Edexcel International AS Chemistry

Your notes

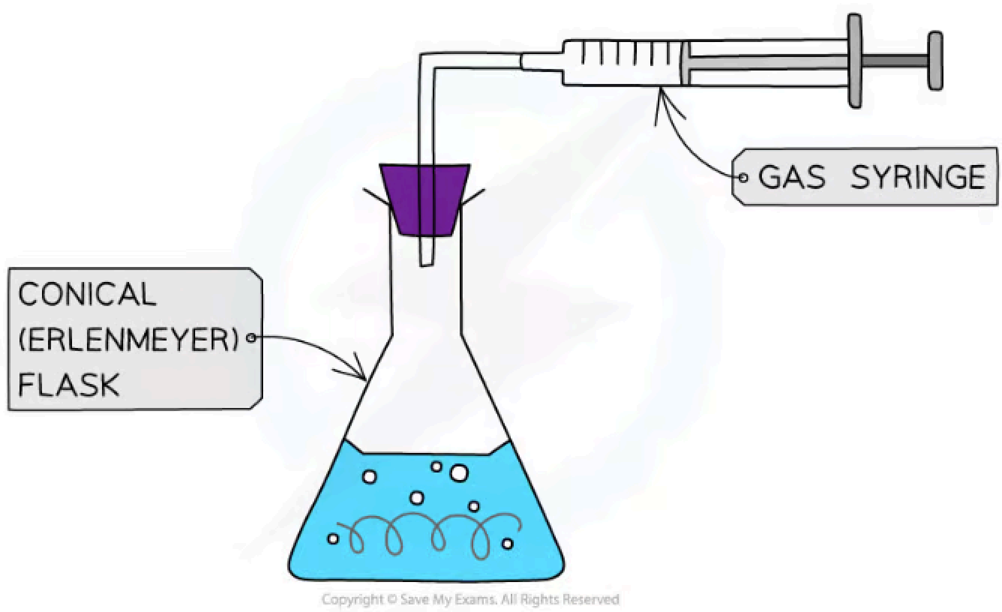
Physical Chemistry Core Practicals

Contents

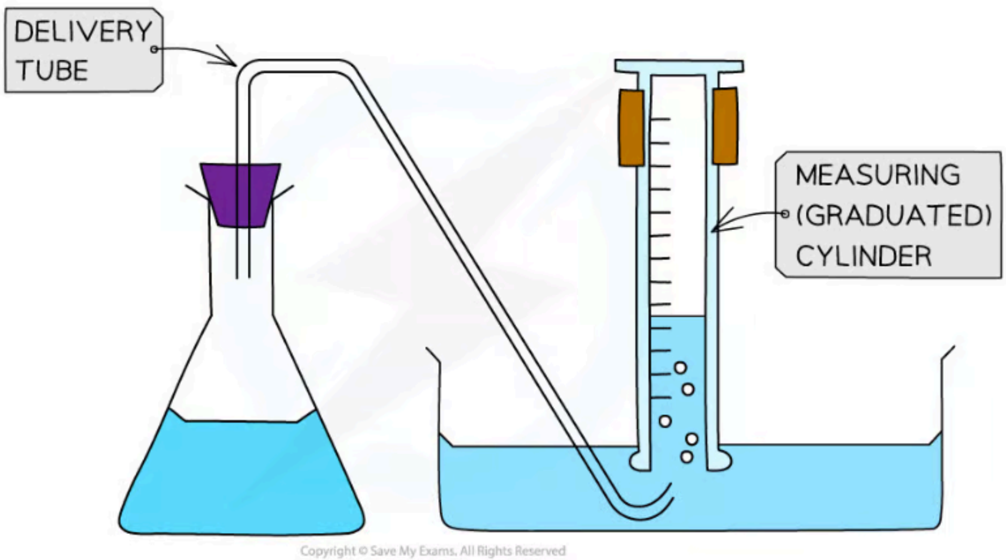
Core Practical 1: Measuring the Molar Volume of a Gas

Measuring gas volumes

- The volume of gas produced in a reaction can be measured by collecting the gas with a gas syringe or by the displacement of water



Gas syringe equipment for collecting the gas produced in a reaction



Displacement of water equipment for collecting the gas produced in a reaction

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Sample method

- For the reaction of hydrochloric acid and sodium carbonate

$$Na_2CO_3\,(s) + 2HCl\,(aq) \rightarrow 2NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$$

