VICTORIA JUNIOR COLLEGE
JC 2 PRELIMINARY EXAMINATION
Higher 2

CANDIDATE NAME ……

CT GROUP ……

CHEMISTRY **9729/02**

Paper 2 Structured **15 September 2020**

Candidates answer on the Question Paper. **2 hours**

Additional Materials: *Data Booklet*

READ THESE INSTRUCTIONS FIRST

Write your name and CT group on all the work you hand in.
Write in dark blue or black pen on both sides of the paper.
You may use a soft pencil for any diagrams, graphs or rough working.
Do not use staples, paper clips, highlighters, glue or correction fluid.

Answer **all** questions.

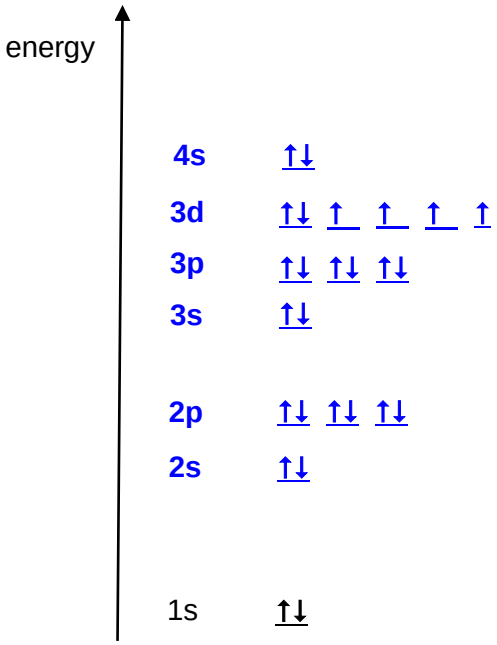
The number of marks is given in brackets [ ] at the end of each question or part question.

For Examiner's Use	
1	/ 20
2	/ 11
3	/ 8
4	/ 19
5	/ 17
Total	/ 75

This document consists of **16** printed pages.

Answer **all** the questions in the spaces provided.

1 (a) (i) Complete the diagram to show the relative energies of all the electrons in an iron atom.



[1]

(ii) Write an equation for the second ionisation energy of iron.

- **$Fe^+(g) \rightarrow Fe^{2+}(g) + e^-$** [1]

(iii) Explain why the second ionisation energy of iron is lower than that of chromium.

- **For the second ionisation energy, 4s electron is removed from iron while 3d electron is removed from chromium. As 4s electron is further away from the nucleus than 3d electron, less energy is needed to remove it, resulting in a lower second ionisation energy.** [1]

(b) A piece of rusted iron was analysed to find out how much of the iron had been oxidised to rust (also known as hydrated iron(III) oxide). A small sample of the rusted iron was dissolved in excess dilute sulfuric acid to give 250 cm^3 of solution. The resultant solution contains Fe^{2+} and Fe^{3+} from iron and rust respectively.

(i) 25.0 cm^3 of this solution required 16.90 cm^3 of 0.0200 mol dm^{-3} $KMnO_4$ for complete oxidation of Fe^{2+}. Calculate the amount of Fe^{2+} in 25.0 cm^3 of the solution.

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

Amount of Fe^{2+} in 25.0 cm^3 of the solution
$= 5 \times 0.0200 \times 16.90 \times 10^{-3}$

- **= 0.00169 mol** [1]

(ii) To another 25.0 cm^3 of the solution, an oxidising agent was added to convert all the Fe^{2+} ions present to Fe^{3+} ions. The Fe^{3+} ions were then titrated with 0.100 mol dm^{-3} $EDTA^{4-}$ solution and 17.60 cm^3 was required.

Assuming 1 mol of $EDTA^{4-}$ reacts with 1 mol of Fe^{3+}, calculate the amount of Fe^{3+} in 25.0 cm^3 of the solution.

Amount of Fe^{3+} in 25.0 cm^3 of the solution
$= 0.100 \times 17.60 \times 10^{-3}$

- **= 0.00176 mol** [1]

(iii) From your answers in **(b)(i)** and **(b)(ii)**, calculate the amount of original Fe^{3+} in 250 cm^3 of the rusted iron solution.

Amount of Fe^{3+} in 250 cm^3 solution
= (0.00176 – 0.00169) × 250/25.0

- **= 7.00×10^{-4} mol** [1]

(iv) Determine the percentage of iron that had been oxidised to rust in the sample.

Percentage of iron that had rusted
= $\frac{0.0007}{0.0176} \times 100\%$

- **= 3.98%** [1]

(c) (i) Explain if a real gas, such as N_2O_3, behaves more or less ideally at:

- high pressures

- **At high pressures, gas particles are closer together. The gas particles occupy significant volume compared to the volume of the container. (OR Forces of attraction between the gas particles are significant / no longer negligible.)** A real gas behaves **less** ideally at high pressures.

