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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

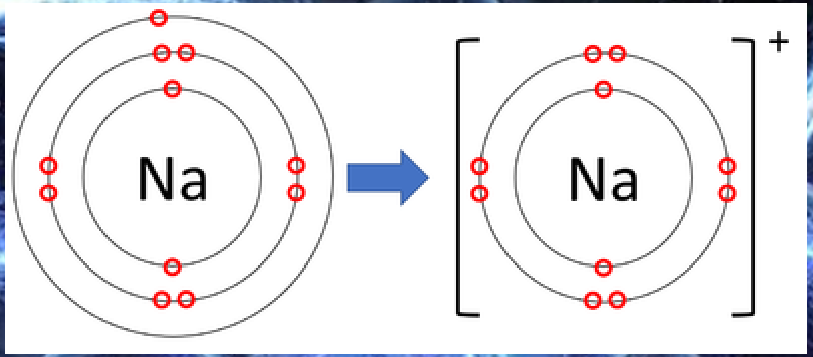
- Heavy overall weightage
- Constitute to **8%** of marks for past 5 year papers

KEY CONCEPT

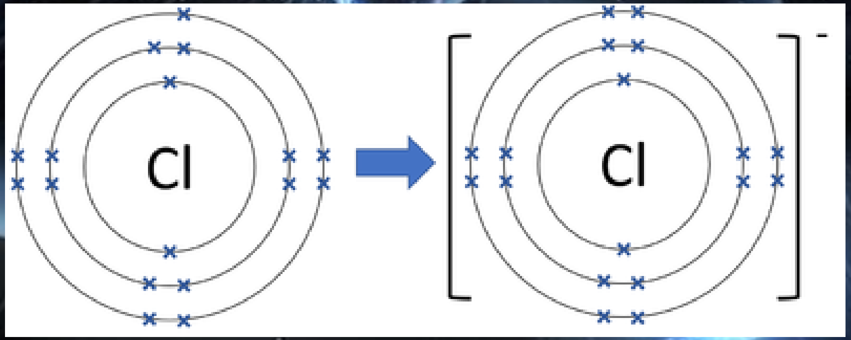
CHEMICAL EQUATION
CHEMICAL FORMULA
BALANCING CHEMICAL EQUATION
IONIC EQUATION

CHEMICAL FORMULA

Cation:



Anion:



IONIC COMPOUNDS

Some common anions:

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

Forming of ionic compounds:

For example,

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

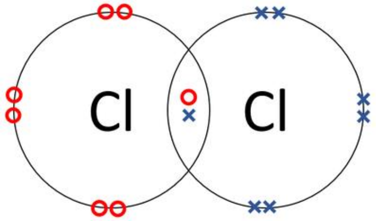
To balance out charges,

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

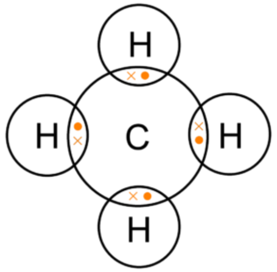
Compound:

$Ca(NO_3)_2$

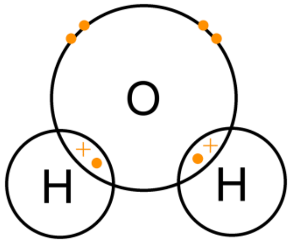
CHEMICAL FORMULA



Chlorine molecule



Methane compound



Water compound

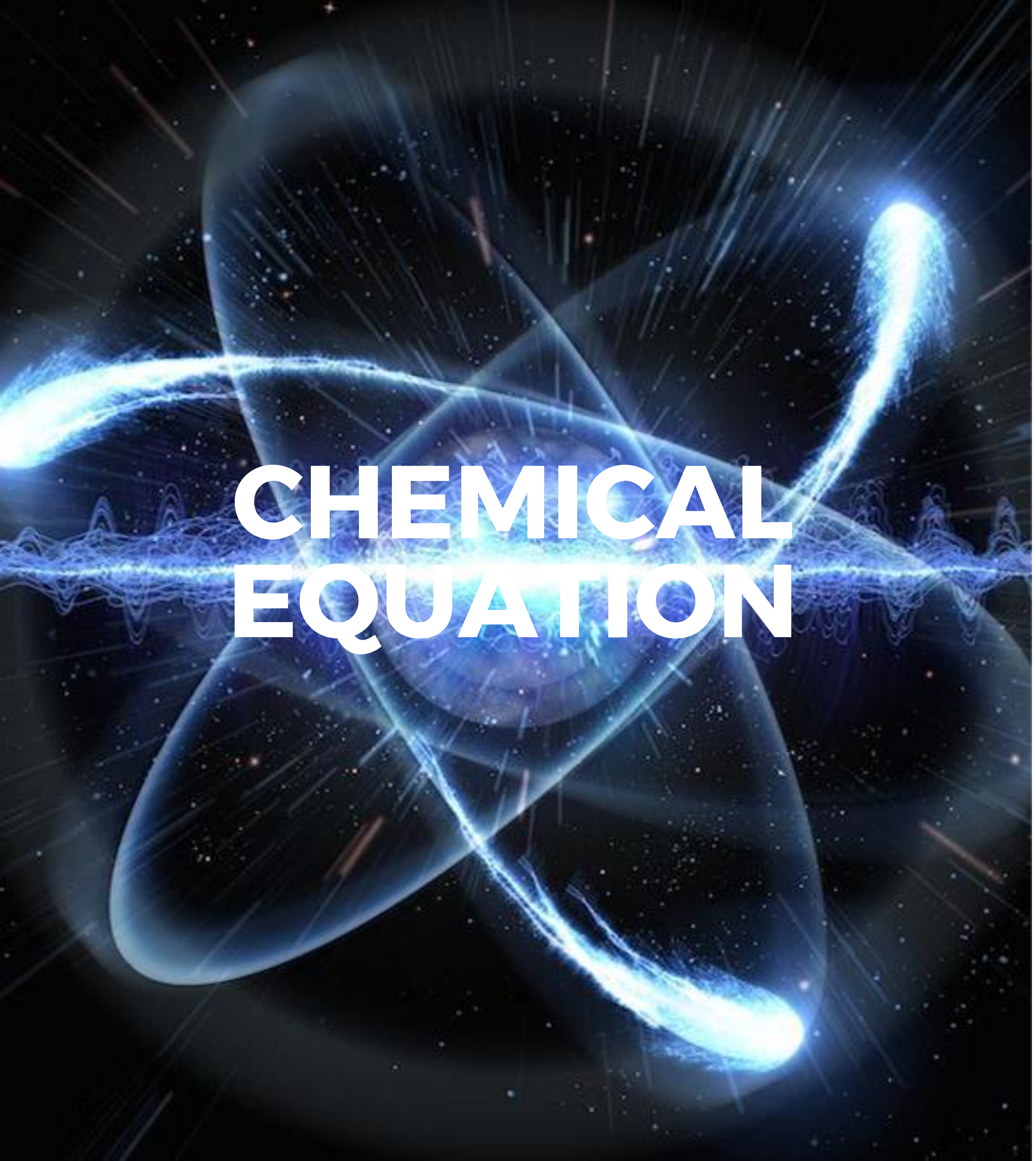
COVALENT COMPOUNDS

Prefixes are generally used to name compounds.

Prefix:

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

For example,
Nitrogen monoxide – NO
Nitrogen dioxide – NO_2



STATE SYMBOLS

Solid (s)
Liquid (l)
Gaseous (g)
Aqueous (aq) – exist as ions in a solution, water was added.

BALANCING EQUATIONS

Check that the number of atoms for each element is equal on both sides of the equation (reactants & products).

To balance the chemical equation, you will need to add a **coefficient** in front of the compounds that are not balanced.

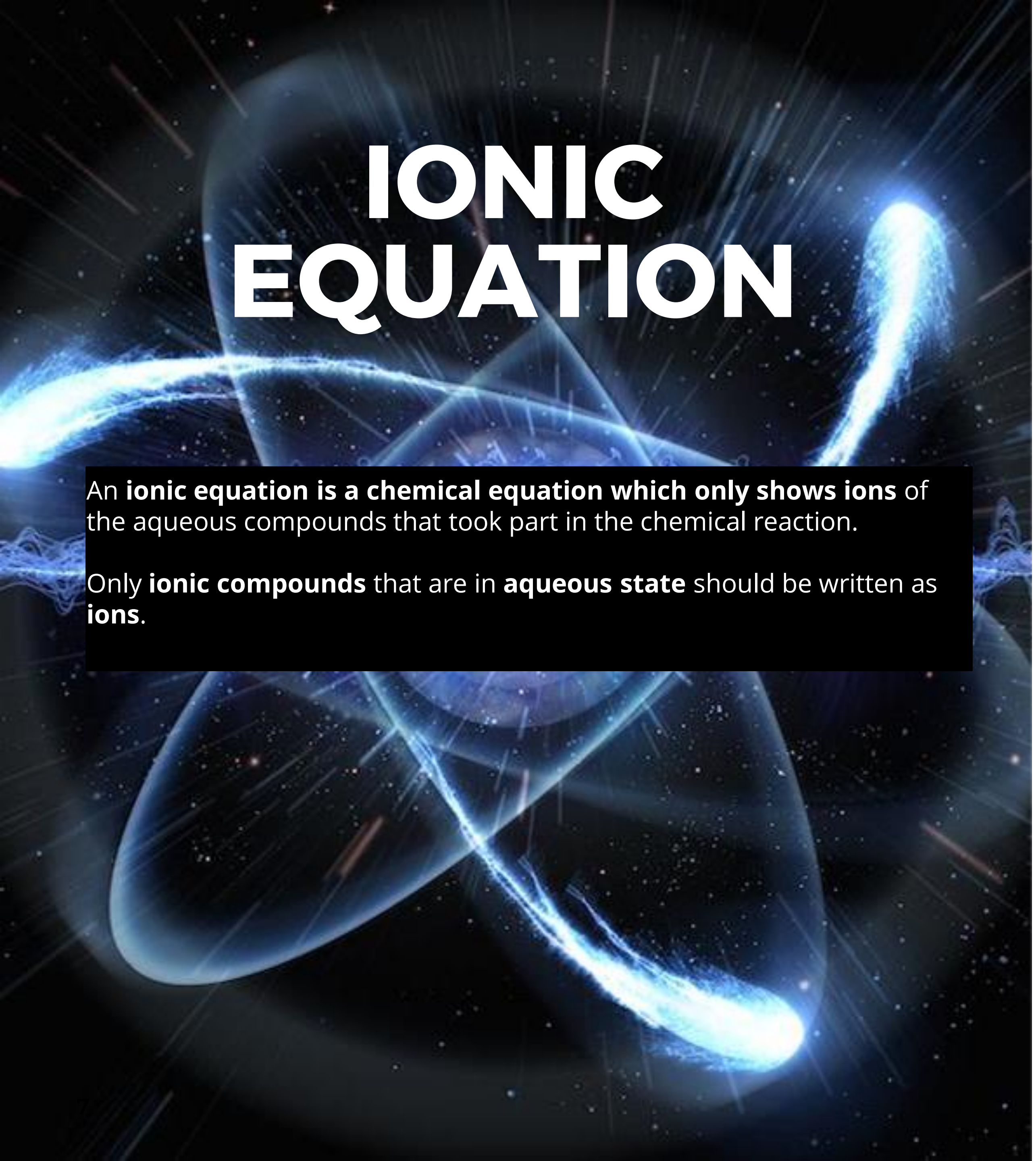
For example,

$$MgCl_2\ (aq) + Na_2CO_3\ (aq) \rightarrow MgCO_3\ (s) + \underline{\mathbf{2}}\ NaCl\ (aq)$$

$$MgCO_3\ (s) + \underline{\mathbf{2}}\ HCl\ (aq) \rightarrow MgCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

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

Practice makes perfect!



