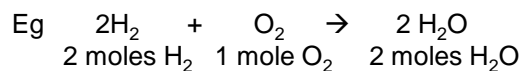


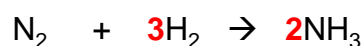
1.24 Calculations and Chemical Reactions

Converting quantities between different substances using a balanced equation

A balanced chemical equation tells us the number of particles of a substance which react with another substance. It also therefore tells us the ratios in moles in which these substances react.



All calculations involving chemical reactions will need quantities given to be converted into amount in moles. Then the balanced equation can be used to identify reacting mole ratios



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

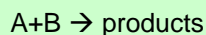
Most calculations will only involve two substances from the reaction and it is only necessary to identify the mole reacting ratio between those two substances.

So if a reaction involved 0.5 mol of N_2 we can say it would react with 1.5 mol of H_2 using the 1 N_2 :3 H_2 ratio

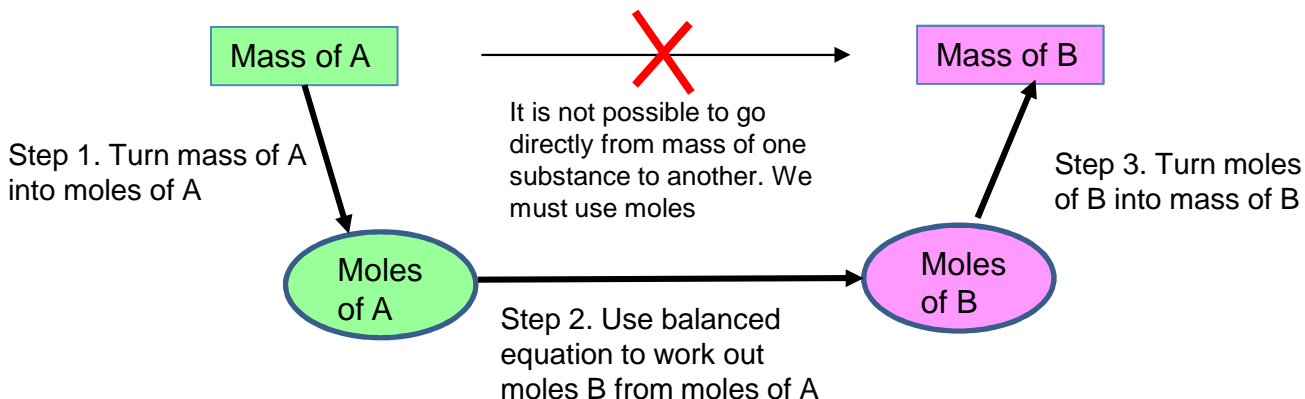
Or if 0.4 moles of NH_3 was produced we would need 0.6 mol of H_2 using the 2 NH_3 :3 H_2 ratio

Reacting Mass Questions

When doing calculations involving masses of different substances in a chemical reactions, to convert from mass of A to mass of substance B we must use moles.



A & B can be any two reactants or products

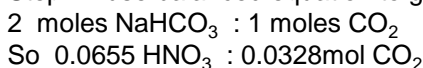


Example 1 Calculate the mass of carbon dioxide produced from heating 5.50 g of sodium hydrogencarbonate.
 $2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$

Step 1: calculate amount, in mol, of sodium hydrogencarbonate

$$\begin{aligned} \text{amount} &= \text{mass} / M_r \\ &= 5.50 / 84.0 \\ &= 0.0655 \text{ mol} \end{aligned}$$

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



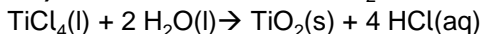
Step 3: calculate mass of CO_2

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

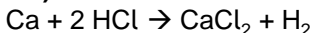
Remember stoichiometric balancing numbers are not used to calculate M_r . The M_r of NaHCO_3 is always 84.0 whatever the balancing number in front of it

Reacting Mass Questions

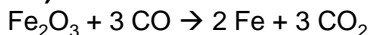
1.1) Calculate the mass of TiO_2 made from reacting 10.0g of TiCl_4 with water.



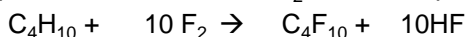
1.2) Calculate the mass of calcium chloride formed when 0.500 g of calcium reacts with hydrochloric acid.



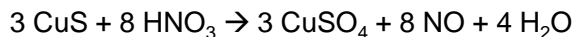
1.3) Calculate the mass of carbon monoxide needed to reduce 3.00 kg of iron oxide to iron.



1.4) Calculate the mass of F_2 needed to produce 150 g of C_4F_{10} .



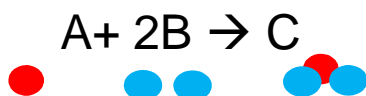
1.5) Calculate the mass of NO produced from reacting 0.500 g of CuS with concentrated nitric acid.



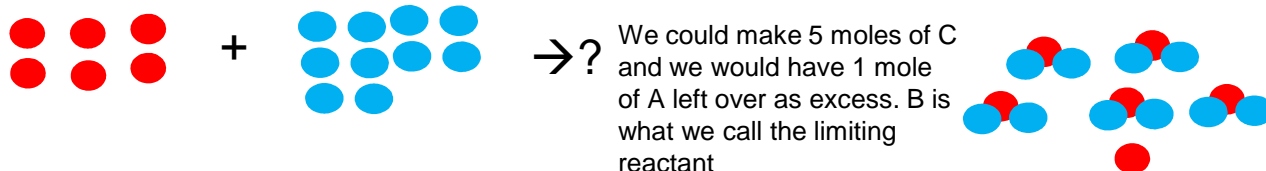
Limiting and excess reactants

If a question gives masses of two or more reactants we will need to consider whether one is in excess. In most cases when there are two reactants, after they have reacted we will be left with some of one reactant. This is the amount that is in excess. Students often find this a complex idea to understand.

Let us consider the reaction of A and B reacting with a ratio of 1A:2B

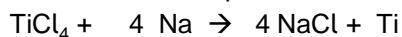


If we had 6 moles of A and 10 of B, which is in excess and how much C would be produced?



So if given masses of two reactants then it is necessary to work out moles of both substances and do the above exercise to work out which one is limited and which one is in excess.

Example 2 Calculate the maximum mass of titanium that could be produced from reacting 100 g of TiCl_4 with 80 g of sodium.



Step 1: work out amount, in mol, TiCl_4

$$\begin{aligned} \text{amount} &= \text{mass} / M_r \\ &= 100 / 189.9 \\ &= 0.527 \text{ mol} \end{aligned}$$

Step 1: work out amount, in mol, Na

$$\begin{aligned} \text{amount} &= \text{mass} / M_r \\ &= 80 / 23.0 \\ &= 3.48 \text{ mol} \end{aligned}$$

Step 2 use balanced equation to work out which reactant is in excess

Using 1 TiCl_4 :4 Na ratio we can see that 0.527 mol of TiCl_4 should react with 2.108 mol of Na. We actually have 3.48 mole of Na which is an excess of 1.372 moles. We can complete calculation using the limiting reactant of TiCl_4

Step 3: use balanced equation to work out amount in mol of Ti formed

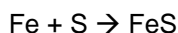
1 mol TiCl_4 : 1 mole Ti

So 0.527 mol TiCl_4 produces 0.527 mole Ti

Step 4: work out mass of Ti formed

$$\begin{aligned} \text{Mass} &= \text{amount} \times M_r \\ &= 0.527 \times 47.9 \\ &= 25.24 \text{ g} \end{aligned}$$

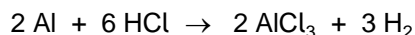
2.1) Calculate the maximum mass of iron sulfide that can be produced from 20.0 g of iron and 30.0g of sulfur



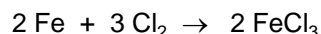
2.2) Calculate the maximum mass of ammonia that can be produced from 400 g of ammonium chloride and 100g of calcium oxide



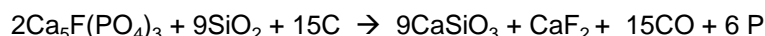
2.3) Calculate the maximum mass of aluminium chloride that can be produced from 25.0 g of aluminium and 50.0 g of HCl



2.4) Calculate the maximum mass of iron (III) chloride that can be produced from 50.0 g of Fe and 100g of chlorine gas



2.5) Calculate the maximum mass of phosphorus that can be produced from reacting 100g of $\text{Ca}_5\text{F}(\text{PO}_4)_3$, 100g of SiO_2 and 100g of Carbon



Calculations involving masses, solutions and gases

Commonly in questions converting between quantities of substances reacting we will use more than just mass data. We might have the volume and concentration of a solution, or the volume of a gas. We need to adapt our existing method or reacting masses to include other quantities. Any of the equations below can be used to convert quantities into moles.

1. For pure solids, liquids 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.

Or use the ideal gas equation to work out gas volumes at other temperatures and pressures

$$PV = nRT$$

3. For solutions

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

Unit of concentration: mol dm⁻³ or M
Unit of Volume: dm³

General method

Step 1:

Use one of the above equations to convert any given quantity into amount in mol
Mass → amount
Volume of gas → amount
Conc and vol of solution → amount

Step 2:

Use balanced equation to convert amount in mol of initial substance into amount in mol of second substance

Step 3

Convert amount, in mol, of second substance into quantity question asked for using relevant equation
e.g. amount, *Mr* → mass
Amount gas → vol gas
amount, vol solution → conc

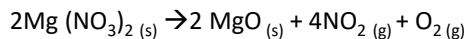
Example 3: Calculate the mass of copper that reacts completely with 150 cm³ of 1.60 mol dm⁻³ nitric acid
 $3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu}(\text{NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}$

Step 1: work out moles of nitric acid
amount = conc x vol
= 1.6 x 0.15
= 0.24 mol

Step 2: use balanced equation to give moles of Cu
8 moles HNO₃ : 3 moles Cu
So 0.24 HNO₃ : 0.09 (0.24 x ³/₈) mol Cu

Step 3: work out mass of Cu
mass = amount x *Mr*
= 0.09 x 63.5
= 5.71 g

Example 4: Calculate the total volume of gas produced in dm³ at 333K and 100kPa when 0.651 g of magnesium nitrate is heated.



Step 1: work out moles of magnesium nitrate
amount = mass / *Mr*
= 0.651 / 148.3
= 0.00439 mol

Step 2: use balanced equation to give moles of gas produced
2 moles Mg(NO₃)₂ : 4NO_{2(g)} + O_{2(g)} ie 5moles of gas
So 0.00439 Mg(NO₃)₂ : 0.01098 (0.00439 x ⁵/₂) moles gas

Step 3: work out volume of gas

$$\begin{aligned} \text{Volume} &= nRT/P \\ &= (0.01098 \times 8.31 \times 333) / 100000 \\ &= 0.000304\text{m}^3 \\ &= 0.303 \text{ dm}^3 \end{aligned}$$

Questions involving masses, solutions and gases

3.1) Hydrogen can be made by the reaction of hydrochloric acid with magnesium according to the equation
 $2\text{HCl} + \text{Mg} \rightarrow \text{MgCl}_2 + \text{H}_2$

Calculate the mass of hydrogen formed when 200cm^3 of hydrochloric acid of concentration 2.50 mol dm^{-3} reacts with an excess of magnesium.

3.2) Lead(II) nitrate can be produced by the reaction between nitric acid and lead(II) oxide as shown by the equation below.

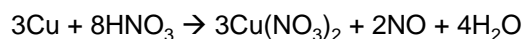


An excess of lead(II) oxide reacted with 250cm^3 of 1.25 mol dm^{-3} nitric acid.

Calculate the maximum mass of lead(II) nitrate which could have been obtained.

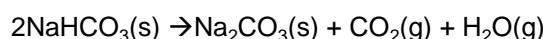
3.3) Calculate the volume, in cm^3 , of 1.50 mol dm^{-3} hydrochloric acid required to react completely with 1.20 g of magnesium hydroxide. $\text{Mg}(\text{OH})_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$

3.4) A volume of 150cm^3 of 1.75 mol dm^{-3} nitric acid was completely reacted with copper metal. The equation for the reaction is shown below. Calculate the mass of copper that would react completely with this amount of nitric acid.

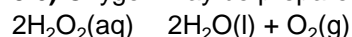


3.5) A sample of Sodium hydrogencarbonate was heated until completely decomposed. The CO_2 formed in the reaction occupied a volume of 352 cm^3 at $1.00 \times 10^5\text{ Pa}$ and 298 K .

Calculate the mass of the NaHCO_3 that has decomposed.



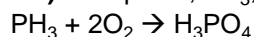
3.6) Oxygen may be prepared by the decomposition of hydrogen peroxide, H_2O_2 , as shown in the equation.



A 160 cm^3 sample of 2.68 mol dm^{-3} aqueous hydrogen peroxide was decomposed completely.

Calculate the mass and volume of oxygen gas produced at $25\text{ }^\circ\text{C}$ and 100 kPa .

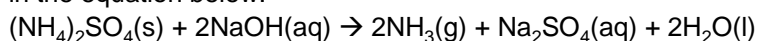
3.7) Phosphine, PH_3 , and oxygen can react to form phosphoric acid, H_3PO_4 , as shown in the equation below.



An excess of oxygen was mixed with 2.70 g of phosphine in a sealed container and allowed to react.

Calculate the mass of phosphoric acid formed in this reaction.

3.8) Ammonium sulfate reacts with sodium hydroxide to form ammonia, sodium sulfate and water as shown in the equation below.



A 3.24 g sample of ammonium sulfate reacted completely with 39.70 cm^3 of a sodium hydroxide solution.

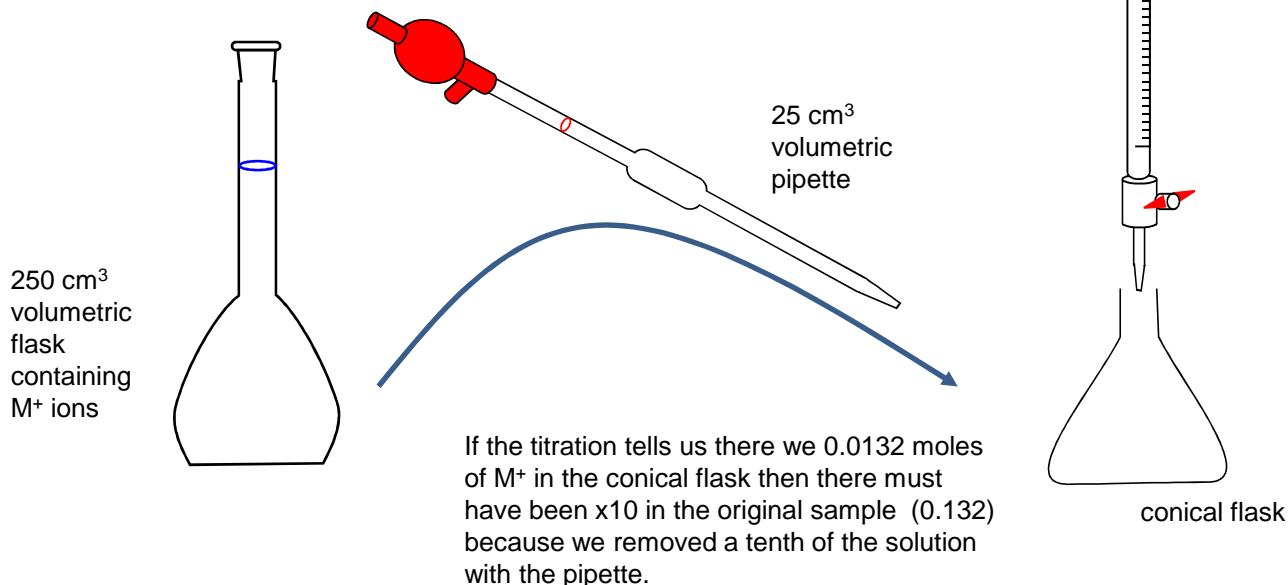
Calculate the concentration, in mol dm^{-3} , of the sodium hydroxide solution used.

Questions involving samples being taken

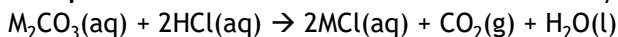
There are many variations on these calculations. It is not possible to learn an exact method for every type of calculation you could ever meet. You need to be able to go back to the chemistry and picture what is happening. Sometimes it is actually a good idea to draw out what is happening. Let's look at some common 'extra' steps.

Samples of solutions taken

Often in the laboratory we might prepare a solution of a reactant in a volumetric flask which we then want to react with another substance. Rather than reacting all the substance at once we often take out a sample to react. We need to take that into account when doing the calculation. This can be done very simply.



Example 5. An unknown metal carbonate reacts with hydrochloric acid according to the following equation.



A 3.96 g sample of M₂CO₃ was dissolved in distilled water to make 250 cm³ of solution. A 25.0 cm³ portion of this solution required 32.8 cm³ of 0.175 mol dm⁻³ hydrochloric acid for complete reaction.

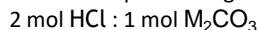
Calculate the *Mr* of M₂CO₃ and identify the metal M

1. Calculate the number of moles of HCl used

$$\begin{aligned} \text{amount} &= \text{conc} \times \text{vol} \\ &= 0.175 \times 0.0328 \\ &= 0.00574 \text{ mol} \end{aligned}$$

2. Work out number of moles of M₂CO₃ in 25.0 cm³ put in conical flask

use balanced equation to give moles of M₂CO₃



So 0.00574 NaOH : 0.00287 moles M₂CO₃

3. Calculate the number of moles M₂CO₃ acid in original 250 cm³ of solution

$$\begin{aligned} \text{Moles in } 250\text{cm}^3 &= 0.00287 \times 10 \\ &= 0.0287 \end{aligned}$$

4. work out the *Mr* of M₂CO₃

$$\begin{aligned} \text{Mr} &= \text{mass} / \text{amount} \\ &= 3.96 / 0.0287 = 138.0 \end{aligned}$$

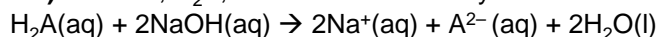
5. Work out Ar of M = $\frac{138-12-16 \times 3}{2}$

$$\text{Ar of M} = 39$$

M = potassium

Questions involving samples being taken

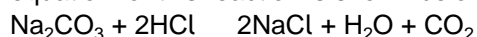
4.1) An acid, H_2A , reacts with sodium hydroxide as shown in the equation below.



A solution of this acid was prepared by dissolving 2.02 g of H_2A in water and making the volume up to 250 cm^3 in a volumetric flask.

A 25.0 cm^3 sample of this solution required 22.80 cm^3 of 0.150 $mol\ dm^{-3}$ aqueous NaOH for complete reaction. Calculate the relative molecular mass, M_r , of H_2A

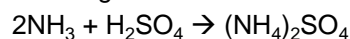
4.2) Sodium carbonate forms several hydrates of general formula $Na_2CO_3 \cdot xH_2O$. A 2.98 g sample of one of these hydrates was dissolved in water and the solution made up to 250 cm^3 . In a titration, a 25.0 cm^3 portion of this solution required 28.3 cm^3 of 0.170 $mol\ dm^{-3}$ hydrochloric acid for complete reaction. The equation for this reaction is shown below.



Calculate the relative molecular mass, M_r of the hydrated sodium carbonate and therefore the value of x .

4.3) 10.8 g of a solid monoprotic acid, HA was dissolved in water and made up to 250 cm^3 in a volumetric flask. 25.0 cm^3 portions of this were titrated against 0.200 $mol\ dm^{-3}$ sodium hydroxide, requiring 23.0 cm^3 . Calculate the relative molecular mass, M_r of the acid.

4.4) A 20.0 g sample of a domestic cleaning chemical containing ammonia (NH_3) was dissolved in water and the solution was made up to 500 cm^3 in a volumetric flask. A 25.0 cm^3 portion of this solution was then reacted with 26.8 cm^3 of 0.20 $mol\ dm^{-3}$ sulfuric acid. Calculate the percentage by mass of ammonia in the cleaning solution



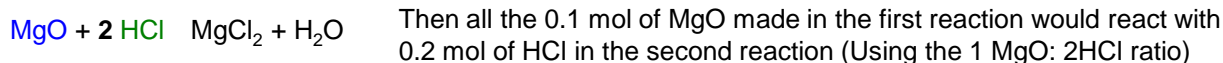
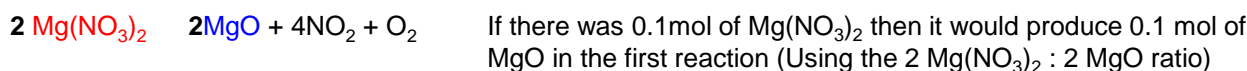
4.5) 1.70 g of a metal carbonate, M_2CO_3 , was dissolved in water and the solution was made up to 250 cm^3 in a volumetric flask. 25.0 cm^3 of this solution was then reacted with 24.6 cm^3 of 0.100 $mol\ dm^{-3}$ hydrochloric acid. Calculate the relative formula mass of M_2CO_3 and hence the relative atomic mass and identity of the metal M.

4.6) 1.00 cm^3 of concentrated hydrochloric acid was transferred with a graduated pipette to a 100 cm^3 volumetric flask. The volume was made up to 100 cm^3 with distilled water.

A 10.0 cm^3 portion of the diluted solution from the volumetric flask was titrated by NaOH and was neutralised by 24.35 cm^3 of sodium hydroxide of concentration 0.0500 $mol\ dm^{-3}$. Calculate the concentration of the original concentrated hydrochloric acid in $mol\ dm^{-3}$.

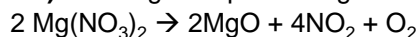
Calculations involving two or more reactions

One type of question is where there is a series of reactions when one occurs after the other. The product of the first reaction becomes the reactant of the second reaction.

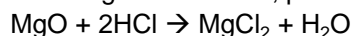


Remember the balancing numbers give the mole ratios for that equation. Don't try to link balancing numbers in one reaction to the balancing numbers in the second

5.1) A 2.23g sample of magnesium nitrate was fully decomposed.



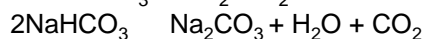
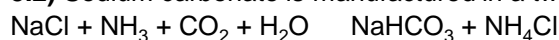
The magnesium oxide, produced was reacted with hydrochloric acid.



This sample of magnesium oxide required 33.2cm³ of hydrochloric acid for complete reaction.

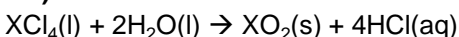
Calculate the concentration, in mol dm⁻³, of the hydrochloric acid.

5.2) Sodium carbonate is manufactured in a two-stage process as shown by the equations below.

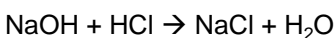


Calculate the maximum mass of sodium carbonate which could be obtained from 600 g of sodium chloride.

5.3) The chloride of an element **X** reacts with water according to the following equation.

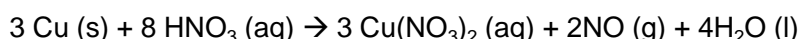


A 1.436 g sample of XCl_4 was added to water. The solid XO_2 was removed by filtration and the resulting solution was made up to 250 cm³ in a volumetric flask. A 25.0 cm³ portion of this solution was titrated against a 0.120 mol dm⁻³ solution of sodium hydroxide, of which 22.3 cm³ were required to reach the end point.

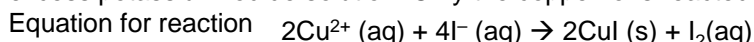


Calculate the relative molecular mass, M_r , of XCl_4 and then deduce the relative atomic mass, A_r , of element **X** and its identity.

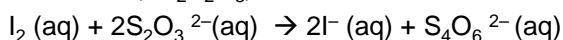
5.4) A old bronze coin of mass 2.15g containing copper and tin was dissolved in concentrated nitric acid and the NO gas escaped through boiling. The equation below shows the reaction of the copper with the acid



The solution was diluted with water to make up to 250cm³. 25cm³ portions were neutralised and added to excess potassium iodide solution. Only the copper ions reacted with the iodide ions to produce iodine.



The liberated iodine was then titrated with sodium thiosulfate. The iodine required 30.40cm³ of 0.100 mol dm⁻³ sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) to react.



Calculate the percentage of copper in the brass coin.

5.5) The amount of ozone in the atmosphere may be determined by passing air through a solution of acidified potassium iodide to form iodine in the following reaction $\text{O}_3 + 2 \text{I}^{-} + 2 \text{H}^{+} \rightarrow \text{O}_2 + \text{H}_2\text{O} + \text{I}_2$

The amount of iodine formed can be determined by titration with a solution of sodium thiosulfate of known concentration in the following reaction $\text{I}_2 + 2 \text{S}_2\text{O}_3^{2-} \rightarrow 2 \text{I}^{-} + \text{S}_4\text{O}_6^{2-}$

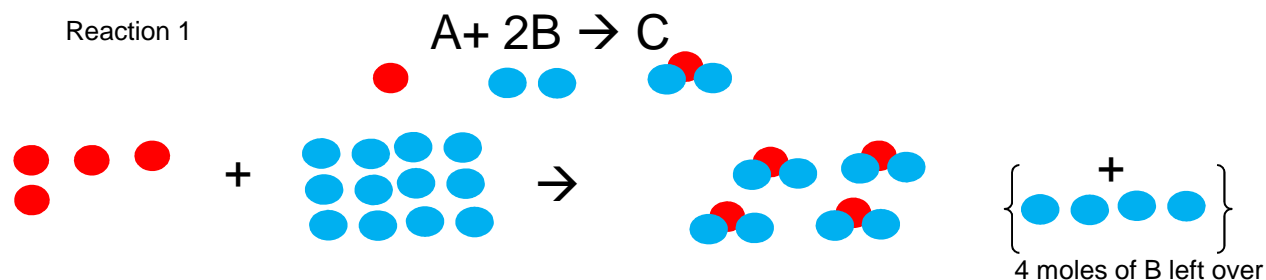
In an experiment to determine the amount of ozone in air, 100 m³ of air was bubbled through 100 cm³ of a solution containing an excess of acidified potassium iodide. The resulting solution was titrated against a solution of sodium thiosulfate of concentration 0.0167 mol dm⁻³. The volume of sodium thiosulfate solution required for the complete reaction was 24.80 cm³.

Calculate the amount in moles of ozone in the 100 m³ sample and calculate the volume of Ozone in m³. (assuming the volume is measured at room temperature and pressure). What percentage by volume of air is ozone?

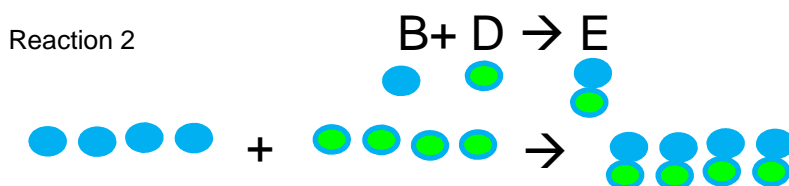
The second type of calculation involving two different reactions occurs when the first reaction leaves an excess reactant. The second reaction then involves the reaction of the excess reactant.

Think about the following series of reactions with A reacting with excess B in a first reaction. Then the remaining B reacts in a second reaction with reactant D. (equations below)

If we had 4 moles of A and 12 of B, which is in excess and how much B would be left over?



The 4 moles of B left over would react with 4 moles of D



Now try this problem. If 17 moles of B reacted with a unknown quantity of A and the unreacted B reacted with 3 moles of D. How many moles of A would there have been? *The answer is 7*

This logic is used to solve a type of calculation we call a **back titration**

Example 6: 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 11.1 cm³ of 0.300 mol dm⁻³ sodium hydroxide solution. Calculate the percentage of CaCO₃ by mass in the tablet.

950 mg of impure calcium carbonate

50.0 cm³ of 1.00 mol dm⁻³ HCl

100 cm³ volumetric flask containing unreacted HCl made up to 100 cm³ distilled water.

10 cm³ volumetric pipette transfers 10 cm³ of solution containing unreacted HCl

11.1 cm³ of 0.300 mol dm⁻³ sodium hydroxide

conical flask

This reaction occurs
 $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 All the CaCO₃ is reacted
 some HCl is left over

This reaction occurs
 $\text{NaOH} + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 The NaOH reacts with left over HCl in 10 cm³ solution

The calculation works backwards using the titration results first.

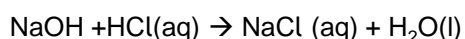
950 mg of impure calcium carbonate tablet was crushed. 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid, an excess, was then added and the mixture was transferred to a volumetric flask. The volume was made up to exactly 100 cm³ with distilled water. 10.0 cm³ of this solution was titrated with 11.1 cm³ of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

Calculate the percentage of CaCO₃ by mass in the tablet.

1. Calculate the amount in moles of sodium hydroxide used.

$$\begin{aligned}\text{amount} &= \text{conc} \times \text{vol} \\ &= 0.30 \times 0.0111 \\ &= 0.00333 \text{ mol}\end{aligned}$$

2. Work out the amount in moles of hydrochloric acid that reacted with the NaOH



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

This is the amount of moles of HCl that were in the conical flask and hence in the 10 cm³ pipette

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

$$\begin{aligned}\text{Moles in } 100\text{cm}^3 &= 0.00333 \times 10 \\ &= 0.0333 \text{ mol}\end{aligned}$$

The 10 cm³ pipetted sample came from the 100 cm³ volumetric flask so to work out the amount of moles in the flask multiply by 10

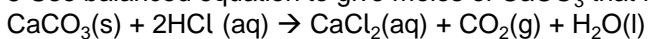
4. Calculate the amount in moles of HCl that reacted with the indigestion tablet.

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

moles of HCl that reacted with the indigestion tablet.
 $= 0.05 - 0.0333$
 $= 0.0167 \text{ mol}$

The amount of moles of HCl that reacted the CaCO₃ is the difference between the total amount of moles of HCl at the start and the moles of HCl left over at the end of this reaction.

- 5 Use balanced equation to give moles of CaCO₃ that reacted with the HCl



2 mol HCl : 1 mol CaCO₃

So 0.0167 HCl : 0.00835 moles CaCO₃

6. work out the mass of CaCO₃ in original tablet

$$\begin{aligned}\text{mass} &= \text{amount} \times M_r \\ &= 0.00835 \times 100.1 = 0.836 \text{ g}\end{aligned}$$

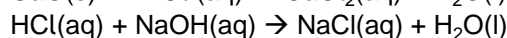
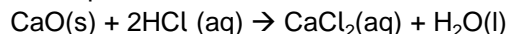
7. Calculate percentage of CaCO₃ by mass in the tablet
 $= 0.836 / 0.950 \times 100$
 $= 88.0 \%$

Divide the mass of calcium carbonate that reacted by the mass of the original impure tablet X100

Back Titration questions

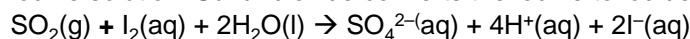
6.1) A 0.132 g of an impure sample of quicklime was dissolved in 50.0 cm³ of hydrochloric acid, concentration 0.100 mol dm⁻³. The excess hydrochloric acid was titrated with sodium hydroxide solution, concentration 0.100 mol dm⁻³, and 17.7 cm³ was needed to just neutralize the acid.

The equations for the reactions involved are shown below.



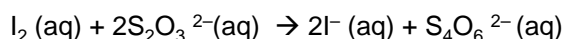
Calculate the percentage purity of the calcium oxide.

6.2) The amount of sulfur dioxide in the air can be measured by bubbling a known volume of air through iodine solution. Sulfur dioxide converts the iodine to iodide ions according to the following equation.



In an experiment, 75 m³ of air were passed through 50 cm³ of iodine, of concentration 0.0200 mol dm⁻³.

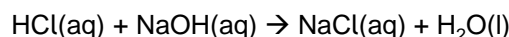
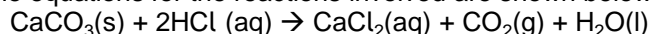
The remaining iodine was titrated with sodium thiosulfate solution and reacted with 14.70 cm³ of sodium thiosulfate, concentration 0.100 mol dm⁻³.



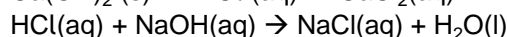
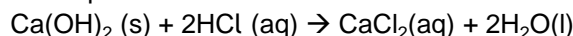
Calculate the moles of sulfur dioxide that were in the original 75 m³ sample.

6.3) 9.8 g of a indigestion remedy containing calcium carbonate was reacted with 100 cm³ of 2.00 mol dm⁻³ hydrochloric acid. (Only the calcium carbonate in the remedy reacts). The resulting solution was made up to 250 cm³ with water in a volumetric flask. A 25.0 cm³ portion of this solution required 28.9 cm³ of 0.200 mol dm⁻³ sodium hydroxide for neutralisation. Calculate the percentage by mass of the calcium carbonate in the indigestion remedy.

The equations for the reactions involved are shown below.



6.4) 1.13 g of an impure sample of calcium hydroxide was dissolved in 50.0 cm³ of hydrochloric acid, concentration 1.00 mol dm⁻³. The resulting solution was made up to 250 cm³ with water in a volumetric flask. A 25.0 cm³ portion of this solution required 30.7 cm³ of 0.100 mol dm⁻³ sodium hydroxide for neutralisation. The equations for the reactions involved are shown below.

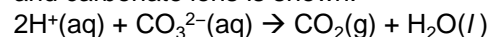


Calculate the percentage purity of the calcium hydroxide.

6.5) 1.50g of Bordeaux mixture containing a mixture of calcium hydroxide (Ca(OH)₂) and copper(II) sulfate-5-water (CuSO₄·5H₂O) was added to 25.0 cm³ of hydrochloric acid, concentration 2.00 mol dm⁻³. The calcium hydroxide in the mixture will neutralise some of the acid. The resulting solution was made up to 250 cm³ with water in a volumetric flask. A 25.0 cm³ portion of this solution required 26.7 cm³ of 0.100 mol dm⁻³ sodium hydroxide for neutralisation. Calculate the percentage, by mass, of calcium hydroxide in the sample of Bordeaux mixture.

6.6) Dolomite is a carbonate-containing mineral with formula CaX(CO₃)₂ where X is a metal ion.

The carbonate ions in Dolomite react with acid. The ionic equation for the reaction between hydrogen ions and carbonate ions is shown.



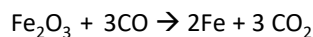
A 6.00 g sample of dolomite is dissolved in 35.0 cm³ of 5.00 mol dm⁻³ hydrochloric acid, which is an excess. The resulting solution is made up to 100 cm³ in a volumetric flask, using distilled water. 10.0 cm³ portions of this solution are titrated against sodium hydroxide. 22.40 cm³ of 0.200 mol dm⁻³ NaOH is required for neutralisation. Calculate the relative formula mass of CaX(CO₃)₂ and hence the relative atomic mass and identity of the metal ion X

% 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 7: 25g of Fe_2O_3 was reacted and it produced 10g of Fe. Calculate the percentage yield.

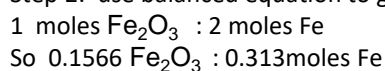


First calculate the maximum/theoretical mass of Fe that could be produced.

Step 1: work out amount in mol of Iron oxide

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

Step 2: use balanced equation to give moles of Fe



Step 3: work out mass of Fe

$$\begin{aligned} \text{Mass} &= \text{amount} \times M_r \\ &= 0.313 \times 55.8 \\ &= 17.48\text{g} \end{aligned}$$

$$\% \text{ yield} = (\text{actual yield} / \text{theoretical yield}) \times 100$$

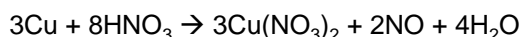
$$= (10 / 17.48) \times 100$$

$$= 57.2\%$$

7.1) 10.0 g of CaCO_3 was reacted with H_2SO_3 . The reaction was incomplete and only 10.7g of CaSO_3 was formed. Calculate the percentage yield of the reaction.



