

5. Formulae, equations and amounts of substance

The mole is the key concept for chemical calculations

DEFINITION: **The mole** is the amount of substance in grams that has the same number of particles as there are atoms in 12 grams of carbon-12.

DEFINITION: **Relative atomic mass** is the **average mass** of one atom compared to one twelfth of the mass of one atom of carbon-12

DEFINITION: **Molar Mass** is the mass in grams of 1 mole of a substance and is given the unit of g mol^{-1}

Molar Mass for a compound can be calculated by adding up the mass numbers (from the periodic table) of each element in the compound
eg $\text{CaCO}_3 = 40.1 + 12.0 + 16.0 \times 3 = 100.1$

For most calculations we will do at AS we will use the following 3 equations

Learn these equations carefully and what units to use in them.

1. For pure solids and gases

$$\text{amount} = \frac{\text{mass}}{\text{MolarMass}}$$

Unit of Mass: grams
Unit of amount : mol

2. For Gases

$$\text{Gas Volume (dm}^3\text{)} = \text{amount} \times 24$$

This equation give the volume of a gas at room pressure (1atm) and room temperature 25°C.

It is usually best to give your answers to 3sf

3. For solutions

$$\text{Concentration} = \frac{\text{amount}}{\text{volume}}$$

Unit of concentration: mol dm^{-3} or M
Unit of Volume: dm^3

Converting volumes

$$\begin{aligned} \text{cm}^3 &\rightarrow \text{dm}^3 \div 1000 \\ \text{cm}^3 &\rightarrow \text{m}^3 \div 1000\ 000 \\ \text{dm}^3 &\rightarrow \text{m}^3 \div 1000 \end{aligned}$$

Typical mole calculations

Some Simple calculations using above equations

Example 1: What is the amount, in mol, in 35.0g of CuSO_4 ?

$$\begin{aligned} \text{amount} &= \text{mass}/M_r \\ &= 35 / (63.5 + 32 + 16 \times 4) \\ &= 0.219 \text{ mol} \end{aligned}$$

Example 2: What is the concentration of solution made by dissolving 5.00g of Na_2CO_3 in 250 cm^3 water?

$$\begin{aligned} \text{amount} &= \text{mass}/M_r \\ &= 5 / (23.1 \times 2 + 12 + 16 \times 3) \\ &= 0.0472 \text{ mol} \\ \text{conc} &= \text{amount}/\text{Volume} \\ &= 0.0472 / 0.25 \\ &= 0.189 \text{ mol dm}^{-3} \end{aligned}$$

Example 3 : What is the volume in dm^3 at room temperature and pressure of 50.0g of Carbon dioxide gas ?

$$\begin{aligned} \text{amount} &= \text{mass}/M_r \\ &= 50 / (12 + 16 \times 2) \\ &= 1.136 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Gas Volume (dm}^3\text{)} &= \text{amount} \times 24 \\ &= 1.136 \times 24 \\ &= 27.26 \text{ dm}^3 \text{ (or } 27.3 \text{ dm}^3 \text{ to 3 sig fig)} \end{aligned}$$

Avogadro's Constant

The mole is the amount of substance in grams that has the same number of particles as there are atoms in 12 grams of carbon-12.

Avogadro's Constant

There are 6.02×10^{23} atoms in 12 grams of carbon-12. Therefore explained in simpler terms 'One mole of any specified entity contains 6.02×10^{23} of that entity':

Avogadro's Constant can be used for atoms, molecules and ions

1 mole of copper atoms will contain 6.02×10^{23} atoms
1 mole of carbon dioxide molecules will contain 6.02×10^{23} molecules
1 mole of sodium ions will contain 6.02×10^{23} ions

No of particles = amount of substance (in mol) X Avogadro's constant

Example 4 : How many atoms of Tin are there in a 6.00 g sample of Tin metal?

$$\begin{aligned}\text{amount} &= \text{mass}/A_r \\ &= 6/118.7 \\ &= 0.05055 \text{ mol} \\ \text{Number atoms} &= \text{amount} \times 6.02 \times 10^{23} \\ &= 0.05055 \times 6.02 \times 10^{23} \\ &= 3.04 \times 10^{22}\end{aligned}$$

Example 5 : How many chloride ions are there in a 25.0 cm³ of a solution of magnesium chloride of concentration 0.400 mol dm⁻³ ?

$$\begin{aligned}\text{amount} &= \text{concentration} \times \text{Volume} \\ \text{MgCl}_2 &= 0.400 \times \mathbf{0.025} \\ &= 0.0100 \text{ mol} \\ \text{Amount of chloride ions} &= 0.0100 \times 2 \\ &= 0.0200 \quad \text{There are two moles of chloride ions for every one mole of MgCl}_2 \\ \text{Number ions of Cl}^- &= \text{amount} \times 6.02 \times 10^{23} \\ &= 0.0200 \times 6.02 \times 10^{23} \\ &= 1.204 \times 10^{22}\end{aligned}$$

