IB Chemistry DP

YOUR NOTES

1. Stoichiometric Relationships

CONTENTS

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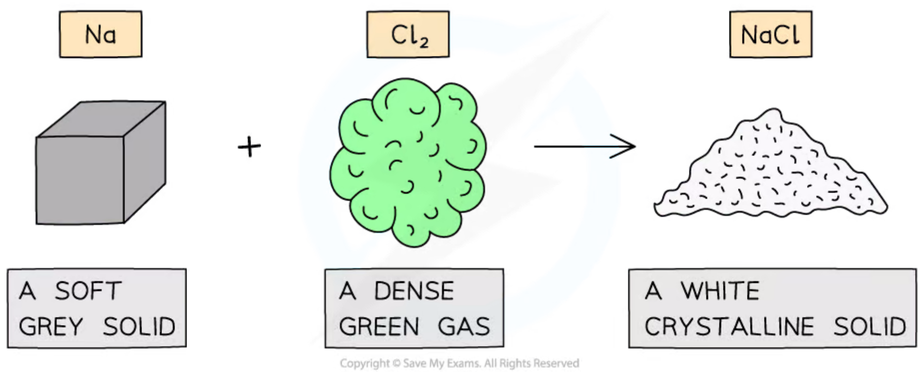
1.1 Matter, Chemical Change & the Mole Concept

1.1.1 Elements, Compounds & Mixtures

Elements & Compounds

- Elements are substances made from one kind of atom
- Compounds are made from two or more elements **chemically combined**
- Elements take part in chemical reactions in which new substances are made in processes that most often involve an energy change
- In these reactions, atoms combine together in **fixed ratios** that will give them full **outer shells** of electrons, producing **compounds**
- The properties of compounds can be quite different from the elements that form them



The properties of sodium chloride are quite different from sodium and chlorine

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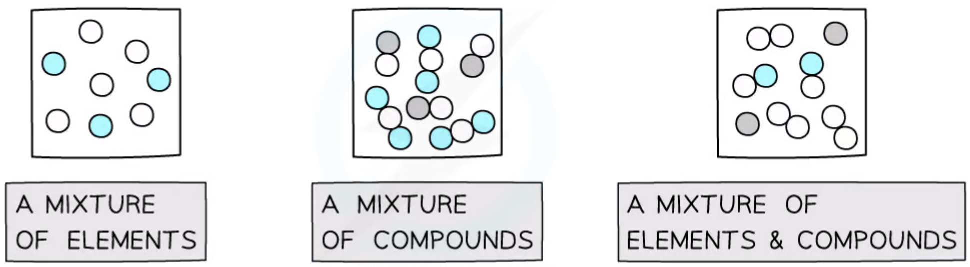
Mixtures

- In a mixture, elements and compounds are interspersed with each other, but are **not** chemically combined
- This means the components of a mixture retain the **same** characteristic properties as when they are in their pure form
- So, for example, the gases nitrogen and oxygen when mixed in air, retain the same characteristic properties as they would have if they were separate
- Substances will burn in air because the oxygen present in the air supports **combustion**



Mixtures at the molecular level

Homogeneous or heterogeneous

- A **homogeneous** mixture has uniform composition and properties throughout
- A **heterogeneous** mixture has non-uniform composition, so its properties are not the same throughout
- It is often possible to see the separate components in a **heterogeneous mixture**, but not in a **homogeneous mixture**

Types of Mixtures

Mixture	Homogeneous or heterogeneous
Air	Homogeneous
Bronze (an alloy)	Homogeneous
Concrete	Heterogeneous
Orange juice with pulp	Heterogeneous

Separating Mixtures

- The components retain their individual properties in a mixture and we can often separate them relatively easily. The technique we choose to achieve this will take advantage of a suitable difference in the physical properties of the components

Mixtures & Separation Techniques

Mixture	What technique can be used to separate the components	The property that is different in the components
Air	Fractional distillation (of liquid air)	Boiling points
Salt and sand	Solution and filtration	Solubility in water
Pigments in food colours	Paper chromatography	Adsorption (on cellulose)
Sulfur and iron	Use a magnet	Magnetism

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Balancing Equations

- A **symbol** equation is a shorthand way of describing a chemical reaction using **chemical symbols** to show the number and type of each atom in the reactants and products
- A **word** equation is a longer way of describing a chemical reaction using only **words** to show the reactants and products

Balancing equations

- During chemical reactions, atoms cannot be **created** or **destroyed**
- The number of each atom on each side of the reaction must therefore be the **same**
  - E.g. the reaction needs to be **balanced**
- When balancing equations remember:
  - Not to change any of the formulae
  - To put the numbers used to balance the equation **in front** of the formulae
  - To balance firstly the carbon, then the hydrogen and finally the oxygen in **combustion reactions** of organic compounds
- When balancing equations follow the following the steps:
  - Write the formulae of the reactants and products
  - Count the numbers of atoms in each reactant and product
  - Balance the atoms one at a time until all the atoms are balanced
  - Use appropriate state symbols in the equation
- The **physical state** of reactants and products in a chemical reaction is specified by using **state symbols**
  - **(s)** solid
  - **(l)** liquid
  - **(g)** gas
  - **(aq)** aqueous

Ionic equations

- In aqueous solutions ionic compounds **dissociate** into their ions
- Many chemical reactions in aqueous solutions involve ionic compounds, however only some of the ions in solution take part in the reactions
- The ions that do **not** take part in the reaction are called **spectator ions**
- An **ionic equation** shows **only** the ions or other particles taking part in a reaction, without showing the spectator ions

Worked Example

Balance the following equation:

Magnesium + Oxygen → Magnesium oxide

Answer:

Step 1: Write out the symbol equation showing reactants and products

$$Mg + O_2 \rightarrow MgO$$

Step 2: Count the numbers of atoms in each reactant and product

	Mg	O
Reactants	1	2
Products	1	1

Step 3: Balance the atoms one at a time until all the atoms are balanced

$$2Mg + O_2 \rightarrow 2MgO$$

This is now showing that 2 moles of magnesium react with 1 mole of oxygen to form 2 moles of magnesium oxide

Step 4: Use appropriate **state symbols** in the fully balanced equation

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Worked Example

1. Balance the following equation

   Zinc + copper(II) sulfate → Zinc sulfate + copper
2. Write down the ionic equation for the above reaction

Answer 1:

Step 1: To balance the equation, write out the symbol equation showing reactants and products

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Step 2: Count the numbers of atoms in each reactant and product. The equation is already balanced

	Zn	Cu	S	O
Reactants	1	1	1	4
Products	1	1	1	4

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Step 3: Use appropriate **state symbols** in the equation

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

Answer 2:

Step 1: The full chemical equation for the reaction is

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

Step 2: Break down reactants into their respective ions

$$Zn\,(s) + Cu^{2+}\,SO_4^{\,2-}\,(aq) \rightarrow Zn^{2+}SO_4^{\,2-}\,(aq) + Cu\,(s)$$

Step 3: Cancel the spectator ions on both sides to give the ionic equation

$$Zn\,(s) + Cu^{2+}\cancel{SO_4^{\,2-}}\,(aq) \rightarrow Zn^{2+}\cancel{SO_4^{\,2-}}\,(aq) + Cu\,(s)$$

$$Zn\,(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}\,(aq) + Cu\,(s)$$

1.1.3 State Changes

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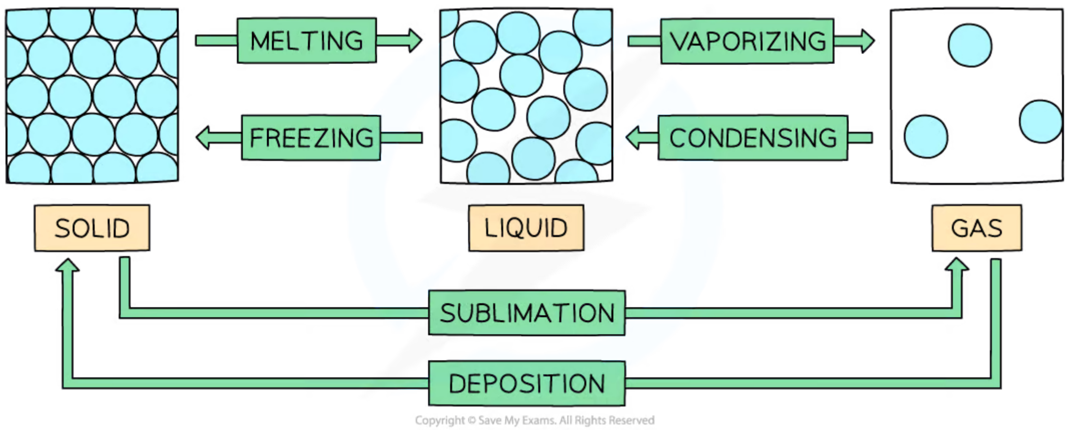
State Changes

- Changes of state are **physical changes** that are reversible
- These changes do not change the chemical properties or chemical makeup of the substances involved
- **Vaporization** includes **evaporation** and **boiling**
- **Evaporation** involves the change of liquid to gas, but unlike boiling, **evaporation** occurs only at the surface and takes place at temperatures below the **boiling point**
- **Boiling** occurs at a specific temperature and takes place when the **vapour pressure** reaches the external atmospheric pressure



State Changes

- The relationship between temperature and energy during state changes can be represented graphically

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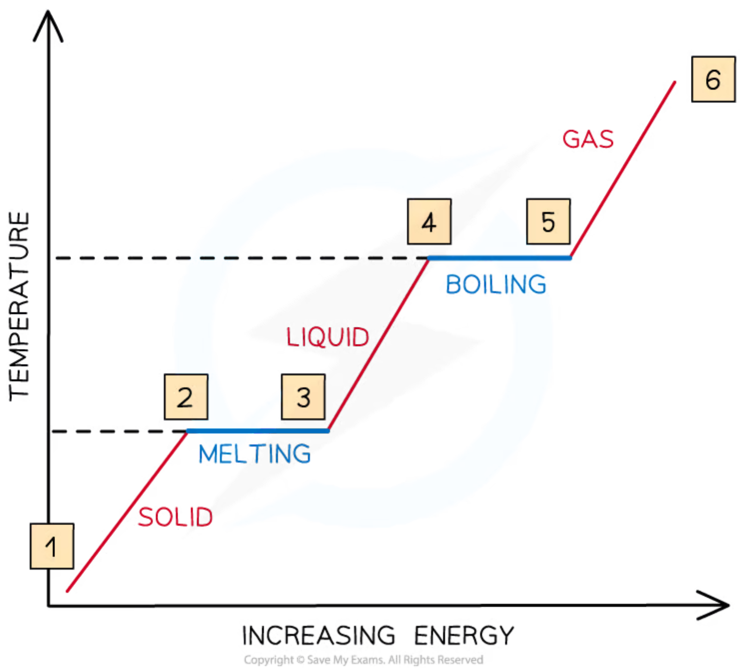