1. Measure out a fixed volume of hydrochloric acid, e.g. 25.0 cm^3, into a conical flask
2. Add a known mass of sodium carbonate, e.g. 0.05 g, to the conical flask
3. Immediately connect the gas syringe delivery tube
4. Allow the reaction to go to completion
5. Record the volume of carbon dioxide produced
6. Repeat the experiment with different masses of sodium carbonate, e.g. 0.10 g, 0.15 g, 0.20 g, 0.25 g... 0.50 g
7. Some assumptions are made about the experiment:
    - The amount of gas lost between adding the sodium carbonate and connecting the delivery tube is negligible
    - The delivery tube set up is airtight so no gas is lost
    - The reaction does go to completion

Sample results

Mass Volume Results Table

Mass of sodium carbonate/g	Volume of carbon dioxide produced/cm^3
0.05	9.6
0.10	24.1
0.15	29.2
0.20	47.3
0.25	60.7
0.30	62.5
0.35	79.0
0.40	84.6
0.45	105.4
0.50	110.1

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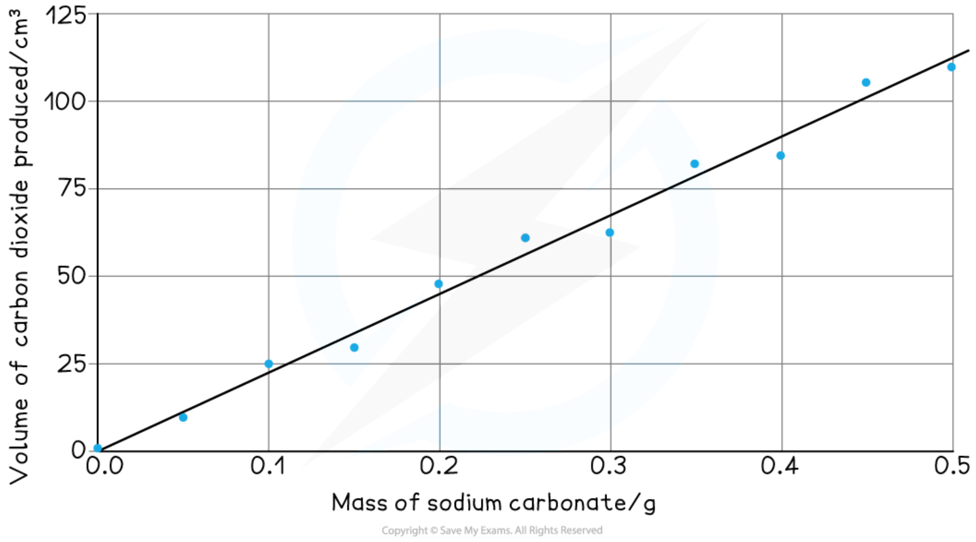
- The results are then plotted on to a graph
  - Mass of sodium carbonate on the *x*-axis and volume or carbon dioxide produced on the *y*-axis
- Anomalous results are ignored and one straight line (or one smooth curve) of best fit is added



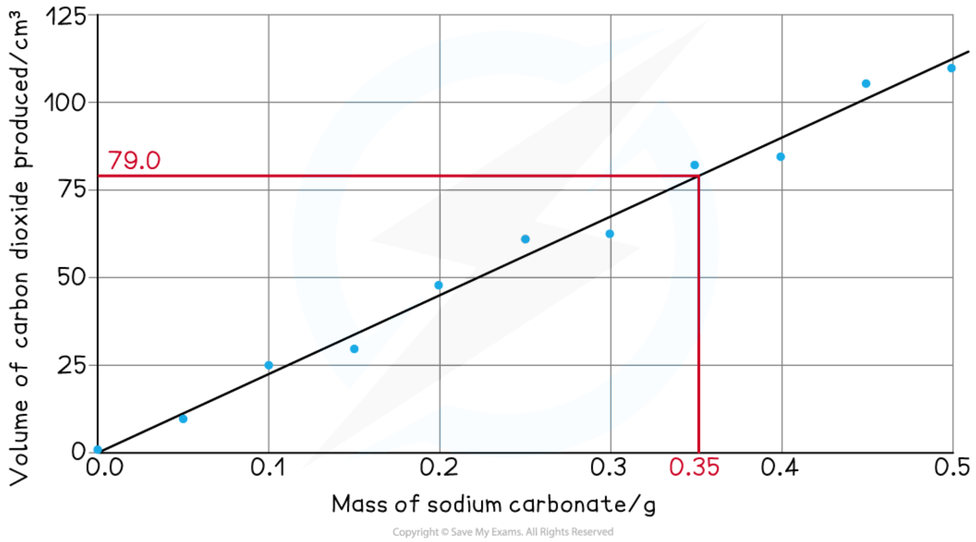
Analysis

- Read off the volume of gas produced for a sensible mass of sodium carbonate, e.g. 0.35 g produces 79.0 cm^3
  - The mass of sodium carbonate may be specified in an exam question

$$Na_2CO_3\,(s) + 2HCl\,(aq) \rightarrow 2NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$$

