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# SCHOLARSHIP EXEMPLAR



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QUALIFY FOR THE FUTURE WORLD  
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## Scholarship 2016 Chemistry

2.00 p.m. Monday 14 November 2016

Time allowed: Three hours

Total marks: 32

Check that the National Student Number (NSN) on your admission slip is the same as the number at the top of this page.

Pull out Resource Sheet S–CHEMR from the centre of this booklet.

You should answer ALL the questions in this booklet.

If you need more room for any answer, use the extra space provided at the back of this booklet.

Check that this booklet has pages 2–20 in the correct order and that none of these pages is blank.

You are advised to spend approximately 45 minutes on each question.

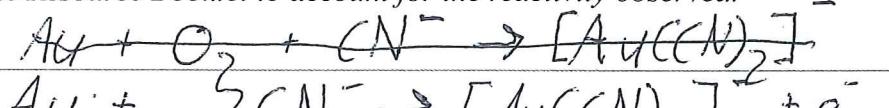
**YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.**

**QUESTION ONE**

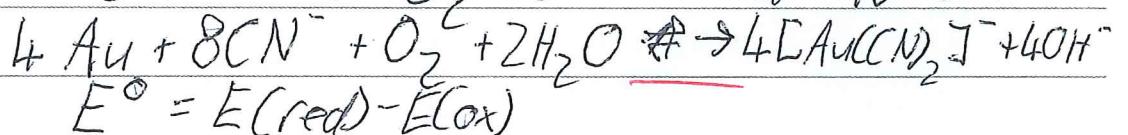
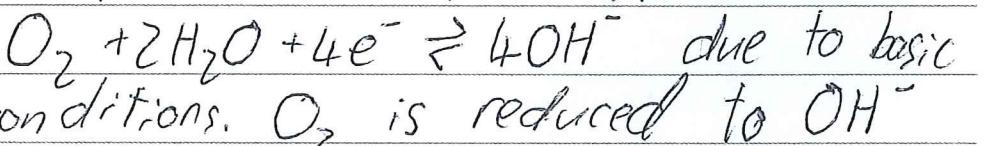
(a) A very dilute solution of sodium cyanide, NaCN, can be used to extract gold from its ore. Gold reacts with oxygen gas to form, in the presence of cyanide ions, the  $[\text{Au}(\text{CN})_2]^-$  ion. This process is carried out at pH 10–11 in order to ensure that  $\text{CN}^-$  ions are not converted to toxic hydrogen cyanide gas, HCN. The gold is then recovered as a metal by reacting the solution with zinc, which is converted to  $[\text{Zn}(\text{CN})_4]^{2-}$ . Sometimes the recovered gold also contains elemental silver as an impurity. This silver impurity can be removed from the gold by reacting the mixture with nitric acid.

- (i) Discuss the chemistry occurring in the extraction process described above by identifying the oxidation-reduction processes that take place, using electrode potentials to explain why gold can be extracted in this way, including the use of nitric acid to separate gold and silver.

*Include balanced equations for each step in the procedure, and use the electrode potentials in the Resource Booklet to account for the reactivity observed.*



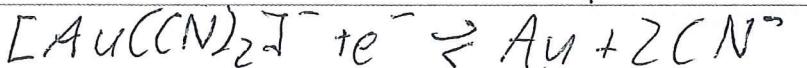
Au is oxidised to  $\text{Au}^{2+}$



$$E^\circ = E(\text{red}) - E(\text{ox})$$

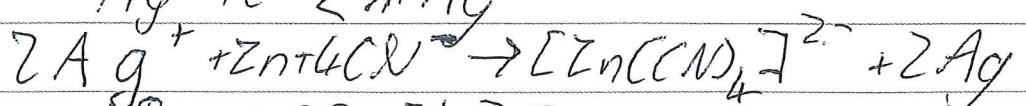
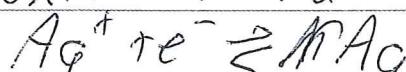
$$E^\circ = 0.40 - 0.60$$

$E^\circ = 1.00\text{V}$   $E^\circ > 0\text{V}$  and thus this reaction is spontaneous.



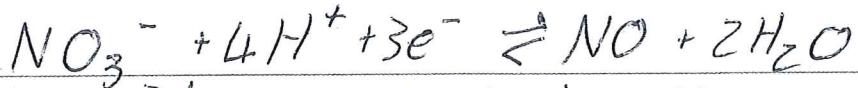
$$E^\circ = -0.60 - 1.26$$

$E^\circ = 0.66\text{V}$   $E^\circ > 0\text{V}$  and hence this reaction is spontaneous when Zn is oxidised and  $\text{Au}^{2+}$  is reduced.



$$E^\circ = 0.80 - 1.26$$

$E^\circ = 2.06\text{V}$  and hence  $\text{Ag(s)}$  is deposited upon reaction. This is a spontaneous and reaction. //



$$E^\circ = 0.94 - 0.80$$

$E^\circ = 0.14\text{V}$  and thus this reaction is spontaneous and puts Ag metal into solution as  $\text{Ag}^+$  ions. Thus the nitrate ion is reduced whilst Ag is oxidised. and Also Zn will be oxidised to  $[\text{Zn}(\text{CN})_4]^{2-}$  and thus remove  $\text{CN}^-$  ions from solution which will cause  $\text{Au}^{2+}$  to form along with  $\text{Gd}(\text{CN})_2$  which without silver being deposited. //

- (ii) The stability of the complex ions  $[\text{Au}(\text{CN})_2]^-$  is reflected in the size of the formation constant  $K_f$ , the equilibrium constant for the following equation:



Determine the concentration of sodium cyanide, NaCN (in mol L<sup>-1</sup>) that is required to keep 99% of the soluble gold in solution as the complex ion (assume that all the soluble gold is Au(I)).

