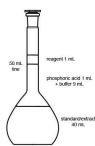
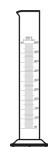
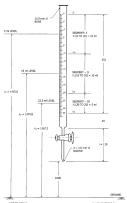
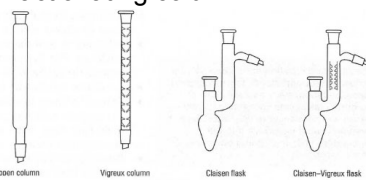
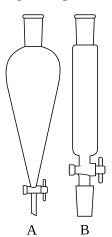
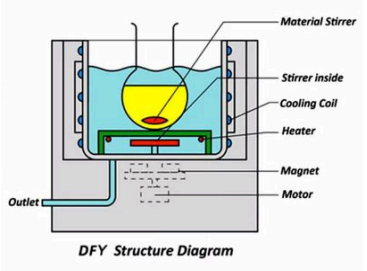


Practical tools

Electronic balance - Proximity 0.01g Beam balance - Dependent on calibration of the basis weights		Metre rule - Proximity 0.1cm Measuring tape - Proximity 0.1 or 0.5cm	Thermometers 1. Mercury 2. Alcohol 3. Digital 4. Data loggers
Volumetric flask Fixed volume of 100cm^3 or 250cm^3 Usage: to acquire an accurate volume of a chemical. 	Conical flask Usage: to mix chemicals without spillage	Measuring cylinder Proximity of 0.5cm^3 Usage: to measure volume 	
Burette Proximity of 0.05cm^3 Usage: dispense a specific volume of acid 	Pipette Fixed volume of 10.0cm^3, 20.0cm^3 or 25.0cm^3 Usage: acquire an accurate volume of chemical	Gas syringe Fixed volume of 100cm^3 Usage: measure gas volume	
Fractionating column  Small solid objects (s.a., glass beads) to provide a larger surface area for vapour condensation		Separating funnel 	Filter funnel Filter paper Stop watch Breaker
Test tube 	Boiling tube 	Water bath  DFY Structure Diagram The water bath keeps the given solution/mixture at a constant temperature, allowing for more accurate experiments.	

Calculation

D.P. rule Whenever calculating two values that <u>only</u> involve addition and subtraction, the calculated value adopts the least d.p., with a minimum of 0d.p.	For s.f., For the final answer, round to 3 s.f. For method calculation, use 4 s.f. To 6 s.f.
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Graphical analysis

Plot Y against X , based on the given equation and determine the relationship of the graph.		
<u>Proportional relationship</u> When two quantities are directly or inversely proportional to each other. - The graph is a straight line - Passes through the origin (0,0)	<u>Linear relationship</u> When two quantities are linearly related. The graph is a straight line	<u>Non-linear relationship</u> Two quantities create a graph that is curved.
$Y = kX$	$Y = mX + C$	$Y = f(X)$
<u>Constant relationship</u> When two quantities that are independent of each other, the graph produces a straight horizontal line.	Note: Some graphs are expected to pass through the origin based on the physical relationship between the variables. For example, if the dependent variable represents the change caused by a limiting reactant, adding 0 volume of the reactant should produce 0 change in the system. Therefore, the graph should pass through the origin (0, 0).	
$Y = k$		

Quantity Algebra, both the numerical magnitude and unit of a quantity are substituted for its symbol in the equation.	Number Algebra, only the numerical magnitude of a quantity is substituted for its symbol in the equation, converted to the usual units by applying the appropriate multiplying factor.
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Errors and uncertainty

Systemic error Constant deviations of the readings in one direction from the true value. - Zero error - Reaction time - Wrongly calibrated scale Systematic errors can be reduced by calibration curves and control experiments.
Random error Measurements are scattered around a mean value. - Reading a scale about interpolation - Timing oscillations without a reference marker - Taking readings of a quantity that varies with time. - Parallax error. Random error can be mitigated by averaging the values, repeating the experiment or by drawing a graph.
Accuracy: how close a measurement is to the actual value Note: accuracy has a 10% range from the true value. Precision: how close the measurements are to each other, or to a mean value
Uncertainty is a range of values in which a measurement can fall. - Uncertainty of a digital equipment is one unit of the smallest scale of the instrument - Uncertainty of an analogue instrument is half the smallest division of the scale. $M_1 = M_0 \pm \delta_M$ Average of values, $\sigma_m = \frac{\sum^n x}{n}$ $\sigma_{m1} = \sigma_m \pm \frac{\max(x) - \min(x)}{2}$ - Note: when calculating the mean of values, it follows the d.p. rule

Absolute uncertainty, δ

Percentage uncertainty, $P_M = \frac{\delta_M}{M_0} \times 100\%$

Addition or subtraction of uncertainties, $(A_0 + B_0) \pm (\delta_A + \delta_B)$

Multiplication or division of uncertainties, $A_0 B_0 \pm A_0 B_0 (P_A + P_B)$

Power or root of uncertainties, $A_0^K B_0^J \pm A_0^K B_0^J (|K|P_A + |J|P_B)$

Note: d.p. Of both must be the same.

Increasing accuracy

When measuring the diameter of an object, take repeated readings at different position across the diameter, finding the average of all the readings.

When measuring the time taken, take repeated readings, finding the average of all the readings.

Precision of instrument

For digital measuring instruments, the measurement uncertainty is typically ± 1 in the last displayed digit. Therefore, record all the digits shown by the display.

- Example: If a digital micrometre displays 12.3456 mm, record 12.3456 mm $\pm$ 0.0001 mm. Do not round it to 12.346 mm.
- Includes: Digital multimeter, digital micrometre screw gauge, digital vernier callipers.

For analogue measuring instruments with discrete scale divisions, the measurement uncertainty is typically $\pm \frac{1}{2}$ of the smallest scale division.

- Example: A metre rule has the smallest division of 0.1 cm (1 mm), so the uncertainty is ± 0.05 cm (± 0.5 mm). Measurements are therefore estimated to the nearest 0.05 cm.
- Includes: voltmeter, ammeter, thermometer, ruler, vernier callipers, micrometre, screw gauge.