- From the reaction equation, one mole of sodium carbonate produces one mole of carbon dioxide
- Calculate the molar mass of sodium carbonate
  - (2 x 23.0) + 12.0 + (3 x 16.0) = 106.0
- Calculate the number of moles of sodium carbonate, using the mass from your graph reading
  - Moles $= \frac{\text{mass}}{\text{molar mass}} = \frac{0.35}{106.0} =$ 0.0033 moles
- Convert the volume of carbon dioxide from your graph reading from cm^3 to dm^3
  - $\frac{79.0\ cm^3}{1000}$ = 0.079 dm^3
- Calculate the molar volume of gas produced:

- Molar gas volume $= \frac{\text{volume}}{\text{moles}} = \frac{0.079}{0.0033} = 23.93\ dm^3$



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Application

- This experiment can be used to determine the identity of an unknown metal, M, in a metal carbonate, MCO_3
- This process can be applied to thermal decomposition of metal carbonates as well as their reaction with acid

Worked Example

At room temperature and pressure, 0.950 g of a Group 2 metal carbonate, MCO_3, reacted with hydrochloric acid to produce 226.0 cm^3 of carbon dioxide.

Deduce the identity of the metal **M**.

Answer:

Step 1: Find the number of moles of carbon dioxide released using the volume produced at room temperature and pressure:

- number of moles of $CO_2 = \frac{\text{volume of gas } (dm^3)}{\text{molar gas volume } (dm^3)}$
  - $n(CO_2) = \frac{0.226}{24} = \mathbf{0.009417\ mol}$

Step 2: Find the number of moles of metal carbonate, MCO_3

- One mole of metal carbonate will release one mole of carbon dioxide
  - Number of moles of CO_2 = number of moles of MCO_3

- $n(MCO_3)$ = **0.009417 mol**

Step 3: Calculate the molar mass of MCO_3

- $M_r = \frac{\text{mass}}{\text{moles}}$
  - $M_r = \frac{0.950}{0.009417}$ = **100.9 g mol^{-1}**

Step 4: Calculate the atomic mass of M in MCO_3 and deduction of the Group 2 metal

- M_r = Σ(atomic masses)
  - 100.9 = M + 12.0 + (3 x 16.0)
  - M = 100.9 - 60.0 = **40.9 g mol^{-1}**
  - The closest Group 2 atomic mass is calcium at 40.1 g mol^{-1}, therefore the metal M is calcium

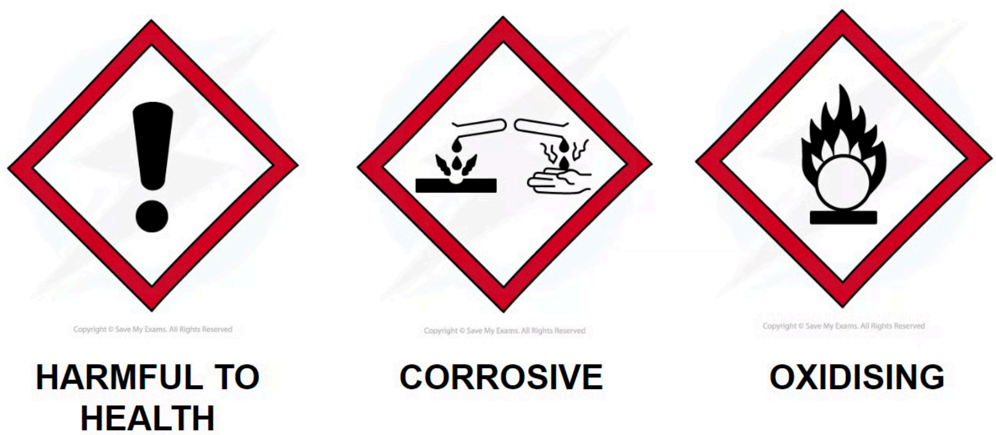
Examiner Tips and Tricks

Careful: Examiners can write these questions to include the following distractions:

- The molar mass of the metal carbonate / MCO_3 is close to the mass of a Group 2 metal
  - The mass of the carbonate ion needs to be subtracted from the molar mass in order to deduce the identity of the metal
- The atomic mass of the metal is close the atomic mass of another metal, not necessarily a Group 2 metal
  - Read the question as it will provide information about the metal

