

F325: Equilibria, Energetics and Elements

5.2.3 Electrode Potentials and Fuel Cells

59 marks

1. Use the standard electrode potentials in the table below to answer the questions that follow.

I	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Fe}(\text{s})$	$E^{\ominus} = -0.44 \text{ V}$
II	$\text{V}^{3+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{V}^{2+}(\text{aq})$	$E^{\ominus} = -0.26 \text{ V}$
III	$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{H}_2(\text{g})$	$E^{\ominus} = 0.00 \text{ V}$
IV	$\text{O}_2(\text{g}) + 4\text{H}^{+}(\text{aq}) + 4\text{e}^{-} \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +0.40 \text{ V}$

An electrochemical cell was set up based on systems I and II.

- (i) Write half-equations to show what has been oxidised and what has been reduced in this cell.

oxidation:

reduction:

[2]

- (ii) Determine the cell potential of this cell.

$$E_{\text{cell}} = \dots\dots\dots \text{ V}$$

[1]

[Total 3 marks]

2. Use the standard electrode potentials in the table below to answer the questions that follow.

I	$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Fe}(\text{s})$	$E^{\circ} = -0.44 \text{ V}$
II	$\text{V}^{3+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{V}^{2+}(\text{aq})$	$E^{\circ} = -0.26 \text{ V}$
III	$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{H}_2(\text{g})$	$E^{\circ} = 0.00 \text{ V}$
IV	$\text{O}_2(\text{g}) + 4\text{H}^{+}(\text{aq}) + 4\text{e}^{-} \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	$E^{\circ} = +0.40 \text{ V}$

An electrochemical fuel cell was set up based on systems **III** and **IV**.

- (i) Construct an equation for the spontaneous cell reaction. Show your working.

[2]

- (ii) Fuels cells based on systems such as **III** and **IV** are increasingly being used to generate energy.

Discuss **two** advantages and **two** disadvantages of using fuels cells for energy rather than using fossil fuels.

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[4]

[Total 6 marks]

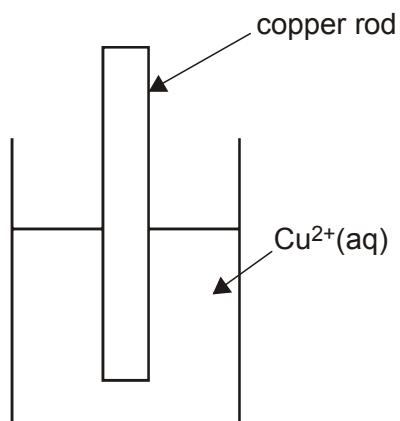
3. The standard electrode potential of $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$ is +0.34 V.

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

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[3]

(b) Complete the diagram to show how the standard electrode potential of $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$ could be measured.



[3]

[Total 6 marks]

4. Chromium is an important metallic element. Its compounds have a number of different oxidation states.

(a) (i) Give one use of chromium metal and state the property of chromium that makes it suitable for this use.

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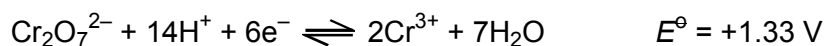
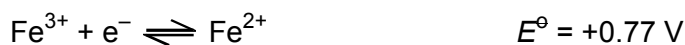
[1]

(ii) Complete the electronic configuration of a chromium atom.

$1s^2 2s^2 2p^6$

[1]

(b) The following equations relate to half-cells involving iron and chromium ions.



A cell was set up by combining these two half-cells.

(i) Derive a balanced equation for the reaction that would occur when the cell is in use.
Explain your reasoning in terms of oxidation and reduction.

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[3]

(ii) Determine the emf of the cell under standard conditions.

emf = V

[1]

[Total 6 marks]

5. The standard electrode potentials for some redox systems involving vanadium are shown below.

These are labelled **A**, **B**, **C** and **D**.

	E^\ominus / V
A $\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{VO}^{2+} + \text{H}_2\text{O}$	+1.00
B $\text{V}^{3+} + \text{e}^- \rightleftharpoons \text{V}^{2+}$	-0.26
C $\text{V}^{2+} + 2\text{e}^- \rightleftharpoons \text{V}$	-1.20
D $\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{V}^{3+} + \text{H}_2\text{O}$	+0.34

(a) Which of the vanadium species shown in **A**, **B**, **C** and **D** is the most powerful oxidising agent?

