Exam Papers Practice

Boost your performance and confidence with these topic-based exam questions

Practice questions created by actual examiners and assessment experts

Detailed mark schemes

Suitable for all boards

Designed to test your ability and thoroughly prepare you

1.2 Reacting Masses & Volumes

IB Chemistry - Revision Notes

www.exampaperspractice.co.uk

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

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

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

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

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

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

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

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

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

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

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**

1.2.2 Reaction Yields

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**)

mass = mol x *M*

= 0.10 mol x 63.55 g mol^{-1}

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

Step 5: Calculate the percentage yield of copper

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

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 = <u>70 cm</u>^{3} of gases**

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**

Molar Gas Volume

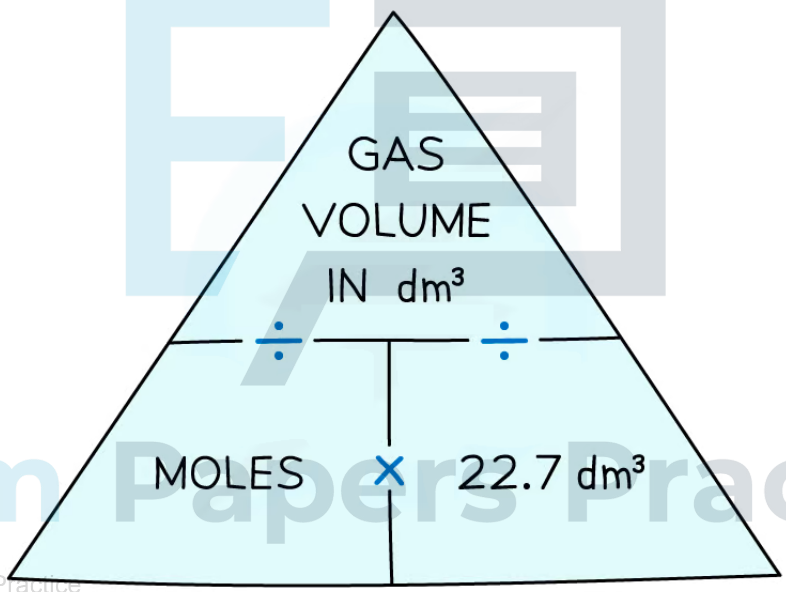
- The **molar gas volume** of 22.7 $dm^3 mol^{-1}$ can be used to find:
  - The volume of a given number of moles of gas:

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

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

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

- 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:

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

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

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

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**

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)

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 ^{o}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