Comment on the pH dependence of the composition of the equilibrium mixture.

$$pK_a(\text{HCN}) = 9.5$$

$$K_f = \frac{[\text{Au}(\text{CN})_2]^-}{[\text{Au}^{2+}][\text{CN}^-]^2}$$

$$4\sqrt[4]{4 \times 10^{28}} = \frac{0.99}{0.01[\text{CN}^-]^2}$$

$$[\text{CN}^-]^2 = \frac{0.99}{0.01 \times 4 \times 10^{28}}$$

$$[\text{CN}^-] = 4.97 \times 10^{-14} \text{ mol dm}^{-3}$$

$$[\text{NaCN}] = 4.97 \times 10^{-14} \text{ mol L}^{-1}$$

$$\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^- \quad pK_a = 9.5$$

below 9.5 pH of 9.5 the majority of the solution will be HCN instead of CN<sup>-</sup> and hence less  $[\text{Au}(\text{CN})_2]^-$  will be present. If the pH is higher than 9.5 then more CN<sup>-</sup> will be present and more  $[\text{Au}(\text{CN})_2]^-$  will be present. Thus the reaction should be done at high pH. //

- (b) Iron ore,  $\text{Fe}_2\text{O}_3$ , in a rock sample can be analysed by titration with acidified potassium dichromate solution,  $\text{K}_2\text{Cr}_2\text{O}_7/\text{H}^+$ .

A 2.8351 g sample of haematite rock containing  $\text{Fe}_2\text{O}_3$ , was dissolved in hot concentrated hydrochloric acid, HCl, and the solution diluted to 250.0 mL in a volumetric flask.

20.00 mL samples were pipetted into conical flasks, and a small excess of tin(II) chloride solution was added to change the colour from yellow to pale green. A saturated solution of mercury(II) chloride was then added until a small amount of white precipitate appeared. A few drops of diphenylamine sulfonate indicator were added. The resulting mixture was titrated with a standard solution of potassium dichromate. The average titre value recorded was 18.56 mL.

The standard solution of potassium dichromate was made by dissolving 2.5077 g of  $\text{K}_2\text{Cr}_2\text{O}_7$  in sufficient water to give 500.0 mL of solution.

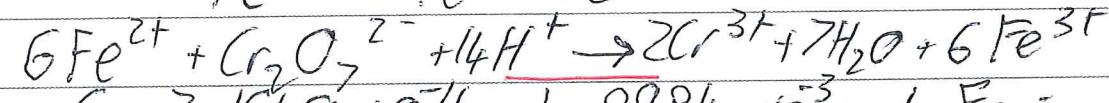
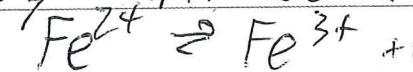
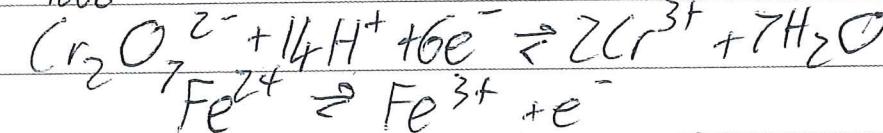
- (i) Calculate the % composition of  $\text{Fe}_2\text{O}_3$  in the haematite sample.

$$\text{M}(\text{K}_2\text{Cr}_2\text{O}_7) = 2 \times 39.1 + 2 \times 52 + 7 \times 16 \\ = 294.2 \text{ g mol}^{-1}$$

$$\frac{2.5077}{294.2} = 8.52 \times 10^{-3} \text{ mol s}$$

$$\frac{8.52 \times 10^{-3}}{(500/1000)} = 0.017048 \text{ mol dm}^{-3} \text{ dichromate}$$

$$\frac{18.56}{1000} \times 0.017048 = 3.1640 \times 10^{-4} \text{ mol K}_2\text{Cr}_2\text{O}_7$$



$$6 \times 3.1640 \times 10^{-4} = 1.8984 \times 10^{-3} \text{ mol Fe ions}$$

$$\text{in } 20\text{ mL} \quad \frac{1.8984 \times 10^{-3}}{(20/1000)} = 0.094921 \text{ mol L}^{-1} \text{ Fe}$$

$$\text{ions. } C = \frac{n}{V}$$

$$0.094921 = \frac{n}{0.25}$$

$$n = 0.023730 \text{ mol s in the ore}$$

$$0.023730 \times (2 \times 55.9 + 3 \times 16) = //$$

- (ii) Explain why it is not appropriate to carry out this titration using potassium permanganate,  $\text{KMnO}_4$ , instead of potassium dichromate, as the oxidant.

