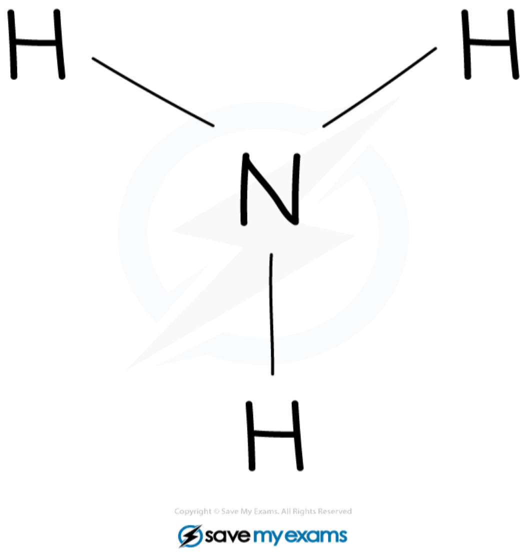
4. Stoichiometry

Symbols & Formulae of Elements & Compounds

Element symbols

- Each element is represented by its own unique symbol as seen on the Periodic Table e.g. H is hydrogen
- Where a symbol contains two letters, the first one is always in **capital** letters and the other is **small,** eg. sodium is Na, not NA
- Atoms combine together in **fixed ratios** that will give them full outer shells of electrons
- The chemical formula tells you the ratio of atoms
- Eg. H_2O is a compound containing 2 hydrogen atoms which combine with 1 oxygen atom
- The chemical formula can be deduced from the relative number of atoms present
- Eg. if a molecule contains 3 atoms of hydrogen and 1 atom of nitrogen then the formula would be NH_3
- Diagrams or models can also be used to represent the chemical formula



The ammonia molecule consists of a central nitrogen atom bonded to 3 hydrogen atom

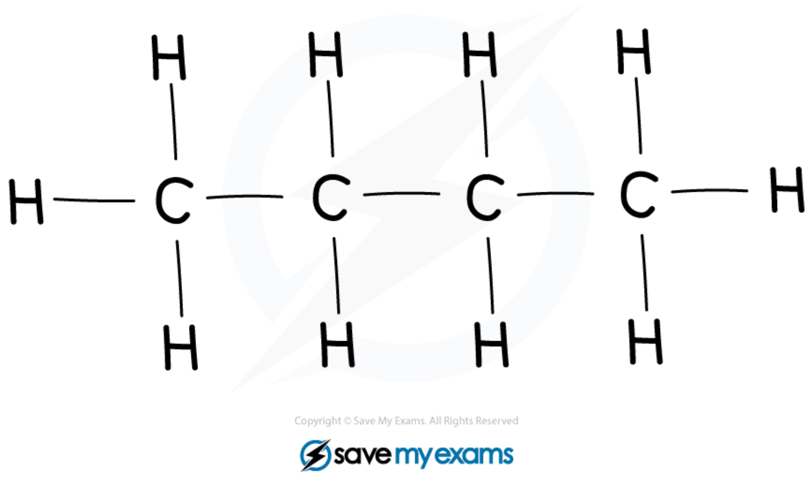
Chemical formulae

- The **structural** formula tells you the way in which the atoms in a particular molecule are bonded. This can be done by either a **diagram** (displayed formula) or **written** (simplified structural formula)
- The **empirical** formula tells you the simplest whole-number ratio of atoms in a compound

- The **molecular** formula tells you the actual number of atoms of each element in one molecule of the compound or element e.g. H_2 has 2 hydrogen atoms, HCl has 1 hydrogen atom and 1 chlorine atom

Example: <u>Butane</u>

- **Structural formula (displayed)**



- **Structural formula (simplified)** $CH_3CH_2CH_2CH_3$

- **Molecular formula** C_4H_{10}

- **Empirical formula** C_2H_5

Deducing formulae by combining power

- The concept of valency is used to deduce the formulae of compounds
- Valency or combing power tells you how many bonds an atom can make with another atom
- Eg. carbon is in Group IV so a single carbon atom can make 4 single bonds or 2 double bonds
- The following valencies apply to elements in each group:

GROUP	VALENCY
I	1
II	2
III	3
IV	4
V	3
VI	2
VII	1
VIII	0

- We can use the combining power of each atom to work out a formula
- Example: what is the formula of aluminium sulfide?
    - Write out the symbols of each element and write their combining powers underneath:
    - Al S
3 2

- The formula is then calculated by cross multiplying each atom with the number opposite, hence the formula for aluminium sulfide is Al_2S_3

Extended Only

Deducing Formulae of Ionic Compounds

- The formulae of these compounds can be calculated if you know the charge on the ions.
- Below are some common ions and their charges:

ION	FORMULA AND CHARGE
IRON(II)	Fe^{2+}
COPPER(II)	Cu^{2+}
CHROMIUM(III)	Cr^{3+}
AMMONIUM	NH^{4+}
HYDROXIDE	OH^-
NITRATE	NO_3^-
SULFATE	SO_4^{2-}
CARBONATE	CO_3^{2-}
HYDROGEN CARBONATE	HCO_3^-

- For ionic compounds you have to balance the charge of each part by multiplying each ion until the sum of the charges = 0
- Example: what is the formula of aluminium sulfate?
  - Write out the formulae of each ion, including their charges
  - Al^{3+} SO_4^{2-}
- Balance the charges by multiplying them out:
  Al^{3+} x **2** = +6 and SO_4^{2-} x **3** = -6; so +6 – 6 = 0
- So the formula is $Al_2(SO_4)_3$

Writing Word Equations & Balanced Equations

Word equations

- These show the reactants and products of a chemical reaction using their full chemical names
- The arrow (which is spoken as "goes to" or "produces") implies the conversion of reactants into products
- Reaction conditions or the name of a catalyst can be written above the arrow

Names of compounds

For compounds consisting of 2 atoms:

- If one is a **metal** and the other a **nonmetal**, then the name of the metal atom comes first and the ending of the second atom is replaced by adding –**ide**
  - Eg. NaCl which contains sodium and chlorine thus becomes sodium chlor**ide**
- If both atoms are **nonmetals** and one of those is **hydrogen**, then hydrogen comes first
  - Eg. hydrogen and chlorine combined is called hydrogen chlor**ide**
- For other combinations of nonmetals as a general rule, the element that has a lower Group number comes first in the name
  - Eg. carbon and oxygen combine to form CO_2 which is carbon dioxide since carbon is in Group 4 and oxygen in Group 6

-

For compounds that contain certain groups of atoms:

- There are **common groups** of atoms which occur regularly in chemistry
- Examples include the carbonate ion(CO_3^{2-}), sulfate ion (SO_4^{2-}), hydroxide ion (OH^-) and the nitrate ion (NO_3^-)
- When these ions form a compound with a metal atom, the name of the **metal** comes **first**
- Eg. KOH is potassium hydroxide, $CaCO_3$ is calcium carbonate

Writing and balancing chemical equations

- These use the chemical symbols of each reactant and product
- When balancing equations, there needs to be the **same number of atoms** of each element on either side of the equation
- The following nonmetals must be written as molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2
- Work across the equation from left to right, checking one element after another
- If there is a group of atoms, for example a nitrate group (NO_3^-) that has not changed from one side to the other, then count the whole group as one entity rather than counting the individual atoms. For example:

- $NaOH + HCl \rightarrow NaCl + H_2O$

- There are equal numbers of each atom on either side of the reaction arrow so the equation is balanced

Extended Only

Equations with State Symbols & Deducing Balanced Equations

Using state symbols:

State symbols are written after formulae in chemical equations to show which physical state each substance is in:

SOLID	LIQUID	GAS	AQUEOUS
(s)	(l)	(g)	(aq)

Example 1

Aluminium (s) + Copper (II) Oxide (s) →

Aluminium Oxide (s) + Copper (s)

Unbalanced symbol equation:

$$Al + CuO \rightarrow Al_2O_3 + Cu$$

ALUMINIUM:

There is 1 aluminium atom on the left and 2 on the right so if you end up with 2, you must start with 2. To achieve this, it must be 2Al

$$2Al + CuO \rightarrow Al_2O_3 + Cu$$

OXYGEN:

There is 1 oxygen atom on the left and 3 on the right so if you end up with 3, you must start with 3. To achieve this, it must be 3CuO

$$2Al + 3CuO \rightarrow Al_2O_3 + Cu$$

COPPER:

There is 3 copper atoms on the left and 1 on the right. The only way of achieving 3 on the right is to have 3Cu

$$2Al + 3CuO \rightarrow Al_2O_3 + 3Cu$$

Example 2

Magnesium Oxide (s) + Nitric Acid (aq) →

Magnesium Nitrate (aq) + Water (l)

Unbalanced symbol equation:

$$MgO + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$$

MAGNESIUM:

There is 1 magnesium atom on the left and 1 on the right so there are equal numbers of magnesium atoms on both sides so these are kept the same

$$MgO + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$$

OXYGEN:

There is 1 oxygen atom on the left and 1 on the right so there is an equal number of oxygen atoms on both sides. It is therefore kept the same (remember that you are counting the nitrate group as a separate group, so do not count the oxygen atoms in this group)

$$MgO + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$$

HYDROGEN:

There is 1 hydrogen atom on the left and 2 on the right. Therefore you must change HNO_3 to $2HNO_3$

$$MgO + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$$

Balancing ionic equations

- In aqueous solutions ionic compounds **dissociate** into their ions, meaning they separate into the component atoms or ions that formed them
- g. hydrochloric acid and potassium hydroxide dissociate as follows:
    - $HCl \rightarrow H^+ + Cl^-$
    - $KOH \rightarrow K^+ + OH^-$
- It is important that you can recognise common ionic compounds and their constituent ions
- These include:
    - Acids such as HCl and H_2SO_4
    - Group I and Group II hydroxides e.g. sodium hydroxide
    - Soluble salts e.g. potassium sulfate, sodium chloride

- Follow the example below to write Error: you must enter a valid popover post ID

Example 3

Write the ionic equation for the reaction of aqueous chlorine and aqueous potassium iodide.

Step 1: Write out the full balanced the equation:

$$2KI(aq) \quad + \quad Cl_2(aq) \quad \rightarrow$$
$$2KCl(aq) \quad + \quad I_2(aq)$$

Step 2: Identify the ionic substances and write down the ions separately:

$$2K^+(aq) + 2I^-(aq) \quad + \quad Cl_2(aq) \quad \rightarrow$$
$$2K^+(aq) + 2Cl^-(aq) \quad + \quad I_2(aq)$$

Step 3: Rewrite the equation eliminating the ions which appear on both sides of the equation (Error: you must enter a valid popover post ID) which in this case are the K^+ ions:

$$2I^-(aq) \quad + \quad Cl_2(aq) \quad \rightarrow$$
$$2Cl^-(aq) \quad + \quad I_2(aq)$$

Exam Tip

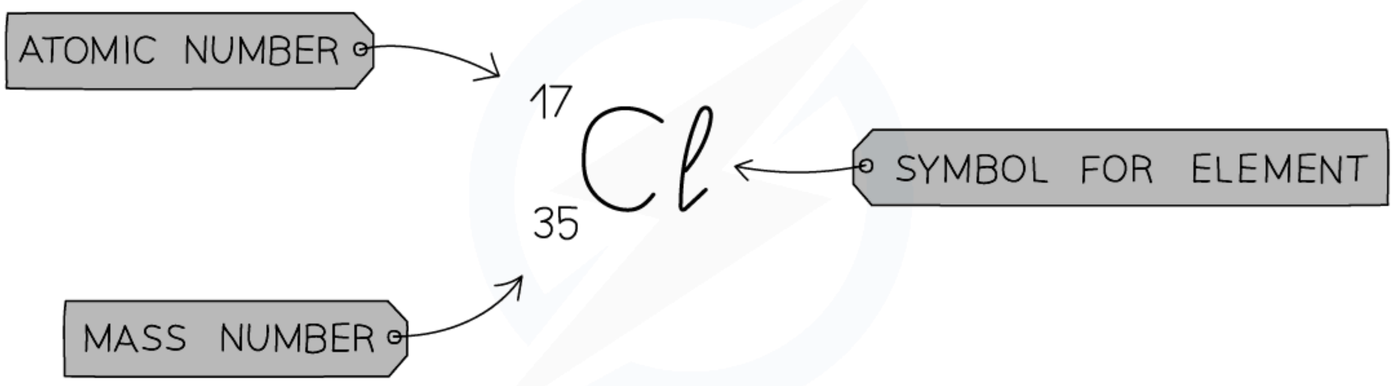
When balancing equations you cannot change any of the formulae, only the amount of each atom or molecule. This is done by changing the numbers that go in **front** of each chemical species.