The relationship between temperature and energy during state changes

- Between 1 & 2, the particles are vibrating and gaining **kinetic energy** and the temperature rises
- Between 2 & 3, all the energy goes into breaking bonds – there is **no** increase in **kinetic energy** or **temperature**
- Between 3 & 4, the particles are moving around and gaining in **kinetic energy**
- Between 4 & 5, the substance is boiling, so bonds are breaking and there is **no** increase in **kinetic energy** or **temperature**
- From 5 & 6, the particles are moving around rapidly and increasing in **kinetic energy**

Exam Tip

Be careful to match the bond breaking or making processes to the flow of energy during state changes. Remember that to **break** bonds, energy is always **needed** to overcome the **forces of attraction** between the particles

1.1.4 The Mole Concept

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The Mole

- The **Avogadro constant** (N_A or L) is the number of particles equivalent to the relative **atomic mass** or **molecular mass** of a substance in grams
  - The Avogadro constant applies to atoms, molecules and ions
  - The value of the Avogadro constant is **6.02×10^{23} g mol^{-1}**
- The mass of a substance with this number of particles is called the **molar mass**
  - **One mole** of a substance contains the same number of fundamental units as there are atoms in exactly 12.00 g of ^{12}C
  - If you had 6.02×10^{23} atoms of carbon-12 in your hand, you would have a mass of exactly 12.00 g
  - One mole of water would have a mass of (2 x 1.01 + 16.00) = 18.02 g

Worked Example

Determine the number of atoms, molecules and the relative mass of 1 mole of:

1. Na
2. H_2
3. NaCl

Answer 1:

- The relative atomic mass of Na is 22.99
- Therefore, 1 mol of Na has a mass of 22.99 g mol^{-1}
- 1 mol of Na will contain **6.02×10^{23} atoms of Na** (Avogadro's constant)

Answer 2:

- The relative atomic mass of H is 1.01
- Since there are 2 H atoms in H_2, the mass of 1 mol of H_2 is (2 x 1.01) 2.02 g mol^{-1}
- 1 mol of H_2 will contain **6.02×10^{23} molecules of H_2**
- However, since there are 2 H atoms in each molecule of H_2, 1 mol of H_2 molecules will contain **1.204×10^{24} H atoms**

Answer 3:

- The relative atomic masses of Na and Cl are 22.99 and 35.45 respectively
- Therefore, 1 mol of NaCl has a mass of (22.99 + 35.45) 58.44 g mol^{-1}
- 1 mol of NaCl will contain **6.02×10^{23}** formula units of NaCl
- Since there is both an Na and a Cl atom in NaCl, 1 mol of NaCl will contain **1.204×10^{24} atoms** in total

1 mole of	Number of atoms	Number of molecules/ formula units	Relative mass
Na	6.02×10^{23}	–	22.99
H_2	1.204×10^{24}	6.02×10^{23}	2.02
NaCl	1.204×10^{24}	6.02×10^{23}	58.44

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Relative Mass

Relative atomic mass, A_r

- The **relative atomic mass** (A_r) of an element is the weighted average mass of one atom compared to one twelfth the mass of a carbon-12 atom
- The relative atomic mass is determined by using the weighted average mass of the **isotopes** of a particular element
- The A_r has **no units** as it is a ratio and the units cancel each other out

$$A_r = \frac{weighted\ average\ mass\ of\ one\ atom\ of\ an\ element}{\frac{1}{12} mass\ of\ one\ atom\ of\ carbon\!-\!12}$$

Relative isotopic mass

- The **relative isotopic mass is** the mass of a particular atom of an **isotope** compared to one twelfth the mass of a carbon-12 atom
- Atoms of the same element with a different number of neutrons are called **isotopes**
- **Isotopes** are represented by writing the **mass number** as ^{20}Ne, or neon-20 or Ne-20
  - To calculate the average atomic mass of an element the **percentage abundance** is taken into account
  - Multiply the atomic mass by the percentage abundance for each isotope and add them all together
  - Divide by 100 to get average relative atomic mass
  - This is known as the **weighted average** of the masses of the isotopes

$$Relative\ atomic\ mass = \frac{\Sigma(isotope\ abundance \times relative\ isotopic\ mass)}{100}$$

Relative molecular mass, M_r

- The **relative molecular mass** (M_r) is the weighted average mass of a molecule compared to one twelfth the mass of a carbon-12 atom
- The M_r has **no units**

$$M_r = \frac{weighted\ average\ mass\ of\ one\ molecule\ of\ a\ compound}{\frac{1}{12} mass\ of\ one\ atom\ of\ carbon\!-\!12}$$

- The M_r can be found by adding up the **relative atomic masses** of all atoms present in one molecule
- When calculating the M_r the **simplest formula** for the compound is used, also known as the **formula unit**
  - E.g. Silicon dioxide has a giant covalent structure, but the simplest formula (the **formula unit**) is SiO_2

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Substance	Atoms present	Mr
Hydrogen (H_2)	2 × H	(2 × 1.01) = 2.02
Water (H_2O)	(2 × H) + (1 × O)	(2 × 1.01) + 16.00 = 18.02
Potassium Carbonate (K_2CO_3)	(2 × K) + (1 × C) + (3 × O)	(2 × 39.10) + 12.01 + (3 × 16.00) = 138.21
Calcium Hydroxide ($Ca(OH)_2$)	(1 × Ca) + (2 × O) + (2 × H)	40.08 × (2 × 16.00) + (2 × 1.01) = 74.10
Ammonium Sulfate ($(NH_4)_2SO_4$)	(2 × N) + (8 × H) + (1 × S) + (4 × O)	(2 × 14.01) + (8 × 1.01) + 32.07 + (4 × 16.00) = 132.17

Relative formula mass, M_r

- The **relative formula mass** (M_r) is used for compounds containing **ions**
- It has the same units and is calculated in the same way as the **relative molecular mass**
- In the table above, the M_r for potassium carbonate, calcium hydroxide and ammonium sulfate are relative formula masses

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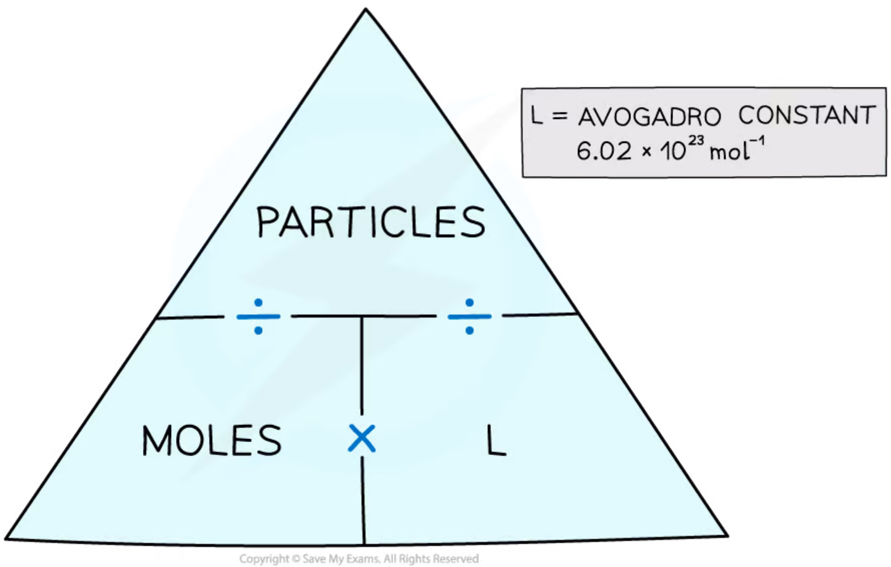
Moles, Particles & Masses

- Since atoms are so small, any sensible laboratory quantity of substance must contain a huge number of atoms
- Such numbers are not convenient to work with, so using **moles** is a better unit to deal with the sort of quantities of substance normally being measured
- When we need to know the number of particles of a substance, we usually count the number of **moles**
- The number of **moles** or particles can be calculated easily using a formula triangle



The moles and particles formula triangle – cover with your finger the one you want to find out and follow the directions in the triangle

Worked Example
How many hydrogen atoms are in 0.010 moles of CH_3CHO?

Answer:

- There are 4 H atoms in 1 molecule of CH_3CHO
- So, there are 0.040 moles of H atoms in 0.010 moles of CH_3CHO
- The number of H atoms is the **amount in moles x L**
- This comes to 0.040 x (6.02 x 10^{23}) = **2.4 x 10^{22} atoms**

Worked Example
How many moles of hydrogen atoms are in 3.612 x 10^{23} molecules of H_2O_2?

Answer:

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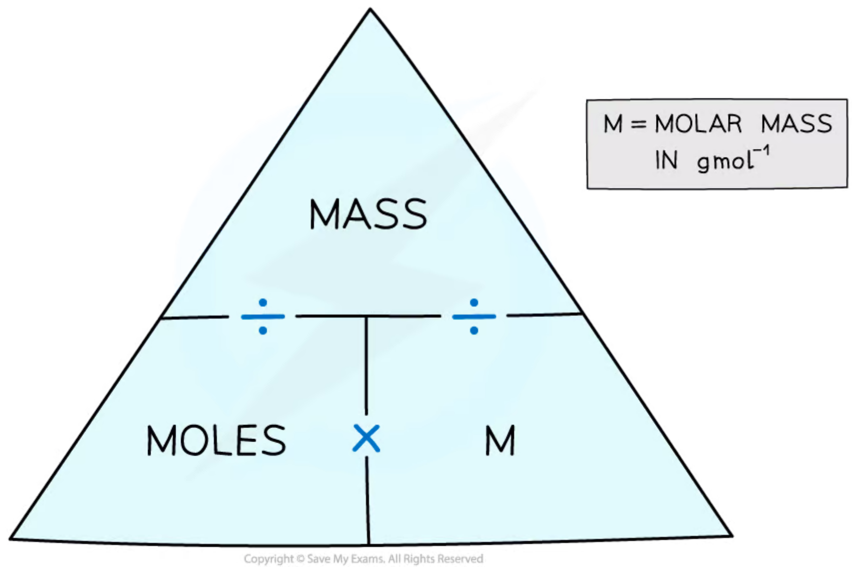
- In 3.612×10^{23} molecules of H_2O_2 there are $2 \times (3.612 \times 10^{23})$ atoms of H
- So, there are 7.224×10^{23} atoms of H
- The number of moles of H atoms is the **number of particles ÷ *L***
- This comes to $7.224 \times 10^{23} \div (6.02 \times 10^{23})$ = **1.20 moles of H atoms**

Moles and Mass

- We count in **moles** by weighing the mass of substances
- The number of **moles** can be calculated by using a formula triangle



The moles and mass formula triangle – cover with your finger the one you want to find out and follow the directions in the triangle

Worked Example

What is the mass of 0.250 moles of zinc?

Answer:

- From the periodic table the relative atomic mass of Zn is 65.38
- So, the molar mass is 65.38 g mol^{-1}
- The mass is calculated by **moles x molar mass**
- This comes to 0.250 mol x 65.38 g mol^{-1} = **16.3 g**

Worked Example

How many moles are in 2.64 g of sucrose, $C_{12}H_{11}O_{22}$ (M_r = 342.3)?

