Class | Adm No

Candidate Name: _____

H

millennia institute

2021 Preliminary Examinations
Pre-University 3

H2 CHEMISTRY **9729/03**

Paper 3 Free Response **17 Sep 2021**

2 hours

Additional materials: Data Booklet

READ THESE INSTRUCTIONS FIRST

Do not turn over this question paper until you are told to do so

Write your name, class and admission number on all the work you hand in.

Write in dark blue or black pen.

You may use an HB pencil for any diagrams or graphs.

Do not use staples, paper clips, glue or correction fluid.

Section A: Answer **all** questions

Section B: Answer any **1** question

A Data Booklet is provided.

The use of an approved scientific calculator is expected, where appropriate.

You are reminded of the need for good English and clear presentation in your answers.

At the end of the examination, fasten all your work securely together.

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

	Section A			Section B		Total
Question	1	2	3	4	5	
Marks	/24	/20	/16	/20	/20	/80

Section A

Answer **all** questions from this section.

1 (a) A metallic element, **M**, exists as three isotopes. The relative isotopic masses and their relative abundances are shown in Table 1.1.

Relative Isotopic Mass	Relative Abundance
53.94	3
55.93	46
56.94	1

Table 1.1

(i) Define the term *relative isotopic mass*. [1]

Relative isotopic mass is defined as the mass of one atom of an isotope compared to $\frac{1}{12}$ of the mass of one atom of ^{12}C.

(ii) Using the information in Table 1.1, calculate the relative atomic mass of **M**, leaving your answer to 2 decimal places. [1]

Relative atomic mass = $\frac{(53.94\times3)+(55.93\times46)+(56.94\times1)}{(3+46+1)} = 55.83$

A dilute aqueous solution of the chloride salt of **M**, **M**Cl_n, was electrolysed using a current of 0.5 A for 1 hour as shown in the setup in Fig. 1.1. The experiment was carried out at 32 °C and 1 atm.

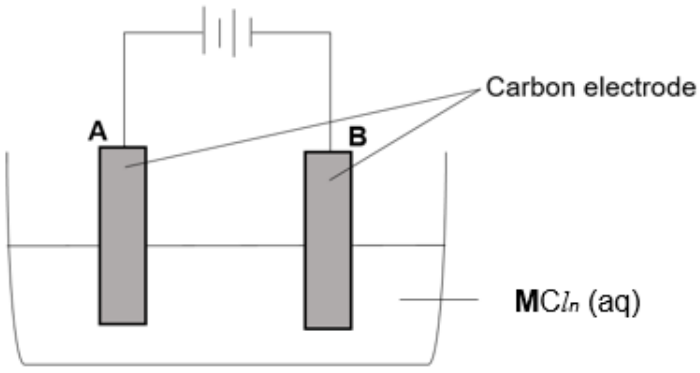


Fig. 1.1

0.3471 g of metal **M** was deposited at electrode **B**.

(iii) Using your answer in **(a)(ii)**, determine *n*. [3]

Amount of **M** deposited = $\frac{0.3471}{55.83} = 0.006217\ mol$;

Q = 0.5 × 1 × 60 × 60 = 1800 C

Amount of electrons = $\frac{1800}{96500} = 0.01865\ mol$;

$\mathbf{M}^{n+} + ne^- \rightarrow \mathbf{M}$

n = $\frac{0.01865}{0.006217} = 3$;

(iv) A gas is produced at electrode **A**.
Write the ion-electron equation for the reaction taking place at electrode **A**. [1]

$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

(v) Calculate the volume of gas produced at electrode **A**. [2]

Amount of O_2 = 0.01865 ÷ 4 = 0.004663 mol ;

V = $\frac{nRT}{p} = \frac{0.004663 \times 8.31 \times (32+273)}{101325} = 1.166 \times 10^{-4}\ m^3 = 117\ cm^3$;

(vi) Suggest why electrode **A** will decrease in mass over time. [1]

The oxygen gas produced from the electrolysis reacts with the carbon electrode to form carbon dioxide.

(b) Chromium(III) chloride, $CrCl_3$, reacts with water to form a violet complex, $[Cr(H_2O)_6]Cl_3$. This violet complex is isomeric to two other complexes as shown in Table 1.2.

Structural Formula of Complex	Colour of Complex
$[Cr(H_2O)_6]Cl_3$	violet
$[CrCl(H_2O)_5]Cl_2 \cdot H_2O$	pale green
$[CrCl_2(H_2O)_4]Cl \cdot 2H_2O$	dark green

Table 1.2

(i) Give one characteristic properties of transition elements as shown by chromium in the isomers in Table 1.2. [1]

Can form coloured compounds

(ii) Using $[Cr(H_2O)_6]Cl_3$, explain what is meant by the term *ligand*. [1]

H_2O is a molecule that has lone pairs of electrons on O atom, which can be used to form a dative bond with the central metal ion Cr^{3+}.

