

1 Redox reactions can be used to generate electrical energy from electrochemical cells.

(a) A student carries out an investigation based on the redox systems shown in **Table 5.1** below.

	redox system	$E^\ominus/V$
1	$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	-0.25
2	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
3	$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	-0.74

**Table 5.1**

The student sets up two standard cells to measure two standard cell potentials.

- **Cell A** is based on redox systems **1** and **2**.
- **Cell B** is based on redox systems **1** and **3**.

(i) Draw a labelled diagram to show how the student could have set up **Cell A**, based on redox systems **1** and **2**, to measure the standard cell potential.

[3]

- (ii) For each standard cell below,
- what would be the standard cell potential?
  - what would be the sign of the Ni electrode?

**Cell A** based on redox system **1** and **2**:

standard cell potential = ..... V

sign of Ni electrode, + or - = .....

**Cell B** based on redox system **1** and **3**:

standard cell potential = ..... V

sign of Ni electrode, + or - = .....

[2]

(b) The student left each cell in (a) connected for a length of time.

For each cell, the student weighed the nickel electrode before connecting the cell and after the cell had been disconnected.

The student made the following observations.

- In **Cell A**, the nickel electrode lost mass.
- In **Cell B**, the nickel electrode gained mass.
- In **both** cells, the measured cell potential slowly changed.

Explain these observations. Include equations in your answer.

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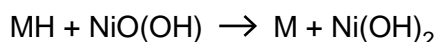
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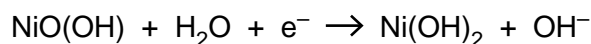
..... [3]

(c) Nickel metal hydride cells (NiMH cells) are being developed for possible use in cars. In a NiMH cell, an alloy is used to absorb hydrogen as a metal hydride. For simplicity, the alloy can be represented as M and the metal hydride as MH.

The overall cell reaction in a NiMH cell is shown below.



The half-equation at one electrode is shown below.



(i) Deduce the half-equation at the other electrode.

..... [1]

(ii) State a method, other than absorption, that is being developed to store hydrogen for possible use as a fuel in cars.

..... [1]

[Total: 10]

- 2 Haematite is the main ore of iron. The percentage of iron in a sample of haematite can be determined using the method below.

**Method**

**Stage 1.** An excess of concentrated hydrochloric acid is added to a 3.25 g sample of haematite. The iron(III) oxide in the haematite reacts to form a solution containing  $\text{Fe}^{3+}$  ions.

**Stage 2.** An excess of aqueous tin(II) chloride is added.  $\text{Sn}^{2+}$  reduces the  $\text{Fe}^{3+}$  present to  $\text{Fe}^{2+}$ . Excess  $\text{Sn}^{2+}$  is removed.

**Stage 3.** The solution is diluted and made up to  $250.0\text{ cm}^3$  in a volumetric flask.

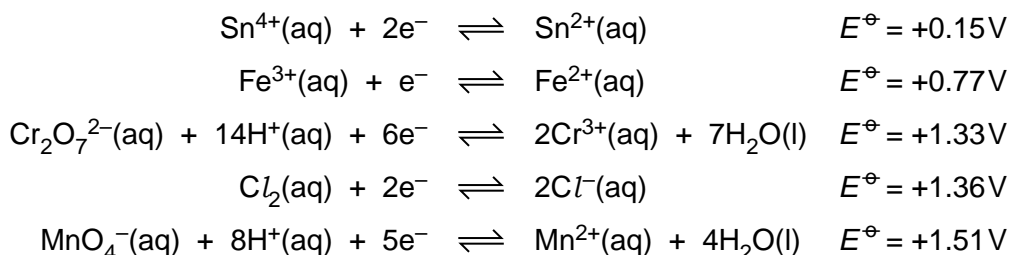
**Stage 4.** A  $25.0\text{ cm}^3$  sample of this solution is pipetted into a conical flask.

**Stage 5.** The solution in the conical flask is titrated with  $0.0200\text{ mol dm}^{-3}$  aqueous potassium dichromate(VI),  $\text{K}_2\text{Cr}_2\text{O}_7$ . The  $\text{Fe}^{2+}$  ions are oxidised to  $\text{Fe}^{3+}$  ions.

**Stage 6.** Stages 4 and 5 are repeated to obtain an average titre of  $26.5\text{ cm}^3$ .

You are provided with the following electrode potentials.

You may need to use this information throughout this question.



- (a) Write an equation for the reaction between iron(III) oxide and concentrated hydrochloric acid, occurring in **Stage 1**.

..... [1]

- (b) Write equations for the reactions involving iron ions in **Stages 2 and 5**.

**Stage 2** .....

**Stage 5** ..... [2]

(c) Calculate the percentage by mass of iron in the haematite ore.

percentage iron = ..... % [5]

(d) Aqueous potassium manganate(VII),  $\text{KMnO}_4(\text{aq})$ , is **not** suitable for titrating the solution in this method. Aqueous potassium dichromate(VI),  $\text{K}_2\text{Cr}_2\text{O}_7(\text{aq})$ , is used instead.

Suggest and explain why potassium dichromate(VI),  $\text{K}_2\text{Cr}_2\text{O}_7$ , is suitable for this titration whereas potassium manganate(VII),  $\text{KMnO}_4$ , is not suitable.

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

[Total: 10]

3 Redox reactions are used to generate electrical energy from electrochemical cells.

(a) **Table 4.1** shows three redox systems, and their standard redox potentials.

redox system	$E^\ominus/V$
$\text{Cu}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.52
$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Cr}(\text{s})$	-0.74
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}(\text{aq})$	+0.15

**Table 4.1**

(i) Draw a labelled diagram to show how the standard electrode potential of a  $\text{Sn}^{4+}/\text{Sn}^{2+}$  redox system could be measured.

(ii) Using the information in **Table 4.1**, write equations for the reactions that are feasible. [3]  
Suggest **two** reasons why these reactions may **not** actually take place.

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- (b) Modern fuel cells are being developed as an alternative to the direct use of fossil fuels. The 'fuel' can be hydrogen but many other substances are being considered. In a methanol fuel cell, the overall reaction is the combustion of methanol.

As with all fuel cells, the fuel (methanol) is supplied at one electrode and the oxidant (oxygen) at the other electrode.

Oxygen reacts at the negative electrode of a methanol fuel cell:



- (i) Write an equation for the complete combustion of methanol.

..... [1]

- (ii) Deduce the half-equation for the reaction that takes place at the positive electrode in a methanol fuel cell.

..... [1]

- (iii) State **two** advantages of vehicles using fuel cells compared with the combustion of conventional fossil fuels.

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

- (iv) Suggest **one** advantage of using methanol, rather than hydrogen, in a fuel cell for vehicles. Justify your answer.

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

[Total: 13]

4 Iron(II) iodide,  $\text{FeI}_2$ , is formed when iron metal reacts with iodine.

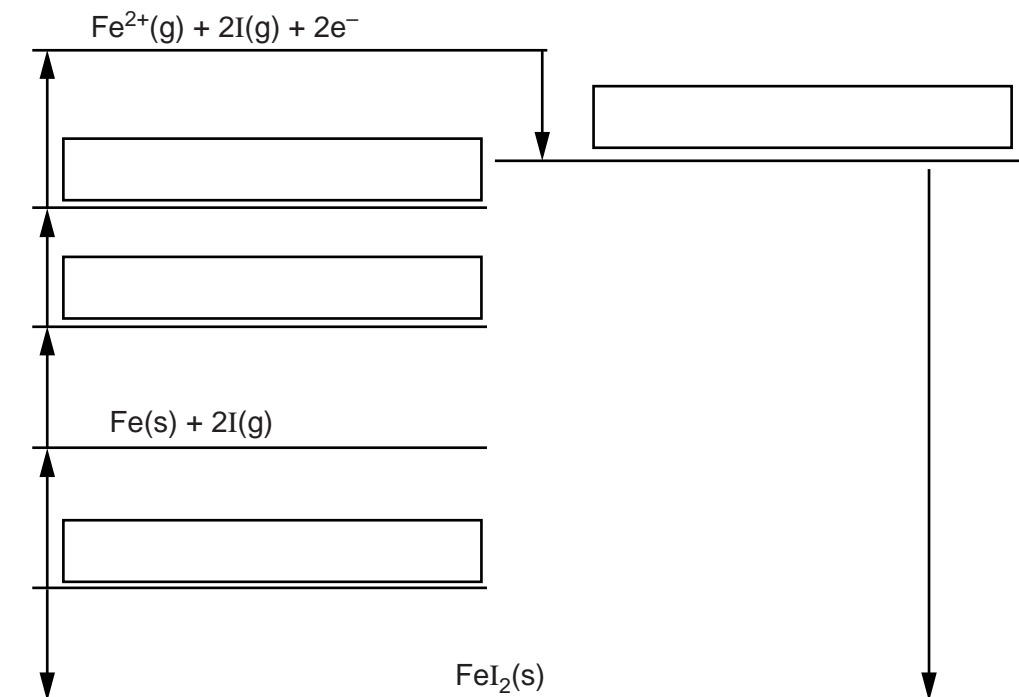
(a) The table below shows enthalpy changes involving iron, iodine and iron(II) iodide.

	Enthalpy change / $\text{kJ mol}^{-1}$
Formation of iron(II) iodide	-113
1st electron affinity of iodine	-295
1st ionisation energy of iron	+759
2nd ionisation energy of iron	+1561
Atomisation of iodine	+107
Atomisation of iron	+416

(i) The incomplete Born–Haber cycle below can be used to determine the lattice enthalpy of iron(II) iodide.

In the boxes, write the species present at each stage in the cycle.

Include state symbols for the species.



[4]

**(ii)** Define the term *lattice enthalpy*.

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

**(iii)** Calculate the lattice enthalpy of iron(II) iodide.

lattice enthalpy = .....  $\text{kJ mol}^{-1}$  [2]





(c)  $\text{Fe}^{2+}$  ions can be used to test for  $\text{NO}_3^-$  ions.

In this test, aqueous iron(II) sulfate is added to a solution containing  $\text{NO}_3^-$  ions, followed by slow addition of concentrated sulfuric acid. The sulfuric acid forms a layer below the aqueous solution.

In the presence of  $\text{NO}_3^-$  ions, a brown ring forms between the two layers.

Two reactions take place.

Reaction 1: In the acid conditions  $\text{Fe}^{2+}$  ions reduce  $\text{NO}_3^-$  ions to NO.

$\text{Fe}^{2+}$  ions are oxidised to  $\text{Fe}^{3+}$  ions.

Water also forms.

Reaction 2: A ligand substitution reaction of  $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$  takes place in which one NO ligand exchanges with one water ligand. A deep brown complex ion forms as the brown ring.

Construct equations for these two reactions.

Reaction 1:

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Reaction 2:

..... [3]

[Total: 16]