

8 • Chemical analysis

AQA • GCSE • Chemistry

Chemical analysis: choosing evidence

A bottle may say “pure milk”, yet milk contains several substances. A green flame may suggest copper, yet a second colour can be hidden in a mixture. **Chemical analysis** combines a suitable test, a recorded observation and a justified conclusion. An observation is what you see; a conclusion is what that observation supports.

This topic covers specification 4.8. Both tiers study purity, formulations, chromatography and the four gas tests. All ion tests and instrumental methods here are **Separate Chemistry, both tiers**, including flame emission spectroscopy. Required practicals 6 and 7 need actual supervised practical work and recorded results; paper examples prepare you but do not certify practical completion.

Purity and designed mixtures (4.8.1.1–4.8.1.2)

A **pure substance** is a single element or compound, not mixed with another substance. Pure oxygen is an element; pure water is a compound, despite containing two chemically combined elements. Milk with nothing added is still a mixture. “Natural” and “chemically pure” answer different questions.

Pure elements and compounds melt and boil at specific temperatures under stated conditions. Compare a measured melting interval with a reference using calibrated apparatus and comparable conditions. A mixture often melts over a wider interval, and an impurity often lowers a solid’s melting point. Do not claim that every mixture has the same temperature pattern or that every impurity changes a boiling point in the same direction; pressure also affects boiling temperature.

Worked evidence. A reference melts at 80.0 °C. Sample A melts from 79.8 to 80.2 °C; B from 72.0 to 78.0 °C using the same apparatus. A’s narrow interval near the reference supports greater purity. B’s lower, broader interval supports impurity. A’s matching interval alone does not uniquely prove identity; another substance could have similar data.

A **formulation** is a mixture designed as a useful product. Each component has a purpose and its amount is carefully measured to achieve the required properties. Paint can contain pigment for colour and a binder to hold it on a surface. Fuels, cleaning agents, paints, medicines, alloys, fertilisers and foods can be formulations. Use supplied product information: there is no need to memorise proprietary ingredients. A mixture produced accidentally is not automatically a formulation.

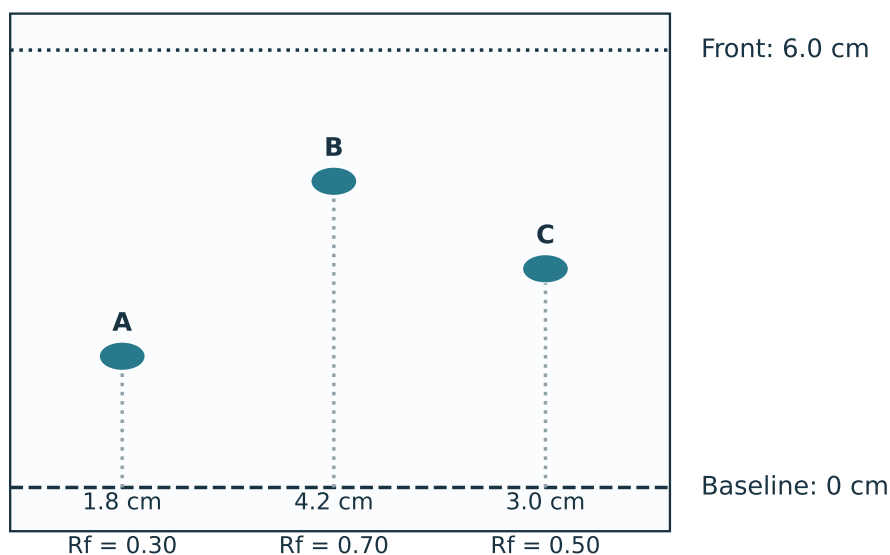
Chromatography and a complete Rf method (4.8.1.3)

Chromatography separates a mixture because its substances distribute differently between a **stationary phase** and a **mobile phase**. In this GCSE paper model, paper is stationary and solvent moves through it. Substances carried more readily by the solvent travel farther; substances held more strongly by the stationary phase travel less. Travel distance is not decided by particle mass alone. For required practical 6, use approved coloured substances and solvent under school supervision. Draw a pencil **baseline** above the solvent level; ink could dissolve and create extra spots. Place small concentrated sample/reference spots on the line and let them dry. Stand the paper in the solvent with the spots above its surface. Let solvent rise; remove the paper before the front reaches the top and mark the **solvent front** immediately, before evaporation hides it. Record the chromatogram and actual distances.

$$R_f = \frac{\text{distance from baseline to centre of spot}}{\text{distance from baseline to solvent front}}$$

Both distances use the same baseline and unit. **Rf value** has no unit. Measure to the centre of a spot, not its edge or from the bottom of the paper. A spot cannot travel beyond the solvent front in this model, so a value above 1 flags a measurement or division error.