Ideal Gas Equation

The ideal gas equation is a slightly more difficult way of working out a gas volume. It can calculate the volume for any temperature and pressure though so is a more useful method.

$$\mathbf{PV = nRT}$$

Unit of Pressure (P): Pa
Unit of Volume (V): m³
Unit of Temp (T): K
n = moles
R = 8.31 JK⁻¹mol⁻¹

Example 6: What is the mass of Cl₂ gas that has a pressure of 100kPa, temperature 293K, volume 500cm³.
(R = 8.31 JK⁻¹mol⁻¹)

$$\begin{aligned}\text{moles} &= \frac{PV}{RT} \leftarrow \begin{array}{l} 100 \text{ kPa} = 100\,000 \text{ Pa} \\ 500 \text{ cm}^3 = 0.0005 \text{ m}^3 \end{array} \\ &= 100\,000 \times 0.0005 / (8.31 \times 293) \\ &= 0.0205 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Mass} &= \text{moles} \times M_r \\ &= 0.0205 \times (35.5 \times 2) \\ &= 1.46 \text{ g}\end{aligned}$$

Converting temperature

$$\text{°C} \rightarrow \text{K add 273}$$

Empirical Formula

Definition: An empirical formula is the **simplest** ratio of atoms of each **element** in the compound.

General method

Step 1 : Divide each mass (or % mass) by the atomic mass of the element

Step 2 : For each of the answers from step 1 divide by the smallest one of those numbers.

Step 3: sometimes the numbers calculated in step 2 will need to be multiplied up to give whole numbers.

These whole numbers will be the empirical formula.

The same method can be used for the following types of data:

1. masses of each element in the compound
2. percentage mass of each element in the compound

Example 7 : Calculate the empirical formula for a compound that contains 1.82g of K, 5.93g of I and 2.24g of O

Step1: Divide each mass by the atomic mass of the element

$$\begin{array}{lll} \text{K} = 1.82 / 39.1 & \text{I} = 5.93/126.9 & \text{O} = 2.24/16 \\ = 0.0465 \text{ mol} & = 0.0467\text{mol} & = 0.14 \text{ mol} \end{array}$$

Step 2 For each of the answers from step 1 divide by the smallest one of those numbers.

$$\begin{array}{lll} \text{K} = 0.0465/0.0465 & \text{I} = 0.0467/0.0465 & \text{O} = 0.14 / 0.0465 \\ =1 & = 1 & = 3 \end{array}$$

Empirical formula = KIO_3

Molecular formula from empirical formula

Definition: A molecular formula is the **actual** number of atoms of each element in the compound.

From the relative molecular mass (M_r) work out how many times the mass of the empirical formula fits into the M_r .

Example 8 : work out the molecular formula for the compound with an empirical formula of $\text{C}_3\text{H}_6\text{O}$ and a M_r of 116

$\text{C}_3\text{H}_6\text{O}$ has a mass of 58

The empirical formula fits twice into M_r of 116

So molecular formula is $\text{C}_6\text{H}_{12}\text{O}_2$

The M_r does not need to be exact to turn an empirical formula into the molecular formula because the molecular formula will be a whole number multiple of the empirical formula

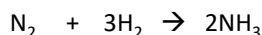
Example 9: 0.150g of a volatile liquid was injected into a sealed gas syringe. The gas syringe was placed in an oven at 70°C at a pressure of 100kPa and a volume of 80cm^3 was measured. What is the M_r of the volatile liquid ? ($R = 8.31$)

$$\begin{aligned} \text{moles} &= PV/RT \\ &= 100\,000 \times 0.00008 / (8.31 \times 343) \\ &= 0.00281 \text{ mol} \end{aligned}$$

$$\begin{aligned} M_r &= \text{mass/moles} \\ &= 0.15 / 0.00281 \\ &= 53.4 \text{ g mol}^{-1} \end{aligned}$$

