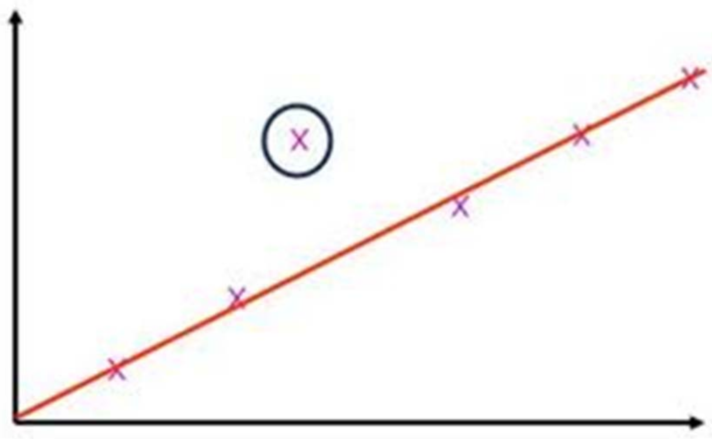
Quantitative Chemistry

When we do experiments to measure something in Chemistry, we:

- **Repeat** experiments (usually 3 times) to **improve the reliability** of the results, by **calculating an average** of our results.
- Repeats also allow us to spot an **anomaly**: a result that does not fit the pattern of the others.

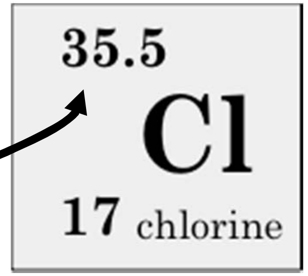
	Maximum temperature reached (°C)			
Mass of Mg (g)	**Expt 1**	**Expt 2**	**Expt 3**	**Average (°C)**
0.2	15.5	16.0	16.5	16.0
0.4	22.5	22.5	24.0	23.0
0.6	26.0	35.5	26.0	**26.0**
0.8	32.5	32.0	32.0	32.2
1.0	40.0	39.0	41.0	40.0

- If we find an anomaly, we may repeat the experiment. We usually draw a circle around the anomalous result to remind us that it **should not be included in the average**.
- We can also see anomalies on graphs. These should also be ringed and then excluded from any line of best fit.
- When we look at the values from the repeats, the closer they are together, the more reliable the results are.



Relative Atomic Masses

The periodic table does not show mass numbers, but relative atomic masses.



The **relative atomic mass** of an element ($\boldsymbol{A_r}$) is the average mass of an atom, on a scale where one atom of the **^{12}C isotope** weighs 12 exactly.

It is an **average** value, taking into account all the isotopes of the element.

Chlorine atoms can't have half a neutron. In reality some isotopes are **^{35}Cl** and some are **^{37}Cl**. The average mass is 35.5

Relative Formula Mass

The **relative formula mass (M_r)**, **sometimes written RFM**, is the **sum** of the Relative Atomic Masses of all the atoms in the substance's formula.

The **units** of relative formula mass are **g/mol**. (We'll see why later)

e.g. The formula for water is H_2O. What is its M_r ?

Add the Relative Atomic Masses of the atoms involved:

O=16 H=1 so we have 2 x H and 1 x 0 = (2 x 1) + (1 x 16)

therefore the relative formula mass of H_2O = 18 g/mol

Practice:

Work out the Relative Formula Mass, M_r, of:

Hydrochloric acid	**HCl**
Methane	**CH_4**
Carbon disulphide	**CS_2**
Copper sulphate	**$CuSO_4$**
Magnesium hydroxide	**$Mg(OH)_2$**
Ammonium sulphate	**$(NH_4)_2SO_4$**

In equations, we often need to use a number <u>in front</u> of a formula to tell us how many molecules are reacting e.g. $\mathbf{2H_2 + O_2 \rightarrow 2H_2O}$

These numbers in front are NOT part of the M_r. The M_r of water is 18 g/mol, not 36 !

Practice:

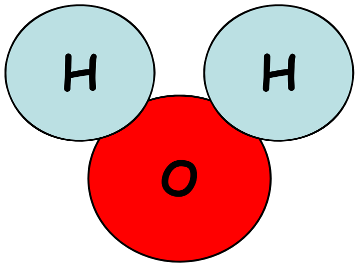
Write down the M_r (or A_r) of the substances underneath the formulae in the following equations:

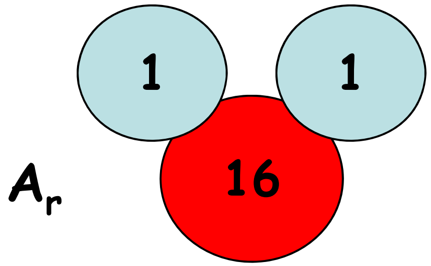
$$3H_2 + N_2 \rightarrow 2NH_3$$

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

Percentage composition (% by mass)

A water molecule contains 1 oxygen atom and 2 hydrogen atoms.





