Cambridge Assessment International Education

Cambridge International AS & A Level

CANDIDATE NAME

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1284515204

CHEMISTRY **9701/23**

Paper 2 AS Level Structured Questions **May/June 2022**

1 hour 15 minutes

You must answer on the question paper.

No additional materials are needed.

INSTRUCTIONS

- Answer **all** questions.
- Use a black or dark blue pen. You may use an HB pencil for any diagrams or graphs.
- Write your name, centre number and candidate number in the boxes at the top of the page.
- Write your answer to each question in the space provided.
- Do **not** use an erasable pen or correction fluid.
- Do **not** write on any bar codes.
- You may use a calculator.
- You should show all your working and use appropriate units.

INFORMATION

- The total mark for this paper is 60.
- The number of marks for each question or part question is shown in brackets [ ].
- The Periodic Table is printed in the question paper.
- Important values, constants and standards are printed in the question paper.

This document has **16** pages. Any blank pages are indicated.

1 **(a)** Define first ionisation energy.

.....

.....

..... [2]

(b) Successive ionisation energies for element **A** are shown in Table 1.1.

Table 1.1

ionisation	1st	2nd	3rd	4th	5th	6th	7th	8th
ionisation energy / kJ mol^{-1}	1310	3390	5320	7450	11 000	13 300	71 000	84 100

Use Table 1.1 to deduce the group of the Periodic Table that **A** belongs to. Explain your answer.

Group .....

.....
[1]

(c) Across Period 3 there is a general trend for first ionisation energies to increase due to the increase in attraction between the nucleus and the outer electron.

Explain why the first ionisation energy of sulfur is less than the first ionisation energy of phosphorus.

.....

.....

..... [2]

(d) In an Al^{2+} ion the nuclear attraction for the outer electron is stronger than in an atom of Na.

Compare the electronic structures of Al^{2+} and an atom of Na and explain why the third ionisation energy of aluminium is greater than the first ionisation energy of sodium.

.....

.....

.....

..... [2]

(e) An isotope of copper has a relative isotopic mass of 65.

Complete Table 1.2 for an atom of copper-65.

Table 1.2

	atomic number	nucleon number	number of neutrons	electronic arrangement
copper-65				

[3]

(f) (i) The element copper has a relative atomic mass of 63.5.

Calculate how many atoms are present in 1.05 g of copper.

atoms of copper present = ... [1]

(ii) Copper has a melting point of 1085 °C and a high electrical conductivity.

Explain these properties of copper by referring to its structure and bonding.

.....

.....

..... [2]

[Total: 13]

[Turn over

2 **(a)** Period 3 elements and their compounds show trends in their physical properties.

(i) On Fig. 2.1 sketch a graph to show the melting points of the first five elements in Period 3.

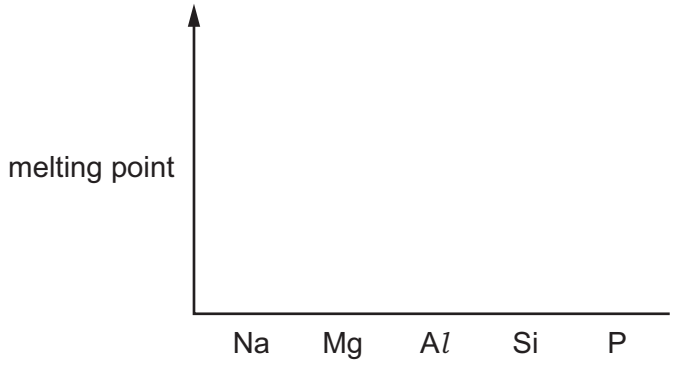

Fig. 2.1

[3]

(ii) Complete Table 2.1 with information for sodium chloride and phosphorus(V) chloride.

Table 2.1

	sodium chloride	phosphorus(V) chloride
state at room temperature		
name of change which occurs on addition of water		
pH of final solution		

[3]

(b) Tennessine, Ts, is an unstable man-made element. It is found below astatine, At, in Group 17. The chemical properties of Ts and its compounds have only been predicted.

(i) Suggest an equation for the reaction of NaTs and Br_2. Assume that Ts follows the same trends as the other elements in Group 17. Explain your answer.

equation .....

explanation .....

.....

[2]

(ii) Some scientists predict that Ts has properties typical of metals like copper.

Complete Table 2.2 with:

- the predicted melting point of tennessine
- the lattice structure of solid chlorine, bromine and tennessine.

Assume that Ts has properties typical of metals like copper.

Table 2.2

element	chlorine	bromine	tennessine
melting point / °C	–101	–7.2	
lattice structure of crystalline solid			

[2]

[Total: 10]

3 **G** belongs to a group of compounds called ethers.

G

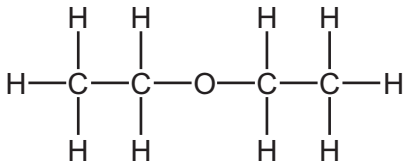

Fig. 3.1

(a) When **G**, $C_4H_{10}O$, is heated, thermal decomposition occurs.

$$C_4H_{10}O(g) \rightarrow C_2H_6(g) + CO(g) + CH_4(g) \qquad \Delta H = -7.00\,kJ\,mol^{-1}$$

(i) The atoms in a molecule of CO are held together by a triple covalent bond. One of these bonds is a coordinate (dative covalent) bond.