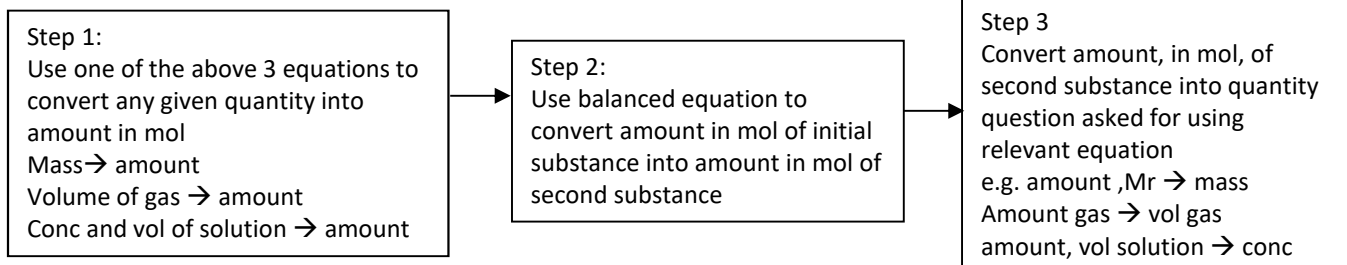
$$\begin{aligned} 100 \text{ kPa} &= 100\,000 \text{ Pa} \\ 80 \text{ cm}^3 &= 0.00008 \text{ m}^3 \end{aligned}$$

Converting quantities between different substances using a balanced equation

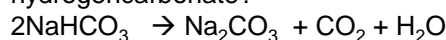


The balancing (stoichiometric) numbers are mole ratios
e.g. 1 mol of N_2 reacts with 3 mol of H_2 to produce 2 mol of NH_3

Typically we are given a quantity of one substance and are asked to work out a quantity for another substance in the reaction. Any of the above three equations can be used.



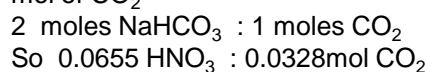
Example 10: What mass of Carbon dioxide would be produced from heating 5.50 g of sodium hydrogencarbonate?



Step 1: work out amount, in mol, of sodium hydrogencarbonate

$$\begin{aligned}\text{amount} &= \text{mass} / \text{Mr} \\ &= 5.5 / 84 \\ &= 0.0655 \text{ mol}\end{aligned}$$

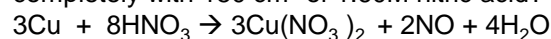
Step 2: use balanced equation to give amount in mol of CO_2



Step 3: work out mass of CO_2

$$\begin{aligned}\text{Mass} &= \text{amount} \times \text{Mr} \\ &= 0.0328 \times 44.0 \\ &= 1.44 \text{ g}\end{aligned}$$

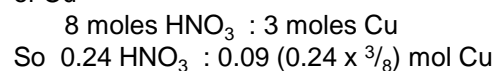
Example 11: What mass of Copper would react completely with 150 cm^3 of 1.60M nitric acid?



Step 1: work out moles of nitric acid

$$\begin{aligned}\text{amount} &= \text{conc} \times \text{vol} \\ &= 1.6 \times 0.15 \\ &= 0.24 \text{ mol}\end{aligned}$$

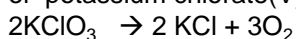
Step 2: use balanced equation to give moles of Cu



Step 3: work out mass of Cu

$$\begin{aligned}\text{Mass} &= \text{amount} \times \text{Mr} \\ &= 0.09 \times 63.5 \\ &= 5.71 \text{ g}\end{aligned}$$

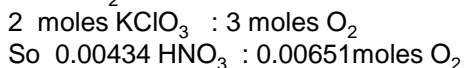
Example 12: What volume in cm^3 of oxygen gas would be produced from the decomposition of 0.532 g of potassium chlorate(V)?



Step 1: work out amount, in mol, of potassium chlorate(V)?

$$\begin{aligned}\text{amount} &= \text{mass} / \text{Mr} \\ &= 0.532 / 122.6 \\ &= 0.00434 \text{ mol}\end{aligned}$$

Step 2: use balanced equation to give amount in mol of O_2



Step 3: work out volume of O_2

$$\begin{aligned}\text{Gas Volume (dm}^3\text{)} &= \text{amount} \times 24 \\ &= 0.00651 \times 24 \\ &= 0.156 \text{ dm}^3 \\ &= 156 \text{ cm}^3\end{aligned}$$

Hydrated salt

A Hydrated salt contains water of crystallisation

$\text{Cu}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$
hydrated copper (II) nitrate(V).

$\text{Cu}(\text{NO}_3)_2$
Anhydrous copper (II) nitrate(V).

Example 13

$\text{Na}_2\text{SO}_4 \cdot x\text{H}_2\text{O}$ has a molar mass of 322.1, Calculate the value of x

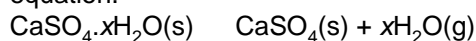
$$\text{Molar mass } x\text{H}_2\text{O} = 322.1 - (23 \times 2 + 32.1 + 16 \times 4) \\ = 180$$

$$X = 180/18 \\ = 10$$

Heating in a crucible

This method could be used for measuring mass loss in various thermal decomposition reactions and also for mass gain when reacting magnesium in oxygen.

The water of crystallisation in calcium sulphate crystals can be removed as water vapour by heating as shown in the following equation.