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

T = 23 ^{o}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

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

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

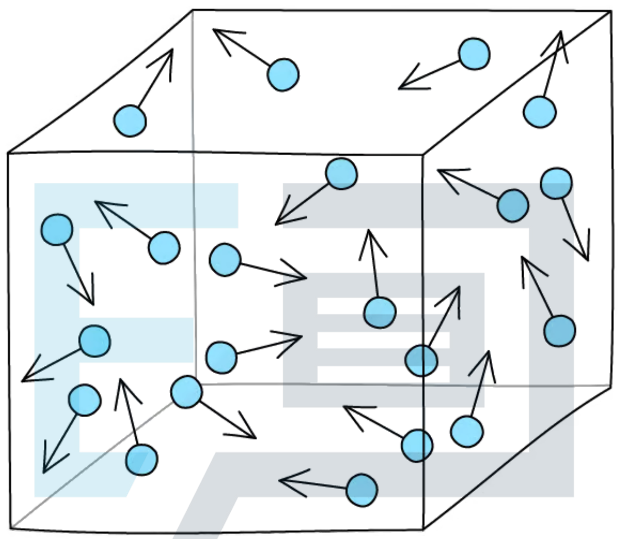
Exam Tip

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

1.2.5 Gas Law Relationships

Gas Law Relationships

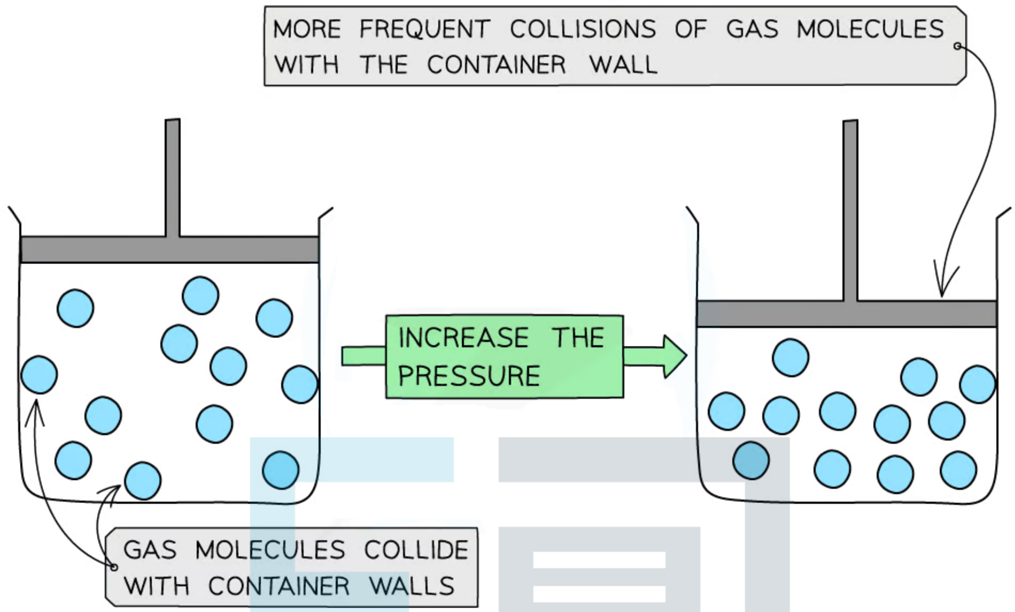
- **Gases** in a container exert a **pressure** as the gas molecules are constantly **colliding** with the walls of the container



