

Electrode Potentials & Redox

AS & A Level

Question Paper 1

Level	A Level
Subject	Chemistry
Exam Board	OCR
Module	Physical Chemistry & Transition Elements
Topic	Electrode Potentials & Redox
Paper	AS & A Level
Booklet	Question Paper 1

Time allowed: 78 minutes

Score: /58

Percentage: /100

Grade Boundaries:

A*	A	B	C	D	E
>85%	73%	60%	47%	34%	21%

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. [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]

- (i) 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. [1]

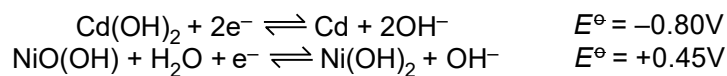
- (i) Suggest **two** advantages of using methanoic acid as the fuel in a fuel cell rather than hydrogen. [2]

[Total 14 Marks]

Question 2

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.



- (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? [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. [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 Marks]

Question 3

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	$Ni^{2+}(aq) + 2e^{-} \rightleftharpoons Ni(s)$	-0.25
2	$Fe^{3+}(aq) + e^{-} \rightleftharpoons Fe^{2+}(aq)$	+0.77
3	$Cr^{3+}(aq) + 3e^{-} \rightleftharpoons Cr(s)$	-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**:

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

[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.

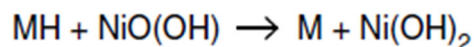
The student made the following observations.

- 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.

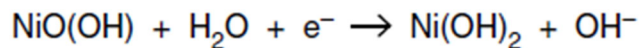
Explain these observations. Include equations in your answer. [3]

- (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]

- (i) State a method, other than absorption, that is being developed to store hydrogen for possible use as a fuel in cars. [1]

[Total 10 Marks]

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 Fe^{3+} ions.

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

Stage 3. The solution is diluted and made up to 250.0 cm^3 in a volumetric flask.

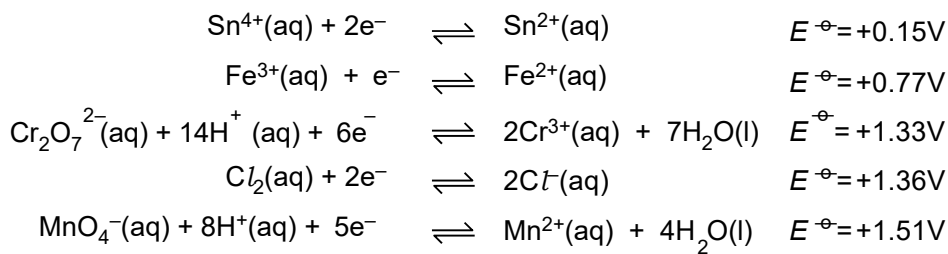
Stage 4. A 25.0 cm^3 sample of this solution is pipetted into a conical flask.

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

Stage 6. Stages 4 and 5 are repeated to obtain an average titre of 26.5 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**. [2]

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

[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), KMnO_4 , is not suitable.

[2]

[Total 10 Marks]

Question 5

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

A student investigates the redox systems shown in **Table 6.1** below.

	Redox system	E^{\ominus} /V
1	$Al^{3+}(aq) + 3e^{-} \rightleftharpoons Al(s)$	-1.66
2	$Cr^{3+}(aq) + e^{-} \rightleftharpoons Cr^{2+}(aq)$	-0.41
3	$V^{3+}(aq) + e^{-} \rightleftharpoons V^{2+}(aq)$	-0.26
4	$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$	+0.34
5	$Cr_2O_7^{2-}(aq) + 14H^{+}(aq) + 6e^{-} \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
6	$FeO_4^{2-}(aq) + 8H^{+}(aq) + 3e^{-} \rightleftharpoons Fe^{3+}(aq) + 4H_2O(l)$	+2.20

Table 6.1

(a) Define the term *standard electrode potential*.

Include all standard conditions in your answer.

[2]

(b) The student sets up a standard cell based on redox systems **3** and **4**.

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

On your diagram,

- show the charge carriers in the circuit joining the two half cells
- label the signs of the electrodes.

[5]

(ii) What is the standard cell potential of this cell?

[1]

(c) Explain what is meant by the terms *oxidising agent* and *reducing agent*.

Illustrate your answer by predicting reactions from redox systems **1–6** in which $\text{Cr}^{3+}(\text{aq})$ behaves as an oxidising agent and as a reducing agent.

Include overall equations and explain why you have made your predictions.

State symbols are **not** required.

[6]

[Total: 14 Marks]