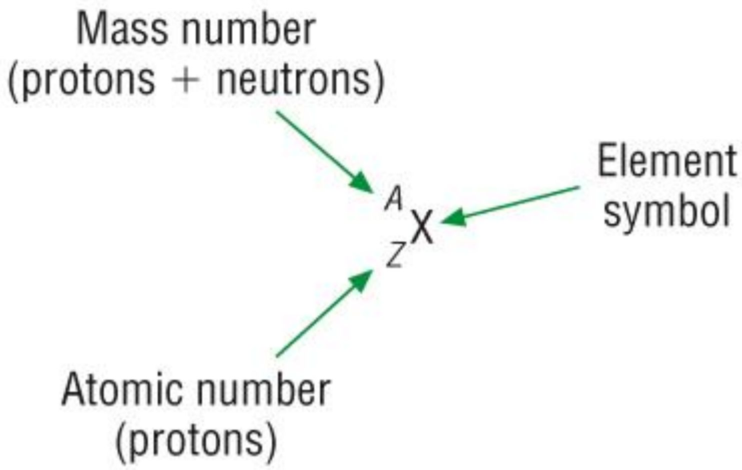
1B – The Mole

Atomic Masses
Protons, electrons and neutrons

Sub-Atomic Particle	Atomic Mass	Atomic Charge	working it out
Proton	1	+1	**bottom**
Electron	1/2000	-1	**bottom**
Neutron	1	0	**top - bottom**

- From GCSE this table shows that only protons and neutrons have a mass.
- Since different elements have different atoms.
- These atoms have different numbers of protons and neutrons (and electrons).
- This means that different elements atoms **must** have different masses.



- So an atom of one element **must** have a different mass from another element, we call this the **Mass Number**.
- The number of protons determines which element an atom is and the bottom number tell us this, we call this the **Atomic number**.

Examples:-

1) Lithium $^{7}_{3}\mathrm{Li}$

- The top number is the **Mass number**. This means that the total number of **protons and neutrons are 7**.
- The bottom number is the **Atomic number. This is the number of protons**.
- Because an **atom is neutral**, this means that this is also the number of electrons. This atom has **3 protons and 3 electrons**.
- If we take the **Atomic number (Z) from the Mass number (A)** we get the number of neutrons. **7-3=4 neutrons**.

2) Nitrogen $^{14}_{7}\mathrm{N}$

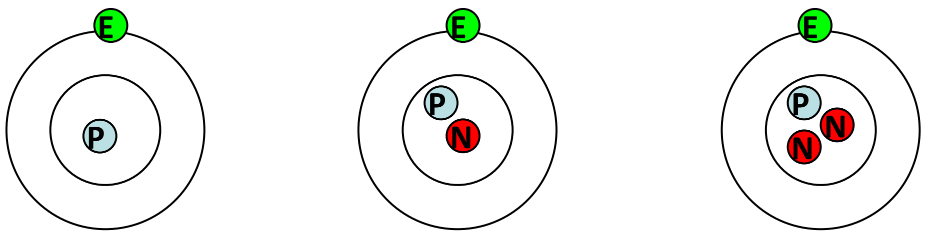
- An atom of nitrogen is twice as heavy as an atom of Lithium.
- The top number is the **mass number**. This means that the total number of **protons and neutrons are 14**.
- The bottom number is the **Atomic number. This is the number of protons**.
- Because an **atom is neutral**, this means that this is also the number of electrons. This atom has **7 protons and 7 electrons**.

- If we take the **Atomic number (Z) from the Mass number (A)** we get the number of neutrons. **14-7=7 neutrons**.

Isotopes

An atom of the same element that has the same number of protons and electrons but a different number of neutrons.

- As the atom has the same number of protons and electrons it will have the same chemical properties.
- They are all hydrogen atoms because they all have the **same number of protons**
- Hydrogen can be used as an example:-



Hydrogen – 0 Neutrons

Hydrogen – 1 Neutron (deuterium)

Hydrogen – 2 Neutrons (tritium)

Measurement of relative masses

- Because Chemistry is about reacting ratios we have to measure the amount of reacting particles of each reactant.
- We now know that each atom (and therefore molecules) have different masses (because they have different numbers of protons and neutrons), we have to be able to weigh out a number of particles of reactants to react with each other.
- Since atoms are so small we give them a mass scale of their own.
- This scale is called:

Unified atomic mass unit, u: 1u = 1.66 x 10^{-24}g

- This is basically the masses of a proton / neutron.

