CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

CONTENTS:

1. Introduction

Mass is a fundamental property of matter, representing the amount of material in an object. Calculations involving mass are essential in physics, chemistry, engineering, and everyday applications. These computations often include:

- **Basic Mass Measurements** – Determining mass using balances or scales, typically in grams (g) or kilograms (kg).

- **Molar Mass Calculations** – Relating mass to the number of particles (atoms, molecules) using Avogadro's number (6.022×10^{23} mol^{-1}).

- **Conservation of Mass** – Balancing mass in chemical reactions or physical processes.

These principles form the basis for more advanced topics like **Stoichiometry.**

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

2. Mole Concept

A Mole (mol) is the standard unit in chemistry for counting tiny particles like atoms, molecules, or ions.

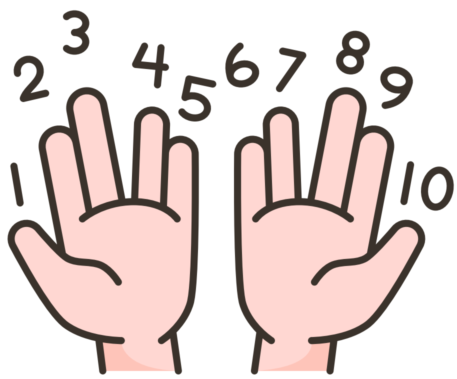
- 1 mole = 6.022×10^{23} particles (Avogadro's number)

Why Use Moles?

Atoms are too small to count individually (e.g., a drop of water has ~10^{21} molecules!). Moles let us:

1. Counting Particles Easily

- A single atom weighs ~10^{-23} grams—too small to measure.
- **1 mole = 6.022×10^{23} particles** (Avogadro's number), allowing us to count atoms by weighing.



2. Simplifying Chemical Reactions

- Reactions depend on **ratios of atoms/molecules**, not mass.
- Example:
  - **$2H_2 + O_2 \rightarrow 2H_2O$** means **2 molecules of H_2 react with 1 molecule of O_2**.
  - But in labs, we measure grams, not molecules.
  - **Moles convert grams → molecules**, making reactions practical.

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

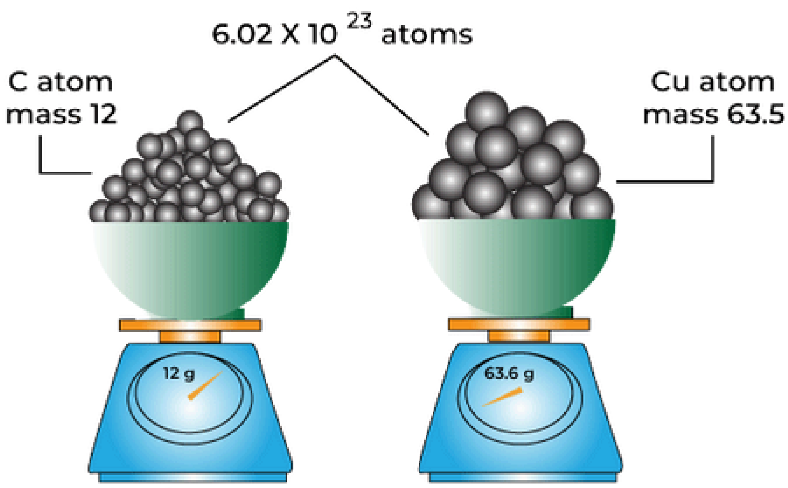
3. Connecting Mass to Number of Particles

Molar mass (mass of 1 mole) links the periodic table to real-world measurements.

Example:

Carbon (C) has a molar mass of 12 g/mol.

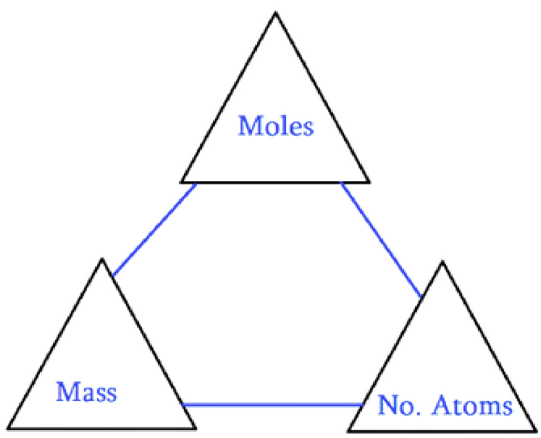
So, 12 grams of carbon = 1 mole = 6.022×10^{23} atoms.



4. Standardizing Measurements

Allows chemists worldwide to use the same scale (grams → moles → particles).

