Which species is not produced by a redox reaction between solid sodium iodide and concentrated sulfuric acid?

A  \( \text{Na}_2\text{SO}_4 \)
B  \( \text{H}_2\text{S} \)
C  \( \text{S} \)
D  \( \text{SO}_2 \)

(Total 1 mark)
This question is about compounds containing ethanedioate ions.

(a) A white solid is a mixture of sodium ethanedioate (Na$_2$C$_2$O$_4$), ethanedioic acid dihydrate (H$_2$C$_2$O$_4$.2H$_2$O) and an inert solid. A volumetric flask contained 1.90 g of this solid mixture in 250 cm$^3$ of aqueous solution.

Two different titrations were carried out using this solution.

In the first titration 25.0 cm$^3$ of the solution were added to an excess of sulfuric acid in a conical flask. The flask and contents were heated to 60 °C and then titrated with a 0.0200 mol dm$^{-3}$ solution of potassium manganate(VII). When 26.50 cm$^3$ of potassium manganate(VII) had been added the solution changed colour.

The equation for this reaction is

$$2\text{MnO}_4^- + 5\text{C}_2\text{O}_4^{2-} + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 10\text{CO}_2$$

In the second titration 25.0 cm$^3$ of the solution were titrated with a 0.100 mol dm$^{-3}$ solution of sodium hydroxide using phenolphthalein as an indicator. The indicator changed colour after the addition of 10.45 cm$^3$ of sodium hydroxide solution.

The equation for this reaction is

$$\text{H}_2\text{C}_2\text{O}_4 + 2\text{OH}^- \rightarrow \text{C}_2\text{O}_4^{2-} + 2\text{H}_2\text{O}$$

Calculate the percentage by mass of sodium ethanedioate in the white solid.

Give your answer to the appropriate number of significant figures.

Show your working.

Percentage by mass of sodium ethanedioate ____________________ %

(8)
(b) Ethanedioate ions react with aqueous iron(III) ions in a ligand substitution reaction. Write an equation for this reaction. Suggest why the value of the enthalpy change for this reaction is close to zero.___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
(2)

(c) Draw the displayed formula of the iron complex produced in the reaction in part (b) Indicate the value of the O—Fe—O bond angle State the type of isomerism shown by the iron complex. Bond angle _________________________________________________________ Type of isomerism ____________________________________________________ (3)

(d) Ethanedioate ions are poisonous because they react with iron ions in the body. Ethanedioate ions are present in foods such as broccoli and spinach. Suggest one reason why people who eat these foods do not suffer from poisoning.___________________________________________________________________
___________________________________________________________________
___________________________________________________________________ (1)

(Total 14 marks)

Iron forms many complexes that contain iron in oxidation states +2 and +3.

(a) Hexaaquairon(III) ions react with an excess of hydrochloric acid in a ligand substitution reaction. Write an equation for this reaction.
___________________________________________________________________
(1)
(b) Explain why the initial and final iron(III) complexes in the equation above have different shapes.

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___________________________________________________________________
___________________________________________________________________

(2)

(c) Hexaaquairon(II) ions react with an excess of H₂NCH₂CH₂NH₂ in a ligand substitution reaction.

Draw the structure of the iron(II) complex formed showing its charge.

(2)

(d) Hexaaquairon(II) ions react with an excess of H₂NCH₂CH₂NH₂ in a ligand substitution reaction.

Which of the following shows the correct change in entropy for a reaction of hexaaquairon(II) ions with H₂NCH₂CH₂NH₂?

Tick (✔) one box.

- change in entropy is negative
- change in entropy is close to zero
- change in entropy is positive
(e) The percentage of iron(II) sulfate in iron tablets can be determined by titration with potassium manganate(VII) in acidic solution.

Deduce an ionic equation for the reaction of iron(II) ions with manganate(VII) ions.

___________________________________________________________________

(1)

(f) A student dissolved 1980 mg of iron tablets in an excess of dilute sulfuric acid. The solution was titrated with 0.0200 mol dm$^{-3}$ potassium manganate(VII) solution. A 32.50 cm$^3$ volume of potassium manganate(VII) solution was required to reach the end point in the titration.

Calculate the percentage of iron in the sample of iron tablets. Give your answer to the appropriate number of significant figures.