Factors that impact yield

1. Incomplete reaction $\rightarrow$ some reactants remain unreacted.
2. Side reactions $\rightarrow$ reactants form unwanted products.
3. Product lost during transfer $\rightarrow$ some sticks to apparatus.
4. Product lost during filtration $\rightarrow$ some passes through or remains on the apparatus.
5. Product lost during washing $\rightarrow$ some soluble product dissolves away.
6. Product lost during crystallisation $\rightarrow$ some remains dissolved in the solution.
7. Product lost during heating $\rightarrow$ decomposition may occur if overheated.
8. Accidental losses $\rightarrow$ crystals/precipitate may be spilled or left behind.

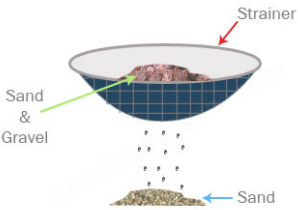
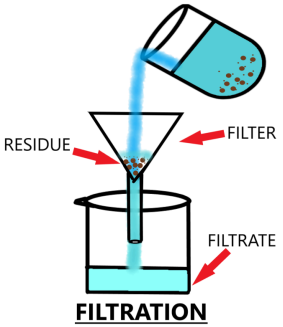
Factors that impact purity

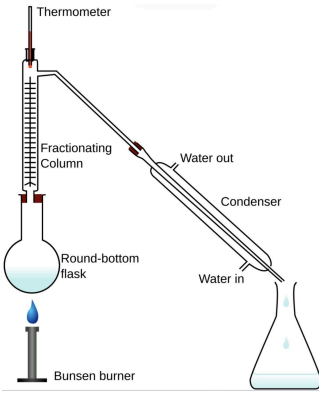
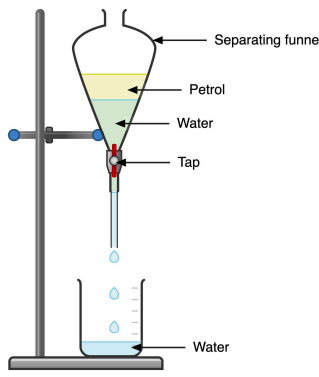
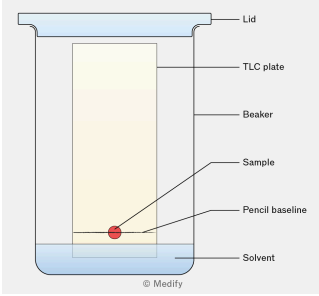
1. Unreacted reactants remain in the product.
2. Side products/by-products are present.
3. Impurities in the starting materials.
4. Insufficient washing $\rightarrow$ soluble impurities remain.
5. Incomplete filtration $\rightarrow$ unwanted solid remains.
6. Incomplete separation $\rightarrow$ another substance remains mixed with the product.
7. Insufficient drying $\rightarrow$ water/solvent remains in the product.

Experiments

1. Separation techniques

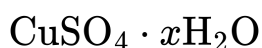
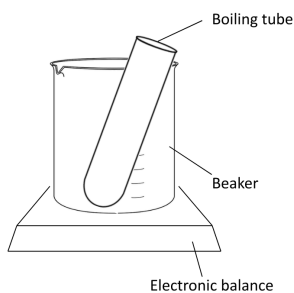
Magnetic

	Experimental procedure 1. Slowly pass a magnet through the sample, sweeping the entire mixture. 2. The magnetic materials would be attracted to the magnet, and can be collected as such; the remaining non-magnetic substances would remain in the container. 3. After the first pass, mix the remaining material and pass the magnet through again. - Repeated separation is particularly useful when the magnetic particles are mixed with a lot of non-magnetic material.
Sieving	
	Experimental procedure 1. Place the mixture on a sieve. 2. Rapidly shake the sieve, moving it back and forth, for a while. 3. The finer particles would be sieved out of the mixture; the larger particles would remain in the sieve.
Filtration	
	Experimental procedure 1. Add the mixture to a beaker. 2. Add distilled water in excess. 3. Stir the mixture with a glass rod, ensuring soluble substance dissolves. 4. Place filter paper inside a filter funnel. 5. Wet the filter paper slightly so it sticks to the funnel. 6. Place the funnel over a clean beaker. 7. Pour the mixture slowly into the filter paper, using a glass rod to guide the liquid. 8. The insoluble solid remains on the filter paper as the residue. 9. The dissolved substance passes through with the water as the filtrate. Obtaining dry residue a. Wash the residue with a small amount of distilled water to remove any remaining soluble substance. b. Dry the residue between a few pieces of filter paper. Obtaining dry filtrate a. Heat the filtrate gently, obtaining a saturated solution. b. Cool the solution to allow crystal formation. c. Filter to collect crystals. d. Wash the crystals with a little bit of deionised water to remove soluble impurities. e. Dry the crystals between a few pieces of filter paper.
Fractional distillation	

	Experimental procedure 1. Set up the apparatus as shown in the diagram. 2. Heat the mixture. 3. When the temperature reaches T_1, [substance] boils, the vapour condenses and is collected in the beaker. 4. Temperature remains constant at T_1 as [substance] distills over and rises once all the [substance] is collected.
	Experimental procedure 1. Pour the mixture of immiscible liquids into the separating funnel. 2. Allow the mixture to stand until two distinct layers form. 3. The denser liquid forms the bottom layer; The less dense liquid forms the top layer. 4. Place a clean beaker beneath the separating funnel. 5. Open the stopcock slowly and allow the bottom layer to flow out. 6. Close the stopcock when the boundary between the two liquids reaches the stopcock. 7. Place another clean container underneath and collect the upper layer by opening the stopper at the top and pouring it out.
	Experimental procedure 1. Set up the apparatus as shown in the diagram. 2. At the end of the experiment remove the chromatogram from the setup. 3. Mark the solvent front. 4. [For invisible spots] Spray the chromatogram with a locating agent so that the colourless substances become visible. 5. Measure the distance travelled by the spots, Δs, and the distance travelled by the solvent from the start line, ΔD. 6. Calculate the R_f value using the formula $R_f = \frac{\Delta s}{\Delta D}$ 7. Use the R_f value of the spots and compare it within known data values.

2.Stoichiometry

Water of crystallisation



Experimental procedure

1. Measure and record the mass of an empty boiling tube using an electronic balance, M_0 .
2. Add the solid **P** to the boiling tube. Measure and record the total mass of the boiling tube and **P**, M_1 .
3. Gently heat the boiling tube.
4. Leave the boiling tube to cool to room temperature.
5. Reweigh the boiling tube and its contents and record the mass, M_2 .
6. Repeat steps 3 to 5 until the mass remains constant M_2 .

