moldm^{-3}$) × Volume of solution (dm^3)
= 0.5 × 0.30
= 0.15 mol

Mass of $CuSO_4$
= Number of moles (mol) × Molar mass ($gmol^{-1}$)
= 0.15 × [64 + 32 + 4(16)]
= 24 g

KEY CONCEPT

STOICHIOMETRY
LIMITING REAGENT
PERCENTAGE YIELD & PERCENTAGE PURITY
EMPIRICAL/MOLECULAR FORMULA

STOICHIOMETRY

STOICHIOMETRY FOR GAS

Avogadro's Law states that the **volume of a gas is directly proportional to the number of moles** if temperature and pressure are constant.

Thus, the **ratio of moles of substances in a chemical equation** also provides information about the **ratio of the volumes of gases in the reaction**.

$$N_2(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

100 cm^3 of N_2 will react with **300 cm^3 of H_2** to produce **200 cm^3 of NH_3**.

STOICHIOMETRY

Example:
Find the mass of hydrogen gas formed when 48g of magnesium metal is reacted with excess hydrochloric acid.

Step 1: Write out the balanced equation.

$$Mg\,(s) + 2\,HCl\,(aq) \rightarrow MgCl_2\,(aq) + H_2\,(g)$$

Step 2: Calculate the number of moles of Mg reacted.

Number of moles of Mg reacted = mass / Mr
= 48 / 24
= 2

Step 3: Determine the molar ratio.

Number of moles of Mg reacted	:	Number of moles of H_2 produced
1	:	1
2	:	2

Step 4: Calculate the mass of H_2 produced.

Mass of H_2 produced = Mole x Mr
= 2 x 2
= 4.0 g

LIMITING REAGENT

VISUALISE THIS

For a car to be assembled, each car body must be assembled with 4 wheels.

1 car body + 4 wheels → 1 full car

If I have 10 car bodies & 12 car wheels, how many full car can I form?

Answer: 3 full cars

Hence, **the wheels are the limiting reagent as it 'limits' further reaction as there is an 'excess' of car bodies.**

LIMITING AND EXCESS REACTANTS

Reactions do not always use the exact amount of reactants as given by the equation.

A reaction is unable to proceed if one reactant is used up even if the other reactants are in excess.

The **limiting reactant** (or reagent) is the reactant that is **completely used up** in a chemical reaction. It determines or limits the amount of product formed.

The **excess reactant** (or reagent) is the reactant that **still remains** when the limiting reactant has been completely reacted away and the chemical reaction stops.

Example:

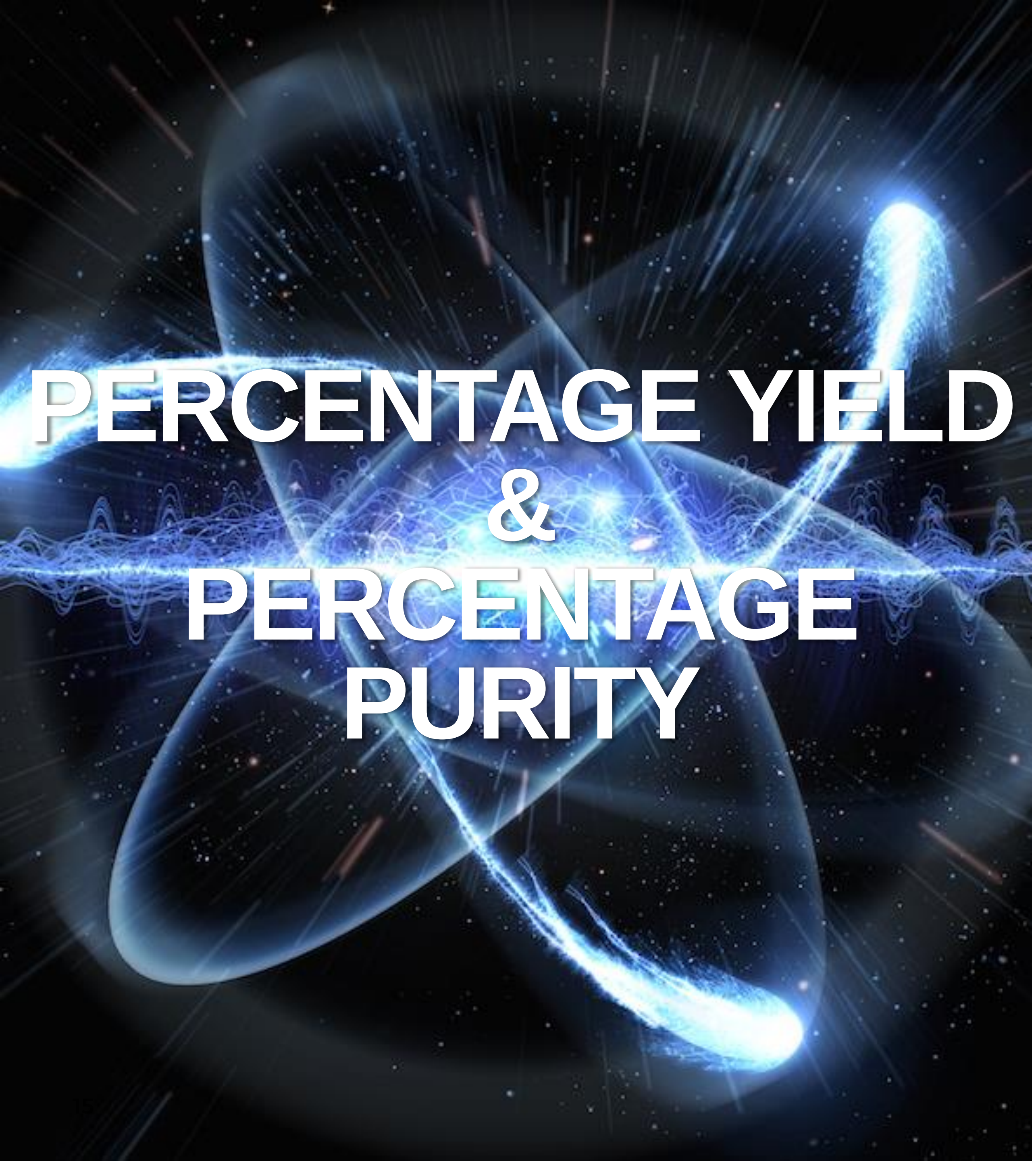
$$Pb(NO_3)_2\,(aq) + 2\ NaCl\,(aq) \rightarrow PbCl_2\,(s) + 2\ NaNO_3\,(aq)$$