Percentage __________________ %

(4)

(g) State the colour change at the end point in this titration.

___________________________________________________________________

(1)

(Total 12 marks)

4 Which of these species is the best reducing agent?

A Cl$_2$  
B Cl$^-$  
C I$_2$  
D I$^-$

(Total 1 mark)
A student oxidised a solution of hydrochloric acid with a few drops of sodium chlorate(l) solution. The reaction mixture effervesced and turned pale green. The gas formed bleached universal indicator paper.

(a) Write a half-equation for the oxidation of chloride ions.

______________________________________________________________

(b) Write a half-equation for the reduction of chlorate(l) ions to chlorine in acidic conditions.

______________________________________________________________

(c) Write an overall equation for the redox reaction of chlorate(l) ions with hydrochloric acid.

______________________________________________________________

(d) A solution of sodium chlorate(l) was added to a colourless solution of potassium iodide. Suggest what is observed.

Explain the reaction that leads to this observation.

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(Total 6 marks)
The halogens are the elements in Group 7.

(a) The electronegativities of the halogens are shown in the table.

<table>
<thead>
<tr>
<th>Halogen</th>
<th>Fluorine</th>
<th>Chlorine</th>
<th>Bromine</th>
<th>Iodine</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electronegativity</td>
<td>4.0</td>
<td>3.0</td>
<td>2.8</td>
<td>2.5</td>
</tr>
</tbody>
</table>

Explain the trend in electronegativities shown by the halogens.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

(2)

(b) The halogens can all behave as oxidising agents in reactions.

(i) Explain, in terms of electron transfer, the meaning of the term oxidising agent.

___________________________________________________________________
___________________________________________________________________

(1)

(ii) An equation for the reaction that takes place when chlorine gas is bubbled through aqueous potassium bromide is shown.

\[ \text{Cl}_2(\text{g}) + 2\text{KBr(aq)} \rightarrow \text{Br}_2(\text{aq}) + 2\text{KCl(aq)} \]

Explain, with reference to the oxidation states, why this is a redox reaction.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

(1)
(c) Solid sodium halides react with concentrated sulfuric acid.

(i) A sample of solid sodium iodide is reacted with concentrated sulfuric acid. A black solid forms and hydrogen sulfide gas is produced.

Write a half-equation for the reaction of sulfuric acid to form hydrogen sulfide.

__________________________________________________________________________ (1)

(ii) Write a half-equation for the formation of the black solid.

__________________________________________________________________________ (1)

(iii) Use your answers to parts (c) (i) and (c) (ii) to write an overall equation for the reaction of sodium iodide with concentrated sulfuric acid.

__________________________________________________________________________ (1)

(iv) Give the role of sulfuric acid in its reaction with sodium iodide.

Tick (✓) one box.

- Acid
- Oxidising agent
- Reducing agent
- Electrophile

(1)

(v) Write an equation for the reaction of concentrated sulfuric acid with solid sodium fluoride.

__________________________________________________________________________ (1)

(vi) Suggest one reason why the reaction of sodium fluoride with concentrated sulfuric acid is different from the reaction with sodium iodide.

__________________________________________________________________________ (1)
Chlorine is an important industrial chemical.

(a) Chlorine is formed when KMnO$_4$ reacts with hydrochloric acid. The ionic equation for this redox reaction is

$$16H^+ + 2MnO_4^- + 10Cl^- \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$$

(i) Deduce the half-equation for the oxidation of chloride ions to chlorine.

(ii) Give the oxidation state of manganese in the MnO$_4^-$ ion.

(iii) Deduce the half-equation for the reduction of the MnO$_4^-$ ions in acidified solution to manganese(II) ions and water.

(Total 13 marks)
(b) Chlorine behaves as an oxidising agent in the extraction of bromine from seawater. In this process, chlorine gas is bubbled through a solution containing bromide ions.

(i) Write the simplest ionic equation for the reaction of chlorine with bromide ions.
__________________________________________________________________________________________

(1)

(ii) Give one observation that would be made during this reaction.
__________________________________________________________________________________________
__________________________________________________________________________________________

(1)

(iii) In terms of electrons, state the meaning of the term oxidising agent.
__________________________________________________________________________________________
__________________________________________________________________________________________

(1)

(c) In sunlight, chlorine can also oxidise water slowly to form oxygen.

