
1.2 Amount of Substance

Relative masses, Mr:

- Where $1/12^{\text{th}}$ of Carbon 12 is essentially the mass of a proton / neutron.
- The first 2 definitions you've come across in the last unit.
- Additional ones are added for covalent and ionic compounds

For atoms:

RAM, Relative Atomic Mass: the weighted mean mass of an atom compared with $1/12^{\text{th}}$ of the mass of carbon -12

For isotopes:

Relative isotopic mass: the mass of an isotope compared with $1/12^{\text{th}}$ of the mass of carbon -12

For covalent molecules (non metal & non metal):

Relative molecular mass, Mr: the weighted mean mass of a molecule compared with $1/12^{\text{th}}$ of the mass of carbon -12

For ionic compounds (metal & non metal):

Relative formula mass, Mr: the weighted mean mass of a formula unit compared with $1/12^{\text{th}}$ of the mass of carbon -12

The Mole and Avogadro's constant

- Since atoms are so small and therefore have such a small mass we have to measure them in large numbers which we call a mole.
- The simplest way to understand the mole is to treat it as a word to describe a number, such as:

Dozen	Tonne	Grand	Mole
12	100	1000	6.02×10^{23}

- It does appear to be quite an unusual number but it was carefully thought out by **Avogadro**
- It is the number of atoms of an element to make its **atomic mass number**
- It is called **Avogadro's constant, N_A**

Avogadro's constant, N_A : 6.02×10^{23}

Basically in 12g of carbon-12 you would find 6×10^{23} atoms of carbon.

1g of ^1H atoms would have 6×10^{23} atoms of H

16g of ^{16}O atoms would have 6×10^{23} atoms of O (atom is 16 x heavier than H)

32g of ^{32}S atoms would have 6×10^{23} atoms of S (atom is 32 x heavier than H)

When you think about it like this it actually makes sense!!!

- In fact if you were to measure out **6×10^{23} (A Mole)** atoms of any element you would find that its mass is the same as its **Relative Atomic Mass**

1 Mole of Sodium ^{23}Na **23g** mol^{-1}

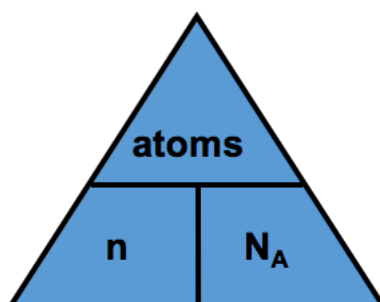
1 Mole of Magnesium ^{24}Mg **24g** mol^{-1}

1 Mole of Iron ^{56}Fe **56g** mol^{-1}

- A molecule is made up from more than 1 atom so the mass of 1 mole of that molecule will be the sum of the **Relative Atomic Masses**
1 Mole of water H_2O **18g** mol^{-1}
1 Mole of Sodium Chloride NaCl **58.5g** mol^{-1}

Number of particles = Number of Moles x Avogadro's constant

$N^\circ \text{ particles} = \text{Moles} \times N_A$



Using Moles

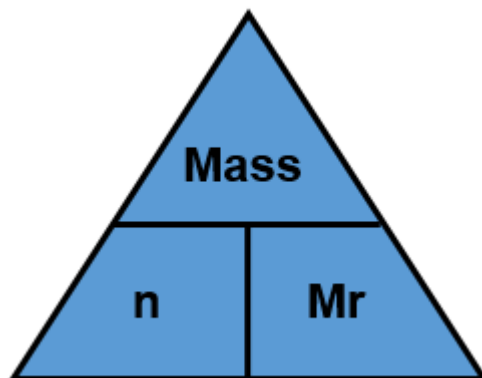
- There are 3 mole formulas you require depending on the units you are working in:

1) Moles and mass, (s) – grams, g

- If 1 Mole of water is 18g then 2 moles would be 36g. 3 moles would be 54g and 0.5 moles would be 9g.

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Mr}}$$

$$n = \frac{m}{\text{Mr}}$$



TIP: mass must be in g so make sure you can convert to this.
k means 1000's of, ie 1000 of grams:
m means 1000th's of a gram, 0.001g

$$1000 \text{ mg} \xrightarrow[\times 10^{-3}]{/ 1000} 1 \text{ g} \xrightarrow[\times 10^{-3}]{/ 1000} 0.001 \text{ kg} \xrightarrow[\times 10^{-3}]{/ 1000} 0.000001 \text{ Tonne}$$

$$0.000001 \text{ Tonne} \xrightarrow[\times 10^3]{\times 1000} 0.001 \text{ kg} \xrightarrow[\times 10^3]{\times 1000} 1 \text{ g} \xrightarrow[\times 10^3]{\times 1000} 1000 \text{ mg}$$

Example:

- a) How many moles of water in 36g of H₂O

$$n = \frac{m}{\text{Mr}}$$

$$n = \frac{36}{18}$$

$$n = 2 \text{ moles}$$

- b) What is the mass of 0.5 moles of NaCl

$$m = n \times \text{Mr}$$

$$m = 0.5 \times 58.5$$

$$m = 29.25\text{g}$$

Questions:

- 1) Calculate the number of moles of 3.45g of Lithium

- 2) Calculate the number of moles of 243mg of Magnesium

- 3) Calculate the number of moles of 2.79Kg of Iron

Numbers of particles and masses

- 4) Calculate the number atoms of Calcium in 4.01g

- 5) Calculate the number atoms of sodium in 575mg

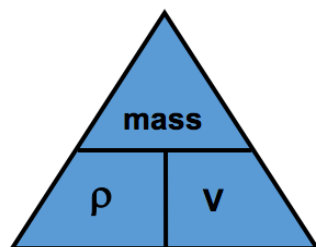
- 6) Calculate the number atoms of Silicon in 5.62Kg

Density

Density, ρ : is the amount of substance (g) per unit volume (cm^3)

- The whole density scale is measured against water which has a density of 1 g cm^{-3}
- This means that in 1 cm^3 of water there is 1g of water
- It is usually used for organic liquids (not solutions, later).

The equation:



- The mass, volume and density could be in any combination of mass or volume units so make sure your units are consistent in the question

Example calculations:

- 1) Calculate the density of ethanol in g cm^{-3} given a mass of 19g in a volume of 25 cm^3 ?

$$\rho = m / V$$

$$\rho = 19 / 25$$

$$\rho = 0.76 \text{ g cm}^{-3}$$

- 2) Calculate the density of gold in g cm^{-3} given that a gold bar contains 5 moles of gold and is 5.3cm wide, 11.8cm long and 0.8cm thick?

$$m = n \times M_r$$

$$V = 5.3 \times 11.8 \times 0.8$$

$$m = 5 \times 197$$

$$V = 50 \text{ cm}^3$$

$$m = 1000\text{g}$$

$$\rho = m / V$$

$$\rho = 1000 / 50$$

$$\rho = 20 \text{ g cm}^{-3}$$

Questions

- 1) Calculate the number of moles of 5cm^3 of Bromine, Br_2 if its density is 1.35 gcm^{-3}
- 2) Calculate the number of moles of 2cm^3 of Ammonia, NH_3 if its density is 1.12 gcm^{-3}
- 3) Calculate the number of moles of 12cm^3 of Hydrogen fluoride, HF if its density is 1.11 gcm^{-3}
- 4) Calculate the number of molecules in 2cm^3 of water, H_2O if its density is 1.00 gcm^{-3}
- 5) Calculate the number of atoms in 6cm^3 of Mercury, Hg if its density is 3.15 gcm^{-3}
- 6) Calculate the number of atoms in 35cm^3 of Neon, Ne if its density is 0.31 gcm^{-3}

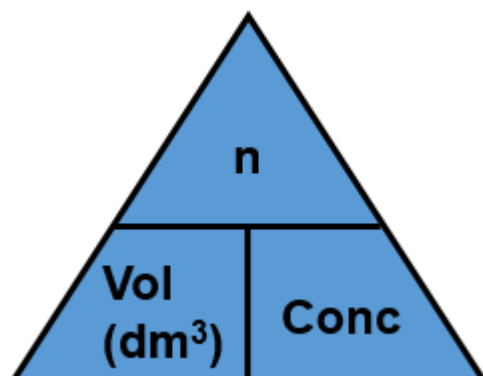
2) Moles and Solutions, (aq) – mol dm⁻³

- A solution is expressed as a number of moles in 1dm³ of solution .

$$\text{Number of moles} = \text{Concentration} \times \text{Volume}$$

(mol dm⁻³) (dm³)

$$n = C \times V \text{ (dm}^3\text{)}$$



TIP: Volume must be in dm³ so make sure you can convert to this:

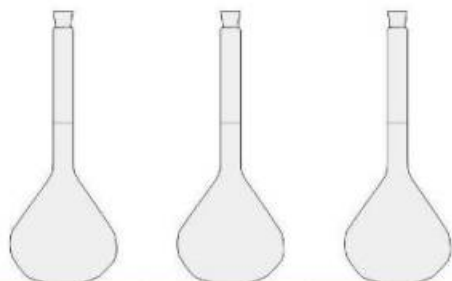
$$1\text{dm}^3 = 1000\text{cm}^3 = 1\text{litre}$$

$$\text{X } 1000$$

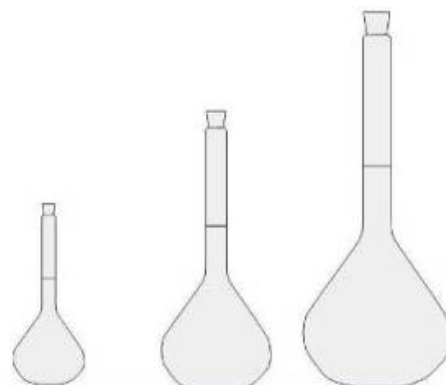
$$1\text{dm}^3 \rightarrow 1000\text{cm}^3$$

$$\text{/ } 1000$$

$$1000\text{cm}^3 \rightarrow 1\text{dm}^3$$



Volume (dm ³)	1.0 dm ³	1.0 dm ³	1.0 dm ³
Moles dissolved	2.0 mole	1.0 mole	0.5 mole
Concentration	2.0 mol dm ⁻³	1.0 mol dm ⁻³	0.5 mol dm ⁻³
Calculated by	Moles / Vol (dm ³)	Moles / Vol (dm ³)	Moles / Vol (dm ³)
	2.0 / 0.5	1.0 / 1.0	1.0 / 2.0



Volume (dm ³)	0.5 dm ³	1.0 dm ³	2.0 dm ³
Moles dissolved	1.0 mole	1.0 mole	1.0 mole
Concentration	2.0 mol dm ⁻³	1.0 mol dm ⁻³	0.5 mol dm ⁻³
Calculated by	Moles / Vol (dm ³)	Moles / Vol (dm ³)	Moles / Vol (dm ³)
	2.0 / 0.5	1.0 / 1.0	1.0 / 2.0

Example:

Calculate the number of moles of NaOH in 50cm³ of a 0.30 Mol dm⁻³ solution

$$n = C \times V \text{ (dm}^3\text{)} \quad \text{V is in dm}^3 \quad 50/1000 = 0.05$$

$$n = 0.3 \times 0.05$$

$$n = 0.015 \text{ moles}$$

Standard solutions calculations

- These are calculations to make a smaller volume of a specific concentration
- These combine both of the mole formulas in a calculation:

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Mr}}$$

$$\text{Number of moles} = \text{Concentration} \times \text{Volume}$$

Example:

What mass of NaOH is required to make 250cm³ of 0.1 mol dm⁻³ solution of sodium hydroxide?

- To calculate the mass, we need moles
- So, we have to calculate the moles from volume and concentration first:

$$n = C \times V (\text{dm}^3) \quad \text{V is in dm}^3 \quad 250/1000 = 0.25$$

$$n = 0.1 \times 0.25$$

$$n = 0.025 \text{ moles}$$

- Now calculate the mass from moles

$$m = n \times \text{Mr}$$

$$m = 0.025 \times 40$$

$$m = 1.00\text{g}$$

Questions

1) Calculate the number of moles in the following.

a) 2 dm³ of 0.05 mol dm⁻³ HCl

b) 50 cm³ of 5 mol dm⁻³ H₂SO₄

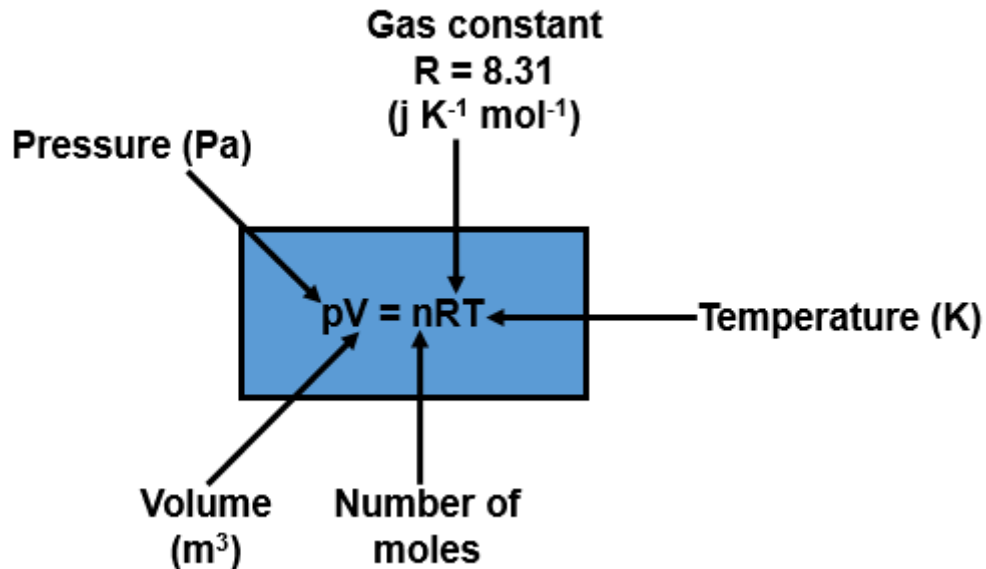
c) 10 cm³ of 0.25 mol dm⁻³ KOH

- 2) Calculate the concentration of the following in **both** mol dm^{-3} and g dm^{-3}
- 0.400 moles of HCl in 2.00 dm^3 of solution
 - 12.5 moles of H_2SO_4 in 5.00 dm^3 of solution
 - 1.05 g of NaOH in 500 cm^3 of solution
- 3) Calculate the volume of each solution that contains the following number of moles.
- 0.00500 moles of NaOH from $0.100 \text{ mol dm}^{-3}$ solution
 - 1.00×10^{-5} moles of HCl from $0.0100 \text{ mol dm}^{-3}$ solution
- 4) Calculate the concentration if 561mg of KOH was dissolves in 50cm^3 of water
- 5) Calculate the concentration if 583mg of $\text{Mg}(\text{OH})_2$ was dissolves in 25cm^3 of water
- 6) Calculate the concentration if 0.0126g of HNO_3 was dissolves in 100cm^3 of water

1) Moles and gases, (g) – M³

- The volume of a gas can vary depending on **temperature** and **pressure**.
- These need to be taken into account when dealing with **moles** and **gases**

The Ideal gas equation



2 assumptions:

- The volume of the molecules is negligible
- The molecules have no intermolecular forces of attraction

TIP: Volume must be in m³ so make sure you can convert to this:

$$1000000\text{cm}^3 = 1000\text{dm}^3 = 1\text{m}^3$$

$$1\text{m}^3 \xrightarrow{\times 1000} 1000\text{dm}^3 \quad 1000\text{dm}^3 \xrightarrow{\times 1000} 1000000\text{cm}^3 \quad 1\text{m}^3 \xrightarrow{\times 1000000} 1000000\text{cm}^3$$

$$1000000\text{cm}^3 \xrightarrow{/ 1000} 1000\text{dm}^3 \quad 1000\text{dm}^3 \xrightarrow{/ 1000} 1\text{m}^3 \quad 1000000\text{cm}^3 \xrightarrow{/ 1000000} 1\text{m}^3$$

TIP: Temperature must be in kelvin, K so make sure you can convert to this:

$$0^\circ\text{C} \xrightarrow{+273} 273\text{K} \quad 100^\circ\text{C} \xrightarrow{+273} 373\text{K} \quad 25^\circ\text{C} \xrightarrow{+273} 298\text{K}$$

$$273\text{K} \xrightarrow{-273} 0^\circ\text{C} \quad 373\text{K} \xrightarrow{-273} 100^\circ\text{C} \quad 298\text{K} \xrightarrow{-273} 25^\circ\text{C}$$

Examples:

a) How many moles are there in 0.05m^3 of Nitrogen gas, at 273K and 100000Pa

$$P V = n R T$$

Rearrange to get n on its own, divide both sides by RT

$$n = \frac{P V}{R T}$$

$$n = \frac{100000 \times 0.05}{8.31 \times 273}$$

$$n = 2.20 \text{ moles}$$

b) What is the volume occupied when 4 moles of Chlorine gas is at 27°C and 100 kPa ?

Convert units to SI units first:

$$T = 27 + 273 = 300\text{K}$$

$$P = 100 \times 1000 = 100000\text{Pa}$$

$$P V = n R T$$

Rearrange to get V on its own, divide both sides by p

$$V = \frac{n R T}{P}$$

$$V = \frac{4 \times 8.31 \times 300}{100000}$$

$$V = 0.0997 \text{ m}^3$$

c) What mass of oxygen gas, O_2 that has a volume of 1200cm^3 at 25°C and 200 kPa ?

To get mass, we need moles and M_r . We have to use $PV = nRT$ first to get moles, n

Convert units to SI units first:

$$T = 25 + 273 = 298\text{K} \quad P = 200 \times 1000 = 200000\text{Pa} \quad V = 1200 / 1000000 = 0.0012\text{m}^3$$

$$P V = n R T$$

Rearrange to get n on its own, divide both sides by RT

$$n = \frac{P V}{R T}$$

$$n = \frac{200000 \times 0.0012}{8.31 \times 298}$$

$$n = 0.0969 \text{ moles}$$

Now use the moles in the mass equation

$$\text{mass} = n \times M_r$$

$$\text{mass} = 0.0969 \times 32$$

$$\text{mass} = 3.10\text{g}$$

Questions

- 1) Calculate the number of moles of a gas that occupies 10m^3 at 373K and 100000Pa of pressure
- 2) Calculate the number of moles of a gas that occupies 150dm^3 at 100°C and 250KPa of pressure
- 3) Calculate the number of moles of a gas that occupies 250cm^3 at 250°C and 25KPa of pressure
- 4) Calculate the mass of N_2 gas that occupies 500cm^3 at 350°C and 150KPa of pressure
- 5) Calculate the mass of NH_3 gas that occupies 2dm^3 at 400°C and 350KPa of pressure

6) a) Calculate the density in gcm^{-3} of carbon dioxide gas at 25°C at 101.325KPa

b) The Density of air is 0.997gcm^{-3} . State whether carbon dioxide would sink or float in an enclosed room.

Empirical and Molecular formula

Empirical formula: is the simplest whole number ratio of atoms of elements in a molecule

Molecular formula: is the actual number ratio of atoms of elements in a molecule

Examples:

- a) A sample of iron oxide was found to have 11.2g of iron and 4.8g of oxygen. Calculate the formula of this compound

Element	Fe		O
Masses	11.2		4.8
Divide by Ar	11.2 / 55.8		4.8 / 16
Moles	0.2	:	0.3
Divide by smallest	0.2 / 0.2	:	0.3 / 0.2
Ratio	1	:	1.5
Whole No Ratio	2	:	3
Empirical formula	Fe₂O₃		

- b) A sample of hydrocarbon was found to have 1.20g of carbon and 0.25g of hydrogen. Calculate the Empirical formula of this compound. Then find out the molecular formula if the Mr = 58

Element	C		H
Masses	1.20		0.25
Divide by Ar	1.20 / 12		0.25 / 1
Moles	0.10	:	0.25
Divide by smallest	0.10 / 0.10	:	2.5 / 0.10
Ratio	1	:	2.5
Whole No Ratio	2	:	5
Empirical formula	C₂H₅ (29 x 2 = 58)		
Molecular formula	C₄H₁₀		

TIP:

%'s may be used instead of masses, treat the calculation in the same way as %'s, these could be thought of as masses in 100g

You may have to calculate the mass or % of an element in a sample by taking the mass of one element from the total mass of the compound

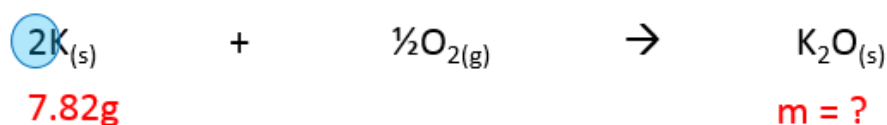
Moles and reactions

- Mole calculations can now be used to calculate reacting amounts / product amounts.
- This is done by using the balanced chemical equation and moles calculations using masses, gas volumes and concentrations.
- **ALL 3** of these require the use of the mole:

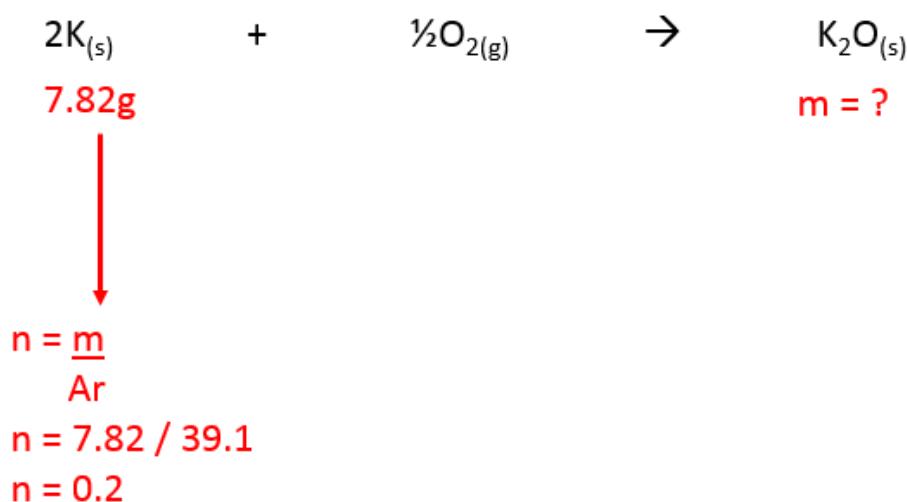
A) Mass / mole calculations:

Example: 7.82g of potassium reacts in air to form potassium oxide. Calculate the mass of potassium oxide made:

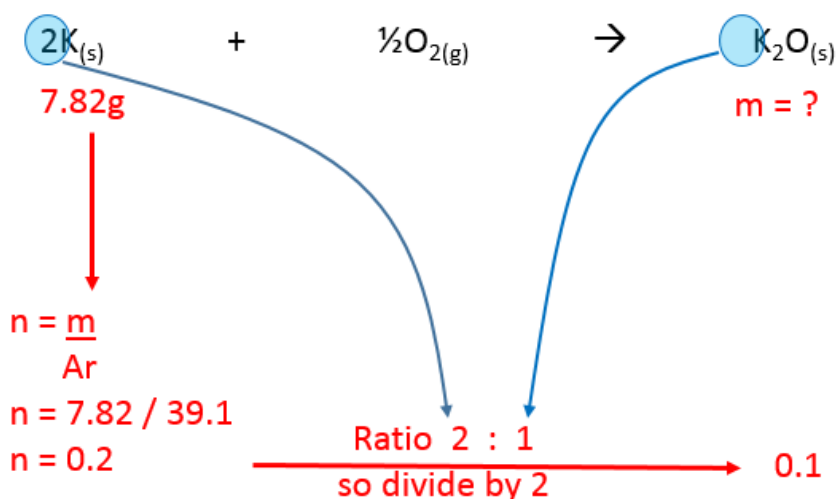
STEP1: Write a balanced chemical equation and add the amounts given and question mark what you are asked to work out:



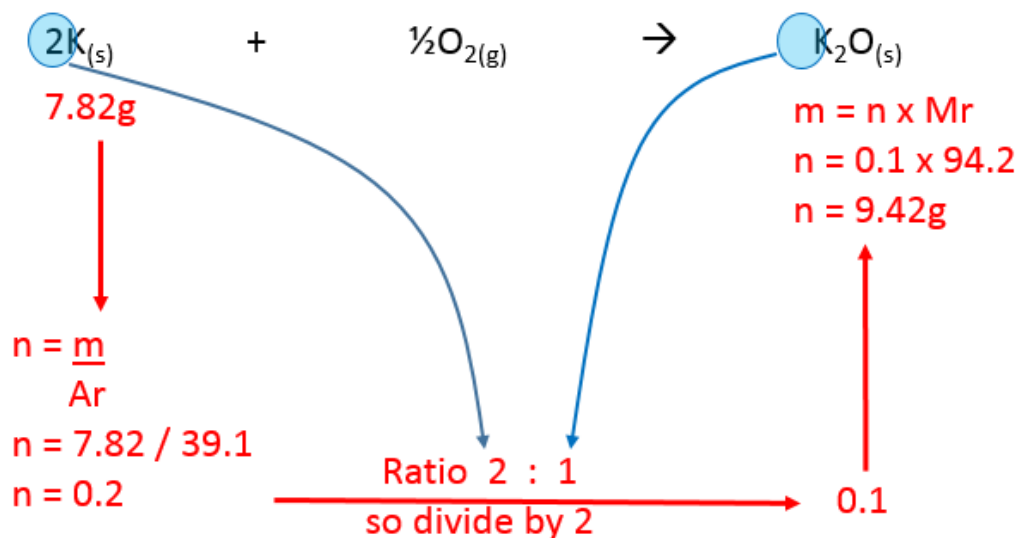
STEP2: Check the state symbol of your starting mass to decide which moles equation you will use
- (s) - means you use **Moles = mass / Ar**



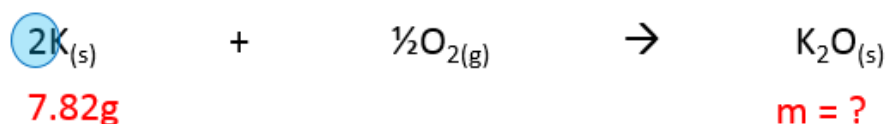
STEP3: Use the reacting **ratios** to work out how many moles you have made (or need):



STEP4: Check the question/ state symbol to decide whether to convert it to mass / concentration / volume - (s) = mass



- These are usually done as a series of steps but the process is the same:



Calculate moles of potassium

$$n \text{ of K} = \text{mass} / Ar$$

$$n \text{ of K} = 7.82 / 39.1$$

$$n \text{ of K} = 0.2$$

Calculate moles of potassium oxide

$$n \text{ of K}_2\text{O} = 0.2 / 2 \text{ (ratio 2:1, divide by 2)}$$

$$n \text{ of K}_2\text{O} = 0.1$$

Calculate mass of potassium oxide

$$\text{mass of K}_2\text{O} = n \times Mr$$

$$\text{mass of K}_2\text{O} = 0.1 \times 94.2$$

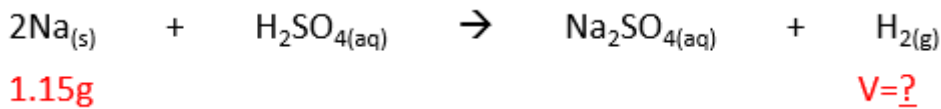
$$\text{mass of K}_2\text{O} = 9.42\text{g}$$

B) Gas / mole calculations:

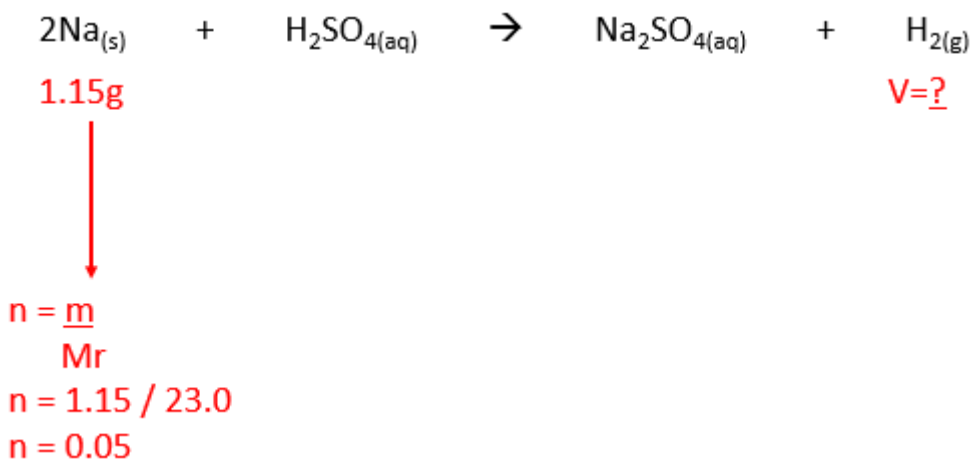
Example: 1.15g sodium reacts with excess sulphuric acid to form sodium sulphate and hydrogen gas.

Calculate the volume of hydrogen made in m^3 if the reaction was carried out at 25°C and 100 kPa:

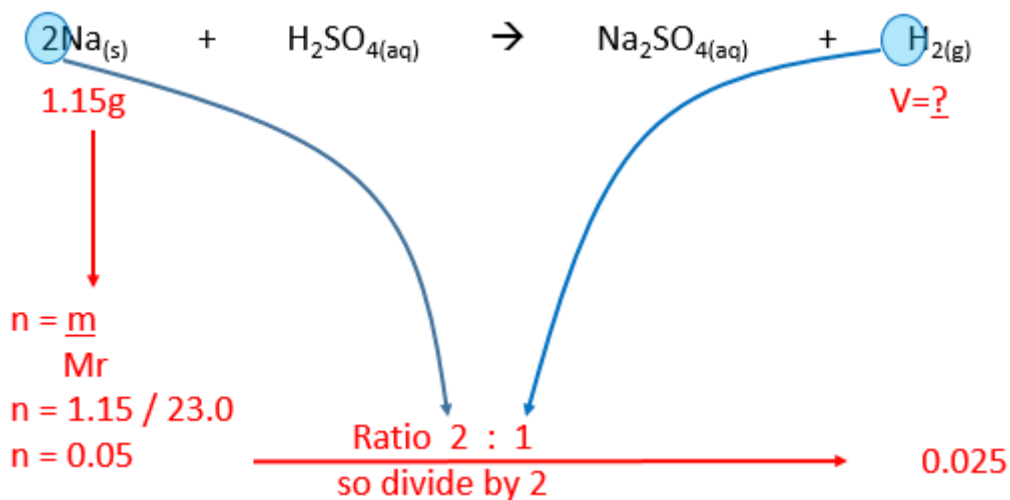
STEP1: Write a balanced chemical equation and add the amounts given and question mark what you are asked to work out:



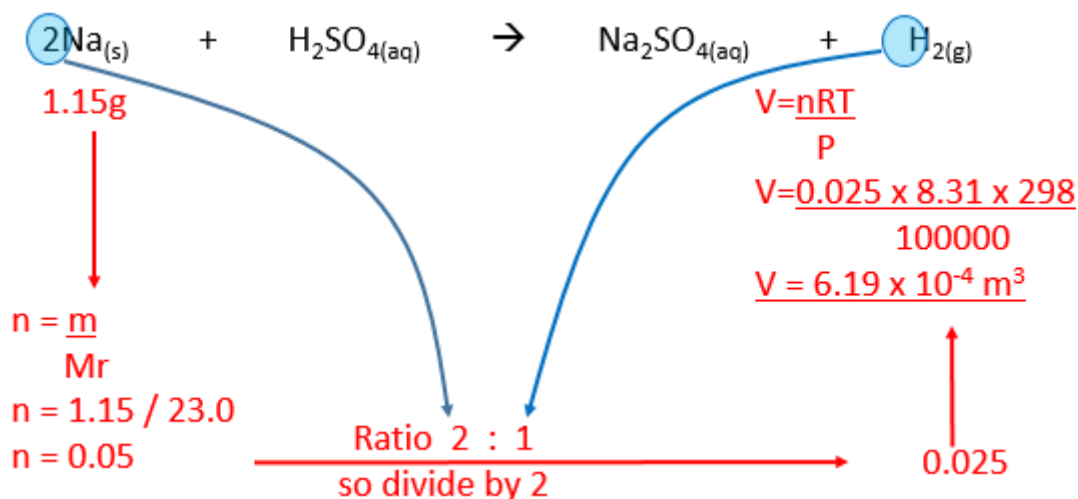
STEP2: Check the state symbol of your starting mass to decide which moles equation you will use
- (s) - means you use **Moles = mass / Ar**



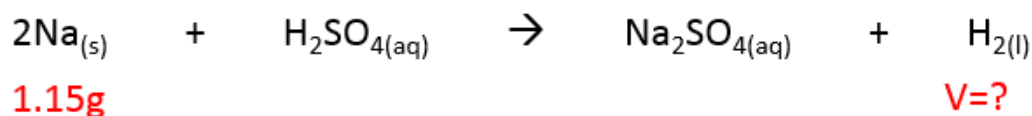
STEP3: Use the reacting **ratios** to work out how many moles you have made (or need):



STEP4: Check the question/ state symbol to decide whether to convert it to mass / concentration / volume - (g) = volume



- Again, these can be done as a series of steps:



Calculate moles of sodium

$$n \text{ of Na} = \text{mass} / A_r$$

$$n \text{ of Na} = 1.15 / 23$$

$$n \text{ of Na} = 0.05$$

Calculate moles of hydrogen

$$n \text{ of H}_2 = 0.05 / 2 \quad (\text{ratio } 2:1, \text{ divide by } 2)$$

$$n \text{ of H}_2 = 0.025$$

Calculate volume of H₂

$$\text{Vol of H}_2 = nRT / P$$

$$\text{Vol of H}_2 = 0.025 \times 8.31 \times 298 / 100000$$

$$\text{Vol of H}_2 = 6.19 \times 10^{-4} \text{ m}^3$$

Required Practical 1 - Titrations

- This technique can be used to find:

Concentration	Mr	Formula	Water of crystallisation
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- To do this you react a certain volume of a solution with an unknown concentration with a solution of **known concentration**.
- The concentration of the known solution must be accurate and is known as a **standard solution**.

