

Assessment Schedule – 2025

Scholarship Chemistry (93102)

Evidence Statement

Q	Evidence	Scholarship Not Awarded	Scholarship	Outstanding Scholarship
ONE (a)(i)	$\text{H}_3\text{Cit}(aq) + 3\text{NaHCO}_3(aq) \rightarrow \text{Na}_3\text{Cit}(aq) + 3\text{H}_2\text{O}(\ell) + 3\text{CO}_2(g)$ For a reaction to be spontaneous, there must be a positive change / net increase in total entropy, which can be represented as: $\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$. The reaction between H_3Cit and NaHCO_3 is endothermic. Energy is absorbed from the environment during the reaction, resulting in it becoming significantly colder. This decrease in random movement of the surrounding particles results in a decrease in the entropy of the surroundings, $\Delta S_{\text{surroundings}}$. As the reaction progresses, four moles of aqueous reactants react to form seven moles of products, three of which are CO₂ gas molecules that have significantly greater entropy due to random movement than particles in solution. This reorganisation of matter from a more ordered state into a less ordered state results in an increase in the entropy of the system, ΔS_{system}. As this reaction proceeds spontaneously the increase in ΔS_{system} must be sufficiently greater than the decrease in $\Delta S_{\text{surroundings}}$ for ΔS_{total} to be positive.	Any of the following: - Explains entropy changes to the system. - Explains entropy changes to the surroundings. - Explains the role of water in the reaction. - Draws some correct Lewis structures, with shapes and bond angles. - Give reasons for the likely reaction. - Explains the correct shape or polarity of BrF_3 molecules. - Completes pH calculations. - Explains buffering capacity of a solution.	A range of the following: - Discusses entropy changes occurring during the reaction and gives reasons for the role of water in the reaction. - Draws most correct Lewis structures, with associated shapes and bond angles for the ions, and gives reasons for the likely reaction and molecular shape and polarity of BrF_3 molecules. - Completes calculations for aspects of the titration curve and evaluates the buffering capacity of one solution.	A range of the following: - Comprehensively discusses the entropy changes occurring during the reaction and explains the role of water in the reaction. - Draws all correct Lewis structures, shapes, and bond angles for all four ions, justifies the likely reaction, and explains the molecular shape and polarity of BrF_3 molecules. - Completes relevant calculations to complete the titration curve and evaluates the buffering capacity of the two solutions.
(ii)	Explanation may include: - The reactant particles are held in a fixed lattice structure in the solid state. Only once dissolved, the reactant particles are free to move and collide with other reactant particles. Reactant particles must collide with sufficient energy and the right orientation for a reaction to occur. - Proton transfer must occur between the H_3Cit molecules and the HCO_3^- ions. This cannot readily occur without water present. Once dissolved, H_3Cit molecules can donate H^+ ions to water molecules, forming H_3O^+ ions, which can then react with the HCO_3^- ions in solution. - Other relevant and chemically accurate discussion.			