1 mole of $Pb(NO_3)_2$ reacts with 2 mole of NaCl.

Hypothetically, let's say there is 1 mole of $Pb(NO_3)_2$ & 5 moles of NaCl.

As there is only 1 mole of $Pb(NO_3)_2$, even if there is 5 moles of NaCl, only 2 mole of NaCl will react.

$Pb(NO_3)_2$ is the limiting reagent while NaCl is the excess reactant.



PERCENTAGE YIELD

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Actual yield refers to the actual amount of product obtained from the experiment.

Theoretical yield refers to the maximum amount of yield that is expected if the reaction went to completion and there were no experimental errors.

PERCENTAGE PURITY

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

EMPIRICAL FORMULA

EMPIRICAL FORMULA

The empirical formula of compounds can be determined from the **mass of the constituent elements** that form the compound.

If values of Mr is given, the **molecular formula** can be determined.

→ Multiply by appropriate ratio to increase empirical formula to match the Mr.

Example (by mass):

A 0.80 g sample of calcium was burnt in the presence of oxygen to give an oxide of calcium.
When the calcium was completely burnt, the oxide was weighed and has a mass of 1.12 g.

Determine the empirical formula of this oxide.

Mass of calcium = 0.80 g
Mass of calcium oxide produced = 1.12 g
Mass of oxygen reacted = 1.12 – 0.80 = 0.32 g

	Calcium (Ca)	Oxygen (O)
Mass in sample/g	0.80	0.32
Molar mass/g mol^{-1}	40	16
Number of moles	0.80 / 40 = 0.02	0.32 / 16 = 0.02
Simplest ratio	1	1

Therefore, empirical formula is CaO.

If mass of molecule is 112, n = 112 / (40+16) = 2

Molecular formula is Ca_2O_2

EMPIRICAL FORMULA

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The empirical formula of compounds can be determined from the **mass of the constituent elements** that form the compound.

If values of Mr is given, the **molecular formula** can be determined.

→ Multiply by appropriate ratio to increase empirical formula to match the Mr.

Example (by percentage composition):

A compound contains 27.1% sodium, 16.5% nitrogen and 56.5% oxygen. Determine the empirical formula of this compound.

	Sodium (Na)	Nitrogen (N)	Oxygen (O)
Mass in 100 g sample/g	27.1	16.5	56.5
Molar mass/g mol^{-1}	23	14	16
Number of moles	27.1 / 23 = 1.178	16.5 / 14 = 1.179	56.5 / 16 = 3.531
Simplest ratio	1	1	3

Therefore, empirical formula is $NaNO_3$.