16g of every 18g of water is the oxygen atoms, 2g of every 18g of water is hydrogen atoms.

We can therefore say that water contains (16 ÷ 18) x 100 = 89% oxygen
water contains (2 ÷ 18) x 100 = 11% hydrogen

In general, the % of X in a substance is:

$$\textbf{\% of X} = \frac{\text{number of X atoms} \times A_r \text{ of X}}{M_r \text{ of substance}} \times 100$$

e.g. % of H in CH_4

$$\text{\% of H} = \frac{4 \times 1}{16} \times 100 \quad \mathbf{= 25\%}$$

What percentage (by mass) of sodium oxide (Na_2O) is sodium?

Step 1 – get the Relative Atomic Masses   Na = 23   O = 16

Step 2 – work out Relative Formula Mass of Na_2O $M_r = 23 + 23 + 16 = 62$ g/mol

Step 3 – use the equation for % composition $\% \text{ Na} = \frac{2 \times 23}{62} \times 100$ **= 74% Na**

<u>Using</u> percentages by mass

e.g. What mass of sodium is there in a 20g chunk of sodium oxide? (Answer to 1 d.p.)
(you previously worked out that the % of Na in Na_2O is 74%)

Mass of sodium = 74% of 20g $= \frac{74}{100} \times 20\text{g}$ **= 14.8g to 1 d.p.**

Practice: Work out the % by mass of **oxygen** in:

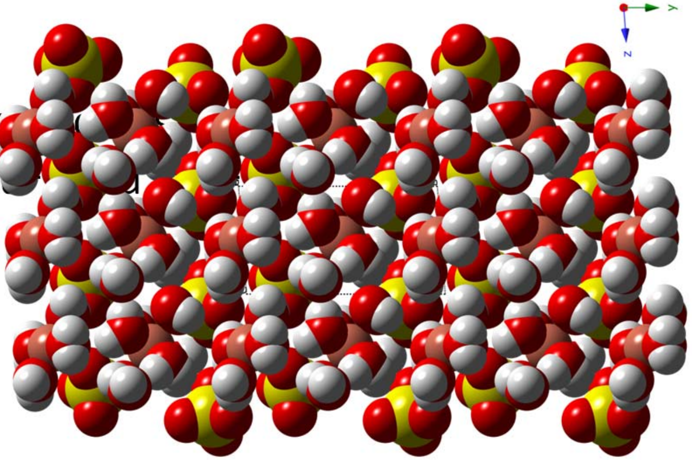
- MgO
- CO_2
- Na_2CO_3

What is the mass of oxygen in 5.5g of magnesium oxide (MgO) ?

(answers at the end of the topic)

Water of crystallisation

Crystals of ionic substances can contain fixed numbers of water molecules as part of the giant ionic lattice. e.g. the formula of copper(II) sulphate crystals shown here is **$CuSO_4 \cdot 5H_2O$**



This means there are 5 water

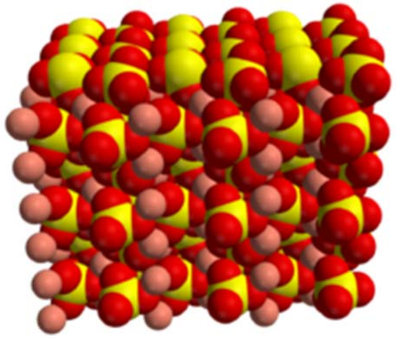
the lattice for every one

ion.

The mass of these water molecules must be included in M_r: **$CuSO_4.5H_2O$ = 63.5 + 32 + (4x16) + 5x(1+1+16)**

= 249.5 g/mol

Without the water of crystallisation, the substance is said to be **anhydrous**:



Practice: Work out the Relative Formula Mass of:

alabaster	**$CaSO_4.2H_2O$**
iron(II) sulphate	**$FeSO_4.7H_2O$**
cobalt(II) chloride	**$CoCl_2.6H_2O$**
alum	**$KAl(SO_4)_2.12H_2O$**

We can also get asked for the % by mass of water in crystals having water of crystallisation:

What is the % by mass of water in **$CuSO_4.5H_2O$** ?