(iii) When 2.665 g of one of the complexes in Table 2 was reacted with an excess of aqueous $AgNO_3$, 2.868 g of $AgCl$ was obtained.

Deduce the structural formula of the complex. [2]

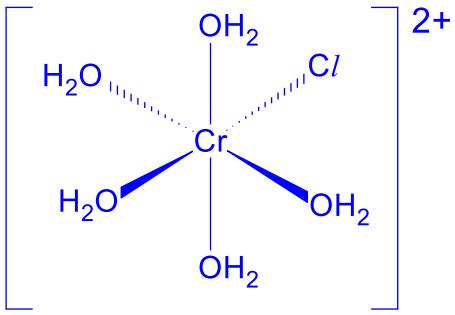
Amount of complex = $\frac{2.665}{52.0+18.0\times6+35.5\times3} = 0.0100\ mol$

Amount of $AgCl$ = $\frac{2.868}{(107.9+35.5)} = 0.0200\ mol$ **both amounts ;**

1 mole of complex contains 2 moles of free chloride ions

Structural formula = $[CrCl(H_2O)_5]Cl_2 \bullet H_2O$;

(iv) Draw the full structure of the cation in the pale green complex, $[CrCl(H_2O)_5]Cl_2 \bullet H_2O$, showing the shape of the cation. [1]



(c) Potassium dichromate(VI) can be produced from chromium(III) chloride via a 3-step reaction.

Step 1:	Chromium(III) chloride is boiled with hydrogen peroxide in an alkaline medium, forming a bright yellow solution, CrO_4^{2-}.
Step 2:	Boiling is continued until excess hydrogen peroxide is destroyed.
Step 3:	The yellow solution is then cooled and acidified with ethanoic acid, forming the orange dichromate(VI) solution.

(i) Write the ion-electron equation for the oxidation of Cr^{3+} in step 1 and hence the overall balanced equation for the reaction between Cr^{3+} and H_2O_2 under alkaline conditions. [2]

$Cr^{3+} + 8OH^- \rightarrow CrO_4^{2-} + 4H_2O + 3e^-$;

$2Cr^{3+} + 10OH^- + 3H_2O_2 \rightarrow 2CrO_4^{2-} + 8H_2O$;

(ii) Suggest how would you know when all the excess hydrogen peroxide has been destroyed in step 2. [1]

No more effervescence

(iii) The reaction in step 3 has the following equation:

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$$

State and explain if the reaction in step 3 is a redox reaction. [1]

No. The oxidation state of chromium remains the same at +6.

(iv) The reaction in step 3 is an exothermic reaction.
State and explain the effect of increasing the temperature of the system on the yield of $Cr_2O_7^{2-}(aq)$. [2]

When temperature is increased, position of equilibrium will shift left to remove the excess heat by favouring the endothermic reaction. ;
Yield of $Cr_2O_7^{2-}$ will decrease. ;

(d) A solution of sodium dichromate(VI) is acidified with dilute sulfuric acid before adding to an organic compound. The solution slowly turns green. Boiling the green solution with more of the organic compound produces a pale blue solution.
0.10 mol of the pale blue ions was found to require 0.020 mol of acidified potassium manganate(VII) solution to oxidise it back to the green ions.

(i) State the identity of the green ion. [1]

Cr^{3+} or $[Cr(H_2O)_6]^{3+}$

(ii) Deduce the oxidation number of the blue ion. [2]

$MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4H_2O$
$Cr^{n+} \rightarrow Cr^{3+} + (3-n)e$
Amount of electrons = 5 x 0.020 = 0.100 mol ;
Ratio of blue ion : e = 0.10 : 0.100 = 1:1
n = 3 – 1 = +2 ;

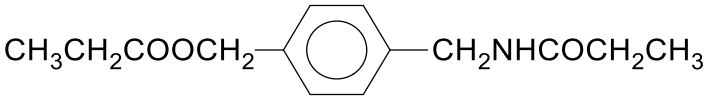
(iii) State, with a reason, which reagent, acidified H_2O_2 or aqueous Fe^{3+} could be used to convert the green ions back to dichromate(VI) ions. [1]

Acidified H_2O_2
Only $E°(H_2O_2/H_2O)$ is more positive than $E°(Cr_2O_7^{2-}/Cr^{3+})$ OR calculate $E°_{cell}$

[Total: 24]

2 (a) Compound **P**, C_8H_9NO, can react with Tollen's reagents but not with Fehling's solution. **P** is also soluble in dilute hydrochloric acid.