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

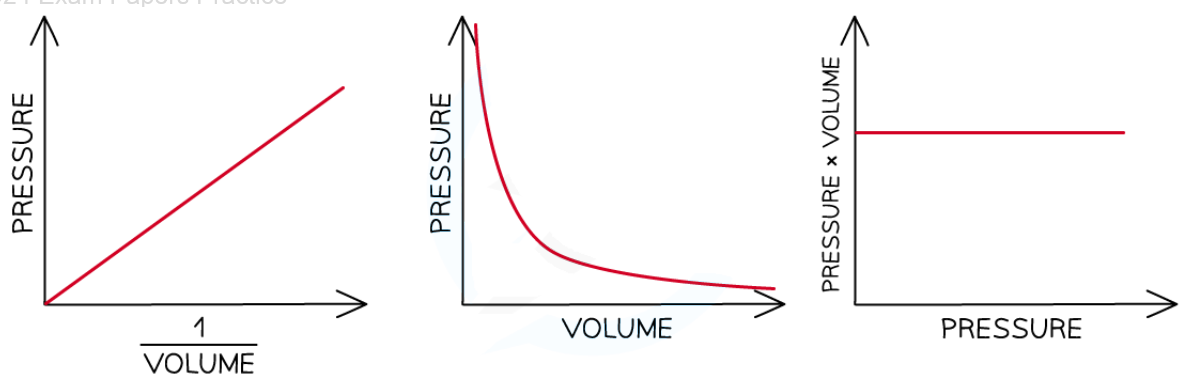
Changing gas volume

- **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

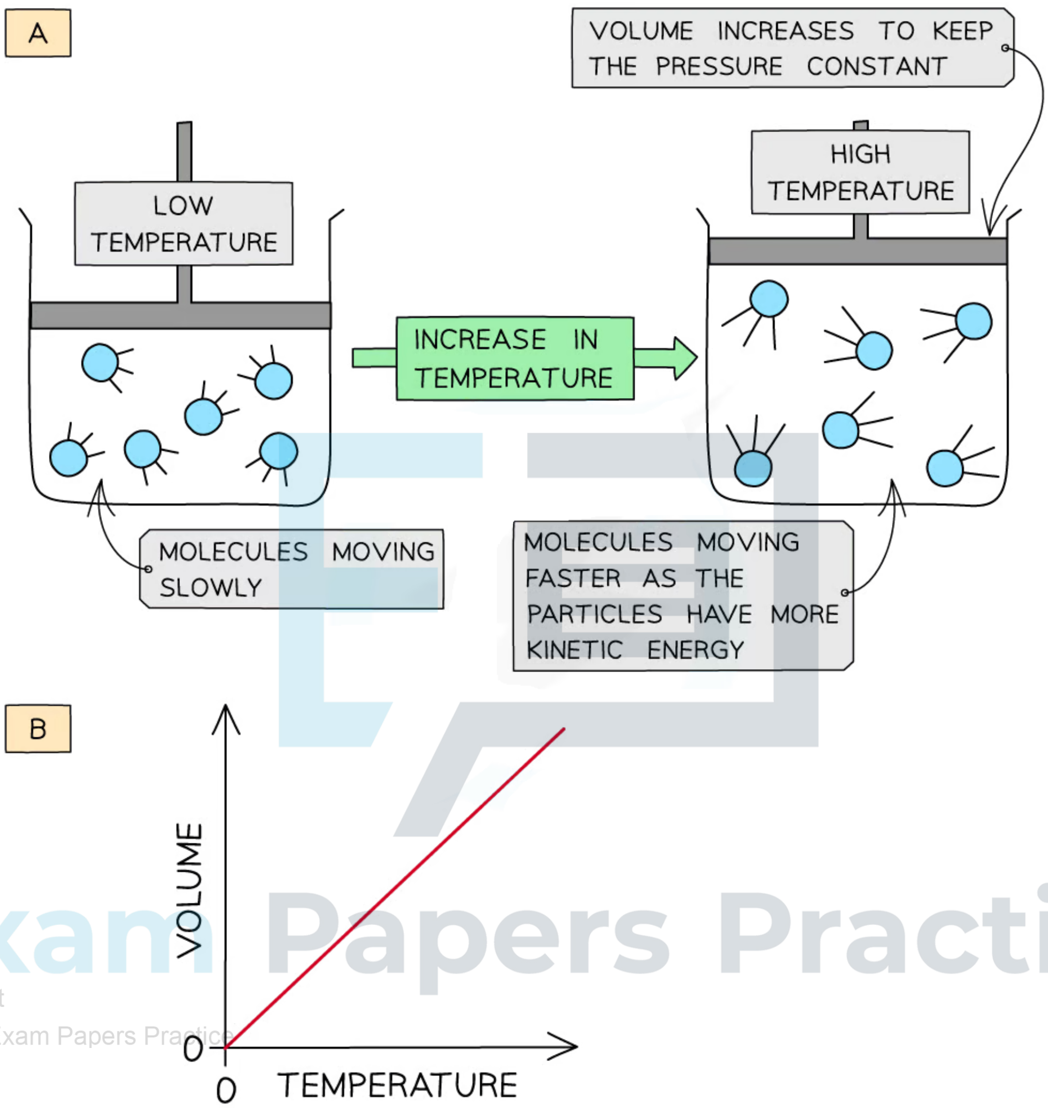
- The **pressure** is therefore **inversely proportional** to the **volume** (at constant temperature)
- This is known as **Boyle's Law**
- Mathematically, we say P ∝ 1/V or **PV = a constant**
- We can show a graphical representation of **Boyle's Law** in three different ways:
  - A graph of pressure 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