Write an equation for this reaction.
Give the oxidation state of chlorine in the chlorine-containing species that is formed.

Equation
__________________________________________________________________________________________

Oxidation state of chlorine in the species formed __________________________________________________________________________

(2)

(d) Explain why chlorine has a lower boiling point than bromine.
__________________________________________________________________________________________
__________________________________________________________________________________________
__________________________________________________________________________________________
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__________________________________________________________________________________________

(2)

(Total 10 marks)
A green solution, X, is thought to contain \([\text{Fe(H}_2\text{O)}_6]^{2+}\) ions.

(a) The presence of these ions can be confirmed by reacting separate samples of solution X with aqueous ammonia and with aqueous sodium carbonate.

Write equations for each of these reactions and describe what you would observe.

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(b) A 50.0 cm\(^3\) sample of solution X was added to 50 cm\(^3\) of dilute sulfuric acid and made up to 250 cm\(^3\) of solution in a volumetric flask.

A 25.0 cm\(^3\) sample of this solution from the volumetric flask was titrated with a 0.0205 mol dm\(^{-3}\) solution of KMnO\(_4\).

At the end point of the reaction, the volume of KMnO\(_4\) solution added was 18.70 cm\(^3\).

(i) State the colour change that occurs at the end point of this titration and give a reason for the colour change.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________

(4)

(2)
For many years, swimming pool water has been treated with chlorine gas. The chlorine is added to kill any harmful bacteria unintentionally introduced by swimmers. Pool managers are required to check that the chlorine concentration is high enough to kill the bacteria without being a health hazard to the swimmers.

When chlorine reacts with water in the absence of sunlight, the chlorine is both oxidised and reduced and an equilibrium is established.

(a) Write an equation for this equilibrium.

For each chlorine-containing species in the equation, write the oxidation state of chlorine below the species.
(b) The pool manager maintains the water at a pH slightly greater than 7.0

Explain how this affects the equilibrium established when chlorine is added to water.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
(1)

(c) Explain why chlorine is used to kill bacteria in swimming pools, even though chlorine is toxic.

___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
___________________________________________________________________
(2)

This question is about Group 7 chemistry.

(a) Sea water is a major source of iodine.

The iodine extracted from sea water is impure. It is purified in a two-stage process.

| Stage 1 | I₂ + 2H₂O + SO₂ → 2HI + H₂SO₄ |
| Stage 2 | 2HI + Cl₂ → I₂ + 2HCl |

(i) State the initial oxidation state and the final oxidation state of sulfur in Stage 1.

Oxidation state of S in SO₂_____________________________________________________

Oxidation state of S in H₂SO₄___________________________________________________

(2)

(ii) State, in terms of electrons, what has happened to chlorine in Stage 2.

____________________________________________________________________________
____________________________________________________________________________
(1)
(b) When concentrated sulfuric acid is added to potassium iodide, iodine is formed in the following redox equations.

\[
\text{KI} + \text{H}_2\text{SO}_4 \rightarrow \text{KHSO}_4 + \text{I}_2 + \text{S} + \text{H}_2\text{O}
\]

\[
8\text{KI} + 9\text{H}_2\text{SO}_4 \rightarrow 8\text{KHSO}_4 + 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}
\]

(i) Balance the equation for the reaction that forms sulfur.

(ii) Deduce the half-equation for the formation of iodine from iodide ions.

(iii) Deduce the half-equation for the formation of hydrogen sulfide from concentrated sulfuric acid.

(c) A yellow precipitate is formed when silver nitrate solution, acidified with dilute nitric acid, is added to an aqueous solution containing iodide ions.

(i) Write the simplest ionic equation for the formation of the yellow precipitate.

(ii) State what is observed when concentrated ammonia solution is added to this yellow precipitate.

(iii) State why the silver nitrate solution is acidified when testing for iodide ions.
(iv) Explain why dilute hydrochloric acid is **not** used to acidify the silver nitrate solution in this test for iodide ions.

______________________________________________________________

______________________________________________________________

______________________________________________________________ (1)

(d) Chlorine is toxic to humans. This toxicity does not prevent the large-scale use of chlorine in water treatment.

(i) Give one reason why water is treated with chlorine.