On reacting **P** with lithium aluminium hydride, **Q**, $C_8H_{11}NO$, is formed. **Q** reacts readily with sodium metal. 1 mole of **Q** reacts with 2 moles of propanoyl chloride to give the following compound.

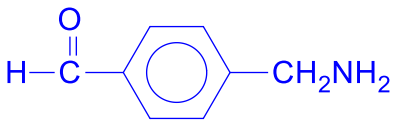


Q also reacts with hot acidified $KMnO_4$ to give **R**, $C_8H_6O_4$. 1 mole of **R** reacts with excess $Na_2CO_3(s)$ to give 1 mole of $CO_2(g)$.

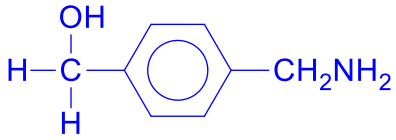
When **P** is heated with excess CH_3Br, a crystalline solid **S**, $C_{11}H_{16}NOBr$, is formed. **S** has very high melting point.

Deduce the structure of compounds **P**, **Q**, **R** and **S**, explaining the chemistry of the reactions described. [10]

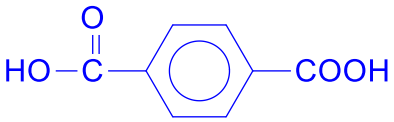
Information	Deductions
P, C_8H_9NO, can react with Tollen's reagents but not with Fehling's solution.	**P** is a benzaldehyde ; Oxidation reaction with Tollen's reagents ;
P is also soluble in dilute hydrochloric acid.	**P** has an alkaline amine group ; Acid-base reaction ;
On reacting **P** with lithium aluminium hydride, **Q**, $C_8H_{11}NO$, is formed.	Reduction reaction ; **Q** contains primary alcohol ;
Q reacts readily with sodium metal.	Acid-metal (or redox) reaction; **Q** contains –OH group ;
1 mole of **Q** reacts with 2 moles of propanoyl chloride to give the following compound.	Condensation reaction;
Q also reacts with hot acidified $KMnO_4$ to give **R**, $C_8H_6O_4$.	Side chain oxidation ;
1 mole of **R** reacts with excess $Na_2CO_3(s)$ to give 1 mole of $CO_2(g)$.	**R** contains 2 carboxylic acid group ; Acid-carbonate reaction ;
When **P** is heated with excess CH_3Br, a crystalline solid **S**, $C_{11}H_{16}NOBr$, is formed. **S** has very high melting point.	Nucleophilic substitution ; **S** has a giant ionic lattice structure ;



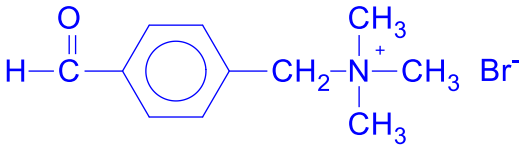
P



Q



R

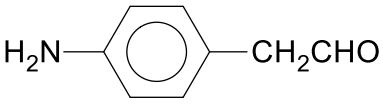


S

Each structure 1m ;;;;

Every 2 deductions 1m, capped at 6m ;;;;;;

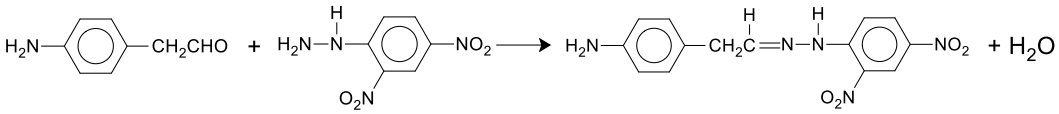
(b) **T** is an isomer of **P**.



T

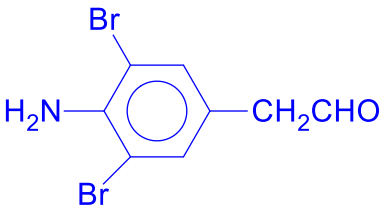
(i) State the type of reaction when **T** reacts with 2,4-dinitrophenylhydrazine. Write a balanced equation for the reaction. [2]

Condensation ;



;

(ii) Draw the structure of the organic product formed when **T** reacts with $Br_2(aq)$. [1]