The mass of a carbon - 12 atom = 12u

The mass of 1/12th of a carbon - 12 atom = 1u

- We have to state the Atomic mass number of the atom as elements usually have isotopes:

There is a carbon - 13 atom containing an extra neutron, this would have an mass of 13u!!

- This is not the only thing we have to be careful of, there are 4 different types of masses and we have to use the correct one depending what we are referring to:

1) Relative isotopic mass:

- We use this one when we are only referring to **one isotope of an element ^{16}O has a mass of 16u**

2) Relative atomic mass:

- We use this one when we are referring to **the mixture of naturally occurring isotopes in of an element (the average by % - later)**

3) Relative molecular mass, (Mr):

- We use this one when we are referring to **simple molecules. Basically covalently bonded molecules, Cl_2, H_2O:**
- **Water, H_2O has a mass of 18u**

H 1 x 2 = 2
O 16 x 1 = <u>16</u>
18

4) Relative formula mass:

- We use this one when we are referring to **ionic compounds and giant covalent compounds. Such as $CaBr_2$, SiO_2**
- **Calcium bromide, $CaBr_2$ has a mass of 199.9u**

Ca 40.1 x 1 = 40.1
Br 79.9 x 2 = <u>159.8</u>
199.9u

- However for convenience we often use relative molecular mass instead.

<u>Calculating Relative Atomic Mass (RAM) from % abundance)</u>

- Use the formula:

$$\text{RAM} = \frac{(\% \times Ar) + (\% \times Ar) + \ldots}{100}$$

- Mass spectroscopy is the analytical method we use to obtain the % abundances for the isotopes.

Worked example:

RAM of Neon isotopes	Abundance %
20	**90.9**
21	**0.2**
22	**8.9**

$$\text{RAM} = \frac{(90.9 \times 20) + (0.2 \times 21) + (8.9 \times 22)}{100}$$

RAM = 20.18

Have a go for Si:

RAM of silicon isotopes	% Abundance
28	92.2
29	4.7
30	3.1

Questions 1-3 p9 / 2,3 p35

Amount of substance and the mole

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

Dozen 12
Ton 100
Pony 20
Grand 1000
Mole 6.02×10^{23} (Avogadro's constant, N_A)

- The mole is such a large number as it takes that many atoms to be able to measure a mass in g.
- It does appear to be quite an unusual number but it has been thought out:

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

- This number has been chosen to make it fit with the Atomic Masses from the Periodic Table:

1g of ^{1}H 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 **RAM**:-

1 Mole of Sodium ^{23}Na **23**g mol^{-1}
1Mole of Magnesium ^{24}Mg **24**g mol^{-1}
1 Mole of Iron ^{56}Fe **56**g 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 RAM

1 Mole of water H_2O 18g mol^{-1}
1Mole of Sodium Chloride NaCl 58.5g mol^{-1}

Using Moles

- Because we can't actually count out molecules or atoms (moles) we convert it to something we can measure i.e. mass.
- 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.
- The numbers are not always this simple so a formula helps.

$$\boxed{\text{No. Moles} = \frac{\text{Mass}}{\text{Ar(Mr)}}} \quad \textbf{OR} \quad \boxed{\text{Moles} = \frac{m}{\text{Mr}}}$$

Questions 1,2 p11 / 4,5 p35 / 2 p36

Types of Formula

Empirical Formula is the simplest ratio of atoms in a molecule. Molecular formulae is the actual ratio of atoms in a molecule

- This can be calculated using moles from percentage composition:-

Example 1
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	**Fe_2O_3**		

Example 2
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	**C_2H_5 (29 x 2 = 58)**		
Molecular formula	**C_4H_{10}**		

Questions 1,2 p13 / 6,7 p35

Moles and gas volumes

Avogadro's hypothesis

Equal volumes of gases will have the same number of atoms / molecules

- This makes gases particularly easy to calculate 1 mole of any gas, no matter what it is, occupies the same volume (at RTP)

1 mole of any gas occupies 24dm^3 (24000cm^3) at room temperature and pressure.