Answer:

- The molar mass of sucrose is 342.3 g mol^{-1}
- The number of moles is found by **mass ÷ molar mass**

- This comes to 2.64 g ÷ 342.3 g mol^{-1} = **7.71 x 10^{-3} mol**

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Exam Tip

Always show your workings in calculations as its easier to check for errors and you may pick up credit if you get the final answer wrong.

1.1.6 Empirical Formulae

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Empirical & Molecular Formulae

- The **molecular formula** is the formula that shows the **number** and **type** of each atom in a molecule
  - E.g. the molecular formula of ethanoic acid is $C_2H_4O_2$
- The **empirical formula** is the simplest whole number ratio of the atoms of each element present in one molecule or formula unit of the compound
  - E.g. the empirical formula of ethanoic acid is CH_2O
- **Organic molecules** often have **different** empirical and molecular formulae
- The formula of an **ionic compound** is always an **empirical formula**

Empirical Formula Calculations

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Empirical formula

- The **empirical formula** is the **simplest whole number ratio** of the atoms of each element present in one molecule or formula unit of the compound
- It is calculated from a knowledge of the masses of each element in a sample of the compound
- It can also be deduced from data that give the **percentage composition by mass** of the elements in a compound

Worked Example

Determine the empirical formula of a compound that contains 10 g of hydrogen and 80 g of oxygen

Answer:

	Hydrogen	Oxygen
Note the mass of each element	10 g	80 g
Divide the masses by atomic masses	$= \frac{10}{1.01}$ = 10 mol	$= \frac{80}{16.00}$ = 5.0 mol
Divide by the lowest figure to obtain nearest whole number ratio	$= \frac{10}{5.0}$ = 2	$= \frac{5.0}{5.0}$ = 1
Empirical formula	H_2O	

Worked Example

Determine the empirical formula of a compound that contains 85.7% carbon and 14.3% hydrogen

Answer:

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	Carbon	Hydrogen
Note the % by mass of each element	85.7	14.3
Divide the % by atomic masses	$= \frac{85.7}{12.01}$ $= 7.14$ mol	$= \frac{14.3}{1.01}$ $= 14.2$ mol
Divide by the lowest figure to obtain nearest whole number ratio	$= \frac{7.14}{7.14}$ $= 1$	$= \frac{14.2}{7.14}$ $= 2$
Empirical formula	CH_2	

Molecular formula

- The **molecular formula** gives the actual numbers of each element present in the formula of the compound
- The molecular formula can be found by dividing the **relative molecular mass** by the **relative mass** of the **empirical formula** and finding the multiple that links the empirical formula to the molecular formula
- **Multiply** the empirical formula by this number to find the molecular formula

Worked Example

The empirical formula of X is $C_4H_{10}S$ and the relative molecular mass of X is 180.42. What is the molecular formula of X? **Relative Atomic Mass** Carbon: 12.01 Hydrogen: 1.01 Sulfur: 32.07

Answer:

Step 1: Calculate the relative mass of empirical formula

Relative empirical mass = (C x 4) + (H x 10) + (S x 1)

Relative empirical mass = (12.01 x 4) + (1.01 x 10) + (32.07 x 1)

Relative formula mass = 90.21

Step 2: Divide relative molecular mass of **X** by relative mass of empirical formula

The multiple between X and the empirical formula = 180.42/90.21 = 2

Step 3: Multiply the empirical formula by 2

2 x $C_4H_{10}S$ = $C_8H_{20}S_2$

The molecular formula of **X** is $C_8H_{20}S_2$

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1.2 Reacting Masses & Volumes

1.2.1 Reacting Masses

Reacting Masses & Limiting Reactants

- The number of moles of a substance can be found by using the following equation:

$$\textit{number of moles} = \frac{mass\ of\ substance\ in\ grams}{molar\ mass\ (g\ mol^{-1})}$$

- It is important to be clear about the type of particle you are referring to when dealing with moles
  - Eg. 1 mole of CaF_2 contains one mole of CaF_2 **formula units**, but one mole of Ca^{2+} and two moles of F^- **ions**

Reacting masses

- The **masses** of reactants are useful to determine how much of the reactants **exactly** react with each other to prevent waste
- To calculate the reacting masses, the chemical equation is required
- This equation shows the ratio of moles of all the reactants and products, also called the **stoichiometry**, of the reaction
- To find the mass of products formed in a reaction the following pieces of information are needed:
  - The mass of the reactants
  - The molar mass of the reactants
  - The balanced equation

Worked Example

Calculate the mass of magnesium oxide that can be made by completely burning 6.0 g of magnesium in oxygen.

magnesium (s) + oxygen (g) → magnesium oxide (s)

Answer:

Step 1: The symbol equation is:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Step 2: The relative atomic masses are:

magnesium : 24.31 oxygen : 16.00

Step 3: Calculate the moles of magnesium used in reaction

$$\textit{number of moles} = \frac{6.0\ g}{24.31\ g\ mol^{-1}} = \underline{\mathbf{0.25\ mol}}$$

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Step 4: Find the ratio of magnesium to magnesium oxide using the balanced chemical equation

	Magnesium	Magnesium Oxide
Mol	2	2
Ratio	1	1
Change in mol	−0.25	+0.25

Therefore, 0.25 mol of MgO is formed

Step 5: Find the mass of magnesium oxide

$$\textbf{mass} = \textbf{\textit{mol}} \times \textbf{\textit{M}}$$

$$\textbf{\textit{mass}} = \textbf{0.25 \textit{mol}} \times \textbf{40.31 \textit{g mol}}^{-1}$$

$$\textbf{\textit{mass}} = \textbf{10.08 g}$$

Therefore, **mass of magnesium oxide produced is 10 g** (2 sig figs)

Excess & limiting reactants

- Sometimes, there is an **excess** of one or more of the reactants (**excess reactant**)
- The reactant which is not in excess is called the **limiting reactant**
- To determine which reactant is limiting:
  - The number of moles of the reactants should be calculated
  - The ratio of the reactants shown in the equation should be taken into account eg:

$$C + 2H_2 \rightarrow CH_4$$

What is limiting when 10 mol of carbon are reacted with 3 mol of hydrogen?

- Hydrogen is the **limiting reactant** and since the ratio of C : H_2 is 1:2 only 1.5 mol of C will react with 3 mol of H_2

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Exam Tip

An easy way to determine the limiting reactant is to find the moles of each substance and divide the moles by the coefficient in the equationThe **lowest** number resulting is the **limiting reactant**

- In the example above:
  - divide 10 moles of C by 1, giving 10
  - divide 3 moles of H by 2, giving 1.5, so hydrogen is limiting

Worked Example

9.2 g of sodium metal is reacted with 8.0 g of sulfur to produce sodium sulfide, Na_2S.Which reactant is in excess and which is limiting?

Answer:

Step 1: Calculate the moles of each reactant

$$\text{number of moles (Na)} = \frac{9.2\ g}{22.99\ g\ mol^{-1}} = 0.40\ \text{mol}$$

$$\text{number of moles (S)} = \frac{8.0\ g}{32.07\ g\ mol^{-1}} = 0.25\ \text{mol}$$

Step 2: Write the balanced equation and determine the coefficients

$$2Na + S \rightarrow Na_2S$$

Step 3: Divide the moles by the coefficient and determine the limiting reagent

- divide 0.40 moles of Na by 2, giving 0.20 – lowest
- divide 0.25 moles of S by 1, giving 0.25

Therefore, **sodium is limiting** and **sulfur is in excess**

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Reaction Yields

Percentage yield

- In a lot of reactions, not all reactants react to form products which can be due to several factors:
  - Other reactions take place simultaneously
  - The reaction does not go to **completion**
  - Products are **lost** during separation and purification

- The **percentage yield** shows how much of a particular product you get from the reactants compared to the maximum theoretical amount that you can get:

$$\textit{percentage yield} = \frac{actual\ yield}{theorectical\ yield} \times 100$$

- The **actual yield** is the number of moles or mass of product obtained **experimentally**
- The **theoretical yield** is the number of moles or mass obtained by a reacting mass calculation

Worked Example

In an experiment to displace copper from copper(II)sulfate, 6.5 g of zinc was added to an excess of copper(II)sulfate solution.The resulting copper was filtered off, washed and dried.The mass of copper obtained was 4.8 g.Calculate the percentage yield of copper.

Answer:

Step 1: The symbol equation is:

$$Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

Step 2: Calculate the amount of zinc reacted in moles

$$\textit{number of moles} = \frac{6.5\ g}{65.38\ g\ mol^{-1}} = 0.10\ \text{mol}$$

Step 3: Calculate the maximum amount of copper that could be formed from the molar ratio:

Since the ratio of Zn(s) to Cu(s) is 1:1 a maximum of 0.10 moles can be produced

Step 4: Calculate the maximum mass of copper that could be formed (**theoretical yield**)

$$\text{mass} = \text{mol} \times M$$

$$= 0.10\ \text{mol} \times 63.55\ \text{g mol}^{-1}$$

= **6.4 g** (2 sig figs)

Step 5: Calculate the percentage yield of copper

$$\textit{percentage yield} = \frac{4.8\ g}{6.4\ g} \times 100 = \underline{\mathbf{75\%}}$$

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1.2.3 Avogadro's Law & Molar Gas Volume

Avogadro's Law

Volumes of gases

- In 1811 the Italian scientist Amedeo **Avogadro** developed a theory about the volume of gases
- **Avogadro's law** (also called **Avogadro's hypothesis)** enables the mole ratio of reacting gases to be determined from volumes of the gases
- **Avogadro** deduced that equal volumes of gases must contain the same number of molecules
- At standard temperature and pressure(**STP**) **one mole** of any gas has a volume of **22.7 dm^3**
- The units are normally written as **dm^3 mol^{-1}**(since it is 'per mole')
- The conditions of **STP** are
  - a temperature of **0°C** (**273 K**)
  - pressure of **100 kPa**

Stoichiometric relationships

- The stoichiometry of a reaction and **Avogadro's Law** can be used to deduce the **exact volumes** of gaseous reactants and products
  - Eg. in the **combustion** of 50 cm^3 of propane, the volume of oxygen needed is (5 x 50) 250 cm^3, and (3 x 50) 150 cm^3 of carbon dioxide is formed, using the ratio of propane: oxygen: carbon dioxide, which is 1: 5: 3 respectively, as seen in the equation

$$C_3H_8\ (g) + 5O_2\ (g) \rightarrow 3CO_2\ (g) + 4H_2O\ (l)$$

- Remember that if the gas volumes are not in the same ratio as the coefficients then the amount of product is determined by the limiting reactant so it is essential to identify it first

Worked Example

What is the total volume of gases remaining when 70 cm^3 of ammonia is combusted completely with 50 cm^3 of oxygen according to the equation shown?

$$4NH_3\ (g) + 5O_2\ (g) \rightarrow 4NO\ (g) + 6H_2O\ (l)$$

Answer:

Step 1: From the equation deduce the molar ratio of the gases, which is NH_3:O_2:NO or 4:5:4 (water is not included as it is in the liquid state)