- high temperatures

- **At high temperatures, forces of attraction between the gas particles are insignificant / negligible as the gas particles have sufficient (high kinetic) energy to overcome them.** A real gas behaves **more** ideally at high temperatures.

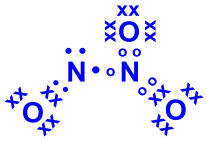
[2]

(ii) At room temperature, N_2O_3 dissociates as shown.

$$N_2O_3(g) \rightleftharpoons NO(g) + NO_2(g)$$

This dissociation involves homolytic fission of N–N bond.

Draw a 'dot-and-cross' diagram for the N_2O_3 molecule.

- 

[1]

(iii) Hence, state the shape around **each** N atom in the N_2O_3 molecule.

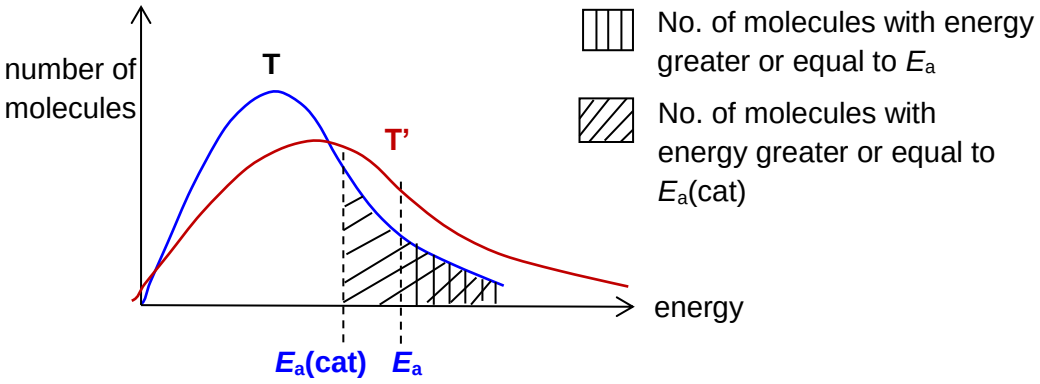
- **Bent** and **trigonal planar**

[1]

(iv) Suggest the common feature of NO and NO_2 that allows each of them to undergo dimerisation to form N_2O_2 and N_2O_4 respectively.

- **Both molecules have a single unpaired electron on N atom. (OR Both are free radicals.)** [1]

(d) (i) With an appropriate sketch of Boltzmann distribution, explain how a catalyst is able to increase the rate of a chemical reaction for a given temperature T.



- **A catalyst speeds up the rate of reaction by providing a different/alternative reaction path which has a lower activation energy. As shown on the diagram, the number of molecules with energy greater or equal to the lowered activation energy E_a(cat) will increase. This results in an increase in the frequency of effective collisions. Hence, the rate of reaction increases.** [2]

(ii) On the sketch in **d(i)**, draw a new distribution curve at a higher temperature T', **clearly labelled T'.** [1]

(e) **X** is an organic liquid that is found as a contaminant in crude oil. It contains only C, H and S. When 0.841 g of **X** was subjected to complete combustion, 1.76 g of CO_2, 0.360 g of H_2O and a gas which can decolourise acidified potassium manganate (VII) were formed.

(i) What is the empirical formula of **X**?

$m_C + m_H + m_S = 0.841$ g

$m_C = \frac{12.0}{44.0} \times 1.76 = 0.480$ g

$m_H = \frac{2 \times 1.0}{18.0} \times 0.360 = 0.040$ g

$m_S = 0.841 - 0.480 - 0.040 = 0.321$ g

Element	C	H	S
• Mass	0.480	0.040	0.321
Mole ratio	0.480 ÷ 12.0 = 0.04	0.040 ÷ 1.0 = 0.04	0.321 ÷ 32.1 = 0.01
Simplest mole ratio	4	4	1

- **Empirical formula is C_4H_4S.** [2]

(ii) **X** is unsaturated and has a structure in which all the carbon and sulfur atoms are connected together in a five-membered ring. Draw the structure of **X**.

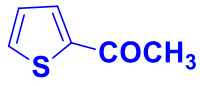
- 

[1]

(iii) Hence, construct the equation for the complete combustion of **X**.

- $C_4H_4S + 6O_2 \rightarrow 4CO_2 + 2H_2O + SO_2$ [1]

(iv) **X** behaves similarly like benzene in its reactions with electrophile. Suggest a possible structure for the product obtained when **X** is reacted with CH_3COCl in the presence of a Lewis acid catalyst.

- 

[1]

[Total: 20]

2 (a) Water is a very stable molecule and was thought to be a chemical element. In 1800, one of the earliest form of electrochemical cells was invented and scientists, Nicholson and Carlyle, caused a big stir by decomposing water into hydrogen and oxygen by electrolysis.

