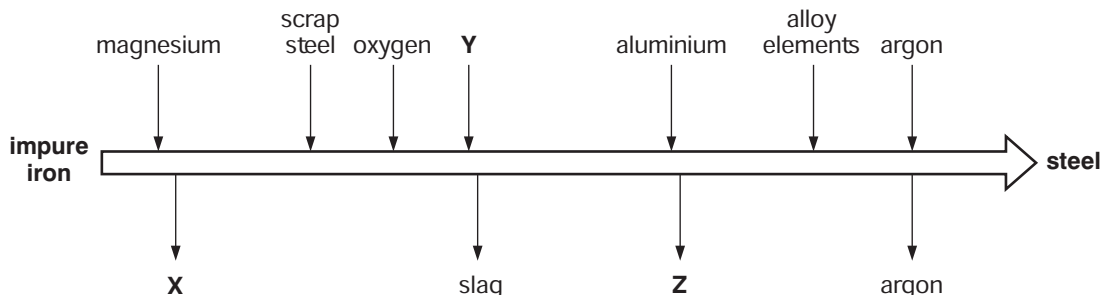


**'THE STEEL STORY' TEST**

60 marks (1 hour)

- 1 The Basic Oxygen Steelmaking process (BOS) can be represented by the following flow diagram.



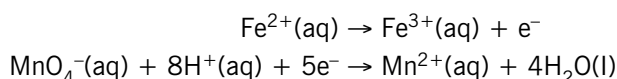
- a Identify the substances **X**, **Y**, and **Z** by name. (3 marks)
- b During the oxygen blast, several non-metal elements besides carbon are converted into oxides, which then react to form a molten 'slag'.
- Give the formula of **one** oxide produced in this way. (1 mark)
  - What **type** of reaction results in the formation of slag? (1 mark)
  - During the oxidation process, carbon monoxide escapes as a gas and is collected. Suggest **two** reasons why the carbon monoxide is not released into the air. (2 marks)
- c The process can produce mild steel. Explain what *mild* steel is, and give **one** important property and **one** use. (3 marks)

[TOTAL: 10 MARKS]

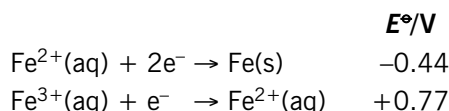
- 2 Samples of iron containing carbon as impurity can be analysed by titration with potassium manganate(VII) in acid solution.

In one such experiment, 1.00 g of impure iron was completely dissolved in dilute sulphuric acid, and the resulting solution containing iron(II) ions made up to 250 cm<sup>3</sup>.

A 25.0 cm<sup>3</sup> portion of this solution reacted exactly with 12.9 cm<sup>3</sup> of a 0.0250 mol dm<sup>-3</sup> solution of potassium manganate(VII). The half-equations for the redox reaction used in the titration are given below.

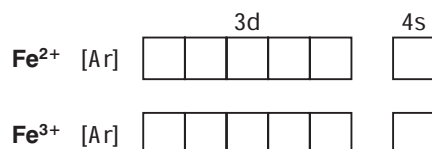


- a State how the end-point of the titration is determined. (1 mark)
- b Calculate the percentage by mass of iron in the sample, using the following steps. (You may assume that the potassium manganate(VII) reacts only with the iron(II) ions.)
- Calculate the amount in moles of  $\text{MnO}_4^{-}$  ions added from the burette in the titration and used to oxidise the iron(II) ions present in 25.0 cm<sup>3</sup>. (2 marks)
  - Use the half-equations to deduce the amount in moles of  $\text{Fe}^{2+}$  present in the 25.0 cm<sup>3</sup> portion. (1 mark)
  - Work out the mass of iron and hence the percentage of pure iron in the 1.00 g sample of impure iron. ( $A_r$ : Fe, 56) (3 marks)
- c Use the following half-equations and standard electrode potentials to explain why iron(II) ions, and *not* iron(III) ions, are formed when iron metal is added to dilute sulphuric acid. (4 marks)



- d Iron has an atomic number of 26. Draw out the 'electrons in a box' arrangements for the  $\text{Fe}^{2+}$  and

$\text{Fe}^{3+}$  ions. Use these to suggest why  $\text{Fe}^{2+}$  ions are readily oxidised to  $\text{Fe}^{3+}$  ions. (4 marks)



e Describe what you would see if excess sodium hydroxide solution were added to the solution of iron(II) ions in sulphuric acid. (2 marks)

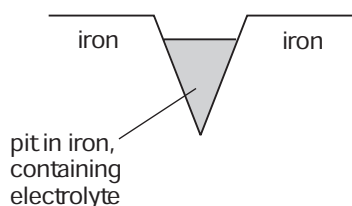
[TOTAL: 17 MARKS]

3 In the early 19th century, cans made of tin-plated iron were first used to preserve food. Tin has the advantages that it corrodes much less readily than iron and can form a barrier that protects the iron. When the coating is scratched, however, the iron rusts faster when in contact with tin. Fortunately neither  $\text{Fe}^{2+}$  nor  $\text{Sn}^{2+}$  ions are toxic.