You need to be able to identify the products which are not ions in ionic equations. These are usually molecules such as water or bromine but they may also be precipitated solids.

Relative Atomic Mass & Relative Molecular Mass

Relative atomic mass

- The symbol for the relative atomic mass is A_r
- This is calculated from the **mass number** and **relative abundances of all the isotopes** of a particular element



Symbol, mass number and atomic number of chlorine

Equation:

$$A_r = \frac{\text{(\% of isotope a x mass of isotope a) + (\% of isotope b x mass of isotope b)}}{100}$$

The top line of the equation can be extended to include the number of different isotopes of a particular element present.

Example for Isotopes:

The table shows information about the Isotopes in a sample of rubidium

ISOTOPE	NUMBER OF PROTONS	NUMBER OF NEUTRONS	PERCENTAGE OF ISOTOPE IN SAMPLE
1	37	48	72
2	37	50	28

(72 x 85) + (28 x 87) ÷ 100 = 85.6

Relative Atomic Mass = 85.6

Relative formula (molecular) mass

- The symbol for the relative molecular mass is M_r and it refers to the total mass of the molecule.
- To calculate the M_r of a substance, you have to add up the Relative Atomic Masses of all the atoms present in the formula

Example:

SUBSTANCE	ATOMS PRESENT	M_r
HYDROGEN (H_2)	2 × H	(2 × 1) = 2
WATER (H_2O)	(2 × H) + (1 × O)	(2 × 1) + 16 = 18
POTASSIUM CARBONATE (K_2CO_3)	(2 × K) + (1 × C) + (3 × O)	(2 × 39) + 12 + (3 × 16) = 138
CALCIUM HYDROXIDE ($Ca(OH)_2$)	(1 × Ca) + (2 × O) + (2 × H)	40 + (2 × 16) + (2 × 1) = 74
AMMONIUM SULFATE ($(NH_4)_2SO_4$)	(2 × N) + (8 × H) + (1 × S) + (4 × O)	(2 × 14) + (8 × 1) + 32 + (4 × 16) = 132

4.2 The Mole Concept

Extended Only

The Mole & Avogadro's Constant

The mole

- This is the mass of a substance containing the same number of fundamental units as there are atoms in exactly 12.000 g of ^{12}C
- The mole is the unit representing the amount of atoms, ions, or molecules
- One mole is the amount of a substance that contains **6.02 x 10^{23}** particles (Atoms, Molecules or Formulae) of a substance (6.02 x 10^{23} is known as the **Avogadro Number)**

Examples

- 1 mole of Sodium (Na) contains 6.02 x 10^{23} **Atoms** of Sodium
- 1 mole of Hydrogen (H_2) contains 6.02 x 10^{23} **Molecules** of Hydrogen

- 1 mole of Sodium Chloride (NaCl) contains 6.02×10^{23} **Formula units** of Sodium Chloride

Linking the mole and the atomic mass

- One mole of any element is equal to the relative atomic mass of that element in grams
- For example one mole of carbon, that is if you had 6.02×10^{23} atoms of carbon in your hand, it would have a mass of 12g
- So one mole of helium atoms would have a mass of 4g, lithium 7g etc
- For a compound we add up the relative atomic masses
- So one mole of water would have a mass of 2 x 1 + 16 = 18g
- Hydrogen which has an atomic mass of 1 is therefore equal to $^{1}/_{12}$ the mass of a ^{12}C atom
- So one carbon atom has the same mass as 12 hydrogen atoms

Extended Only

The Mole & the Volume of Gases

Molar volume

- This is the volume that one mole of any gas (be it molecular such as CO_2 or monoatomic such as helium) will occupy
- It's value is $24dm^3$ or 24,000 cm^3 at room temperature and pressure (r.t.p.)