IONIC EQUATION

An **ionic equation is a chemical equation which only shows ions** of the aqueous compounds that took part in the chemical reaction.

Only **ionic compounds** that are in **aqueous state** should be written as **ions**.

Step 1

Write the balanced chemical equation for the reaction.

$CaCl_2$ (aq) + $CuSO_4$ (aq) → $CaSO_4$ (s) + $CuCl_2$ (aq)

Step 2

Ionic compounds that are in **aqueous state** should be written as **ions.**

Ca^{2+} (aq) + $2Cl^-$ (aq) + Cu^{2+} (aq) + SO_4^{2-} (aq) → $CaSO_4$ (s) + Cu^{2+} (aq) + $2Cl^-$ (aq)

Step 3

Remove all the spectator ions.

Ca^{2+} (aq) + ~~$2Cl^-$ (aq)~~ + ~~Cu^{2+} (aq)~~ + SO_4^{2-} (aq) → $CaSO_4$ (s) + ~~Cu^{2+} (aq)~~ + ~~$2Cl^-$ (aq)~~

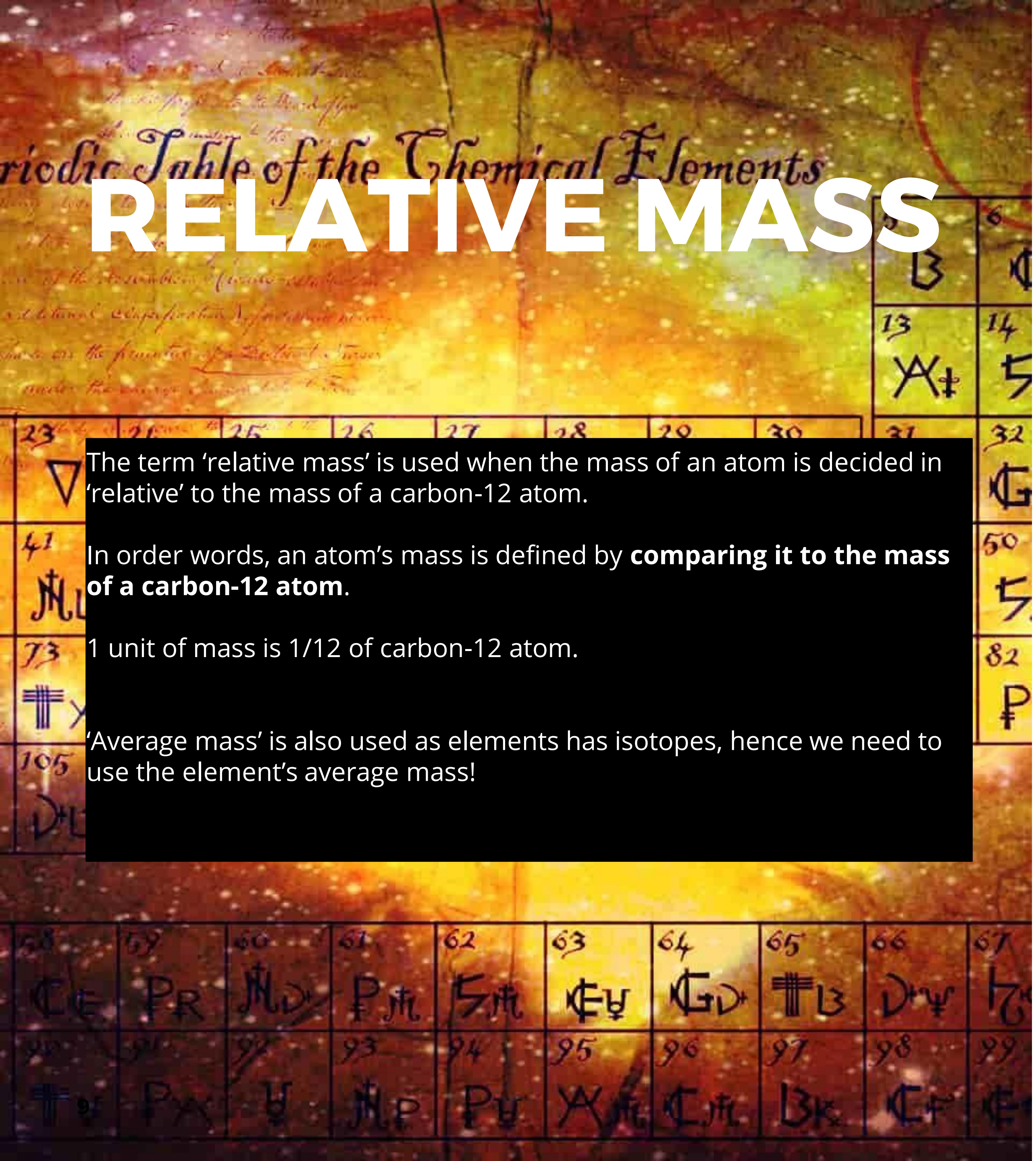
Step 4

Obtain the final ionic equation.

Ca^{2+} (aq) + SO_4^{2-} (aq) → $CaSO_4$ (s)

KEY CONCEPT

MOLE CONCEPT
Ar, Mr
MOLE
CONCENTRATION



RELATIVE MASS

The term 'relative mass' is used when the mass of an atom is decided in 'relative' to the mass of a carbon-12 atom.

In order words, an atom's mass is defined by **comparing it to the mass of a carbon-12 atom**.

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

'Average mass' is also used as elements has isotopes, hence we need to use the element's average mass!

Relative atomic mass **(A_r)**

A_r of an element is defined as the **average mass** of its atom **compared to 1/12 of the mass of one carbon-12 atom.**

Relative molecular mass **(M_r)**

M_r is defined as the **average mass** of a molecule of a substance **compared to 1/12 of the mass of one carbon-12 atom.**

*Carbon–12 is used as a basis of comparison because it is the most commonly available element on Earth.

Percentage by mass of an element present in a compound:

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

MOLE

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

WHAT IS MOLE?

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

