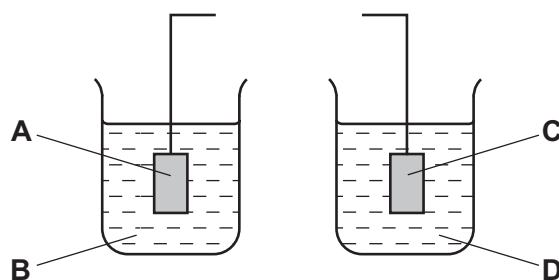


1. Question | 9701 May-June 2016 P4 V1 Q4

- 4 (a) (i) Define the term *standard cell potential*,  $E_{\text{cell}}^{\ominus}$ .

.....  
 .....  
 ..... [1]

The following incomplete diagram shows the apparatus that can be used to measure the  $E_{\text{cell}}^{\ominus}$  for a cell composed of the  $\text{Fe}^{3+}/\text{Fe}^{2+}$  and  $\text{Ag}^{+}/\text{Ag}$  half-cells.



- (ii) Complete the diagram, labelling the components you add. [1]

- (iii) Identify the components **A-D**.

**A** .....  
**B** .....  
**C** .....  
**D** ..... [3]

- (b) (i) Use  $E^{\ominus}$  values to write an equation for the cell reaction that takes place if the two electrodes in (a) are connected by a wire and the circuit is completed.

.....  
 ..... [1]

- (ii) Another electrochemical cell was set up using  $0.31 \text{ mol dm}^{-3} \text{ Ag}^{+}(\text{aq})$  instead of the standard Ag solution.

Use the Nernst equation,  $E = E^{\ominus} + 0.059 \log [\text{Ag}^{+}(\text{aq})]$ , and the relevant  $E^{\ominus}$  values to calculate the new  $E_{\text{cell}}$  in this experiment.

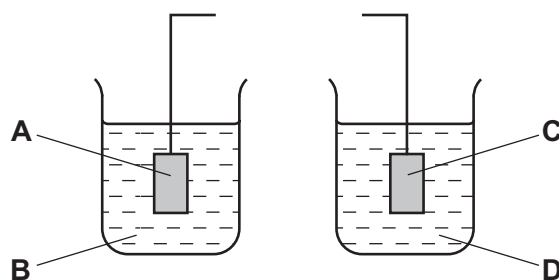
$$E_{\text{cell}} = \dots\dots\dots \text{ V [2]}$$

[Total: 8]

- 4 (a) (i) Define the term *standard cell potential*,  $E_{\text{cell}}^{\ominus}$ .

.....  
 .....  
 ..... [1]

The following incomplete diagram shows the apparatus that can be used to measure the  $E_{\text{cell}}^{\ominus}$  for a cell composed of the  $\text{Fe}^{3+}/\text{Fe}^{2+}$  and  $\text{Ag}^{+}/\text{Ag}$  half-cells.



- (ii) Complete the diagram, labelling the components you add. [1]

- (iii) Identify the components **A-D**.

**A** .....  
**B** .....  
**C** .....  
**D** ..... [3]

- (b) (i) Use  $E^{\ominus}$  values to write an equation for the cell reaction that takes place if the two electrodes in (a) are connected by a wire and the circuit is completed.

.....  
 ..... [1]

- (ii) Another electrochemical cell was set up using  $0.31 \text{ mol dm}^{-3} \text{ Ag}^{+}(\text{aq})$  instead of the standard Ag solution.

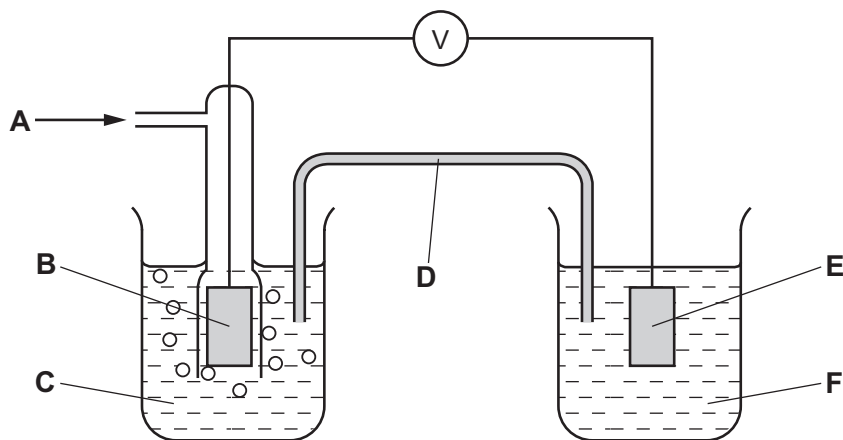
Use the Nernst equation,  $E = E^{\ominus} + 0.059 \log [\text{Ag}^{+}(\text{aq})]$ , and the relevant  $E^{\ominus}$  values to calculate the new  $E_{\text{cell}}$  in this experiment.

$$E_{\text{cell}} = \dots\dots\dots \text{ V [2]}$$

[Total: 8]

3. Question part 1/2 | 9701 Feb-March 2017 P4 V2 Q3

- 3 (a) The diagram shows the apparatus used to measure the standard electrode potential,  $E^\ominus$ , of  $\text{Fe}^{3+}(\text{aq})/\text{Fe}^{2+}(\text{aq})$ .



- (i) Identify what the letters **A** to **F** represent.

|                |                |
|----------------|----------------|
| <b>A</b> ..... | <b>D</b> ..... |
| <b>B</b> ..... | <b>E</b> ..... |
| <b>C</b> ..... | <b>F</b> ..... |

[3]

- (ii) Label the diagram to show

- which is the positive electrode,
- the direction of electron flow in the external circuit.

Use the *Data Booklet* to help you.

[1]

- (b) In another experiment, an  $\text{Fe}^{3+}(\text{aq})/\text{Fe}^{2+}(\text{aq})$  half-cell was connected to a  $\text{Cu}^{2+}(\text{aq})/\text{Cu}(\text{s})$  half-cell.

Determine the standard cell potential,  $E^\ominus_{\text{cell}}$ , when these two half-cells are connected by a wire and the circuit is completed.

Use the *Data Booklet* to help you.

$E^\ominus_{\text{cell}} = \dots\dots\dots \text{V}$  [1]

3. Question part 2/2 | 9701 Feb-March 2017 P4 V2 Q3

(c) (i) The  $E^\circ$  of  $\text{Ni}^{2+}(\text{aq})/\text{Ni}(\text{s})$  is  $-0.25\text{ V}$ .

State and explain how the electrode potential changes if the concentration of  $\text{Ni}^{2+}(\text{aq})$  is decreased.

.....  
.....  
..... [1]

(ii) The  $E^\circ$  of  $\text{Cr}^{3+}(\text{aq})/\text{Cr}^{2+}(\text{aq})$  is  $-0.41\text{ V}$ .

Calculate the electrode potential when  $[\text{Cr}^{3+}(\text{aq})]$  is  $0.60\text{ mol dm}^{-3}$  and  $[\text{Cr}^{2+}(\text{aq})]$  is  $0.15\text{ mol dm}^{-3}$ . Use the Nernst equation.

$$E = E^\circ + \frac{0.059}{z} \log \frac{[\text{oxidised species}]}{[\text{reduced species}]}$$

$E = \dots\dots\dots\text{ V}$  [2]

[Total: 8]

4. Question part 1/2 | 9701 Oct-Nov 2017 P4 V1 Q3

**3 (a)** Define the term *standard cell potential*.

.....

.....

..... [2]

**(b) (i)** Draw a fully labelled diagram of the experimental set-up you could use to measure the standard electrode potential of the  $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$  electrode. Include the necessary chemicals.

[4]

**(ii)** The  $E^\circ$  for a  $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$  electrode is  $-0.13\text{ V}$ .

Suggest how the  $E$  for this electrode would differ from its  $E^\circ$  value if the concentration of  $\text{Pb}^{2+}(\text{aq})$  ions is reduced. Indicate this by placing a tick (✓) in the appropriate box in the table.

| more negative | no change | less negative |
|---------------|-----------|---------------|
|               |           |               |

Explain your answer.

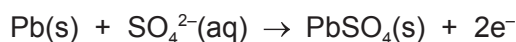
.....

.....

..... [2]

- (c) Car batteries are made up of rechargeable lead-acid cells. Each cell consists of a negative electrode made of Pb metal and a positive electrode made of PbO<sub>2</sub>. The electrolyte is H<sub>2</sub>SO<sub>4</sub>(aq).

When a lead-acid cell is in use, Pb<sup>2+</sup> ions are precipitated out as PbSO<sub>4</sub>(s) at the negative electrode.



- (i) Calculate the mass of Pb that is converted to PbSO<sub>4</sub> when a current of 0.40A is delivered by the cell for 80 minutes.

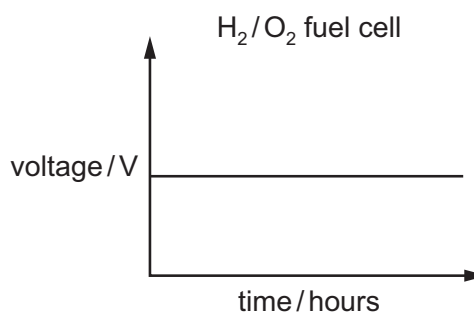
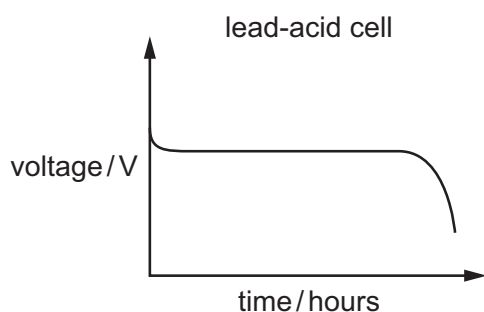
mass of Pb = ..... g [2]

- (ii) Complete the half-equation for the reaction taking place at the positive electrode.



[1]

- (d) The diagrams show how the voltage across two different cells changes with time when each cell is used to provide an electric current.



Suggest a reason why

- the voltage of the lead-acid cell changes after several hours,

.....

.....

- the voltage of the fuel cell remains constant.

.....

.....

[2]

[Total: 13]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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5. Question part 1/2 | 9701 Oct-Nov 2017 P4 V2 Q4

- 4** An electrochemical cell consists of a half-cell containing  $V^{3+}(aq)$  and  $V^{2+}(aq)$  ions and another half-cell containing  $VO_2^+(aq)$  and  $VO^{2+}(aq)$  ions.

**(a) (i)** Use data from the *Data Booklet* to calculate a value for the  $E_{cell}^\ominus$ .

$$E_{cell}^\ominus = \dots\dots\dots V \quad [1]$$

**(ii)** Write the ionic equation for the cell reaction.

..... [1]

**(iii)** Draw a fully labelled diagram of the apparatus you could use to measure the potential of this cell. Include the necessary chemicals.

[4]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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5. Question part 2/2 | 9701 Oct-Nov 2017 P4 V2 Q4

(b) Use data from the *Data Booklet* to predict whether a reaction might take place when the following pairs of aqueous solutions are mixed. If a reaction occurs, write an equation for it and calculate the  $E_{\text{cell}}^{\ominus}$ .

- $\text{V}^{2+}(\text{aq})$  and  $\text{Sn}^{4+}(\text{aq})$

Does a reaction occur? .....

equation .....

$E_{\text{cell}}^{\ominus}$  .....

- $\text{VO}^{2+}(\text{aq})$  and  $\text{Fe}^{3+}(\text{aq})$

Does a reaction occur? .....

equation .....

$E_{\text{cell}}^{\ominus}$  .....

[3]

[Total: 9]

6. Question part 1/2 | 9701 Oct-Nov 2017 P4 V3 Q3

3 (a) Define the term *standard cell potential*.

.....  
.....  
..... [2]

(b) (i) Draw a fully labelled diagram of the experimental set-up you could use to measure the standard electrode potential of the  $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$  electrode. Include the necessary chemicals.

[4]

(ii) The  $E^\circ$  for a  $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$  electrode is  $-0.13\text{ V}$ .

Suggest how the  $E$  for this electrode would differ from its  $E^\circ$  value if the concentration of  $\text{Pb}^{2+}(\text{aq})$  ions is reduced. Indicate this by placing a tick ( $\checkmark$ ) in the appropriate box in the table.

| more negative | no change | less negative |
|---------------|-----------|---------------|
|               |           |               |

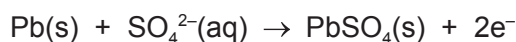
Explain your answer.

.....  
.....  
..... [2]

6. Question part 2/2 | 9701 Oct-Nov 2017 P4 V3 Q3

- (c) Car batteries are made up of rechargeable lead-acid cells. Each cell consists of a negative electrode made of Pb metal and a positive electrode made of PbO<sub>2</sub>. The electrolyte is H<sub>2</sub>SO<sub>4</sub>(aq).

When a lead-acid cell is in use, Pb<sup>2+</sup> ions are precipitated out as PbSO<sub>4</sub>(s) at the negative electrode.



- (i) Calculate the mass of Pb that is converted to PbSO<sub>4</sub> when a current of 0.40A is delivered by the cell for 80 minutes.