Making a standard solution – Making 250cm³ of a 0.1 mol dm⁻³ solution of NaOH

- Weigh a known mass (number of moles) out in a weighing boat recording its mass to the number of decimal places on the balance.

$$n = C \times V \text{ (dm}^3\text{)} \quad (250/1000 = 0.25)$$

$$m = n \times Mr$$

$$n = 0.1 \times 0.25$$

$$m = 0.025 \times 40$$

$$n = \mathbf{0.025 \text{ moles}}$$

$$\mathbf{m = 1.00g}$$

- Transfer to a beaker and reweigh the weighing boat (as there may be some left in the weighing boat). The difference is the **precise** mass added to a beaker:

Mass of weighing boat + calculated mass NaOH	2.62g
Mass of weighing boat	1.63g
Mass of NaOH dissolved	0.99g

- Dissolve in 100cm³ of distilled water and stir with a glass rod.
- Using a funnel, pour into a volumetric flask.
- Use the wash bottle to wash beaker, funnel and glass rod into the volumetric flask.
- Fill the volumetric flask with distilled water so the meniscus sits on the line.
- Stopper the flask and invert several times to ensure mixing.
- Now calculate the **exact concentration**:

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

$$C = \frac{n}{V}$$

$$n = \frac{0.99}{40}$$

$$C = \frac{0.02475}{0.25}$$

$$n = \mathbf{0.02475 \text{ moles}}$$

$$\mathbf{C = 0.099 \text{ mol dm}^{-3}}$$

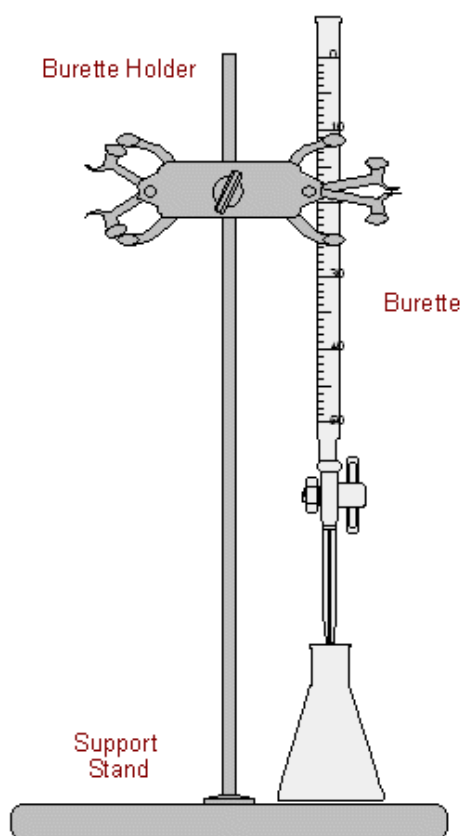
Carrying out a titration:

- Using moles and reacting ratios, you can calculate the concentration of a solution.
- The unknown goes in the conical flask and the known goes in the burette
- The only requirement is that you can tell when one solution has completely reacted with the other.
- Between acids and alkalis, we use indicators to let us know when the resulting solution is neutral.
- An indicator will change colour at the 'end point' (neutral).
- Common indicators are:

Indicator	Acidic colour	Base colour	End point colour
Methyl orange	Red	Yellow	Orange
Phenylphthalein	colourless	Pink	Pale pink

Technique/procedure

Example – finding an unknown concentration of NaOH using $0.10 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$



- 1) Rinse the burette with sulphuric acid, H_2SO_4 .
- 2) Fill the burette to the graduation mark ensuring the air is removed from the tap.
- 3) Rinse a pipette with sodium hydroxide, NaOH fill and transfer 25 cm^3 to a clean, dry conical flask.
- 4) Add 2-3 drops of indicator.
- 5) Run the acid into the alkali and stop when the colour changes. This is your 'trial'.
- 6) Record the burette readings to 2dp ending 0 / 5
- 7) Repeat the titration until you get **2 concordant results**
- 8) Calculate the mean titre to 2dp.

Record results in a table like the one below:

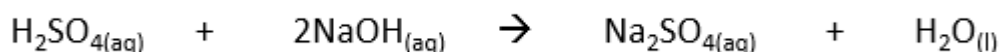
	Trial	1	2	3
Final burette reading / cm^3				
Initial burette reading / cm^3				
Titre / cm^3				
Mean Titre 2dp / cm^3				

C) Aqueous solution / mole calculation – example

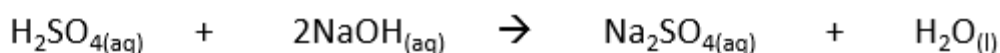
In a titration 0.01M sulphuric acid was added to 25cm³ of sodium hydroxide. Calculate the concentration of the sodium hydroxide given the following results:

	Trial	1	2
Final burette reading /cm ³	22.3	21.8	21.7
Initial burette reading /cm ³	0.00	0.00	0.00
Titre /cm ³	22.3	21.8	21.7
Mean Titre 2dp /cm ³		21.75	

