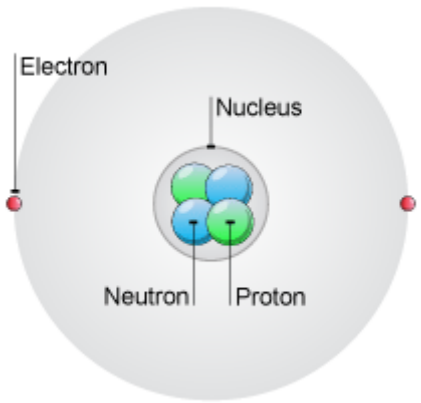
# Unit 3: Quantitative Chemistry (Triple Content)




**BBC Bitesize—Atomic Structure**

Atoms consist of a small, dense nucleus made of protons and neutrons, with electrons arranged in shells around the outside.

The **Atomic Number** is the number of protons an atom of an element has.

The **Relative Atomic Mass** ($A_r$) is the *average* mass of the atom compared to Carbon-12. The mass of one atom only takes into account protons and neutrons (as the mass of an electron is 2000x smaller—so makes very little difference!)

- When atoms chemically combine (or **bond**) they form **compounds**
- The **Relative Formula Mass** ($M_r$) of a compound is found by adding together the Relative Atomic Masses of all the atoms in the compound

E.g. $M_r$ ($C_2H_6$) = (2 x 12) + (6 x 1) = 30g/mol

## Moles

- Chemical amounts are measured in **moles** (symbol: mol).
- A mole is just a big number: like a million! One "mole" of a substance contains **6.02 x $10^{23}$** of the started particle (e.g. atom, molecule, ion or electron). This number is called "**Avogadro's constant**".
- The mass of one mole of a substance in grams is equal to its relative formula mass ($M_r$).

| Mass | = | $M_r$ | x | Mole |
|---|---|---|---|---|
| (g) | | (g/mol) | | (mol) |

E.g. The mass of 5 moles of carbon dioxide = ((1 x 12)+(2 x 16)) x 5 = 220g

## Conservation of Mass

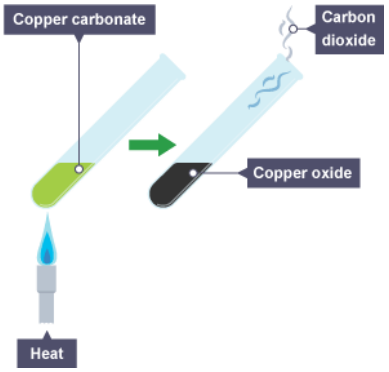
- In a **chemical reaction** no atoms can be **created** (made) or **destroyed** (lost)
- This means the mass of the **reactants** must equal the mass of the **products**
- Chemical reactions can be represented by symbol equations that are balanced in terms of the atoms of each element on both sides of the equation

E.g. $N_2 + 3H_2 \longrightarrow 2NH_3$

| | |
|---|---|
| 2 atoms Nitrogen | 2 atoms Nitrogen |
| 6 atoms Hydrogen | 6 atoms Hydrogen |

The **masses** of both sides of the equation also balance.

E.g. Reactants: (2 x 14) + = 30 Products = (2 x 14) + (6x1) = 30




**BBC Bitesize: Thermal decomposition**

### Reactions involving gases

In a reaction involving gases, the mass may **appear** to change, because not all the atoms are weighed.

e.g. $CuCO_3$ (s) —> CuO (s) + $CO_2$ (g)

Although the $CuCO_3$ weighs the same as the CuO and $CO_2$ put together, when you weigh your reaction vessel the $CO_2$ has already left—so the mass appears to have decreased.

Look out for state symbols!

## Limiting Reactants

In a chemical reaction involving two **reactants**, it is common to use an **excess** of one of the reactants to ensure that all of the other reactant is used.

The reactant that is completely used up is called the limiting reactant because it limits the amount of products.