In this case:

$$\% \text{ of water} = \frac{\text{number of } H_2O \times M_r \text{ of } H_2O}{M_r \text{ of hydrated substance}} \times 100$$

M_r of **$CuSO_4.5H_2O$**

$= 63.5 + 32 + (4 \times 16) + (5 \times 18)$

$= 249.5$ g/mol

$$\text{so } \% \; H_2O = \frac{5 \times 18}{249.5} \times 100 = \underline{36\%}$$

Experiment to calculate number of waters of crystallisation

The change in mass of crystals when they are heated to drive off the water of crystallisation is used to work out how many waters of crystallisation the crystals contain.

e.g. hydrated copper(II) sulphate crystals have the formula:

$CuSO_4.nH_20$ where n = integer

<u>To find 'n':</u>

- Weigh the hydrated crystals.
- Heat until all the water has been given off (the crystals turn white)
- Allow to cool and weigh the anhydrous crystals.

Example:

<u>Results</u>

Mass of hydrated crystals (before heating): 4.99g

Mass of anhydrous crystals (after heating): 3.19g

Mass of water given off = 4.99 – 3.19 = 1.80g

<u>Calculation</u>

M_r of CuSO4 = 63.5 + 32 + (4 x 16) = 159.5 g/mol

M_r of water = 1 + 1 + 16 = 18 g/mol

Use a table format like this to lay out your working:

	$CuSO_4$		**H_2O**	
mass (g)	3.19		1.80	
÷				
M_r (g/mol)	159.5		18	
=				
Ratio	0.02	:	1.00	**(divide all by smallest)**
	1	:	5	

So the formula is $CuSO_4 \cdot 5H_20$

Molecular formula

- tells you how many of each type of atom in a molecule

e.g. NH_3 has one N and three H atoms

Empirical formula

– the **simplest whole-number ratio** of the atoms present

The empirical formula can be the same as the molecular formula, but often is different.

e.g.

Name	Molecular formula	Empirical formula
water	H_2O	H_2O
ethane	C_2H_6	CH_3
glucose	$C_6H_{12}O_6$	CH_2O
benzene	C_6H_6	CH
hydrazine	N_2H_4	NH_2

Empirical formula calculations

e.g. A substance contains 4.0g calcium, 1.2g carbon and 4.8g oxygen. What is its empirical formula? *Use a table like this to lay out your working:*

	Ca	C	O	
mass	4.0	1.2	4.8	
÷ A_r	40	12	16	
= ratio	0.1	0.1	0.3	*divide all by smallest to get whole numbers*
	1 :	1 :	3	**Formula is $CaCO_3$**

We also get questions where we are given **percentage** of each element (or of all but one element – remember the percentages must add up to 100%), rather than **mass**. The method is the same:

e.g. A hydrocarbon contains 25% of hydrogen and 75% of carbon. What is its empirical formula?

	H	**C**	
mass	25%	75%	
A_r	1	12	
Ratio	25	6.25	*÷ all by smallest to get whole numbers*
	4	1	**Formula is CH_4**

Practice:

A compound of phosphorus and fluorine only, contains 24.6% phosphorus. What is its empirical formula?

Sometimes the 'whole numbers' don't come out as perfect integers – usually because of rounding errors e.g. in the masses used. If your answer is NEARLY a whole number e.g. 2.997 or 3.0017 … then you should round it to a whole number.

If your answer is in-between whole numbers, it may be a fraction (e.g. ½):

	Al		**O**	
Ratio	1	:	1.5	… deal with this by multiplying everything up x 2
	2	:	3	**Formula is Al_2O_3**

Experiment to determine empirical formula of a substance

The mass change when an element combines with oxygen can be used to work out the empirical formula of the oxide. e.g. Magnesium ribbon burns in air to form white magnesium oxide.

Method:

- weigh a crucible and lid.
- place magnesium ribbon in crucible with lid, and reweigh.
- heat crucible until magnesium burns.
- lift the lid occasionally until there is no further reaction.
- allow the crucible and lid to cool, and reweigh.
- repeat the heating, cooling and reweighing until two consecutive masses are the same, to make sure that all of the magnesium has reacted
- calculate the mass of magnesium oxide formed.

e.g. Results

Mass of magnesium before burning:	= 1.20g
Mass of magnesium oxide after burning:	= 2.00g
Mass of oxygen that reacted = 2.00 – 1.20	= 0.80g

Calculation	**Mg**		**O**
mass (g)	1.20		0.80
÷ A_r (g/mol)	24		16
= Ratio	0.05	:	0.05
	1	:	1

The formula of magnesium oxide is **MgO**

From empirical to molecular formula

If you have worked out the empirical formula for a substance and you know the relative formula mass, M_r, of the substance, then you can work out the molecular formula.

