1 (a) Define first ionisation energy.

....................................................................................................................................................
....................................................................................................................................................
.............................................................................................................................................. [2]

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

<table>
<thead>
<tr>
<th>ionisation energy/kJ mol⁻¹</th>
<th>1st</th>
<th>2nd</th>
<th>3rd</th>
<th>4th</th>
<th>5th</th>
<th>6th</th>
<th>7th</th>
<th>8th</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>1310</td>
<td>3390</td>
<td>5320</td>
<td>7450</td>
<td>11000</td>
<td>13300</td>
<td>71000</td>
<td>84100</td>
</tr>
</tbody>
</table>

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 Al²⁺ ion the nuclear attraction for the outer electron is stronger than in an atom of Na.

Compare the electronic structures of Al²⁺ 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>
<thead>
<tr>
<th>Table 1.2</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
</tr>
<tr>
<td>copper-65</td>
</tr>
</tbody>
</table>

(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]
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.

![Graph showing melting points of Period 3 elements](image)

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

<table>
<thead>
<tr>
<th></th>
<th>sodium chloride</th>
<th>phosphorus(V) chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>state at room temp.</td>
<td></td>
<td></td>
</tr>
<tr>
<td>name of change</td>
<td></td>
<td></td>
</tr>
<tr>
<td>occurs on addition</td>
<td></td>
<td></td>
</tr>
<tr>
<td>of water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>pH of final</td>
<td></td>
<td></td>
</tr>
<tr>
<td>solution</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

(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 Br₂. Assume that Ts follows the same trends as the other elements in Group 17. Explain your answer.

equation ........................................................................................................................................

explanation ....................................................................................................................................

....................................................................................................................................................
(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**

<table>
<thead>
<tr>
<th>element</th>
<th>chlorine</th>
<th>bromine</th>
<th>tennessine</th>
</tr>
</thead>
<tbody>
<tr>
<td>melting point / °C</td>
<td>–101</td>
<td>–7.2</td>
<td></td>
</tr>
<tr>
<td>lattice structure of crystalline solid</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

[2]

[Total: 10]
3 G belongs to a group of compounds called ethers.

![Diagram of ether molecule]

**Fig. 3.1**

(a) When G, C₄H₁₀O, is heated, thermal decomposition occurs.

\[
\text{C}_4\text{H}_{10}\text{O}(g) \rightarrow \text{C}_2\text{H}_6(g) + \text{CO}(g) + \text{CH}_4(g) \quad \Delta H = -7.00 \text{ kJ mol}^{-1}
\]

(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.

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

<table>
<thead>
<tr>
<th>bond</th>
<th>bond energy / kJ mol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>C–C</td>
<td>350</td>
</tr>
<tr>
<td>C–O (in G)</td>
<td>360</td>
</tr>
<tr>
<td>C–H</td>
<td>410</td>
</tr>
</tbody>
</table>

Bond energy (C≡O) = ....................... kJ mol\(^{-1}\)
(iii) When \( \text{G, C}_4\text{H}_{10}\text{O,} \) is heated in a sealed container, an equilibrium mixture is produced.

\[
\text{C}_4\text{H}_{10}\text{O}(g) \rightleftharpoons \text{C}_2\text{H}_6(g) + \text{CO}(g) + \text{CH}_4(g)
\]

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

\[
K_c = \\
\text{units = .........................}[2]
\]

(iv) Thermal decomposition of \( \text{G} \) in the presence of \( \text{I}_2 \) affects the activation energy, \( E_a \), for the reaction. Table 3.2 shows \( E_a \) for the thermal decomposition of \( \text{G} \) with and without \( \text{I}_2 \).

<table>
<thead>
<tr>
<th>reaction</th>
<th>( E_a ) (with ( \text{I}_2 ))/kJ mol(^{-1} )</th>
<th>( E_a )/kJ mol(^{-1} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{C}<em>4\text{H}</em>{10}\text{O}(g) \rightarrow \text{C}_2\text{H}_6(g) + \text{CO}(g) + \text{CH}_4(g) )</td>
<td>143</td>
<td>224</td>
</tr>
</tbody>
</table>

State what effect adding \( \text{I}_2 \) 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 \( \text{G} \) at constant temperature, \( T \)°C.

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

![Fig. 3.2](image)
(b) The functional group in **G** is an oxygen atom bonded to two carbon atoms.