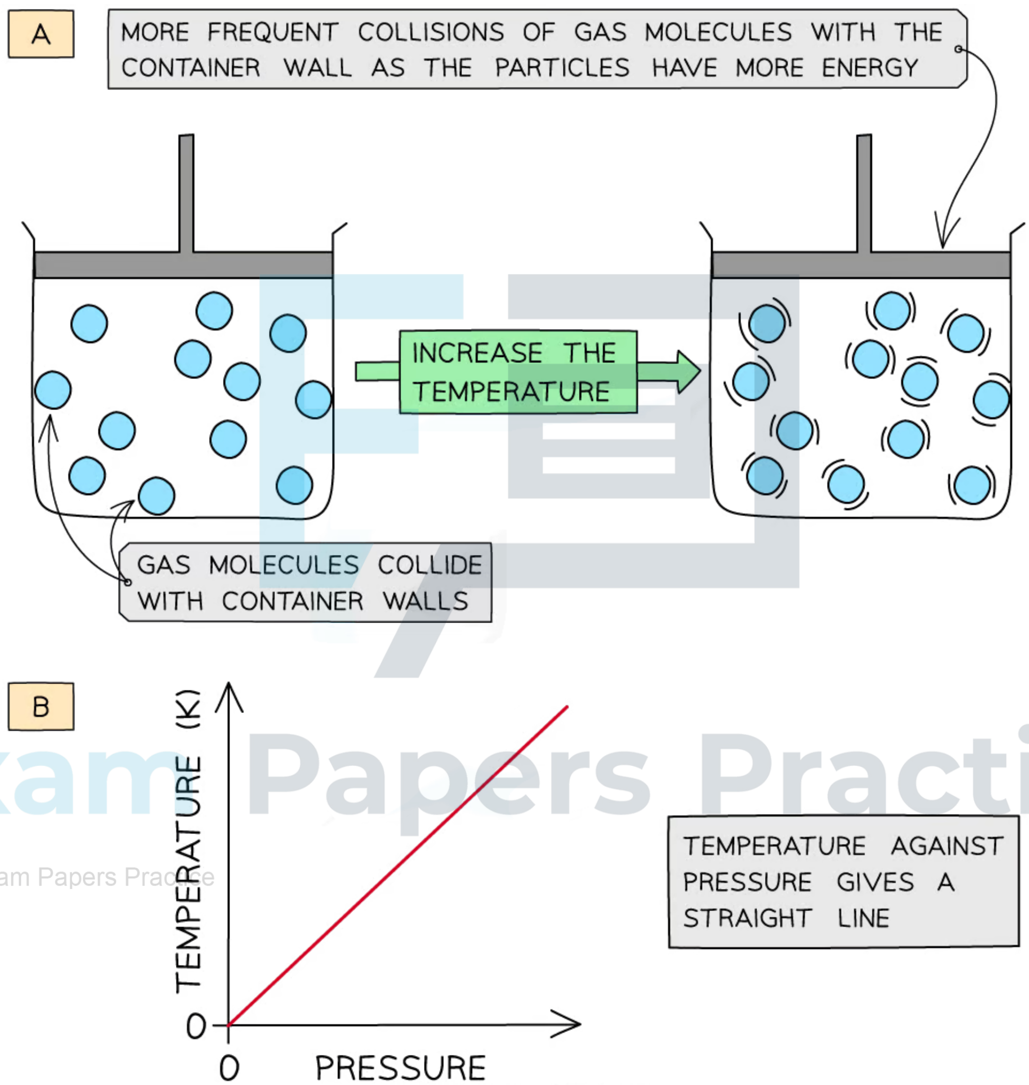


Increasing the temperature of a gas causes an increased collision frequency of the gas particles with the container wall (a); volume is directly proportional to the temperature in Kelvin (b)

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 ∝ T or **P/T = a constant**
- A graph of **temperature in Kelvin** of a gas plotted against **pressure** gives a straight line



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

- 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 ºC to 70 ºC 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

$$P_2 = \frac{P_1 V_1 T_2}{V_2 T_1}$$

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

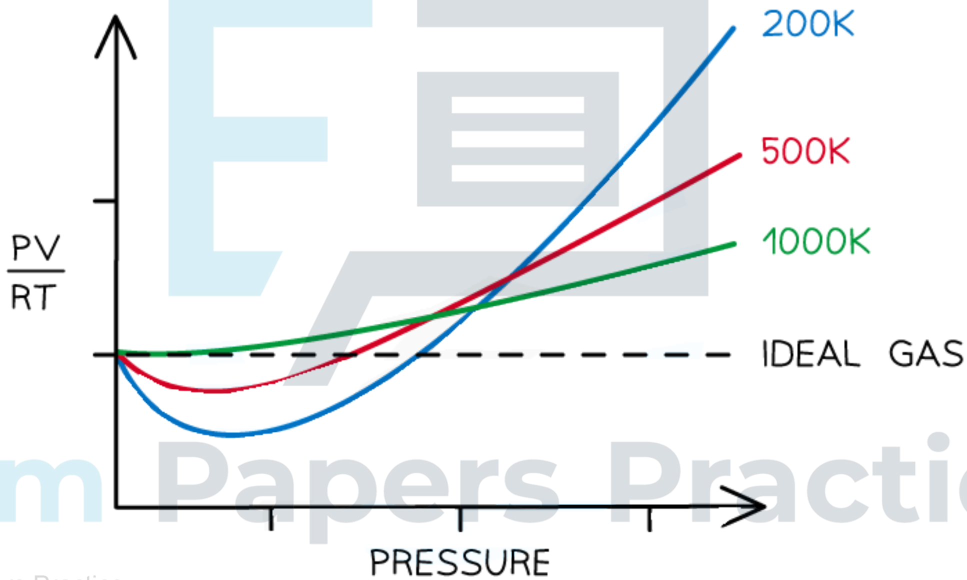
$T_2 = 70 + 273 = 343$ K

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

1.2.6 Real Gases

Real Gas Behaviour

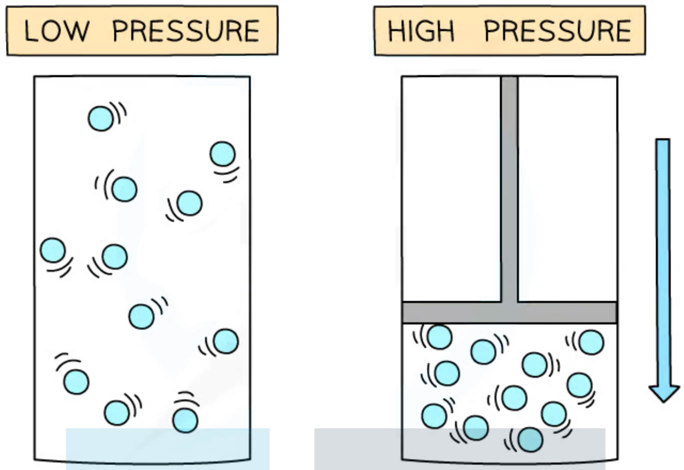
- The **ideal gas equation** does not fit all measurements and observations taken at all conditions with real gases
- The relationship between pressure, volume and temperature shows significant deviation from ***PV = nRT*** when the **temperature is very low** or the **pressure is very high**
- This 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