Calculations involving gases

General equation:

Amount of gas (mol) = Volume of gas (dm^3) ÷ 24

or

Amount of gas (mol) = Volume of gas (cm^3) ÷ 24000

1. Calculating the volume of gas that a particular amount of moles occupies

Equation:

Volume of gas (dm^3) = Amount of gas (mol) x 24

or

Volume of gas (cm^3) = Amount of gas (mol) x 24000

Example:

NAME OF GAS	AMOUNT OF GAS	VOLUME OF GAS
HYDROGEN	3 mol	(3 x 24) = 72 dm^3
CARBON DIOXIDE	0.25 mol	(0.25 x 24) = 6 dm^3
OXYGEN	5.4 mol	(5.4 x 24,000) = 129,600 cm^3
AMMONIA	0.02 mol	(0.02 x 24) = 0.48 dm^3

2. Calculating the moles in a particular volume of gas

Equation:

$$\text{Amount of gas (mol)} = \text{Volume of gas } (dm^3) \div 24$$

or

$$\text{Amount of gas (mol)} = \text{Volume of gas } (cm^3) \div 24000$$

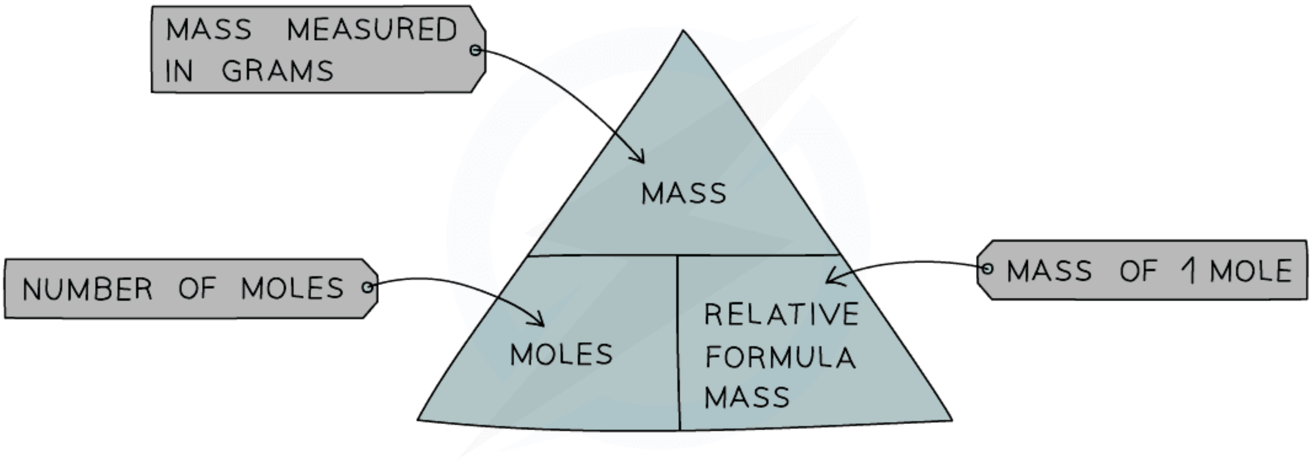
Example:

NAME OF GAS	VOLUME OF GAS	MOLES OF GAS
METHANE	225.6 dm^3	(225.6 : 24) = 9.4 mol
CARBON MONOXIDE	7.2 dm^3	(7.2 : 24) = 0.3 mol
SULFUR DIOXIDE	960 dm^3	(960 : 24) = 40 mol
OXYGEN	1200 cm^3	(1200 : 24000) = 0.05 mol

Extended Only

Calculating Reacting Masses, Solutions & Concentrations of Solutions in g/dm3 & mol/dm3

Calculating percentage composition, moles, mass and relative formula mass



Formula triangle for moles, mass and formula mass

1. Calculating Moles

Equation:

Amount in Moles = Mass of Substance in grams ÷ M_r (or A_r)

Example:

SUBSTANCE	MASS	MR	AMOUNT
NaOH	80 g	40	(80 ÷ 40) = 2 moles
$CaCO_3$	25 g	100	(25 ÷ 100) = 0.25 moles
H_2SO_4	4.9 g	98	(4.9 ÷ 98) = 0.05 moles
H_2O	108 g	18	(108 ÷ 18) = 6 moles
$CuSO_4.5H_2O$	75 g	250	(75 ÷ 250) = 0.3 moles

2. Calculating Mass

Equation:

$$\text{Mass of substance (grams)} = \text{Moles x } M_r \text{ (or } A_r)$$

Example:

SUBSTANCE	AMOUNT	MR	MASS
H_2O	0.5 moles	18	(0.5 x 18) = 9 g
NaCl	3 moles	58.5	(3 x 58.5) = 175.5 g
K_2CO_3	0.2 moles	138	(0.2 x 138) = 27.6 g
$(NH_4)2SO_4$	2.5 moles	132	(2.5 x 132) = 330 g
$MgSO_4.7H_2O$	0.25 moles	246	(0.25 x 246) = 61.5 g

3. Calculating Relative Formula Mass

Equation:

$$M_r \text{ (or } A_r) = \text{Mass of Substance in Grams} \div \text{Moles}$$

Example:

10 moles of Carbon Dioxide has a Mass of 440 g. What is the Relative Formula Mass of Carbon Dioxide?

$$\text{Relative Formula Mass} = \text{Mass} \div \text{Number of Moles}$$

$$\text{Relative Formula Mass} = 440 \div 10 = 44$$

Relative Formula Mass of Carbon Dioxide = 44

4. Calculating Percentage Composition

- The percentage composition is found by calculating the percentage by mass of each particular element in a compound

Example:

Calculate the percentage of oxygen in CO_2

Step 1 – Calculate the molar mass of the compound

Molar mass CO_2 = (2 x 16) + 12 = 44

Step 2 – Add the atomic masses of the element required as in the question (oxygen)

16 + 16 = 32

Step 3 – Calculate the percentage

% of oxygen in CO_2 = 32/44 x 100 = 72.7%

Calculations of solutions: moles, concentration and volume

General Equation:

Concentration (mol / dm^3) = Amount of substance (mol) ÷ Volume of solution (dm^3)

This general equation is rearranged for the term as is asked in the question.

1. Calculating Moles

Equation:

Amount of Substance (mol) = Concentration x Volume of Solution (dm^3)

Example:

Calculate the Moles of Solute Dissolved in 2 dm^3 of a 0.1 mol / dm^3 Solution

Concentration of Solution : 0.1 mol / dm^3

Volume of Solution : 2 dm^3

Moles of Solute = 0.1 x 2 = 0.1 mol (the dm^3 above and below the line cancel out)

Amount of Solute = 0.2 mol

2. Calculating Concentration

Equation:

Concentration (mol / dm^3) = Amount of substance (mol) ÷ Volume of solution (dm^3)

Example:

25.0 cm^3 of 0.050 mol / dm^3 sodium carbonate was completely neutralised by 20.00 cm^3 of dilute hydrochloric acid. Calculate the concentration in mol / dm^3 of the hydrochloric acid.

Step 1 – Calculate the amount, in moles, of sodium carbonate reacted by rearranging the equation for amount of substance (mol) and dividing by 1000 to convert cm^3 to dm^3

Amount of Na_2CO_3 = (25.0 x 0.050) ÷ 1000 = 0.00125 mol

Step 2 – Calculate the amount, in moles, of hydrochloric acid reacted

1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the Molar Ratio is 1 : 2

Therefore 0.00125 moles of Na_2CO_3 react with 0.00250 moles of HCl

Step 3 – Calculate the concentration, in mol / dm^3 of the Hydrochloric Acid

1 dm^3 = 1000 cm^3

Volume of HCl = 20 ÷ 1000 = 0.0200 dm^3

Concentration HCl (mol / dm^3) = 0.00250 ÷ 0.0200 = 0.125

Concentration of Hydrochloric Acid = 0.125 mol / dm^3

-

3. Calculating Volume

Equation:

Volume (dm^3) = Amount of substance (mol) ÷ Concentration (mol / dm^3)

Example:

Calculate the volume of hydrochloric acid of concentration 1.0 mol / dm^3 that is required to react completely with 2.5g of calcium carbonate.

