C3 Quantitative Chemistry. **RP2 = Titrations**

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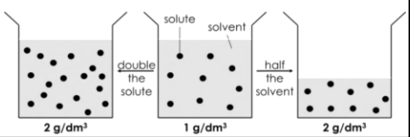
Keywords

chemical formula	Represents chemicals with symbols and numbers
mole	Used to represent the number of particles. One mole is 6.02×10^{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.
titration	Method used to accurately determine an unknown concentration
concordant	Results within 0.1 cm^3 of each other
yield	The amount (mass, volume etc) of product produced in 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

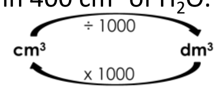
Concentrations of solutions (in g/dm^3)

This is the mass of solute in 1 dm^3 (1 litre) of solution. The concentration will increase if solute is added, or solvent is removed.



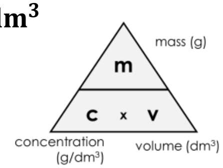
Example: Find the concentration of the solution formed when 2 g of HCl is dissolved in 400 cm^3 of H_2O.

Step 1: Convert cm^3 to dm^3



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

Step 2: Use equation



$$c = \frac{m}{v} = \frac{2}{0.4} = 5\ g/dm^3$$

The mole (mol)

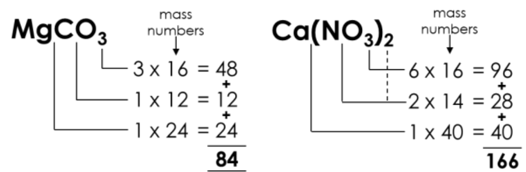
A mole represents the number of particles in a reaction. These particles can be ions, electrons, atoms, molecules etc. The number of particles in one mole is the same as Avogadro's constant – 6.02×10^{23}.

Example:

1 mole of NaBr contains 6.02×10^{23} Na^+ ions

Relative formula mass (M_r)

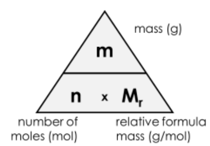
The M_r is the mass of one mole of a chemical in grams. It is found by adding the mass numbers of each atom in a chemical:



Mole equations

Masses

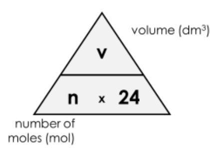
$$n = \frac{m}{M_r}$$



Gases

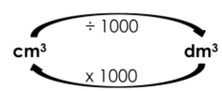
At room temperature and pressure (20 °C, 1 atm) one mole of any gas occupies a volume of 24 dm^3. Therefore this equation is used:

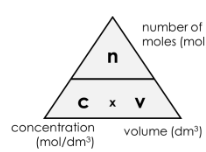
$$n = \frac{v}{24}$$



Solutions

$$n = c \times v$$





Conservation of mass

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

2 Mg	**+**	O_2	→	**2 MgO**
2 x 24		32		2 x 40
M_r of reactants **= 80**				M_r of products = **80**

Therefore:

mass of the products = mass of the reactants

Also **M_r products = M_r reactants**

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. $2\ Ca + O_2 \rightarrow 2\ CaO$).
- If a product is a gas, the mass decreases (eg. $ZnCO_3 \rightarrow ZnO + CO_2$)

Uncertainty in measurements

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

1. Calculate the mean : $\mathbf{90 + 100 + 93 + 95 = \frac{378}{4} = 94.5\ s}$

(add numbers then divide by number of numbers)

2. Calculate the range: $\mathbf{100 - 90 = 10\ s}$

(*highest – lowest*)

3. Calculate the uncertainty: $\mathbf{\frac{10}{2} = 5\ s}$

(range ÷ 2)

4. Result is mean ± uncertainty: **94.5 ± 5 s**

Using the mole equations

Calculations can be performed by inputting numbers into the appropriate equation.

***Example:* Find the mass of 0.75 moles of Na_2O.**

M_r Na_2O = 23 x 2 + 16 = 62

m = n x M_r = 0.75 x 62 = **46.5 g**

Sometimes two equations need to be used. As the equations are all linked by the number of moles:

1. **Find the moles**
2. **Answer the question**

***Example:* What mass of O_2 gas has a volume of 36 dm³?**

1. $n = \frac{v}{24} = \frac{36}{24} = 1.5$ mol

2. m O_2 = n x M_r (M_r O_2 = 2 x 16 = 32)
= 1.5 x 32 = **48 g**

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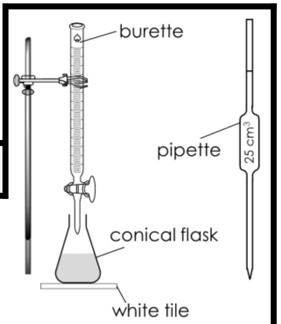


Titrations

Titrations are used to accurately find an unknown concentration of a chemical.