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**Balancing Equations**

Balancing equations is about making sure there are the same number of atoms of each type on each side. An unbalanced equation like:

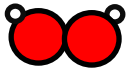
$$H_2 + O_2 \longrightarrow H_2O$$

Suggests that an oxygen atom has been lost: and we know that **never** happens in chemical reactions.

When you balance the equation, you can **only** put numbers **in front** of the chemicals to show that the numbers of them has changed

E.g. $2H_2O$ means there are now two water molecules:

But writing $H_2O_2$ would mean you had made a completely different chemical!



**To Balance Equations**

Step 1: Count the atoms you have on both sides of the equation.

$$H_2 + O_2 \longrightarrow H_2O$$

2 hydrogen, 2 oxygen 2 hydrogen, 1 oxygen

Step 2: Pick an unbalanced element, and add a number in front of the species containing it to balance it. Hint: Look for the **Lowest Common Multiple**.

$$H_2 + O_2 \longrightarrow 2H_2O$$

Step 3: Recount the atoms

2 hydrogen, 2 oxygen 4 hydrogen, 2 oxygen

Step 4: Pick a different element that is now unbalanced and add a number in front of the species containing it to help it balance:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Continue until all the elements are balanced!

**Balancing equations using masses**

It is possible to balance an equation by working out how many moles of each chemical there are, and converting this to the simplest whole number ratio.

E.g. 480g Ti + 1420g $Cl_2$ —> 1900g $TiCl_4$

**Step 1: Calculate $A_r$ or $M_r$**

$A_r$ (Ti) = 48g/mol
$M_r(Cl_2)$ = (2 x 35.5) = 71g/mol
$M_r(TiCl_4)$ = (1 x 48) + (4 x 35.5) = 190g/mol

**Step 2: Calculate number of moles in reaction**

**Ti:** 480g ÷ 48g/mol = 10 mol
**$Cl_2$:** 1420g ÷ 71g/mol = 20 mol
**$TiCl_4$**: 1900g ÷ 190g/mol = 10 mol

**Step 3: Simplify ratio**

10 mol: 20 mol: 10 mol = 1 : 2 : 1

**Step 4: Add numbers to balance equation**

Ti + **2** Cl2 —> TiCl4

**Measurement Uncertainty**

- Whenever a measurement is made there is always some **uncertainty**.
- This is due to the **resolution** of your equipment e.g. your ruler only allows you to measure to the nearest millimetre!
- Uncertainty can be represented using a ± sign e.g. the temperature is 20 °C ± 2°C
- To calculate **percentage uncertainty** for a set of data:

  Percentage Uncertainty = (Range / Mean) x 100%
- You may be asked to do this from a graph or a table of data.

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**Predicted Yield**

The predicted yield (or expected yield) is how much of a chemical product a reaction is expected to make.

**Using Balanced Equations to Calculate Predicted Yield**

Balanced equations tell you the **ratio** of reacting chemicals

E.g. $H_2 + Cl_2 \longrightarrow 2HCl$

tells you that 1 mole of hydrogen molecules react with 1 mole of chlorine molecules to make 2 moles of hydrogen chloride.

If you know the $M_r$ (Relative formula mass) of each chemical, this allows you to calculated **Predicted Yield** (how much product should be made).

E.g. If 5g of $H_2$ react with chlorine, how much hydrogen chloride will be made?

$M_r(H_2) = (2 \times 1) = 2$ $\quad$ $M_r(HCl) = (1 \times 1) + (1 \times 35.5) = 36.5$

$2HCl = 2 \times 36.5 = 73$

This tells us that **2g** of $H_2$ would make **73**g of HCl… But we don't have 2g!

Divide both sides by the $M_r$ of the **known** substance

$2 \div 2 = 1g$ $\quad$ $73 \div 2 = 31.5g$

This tells us that **1g** of $H_2$ would make **31.5g** of HCl

Now multiply each side by the mass of the known substance from the question.

$1g \times 5 = 5g$ $\quad$ $31.5g \times 5 = \underline{157.5g}$

So **5g** of hydrogen will react with chlorine to make **157.5g** of hydrogen chloride.

(The same calculation can be used to calculate how much of a reactant was used to make a set amount of product).