Step 2: We can see that oxygen will run out first (the **limiting reactant**) and so 50 cm^3 of O_2 requires 4/5 x 50 cm^3 of NH_3 to react = 40 cm^3

Step 3: Using Avogadro's Law, we can say 40 cm^3 of NO will be produced

Step 4: There will be of 70–40 = 30 cm^3 of NH_3 left over

Therefore **the total remaining volume will be 40 + 30 = 70 cm^3 of gases**

YOUR NOTES

Exam Tip

Since gas volumes work in the same way as moles, we can use the '**lowest is limiting**' technique in limiting reactant problems involving gas volumes. This can be handy if you are unable to spot which gas reactant is going to run out first.Divide the volumes of the gases by the cofficients and whichever gives the lowest number is the **limiting reactant**

- E.g. in the previous problem we can see that
  - For NH_3 70/4 gives 17.5
  - For O_2 50/5 gives 10, so **oxygen is limiting**

YOUR NOTES
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Molar Gas Volume

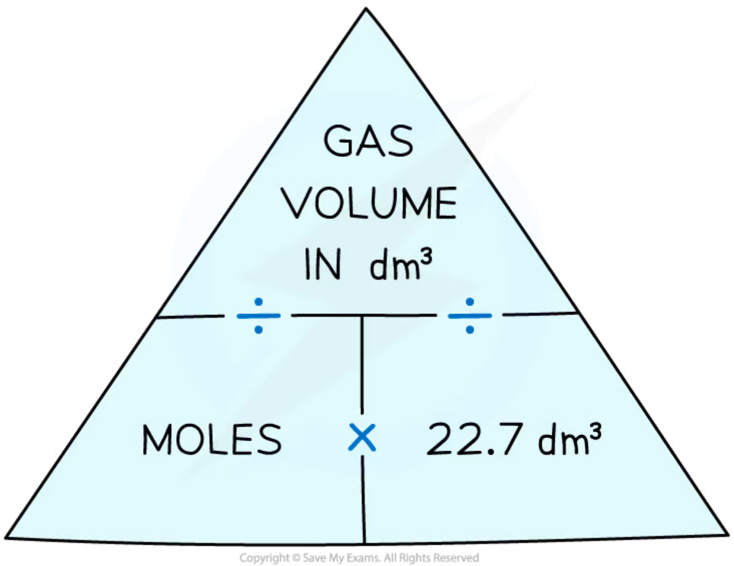
- The **molar gas volume** of 22.7 dm^3 mol^{-1} can be used to find:
  - The volume of a given number of moles of gas:

volume of gas (dm^3) = amount of gas (mol) x 22.7 dm^3 mol^{-1}

  - The number of moles of a given volume of gas:

$$\textit{amount of gas (moles)} = \frac{volume\ of\ gas\ in\ dm^3}{22.7\ dm^3\ mol^{-1}}$$

- The relationships can be expressed using a formula triangle



To use the gas formula triangle cover the one you want to find out about with your finger and follow the instructions

Worked Example

What is the volume occupied by 3.0 moles of hydrogen at stp?

Answer:

volume of gas (dm^3) = amount of gas (mol) x 22.7 dm^3 mol^{-1}

3.0 mol x 22.7 dm^3 mol^{-1} = **68 dm^3**

Worked Example

How many moles are in the following volumes of gases?

1. 7.2 dm^3 of carbon monoxide
2. 960 cm^3 of sulfur dioxide

YOUR NOTES

Answer 1:

Use the formula:

$$\text{amount of gas (moles)} = \frac{volume\ of\ gas\ in\ dm^3}{22.7\ dm^3\ mol^{-1}}$$

$$\text{amount of gas (moles)} = \frac{7.2\ dm^3}{22.7\ dm^3\ mol^{-1}} = \mathbf{0.32\ mol}$$

Answer 2:

Step 1: Convert the volume from cm^3 to dm^3

960/1000 = 0.960 dm^3

Step 2: Use the formula

$$\text{amount of gas (moles)} = \frac{0.960\ dm^3}{22.7\ dm^3\ mol^{-1}} = \mathbf{4.22 \times 10^{-2}\ mol}$$

1.2.4 The Ideal Gas Equation

YOUR NOTES

Ideal Gas Equation

Kinetic theory of gases

- **The kinetic theory of gases** states that molecules in gases are constantly moving
- The theory makes the following assumptions:
  - The gas molecules are moving very fast and randomly
  - The molecules hardly have any volume
  - The gas molecules do not attract or repel each other (**no intermolecular forces**)
  - No kinetic energy is lost when the gas molecules collide with each other (**elastic collisions**)
  - The temperature of the gas is directly proportional to the average kinetic energy of the molecules
- Gases that follow the kinetic theory of gases are called **ideal gases**
- However, in reality gases do not fit this description exactly **but** may come very close and are called **real gases**
- The volume that a gas occupies depends on:
  - Its pressure
  - Its temperature

Ideal gas equation

- The **ideal gas equation** shows the relationship between pressure, volume, temperature and number of moles of gas of an ideal gas:

$$PV = nRT$$

P = pressure (pascals, Pa)

V = volume (m^3)

n = number of moles of gas (mol)

R = gas constant (8.31 J K^{-1} mol^{-1})

T = temperature (Kelvin, K)

- The ideal gas equation can also be used to calculate the **molar mass** (*M*) of a gas

Worked Example

Calculate the volume, in dm^3, occupied by 0.781 mol of oxygen at a pressure of 220 kPa and a temperature of 21 °C.

Answer:

Step 1: Rearrange the ideal gas equation to find volume of the gas

$$V = \frac{nRT}{P}$$

Step 2: Convert into the correct units and calculate the volume the oxygen gas occupies

p = 220 kPa = 220 000 Pa

n = 0.781 mol

R = 8.31 J K^{-1} mol^{-1}

T = 21 °C = 294 K

$$V = \frac{0.781\ mol \times 8.31\ J\ K^{-1} mol^{-1} \times 294\ K}{220\ 000\ Pa}$$

= 0.00867 m^3

= **8.67 dm^3**

YOUR NOTES

Exam Tip

A word about units... Students often mess up gas calculations by getting their unit conversions wrong, particularly from cm^3 to m^3. Think about what a cubic metre actually is - a cube with sides 1 m or 100 cm long. The volume of this cube is 100 x 100 x 100 = 1 000 000 or **10^6** cm^3 So when you convert from **m^3 to cm^3** you **MULTIPLY by 10^6** and when you convert from **cm^3 to m^3** you **DIVIDE by 10^6** (or multiply by 10^{-6} which is the same thing)

Worked Example

Calculate the pressure of a gas, in kPa, given that 0.20 moles of the gas occupy 10.1 dm^3 at a temperature of 25 °C.

Answer:

Step 1: Rearrange the ideal gas equation to find the pressure of the gas

$$P = \frac{nRT}{V}$$

Step 2: Convert to the correct units and calculate the pressure

n = 0.20 mol

V = 10.1 dm^3 = 0.0101 m^3 = 10.1 x 10^{-3} m^3

R = 8.31 J K^{-1} mol^{-1}

T = 25 °C = 298 K

$$P = \frac{0.20\ mol \times 8.31\ J\ K^{-1} mol^{-1} \times 298\ K}{10.1 \times 10^{-3}\ m^3}$$

P = 49 037 Pa = **49 kPa** (2 sig figs)

YOUR NOTES

Worked Example

Calculate the temperature of a gas, in °C, if 0.047 moles of the gas occupy 1.2 dm^3 at a pressure of 100 kPa.

Answer:

Step 1: Rearrange the ideal gas equation to find the temperature of the gas

$$T = \frac{PV}{nR}$$

Step 2: Convert to the correct units and calculate the pressure

n = 0.047 mol

V = 1.2 dm^3 = 0.0012 m^3 = 1.2 x 10^{-3} m^3

R = 8.31 J K^{-1} mol^{-1}

P = 100 kPa = 100 000 Pa

$$T = \frac{100\ 000 \times 1.2 \times 10^{-3}\ m^3}{0.047\ mol \times 8.31\ J\ K^{-1} mol^{-1}}$$

T = 307.24 K = 34.24 °C = **34 °C** (2 sig figs)

Worked Example

A flask of volume 1000 cm^3 contains 6.39 g of a gas. The pressure in the flask was 300 kPa and the temperature was 23 °C. Calculate the molar mass of the gas.

Answer:

Step 1: Rearrange the ideal gas equation to find the number of moles of gas

$$n = \frac{pV}{RT}$$

Step 2: Convert to the correct units and calculate the number of moles of gas

P = 300 kPa = 300 000 Pa

V = 1000 cm^3 = 0.001 m^3 = 1.0 x 10^{-3} m^3

YOUR NOTES
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R = 8.31 J K^{-1} mol^{-1}

T = 23 °C = 296 K

$$n = \frac{300\ 000\ Pa \times 1 \times 10^{-3}\ m^3}{8.31\ J\ K^{-1} mol^{-1} \times 296\ K}$$

n = 0.12 mol

Step 3: Calculate the molar mass using the number of moles of gas

$$\text{molar mass} = \frac{mass}{moles}$$

$$M = \frac{6.39\ g}{0.12\ mol} = \mathbf{53\ g\ mol^{-1}}\ \text{(2 sig figs)}$$

Exam Tip

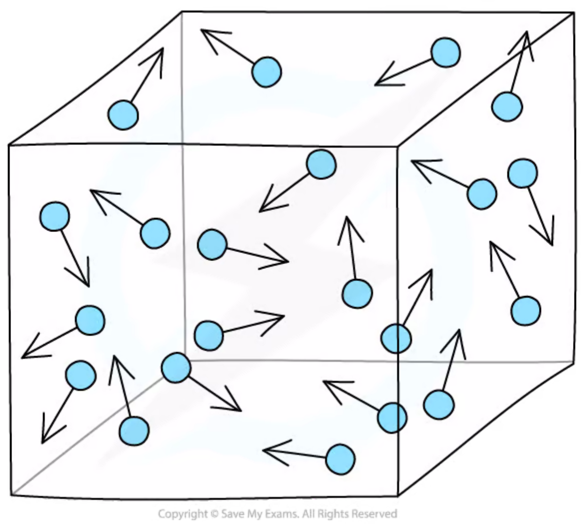
To calculate the temperature in **Kelvin**, add 273 to the Celsius temperature, eg. 100 °C is 373 Kelvin.