The  $\text{MnO}_4^-$  could reduce oxidise  $\text{Cl}^-$  to  $\text{Cl}_2$  as  $\text{Cr}_2\text{O}_7^{2-}$  is not strong enough to do this but  $\text{MnO}_4^-$  is.

$$E^\circ(\text{MnO}_4^-) = 1.51 - 1.40$$

$$E^\circ(\text{MnO}_4^-) = 0.11 \text{ V, spontaneous}$$

$$E^\circ(\text{Cr}_2\text{O}_7^{2-}) = 1.36 - 1.40$$

$$E^\circ(\text{Cr}_2\text{O}_7^{2-}) = 0.04 \text{ V, not spontaneous}$$

$\text{Cl}_2(g)$  is poisonous and thus  $\text{KMnO}_4$  should not be used. //

## QUESTION TWO

(a) Compound A, has the molecular formula  $C_{10}H_{17}O_2Cl$  and has one chiral centre. Compound A does not decolourise bromine water. When Compound A is reacted with dilute acid solution, two products, Compound B and Compound C, are formed.

Compound B does not have a chiral centre. When subjected to mass spectrometry, it is found to have a molecular ion peak at  $m/z = 114$ .

Compound C, when subjected to mass spectrometry, is found to have a molecular ion peak at  $m/z = 108$ . There is also a  $M+2$  peak present, and the ratio of the heights of the M and  $M+2$  peaks is 3:1.

When Compound C is treated with thionyl chloride,  $SOCl_2$ , Compound D is produced.

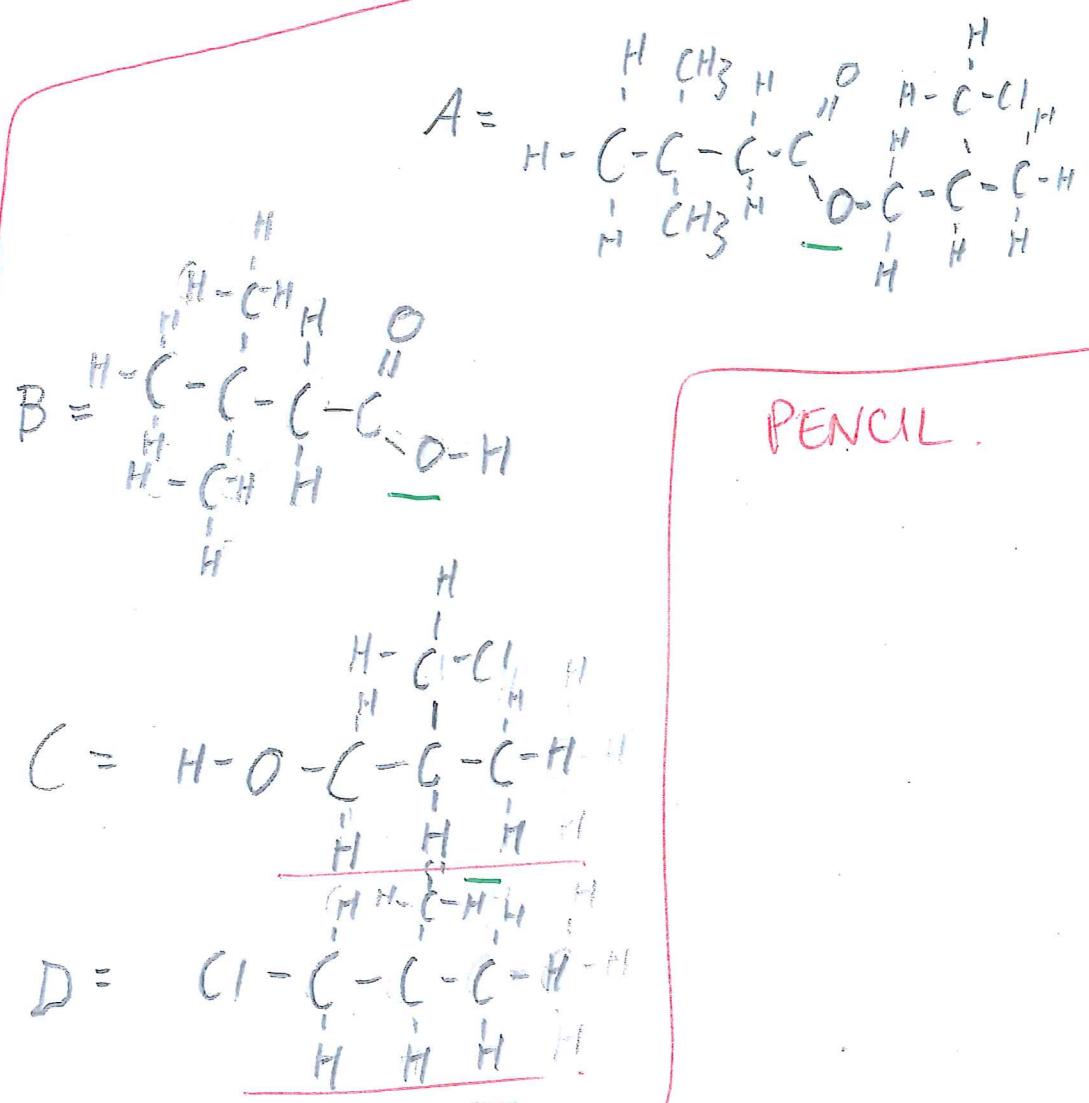
The  $^{13}C$  NMR spectrum of all Compounds A to D are available in the Resource Booklet.

