8 **(a)** **(i)** Define the term *standard cell potential.*

....

....

.... [2]

An electrochemical cell is set up to measure the standard electrode potential of a cell, E°_{cell}, made of a Co^{3+}/Co^{2+} half-cell and a Cl_2/Cl^- half-cell.

(ii) Complete the table with the substance used to make the electrode in each of these half-cells.

half-cell	electrode
Co^{3+}/Co^{2+}	
Cl_2/Cl^-	

[1]

(iii) Use data from the *Data Booklet* to calculate the E°_{cell}.

E°_{cell} = ... V [1]

(iv) Write the equation for the overall cell reaction.

.... [1]

(b) A fuel cell is an electrochemical cell that can be used to generate electrical energy. A methanol-oxygen fuel cell can be used as an alternative to a hydrogen-oxygen fuel cell. When the cell operates, the carbon atoms in the methanol molecules are converted into carbon dioxide.

$$CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$$

Calculate the volume of CO_2, in cm^3, formed when a current of 2.5A is delivered by the cell for 30 minutes. Assume the cell is operated at room conditions.

volume of CO_2 = ... cm^3 [2]

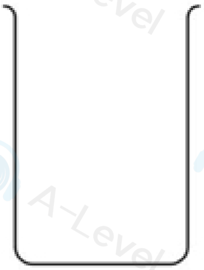
[Total: 7]

3 Some electrode potentials are shown in Table 3.1.

Table 3.1

electrode reaction	$E^{\ominus}$/V
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20
$V^{3+} + e^- \rightleftharpoons V^{2+}$	–0.26
$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$	+0.34
$VO_2^+ + 2H^+ + e^- \rightleftharpoons VO^{2+} + H_2O$	+1.00
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$Fe^{3+} + 3e^- \rightleftharpoons Fe$	–0.04
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89

(a) **(i)** Complete the diagram to show a standard hydrogen electrode.
Label your diagram. Identify all substances. You do **not** need to state standard conditions.



[1]

(ii) An electrochemical cell is set up using an Fe^{3+}/Fe^{2+} electrode and a standard hydrogen electrode.

Identify the positive electrode in the electrochemical cell and the direction of electron flow in the external circuit.

positive electrode ....

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

(b) The vanadium-containing species in the electrode reactions given in Table 3.1 are V, V^{2+}, V^{3+}, VO^{2+} and VO_2^{+}.

(i) Identify **one** vanadium-containing species that does **not** react with Fe^{2+} ions under standard conditions.

Use data from Table 3.1 to explain your answer.

....

.... [1]

(ii) Identify **all** the vanadium-containing species that will react with Fe^{2+} ions under standard conditions.

.... [1]

(iii) Write an equation for **one** of the possible reactions identified in **(ii)**.

.... [1]

(c) Another electrochemical cell is set up using an Fe^{3+}/Fe^{2+} electrode and an alkaline ClO^-/Cl^- electrode.

The concentration of Fe^{3+} is 1000 times greater than the con the Fe^{3+}/Fe^{2+} electrode. All other conditions are standard.

(i) Use the Nernst equation to calculate the E value of the Fe^{3+}/F

Show your working.

E = ... V [2]

(ii) Write an equation for the reaction that occurs in the cell, under these conditions.

.... [1]

(d) Another electrochemical cell is set up using an Fe^{2+}/Fe electrode and an alkaline ClO^-/Cl^- electrode under standard conditions.

Calculate the value of $\Delta G^\ominus$ for the cell.

$\Delta G^\ominus$ = ... $kJ\,mol^{-1}$ [3]

1 **(a)** The most common oxidation states of cobalt are +2 and +3.

Complete the electronic configurations of the following free ions.

- Co^{2+} [Ar] ....
- Co^{3+} [Ar] ....

[1]

(b) Co^{2+} and Co^{3+} both form complexes with $edta^{4-}$.

half-equation	E°/V
$Co^{3+} + e^- \rightleftharpoons Co^{2+}$	+1.82
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23
$[Co(edta)]^- + e^- \rightleftharpoons [Co(edta)]^{2-}$	+0.38
$Co^{2+} + 2e^- \rightleftharpoons Co$	–0.28

Use the data in the table to predict what happens, if anything, when separate aqueous solutions of Co^{3+} and $[Co(edta)]^-$ are left to stand in the air.

aqueous solution of Co^{3+}

....

....

aqueous solution of $[Co(edta)]^-$

....

....

[3]

(b) The Group 1 iodides all form stable ionic lattices and are soluble in water.

(i) Define enthalpy change of solution.

....

....

.... [1]

(ii) Use the data in Table 1.1 to calculate the enthalpy change of solution of potassium iodide, KI.