1.2.5 Gas Law Relationships

YOUR NOTES

Gas Law Relationships

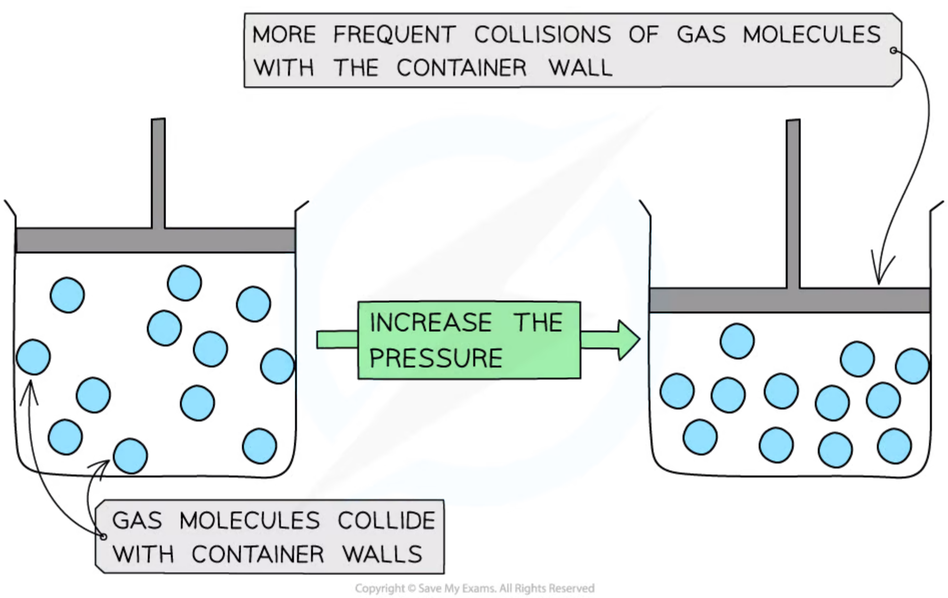
- **Gases** in a container exert a **pressure** as the gas molecules are constantly **colliding** with the walls of the container



Gas particles exert a pressure by constantly colliding with the walls of the container

Changing gas volume

- **Decreasing** the **volume** (at constant temperature) of the container causes the molecules to be **squashed** together which results in more **frequent** collisions with the container wall
- The **pressure** of the gas **increases**



Decreasing the volume of a gas causes an increased collision frequency of the gas particles with the container wall

YOUR NOTES
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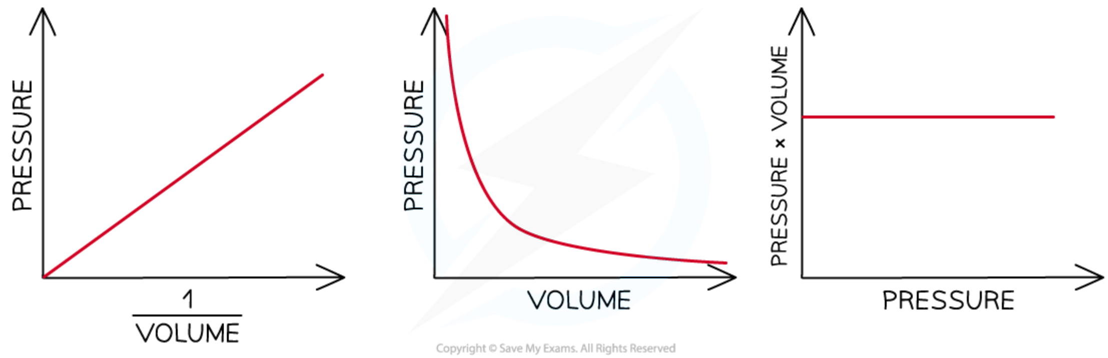
- The **pressure** is therefore **inversely proportional** to the **volume** (at constant temperature)
- This is known as **Boyle's Law**
- Mathematically, we say $P \propto 1/V$ or **PV = a constant**
- We can show a graphical representation of **Boyle's Law** in three different ways:
  - A graph of pressure of gas plotted against 1/ volume gives a straight line
  - A graph of pressure against volume gives a curve
  - A graph of PV versus P gives a straight line



Three graphs that show Boyle's Law

Changing gas temperature

- When a gas is **heated** (at constant pressure) the particles gain more **kinetic energy** and undergo more **frequent collisions** with the container walls
- To keep the **pressure constant**, the molecules must get further apart and therefore the **volume increases**
- The **volume** is therefore **directly proportional** to the **temperature in Kelvin** (at constant pressure)
- This is known as **Charles' Law**
- Mathematically, $V \propto T$ or **V/T = a constant**
- A graph of **volume** against **temperature in Kelvin** gives a straight line

YOUR NOTES
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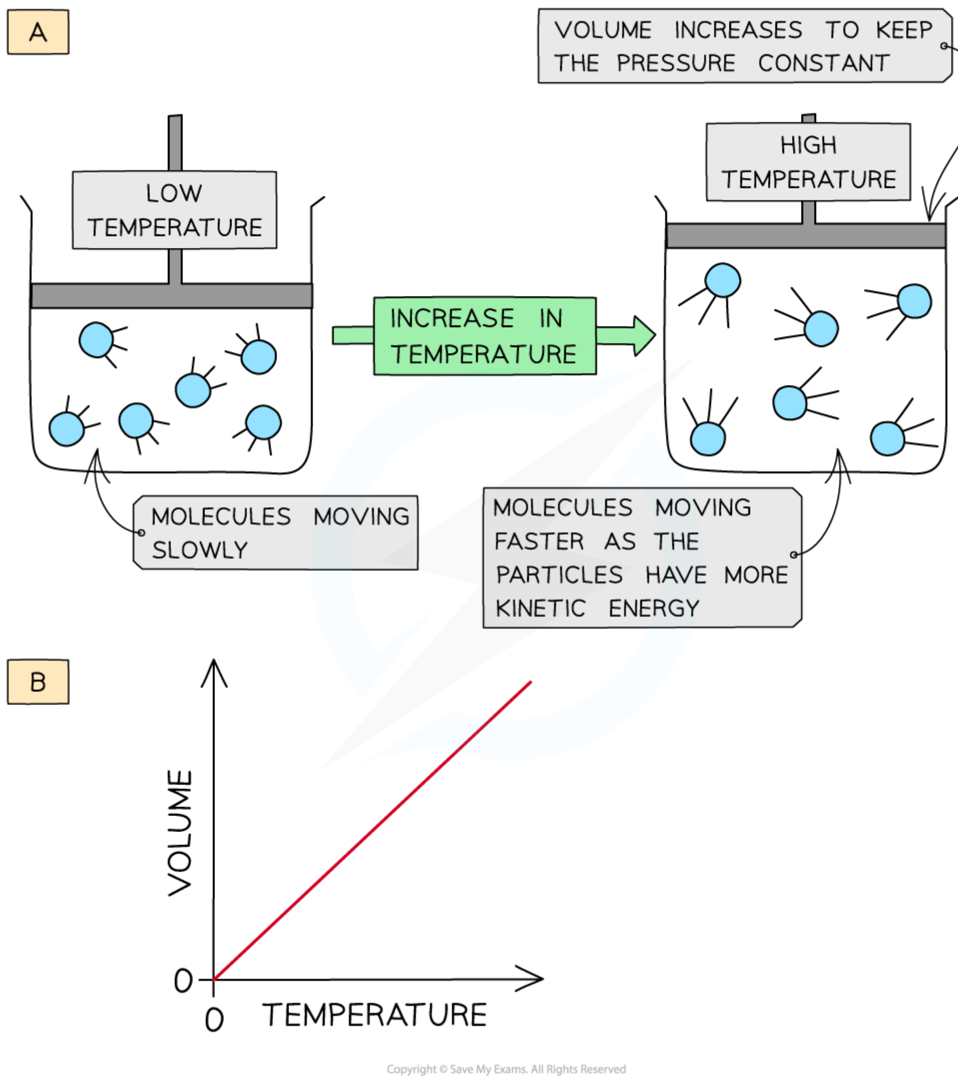

Increasing the temperature of a gas causes an increased collision frequency of the gas particles with the container wall (a); volume is directly proportional to the temperature in Kelvin (b)

Changing gas pressure

- **Increasing** the **temperature** (at constant volume) of the gas causes the molecules to gain more **kinetic energy**
- This means that the particles will move **faster** and **collide** with the container walls more **frequently**
- The **pressure** of the gas **increases**
- The **temperature** is therefore **directly proportional** to the **pressure** (at constant volume)
- Mathematically, we say that $P \propto T$ or **P/T = a constant**
- A graph of **temperature in Kelvin** of a gas plotted against **pressure** gives a straight line

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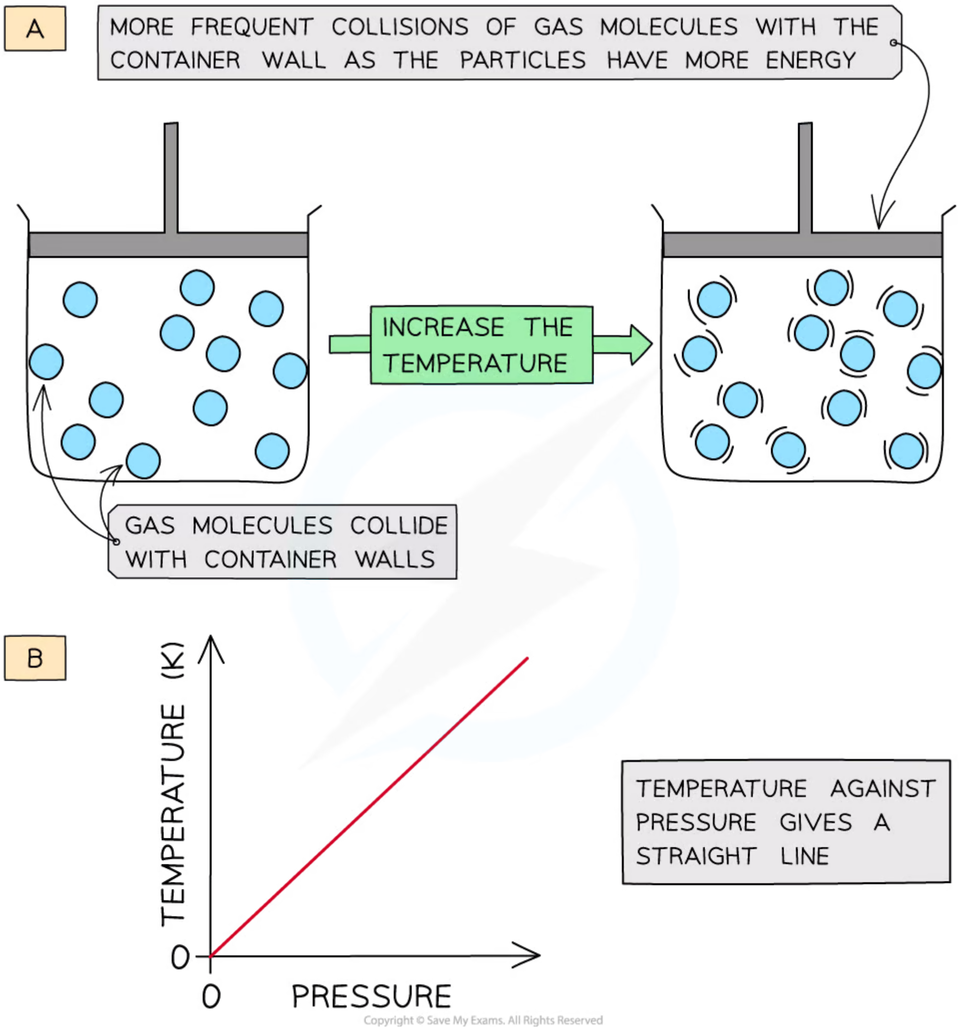

Increasing the temperature of a gas causes an increased collision frequency of the gas particles with the container wall (a); temperature is directly proportional to the pressure (b)

Pressure, volume and temperature

- Combining these three relationships together:
  - P/V = a constant
  - V/T = a constant
  - P/T = a constant
- We can see how the **ideal gas equation** is constructed
  - PV/T = a constant
  - PV = a constant x T
- This constant is made from two components, the number of **moles**, **n**, and the **gas constant**, **R**, resulting in the overall equation:
  - **PV = nRT**

Changing the conditions of a fixed amount of gas

YOUR NOTES
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- For a fixed amount of gas, **n** and **R** will be constant, so if you change the conditions of a gas we can ignore **n** and **R** in the **ideal gas equation**
- This leads to a very useful expression for problem solving

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- Where P_1, V_1 and T_1 are the initial conditions of the gas and P_2, V_2 and T_2 are the final conditions

Worked Example

At 25 °C and 100 kPa a gas occupies a volume of 20 dm^3. Calculate the new temperature, in °C, of the gas if the volume is decreased to 10 dm^3 at constant pressure.