______________________________________________________________

______________________________________________________________

______________________________________________________________ (1)

(ii) Explain why the toxicity of chlorine does **not** prevent this use.

______________________________________________________________

______________________________________________________________

______________________________________________________________ (1)

(iii) Write an equation for the reaction of chlorine with cold water.

______________________________________________________________ (1)

(e) Give the formulas of the **two** different chlorine-containing compounds that are formed when chlorine reacts with cold, dilute, aqueous sodium hydroxide.

Formula 1 __________________________________________________________

Formula 2 __________________________________________________________ (1)

(Total 14 marks)
(a) Moles $\text{MnO}_4^- = \frac{26.50 \times 0.02}{1000} = 5.30 \times 10^{-4}$

Moles in 25 cm$^3$ sample / pipette $\text{C}_2\text{O}_4^{2-}$ (from acid and salt)
$= 5.30 \times 10^{-4} \times \frac{5}{2} = (1.325 \times 10^{-3})$

Moles $\text{NaOH} = \frac{1.045 \times 0.1}{1000} = (1.045 \times 10^{-3})$

So moles $\text{C}_2\text{O}_4^{2-}$ from acid in 25 cm$^3$ sample / pipette
$= 1.045 \times 10^{-3} \div 2 = 5.225 \times 10^{-4}$

Hence moles $\text{C}_2\text{O}_4^{2-}$ in sodium ethanedioate in 25 cm$^3$
$= 1.325 \times 10^{-3} - 5.225 \times 10^{-4} = (8.025 \times 10^{-4})$

So moles $\text{C}_2\text{O}_4^{2-}$ in sodium ethanedioate in original sample
$= 8.025 \times 10^{-4} \times 10 = (8.025 \times 10^{-3})$

Mass $\text{Na}_2\text{C}_2\text{O}_4 = 8.025 \times 10^{-3} \times 134.0 = 1.075(35)$ g

So % sodium ethanedioate in original sample

$\frac{1.075(35)}{1.90} \times 100 = 56.6\%$ to 3 sig fig

The first CE is penalised by 2 marks; further errors are penalised by one mark each

$M2 = M1 \times \frac{5}{2}$

$M4 = M3 \div 2$

$M5 = M2 - M4$ (do not allow if negative and do not allow $= M4 - M2$)

If no subtraction, max = 5 ($M1, M2, M3, M4$ and $M6$)

If incorrect subtraction, max = 6 ($M1, M2, M3, M4, M6$ and $M7$)

$M6 = M5 \times 10$

($M6$ can be scored by multiplying $M2$ and $M4$ by 10 before subtraction (giving $1.325 \times 10^{-2} - 5.225 \times 10^{-3} = 8.025 \times 10^{-3}$)

$M7 = M6 \times 134$

$M8 = (M7/1.90) \times 100$ Allow 56.5 – 56.8%
(b) \[ \text{[Fe(H}_2\text{O)}_6\text{]}^{3+} + 3\text{C}_2\text{O}_4^{2−} \rightarrow \text{[Fe(C}_2\text{O}_4\text{)]}_3^{3−} + 6\text{H}_2\text{O} \]

There are 6 Fe–O bonds broken and then made / same number and type of bond being broken and made.

(c) 

Ignore all charges even if wrong
Ignore absence of square brackets
Candidates do not need to show 3D shape

90° or 180°

optical

(d) The ethanedioic acid is only present in small quantities/low concentration in these foods.

(a) \[ \text{[Fe(H}_2\text{O)}_6\text{]}^{3+} + 4\text{Cl}^{−} \rightarrow \text{FeCl}_4^{−} + 6\text{H}_2\text{O} \]

(b) Cl− is a bigger ligand

So only 4Cl− can fit around the metal

Allow fewer Cl− can fit around the metal
(c) [Structure of complex]

(d) Change in entropy is positive

(e) \[5\text{Fe}^{2+} + \text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}\]

(f) Amount of manganate (VII) = \(6.50 \times 10^{-4}\) mol

\[
\text{Amount of iron(II)} = 3.25 \times 10^{-3}\ \text{mol}
\]

\[
\text{ie } M1 \times 5
\]

\[
\text{Mass of iron} = 0.181 \ \text{g} = 181 \ \text{mg}
\]

\[
\text{Allow } M2 \times 55.8
\]

\[
\text{Percentage Fe} = \frac{181}{1980} \times 100 = 9.14(\%) \ 3\ \text{sf}
\]

(g) Colourless to pale pink

4

5

(a) \[2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^{(-)}\]

\[
\text{Allow } 2\text{Cl}^- - 2e^{(-)} \rightarrow \text{Cl}_2
\]

\[
\text{Allow correct equation forming } \text{ClO}^- \text{ but not } \text{Cl}^+
\]

(b) \[2\text{ClO}^- + 4\text{H}^+ + 2e^{-} \rightarrow \text{Cl}_2 + 2\text{H}_2\text{O}\]

\[
\text{Allow HClO in correctly balanced equation}
\]
(c) \[
\text{ClO}^- + \text{Cl}^- + 2\text{H}^+ \rightarrow \text{Cl}_2 + \text{H}_2\text{O}
\]

allow \(\text{HClO} + \text{HCl} + \rightarrow \text{Cl}_2 + \text{H}_2\text{O}\)

1

(d) Goes brown (or shades of brown)

Allow black ppt/solid but NOT black solution or purple

Due to iodine or \(\text{I}_3^-\)

Correct \(\frac{1}{2}\) equation scores M2 and M3

Because \(\text{I}^-\) oxidised

1

(a) Increasing atomic radius / shielding / number of shells / size (down group) or reverse argument

\text{NOT} \ 'molecules'

1

Decreasing attraction of nucleus/protons for shared (electron) pair / bond electrons

\text{NOT} \ if \ attraction \ for \ single \ electron \ implied

1

(b) (i) Electron acceptor / species that accepts electrons / species that gains electrons

\text{NOT} \ electron \ pair

\text{NOT} \ just \ 'gain \ of \ electrons'

1

(ii) Chlorine 0 to \(-1\) / oxidation state/number of chlorine decreases

\text{AND}

Bromine \(-1\) to 0 / oxidation state/number of bromine increases

Penalise if oxidised for chlorine and/or reduced for bromine

Credit oxidation states if labelled on equation

1

(c) (i) \(\text{H}_2\text{SO}_4 + 8\text{H}^+ + 8\text{e}^{(-)} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O}\)

ALLOW \(\text{SO}_4^{2-} + 10\text{H}^+ + 8\text{e}^{(-)} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O}\)

ALLOW fractions/multiples

IGNORE state symbols

1

(ii) \(2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^{(-)}\)

ALLOW fractions/multiples

IGNORE state symbols

ALLOW \(2\text{I}^- - 2\text{e}^{(-)} \rightarrow \text{I}_2\)

1
(iii) \( \text{H}_2\text{SO}_4 + 8\text{H}^+ + 8\text{I}^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 \)  

\text{ALLOW}  

\( \text{H}_2\text{SO}_4 + 8\text{HI} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 \)  

\( \text{SO}_4^{2-} + 2\text{H}^+ + 8\text{HI} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 \)  

\( \text{SO}_4^{2-} + 10\text{H}^+ + 8\text{I}^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 \)  

\( 9\text{H}_2\text{SO}_4 + 8\text{I}^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 + 8\text{HSO}_4^- \)  

\( 9\text{H}_2\text{SO}_4 + 8\text{NaI} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 + 8\text{NaHSO}_4 \)  

\( \text{H}_2\text{SO}_4 + 8\text{H}^+ + 8\text{NaI} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 + 8\text{Na}^+ \)  

\( 5\text{H}_2\text{SO}_4 + 8\text{I}^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 + 4\text{SO}_4^{2-} \)  

\( 5\text{H}_2\text{SO}_4 + 8\text{NaI} \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} + 4\text{I}_2 + 4\text{Na}_2\text{SO}_4 \)  

1

(iv) ‘Oxidising agent’ box ticked  

1

(v) \( \text{H}_2\text{SO}_4 + 2\text{NaF} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{HF} \)  

\text{OR}  

\( \text{H}_2\text{SO}_4 + \text{NaF} \rightarrow \text{NaHSO}_4 + \text{HF} \)  

1

(vi) Fluoride less powerful reducing agent (than iodide)  

\text{OR}  

Fluoride less easily oxidised than iodide  

Or reverse argument in either case  

\text{NOT general group VII trend statement}  

\text{NOT fluorine/F or iodine/I}  

\text{Must be comparative}  

1

(d) (i) \( \text{Cl}_2 + \text{H}_2\text{O} = 2\text{H}^+ + \text{Cl}^- + \text{ClO}^-/\text{HCl} + \text{HOCl} \)  

\text{ALLOW} \rightarrow \text{for } =  

1

(ii) Equilibrium \text{shifts/moves} left  

1

(Producing) chlorine (which) is toxic/poisonous  

\text{Mark independently}  

1
(a) (i) \[ 2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^- \]

*Ignore state symbols*

*Credit loss of electrons from LHS*

*Credit multiples*

*Do not penalise absence of charge on electron*

(ii) \[ +7 \text{ OR VII OR } +\text{VII} \]

*Allow Mn\(^{+7}\) and 7+

(iii) \[ \text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \]

*Ignore state symbols*

*Credit loss of electrons from RHS*

*Credit multiples*

*Do not penalise absence of charge on electron*

(b) (i) \[ \text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2 \]

*OR*

\[ \frac{1}{2}\text{Cl}_2 + \text{Br}^- \rightarrow \text{Cl}^- + \frac{1}{2}\text{Br}_2 \]

*One of these two equations only*

*Ignore state symbols*

(ii) *(Turns to) yellow / orange / brown (solution)*

*Penalise “red / reddish” as the only colour*

*Accept “red-brown” and “red-orange”*

*Ignore “liquid”*

*Penalise reference to a product that is a gas or a precipitate*

(iii) *(Chlorine) gains electron(s) / takes electron(s) / accepts electron(s) (from the bromide ions)*

*OR*

*(Chlorine) causes another species (Br\(^{-}\)) to lose electron(s)*

*Penalise “electron pair acceptor”*

*Not simply “causes loss of electrons”*
(c) \[ \text{M1} \quad 2\text{Cl}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{HCl} + \text{O}_2 \]

\((4\text{H}^+ + 4\text{Cl}^-)\)

M2 Oxidation state \(-1\)

*Ignore state symbols

*Credit multiples

**M2** consequent [ial on HCl or Cl\(^-\) which **must** be the only chlorine-containing product in the (un)balanced equation.

*For M2 allow Cl\(^-\) or Cl\(^{1-}\) but **not** Cl\(^{2-}\)*

(d) **M1** The relative size (of the molecules / atoms)

Chlorine is smaller than bromine **OR** has fewer electrons / electron shells

*For M1 ignore whether it refers to molecules or atoms.

**OR** It is smaller / It has a smaller atomic radius / it is a smaller molecule / atom (or converse)

\(CE=0\) for the clip for reference to (halide) ions or incorrect statements about relative size

*Ignore molecular mass and M\(_r\)*

M2 How size of the intermolecular force affects energy needed

*Ignore shielding*

The forces between chlorine / Cl\(_2\) molecules are weaker (than the forces between bromine / Br\(_2\) molecules)

(or converse for bromine)

**OR** chlorine / Cl\(_2\) has weaker / fewer / less (VdW) intermolecular forces / forces between molecules

(or converse for bromine)

QoL **in M2** for clear reference to the difference in size of the force between molecules. Reference to Van der Waals forces alone is not enough.

*Penalise M2 if (covalent) bonds are broken*
(a) \[[\text{Fe(H}_2\text{O)}_6]^{2+} + 2\text{NH}_3 \rightarrow \text{Fe(H}_2\text{O)}_4(\text{OH})_2 + 2\text{NH}_4^+\]
Allow equation with OH\(^{-}\) provided equation showing formation of
OH\(^{-}\) from NH\(_3\) given

Green precipitate

\[[\text{Fe(H}_2\text{O)}_6]^{2+} + \text{CO}_3^{2-} \rightarrow \text{FeCO}_3 + 6\text{H}_2\text{O}\]

Green precipitate

*effervescence incorrect so loses M4*

(b) (i) Colourless / (pale) green changes to pink / purple (solution)

*Do not allow pale pink to purple*

Just after the end−point MnO\(_4^-\) is in excess / present

(ii) MnO\(_4^-\) + 8H\(^+\) + 5Fe\(^{2+}\) → Mn\(^{2+}\) + 4H\(_2\)O + 5Fe\(^{3+}\)