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

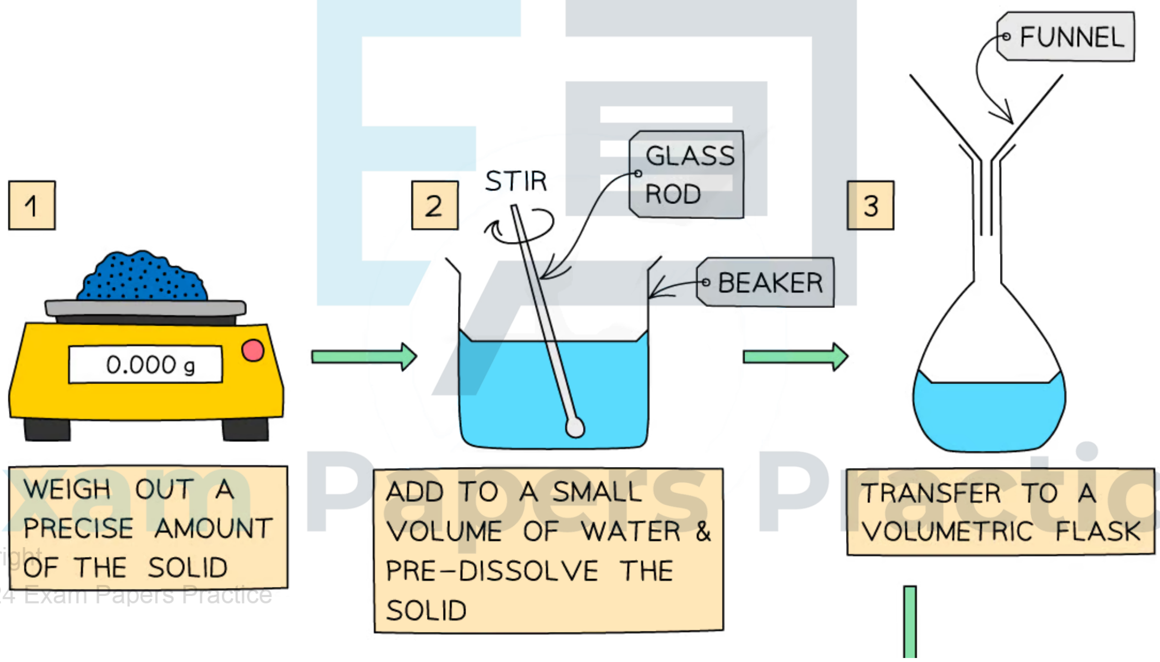
The ideal gas equation and the gas constant are given in the IB Chemistry Data Booklet which can be used in Paper 2, but not in Paper 1.