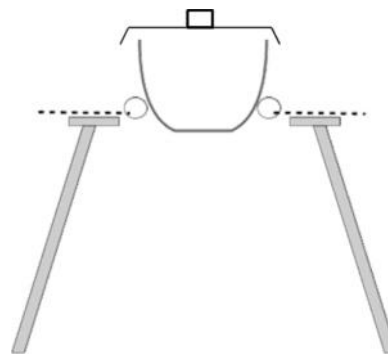
Method.

- Weigh an empty clean dry crucible and lid .
- Add 2g of hydrated calcium sulphate to the crucible and weigh again
- Heat strongly with a Bunsen for a couple of minutes
- Allow to cool
- Weigh the crucible and contents again
- Heat crucible again and reweigh until you reach a constant mass (do this to ensure reaction is complete).

Large amounts of hydrated calcium sulphate, such as 50g, should not be used in this experiment as the decomposition is like to be incomplete.

The crucible needs to be dry otherwise a wet crucible would give an inaccurate result. It would cause mass loss to be too large as water would be lost when heating.

The lid improves the accuracy of the experiment as it prevents loss of solid from the crucible but should be loose fitting to allow gas to escape.



Small amounts the solid , such as 0.100 g, should **not** be used in this experiment as errors in weighing are too high.

Example 14. 3.51 g of hydrated zinc sulphate were heated and 1.97 g of anhydrous zinc sulphate were obtained.

Use these data to calculate the value of the integer x in $\text{ZnSO}_4 \cdot x\text{H}_2\text{O}$

Calculate the mass of $\text{H}_2\text{O} = 3.51 - 1.97 = 1.54\text{g}$

Calculate moles of ZnSO_4	$= \frac{1.97}{161.5}$	Calculate moles of H_2O	$= \frac{1.54}{18}$
	$= 0.0122$		$= 0.085$

Calculate ratio of mole of ZnSO_4 to H_2O	$= \frac{0.0122}{0.0122}$		$= \frac{0.085}{0.0122}$
	$= 1$		$= 7$

$$X = 7$$

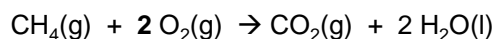
Reacting Volumes of Gas

Equal volumes of any gases measured under the same conditions of temperature and pressure contain equal numbers of molecules (or atoms if the gas is monatomic)

1 mole of any gas at room pressure (1atm) and room temperature 25°C will have the volume of 24dm³

Volumes of gases reacting in a balanced equation can be calculated by simple ratio

Example 15 If one burnt 500 cm³ of methane at 1atm and 300K what volume of Oxygen would be needed and what volume of CO₂ would be given off under the same conditions?



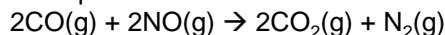
1 mole 2 mole 1 mole

500cm³ 1dm³ 500cm³

Simply multiply
gas volume x2

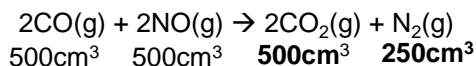
Example 16

An important reaction which occurs in the catalytic converter of a car is



In this reaction, when 500 cm³ of CO reacts with 500 cm³ of NO at 650 °C and at 1 atm.

Calculate the **total** volume of gases produced at the same temperature and pressure ?



total volume of gases produced = 750cm³

Using a gas syringe

Gas syringes can be used for a variety of experiments where the volume of a gas is measured, possibly to work out moles of gas or to follow reaction rates.

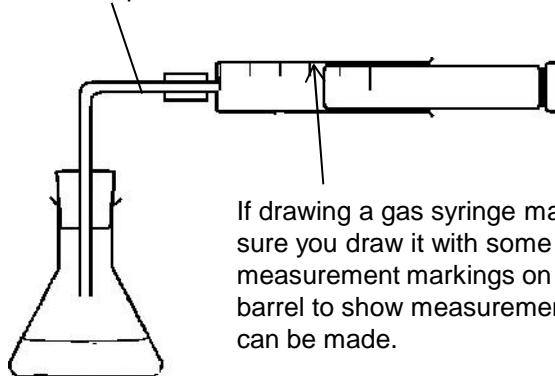
The volume of a gas depends on pressure and temperature so when recording volume it is important to note down the temperature and pressure of the room.

Moles of gas can be calculated from gas volume (and temperature and pressure) using molar gas volume or the ideal gas equation

Potential errors in using a gas syringe

- gas escapes before bung inserted
- syringe sticks
- some gases like carbon dioxide or sulphur dioxide are soluble in water so the true amount of gas is not measured.

Make sure you don't leave gaps in your diagram where gas could escape



Making a solution

- Weigh the sample bottle containing the required mass of solid on a 2 dp balance
- Transfer to beaker and reweigh sample bottle
- Record the difference in mass

- Add 100cm^3 of distilled water to the beaker. Use a glass rod to stir to help dissolve the solid.