Finding the percentage mass of water

- Calculate the mass of crystallised **P**, $M_1 - M_0$.
- Calculate the mass of water of crystallisation lost during the water, $M_1 - M_2$.
- Calculate the percentage decrease in mass of the crystallised **P**

$$P_{\text{H}_2\text{O}\%} = \frac{M_1 - M_2}{M_1 - M_0} \times 100\%$$

Finding the molar ratio, x

- From the molecular formula of the **P**, for every mole of crystallised **P**, there is 1 mole of **P**, and x mole of water of crystallisation.
- Calculate the mass of water of crystallisation lost during the water, $M_1 - M_2$.
- Calculate the number of moles of water of crystallisation,

$$\text{Moles}_{\text{H}_2\text{O}} = \frac{M_1 - M_2}{M_r[\text{H}_2\text{O}]}$$

- Calculate the number of moles of **P** after heating,

$$\text{Moles}_{\text{Sub.}} = \frac{M_2}{M_r[\text{Sub.}]}$$

- Therefore, using the molar ratio,

$$\text{Sub.} : \text{H}_2\text{O}$$

$$1 : x$$

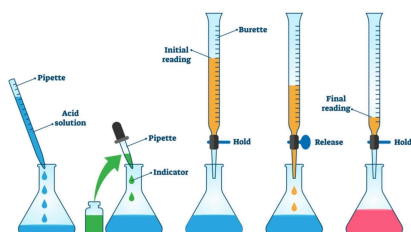
$$\Rightarrow x = \frac{\text{Moles}_{\text{H}_2\text{O}}}{\text{Moles}_{\text{Sub.}}}$$

Titration

Titration table

Titration number	1	2	3
Initial reading/cm ³			
Final reading/cm ³			

Volume of _____ used/cm ³			
Best results			



Experimental procedure

1. Fill up burette with **Q**.
2. Pipette 25.0cm³ of **P** into a conical flask.
3. Add 3 drops of suitable indicator to the conical flask. The solution turns (indicator colour at that pH)
4. Add **Q** from burette slowly until solution turns (transition colour).
5. Record volume of **Q** required for complete neutralisation, V_{τ} .
6. Repeat steps 1 to 5, and record the volume of acid/base for complete neutralisation.
7. Take the best readings, $\pm 0.20\text{cm}^3$, and calculate the average volume of **Q** used, V_{τ} .
8. Pipette 25.0cm³ of **P** into a conical flask, then add V_{τ} of **Q** from the burette.

Perform crystallisation to obtain crystals of the solution

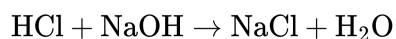
1. Heat the solution until saturated.
2. Allow the saturated solution to cool and to form crystals.
3. Rinse the crystals with a little cold deionised water.
4. Dry the crystals between a few pieces of filter paper.

Determining the concentration of the solution **P**, (Taking: **P** as HCl, and **Q** as NaOH)

1. Calculate the number of moles of **Q** using the known concentration,

$$\text{Moles}_{\text{Q}} = C[\text{Q}] \times V_{\tau}$$

2. From the equation,



1 mole of **P** reacts with 1 mole of **Q**.

Hence, determine the number of moles of **P**,

$$\text{Moles}_{\text{Q}} : \text{Moles}_{\text{P}}$$

$$1\text{mol} : 1\text{mol}$$

$$C[\text{Q}] \times V_{\tau} : C[\text{Q}] \times V_{\tau}$$

3. Calculate the concentration of **P**

$$C[\text{P}] = \frac{C[\text{Q}] \times V_{\tau}}{25\text{cm}^3}$$

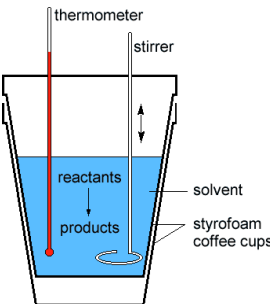
pH indicators

Indicator	Low pH	pH range of colour change	High pH
Litmus	Red	7 (5-8) Purple	Blue
Methyl Orange	Red	4 (3.1-4.4) Orange	Yellow

Bromothymol blue	Yellow	7 (6.0-7.6) Green	Blue
Screened methyl orange	Violet	4 (3.8-4.1) Grey	Green
Thymolphthalein	Colourless	10 (9.3-10.5) Pale Blue	Blue
Phenolphthalein	Colourless	9 (8.3-10.0) Pink	Red

Note: For the intermediate colour change, just use secondary and primary colour theories to guess the middle colour.

Using temperature to find needed volume for complete reaction



Experimental procedure

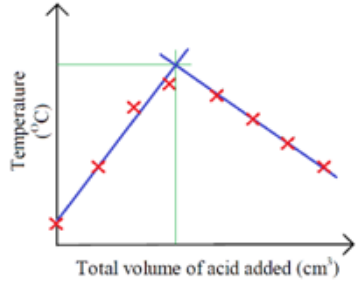
1. Pipette 25.0cm^3 of **P** into styrofoam cup,
2. Using a thermometer, measure and record the initial temperature of **P**, T_0 .
3. Fill up burette with **Q**.
4. Add 5.0cm^3 of **Q** to the styrofoam cup.
5. Stir the mixture thoroughly with a thermometer/glass rod.
6. Record the maximum temperature reached, T_1 .
7. Calculate the maximum temperature rise, $\Delta T = T_1 - T_0$.
8. Repeat steps 1 to 6, with consistent intervals of volume of **Q**, record the new values of ΔT .
9. Tabulate the recordings.

Example of a table

Exp.	Vol. P / cm^3	Vol. Q / cm^3	$T_0/^\circ\text{C}$	$T_1/^\circ\text{C}$	$\Delta T/^\circ\text{C}$
1	25.0	5.00	31.0	x_1	D_1
2	25.0	10.00	31.0	x_2	D_2
⋮	⋮	⋮	⋮	⋮	⋮
n	25.0	$5.00n$	31.0	x_n	D_n

Explanation, usage of styrofoam cup:

- To reduce heat dissipation to the surroundings.