- Again the numbers are not always simple so a general formula will help:-

$$\text{No. Moles} = \frac{\text{Volume (dm}^3\text{)}}{24\text{dm}^3}$$

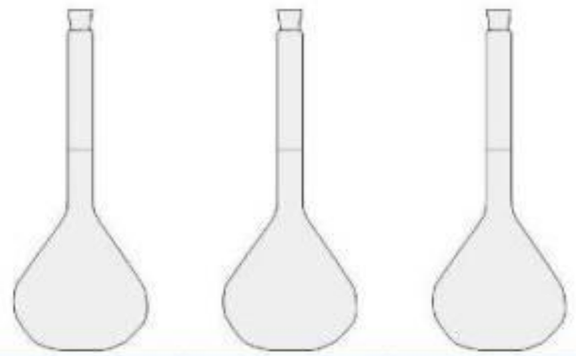
$$\text{Moles} = \frac{\text{V (dm}^3\text{)}}{24}$$

Questions 1-3 p15 / 8 p35 / 3 p36

Moles and solution

Concentration:

- So far when we have looked at reacting quantities we have dealt with the mole.
- Many reactions in chemistry involve solutions.
- A solution is expressed as a number of moles in 1dm^3 (1 litre or 1000cm^3) of solvent (usually water).
- **Concentration is the number of moles of specified entities in 1 dm^3 of solution.**



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

- We use square brackets to denote concentration [X]. (also known as **molarity, M**).

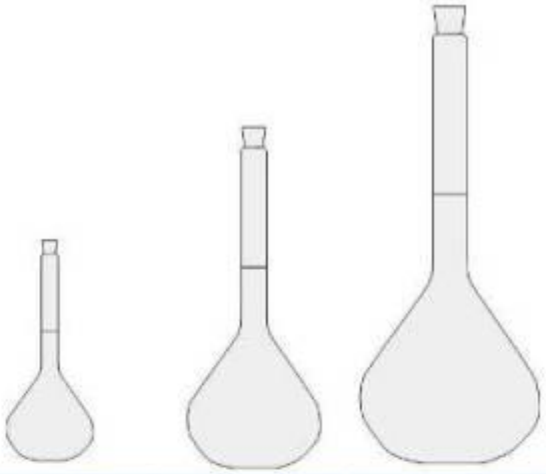
- For example – a solution of sodium hydroxide has a concentration of 1.0 Mol dm^{-3} (1.0M)

$[NaOH_{(aq)}]$ = 1.0M This means there is 1 mole of sodium hydroxide dissolved in $1dm^3$ of water.

$[NaCl_{(aq)}]$ = 2.0M This means there is 2 moles of sodium chloride dissolved in $1dm^3$ of water.

$[KOH_{(aq)}]$ = 0.5M This means there is 0.5 mole of potassium hydroxide dissolved in $1dm^3$ of water.

- This is fine so long as we keep making up $1dm^3$ solutions.
- Most lab experiments only use a $100cm^3$ or so and making up $1000cm^3$ would be waste so we need to scale down:
- For example – $500cm^3$ solution of sodium hydroxide has a concentration of 1.0 Mol dm^{-3} (1.0M)



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

$500cm^3$ $[NaOH_{(aq)}]$ = 1.0M: 0.5 moles of sodium hydroxide dissolved in $0.5dm^3$ water.

$2000cm^3$ $[NaCl_{(aq)}]$ = 2.0M: 4 moles of sodium chloride dissolved in $2dm^3$ of water.

$100cm^3$ $[KOH_{(aq)}]$ = 0.5M: 0.05 moles of sodium hydroxide dissolved in 0.1 dm^3 of water.

- The values are not usually as nice as this so we can use the following formula:-

$$\textbf{Concentration (mol dm}^{-3}\textbf{)} = \frac{\textbf{No. of moles}}{\textbf{Volume(dm}^3\textbf{)}} \qquad C = \frac{n}{V\ (dm^3)}$$

- Basically the number of moles per volume of solution.
- We usually like our formula to have **n, number of moles at the start:**

$$\textbf{No. of moles = Concentration (Mol dm}^{-3}\textbf{) x Volume (dm}^3\textbf{)} \qquad n = C \times V\ (dm^3)$$