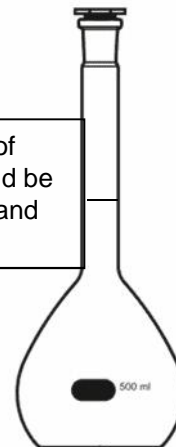
- Sometimes the substance may not dissolve well in cold water so the beaker and its contents could be heated gently until all the solid had dissolved.

- Pour solution into a 250cm^3 graduated flask via a funnel.
- Rinse beaker and funnel and add washings from the beaker and glass rod to the volumetric flask.
- make up to the mark with distilled water using a dropping pipette for last few drops.
- Invert flask several times to ensure uniform solution.

Diluting a solution

- Pipette 25cm^3 of original solution into volumetric flask
- make up to the mark with distilled water using a dropping pipette for last few drops.
- Invert flask several times to ensure uniform solution.

Alternatively the known mass of solid in the weighing bottle could be transferred to beaker, washed and washings added to the beaker.



Remember to fill so the bottom of the meniscus sits on the line on the neck of the flask. With dark liquids like potassium manganate it can be difficult to see the meniscus.

Shake the volumetric flask thoroughly to ensure a uniform concentration

Using a pipette is more accurate than a measuring cylinder because it has a smaller sensitivity error

Use a teat pipette to make up to the mark in volumetric flask to ensure volume of solution accurately measured and one doesn't go over the line

Titration

The method for carrying out the titration

- rinse equipment** (burette with acid, pipette with alkali, conical flask with distilled water)
- pipette 25 cm³ of alkali into conical flask**
- touch surface of alkali with pipette** (to ensure correct amount is added)
- adds acid solution from burette**
- make sure the jet space** in the burette **is filled** with acid
- add a few drops of indicator** and refer to colour change at end point
- phenolphthalein [pink (alkali) to colourless (acid): end point pink colour just disappears] [use if NaOH is used]
- methyl orange [yellow (alkali) to red (acid): end point orange] [use if HCl is used]
- use a white tile underneath the flask to help observe the colour change
- add acid to alkali whilst **swirling the mixture** and **add acid dropwise at end point**
- note burette reading** before and after addition of acid
- repeats titration** until **at least 2 concordant results** are obtained- two readings within 0.1 of each other

There will be a small amount of the liquid left in the pipette when it has been emptied. Do not force this out. The pipette is calibrated to allow for it.

If the jet space in the burette is not filled properly prior to commencing the titration it will lead to errors if it then fills during the titration, leading to a larger than expected titre reading.

Working out average titre results

Only make an average of the concordant titre results

Recording results

- Results should be clearly recorded in a table
- Result should be recorded in full (i.e. both initial and final readings)
- Record titre volumes to 2dp (0.05 cm³)**

Testing batches

In quality control it will be necessary to do titrations/testing on several samples as the amount/concentration of the chemical being tested may vary between samples.

Safety precautions

Acids and alkalis are corrosive (at low concentrations acids are irritants)

Wear eye protection and gloves

If spilled immediately wash affected parts after spillage

If substance is unknown treat it as potentially toxic and wear gloves.

A conical flask is used in preference to a beaker because it is easier to swirl the mixture in a conical flask without spilling the contents.

Distilled water can be added to the conical flask during a titration to wash the sides of the flask so that all the acid on the side is washed into the reaction mixture to react with the alkali. It does not affect the titration reading as water does not react with the reagents or change the number of moles of acid added.

Only distilled water should be used to wash out conical flasks between titrations because it does not add and extra moles of reagents

If **2 or 3 values are within 0.10cm³** and therefore **concordant** or close then we can say results are accurate and **reproducible** and **the titration technique is good/ consistent**

Titrating mixtures

If titrating a mixture to work out the concentration of an active ingredient it is necessary to consider if the mixture contains other substances that have acid base properties.

If they don't have acid base properties we can titrate with confidence.

Titration Calculations

Example 17: 23.6cm³ of H₂SO₄ neutralised 25.0cm³ of 0.150M NaOH. What is the concentration of the H₂SO₄?
 $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

Step 1: work out amount, in mol, of sodium hydroxide
 amount = conc x vol
 = 0.15 x 0.025
 = 0.00375 mol

Step 2: use balanced equation to give moles of H₂SO₄
 2 moles NaOH : 1 moles H₂SO₄
 So 0.00375 NaOH : 0.001875 moles H₂SO₄

Step 3 work out concentration of H₂SO₄
 conc= amount/Volume
 = 0.001875 / 0.0236
 = 0.0794 mol dm⁻³