Draw a dot-and-cross diagram to show the arrangement of outer electrons in a CO molecule.

Use ● to represent electrons from an oxygen atom.
Use × to represent electrons from a carbon atom.

[1]

(ii) Calculate the bond energy of C≡O using the bond energy values in Table 3.1 and the enthalpy change, ΔH, for the thermal decomposition of **G**. Show your working.

Table 3.1

bond	bond energy / kJ mol^{-1}
C–C	350
C–O (in **G**)	360
C–H	410

bond energy (C≡O) = ... kJ mol^{-1}
[2]

(iii) When **G**, $C_4H_{10}O$, is heated in a sealed container, an equilibrium mixture is produced.

$$C_4H_{10}O(g) \rightleftharpoons C_2H_6(g) + CO(g) + CH_4(g)$$

Complete the expression for the equilibrium constant, K_c, for this reaction.
State the units of K_c.

K_c =

units = ...
[2]

(iv) Thermal decomposition of **G** in the presence of I_2 affects the activation energy, E_a, for the reaction. Table 3.2 shows E_a for the thermal decomposition of **G** with and without I_2.

Table 3.2

reaction	E_a (with I_2) / kJ mol^{-1}	E_a / kJ mol^{-1}
$C_4H_{10}O(g) \rightarrow C_2H_6(g) + CO(g) + CH_4(g)$	143	224

State what effect adding I_2 to the reaction mixture has on the value of K_c.
Explain your answer.

.....

.....

.....

..... [2]

(v) Fig. 3.2 shows the Boltzmann distribution of energies for molecules of **G** at constant temperature, T°C.

Sketch, on Fig. 3.2, the Boltzmann distribution of energies for molecules of **G** at a higher temperature, (T+100)°C.

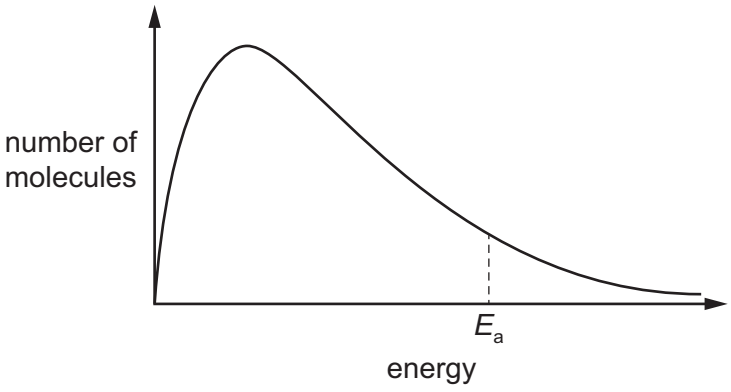

Fig. 3.2

[2]

[Turn over

(b) The functional group in **G** is an oxygen atom bonded to two carbon atoms.

G

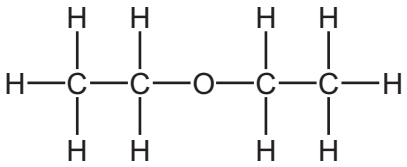

Fig. 3.3

G, **H** and **J** are structural isomers with molecular formula $C_4H_{10}O$.
H and **J** are straight chain molecules.

Table 3.3 shows the boiling points and reactions of **G**, **H** and **J** when heated under reflux with excess acidified $K_2Cr_2O_7$.

Table 3.3

	G	**H**	**J**
boiling point/°C	35	117	98
heat under reflux with excess acidified $K_2Cr_2O_7$	remains orange	orange to green	orange to green

(i) Identify the type of structural isomerism shown between **G** and **H** using the information in Table 3.3.

..... [1]

(ii) Identify the type of structural isomerism shown between **H** and **J** using the information in Table 3.3.

..... [1]

(iii) Draw a possible structure for **H** and for **J**. State the systematic name for each structure.

H

name .....

J

name .....

[3]

(c) **K** has molecular formula C_3H_6O.
When **K** is added to 2,4-dinitrophenylhydrazine, an orange precipitate forms.
When **K** is warmed with Tollens' reagent, a silver mirror forms.

Draw the displayed formula of **K**.

K

[2]

[Total: 16]

4 **(a)** 2-methylpropene reacts with HCl(g) at room temperature. The major organic product is 2-chloro-2-methylpropane.

(i) Complete Fig. 4.1 to show the structure of the intermediate and mechanism for this reaction. Include charges, dipoles, lone pairs of electrons and curly arrows as appropriate.

H—Cl

Fig. 4.1

[3]

(ii) Explain why, in this reaction, 2-chloro-2-methylpropane is produced at a higher yield than 1-chloro-2-methylpropane.

.....

.....

..... [2]

(b) Two bottles labelled **Q** and **M** each contain a straight-chain halogenoalkane with molecular formula C_4H_9**X**, where **X** represents Cl, Br or I.
A sample from each bottle is added to separate samples of equal amounts of aqueous silver nitrate in ethanol. In each reaction, the same organic product, **T**, and a precipitate are made, as shown in Fig. 4.2.