- However this formula assumes we are working in dm^3 and we usually work in cm^3.
- **dm^3 and cm^3** are related by a **factor of 1000 – you must convert into dm^3 by dividing cm^3 by 1000**

Standard solutions

- We can make solutions of known concentration using volumetric flasks. The easiest way of learning this is to try an example.
- We need 250cm^3 of 0.1 mol dm^3 solution of sodium hydroxide.
- Use the formula to calculate the No. of moles of sodium hydroxide –

No. of moles = Concentration (Mol dm^3) x Volume (cm^3)
No. of moles = 0.1 Mol dm^3 x 250 cm^3
No. of moles = 0.025 Mol

- Now we know how many moles of sodium carbonate we need in 250 cm^3 to make a 0.800 Molar solution.
- We now need to convert moles into a mass –

Mass = No. Moles x Mr

Mass = 0.025 x 40

Mass = 1g

Mr (NaOH) = 1x23 = 23
1x16 = 16
1x 1 = $\underline{1}$
40 gMol^{-1}

- Weigh this out in a beaker.
- Dissolve in distilled water and pour into the graduated flask.
- Add more distilled water to the beaker to wash the solute into the graduated flask.
- Repeat the last step several times to ensure all the solute is in the graduated flask.
- Fill the graduated flask with distilled water so the meniscus sits on the line.
- Stopper the flask and invert several times to ensure mixing.

Mass Concentrations:

- This is the mass in g dissolved in 1 dm^3 of solution
- This means that the concentrations are measured in g dm^{-3} instead
- The concentration for the solution made above would be 1 g dm^{-3} as a mass concentration.

Concentrated vs diluted:

- This is down to the amount of solute dissolved in solution:
- **Concentrated** : Lots of solute per dm^3
- **Diluted** : A small amount of solute per dm^3

Questions 1-3 p17 / 9,10 p35 / 4 p36

<u>Moles and reactions</u>

- 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** of these require the use of the mole:

Mass / mole calculations:

Example: 2.4g Magnesium reacts in air to form magnesium oxide. Calculate the mass of magnesium oxide made:

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

$2Mg_{(s)}$	+	$O_{2(g)}$	$\rightarrow$	$2MgO_{(s)}$
2.4g				Mass ?

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

$2Mg_{(s)}$	+	$O_{2(g)}$	$\rightarrow$	$2MgO_{(s)}$
2.4g				Mass ?

Moles = mass / Ar
Moles = 2.4 / 24
Moles = 0.1

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

$2Mg_{(s)}$	+	$O_{2(g)}$	$\rightarrow$	$2MgO_{(s)}$
2.4g				Mass ?

Moles = mass / Ar $\xrightarrow{\text{Ratio is 1:1 (from 2:2)}}$ Moles = 0.1 MgO made
Moles = 2.4 / 24
Moles = 0.1

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

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

$2Mg_{(s)}$: 2.4g

Moles = mass / Ar
Moles = 2.4 / 24
Moles = 0.1

$\xrightarrow{\text{Ratio is 1:1 (from 2:2)}}$

Moles = 0.1 MgO made

$2MgO_{(s)}$:
Mass = moles x Mr
Mass = 0.1 x 40
Mass = 4g

Gas / mole calculations:

Example: 2.3g sodium reacts with excess sulphuric acid to form sodium sulphate and hydrogen gas. Calculate the volume of hydrogen made:

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

$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$

$2Na_{(s)}$: 2.3g

$H_{2(g)}$: Volume = ?

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

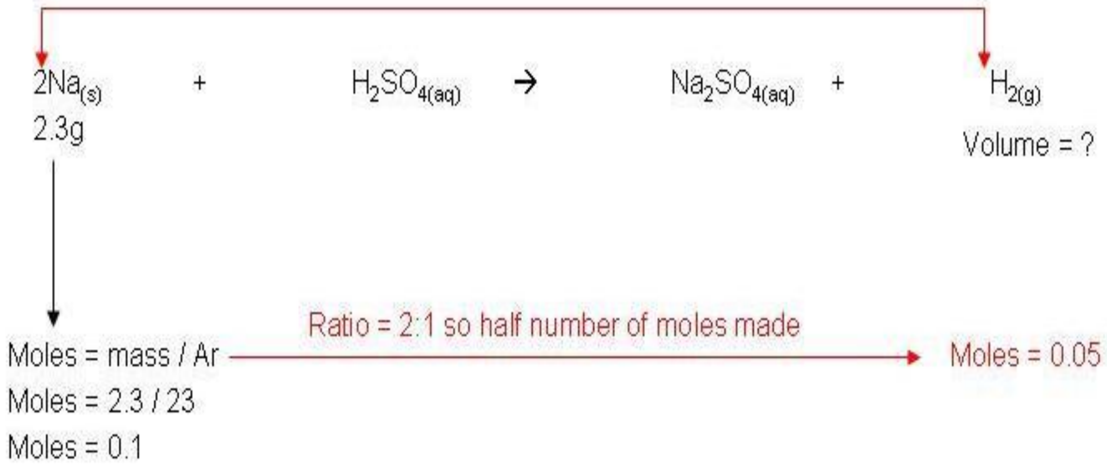
$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$