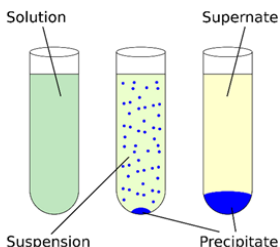
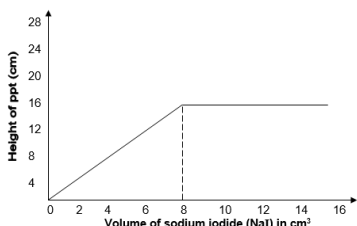
Graphical analysis

- Plot a graph with the rise in temperature against volume of **Q** used.
- Draw two best fit lines, ensuring that the intersection is above or equal to the highest temperature rise.
- Extend horizontal and vertical lines from the axes to the point of intersection, read the volume of **Q** marked.
- This corresponds to the maximum temperature is the volume of **Q** required for complete reaction.

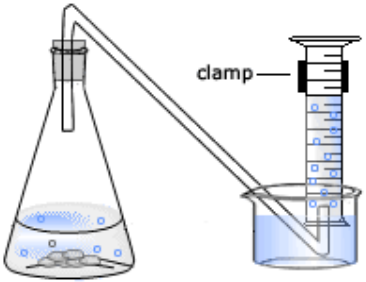
Explanation:

	- The reaction is exothermic. - Before the equivalence point, Q is limiting reactant. Since Q is the limiting reactant, increasing the volume of Q added increases the amount of Q that reacts with P, which increases the amount of heat released. This causes the temperature rise to increase. - After the equivalence point, P is limiting reactant, and Q is in excess. Once P is fully used up, P becomes the limiting reactant and the amount of heat released becomes fixed. - However, the additional unreacted Q absorbs some of the heat produced and adds to the total volume of the mixture, so the temperature rise per unit volume decreases as more Q is added.
--	---

Using height of precipitate to find needed volume for complete reaction

	Experimental procedure 1. Prepare 6 test tubes. 2. Pipette 5.0cm^3 of P into each test tube. 3. Fill up burette with Q. 4. Add Q to each test tube in increasing volumes of 1.00cm^3, starting at 1.00cm^3 (i.e., 1.00cm^3, 2.00cm^3, 3.00cm^3, etc.) 5. Use a glass rod and carefully stir the contents of each of the test-tubes 6. Wait for about 25 minutes, allowing the precipitate to settle. 7. Measure and record the height of the precipitate in each test tube using a ruler. 8. Tabulate the recordings.
	Graphical analysis - Plot a graph of height of precipitate against volume of Q added. - Draw two lines (one horizontal, one diagonal) - The point of intersection corresponds to the volume of Q required for complete reaction. Explanation: - With more Q, the height will increase as more precipitate forms as Q is limiting reactant and P is in excess. - Eventually the graph will remain constant once Q is in excess, and P is limiting reactant

Using gas produced to find molar mass, carbonate compound



Experimental procedure

- Set up the apparatus as shown in the diagram.
- Measure and record the mass of **P** using an electronic balance, *M*.
- Add excess dilute acid to a conical flask containing **P**.
- Once no effervescence is seen, record the volume of carbon dioxide gas collected, *V*.
- Calculate the number of moles of carbon dioxide collected using Avogadro's gas law

$$\text{Moles}_{\text{CO}_2} = \frac{V}{24 \text{ dm}^3}$$

- From the equation,

$$\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

$$\text{X}_n\text{CO}_3 \Rightarrow \begin{matrix} \text{X} : \text{CO}_3^{2-} \\ n : 1 \end{matrix}$$

2 moles of hydrogen ions produces 1 mole of carbon dioxide, and requires 1 mole of carbonate ion.
Hence, calculating for the relative atomic mass of X,

$$A_r[\text{X}] = n \text{ Moles}_{\text{CO}_2}$$

3. Qualitative analysis

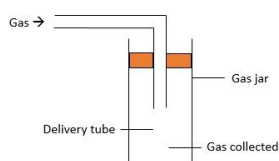
Solubility		Characteristics of gases		
Ions	Solubility	Gas	Density to air	Solubility
Group 1 cations NH_4^+ "Super-solubles"	Soluble	H_2 Hydrogen	Less dense	Insoluble
NO_3^-	Soluble	NH_3 Ammonia	Less dense	Very soluble
SO_4^{2-}	Soluble	CO Carbon monoxide	Slightly Less dense	Insoluble
Exc. Ca^{2+} $\text{Pb}^{2+/4+}$ Ba^{2+}		O_2 Oxygen	Slightly denser	Slightly soluble
Cl^- , Br^- , I^-	Soluble	CO_2 Carbon dioxide	Denser	Slightly soluble
Exc. $\text{Pb}^{2+/4+}$ Ag^+		Cl_2 Chlorine	Denser	Soluble
CO_3^{2-}	Insoluble	HCl Hydrochloric acid	Denser	Very Soluble
Exc. Group 1 NH_4^+		SO_2 Sulfur dioxide	Denser	Very Soluble

Note: Oxygen gas can be used as the estimate benchmark

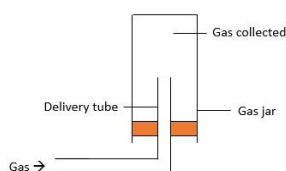
OH^- , O^{2-} Exc. Group 1 NH_4^+ Ca^{2+} Ba^{2+} (sparingly soluble)	Insoluble	to compare which gases are more dense than air.
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Collection of gas

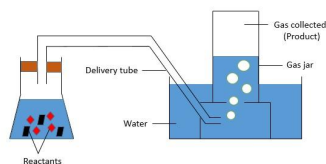
Denser gases (compared to air), downward delivery



Less dense gases (compared to air), upward delivery

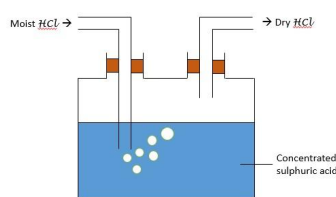


Insoluble gas, water displacement



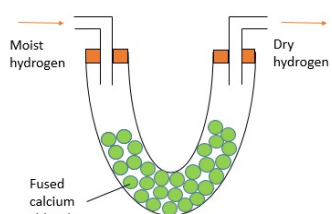
Drying gases

Concentrated Sulfuric acid



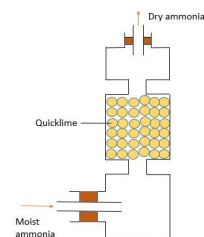
Sulfuric acid is acidic in nature, hence basic gases are not suited, s.a $NH_3(g)$