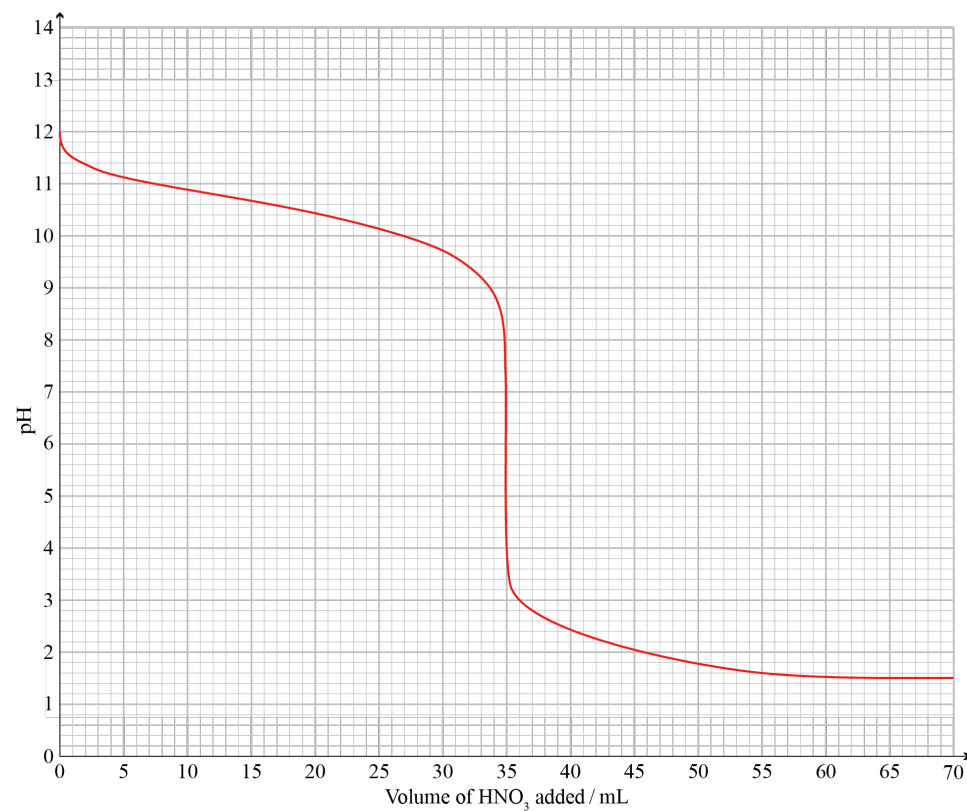
(b)(i)	$\left[:\ddot{\text{F}}-\ddot{\text{Br}}-\ddot{\text{F}}: \right]^+$ Bent – 109.5° $\left[\begin{array}{c} :\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}-\ddot{\text{Br}}-\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}: \end{array} \right]^-$ Square planar – 90° $\left[\begin{array}{c} :\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}-\ddot{\text{Br}}-\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}: \end{array} \right]^+$ See-saw – 90° + 120° $\left[:\ddot{\text{F}}-\ddot{\text{Br}}-\ddot{\text{F}}: \right]^-$ Linear – 180°			
(ii)	When a fluorine ion is transferred between the reactant molecules, it could be transferred as F⁺, as in the first reaction, or as F[–], as in the second reaction. Fluorine atoms have a higher electronegativity than bromine atoms. As fluorine atoms attract bonding electrons more strongly than bromine atoms, it is more likely that the fluorine atom will be transferred with the bonded electrons still attached, as F[–] ions. Equation (2) is therefore more likely to be the correct one.			
(iii)	In BrF₃ molecules, there is a central Br atom surrounded by five regions of electron density, three bonding pairs of electrons, and two non-bonding (lone) pairs of electrons. In a trigonal planar shaped molecule, all three bonding pairs are positioned equatorially, while the two lone pairs are positioned 90° to these in axial positions. In a T-shaped molecule, two bonding pairs are positioned axially, and one is positioned equatorially. The two lone pairs are positioned equatorially at 120° apart from each other, and 90° apart from bonding pairs. Lone-pairs of electrons repel more than bonding-pairs of electrons. The more electrically stable arrangement of electrons involves the greatest bond angles between the lone-pairs and bonding-pairs, as is achieved in the T-shaped molecular geometry. All three bonds are polar due to the electronegativity differences between Br and F atoms, and all three are identical with the same bond polarity. In this shape, the three bond dipoles are not positioned symmetrically and do not cancel, resulting in polar molecules.			

(c)(i)	<i>Initial concentration:</i> $n(\text{HNO}_3)_{\text{equivalence}} = 0.1300 \text{ mol L}^{-1} \times 0.03498 \text{ L} = 0.0045475 \text{ mol}$ $[\text{CH}_3\text{CH}_2\text{NH}_2]_{\text{initial}} = \frac{0.0045475 \text{ mol}}{0.025 \text{ L}} = 0.1819 \text{ mol L}^{-1}$ <i>Initial pH:</i> $[\text{CH}_3\text{CH}_2\text{NH}_2] = 0.1819 \text{ mol L}^{-1}$ $K_b = \frac{10^{-14}}{10^{-10.64}} = 4.365 \times 10^{-4}$ $4.365 \times 10^{-4} = \frac{[\text{OH}^-]^2}{0.1819 \text{ mol L}^{-1}}$ $[\text{OH}^-] = 0.008913 \text{ mol L}^{-1}$ $[\text{H}_3\text{O}^+] = \frac{1 \times 10^{-14}}{0.008913}$ $[\text{H}_3\text{O}^+] = 1.12 \times 10^{-12} \text{ mol L}^{-1}$ $\text{pH} = -\log(1.12 \times 10^{-12} \text{ mol L}^{-1}) = 11.95$ <i>pH at equivalence:</i> $[\text{CH}_3\text{CH}_2\text{NH}_3^+] = 0.1819 \text{ mol L}^{-1} \times \frac{0.02500 \text{ L}}{0.05998 \text{ L}} = 0.07582 \text{ mol L}^{-1}$ $10^{-\text{p}K_a} = \frac{[\text{H}_3\text{O}^+]^2}{[\text{CH}_3\text{CH}_2\text{NH}_3^+]}$ $10^{-10.64} = \frac{[\text{H}_3\text{O}^+]^2}{0.07582 \text{ mol L}^{-1}}$ $[\text{H}_3\text{O}^+] = 1.318 \times 10^{-6} \text{ mol L}^{-1}$ $\text{pH} = -\log(1.318 \times 10^{-6} \text{ mol L}^{-1}) = 5.88$ <i>At halfway to the equivalence point:</i> $V(\text{HNO}_3) = \frac{34.98 \text{ mL}}{2} = 17.49 \text{ mL}$ $\text{pH} = 10.64$			
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After 20 mL excess acid has been added:

$$[\text{H}_3\text{O}^+] = \frac{0.1300 \text{ mol L}^{-1} \times 0.0200 \text{ L}}{0.07998 \text{ L}} = 0.0325 \text{ mol L}^{-1}$$