$2Na_{(s)}$: 2.3g

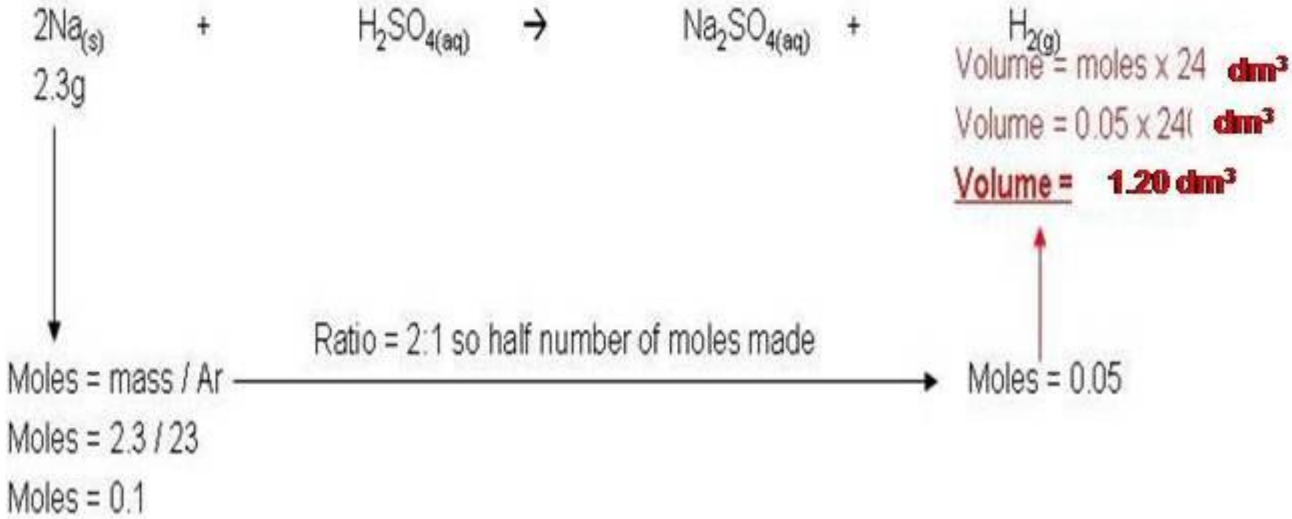
$H_{2(g)}$: Volume = ?

Moles = mass / Ar
Moles = 2.3 / 23
Moles = 0.1

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



Concentration / mole calculations:

Example: 2.3g sodium reacts with $250cm^3$ sulphuric acid to form sodium sulphate and hydrogen gas. Calculate the concentration of sodium sulphate solution made:

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

$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$

2.3g

Concentration ?