Fused Calcium Chloride



Fused Calcium Chloride can react with compounds, s.a $NH_3(g)$

Quicklime, CaO



Quicklime is basic in nature, hence acidic gases are not suited, s.a $CO_2(g)$

Identifying evolved gas

NH_3	$NH_3(g)$ has a pungent smell,	CO_2	$CO_2(g)$ is odourless,
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	colourless. (Alkaline) Place a piece of damp red litmus paper near the sample gas. If litmus paper turns blue, $NH_3(g)$ is present.		colourless. (Acidic) Downward delivery of gas into limewater, $Ca(OH)_2$. If white precipitate forms $CaCO_3$, $CO_2(g)$ is present. (Neutralisation) $CaCO_3$ dissolves in excess $CO_2(g)$
Cl_2	$Cl_2(g)$ has a pungent smell, greenish-yellow. Place a piece of damp red/blue litmus paper near the sample gas. If litmus paper turns white -bleached-, $Cl_2(g)$ is present.	H_2	$H_2(g)$ is odourless, colourless. Place a lighted splint near the sample gas. If splint extinguishes with a pop sound, $H_2(g)$ is present.
SO_2	$SO_2(g)$ has a pungent smell, colourless. Displace gas into purple acidified potassium permanganate $KMnO_4$ If acid becomes colourless, $SO_2(g)$ is present. (Reduction)	O_2	$O_2(g)$ is odourless, colourless. Place a glowing splint near the sample gas. If the splint relights, $O_2(g)$ is present.

Identifying water

Add anhydrous salt to solution 1. $CuSO_4$ 2. $CoCl_2$	If, 1. $CuSO_4 \cdot nH_2O$ Solution turns blue 2. $CoCl_2 \cdot nH_2O$ Solution turns purple
--	---

Identifying cation

Cation	+ $NaOH(aq)$	+ $NH_3(aq)$
NH_4^+	No precipitate Ammonia gas released on warming, leaving a colourless solution $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O(l)$ In excess, no reaction	No reaction
Ca^{2+}	White precipitate: - $Ca(OH)_2$ $Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$ $Ca(OH)_2$ does not dissolve in excess	No precipitate

Al^{3+}	White precipitate: - $Al(OH)_3$ $Al^{3+}(aq) + 3OH^{-}(aq) \rightarrow Al(OH)_3(s)$ $Al(OH)_3$ dissolves in excess giving colourless solution	White precipitate: - $Al(OH)_3$ $Al^{3+}(aq) + 3OH^{-}(aq) \rightarrow Al(OH)_3(s)$ $Al(OH)_3$ does not dissolve in excess
Pb^{2+}	White precipitate: - $Pb(OH)_2$ $Pb^{2+}(aq) + 2OH^{-}(aq) \rightarrow Pb(OH)_2(s)$ $Pb(OH)_2$ dissolves in excess giving colourless solution	White precipitate: - $Pb(OH)_2$ $Pb^{2+}(aq) + 2OH^{-}(aq) \rightarrow Pb(OH)_2(s)$ $Pb(OH)_2$ does not dissolve in excess
Zn^{2+}	White precipitate: - $Zn(OH)_2$ $Zn^{2+}(aq) + 2OH^{-}(aq) \rightarrow Zn(OH)_2(s)$ $Zn(OH)_2$ dissolves in excess giving colourless solution	White precipitate: - $Zn(OH)_2$ $Zn^{2+}(aq) + 2OH^{-}(aq) \rightarrow Zn(OH)_2(s)$ $Zn(OH)_2$ dissolves in excess giving colourless solution
Cu^{2+}	Light Blue precipitate: - $Cu(OH)_2$ $Cu^{+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$ $Cu(OH)_2$ does not dissolve in excess	Light Blue precipitate: - $Cu(OH)_2$ $Cu^{+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$ $Cu(OH)_2$ dissolves in excess giving a dark blue solution
Fe^{2+}	Green precipitate: - $Fe(OH)_2$ $Fe^{+}(aq) + 2OH^{-}(aq) \rightarrow Fe(OH)_2(s)$ $Fe(OH)_2$ does not dissolve in excess	Green precipitate: - $Fe(OH)_2$ $Fe^{+}(aq) + 2OH^{-}(aq) \rightarrow Fe(OH)_2(s)$ $Fe(OH)_2$ does not dissolve in excess
Fe^{3+}	Reddish-brown precipitate: - $Fe(OH)_3$ $Fe^{+}(aq) + 2OH^{-}(aq) \rightarrow Fe(OH)_3(s)$ $Fe(OH)_3$ does not dissolve in excess	Reddish-brown precipitate: - $Fe(OH)_3$ $Fe^{+}(aq) + 2OH^{-}(aq) \rightarrow Fe(OH)_3(s)$ $Fe(OH)_3$ does not dissolve in excess
Group 1	No precipitate	No precipitate

Cation	Test to discern
Al^{3+}	$Al^{3+} + 3I^{-} \rightarrow AlI_3(s)$

	Aluminium ions react with iodide ions to form a white precipitate, aluminium iodide
Pb^{2+}	$Pb^{2+} + 2I^{-} \rightarrow PbI_2$ Lead ions react with iodide ions to form a yellow precipitate, lead iodide.
$Fe(OH)_2$ oxidises over time to become a reddish-brown precipitate, $Fe(OH)_3$	