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[1]

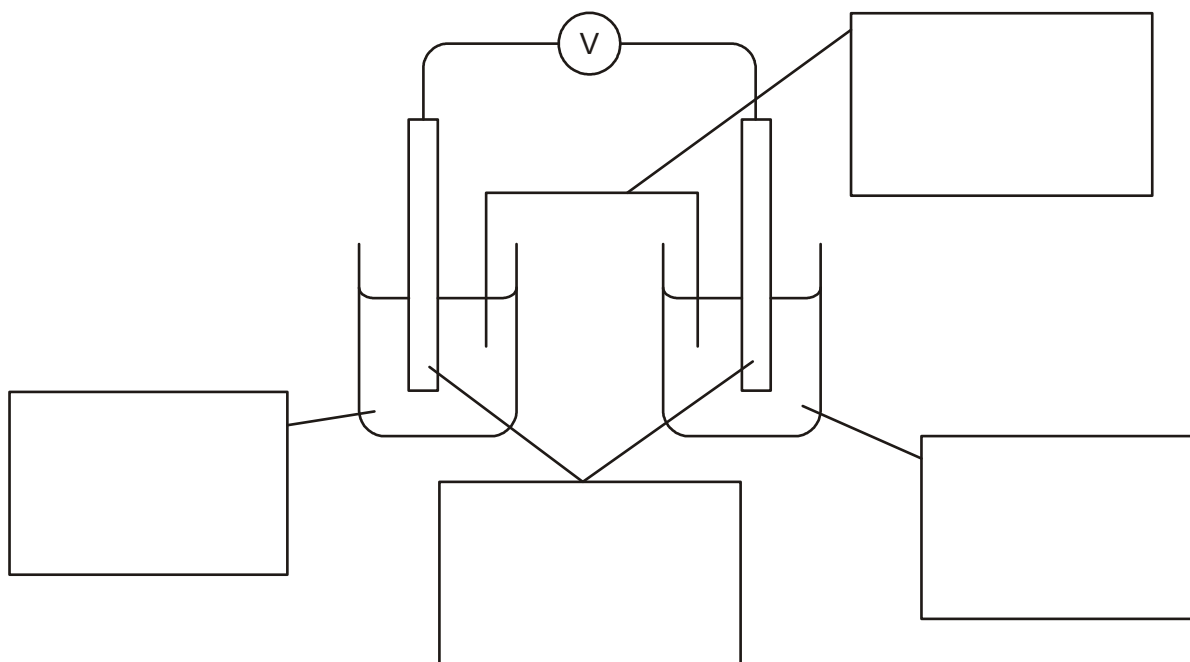
(b) A student wishes to set up a cell with a standard cell potential of 0.60V.

(i) Which two of the redox systems, **A**, **B**, **C** or **D**, should he choose?

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[1]

- (ii) Complete the labelling of the following diagram which shows the cell with a standard cell potential of 0.60V.



[4]

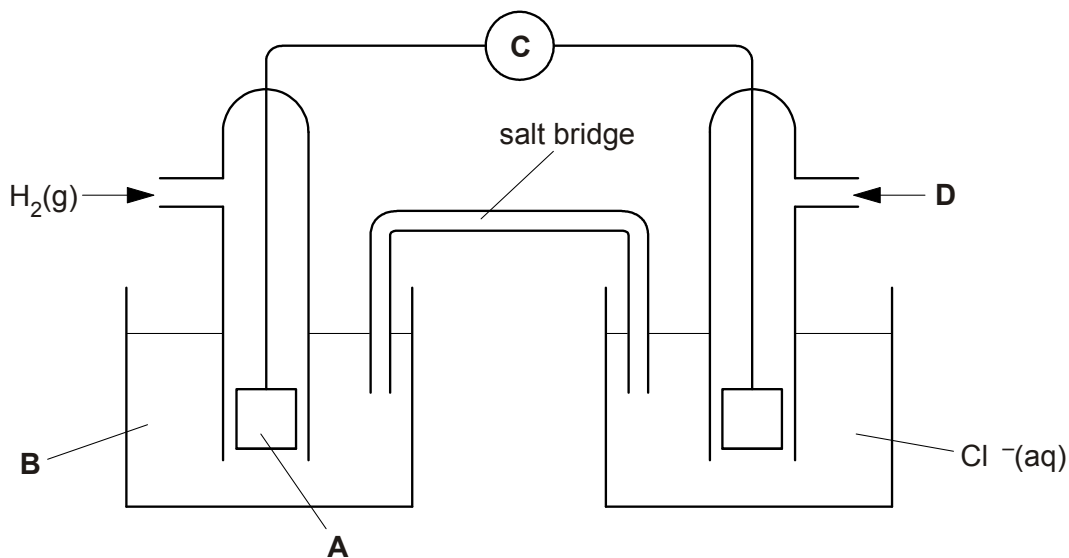
- (iii) The emf of this cell is only 0.60 V under standard conditions. What do you understand by the expression *standard conditions*?

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[1]

[Total 7 marks]

6. The standard electrode potential of the $\frac{1}{2} \text{Cl}_2 / \text{Cl}^-$ half-cell may be measured using the following apparatus.



- (a) Suggest suitable labels for **A**, **B**, **C** and **D**.

A

B

C

D

[2]

- (b) The half cell reactions involved are shown below.



