

6.2.1 Titration analysis and equilibrium constants

Name: _____ Class: _____ Date: _____ **Total: 21 marks**

KEY WORDS concordant titres • endpoint • saturated solution

Objective

Build a titration answer end to end: concordant titres and the mean, the indicator with its colour change, an expected titre from a data-book value, and an equilibrium constant with derived units – then diagnose the discrepancy. Original practice for WCH16 Q2 (Summer 2025) and CP13a/b context.

1 Key knowledge and worked examples

■ Concordant titres only

Titres within 0.2 cm^3 are concordant and enter the mean; a rough or stray first titre does not. Readings: 9.95 (excluded – 0.5 off the others), 9.55, 9.45 → mean = $(9.55 + 9.45)/2 = 9.50 \text{ cm}^3$. Quote titres to 0.05 cm^3 .

■ Indicator and endpoint

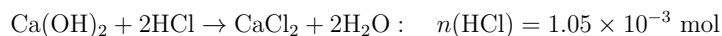
Strong acid–strong base: any screen indicator. Here HCl into $\text{Ca}(\text{OH})_2$: **methyl orange, yellow → orange**, or **phenolphthalein, pink → colourless**. The direction is read from the alkali's colour before to the acid's after.

■ The expected titre chain

Data-book solubility 1.56 g dm^{-3} , $M_r(\text{Ca}(\text{OH})_2) = 74.1$:

$$c = 1.56/74.1 = 0.0211 \text{ mol dm}^{-3}$$

$$n = 0.0211 \times 0.0250 = 5.26 \times 10^{-4} \text{ mol in the pipette}$$



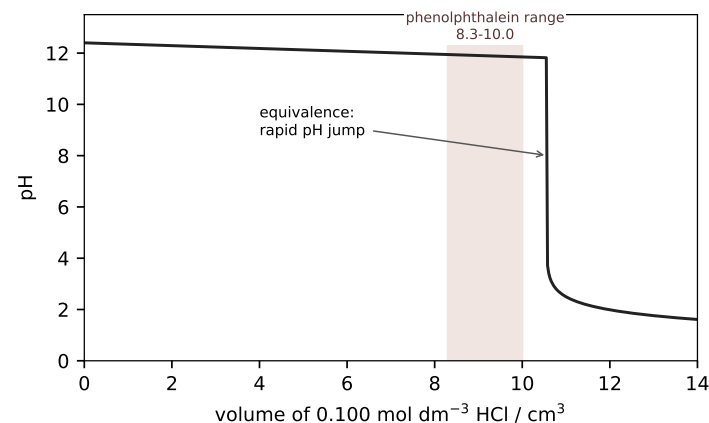
$$V = n/c \times 1000 = 1.05 \times 10^{-3}/0.100 \times 1000 = 10.55 \text{ cm}^3$$

■ Discrepancy and K_c

Mean titre lower than expected, solubility falling as temperature rises → the experiment ran **warmer than 20°C** . Then $K_c = [\text{Ca}^{2+}][\text{OH}^-]^2$: with $[\text{Ca}^{2+}] = 0.0190$, $[\text{OH}^-] = 0.0380$:

$$K_c = 0.0190 \times (0.0380)^2 = 2.74 \times 10^{-5} \text{ mol}^3 \text{ dm}^{-9}$$

Units come from the expression: $\text{mol dm}^{-3} \times (\text{mol dm}^{-3})^2$.



2 Practice

2.1 Titres of 10.20, 9.85, 9.90 and 9.95 cm^3 are recorded. State which titres enter the mean and calculate it. [2]

2.2 Name a suitable indicator for titrating saturated calcium hydroxide with hydrochloric acid, with the colour change. [2]

2.3 25.0 cm³ of 0.0150 mol dm⁻³ Ca(OH)₂ is titrated with 0.100 mol dm⁻³ HCl. Calculate the expected titre. [3]

2.4 With [Ca²⁺] = 0.0150 mol dm⁻³ in a saturated solution, calculate K_c with units. [2]

3 Exam-style questions

3.1 The student's mean titre (9.50 cm³) is smaller than the expected 10.55 cm³. Given that calcium hydroxide's solubility decreases as temperature rises, explain the difference, and state one procedural reason the flask is left stoppered for one day before filtering. [3]

3.2 A redox titration: 25.0 cm³ of acidified iron(II) sulfate is titrated with 0.0200 mol dm⁻³ KMnO₄, needing 22.50 cm³. Given $5\text{Fe}^{2+} + \text{MnO}_4^- + 8\text{H}^+ \rightarrow 5\text{Fe}^{3+} + \text{Mn}^{2+} + 4\text{H}_2\text{O}$, calculate the iron(II) concentration, and state how the endpoint is seen without an indicator. [4]

3.3 In the iodine–thiosulfate titration (CP13b), starch is added near the endpoint, not at the start. Give the reason, and the endpoint colour change. [2]

3.4 The burette used for the calcium hydroxide titration has an uncertainty of ± 0.05 cm³ per reading. Calculate the percentage uncertainty in a delivered volume of 9.50 cm³, and state one change that reduces this percentage uncertainty. [3]

Solutions

Original-task marking: M1 method, A1 accuracy, B1 reasoning.

2.1 9.85, 9.90 and 9.95 are within 0.2 cm^3 ; mean = $(9.85 + 9.90 + 9.95)/3 = 9.90 \text{ cm}^3$.

2.2 Methyl orange: yellow $\rightarrow$ orange. (Phenolphthalein pink $\rightarrow$ colourless also accepted.)

2.3 $n(\text{Ca}(\text{OH})_2) = 0.0150 \times 0.0250 = 3.75 \times 10^{-4} \text{ mol}$; $n(\text{HCl}) = 7.50 \times 10^{-4} \text{ mol}$;
 $V = 7.50 \times 10^{-4} / 0.100 \times 1000 = 7.50 \text{ cm}^3$.

2.4 $[\text{OH}^-] = 2 \times 0.0150 = 0.0300$; $K_c = 0.0150 \times 0.0300^2 = 1.35 \times 10^{-5} \text{ mol}^3 \text{ dm}^{-9}$.

3.1 A smaller titre means less dissolved calcium hydroxide than at 20°C – the solution was titrated warmer than 20°C , and since solubility falls with temperature, fewer moles were present. Standing stoppered for a day lets the dissolution equilibrium complete, giving a saturated solution of constant concentration.

3.2 $n(\text{MnO}_4^-) = 0.0200 \times 0.02250 = 4.50 \times 10^{-4} \text{ mol}$; $n(\text{Fe}^{2+}) = 5 \times 4.50 \times 10^{-4} = 2.25 \times 10^{-3} \text{ mol}$; $c = 2.25 \times 10^{-3} / 0.0250 = 0.0900 \text{ mol dm}^{-3}$. The purple manganate(VII) is decolourised while iron(II) remains; the endpoint is the **first permanent pink** where the next drop of purple has nothing to oxidise.

3.3 Starch added too early adsorbs iodine strongly (or forms an insoluble blue complex) making the endpoint sluggish and inaccurate; added near the endpoint it gives a sharp finish. Colour change at the endpoint: blue-black $\rightarrow$ colourless as the last iodine is reduced.

3.4 Two readings, so the absolute uncertainty is $\pm 0.10 \text{ cm}^3$; percentage uncertainty = $0.10/9.50 \times 100 = 1.05\%$ (about 1.1%). Reduce it by titrating with a larger (deliverable) volume – a bigger titre divides the same absolute uncertainty by a larger value – or by using a burette with finer graduations.