Identifying anion

Anions	Test	Confirmation of test
CO_3^{2-}	+ $HCl(aq)$ Effervescence seen: CO_2 evolved	Pass gas released through limewater, $Ca(OH)_2$. If white precipitate forms, limestone $CaCO_3(s)$, $CO_2(g)$ is present. $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$
Cl^{-}	Acidify with: $HNO_3(aq)$ then + $AgNO_3(aq)$ White precipitate seen: $AgCl(s)$	Acidification should be done sufficiently to ensure removal of other ions, such as CO_3 which also can form a white precipitate with Ag^{+} . $Cl^{-}(aq) + H^{+}(aq) \rightarrow HCl(aq)$ $Cl^{-}(aq) + Ag^{+}(aq) \rightarrow AgCl(s)$
I^{-}	Acidify with: $HNO_3(aq)$ then + $AgNO_3(aq)$ Pale yellow precipitate seen: $AgI(s)$ Acidify with: $HNO_3(aq)$ then + $Pb(NO_3)_2(aq)$ Pale yellow precipitate seen: $PbI_2(s)$	Acidification should be done sufficiently to ensure removal of other ions, such as CO_3^{2-} which also can form a yellow precipitate with Al^{+} and Pb^{2+} . $I^{-}(aq) + H^{+}(aq) \rightarrow HI(aq)$ $I^{-}(aq) + Ag^{+}(aq) \rightarrow AgI(s)$ $I^{-}(aq) + H^{+}(aq) \rightarrow HI(aq)$ $2I^{-}(aq) + Pb^{2+}(aq) \rightarrow PbI_2(s)$ Note: the yellow colour may be faint, use a whiter colour to discern.
SO_4^{2-}	Acidify with: $HNO_3(aq)$ then + $Ba(NO_3)_2(aq)$ White precipitate seen: $BaSO_4$	Acidification should be done sufficiently to ensure removal of other ions, such as CO_3^{2-} which also can form a white precipitate with Ba^{2+} .
NO_3^{-}	+ $NaOH(aq)$ then + $Al(s)$ foil, warm carefully Effervescence seen: $NH_3(g)$ evolved	Placing a damp red litmus paper near the gas sample. If litmus turns blue, NO_3^{-} is present. Same observation will be noted if $NH_4^{+}(g)$ was present in initial sample.

		- Hence, to ensure confirmation of nitrate, add excess $NaOH$ and heat to remove the $NH_3(g)$ due to reaction with $NH_4^+(g)$, before adding the aluminium foil. If observations are the same, NO_3^- is present.
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Note: when describing observed effects, use “dissolve” to describe ppt. Disappearing, “turn”/“giving” to describe colour change, “solution” as long as it has water, “mixture” when it is a mixture.

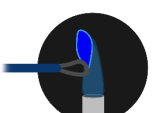
Identifying reducing agent

Oxidising agent	Added reducing agent	
$KMnO_4$ Acidified potassium manganate(VII)	Purple → Colourless +7 to +2	Manganate(VII) ions are reduced from purple MnO_4^- to Mn^{2+} , causing the solution to decolourise.
$K_2Cr_2O_7$ Acidified potassium dichromate(VI)	Orange → Green $Cr: +6 \rightarrow +3$	
$FeCl_3$ Fresh iron(III) chloride solution	Reddish brown → Green	
I_2 Iodine solution	Brown → Colourless	
H_2O_2 Hydrogen peroxide	No visible changes	

Identifying oxidising agent

Reducing agent	Added oxidising agent
KI Acidified potassium iodide solution	Colourless → Brown I_2 confirmed by adding starch, which turns dark blue
$FeSO_4$ Acidified iron sulfate solution	Green → Yellow Fe^{3+} cation can be confirmed by adding $NaOH$, which forms a reddish-brown precipitate.
HCl Concentrated hydrochloric acid	$Cl_2(g)$ evolved
Fe^{2+} Iron(II) solution	Green precipitate turns brown

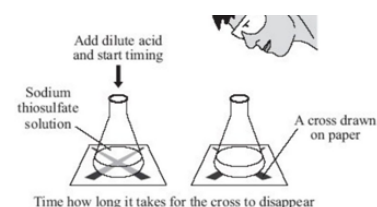
Identifying ion based on flame colour

Group 1 Metals			Group 2 Metals		
					
Lithium, Li^+ RED	Sodium, Na^+ YELLOW-ORANGE	Potassium, K^+ LILAC	Calcium, Ca^{2+} YELLOW-RED	Strontium, Sr^{2+} RED	Barium, Ba^{2+} PALE GREEN
					
Rubidium, Rb^+ RED	Caesium, Cs^+ BLUE		Magnesium, Mg^{2+} NO FLAME COLOUR	Beryllium, Be^{2+} NO FLAME COLOUR	

4. Rate of reaction

Determining rate of reaction

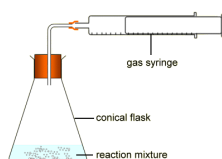
Disappearing cross experiment



Experimental procedure

1. Fill a burette with sodium thiosulfate
2. Add 25.0cm^3 of sodium thiosulfate solution into the flask.
3. Fill another burette with hydrochloric acid.
4. Add 25.0cm^3 of hydrochloric acid to a measuring cylinder from the burette.
5. Add the hydrochloric acid to the flask and start the stopwatch immediately.
6. Swirl the flask once to mix the solutions.
7. Look vertically down through the solution at the cross.
8. Stop the stopwatch when the cross can no longer be seen. Record the time taken, t .
9. Repeat the experiment using different concentrations of sodium thiosulfate or hydrochloric acid, ensuring that the total volume is 50.0cm^3 .

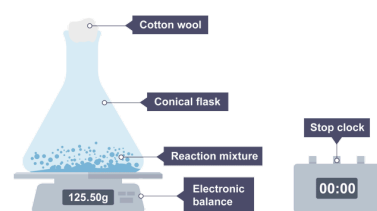
Rate of gas evolution



Experimental procedure

1. Set up the apparatus as shown in the diagram.
2. Add the reaction mixture into the conical flask and quickly seal the conical flask.
3. Connect the conical flask to a gas syringe.
4. Measure and record the volume of gas at discrete intervals of time.

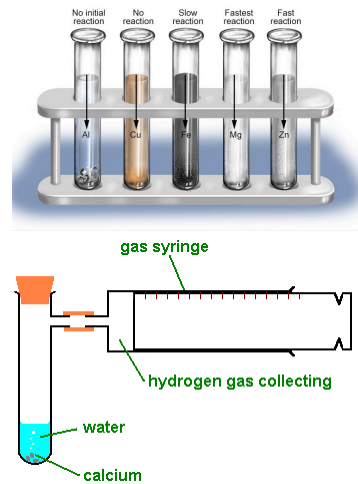
Rate of mass change



Experimental procedure

1. Set up the apparatus as shown in the diagram.
2. Add the reaction mixture into the conical flask and seal the conical flask with the cotton wool.
3. Measure and record the mass of the reaction mixture at discrete intervals of time.