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

$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$

2.3g

Moles = mass / Ar

Moles = 2.3 / 23

Moles = 0.1

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

Next step

$$2Na_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_{2(g)}$$

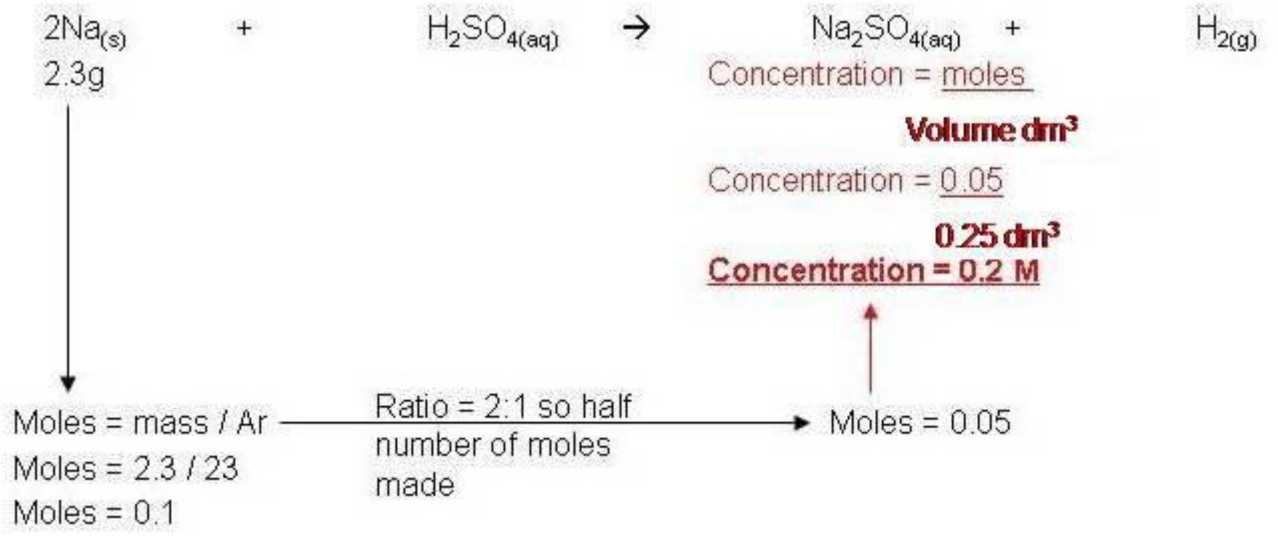
2.3g

Moles = mass / Ar $\xrightarrow{\text{Ratio = 2:1 so half number of moles made}}$ Moles = 0.05

Moles = 2.3 / 23

Moles = 0.1

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



Questions 1-3 p21 / 4 p36

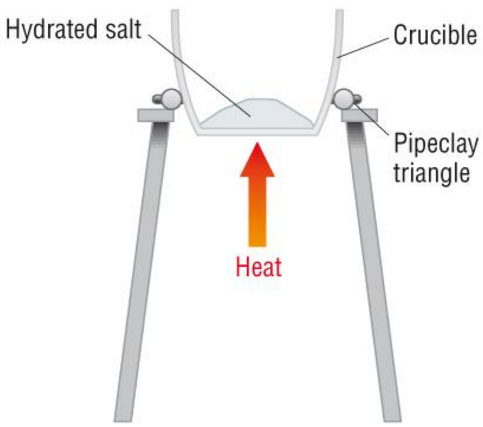
Water of crystallisation

- Coloured crystals such as blue copper sulphate contain that colour due to water locked in the crystalline structure.
- If the water is driven off then the crystals appear white.
- 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:

$CuSO_4.H_2O$

- However most crystals will contain more than 1 mole of water per mole of crystal.

- For copper sulphate, 1 mole of copper sulphate crystals will contain 5 moles of water:

$$CuSO_4.5H_2O$$

- The number of moles of water per mole of crystal depends upon that crystal:

$$CuSO_4.5H_2O$$
$$CoCl_2.6H_2O$$
$$Na_2SO_4.10H_2O$$

- The water of crystallisation can be calculated by **empirical formula** and by **moles calculations:**

From empirical formula

- Count the number of **hydrogen's and divide by 2**. Do not use oxygen as this may be in the crystal formula

$$MgCl_2H_{10}O_5 \rightarrow MgCl_2.5H_2O$$
$$Na_2CH_{20}O_{13} \rightarrow Na_2CO_3.10H_2O$$
$$CaN_2H_8O_{10} \rightarrow Ca(NO_3)_2.{}_4H_2O$$

From mole calculations (example)

- From the experiment done earlier it is possible to calculate the masses of crystal and water.

Mass of hydrated $MgSO_4.xH_2O$ = 4.312g
Mass of anhydrous $MgSO_4$ = 2.107g

- From this we can calculate the masses of the crystal and water:
- Convert these to moles of crystal and water.
- Divide through by the smallest number of moles to get a whole number ratio:

	Crystal, $MgSO_4$	**Water, H_2O**
Masses of each	2.107g	(4.312 - 2.107)
	2.107g	**2.205g**
Moles of each	2.107 / Mr	2.205 / Mr
	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 $MgSO_4.xH_2O$ = $MgSO_4.7H_2O$
Now calculate the Formula from the practical you completed earlier.