Essential for stoichiometry, gas laws, and solution chemistry.



CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

WHY IT MATTERS

1. Making the Invisible Visible

Atoms and molecules are **too small to see or count individually**, but we need to work with them in labs and factories.

- **Example:**
  - A single iron (Fe) atom weighs ~9.3×10^{-23} grams.
  - But **1 mole of iron (55.8 g, about a small nail)** contains **6.022×10^{23} atoms**—a measurable amount!

2. Essential for Chemical Reactions

Chemical equations (like recipes) depend on **ratios of particles**, not mass. Moles make this practical.

- **Example:**
  - The reaction **$2H_2 + O_2 \rightarrow 2H_2O$** requires **2 molecules of H_2 per 1 molecule of O_2**.
  - But in a lab, you measure grams, not molecules.
  - **Moles let you mix 2 g of H_2 (1 mole) with 32 g of O_2 (1 mole)** for the perfect burn.

3. Industry & Manufacturing

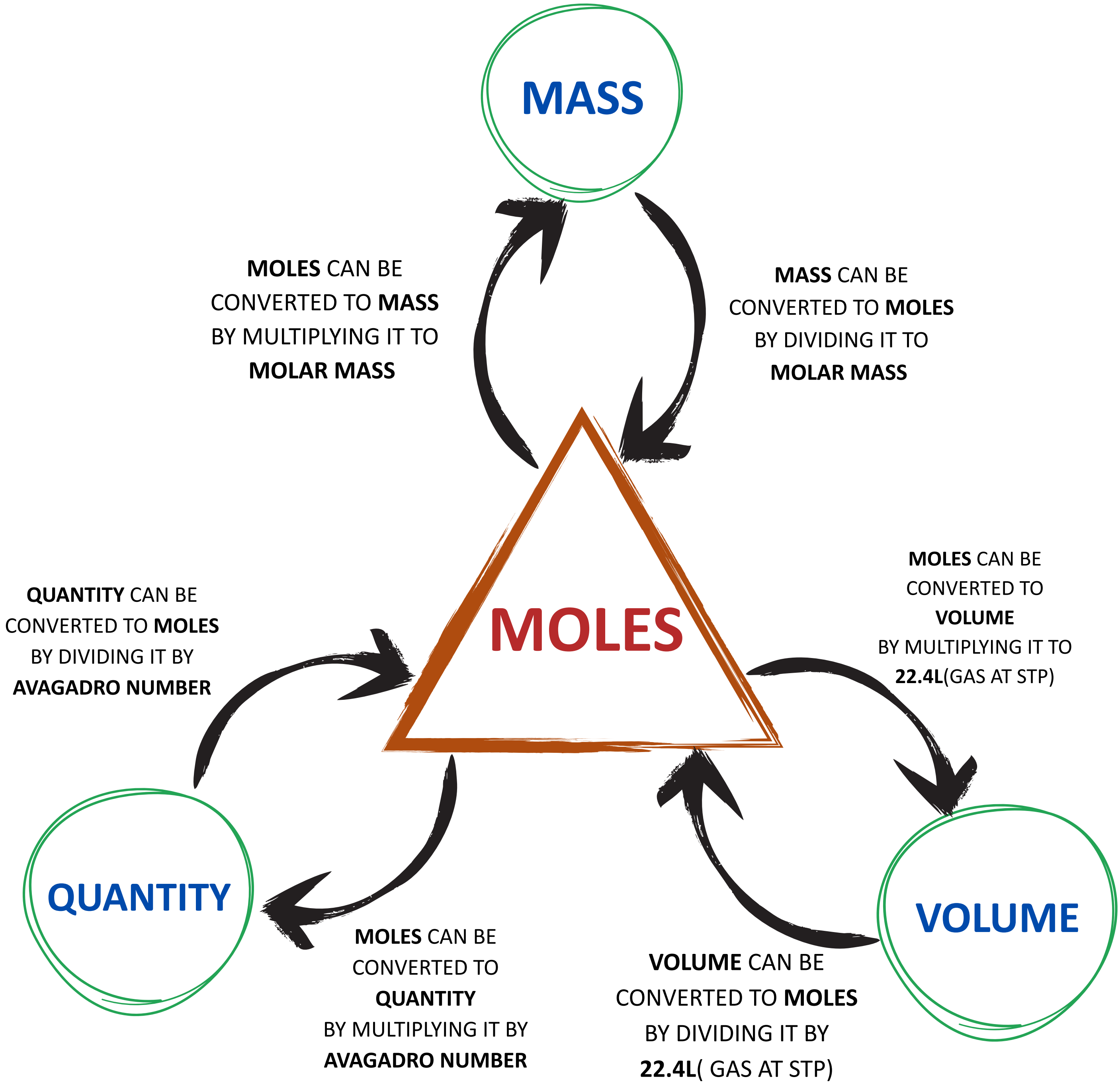
From plastics to fertilizers, factories rely on moles for cost control and quality.

