Chemistry Knowledge Organiser

Quantitative Chemistry (Trilogy Science)

Calculating relative atomic mass (A_r)

Relative atomic mass is the average mass of an atom that takes account of the **abundance** of the **isotopes** of the **element**.

$$A_r = \frac{\text{(mass number} \times \text{percentage) of isotope } \mathbf{A} + \text{(mass number} \times \text{percentage) of isotope } \mathbf{B}}{100}$$

	Mass number	Percentage abundance (%)
Isotope **A**	39	93.3
Isotope **B**	41	6.7

$$A_r = \frac{(39 \times 93.3) + (41 \times 6.7)}{100}$$

$A_r = 39.134$

Relative atomic mass values are given on the periodic table as the number above the symbol.

Relative formula mass

The relative formula mass is the total of the relative atomic masses of all of the atoms in the formula of a compound.

The relative atomic masses will be given in the question or can be found on the periodic table.

Example: $FeCl_3$

$M_r = (1 \times Fe) + (3 \times Cl)$
$M_r = (1 \times 56) + (3 \times 35.5)$
$M_r = 56 + 106.5$
$M_r = 162.5$

Example: $Ca(NO_3)_2$

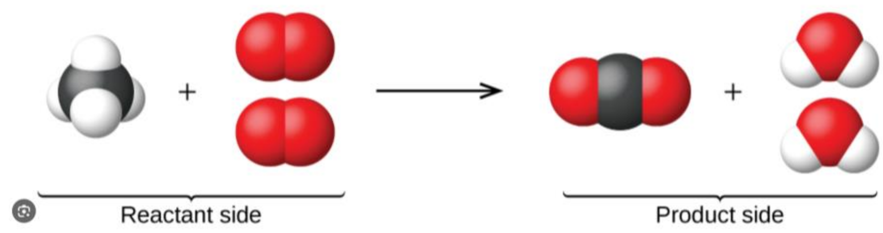
$M_r = (1 \times Ca) + (2 \times N) + (6 \times O)$
$M_r = (1 \times 40) + (2 \times 14) + (6 \times 16)$
$M_r = 40 + 28 + 96$
$M_r = 164$

If there are brackets in the formula, multiply the number of atoms of each element in the brackets by the number after the brackets.

Conservation of mass

The **law of conservation of mass** states that no **atoms** are lost or made during a chemical reaction.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

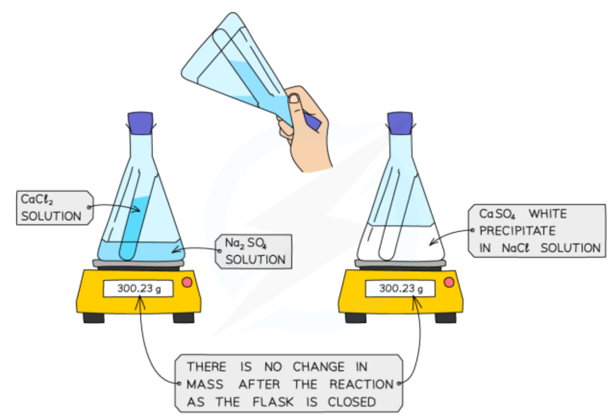


This means that the total mass of the reactants will equal the total mass of the products.

So chemical reactions can be represented by balanced symbol equations that have an equal number of **atoms** of each type of **element** on each side of the equation.

Demonstrating conservation of mass

The law of conservation of mass can be demonstrated by showing that the mass of all the reactants equals the mass of all the products.



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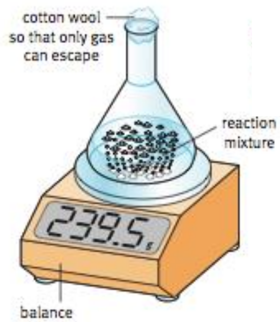
Reactions producing gases

In reactions which produce gases, if the gas is free to escape the reaction container then the mass will **decrease**. This is due to the mass of the gas that has escaped from the container, which would equal the decrease in mass.

