

1 What is the oxidation number of Mn in K_2MnO_4 ?

A +4

B +5

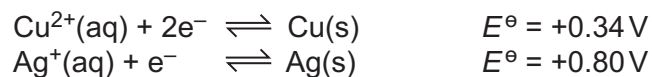
C +6

D +7

Your answer

[1]

- 2 A cell is constructed from the two redox systems below.



Which statement(s) is/are correct for the cell?

- 1 The cell potential is 1.14 V.
 - 2 The reaction at the copper electrode is $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-}$.
 - 3 The silver electrode increases in mass.
- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

[1]

- 3(a)** On gently heating, the compound KClO_3 reacts as shown in the equation.



This reaction is an example of disproportionation.

- (i) State what is meant by *disproportionation* and use oxidation numbers to show that disproportionation has taken place.

.....

.....

.....

.....

.....

..... [3]

- 4 This question is about redox, electrode potentials and feasibility.

Table 22.1 shows standard electrode potentials for four redox systems. You need to use this information to answer the questions below.

Redox system	Equation	E°/V
1	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
2	$\text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.17
3	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
4	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 22.1

- (a) A standard cell is set up in the laboratory based on redox systems **1** and **3** and the standard cell potential is measured.

- (i) Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential.

Include details of the apparatus, solutions and the standard conditions required to measure this standard cell potential.

Standard conditions

 [4]

- (ii) Predict the standard cell potential of this cell.

standard cell potential = ... V [1]

(b) In **Table 22.1**, what is the strongest reducing agent and the strongest oxidising agent?

Strongest reducing agent

Strongest oxidising agent

[2]

(c) Electrode potentials can be used to predict the feasibility of reactions.

Construct an overall equation for the predicted reaction between the species in redox systems **2** and **4**.

..... [2]

- 5 In the compound $[\text{ICl}_2]^+ [\text{SbCl}_6]^-$, the oxidation number of chlorine is -1 .

What are the oxidation numbers of I and Sb in the compound?

	I	Sb
A	+1	+5
B	+1	+7
C	+3	+5
D	+3	+7

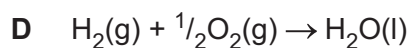
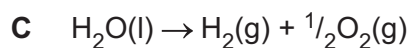
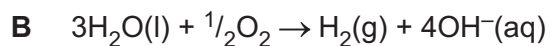
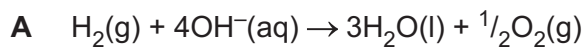
Your answer

[1]

- 6 The redox equilibria for a hydrogen–oxygen fuel cell in alkaline solution are shown below.



What is the equation for the overall cell reaction?



Your answer

[1]

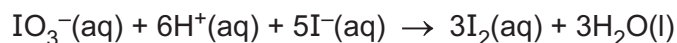
- 7(a) Compound **B** is an iodate(V) salt of a Group 1 metal.
The iodate(V) ion has the formula IO_3^- .

A student carries out a titration to find the formula of compound **B**.

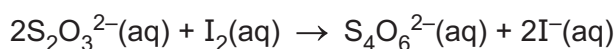
Step 1: The student dissolves 1.55 g of **B** in water and makes up the solution to 250.0 cm^3 in a volumetric flask.

Step 2: The student pipettes 25.00 cm^3 of the solution of **B** into a conical flask, followed by 10 cm^3 of dilute sulfuric acid and an excess of KI(aq) .

The iodate(V) ions are reduced to iodine, as shown below.



Step 3: The resulting mixture is titrated with $0.150 \text{ mol dm}^{-3} \text{ Na}_2\text{S}_2\text{O}_3(\text{aq})$.



The student repeats **step 2** and **step 3** until concordant titres are obtained.

Titration readings

Titration	Trial	1	2	3
Final burette reading / cm^3	24.00	47.40	23.75	47.05
Initial burette reading / cm^3	0.00	24.00	0.00	23.20
Titre / cm^3				

Table 20.1

- (i) Complete **Table 20.1** and calculate the mean titre that the student should use for analysing the results.

mean titre = ... cm^3 [2]

- (ii) The uncertainty in each burette reading is $\pm 0.05 \text{ cm}^3$.

Calculate the percentage uncertainty in the titre obtained from **titration 1**.

Give your answer to **two** decimal places.

percentage uncertainty = ... % [1]

- (iii) Describe and explain how the student should determine the end point of this titration accurately.

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

- (iv) Determine the relative formula mass and formula of the Group 1 iodate(V), **B**.

Show your working.

relative formula mass of **B** = ...

formula of **B** = ... [5]

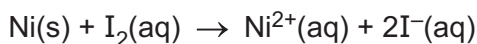
21 This question is about some reactions of d block elements and their ions.

Table 21.1 shows standard electrode potentials which will be needed within this question.

$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-}$	$\rightleftharpoons$	$\text{Zn}(\text{s})$	$E^{\circ} = -0.76\text{V}$
$\text{Cr}^{3+}(\text{aq}) + \text{e}^{-}$	$\rightleftharpoons$	$\text{Cr}^{2+}(\text{aq})$	$E^{\circ} = -0.42\text{V}$
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-}$	$\rightleftharpoons$	$\text{Ni}(\text{s})$	$E^{\circ} = -0.25\text{V}$
$\text{I}_2(\text{aq}) + 2\text{e}^{-}$	$\rightleftharpoons$	$2\text{I}^{-}(\text{aq})$	$E^{\circ} = +0.54\text{V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-}$	$\rightleftharpoons$	$\text{Fe}^{2+}(\text{aq})$	$E^{\circ} = +0.77\text{V}$
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{e}^{-}$	$\rightleftharpoons$	$2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	$E^{\circ} = +1.33\text{V}$
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^{+}(\text{aq}) + 2\text{e}^{-}$	$\rightleftharpoons$	$2\text{H}_2\text{O}(\text{l})$	$E^{\circ} = +1.78\text{V}$

Table 21.1

8(a) A standard cell is set up in the laboratory with the cell reaction shown below.



- (i) Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential.

Include details of apparatus, solutions and the standard conditions required.

Standard conditions

.....

..... **[4]**

- (ii) Predict the standard cell potential of this cell.

standard cell potential = ... V [1]

- (b) Use the information in **Table 21.1** to help you answer both parts of this question.

- (i) Write the overall equation for the oxidation of Fe^{2+} by acidified H_2O_2 .

..... [1]

- (ii) Zinc reacts with acidified $\text{Cr}_2\text{O}_7^{2-}$ ions to form Cr^{2+} ions in two stages.

Explain why this happens in terms of electrode potentials and equilibria.

Include overall equations for the reactions which occur.

.....
.....
.....
.....
.....
.....
.....
.....
.....
.....
..... [4]

9 What is the oxidation number of N in $\text{Mg}(\text{NO}_2)_2 \cdot 3\text{H}_2\text{O}$?

A +2

B +3

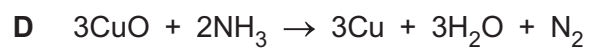
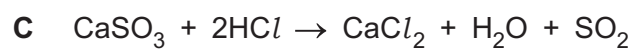
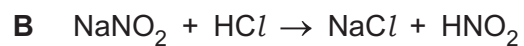
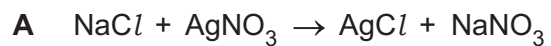
C +4

D +5

Your answer

[1]

10 Which reaction is a redox reaction?



Your answer

[1]

- 11 Four redox systems relevant to hydrogen–oxygen fuel cells are shown below.

	E°/V
$\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightleftharpoons \text{OH}^-(\text{aq}) + \frac{1}{2}\text{H}_2(\text{g})$	-0.83
$\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \frac{1}{2}\text{H}_2(\text{g})$	0.00
$\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{OH}^-(\text{aq})$	+0.40
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}(\text{l})$	+1.23

Which statement(s) is/are correct for an alkaline hydrogen–oxygen fuel cell?

- 1 The reaction at the positive electrode is: $\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}(\text{l})$.
- 2 The overall cell reaction is: $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$.
- 3 The cell potential is 1.23 V.

- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

[1]

- 12(a)** Vanadium ions have four common oxidation states. **Table 18.1** shows the colours of the ions in aqueous solution.

Oxidation state of vanadium	Vanadium ion	Colour
+5	$\text{VO}_2^+(\text{aq})$	yellow
+4	$\text{VO}^{2+}(\text{aq})$	blue
+3	$\text{V}^{3+}(\text{aq})$	green
+2	$\text{V}^{2+}(\text{aq})$	violet

Table 18.1

- (i) The student adds excess iron to a solution containing $\text{VO}^{2+}(\text{aq})$ ions, and observes that the colour of the solution changes from blue to green and then to violet.

Use the relevant standard electrode potentials shown in **Table 18.2** to explain these observations.

Redox system			E°/V
1	$\text{V}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{V}(\text{s})$		-1.18
2	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$		-0.44
3	$\text{V}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{V}^{2+}(\text{aq})$		-0.26
4	$\text{VO}^{2+}(\text{aq}) + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$		+0.34
5	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$		+0.77
6	$\text{VO}_2^+(\text{aq}) + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$		+1.00

Table 18.2

.....

.....

.....

.....

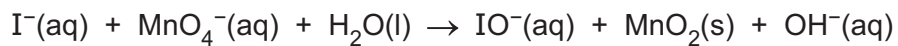
.....

..... **[3]**

- (ii) Construct an equation for the **first** colour change from blue to green.

..... **[1]**

- 13** Iodide ions, $\text{I}^{-}(\text{aq})$, react with $\text{MnO}_4^{-}(\text{aq})$. The unbalanced equation is shown below.



What is the ratio of $\text{MnO}_2(\text{s})$ to $\text{OH}^{-}(\text{aq})$ in the balanced equation?

A 1 : 3

B 1 : 2

C 1 : 1

D 3 : 2

Your answer

[1]

- 14** This question is about reactions and uses of the weak acids methanoic acid, HCOOH , and ethanoic acid, CH_3COOH .

- (a)** A student adds magnesium metal to an aqueous solution of ethanoic acid, CH_3COOH .
A redox reaction takes place.

Write the overall equation for this reaction and explain, in terms of oxidation numbers, which element has been oxidised and which element has been reduced.

Equation

Oxidation

.....

Reduction

.....

[3]

- 15 Standard electrode potentials for four redox systems are shown in **Table 19.1**.

Redox system	Half-equation	E°/V
1	$\text{CO}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{HCOOH}(\text{aq})$	-0.11
2	$\text{HCOOH}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{HCHO}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	-0.03
3	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
4	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 19.1

- (a) A student sets up a standard cell in the laboratory based on redox systems **3** and **4**.

Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential at 298 K.

[3]

- (b)** A student warms a mixture of methanal, HCHO , and acidified potassium manganate(VII).

The student observes gas bubbles.

Explain this observation in terms of electrode potentials and equilibria.

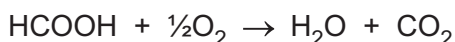
Include overall equations in your answer.

[4]

- (c) Methanoic acid, HCOOH , can be used in a fuel cell. As with all fuel cells, the fuel (HCOOH) is supplied at one electrode and the oxidant (oxygen) at the other electrode.

The standard cell potential for this fuel cell is 1.34 V.

The overall reaction is shown below.



Using the information in **Table 19.1**, deduce the half-equation for the reaction at the oxygen electrode, and calculate the standard electrode potential for the oxygen half-cell.

half-equation

standard electrode potential = ... V

[2]

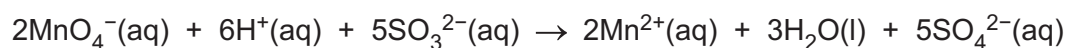
- 16(a)** Sodium sulfite(IV), Na_2SO_3 , is used as a preservative in some foods. Food safety legislation allows a maximum of 850 mg Na_2SO_3 per kg of burger meat.

A chemist determines the amount of Na_2SO_3 in a sample of burger meat using a manganate(VII) titration.

Step 1 The Na_2SO_3 from 525 g of burger meat is extracted to form a solution containing $\text{SO}_3^{2-}(\text{aq})$ ions.

Step 2 The solution from **step 1** is made up to 250.0 cm^3 in a volumetric flask with water. 25.0 cm^3 of this diluted solution is pipetted into a conical flask.

Step 3 The pipetted solution from **step 2** is acidified with dilute sulfuric acid and then titrated with $0.0100\text{ mol dm}^{-3}$ potassium manganate(VII), KMnO_4 .

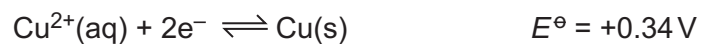
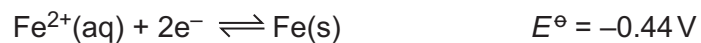


12.60 cm^3 of $\text{KMnO}_4(\text{aq})$ is required to reach the endpoint.

Analyse the results to determine whether the burger meat complies with food safety legislation.

[5]

17 Two redox systems are shown below.



Which species in the two redox systems is the strongest oxidising agent?

A $\text{Fe}^{2+}(\text{aq})$

B $\text{Fe}(\text{s})$

C $\text{Cu}^{2+}(\text{aq})$

D $\text{Cu}(\text{s})$

Your answer

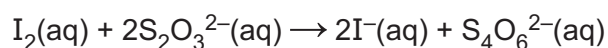
[1]

- 18 A student carries out an experiment to determine the percentage by mass of copper in an ore containing copper in its +2 oxidation state.

The student is provided with a sample of the copper ore, 1 mol dm^{-3} potassium iodide, KI(aq) , and $0.0200 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.

The student's method is outlined below.

- Step 1** Add an excess of warm nitric acid to 2.50 g of the ore.
The copper(II) compounds in the ore react, forming aqueous copper(II) nitrate.
- Step 2** Filter the mixture to remove the unreacted rock. Neutralise the filtrate.
- Step 3** Add an excess of aqueous potassium iodide, KI(aq) .
A precipitate of copper(I) iodide and a solution of iodine, $\text{I}_2(\text{aq})$, forms.
- Step 4** Titrate the mixture from **Step 3** using $0.0200 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$ in the burette.



26.55 cm^3 of $0.0200 \text{ mol dm}^{-3}$ $\text{Na}_2\text{S}_2\text{O}_3$ are required to reach the end point.

- (a) In **Step 1**, the student observed that bubbles of gas were produced.

Suggest the formula of the copper(II) compound which reacted with HNO_3 to form the gas, and write a full equation for the reaction.

Formula:

Equation: [2]

- (b) Write an **ionic** equation, including state symbols, for the reaction in **Step 3**.

..... [1]

- (c) Suggest a suitable indicator for this titration and state the colour change at the end point in **Step 4**.

Indicator:

Colour from to [1]

- (d) Determine the percentage, by mass, of copper in the copper ore.
Give your answer to an **appropriate** number of significant figures.

percentage = ... % [4]

- (e) Explain whether the calculated percentage by mass of copper would be higher, lower or the same if the following changes were made to the method.

- (i) The potassium iodide was not in excess, in **Step 3**.

.....
..... [1]

- (ii) The burette readings were read from the top of the meniscus, in **Step 4**.

.....
..... [1]

- (f) The student then modifies the method in order to obtain a more accurate value for the percentage by mass of copper in the ore. The student decides to use 25.00g of the copper ore in **Step 1**.

What further modifications should the student make to produce a more accurate value for the percentage by mass of copper in the ore?

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

- 19 Storage cells and fuels cells are types of electrochemical cell.

The electrode potentials for five redox systems are shown in **Table 19.1**.

Redox system	Half-equation	$E^{\ominus}/V$
1	$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-} \rightleftharpoons \text{Cr}(\text{s})$	-0.74
2	$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^{-} \rightleftharpoons 4\text{OH}^{-}(\text{aq})$	+0.40
3	$\text{MnO}_4^{-}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{MnO}_4^{2-}(\text{aq})$	+0.56
4	$\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{e}^{-} \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
5	$\text{MnO}_4^{2-}(\text{aq}) + 4\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.70

Table 19.1

- (a) A student sets up an electrochemical cell based on redox systems **1** and **4** in **Table 19.1**.

- (i) Draw a labelled diagram to show how this cell could be set up in the laboratory.

[3]

- (ii) Construct the equation for the overall cell reaction.

..... [1]

(b) In acid conditions, $\text{MnO}_4^{2-}(\text{aq})$ disproportionates to form $\text{MnO}_2(\text{s})$ and $\text{MnO}_4^{-}(\text{aq})$.

(i) Explain, in terms of oxidation numbers, why disproportionation has taken place.

.....

 [2]

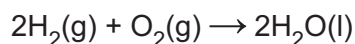
(ii) Explain, in terms of electrode potentials and equilibrium shifts why $\text{MnO}_4^{2-}(\text{aq})$ disproportionates in acid conditions. Use the information in **Table 19.1**.

.....

 [2]

(c) An **alkaline** hydrogen-oxygen fuel cell is set up.

The overall equation for the cell reaction is shown below.



Redox system **2** in **Table 19.1** is the positive electrode of this cell.

(i) Write the half-equation at the negative electrode.

.....
 [1]

(ii) The cell potential is 1.23 V.

Calculate the electrode potential of the negative electrode.

electrode potential = ... V [1]

(iii) State **one** important feature of a fuel cell that is different from a conventional storage cell.

.....
 [1]

20 Standard electrode potentials for two redox systems are shown below.



What is the standard electrode potential for $\text{Cr}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cr}(\text{s})$?

A -0.32 V

B -0.90 V

C -1.16 V

D -1.80 V

Your answer

[1]

- 21** A student adds magnesium to dilute hydrochloric acid in one test tube.
The student adds calcium to dilute hydrochloric acid in a second test tube.

A redox reaction takes place in each test tube.

- (a)** Write half-equations for the reaction of magnesium with hydrochloric acid.

Oxidation half-equation:

Reduction half-equation:

[2]

22

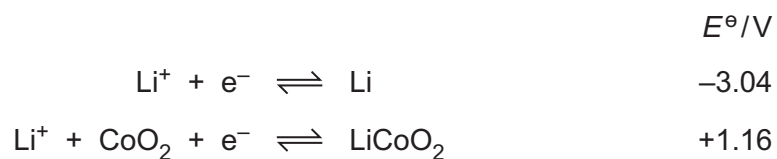
Aqueous sodium hydroxide is added to an aqueous solution of iron(II) sulfate. A pale green precipitate forms which turns brown when left to stand in air.

(a) Use the information below to explain why the pale green precipitate turns brown when left to stand in air and construct an equation for the reaction which occurs.

Redox System	Equation	E°/V
1	$\text{Fe}(\text{OH})_3(\text{s}) + \text{e}^- \rightleftharpoons \text{Fe}(\text{OH})_2(\text{s}) + \text{OH}^-(\text{aq})$	-0.56 V
2	$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightleftharpoons 4\text{OH}^-(\text{aq})$	$+0.40\text{ V}$

[4]

- 23 The equations show the electrode potentials of the half-cells used in a lithium-ion cell.



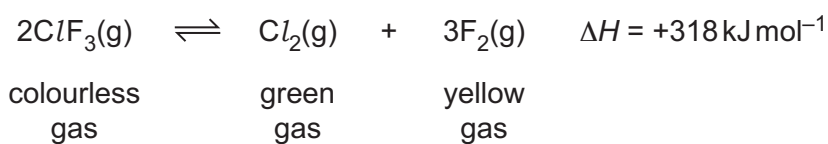
Which statement is correct in a lithium-ion cell?

- A** The cell potential is 2.88 V.
- B** The reaction at the positive electrode is: $\text{LiCoO}_2 \rightarrow \text{Li}^+ + \text{CoO}_2 + \text{e}^-$
- C** The overall cell reaction is: $\text{Li} + \text{CoO}_2 \rightarrow \text{LiCoO}_2$
- D** The oxidation number of Co changes from +2 to +1.

Your answer

[1]

- 24** Chlorine trifluoride can be decomposed into its elements forming the equilibrium mixture below.



Which statement(s) is/are correct?

- 1** The decomposition is a redox reaction.
 - 2** When the equilibrium mixture is cooled, the colour fades.
 - 3** The decomposition has a negative entropy change.
- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

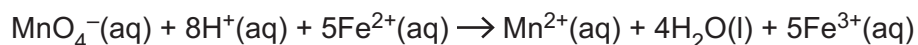
[1]

- 25 Some grass fertilisers contain compounds of iron.

During heavy rain, a fertiliser is washed into a nearby river causing the water to be polluted with a mixture of iron(II) and iron(III) ions.

- (a) A student determines the concentration of iron(II) ions in a sample of river water by titration with potassium manganate(VII).

25.0 cm³ portions of river water are acidified with dilute sulfuric acid. Each portion is titrated with 0.00250 mol dm⁻³ potassium manganate(VII) until a colour change is seen.



- (i) State the colour change seen at the end point of the titration.

from to

[1]

- (ii) The student's titration results are shown in the table below.
The trial titre has been omitted.

	1	2	3
Final volume / cm ³	12.65	25.60	38.35
Initial volume / cm ³	0.00	12.65	25.60
Titre volume / cm ³			

Complete the table above and calculate the mean titre that the student should use to determine the concentration of iron(II) ions in the river water.

mean titre = ... cm³ [2]

- (iii) Determine the concentration, in mol dm⁻³, of iron(II) ions in the river water.

concentration = ... mol dm⁻³ [3]

- (b) The student modifies the experiment in (a) to determine the combined concentration of iron(II) and iron(III) ions in the river water.

The student's method is shown below.

- Step 1** Add excess zinc to a 250.0 cm^3 sample of river water and warm gently.
- Step 2** Cool the solution and remove excess zinc by filtration.
- Step 3** Acidify 25.0 cm^3 portions of the filtrate from **Step 2**. Then titrate each portion with $0.00250\text{ mol dm}^{-3}$ potassium manganate(VII) until a colour change is seen.

The table below shows information about three redox systems.

Redox system	Half-equation	$E^\ominus/\text{V}$
1	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
2	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
3	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Use the information in the table above to explain the reasons for **Step 1** and **Step 2**.

Reason(s) for **Step 1**

.....

.....

.....

Reason(s) for **Step 2**

.....

.....

.....

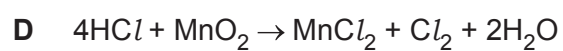
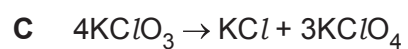
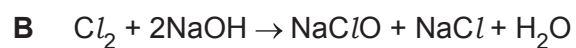
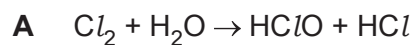
[4]

- 26(a)** An acidified solution containing $\text{Cr}_2\text{O}_7^{2-}$ ions reacts with vanadium(III) ions in a redox reaction to form a solution containing Cr^{3+} ions and VO_2^+ ions.

Construct the overall equation for this reaction.

..... [2]

27 Which equation does **not** represent a disproportionation reaction?



Your answer

[1]

28 Chlorine has the electron configuration $[\text{Ne}]3s^23p^5$.

Which statement(s) about chlorine is/are correct when it reacts in redox reactions?

- 1** It can gain one electron to form $1-$ ions.
- 2** It can lose its $3s^2$ electrons to form $2+$ ions.
- 3** It can lose its $3p^5$ electrons to form $5+$ ions.

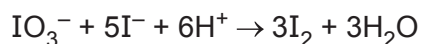
- A** 1, 2 and 3
- B** Only 1 and 2
- C** Only 2 and 3
- D** Only 1

Your answer

[1]

- 29(a)** Potassium iodate tablets prevent the uptake of radioactive iodine in the human body following a nuclear accident.

The mass of potassium iodate(V), KIO_3 , in a tablet can be determined by reaction with an aqueous solution of potassium iodide, KI , in the presence of acid.



A chemist finds that two KIO_3 tablets react with exactly 26.2 cm^3 0.150 mol dm^{-3} KI .

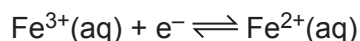
Calculate the mass, in mg, of KIO_3 in **one** tablet.

Give your answer to the nearest whole number.

mass KIO_3 = ... mg **[4]**

- (b)** Standard electrode potentials are measured by comparison with a reference half-cell.

Draw a labelled diagram to show how the standard electrode potential could be measured for the redox system below.



Include details of the apparatus, solutions and the standard conditions needed when measuring this standard electrode potential.

Standard conditions

..... **[4]**

(c) Many electric vehicles are powered by lithium-ion cells.

Hydrogen-oxygen fuel cells can also be used to power vehicles.

Six redox systems are shown in the table. State symbols have been omitted.

Redox system	Half-equation	E°/V
1	$Li^{+} + e^{-} \rightleftharpoons Li$	-3.04
2	$2H_2O + 2e^{-} \rightleftharpoons H_2 + 2OH^{-}$	-0.83
3	$2H^{+} + 2e^{-} \rightleftharpoons H_2$	0.00
4	$O_2 + 2H_2O + 4e^{-} \rightleftharpoons 4OH^{-}$	+0.40
5	$Li^{+} + CoO_2 + e^{-} \rightleftharpoons LiCoO_2$	+1.16
6	$O_2 + 4H^{+} + 4e^{-} \rightleftharpoons 2H_2O$	+1.23

(i) A lithium-ion cell involves **redox systems 1 and 5**.

Construct the overall cell equation for a lithium-ion cell.

..... [1]

(ii) Hydrogen-oxygen fuel cells can operate in acidic or in alkaline conditions.

Show that for acidic and alkaline hydrogen–oxygen fuel cells, the standard cell potentials, and the overall cell equations, are the same.

Acidic

.....

.....

.....

Alkaline.....

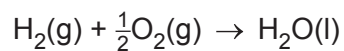
.....

.....

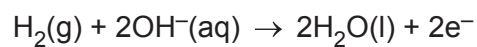
.....

[3]

- 30** The overall equation for a hydrogen–oxygen fuel cell is:



The half-equation at the negative electrode is:



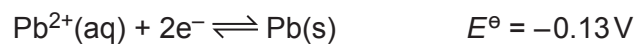
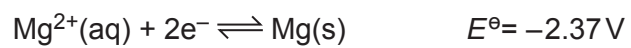
What is the half-equation at the positive electrode?

- A** $\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + 2\text{e}^-$
- B** $\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{OH}^-(\text{aq}) + 2\text{e}^-$
- C** $\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow 2\text{OH}^-(\text{aq})$
- D** $\text{O}_2(\text{g}) + 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$

Your answer

[1]

31 Two redox systems are shown below.



Which species is the strongest reducing agent?

A $\text{Mg}^{2+}(\text{aq})$

B $\text{Mg}(\text{s})$

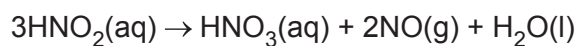
C $\text{Pb}^{2+}(\text{aq})$

D $\text{Pb}(\text{s})$

Your answer

[1]

32(a) When heated, aqueous HNO_2 disproportionates. The equation is shown below.



- What is meant by **disproportionation**?
- Use oxidation numbers to show that disproportionation has taken place.

.....

.....

.....

.....

.....

.....

.....

.....

.....

..... **[3]**

- 33 Standard electrode potentials for seven redox systems are shown in the table below.

Redox system	Half-equation	E°/V
1	$\text{V}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{V}(\text{s})$	-1.18
2	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
3	$\text{Co}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Co}(\text{s})$	-0.28
4	$\text{V}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{V}^{2+}(\text{aq})$	-0.26
5	$2\text{H}^+(\text{aq}) + \text{VO}^{2+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.34
6	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
7	$2\text{H}^+(\text{aq}) + \text{VO}_2^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+1.00

(a)

- (i) A student sets up a standard cell based on redox systems **2** and **4**.

Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential.

Include the conditions required to measure the standard cell potential.

standard conditions

 [4]

- (ii) Calculate the standard cell potential of this cell.

standard cell potential = ... V [1]

(b) A student predicts that cobalt metal will reduce acidified VO^{2+} ions to form V^{2+} ions in two stages.

Explain why this two-stage reduction should happen in terms of electrode potentials and equilibria shifts.

Include overall equations for the predicted reactions.

[4]

34(a) Mixture A is a mixture of copper(II) sulfate crystals, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ and iron(II) sulfate crystals, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.

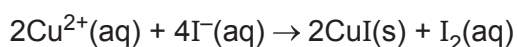
A student carries out a titration to determine the percentage by mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in **Mixture A**.

The student uses the following method.

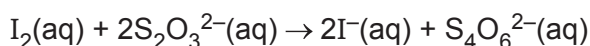
Step 1 Dissolve 9.51 g of **Mixture A** in water and make up the solution to 250.0 cm^3 in a volumetric flask.

Step 2 Pipette 25.0 cm^3 of the solution from **Step 1** into a conical flask and add an excess of aqueous potassium iodide, $\text{KI}(\text{aq})$.

The copper(II) ions react with iodide ions to form a precipitate of copper(I) iodide and a solution of iodine.



Step 3 Titrate the contents of the conical flask from **Step 2** with $0.0800\text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, using starch as the indicator.



Step 4 Repeat **Step 2** and **Step 3** to obtain concordant titres and calculate the mean titre.

Results

Mean titre = 31.15 cm^3

(i) Calculate the percentage by mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in **Mixture A**.

Give your answer to **3** significant figures.

Percentage of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in **Mixture A** = ... % **[5]**

- (ii) The student repeats their method, but after carrying out **Step 1** they leave the solution exposed to the air for a few days.
Some of the iron(II) sulfate is oxidised to iron(III) sulfate.

The student then continues with **Step 2**, **Step 3** and **Step 4** of the method.

This mean titre is greater than 31.15 cm^3 .

Suggest why this mean titre is greater than 31.15 cm^3 .

.....

.....

.....

..... [1]