The IR spectrum of Compound B is also available.

Determine the identity of Compounds A to D.

Justify your answers by linking your choices to the spectra provided and the information given above.

Compounds A to D:



Justifications: A is not an alkene. A must be an ester to form two products upon hydrolysis. B does not contain the Cl as C has a 3:1 split with the  $M:M+2$  peak. A has 8 carbon environments in the NMR. Thus one C must have ~~two~~<sup>three</sup> methyl groups off it to be ~~both~~ chiral and only have 8 environments. This the chiral part contains the of the ester has the Cl in it as it is C. C has 4 distinct carbon environments ~~with one being a methyl~~ the other being the chiral carbon and another being a carbon attached to a Cl. B has an O-H shown by the wide bit of the IR spectrum at around  $3000\text{ cm}^{-1}$  and also a C=O shown by the peak at  $1700\text{ cm}^{-1}$ . This must be the acid part. It has 4 carbon environments with one being the acid carbon.  $114 - 16 - 1 - 16 - 12 = 69$  left in the chain, ~~this indicates around 12~~ The molecule has ~~no~~ ~~one~~ dimethyl carbons. No single methyl carbons. C has one chiral and since it is ~~so~~ must be a straight chain alcohol the Cl must be on the O-H as one & For D to lose an environment from C the alcohol must be on a carbon connected to a carbon that has the chlorine group on it.

- (b) Nitrogen dioxide,  $\text{NO}_2$ , is a brown gas that is often a component of air pollution. It is in equilibrium with a dimeric species, dinitrogen tetroxide,  $\text{N}_2\text{O}_4$ . Dinitrogen tetroxide is used as a rocket fuel in combination with hydrazine,  $\text{N}_2\text{H}_4$ , and the products of the combustion reaction of this mixture are gaseous nitrogen and water.

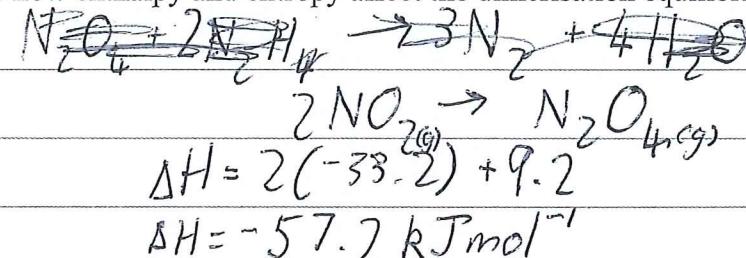
Table 1: Selected thermochemical data

Chemical	Enthalpy of formation/kJ mol <sup>-1</sup>	Boiling point/°C
$\text{NO}_2$	33.2	21
$\text{N}_2\text{O}_4$	9.2	21
$\text{N}_2\text{H}_4$	50.6	114
$\text{N}_2$	0	-196
$\text{H}_2\text{O}$	-286	100

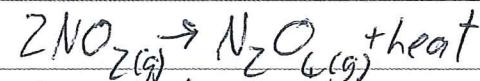
Table 2: Selected bond enthalpies (kJ mol<sup>-1</sup>)

N–N ( $\text{N}_2$ )	945	O–H ( $\text{H}_2\text{O}$ )	499
N–N ( $\text{N}_2\text{O}_4$ )	57	N–O ( $\text{NO}_2$ )	305
N–N ( $\text{N}_2\text{H}_4$ )	298	N–H	391

- (i) Discuss how enthalpy and entropy affect the dimerisation equilibrium.



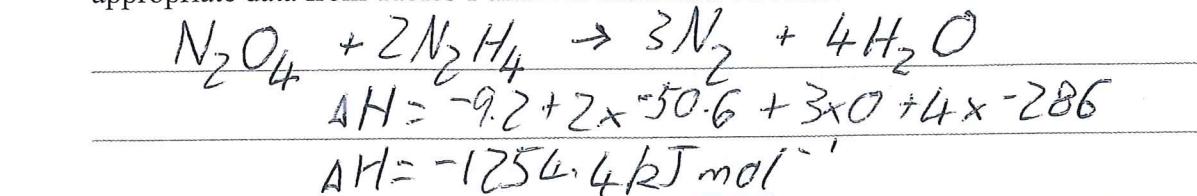
The enthalpy change is negative and thus energy is released in the forward reaction



Thus low temperatures favour the forward reaction as heat will be produced. In terms of entropy the change is negative as two moles of gas become one mole, this is a decrease in disorder of the system and hence a decrease in entropy.