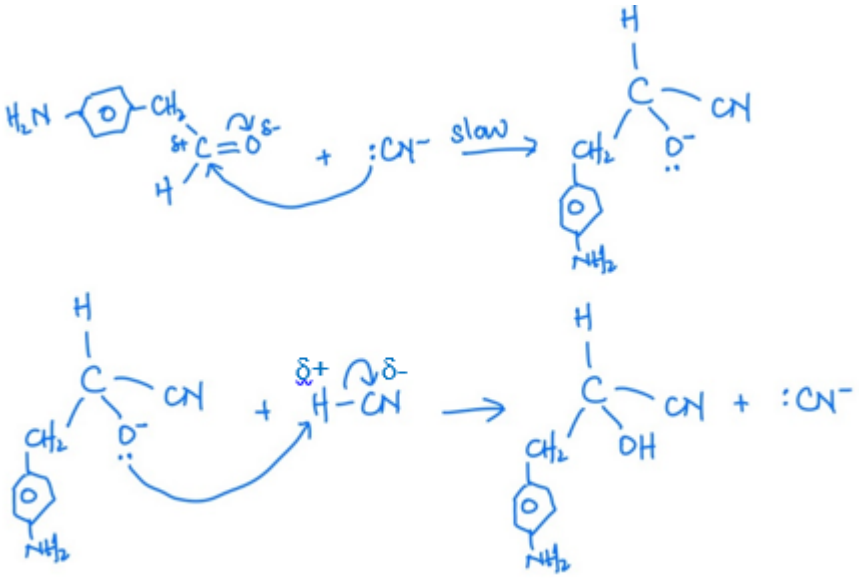
;

(iii) Outline the mechanism for the reaction of **T** with cold alkaline HCN. [3]

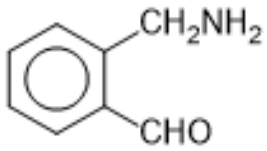
Nucleophilic Addition ;

$HCN \rightleftharpoons H^+ + CN^-$

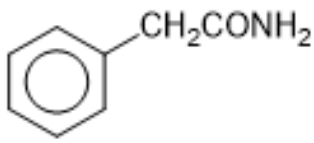
$NaOH + HCN \rightarrow NaCN + H_2O$



(c) Compounds **U** and **W** are constitutional isomers of **T**.



U



W

(i) Explain why **U** has a lower boiling point compared to **W**. [2]

U can form intramolecular hydrogen bond due to the proximity of the $–NH_2$ and the –CHO groups hence its intermolecular forces of attraction is predominantly instantaneous dipole-induced dipole interactions or it forms less extensive intermolecular hydrogen bonds ;

Less energy is required to overcome the weaker id-id between **U** molecules than the stronger hydrogen bond between **W** molecules ;

(ii) State and explain the relative basicity of **U** and **W**. [2]

U is a stronger base than **W** ;

The lone pair on N of **W** is delocalised into the π electron cloud of the C=O bond causing it unavailable for protonation ;

[Total: 20]

3 This question involves sulfur and its compounds.

(a) **(i)** State the full electronic configuration of S^{3+}. [1]

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$;

(ii) Write an equation to represent the fourth ionisation energy of sulfur. [1]

$S^{3+}(g) \rightarrow S^{4+}(g) + e^-$;

(iii) The fourth ionisation energy of sulfur and phosphorus is 4540 kJ mol^{-1} and 4960 kJ mol^{-1} respectively.

Explain why the fourth ionisation energy of sulfur is lower than that of phosphorus. [2]

Less energy is required to remove an electron from the **3p orbital in S^{3+} than to remove an electron from the 3s orbital in P^{3+}** ; as the **3p electron is further away from the nucleus** and experiences **greater shielding** than 3s electron. ;

(iv) A sample of S contains 2 isotopes. In order to identify the mass of these two isotopes, they were first ionised to form singly-charged positive ions and passed through an electric field of constant strength. Their angles of deflection, along with a reference sample of helium nucleus, $^4_2\mathrm{He}$, were recorded in Table 3.1.

Species	Angle of deflection
Isotope **X**	1.60 °
Isotope **Y**	1.51 °
Helium nucleus, $^4_2\mathrm{He}$	25.6 °

Table 3.1

Calculate the relative mass of each isotope of sulfur, giving your answer to the nearest whole number. [3]

$25.6 = k\frac{2}{4}$

k = 51.2 ;

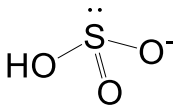
$1.60 = 51.2\left(\frac{1}{mass\ of\ X}\right)$

Mass of X = 32 ;

$1.51 = 51.2\left(\frac{1}{mass\ of\ Y}\right)$

Mass of Y = 34 ;

(b) Hydrogen sulfite ion, HSO_3^-, has the following structure:



Hydrogen sulfite ion reacts with carbonate ion according to the following equation:

$$2HSO_3^- + CO_3^{2-} \rightarrow 2SO_3^{2-} + H_2O + CO_2$$

(i) Define the term *Brønsted-Lowry acid*. [1]