1 Write a balanced equation



2 Calculate the number of moles of acid added from the burette



$$C = 0.01 \text{ mol dm}^{-3}$$

$$V = (21.75 \text{ cm}^3)$$

$$21.75 \times 10^{-3} \text{ dm}^3$$

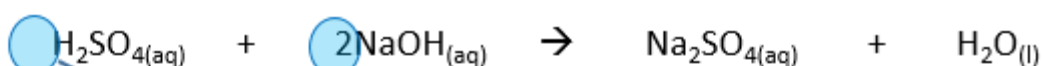


$$n = C \times V$$

$$n = 0.01 \times 0.02175$$

$$n = 2.175 \times 10^{-4} \text{ moles}$$

3 Use the ratio to work out the number of moles in the sample of alkali



$$C = 0.01 \text{ mol dm}^{-3}$$

$$V = (21.75 \text{ cm}^3)$$

$$21.75 \times 10^{-3} \text{ dm}^3$$

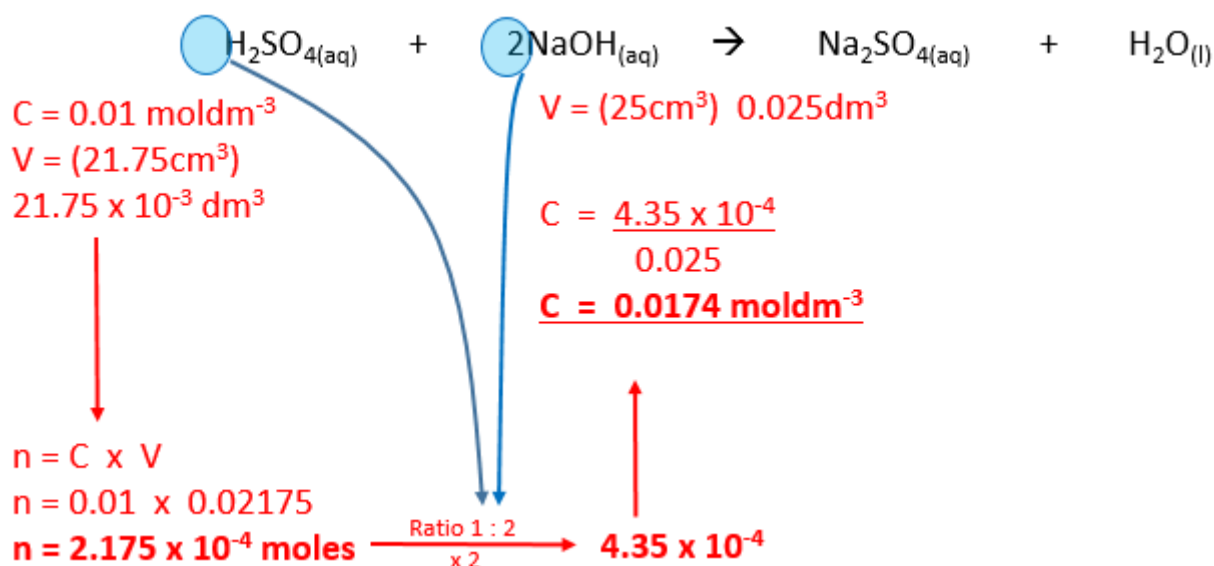


$$n = C \times V$$

$$n = 0.01 \times 0.02175$$

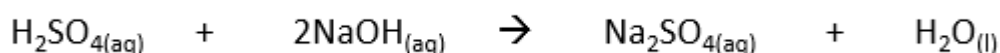
$$n = 2.175 \times 10^{-4} \text{ moles} \xrightarrow[\times 2]{\text{Ratio } 1 : 2} 4.35 \times 10^{-4}$$

4 Calculate the concentration.



- Again, these can be done as a series of steps:

1 Write a balanced equation



2 Calculate the number of moles of H_2SO_4 added from the burette

$$n \text{ of } \text{H}_2\text{SO}_4 = C \times V$$

$$n \text{ of } \text{H}_2\text{SO}_4 = 0.01 \times 0.02175$$

$$n \text{ of } \text{H}_2\text{SO}_4 = 2.175 \times 10^{-4}$$

3 Use the ratio to work out the number of moles of NaOH in the conical flask

$$\text{H}_2\text{SO}_4 : \text{NaOH} \quad 1 : 2$$

$$n \text{ of NaOH} = 2.175 \times 10^{-4} \times 2$$

$$n \text{ of NaOH} = 4.35 \times 10^{-4}$$

4 Calculate the concentration of NaOH

$$C = \frac{4.35 \times 10^{-4}}{0.025}$$

$$C = 0.0174 \text{ moldm}^{-3}$$

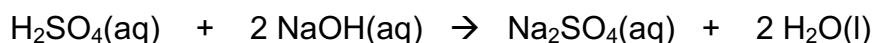
TIP:

Mass, gas and aqueous solution formulas may be used in a combination of ways in these reacting mole calculations

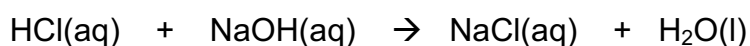
The format remains the same – a starting point – an end point, in the balanced equation

Questions

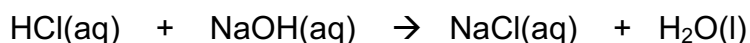
- 1) 25.0 cm³ of a solution of sodium hydroxide solution required 21.5 cm³ of 0.100 moldm⁻³ sulphuric acid for neutralisation. Find the concentration of the sodium hydroxide solution.



- 2) Find the volume of 1.0 moldm⁻³ hydrochloric acid that reacts with 25 cm³ of 1.50 moldm⁻³ sodium hydroxide.



- 3) 25.0 cm³ of 0.100 moldm⁻³ sodium hydroxide neutralises 19.0 cm³ of hydrochloric acid. Find the concentration of the acid.



- 4) What volume of 0.040 moldm⁻³ calcium hydroxide solution just neutralises 25.0 cm³ of 0.100 moldm⁻³ nitric acid?



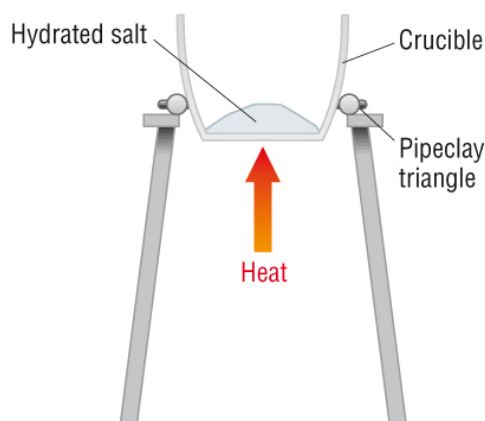
Water of crystallisation

- Coloured crystals such as blue copper sulphate have water molecules attached to the ions.
- The water can be driven by heat, leaving white copper sulphate crystals.
- This water locked in the crystal is called the **water of crystallisation**.

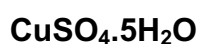
Hydrated - Crystals that contain water

Anhydrous - Crystals that do not contain water

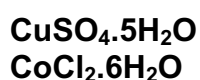
- The water can be evaporated by heat. Some compounds will decompose so a moderate heat must be used:



- The waters in the crystal obviously have a mass and will affect the Mr of the crystal.
- The water must be written in the formula. This is done by following a dot after the crystal formula:



- For copper sulphate, 1 mole of copper sulphate crystals will contain 5 moles of water:
- The number of moles of water per mole of crystal depends upon that crystal:



From mole calculations (example)

Mass of hydrated $\text{MgSO}_4 \cdot x\text{H}_2\text{O} = 4.312\text{g}$

Mass of anhydrous $\text{MgSO}_4 = 2.107\text{g}$

	Crystal, MgSO_4	Water, H_2O
Masses of each	2.107g	(4.312 - 2.107)
	2.107g	2.205g
Moles of each	2.107 / 120.4	2.205 / 18
	0.0175	0.1225
Divide by the smallest	0.0175 / 0.0175	0.1225 / 0.0175
	1	7

So the formula of hydrated $\text{MgSO}_4 \cdot x\text{H}_2\text{O} = \text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

Percentage yield:

Is a measure of how wasteful a chemical process is

- When we think about reactions, we always think of them as going 100% to products.
- This is usually **not** the case due to:

Equilibria Side reactions Purity Transfers Separation / purification

- Percentage yield is like a score in a test. It is an indication of what you achieved out of what you could have got:

$$\% \text{ Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

The rules:

- 1 Write a balanced chemical equation
- 2 Calculate the theoretical amount of product in moles
- 3 Calculate the theoretical amount of product in g
- 4 Calculate % yield using the formula:

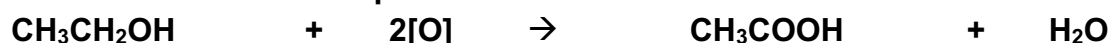
$$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Examples:

A) Preparation of ethanoic acid:

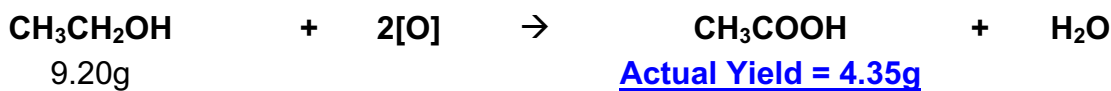
A student reacted 9.20g of ethanol with an excess of sulphuric acid and sodium dichromate (the oxidising agent). The student obtained 4.35g of ethanoic acid. Calculate the % yield:

1) Write a balanced chemical equation:



2) Calculate the theoretical amount of moles of product:

- Calculate the amount of moles of ethanoic acid you could have made:



↓

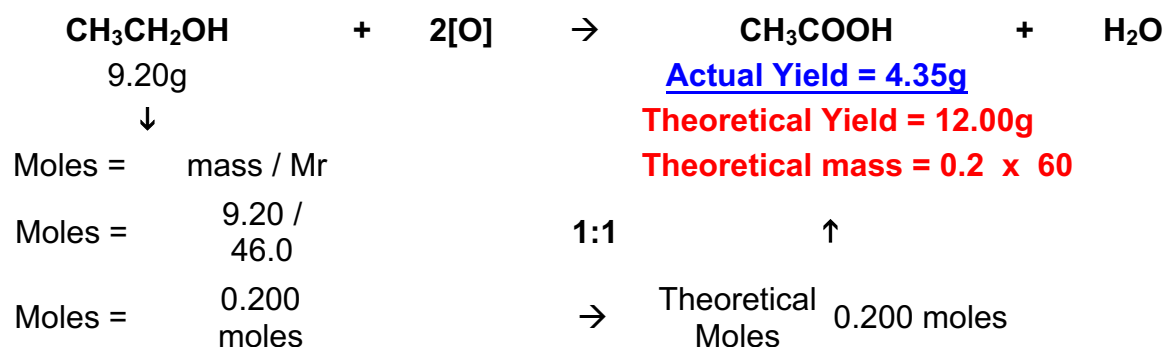
$$\text{Moles} = \frac{\text{mass}}{\text{Mr}}$$

$$\text{Moles} = \frac{9.20}{46.0} \quad 1:1$$

$$\text{Moles} = 0.200 \text{ moles} \quad \rightarrow \quad \begin{array}{l} \text{Theoretical} \\ \text{Moles} \end{array} \quad 0.200 \text{ moles}$$

3) Calculate the theoretical amount of product obtained in g:

- Calculate the number of moles you actually made:
- Calculate the amount of moles of ethanoic acid you could have made:



5) Calculate % yield using the formula:

$$\begin{aligned} \% \text{ Yield} &= \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 \\ &= \frac{4.35}{12.00} \times 100 \\ &= 36.25\% \end{aligned}$$

Atom economy:

Is a measure of how wasteful a reaction is

- Atom economy takes into account any wasteful by products too
- By products are considered wasteful as they are usually disposed of. This is costly and can cause environmental problems.
- A more efficient way of dealing with by products would be to sell them on to companies that would make use of them.

$$\text{Atom economy} = \frac{\text{Mr of the desired product}}{\text{Sum of Mr's of all products}} \times 100$$

Atom economy – Type of reaction:

- Reactions having only one product have a high atom economy. The type of reactions giving only one product are **addition reactions**.
- Reactions giving more than one product have a low atom economy. The type of reactions giving more than one product are **substitution / elimination reactions**.
- To improve the atom economy for **substitution / elimination** reactions, a use for the undesired product should be found.
- If the undesired product is toxic, we have even bigger problems -disposal.

Atom economy – Economic advantage

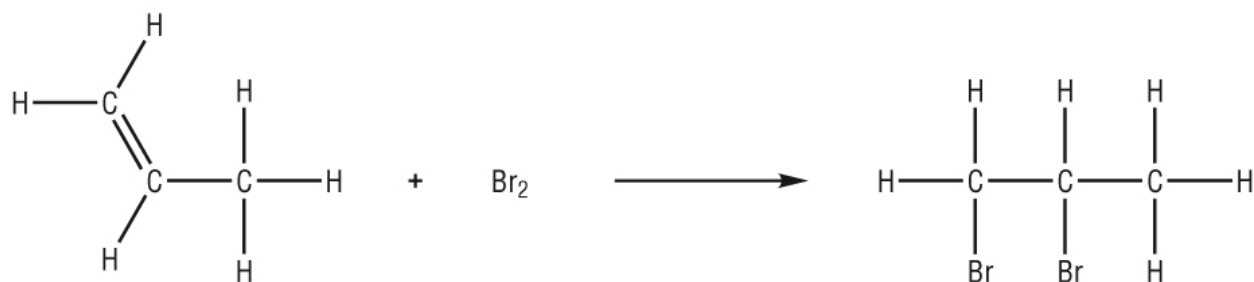
- Reactions that use a lot of starting materials to make a small amount of product has high waste.
- Reactions that give many other products apart from the desired products has high waste.
- Both of these will cost more money to make.
- Reducing the waste reduces cost – eg Ibuprofen has improved from 40% → 77%.

Atom Economy – Environment / ethics

- Raw materials – usually have limited supply so using them more efficiently makes them last longer.
- Waste materials – Disposal can be problematic as chemical waste is often harmful.
- Reducing both of the above 2 points can:
 - Reduce the demand on the worlds resources
 - Reduce the cost making them cheaper and more available

Calculating atom economy:

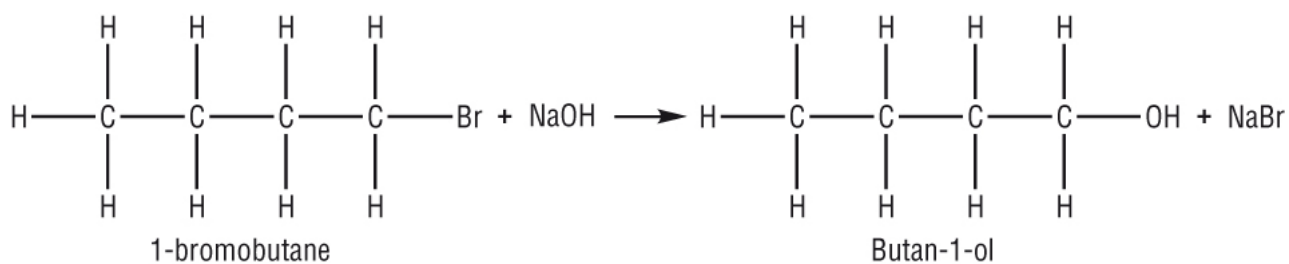
A) Bromination of propene:



$$\begin{aligned}\text{Atom economy} &= \frac{\text{Mr of the desired product}}{\text{Sum of Mr's of all products}} \times 100 \\ &= \frac{201.8}{201.8} \times 100 \\ &= 100\%\end{aligned}$$

- Any reaction that gives only one product is very atom economic, addition reactions for example.

B) Preparation of butan - 1 - ol:



$$\begin{aligned} \text{Atom economy} &= \frac{\text{Mr of the desired product}}{\text{Sum of Mr's of all products}} \times 100 \\ &= \frac{74.0}{176.9} \times 100 \\ &= 41.8\% \end{aligned}$$

- This means that most of the starting materials ended up as waste.