Example 18: A 25.0cm³ sample of vinegar was diluted in a 250cm³ volumetric flask. This was then put in a burette and 23.10cm³ of the diluted vinegar neutralised 25 cm³ of 0.100 M NaOH. What is the concentration of the vinegar in gdm⁻³ ?
 $\text{CH}_3\text{CO}_2\text{H} + \text{NaOH} \rightarrow \text{CH}_3\text{CO}_2\text{Na} + \text{H}_2\text{O}$

Step 1: work out amount, in mol, of sodium hydroxide
 amount = conc x vol
 = 0.10 x 0.025
 = 0.00250 mol

Step 2: use balanced equation to give moles of CH₃CO₂H
 1 moles NaOH : 1 moles CH₃CO₂H
 So 0.00250 NaOH : 0.00250 moles CH₃CO₂H

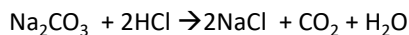
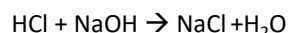
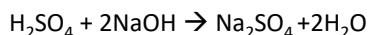
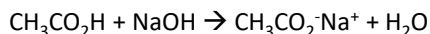
Step 3 work out concentration of diluted CH₃CO₂H in 23.1 (and 250 cm³) in moldm⁻³
 conc= amount/Volume
 = 0.00250 / 0.0231
 = 0.108 mol dm⁻³

Step 4 work out concentration of original concentrated CH₃CO₂H in 25cm³ in moldm⁻³
 conc = 0.108 x 10 = 1.08 mol dm⁻³

Step 5 work out concentration of CH₃CO₂H in original concentrated 25 cm³ in gdm⁻³
 conc in gdm⁻³ = conc in mol dm⁻³ x Mr
 = 1.08 x 60 = 64.8 g dm⁻³

To turn concentration measured in mol dm⁻³ into concentration measured in g dm⁻³ multiply by Mr of the substance
 conc in g dm⁻³ = conc in mol dm⁻³ x Mr
 The concentration in g dm⁻³ is the same as the mass of solute dissolved in 1dm³

Common Titration Equations



Example 19

950 mg of impure calcium carbonate tablet was crushed. 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid, an excess, was then added and the mixture was transferred to a volumetric flask. The volume was made up to exactly 100 cm³ with distilled water. 10.0 cm³ of this solution was titrated with 11.1cm³ of 0.300 mol dm⁻³ sodium hydroxide solution. What is the percentage of CaCO₃ by mass in the tablet

1. Calculate the number of moles of sodium hydroxide used

amount = conc x vol
 = 0.30 x 0.0111
 = 0.00333 mol

2. Work out number of moles of hydrochloric acid left in 10.0 cm³

use balanced equation to give moles of HCl
 1 mol NaOH : 1 mol HCl
 So 0.00333 NaOH : 0.00333 moles HCl

3. Calculate the number of moles of hydrochloric acid left in 100 cm³ of solution

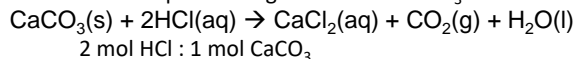
Moles in 100cm³ = 0.00333 x 10
 = 0.0333

4. Calculate the number of moles of HCl that reacted with the indigestion tablet.

In original HCl 50.0 cm³ of 1.00 mol dm⁻³ there is 0.05moles

moles of HCl that reacted with the indigestion tablet. = 0.05 - 0.0333 = 0.0167

5 Use balanced equation to give moles of CaCO₃



2 mol HCl : 1 mol CaCO₃
 So 0.0167 HCl : 0.00835 moles CaCO₃

6. work out the mass of CaCO₃ in original tablet

mass= amount x Mr
 = 0.00835 x 100 = 0.835 g

percentage of CaCO₃ by mass in the tablet = 0.835/0.950 x 100 = 87.9 %

Uncertainty

Readings and Measurements

Readings

the values found from a single judgement when using a piece of equipment

Measurements

the values taken as the difference between the judgements of two values (e.g. using a burette in a titration)

The uncertainty of a reading (one judgement) is at least ± 0.5 of the smallest scale reading.
The uncertainty of a measurement (two judgements) is at least ± 1 of the smallest scale reading.

Calculating Apparatus Uncertainties

Each type of apparatus has a sensitivity uncertainty

- balance ± 0.001 g
- volumetric flask ± 0.1 cm³
- 25 cm³ pipette ± 0.1 cm³
- burette ± 0.05 cm³

Calculate the percentage error for each piece of equipment used by

$$\% \text{ uncertainty} = \pm \frac{\text{uncertainty}}{\text{Measurement made on apparatus}} \times 100$$

e.g. for pipette

$$\% \text{ uncertainty} = 0.05 / 25 \times 100$$

To calculate the maximum percentage apparatus uncertainty in the final result add all the individual equipment uncertainties together.

To decrease the apparatus uncertainties you can either decrease the sensitivity uncertainty by using apparatus with a greater resolution (finer scale divisions) or you can increase the size of the measurement made.

Uncertainty of a measurement using a burette.

If the burette used in the titration had an uncertainty for each reading of ± 0.05 cm³ then during a titration two readings would be taken so the uncertainty on the titre volume would be ± 0.10 cm³.

Reducing uncertainties in a titration

Replacing measuring cylinders with pipettes or burettes which have lower apparatus uncertainty will lower the error

To reduce the uncertainty in a burette reading it is necessary to make the titre a larger volume. This could be done by: increasing the volume and concentration of the substance in the conical flask or by decreasing the concentration of the substance in the burette.

If looking at a series of measurements in an investigation the experiments with the smallest readings will have the highest experimental uncertainties.

Reducing uncertainties in measuring mass

Using a more accurate balance or a larger mass will reduce the uncertainty in weighing a solid
Weighing sample before and after addition and then calculating difference will ensure a more accurate measurement of the mass added.

Calculating the percentage difference between the actual value and the calculated value

If we calculated an Mr of 203 and the real value is 214, then the calculation is as follows:

Calculate difference $214 - 203 = 11$

$$\% = 11 / 214 \times 100$$
$$= 5.41\%$$

If the %**uncertainty** due to the apparatus $<$ percentage difference between the actual value and the calculated value then there is a discrepancy in the result due to other errors.

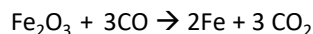
If the %**uncertainty** due to the apparatus $>$ percentage difference between the actual value and the calculated value then there is no discrepancy and all errors in the results can be explained by the sensitivity of the equipment.

% Yield

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

% yield in a process can be lowered through incomplete reactions, side reactions, losses during transfers of substances, losses during purification stages.

Example 20: 25.0g of Fe_2O_3 was reacted and it produced 10.0g of Fe. What is the percentage yield?

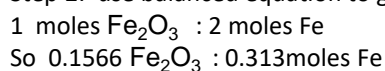


First calculate maximum mass of Fe that could be produced

Step 1: work out amount in mol of Iron oxide

$$\begin{aligned}\text{amount} &= \text{mass} / \text{Mr} \\ &= 25 / 159.6 \\ &= 0.1566 \text{ mol}\end{aligned}$$

Step 2: use balanced equation to give moles of Fe



Step 3: work out mass of Fe

$$\begin{aligned}\text{Mass} &= \text{amount} \times \text{Mr} \\ &= 0.313 \times 55.8 \\ &= 17.5\text{g}\end{aligned}$$

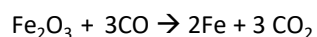
$$\begin{aligned}\% \text{ yield} &= (\text{actual yield} / \text{theoretical yield}) \times 100 \\ &= (10 / 17.48) \times 100 \\ &= 57.2\%\end{aligned}$$

% Atom Economy

$$\text{percentage atom economy} = \frac{\text{Mass of useful products}}{\text{Mass of all reactants}} \times 100$$

Do take into account balancing numbers when working out % atom economy.

Example 21: What is the % atom economy for the following reaction where Fe is the desired product assuming the reaction goes to completion?



$$\begin{aligned}\% \text{ atom economy} &= \frac{(2 \times 55.8)}{(2 \times 55.8 + 3 \times 16) + 3 \times (12 + 16)} \times 100 \\ &= 45.8\%\end{aligned}$$

Sustainable chemistry requires chemists to design processes with high atom economy that minimise production of waste products.

Reactions where there is only one product where all atoms are used making product are ideal and have 100% atom economy.
e.g. $\text{CH}_2=\text{CH}_2 + \text{H}_2 \rightarrow \text{CH}_3\text{CH}_3$

If a process does have a side, waste product the economics of the process can be improved by selling the bi-product for other uses

Reactions of Acids

A **Salt** is formed when the H^+ ion of an acid is replaced by a metal ion or an ammonium ion

The most common strong acids are :
Hydrochloric (HCl), sulphuric (H_2SO_4) and nitric (HNO_3) acid;

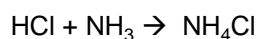
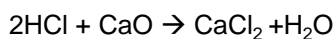
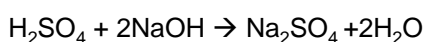
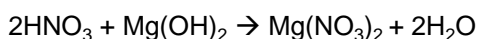
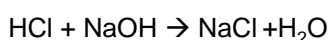
Neutralisation reactions form salts

Bases neutralise acids. Common bases are metal oxides, metal hydroxides and ammonia.