1. Add up the masses of the atoms in the empirical formula
2. Divide by the relative formula mass
3. Multiply all the numbers in the empirical formula by this amount

e.g. The empirical formula of a substance is found to be NO_2. The relative formula mass is found by mass spectrometry experiments, and found to be 92. What is the molecular formula of this substance?

Empirical formula = NO_2 $M_r = 92$
Step 1: $NO_2 = 14 + 16 + 16 =$ **46**
Step 2: $92/46 =$ **2**
Step 3: NO_2 x 2 so **molecular formula is N_2O_4**

Practice:

Determine the molecular formula of these substances:

empirical formula HO $M_r = 34$
empirical formula CH_2 $M_r = 56$
empirical formula CH $M_r = 78$

The Mole

The relative formula mass of a substance, weighed out in grams, is known as one mole of that substance.

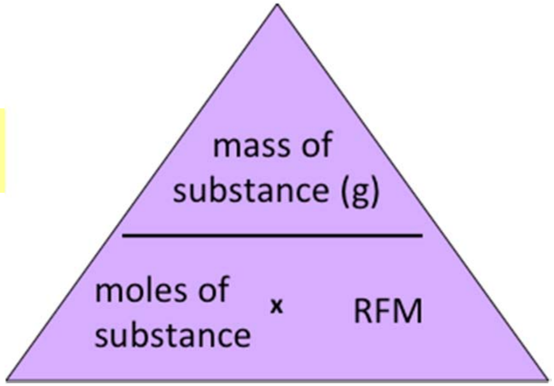
This means that the **units** of relative atomic mass, and relative formula mass, are grams per mole: **g/mol.**

One mole of any substance contains exactly the same AMOUNT of that substance as one mole of any OTHER substance. e.g. 18g of water (H_2O) contains exactly the same number of molecules as 2g of hydrogen (H_2) or 32g of oxygen (O_2).

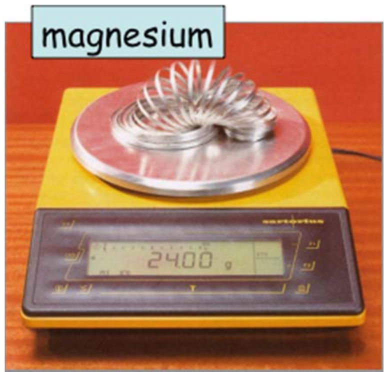
Converting between mass and moles:

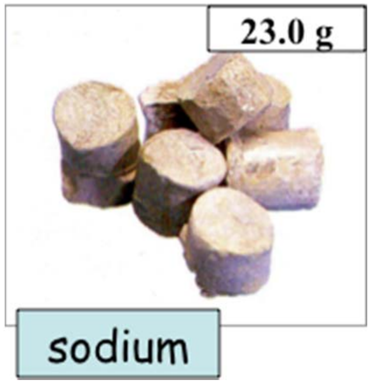
mass (g) = moles of substance x relative formula mass

$$\textbf{moles} = \frac{\textbf{mass of substance (g)}}{\textbf{relative formula mass}}$$



One mole of different elements





Practice:

How many moles are there in:

- 4g of hydrogen (H_2)
- 36g of carbon atoms (C)
- 160g of ozone (O_3)

What is the mass of:

- 0.1 moles of copper oxide (CuO)
- 2 moles of water
- 10 moles of ammonia (NH_3)
- 0.2 moles of ethane (C_2H_6)

Calculating Reacting Quantities

Using the chemical equation for a reaction we can use moles to work out what mass of product we might make, or what mass of reactants we need.

e.g. "42g of nitrogen (N_2) are reacted with hydrogen (H_2) to form ammonia (NH_3) according to the equation $N_2 + 3H_2 \rightarrow 2NH_3$. What mass of ammonia will be made?"

