1 Standard electrode potentials for seven redox systems are shown in **Table 7.1**.
You may need to use this information in parts **(a)–(d)** of this question.

Redox system				$E^{\ominus}$ / V
1	$Mg^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Mg(s)$	–2.37
2	$Cu^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Cu(s)$	+0.34
3	$Al^{3+}(aq) + 3e^-$	$\rightleftharpoons$	$Al(s)$	–1.66
4	$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$	+0.77
5	$I_2(aq) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$	+0.54
6	$Cl_2(g) + 2e^-$	$\rightleftharpoons$	$2Cl^-(aq)$	+1.36
7	$ClO^-(aq) + 2H^+(aq) + e^-$	$\rightleftharpoons$	$\frac{1}{2}Cl_2(g) + H_2O(l)$	+

Table 7.1

(a) Define the term *standard electrode potential*.
Include all standard conditions in your answer.

....

....

....

....

.... **[2]**

(b) An electrochemical cell can be made based on redox systems **1** and **2**.

Write down the standard cell potential of this cell.

standard cell potential = ... V **[1]**

(c) Using redox systems **3**, **4** and **5 only** in **Table 7.1**, predict **three** reactions that might be feasible.

(i) Write the overall equation for each predicted reaction.

....

....

.... **[3]**

(ii) Give **two** reasons why it is uncertain whether reactions predicted from $E^\ominus$ values may actually take place.

....

....

....

.... **[2]**

(d) In aqueous acid, Cl^-(aq) ions react with ClO^-(aq) ions to form chlorine gas, Cl_2(g).
In aqueous alkali, chlorine gas, Cl_2(g), reacts to form Cl^-(aq) and ClO^-(aq) ions.

Explain this difference.
Use **Table 7.1** to help you with your answer.

....

....

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.... **[4]**

(e) In acidic conditions, Sn^{2+} ions react with IO_3^- ions to produce iodine and Sn^{4+} ions.

(i) What is the oxidising agent in this reaction?
Explain your answer.

....

....

.... **[1]**

(ii) Construct an equation for this reaction.

.... **[2]**

[Total: 15]

Turn over

2 Standard electrode potentials for eight redox systems are shown in **Table 6.1**.

You will need to use this information throughout this question.

redox system	**half-equation**	$E^{\ominus}$/V
1	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
2	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
3	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
4	$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
5	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
6	$CO_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons HCOOH(aq)$	–0.22
7	$HCOOH(aq) + 2H^+(aq) + 2e^- \rightleftharpoons HCHO(aq) + H_2O(l)$	+0.06
8	$Cr^{3+}(aq) + 3e^- \rightleftharpoons Cr(s)$	–0.74

Table 6.1

(a) A student sets up a standard cell in the laboratory based on redox systems **2** and **8**. His circuit allows him to measure the standard cell potential.

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

[3]

(ii) Write down the overall cell reaction.

.... **[1]**

(iii) Write down the standard cell potential.

standard cell potential = ... V **[1]**

(b) Select from **Table 6.1**, the strongest oxidising agent.

.... **[1]**

(c) Using the redox systems in **Table 6.1**, construct an equation for a reaction between acidified dichromate(VI) ions and methanoic acid, HCOOH.

Rather than using [O] or [H], your equation must show the actual reactants and products.

[2]

(d) A student added some chromium metal to an acidified solution containing copper(II) ions. A reaction took place. The student concluded that 'chromium is more reactive than copper'.

(i) Explain, in terms of their electrode potentials, why 'chromium is more reactive than copper' in this reaction.

....

....

....

.... **[2]**

(ii) When this experiment was carried out, the student observed some bubbles of a gas.

Suggest an explanation for this observation.

....

.... **[1]**

(e) Methanoic acid, HCOOH, has the common name of 'formic acid'. Direct-Formic Acid Fuel Cells (DFAFCs) are being developed for use in small, portable electronics such as phones and laptop computers.

In this fuel cell, methanoic acid (the fuel) reacts with oxygen to generate a cell potential.

(i) Predict the standard cell potential of a DFAFC.

standard cell potential = ... V **[1]**

(ii) Suggest **two** advantages of using methanoic acid as the fuel in a fuel cell rather than hydrogen.

....

.... **[2]**

[Total: 14]

3 Nickel–cadmium cells (NiCd cells) have been extensively used as rechargeable storage cells. NiCd cells have been a popular choice for many electrical and electronic applications because they are very durable, reliable, easy-to-use and economical.