The above points can be applied to any metal carbonate, not just Group 2 metal carbonates although they are the most common

Hazards, risks and precautions



- The hazards associated with acids depend on the type and concentration of the acid

- Most dilute acids either require no hazard symbol or they are an irritant, so require the symbol to show they are harmful to health
    - Eye protection should be worn when handling
- Moderately concentrated acids are often corrosive
    - In addition to eye protection, gloves should also be worn
- Some concentrated acids, e.g. nitric acid, are oxidising which can cause or intensify a fire in contact with combustible materials
    - Eye protection and gloves are necessary when handling concentrated acids and the use of a fume cupboard is often required

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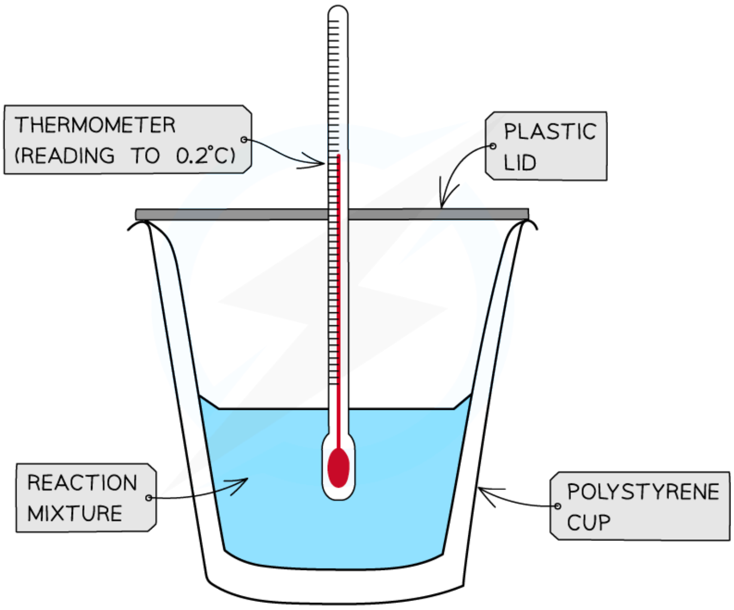
Core Practical 2: Determining Enthalpy Change of Reaction

Measuring enthalpy changes

- **Calorimetry** is a technique used to measure changes in enthalpy of chemical reactions
- A **calorimeter** can be made up of a **polystyrene drinking cup**, a **vacuum flask** or **metal can**



A polystyrene cup can act as a calorimeter to find enthalpy changes in a chemical reaction

The energy needed to raise the temperature of 1 g of a substance by 1 K is called the **specific heat capacity** (c) of the liquid The **specific heat capacity** of water is 4.18 J g^{-1} K^{-1} The energy transferred as heat can be calculated by:

$$q = m \times c \times \Delta T$$

q = THE HEAT TRANSFERRED, J
m = THE MASS OF WATER, g
c = THE SPECIFIC HEAT CAPACITY, J g^{-1} K^{-1}
ΔT = THE TEMPERATURE CHANGE, K

Equation for calculating energy transferred in a calorimeter

Enthalpy changes for reactions in solution

- The principle of these calorimetry experiments is to carry out the reaction with an excess of one reagent and measure the temperature change over the course of a few minutes
- The apparatus needed to carry out an enthalpy of reaction in solution calorimetry experiment is shown above

Sample method for a displacement reaction

1. Using a measuring cylinder place 25 cm^3 of the 1.0 mol dm^{-3} copper(II) sulphate solution into the polystyrene cup
2. Weigh about 6 g of zinc powder - as this is an excess there is no need to be very accurate
3. Draw a table to record the initial temperature and then the temperature and time every half minute up to 9.5 minutes
4. Put a thermometer or temperature probe in the cup, stir, and record the temperature every half minute for 2.5 minutes
5. At precisely 3 minutes, add the zinc powder to the cup (DO NOT RECORD THE TEMPERATURE AT 3 MINUTES)
6. Continue stirring and record the temperature for an additional 6 minutes

- For the purposes of the calculations, some assumptions are made about the experiment:
  - That the specific heat capacity of the solution is the same as pure water, i.e. **4.18 J g^{-1} K^{-1}**
  - That the density of the solution is the same as pure water, i.e. **1 g cm^{-3}**
  - The specific heat capacity of the container is ignored
  - The reaction is complete
  - There are negligible heat losses