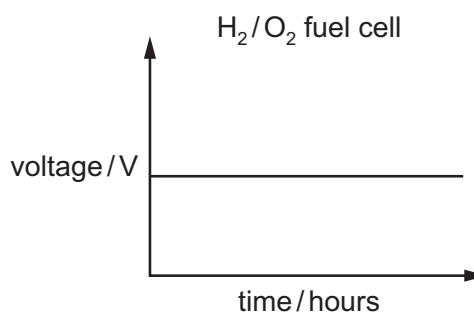
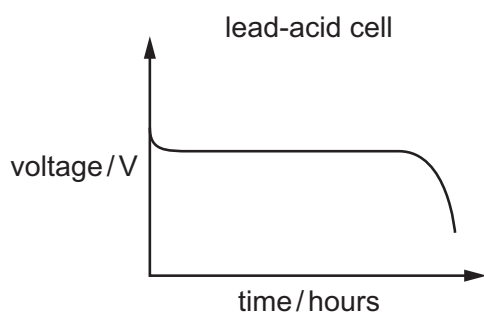
mass of Pb = ..... g [2]

- (ii) Complete the half-equation for the reaction taking place at the positive electrode.



[1]

- (d) The diagrams show how the voltage across two different cells changes with time when each cell is used to provide an electric current.



Suggest a reason why

- the voltage of the lead-acid cell changes after several hours,

.....

.....

- the voltage of the fuel cell remains constant.

.....

.....

[2]

[Total: 13]

7. Question part 1/2 | 9701 Feb-March 2018 P4 V2 Q2

- 2 (a) Describe the trend in the reactivity of the halogens  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$  as oxidising agents. Explain this trend using values of  $E^\circ (\text{X}_2/\text{X}^-)$  from the *Data Booklet*.

.....

.....

.....

.....

..... [2]

- (b) (i) Write an equation for the reaction between chlorine and water.

..... [1]

- (ii) Use standard electrode potential,  $E^\circ$ , data from the *Data Booklet* to calculate the  $E^\circ_{\text{cell}}$  for the following reaction.



$$E^\circ_{\text{cell}} = \dots\dots\dots \text{V} \quad [2]$$

- (iii) The  $[\text{OH}^-]$  was increased and the  $E_{\text{cell}}$  was measured.

Indicate how the value of the  $E_{\text{cell}}$  measured would compare to the  $E^\circ_{\text{cell}}$  calculated in (ii) by placing **one** tick (✓) in the table.

|                                                                        |  |
|------------------------------------------------------------------------|--|
| $E_{\text{cell}}$ becomes less positive than $E^\circ_{\text{cell}}$ . |  |
| $E_{\text{cell}}$ stays the same as $E^\circ_{\text{cell}}$ .          |  |
| $E_{\text{cell}}$ becomes more positive than $E^\circ_{\text{cell}}$ . |  |

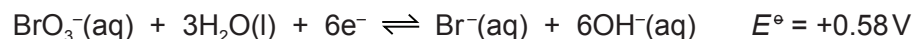
Explain your answer.

.....

.....

..... [2]

(c) A half-equation involving bromate(V) ions,  $\text{BrO}_3^-$ , and bromide ions is shown.



(i) An alkaline solution of chlorate(I),  $\text{ClO}^-$ , can be used to oxidise bromide ions to bromate(V) ions.

Use the *Data Booklet* and the half-equation shown to write an equation for this reaction.

..... [1]

(ii) Calculate the  $E^\circ_{\text{cell}}$  for the reaction in (i).

$$E^\circ_{\text{cell}} = \text{..... V [1]}$$

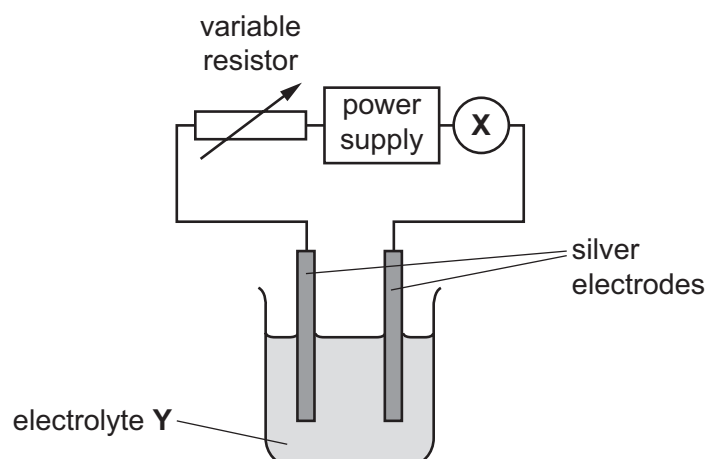
(iii) When a concentrated solution of bromic(V) acid,  $\text{HBrO}_3$ , is warmed, it decomposes to form bromine, oxygen and water only.

Write an equation for this reaction. The use of oxidation numbers may be helpful.

..... [1]

[Total: 10]

- 6 The apparatus shows a cell which can be used to determine a value of the Avogadro constant,  $L$ .



- (a) (i) Name component **X**.

..... [1]

- (ii) Suggest a suitable electrolyte **Y**.

..... [1]

- (b) In an experiment, a current of 0.200 A was passed through the cell for 40.0 minutes. The mass of the silver cathode increased by 0.500 g.

The charge on the electron is  $-1.60 \times 10^{-19} \text{ C}$ .

Calculate the:

- number of moles of silver deposited on the cathode
- number of coulombs of charge passed
- number of electrons passed
- number of electrons needed to deposit 1 mol of silver at the cathode.

[3]

[Total: 5]

- 3 (a) Complete the table, identifying the substance liberated at each electrode during electrolysis with inert electrodes.

| electrolyte                           | substance liberated at the anode | substance liberated at the cathode |
|---------------------------------------|----------------------------------|------------------------------------|
| $\text{AgNO}_3(\text{aq})$            |                                  |                                    |
| concentrated $\text{NaCl}(\text{aq})$ |                                  |                                    |
| $\text{CuSO}_4(\text{aq})$            |                                  |                                    |

[3]

- (b) Molten calcium iodide,  $\text{CaI}_2$ , is electrolysed in an inert atmosphere with inert electrodes.

- (i) Write ionic equations for the reactions occurring at the electrodes.

- .....
- .....

[2]

- (ii) The electrolysis of molten  $\text{CaI}_2$  is a redox process.

Identify the ion that is oxidised and the ion that is reduced, explaining your answer by reference to oxidation numbers.

.....  
 .....  
 ..... [2]

- (iii) Describe **two** visual observations that would be made during this electrolysis.

- 1 .....
- 2 ..... [1]

- (c) An oxide of iron dissolved in an inert solvent is electrolysed for 2.00 hours using a current of 0.800A. The electrolysis products are iron and oxygen. The mass of iron produced is 1.11 g.

Calculate the oxidation number of Fe in the oxide of iron. Show **all** your working.

oxidation number of Fe = ..... [3]

[Total: 11]

- 3 (a) Complete the table by predicting the identity of the substance liberated at each electrode during electrolysis with inert electrodes.

| electrolyte                             | substance liberated at the anode | substance liberated at the cathode |
|-----------------------------------------|----------------------------------|------------------------------------|
| NaOH(aq)                                |                                  |                                    |
| dilute $\text{CuCl}_2(\text{aq})$       |                                  |                                    |
| concentrated $\text{MgCl}_2(\text{aq})$ |                                  |                                    |

[3]

- (b) (i) The electrolysis of molten  $\text{ZnBr}_2$  is a redox process.

Identify the ion that is oxidised and the ion that is reduced.

Use ionic half-equations to explain your answer.

.....  
 .....  
 .....  
 .....  
 ..... [3]

- (ii) Describe **one** visual observation that would be made during this electrolysis.

..... [1]

- (c) Dilute sulfuric acid is electrolysed for 50.0 minutes using inert electrodes and a current of 1.20A. A different gas is collected above each electrode. The volumes of the two gases are measured under room conditions.

Calculate the maximum volume of gas that could be collected at the **cathode**.

volume = .....  $\text{cm}^3$  [3]

[Total: 10]

# 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

11. Question | 9701 May-June 2018 P4 V3 Q3

- 3 (a)** Complete the table, identifying the substance liberated at each electrode during electrolysis with inert electrodes.

| electrolyte                           | substance liberated at the anode | substance liberated at the cathode |
|---------------------------------------|----------------------------------|------------------------------------|
| $\text{AgNO}_3(\text{aq})$            |                                  |                                    |
| concentrated $\text{NaCl}(\text{aq})$ |                                  |                                    |
| $\text{CuSO}_4(\text{aq})$            |                                  |                                    |

[3]

- (b)** Molten calcium iodide,  $\text{CaI}_2$ , is electrolysed in an inert atmosphere with inert electrodes.

- (i)** Write ionic equations for the reactions occurring at the electrodes.

- .....
- ..... [2]

- (ii)** The electrolysis of molten  $\text{CaI}_2$  is a redox process.

Identify the ion that is oxidised and the ion that is reduced, explaining your answer by reference to oxidation numbers.

.....  
 .....  
 ..... [2]

- (iii)** Describe **two** visual observations that would be made during this electrolysis.

- 1 .....  
 2 ..... [1]

- (c)** An oxide of iron dissolved in an inert solvent is electrolysed for 2.00 hours using a current of 0.800A. The electrolysis products are iron and oxygen. The mass of iron produced is 1.11 g.

Calculate the oxidation number of Fe in the oxide of iron. Show **all** your working.

oxidation number of Fe = ..... [3]

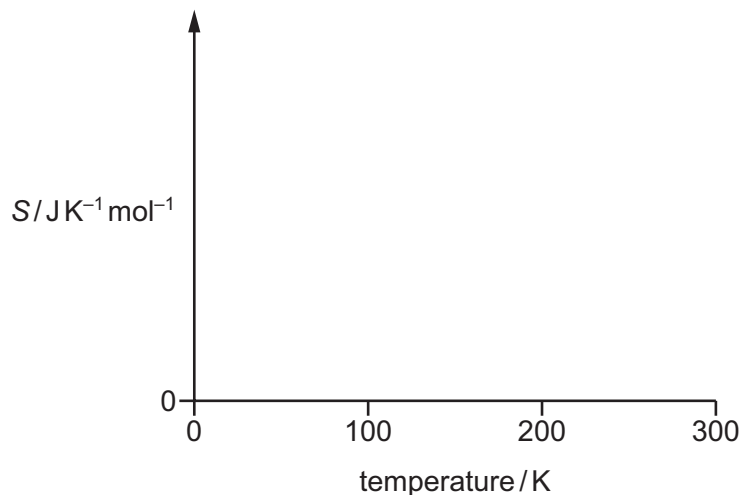
[Total: 11]

12. Question part 1/3 | 9701 Oct-Nov 2018 P4 V1 Q8

8 Entropy is a measure of the disorder of a system.

(a) Assume the entropy,  $S$ , for  $\text{H}_2\text{O}$  is zero at 0 K.

Sketch a graph on the axes to show how the entropy changes for  $\text{H}_2\text{O}$  between 0 K and 300 K.



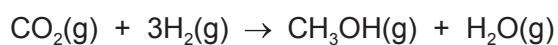
[2]

(b) Place **one tick (✓)** in **each row** of the table to show the sign of the entropy changes,  $\Delta S$ .

|                           | $\Delta S$ is negative | $\Delta S$ is positive |
|---------------------------|------------------------|------------------------|
| solid dissolving in water |                        |                        |
| water boiling to steam    |                        |                        |

[1]

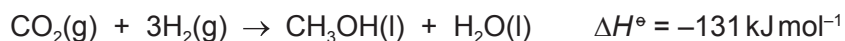
(c) The equation for a reaction that produces methanol is shown.



Use relevant bond energies from the *Data Booklet* to calculate the enthalpy change,  $\Delta H$ , for this gas phase reaction.

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

(d) At 298 K, both products of this reaction are liquid.



Standard entropies are shown in the table.

| substance                               | $\text{CO}_2(\text{g})$ | $\text{H}_2(\text{g})$ | $\text{CH}_3\text{OH}(\text{l})$ | $\text{H}_2\text{O}(\text{l})$ |
|-----------------------------------------|-------------------------|------------------------|----------------------------------|--------------------------------|
| $S^\circ/\text{JK}^{-1}\text{mol}^{-1}$ | +214                    | +131                   | +127                             | +70                            |

(i) Calculate the standard entropy change,  $\Delta S^\circ$ , for this reaction.

$$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1}\text{mol}^{-1} \quad [2]$$

(ii) Calculate the standard Gibbs free energy change,  $\Delta G^\circ$ , for this reaction at 298 K.

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

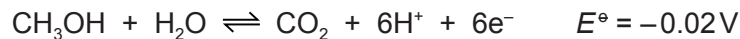
(iii) Predict the effect of increasing the temperature on the feasibility of this reaction.

.....  
..... [1]

12. Question part 3/3 | 9701 Oct-Nov 2018 P4 V1 Q8

- (e) In a methanol-oxygen fuel cell,  $\text{CH}_3\text{OH}(\text{l})$  and  $\text{O}_2(\text{g})$  are in contact with two inert electrodes immersed in an acidic solution.

The half-equation for the reaction at the methanol electrode is shown.



- (i) Use the *Data Booklet* to write an equation for the overall cell reaction.

.....  
.....  
..... [1]

- (ii) Use  $E^\circ$  values to calculate the  $E^\circ_{\text{cell}}$  for this reaction.

$$E^\circ_{\text{cell}} = \dots\dots\dots \text{V} \quad [1]$$

[Total: 12]

13. Question part 1/2 | 9701 May-June 2019 P4 V1 Q3

**3** Chlorate(V) ions are powerful oxidising agents.

- (a)** The reduction of chlorate(V) ions,  $\text{ClO}_3^-$ , with  $\text{SO}_2$  forms chlorine dioxide,  $\text{ClO}_2$ , and sulfate ions,  $\text{SO}_4^{2-}$ , as the only products.

Construct an equation for this reaction.

..... [1]

- (b) (i)** Chlorine dioxide,  $\text{ClO}_2$ , disproportionates with hydroxide ions,  $\text{OH}^-(\text{aq})$ , to form a mixture of  $\text{ClO}_2^-$  and  $\text{ClO}_3^-$  ions.