As  $\Delta G = \Delta H - T\Delta S$ , a decrease in temperature makes  $\Delta G$  more negative as  $-T\Delta S$  will be less positive. Thus a low temperature favours the forward reaction in terms of both enthalpy and entropy. //

- (ii) Discuss why a  $\text{N}_2\text{O}_4 / \text{N}_2\text{H}_4$  mixture is suitable for use as a rocket fuel, linking appropriate data from Tables 1 and 2 to molecular structures.



$$\begin{aligned} 57 + 4(945) + 4(499) + 2 \times 298 + 8 \times 391 \\ -(3 \times 945 + 8 \times 499) = -1256.4 \end{aligned}$$

$$\text{N}-\text{OCN}_2\text{O}_4 = 447.9 \text{ kJ mol}^{-1}$$

Thus it is favourable for  $\text{NO}_2$  to form  $\text{N}_2\text{O}_4$  and the reaction of  $\text{N}_2\text{H}_4 + \text{N}_2\text{O}_4$  is exothermic. The reactants have more energy in their forward reaction releases energy and thus // a mixture can be used as rocket fuel. The energy required to break the bonds is less than the energy released from forming new bonds. //

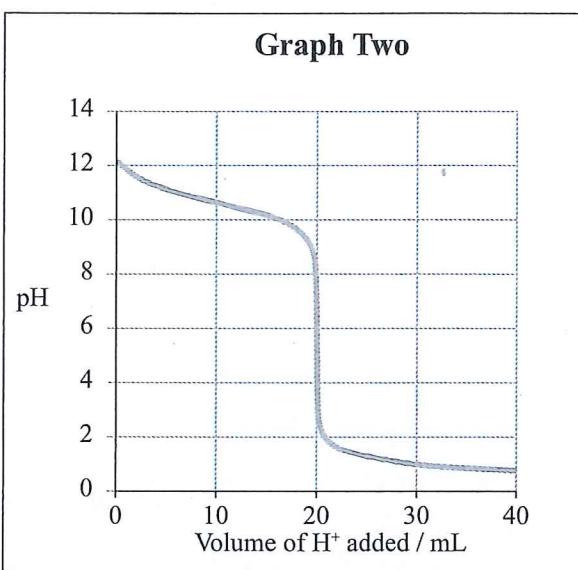
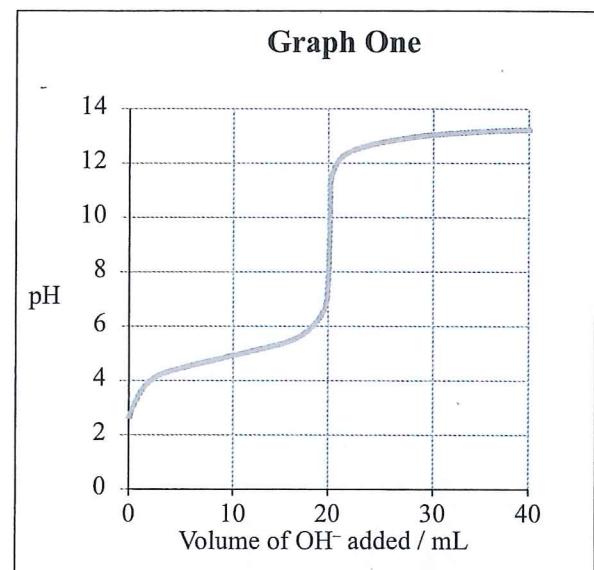
**QUESTION THREE**

- (a) The pH titration curves for the following reactions are shown below.

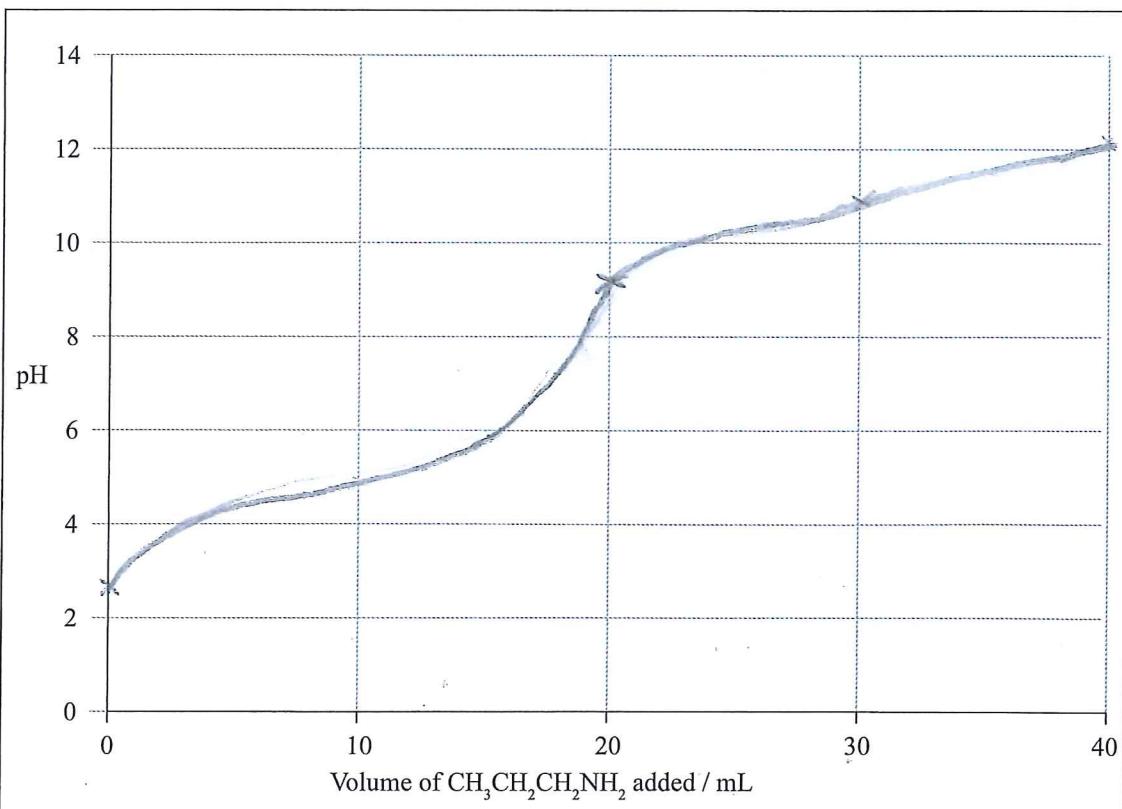
20 mL of 0.5 mol L<sup>-1</sup> propanoic acid, CH<sub>3</sub>CH<sub>2</sub>COOH, titrated with 0.5 mol L<sup>-1</sup> sodium hydroxide (Graph One)