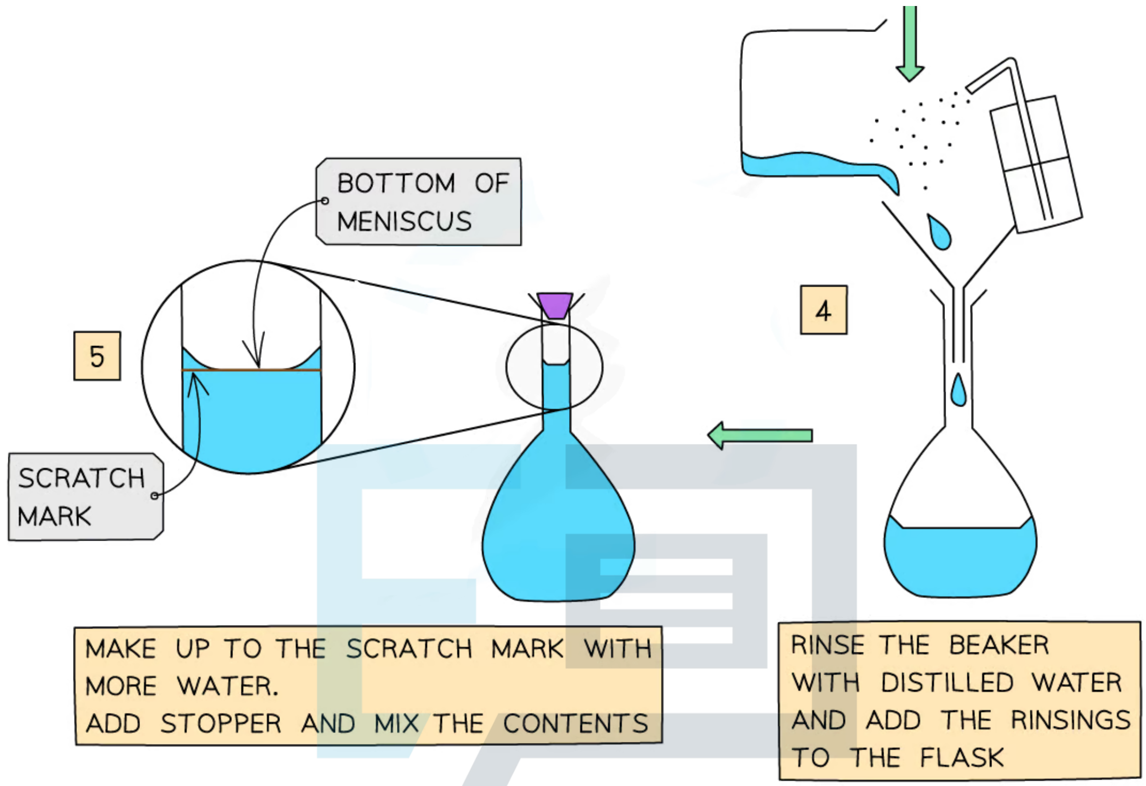
1.2.7 Standard Solutions

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:





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

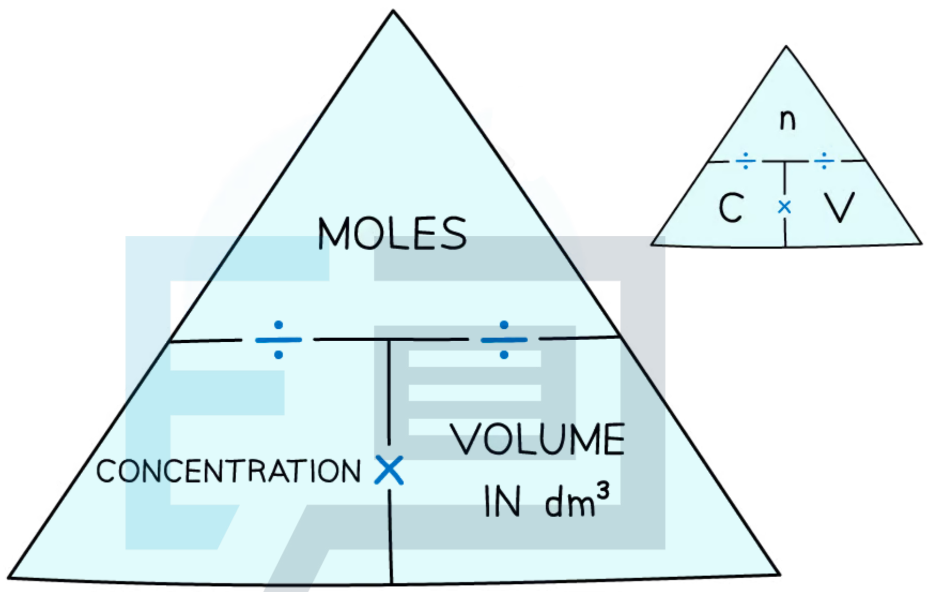
- A **concentrated** solution is a solution that has a **high** concentration of solute
- A **dilute** solution is a solution with a **low** concentration of solute
- Concentration is usually expressed in one of three ways:
  - moles per unit volume
  - mass per unit volume
  - parts per million

Moles per unit volume

- The formula for expressing concentration using moles is:

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

- 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 (dm³)

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

n = **0.0500 mol**

Step 2: Find the molar mass of NaOH

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

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:

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

- 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)

$3 \times 2.5 \times 10^6 = 7.5 \times 10^6$ mg

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

$7.5 \times 10^6 \times 10^{-6}$ kg = **7.5 kg**

1.2.8 Concentration Calculations

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

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

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

$$\textbf{amount } (Na_2CO_3) = 0.0250 \text{ dm}^3 \times 0.050 \text{ mol dm}^{-3} = 0.00125 \text{ mol}$$