Answer:

Step 1: Rearrange the formula to change the conditions of a fixed amount of gas. Pressure is constant so it is left out of the formula

$$T_2 = \frac{V_2T_1}{V_1}$$

Step 2: Convert the temperature to Kelvin. There is no need to convert the volume to m^3 because the formula is using a **ratio** of the two volumes

$V_1 = 20$ dm^3

$V_2 = 10$ dm^3

$T_1 = 25 + 273 = 298$ K

Step 3: Calculate the new temperature

$$T_2 = \frac{10\ dm^3 \times 298\ K}{20\ dm^3} = 149\ \text{K} = \mathbf{-124\ °C}$$

Worked Example

A 2.00 dm^3 container of oxygen at a pressure of 80 kPa was heated from 20 °C to 70 °C The volume expanded to 2.25 dm^3. What was the final pressure of the gas?

Answer:

Step 1: Rearrange the formula to change the conditions of a fixed amount of gas

$$P_2 = \frac{P_1V_1T_2}{V_2T_1}$$

YOUR NOTES
⬇

Step 2: Substitute in the values and calculate the final pressure

$P_1 = 80$ kPa

$V_1 = 2.00$ dm^3

$V_2 = 2.25$ dm^3

$T_1 = 20 + 273 = 293$ K

$T_2 = 70 + 273 = 343$ K

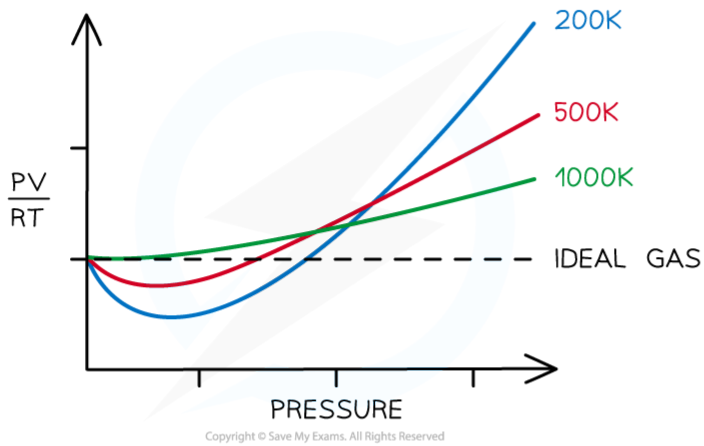
$$P_2 = \frac{80\ kPa \times 2.00\ dm^3 \times 343\ K}{293\ K \times 2.25\ dm^3} = \mathbf{83\ kPa}$$

1.2.6 Real Gases

YOUR NOTES

Real Gas Behaviour

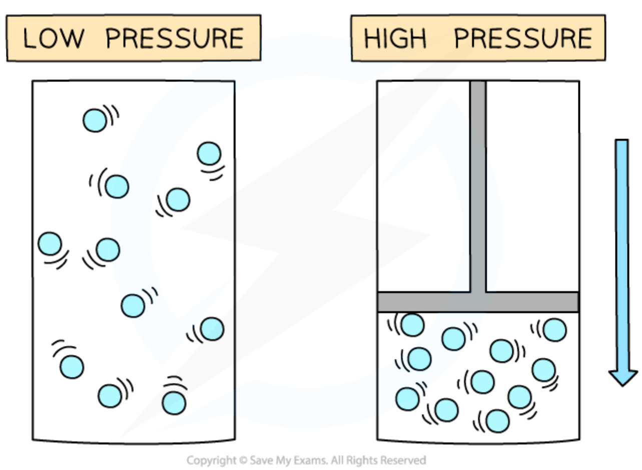
- The **ideal gas equation** does not fit all measurements and observations taken at all conditions with real gases
- The relationship between pressure, volume and temperature shows significant deviation from ***PV = nRT*** when the **temperature is very low** or the **pressure is very high**
- This is because the **ideal gas equation** is built on the **kinetic theory of matter**
- The **kinetic theory of matter** makes some key assumptions about the behaviour of gases



At low temperatures and high pressures real gases deviate significantly from the ideal gas equation. The higher the pressure and the lower the temperature the greater the deviation

Assumptions about volume

- The **kinetic theory** assumes that the volume the actual gas molecules themselves take up is tiny compared to the volume the gas occupies in a container and can be ignored
- This is broadly true for gases at normal conditions, but becomes increasingly inaccurate at low temperatures and high pressures
- At these conditions the gas molecules are very close together, so the **fraction of space** taken up by the molecules is **substantial** compared to the total volume



At low temperatures and high pressures, the fraction of space taken up by the molecules is substantial

YOUR NOTES
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Assumptions about attractive forces

- Another assumption about gases is that when gas molecules are far apart there is very little interaction between the molecules
- As the gas molecules become closer to each other **intermolecular forces** cause **attraction** between molecules
- This reduces the number of collisions with the walls of the container
- The pressure is less than expected by the **ideal gas equation**

Exam Tip

The ideal gas equation and the gas constant are given in the IB Chemistry Data Booklet which can be used in Paper 2, but not in Paper 1.

1.2.7 Standard Solutions

YOUR NOTES

Concentrations of Solutions

Standard solutions

- Chemists routinely prepare solutions needed for analysis, whose concentrations are known precisely
- These solutions are termed **standard solutions**
- They are made as accurately and precisely as possible using three decimal place balances and volumetric flasks to reduce the impact of measurement uncertainties
- The steps are:

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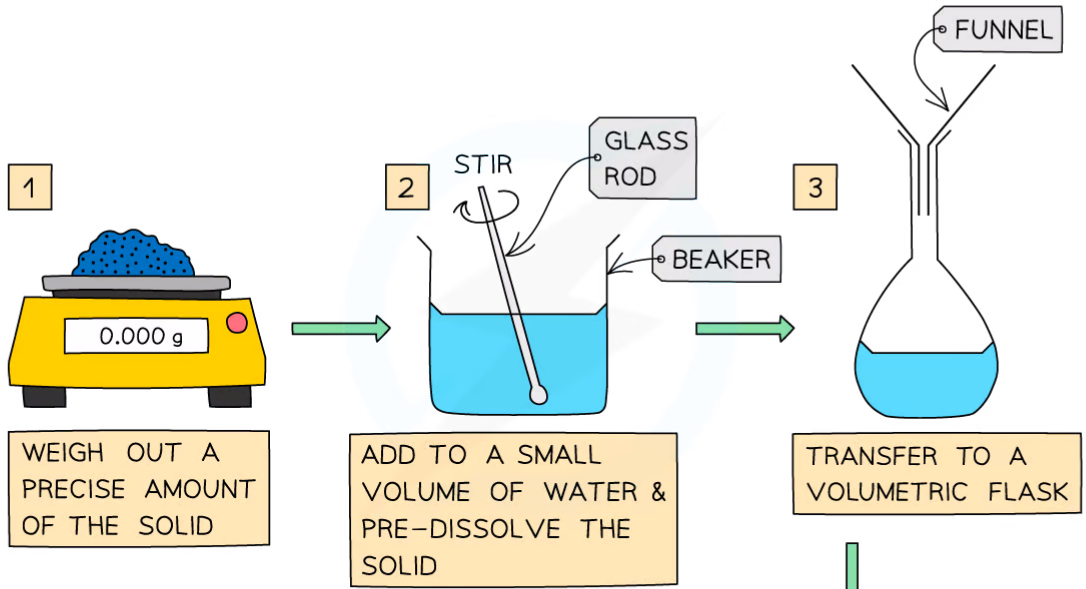

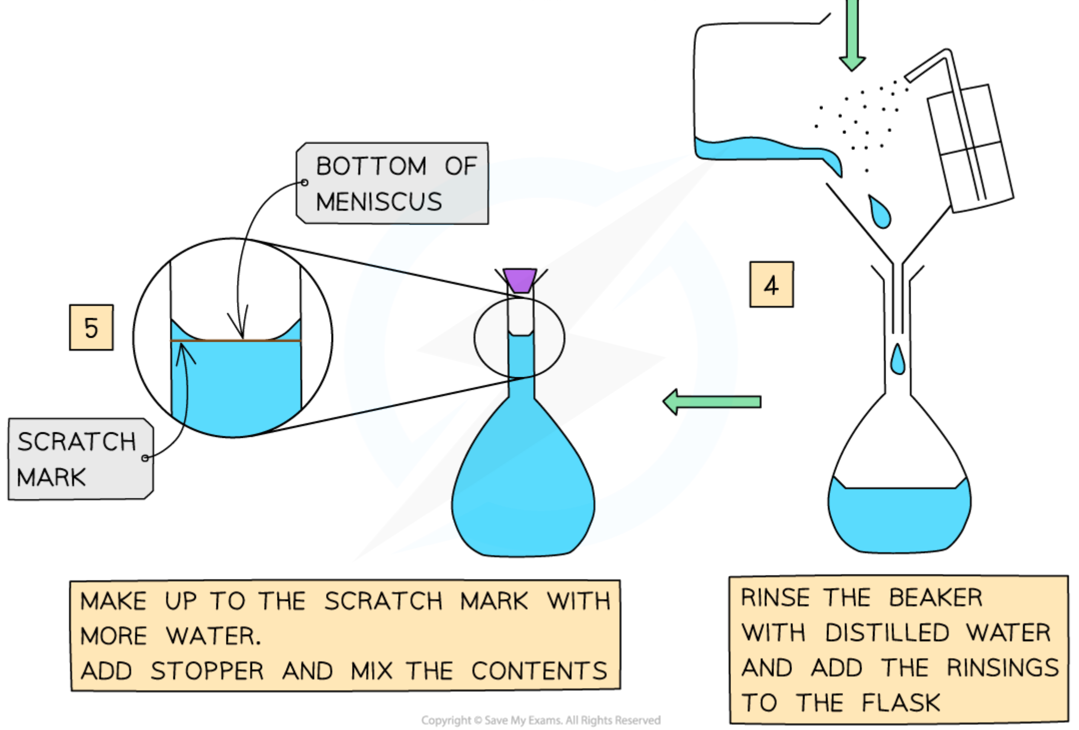

Volumes & concentrations of solutions

- The **concentration** of a solution is the amount of **solute** dissolved in a **solvent** to make 1 dm^3 of **solution**
  - The solute is the substance that dissolves in a solvent to form a solution
  - The solvent is often water

- A **concentrated** solution is a solution that has a **high** concentration of solute
- A **dilute** solution is a solution with a **low** concentration of solute
- Concentration is usually expressed in one of three ways:

YOUR NOTES
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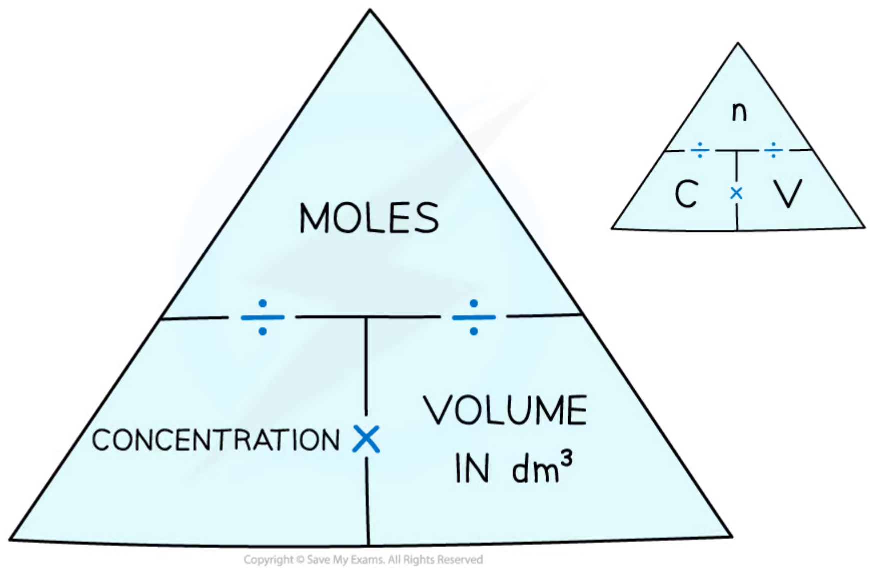
- moles per unit volume
- mass per unit volume
- parts per million

Moles per unit volume

- The formula for expressing concentration using moles is:

$$\textit{concentration} (mol\ dm^{-3}) = \frac{number\ of\ moles\ of\ solute\ (mol)}{volume\ of\ solution\ (dm^3)}$$

- You must make sure you change cm^3 to dm^3 (by dividing by 1000)
- The relationships can be expressed using this formula triangle:



To use the concentration formula triangle cover the one you want to find out about with your finger and follow the instructions

Worked Example

Calculate the mass of sodium hydroxide, NaOH, required to prepare 250 cm^3 of a 0.200 mol dm^{-3} solution

Answer:

Step 1: Use the formula triangle to find the number of moles of NaOH needed

number of moles = concentration (mol dm^{-3}) x volume (dm^3)

n = 0.200 *mol dm^{-3}* x 0.250 *dm^3*

n = **0.0500 mol**

Step 2: Find the molar mass of NaOH

M = 22.99 + 16.00 + 1.01 = 40.00 g mol^{-1}

YOUR NOTES
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Step 3: Calculate the mass required

mass = moles x molar mass

mass = 0.0500 mol x 40.00 g mol^{-1} = **2.00 g**

Mass per unit volume

- Sometimes it is more convenient to express concentration in terms of mass per unit volume
- The formula is:

$$\textit{concentration}\ (g\ dm^{-3}) = \frac{mass\ of\ solute\ (g)}{volume\ of\ solution\ (dm^3)}$$

- To change a concentration from mol dm^{-3} to g dm^{-3}
  - Multiply the moles of solute by its molar mass

mass of solute (g) = number of moles (mol) x molar mass (g mol^{-1})

Parts per million

- When expressing extremely low concentrations a unit that can be used is **parts per million** or **ppm**
- This is useful when giving the concentration of a pollutant in water or the air when the absolute amount is tiny compared the the volume of water or air
- **1 ppm** is defined as
  - A mass of **1 mg** dissolved in **1 dm^3** of water
- Since 1 dm^3 weighs 1 kg we can also say it is
  - A mass of **1 mg** dissolved in **1 kg** of water, or 10^{-3} g in 10^3 g which is the same as saying the concentration is **1 in 10^6** or **1 in a million**

Worked Example

The concentration of chlorine in a swimming pool should between between 1 and 3 ppm. Calculate the maximum mass, in kg, of chlorine that should be present in an olympic swimming pool of size 2.5 million litres.

Answer:

Step 1: calculate the total mass in mg assuming 3ppm(1 litre is the same as 1 dm^3)

3 x 2.5 x 10^6 = 7.5 x 10^6 mg

Step 2: convert the mass into kilograms (1 mg = 10^{-6} kg)

7.5 x 10^6 x 10^{-6} kg = **7.5 kg**

1.2.8 Concentration Calculations

YOUR NOTES
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Concentration Calculations

Step by step

- Concentration calculations involve bringing together the skills and knowledge you have acquired previously and applying them to problem solving
- You should be able to easily convert between moles, mass, concentrations and volumes (of solutions and gases)
- The four steps involved in problem solving are:
  - write the balanced equation for the reaction
  - determine the mass/ moles/ concentration/ volume of the of the substance(s) you know about
  - use the balanced equation to deduce the mole ratios of the substances present
  - calculate the mass/ moles/ concentration/ volume of the of the unknown substance(s)

Worked Example

25.0 cm^3 of 0.050 mol dm^{-3} sodium carbonate was completely neutralised by 20.0 cm^3 of dilute hydrochloric acid.Calculate the concentration in mol dm^{-3} of the hydrochloric acid.

Answer:

Step 1: Write the balanced equation for the reaction

$$\mathbf{Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2}$$

Step 2: Determine the moles of the known substance, in this case sodium carbonate. Don't forget to divide the volume by 1000 to convert cm^3 to dm^3

moles = volume x concentration

amount (Na_2CO_3) = 0.0250 dm^3 x 0.050 mol dm^{-3} = 0.00125 mol

Step 3: Use the balanced equation to deduce the mole ratio of sodium carbonate to hydrochloric acid:

1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the mole ratio is 1 : 2

Therefore 0.00125 moles of Na_2CO_3 react with 0.00250 moles of HCl

Step 4: Calculate the concentration of the unknown substance, hydrochloric acid

$$\text{concentration} = \frac{moles}{volume}$$

$$\text{concentration(HCl)} = \frac{0.00250\ mol}{0.0200\ dm^3} = \mathbf{0.125\ mol\ dm^{-3}}$$

YOUR NOTES

Worked Example

Calculate the volume of hydrochloric acid of concentration 1.0 mol dm^{-3} that is required to react completely with 2.5 g of calcium carbonate.

Answer:

Step 1: Write the balanced equation for the reaction

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

Step 2: Determine the moles of the known substance, calcium carbonate

$$\text{amount of } CaCO_3 = \frac{2.5\ g}{100.09\ g\ mol^{-1}} = 0.025 \text{ mol}$$

Step 3: Use the balanced equation to deduce the mole ratio of calcium carbonate to hydrochloric acid:

1 mol of $CaCO_3$ requires 2 mol of HCl

So 0.025 mol of $CaCO_3$ requires 0.050 mol of HCl

Step 4: Calculate the volume of HCl required

$$\text{Volume of HCl} = \frac{moles}{concentration} = \frac{0.050\ mol}{1.0\ mol\ dm^{-3}} = \mathbf{0.050\ dm^3}$$

YOUR NOTES
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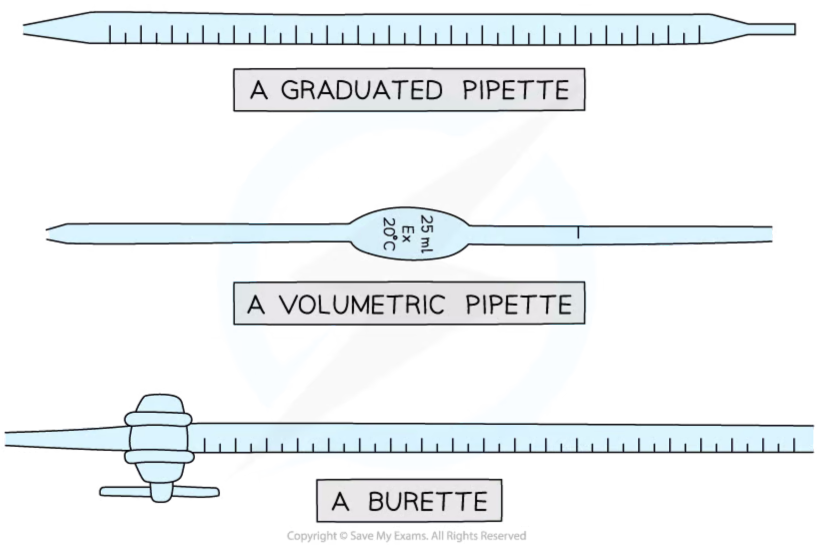
Titrations

- **Volumetric analysis** is a process that uses the volume and concentration of one chemical reactant (**a standard solution**) to determine the concentration of another unknown solution
- The technique most commonly used is a **titration**
- The volumes are measured using two precise pieces of equipment, a **volumetric** or **graduated pipette** and a **burette**



Equipment used to measure volumes precisely in titrations

- **Burettes** are usually marked to a precision of 0.10 cm^3
  - Since they are analogue instruments, the uncertainty is recorded to **half** the smallest marking, in other words to ±0.05 cm^3
- The **end point** or **equivalence point** occurs when the two solutions have reacted completely and is shown with the use of an **indicator**