Species which donate a proton ;

(ii) State which species in the reaction between hydrogen sulfite ion and carbonate ion acts as a Brønsted-Lowry acid. [1]

HSO_3^-

(iii) The carbon dioxide gas produced from the reaction is a non-ideal gas. Fig. 3.1 shows the experimental values of the compressibility factor (Z) plot for 1 mole of carbon dioxide gas at 298 K. (Z = *p*V/RT)

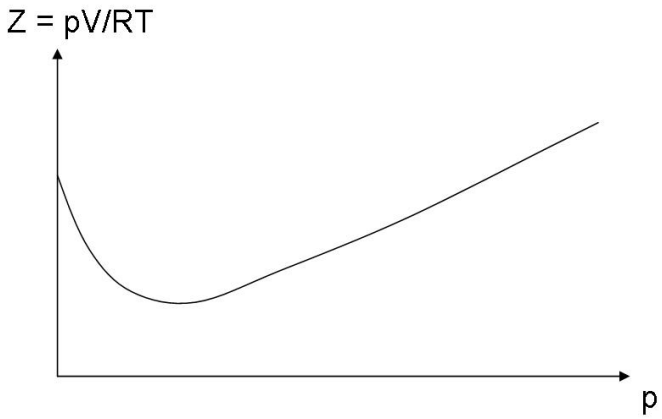
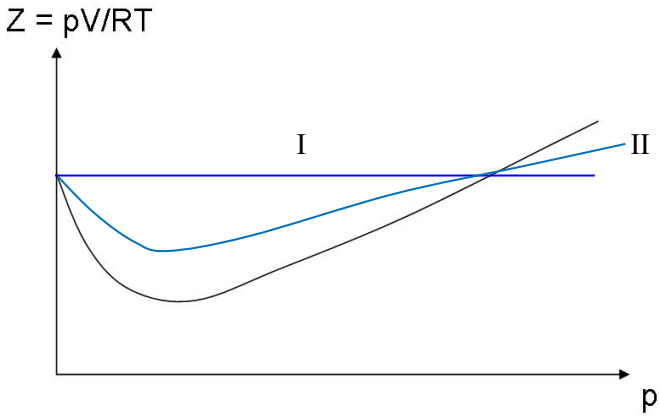


Fig. 3.1

On the plot given in Fig. 3.1, sketch the graph to illustrate the behavior of

I. 1 mole of an ideal gas at 298 K,
II. 1 mole of carbon dioxide gas at 398 K.

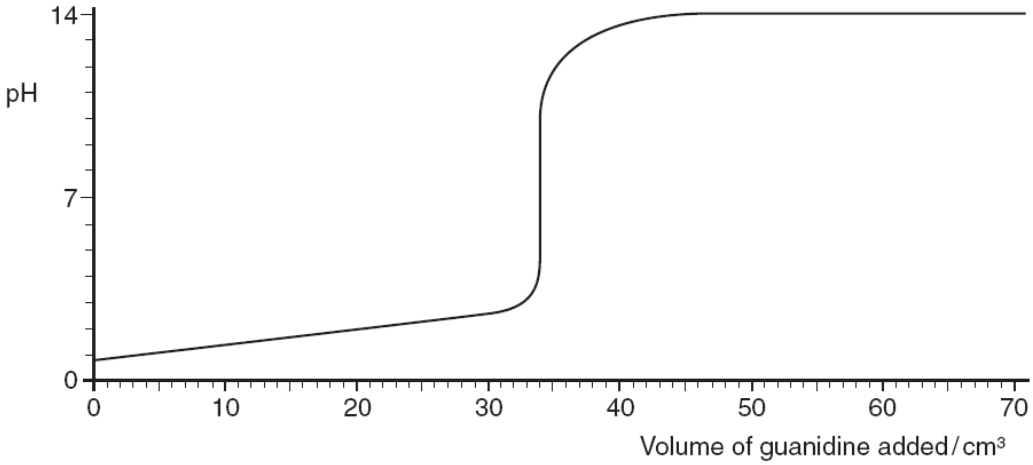
Label your answer clearly. [2]



(c) Sulfuric acid reacts with guanidine, $(H_2N)_2CNH_2$, as follows:

$$H_2SO_4 + 2(H_2N)_2CNH \rightarrow SO_4^{2-} + 2(H_2N)_2CNH_2^+$$

When a 25.0 cm^3 sample of dilute sulfuric acid was titrated against a solution of guanidine, the following titration curve was obtained.



(i) Deduce if guanidine is a strong or a weak base. [1]

Guanidine is a strong base.

