
There is no practical exam: Papers 1 and 2 ask you to describe methods, record observations, process results and suggest improvements for the practicals named in the specification. Learn each method as numbered steps, and each observation as what you SEE.

Apparatus and readings

Measuring cylinder – read the bottom of the meniscus at eye level; good enough for a reagent added in excess, not for an exact volume

Pipette and burette – for exact volumes in a titration. Read a burette to the nearest 0.05 cm^3 , and record both readings to two decimal places

Gas syringe – collects a gas to the nearest 1 cm^3 ; an upturned measuring cylinder over water works too, but some carbon dioxide dissolves in the water, so the syringe is better for it

Balance – reads to 0.01 g . Zero it with the empty container on it, or weigh before and after and subtract

Thermometer – stir, then read the highest (or lowest) temperature reached; a polystyrene cup with a lid reduces heat transfer to the surroundings

Observations are what you see, hear or smell: ‘bubbles’, ‘the solid disappears’, ‘a white precipitate’. ‘Hydrogen is given off’ is a deduction, not an observation.

Making salts

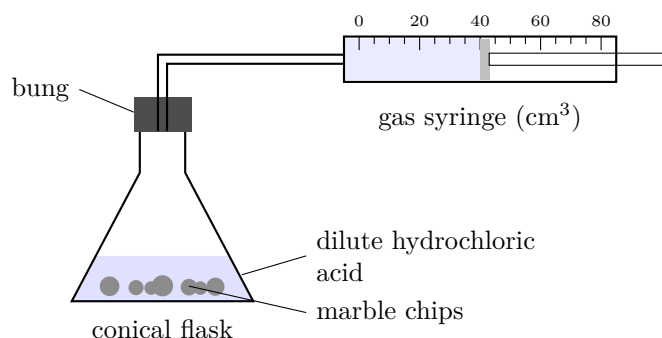
Soluble salt from an insoluble base – e.g. hydrated copper(II) sulfate from copper(II) oxide: warm the dilute sulfuric acid, add the oxide until some is left unreacted (so all the acid is used up), filter off the excess, heat the filtrate to its crystallisation point, leave to cool and crystallise, then pat the crystals dry with filter paper

Soluble salt from an alkali – both reactants are solutions, so neither can be added in excess and filtered off: titrate with an indicator to find the volumes, then mix the same volumes again WITHOUT the indicator and crystallise

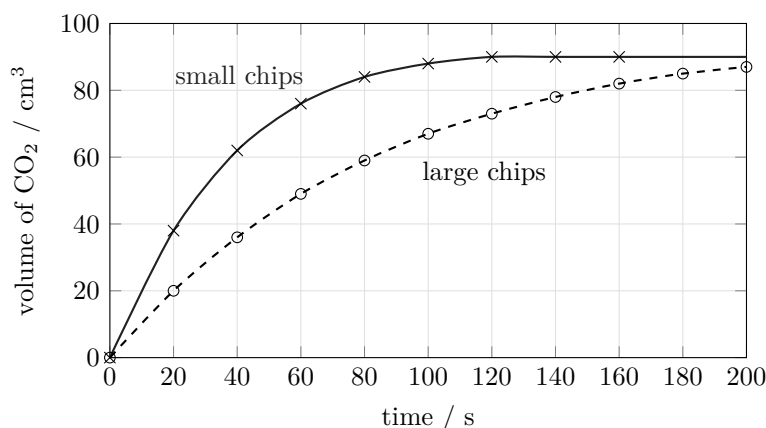
Insoluble salt by precipitation – e.g. lead(II) sulfate: mix solutions of lead(II) nitrate and sodium sulfate, filter, wash the precipitate on the filter paper with distilled water, and dry it in a warm oven

Worked analysis: rate of reaction with marble chips :

Excess marble chips (calcium carbonate) were added to 50 cm³ of dilute hydrochloric acid in a flask connected to a gas syringe, and the volume of carbon dioxide was read every 20 s. The run was repeated with the same mass of smaller chips.



The bung goes in as soon as the chips are added; the gas pushes the plunger out, and the volume is read from the scale on the barrel.



Both runs use the same acid, so both end at the same volume (90 cm³); the small chips get there sooner.

Describe: the small chips gave a steeper curve at the start, so a faster rate; both curves level off as the acid is used up.

Explain: smaller chips have a larger surface area for the same mass, so more particles are exposed and there are more frequent collisions with acid particles.

Mean rate over the first 20 s with small chips = $38 \text{ cm}^3 \div 20 \text{ s} = 1.9 \text{ cm}^3/\text{s}$. The initial rate is the gradient of the curve at the start.

Controlled variables: mass of chips, volume, concentration and temperature of the acid. The same mass of chips in excess, and the same acid, make the final volumes equal – check the graph agrees.

Worked analysis: temperature change on neutralisation :

25.0 cm³ of 2.0 mol/dm³ hydrochloric acid was put in a polystyrene cup and its temperature measured. 25.0 cm³ of 2.0 mol/dm³ sodium hydroxide solution at the same temperature was added and stirred; the highest temperature was 13.0 °C above the starting temperature.

1. Heat energy change $Q = m \times c \times \Delta T = 50.0 \text{ g} \times 4.2 \text{ J/g } ^\circ\text{C} \times 13.0 \text{ } ^\circ\text{C} = 2730 \text{ J}$ (1 cm³ of solution taken as 1 g).
2. Moles of water formed = moles of HCl = $2.0 \times 25.0 \div 1000 = 0.050 \text{ mol}$.

3. $\Delta H = -Q \div \text{moles} = -2730 \text{ J} \div 0.050 \text{ mol} = -54\,600 \text{ J/mol} = -55 \text{ kJ/mol}$. Negative: the reaction is exothermic.

The data-book value is about -57 kJ/mol . Heat lost to the cup, the thermometer and the air makes ΔT , and so the size of ΔH , too small. Improvements: a lid, a second cup as insulation, and a thermometer reading to 0.1°C .

Worked calculation: the formula of magnesium oxide :

Magnesium ribbon was heated in a lidded crucible, lifting the lid now and then to let air in, until it no longer glowed; it was heated to constant mass.

	mass / g
crucible and lid	25.40
crucible, lid and magnesium	25.64
crucible, lid and magnesium oxide	25.80

1. Mass of magnesium = $25.64 - 25.40 = 0.24 \text{ g}$; mass of oxygen = $25.80 - 25.64 = 0.16 \text{ g}$.
2. Moles: Mg $0.24 \div 24 = 0.010$; O $0.16 \div 16 = 0.010$.
3. Ratio 1 : 1, so the empirical formula is MgO .

Errors: white smoke (magnesium oxide) escaping when the lid is lifted makes the oxygen mass too small; magnesium that has not all reacted does the same. Heating to constant mass checks the reaction is complete.

The specification's practicals – what each one needs

Paper chromatography – start line in PENCIL (ink would run), spots small, solvent below the line, lid on. A pure substance gives one spot; $R_f = \text{distance moved by the spot} \div \text{distance moved by the solvent}$, both from the start line

Oxygen in air – pass 100 cm^3 of air between two gas syringes over heated copper until the volume stops falling (or leave damp iron filings in a tube over water); about 21 cm^3 is used up. Let the apparatus cool before the final reading

Metals with dilute acids – compare the rate of bubbling of magnesium, zinc and iron with the same acid; hydrogen gives a squeaky pop with a lighted splint

Conditions for rusting – iron nails in three tubes: water and air; boiled water under a layer of oil (no air); air dried with anhydrous calcium chloride (no water). Only the first rusts

Electrolysis of aqueous solutions – inert electrodes; test the gas at each electrode. At the cathode, copper or hydrogen; at the anode, oxygen, or chlorine (bleaches damp litmus paper) from a concentrated chloride. Do it in a fume cupboard if chlorine forms

Acid–alkali titration – pipette the alkali into a conical flask with a few drops of indicator, add acid from the burette, swirling, dropwise near the end point. Repeat until two titres agree within 0.20 cm^3 and average those

Tests for ions – flame tests (Li^+ red, Na^+ yellow, K^+ lilac, Ca^{2+} orange-red, Cu^{2+} blue-green); sodium hydroxide solution for metal ions (Cu^{2+} blue, Fe^{2+} green, Fe^{3+} brown precipitate; NH_4^+ gives ammonia on warming); halides with nitric acid then silver nitrate; sulfate with hydrochloric acid then barium chloride; carbonate fizzes with acid, giving CO_2 that turns limewater milky

Catalysts for hydrogen peroxide

– the same mass of each solid, the same volume and concentration of hydrogen peroxide; measure the oxygen in a gas syringe over time. A catalyst can be filtered off, dried and weighed afterwards: its mass is unchanged

Temperature changes

– dissolving salts, neutralisation and displacement in a polystyrene cup; combustion of a fuel heating water in a copper can. The fuel's ΔH comes out far too small, because most heat goes to the air

What examiners keep reporting

An observation written as a name: 'carbon dioxide is produced' instead of 'fizzing' or 'bubbles'.

A colour change with one colour: write both, from $\rightarrow$ to – 'blue solution turns colourless'.

'Heat the solution until all the water evaporates' for a hydrated salt. Heat only to the crystallisation point, then leave it to cool: dry heating removes the water of crystallisation.

'Add the acid' with no volume or concentration. A method names how much, of what, measured with what.