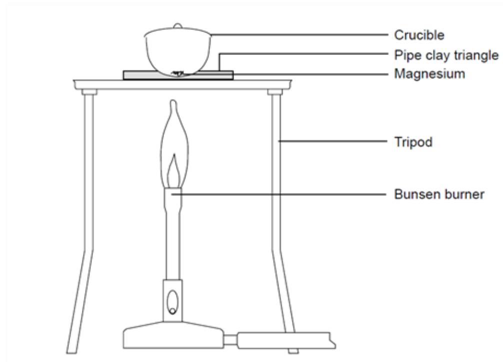


The cotton wool **lets** the gas escape but stops droplets of liquid escaping.

If the gas is trapped, the mass will remain constant.

Reactions where a reactant is a gas

If one of the reactants is a gas, in particular a gas in the air such as oxygen, then the mass will **increase**. This is due to the gas from the air adding onto the reactants, which would have equal mass to the increase in mass.



The lid on the crucible **stops** the magnesium oxide escaping but must be raised to let more oxygen in.

Uncertainty in measurements

Often when a reading is repeated the value obtained will be slightly different each time.

It is usual to calculate a **mean value** and to use this in analysing the data.

How far the largest and smallest readings were from the mean value is called the **uncertainty**, and is sometimes expressed in the following form:

$$7.2 \pm 0.3 \text{ g}$$

This means that the lowest reading was 7.2 - 0.3 = 6.9 g and the highest was 7.2 + 0.3 = 7.5g.

Moles (higher tier only)

Amounts in chemistry are measured in **moles**, sometimes abbreviated to mol.

LEARN:

$$\text{Moles} = \frac{\text{mass (in g)}}{M_r}$$

So for a substance 1 mole = relative formula mass in grams.

1 mole of a substance contains the same number of particles – Avogadro's number (6.02×10^{23}).

Example: How many molecules in 1.2 moles of water?

Number of molecules = moles x Avogadro's number
Number of molecules = $1.2 \times 6.02 \times 10^{23}$
Number of molecules = 7.224×10^{23}

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Using balanced equations (higher tier)

The big number in a balanced equation can be related to the number of **moles** of that substance that reacts.

For example: $Mg + 2HCl \rightarrow MgCl_2 + H_2$

shows that one **mole** of Mg reacts with two moles of HCl to produce one **mole** of $MgCl_2$ and one **mole** of H_2.

The big numbers in the balanced equation form a ratio of the number of **moles** of each substance reacting. Therefore, in the example above, the ratio between Mg and HCl will be 1:2 and if there are 2 **moles** of Mg then 4 **moles** of HCl will react.

Reacting mass calculations (higher tier)

This can be used to calculate the mass of a reactant or product:

Step:	Example: What mass of O_2 is is needed for the complete combustion of 80 g of CH_4?
1. Work out moles of substance you are given the mass of	Mass CH_4 = 80 g, M_r = 16 Moles Na_2CO_3 = 80/16 = 5
2. Use ratio from equation to work out moles of substance you are asked about	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ Ratio CH_4:O_2 = 1:2 So moles O_2 = 5 x 2 = 10
3. Work out mass of substance you are asked about	Moles O_2 = 10, M_r = 2 x 16 = 32 Mass O_2 = moles x M_r = 10 x 32 = 320 g

Finding reacting ratios (higher tier)

The balancing numbers (big numbers) in a symbol equation can be calculated from the masses of reactants and products by converting the masses in grams to amounts in **moles** and converting the numbers of moles to simple whole number ratios.

Example:

140 g of nitrogen (N_2) reacts with 30 g of hydrogen (H_2) to form ammonia.

$$\text{Moles of } N_2 = \frac{140}{28} = 5.0 \text{ mol} \qquad \text{Moles of } H_2 = \frac{30}{2} = 15 \text{ mol}$$

$$\text{Reacting ratio} \quad N_2 : H_2 = 5.0 : 15.0 = \frac{5.0}{5.0} : \frac{15.0}{5.0} = 1 : 3$$

$$\therefore \quad N_2 + 3H_2 \rightarrow$$

Limiting reagents (higher tier only)

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.

Concentration of solutions

The concentration of a solution can be measured in mass per given volume of solution, eg grams per dm^3 (g/dm^3).

LEARN:

$$\text{concentration } (g/dm^3) = \frac{\text{mass } (g)}{\text{volume } (dm^3)}$$