For the titration of H_2SO_4 (a strong acid) against guanidine, the pH of equivalence point is about 7. (also accept pH is approaching 14 when guanidine is in excess)

Will not accept no inflexion point, no buffer region.

(ii) The initial pH of the sulfuric acid is 0.61.

Calculate $[H^+]$, and hence the initial concentration of sulfuric acid. [2]

$[H^+] = 10^{-0.61}$ = **0.245 mol dm^{-3}**

$[H_2SO_4]$ = **0.123 mol dm^{-3}**

(iii) Calculate the concentration, in mol dm^{-3}, of guanidine in the solution. [2]

Amount of guanidine = 2 × amount of H_2SO_4 = $2 \times \frac{25}{1000} \times 0.123 = 0.00614\ mol$;

[guanidine] = $\frac{0.00614}{\left(\frac{34}{1000}\right)} = 0.180\ mol\ dm^{-3}$;

Accept between 33.5 to 34.5 cm^3 as end point (same numerical answer)

[Total: 16]

Section B

Answer **one** question from this section.

4 In the stratosphere, there is an equilibrium established between O_2 and O_3:

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

When UV radiation from the sun breaks up O_3 into an oxygen atom and an oxygen molecule, the oxygen atom quickly recombines with an oxygen molecule to form back ozone. However, this equilibrium is disrupted by the radicals produced from the chlorofluorocarbons (CFCs) due to human activities.

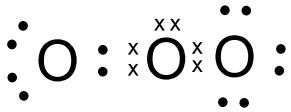
(a) One such CFC is CF_2Cl_2.

Write two equations to illustrate how CF_2Cl_2, in the presence of UV light, causes the decomposition of ozone layer. [2]

$CF_2Cl_2 \xrightarrow{UV} CF_2Cl\bullet + Cl\bullet$

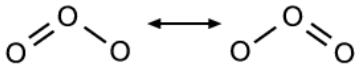
$Cl\bullet + O_3 \rightarrow ClO\bullet + O_2$

(b) The dot-and-cross diagram of O_3 is shown below.

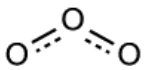


With the aid of suitable diagram, explain why the two oxygen-oxygen bonds in the ozone molecule have the same bond strength. [2]

O_3 forms resonance structures as shown below. ;



OR



;

(dative bonds acceptable)

Thus, the two oxygen-oxygen bonds are equivalent in length and strength

(c) One method to determine the concentration of the ozone layer is to pass air through acidified potassium iodide and to measure the amount of iodine liberated. The following reaction takes place:

$$O_3 + 2H^+ + 2I^- \rightarrow O_2 + H_2O + I_2$$

The concentration of iodine liberated, in mol dm^{-3}, is measured using platinum / iodine electrode against standard silver / silver chloride reference electrode. The E_{cell} of the system, in volts, is given by the following equation:

$$E_{cell} = 0.32 + 0.029 \log_{10}[I_2]$$

To determine concentration of ozone in the atmosphere, a balloon was used to carry a sampling device. A 1.00 m^3 sample of air, measured at 25 °C and 0.24 atm was passed through 10 cm^3 of acidified potassium iodide. The E_{cell} of the platinum electrode dipping into the solution against the standard silver / silver chloride electrode was 0.21 V.

(i) Calculate the amount of ozone present in the sample. [3]

$0.21 = 0.32 + 0.029 \log_{10}[I_2]$

$[I_2] = 0.0001610$ mol dm^{-3} ;

amount of $I_2 = 0.0001610 \times 0.0100 = 1.61 \times 10^{-6}$ mol ;

Amount of $O_3 = 1.61 \times 10^{-6}$ mol ;

(ii) Hence calculate the volume of O_3 present in the 1.00 m^3 sample of air. [2]

$pV = nRT$

$V = \frac{nRT}{p} = \frac{1.61\times10^{-6}\times8.31\times(25+273)}{0.24\times101325} = 1.64 \times 10^{-7} m^3 = 0.164\ cm^3$

Correct conversion of units;

Correct substitution of values using ideal gas equation ;

(iii) The actual volume of O_3 present in the 1.00 m^3 sample of air would be less than the calculated value in **(c)(ii)**. Suggest a reason for this discrepancy. [1]

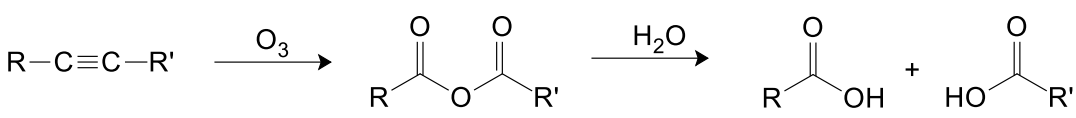
O_3 has significant intermolecular forces of attraction, which causes the molecule to be closer together, hence the volume decreases. ;

(iv) Explain why O_2 is a gas but I_2 is a solid at room temperature. [2]

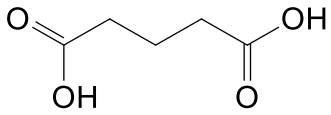
Both are simple molecular structures ;

However I_2 has stronger instantaneous dipole-induced dipole interactions due to its larger electron cloud size ; hence requiring more energy to overcome its IMF.