Temperature correction graphs

- For reactions which are not instantaneous there may be a delay before the maximum temperature is reached
- During that delay the substances themselves may be losing heat to the surroundings, so that the true maximum temperature is never actually reached
- To overcome this problem we can use graphical analysis to determine the maximum enthalpy change

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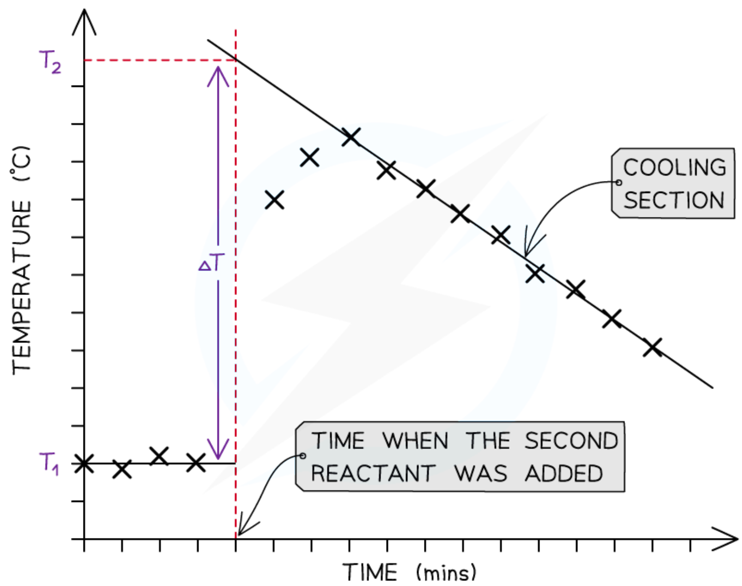

A temperature correction graph for a metal displacement reaction between zinc and copper sulfate solution. The zinc is added after 4 minutes

The steps to make a temperature correction graph are:

1. Take a temperature reading before adding the reactants for a few minutes to get a steady value
2. Add the second reactant and continue recording the temperature and time
3. Plot the graph and **extrapolate** the cooling part of the graph until you intersect the time at which the second reactant was added

Analysis

- Use both extrapolated lines to calculate ΔT as shown on the graph
- Use the equation $q = mc\Delta T$ to calculate the energy transferred
  - q = energy transferred
  - m = mass - this will be the mass of the 25 cm^3 solution which will be 25 g (assuming a density of 1 g cm^{-3})
  - c = specific heat capacity - this will be assumed to be 4.18 J g^{-1} K^{-1}, which is the specific heat capacity of water
  - ΔT = the temperature change from the graph
- Convert your value for energy transferred from J into kJ
- Then use the equation $\Delta H = \frac{q}{n}$ to calculate the enthalpy change for the reaction
  - q = energy transferred

    - *n* = number of moles - this would be the number of moles of the **limiting reagent**, which means that you will have an extra calculation to do to determine whether this is the zinc or the copper sulfate
- Remember that in the example above, the temperature of the reaction mixture increased which means that the reaction is exothermic and should, therefore, have a negative value

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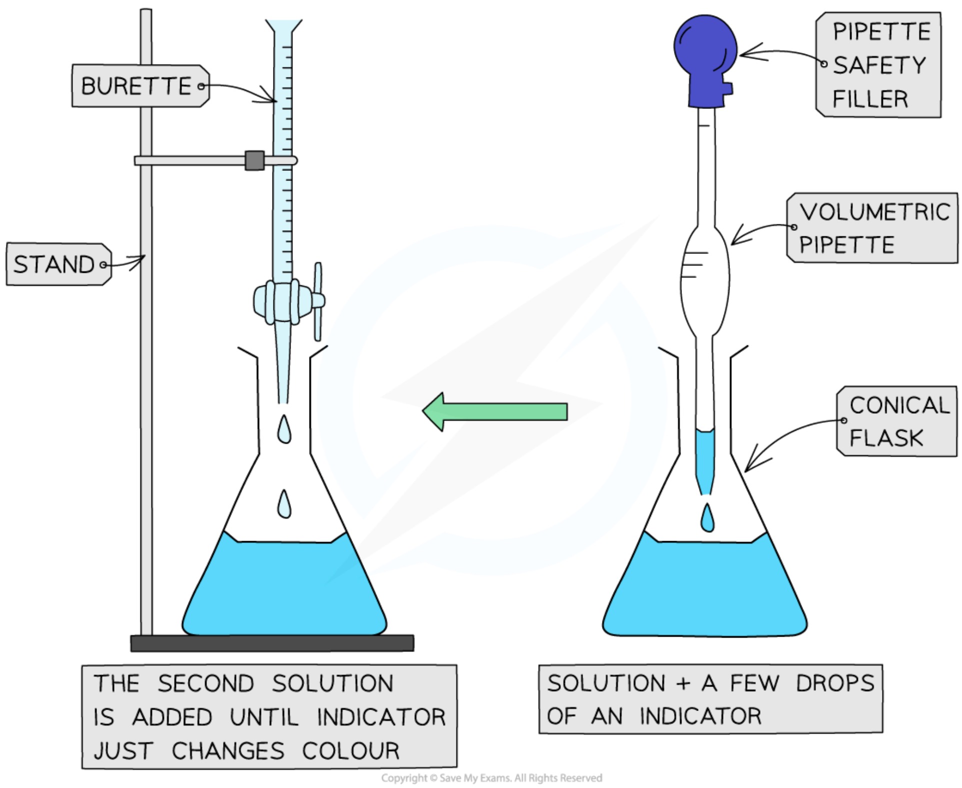
Core Practical 3: Hydrochloric Acid Concentration