Explain, using this reaction as an example, what is meant by *disproportionation*.

.....  
 ..... [1]

- (ii)** Deduce the ionic half-equations for the reaction in **(b)(i)**.

.....  
 ..... [2]

- (c)** A lithium-iodine electrochemical cell can be used to generate electricity for a heart pacemaker. The cell consists of a lithium electrode and an inert electrode immersed in body fluids. When current flows lithium is oxidised and iodine is reduced.

- (i)** Use the *Data Booklet* to write half-equations for the reactions taking place at the two electrodes. Hence write the overall equation for when a current flows.

- .....
- .....

overall equation ..... [2]

- (ii)** Use the *Data Booklet* to calculate the  $E_{\text{cell}}^\ominus$  for this cell.

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

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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13. Question part 2/2 | 9701 May-June 2019 P4 V1 Q3

(iii) A current of  $2.5 \times 10^{-5} \text{ A}$  is drawn from this cell.

Calculate the time taken for 0.10 g of lithium electrode to be used up. Assume the current remains constant throughout this period.

time = ..... s [3]

[Total: 10]

14. Question part 1/2 | 9701 May-June 2019 P4 V3 Q3

**3** Chlorate(V) ions are powerful oxidising agents.

- (a)** The reduction of chlorate(V) ions,  $\text{ClO}_3^-$ , with  $\text{SO}_2$  forms chlorine dioxide,  $\text{ClO}_2$ , and sulfate ions,  $\text{SO}_4^{2-}$ , as the only products.

Construct an equation for this reaction.

..... [1]

- (b) (i)** Chlorine dioxide,  $\text{ClO}_2$ , disproportionates with hydroxide ions,  $\text{OH}^-(\text{aq})$ , to form a mixture of  $\text{ClO}_2^-$  and  $\text{ClO}_3^-$  ions.



Explain, using this reaction as an example, what is meant by *disproportionation*.

.....  
 ..... [1]

- (ii)** Deduce the ionic half-equations for the reaction in **(b)(i)**.

.....  
 ..... [2]

- (c)** A lithium-iodine electrochemical cell can be used to generate electricity for a heart pacemaker. The cell consists of a lithium electrode and an inert electrode immersed in body fluids. When current flows lithium is oxidised and iodine is reduced.

- (i)** Use the *Data Booklet* to write half-equations for the reactions taking place at the two electrodes. Hence write the overall equation for when a current flows.

- .....
- .....

overall equation ..... [2]

- (ii)** Use the *Data Booklet* to calculate the  $E_{\text{cell}}^\ominus$  for this cell.

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

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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14. Question part 2/2 | 9701 May-June 2019 P4 V3 Q3

(iii) A current of  $2.5 \times 10^{-5} \text{ A}$  is drawn from this cell.

Calculate the time taken for 0.10 g of lithium electrode to be used up. Assume the current remains constant throughout this period.

time = ..... s [3]

[Total: 10]

15. Question part 1/2 | 9701 Oct-Nov 2019 P4 V1 Q1

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

- an  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell
- an  $\text{Al}^{3+}/\text{Al}$  half-cell

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

- $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell .....
- $\text{Al}^{3+}/\text{Al}$  half-cell .....

[1]

(b) The cell is operated at 298 K.

The  $\text{Al}^{3+}/\text{Al}$  half-cell has standard concentrations.

The  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell has  $[\text{Sn}^{4+}] = 0.300 \text{ mol dm}^{-3}$  and  $[\text{Sn}^{2+}] = 0.150 \text{ mol dm}^{-3}$ .

(i) Use the Nernst equation to calculate the electrode potential,  $E$ , of the  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell under these conditions.

$E = \dots\dots\dots \text{ V}$  [2]

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

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

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

..... [2]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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15. Question part 2/2 | 9701 Oct-Nov 2019 P4 V1 Q1

- (c) Aluminium is produced industrially by electrolysis of a melt containing large amounts of  $\text{Al}^{3+}$  ions.

Calculate the mass of aluminium that is obtained when a current of 300 000 A is passed for 24 hours. Give your answer to **three** significant figures.

mass = ..... units = .....  
[4]

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

.....  
.....  
.....  
..... [2]

[Total: 12]

16. Question part 1/2 | 9701 Oct-Nov 2019 P4 V2 Q1

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

- a  $\text{Br}_2/\text{Br}^-$  half-cell
- an  $\text{Mn}^{3+}/\text{Mn}^{2+}$  half-cell

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

$\text{Br}_2/\text{Br}^-$  half-cell .....

$\text{Mn}^{3+}/\text{Mn}^{2+}$  half-cell .....

[1]

(b) The cell is operated at 298 K.

The  $\text{Br}_2/\text{Br}^-$  half-cell has standard concentrations.

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

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

$E = \dots\dots\dots \text{ V}$  [2]

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

$E_{\text{cell}} = \dots\dots\dots \text{ V}$  [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  $-1.60 \times 10^{-19} \text{ C}$ .

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.

$L = \dots\dots\dots \text{ mol}^{-1}$  [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]

17. Question part 1/2 | 9701 Oct-Nov 2019 P4 V3 Q1

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

- an  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell
- an  $\text{Al}^{3+}/\text{Al}$  half-cell

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

- $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell .....
- $\text{Al}^{3+}/\text{Al}$  half-cell .....

[1]

(b) The cell is operated at 298 K.

The  $\text{Al}^{3+}/\text{Al}$  half-cell has standard concentrations.

The  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell has  $[\text{Sn}^{4+}] = 0.300 \text{ mol dm}^{-3}$  and  $[\text{Sn}^{2+}] = 0.150 \text{ mol dm}^{-3}$ .

- (i) Use the Nernst equation to calculate the electrode potential,  $E$ , of the  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell under these conditions.

$E = \dots\dots\dots \text{ V}$  [2]

- (ii) Calculate the  $E_{\text{cell}}$  under these conditions.

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

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

..... [2]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

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17. Question part 2/2 | 9701 Oct-Nov 2019 P4 V3 Q1

- (c) Aluminium is produced industrially by electrolysis of a melt containing large amounts of  $\text{Al}^{3+}$  ions.

Calculate the mass of aluminium that is obtained when a current of 300 000 A is passed for 24 hours. Give your answer to **three** significant figures.

mass = ..... units = .....  
[4]

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

.....  
.....  
.....  
..... [2]

[Total: 12]

18. Question part 1/5 | 9701 Feb-March 2020 P4 V2 Q3

**3** Gold is an unreactive metal that can only be oxidised under specific conditions.

**(a)** The standard electrode potential,  $E^\ominus$ , of  $\text{Au}^{3+}(\text{aq})/\text{Au}(\text{s})$  is +1.50 V.

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

.....  
.....  
..... [2]

**(ii)** Draw a fully labelled diagram of the apparatus that should be used to measure the standard cell potential,  $E^\ominus_{\text{cell}}$ , of  $\text{Au}^{3+}(\text{aq})/\text{Au}(\text{s})$  and  $\text{HNO}_3(\text{aq})/\text{NO}(\text{g})$ .

Include all necessary chemicals.

[4]

Some relevant half-equations and their standard electrode potentials are given.

|   | half-equation                                                                                                                              | $E^\ominus/\text{V}$ |
|---|--------------------------------------------------------------------------------------------------------------------------------------------|----------------------|
| 1 | $\text{Au}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Au}(\text{s})$                                                           | +1.50                |
| 2 | $[\text{AuCl}_4]^- (\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Au}(\text{s}) + 4\text{Cl}^- (\text{aq})$                            | +1.00                |
| 3 | $\text{NO}_3^- (\text{aq}) + 4\text{H}^+ (\text{aq}) + 3\text{e}^- \rightleftharpoons \text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ | +0.96                |

- (iii) Write an ionic equation to show the spontaneous reaction that occurs when an electric current is drawn from the cell in (a)(ii).

..... [1]

- (iv) Calculate the  $E^\ominus_{\text{cell}}$  of the reaction in (a)(iii).

$E^\ominus_{\text{cell}} = \dots\dots\dots \text{V}$  [1]

- (v) Gold can be oxidised by a mixture of concentrated hydrochloric acid and concentrated nitric acid, known as *aqua regia*. Concentrated hydrochloric acid is  $12\text{mol dm}^{-3}$ . Concentrated nitric acid is  $16\text{mol dm}^{-3}$ .

Explain why *aqua regia* is able to dissolve gold.

In your answer, state and explain what effect the use of concentrated hydrochloric acid and concentrated nitric acid have on the  $E$  values of half-equations 2 and 3.

.....

.....

.....

.....

.....

..... [3]

- (b) Aqueous gold(III) chloride,  $\text{AuCl}_3$ , reacts with aqueous hydrogen peroxide,  $\text{H}_2\text{O}_2$ , under certain conditions, forming  $\text{Au}$ ,  $\text{O}_2$  and  $\text{HCl}$ .

A student carries out separate experiments using different initial concentrations of  $\text{AuCl}_3$  and  $\text{H}_2\text{O}_2$ . The initial rate of each reaction is measured.

The table shows the results that are obtained.

| experiment | $[\text{AuCl}_3]$<br>/ $\text{mol dm}^{-3}$ | $[\text{H}_2\text{O}_2]$<br>/ $\text{mol dm}^{-3}$ | rate of production of $\text{O}_2(\text{g})$<br>/ $\text{dm}^3 \text{ minute}^{-1}$ |
|------------|---------------------------------------------|----------------------------------------------------|-------------------------------------------------------------------------------------|
| 1          | 0.05                                        | 0.50                                               | $7.66 \times 10^{-2}$                                                               |
| 2          | 0.10                                        | 0.50                                               | $1.53 \times 10^{-1}$                                                               |
| 3          | 0.15                                        | 1.00                                               | $4.60 \times 10^{-1}$                                                               |

- (i) Write an equation for the reaction of  $\text{AuCl}_3$  with  $\text{H}_2\text{O}_2$ .

..... [1]

- (ii) Determine the rate equation of the reaction.

Show your reasoning, quoting data from the table.

.....  
 .....  
 .....  
 .....  
 ..... [3]

- (iii) Use the results of experiment 2 to calculate the value of the rate constant,  $k$ , for this reaction.

Include the units of  $k$ .

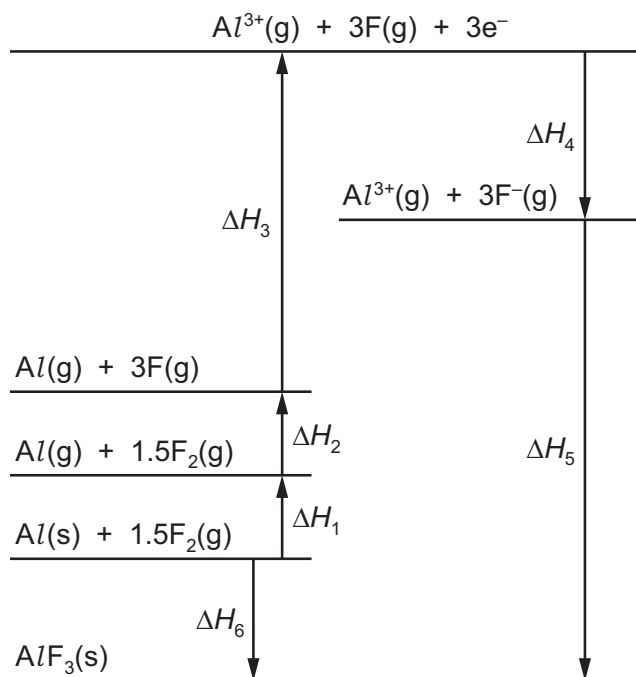
rate constant,  $k$  = .....

units = .....

[2]

(c)  $\text{AlF}_3$  is an ionic compound.

The Born–Haber cycle for the formation of  $\text{AlF}_3$  is shown.



(i) Name the enthalpy changes labelled  $\Delta H_4$  and  $\Delta H_6$ .

$\Delta H_4 =$  .....

$\Delta H_6 =$  .....

[2]

(ii) Use the data in the table and data from the *Data Booklet* to calculate the lattice energy of  $\text{AlF}_3$ .

| process                                                                     | enthalpy change<br>/ $\text{kJ mol}^{-1}$ |
|-----------------------------------------------------------------------------|-------------------------------------------|
| $\text{Al(s)} \rightarrow \text{Al(g)}$                                     | +326                                      |
| $\text{Al(g)} \rightarrow \text{Al}^{3+}(\text{g})$                         | +5137                                     |
| $\text{F(g)} \rightarrow \text{F}^-(\text{g})$                              | −328                                      |
| $\text{Al(s)} + 1.5\text{F}_2(\text{g}) \rightarrow \text{AlF}_3(\text{s})$ | −1504                                     |

lattice energy of  $\text{AlF}_3 =$  .....  $\text{kJ mol}^{-1}$  [2]

- (iii) Scandium fluoride,  $\text{ScF}_3$ , is an ionic compound.

Use data from the *Data Booklet* to suggest how the lattice energy of  $\text{AlF}_3$  compares with the lattice energy of  $\text{ScF}_3$ .

Explain your answer.

.....  
.....  
..... [2]

- (d)  $\text{AlF}_3$  is sparingly soluble in water. The concentration of its saturated solution at 298 K is  $6.5 \times 10^{-2} \text{ mol dm}^{-3}$ .

- (i) Write an expression for the solubility product,  $K_{\text{sp}}$ , of  $\text{AlF}_3$ .

$K_{\text{sp}} =$  ..... [1]

- (ii) Calculate the numerical value of  $K_{\text{sp}}$  for  $\text{AlF}_3$  at 298 K.