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

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

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

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

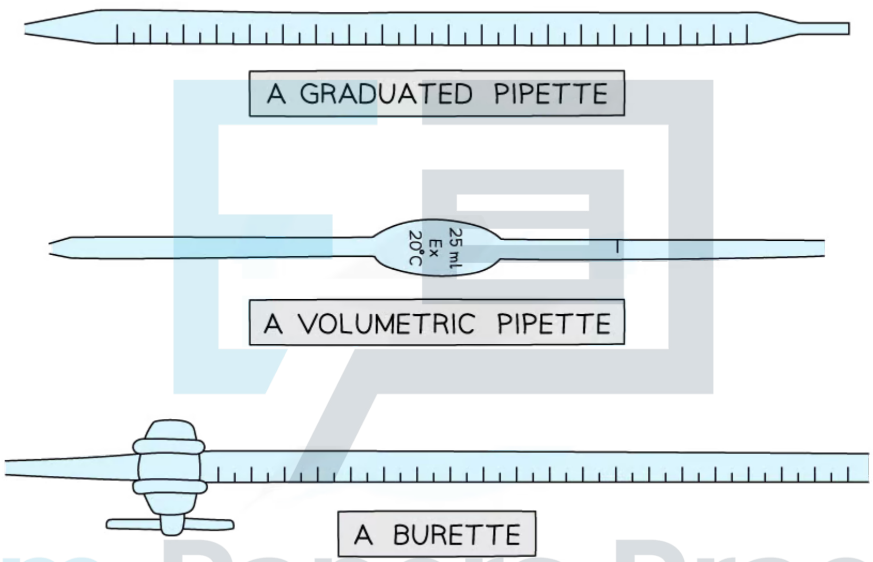
Step 4: Calculate the volume of HCl required

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

1.2.9 Titrations

Titrations

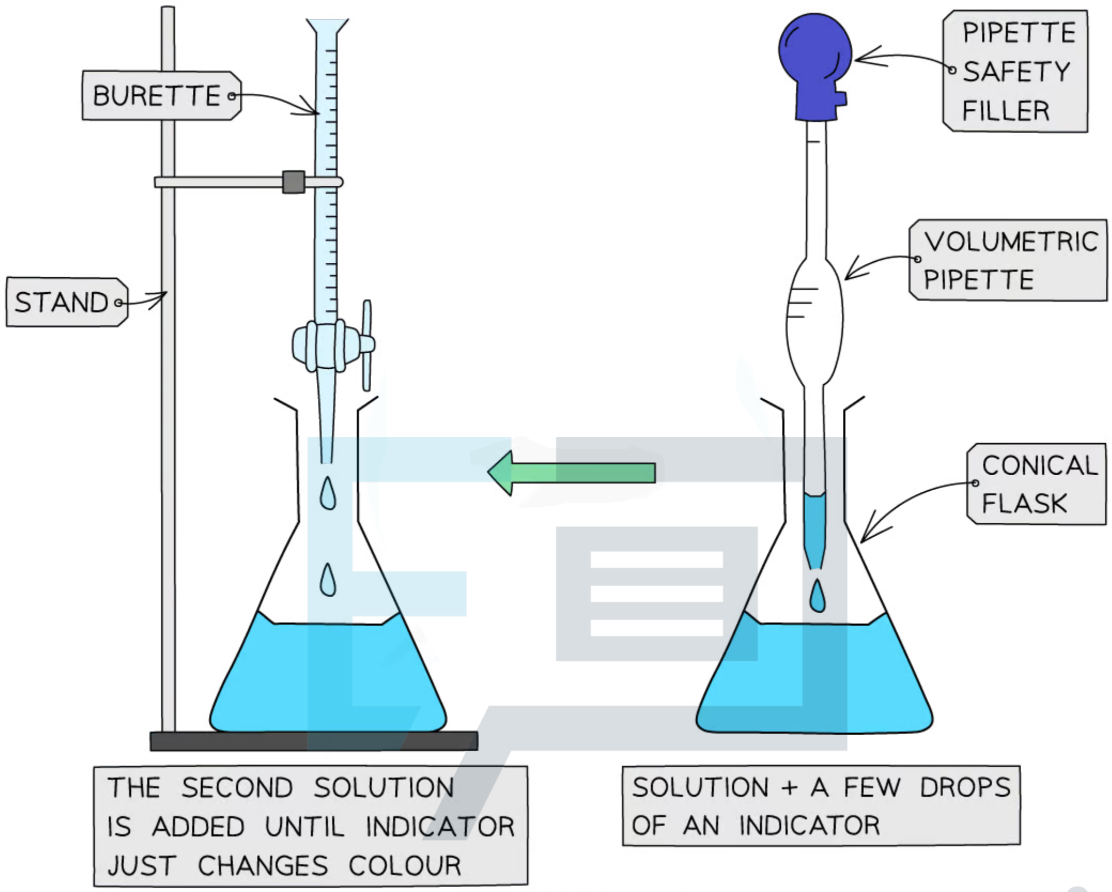
- **Volumetric analysis** is a process that uses the volume and concentration of one chemical reactant (**a standard solution**) to determine the concentration of another unknown solution
- The technique most commonly used is a **titration**
- The volumes are measured using two precise pieces of equipment,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