- **Example:**
  - The Haber process **($N_2 + 3H_2 \rightarrow 2NH_3$)** uses moles to produce ammonia efficiently.
  - Too little hydrogen? Wasted nitrogen. Too much? Explosive risk.

→ Moles optimize **raw materials** and **prevent disasters.**

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

USEFUL CONVERSIONS FOR PROBLEM SOLVING



NOTE: **As a practical idea and in order to make the calculation easy, we can perform these mid steps for the easier solvation of questions.**

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

Example Calculation:

Problem: How many moles are in 36 g of water (H_2O)?

Solution:

Molar mass of H_2O = (2×1) + 16 = 18 g/mol

Moles = Mass ÷ Molar mass = 36 g ÷ 18 g/mol

= 2 moles

Problem: A recipe uses 8.4 grams of baking soda ($NaHCO_3$). How many formula units (particles) is this?

Solution:

Molar mass of $NaHCO_3$ = 23 (Na) + 1 (H) + 12 (C) + 3×16 (O)

= 84 g/mol.

Moles of $NaHCO_3$ = 8.4 g / 84 g/mol

= 0.1 moles.

Number of particles = $0.1 \times 6.022 \times 10^{23}$

= 6.022×10^{22} formula units.

Problem: 1g piece of aluminum foil contains how many aluminum atoms?

Solution:

Molar mass of Aluminium = 27 g/mol

Moles = 1g ÷ 27g/mol

≈ 0.037 mol

Atoms = $0.037 \times 6.022\times10^{23}$

≈ 2.2×10^{22} atoms

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

Let us understand it with a fun example:

Fun Fact:
If you had a mole of pennies, you could cover Earth's surface 5 km deep in coins!

1 Mole of pennies
= 6.023 X 10^{23}

Earth covered with 1 mole of pennies

Now we can clearly see that **1 Mole** is such a big value.......

Facts about Mole Concept:

- **(1 mole = 6.022 × 10^{23} particles** (Avogadro's number)).
- Molar mass = Mass of 1 mole of a substance (in grams).
- Connects **microscopic atoms** to measurable grams.
- Essential for:
  a. Balancing chemical equations
  b. Medicine dosages
  c. Industrial production
  d. Environmental science

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

3. Relative Formula Mass and Empirical Formula

RELATIVE FORMULA

The relative formula mass (M_r) of a compound is the sum of the relative atomic masses (A_r) of all the atoms in its chemical formula.

- A_r values are found on the periodic table (e.g., C = 12, O = 16, H = 1).
- If an element appears multiple times in a formula, multiply its A_r by the subscript.

Example:

- Water (H_2O):

$$\mathbf{M_r} = (2 \times A_r \text{ of H}) + (1 \times A_r \text{ of O})$$
$$= (2 \times 1) + 1$$
$$= 18$$

- Calcium chloride ($CaCl_2$):

$$\mathbf{M_r} = (1 \times A_r \text{ of Ca}) + (2 \times A_r \text{ of Cl})$$
$$= 40 + (2 \times 35.5)$$
$$= 111$$

Key Applications:

- **Stoichiometry:** Calculating reactant/product masses in reactions.
- **Empirical/Molecular Formulas:** Determining compound formulas from mass data.

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

EMPIRICAL FORMULA

The empirical formula of a compound shows the simplest whole-number ratio of the atoms of each element present.

Key Points:

- It's the reduced form of the molecular formula

  (e.g., $C_6H_{12}O_6 \rightarrow CH_2O$).
- Found using experimental data (mass or % composition).
- Does not show the actual number of atoms (unlike molecular formula).

Example:

- A compound has 2.4g Carbon & 0.6g Hydrogen.

Moles of C = 2.4 ÷ 12 = 0.2

Moles of H = 0.6 ÷ 1 = 0.6

$$\text{Ratio} = \frac{0.2}{0.2} = \frac{0.6}{0.2}$$
$$= \frac{1}{3}$$

Empirical formula = CH_3

Key Applications:

- **Compound Identification –** Simplest ratio for unknown substances.
- **Molecular Formula Basis –** Used with molar mass to find true formula.
- **Stoichiometry** – Balances reactions & calculates reactant/product masses.