In one experiment to electrolyse water, copper was used at both electrodes. A gas was generated only at one electrode during the electrolysis.

Write half-equations for the reaction that took place during the electrolysis.

anode: **$Cu \rightarrow Cu^{2+} + 2e^-$**
cathode: **$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$** [1]

(b) In another electrolysis experiment with copper electrodes, the electrolyte was replaced with an acidified solution of copper(II) sulfate, $CuSO_4$.

During this electrolysis, the mass of cathode increased with simultaneous evolution of H_2 gas.

(i) When a constant current of 2.0 A was applied for 30 min, the mass of cathode increased by 0.635 g. Calculate the volume of H_2 evolved, measured at r.t.p.

$Q = I \times t = 2.0 \times 30 \times 60 = 3600$ C

- **Amount of e^- passed through = $\frac{3600}{96500}$ = 0.0373 mol**

Amount of Cu formed = $\frac{0.635}{63.5}$ = 0.0100 mol
At the cathode, simultaneous discharge of Cu and H_2 occurred.
$Cu^{2+} + 2e^- \rightarrow Cu$; $2H^+ + 2e^- \rightarrow H_2$
Amount of e^- passed through = 2 × amount of Cu formed + 2 × amount of H_2 formed

- **Amount of H_2 formed = $\frac{0.0373 - 2(0.0100)}{2}$ = 8.65×10^{-3} mol (ecf)**
- **Maximum volume of H_2 = $8.65 \times 10^{-3} \times 24 = 0.208$ dm^3 (ecf)**

[3]

(ii) When $Cu(OH)_2(s)$ started to precipitate, the concentration of Cu^{2+} in the electrolyte was 0.100 mol dm^{-3} and the pH is 4.84.

Write the expression for the solubility product, K_{sp}, of $Cu(OH)_2(s)$.

- **$K_{sp} = [Cu^{2+}][OH^-]^2$** [1]

(iii) Calculate the K_{sp} of $Cu(OH)_2(s)$, stating the units.
pOH = 14 – 4.84 = 9.16

- **$[OH^-] = 10^{-9.16} = 6.92 \times 10^{-10}$ mol dm^{-3}**
- **$K_{sp} = [Cu^{2+}][OH^-]^2 = 4.79 \times 10^{-20}$ mol^3 dm^{-9}** [2]

(c) For a reversible reaction, $M^{3+}(aq) + ne^- \rightleftharpoons M^{(3-n)+}(aq)$, the Nernst equation below shows the dependence of its reduction potential, $E_{M^{3+}/M^{(3-n)+}}$, to the concentration of M^{3+} and $M^{(3-n)+}$ as expressed in Q, at 298 K.

$$E_{M^{3+}/M^{(3-n)+}} = E^{\ominus}_{M^{3+}/M^{(3-n)+}} - \frac{0.0252}{n} \ln Q$$

where $E^{\ominus}_{M^{3+}/M^{(3-n)+}}$ is the standard reduction potential, n is the amount of electron transferred and Q is the reaction quotient, a measure of the relative concentrations of products and reactants present in a reaction at a given time.

For the reversible reaction, $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$,

$$Q = \frac{1}{[Cu^{2+}]}$$

Q is numerically equal to the equilibrium constant, K_c, when the reaction achieves dynamic equilibrium.

(i) Using Nernst equation, calculate the $E_{Cu^{2+}/Cu}$ at the start of the precipitation during the electrolysis in **(b)**.

- $E_{Cu^{2+}/Cu} = E^{\ominus}_{Cu^{2+}/Cu} - \frac{0.0252}{n} \ln Q$

  $= +0.34 - \frac{0.0252}{2} \ln \frac{1}{0.1} = +0.311$ V [1]

(ii) Your answer to **(c)(i)** is also the reduction potential of equation **1** when the precipitation of $Cu(OH)_2$ just started as described in **b(ii).**

$$Cu(OH)_2(s) + 2e^- \rightleftharpoons Cu(s) + 2OH^-(aq) \qquad \text{equation 1}$$

Calculate the standard reduction potential of equation **1** using the Nernst equation.

$$E_{Cu(OH)_2/Cu} = E^{\ominus}_{Cu(OH)_2/Cu} - \frac{0.0252}{n} \ln Q$$

$$E^{\ominus}_{Cu(OH)_2/Cu} = E_{Cu(OH)_2/Cu} + \frac{0.0252}{n} \ln [OH^-]^2$$

- $= +0.311 + \frac{0.0252}{2} \ln (6.92 \times 10^{-10})^2$
- $= -0.221$ V (ecf if $[OH^-]$ calculated in (b)(iii) was wrong) [2]

(iii) Suggest why the electrolysis in **(b)** might stop if the electrolyte pH is allowed to rise above 4.84.