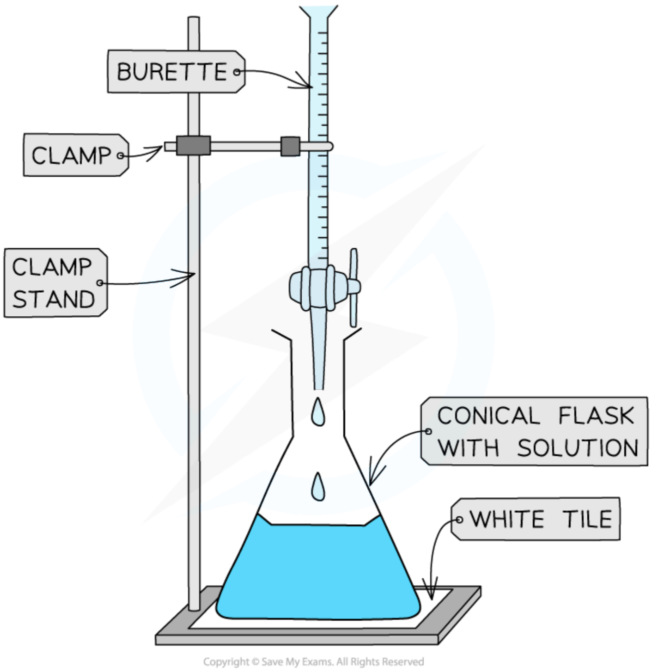
Performing the Titration

- The key piece of equipment used in the titration is the burette
- **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

- A white tile is placed under the conical flask while the titration is performed, to make it easier to see the colour change



Titrating

- The steps in a titration are:
  - Measuring a known volume (usually 20 or 25 cm^3) of one of the solutions with a **volumetric pipette** and placing it into a **conical flask**
  - The other solution is placed in the **burette**
    - To start with, the burette will usually be filled to 0.00 cm^3
  - A few drops of the **indicator** are added to the solution in the conical flask
  - The tap on the **burette** is carefully opened and the solution added, portion by portion, to the **conical flask** until the **indicator** starts to change colour
  - As you start getting near to the end point, the flow of the burette should be slowed right down so that the solution is added dropwise
    - You should be able to close the tap on the burette after one drop has caused the colour change
  - Multiple runs are carried out until **concordant** results are obtained
    - Concordant results are within 0.1 cm^3 of each other