Moles KMnO\(_4\) = 18.7 × 0.0205 / 1000 = (3.8335 × 10\(^-4\))

*Process mark*

Moles Fe\(^{2+}\) = 5 × 3.8335 × 10\(^-4\) = 1.91675 × 10\(^{-3}\)

*Mark for M2 × 5*

Moles Fe\(^{2+}\) in 250 cm\(^3\) = 10 × 1.91675 × 10\(^{-3}\) = 0.0191675 moles in 50 cm\(^3\)

*Process mark for moles of iron in titration (M3) × 10*

Original conc Fe\(^{2+}\) = 0.0191675 × 1000 / 50 = 0.383 mol dm\(^{-3}\)

*Answer for moles of iron (M4) × 1000 / 50*

*Answer must be to at least 2 sig. figs. (0.38)*

[11]

(a) Cl\(_2\) + H\(_2\)O = HOCl + HCl

*Allow the products shown as ions.*

Cl\(_2\) = 0, HOCl = +1 and HCl = −1

*1 mark for all three oxidation states correct. Allow a reaction arrow in this equation. Oxidation states must match the species*
(b) Hydroxide / alkali ions react with the acids

Mark independently

Equilibrium moves to the right

(c) Only used in small amounts

The health benefits outweigh the risks

(a) (i) \[ M1^{+} 4 \text{ OR } IV \]

\[ M2^{+} 6 \text{ OR } VI \]

(ii) It / Chlorine has gained / accepted electron(s)

OR

Correctly balanced half-equation eg \[ \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \]

Credit 1 or 2 electrons but not lone pair.

The idea of 'reduction' alone is not enough.

(b) (i) \[ 6\text{KI} + 7\text{H}_2\text{SO}_4 \rightarrow 6\text{KHSO}_4 + 3\text{I}_2 + \text{S} + 4\text{H}_2\text{O} \]

(ii) \[ 2\text{I}^- \rightarrow \text{I}_2 + 2e^- \]

OR

\[ 8\text{I}^- \rightarrow 4\text{I}_2 + 8e^- \]

Ignore charge on the electron unless incorrect.

Or multiples.

Credit the electrons being subtracted on the LHS.

Ignore state symbols.

(iii) \[ \text{H}_2\text{SO}_4 + 8\text{H}^+ + 8e^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} \]

OR

\[ \text{SO}_4^{2-} + 10\text{H}^+ + 8e^- \rightarrow \text{H}_2\text{S} + 4\text{H}_2\text{O} \]

Ignore charge on the electron unless incorrect.

Or multiples.

Credit the electrons being subtracted on the RHS.

Ignore state symbols.
(c) (i) \[ \text{Ag}^+ + \text{I}^- \rightarrow \text{AgI} \] ONLY

*Ignore state symbols.*

*Not multiples.*

(ii) The precipitate / solid / it does not dissolve / is insoluble / remains

OR a white / cream / yellow solid / precipitate

OR stays the same

OR no (visible / observable) change

OR no effect / no reaction

*Ignore ‘nothing (happens)’.*

*Ignore ‘no observation’.*

(iii) The silver nitrate is acidified to

• react with / remove (an)ions that would interfere with the test

*Credit a correct reference to ions that give a ‘false positive’.*

• prevent the formation of other silver precipitates / insoluble silver compounds that would interfere with the test

*Do not penalise an incorrect formula for an ion that is written in addition to the name.*

• remove (other) ions that react with the silver nitrate

*If only the formula of the ion is given, it must be correct.*

• react with / remove carbonate / hydroxide / sulfite (ions)

*Ignore ‘sulfate’.*

(iv) HCl would form a (white) precipitate / (white) solid (with silver nitrate and this would interfere with the test)

*It is not sufficient simply to state either that it will interfere or simply that the ions / compounds react to form AgCl*

(d) (i) Any one from

*Ignore ‘to clean water’.*

• to sterilise / disinfect water

*Ignore ‘water purification’ and ‘germs’.*

• to destroy / kill microorganisms / bacteria / microbes / pathogens

*Credit ‘remove bacteria etc’ / prevent algae.*
(ii) The (health) benefit outweighs the risk

OR

a clear statement that once it has done its job, little of it remains

OR

used in (very) dilute concentrations / small amounts / low doses

(iii) \[ \text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{HClO} + \text{HCl} \]

OR

\[ \text{Cl}_2 + \text{H}_2\text{O} \rightarrow 2\text{H}^+ + \text{ClO}^- + \text{Cl}^- \]

OR

\[ 2\text{Cl}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{HCl} + \text{O}_2 \]

Credit HOCl or ClOH

Or multiples.