The value 6.02×10^{23} is referred to as the Avogadro's constant.

$$\text{No. of particles} = \text{mole} \times 6.02 \times 10^{23}$$

MOLAR VOLUME OF GASES

At room temperature and conditions, one mole of gas has a volume of **24 dm^3 or 24 000 cm^3**.

Any type of gas, regardless of their chemical formula & M_r, all have the same volume.

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

Concentration

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

No. of moles = Concentration x volume

<u>CONCENTRATION</u>

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

There are two ways to measure concentration:

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:

<u>Calculate the mass of solute in 600 cm^3 of 0.4 $moldm^{-3}$ copper(II) sulfate solution.</u>

Volume of solution = 600 cm^3 = 0.60 dm^3

Number of moles of $CuSO_4$
= Concentration ($moldm^{-3}$) × Volume of solution (dm^3)
= 0.4 × 0.60
= 0.24 mol

Mass of $CuSO_4$
= Number of moles (mol) × Molar mass ($gmol^{-1}$)
= 0.24 × [64 + 32 + 4(16)]
= 38.4 g

KEY CONCEPT

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

CHEMICAL CALCULATIONS

STOICHIOMETRY FOR GAS

Since one mole of all gases share the same volume (1 mol = 24dm^3), assuming temperature and pressure are constant, volume of a gas is directly proportional to the number of moles.

Hence, the mole ratio of gases in a chemical equation can also let us know the **ratio of the volumes of gases in the chemical reaction**.

$$N_2\,(g) + 2\,O_2\,(g) \rightarrow 2\,NO_2\,(g)$$

10 cm^3 of N_2 will react with **20 cm^3 of O_2** to produce **20 cm^3 of NO_2**.

CHEMICAL CALCULATIONS

Example:
Find the mass of hydrogen gas formed when 80g of calcium metal is reacted with excess hydrochloric acid.

Step 1: Write out the balanced equation.

$$Ca\,(s) + 2\,HCl\,(aq) \rightarrow CaCl_2\,(aq) + H_2\,(g)$$

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

Number of moles of Ca reacted = mass / Mr
= 80 / 40
= 2

Step 3: Determine the molar ratio.