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**

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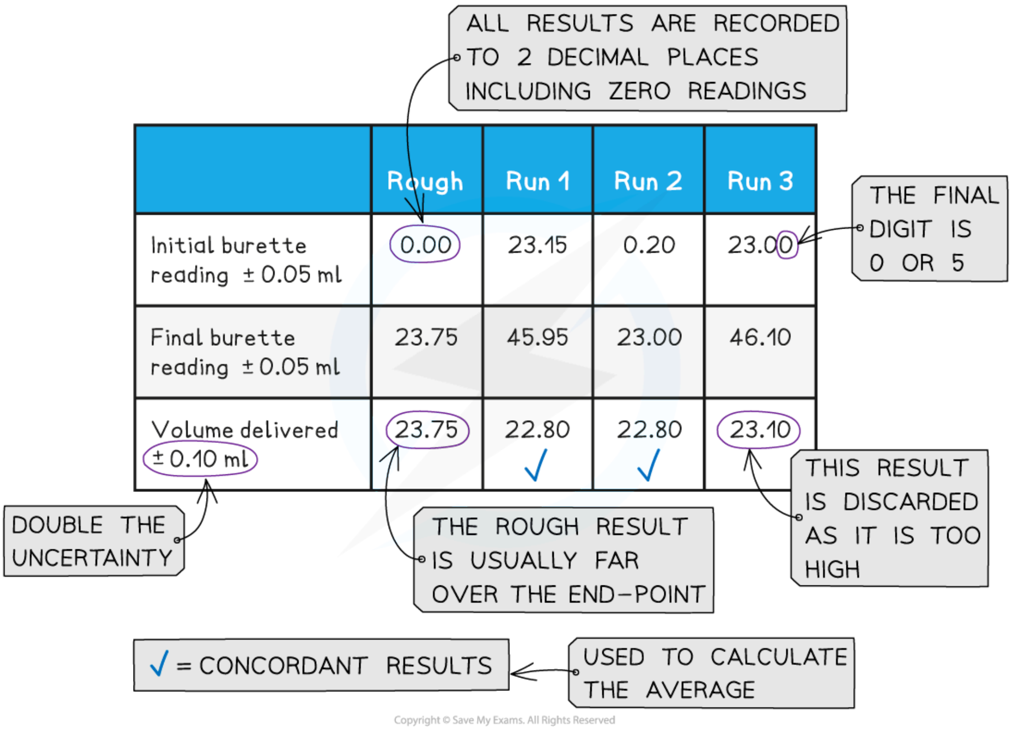

A typical layout and set of titration results

- 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**
- **Concordant** results are then averaged, and non-concordant results are discarded
- The appropriate calculations are then done

Worked Example

25.0 cm^3 of hydrochloric acid was titrated with a 0.200 mol dm^{-3} solution of sodium hydrogencarbonate, $NaHCO_3$.

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

Use the following results to calculate the concentration of the acid, to 3 significant figures.

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	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

Answer

Step 1: Calculate the average titre

- Average titre $= \dfrac{22.80 \ + \ 22.80}{2}$ = 22.80 cm^3

Step 2: Calculate the number of moles of sodium hydrogencarbonate

- Moles = $\dfrac{22.80}{1000}$ x 0.200 = 4.56 x 10^{-3} moles

Step 3: Calculate (or deduce) the number of moles of hydrochloric acid

- The stoichiometry of $NaHCO_3$: HCl is 1 : 1
  - Therefore, the number of moles of sodium hydrogencarbonate is also 4.56 x 10^{-3} moles

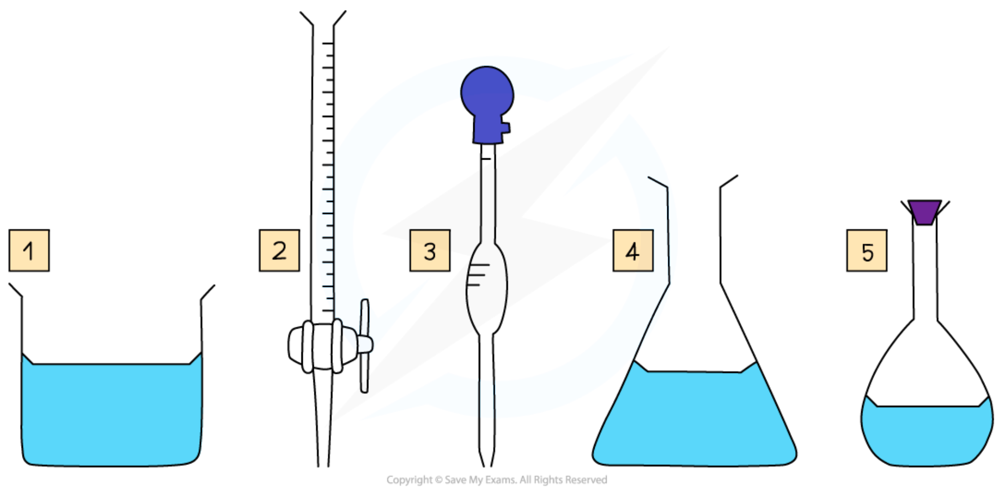
Step 4: Calculate the concentration of hydrochloric acid

- Concentration = $= \dfrac{moles}{volume} = \dfrac{4.56 \ \times \ 10^{-3}}{(25.0 \ / \ 1000)} =$ **0.182 mol dm^{-3}**

Core Practical 4: Preparing a Standard Solution

Volumetric Analysis

- **Volumetric analysis** is a process that uses the volume and concentration of one chemical reactant (**a volumetric 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**
- Before the titration can be done, the standard solution must be prepared
- Specific apparatus must be used both when preparing the standard solution and when completing the titration, to ensure that volumes are measured precisely

