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’.
   i. Give the formula of one oxide produced in this way. (1 mark)
   ii. What type of reaction results in the formation of slag? (1 mark)
   iii. 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 cm³. A 25.0 cm³ portion of this solution reacted exactly with 12.9 cm³ of a 0.0250 mol dm⁻³ solution of potassium manganate(VII). The half-equations for the redox reaction used in the titration are given below.

   \[
   \begin{align*}
   \text{Fe}^{2+}(aq) & \rightarrow \text{Fe}^{3+}(aq) + e^- \\
   \text{MnO}_4^- (aq) + 8\text{H}^+ (aq) + 5e^- & \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l)
   \end{align*}
   \]

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.)
   i. Calculate the amount in moles of MnO₄⁻ ions added from the burette in the titration and used to oxidise the iron(II) ions present in 25.0 cm³. (2 marks)
   ii. Use the half-equations to deduce the amount in moles of Fe²⁺ present in the 25.0 cm³ portion. (1 mark)
   iii. Work out the mass of iron and hence the percentage of pure iron in the 1.00 g sample of impure iron. (Aᵣ: 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)

\[
\begin{align*}
\text{Fe}^{2+}(aq) + 2e^- & \rightarrow \text{Fe(s)} & -0.44 \\
\text{Fe}^{3+}(aq) + e^- & \rightarrow \text{Fe}^{2+}(aq) & +0.77
\end{align*}
\]

d. Iron has an atomic number of 26. Draw out the ‘electrons in a box’ arrangements for the Fe²⁺ and
Fe$^{3+}$ ions. Use these to suggest why Fe$^{2+}$ ions are readily oxidised to Fe$^{3+}$ ions. (4 marks)

![Electron Configuration Diagram]

Fe$^{2+}$ [Ar] 

Fe$^{3+}$ [Ar] 

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

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 Fe$^{2+}$ nor Sn$^{2+}$ ions are toxic.

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

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

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 $E_{cell}$. (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: 17 MARKS]

[Adapted from OCR Chemistry (Salters), Paper 1, question 6, 1994]
Chlorine gas used to be manufactured from hydrogen chloride and oxygen, in the presence of a copper(I) compound. This involved two stages:

**Stage A**

\[ 4\text{HCl} + O_2 + 4\text{CuCl} \rightarrow 4\text{CuCl}_2 + 2\text{H}_2\text{O} \]

**Stage B**

\[ 2\text{CuCl}_2 \rightarrow 2\text{CuCl} + \text{Cl}_2 \]

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 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, \(K_{stab}\).

<table>
<thead>
<tr>
<th>Complex ion of copper</th>
<th>(K_{stab})</th>
<th>(\log K_{stab})</th>
</tr>
</thead>
<tbody>
<tr>
<td>([\text{Cu(NH}_3)_4(\text{H}_2\text{O})_2]^{2+})</td>
<td>(1.3 \times 10^{13})</td>
<td>13.1</td>
</tr>
<tr>
<td>([\text{Cu(en)}_2(\text{H}_2\text{O})_2]^{2+})</td>
<td>(4.0 \times 10^{19})</td>
<td>19.6</td>
</tr>
<tr>
<td>([\text{Cu(CN)}_4(\text{H}_2\text{O})_2]^{2-})</td>
<td>(2.0 \times 10^{27})</td>
<td>27.3</td>
</tr>
<tr>
<td>([\text{Cu(edta)}]^{2-})</td>
<td>(6.3 \times 10^{18})</td>
<td>18.8</td>
</tr>
</tbody>
</table>

\(\text{en} = 1,2\text{-diaminoethane; edta = ligand produced from ethylenediamine-tetra-acetic acid.}\)

(For simplicity, the units of \(K_{stab}\) have been omitted from the table.)

i. Choose from the examples above:
   1. the formula of a neutral ligand; (1 mark)
   2. 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 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 \([\text{Cu(NH}_3)_4(\text{H}_2\text{O})_2]^{2+}\). 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.
   i. Explain why a blue solution looks blue when illuminated by white light. (1 mark)
   ii. You are given a colorimeter and a 0.10 mol dm\(^{-3}\) solution of Cu\(^{2+}\) ions. Describe what you would do to plot a calibration curve and measure the concentration of an unknown solution of Cu\(^{2+}\) (aq). (4 marks)

[**TOTAL: 17 MARKS**]