Step 1 – Calculate the amount, in moles, of calcium carbonate that reacts

M_r of $CaCO_3$ is 100

Amount of $CaCO_3$ = $(2.5 \div 100)$ = 0.025 mol

Step 2 – Calculate the moles of hydrochloric acid required

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

1 mol of $CaCO_3$ requires 2 mol of HCl

So 0.025 mol of $CaCO_3$ Requires 0.05 mol of HCl

Step 3 – Calculate the volume of HCl Required

Volume = (Amount of Substance(mol) $\div$ Concentration (mol / dm^3)

= $0.05 \div 1.0$

= 0.05 dm^3 (the moles cancel out above and below the line)

Volume of Hydrochloric Acid = 0.05 dm^3

-

The limiting reactant and reacting masses

Limiting reactant

- The limiting reactant is the reactant which is **not present in excess** in a reaction
- It is always the first reactant to be used up which then causes the reaction to stop
- In order to determine which reactant is the limiting reagent in a reaction, we have to consider the ratios of each reactant in the balanced equation

Example:

9.2g of sodium is reacted with 8.0g of sulfur to produce sodium sulfide, NaS. Which reactant is in excess and which is the limiting reactant?

Step 1 – Calculate the moles of each reactant

Moles = Mass $\div$ A_r

Moles Na = 9.2/23 = 0.40

Moles S = 8.0/32 = 0.25

Step 2 – Write the balanced equation and determine the molar ratio

$2Na + S \rightarrow Na_2S$ so the molar ratios is 2 : 1

Step 3 – Compare the moles. So to react completely 0.40 moles of Na require 0.20 moles of S and since there are 0.25 moles of S, then S is in excess. Na is therefore the limiting reactant.

<u>Calculating reacting masses</u>

- Chemical equations can be used to calculate the moles or masses of reactants and products
- Use information from the question to find the amount in moles of the substances being considered
- Identify the ratio between the substances using the balanced chemical equation
- Apply mole calculations to find answer

<u>Example 1:</u>

Calculate the Mass of Magnesium Oxide that can be made by completely burning 6 g of Magnesium in Oxygen

Magnesium (s) + Oxygen (g) → Magnesium Oxide (s)

Symbol Equation:

$$2Mg + O_2 \rightarrow 2MgO$$

Relative Formula Mass: Magnesium : 24 Magnesium Oxide : 40

Step 1 – Calculate the moles of Magnesium Used in reaction

Moles = Mass ÷ M_r Moles = 6 ÷ 24 = 0.25

Step 2 – Find the Ratio of Magnesium to Magnesium Oxide using the balanced Chemical Equation

	MAGNESIUM	MAGNESIUM OXIDE
MOL	2	2
RATIO	1	1
MOL	0.25	0.25
MOLES OF MAGNESIUM OXIDE	= 0.25	

Step 3 – Find the Mass of Magnesium Oxide

Moles of Magnesium Oxide = 0.25

Mass = Moles x M_r Mass = 0.25 x 40 = 10 g

Mass of Magnesium Oxide Produced = 10 g

<u>Example 2:</u>

Calculate the Mass, in Tonnes, of Aluminium that can be Produced from 51 Tonnes of Aluminium Oxide

Aluminium Oxide (s) → Aluminium (s) + Oxygen (g)

Symbol Equation:

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

A_r and M_r: Aluminium : 27   Oxygen : 16   Aluminium Oxide : 102

1 Tonne = 10^6 g

Step 1 – Calculate the moles of aluminium oxide used

Mass of Aluminium Oxide in Grams = 51 x 10^6 = 51,000,000 g

Moles = Mass ÷ A_r Moles = 51,000,000 ÷ 102 = 500,000

Step 2 – Find the ratio of aluminium oxide to aluminium using the balanced chemical equation

	ALUMINIUM OXIDE	ALUMINIUM
MOL	1	2
RATIO	1	2
MOL	500,000	1,000,000
MOLES OF ALUMINIUM = 1,000,000		

Step 3 – Find the mass of aluminium

Moles of aluminium = 1,000,000

Mass in grams = Moles x A_r Mass = 1,000,000 x 27 = 27,000,000

Mass in Tonnes = 27,000,000 ÷ 10^6 = 27 Tonnes

Mass of Aluminium Produced = 27 Tonnes

<u>Extended Only</u>

Using the Mole to Determine Empirical & Molecular Formulae

Empirical formula:

It gives the **simplest whole number ratio** of atoms of each element in the compound

- Calculated from knowledge of the ratio of masses of each element in the compound

Example:

A compound that contains 10 g of Hydrogen and 80 g of Oxygen has an Empirical Formula of H_2O. This can be shown by the following calculations:

Amount of Hydrogen Atoms = Mass in grams ÷ A_r of Hydrogen = (10 ÷ 1) = **10 moles**