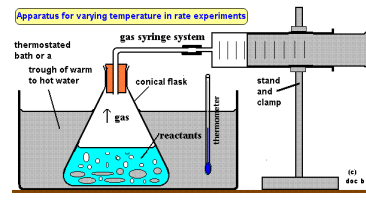
Using rate of reaction to determine reactivity of metal



Experimental procedure

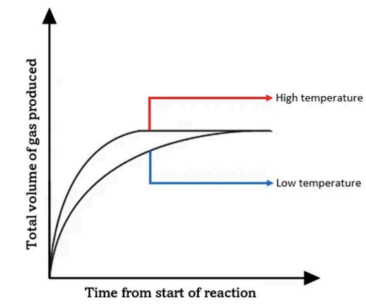
1. Pipette 25.0cm^3 of excess of hydrochloric acid, of equal concentration, into each test tube.
2. Prepare the sample metals, ensuring that they have roughly equal surface area.
3. Add each metal to its respective test tube. Immediately connect the test tube to a gas syringe and start the stopwatch.
4. Once the volume measured on the gas syringe is greater than 50.0cm^3 , stop the stopwatch.
5. Measure and record the time taken on the stopwatch, t .
6. Tabulate the results and compare the time taken for each metal to produce 50.0cm^3 of hydrogen gas.

Effect of changing temperature on the rate of reaction



Experimental procedure

1. Set up the apparatus as shown in the diagram.
2. Set the water bath to a fixed temperature.
3. Measure and record the temperature using a thermometer, T .
4. Add the reaction mixture into the conical flask. Immediately connect the test tube to a gas syringe and start the stopwatch.
5. Measure and record the volume of gas produced at even time intervals, V and Δt respectively.
6. Repeat steps 2 to 5 for different initial temperatures T in increasing temperatures of 5.0°C , starting above room temperature, i.e. 35.0°C , 40.0°C , etc.
7. Tabulate the results.



Graphical analysis

- Plot the volume of gas produced against the time from the start of reaction.
- As reaction progresses, the gradient of the graph decreases.
- As the temperature increases, the initial gradient of the graph increases.

Explanation, gradual decrease in gradient as reaction progresses

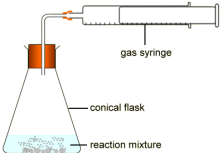
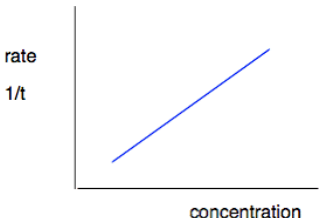
- As reaction progress, more of the molecules are used up in reaction. Thus the number of reactant molecules decreases, and the concentration of the reactant molecules decreases.
- Since there is less number of reactant molecules per unit volume, the effective collision frequency decreases, which thereby causes the rate of reaction to decrease.
- Hence, this causes the gradient to become less steep.

Explanation, temperature and gradient relation

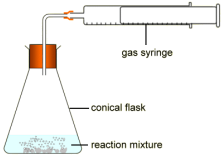
- As the temperature increases, the average kinetic energy of the reactant molecules increases.
- Thus, the reactant molecules are moving at a higher speed, and possess more energy.
- Hence, more molecules collide frequently, and more molecules possess energy greater or equal to activation.

	Therefore, the effective collision frequency increases, causing the rate of reaction to increase. - This causes the gradient to become more steep.
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Effect of changing concentration of reactant on the rate of reaction

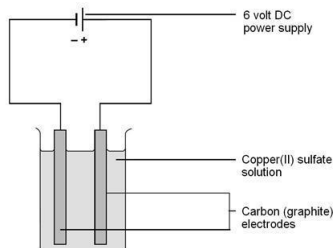
Effect of changing concentration of reactant on the rate of reaction.																															
	Experimental procedure 1. Set up the apparatus as shown in the diagram. 2. Fill a burette with P. 3. Fill another burette with deionised water. 4. Add 25.0cm³ of P to the conical flask from the first burette. 5. Add 5.0cm³ of deionised water to a measuring cylinder from the second burette. 6. Pour all the contents from the measuring cylinder into the conical flask. 7. Pour excess Q into the conical flask. Immediately connect the test tube to a gas syringe and start the stopwatch. 8. Once the volume measured on the gas syringe is greater than 50.0cm³, stop the stopwatch. 9. Measure and record the time taken on the stopwatch, <i>t</i>. 10. Calculate the average rate of reaction, $\frac{1}{t}$. 11. Calculate the concentration of P, $C'[P] = \frac{Vol.P \times C[P]}{Vol.P \times C[P] + Vol.water}$ 12. Repeat steps 4 to 7 for different volumes of P and deionised water, ensuring that the total volume adds up to 30.0cm³. 13. Tabulate the results Note: Q has known constant concentration and is reacting with P Example of a table <table><tr><th>Exp.</th><th>Vol.P/cm³</th><th>Vol.water/cm³</th><th>Conc.P/mol dm⁻³</th><th><i>t</i>/s</th><th>$\frac{1}{t}$/s⁻¹</th></tr><tr><td>1</td><td>30.00</td><td>0.00</td><td><i>C</i>₁</td><td><i>t</i>₁</td><td><i>t</i>₁⁻¹</td></tr><tr><td>2</td><td>25.00</td><td>5.00</td><td><i>C</i>₂</td><td><i>t</i>₂</td><td><i>t</i>₂⁻¹</td></tr><tr><td>⋮</td><td>⋮</td><td>⋮</td><td>⋮</td><td>⋮</td><td>⋮</td></tr><tr><td>7</td><td>0.00</td><td>30.00</td><td><i>C</i>₇</td><td><i>t</i>₇</td><td><i>t</i>₇⁻¹</td></tr></table>	Exp.	Vol. P /cm ³	Vol.water/cm ³	Conc. P /mol dm ⁻³	<i>t</i> /s	$\frac{1}{t}$ /s ⁻¹	1	30.00	0.00	<i>C</i> ₁	<i>t</i> ₁	<i>t</i> ₁ ⁻¹	2	25.00	5.00	<i>C</i> ₂	<i>t</i> ₂	<i>t</i> ₂ ⁻¹	⋮	⋮	⋮	⋮	⋮	⋮	7	0.00	30.00	<i>C</i> ₇	<i>t</i> ₇	<i>t</i> ₇ ⁻¹
Exp.	Vol. P /cm ³	Vol.water/cm ³	Conc. P /mol dm ⁻³	<i>t</i> /s	$\frac{1}{t}$ /s ⁻¹																										
1	30.00	0.00	<i>C</i> ₁	<i>t</i> ₁	<i>t</i> ₁ ⁻¹																										
2	25.00	5.00	<i>C</i> ₂	<i>t</i> ₂	<i>t</i> ₂ ⁻¹																										
⋮	⋮	⋮	⋮	⋮	⋮																										
7	0.00	30.00	<i>C</i> ₇	<i>t</i> ₇	<i>t</i> ₇ ⁻¹																										
	Graphical analysis - Plot the average rate of reaction against the concentration. - As the concentration increases, the rate of reaction increases. Explanation, concentration and rate of reaction - As concentration increases, there is more number of reactant molecules per unit volume, the effective collision frequency increases, which thereby causes the rate of reaction to increase.																														

Effect of presence of catalyst on the rate of reaction

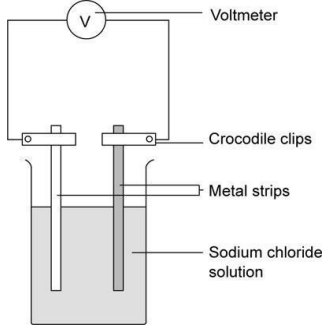
	Experiment procedure: 1. Set up the apparatus as shown in the diagram. 2. Add the reaction mixture into the conical flask without the catalyst. Immediately connect the test tube to a gas syringe and start the stopwatch. 3. Once the volume measured on the gas syringe is greater than 4. 50.0 cm^3, stop the stopwatch. 5. Measure and record the time taken on the stopwatch, t. 6. Repeat steps 2 to 5, but include the catalyst in step 2, and record the new time taken, t'. Explanation, catalyst on the rate of reaction - The catalyst provides an alternative path of lower activation energy, thus more particles possess energy greater or equal to activation. Hence, the effective collision frequency increases, and rate of reaction increases.
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5. Electrolysis

Electroplating

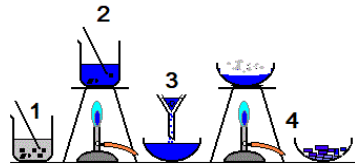
	Experimental procedure 1. Set up the apparatus as shown in the diagram. 2. Close the circuit. Observation: - A grey($Zn/Fe/Ag$)/Red-brown(Cu) deposit is seen on the surface of the cathode. - [If the anode is the element deposited] • The size of the anode gradually decreases, dissolving into the electrolyte - [If the anode is not the element deposited, and element is Cu] • The blue colour of the solution gradually fades. Explanation, Copper deposit on cathode: - Copper(II) ions preferentially discharge at the cathode, and reduces to become Copper. The accumulation of the Copper forms the red-brown deposit. Explanation, reactive Copper anode: - At the anode, the copper anode oxidises to become Copper(II) ions, causing the anode to dissolve into the electrolyte. - This replenishes the number of moles of Copper(II) ions in the electrolyte, Copper(II) ions would be continuously preferentially discharged at the cathode.
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Using Electrochemical Cell (Galvanic Cell) set up to determine reactivity

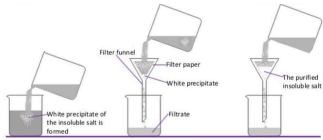
 Voltmeter Crocodile clips Metal strips Sodium chloride solution	Experimental procedure 1. Set up the apparatus as shown in the diagram. 2. Connect the metal, P, to one of the crocodile clips; connected to the positive terminal of the voltmeter. 3. Connect one of the sample metals that is being compared to the other crocodile clip; connected to the negative terminal of the voltmeter. 4. Submerge both metals into the electrolyte. 5. Measure and record the reading on the voltmeter, <i>V</i>. - If the reading is positive, P is more reactive than the sample metal. - If the reading is negative, P is less reactive than the sample metal. 6. Take the two metals out of the electrolyte. 7. Repeat steps 3 to 6 for different sample metals. 8. Tabulate the results and compare the voltage produced. - The more positive the value, the more reactive the metal is relative to metal, P. - The more negative the value, the less reactive the metal is relative to metal, P. Explanation: - The more reactive metal is more readily oxidised and loses electrons: $M \rightarrow M^{n+} + ne^{-}$ - These electrons flow through the external wire from the more reactive metal to the less reactive metal, producing a potential difference between the two metals. - Therefore, the more reactive metal acts as the negative electrode, while the less reactive metal acts as the positive electrode. - Since metal P is connected to the positive terminal of the voltmeter, a negative reading means that P is the negative electrode and P is therefore more reactive than the sample metal. - Conversely, a positive reading means that P is the positive electrode and P is therefore less reactive than the sample metal.
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6.Preparation of salts

Acid to insoluble base

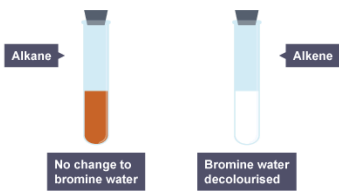
	Experimental procedure 1. Warm the acid, Q. 2. Add the insoluble base, P, in small portions with stirring. 3. Continue adding more P until no more P can dissolve. 4. Filter to remove excess solid P. Perform crystallisation to collect crystallised solution 5. Heat the solution until saturated. 6. Allow the saturated solution to cool and to form crystals. 7. Rinse the crystals with a little cold deionised water. 8. Dry the crystals between a few pieces of filter paper.
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Precipitation

	Experimental procedure 1. Pipette 25.0 cm^3 of P to a beaker. 2. Add excess Q to the beaker with stirring. 3. Filter the mixture using filter funnel lined with filter paper, obtaining precipitate as residue. 4. Wash residue with deionised water, removing any soluble impurities. 5. Dry the precipitate between a few pieces of filter paper.
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7. Unsaturated hydrogen

Bromination

	Experimental procedure 1. Prepare a clean test tube. 2. Pour about 1 cm depth of the unknown hydrocarbon, P, into the clean test tube. 3. Add a few drops of aqueous bromide solution to the test tube. 4. Observe the colour change: - If the reddish-brown solution rapidly decolourises, the hydrocarbon is unsaturated. - If the reddish-brown solution retains its colour, the hydrocarbon is saturated. Explanation: - The $\text{C}=\text{C}$ bonds are highly reactive and reacts with bromide.
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