

8.3 Identification of ions

Name: _____ Class: _____ Date: _____

Total: 33 marks**KEY WORDS** flame test • precipitate • spectroscopy

Objective

Identify the specified ions, write insoluble hydroxide equations and interpret supplied instrumental references. Original questions and authored marking guidance for 4.8.3.1–4.8.3.7: Separate Chemistry, **both tiers**, including spectroscopy. Actual supervised RP7 remains separate.

1 Key knowledge and worked examples

Flames: lithium crimson; sodium yellow; potassium lilac; calcium orange-red; copper green. Clean apparatus and compare references; mixture colours can mask one another.

NaOH produces white hydroxides with Al^{3+} , Ca^{2+} and Mg^{2+} ; only aluminium hydroxide dissolves in excess among these three. Cu^{2+} gives blue, Fe^{2+} green and Fe^{3+} brown precipitate. White remaining in excess does not distinguish calcium from magnesium.

Carbonate + dilute acid releases gas that makes limewater cloudy. Halides: dilute nitric acid then silver nitrate $\rightarrow$ chloride white, bromide cream, iodide yellow. Sulfate: dilute hydrochloric acid then barium chloride $\rightarrow$ white. Use fresh portions to avoid contamination.

Worked equation. $\text{CuCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + 2\text{NaCl}(\text{aq})$. The insoluble hydroxide is solid; two OH groups require 2NaOH. For FeCl_3 , use 3NaOH, $\text{Fe}(\text{OH})_3$ and 3NaCl. An aluminium-dissolution equation is not required.

Worked interpretation. A white precipitate dissolving in excess NaOH identifies Al^{3+} among the listed ions. Its initial white colour alone would not suffice.

In flame emission spectroscopy, light from the sample in a flame passes through a spectroscope. Compare line positions to supplied references; use calibrated intensity for concentration. Accurate, sensitive and rapid are different advantages. These instrumental methods are both-tier here.

Worked interpolation. A stated linear calibration gives intensity 20 at 2.0 mg/dm^3 and 60 at 6.0 mg/dm^3 . At intensity 35, $c = 2.0 + (35 - 20)/(60 - 20) \times 4.0 = 3.5 \text{ mg/dm}^3$. Do not extrapolate beyond the calibrated range without evidence.

2 Practice

2.1 Give potassium's flame colour. [1]

2.2 Give the acid and following reagent used for sulfate ions. [2]

2.3 What can calibrated spectral-line intensity be used to measure? [1]

2.4 Complete and balance, with states: $\text{FeCl}_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \dots + \dots$. Name the precipitate colour. [3]

2.5 A white precipitate remains in excess NaOH. Give the two possible listed metal ions and one further positive result that distinguishes them. [3]

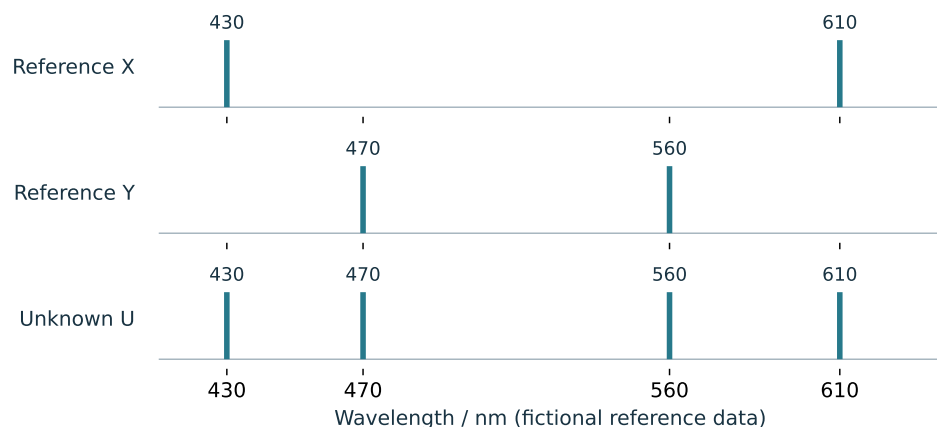
3 Exam-style questions

3.1 Describe the test for iodide ions, giving both reagents and the observation. [3]

3.2 An unknown gives a green flame, but a separate NaOH test records no precipitate. Explain the inference and discrepancy; suggest how to improve the evidence. [4]

3.3 Explain two advantages of instrumental analysis over the specified simple tests, developing each with a measurement purpose. [4]

3.4 Fictional metal references X and Y and unknown U have the spectra below. Identify the supported components of U, justify your inference and describe how the spectrum is obtained. [4]



3.5 Fresh portions of one pure soluble salt give a crimson flame, effervescence with dilute acid and gas making limewater cloudy. Identify the cation and anion and explain why effervescence alone is insufficient. [4]

3.6 A linear calibration gives intensities 12 and 52 at concentrations 1.0 and 5.0 mg/dm³. An unknown gives intensity 32. Calculate its concentration and explain two controls/limits needed before accepting the result. [4]

Solutions

2.1 Lilac, not calcium's orange-red.

2.2 Dilute hydrochloric acid, then barium chloride solution.

2.3 Metal-ion concentration, using an appropriate calibration.

2.4 $\text{FeCl}_3(\text{aq}) + 3\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s}) + 3\text{NaCl}(\text{aq})$: correct products/states, balanced coefficients. Brown precipitate. Check Fe1,Cl3,Na3,O3,H3 each side; $\text{Fe}(\text{OH})_2$ would describe a different iron ion.

2.5 Calcium and magnesium ions. An orange-red flame supports calcium; absence of that colour alone is weaker evidence if contamination/masking is possible. Aluminium's white precipitate would dissolve under the specified excess test.

3.1 Add dilute nitric acid, then silver nitrate solution; yellow precipitate, silver iodide. Hydrochloric acid would introduce chloride and contaminate this test.

3.2 Green flame suggests copper ions, but contamination or mixture masking limits the result. Copper(II) should give a blue NaOH precipitate, so the inconsistent result requires review rather than a confident identification. Repeat with clean apparatus/fresh reference and sample portions, or compare a supplied reference emission spectrum. No precipitate alone does not establish that copper is absent at every concentration.

3.3 Any two developed advantages: sensitive detects small amounts that a simple visible test may miss; rapid obtains results more quickly for many samples; accurate gives measurements close to accepted values with calibration. A spectrum can distinguish supplied mixed-metal line patterns rather than an obscured flame colour. Four marks maximum; do not count sensitive and accurate as synonyms.

3.4 X and Y: U includes X's 430/610 nm and Y's 470/560 nm reference lines. Put sample in a flame and pass emitted light through a spectroscope to record the line spectrum. Positions are supplied fictional data. A missing faint line would not necessarily prove an ion absent.

3.5 Lithium ion from the crimson flame; carbonate ion, supported by the evolved gas making limewater cloudy and therefore identified as carbon dioxide. Effervescence only shows gas evolution and does not name the gas; the positive gas test is needed. The task specifies one salt; do not generalise the combination to every mixture.

3.6 Fraction of calibrated interval = $(32 - 12)/(52 - 12) = 0.50$. Concentration = $1.0 + 0.50 \times (5.0 - 1.0) = \mathbf{3.0 \text{ mg/dm}^3}$. Any two controls/limits: retain the stated linear model within its calibration range; compare using the same measurement conditions and suitable references; use clean apparatus/blank to check contamination; repeat measurements to assess consistency. Intensity alone is not an absolute concentration scale without calibration.