20 mL of 0.5 mol L<sup>-1</sup> propan-1-amine, CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub>, titrated with 0.5 mol L<sup>-1</sup> hydrochloric acid (Graph Two)

$$pK_a(\text{CH}_3\text{CH}_2\text{COOH}) = 4.88 \quad pK_a(\text{CH}_3\text{CH}_2\text{CH}_2\text{NH}_3^+) = 10.71$$



- (i) Sketch the titration curve for the reaction of 20 mL of 0.5 mol L<sup>-1</sup> propanoic acid, CH<sub>3</sub>CH<sub>2</sub>COOH, titrated with 0.5 mol L<sup>-1</sup> propan-1-amine, CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub>.



- (ii) Aqueous solutions containing equal concentrations of propanoic acid and propan-1-amine are mixed.

Complete the table below by giving the **dominant** species that would be present in the mixture at the given pH.

Give reasons linked to the graphs, and show relevant equations.

pH	1	8	13
Dominant species	CH <sub>3</sub> CH <sub>2</sub> COOH CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> NH <sub>3</sub> <sup>+</sup>	CH <sub>3</sub> CH <sub>2</sub> COO <sup>-</sup> CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> NH <sub>3</sub>	CH <sub>3</sub> CH <sub>2</sub> COO <sup>-</sup> CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> NH <sub>2</sub>

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\frac{[\text{base}]}{[\text{acid}]} = 10^{pH - pK_a}$$

propanoic acid

pH 1

$$\frac{[\text{base}]}{[\text{acid}]} = 10^{1-4.88} \\ = 1.32 \times 10^{-4}$$

Thus most dominant form

is CH<sub>3</sub>CH<sub>2</sub>COOH as ratio of base to acid is low

pH 8

$$\frac{[\text{base}]}{[\text{acid}]} = 10^{8-4.88} \\ = 1318$$

Thus base is

dominant as CH<sub>3</sub>CH<sub>2</sub>COO<sup>-</sup>

pH 13

$$\frac{[\text{base}]}{[\text{acid}]} = 10^{13-4.88} \\ = 131 \times 10^6$$

Thus base is significantly more present and most dominant form is CH<sub>3</sub>CH<sub>2</sub>COO<sup>-</sup>

Propan-1-amine

pH 1 [base] = 10<sup>1-10.71</sup>

$$\frac{[\text{base}]}{[\text{acid}]} = 1.9498 \times 10^{-10}$$

Thus dominant form is acid which is

CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>3</sub><sup>+</sup>

$$pH 8 \frac{[\text{base}]}{[\text{acid}]} = 10^{8-10.71} \\ = 1.95 \times 10^{-3}$$

Thus acid is still dominant in form

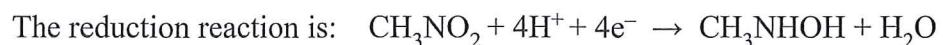
of CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>3</sub><sup>+</sup>

$$pH 13 \frac{[\text{base}]}{[\text{acid}]} = 10^{13-10.71} \\ = 195$$

Thus dominant form is base as CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>NH<sub>2</sub>

- (b) An organic nitro-compound,  $\text{CH}_3\text{NO}_2$ , is reduced by electrolysis in an aqueous ethanoate buffer solution at pH 5.0. In the solution, the total ethanoate concentration,  $[\text{CH}_3\text{COOH}] + [\text{CH}_3\text{COO}^-]$  is known to be 0.500 mol L<sup>-1</sup>.

A 300 mL sample of the buffered solution contains 0.0100 mol L<sup>-1</sup>  $\text{CH}_3\text{NO}_2$  that has been completely reduced by the electrolytic process.