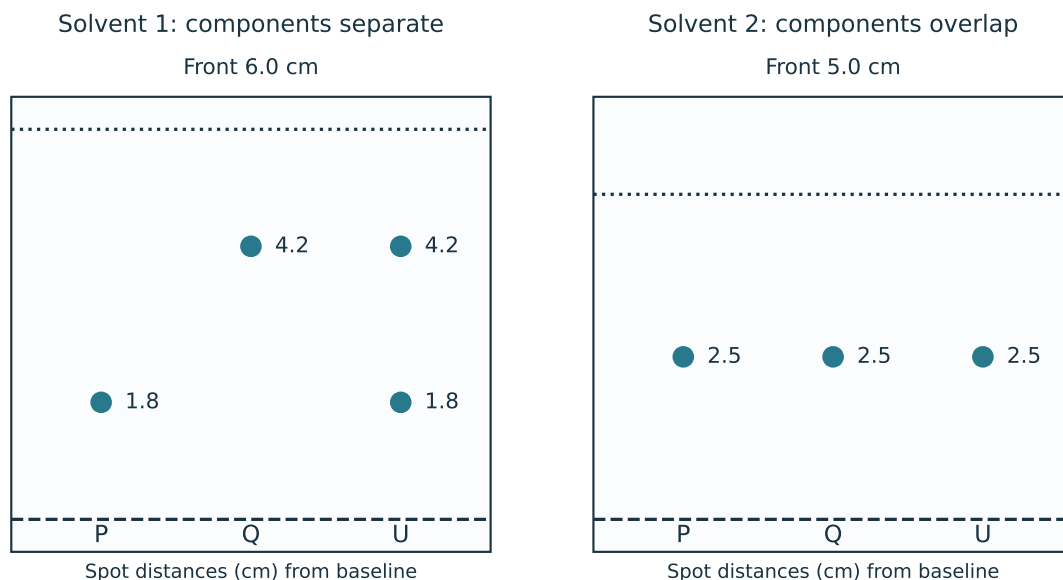
Distances measured from baseline to spot centre



Worked calculation. A spot travels 3.6 cm and the solvent front travels 6.0 cm:

$$R_f = \frac{3.6 \text{ cm}}{6.0 \text{ cm}} = 0.60$$

A matching spot/R_f with a known reference under the **same conditions** supports identification. Changing solvent can change R_f and whether substances separate. Two spots show at least two detected components. A pure compound gives one spot in all suitable solvents, but the reverse inference is limited: one observed spot can contain two substances that travel together, or miss an undetected component. Use another solvent or another test when the evidence is insufficient.



Transfer. In solvent 1, P has R_f 0.30, Q 0.70, and unknown U has both spots. In solvent 2, P and Q each have R_f 0.50, and U shows one spot at 0.50. Solvent 1 supports P and Q in U; solvent 2's single spot does not overturn that evidence or prove purity. These are fictional reference substances used to teach interpretation.

Four gas tests: pair the action and observation (4.8.2.1–4.8.2.4)

Gas	Test	Positive observation
hydrogen	Hold a burning splint at the open end of a test tube of gas.	A pop sound, often called a squeaky pop.
oxygen	Insert a glowing splint into the gas.	The glowing splint relights.
carbon dioxide	Bubble through or shake with limewater, aqueous calcium hydroxide.	Limewater turns milky/cloudy.
chlorine	Expose damp litmus paper to the gas in the approved setup.	Litmus is bleached white.

For oxygen, a splint that is already burning cannot demonstrate **relighting**. Let the flame go out while retaining a glow, then use the specified test. Do not claim this observation proves oxygen uniquely among every possible gas; it identifies oxygen

among the four gases in this syllabus context.