4. Deduce Molecular Formula from Empirical Formula.

Simple Steps:

1. Find the mass of the empirical formula
   - Add up the atomic masses of all atoms in the empirical formula.
2. Divide the actual molecular mass by the empirical formula mass
   - This gives a multiplier (n).
3. Multiply each atom in the empirical formula by this number (n)
   - This gives the true molecular formula.

Example:

- Empirical formula = Fe_2O_3
- Given molecular mass = 400 g/mol

Step 1: Mass of Fe_2O_3

- Fe = 56, O = 16

$$= (2 \times 56) + (3 \times 16)$$
$$= 112 + 48$$
$$= 160 \text{ g/mol}$$

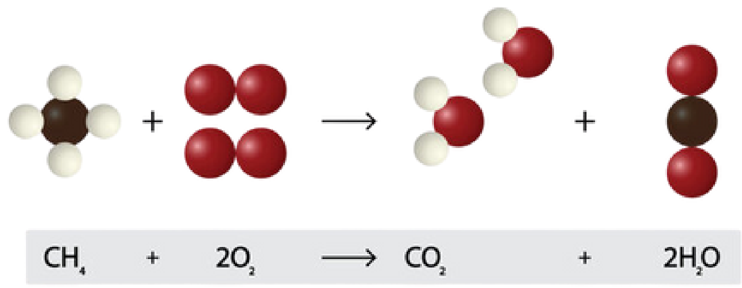
Step 2: Find multiplier (n)

- **n = Molecular Mass / Empirical Mass**

$$= 400 / 160$$
$$= 2.5$$

- Since formulas need whole numbers, check for calculation errors or possible non-integer ratios (e.g., if given data is approximate).

5. LAW OF CONSERVATION OF MASS.



In any chemical reaction, the total mass of the reactants (starting materials) always equals the total mass of the products (what you end up with). No atoms magically appear or disappear—they just rearrange!"

How It Works in Different Situations:

1. In a Closed System (Sealed Container)

- Mass stays exactly the same
- Example: Mixing two solutions in a closed flask that forms a solid (precipitate).
- Even though liquids turn into a solid, all atoms are still inside the flask—just in a new form.

2. In an Open System (Unsealed Container)

- Mass can seem to change
- Example: Baking soda + vinegar in an open beaker produces bubbles (CO_2 gas).
- The gas escapes into the air, so if you weigh it, the mass decreases.
- If a reaction absorbs gas (like iron rusting), the mass increases.

CALCULATIONS INVOLVING MASSES – GCSE CHEMISTRY

6. Stoichiometry

Stoichiometry uses mass measurements from experiments to determine the balancing coefficients in chemical equations. The process converts:

Mass (g) → Moles → Simple Ratio → Balanced Equation

- Balancing numbers in a symbol equation can be calculated from the masses of reactants and products: ○ convert the masses in grams to amounts in moles (moles = mass/Mr) ○ convert the numbers of moles to simple whole number ratios

Example:

- For the reaction:

  $Cu + O_2 \text{->} CuO$ (not balanced),

- 127 g Cu react, 32g of oxygen react and 159g of CuO are formed.

- Work out the balanced equation using the masses given:

  (Moles = Mass/Molar Mass)

- Cu: moles = 127 / 63.5 = 2
- O_2 : moles= 32 / (16 x 2) = 32/32 = 1
- CuO moles = 159 / (16 + 63.5) = 2

Therefore you have a ratio of 2 : 1 : 2 for $Cu:O_2:CuO$, making the overall balanced equation **[$2Cu + O_2 \text{->} 2CuO$]**

7. FAQs (Frequently Asked Questions)

1. What is a Mole?

A mole (mol) is the SI unit for counting particles (atoms, molecules, ions).

- 1 mole = 6.022×10^{23} particles (Avogadro's number).
- Analogy: Like a "dozen" (12 items), but for atoms!

2. What if two reactants are given? How do I know which one limits the product?

Steps that must be followed:

1. Calculate moles of both reactants.
2. Compare to the mole ratio in the balanced equation.
3. The reactant that produces fewer moles of product is the limiting reactant.

3. Why does the total mass of reactants equal the mass of products?

- Law of Conservation of Mass: Atoms are rearranged, not destroyed.
- Exception: In open systems, gases may escape (e.g., CO_2 in baking soda/vinegar reactions).

4. What's the difference between molar mass and relative atomic mass?

- Relative atomic mass (A_r): Average mass of an atom (unitless, from periodic table).
- Molar mass: Mass of 1 mole of a substance (units: g/mol).
  - Example: Carbon's A_r = 12 → Molar mass = 12 g/mol.

.

5. How do I measure mass experimentally in the lab?

- Electronic balance (for solids/liquids)
- Gas syringe (for gas volumes → convert to mass using moles)