Step 1 – What can I work out moles of? *(which substance do I know formula and mass for?)*

Moles of N_2 = mass of $N_2 \div M_r$ *M_r of N_2 = 14 + 14 = 28 g/mol*

Moles of N_2 *= 42 ÷ 28* ***= 1.5***

Step 2 – use the equation to get mole ratios **$N_2 + 3\ H_2 \rightarrow 2\ NH_3$**

	N_2		$3\ H_2$		$2\ NH_3$
Ratio :	*1*	*:*	*3*	*:*	*2*
Moles:	***1.5***	*:*	*4.5*	*:*	***3.0***

Write in the number of moles you calculated in step 1, then use the ratios to work out the moles of the other substances in the equation.

Step 3 – work out the mass you were asked for in the question

Mass of NH_3 = moles of NH_3 x M_r *M_r of NH_3 = 14 + (3 x 1) = 17 g/mol*

The mass of ammonia *=* ***3.0*** *x 17* ***= 51g***

% Yield

Reactions often do not go all the way to completion, or we don't manage to recover all the product made.

- product left behind in the apparatus
- difficulty separating product from reaction mixture

Our calculations assumed **all** our reactants would end up as products. The amount of product we **actually** obtain is called the **yield.**

$$\text{Percentage yield (\% yield)} = \frac{\text{amount actually made}}{\text{maximum amount possible}} \times 100$$

The maximum amount possible is **calculated**, assuming all the reactants react completely. The amount you actually make is **measured**.

Example: Carbon burns in oxygen to make carbon dioxide, but other combustion products can also be produced. 12g of carbon was burnt, and produced 33g of CO_2. What was the percentage yield of carbon dioxide? $C + O_2 \rightarrow CO_2$

Step 1: Work out the maximum amount of CO_2 which can be made. (using moles)

12g carbon = 1 mole (moles = mass/M_r = 12/12)

mole ratio: 1 mole C makes one mole of CO_2

mass CO_2 = moles x M_r of CO_2 = 1 x 44 **= 44 g**

Step 2: Work out % yield using actual mass and maximum mass

$$\text{\% yield} = \frac{\text{actual amount (33g)}}{\text{maximum amount (44g)}} \times 100 = 75\%$$

Using Avogadro's number

602,000,000,000,000,000,000,000 atoms

Avogadro's number is used to convert between moles of a substance and the actual number of particles (atoms, molecules, ions etc.)

Symbol: $\mathbf{N_A}$ Value: $\mathbf{6.0 \times 10^{23}}$

"number of…" = "moles of…" x N_A

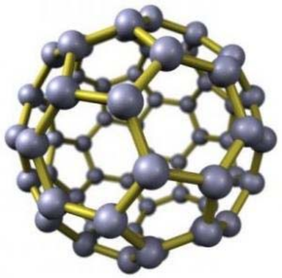
e.g. How many molecules are there in 90g of water?

Step 1: **convert mass of water to moles of water**

moles of H_2O = mass of H_20 ÷ M_r of H_2O = 90/18 **= 5.0 moles**

Step 2: **multiply moles by N_A to get number of particles** (molecules in this case)

molecules = moles x N_A = 5.0 x 6.0×10^{23} **= 3.0×10^{24} molecules**



e.g. A buckyball contains 60 carbon atoms. What does it weigh ?

Step 1: use N_A to work out how many moles of C_{60} one molecule represents

moles of C_{60} = molecules of C_{60} ÷ N_A = 1 ÷ 6.0×10^{23} **= 1.67×10^{-24} moles**

Step 2: convert moles to mass M_r of C_{60} = 60 x 12 = 720 g/mol

Mass of buckyball = moles of C_{60} x M_r of C_{60} = 1.67×10^{-24} x 720 **= 1.2×10^{-21} g**

Moles of gases

It is not always convenient to work with masses of a gas. We usually measure volumes instead.

1 mole of **any gas** has a volume of **24.0 dm^3** at room temperature and pressure. This is called the **molar volume**. **N.B. 1 dm^3 = 1000cm^3**

number of moles = $\dfrac{\text{volume of gas } (dm^3)}{\text{molar volume}}$

volume (dm^3) = number of moles x molar volume

Practice converting volume to moles:

How many moles of gas in:

i) 6.0 dm^3 of CO_2
ii) 2.4 dm^3 of NH_3
iii) 240cm^3 of O_2

What volume would be occupied by:

i) 0.5 moles of CH_4
ii) 2.0 moles of H_2
iii) 0.0125 moles of N_2

We can come across volumes of gas within reacting quantity (mole) calculations:

More practice: "What volume of gas would be collected if 10g of calcium carbonate was heated until it thermally decomposed: **$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$** "

Hint: calculate moles of $CaCO_3$ decomposing, then use the 1:1:1 mole ratio in the equation to work out moles of CO_2 produced, then convert moles of CO_2 to volume of CO_2.