$K_{\text{sp}} =$  ..... [1]

[Total: 25]

19. Question | 9701 May-June 2020 P4 V1 Q3

- 3 The overall reaction for photosynthesis is shown.



Water is oxidised in this process according to the following half-equation.



- (a) (i) Use these equations to deduce the half-equation for the reduction of carbon dioxide in this process.

[2]

- (ii) Draw a fully labelled diagram of the apparatus that should be used to measure the standard electrode potential,  $E^\ominus$ , of  $\text{O}_2(\text{g})$  in half-equation 1 under standard conditions. Include all necessary chemicals.

[4]

- (iii) For the cell drawn in (a)(ii), use the *Data Booklet* to calculate the  $E^\ominus_{\text{cell}}$  and deduce which electrode is positive.

$$E^\ominus_{\text{cell}} = \dots\dots\dots \text{ V}$$

identity of the positive electrode =  $\dots\dots\dots$   
[1]

[Total: 7]

- 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,  $E_{\text{cell}}^{\ominus}$ , made of a  $\text{Co}^{3+}/\text{Co}^{2+}$  half-cell and a  $\text{Cl}_2/\text{Cl}^-$  half-cell.

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

| half-cell                       | electrode |
|---------------------------------|-----------|
| $\text{Co}^{3+}/\text{Co}^{2+}$ |           |
| $\text{Cl}_2/\text{Cl}^-$       |           |

[1]

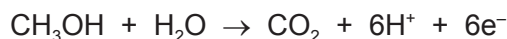
- (iii) Use data from the *Data Booklet* to calculate the  $E_{\text{cell}}^{\ominus}$ .

$$E_{\text{cell}}^{\ominus} = \dots\dots\dots \text{V} \quad [1]$$

- (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.



Calculate the volume of  $\text{CO}_2$ , in  $\text{cm}^3$ , formed when a current of 2.5A is delivered by the cell for 30 minutes. Assume the cell is operated at room conditions.

$$\text{volume of CO}_2 = \dots\dots\dots \text{cm}^3 \quad [2]$$

[Total: 7]

21. Question | 9701 May-June 2020 P4 V3 Q3

- 3 The overall reaction for photosynthesis is shown.



Water is oxidised in this process according to the following half-equation.



- (a) (i) Use these equations to deduce the half-equation for the reduction of carbon dioxide in this process.

[2]

- (ii) Draw a fully labelled diagram of the apparatus that should be used to measure the standard electrode potential,  $E^\ominus$ , of  $\text{O}_2(\text{g})$  in half-equation 1 under standard conditions. Include all necessary chemicals.

[4]

- (iii) For the cell drawn in (a)(ii), use the *Data Booklet* to calculate the  $E^\ominus_{\text{cell}}$  and deduce which electrode is positive.

$$E^\ominus_{\text{cell}} = \dots\dots\dots \text{ V}$$

identity of the positive electrode =  $\dots\dots\dots$   
[1]

[Total: 7]

22. Question part 1/2 | 9701 May-June 2020 P4 V3 Q7

- 7 (a) Silver carbonate,  $\text{Ag}_2\text{CO}_3$ , is sparingly soluble in water. The numerical value of the solubility product,  $K_{\text{sp}}$ , for silver carbonate is  $6.3 \times 10^{-12}$  at  $25^\circ\text{C}$ .

(i) Write an expression for the solubility product,  $K_{\text{sp}}$ , of  $\text{Ag}_2\text{CO}_3$ , and state its units.

$$K_{\text{sp}} =$$

units = ..... [2]

(ii) Calculate the equilibrium concentration of  $\text{Ag}^+$  in a saturated solution of  $\text{Ag}_2\text{CO}_3$  at  $25^\circ\text{C}$ .

$$[\text{Ag}^+] = \dots\dots\dots \text{mol dm}^{-3} \quad [1]$$

(iii) Solid  $\text{Ag}_2\text{CO}_3$  is stirred at  $25^\circ\text{C}$  with  $0.050 \text{ mol dm}^{-3} \text{ AgNO}_3$  until no more  $\text{Ag}_2\text{CO}_3$  dissolves.

Calculate the concentration of carbonate ions,  $[\text{CO}_3^{2-}]$ , in this solution.

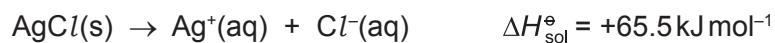
$$[\text{CO}_3^{2-}] = \dots\dots\dots \text{mol dm}^{-3} \quad [1]$$

(iv) An electrochemical cell is set up to measure the electrode potential,  $E$ , for the  $\text{Ag}^+/\text{Ag}$  half-cell using the saturated  $\text{Ag}_2\text{CO}_3(\text{aq})$  with a standard hydrogen electrode.

Use the *Data Booklet*, your answer to (a)(ii), and the Nernst equation to calculate the electrode potential,  $E$ , for this  $\text{Ag}^+/\text{Ag}$  half-cell.

$$E \text{ for } \text{Ag}^+/\text{Ag} \text{ half-cell} = \dots\dots\dots \text{V} \quad [2]$$

- (b) Silver chloride,  $\text{AgCl}$ , is sparingly soluble in water. The equation for the enthalpy change of solution is shown.



Standard entropies are shown in the table.

| species                                          | $\text{AgCl(s)}$ | $\text{Ag}^{\text{+}}(\text{aq})$ | $\text{Cl}^{-}(\text{aq})$ |
|--------------------------------------------------|------------------|-----------------------------------|----------------------------|
| $S^{\ominus} / \text{J K}^{-1} \text{ mol}^{-1}$ | +96.2            | +72.7                             | +56.5                      |

- (i) Calculate the standard entropy change of solution,  $\Delta S^{\ominus}$ .

$$\Delta S^{\ominus} = \dots\dots\dots \text{J K}^{-1} \text{ mol}^{-1} \quad [1]$$

- (ii) Explain, with the aid of a calculation, why  $\text{AgCl}$  is insoluble in water at  $25^{\circ}\text{C}$ .

You should use data from this question and your answer to (b)(i).

.....  
 ..... [3]

[Total: 10]

## 9701 Chemistry Paper 4 - Electrochemistry

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23. Question part 1/2 | 9701 Oct-Nov 2020 P4 V1 Q3

- 3 (a)** Identify the substances liberated at the anode and at the cathode during the electrolysis of aqueous sodium sulfate,  $\text{Na}_2\text{SO}_4(\text{aq})$ .

anode .....

cathode .....

[1]

- (b)** When molten sodium chloride is electrolysed, chlorine is liberated at the anode and sodium is liberated at the cathode.

A sample of molten sodium chloride is electrolysed for 1.50 hours using a current of 4.50A.

Calculate the volume of chlorine and the mass of sodium that are liberated under room conditions.

volume of chlorine = .....  $\text{dm}^3$

mass of sodium = ..... g

[4]

- (c) The equation representing the standard electrode potential,  $E^\circ$ , for the reduction of  $\text{MnO}_4^-(\text{aq})$  to  $\text{Mn}^{2+}(\text{aq})$  in acid solution is given.



- (i) Draw a diagram of the apparatus that would be used to measure the  $E^\circ$  value of this half-cell. Your diagram should be fully labelled to identify all apparatus, substances and conditions.

[4]

- (ii) Use the *Data Booklet* to identify a substance that could be used to oxidise  $\text{Mn}^{2+}$  ions to  $\text{MnO}_4^-$  ions under standard conditions.

Write an equation for the reaction.

.....

.....

..... [2]

[Total: 11]

24. Question part 1/2 | 9701 Oct-Nov 2020 P4 V2 Q4

- 4 (a) Identify the substances liberated at the anode and at the cathode during the electrolysis of saturated  $\text{KCl(aq)}$ .

at the anode .....

at the cathode .....

[1]

- (b) When dilute sulfuric acid is electrolysed, oxygen is liberated at the anode.

Dilute sulfuric acid is electrolysed for 15.0 minutes using a current of 0.750 A.

Calculate the volume of oxygen that is liberated under room conditions.

volume of oxygen = .....  $\text{cm}^3$  [3]

- (c) The halogens chlorine, bromine and iodine differ in their strengths as oxidising agents. These strengths are indicated by the  $E^\circ$  values for these halogens.

- (i) Give the  $E^\circ$  values for chlorine, bromine and iodine acting as oxidising agents.

..... [1]

- (ii) Deduce which of chlorine, bromine and iodine will react with a solution of  $\text{Sn}^{2+}(\text{aq})$  under standard conditions.

Explain your answer. Include a relevant equation in your explanation.

.....

.....

..... [3]

- (iii) An excess of chlorine is added to a solution of acidified  $\text{Mn}^{2+}(\text{aq})$  under standard conditions.

Give the formula of the product of this reaction that contains manganese.

..... [1]

## 9701 Chemistry Paper 4 - Electrochemistry

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24. Question part 2/2 | 9701 Oct-Nov 2020 P4 V2 Q4

- (d) An electrochemical cell can be made by connecting an  $\text{Fe}^{3+}/\text{Fe}^{2+}$  half-cell to an  $\text{S}_2\text{O}_8^{2-}/\text{SO}_4^{2-}$  half-cell under standard conditions.

- (i) Calculate the standard cell potential of this electrochemical cell.

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

- (ii) State the material that should be used as the electrode in each half-cell.

in the  $\text{Fe}^{3+}/\text{Fe}^{2+}$  half-cell .....

in the  $\text{S}_2\text{O}_8^{2-}/\text{SO}_4^{2-}$  half-cell .....  
[1]

- (iii) Describe **one** change to each half-cell that would **increase** the value of the cell potential. The temperature should remain at 298 K.

$\text{Fe}^{3+}/\text{Fe}^{2+}$  half-cell .....

.....

$\text{S}_2\text{O}_8^{2-}/\text{SO}_4^{2-}$  half-cell .....

.....

[1]

[Total: 12]

## 9701 Chemistry Paper 4 - Electrochemistry

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25. Question part 1/2 | 9701 Oct-Nov 2020 P4 V3 Q3

- 3 (a)** Identify the substances liberated at the anode and at the cathode during the electrolysis of aqueous sodium sulfate,  $\text{Na}_2\text{SO}_4(\text{aq})$ .

anode .....

cathode .....

[1]

- (b)** When molten sodium chloride is electrolysed, chlorine is liberated at the anode and sodium is liberated at the cathode.

A sample of molten sodium chloride is electrolysed for 1.50 hours using a current of 4.50 A.

Calculate the volume of chlorine and the mass of sodium that are liberated under room conditions.

volume of chlorine = .....  $\text{dm}^3$

mass of sodium = ..... g

[4]

- (c) The equation representing the standard electrode potential,  $E^\circ$ , for the reduction of  $\text{MnO}_4^-(\text{aq})$  to  $\text{Mn}^{2+}(\text{aq})$  in acid solution is given.



- (i) Draw a diagram of the apparatus that would be used to measure the  $E^\circ$  value of this half-cell. Your diagram should be fully labelled to identify all apparatus, substances and conditions.

[4]

- (ii) Use the *Data Booklet* to identify a substance that could be used to oxidise  $\text{Mn}^{2+}$  ions to  $\text{MnO}_4^-$  ions under standard conditions.

Write an equation for the reaction.

.....

.....

..... [2]

[Total: 11]

26. Question part 1/3 | 9701 May-June 2021 P4 V2 Q3

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

.....

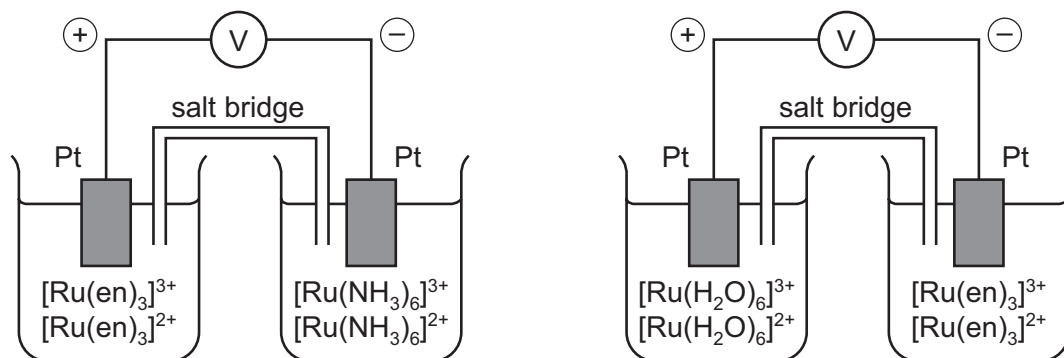
.....

..... [2]

Three redox systems, **A**, **B** and **C**, are shown. The ligand 1,2-diaminoethane,  $\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$ , is represented by en.

|          |                                                                                                                 |
|----------|-----------------------------------------------------------------------------------------------------------------|
| <b>A</b> | $[\text{Ru}(\text{H}_2\text{O})_6]^{3+} + \text{e}^- \rightleftharpoons [\text{Ru}(\text{H}_2\text{O})_6]^{2+}$ |
| <b>B</b> | $[\text{Ru}(\text{NH}_3)_6]^{3+} + \text{e}^- \rightleftharpoons [\text{Ru}(\text{NH}_3)_6]^{2+}$               |
| <b>C</b> | $[\text{Ru}(\text{en})_3]^{3+} + \text{e}^- \rightleftharpoons [\text{Ru}(\text{en})_3]^{2+}$                   |

Two electrochemical cells are set up to compare the standard electrode potentials,  $E^\circ$ , of three half-cells. The diagrams show the relative potential of each electrode.

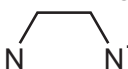


(ii) Use this information to complete the table by adding the labels **A**, **B** and **C** to deduce the order of  $E^\circ$  for the three half-cells.