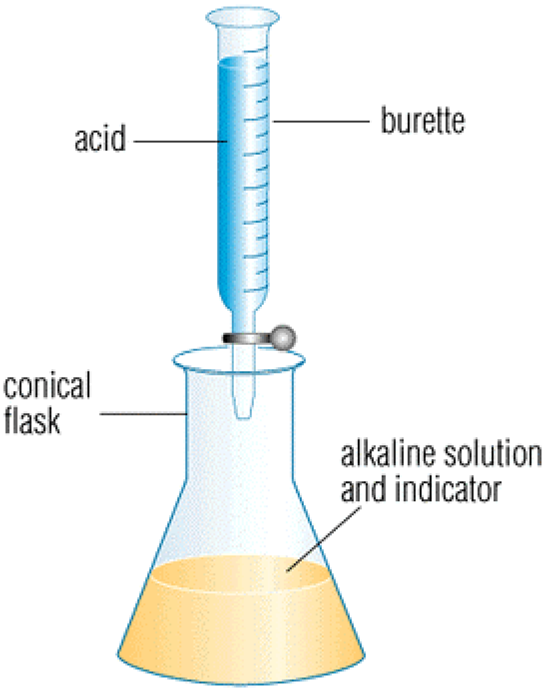
Questions 1-2 p27 / 15 p35

Titrations - Volumetric analysis

- This technique can be used to find:
  - **Concentration**
  - **Mr**
  - **Formula**
  - **Water of crystalisation**
- To do this you react a certain volume of a solution with an unknown concentration with a solution of known concentration.
- Using moles and reacting ratios, you can calculate the concentration of this solution. This is known as titration.
- 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
Bromothymol Blue	Yellow	Blue	Green
Phenylphthalein	colourless	Pink	Pale pink

Technique/procedure



- Rinse the burette with distilled water then acid.
- Fill the burette to the graduation mark ensuring the air bubble is removed from the tap.
- Rinse a pipette with alkali, fill and transfer a known volume to a clean, dry conical flask.
- Add some indicator.
- Run the acid into the alkali and stop when the colour changes. This is your **'range finder'**.
- Wash the conical flask with pure water.
- Repeat the previous steps but stop a few cm^3 before the colour change.
- Add drop wise until the first drop changes the colour of the indicator.
- Repeat until you get 2 results that agree within $0.10 cm^3$.

- Record results in a table like the one below:

	Range finder / titration / cm3	**Run 1 / cm3**	**Run 2 / cm3**
2nd burette reading			
1st burette reading			
Volume added			

Calculation – example
25cm^3 of 0.01M sulphuric acid was added to exactly neutralize 50cm^3 of sodium hydroxide. Calculate the concentration of the sulphuric acid:

1 Write a balanced equation

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_20_{(l)}$$

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

No. of moles = Concentration (Mol dm^3) x Volume (cm^3)
(H_2SO_4)

No. of moles = 0.01 Mol dm^3 x 25 cm^3
(H_2SO_4) 1000

No. of moles = 25 x 10^{-3} Moles
(H_2SO_4)

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

2 : 1 NaOH : H_2SO_4

2x the number of moles for H_2SO_4

No. of moles = 25 x 10^{-3} x 2
(NaOH)

No. of moles = 50 x 10^{-3} Mol
(NaOH)

4 Calculate the concentration. (Convert to g dm^3 if necessary)

Concentration (Mol dm^3) = $\frac{\text{No Moles}}{\text{Volume (dm}^3\text{)}}$

Concentration (Mol dm^3) = $\frac{50 \times 10^{-3}}{0.05 \text{ (dm}^3\text{)}}$

Concentration (Mol dm^3) = **1 Mol dm^{-3}**

5 Convert to g dm^3 if required (usually for solubility)

Mass = No. moles x Mr (NaOH)

Mass = 1 x 40

Concentration = 40 g dm^{-3}

Questions 1-2 p29 / 17 p35

Questions 1-2 p33 / Remaining questions p35 - 37