Carbon dioxide forms a white solid in limewater. The required positive observation is cloudiness; clearing on prolonged addition is not needed to identify it in this test. Chlorine gas and chloride ions need different tests. Litmus turning red alone is not the specified chlorine result: bleaching white is the key observation.

Worked decision. An unknown relights a glowing splint. Conclude oxygen from the positive test, rather than guessing from the reaction that produced it. A failed pop test alone does not identify which of the other gases is present. Use separate samples and the appropriate positive test. Actual gas tests use small approved samples and teacher-controlled apparatus; chlorine is handled only in the school's approved contained setup, never generated as an unsupervised exercise.

Metal ions: flames, precipitates and excess reagent (4.8.3.1–4.8.3.2)

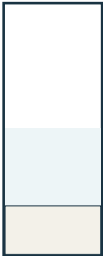
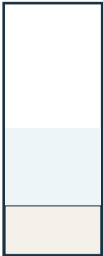
For a **flame test** , use the school's approved clean wire/reference procedure. Compare unknown and known samples, cleaning between tests to prevent contamination. Record the observed colour before naming the ion. A mixture can mask a colour, so an absent visible colour does not establish an absent ion.

Metal ion	Flame colour
lithium	crimson
sodium	yellow
potassium	lilac
calcium	orange-red
copper	green

Add sodium hydroxide solution to a fresh portion to look for an insoluble hydroxide **precipitate** . Then add excess sodium hydroxide when the white-precipitate distinction is needed.

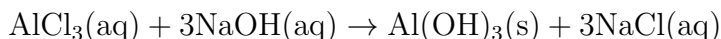
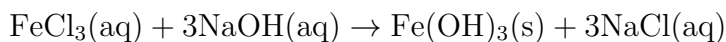
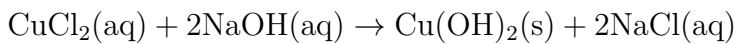
Cation	Initial hydroxide observation	In excess sodium hydroxide
Al^{3+}	white precipitate	dissolves
Ca^{2+}	white precipitate	remains
Mg^{2+}	white precipitate	remains
Cu^{2+}	blue precipitate	remains
Fe^{2+}	green precipitate	remains
Fe^{3+}	brown precipitate	remains

NaOH: initial precipitate colour; label states behaviour in excess

Al^{3+}	Ca^{2+}	Mg^{2+}	Cu^{2+}	Fe^{2+}	Fe^{3+}
					
Dissolves	Remains	Remains	Remains	Remains	Remains

Worked identification. A white precipitate alone leaves Al^{3+} , Ca^{2+} and Mg^{2+} possible. If it dissolves in excess NaOH, identify aluminium among these listed ions. If it remains, calcium and magnesium are not distinguished by this test alone; an orange-red flame supports calcium. Do not confuse copper's green flame with its blue hydroxide precipitate, or iron(II)'s green precipitate with iron(III)'s brown one.

Write balanced **full symbol equations** for the specified insoluble hydroxides. Balance the hydroxide groups and salt products; brackets show that each metal can need more than one OH group:



For CaCl_2 , MgCl_2 or FeCl_2 , the corresponding $\text{M}(\text{OH})_2$ needs 2NaOH and forms 2NaCl . These formation equations are both-tier knowledge here. An equation for the aluminium precipitate dissolving to sodium aluminate is not required. Ionic/half equations retain their separate Higher Tier scope elsewhere.

Anions: choose the right acid and fresh sample (4.8.3.3–4.8.3.5)

Ion	Reagents and procedure	Positive evidence
carbonate	Add dilute acid; test the gas with limewater.	Effervescence and gas turning limewater milky support carbonate.
chloride	Dilute nitric acid, then silver nitrate solution.	White silver chloride precipitate.
bromide	Dilute nitric acid, then silver nitrate solution.	Cream silver bromide precipitate.
iodide	Dilute nitric acid, then silver nitrate solution.	Yellow silver iodide precipitate.
sulfate	Dilute hydrochloric acid, then barium chloride solution.	White barium sulfate precipitate.

Use **fresh sample portions** for different tests, so a previous reagent does not add an ion that creates a false result. In the halide test, use nitric acid rather than hydrochloric acid, which would introduce chloride. Do not swap the acid/reagent pairs between the halide and sulfate tests. “A white precipitate” without its reagent is insufficient to distinguish these several possibilities.