An Alkali is a soluble base that releases OH^- ions in aqueous solution;
The most common alkalis are sodium hydroxide (NaOH), potassium hydroxide (KOH) and aqueous ammonia (NH_3)

Common Acid Reaction Equations

ACID + BASE → SALT + WATER



Acid + Carbonate → Salt + Water + Carbon Dioxide



Observations : In carbonate reactions there will be Effervescence due to the CO_2 gas evolved and the solid carbonate will dissolve. The temperature will also increase.

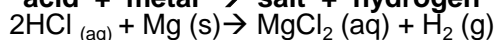
acid + metal → salt + hydrogen



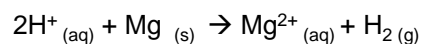
Observations: These reaction will effervesce because H_2 gas is evolved and the metal will dissolve

Ionic equations for reactions of acids with metals, carbonates, bases and alkalis

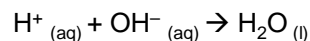
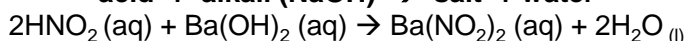
acid + metal → salt + hydrogen



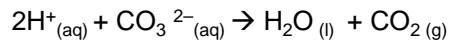
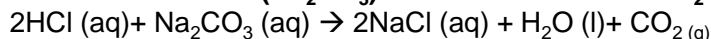
Ionic Equations



acid + alkali (NaOH) → salt + water

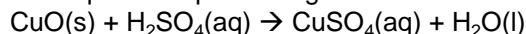


acid + carbonate (Na_2CO_3) → salt + water + CO_2



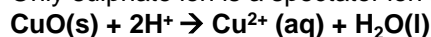
Example 22

The equation representing the reaction between copper(II) oxide and dilute sulphuric acid is



What is the ionic equation?

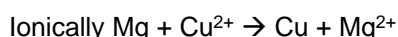
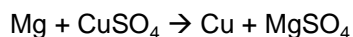
Only sulphate ion is a spectator ion in this case because it's the only ion not changing state



Displacement Reactions

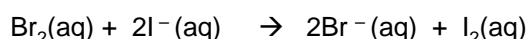
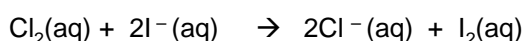
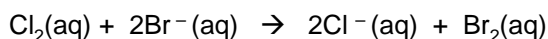
Metal displacement reactions

More reactive metals will displace less reactive metals from their compounds



Halogen displacement reactions

A halogen that is a strong oxidising agent will displace a halogen that has a lower oxidising power from one of its compounds



See topic 4b Halogens for more detail

Precipitation Reactions

Insoluble salts can be made by mixing appropriate solutions of ions so that a **precipitate** is formed

Lead nitrate (aq) + sodium chloride (aq) → **lead chloride (s)** + sodium nitrate (aq)

These are called **precipitation** reactions. A **precipitate is a solid**

When making an insoluble salt, normally the salt would be removed by **filtration**, washed with distilled water to remove soluble impurities and then **dried on filter paper**

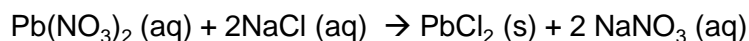
Writing ionic equations for precipitation reactions

We usually write ionic equations to show precipitation reactions. Ionic equations only show the ions that are reacting and leave out spectator ions.

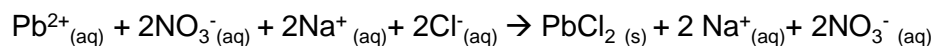
Spectator ions are ions that are not

- Not changing state
- Not changing oxidation number

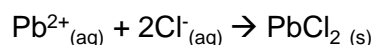
Take full equation



Separate (aq) solutions into ions



Cancel out spectator ions leaving ionic equation



Hazards and Risks

A **hazard** is a substance or procedure that can have the potential to do harm.

Typical hazards are toxic/flammable/harmful/irritant/corrosive/oxidizing/carcinogenic

RISK: This is the probability or chance that harm will result from the use of a hazardous substance or a procedure

In the laboratory we try to minimise the risk

Irritant - dilute acid and alkalis- wear goggles

Corrosive- stronger acids and alkalis wear goggles

Flammable – keep away from naked flames

Toxic – wear gloves- avoid skin contact- wash hands after use

Oxidising- Keep away from flammable / easily oxidised materials

Hazardous substances in low concentrations or amounts will not pose the same risks as the pure substance.

Safely dealing with excess acid

Sodium hydrogen carbonate (NaHCO_3) and calcium carbonate (CaCO_3) are good for neutralising excess acid in the stomach or acid spills because they are not corrosive and will not cause a hazard if used in excess. They also have no toxicity if used for indigestion remedies but the CO_2 produced can cause wind. Magnesium hydroxide is also suitable for dealing with excess stomach acid as it has low solubility in water and is only weakly alkaline so not corrosive or dangerous to drink (unlike the strong alkali sodium hydroxide). It will also not produce any carbon dioxide gas.