- (i) Calculate the pH of the solution after reduction of  $\text{CH}_3\text{NO}_2$  is complete.

Assume that the effect of the hydrolysis of the methylhydroxylamine,  $\text{CH}_3\text{NHOH}$ , is negligible at this pH.

$$K_a(\text{CH}_3\text{COOH}) = 1.75 \times 10^{-5} \quad pK_a(\text{CH}_3\text{COOH}) = 4.76$$

$$\text{let } x = [\text{CH}_3\text{COH}]$$

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

$$5 = 4.76 + \log \frac{0.5-x}{x}$$

$$\underline{0.5-x = 1.738}$$

$$\underline{x}$$

$$\underline{0.5-x = 1.738x}$$

$$\underline{x = 0.1826 \text{ mol L}^{-1}}$$

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.75 \times 10^{-5} = \frac{(0.5 - 0.1826)(0.1826)}{0.1826}$$

$$\underline{[\text{H}^+] = 1.007 \times 10^{-5}}$$

$$\underline{[\text{H}^+] = 10^{-5}}$$

$$\underline{n(\text{H}^+) = \frac{300}{1000} \times 10^{-5}}$$

$$\underline{0.01 \times \frac{300}{1000} = 3 \times 10^{-3} \text{ mol CH}_3\text{NO}_2}$$

$$\underline{[\text{CH}_3\text{COOH}] = 0.1826 \text{ mol L}^{-1}}$$

$$\underline{\frac{300}{1000} \times 0.1826 = 0.05478 \text{ mol CH}_3\text{COOH}}$$

$$\underline{0.05478 - 3 \times 10^{-3} = 0.05178 \text{ mol CH}_3\text{COOH}}$$

$$\underline{[\text{CH}_3\text{COOH}] = 0.1726 \text{ mol L}^{-1}}$$

$$\underline{n(\text{CH}_3\text{OO}^-) = 0.5 - 0.1826 + 3 \times 10^{-3}} \\ \underline{= 0.3204 \text{ mol}}$$

$$\underline{[\text{CH}_3\text{OO}^-] = 1.068}$$

$$pH = pK_a + \log \frac{[\text{CH}_3\text{OO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$pH = 4.76 + \log \frac{1.068}{0.1726}$$

$$pH = 5.55 \text{ after reaction} \checkmark$$

- (ii) Evaluate the validity of the assumption that hydrolysis of methylhydroxylamine does not occur at this pH.

$$K_a(\text{CH}_3\text{NH}_2\text{OH}^+) = 1.09 \times 10^{-6} \quad pK_a(\text{CH}_3\text{NH}_2\text{OH}^+) = 5.96$$

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$5.55 = 5.96 + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\frac{[\text{base}]}{[\text{acid}]} = 0.3904$$

There is more acid than base at pH of 5.55 and thus the assumption of no hydrolysis is incorrect as there is more of the acid form than the basic form //

**QUESTION FOUR**

The melting points of a selection of hydrides are given in the table below. The position of the hydrides in the table reflects the relative position of the non-hydrogen element in the Periodic Table.

Period	1	2	13	14	15	16	17
1							
2				$\text{CH}_4$ -182°C	$\text{NH}_3$ -78°C	$\text{H}_2\text{O}$ 0°C	$\text{HF}$ -83°C
3				$\text{SiH}_4$ -185°C	$\text{PH}_3$ -133°C	$\text{H}_2\text{S}$ -86°C	
4					$\text{AsH}_3$ -116°C	$\text{H}_2\text{Se}$ -66°C	
5					$\text{SbH}_3$ -88°C	$\text{H}_2\text{Te}$ -49°C	

- (a) (i) Discuss the variation in the melting points of the hydrides of the elements of Period 3 ( $\text{SiH}_4$ ,  $\text{PH}_3$ , and  $\text{H}_2\text{S}$ ).

The melting point increases across the period from  $\text{SiH}_4$  to  $\text{H}_2\text{S}$ . The three compounds all have covalent bonds and thus their boiling points are determined by the strength of their intermolecular forces. More electrons are added to the non-hydrogen element from  $\text{SiH}_4$  to  $\text{H}_2\text{S}$  and thus the strength of Van der Waal's forces increases. This causes an increased melting point across the three hydrides. All 3 ~~the~~ hydrides have relatively similar ~~electronegativity~~ across the period. Also across the period the electronegativity of the non-hydrogen atoms increases. This causes the

hydrides to become more polar and thus increases the strength of the intermolecular forces and hence the melting point. None of the three hydrides can form hydrogen bonds and thus there is no contribution to the strength of intermolecular forces from this.

- (ii) Discuss the trend in the melting points of the hydrides of the elements of Group 15 ( $\text{NH}_3$  to  $\text{SbH}_3$ ).