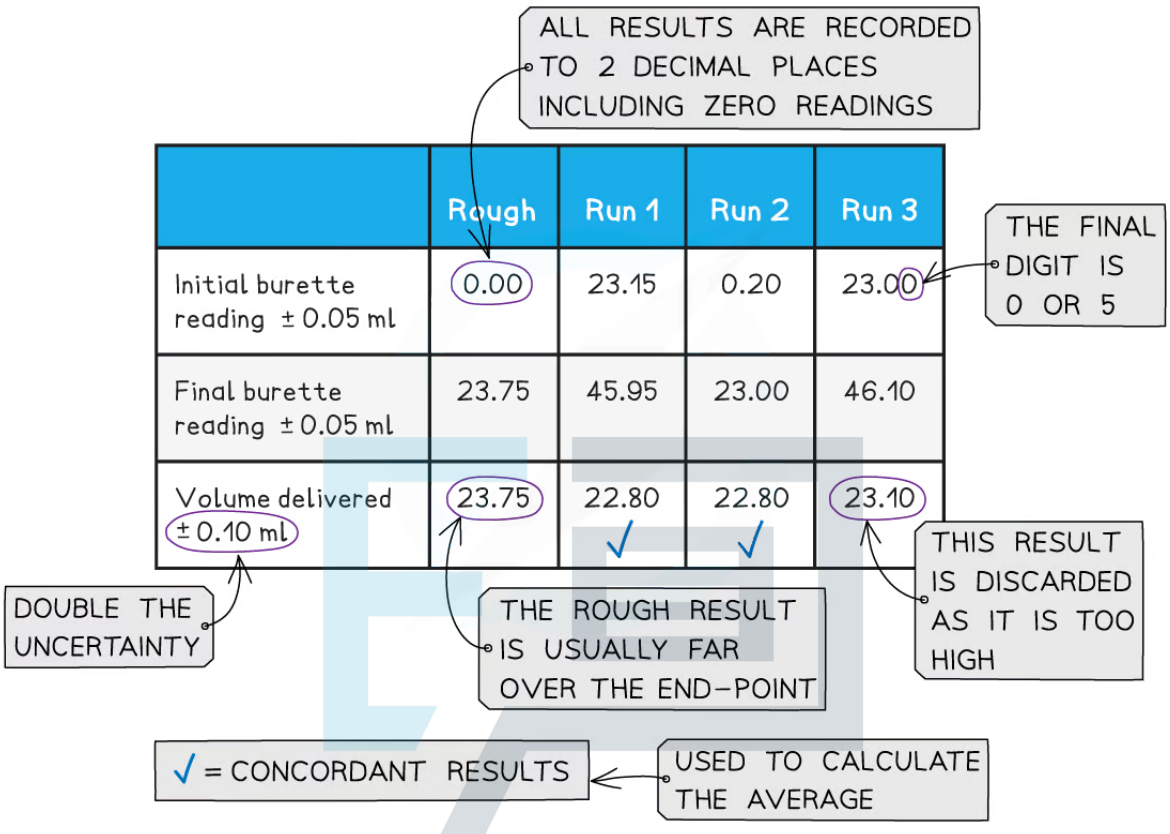
The steps in a titration

- 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 ±0.05 cm^3, the same as the **uncertainty**

	Rough	Run 1	Run 2	Run 3
Initial burette reading ± 0.05 ml	0.00	23.15	0.20	23.00
Final burette reading ± 0.05 ml	23.75	45.95	23.00	46.10
Volume delivered ± 0.10 ml	23.75	22.80 ✓	22.80 ✓	23.10



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.

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

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

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

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

Step 5: Calculate the molar mass

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

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

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

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 (3.060×10^{-3} mol) ÷ 2

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

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

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

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

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

Exam Tip

Rounding off when you take averages When you have an average of burette readings that comes to three decimal places, e.g.($23.20 \text{ cm}^3 + 23.25 \text{ cm}^3$) ÷ 2 = 23.225 cm^3 You 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