Amount of Oxygen Atoms = Mass in grams ÷ A_r of Oxygen = (80 ÷ 16) = **5 moles**

The ratio of moles of hydrogen atoms to moles of oxygen atoms:

	HYDROGEN	OXYGEN
MOLES	10	5
RATIO	2	1

Since equal numbers of moles of atoms contain the same number of atoms, the ratio of hydrogen atoms to oxygen atoms is 2:1

Hence the **empirical formula** is H_2O

Molecular formula:

It gives the exact numbers of atoms of each element present in the formula of the compound

- Divide the relative formula mass of the molecular formula by the relative formula mass of the Empirical Formula
- Multiply the number of each element present in the Empirical Formula by this number to find the Molecular Formula

Relationship between Empirical and Molecular Formula:

NAME OF COMPOUND	EMPIRICAL FORMULA	MOLECULAR FORMULA
METHANE	CH_4	CH_4
ETHANE	CH_3	C_2H_6
ETHENE	CH_2	C_2H_4
BENZENE	CH	C_6H_6

Example:

The Empirical Formula of X is $C_4H_{10}S_1$ and the Relative Formula Mass of X is 180. What is the Molecular Formula of X?

Relative Formula Mass: Carbon : 12 Hydrogen : 1 Sulfur : 32

Step 1 – Calculate Relative Formula Mass of Empirical Formula

$(C \times 4) + (H \times 10) + (S \times 1) = (12 \times 4) + (1 \times 10) + (32 \times 1) = 90$

Step 2 – Divide Relative Formula Mass of X by Relative Formula Mass of Empirical

Formula

$180 / 90 = 2$

Step 3 – Multiply Each Number of Elements by 2

$(C_{4 \times 2}) + (H_{10 \times 2}) + (S_{1 \times 2}) = (C_8) + (H_{20}) + (S_2)$

Molecular Formula of X = $C_8H_{20}S_2$

Extended Only

Calculating Percentage Yield & Percentage Purity of the Product

Percentage yield

- This is the calculation of the percentage yield obtained from the theoretical yield
- In practice, you never get 100% yield in a chemical process for several reasons
- These include some reactants being left behind in the equipment, the reaction may be reversible or product may also be lost during separation stages

Equation:

Percentage Yield = (Yield Obtained / Theoretical Yield) x 100

Example:

In an experiment to displace copper from copper sulfate, 6.5 g of Zinc was added to an excess of copper (II) sulfate solution. The copper was filtered off, washed and dried. The mass of copper obtained was 4.8 g. Calculate the percentage yield of copper.

Equation Of Reaction:

$$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$$

Step 1: Calculate the Amount, in Moles of Zinc Reacted

Moles of Zinc = 6.5 ÷ 65 = 0.10 moles

Step 2: Calculate the Maximum Amount of Copper that could be formed from the

Molar ratio

Maximum Moles of Copper = 0.10 moles (Molar ratio is 1:1)

Step 3: Calculate the Maximum Mass of Copper that could be Formed

Maximum Mass of Copper = (0.10 x 64) = 6.4 g

Step 4: Calculate the Percentage of Yield of Copper

Percentage Yield = (4.8 ÷ 6.4) x 100 = 75%

Percentage Yield of Copper = 75%

Percentage purity

- Often the product you are trying to fabricate may become contaminated with unwanted substances such as unreacted reactants, catalysts etc.

Equation:

Percentage Purity = (Mass of pure substance / Mass of impure substance) x 100

Example:

In an experiment 7.0g of impure calcium carbonate were heated to a very high temperature and 2.5g of carbon dioxide were formed. Calculate the percentage purity of the calcium carbonate.

Equation Of Reaction:

$$CaCO_3 (s) \rightarrow CaO(s) + CO_2 (g)$$

Step 1: Calculate the relative formula masses

1 mole $CaCO_3 \rightarrow$ 1 mole CO_2

$40+12+(3\times16)$ $12+(2\times16)$

$100 \rightarrow 44$

Step 2: Calculate the theoretical mass of calcium carbonate used if pure

From 2.5g CO_2 we would expect 2.5/44 x 100 = 5.68g

Step 3: Calculate the percentage purity

(Mass of pure substance / mass of impure substance) x 100

= 5.68/7.0 x 100

= 81.1%

****