The melting points ~~decrease~~ decrease from  $\text{NH}_3$  to  $\text{PH}_3$  but then ~~increase~~ increase from  $\text{PH}_3$  to  $\text{SbH}_3$ . This is primarily due to variations in the polarity of the hydride. Electronegativity decreases down the group from N to Sb but Hydrogen's electronegativity is constant. The difference in the two electronegativities indicates how polar the molecule is and hence how strong the intermolecular forces are. The electronegativity difference between P and H must be less than the other 3 hydrides and hence it has the lowest melting point. After P, the difference increases and hence so does the melting point.  $\text{NH}_3$  is also most likely to have a higher melting point than the other 3 and as it can form hydrogen bonds, the N has a strong enough pull on the bonding electrons for the H to have a partial charge and hence it will be attracted to the N of other  $\text{NH}_3$  molecules. Down the group there are more electrons and hence stronger Van der Waal's forces which also increases the melting point down the group with  $\text{NH}_3$  being an exception to the trend.

Question Four continues on the following page.

- (iii) The acidic properties of the non-metal hydrides increase left to right across a period, and down a group.

Account for these trends using the hydrides of Period 2 elements ( $\text{CH}_4$  to  $\text{HF}$ ) and those of Group 16 elements ( $\text{H}_2\text{O}$  to  $\text{H}_2\text{Te}$ ).

Across period 2 electronegativity increases. This makes the H more positive and hence more likely to be attracted to a negative charge and thus it is more acidic. Down the group the bond lengths increase due to electrons being added to orbitals farther from the nucleus. This makes the bond to the H weaker and hence the H is more likely to be donated and thus acidity increases down the group. //

- (b) The hydrides of the elements of Groups 1 and 2 are all solids at room temperature, but apart from  $\text{LiH}$ , which melts at  $689^\circ\text{C}$ , the metallic hydrides decompose at temperatures around  $250^\circ\text{C}$ . Because of this it is not possible to use melting points to compare the strength of the metal hydride bond. To compare the strength of these bonds, some enthalpy data is provided in the table below.

Note: Lattice enthalpy is the standard enthalpy change for the conversion of an ionic solid into a gas of ions.

	$\Delta_f H$ / kJ mol <sup>-1</sup>	Lattice enthalpy / kJ mol <sup>-1</sup>
$\text{LiH}$	-90.65	912.5
$\text{NaH}$	-56.4	845.7
$\text{KH}$	-57.8	741.9

	$\Delta_f H$ / kJ mol <sup>-1</sup>	Lattice enthalpy / kJ mol <sup>-1</sup>
$\text{MgH}_2$	-75.2	2721
$\text{CaH}_2$	-186	2458

Compare and contrast the expected strength of the metal hydride bonds of the Group 1 and 2 elements, and justify whether the lattice enthalpy data or the enthalpy of formation data is the better predictor of the expected trends in the strength of the metal hydride bonds.

Metal hydride bond strength would be expected to decrease down the group as the ions become less charge dense due to their increased volume as electrons are added to shells further from the nucleus. Group 2 would have stronger bonds as they are more charge dense due to losing two electrons which also decreases radius and hence volume, the cation is more attracted to anion the anion. Lattice enthalpy is not the best indicator as it involves changing into gaseous ions rather than melting which is the change between solid and liquid.  $\Delta H_f$  is more indicative as it shows how much stronger the hydride bonds are compared to the metallic bonds of isolated atoms. //

QUESTION  
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**Scholarship paper – Total score = 24.**

<b>Q</b>	<b>Score</b>	<b>Annotation notes</b>
1	6	<p>The candidate has interpreted the question and applied their chemistry knowledge to answer the question. Their answer includes appropriate chemistry symbols and terms. Calculations are clearly set out and organised.</p> <p>In part (a)(i), the candidate has identified some of the oxidation-reduction processes taking place and calculated electrode potentials for them.</p> <p>In part (a)(ii), the candidate has determined the concentration of sodium cyanide and commented on the pH dependence of the composition of the equilibrium mixture.</p> <p>In part (b)(i), the concentration of the standard solution is calculated and the titration reaction equation is given.</p> <p>Finally, in part (b)(ii), the candidate has explained why dichromate is used in the titration but not permanganate; supported by calculations.</p>
2	7	<p>Structures have been clearly drawn and the candidate has selected relevant information and used it to justify the identification of the compounds. The candidate has applied appropriate chemistry principles in relation to enthalpy and entropy and given reasoned explanations.</p> <p>In part (a), structures for compounds C &amp; D are correct with some justification for their identification.</p> <p>In part (b), enthalpies for reactions in both parts (i) and (ii) are correct and there is some discussion of both enthalpy and entropy changes.</p>
3	6	<p>The candidate has selected and applied their chemistry knowledge to a complex situation and given clear, succinct justifications for their answer. They have demonstrated understanding of a complex problem, and shown their thinking, despite failing to complete the calculation.</p> <p>In part (a)(i), the candidate sketched an appropriate titration curve for the weak acid weak base titration.</p> <p>In part (a)(ii), the candidate identified the dominant species with justifications.</p> <p>In part (b)(i), the candidate attempted the calculation and could determine the concentration of the ethanoic acid.</p>
4	5	<p>The candidate has communicated key chemistry concepts in relation to intermolecular forces. They have analysed the data given and applied their knowledge to explain the trends. The candidate has logically developed their ideas around periodic trends.</p>