Number of moles of Ca 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

How many full cars can I assemble if I have 10 car bodies & 12 car wheels?

Answer: 3 full cars

Hence, **the wheels are the limiting reagent as it 'limits' further reaction** to assemble more cars even though there is an **'excess' of car bodies.**

LIMITING AND EXCESS REACTANTS

Not all the reactants are always fully used up in a chemical reaction.

The reaction will stop when one reactant is fully used up, even if the other reactants are still available.

The **limiting reactant** is the reactant that is **completely used up** first. It **limits the amount of product** that can be formed.

The **excess reactant** is the reactant that would **still remain** in excess even when the limiting reactant has been completely reacted.

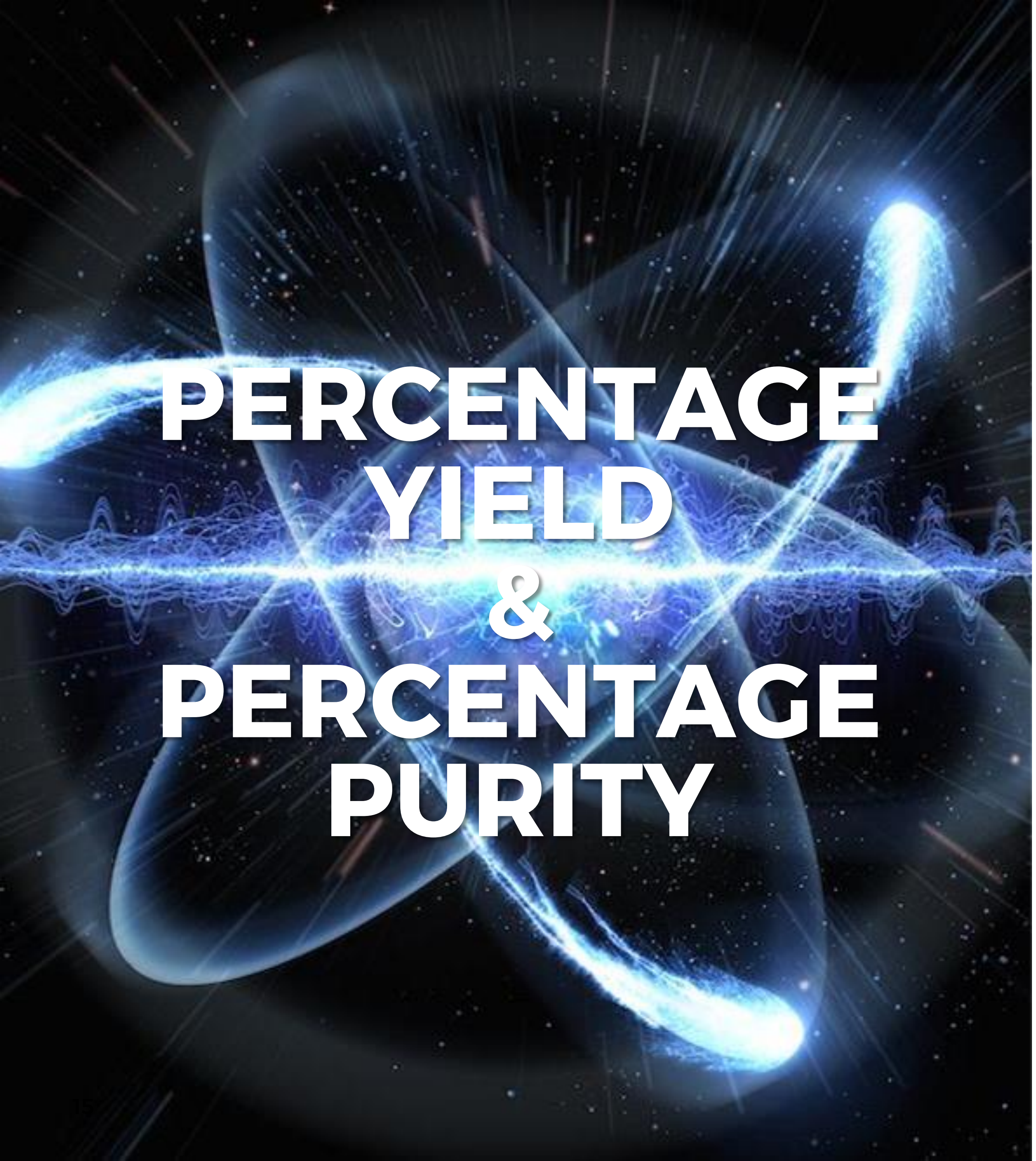
Example:

$$Cu(NO_3)_2\,(aq) + 2\,KCl\,(aq) \rightarrow CuCl_2\,(s) + 2\,KNO_3\,(aq)$$

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

As there is only 1 mole of $Cu(NO_3)_2$, so even if there are 5 moles of KCl, only 2 moles of KCl will react.

$Cu(NO_3)_2$ is the limiting reactant while **KCl is the excess reactant.**



PERCENTAGE YIELD & PERCENTAGE PURITY

PERCENTAGE YIELD

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

Actual yield refers to the actual amount of product obtained.

Theoretical yield refers to the maximum amount of products formed based on chemical calculation.

PERCENTAGE PURITY

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

EMPIRICAL FORMULA

EMPIRICAL FORMULA

The empirical formula is the **simplest ratio of the constituent elements of a compound**.

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

→ Just multiply by appropriate ratio to increase empirical formula to match the M_r.

Example (by mass):

Calcium metal of mass 1.6g was burnt in oxygen to form calcium oxide. When the calcium was completely burnt, the oxide produced had a mass of 2.24 g.

Determine the empirical formula & molecular formula of this oxide. (Mr is 102)

Mass of calcium = 1.60 g
Mass of calcium oxide produced = 2.24 g
Mass of oxygen reacted = 2.24 – 1.60 = 0.64 g

	Calcium (Ca)	Oxygen (O)
Mass in sample/g	1.6	0.64
Molar mass/g mol^{-1}	40	16
Number of moles	1.6 / 40 = 0.04	0.64 / 16 = 0.04
Simplest ratio	1	1

Hence, the **empirical formula of the oxide is CaO**.

Since M_r of oxide is 102,
n(40+16) = 102

Hence, **molecular formula is Ca_2O_2**.

Try it yourself! (TYS Question)

24. Which compound contains the highest percentage of sulfur by mass?

(N2020/P1/Q9)

A SO_2 B H_2SO_4
C Na_2S D PbS_2

( )

Answer:

24. **A**

Percentage of S in $SO_2 = \frac{32}{(32 + 2 \times 16)} \times 100\%$

$= 50\%$

Percentage of S in H_2SO_4

$= \frac{32}{(2 \times 1 + 32 + 4 \times 16)} \times 100\%$

$= 32.7\%$

Percentage of S in $Na_2S = \frac{32}{(2 \times 23 + 32)} \times 100\%$

$= 41\%$

Percentage of S in $PbS_2 = \frac{(2 \times 32)}{(207 + 2 \times 32)} \times 100\%$

$= 23.6\%$

SO_2 contains the highest percentage of sulfur by mass.

Try it yourself! (TYS Question)

25. On heating, the carbonate of element X decomposes.

$$XCO_3(s) \rightarrow XO(s) + CO_2(g)$$

6.25 g of XCO_3 is heated and 1.2 dm^3 of carbon dioxide, measured at room temperature and pressure, is produced.
What is the relative atomic mass, A_r, of X? (N2020/P1/Q10)

A 57 B 65
C 125 D 150 ( )

Answer:

25. **B**
No. of moles of $CO_2 = \frac{1.2}{24}$
$= 0.05$ mol
By comparing mole ratio,
no. of moles of $XCO_3 = 0.05$ mol
M_r of $XCO_3 = \frac{6.25}{0.05}$
$= 125$
A_r of X $= 125 - 12 - 3(16)$
$= 65$

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