Table 1.1

process	enthalpy change, ΔH / kJ mol^{-1}
$K^+(g) + I^-(g) \rightarrow KI(s)$	–629
$K^+(g) \rightarrow K^+(aq)$	–322
$I^-(g) \rightarrow I^-(aq)$	–293

enthalpy change of solution = ... kJ mol^{-1} [1]

(iii) Suggest the trend in the magnitude of the lattice energies of the Group 1 iodides, LiI, NaI, KI.
Explain your answer.

....

....

....

.... [2]

1 An electrochemical cell is constructed using two half-cells.

- a Br_2/Br^- half-cell
- an Mn^{3+}/Mn^{2+} half-cell

(a) State the material used for the electrode in each half-cell.

Br_2/Br^- half-cell ....

Mn^{3+}/Mn^{2+} half-cell ....
[1]

(b) The cell is operated at 298 K.

The Br_2/Br^- half-cell has standard concentrations.

The Mn^{3+}/Mn^{2+} half-cell has $[Mn^{3+}] = 0.500\,mol\,dm^{-3}$ and $[Mn^{2+}] = 0.100\,mol\,dm^{-3}$.

(i) Use the Nernst equation to calculate the electrode potential, *E*, of the Mn^{3+}/Mn^{2+} half-cell under these conditions.

E = ... V [2]

(ii) Calculate the E_{cell} under these conditions.

E_{cell} = ... V [1]

(iii) Write an equation for the overall cell reaction that occurs.

.... [2]

(c) An aqueous solution of copper(II) sulfate is electrolysed using copper electrodes. A current of 1.50 A is passed for 3.00 hours. 5.09 g of copper is deposited on the cathode.

The charge on one electron is -1.60×10^{-19} C.

The relative atomic mass of copper is 63.5.

Use these data to calculate an experimentally determined value for the Avogadro constant, *L*. Give your answer to **three** significant figures.

L = ... mol^{-1} [5]

(d) Explain why magnesium metal cannot be obtained by the electrolysis of dilute aqueous magnesium sulfate. Your answer should include data from the *Data Booklet*.

....

....

....

.... [2]

[Total: 13]

3 Vanadium is a transition element in Period 4 of the Periodic Table.

(a) Define transition element.

....

.... [1]

(b) Vanadium shows typical chemical properties of transition elements, including variable oxidation states.

(i) State **two** other typical chemical properties of transition elements.

1 ....

2 ....
[1]

(ii) Explain why transition elements have variable oxidation states.

....

.... [1]

(ii) An excess of $C_2O_4^{2-}$ reacts with VO_2^+ to form a mixture of two octahedral complex ions. The complex ions are stereoisomers of each other.

Each complex ion contains a V^{2+} cation and three $C_2O_4^{2-}$ ions.

Complete the diagram to show the three-dimensional structure of **one** of the complex ions.

Include the charge of the complex ion.

Use O⌒O to represent a $C_2O_4^{2-}$ ion.

[2]

(d) V^{2+}(aq) can be oxidised by H_2O_2(aq).

Table 3.2 gives some relevant data.

Table 3.2

	half-equation	$E^\ominus$ /V
1	$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
2	$VO_2^+(aq) + 2H^+(aq) + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	+1.00
3	$VO^{2+}(aq) + 2H^+(aq) + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
4	$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	–0.26

(i) Identify the vanadium species that forms when an excess of H_2O_2(aq) reacts with V^{2+}(aq) under standard conditions.

Explain your answer with reference to the data in Table 3.2.

....

....

.... [1]

(ii) Concentrated acidified H_2O_2 can react with V^{2+} to form red VO^{3+} ions

9 **(a)** Define standard cell potential, $E^{\ominus}_{\text{cell}}$.

....

.... [1]

(b) An electrochemical cell is set up to measure $E^{\ominus}_{\text{cell}}$ of a cell consisting of an Fe^{3+}/Fe^{2+} half-cell and a Cl_2/Cl^- half-cell.

Draw a labelled diagram of this electrochemical cell.

Include all necessary substances. It is **not** necessary to state conditions used.

[3]

(c) The cell reaction for the electrochemical cell in **(b)** is shown.

$$Cl_2 + 2Fe^{2+} \rightarrow 2Fe^{3+} + 2Cl^- \qquad E^{\ominus}_{\text{cell}} = +0.59\,\text{V}$$

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

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

2 **(a)** **(i)** Lithium nitrate, $LiNO_3$, decomposes on heating in a similar way to Group 2 nitrates to give the metal oxide, a brown gas and oxygen.

Write an equation for the decomposition of $LiNO_3$.

.... [1]

(ii) The other Group 1 nitrates, MNO_3, decompose on heating to form the metal nitrite, MNO_2, and oxygen.

The thermal stability of these nitrates increases down the group.