The electrolyte in NiCd cells is aqueous KOH. The standard electrode potentials for the redox systems that take place in NiCd cells are shown below.

$$Cd(OH)_2 + 2e^- \rightleftharpoons Cd + 2OH^- \qquad E^\ominus = -0.80\,V$$
$$NiO(OH) + H_2O + e^- \rightleftharpoons Ni(OH)_2 + OH^- \qquad E^\ominus = +0.45\,V$$

(a) Define the term *standard electrode potential*, including all standard conditions in your answer.

....

....

....

....

.... **[2]**

(b) What is the standard cell potential of a NiCd cell?

answer = ...V **[1]**

(c) When a NiCd cell is being used for electrical energy, it is being discharged.

(i) Construct the overall cell reaction that takes place during discharge of a NiCd cell.

....

....

.... **[2]**

(ii) Using oxidation numbers, show the species that have been oxidised and reduced during discharge of a NiCd cell.

oxidation ....

....

reduction ....

.... **[2]**

(d) NiCd cells are recharged using a battery charger.

(i) Suggest the reactions that take place in the NiCd cell during the recharging process.

....

.... **[1]**

(ii) As the cell approaches full charge, the aqueous KOH electrolyte starts to decompose, forming hydrogen gas at one electrode and oxygen gas at the other electrode.

Predict half-equations that might take place at each electrode for the decomposition of the electrolyte to form hydrogen and oxygen.

....

.... **[2]**

[Total: 10]

4 Electrochemical cells have been developed as a convenient and portable source of energy.

The essential components of any electrochemical cell are two redox systems, one providing electrons and the other accepting electrons. The tendency to lose or gain electrons can be quantified using values called standard electrode potentials.

Standard electrode potentials for seven redox systems are shown in **Table 4.1**.
You may need to use this information throughout this question.

Table 4.1

redox system	**equation**	$E^{\ominus}$**/V**
1	$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0
2	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
3	$SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + H_2O(l)$	+0.17
4	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.34
5	$Cl_2(aq) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
6	$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
7	$I_2(aq) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54

(a) An electrochemical cell can be made based on redox systems 2 and 4.

(i) Draw a labelled diagram to show how this cell can be set up in the laboratory. **[3]**

(ii) State the charge carriers that transfer current

through the wire, ...

through the solution. .... **[1]**

(iii) Write down the overall cell reaction.

.... **[1]**

(iv) Write down the cell potential.

cell potential .... V **[1]**

(b) Select from **Table 4.1**,

(i) a species which oxidises $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$,

.... **[1]**

(ii) a species which reduces $Fe^{3+}(aq)$ to $Fe^{2+}(aq)$ but does **not** reduce $Ag^{+}(aq)$ to Ag(s).

.... **[1]**

(c) Fuel cells are a type of electrochemical cell being developed as a potential source of energy in the future.

- State **one** important difference between a fuel cell and a conventional electrochemical cell.
- Write the equation for the overall reaction that takes place in a hydrogen fuel cell.
- State **two** ways that hydrogen might be stored as a fuel for cars.
- Suggest why some people consider that the use of hydrogen as a fuel for cars consumes more energy than using fossil fuels such as petrol and diesel.

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.... **[5]**

[Total: 13]

5 Redox titrations using $KMnO_4$ in acidic conditions can be used to analyse reducing agents. Acidified $KMnO_4$ is a strong oxidising agent, readily removing electrons:

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

A student analysed a solution of hydrogen peroxide, H_2O_2(aq), using a redox titration with $KMnO_4$ under acidic conditions. Under these conditions, H_2O_2 is a reducing agent.

The overall equation for the reaction is given below.

$$2MnO_4^- + 6H^+ + 5H_2O_2 \longrightarrow 2Mn^{2+} + 8H_2O + 5O_2$$

(a) Deduce the simplest whole number half-equation for the oxidation of H_2O_2 under these conditions.

[2]

(b) The student diluted 25.0 cm^3 of a solution of hydrogen peroxide with water and made the solution up to 250.0 cm^3. The student titrated 25.0 cm^3 of this solution with 0.0200 mol dm^{-3} $KMnO_4$ under acidic conditions. The volume of $KMnO_4$(aq) required to reach the end-point was 23.45 cm^3.

- Calculate the concentration, in g dm^{-3}, of the **undiluted** hydrogen peroxide solution.
- What volume of oxygen gas, measured at RTP, would be produced during this titration?

[6]

[Total: 8]