**Percentage Yield**

Percentage Yield is a measure of how much product you made (Actual Yield) compared to how much expected to make (Predicted Yield)

$$\text{Percentage Yield} = \frac{\text{Mass of Product Made (AY)}}{\text{Maximum Theoretical Mass (PY)}} \times 100\%$$

E.g. If you expected to make 600g of aluminium, but you actually made 450g, your Percentage Yield is $450 \div 600g \times 100\% = \underline{75\%}$

**Why is Percentage Yield not 100%?**

Percentage Yield is never 100%. This may be for many reasons:

- The reaction does not go to completion (e.g. there is insufficient time for the reaction to finish).
- Some reactants take part in alternative reactions (e.g. reacting with oxygen in the atmosphere).
- Some product is lost during extraction (e.g. because it stays in the container the reaction was carried out in).
- The reactants were impure (e.g. metal was extracted from an ore that also contained many other minerals)
- The reaction was **reversible**. (If this is the case, the equation will have this symbol: $\rightleftharpoons$ instead of a straight arrow

**Atom Economy**

What percentage of all atoms in the products are in the **useful** product?

$$\text{Atom Economy} = \frac{\text{Mass of Useful Product} \times 100\%}{\text{Mass of All Products}}$$

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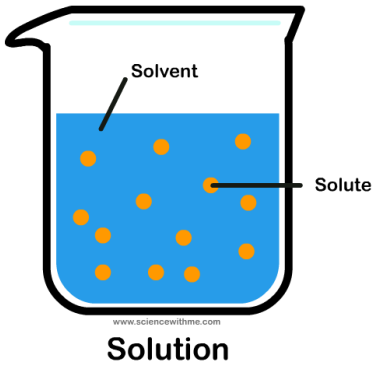



## Concentration of Solutions

Many reactions take place in a **solution** (i.e. one substance dissolved in another)

**Solute** - a substance dissolved in another (e.g. salt)

**Solvent** - a liquid used to dissolve a solute (e.g. water)

**Solution** - a mixture of solute and solvent

**Soluble** (adj.) - able to be dissolved

The **ratio** of solute and solvent can be described as the solution's **concentration**.

Concentration = Mass of Solute ÷ Volume of Solvent

($g/dm^3$) (g) ($dm^3$)

E.g. A salt solution containing 100g of salt dissolved in $2dm^3$ has a concentration of 100 ÷ 2 = <u>$50g/dm^3$</u>

You may also be asked to rearrange this equation, and use it to work out the Mass of Solute in a given volume of solution

Mass of Solute = Concentration x Volume of Solvent

(g) ($g/dm^3$) ($dm^3$)

E.g. $5dm^3$ of a $20g/dm^3$ solution contains 20 x 5 = <u>100g of solute</u>.

## Unit Conversion

- A decimetre cubed ($dm^3$) is the same size as a litre (l)
- There are 1000ml or $1000cm^3$ in 1 litre or 1 $dm^3$
- E.g. $5dm^3$ = 5l = 5000ml = $5000cm^3$

## Concentration in $mol/dm^3$

It is often more useful to give concentrations in $mol/dm^3$ rather than $g/dm^3$.

To calculate concentration with these units use:

Concentration = Moles of Solute ÷ Volume of Solvent

($mol/dm^3$) (mol) ($dm^3$)

E.g. A salt solution containing 0.8 mol of salt dissolved in 0.5 $dm^3$ of water has a concentration of 0.8 ÷ 0.5 = $1.6mol/dm^3$

If you need to convert between concentrations in $g/dm^3$ and $mol/dm^3$, use **exactly the same approach you would use to convert between g and mol!**