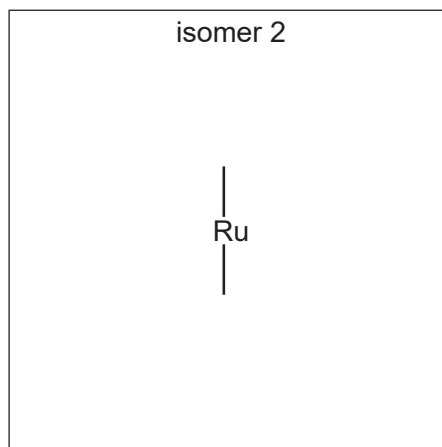
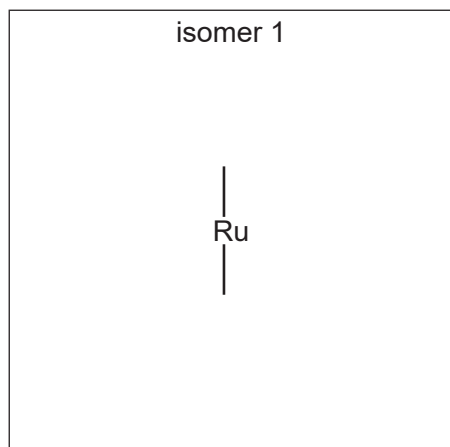
| $E^\circ$      | redox system |
|----------------|--------------|
| most negative  |              |
| ↑              |              |
| least negative |              |

[1]

- (iii) The complex  $[\text{Ru}(\text{en})_3]^{3+}$  shows stereoisomerism. The ligand en is bidentate.

Draw three-dimensional diagrams to show the two isomers of  $[\text{Ru}(\text{en})_3]^{3+}$ . Represent the ligand en by using .

Name the type of stereoisomerism.



type of stereoisomerism ..... [3]

- (b) (i) An electrochemical cell consists of a  $\text{Br}_2/\text{Br}^-$  half-cell and a  $\text{Ag}^+/\text{Ag}$  half-cell, under standard conditions.

Use the *Data Booklet* to calculate the  $E^\ominus_{\text{cell}}$ . Deduce the direction of electron flow in the wire through the voltmeter between these two half-cells.

$$E^\ominus_{\text{cell}} = \dots\dots\dots \text{ V}$$

direction of electron flow from ..... to ..... [1]

- (ii) Water is added to the  $\text{Ag}^+/\text{Ag}$  half-cell in (b)(i).

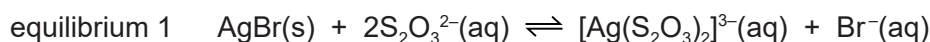
Suggest the effect of this addition on the  $E_{\text{cell}}$ . Place a tick (✓) in the appropriate box.

| less positive | no change | more positive |
|---------------|-----------|---------------|
|               |           |               |

Explain your answer.

.....  
 .....  
 ..... [2]

- (c) Silver bromide, AgBr, dissolves in an aqueous solution of  $\text{S}_2\text{O}_3^{2-}$  ions to form the complex ion  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$ . The  $\text{S}_2\text{O}_3^{2-}$  ions act as monodentate ligands.



- (i) Define the term *ligand*.

.....  
 ..... [1]

- (ii) Write an expression for the equilibrium constant,  $K_c$ , for equilibrium 1.

$K_c =$  ..... [1]

- (iii) Some additional data are given about the dissolution of AgBr in  $\text{S}_2\text{O}_3^{2-}(\text{aq})$ .

| equilibrium constant                                                                  | numerical value       |
|---------------------------------------------------------------------------------------|-----------------------|
| solubility product, $K_{\text{sp}}$ , of AgBr                                         | $5.4 \times 10^{-13}$ |
| stability constant, $K_{\text{stab}}$ , of $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$ | $2.9 \times 10^{13}$  |

Use your answer to (c)(ii) and these data to calculate  $K_c$  for equilibrium 1. Include the units for  $K_c$ .

$K_c =$  ..... units ..... [2]

- (d) The numerical values for the stability constants,  $K_{\text{stab}}$ , of two other silver(I) complexes are given.

| silver(I) complex                | numerical value of $K_{\text{stab}}$ |
|----------------------------------|--------------------------------------|
| $[\text{Ag}(\text{CN})_2]^{-}$   | $5.3 \times 10^{18}$                 |
| $[\text{Ag}(\text{NH}_3)_2]^{+}$ | $1.6 \times 10^7$                    |

An aqueous solution containing  $\text{Ag}^{+}$  is added to a solution containing equal concentrations of  $\text{CN}^{-}(\text{aq})$ ,  $\text{NH}_3(\text{aq})$  and  $\text{S}_2\text{O}_3^{2-}(\text{aq})$ . The mixture is left to reach equilibrium.

Deduce the relative concentrations of  $[\text{Ag}(\text{CN})_2]^{-}$ ,  $[\text{Ag}(\text{NH}_3)_2]^{+}$  and  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$  present in the resulting mixture. Explain your answer.

..... > ..... > .....  
 highest concentration ..... lowest concentration  
 .....  
 ..... [2]

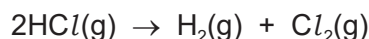
[Total: 15]

27. Question part 1/2 | 9701 Oct-Nov 2021 P4 V2 Q3

- 3** Pure water is a very poor conductor of electricity. However, when hydrogen chloride gas is dissolved in water, ions are formed and a current flows during electrolysis.



The overall change after electrolysis is that hydrogen chloride gas is converted into hydrogen and chlorine.



When a current of 3.10A is passed through the solution for  $Y$  minutes,  $351\text{ cm}^3$  of chlorine are produced at the anode, measured under room conditions.

- (a)** Calculate the number of chlorine molecules produced during the electrolysis.

number of chlorine molecules = ..... [2]

- (b)** Calculate the total number of electrons transferred to produce this number of chlorine molecules.

total number of electrons = ..... [1]

- (c)** Calculate the quantity of charge, in coulombs, of the total number of electrons calculated in **(b)**.

quantity of charge = ..... C [1]

- (d)** Calculate the time,  $Y$ , in minutes, for which the current flows.

$Y$  = ..... minutes [1]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

27. Question part 2/2 | 9701 Oct-Nov 2021 P4 V2 Q3

(e) The standard entropies,  $S^\circ$ , of three species are given in the table.

| species                 | $S^\circ / \text{JK}^{-1} \text{mol}^{-1}$ |
|-------------------------|--------------------------------------------|
| $\text{HCl(g)}$         | +187                                       |
| $\text{H}_2(\text{g})$  | +131                                       |
| $\text{Cl}_2(\text{g})$ | +223                                       |

(i) Calculate  $\Delta S^\circ$  for the reaction  $2\text{HCl(g)} \rightarrow \text{H}_2(\text{g}) + \text{Cl}_2(\text{g})$ .

$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1} \text{mol}^{-1}$  [1]

(ii)  $\Delta H^\circ$  for the reaction  $2\text{HCl(g)} \rightarrow \text{H}_2(\text{g}) + \text{Cl}_2(\text{g})$  is  $+185 \text{ kJ mol}^{-1}$ .

Calculate  $\Delta G^\circ$  for this reaction at 298 K.

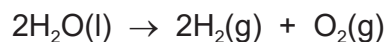
$\Delta G^\circ = \dots\dots\dots \text{kJ mol}^{-1}$  [2]

(iii) Predict the effect of increasing temperature on the spontaneity of this reaction. Explain your answer.

.....  
..... [1]

[Total: 9]

- 1** When dilute sulfuric acid is electrolysed, water is split into hydrogen and oxygen.



A current of **x**A is passed through the solution for 14.0 minutes. 462 cm<sup>3</sup> of hydrogen are produced at the cathode, measured under room conditions.

- (a)** Calculate the number of hydrogen molecules produced during the electrolysis.

number of hydrogen molecules = ..... [2]

- (b)** Calculate the total number of electrons transferred to produce this number of hydrogen molecules.

total number of electrons = ..... [1]

- (c)** Calculate the quantity of charge, in coulombs, of the total number of electrons calculated in **(b)**.

quantity of charge = ..... C [1]

- (d)** Calculate the current, **x**, passed during this experiment.

**x** = ..... A [1]

## 9701 Chemistry Paper 4 - Electrochemistry

Topical questions | 42 questions

28. Question part 2/2 | 9701 Oct-Nov 2021 P4 V3 Q1

(e) The standard entropies,  $S^\circ$ , of three species are given in the table.

| species                        | $S^\circ / \text{JK}^{-1} \text{mol}^{-1}$ |
|--------------------------------|--------------------------------------------|
| $\text{H}_2\text{O}(\text{l})$ | +70                                        |
| $\text{H}_2(\text{g})$         | +131                                       |
| $\text{O}_2(\text{g})$         | +205                                       |

(i) Calculate  $\Delta S^\circ$  for the reaction  $2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$ .

$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1} \text{mol}^{-1}$  [1]

(ii)  $\Delta H^\circ$  for the reaction  $2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$  is  $+572 \text{ kJ mol}^{-1}$ .

Calculate  $\Delta G^\circ$  for this reaction at 298 K.

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

(iii) Predict the effect of increasing temperature on the spontaneity of this reaction.  
Explain your answer.

.....  
..... [1]

[Total: 9]

29. Question part 1/3 | 9701 May-June 2022 P4 V1 Q3

- 3 (a) The rate of reaction between 2-chloro-2-methylpropane,  $(\text{CH}_3)_3\text{CCl}$ , and methanol is investigated. When a large excess of methanol is used, the overall reaction is first order.



Fig. 3.1 shows the results obtained.

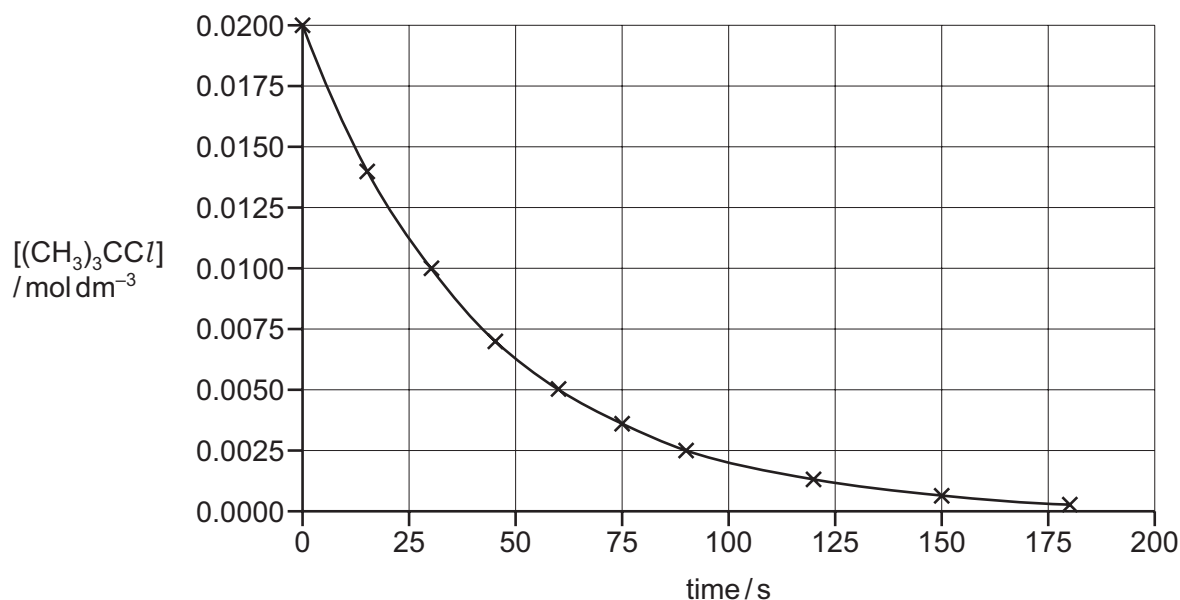


Fig. 3.1

- (i) Use the graph to determine the rate of reaction at 40 s. Show all your working.

rate = .....  $\text{mol dm}^{-3} \text{s}^{-1}$  [1]

- (ii) Use the graph to show that the overall reaction is first order. Explain your answer.

.....  
 .....  
 ..... [2]

- (b) In a different reaction, which is also a first order reaction, 75% of the reactant is consumed in 320 s.

Calculate the rate constant,  $k$ , for this reaction. State the units for  $k$ .

$k$  = ..... units = ..... [2]

- (c) (i) Define standard electrode potential,  $E^\circ$ .

.....  
.....  
..... [1]

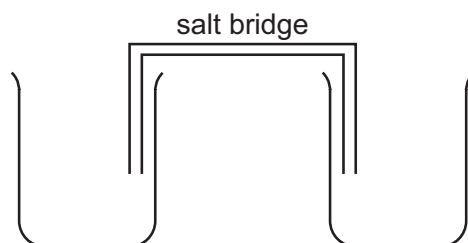
- (ii) A salt bridge is used in an electrochemical cell.

State the function of the salt bridge. Explain your answer.

.....  
..... [1]

- (iii) Complete the diagram of the apparatus that can be used to measure the  $E^\circ$  of the  $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ ,  $\text{H}^+(\text{aq})/\text{Cr}^{3+}(\text{aq})$  electrode against the standard hydrogen electrode.

Your diagram should be fully labelled to identify all apparatus, substances and conditions.



[3]

- (iv) The  $E^\circ$  of the  $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ ,  $\text{H}^+(\text{aq})/\text{Cr}^{3+}(\text{aq})$  electrode is +1.33 V.

Label the negative electrode and the direction of electron flow in the external circuit when the current flows in your diagram in (c)(iii). [1]

- (d) Table 3.1 lists relevant electrode potentials for some electrode reactions for use in (d)(i) and (d)(ii).

Table 3.1

| electrode reaction                                                                                                 | $E^\circ/V$ |
|--------------------------------------------------------------------------------------------------------------------|-------------|
| $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ | +1.33       |
| $\text{CH}_3\text{CHO} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CH}_3\text{CH}_2\text{OH}$             | -0.61       |
| $\text{CH}_3\text{COOH} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CH}_3\text{CHO} + \text{H}_2\text{O}$ | -0.94       |
| $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$                                    | +1.23       |

- (i) Ethanal is oxidised to ethanoic acid in the presence of  $\text{Cr}_2\text{O}_7^{2-}$  ions.

Construct the ionic equation for the oxidation of ethanal to ethanoic acid using dichromate(VI) in acid conditions. Calculate the  $E^\circ_{\text{cell}}$  for this reaction.

ionic equation .....

$$E^\circ_{\text{cell}} = \dots\dots\dots \text{V} \quad [2]$$

- (ii) In an ethanol-oxygen fuel cell,  $\text{CH}_3\text{CH}_2\text{OH}(\text{l})$  and  $\text{O}_2(\text{g})$  are in contact with two inert electrodes immersed in an acidic solution.

The cell reaction for the oxidation of ethanol by oxygen is shown.



Calculate  $\Delta G^\circ$ , in  $\text{kJ mol}^{-1}$ , for the oxidation of ethanol by oxygen.

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

[Total: 15]

- 5 (a)** Complete Table 5.1 to predict the substance liberated at each electrode during electrolysis of the indicated electrolyte with inert electrodes.

**Table 5.1**

| electrolyte                           | substance liberated at the anode | substance liberated at the cathode |
|---------------------------------------|----------------------------------|------------------------------------|
| $\text{PbBr}_2(\text{l})$             |                                  |                                    |
| concentrated $\text{NaCl}(\text{aq})$ |                                  |                                    |
| $\text{Cu}(\text{NO}_3)_2(\text{aq})$ |                                  |                                    |

[3]

- (b)** An electrolytic cell is set up to determine a value for the Avogadro constant,  $L$ . The electrolyte is dilute sulfuric acid and both electrodes are copper.

When a current of 0.600A is passed through the acid for 30.0 minutes, the anode decreases in mass by 0.350g.

- (i)** State the relationship between the Faraday constant,  $F$ , and the Avogadro constant,  $L$ .

..... [1]

- (ii)** Use the experimental information in **(b)** and data from the table on page 23 to calculate a value for the Avogadro constant,  $L$ .

Show all working.

Avogadro constant,  $L =$  ..... [4]

[Total: 8]

31. Question part 1/3 | 9701 May-June 2022 P4 V3 Q3

- 3 (a) The rate of reaction between 2-chloro-2-methylpropane,  $(\text{CH}_3)_3\text{CCl}$ , and methanol is investigated. When a large excess of methanol is used, the overall reaction is first order.



Fig. 3.1 shows the results obtained.

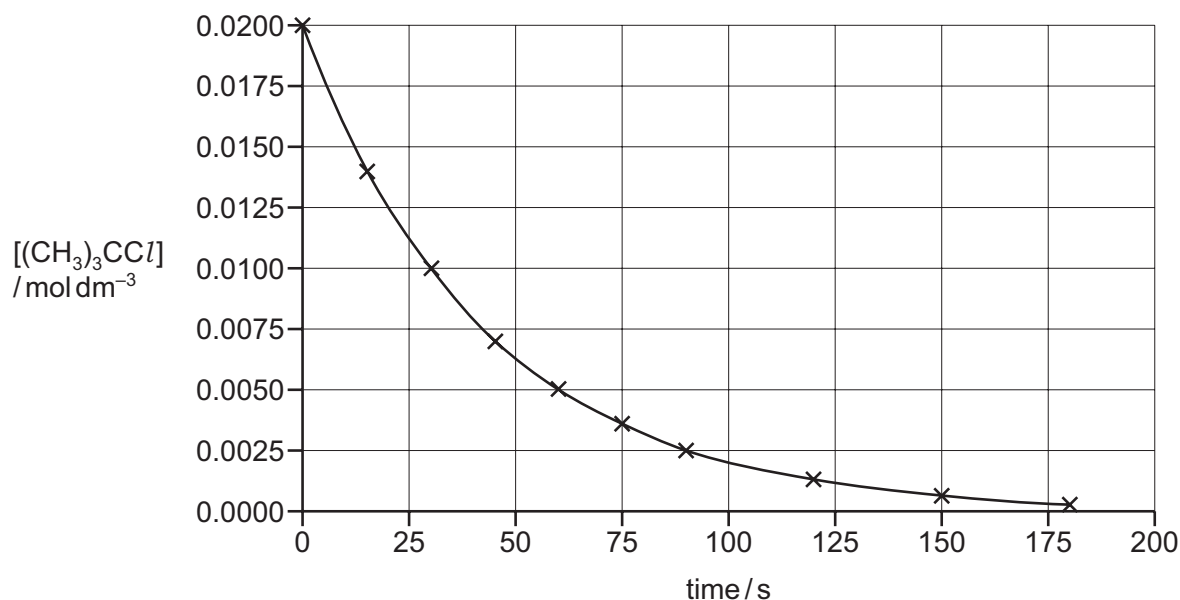


Fig. 3.1

- (i) Use the graph to determine the rate of reaction at 40 s. Show all your working.

rate = .....  $\text{mol dm}^{-3} \text{s}^{-1}$  [1]

- (ii) Use the graph to show that the overall reaction is first order. Explain your answer.

.....  
 .....  
 ..... [2]

- (b) In a different reaction, which is also a first order reaction, 75% of the reactant is consumed in 320 s.

Calculate the rate constant,  $k$ , for this reaction. State the units for  $k$ .

$k$  = ..... units = ..... [2]

31. Question part 2/3 | 9701 May-June 2022 P4 V3 Q3

- (c) (i) Define standard electrode potential,  $E^\circ$ .

.....  
.....  
..... [1]

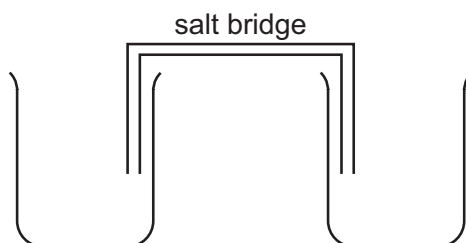
- (ii) A salt bridge is used in an electrochemical cell.

State the function of the salt bridge. Explain your answer.

.....  
..... [1]

- (iii) Complete the diagram of the apparatus that can be used to measure the  $E^\circ$  of the  $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ ,  $\text{H}^+(\text{aq})/\text{Cr}^{3+}(\text{aq})$  electrode against the standard hydrogen electrode.

Your diagram should be fully labelled to identify all apparatus, substances and conditions.



[3]

- (iv) The  $E^\circ$  of the  $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$ ,  $\text{H}^+(\text{aq})/\text{Cr}^{3+}(\text{aq})$  electrode is +1.33 V.

Label the negative electrode and the direction of electron flow in the external circuit when the current flows in your diagram in (c)(iii). [1]

31. Question part 3/3 | 9701 May-June 2022 P4 V3 Q3

- (d) Table 3.1 lists relevant electrode potentials for some electrode reactions for use in (d)(i) and (d)(ii).

Table 3.1

| electrode reaction                                                                                                 | $E^\circ/V$ |
|--------------------------------------------------------------------------------------------------------------------|-------------|
| $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ | +1.33       |
| $\text{CH}_3\text{CHO} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CH}_3\text{CH}_2\text{OH}$             | -0.61       |
| $\text{CH}_3\text{COOH} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CH}_3\text{CHO} + \text{H}_2\text{O}$ | -0.94       |
| $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$                                    | +1.23       |

- (i) Ethanal is oxidised to ethanoic acid in the presence of  $\text{Cr}_2\text{O}_7^{2-}$  ions.

Construct the ionic equation for the oxidation of ethanal to ethanoic acid using dichromate(VI) in acid conditions. Calculate the  $E^\circ_{\text{cell}}$  for this reaction.

ionic equation .....

$$E^\circ_{\text{cell}} = \dots\dots\dots \text{V} \quad [2]$$

- (ii) In an ethanol-oxygen fuel cell,  $\text{CH}_3\text{CH}_2\text{OH}(\text{l})$  and  $\text{O}_2(\text{g})$  are in contact with two inert electrodes immersed in an acidic solution.

The cell reaction for the oxidation of ethanol by oxygen is shown.



Calculate  $\Delta G^\circ$ , in  $\text{kJ mol}^{-1}$ , for the oxidation of ethanol by oxygen.

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

[Total: 15]

- 3 Data should be selected from Table 3.1 in order to answer some parts of this question.

Table 3.1

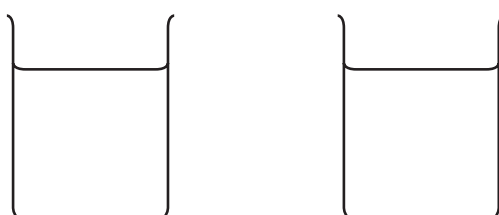
| electrode reaction                                                                                | $E^\circ / \text{V}$ |
|---------------------------------------------------------------------------------------------------|----------------------|
| $\text{Cl}_2 + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$                                       | +1.36                |
| $2\text{HOCl} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Cl}_2 + 2\text{H}_2\text{O}$   | +1.64                |
| $\text{ClO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{Cl}^- + 2\text{OH}^-$   | +0.89                |
| $\text{Sn}^{4+} + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}$                                  | +0.15                |
| $\text{Sn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sn}$                                       | -0.14                |
| $\text{V}^{2+} + 2\text{e}^- \rightleftharpoons \text{V}$                                         | -1.20                |
| $\text{V}^{3+} + \text{e}^- \rightleftharpoons \text{V}^{2+}$                                     | -0.26                |
| $\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{V}^{3+} + \text{H}_2\text{O}$ | +0.34                |
| $\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{VO}^{2+} + \text{H}_2\text{O}$ | +1.00                |

- (a) Standard electrode potentials are measured under standard conditions.

- (i) Describe the standard conditions used in the  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell.

.....  
 ..... [1]

- (ii) Complete the diagram below to show how  $E^\circ (\text{Sn}^{4+}/\text{Sn}^{2+})$  can be measured experimentally. Your diagram should be fully labelled to identify all apparatus and substances.



[3]

- (iii) Equal volumes of  $1.0 \text{ mol dm}^{-3} \text{ Sn}^{2+}(\text{aq})$  and  $1.0 \text{ mol dm}^{-3} \text{ Cl}^{-}(\text{aq})$  are mixed.

Use relevant  $E^\circ$  values to explain whether a reaction occurs between these two ions.

.....  
.....  
.....  
..... [2]

- (iv) Equal volumes of  $1.0 \text{ mol dm}^{-3}$  of  $\text{Sn}^{2+}(\text{aq})$  and acidified  $1.0 \text{ mol dm}^{-3} \text{ VO}^{2+}(\text{aq})$  are mixed.

Write an equation for the reaction that takes place in the resulting mixture.

..... [2]

- (b) A solution of  $\text{SnCl}_2(\text{aq})$  is electrolysed for a measured time using a steady current.

A mass of 2.95 g of tin metal is produced at the cathode.

$\text{Al}_2\text{O}_3(\text{l})$  is electrolysed for the same time by the same current.

Calculate the mass of aluminium metal produced at the cathode. Give your answer to **three** significant figures. Show your working.

mass of aluminium metal = ..... g [2]

[Total: 10]

- 3 Data should be selected from Table 3.1 in order to answer some parts of this question.

Table 3.1

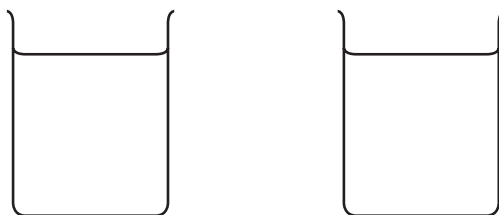
| electrode reaction                                                                                   | $E^\circ/\text{V}$ |
|------------------------------------------------------------------------------------------------------|--------------------|
| $\text{Mg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mg}$                                          | -2.38              |
| $\text{Mn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mn}$                                          | -1.18              |
| $\text{Mn}^{3+} + \text{e}^- \rightleftharpoons \text{Mn}^{2+}$                                      | +1.49              |
| $\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 2\text{H}_2\text{O}$   | +1.23              |
| $\text{MnO}_4^- + \text{e}^- \rightleftharpoons \text{MnO}_4^{2-}$                                   | +0.56              |
| $\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{MnO}_2 + 2\text{H}_2\text{O}$   | +1.67              |
| $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$ | +1.52              |

- (a) An electrochemical cell can be constructed from a  $\text{Mg}^{2+}/\text{Mg}$  half-cell and a  $\text{MnO}_4^-/\text{Mn}^{2+}$  half-cell. The standard cell potential of this cell can be calculated using the standard electrode potentials of the two half-cells.

- (i) Define standard electrode potential. Include details of the standard conditions used.

.....  
 .....  
 .....  
 ..... [2]

- (ii) Complete the diagram below to show an electrochemical cell constructed from a  $\text{Mg}^{2+}/\text{Mg}$  half-cell and a  $\text{MnO}_4^-/\text{Mn}^{2+}$  half-cell. Label your diagram.



[3]

- (iii) Use a positive (+) sign and a negative (–) sign to identify the polarity of each of the two electrodes in your diagram.  
 Use an arrow and the symbol 'e' to show the direction of electron flow in the external circuit.

[1]

- (iv) Calculate the standard cell potential,  $E_{\text{cell}}^{\ominus}$ , of this cell.

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

- (v) Construct an equation for the cell reaction.

..... [1]

- (vi) Predict how the cell reaction will change, if at all, when the solution in the  $\text{Mg}^{2+}/\text{Mg}$  half-cell is diluted by the addition of a large volume of water. Explain your answer.

.....

.....

..... [1]

- (b) A molten magnesium salt is electrolysed for 15.0 minutes by a constant current.

$4.75 \times 10^{22}$  magnesium atoms are produced at the cathode.

Calculate the value of the current used.

$$\text{current} = \dots\dots\dots \text{ A } [2]$$

[Total: 11]

34. Question part 1/2 | 9701 Oct-Nov 2022 P4 V3 Q3

- 3 Data should be selected from Table 3.1 in order to answer some parts of this question.

**Table 3.1**

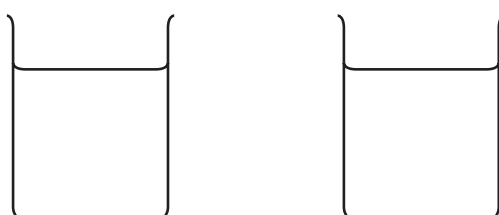
| electrode reaction                                                                                | $E^\circ / \text{V}$ |
|---------------------------------------------------------------------------------------------------|----------------------|
| $\text{Cl}_2 + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$                                       | +1.36                |
| $2\text{HOCl} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Cl}_2 + 2\text{H}_2\text{O}$   | +1.64                |
| $\text{ClO}^- + \text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{Cl}^- + 2\text{OH}^-$   | +0.89                |
| $\text{Sn}^{4+} + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}$                                  | +0.15                |
| $\text{Sn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sn}$                                       | -0.14                |
| $\text{V}^{2+} + 2\text{e}^- \rightleftharpoons \text{V}$                                         | -1.20                |
| $\text{V}^{3+} + \text{e}^- \rightleftharpoons \text{V}^{2+}$                                     | -0.26                |
| $\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{V}^{3+} + \text{H}_2\text{O}$ | +0.34                |
| $\text{VO}_2^+ + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{VO}^{2+} + \text{H}_2\text{O}$ | +1.00                |

- (a) Standard electrode potentials are measured under standard conditions.

- (i) Describe the standard conditions used in the  $\text{Sn}^{4+}/\text{Sn}^{2+}$  half-cell.

.....  
 ..... [1]

- (ii) Complete the diagram below to show how  $E^\circ (\text{Sn}^{4+}/\text{Sn}^{2+})$  can be measured experimentally. Your diagram should be fully labelled to identify all apparatus and substances.



[3]

34. Question part 2/2 | 9701 Oct-Nov 2022 P4 V3 Q3

- (iii) Equal volumes of  $1.0 \text{ mol dm}^{-3} \text{ Sn}^{2+}(\text{aq})$  and  $1.0 \text{ mol dm}^{-3} \text{ Cl}^{-}(\text{aq})$  are mixed.

Use relevant  $E^{\circ}$  values to explain whether a reaction occurs between these two ions.

.....  
.....  
.....  
..... [2]

- (iv) Equal volumes of  $1.0 \text{ mol dm}^{-3}$  of  $\text{Sn}^{2+}(\text{aq})$  and acidified  $1.0 \text{ mol dm}^{-3} \text{ VO}^{2+}(\text{aq})$  are mixed.

Write an equation for the reaction that takes place in the resulting mixture.

..... [2]

- (b) A solution of  $\text{SnCl}_2(\text{aq})$  is electrolysed for a measured time using a steady current.

A mass of 2.95 g of tin metal is produced at the cathode.

$\text{Al}_2\text{O}_3(\text{l})$  is electrolysed for the same time by the same current.

Calculate the mass of aluminium metal produced at the cathode. Give your answer to **three** significant figures. Show your working.

mass of aluminium metal = ..... g [2]

[Total: 10]

35. Question part 1/4 | 9701 Feb-March 2023 P4 V2 Q2

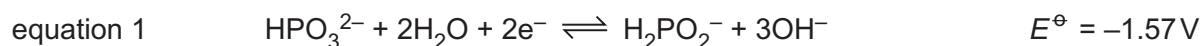
**2** Hypophosphorous acid is an inorganic acid.

The conjugate base of hypophosphorous acid is  $\text{H}_2\text{PO}_2^-$ .

**(a)** Give the formula of hypophosphorous acid.

..... [1]

**(b)**  $\text{H}_2\text{PO}_2^-$  is a strong reducing agent. It can be used to reduce metal cations without the need for electrolysis.



**(i)** In an experiment, an alkaline  $\text{HPO}_3^{2-}/\text{H}_2\text{PO}_2^-$  half-cell is constructed with  $[\text{H}_2\text{PO}_2^-] = 0.050\text{ mol dm}^{-3}$ .

All other ions are at their standard concentration.

Predict how the value of  $E$  of this half-cell differs from its  $E^\ominus$  value.

Explain your answer.

.....  
 .....  
 .....  
 ..... [2]

**(ii)** The  $\text{Cr}^{3+}/\text{Cr}$  half-cell has a standard electrode potential of  $-0.74\text{V}$ .

An electrochemical cell consists of an alkaline  $\text{HPO}_3^{2-}/\text{H}_2\text{PO}_2^-$  half-cell and a  $\text{Cr}^{3+}/\text{Cr}$  half-cell.

Calculate the standard cell potential,  $E_{\text{cell}}^\ominus$ .

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

- (iii) Complete the diagram in Fig. 2.1 to show how the standard electrode potential of the  $\text{Cr}^{3+}/\text{Cr}$  half-cell can be measured relative to that of the standard hydrogen electrode.

Identify the chemicals, conditions and relevant pieces of apparatus.

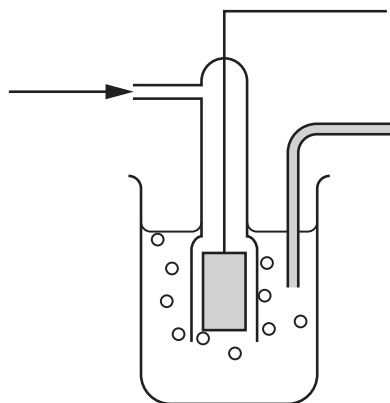


Fig. 2.1

[3]

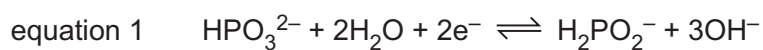
- (iv) Label Fig. 2.1 to show:

- which is the positive electrode
- the direction of electron flow in the external circuit.

[1]

- (v)  $\text{H}_2\text{PO}_2^-$  reduces  $\text{Ni}^{2+}$  to Ni in alkaline conditions.

Use equation 1 to construct the ionic equation for this reaction.



..... [1]

(c)  $\text{H}_2\text{PO}_2^-(\text{aq})$  reacts with  $\text{OH}^-(\text{aq})$ .

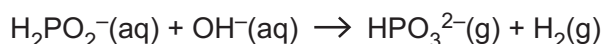


Table 2.1 shows the results of a series of experiments used to investigate the rate of this reaction.

**Table 2.1**

| experiment | $[\text{H}_2\text{PO}_2^-(\text{aq})]$<br>/ $\text{mol dm}^{-3}$ | $[\text{OH}^-(\text{aq})]$<br>/ $\text{mol dm}^{-3}$ | volume of $\text{H}_2$<br>produced in 60 s<br>/ $\text{cm}^3$ |
|------------|------------------------------------------------------------------|------------------------------------------------------|---------------------------------------------------------------|
| 1          | 0.40                                                             | 2.00                                                 | 6.4                                                           |
| 2          | 0.80                                                             | 2.00                                                 | 12.8                                                          |
| 3          | 1.20                                                             | 1.00                                                 | 4.8                                                           |

(i) The volume of  $\text{H}_2$  was measured under room conditions.

Use the molar volume of gas,  $V_m$ , and the data from experiment 1 to calculate the rate of reaction in  $\text{mol dm}^{-3} \text{s}^{-1}$ .

rate of reaction = .....  $\text{mol dm}^{-3} \text{s}^{-1}$  [1]

(ii) The rate equation was found to be:

$$\text{rate} = k [\text{H}_2\text{PO}_2^-(\text{aq})] [\text{OH}^-(\text{aq})]^2$$

Show that the data in Table 2.1 is consistent with the rate equation.

.....  
 .....  
 .....  
 .....  
 ..... [2]

(iii) State the units of the rate constant,  $k$ , for the reaction.

..... [1]

- (iv) The experiment is repeated using a large excess of  $\text{OH}^-(\text{aq})$ .

Under these conditions, the rate equation is:

$$\text{rate} = k_1 [\text{H}_2\text{PO}_2^-(\text{aq})]$$

$$k_1 = 8.25 \times 10^{-5} \text{ s}^{-1}$$

Calculate the value of the half-life,  $t_{\frac{1}{2}}$ , of the reaction.

$$t_{\frac{1}{2}} = \dots\dots\dots \text{ s [1]}$$

- (v) Describe how an increase in temperature affects the value of the rate constant,  $k_1$ .

.....  
..... [1]

- (d) A student suggests that the reaction between  $\text{H}_2\text{PO}_2^-(\text{aq})$  and  $\text{OH}^-(\text{aq})$  might happen more quickly in the presence of a heterogeneous catalyst.

Describe the mode of action of a heterogeneous catalyst.

.....  
.....  
.....  
..... [2]

[Total: 17]

36. Question part 1/3 | 9701 May-June 2023 P4 V2 Q9

- 9 (a) Define standard cell potential,  $E_{\text{cell}}^{\ominus}$ .

.....  
..... [1]

- (b) An electrochemical cell is set up to measure  $E_{\text{cell}}^{\ominus}$  of a cell consisting of an  $\text{Fe}^{3+}/\text{Fe}^{2+}$  half-cell and a  $\text{Cl}_2/\text{Cl}^-$  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.



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

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

(d) Another experiment is set up using the same electrochemical cell.

In this experiment the  $\text{Fe}^{2+}$  concentration is  $0.15 \text{ mol dm}^{-3}$ . All other concentrations remain at their standard values.

The Nernst equation is shown.

$$E = E^\ominus + (0.059/z) \log \frac{[\text{oxidised species}]}{[\text{reduced species}]}$$

(i) Use the Nernst equation to calculate the electrode potential,  $E$ , for the  $\text{Fe}^{3+}/\text{Fe}^{2+}$  half-cell in this experiment.

$$[E^\ominus: \text{Fe}^{3+}/\text{Fe}^{2+} = +0.77 \text{ V}]$$

$$E = \dots\dots\dots \text{ V [1]}$$

(ii) Use your answer to (d)(i) to calculate  $E_{\text{cell}}$  for this electrochemical cell.

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

[Total: 8]

### Important values, constants and standards

|                                 |                                                                                                                                                             |
|---------------------------------|-------------------------------------------------------------------------------------------------------------------------------------------------------------|
| molar gas constant              | $R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$                                                                                                                |
| Faraday constant                | $F = 9.65 \times 10^4 \text{ C mol}^{-1}$                                                                                                                   |
| Avogadro constant               | $L = 6.022 \times 10^{23} \text{ mol}^{-1}$                                                                                                                 |
| electronic charge               | $e = -1.60 \times 10^{-19} \text{ C}$                                                                                                                       |
| molar volume of gas             | $V_{\text{m}} = 22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p. (101 kPa and 273 K)<br>$V_{\text{m}} = 24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions |
| ionic product of water          | $K_{\text{w}} = 1.00 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ (at 298 K (25 °C))                                                                      |
| specific heat capacity of water | $c = 4.18 \text{ kJ kg}^{-1} \text{ K}^{-1}$ (4.18 $\text{J g}^{-1} \text{ K}^{-1}$ )                                                                       |

Topical questions | 42 questions

36. Question part 3/3 | 9701 May-June 2023 P4 V2 Q9

|    |    |           |       |    |    |         |       |    |    |              |       |    |    |           |       |    |    |            |   |    |    |           |       |    |    |           |       |    |    |            |       |    |    |           |       |    |    |             |       |    |    |             |       |     |    |         |       |     |    |            |       |     |    |          |       |     |    |            |       |
|----|----|-----------|-------|----|----|---------|-------|----|----|--------------|-------|----|----|-----------|-------|----|----|------------|---|----|----|-----------|-------|----|----|-----------|-------|----|----|------------|-------|----|----|-----------|-------|----|----|-------------|-------|----|----|-------------|-------|-----|----|---------|-------|-----|----|------------|-------|-----|----|----------|-------|-----|----|------------|-------|
| 57 | La | lanthanum | 138.9 | 58 | Ce | cerium  | 140.1 | 59 | Pr | praseodymium | 140.9 | 60 | Nd | neodymium | 144.4 | 61 | Pm | promethium | — | 62 | Sm | samarium  | 150.4 | 63 | Eu | euroium   | 152.0 | 64 | Gd | gadolinium | 157.3 | 65 | Tb | terbium   | 158.9 | 66 | Dy | dysprosium  | 162.5 | 67 | Ho | holmium     | 164.9 | 68  | Er | erbium  | 167.3 | 69  | Tm | thulium    | 168.9 | 70  | Yb | yterbium | 173.1 | 71  | Lu | lutetium   | 175.0 |
| 89 | Ac | actinium  | —     | 90 | Th | thorium | 232.0 | 91 | Pa | protactinium | 231.0 | 92 | U  | uranium   | 238.0 | 93 | Np | neptunium  | — | 94 | Pu | plutonium | —     | 95 | Am | americium | —     | 96 | Cm | curium     | —     | 97 | Bk | berkelium | —     | 98 | Cf | californium | —     | 99 | Es | einsteinium | —     | 100 | Fm | fermium | —     | 101 | Md | merendeium | —     | 102 | No | nobelium | —     | 103 | Lr | lawrencium | —     |

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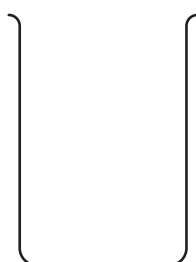
37. Question part 1/4 | 9701 Oct-Nov 2023 P4 V2 Q3

- 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 concentration of  $Fe^{2+}$  in the  $Fe^{3+}/Fe^{2+}$  electrode. All other conditions are standard.

- (i) Use the Nernst equation to calculate the  $E$  value of the  $Fe^{3+}/Fe^{2+}$  electrode.

Show your working.

$$E = \dots\dots\dots V \quad [2]$$

- (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.

$$\Delta G^\ominus = \dots\dots\dots \text{kJ mol}^{-1} \quad [3]$$

- (e) A solution of iron(II) sulfate,  $\text{FeSO}_4(\text{aq})$  is electrolysed with iron electrodes. Under the conditions used, no gas is evolved at the cathode.

A current of 0.640 A is passed for 17.0 minutes. The mass of the cathode increases by 0.185 g.

Use these results to calculate an experimental value for the Avogadro constant,  $L$ .

Show your working.

$$L = \dots\dots\dots \text{mol}^{-1} \quad [3]$$

- (f) Iron(II) chloride,  $\text{FeCl}_2$ , is oxidised by chlorine to form iron(III) chloride,  $\text{FeCl}_3$ , under standard conditions.

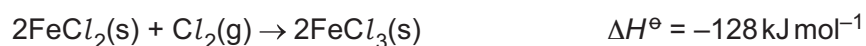


**Table 3.2**

| species                   | $S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$ |
|---------------------------|-----------------------------------------------|
| $\text{Cl}_2(\text{g})$   | 223                                           |
| $\text{FeCl}_2(\text{s})$ | 120                                           |
| $\text{FeCl}_3(\text{s})$ | 142                                           |

- (i) Use Table 3.2 and other data to calculate the Gibbs free energy change,  $\Delta G^\ominus$ , for this reaction.

Show your working.

$$\Delta G^\ominus = \dots\dots\dots \text{kJ mol}^{-1} \quad [3]$$

37. Question part 4/4 | 9701 Oct-Nov 2023 P4 V2 Q3

(ii) Predict whether this reaction becomes more or less feasible at a higher temperature.

Explain your answer.

The reaction becomes ..... feasible.

explanation .....

.....  
[1]

[Total: 18]

- 4 (a) Define standard electrode potential,  $E^\ominus$ , including a description of standard conditions.

.....  
.....  
..... [2]

- (b) (i) An electrochemical cell is set up to measure  $E^\ominus$  of the  $\text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$  electrode.

Draw a labelled diagram of this electrochemical cell.

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

[3]

- (ii) A separate electrochemical cell is set up using a **lower** concentration of  $\text{Ag}^+(\text{aq})$  than that used in (b)(i).

Suggest how the electrode potential,  $E$ , for the  $\text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$  electrode would change from its  $E^\ominus$  value. Explain your answer.

.....  
.....  
..... [1]

- (c) Define enthalpy change of solution,  $\Delta H_{\text{sol}}^\ominus$ .

.....  
..... [1]

- (d) Some relevant energy changes for  $\text{AgNO}_3$  are shown in Table 4.1.

Table 4.1

| energy change                                            | value / $\text{kJ mol}^{-1}$ |
|----------------------------------------------------------|------------------------------|
| enthalpy change of solution of $\text{AgNO}_3(\text{s})$ | +22.6                        |
| enthalpy change of hydration of silver ions              | -475                         |
| enthalpy change of hydration of nitrate ions             | -314                         |

- (i) Complete the energy cycle in Fig. 4.1 to show the relationship between the lattice energy,  $\Delta H_{\text{latt}}^{\ominus}$ , of  $\text{AgNO}_3(\text{s})$  and the energy changes shown in Table 4.1.

Include state symbols for all the species.

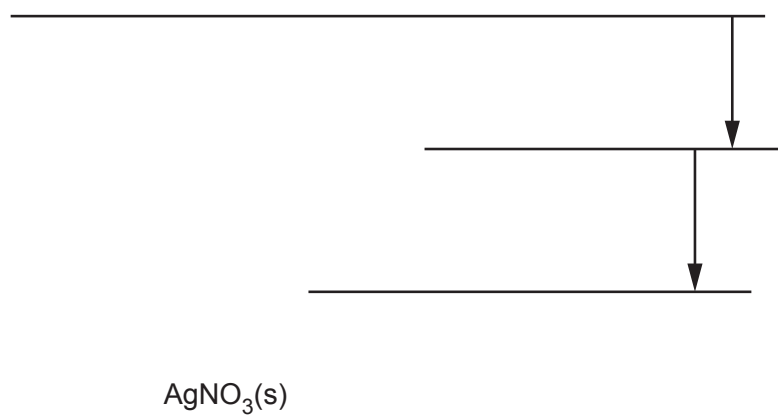


Fig. 4.1

[2]

- (ii) Calculate the lattice energy,  $\Delta H_{\text{latt}}^{\ominus}$ , of  $\text{AgNO}_3(\text{s})$ .

$$\Delta H_{\text{latt}}^{\ominus} = \dots\dots\dots \text{kJ mol}^{-1} \quad [1]$$

38. Question part 3/3 | 9701 May-June 2024 P4 V1 Q4

- (e) Suggest the trend in the magnitude of the lattice energies of the metal nitrates,  $\text{NaNO}_3(\text{s})$ ,  $\text{Mg}(\text{NO}_3)_2(\text{s})$  and  $\text{RbNO}_3(\text{s})$ .

Explain your answer.

.....

most exothermic ..... least exothermic

.....

.....

.....

.....

.....

[3]

[Total: 13]

39. Question part 1/2 | 9701 May-June 2024 P4 V2 Q2

- 2 (a) (i)** Lithium nitrate,  $\text{LiNO}_3$ , 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  $\text{LiNO}_3$ .

..... [1]

- (ii)** The other Group 1 nitrates,  $\text{MNO}_3$ , decompose on heating to form the metal nitrite,  $\text{MNO}_2$ , and oxygen.

The thermal stability of these nitrates increases down the group.

Suggest why the thermal stability of  $\text{MNO}_3$  increases down the group.

.....  
 .....  
 ..... [2]

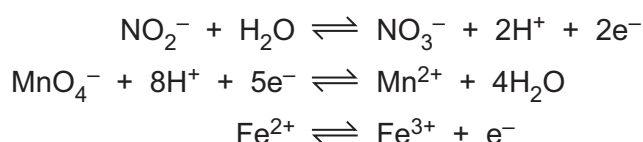
- (b)** Acidified manganate(VII) ions,  $\text{MnO}_4^-$ , can be used to analyse solutions containing nitrite ions,  $\text{NO}_2^-$ , by titration.

**X** is a solution of  $\text{NaNO}_2$ .

$250.0\text{ cm}^3$  of **X** is added to  $50.0\text{ cm}^3$  of  $0.125\text{ mol dm}^{-3}$  acidified  $\text{MnO}_4^-$ (aq). The  $\text{MnO}_4^-$ (aq) ions are in excess; all the  $\text{NO}_2^-$  ions are oxidised in the reaction.

The unreacted  $\text{MnO}_4^-$ (aq) required  $22.50\text{ cm}^3$  of  $0.0400\text{ mol dm}^{-3}$   $\text{Fe}^{2+}$ (aq) to reach the end-point.

The relevant half-equations are shown.



Calculate the concentration, in  $\text{mol dm}^{-3}$ , of  $\text{NaNO}_2$  in **X**.



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

Table 2.1

| electrode reaction                                                                                     | $E^\ominus/V$ |
|--------------------------------------------------------------------------------------------------------|---------------|
| $\text{Mn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mn}$                                            | -1.18         |
| $\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 2\text{H}_2\text{O}$     | +1.23         |
| $\text{MnO}_4^- + \text{e}^- \rightleftharpoons \text{MnO}_4^{2-}$                                     | +0.56         |
| $\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{MnO}_2 + 2\text{H}_2\text{O}$     | +1.67         |
| $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$   | +1.52         |
| $\text{MnO}_4^- + 2\text{H}_2\text{O} + 3\text{e}^- \rightleftharpoons \text{MnO}_2 + 4\text{OH}^-$    | +0.59         |
| $\text{MnO}_4^{2-} + 2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{MnO}_2 + 4\text{OH}^-$ | +0.60         |
| $\text{MnO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{MnO}_2 + 2\text{H}_2\text{O}$  | +1.70         |

- (i) Aqueous manganate(VI) ions,  $\text{MnO}_4^{2-}$ , are unstable in acidic conditions and undergo a disproportionation reaction.

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

Construct an overall ionic equation for this disproportionation reaction.

..... [2]

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

.....  
 .....  
 ..... [1]

[Total: 9]

- 4 (a) Define standard electrode potential,  $E^\ominus$ , including a description of standard conditions.

.....  
.....  
..... [2]

- (b) (i) An electrochemical cell is set up to measure  $E^\ominus$  of the  $\text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$  electrode.

Draw a labelled diagram of this electrochemical cell.

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

[3]

- (ii) A separate electrochemical cell is set up using a **lower** concentration of  $\text{Ag}^+(\text{aq})$  than that used in (b)(i).

Suggest how the electrode potential,  $E$ , for the  $\text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$  electrode would change from its  $E^\ominus$  value. Explain your answer.

.....  
.....  
..... [1]

- (c) Define enthalpy change of solution,  $\Delta H_{\text{sol}}^\ominus$ .

.....  
..... [1]

- (d) Some relevant energy changes for  $\text{AgNO}_3$  are shown in Table 4.1.

Table 4.1

| energy change                                            | value / $\text{kJ mol}^{-1}$ |
|----------------------------------------------------------|------------------------------|
| enthalpy change of solution of $\text{AgNO}_3(\text{s})$ | +22.6                        |
| enthalpy change of hydration of silver ions              | -475                         |
| enthalpy change of hydration of nitrate ions             | -314                         |

- (i) Complete the energy cycle in Fig. 4.1 to show the relationship between the lattice energy,  $\Delta H_{\text{latt}}^\ominus$ , of  $\text{AgNO}_3(\text{s})$  and the energy changes shown in Table 4.1.

Include state symbols for all the species.

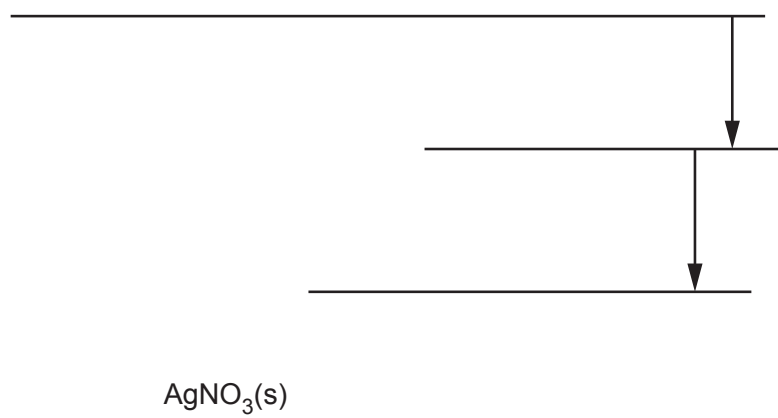


Fig. 4.1

[2]

- (ii) Calculate the lattice energy,  $\Delta H_{\text{latt}}^\ominus$ , of  $\text{AgNO}_3(\text{s})$ .

$$\Delta H_{\text{latt}}^\ominus = \dots\dots\dots \text{kJ mol}^{-1} \quad [1]$$

40. Question part 3/3 | 9701 May-June 2024 P4 V3 Q4

- (e) Suggest the trend in the magnitude of the lattice energies of the metal nitrates,  $\text{NaNO}_3(\text{s})$ ,  $\text{Mg}(\text{NO}_3)_2(\text{s})$  and  $\text{RbNO}_3(\text{s})$ .

Explain your answer.

.....

most exothermic ..... least exothermic

.....

.....

.....

.....

.....

[3]

[Total: 13]

41. Question part 1/2 | 9701 Oct-Nov 2025 P4 V1 Q4

- 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  $-0.76\text{ V}$ .

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  $0.25\text{ mol dm}^{-3}$ .

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

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

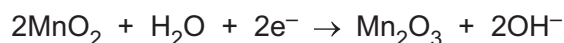
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- (c) An electrochemical cell consists of a  $\text{ZnO}/\text{Zn}$  electrode and a  $\text{MnO}_2/\text{Mn}_2\text{O}_3$  electrode in an alkaline electrolyte.

The standard cell potential,  $E^\ominus_{\text{cell}}$ , for this cell is +1.47 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  $\text{ZnO}/\text{Zn}$  electrode is  $-1.28\text{ V}$ .

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]

42. Question part 1/2 | 9701 Oct-Nov 2025 P4 V3 Q4

- 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  $-0.76\text{ V}$ .

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  $0.25\text{ mol dm}^{-3}$ .

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

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

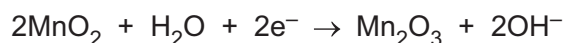
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- (c) An electrochemical cell consists of a  $\text{ZnO}/\text{Zn}$  electrode and a  $\text{MnO}_2/\text{Mn}_2\text{O}_3$  electrode in an alkaline electrolyte.

The standard cell potential,  $E^\ominus_{\text{cell}}$ , for this cell is +1.47 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  $\text{ZnO}/\text{Zn}$  electrode is  $-1.28\text{ V}$ .

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]