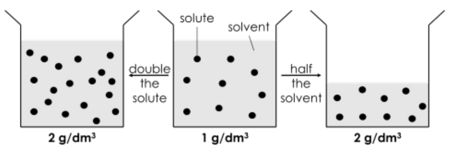
C3 Quantitative Chemistry
Combined Higher

Keywords

Keywords	
chemical formula	Represents chemicals with symbols and numbers
atom	The smallest particle of an element
element	A type of atom found on the periodic table
mole	Used to represent the number of particles. One mole is 6.02 x 10^{23} particles.
limiting reactant	The reactant that is completely used up during a reaction – this limits the amount of product.
excess	More of a chemical than you need for a reaction.
mean	The average of a set of numbers
range	The highest minus the lowest number
uncertainty	How close the highest and lowest numbers are to the mean

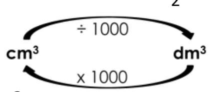
Concentrations of solutions

The concentration is the mass of the solute in 1 dm^3 (≈1 litre) of solution. If more solute is added, the concentration of the solution will increase. If there is less solvent, the concentration of a solution will also increase .



Example: Find the concentration of the solution formed when 2 g of HCl is dissolved in 400 cm^3 of H_2O.

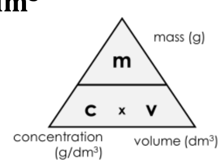
Step 1: Convert cm^3 to dm^3



$$\textbf{volume in dm}^3 = \frac{400}{1000} = 0.4\ dm^3$$

Step 2: Use equation

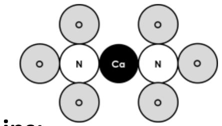
$$c = \frac{m}{v} = \frac{2}{0.4} = 0.05\ g/dm^3$$



Chemical formulae

In a chemical formula, each new symbol starts with a capital letter. The number after each symbol shows how many of that element are in the chemical. No number after the symbol means there is one. Brackets are used to group elements - the number after the brackets shows how many of that group there are in the chemical.

Example: Calcium nitrate (has two lots of NO_3)



$Ca(NO_3)_2$
- 3 x 2 = 6 x O (oxygen)
- 1 x 2 = 2 x N (nitrogen)
- 1 x Ca (calcium)

$Ca(NO_3)_2$ contains:
- 3 elements - Ca, N and O
- 9 atoms - 1 Ca, 2 Ns and 6 Os

Conservation of mass

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

C	**+**	**O_2**	⟶	**CO_2**
3 g		8 g		?
mass of reactants **= 11 g**				so mass of products **= 11 g**

mass of the products = mass of the reactants

Also **M_r products = M_r reactants**

Example:

	2 Mg	**+**	**O_2**	⟶	**2 MgO**
M_r x balancing number	2 x 24		1 x 32		2 x 40
	M_r of reactants **= 80**				M_r of products **= 80**

In an open system, the mass can change during an experiment involving gases as the gas can be 'added' or 'escape'.
- If a reactant is a gas, the mass increases (eg. 2 Ca + O_2 → 2 CaO).
- If a product is a gas, the mass decreases (eg. $ZnCO_3$ → ZnO + CO_2)

The mole (mol)

A mole represents the number of particles in a reaction. These particles can be ions, electrons, atoms, molecules etc. The number of particles in one mole is the same as Avagadro's constant – 6.02 x 10^{23}.

Examples: 1 mole of He contains 6.02 x10^{23} atoms
1 mole of H_2O contains 6.02 x10^{23} molecules
1 mole of NaBr contains 6.02 x10^{23} Na^+ ions

Uncertainty in measurements

When we repeat an experiment, we can use the results to tell us how reliable the experiment was.

Example: A beaker containing 100 cm^3 water was heated by 20 °C. The experiment was repeated, and the times taken in seconds were **90**, **100**, **93, 95**.

1. Calculate the mean
   (add numbers then divide by number of numbers)
   $$\frac{90 + 100 + 93 + 95}{4} = \frac{378}{4} = 94.5\ s$$
2. Calculate the range (*highest – lowest*)
   $$100 - 90 = 10\ s$$
3. Calculate the uncertainty *(range ÷ 2)*
   $$\frac{10}{2} = 5\ s$$
4. Result is mean ± uncertainty = **94.5 ± 5 s**

Relative formula mass (M_r)

The M_r is the mass of one mole of a chemical in grams. It is found by adding the mass numbers of each atom in a chemical:

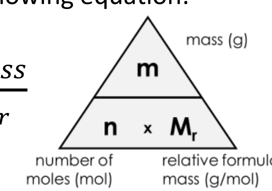
$MgCO_3$ (mass numbers)
- 3 x 16 = 48
- \+ 1 x 12 = 12
- \+ 1 x 24 = 24
- **84**