- **Concentration of OH^- increases when pH rises**. This will cause $E_{Cu(OH)_2/Cu}$ **to be less positive since position of equilibrium of equation 1 shift left but $E_{Cu^{2+}/Cu}$ is unchanged. Hence Cu(s) might be preferentially oxidised to insoluble $Cu(OH)_2$ at the anode, instead of oxidation of Cu(s) to Cu^{2+} and so the electrolysis in (b) stops.** [1]

[Total: 11]

3 (a) A yellow solution of chromate(VI), CrO_4^{2-}, is slowly reduced to green tetrahydroxochromate(III) ion, $[Cr(OH)_4]^-$, in an alkaline medium by arsenic(III) oxide, As_2O_3. The relevant half-equation involving arsenic oxides is given below.

$$2AsO_4^{3-} + 5H_2O + 4e^- \rightleftharpoons As_2O_3 + 10OH^- \qquad E^\ominus = -0.74V$$

(i) Write a balanced equation for the reaction between chromate(VI) and arsenic(III) oxide in an alkaline medium.

$CrO_4^{2-} + 4H_2O + 3e^- \rightarrow [Cr(OH)_4]^- + 4OH^-$

- **$4CrO_4^{2-} + 3As_2O_3 + H_2O + 14OH^- \rightarrow 4[Cr(OH)_4]^- + 6AsO_4^{3-}$** [1]

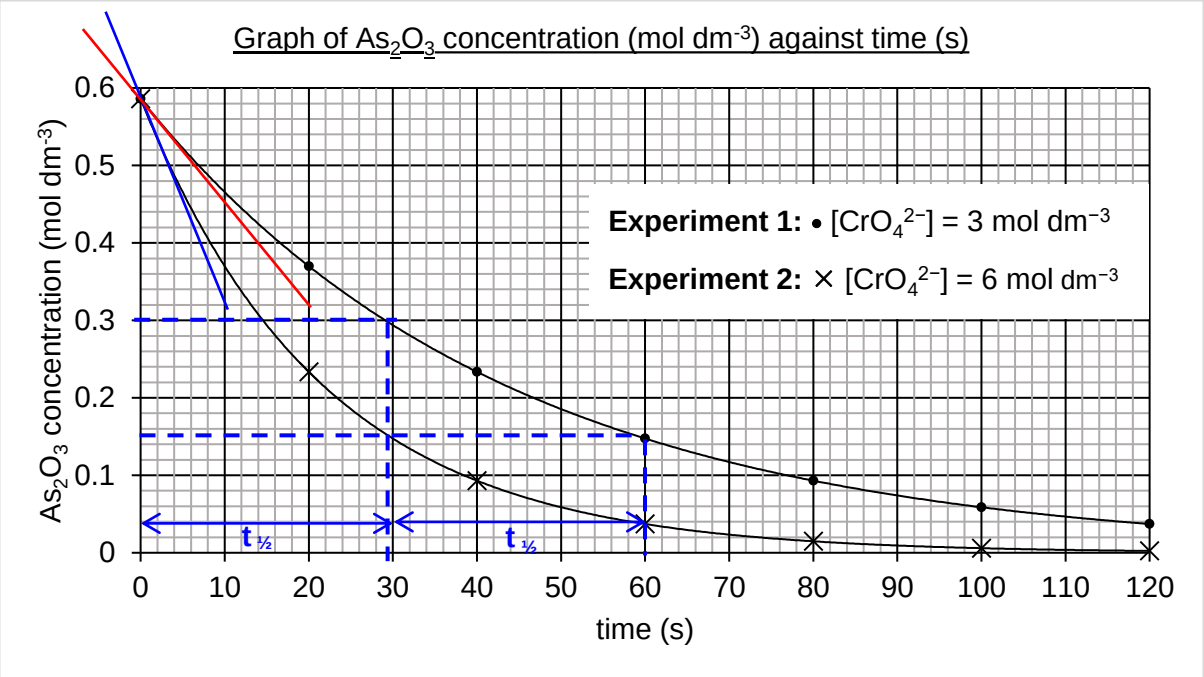
(ii) Briefly suggest, with explanation, one experimental method by which the rate of the reaction can be followed.

- **colorimetric method because CrO_4^{2-} and $[Cr(OH)_4]^-$ have different colours**
- **OR use a pH meter because the mixture will become less alkaline as the reaction progresses**
- **OR measure the electrical conductivity of the mixture because the mixture will have less ions as reaction progresses leading to lower electrical conductivity.**

[1]

(b) Two experiments were carried out, starting with different initial concentrations of chromate(VI) and the same initial concentration of As_2O_3 of 0.58 mol dm^{-3}.

At a fixed time interval, a portion of the reaction mixture was quenched and analysed to determine the concentration of As_2O_3. The following graphs were obtained.