Standard electrode potentials for some relevant half-reactions are given in the list below.

	$E^\circ/V$	
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s})$	-0.44	<b>Equation 3.1</b>
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Sn}(\text{s})$	-0.14	<b>Equation 3.2</b>
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$	+0.40	<b>Equation 3.3</b>
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$	+0.77	<b>Equation 3.4</b>

- a State **two** substances that are necessary for iron to rust, and from which the iron is protected by the tin layer. (2 marks)
- b Any tin ions present encourage the iron to corrode because a reaction occurs involving the half-reactions represented by **Equations 3.1** and **3.2**. Write a balanced overall equation for this reaction. (1 mark)
- c i When iron starts to corrode, a reaction occurs involving the half-reactions represented by **Equations 3.1** and **3.3**. Write a balanced overall equation for this reaction, with state symbols. (3 marks)
- ii When the surface of the iron is pitted, the half-reactions represented by **Equations 3.1** and **3.3** occur at different sites, with electrons flowing between these sites.



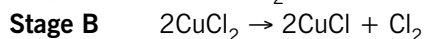
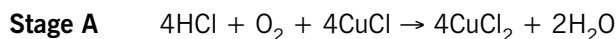
Mark on the diagram above the places where the half-reactions represented by **Equations 3.1** and **3.3** occur, and show the direction of flow of electrons between the two sites. (3 marks)

- d Two electrodes, whose half-reactions are represented by **Equations 3.2** and **3.4**, are put together to form an electrochemical cell.
- i Draw a diagram of this cell, labelling the reagents and indicating how standard conditions are achieved. (4 marks)
- ii Calculate the value of  $E_{\text{cell}}$ . (1 mark)
- iii Describe how you would modify the cell to measure the *standard electrode potential* of the electrode whose half-reaction is represented by **Equation 3.2**. (2 marks)

[TOTAL: 16 MARKS]

(Adapted from OCR Chemistry (Salters), Paper 1, question 6, 1994)

- 4 Chlorine gas used to be manufactured from hydrogen chloride and oxygen, in the presence of a copper(I) compound. This involved two stages:



- a What is the role of the copper(I) compound in the production of chlorine? (1 mark)
- b The copper(I) ion has an outer electronic configuration of  $3d^{10}$ . Use this information to explain why copper(I) chloride is not coloured like copper(II) chloride. (2 marks)
- c The  $\text{Cu}^{2+}$  ion forms stable complexes with many ligands, some of which are shown in the table below, together with information about their stability constants,  $K_{\text{stab}}$ .

Complex ion of copper	$K_{\text{stab}}$	$\lg K_{\text{stab}}$
$[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$	$1.3 \times 10^{13}$	13.1
$[\text{Cu}(\text{en})_2(\text{H}_2\text{O})_2]^{2+}$	$4.0 \times 10^{19}$	19.6
$[\text{Cu}(\text{CN})_4(\text{H}_2\text{O})_2]^{2-}$	$2.0 \times 10^{27}$	27.3
$[\text{Cu}(\text{edta})]^{2-}$	$6.3 \times 10^{18}$	18.8

en = 1,2-diaminoethane; edta = ligand produced from ethylenediamine-tetra-acetic acid. (For simplicity, the units of  $K_{\text{stab}}$  have been omitted from the table.)

- i Choose from the examples above:
- the formula of a neutral ligand; (1 mark)
  - the formula of a negatively charged ligand. (1 mark)
- ii The ligand 'edta' is described as *polydentate*. Explain what this means. (1 mark)
- iii What is the significance of the size of the stability constant? (2 marks)
- iv If each of the ligands in the above table, other than water, were added in turn to an aqueous solution of  $\text{Cu}^{2+}$  ions, what would be the formula of the resulting complex ion? (1 mark)
- v Sketch the shape you would expect for the complex ion  $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ . Label the ions and molecules that make up the structure. (3 marks)
- d The concentration of a solution of aqueous copper(II) ions can be found using a colorimeter.
- Explain why a blue solution looks blue when illuminated by white light. (1 mark)
  - You are given a colorimeter and a  $0.10 \text{ mol dm}^{-3}$  solution of  $\text{Cu}^{2+}$  ions. Describe what you would do to plot a calibration curve and measure the concentration of an unknown solution of  $\text{Cu}^{2+}(\text{aq})$ . (4 marks)

[TOTAL: 17 MARKS]