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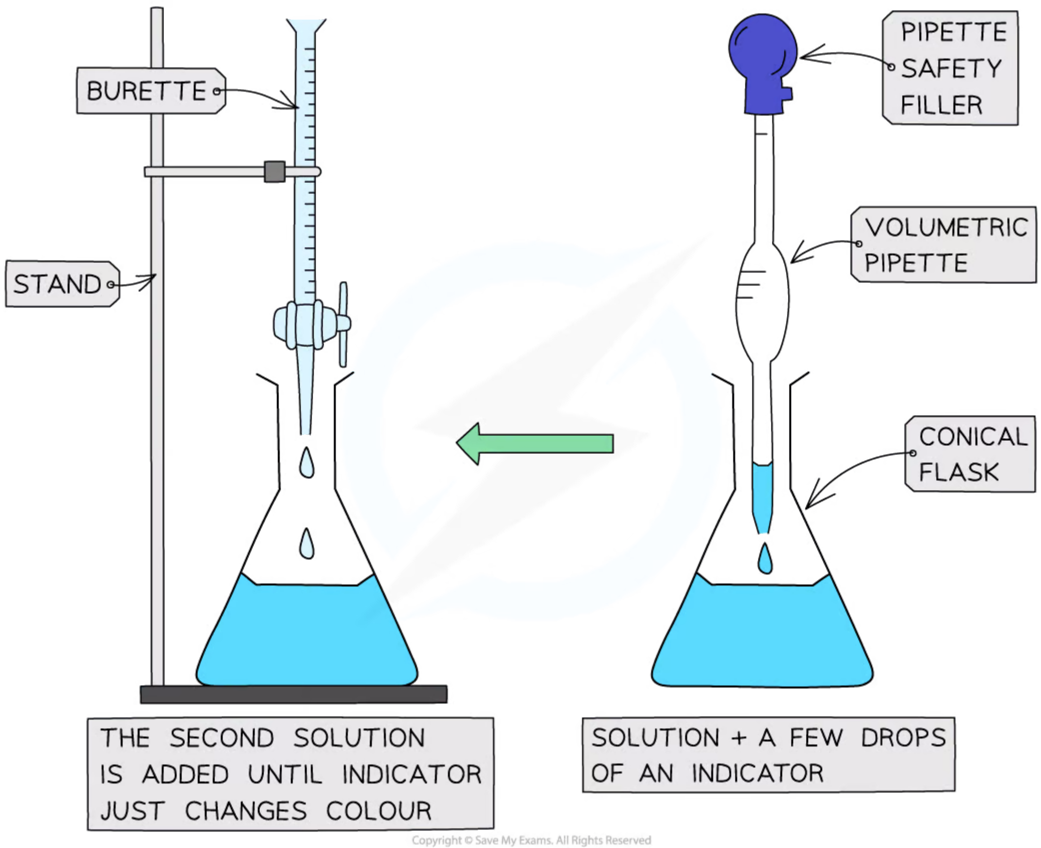

The steps in a titration

- The steps in a titration are:
  - Measuring a known volume (usually 20 or 25 cm^3) of one of the solutions with a **volumetric** or **graduated pipette** and placing it into a **conical flask**
  - The other solution is placed in the **burette**
  - A few drops of the **indicator** are added
  - The tap on the **burette** is carefully opened and the solution added, portion by portion, to the **conical flask** until the **indicator** just changes colour
  - Multiple trials are carried out until **concordant** results are obtained

Recording and processing titration results

- Both the initial and final **burette** readings should be recorded and shown to a **precision** of $\pm 0.05\ cm^3$, the same as the **uncertainty**

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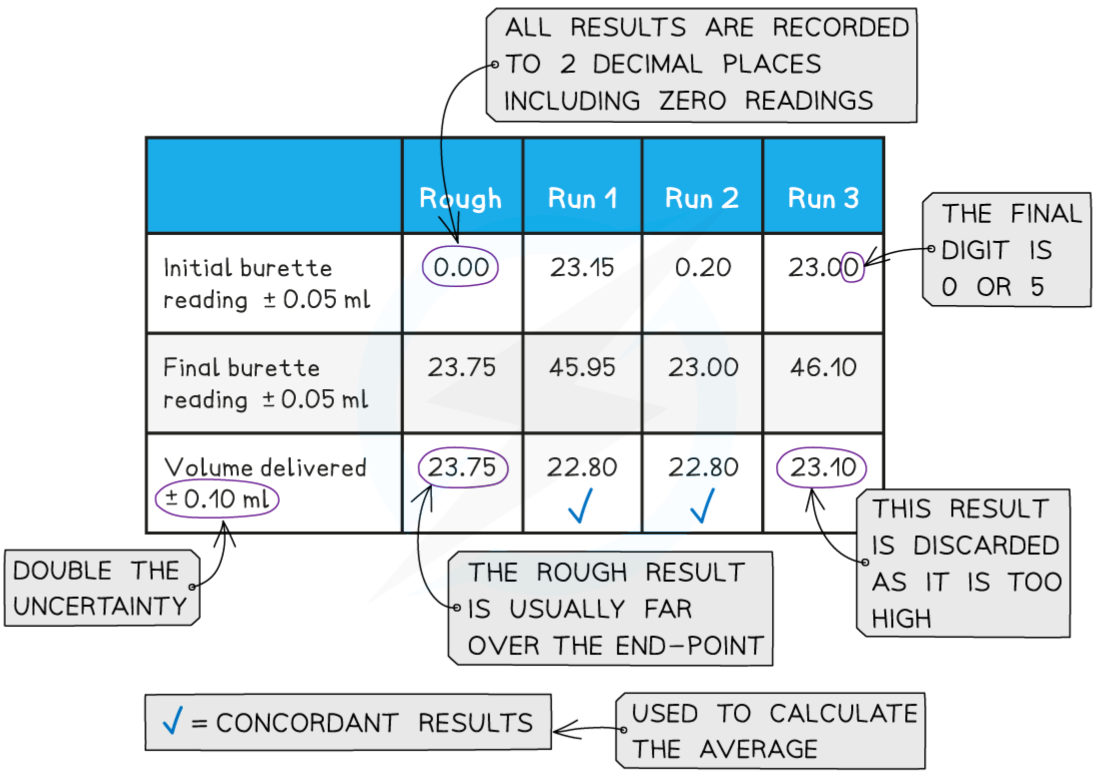

A typical layout and set of titration results

- The volume delivered (**titre**) is calculated and recorded to an **uncertainty** of ±0.10 cm^3
  - The **uncertainty** is doubled, because two **burette** readings are made to obtain the **titre** (V final – V initial), following the rules for **propagation of uncertainties** (you can find more about this in Topic 11)
- **Concordant** results are then averaged, and non-concordant results are discarded
  - Concordance is usually considered to be a consistency of ±0.05 between results, depending on the quality of the **burette**
- The calculation then follows the steps given in 1.2.8 Concentration calculations

Exam Tip

When performing titration calculations using **monoprotic** acids (meaning one H^+) such as HCl, the number of moles of the acid and alkali will be the same. This allows you to use the relationship

$$C_1V_1 = C_2V_2$$

where C_1 and V_1 are the concentration and volume of the acid and C_2 and V_2 are the concentration and volume of the alkali. There is no need to convert the units of volume to dm^3 as this is a ratio. Simply re-arrange the formula to solve for the unknown quantity.

YOUR NOTES
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Worked Example

A 0.675 g sample of a solid acid, HA, was dissolved in distilled water and made up to 100.0 cm^3 in a volumetric flask. 25.0 cm^3 of this solution was titrated against 0.100 mol dm^{-3} NaOH solution and 12.1 cm^3 were required for complete reaction. Determine the molar mass of the acid.

Answer:

Step 1: Write the equation for the reaction:

$$HA\,(aq) + NaOH\,(aq) \rightarrow NaA\,(aq) + H_2O\,(l)$$

Step 2: Calculate the number of moles of the NaOH

$$n(\mathrm{NaOH})_{\mathrm{sample}} = \left(\frac{12.1}{1000}\right)\ \mathrm{dm^3 \times 0.100\ mol\ dm^{-3} = 1.21 \times 10^{-3}\ mol}$$

Step 3: Deduce the number of moles of the acid

Since the acid is monoprotic the number of moles of HA is also 1.21×10^{-3} mol

This is present in 25.0 cm^3 of the solution

Step 4: Scale up to find the amount in the original solution

$$n(\mathrm{NaOH})_{\mathrm{original}} = \frac{1.21 \times 10^{-3}\ \mathrm{mol} \times 100.0\ \mathrm{cm^3}}{25.0\ \mathrm{cm^3}} = 4.84 \times 10^{-3}\ \mathrm{mol}$$

Step 5: Calculate the molar mass

$$\text{moles} = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{molar mass} = \frac{\text{mass}}{\text{moles}} = \frac{0.675\ \mathrm{g}}{4.84 \times 10^{-3}\ \mathrm{mol}} = \mathbf{139\ g\ mol^{-1}}$$

Back titration

- A **back titration** is a common technique used to find the **concentration** or amount of an unknown substance indirectly
- The principle is to carry out a reaction with the unknown substance and an **excess** of a further reactant such as an acid or an alkali
- The **excess** reactant, after reaction, is then analysed by **titration** and the **mole ratios** are used to deduce the **moles** or **concentration** of the original substance being analysed

YOUR NOTES
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Worked Example

The percentage by mass of calcium carbonate, $CaCO_3$, in a sample of marble was determined by adding excess hydrochloric acid to ensure that all the calcium carbonate had reacted. The excess acid left was then titrated with aqueous sodium hydroxide. A student added 27.20 cm^3 of 0.200 mol dm^{-3} HCl to 0.188 g of marble. The excess acid required 23.80 cm^3 of 0.100 mol dm^{-3} NaOH for neutralization. Calculate the percentage of calcium carbonate in the marble.

Answer:

Step 1: Write the equation for the titration reaction:

$HCl (aq) + NaOH (aq) \rightarrow NaCl (aq) + H_2O (l)$

Step 2: Calculate the number of moles of the NaOH

$n(NaOH) = 0.02380\ dm^3 \times 0.100\ mol\ dm^{-3} = 2.380 \times 10^{-3}\ mol$

Step 3: Deduce the number of moles of the excess acid

Since the reacting ratio is 1:1 the number of moles of HCl is also 2.380×10^{-3} mol

Step 4: Find the amount of HCl in the original solution and then the amount reacted

$n(HCl)_{original} = 0.02720\ dm^3 \times 0.200\ mol\ dm^{-3} = 5.440 \times 10^{-3}\ mol$

$n(HCl)_{reacted} = 5.440 \times 10^{-3}\ mol - 2.380 \times 10^{-3}\ mol = 3.060 \times 10^{-3}\ mol$

Step 5: Write the equation for the reaction with the calcium carbonate

$2HCl (aq) + CaCO_3 (s) \rightarrow CaCl_2 (aq) + CO_2 (g) + H_2O (l)$

Step 6: Deduce the number of moles of the calcium carbonate that reacted

Since the reacting ratio is 2:1 the number of moles of $CaCO_3$ is $(3.060 \times 10^{-3}\ mol) \div 2$

$n(CaCO_3) = 1.530 \times 10^{-3}\ mol$

Step 7: Calculate the mass of calcium carbonate in the sample of marble

mass = moles x molar mass = $1.530 \times 10^{-3}\ mol \times 100.09\ g\ mol^{-1} = 0.1531 g$

Step 8: Calculate the percentage of calcium carbonate in the marble

$$\text{Percentage of } CaCO_3 \text{ in marble} = \frac{0.1531 \times 100}{0.188} = \mathbf{81.5\ \%}$$

YOUR NOTES

Exam Tip

Rounding off when you take averagesWhen you have an average of burette readings that comes to three decimal places, e.g.(23.20 cm^3 + 23.25 cm^3) ÷ 2 = 23.225 cm^3You CANNOT show more than two decimal places because that would make the average more precise than the readings. To manage this situation you need to follow a simple rule. If the last digit is between a 5 and 9 then you round up; if the digit is between 0 and 4 you round down.So in this case the value recorded would be 23.23 cm^3