(i) Use your graphs to determine the order of reaction with respect to As_2O_3 and CrO_4^{2-}, showing your working clearly.

- **<u>The reaction is first order wrt As_2O_3</u> since half-life of the reaction is constant (30s for experiment 1)**

Expt 1: Gradient = $\frac{0.58 - 0.32}{0 - 20}$ = –0.013

Expt 2: Gradient = $\frac{0.58 - 0.32}{0 - 10}$ = –0.026

- **Since the gradients represent the rates of the reaction, when concentration of CrO_4^{2-} doubled, the rate doubled. The reaction is first order wrt to CrO_4^{2-}.**

[3]

(ii) Write the rate equation for the reaction.

- **rate = k[As_2O_3][CrO_4^{2-}]** [1]

(iii) The experiment was repeated one more time using 0.58 mol dm^{-3} aqueous chromate(VI) and 0.58 mol dm^{-3} aqueous As_2O_3.

Suggest **two** differences to the graph of concentration of As_2O_3 (mol dm^{-3}) against time (s) obtained in this repeated experiment compared to that of experiment **1**.

- **Initial rate of reaction will be much slower. (OR gentler gradient at t= 0s) since a smaller concentration of CrO_4^{2-} is used.**
- **Half-lives of reaction are no longer constant as CrO_4^{2-} is not in excess. (OR reaction is no longer pseudo first order.)**

[2]
[Total: 8]

4 (a) MgO is formed from the thermal decomposition of $MgSiO_3$, a major component in mineral rock. In this process, SiO_2 is also produced.

(i) During thermal decomposition of mineral rocks, SO_3 and Al_2O_3 are also produced. Arrange MgO, SO_2 and Al_2O_3 in the order of increasing pH when dissolved in water.

- **SO_2 < Al_2O_3 < MgO** [1]

(ii) The melting points of the oxides are given below.

Compound	Melting point / °C
MgO	more than 2000
Al_2O_3	more than 2000
SiO_2	1710
SO_2	–72

Explain the melting point of these oxides in terms of their structure and bonding.

- **MgO and Al_2O_3 have giant ionic structures with strong ionic bonds between the cations and anion. A lot of energy is needed to break the strong ionic bonds. Thus they are both solid with very high melting point.**
- **SiO_2 has a giant molecular structure with an extensive network of strong covalent bonds between Si and O atoms. A lot of energy is needed to break the numerous strong covalent bonds. Thus, melting point of SiO_2 is high.**
- **SO_2 has a simple molecular structure with strong covalent bonds between atoms and weak permanent dipole-permanent dipole attractions (intermolecular forces) between SO_2 molecules. Much less energy is needed to overcome the weak intermolecular forces, hence SO_2 has very low boiling point and exists as a gas.**

[3]

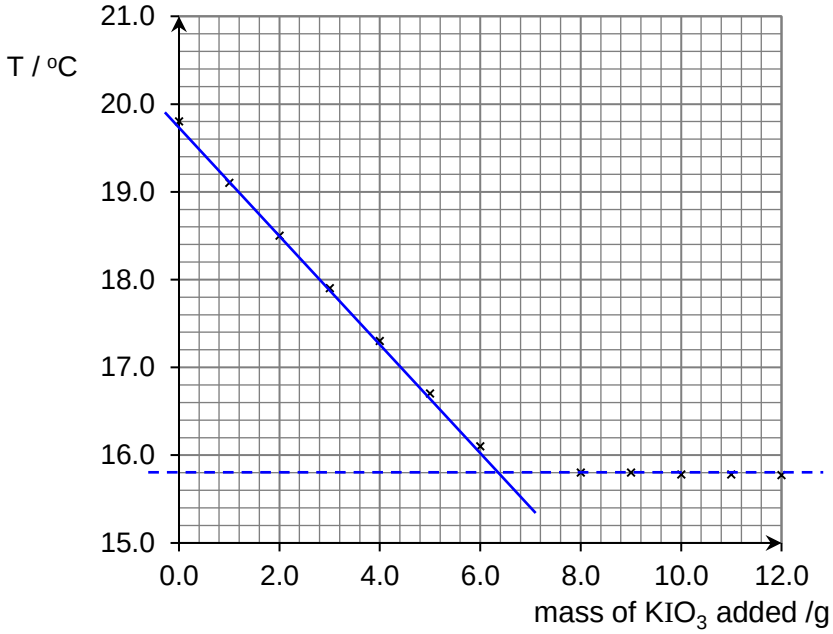
(iii) MgO is sparingly soluble in water but much more soluble in magnesium chloride solution. Explain the higher solubility of magnesium oxide in magnesium chloride solution with relevant equations.