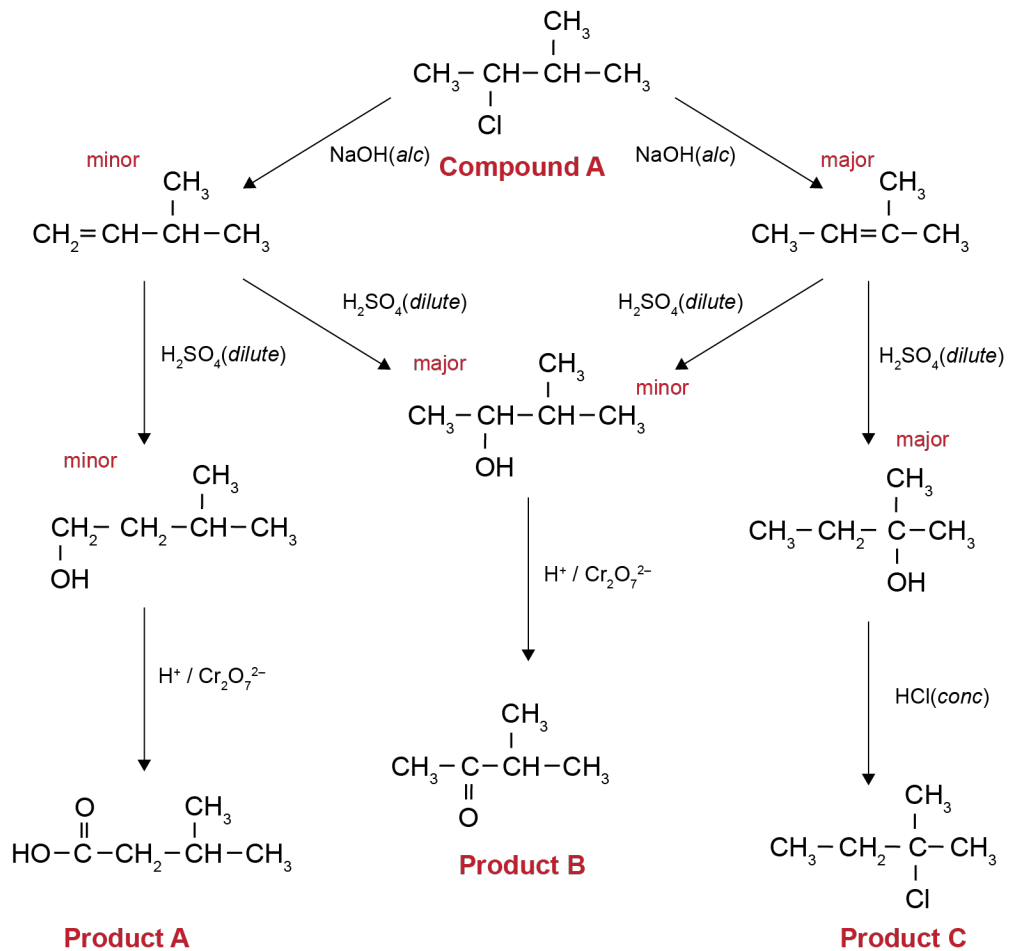
$$\text{pH} = -\log(0.0325 \text{ mol L}^{-1}) = 1.49$$



- (ii) After 18 mL of HNO_3 has been added, the solution will still be effective at buffering against small amounts of strong acid and strong base, with roughly equivalent amounts of weak base and conjugate acid present.
- After 28 mL of HNO_3 has been added, the solution will have a much higher concentration of conjugate acid, and a much lower concentration of weak base. It will have reduced buffering capacity against small amounts of strong acid but can still buffer against small amounts of strong base.

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TWO (a)(i)	$[\text{CoCl}_4]^{2-}(\text{aq}) + 6\text{H}_2\text{O}(\ell) \rightleftharpoons [\text{Co}(\text{H}_2\text{O})_6]^{2+}(\text{aq}) + 4\text{Cl}^{-}(\text{aq})$ Addition of concentrated $\text{HCl}(\text{aq})$ will increase the concentration of the Cl^{-} ions in the solution, favouring the reverse direction. This will result in an increase in concentration of $[\text{CoCl}_4]^{2-}(\text{aq})$, the blue species, and a decrease in concentration of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$, the pink species. The solution will change colour from purple coloured solution to one which is more blue in colour. Following addition of solid $\text{AgNO}_3(\text{s})$, it will dissolve to release Ag^{+} ions, which will bond the Cl^{-} ions in the solution, forming a white AgCl precipitate. A decrease in the concentration of Cl^{-} ions will result in the equilibrium (above) favouring the forwards direction. This will result in an increase in concentration of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$, the pink species, and a decrease in concentration of $[\text{CoCl}_4]^{2-}(\text{aq})$, the blue species. The solution will change colour from purple coloured solution to one which is more pink in colour, but cloudy due to the white precipitate present.	Any of the following: - Writes a balanced equilibrium equation. - Explains observed changes to the CoCl_2 solution from addition of HCl. - Explains observed changes to the CoCl_2 solution from addition of AgNO_3. - Explains observed changes to the CoCl_2 paper from addition of H_2O. - Identifies Compound A - Completes parts of the reaction scheme correctly. - Explains the use of ^{13}C NMR in distinguishing compounds. - Explains the use of IR in distinguishing compounds. - Discusses the identity of an element correctly.	A range of the following: - Uses a balanced equilibrium equation to justify the impact of changes in Cl^{-} or H_2O on the equilibrium position and resulting observations on the solution and / or CoCl_2 paper. - Identifies Compound A, completes most of the reaction scheme correctly, and explains the use of ^{13}C NMR and IR in distinguishing the proposed final products. - Discusses the identity of two of the three elements correctly with clear links to the atomic properties outlined in the table.	A range of the following: - Discusses, with support of the balanced equilibrium equation, the impact of changes in Cl^{-} and H_2O on the equilibrium position and resulting observations for the solution and CoCl_2 paper. - Identifies Compound A, correctly completes the reaction scheme with all three final products, all major / minor products labelled, and provides a clear explanation for the use of ^{13}C NMR and IR spectroscopy in distinguishing the three final products. - Identifies and discusses the identity of three elements with clear links to the atomic properties outlined in the table.
(ii)	As water, H_2O, is a reagent in the equilibrium between the two coloured species, and there is a reference in the question to drying the paper, then it can be used to test for the presence of water. When dried, the paper is blue, due to the presence of primarily $[\text{CoCl}_4]^{2-}(\text{aq})$, the blue species, due to the equilibrium shifting entirely to the reactant side during drying. Following the addition of water to the test paper, the equilibrium will shift in the forwards direction, forming more of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}(\text{aq})$, the pink species, resulting in a colour change from blue to pink of the indicator test paper.			