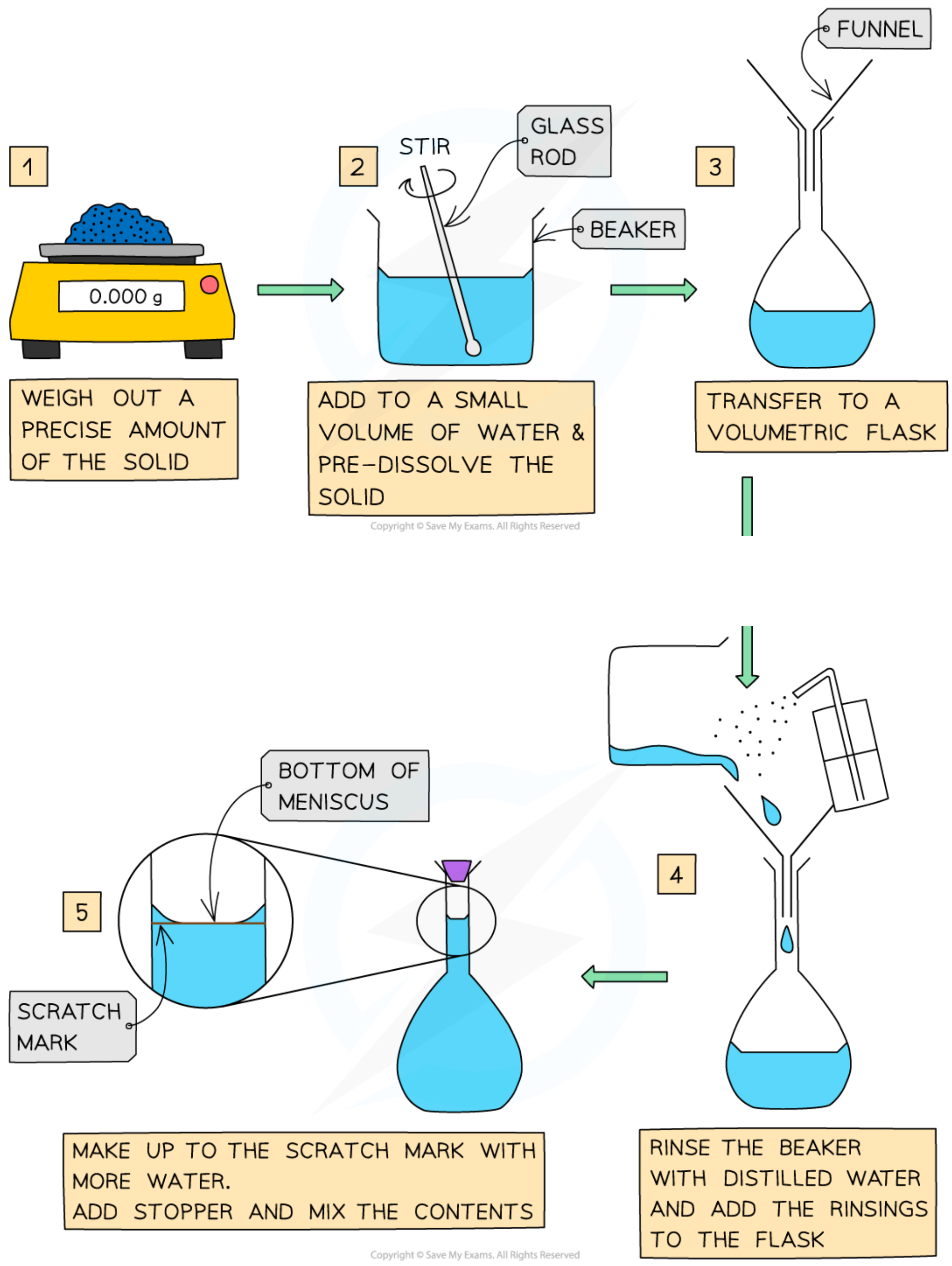


Some key pieces of apparatus used to prepare a volumetric solution and perform a simple titration

1. Beaker
2. Burette
3. Volumetric Pipette
4. Conical Flask
5. Volumetric Flask

Making a Volumetric Solution

- Chemists routinely prepare solutions needed for analysis, whose concentrations are known precisely
- These solutions are termed **volumetric solutions** or **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

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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: Find the number of moles of NaOH needed from the concentration and volume:

- *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}

Step 3: Calculate the mass required

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