7.2) 5.70g of Copper metal was reacted with HNO_3 . The reaction was incomplete and only 14.80g of $\text{Cu}(\text{NO}_3)_2$ was formed. Calculate the percentage yield of the reaction.

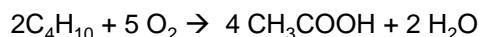


7.3) 5.78 g of magnesium carbonate were added to an excess of sulfuric acid. The following reaction occurred.

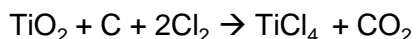


Calculate the actual mass of MgSO_4 produced in this reaction assuming a 85% yield.

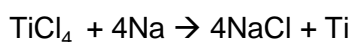
7.4) The reaction from ethanoic acid (CH_3COOH) has a percentage yield of 82%. Calculate the minimum mass of butane needed to actually produce 10.2g of CH_3COOH



7.5) Titanium can be produced from TiO_2 from the following two reactions



The first reaction has a percentage yield of 90%



The second reaction has a percentage yield of 95%

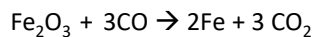
Calculate the actual mass of titanium that can be produced from 1.00kg of TiO_2 taking into account the percentage yields.

% 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 8 : Calculate the % atom economy for the following reaction where iron is the desired product assuming the reaction goes to completion.



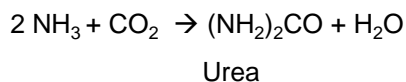
$$\begin{aligned} \text{\% atom economy} &= \frac{(2 \times 55.8)}{(2 \times 55.8 + 3 \times 16) + 3 \times (12 + 16)} \times 100 \\ &= 45.8\% \end{aligned}$$

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

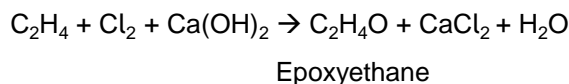
Reactions where there is only one product where all atoms are used making product are ideal and have 100% atom economy.
e.g. $\text{CH}_2=\text{CH}_2 + \text{H}_2 \rightarrow \text{CH}_3\text{CH}_3$

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.

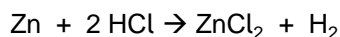
8.1) Calculate the atom economy to make urea from the following equation



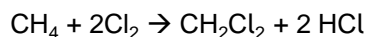
8.2) Calculate the atom economy to make epoxyethane from the following equation



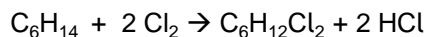
8.3) Calculate the atom economy to make hydrogen from the reaction of zinc with HCl.



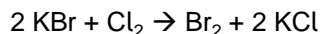
8.4) Calculate the % atom economy for the formation of CH_2Cl_2 in this reaction



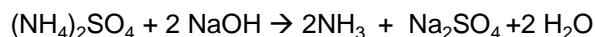
8.5) Calculate the % atom economy for the formation of $\text{C}_6\text{H}_{12}\text{Cl}_2$ in this reaction



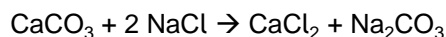
8.6) Calculate the % atom economy for the formation of Br_2 in this reaction



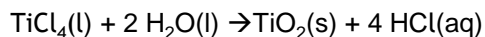
8.7) Calculate the % atom economy for the formation of NH_3 in this reaction



8.8) Calculate the % atom economy to make sodium carbonate from this reaction



8.9) Calculate the % atom economy for the formation of TiO_2 in this reaction



Answers

- | | |
|-------------|------------|
| 1.1) 4.21g | 2.1) 31.5g |
| 1.2) 1.39g | 2.2) 60.6g |
| 1.3) 1579g | 2.3) 61.0g |
| 1.4) 239g | 2.4) 145g |
| 1.5) 0.418g | 2.5) 18.4g |

- | | |
|--------------------------------|--------------------------------|
| 3.1) 0.500g | 4.1) 118 |
| 3.2) 51.8g | 4.2) $M_r = 124$ $x=1$ |
| 3.3) 27.4cm ³ | 4.3) 235 |
| 3.4) 6.25g | 4.4) 18% |
| 3.5) 2.38g | 4.5) 138 $M=39$, K |
| 3.6) 6.86g 5.14dm ³ | 4.6) 12.2 mol dm ⁻³ |
| 3.7) 7.78g | |
| 3.8) 1.24 mol dm ⁻³ | |

- 5.1) 0.906
5.2) 544
5.3) 72.6, Ge
5.4) 89.8%
5.5) 5.13 x10⁻⁶ m³ and % (if using ideal gas equation)

- 6.1) 68.6%
6.2) 2.65 x10⁻⁴
6.3) 72.6%
6.4) 63.3%
6.5) 57.6%
6.6) M_r 184.3 , Ar 24.2 , Mg

- | | |
|------------|------------|
| 7.1) 89% | 8.1) 76.9% |
| 7.2) 88% | 8.2) 25.4% |
| 7.3) 7.02g | 8.3) 1.44% |
| 7.4) 6.01g | 8.4) 53.8% |
| 7.5) 513g | 8.5) 68.0% |
| | 8.6) 51.7% |
| | 8.7) 16% |
| | 8.8) 48.8% |
| | 8.9) 35.4% |