(b)(i)



(ii)	The ^{13}C NMR spectra for all three compounds will produce four peaks, as all three compounds contain four unique carbon bonding environments (all three contain two symmetrically bonded CH_3 groups). Therefore, the number of peaks present will not distinguish them. All three will produce three peaks situated close to 0–40 ppm for the CH_x bonding environments. The fourth peak position, however, will vary for each compound. For Product A, the fourth peak will be situated around 170–190 ppm representing the COOH carbon bonding environment. For Product B, the fourth will be situated above 200 ppm representing the $\text{C}=\text{O}$ carbon bonding environment. For Product C, the fourth will be situated in the same vicinity as the CH_x bonding environments for the $\text{C}-\text{Cl}$ carbon bonding environment. The IR spectra for all three compounds will be unique. Each will contain a peak around 2900 cm^{-1} for the $\text{C}-\text{H}$ bonds present. Product A will produce two additional distinguishing peaks; one around 1700 cm^{-1} for the $\text{C}=\text{O}$ bond, and a broad peak around $2500\text{--}3000\text{ cm}^{-1}$ for the $\text{O}-\text{H}$ bond. Product B will only produce one additional distinguishing peak around 1700 cm^{-1} for the $\text{C}=\text{O}$ bond. Product C will not produce any further distinguishing peaks other than a potential $\text{C}-\text{Cl}$ peak in the fingerprint region.			
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(c)(i)	<table><tr><th>Description</th><th>Element Symbol</th></tr><tr><td>This element has one electron in the 1s shell and is the lightest element</td><td>H</td></tr><tr><td>Found in Period 2, this element has no paired p electrons and does not follow the general trend in ionisation energy across a period.</td><td>B</td></tr><tr><td>This element has the highest ionisation energy of any element.</td><td>He</td></tr><tr><td>Ions of this Period 3 element have the smallest ionic radius of all ions in this period.</td><td>Al</td></tr><tr><td>This element exists as a gaseous diatomic molecule, with the highest bond enthalpy of any of the elemental diatomic molecules.</td><td>N</td></tr></table>	Description	Element Symbol	This element has one electron in the 1s shell and is the lightest element	H	Found in Period 2, this element has no paired p electrons and does not follow the general trend in ionisation energy across a period.	B	This element has the highest ionisation energy of any element.	He	Ions of this Period 3 element have the smallest ionic radius of all ions in this period.	Al	This element exists as a gaseous diatomic molecule, with the highest bond enthalpy of any of the elemental diatomic molecules.	N			
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(ii)	• B: Boron is found in Period 2 with an electron configuration of $1s^2, 2s^2, 2p^1$, therefore no paired p electrons. Ionisation energies generally increase across a period, but B has a lower IE than Be, the element with one fewer electron. Be: $1s^2, 2s^2$. As Be has no p electrons, B has an electron in a subshell further from the nucleus, so despite B having five protons compared to Be's four, the higher nuclear charge is offset by the electron in the p shell which takes less energy to remove than B's least tightly held electron. • He: Ionisation energy is the energy required to remove a mole of the least tightly held electron from a mole of gaseous atoms and depends on the number of protons, nuclear charge and the electron arrangement. Despite He having only two protons and therefore a reasonably weak nuclear charge, its outer electrons are in the first shell, so are very close to the nucleus and there is no shielding (inner shell e-e repulsions), as there are no inner shells. This results in He having the highest IE of any element. • Al: In Period 3, the ions formed are: $\text{Na}^+, \text{Mg}^{2+}, \text{Al}^{3+}, \text{P}^{3-}, \text{S}^{2-}, \text{Cl}^-$. The three anions are all larger than the atoms they form from, due to extra valence electron, resulting in more valence shell repulsion and a larger radius. The three cations are all smaller than the atom they formed from due to losing their valence shell (they all go from three to two shells). Of the three cations, Al has the most protons and therefore the greatest nuclear charge, resulting in the smallest radius of the three cations. • N: The gaseous diatomic molecules are: $\text{H}_2, \text{N}_2, \text{O}_2, \text{Cl}_2, \text{F}_2$. Bond enthalpy depends largely on the number of shared electrons in the covalent bond(s) between atoms, which is dependent on the number of valence electrons each atom has. In terms of the number of shared electrons: H, F and Cl form single bonds between atoms, O forms double bonds between atoms, whilst N with five valence electrons forms triple bonds between atoms. As N_2 is the only diatomic gas with a triple bond, it therefore has the highest bond enthalpy.															