Credit other ionic or mixed representations.

Ignore state symbols.

(e) **In either order - Both required for one mark only**

Credit correct ionic formulae.

NaClO (OR NaOCl) and NaCl

Give credit for answers in equations unless contradicted.
Only 42.2% of students selected the correct answer. Perhaps many students did not appreciate that a range of sulfur-based products are formed in the reaction.

This was a challenging question with students finding most parts of it difficult.

(a) Only the best students got to the end of the calculation. Many students scored the first four marks, and some of these then realised that they needed to multiply by ten because of the dilution used. Very few students appreciated that the amount in moles of ethanedioic acid had to be taken away from the total amount in moles of ethanedioate ions to find the amount in moles of sodium ethanedioate. Those students who did, usually went on to get the right answer.

(b) Students often failed to give a balanced equation; charges on the complex ions were often incorrect, and many did not give water as one of the products. Many students understood that the same number of bonds were broken and formed, but few could also state that they were the same type of bonds. Some stated that the bond enthalpies were similar, but could not explain why.

(c) Many students struggled to draw the displayed formula; those who got close often omitted the carbon atoms from the ligand. Most students knew the bond angle, but some incorrectly gave 120° or 109.5°. The most common incorrect type of isomerism given was stereoisomerism or cis/trans.

(d) Common answers that did not score here included references to ethanedioate ions being indigestible, or being so firmly chelated to iron ions already that they were not free to be absorbed.

This question was answered poorly. Even the half equation in (a) was usually not answered correctly and the more challenging equations in (b) and (c) were only seen from the better students. In (d) the explanation that iodide ions are oxidised wasn’t seen very often but a reasonable proportion of students scored some marks on this question.

In a definition of electronegativity students must make it clear that they are referring to the attraction between the nucleus and the shared pair of bond electrons. Most students were awarded the first mark but many answers were too vague in their references to gain the second mark. Most of the rest of this question was well answered although there were some issues with balancing charges in the redox equations.

Very few students showed the acids formed in part (d)(i) as being dissociated and this, in turn, seemed to make part (d)(ii) harder than expected because many students failed to spot that addition of acid would shift the equilibrium to the left and so produce toxic chlorine gas. Although there was no specific reference to le Chatelier in the question (which would have given a more specific clue) the stem of part (d) did refer to an equilibrium mixture and this was intended to point students in the right direction.

This was a relatively straightforward start to the paper, but the usual errors were made in balancing equations. The idea of oxidation state is relatively well understood. In part (d), students demonstrated a lack of clarity in what happens during the process of boiling, often alluding to bond breakage and rarely referring to the forces between molecules.
Most students gave correct observations in part (a) but their attempts at equations were less successful and often unbalanced. Very few correct answers to part (b)(i) were seen. The most usual response was to suggest, without any explanation, that the colour change was from purple to colourless. Part (b)(ii) produced a whole range of responses. Good students were able to score all 5 marks in impressive style but others lost marks. The most common mistake was an incorrect equation and the use of an incorrect molar ratio for moles of iron(II) to manganate(VII). Many students were also unable to calculate the concentration of iron(II) ions in the original solution from the number of moles of iron(II) ions in the 250 cm³ flask. A factor of 1000/50 was required but factors such as 1000/250 or 250/25 were common mistakes.

In part (a) the equation and oxidation states were frequently seen but, for some, the state for chlorine itself was omitted. In part (b) there was often no reference to the equilibrium shift and over-complicated answers led to contradictions in responses. Part (c) was well known.

Part (a) was answered well but part (b) proved more demanding. Only about one-third of students were able to balance the equation in part (b)(i) and only about half of students gave correct half-equations in parts (b)(ii) and (b)(iii). The halide ion analysis in part (c) was well known with the lowest scoring part being (c)(iv), with many having failed to recognise that it is insufficient to simply state that AgCl is formed. Any question involving NaClO is always demanding and part (e) was no exception with less than half of the students scoring the mark.