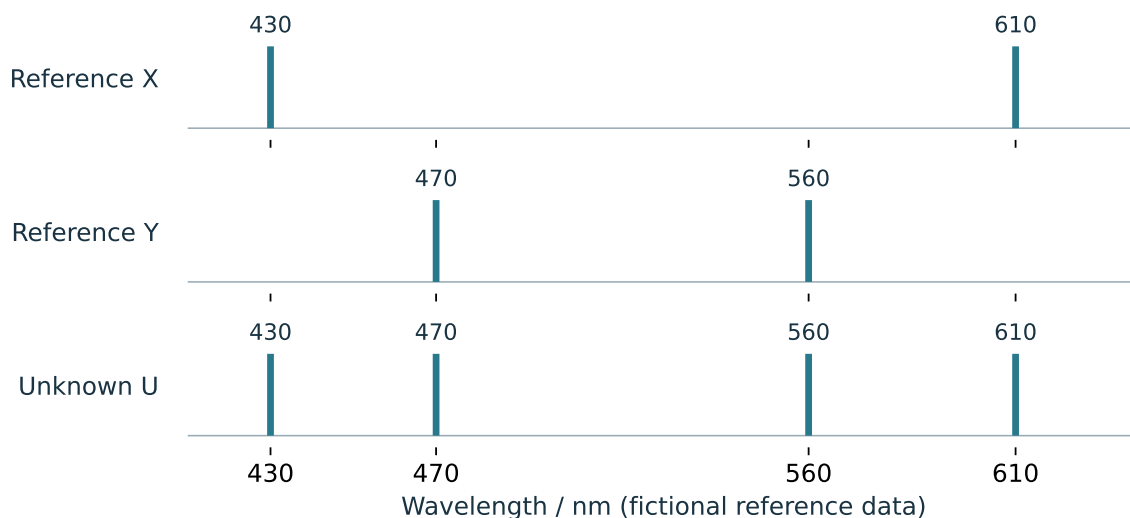
Worked combined result. Fresh portions of a soluble unknown give a yellow flame and a white precipitate with dilute nitric acid followed by silver nitrate. These observations support sodium ions and chloride ions respectively; sodium chloride is consistent with both if the task specifies one pure salt. The halide test does not identify sodium, and the flame test does not identify chloride. Record uncertainty if the sample may be a mixture.

Required practical 7 involves actual school-supervised flame and ion tests on known/unknown ionic compounds. Use approved small quantities, clean apparatus and fresh portions; record reagent, observation and inference separately. Class paper results and the tables here do not replace doing and documenting the practical.

Instrumental results and reference spectra (4.8.3.6–4.8.3.7)

An **instrumental method** can identify substances accurately, detect small amounts (**sensitivity**), and give rapid results. Accurate means close to the accepted value; sensitive means detecting small quantities. They describe different advantages. Instruments still need suitable references, calibration and controlled measurements.

Spectroscopy analyses light. In **flame emission spectroscopy**, put the sample into a flame and pass its emitted light through a spectroscope. The output is a **line spectrum**. Compare line positions with supplied reference spectra to identify metal ions. A mixture can contain lines from more than one metal. Use supplied intensity/concentration calibration to measure concentration; brightness without a calibrated comparison does not supply an absolute concentration.



Worked spectrum. Reference X has lines at 430 and 610 nm; Y has lines at 470 and 560 nm. The unknown has all four lines, supporting X and Y within these references. These fictional positions are given data, not wavelengths to memorise. A missing faint line can reflect the detection limit; do not assert that every absent line proves zero concentration.

Worked calibration. Under a stated linear calibration, 2.0 mg/dm³ gives intensity 20 and 6.0 mg/dm³ gives intensity 60. Intensity 35 lies 15/40 of the way between them:

$$c = 2.0 \text{ mg/dm}^3 + \frac{35 - 20}{60 - 20} \times (6.0 - 2.0) \text{ mg/dm}^3 = 3.5 \text{ mg/dm}^3$$

Stay inside the supplied calibrated range. This calculation uses an explicitly given model; do not assume every instrument is linear at every concentration. Choose an instrument when the task needs detection of a small amount, several metals or a measured concentration; choose a simple chemical test when an appropriate qualitative observation is sufficient.