

- 4 (a) Define standard cell potential, $E_{\text{cell}}^{\ominus}$. Include a description of standard conditions.

.....

 [2]

- (b) The Daniell cell is an electrochemical cell consisting of a $\text{Cu}^{2+}(\text{aq})/\text{Cu}(\text{s})$ electrode and a $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$ electrode.

- (i) Draw a labelled diagram of this electrochemical cell.

Include all necessary substances and relevant pieces of apparatus needed to measure the $E_{\text{cell}}^{\ominus}$.

It is **not** necessary to state the conditions used.

[3]

- (ii) State the charge carriers that transfer current through the solutions and through the wire.

the solutions the wire [1]

- (iii) The standard electrode potential, $E^{\ominus}$, for the $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$ electrode is -0.76 V .

Water is added to a standard $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$ electrode.

The new concentration of $\text{Zn}^{2+}(\text{aq})$ is 0.25 mol dm^{-3} .

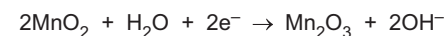
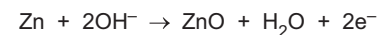
Use the Nernst equation to calculate the electrode potential, E , for this new $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$ electrode.

① $E(\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})) = \dots\dots\dots \text{ V}$ [2]

- (c) An electrochemical cell consists of a ZnO/Zn electrode and a $\text{MnO}_2/\text{Mn}_2\text{O}_3$ electrode in an alkaline electrolyte.

The standard cell potential, $E_{\text{cell}}^{\ominus}$, for this cell is $+1.47\text{ V}$.

The half-equation at each electrode when this cell is discharging is shown.



- (i) Use this information to determine the change in oxidation state of manganese when this cell is discharging.

from to [1]

- (ii) Write the equation for the overall reaction that occurs when this cell is discharging.

..... [1]

- (iii) The $E^{\ominus}$ for the ZnO/Zn electrode is -1.28 V .

Calculate the standard electrode potential, $E^{\ominus}$, for the $\text{MnO}_2/\text{Mn}_2\text{O}_3$ electrode.

$$E^{\ominus}(\text{MnO}_2/\text{Mn}_2\text{O}_3) = \dots\dots\dots \text{ V} [1]$$

[Total: 11]

②

5 Cobalt is a transition element which forms compounds containing Co^{2+} and Co^{3+} ions. Cobalt(II) sulfate dissolves in water to form a solution containing the $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ complex ion.

(a) (i) Complete the electronic configurations of a Co^{2+} ion and a Co^{3+} ion.

$\text{Co}^{2+} = [\text{Ar}] \dots\dots\dots$

$\text{Co}^{3+} = [\text{Ar}] \dots\dots\dots$

[1]

(ii) Explain why transition elements can form complex ions.

.....

..... [1]

(iii) An excess of concentrated HCl is added to a solution containing $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

Describe the colour change observed and the state of the cobalt-containing product.

The colour changes from to

The state of the cobalt-containing product is

[2]

(iv) Write an equation for the reaction occurring in (a)(iii).

..... [1]

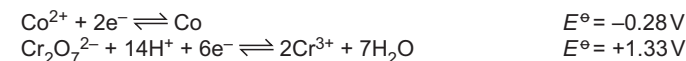
(v) Name the type of reaction occurring in (a)(iii).

..... [1]

(vi) Write an equation for the reaction that occurs when an excess of $\text{NaOH}(\text{aq})$ is added to a solution containing $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

..... [1]

(b) Cobalt metal can be oxidised by acidified $\text{K}_2\text{Cr}_2\text{O}_7$. The relevant half-equations, and their $E^\ominus$ values, are shown.



(i) A Co^{2+}/Co electrode is constructed in which $[\text{Co}^{2+}]$ is $0.020 \text{ mol dm}^{-3}$ at 298 K .

Use the Nernst equation to show that the E value for this Co^{2+}/Co electrode is -0.33 V .

[2]

(ii) An electrochemical cell is constructed using the Co^{2+}/Co electrode described in (b)(i) and a $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$ electrode in which all conditions are standard.

Calculate the value of E_{cell} .

$E_{\text{cell}} = \dots\dots\dots$ [1]

(iii) A current is drawn from the electrochemical cell described in (b)(ii).

Write an equation for the reaction taking place in the cell.

..... [1]

(iv) Complete the sentences to identify the negative electrode and the direction of electron flow when a current is drawn from the cell described in (b)(ii).

The electrode is the negative electrode.

Electrons flow from the electrode to the electrode. [1]

(c) A molten Co^{2+} salt is electrolysed using a current of 0.500 A .

0.547 g of cobalt metal forms at the cathode. Under the conditions used no other reduction reaction occurs at the cathode.

Calculate the time in minutes for which the current flows to produce this mass of cobalt.

Give your answer to **three** significant figures.

time = min [3]

2 (a) Iron can form stable ions in the +2 and +3 oxidation states.

Explain why transition elements have variable oxidation states.

.....
 [1]

(b) Aqueous solutions of iron(II) salts contain the complex ion $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$.

Define complex ion.

.....
 [1]

(c) $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$ can be converted into $[\text{Fe}(\text{H}_2\text{O})_4(\text{OH})_2]$.

(i) Suggest a suitable reagent for this conversion. State the type of reaction.

reagent

type of reaction

[1]

(ii) $[\text{Fe}(\text{H}_2\text{O})_4(\text{OH})_2]$ is a green precipitate that turns brown on standing in air.

Table 2.1 shows electrode potentials for some electrode reactions.

Table 2.1

electrode reaction	$E^\ominus / \text{V}$
$\text{Fe}(\text{H}_2\text{O})_3(\text{OH})_3 + \text{H}_2\text{O} + \text{e}^- \rightleftharpoons \text{Fe}(\text{H}_2\text{O})_4(\text{OH})_2 + \text{OH}^-$	-0.56
$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-$	+0.40

Use the information in Table 2.1 to explain why $[\text{Fe}(\text{H}_2\text{O})_4(\text{OH})_2]$ turns brown on standing in air.

Include an equation for this reaction.

.....

 [3]

(d) The complex $[\text{Co}(\text{NH}_3)_6]^{2+}$ reacts with hydrogen peroxide as shown.



Calculate $\Delta G^\ominus$, in kJ mol^{-1} , for reaction 1.

$\Delta G^\ominus = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

[Total: 8]