Moles in solution

Concentrated: a concentrated acid (or alkali) has a large number of acid molecules per cm^3 of aqueous solution.

Dilute: a dilute acid (or alkali) has a small number of acid molecules per cm^3 of aqueous solution.

The units of concentration are **mol/dm³** Note: $1 dm^3 = 1000 cm^3$ and $1 dm^3 = 1$ litre

A solution with a concentration of **1 mol/dm³** has one mole of the solute dissolved in 1 dm^3 of the solution.

A solution of 0.1 mol/dm^3 is only a tenth of the concentration, i.e. it is ten times more dilute.

$$\textbf{concentration (mol/dm}^3\textbf{)} = \frac{\textbf{moles}}{\textbf{volume (in dm}^3\textbf{)}}$$

e.g. 7.3g of HCl are dissolved in 0.1 dm^3 (100cm^3) of water. What is the concentration of the HCl solution?

M_r of HCl = 1 + 35.5 = 36.5
moles of HCl = mass of HCl / M_r of HCl = 7.3 ÷ 36.5 ***= 0.2 moles***
concentration = moles of HCl ÷ volume of solution in dm^3 = 0.2 ÷ 0.1 = ***2 mol/dm³***

We can also work out how many moles are in a solution if we know its concentration and its volume:

moles = concentration (in mol/dm^3) x volume (in dm^3)

e.g. How many moles of sodium hydroxide (NaOH) would I need to dissolve to make up 500cm^3 of solution with 0.1 mol/dm^3 concentration ? **Remember 1000cm^3 = 1dm^3**

Step 1: Work out how many moles of NaOH would be needed (and convert volume to dm^3)

moles of NaOH = concentration x volume in dm^3 = 0.1 x 0.5 ***= 0.05 mol***

Step 2: Work out the mass of sodium hydroxide needed for this number of moles

M_r of NaOH = 23 + 16 + 1 = 40

mass of NaOH = moles of NaOH x M_r of NaOH = 0.05 x 40 ***= 2.0g***

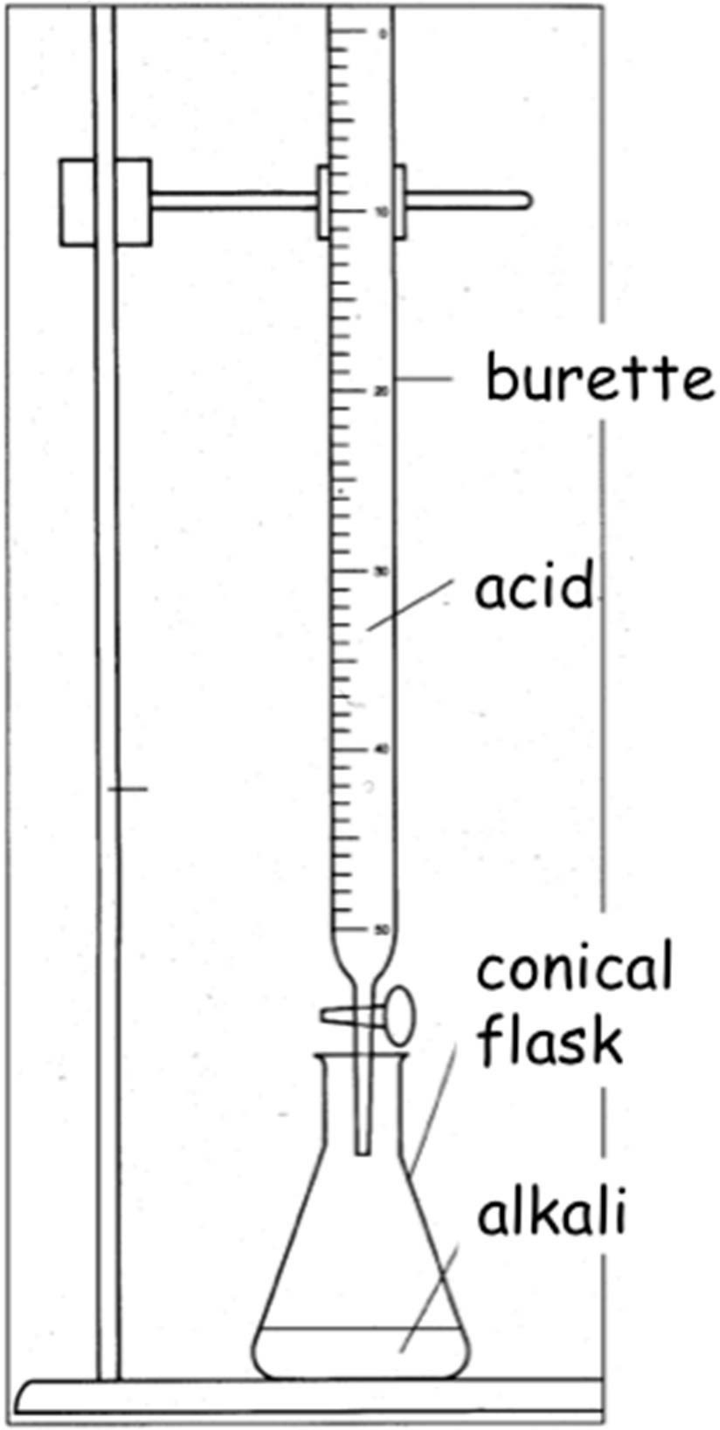
Practice:

How many moles of sodium chloride would I need in order to make 250cm^3 of solution with concentration 2 mol/dm^3 ?

What would the mass of the sodium chloride be ?

If I only had 5.85 g of sodium chloride, what concentration solution would I get if I dissolved it to make 100cm^3 of solution ?

(answers for all questions at the end of the topic)



Titration

Titration is a technique used to measure how much of an acid is needed to exactly neutralise an alkali.

If we know the concentration of either the acid or the alkali, we can use titration to find the concentration of the other.

How to do a titration:

A volume of alkali is measured into the flask, using a **pipette**. (more accurate than a measuring cylinder)

A few drops of **indicator** are mixed with the alkali.

The level of the acid in the **burette** is noted; burettes are read to the **nearest 0.05cm^3** (half a division)

Acid is added **dropwise** until the indicator just changes colour. This is called the **endpoint**.

The acid level is noted again, and the volume of acid that has been added (the **titre**) is worked out.

Repeat titrations are done until you get two **consistent** titres (within 0.2cm^3 of each other), which are then **averaged**.

Finding the concentration of an acid:

e.g. "A conical flask contained 25.0cm^3 of NaOH solution and its concentration was 0.10 mol/dm^3. When titrated, the indicator changed colour after 16.50cm^3 of HCl of unknown concentration had been added. Work out the concentration of the acid."

$$HCl + NaOH \rightarrow NaCl + H_2O$$

Step 1: Work out the number of moles of alkali (NaOH) in the flask

moles of NaOH = concentration of NaOH x volume of NaOH ***in dm^3***

= 0.10 x 0.025 dm^3 (25cm^3 = 0.025dm^3)

= 0.0025 moles of NaOH

Step 2: Use the mole ratio from a balanced equation

Knowing the moles of alkali, you can work out how many moles of acid must have been added to exactly neutralise the alkali, using the mole ratio from the equation:

$$NaOH + HCl \rightarrow NaCl + H_2O$$

1 : 1 : 1 : 1

0.0025 : 0.0025 *Therefore 0.0025 moles of acid was added*

Step 3: Calculate concentration of the acid

Use the volume of acid and the moles of acid to work out the concentration of the acid:

concentration of HCl = moles of HCl ÷ volume of HCl (in dm^3)

= 0.0025 ÷ 0.0165 dm^3 (16.50 cm^3 = 0.0165dm^3)

= 0.15 mol/dm^3

Finding the concentration of an alkali:

e.g. A conical flask contained 25.0cm^3 of NaOH solution and its concentration was unknown. The indicator changed colour after 10.00cm^3 of H_2SO_4 of 0.10 mol/dm^3 concentration had been added. Work out the concentration of the alkali.

THE STEPS ARE ESSENTIALLY THE SAME:

1) Show that the moles of acid used (conc x vol) = 0.001 mol
2) Use the mole ratio to show that moles of alkali used = 0.002 mol
3) convert moles of alkali to concentration of alkali (mol/vol) = 0.08 mol/dm^3

A table format can be used to lay out the calculation:
Calculated values in purple, given values in black.