$Ca(NO_3)_2$ (mass numbers)
- 6 x 16 = 96
- \+ 2 x 14 = 28
- \+ 1 x 40 = 40
- **166**

Masses and moles

As the M_r is the mass of one mole of a chemical, the number of moles can be found using the following equation:

$$moles = \frac{mass}{M_r}$$



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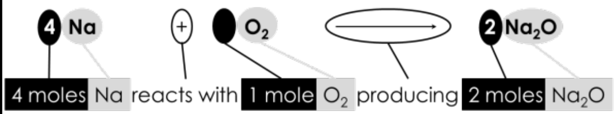
Balanced chemical equations

Balancing numbers

The 'balancing number' before each chemical shows the ratio of the number of moles of each chemical that reacts. No number before the chemical means there is only one mole reacting.



Mole ratios

The **mole ratios** from a balanced chemical equation can be used to find the moles of one chemical from the moles of another.

Example: $N_2 + 3H_2 \longrightarrow 2NH_3$

moles = 1.5 : ?

÷3 ×2

If 1.5 moles H_2 react with excess N_2 how many moles of NH_3 are formed?

$$moles\ NH_3 = \frac{1.5}{3}\ x\ 2 = 1\ \text{mol}$$

Masses to balanced equations

The mass can be used to find the mole ratio and write a balanced chemical equation

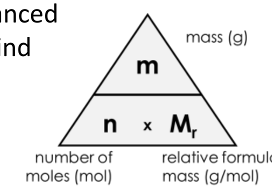
Example: 36 g Mg reacts with 24 g of O_2 producing 60 g of MgO. Find the balanced chemical equation for this reaction.

	Mg	+	O_2	→	MgO
1. Write mass	36		24		60
2. Find moles	$n = \frac{m}{M_r} = \frac{36}{24} = 1.5$		$n = \frac{m}{M_r} = \frac{24}{32} = 0.75$		$n = \frac{m}{M_r} = \frac{60}{40} = 1.5$
3. ÷ smallest number	$\frac{1.5}{0.75} = 2$		$\frac{0.75}{0.75} = 1$		$\frac{1.5}{0.75} = 2$
4. Balanced equation	**2** Mg	+	O_2	→	**2** MgO

Always check the equation is balanced!

Finding the mass of one chemical from the mass of another (mole calculations!)

By calculating the number of moles of one chemical, we can use the mole ratio from the balanced chemical equation to calculate the number of moles of another chemical. From this we can find the mass we need or would expect to produce. These calculations have three steps:



1 - Find the moles. 2 - Use mole ratios. 3 - Answer the question.

Note *– it is useful to keep track of the moles by writing them under the balanced equation*

Example

126 g of magnesium carbonate thermally decomposed forming magnesium oxide and carbon dioxide. What mass of magnesium oxide was produced?

$MgCO_3 \rightarrow MgO + CO_2$

69 g of sodium was reacted with chlorine gas to make sodium chloride. What mass of chlorine gas was needed for the reaction?

$2\ Na + Cl_2 \rightarrow 2\ NaCl$

1. Find the moles

$\text{moles} = \frac{\text{mass}}{M_r}$ M_r of $MgCO_3$ = 84 (see page 1!)

$\text{moles } MgCO_3 = \frac{126}{84} = 1.5$ mol

$\text{moles} = \frac{\text{mass}}{M_r}$ M_r of Na = 23

$\text{moles Na} = \frac{69}{23} = 3$ mol

2. Use mole ratios

ratio $MgCO_3$: MgO = 1 : 1

moles $MgCO_3$ = 1.5 mol

$MgCO_3 \longrightarrow MgO + CO_2$

moles = 1.5 ... 1.5 (÷ 1, x 1)

ratio Na : Cl_2 = 2 : 1

moles Cl_2 = 1.5 mol

$2\ Na + Cl_2 \longrightarrow 2\ NaCl$

moles = 3 : 1.5 (÷ 2, x 1)

3. Answer the question

mass = moles x M_r

M_r of MgO = 24 + 16 = 40

mass = 1.5 x 40 = **<u>60 g</u>**

mass = moles x M_r

M_r of Cl_2 = 2 x 35.5 = 71

mass = 1.5 x 71 = **<u>142 g</u>**

Excess and limiting reactants

Chemists often use an excess of one reactant to make sure that all of the other reactant is used up. The reactant which is used up is called the limiting reactant as this limits the amount of product formed in the reaction. Therefore, when performing mole calculations it is important to use the moles of the limiting reactant to calculate the moles of product formed.