Mass = $M_r$ x mol

Mol = Mass ÷ $M_r$

**Worked Example 1: What is the concentration of a 3 $mol/dm^3$ $HNO_3$ solution, in $g/dm^3$?**

Mass = $M_r$ x mol

Therefore multiply your concentration in $mol/dm^3$ by $M_r$

$M_r$ ($HNO_3$) = (1 x 1) + (1 x 14) + (3 x 16) = 1 + 14 + 48 = 63g/mol

3 $mol/dm^3$ x 63g / mol = <u>$189g/dm^3$</u>

**Worked Example 2: What is the concentration of a 9.8$g/dm^3$ $H_2SO_4$ solution, in $mol/dm^3$?**

Mol = Mass ÷ $M_r$

Therefore divide your concentration in $g/dm^3$ by $M_r$

$M_r$ ($H_2SO_4$) = (2 x 1) + (1 x 32) + (4 x 16) = 2 + 32 + 64 = 98g/mol

9.8$g/dm^3$ ÷ 98g/mol = <u>0.1 $mol/dm^3$</u>

# Unit 3: Quantitative Chemistry (Triple Content)

**Reactions with Gases**

- Room temperature = 20°C
- Room pressure = 1 atmospheric pressure
- Equal amounts in moles of gases occupy the same volume under the same conditions of temperature and pressure.
- E.g. 5 mole hydrogen gas take up the same amount of room as 5 mole oxygen gas, as long as the temperature and pressure remain the same.
- 1 mole of any gas at room temperature and pressure takes up 24 $dm^2$.
- E.g. 5 mole hydrogen take up 5 x 24 = 120$dm^3$
- The volumes of gaseous reactants and products can be calculated from the balanced equation for the reaction.

**Worked Example:**

247g of copper carbonate decompose thermally decompose. What volume of carbon dioxide is produced?

$$CuCO_3\,(s) \longrightarrow CuO\,(s) + CO_2\,(g)$$

**Step 1: Calculate the $M_r$ of copper carbonate**

$M_r(CuCO_3)$ = (1 x 63.5) + (1 x 12) + (3 x 16) = 123.5g/mol

**Step 2: Calculate how many moles of copper carbonate there are**

Mass = $M_r$ x mole

247g = 123.5 g/mol x mole

Mole = 2 mol

**Step 3: Use the balanced equation to work out how many moles of $CO_2$ are produced.**

Ratio = 1:1, therefore 2 moles of $CO_2$ also producted

**Step 4: Calculate Volume**

1 mole takes up 24$dm^3$

2 moles takes up 2 x 24$dm^3$ = 48$dm^3$

**Titration**

If you have two solutions that react completely, and you know EITHER the concentrations of both and the volume of one OR the volumes of both and the concentration of the other, you can work out the missing information.

Concentration 1 x Volume 1 = Concentration 2 x Volume 2

E.g. If 0.1$dm^3$ of 0.5mol/$dm^3$ HCl react with 0.2$dm^3$ of NaOH, what is the concentration of the NaOH?

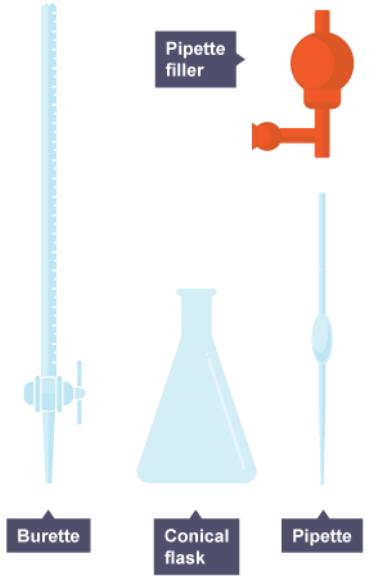
C1 x V1 = C2 x V2

0.1 x 0.5 = C2 x 0.2

0.1 x 0.5 ÷ 0.2 = C2

0.25 mol /$dm^3$ = C2

Make sure the units are the same on each side (e.g. both in $dm^3$ or in ml!)




**Titration Technique**

- Fill the **burette** with the solution of **known concentration**.
- Use the **pipette** to measure exactly 25ml of the unknown concentration solution into the **conical flask**.
- Add an appropriate **indicator**.
- Place the flask on a **white tile** to help you observe the colour change.
- Slowly add from the burette to the pipette, regularly **swirling** until you reach the **endpoint**.

**BBC Bitesize: Titration**