Method: *(can swap the acid and alkali)*

1. Use the pipette and pipette filler to add 25 cm^3 of alkali to a clean conical flask.
2. Add a few drops of indicator and put the conical flask on a white tile.
3. Fill the burette with acid and record the volume.
4. Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.
5. Stop adding the acid when the end-point is reached (the appropriate colour change happens) and record the final volume of acid.
6. Repeat until you have concordant results (within 0.1 cm^3).

Results: Use two decimal places.

Volume acid / cm^3	Rough	1st titration	2nd titration	3rd titration
Initial volume (cm^3)	0.30	5.80	10.90	15.80
Final volume (cm^3)	5.80	10.90	15.80	21.00
Titre (cm^3) (final vol. – initial vol.)	**5.50**	**5.10**	**4.90**	**5.20**

Concordant results: only use these to calculate the mean

Atom economy

This measures the proportion of starting material that is converted into useful product.

$$atom\ economy = \frac{M_r\ of\ desired\ product}{M_r s\ of\ all\ reactants}\ x\ 100$$

Example: What is the atom economy for making hydrogen by reacting coal with steam?

$$C + 2\,H_2O \longrightarrow CO_2 + 2\,H_2$$

12   2 x 18   2 x 2

M_r of reactants **= 48** M_r of desired products = **4**

$$\frac{4}{48}\ x\ 100 = \mathbf{8.3\ \%}$$

Note - If only useful products are produced then the atom economy is 100 %.

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.

4 Na + O_2 → **2** Na_2O

4 moles Na reacts with 1 mole O_2 producing 2 moles Na_2O

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

moles = 1.5 : ?

÷3 → ×2

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

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

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.

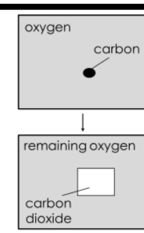
	Mg	+	O_2	→	MgO
1. Write mass	36		24		60
2. Find moles	$n = \frac{m}{M_r} = \frac{36}{24} = 1.5$		$n = \frac{m}{M_r} = \frac{24}{32} = 0.75$		$n = \frac{m}{M_r} = \frac{60}{40} = 1.5$
3. ÷ smallest number	$\frac{1.5}{0.75} = 2$		$\frac{0.75}{0.75} = 1$		$\frac{1.5}{0.75} = 2$
4. Balanced equation	**2**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

Always check the equation is balanced!

Excess and limiting reactants

$$C + O_2 \longrightarrow CO_2$$

C	O_2	CO_2
Small amount of carbon	Room full of oxygen	moles produced depends on the moles of the limiting reactant (carbon)
limiting reactant	**excess**	



Mole calculations

These calculations have three steps:

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

$n = \frac{m}{M_r}$ (masses) $n = v\ x\ c$ (solutions) $n = \frac{v}{24}$ (gas volumes) → **mole ratios** (from balanced equation)

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

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

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

1

$n = \frac{m}{M_r}$ M_r of Na = 23

$n\ Na = \frac{69}{23} = 3\ mol$

2

ratio Na : Cl_2 = 2 : 1

n Cl_2 = 1.5 mol

2 Na + Cl_2 → 2 NaCl

n = 3 : 1.5 (÷ 2, x 1)

3

$v = n\ x\ 24$ = 1.5 x 24 = **<u>136 dm^3</u>**

Percentage yield *(should always be less than 100!)*

In a reaction, not all of the reactants necessarily turn into the expected products. This could be because:

- a reaction is reversible
- product is lost when separated from reaction mixture
- the reactants reacted in a different way than expected

The % yield compares the amount of product obtained with the amount you'd expect in a 'perfect' reaction.

$$\%\ yield = \frac{actual\ mass}{theoretical\ mass}\ x\ 100$$

Example: Rob expected to get 186g of copper sulfate crystals, but got 158g. Calculate the percentage yield.

$$\%\ yield = \frac{158}{186}\ x\ 100 = \mathbf{84.9}\ \%$$

Note - The theoretical mass can be calculated by using mole calculations