Suggest why the thermal stability of MNO_3 increases down the group.

....

....

.... [2]

(b) Acidified manganate(VII) ions, MnO_4^-, can be used to analyse solutions containing nitrite ions, NO_2^-, by titration.

X is a solution of $NaNO_2$.

250.0 cm³ of **X** is added to 50.0 cm³ of 0.125 mol dm⁻³ acidified MnO_4^-(aq). The MnO_4^-(aq) ions are in excess; all the NO_2^- ions are oxidised in the reaction.

The unreacted MnO_4^-(aq) required 22.50 cm³ of 0.0400 mol dm⁻³ Fe^{2+}(aq) to reach the end-point.

The relevant half-equations are shown.

$$NO_2^- + H_2O \rightleftharpoons NO_3^- + 2H^+ + 2e^-$$
$$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$$
$$Fe^{2+} \rightleftharpoons Fe^{3+} + e^-$$

Calculate the concentration, in mol dm⁻³, of $NaNO_2$ in **X**.

(c) Table 2.1 shows electrode potentials for some electrode reactions involving manganese compounds.

Table 2.1

electrode reaction	$E^\ominus$/V
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	–1.18
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + e^- \rightleftharpoons MnO_4^{2-}$	+0.56
$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.67
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.52
$MnO_4^- + 2H_2O + 3e^- \rightleftharpoons MnO_2 + 4OH^-$	+0.59
$MnO_4^{2-} + 2H_2O + 2e^- \rightleftharpoons MnO_2 + 4OH^-$	+0.60
$MnO_4^{2-} + 4H^+ + 2e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.70

(i) Aqueous manganate(VI) ions, MnO_4^{2-}, are unstable in acidic conditions and undergo a disproportionation reaction.

The $E^\ominus_{cell}$ for this reaction is +1.14 V.

Construct an overall ionic equation for this disproportionation reaction.

.... [2]

(ii) Suggest and explain how the E_{cell} value of the disproportionation reaction changes with an increase in pH.

....

....

.... [1]

[Total: 9]

1 **(a)** **(i)** Complete the electronic configuration of the copper(II) ion.

$1s^2 2s^2 2p^6$ .... [1]

(ii) State the colour of the solutions containing the following ions.

- $[Cu(H_2O)_6]^{2+}(aq)$ ....
- $[CuCl_4]^{2-}(aq)$ ....

[1]

(iii) Octahedral complexes of Cu^{2+} with different ligands can have different colours.

Explain why.

....

....

....

....

.... [2]

(b) Copper(I) and silver(I) salts are colourless.

Suggest why.

....

....

....

.... [2]

(c) Consider the following two equilibria and associated data values at 298 K.

$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$ equilibrium **1** $K_{sp} = 5.0 \times 10^{-13}\,mol^2\,dm^{-6}$

$Ag^+(aq) + 2NH_3(aq) \rightleftharpoons [Ag(NH_3)_2]^+(aq)$ equilibrium **2** $K_{stab} = 1.7 \times 10^7\,mol^{-2}\,dm^6$

The equilibrium constant for equilibrium **1** is the solubility product, K_{sp}, of AgBr(s). The equilibrium constant for equilibrium **2** is the stability constant, K_{stab}, for the formation of $[Ag(NH_3)_2]^+(aq)$.

(i) Calculate the solubility of AgBr at 298 K in mol dm^{-3}.

solubility of AgBr = ... mol dm^{-3} [1]

(ii) Use Le Chatelier's principle as applied to equilibria **1** and **2** to suggest why AgBr(s) dissolves in concentrated $NH_3(aq)$.

....

....

....

.... [2]

(iii) Use equilibria **1** and **2** to construct an equation for the reaction of AgBr(s) with concentrated $NH_3(aq)$. This is equilibrium **3**.

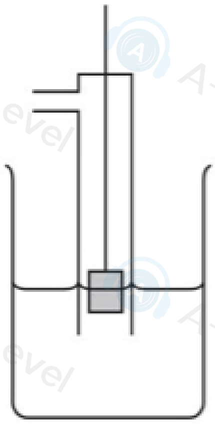
.... equilibrium **3** [1]

(iv) Write an expression for the equilibrium constant of equilibrium **3**, K_{eq3}, in terms of K_{sp} for equilibrium **1** and K_{stab} for equilibrium **2**.

$K_{eq3} =$

[1]

(e) **(i)** Complete and **label** the diagram to show how the standard electrode potential, E°, of $Ag^+(aq)/Ag(s)$ could be measured under **standard conditions**.



[4]

(ii) Use the *Data Booklet* to label the diagram in **(e)(i)** to show

- which is the positive electrode,
- the direction of electron flow in the external circuit when a current flows.

[1]

[Total: 17]