(d) Ozone reacts with alkynes to form carboxylic acids in a two-step reaction as shown below.



When alkyne **E** undergoes the same two-step reaction, only compound **F** is formed.



F

Suggest the structure of **E**. [1]

(e) Nitrogen monoxide in the air can be converted to nitric acid, which results in acid rain. Both nitrogen monoxide and nitrogen dioxide also cause ozone layer depletion. One way of forming nitrogen monoxide is through the dissociation of nitrogen dioxide.

$$2NO_2\ (g) \rightleftharpoons 2NO\ (g) + O_2\ (g) \qquad \Delta H = +114.2\ kJ\ mol^{-1}$$

At 494 °C, the value of K_p for the above reaction is 36.9 kPa.

When a certain partial pressure of nitrogen dioxide is put into an empty vessel at 494 °C, equilibrium is reached when 40% of the original nitrogen dioxide has dissociated.

(i) Write an expression for the equilibrium constant, K_p, for the reaction. [1]

$K_p = \frac{P_{NO}{}^2 P_{O_2}}{P_{NO_2}{}^2}$;

(ii) Calculate the initial partial pressure of nitrogen dioxide before any dissociation occurred. [2]

Let the initial pressure of NO_2 be ***x*** mol.

	$2\ NO_2\ (g)$	$\rightleftharpoons$	$2\ NO\ (g)$	$+\ O_2\ (g)$
Initial pressure (kPa)	***x***		0	0
Change in Pressure	-0.40***x***		+0.40***x***	+0.20***x***
Equilibrium pressure (kPa)	0.60***x***		0.40***x***	0.20***x***

$K_p = \frac{(0.40x)^2(0.20x)}{(0.60x)^2} = 36.9$;

x = 415 kPa

Hence, initial pressure of NO_2 = 415 kPa ;

(iii) Given that the standard enthalpy change of formation of NO(g) is +90.3 kJ mol^{-1}, determine the standard enthalpy change of formation of NO_2(g). [2]

$\Delta H = \Delta H_f$ (products) – ΔH_f (reactants)

$114.2 = 2(90.3) + 0 – (2 \times \Delta H_f\ NO_2)$;

$\Delta H_f\ NO_2 = +33.2\ kJ\ mol^{-1}$;

(iv) With reference to the equation given in **(e)**, state and explain the sign of entropy change, ΔS, for the dissociation of $NO_2(g)$. [1]

ΔS is positive since there is an increase in disorderliness due to increase in amount of gas particles. ;

(v) Using your answer in **(e)(iv)**, state whether the reaction would be spontaneous at high or low temperatures. [1]

[Total: 20]

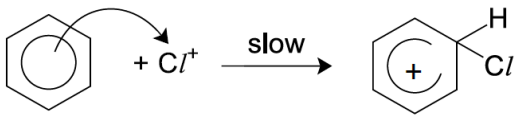
High temperature ;

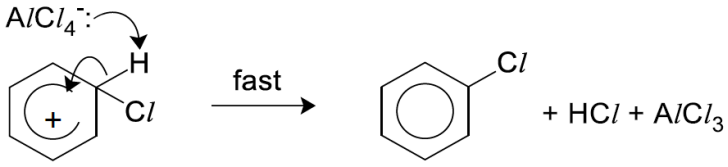
5 (a) Aluminium chloride is used in the production of chlorobenzene from the reaction between benzene and chlorine.

(i) Outline the mechanism for the reaction above. [3]

Electrophilic substitution ;

$AlCl_3 + Cl_2 \rightarrow AlCl_4^- + Cl^+$





(ii) Explain, with the aid of relevant equations, why the aluminium chloride used in the above reaction must be anhydrous. [2]