- **Aqueous $MgCl_2$ is slightly acidic.**
- **When magnesium chloride dissolves in water, it forms $[Mg(H_2O)_6]^{2+}(aq)$ which undergoes partial hydrolysis with water to give a slightly acidic solution.(OR high charge density Mg^{2+} polarises the water molecule in $[Mg(H_2O)_6]^{2+}$, weaken the O-H bond, releases the H^+ ions)**
- **$[Mg(H_2O)_6]^{2+}(aq) + H_2O(l) \rightleftharpoons [Mg(H_2O)_5(OH)]^+(aq) + H_3O^+(aq)$**
- **H_3O^+ formed undergo acid-base reaction with MgO(s) to form a soluble salt, leading to the higher solubility of MgO.**
- **$MgO(s) + 2H_3O^+(aq) \rightarrow Mg^{2+}(aq) + 3H_2O(l)$** [3]

(b) Potassium iodate (KIO_3, M_r = 214.0) dissolves to a limited extent in water.

$$KIO_3(s) \rightleftharpoons K^+(aq) + IO_3^-(aq) \qquad \Delta H_1$$

To find ΔH_1, small portions of solid were successively added to 100 cm^3 of water in a well-insulated container. After each addition, the mixture was stirred extensively and the lowest temperature was recorded. The results were plotted below.



You may assume that
- the specific heat capacity of solution is 4.18 J cm^{-3} K^{-1},
- the volume of solution remains the same during the experiment, and
- the heat capacity of the insulated container is zero.

(i) Use the graph above to calculate the solubility of KIO_3 in mol dm^{-3}.

- **Amount of KIO_3 dissolved = $\frac{6.4}{214.0}$ = 0.0299 mol (ecf)**
- **Solubility = $[K^+] = [IO_3^-] = \frac{0.0299}{100} \times 1000$ = 0.299 mol dm^{-3} (ecf)**

[3]

(ii) Calculate the enthalpy change of dissolution of KIO_3, ΔH_1.

- **Heat absorbed = $100 \times 4.18 \times (19.8 - 15.8)$**
  = 1670 J
- **ΔH_1 = $+\dfrac{1670}{\text{Amount of } KIO_3 \text{ dissolved}}$ (ecf)**
  = $+\dfrac{1670}{0.0299}$
  = +55900 J mol^{-1} = +55.9 kJ mol^{-1}

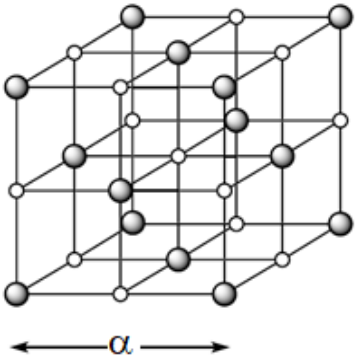
[2]

(iii) ΔS for the dissolution of some ionic salts could have a **negative** sign. Suggest why this is so.

- **In salt solutions, the water molecules hydrating the ions are more orderly arranged than those in pure water, and this causes a decrease in entropy, and so ΔS of dissolution has a negative sign.**

[1]

(c) The binary compounds of lithium, LiX (X = H, F, C*l*, Br, and I), all crystallise in the cubic lattice similar to NaC*l* lattice. One unit cell of this structure (small circles = Li^+, large circles = X^-) is shown below. The length of the edge of the unit cell is abbreviated α.



compound	α / pm
LiH	408.3
LiF	403.5
LiC*l*	513.6
LiBr	548.9
LiI	601.9

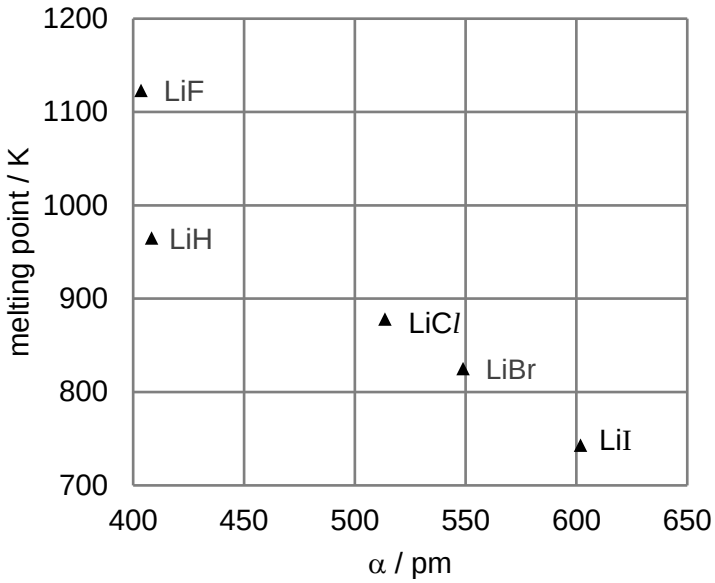
1 pm = 1×10^{-10} cm

(i) Given that there are four Li^+ ions and four X^- ions in one unit cell, calculate the mass of one unit cell and hence the density of LiC*l* in g cm^{-3}.
[Density, ρ = mass / volume]

- **Mass of one unit cell of LiC*l* = $4 \times (6.9 + 35.5) / (6.02 \times 10^{23}) = 2.82 \times 10^{-22}$ g**
- **Density = $\dfrac{2.82 \times 10^{-22}}{(513.6 \times 10^{-10})^3}$ = 2.08 g cm^{-3} (ecf)**

[2]

The melting points of the lithium halides and of lithium hydride are plotted against the unit cell edge length, α, as shown below.