![Diagram of G molecule]

**Fig. 3.3**

**G**, **H** and **J** are structural isomers with molecular formula **C₄H₁₀O**. **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 **K₂Cr₂O₇**.

<table>
<thead>
<tr>
<th></th>
<th><strong>G</strong></th>
<th><strong>H</strong></th>
<th><strong>J</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>boiling point/°C</td>
<td>35</td>
<td>117</td>
<td>98</td>
</tr>
<tr>
<td>heat under reflux with excess acidified <strong>K₂Cr₂O₇</strong> remains orange</td>
<td>orange to green</td>
<td>orange to green</td>
<td></td>
</tr>
</tbody>
</table>

(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 \( \text{C}_3\text{H}_6\text{O} \).
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 \( K \).
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.

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

(b) Two bottles labelled Q and M each contain a straight-chain halogenoalkane with molecular formula C₄H₉X, 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.

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

Table 4.1

<table>
<thead>
<tr>
<th>halogenoalkane added to AgNO₃(aq) in ethanol</th>
<th>colour of precipitate</th>
</tr>
</thead>
<tbody>
<tr>
<td>Q</td>
<td>white</td>
</tr>
<tr>
<td>M</td>
<td>yellow</td>
</tr>
</tbody>
</table>

(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.
(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 KMnO₄ to form organic product, Y. Y does not react with Na₂CO₃. A gas is produced when an excess of Na is added to Y.

(i) Describe the conditions for the KMnO₄ used in the reaction to form Y from but-2-ene.
........................................................................................................................................................................ [1]

(ii) 24.0 cm³ 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 dm³ 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.

(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]
### Important values, constants and standards

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

<table>
<thead>
<tr>
<th>Group</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
<th>9</th>
<th>10</th>
<th>11</th>
<th>12</th>
<th>13</th>
<th>14</th>
<th>15</th>
<th>16</th>
<th>17</th>
<th>18</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>H</td>
<td></td>
<td></td>
<td></td>
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<td></td>
<td></td>
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<tr>
<td>2</td>
<td>He</td>
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</tr>
<tr>
<td>3</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
<td></td>
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<td></td>
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</tr>
<tr>
<td>4</td>
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<td>Al</td>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td>Ar</td>
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</tr>
<tr>
<td>5</td>
<td>K</td>
<td>Ca</td>
<td>Sc</td>
<td>Ti</td>
<td>V</td>
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<td>Mn</td>
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<td>Ge</td>
<td>As</td>
<td>Se</td>
<td>Br</td>
<td>Kr</td>
</tr>
<tr>
<td>6</td>
<td>Rb</td>
<td>Sr</td>
<td>Y</td>
<td>Zr</td>
<td>Nb</td>
<td>Mo</td>
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<td>Ru</td>
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<td>Cd</td>
<td>In</td>
<td>Sn</td>
<td>Sb</td>
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<td>I</td>
<td>Xe</td>
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<tr>
<td>7</td>
<td>Cs</td>
<td>Ba</td>
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<td>Hf</td>
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<td>W</td>
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<td>Pt</td>
<td>Au</td>
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<td>Tl</td>
<td>Pb</td>
<td>Bi</td>
<td>Po</td>
<td>At</td>
<td>Rn</td>
</tr>
<tr>
<td>8</td>
<td>Fr</td>
<td>Ra</td>
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<td>Ra</td>
<td>Db</td>
<td>Sg</td>
<td>Bh</td>
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<td>Nh</td>
<td>F</td>
<td>Mc</td>
<td>Lv</td>
<td>Tc</td>
<td>Ti</td>
<td>Zr</td>
<td>Nb</td>
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<tr>
<td>lanthanoids</td>
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<td>Ce</td>
<td>Pr</td>
<td>Nd</td>
<td>Pm</td>
<td>Sm</td>
<td>Eu</td>
<td>Gd</td>
<td>Tb</td>
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<tr>
<td>actinoids</td>
<td>Ac</td>
<td>Th</td>
<td>Pa</td>
<td>U</td>
<td>Np</td>
<td>Pu</td>
<td>Am</td>
<td>Cm</td>
<td>Bk</td>
<td>Cf</td>
<td>Es</td>
<td>Fm</td>
<td>Md</td>
<td>No</td>
<td>Lr</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Key:
- Atomic number
- Atomic symbol
- Name
- Relative atomic mass

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