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THREE (a)	$n(\text{Na}_2\text{CO}_3) = \frac{m}{M} = \frac{2.691 \text{ g}}{106.0 \text{ g mol}^{-1}} = 0.2539 \text{ mol}$ $c(\text{Na}_2\text{CO}_3) = \frac{n}{V} = \frac{0.2539 \text{ mol}}{0.25 \text{ L}} = 0.10155 \text{ mol L}^{-1}$ $n(\text{Na}_2\text{CO}_3)_{\text{titration}} = cV = 0.01 \text{ L} \times 0.10155 \text{ mol L}^{-1} = 1.0155 \times 10^{-3} \text{ mol}$ $n(\text{HCl}) = 2 \times 1.0155 \times 10^{-3} \text{ mol} = 2.031 \times 10^{-3} \text{ mol}$ Student A has concordant (reliable data). Using this average titre $\frac{15.80 + 15.70 + 15.65 + 15.70}{4} = 15.7125 \text{ mL}$ $c(\text{HCl}) = \frac{2.031 \times 10^{-3} \text{ mol}}{0.0157125 \text{ L}} = 0.1293 \text{ mol L}^{-1}$ This result is significantly higher than the teacher's concentration of HCl. While student A's method appears valid (reliable concordant titres), the results are not accurate. Student B does not have concordant titres. Averaging the closest 4, we get $\frac{20.20 + 20.45 + 19.80 + 19.50}{4} = 19.99 \text{ mL}$ $c(\text{HCl}) = \frac{2.031 \times 10^{-3} \text{ mol}}{0.01999 \text{ L}} = 0.1016 \text{ mol L}^{-1}$ <i>Depending on the titres averaged, a range of answers 0.100–0.102 mol L⁻¹ were accepted.</i> This result is close to the teacher's concentration of HCl but appears to be more due to luck. Student B has either carried out a method that is not valid, or followed a valid method with poor technique, resulting in non-concordant titres that have incidentally given results that are accurate. A method can be valid but give inaccurate results if there has been a single significant error, or a similar mistake is made repeatedly. Reasons for this may include – incorrect weighing of the primary standard, poor transfer of the primary standard to the flask, using an impure sample of the primary standard, using the wrong type of flask, accidental dilution of a solution, repeatedly reading the burette incorrectly. A method can be low in validity (ie. unreliable) but give accurate results due to random experimental errors that have cancelled out. Reasons for this can include – inconsistent pipetting, not reading the burette correctly, not using cleaned equipment, refilling a burette mid-experiment, variations in end point or going past the end point of the titration on some titres.	Any of the following: • Calculates HCl concentration for one student. • Discusses the validity of a method. • Discusses the accuracy of one set of results. • Calculates an enthalpy change correctly. • Provides reasons for variation in enthalpy change values. • Identifies the structure of ethyl butanoate and explains how it can be synthesised. • Explains some safety considerations during ester synthesis. • Explains the use of a distillation apparatus during hydrolysis. • Explains the use of a reflux apparatus during hydrolysis.	A range of the following: • Calculates HCl concentrations for both students and discusses the validity of their methods or the accuracy of their results, with sources of error outlined. • Calculates two of the three enthalpy changes correctly and provides a reason for variation in the different values. • Discusses two reactions for the synthesis of ethyl butanoate, with structural equations and some ideas provided towards the conditions and safety consideration, and gives an outline of the use of reflux and distillation during hydrolysis.	A range of the following: • Correctly calculates the concentration of the HCl solution for both students, and discusses the validity of the method and accuracy of the results for both students with clearly articulated sources of error. • Calculates the correct enthalpy changes for the three methods and provides clear justification for the variation in the different values. • Fully discusses two reactions for the synthesis of ethyl butanoate, with structural equations, conditions and safety considerations well-articulated, and explains the use of reflux and distillation equipment in the hydrolysis of the product ester.

(b)(i)	Method 1: Using experimental data $q = mc\Delta T$ $= 300 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} \times 29.2 \text{ }^{\circ}\text{C}$ $= 36\,617 \text{ J}$ $= 36.617 \text{ kJ}$ $n = \frac{m}{M} = \frac{1.35 \text{ g}}{102.2 \text{ g mol}^{-1}} = 0.013209 \text{ mol}$ $\Delta_c H^{\circ} (\text{C}_6\text{H}_{13}\text{OH}(\ell)) = \frac{-36.617 \text{ kJ}}{0.013209 \text{ mol}} = -2772 \text{ kJ mol}^{-1}$ Method 2: Using enthalpies of formation data: $\Delta_c H^{\circ} (\text{C}_6\text{H}_{13}\text{OH}(\ell)) = (6 \times -393 + 7 \times -286) - 320$ $= (-2358 + -2002) + 320 = -4040 \text{ kJ mol}^{-1}$ Method 3: using bond enthalpies and enthalpies of reaction data $\Delta_c H^{\circ} (\text{C}_6\text{H}_{13}\text{OH}(\ell)) = \text{C}_6\text{H}_{13}\text{OH}(\ell) + 9\text{O}_2(\text{g}) \rightarrow 6\text{CO}_2(\text{g}) + 7\text{H}_2\text{O}(\ell)$ Bond enthalpy for C–H = $-\frac{-1652}{4} = 413 \text{ kJ mol}^{-1}$ Bond enthalpy for C–C = $-(2826 - (6 \times 413)) = 348 \text{ kJ mol}^{-1}$ Bonds broken = $5 \times \text{C–C} + 13 \times \text{C–H} + 1 \times \text{C–O} + 1 \times \text{O–H} + 9 \times \text{O=O}$ $= 1740 + 5369 + 358 + 463 + 4482 = 12\,412 \text{ kJ mol}^{-1}$ Bonds formed = $12 \times \text{C=O} + 14 \times \text{O–H}$ $= 8940 + 6482 = 15\,422 \text{ kJ mol}^{-1}$ $\Delta_c H^{\circ} (\text{C}_6\text{H}_{13}\text{OH}(\ell)) = 12\,412 - 15\,422 = -3010 \text{ kJ mol}^{-1}$			
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(ii)	The three enthalpy of combustion values were different because each method has its own limitations. The value calculated using bond enthalpy data can be less accurate because bond enthalpies are average values that don't consider the specific environment of each bond in a molecule, and also do not take into account intermolecular forces between molecules. The method using enthalpies of formation is generally more reliable, as it uses data measured under standard conditions for specific substances. It is more negative than the value calculated using bond enthalpies because it does take into account that two of the substances are in the liquid state. The experimental value based on temperature change is the least negative and least accurate, because of heat loss to the surroundings, incomplete combustion, or measurement errors.			
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(c)(i)	$ \begin{array}{c} \text{H} & \text{H} & \text{H} & & \text{O} & & \text{H} & \text{H} \\ & & & & // & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C} & & \text{O} & & \text{C}-\text{C}-\text{H} \\ & & & & & & & \\ \text{H} & \text{H} & \text{H} & & \text{O} & & \text{H} & \text{H} \\ & & & & & & & \\ & & & & & & \text{H} & \text{H} \end{array} $ Ethyl butanoate			
(ii)	$ \begin{array}{c} \text{H} & \text{H} & \text{H} & & & & \text{H} & \text{H} \\ & & & & & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{OH} & + & \text{HO}-\text{C}-\text{C}-\text{H} & \rightarrow & \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{O}-\text{C}-\text{C}-\text{H} & + & \text{H}_2\text{O} \\ & & & & & & & \\ \text{H} & \text{H} & \text{H} & & \text{O} & & \text{H} & \text{H} \\ & & & & & & & \\ & & & & \text{O} & & & \end{array} $ $ \begin{array}{c} \text{H} & \text{H} & \text{H} & & & & \text{H} & \text{H} \\ & & & & & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{Cl} & + & \text{HO}-\text{C}-\text{C}-\text{H} & \rightarrow & \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{O}-\text{C}-\text{C}-\text{H} & + & \text{HCl} \\ & & & & & & & \\ \text{H} & \text{H} & \text{H} & & \text{O} & & \text{H} & \text{H} \\ & & & & & & & \\ & & & & \text{O} & & & \end{array} $ The two methods are using butanoic acid and ethanol, and using butanoyl chloride and ethanol. - Butanoic acid and ethanol will react to form the ester under reflux using concentrated sulfuric acid as a catalyst. It is an equilibrium reaction, so it will not go to completion and will not have a high yield. The yield, however, can be increased by using excess ethanol to favour product formation. Concentrated sulfuric acid is very corrosive and should be handled with care. Heating / reflux can form flammable vapours – particularly gaseous ethanol – and should be done far away from ignition sources. Hot equipment also poses a burn risk. - An alternative method is to use butanoyl chloride and ethanol. Butanoyl chloride is a more reactive reagent, so it is important that the conditions are water free. The reaction occurs more quickly, does not require reflux, and goes to completion without a catalyst, but the byproduct is HCl, which is harmful and will need to be removed safely. The reaction may be carried out in a fume hood to avoid exposure to HCl fumes.			

(iii)	First, the ester is heated under reflux with water and a strong acid catalyst, such as sulfuric acid. Apparatus 2 allows the reaction to be heated to increase the rate of reaction without losing volatile reactants or products. As the mixture boils, any vapours will rise into the condenser, cool, and condense back into the reaction flask. This ensures that the ester is fully hydrolysed into the alcohol and carboxylic acid. After hydrolysis is complete, the reaction mixture contains a mixture of ethanol, butanoic acid, and acid catalyst. To separate and isolate the products, distillation is carried out using Apparatus 1. As the mixture is heated, ethanol, which has a lower boiling point than butanoic acid, will vaporise first. It is condensed using the cooled condenser, and collected in a separate container. The remaining mixture, now containing mostly butanoic acid, can be separated further using the same equipment.			
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Q	Evidence	Scholarship Not Awarded	Scholarship	Outstanding Scholarship
FOUR (a)(i)	For precipitation of silver chromate: $[\text{CrO}_4^{2-}] = 1.39 \times 10^{-3} \text{ mol L}^{-1}$ $K_s(\text{Ag}_2\text{CrO}_4) = [\text{Ag}^+]^2[\text{CrO}_4^{2-}]$ $2.60 \times 10^{-12} = [\text{Ag}^+]^2(1.39 \times 10^{-4})$ $[\text{Ag}^+] = \mathbf{1.37 \times 10^{-4} \text{ mol L}^{-1}}$ For precipitation of silver chloride: $[\text{Cl}^-] = 1.68 \times 10^{-4} \text{ mol L}^{-1}$ $K_s(\text{AgCl}) = [\text{Ag}^+][\text{Cl}^-]$ $1.80 \times 10^{-10} = [\text{Ag}^+](1.68 \times 10^{-4})$ $[\text{Ag}^+] = \mathbf{1.07 \times 10^{-6} \text{ mol L}^{-1}}$ As the concentration of Ag^+ ions required to precipitate silver chloride is lower than that required to precipitate silver chromate, AgCl will precipitate first.	Any of the following: • Completes calculations to determine which solid precipitates first from the solution. • Calculates the percentage of Cl^- remaining in the solution during precipitation. • Partially completes the calculation to determine the concentration of the AgNO_3 solution. • Discusses observations for one cell. • Explains oxidation-reduction, flow of charge, movement of particles, for one cell. • Calculates E° value for one of the two half cells. • Explains differences in Na^+ and Mg^{2+} ions in relation to their attractive forces with Cl^- ions.	A range of the following: • Carries out calculations to determine the first solid to precipitate from the solution, and either whether the process can be used to separate the anions, or the concentration of the AgNO_3 solution used in the second experiment. • Discusses a range of observations, oxidation-reduction processes, flow of charge, and movement of particles for the two cells, and E° values for the two half cells described. • Explains differences in energy requirements for the formation of gaseous ions from NaCl and MgCl_2.	A range of the following: • Carries out calculations to identify the first solid to precipitate from the solution, evaluate whether the process can be used to separate the two ions, and the concentration of the AgNO_3 solution utilised in the second experiment. • Discusses the observed changes, oxidation-reduction processes, flow of charge, and movement of particles, in the two cells, and carries out cell potential calculations to determine the spontaneity of the reaction between Zn and $\text{H}^+ / \text{MnO}_4^-$ solution. • Comprehensively justifies the differences in energy requirements for the formation of gaseous ions from NaCl and MgCl_2 crystals with clear communication.
(ii)	When silver chromate begins to precipitate, $[\text{Ag}^+] = 1.37 \times 10^{-4} \text{ mol L}^{-1}$ At this point: $K_s(\text{AgCl}) = [\text{Ag}^+][\text{Cl}^-]$ $1.80 \times 10^{-10} = (1.37 \times 10^{-4})[\text{Cl}^-]$ $[\text{Cl}^-] = \mathbf{1.32 \times 10^{-6} \text{ mol L}^{-1}}$ $\% \text{Cl}^- \text{ remaining} = \frac{1.32 \times 10^{-6} \text{ mol L}^{-1}}{1.68 \times 10^{-4} \text{ mol L}^{-1}} = 0.783 \%$ $\% \text{Cl}^- \text{ precipitated} = 100\% - 0.783\% = \mathbf{99.2\%}$ As only 99.2% has been precipitated, and 99.9% is required, this procedure does not meet the criteria provided.			

(iii)	The second solid to precipitate in the titration is Ag_2CrO_4. When Ag_2CrO_4 begins to precipitate, 99.2% of Cl^- has precipitated (calculated in (ii)): $n(\text{Cl}^-)_{\text{precipitated}} = 99.2\% \times 1.68 \times 10^{-4} \text{ mol L}^{-1} \times 0.125 \text{ L}$ $= 2.0832 \times 10^{-5} \text{ mol}$ $= n(\text{Ag}^+)_{\text{precipitated}}$ $[\text{CrO}_4^{2-}] = \frac{1.39 \times 10^{-4} \text{ mol L}^{-1} \times 125 \text{ mL}}{141.8 \text{ mL}}$ $= 1.225 \times 10^{-4} \text{ mol L}^{-1}$ $K_s[\text{Ag}_2\text{CrO}_4] = [\text{Ag}^+]^2 [\text{CrO}_4^{2-}]$ $2.60 \times 10^{-12} = [\text{Ag}^+]^2 (1.225 \times 10^{-4})$ $[\text{Ag}^+] = 1.457 \times 10^{-4} \text{ mol L}^{-1}$ $n(\text{Ag}^+)_{\text{solution}} = 1.457 \times 10^{-4} \text{ mol L}^{-1} \times 0.1418 \text{ L}$ $= 2.065 \times 10^{-5} \text{ mol}$ $n(\text{Ag}^+)_{\text{total}} = n(\text{Ag}^+)_{\text{precipitated}} + n(\text{Ag}^+)_{\text{solution}}$ $= 2.0832 \times 10^{-5} \text{ mol} + 2.065 \times 10^{-5} \text{ mol}$ $= 4.148 \times 10^{-5} \text{ mol}$ $[\text{AgNO}_3] = \frac{4.148 \times 10^{-5} \text{ mol}}{0.0168 \text{ L}}$ $= \mathbf{2.469 \times 10^{-3} \text{ mol L}^{-1}}$ <i>If it is assumed that all the Cl^- has precipitated by the time the second solid precipitates, or (ii) has not been solved, then:</i> $n(\text{Ag}^+)_{\text{precipitated}} = 2.10 \times 10^{-5} \text{ mol}$ $n(\text{Ag}^+)_{\text{total}} = 4.165 \times 10^{-5} \text{ mol}$ $[\text{AgNO}_3] = \mathbf{2.479 \times 10^{-3} \text{ mol L}^{-1}}$			
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(b)(i) Electrons will flow spontaneously from the negative electrode to the positive electrode. Cell 1 has a positive voltage, indicating electrons will flow spontaneously from the left-hand Cu beaker as drawn (the negative electrode where oxidation occurs) to the MnO_4^- beaker as drawn (the positive electrode where reduction occurs). • Left cell (oxidation): $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$. The brown / red Cu electrode loses mass as it is oxidised. The solution becomes darker blue as $[\text{Cu}^{2+}]$ increases. SO_4^{2-} ions from the salt bridge move into the solution to balance the charge. • Right cell (reduction): $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$. The solution fades from purple to pale pink as MnO_4^- is reduced. The inert Pt electrode shows no visible change. As $[\text{H}^+]$ decreases, K^+ ions move into the solution from the salt bridge to balance the charge. Cell 2 has a negative voltage, indicating the electrons are flowing the other way around, oxidation is occurring in the right-hand Zn beaker (the negative electrode) and are flowing to the Cu beaker (positive electrode) where reduction occurs. • Right cell (oxidation): $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$. The grey Zn electrode loses mass as it is oxidised. No visible change in solution as Zn^{2+} is colourless. SO_4^{2-} ions move into the solution from the salt bridge to balance the charge. • Left cell (reduction): $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$. The Cu electrode gains mass with a red / brown coating. The solution turns lighter blue as $[\text{Cu}^{2+}]$ decreases. K^+ ions move into the solution from the salt bridge to balance the charge. (ii) Cell 2: $E^\circ_{\text{cell}} = E^\circ_{\text{red}} - E^\circ_{\text{ox}}$ $1.10 \text{ V} = E^\circ(\text{Cu}^{2+} / \text{Cu}) - (-0.76 \text{ V})$ $E^\circ(\text{Cu}^{2+} / \text{Cu}) = +0.34 \text{ V}$ Cell 1: $1.17 \text{ V} = E^\circ(\text{MnO}_4^- / \text{Mn}^{2+}) - 0.34 \text{ V}$ $E^\circ(\text{MnO}_4^- / \text{Mn}^{2+}) = +1.51 \text{ V}$ For the reaction $\text{Zn} + \text{MnO}_4^-$, Zn would be oxidised and MnO_4^- reduced: $E^\circ_{\text{cell}} = 1.51 - (-0.76) = +2.27 \text{ V}$ The cell potential is positive, so a reaction will occur. Zn metal would disappear as it gets oxidised to Zn^{2+} ions, and the purple solution would lighten as MnO_4^- ions react to form Mn^{2+} ions. AND / OR Zn could react with H^+ ions to form H_2 gas and Zn^{2+} ions. $E^\circ_{\text{cell}} = 0 - (-0.76) = +0.76 \text{ V}$ The cell potential is positive, so this reaction would also occur. The Zn grey solid would disappear and bubbles of a colourless gas would be produced.			
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(c)	The energy required to turn one mol of an ionic solid into its gaseous ions under standard conditions reflects the strength of the electrostatic forces between oppositely charged ions in the crystal lattice. A higher energy requirement indicates stronger attractions and more energy required to separate the ions. Both MgCl_2 and NaCl are ionic compounds composed of Cl^- anions and metal cations. Since Cl^- is common to both, the difference in energy required arises from the properties of the cations. The Mg^{2+} ion has a smaller ionic radius and a greater charge than the Na^+ ion. The Mg^{2+} and Cl^- ions can therefore pack more closely together in MgCl_2, causing a greater strength of electrostatic attractions between them, and contributing to the higher energy requirement. Both cations are isoelectric, they have the same electron configuration ($1s^2 2s^2 2p^6$), but Mg^{2+} has 12 protons compared to Na^+ with 11, resulting in a higher charge density for Mg^{2+}. This leads to stronger electrostatic attractions between Mg^{2+} and Cl^- ions in MgCl_2, contributing to the higher energy requirement than NaCl.			
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Cut Scores

Scholarship	Outstanding Scholarship
17 – 24	25 – 32