- (i) Use an arrow to show the direction of flow of electrons in the diagram of the apparatus. Explain your answer.

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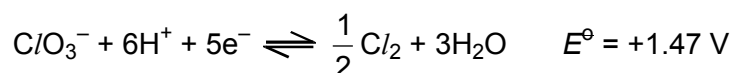
[2]

- (ii) The values of E^\ominus are measured under standard conditions. What are the standard conditions?

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[2]

- (c) The half cell reaction for $\text{ClO}_3^- / \frac{1}{2} \text{Cl}_2$ is shown below.



What does this tell you about the oxidising ability of ClO_3^- compared with Cl_2 ?

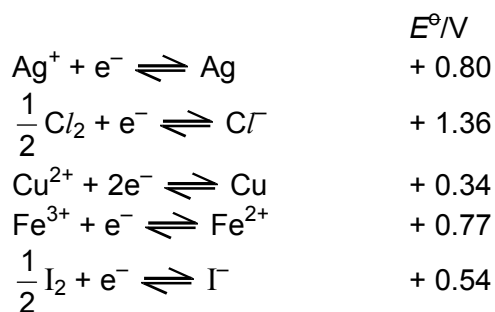
Explain your answer.

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[2]

[Total 8 marks]

7. Some standard electrode potentials are shown below.

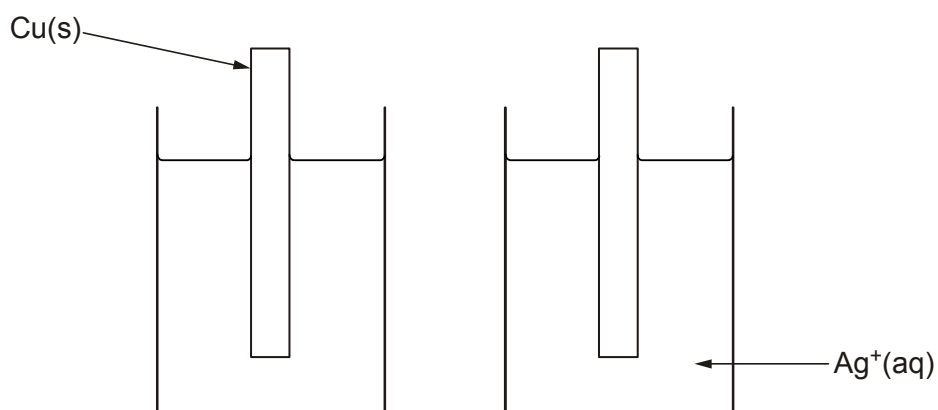


(a) Define the term *standard electrode potential*.

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[3]

(b) The diagram below shows an incomplete cell consisting of Cu/Cu^{2+} and Ag/Ag^+ half-cells.



(i) Complete and label the diagram to show how the cell potential of this cell could be measured.

[2]

(ii) On the diagram, show the direction of **electron** flow in the circuit if a current was allowed.

[1]

(iii) Calculate the standard cell potential.

standard cell potential =V

[1]

(iv) Write the overall cell reaction.

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[1]

(c) Chlorine will oxidise Fe^{2+} to Fe^{3+} but iodine will not. Explain why, using the electrode potential data.

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[2]

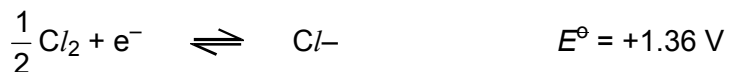
[Total 10 marks]

8. NO_2 reacts with oxygen and water to form nitric acid, HNO_3 . In the atmosphere, this contributes to acid rain. Construct a balanced equation for this formation of nitric acid and use oxidation numbers to show that this is a redox reaction.

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[Total 2 marks]

9. Chlorine gas may be prepared in the laboratory by reacting hydrochloric acid with potassium manganate(VII). The following standard electrode potentials relate to this reaction.



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

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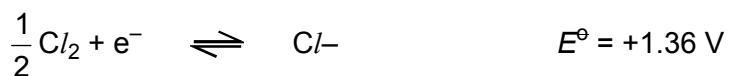
[3]

- (b) Determine the standard cell potential for a cell constructed from these two redox systems.

[1]

[Total 4 marks]

10. Chlorine gas may be prepared in the laboratory by reacting hydrochloric acid with potassium manganate(VII). The following standard electrode potentials relate to this reaction.



- (a) Use the half-equations above to:

- (i) construct an ionic equation for the reaction between hydrochloric acid and potassium manganate(VII);

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[2]

(ii) determine the oxidation numbers of chlorine and manganese before and after the reaction has taken place;

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[2]

(iii) state what is oxidised and what is reduced in this reaction.

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[2]

(b) If potassium manganate(VII) and very dilute hydrochloric acid are mixed, there is no visible reaction. Suggest why there is no visible reaction in this case.

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[1]

Total 7 marks]