(ii) Explain the relationship observed between α values and melting points of the lithium halides.

Melting points of lithium halides depends on the strength of ionic bonds.

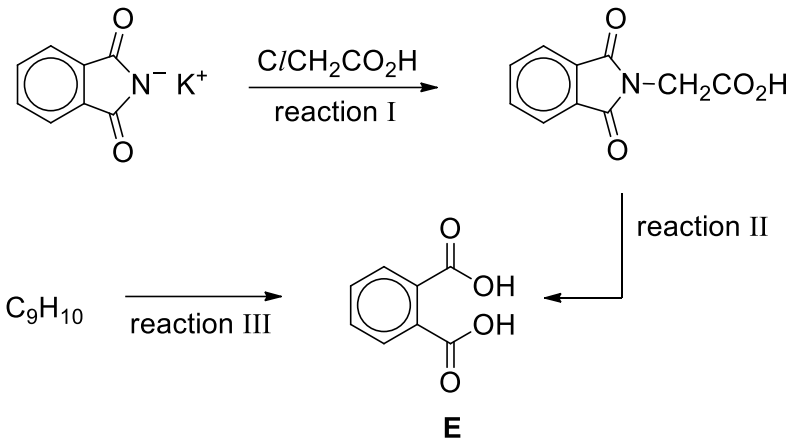
- **Lattice energy, $\Delta H_{LE} \propto \frac{q^+q^-}{r^+ + r^-}$, indicates the strength of ionic bonds.**
- **Since $r^+ + r^- = \alpha$, the smaller the α, the stronger the ionic bond, leading to higher melting points.** [2]

(iii) Lithium hydride and lithium fluoride have similar α values.
Suggest an explanation on why the melting point of lithium hydride is much lower than that of lithium fluoride.

- **H is much less electronegative than F, so LiH is much less ionic than LiF.**
- **Since $\Delta H_{LE} \propto \frac{q^+q^-}{r^+ + r^-}$ and the charges on Li^+ and H^- in LiH are smaller than those on Li^+ and F^- in LiF, LiH has a lower melting point.** [2]

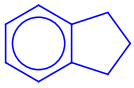
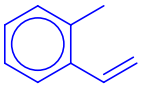
[Total: 19]

5 (a) The following reaction scheme involves compound **E**, benzene-1,2-dicarboxylic acid.



E

(i) Suggest **two** structures of C_9H_{10} that could be used in reaction III to form compound **E**.

-  OR  [1]

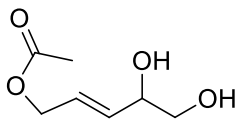
(ii) Suggest the reagents and conditions for reaction III.

- **$KMnO_4$(aq), H_2SO_4(aq), heat** [1]

(iii) State the types of reaction for reactions I and II.

reaction I: nucleophilic substitution
reaction II: (acid) hydrolysis [1]

(b) The structure of compound **G** is shown below.

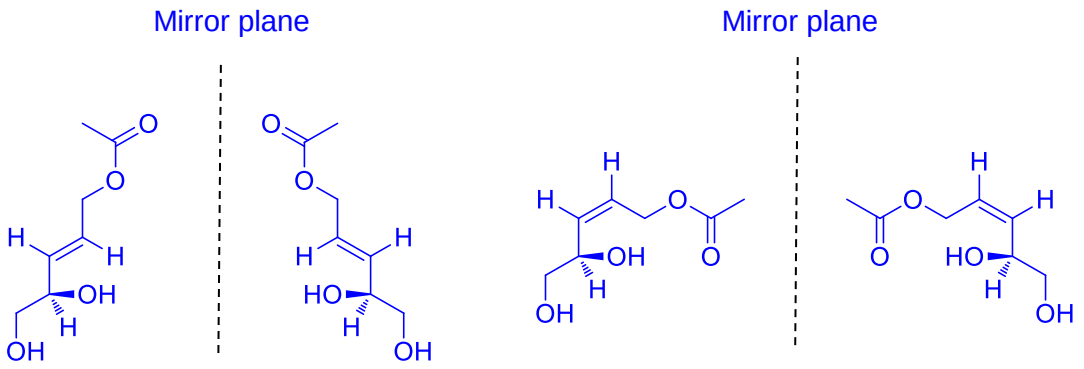


compound **G**

(i) State the types of stereoisomerism exhibited by compound **G** and the total number of stereoisomers of **G**.