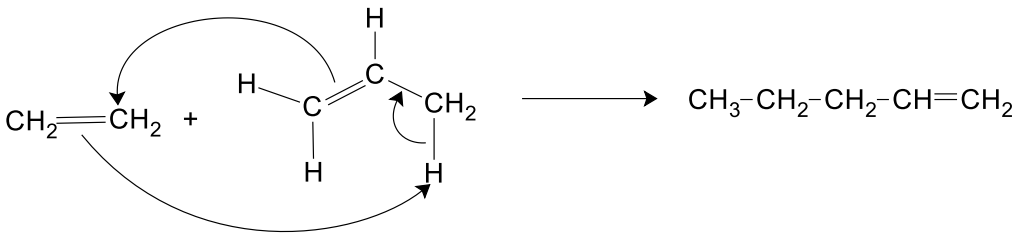
In the presence of water, Al^{3+}, having a high charge density is able to hydrolyse in water, causing $AlCl_3$ unable to act as a halogen carrier. ;

$AlCl_3 + 6H_2O \rightarrow [Al(H_2O)_6]^{3+} + 3Cl^-$

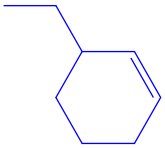
$[Al(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H_3O^+$;

(iii) Aluminium chloride can also act as a catalyst for the Alder-ene reaction.

One example of Alder-ene reaction is illustrated below using propene and ethene.



Suggest the structure of the organic product formed when cyclohexene reacts with ethene in an Alder-ene reaction. [1]



(b) Aluminium chloride reacts with CH_3NH_2 to form an adduct $AlCl_3 \cdot CH_3NH_2$.
Explain how this adduct is formed. [2]

Al in $AlCl_3$ is electron-deficient and is able to accept the lone pair on N of CH_3NH_2, ;
forming a dative bond. ;

(c) Describe the reaction, if any, of aluminium oxide with water, and sodium oxide with water. Write equations, with state symbols, where appropriate and predict the pH of the resultant solution. [3]

Aluminium oxide cannot dissolve in water;

Sodium oxide dissolves in water to form an alkaline solution of pH 13 ;

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$;

(d) 100 cm^3 of 0.100 mol dm^{-3} of aqueous magnesium nitrate was mixed with 100 cm^3 of aqueous sodium hydroxide and the reaction mixture was filtered. Solid magnesium hydroxide was collected as residue, and the filtrate contained a saturated solution of magnesium hydroxide, $Mg(OH)_2$, at 298 K and pH 9.5.

$$Mg(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaNO_3(aq)$$

The value of K_{sp} of magnesium hydroxide is 1.5×10^{-11}.

(i) Write an expression for the K_{sp} of magnesium hydroxide, stating its units. [2]

$K_{sp} = [Mg^{2+}] [OH^-]^2$;

$mol^3\ dm^{-9}$;

(ii) Calculate the concentration of hydroxide ions that remained in the **filtrate**. [2]

$pOH = pK_w - pH = 14 - 9.5 = 4.5$

$[OH^-] = 10^{-pOH} = 10^{-4.5} = 0.0000316\ mol\ dm^{-3}$

(iii) Calculate the concentration of magnesium ions that remained in the **filtrate**. [1]

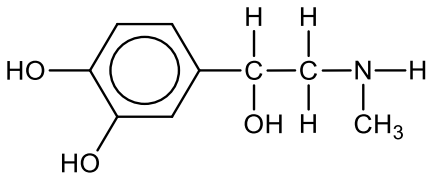
$[Mg^{2+}] = \frac{K_{sp}}{[OH^-]^2} = \frac{1.5\times10^{-11}}{(0.0000316)^2} = 0.0150\ mol\ dm^{-3}$;

(iv) Hence, calculate the mass of solid magnesium hydroxide formed. [2]

Amount of Mg^{2+} in solid = $(0.100 \times \frac{100}{1000}) - (0.0150 \times \frac{200}{1000}) = 7.00 \times 10^{-3}$ mol ;

Mass of $Mg(OH)_2$ formed = $7.00 \times 10^{-3} \times [24.3+2(1.0 + 16.0)] = 0.408$ g ;

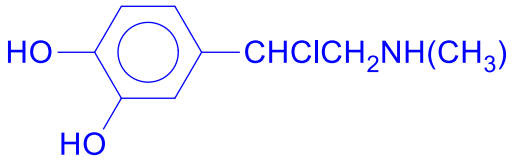
(e) The structure of compound **L** is shown below.



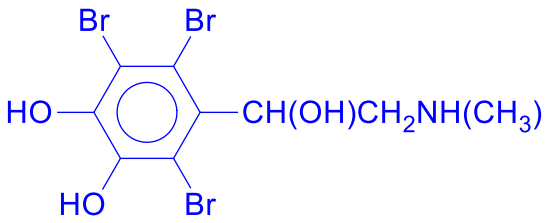
Compound **L**

Draw the structure of the organic product formed when compound **L** reacts with the following reagents.

(i) $SOCl_2$ [1]



(ii) Bromine water [1]



[Total: 20]