Equation	$2NaOH$	$+ H_2SO_4$	$\rightarrow$ $NaCl + H_2O$	
Concentration (mol/dm^3)	*0.08*	*0.10*		
Volume (dm^3)	*0.025*	*0.010*	*N.B*	*10cm3 = 0.010dm^3*
Moles	*0.002*	*0.001*		*25cm3 = 0.025dm^3*
Ratio	*2 :*	*1*		

The concentration of the alkali was 0.08 mol/dm^3

Answers:

Work out the Relative Formula Mass, M_r, of:

Hydrochloric acid	**HCl**	36.5 g/mol
Methane	**CH_4**	16 g/mol
Carbon disulphide	**CS_2**	76 g/mol
Copper sulphate	**$CuSO_4$**	160 g/mol
Magnesium hydroxide	**$Mg(OH)_2$**	58 g/mol
Ammonium sulphate	**$(NH_4)_2SO_4$**	132 g/mol

Write down the M_r (or A_r) of the substances underneath the formulae in the following equations:

$$\underset{2}{3H_2} + \underset{28}{N_2} \rightarrow \underset{17}{2NH_3}$$

$$\underset{27}{4Al} + \underset{32}{3O_2} \rightarrow \underset{102}{2Al_2O_3}$$

Work out the % by mass of **oxygen** in:

MgO = 1 x 16 x 100/(24+16) **= 40%**

CO_2 = 2 x 16 x 100/(12+16+16) **= 73%**

Na_2CO_3 = 3 x 16 x 100/106 **= 45%**

What is the mass of oxygen in 5.5g of magnesium oxide (MgO) ?

% O in MgO = 40% (see above)

$$40\% \text{ of } 5.5g = \frac{5.5 \times 40}{100} = \mathbf{2.2g}$$

Work out the Relative Formula Mass of:

alabaster	$\mathbf{CaSO_4.2H_2O}$	**172 g/mol**
iron(II) sulphate	$\mathbf{FeSO_4.7H_2O}$	**278 g/mol**
cobalt(II) chloride	$\mathbf{CoCl_2.6H_2O}$	**238 g/mol**
alum	$\mathbf{KAl(SO_4)_2.12H_2O}$	**474 g/mol**

A compound of phosphorus and fluorine only, contains 24.6% phosphorus. What is its empirical formula?

	P	**F**	
mass	24.6	(100 – 24.6) = 75.4	
A_r	31	19	
ratio	0.793548…	3.96842…	
	1	5	**Formula is PF_5**

Determine the molecular formula of these substances:

empirical formula HO	$M_r = 34$	H_2O_2
empirical formula CH_2	$M_r = 56$	C_4H_8
empirical formula CH	$M_r = 78$	C_6H_6

How many moles are there in**:**

4g of hydrogen (H_2)	2
36g of carbon atoms (C)	3
160g of ozone (O_3)	3.33..

What is the mass of:

0.1 moles of copper oxide (CuO)	8g
2 moles of water	36g
10 moles of ammonia (NH_3)	170g
0.2 moles of ethane (C_2H_6)	6g

Practice converting volume to moles:

How many moles of gas in:

i)	6.0 dm^3 of CO_2	**0.25 mol**
ii)	2.4 dm^3 of NH_3	**0.1 mol**
iii)	240cm^3 of O_2	**0.01 mol**

What volume would be occupied by:

i)	0.5 moles of CH_4	**12 dm^3**
ii)	2.0 moles of H_2	**48 dm^3**
iii)	0.0125 moles of N_2	**300 cm^3**

What volume of gas would be collected if 10g of calcium carbonate was heated until it thermally decomposed **$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$**

M_r of $CaCO_3$ = 40 + 12 + (3 x 16) = 100
Moles of $CaCO_3$ = mass ÷ Mr = 10/100 = 0.1 moles
Mole ratio is 1:1 so 0.1 moles of $CaCO_3$ produces 0.1 moles of CO_2
Volume of CO_2 (dm^3) = moles x molar volume = 0.1 x 24 **= 2.4 dm^3**

How many moles of sodium chloride would I need in order to make $250cm^3$ of solution with concentration 2 mol/dm^3 ?

$250cm^3 = 0.25\ dm^3$

Moles pf NaCl = conc x vol (dm^3) = 2.0 x 0.25 **= 0.5 moles**

What would the mass of the sodium chloride be ?

M_r of NaCl = 23 + 35.5 = 58.5

Mass of NaCl = moles x M_r = 0.5 x 58.5 **= 29.25g**

If I only had 5.85 g of sodium chloride, what concentration solution would I get if I dissolved it to make $100cm^3$ of solution ?

$100cm^3 = 0.1dm^3$

Moles of NaCl = mass of NaCl / M_r = 5.85 / 58.5 = 0.1 moles

Concentration = moles / volume (dm^3) = 0.1 / 0.1 **= 1.0 mol/dm3**