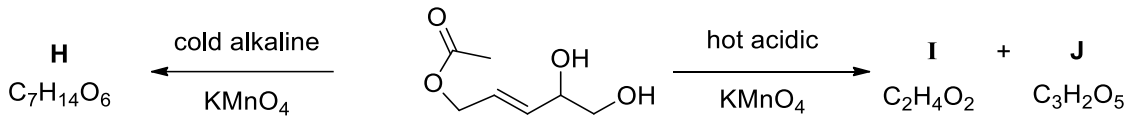
- ***cis-trans* isomerism and enantiomerism** [1]
- **4 stereoisomers**

(ii) Draw all the possible stereoisomers of compound **G**.

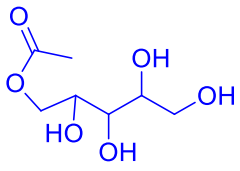


[2]

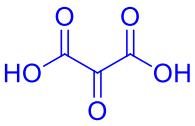
(iii) Compound **G** reacts with $KMnO_4$ under different conditions to give different products.



Suggest structures for compounds **H,** **I** and **J**.

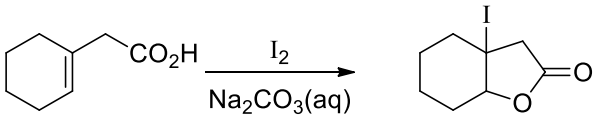
H: 

I: CH_3CO_2H

J: 

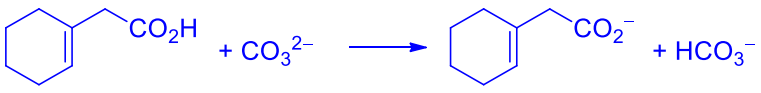
[3]

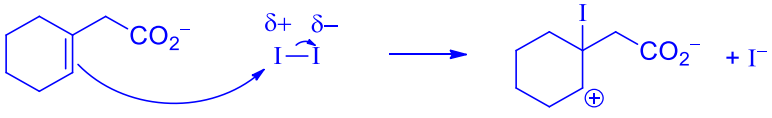
(c) In the presence of a base such as aqueous sodium carbonate, compound **X** can react with iodine to form compound **P** as shown below.

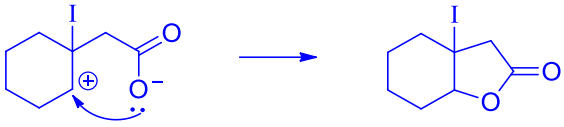


compound **X** compound **P**

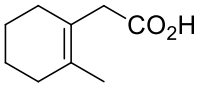
(i) Propose a three-step mechanism for this reaction. The first step of the mechanism involves the neutralisation of **X** by carbonate. Show clearly any intermediates that may be formed, all charges and use curly arrows to indicate the movement of electron pairs, and state the role of iodine in the reaction.

- **I_2 acts an electrophile.**
- **Step 1:** 

- **Step 2:** 

- **Step 3:**  **[3]**

(ii) Describe a chemical method with not more than two steps for distinguishing between compounds **X** and **Y**. State the observations clearly.



compound **Y**

- **Add compounds X and Y to two separate test-tubes and then add $KMnO_4$(aq), H_2SO_4(aq) and then heat both test-tubes.**
- **Add (excess) NaOH(aq) and then I_2(aq) to each test-tube and then warm. Both compounds will decolorise purple $KMnO_4$ in step 1.**
- **Compound Y gives a pale-yellow ppt while compound X will not give a pale-yellow ppt in step 2.**

[3]

(d) When the compounds shown below were reacted with aqueous sodium hydroxide at 55 °C, the rate of reaction decreases in the following order.

$$C_6H_5CH_2Br > (CH_3)_3CBr > C_6H_5Br$$

(i) Both $C_6H_5CH_2Br$ and $(CH_3)_3CBr$ undergo unimolecular nucleophilic substitution with aqueous sodium hydroxide. Explain their relative rates of reaction.

Both $C_6H_5CH_2Br$ and $(CH_3)_3CBr$ undergo S_N1 with aqueous NaOH.

- **$C_6H_5CH_2Br$ gives a more stable carbocation, $C_6H_5CH_2^+$, than $(CH_3)_3C^+$ because the positive charge can be dispersed into the benzene ring. Thus, the formation of $C_6H_5CH_2^+$ is favoured leading to a higher rate.** [1]

(ii) Explain why C_6H_5Br has almost no reaction with aqueous sodium hydroxide at 55 °C.

- **In C_6H_5Br, the p orbital of the Br overlaps with the p orbitals of C atoms in the benzene ring, leading to a partial double bond character. Thus, the C-Br bond is stronger and more difficult to be broken.** [1]

[Total: 17]