Fig. 4.2

Table 4.1 describes the colour of each of the precipitates made.

Table 4.1

halogenoalkane added to $AgNO_3$(aq) in ethanol	colour of precipitate
Q	white
M	yellow

(i) Identify the functional group present in **T** and name the type of reaction that occurs using the information in Fig. 4.2 and Table 4.1.

functional group in **T** .....

type of reaction .....
[2]

(ii) Construct an ionic equation to describe the formation of the yellow precipitate produced when **M** reacts with $AgNO_3(aq)$ in ethanol.

..... [1]

(iii) Describe which reagent, **Q** or **M**, will produce a precipitate more quickly when each is added to $AgNO_3(aq)$ in ethanol. Explain your answer.

reagent .....

.....

.....
[1]

(iv) When pure **T** is added to alkaline $I_2(aq)$, a yellow precipitate and an anion, **L**, are made.

Identify the anion **L**.

..... [1]

(v) Deduce the structure of the straight-chain halogenoalkane **M**.

[2]

[Total: 12]

5 **(a)** But-2-ene reacts with $KMnO_4$ to form organic product, **Y**.
Y does not react with Na_2CO_3.
A gas is produced when an excess of Na is added to **Y**.

(i) Describe the conditions for the $KMnO_4$ used in the reaction to form **Y** from but-2-ene.

..... [1]

(ii) 24.0 cm^3 of gas is produced when an excess of Na is added to 0.001 mol of **Y**, when measured under room conditions.
Assume that 1 mol of gas occupies 24.0 dm^3 under room conditions.

Deduce a possible structure of **Y**. Explain your answer.

.....

.....

.....
[3]

(b) **Z** contains three types of atom: carbon, hydrogen and a halogen. The mass spectrum of **Z** is recorded. Fig. 5.1 shows a section of the mass spectrum at *m*/*e* greater than 63.
The fragment at *m*/*e* = 64 is the molecular ion peak.

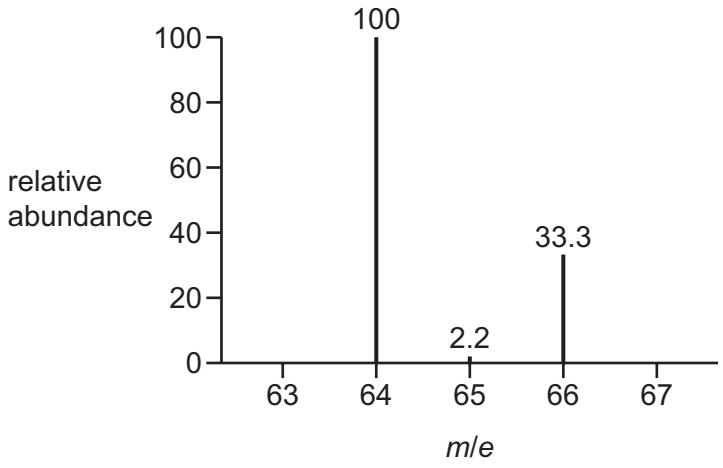

Fig. 5.1

(i) Deduce the number of carbon atoms present in a molecule of **Z** using Fig. 5.1.
Show your working.

..... [1]

(ii) Deduce which halogen is present in **Z** using Fig. 5.1. Explain your answer.

.....

..... [1]

(iii) There are also peaks at *m*/*e* = 29 and *m*/*e* = 49.

Suggest the formulae of these fragments.
Deduce the name of **Z**.

m/*e* = 29 .....

m/*e* = 49 .....

name of **Z** .....
[3]

[Total: 9]

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Important values, constants and standards

molar gas constant	$R = 8.31\,\text{J K}^{-1}\,\text{mol}^{-1}$
Faraday constant	$F = 9.65 \times 10^{4}\,\text{C mol}^{-1}$
Avogadro constant	$L = 6.022 \times 10^{23}\,\text{mol}^{-1}$
electronic charge	$e = -1.60 \times 10^{-19}\,\text{C}$
molar volume of gas	$V_m = 22.4\,\text{dm}^3\,\text{mol}^{-1}$ at s.t.p. (101 kPa and 273 K) $V_m = 24.0\,\text{dm}^3\,\text{mol}^{-1}$ at room conditions
ionic product of water	$K_w = 1.00 \times 10^{-14}\,\text{mol}^2\,\text{dm}^{-6}$ (at 298 K (25 °C))
specific heat capacity of water	$c = 4.18\,\text{kJ kg}^{-1}\,\text{K}^{-1}$ ($4.18\,\text{J g}^{-1}\,\text{K}^{-1}$)

The Periodic Table of Elements

Key

atomic number
atomic symbol
name
relative atomic mass

Group 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0											5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3											13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –	113 Nh nihonium –	114 Fl flerovium –	115 Mc moscovium –	116 Lv livermorium –	117 Ts tennessine –	118 Og oganesson –

lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.4	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
actinoids	89 Ac actinium –	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –