1. 9701/41/O/N/16 Q4 c,d,e,f

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

....

....

.... [2]

(d) An electrochemical cell was set up to measure the standard electrode potential, $E^\ominus_{cell}$, of a cell made of a Co^{2+}/Co half-cell and a Fe^{3+}/Fe^{2+} half-cell.

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

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

[1]

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

.... [1]

(iii) Use the *Data Booklet* to calculate the $E^\ominus_{cell}$.

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

(e) The electrochemical cell in **(d)** was set up again but this time the concentration of Co^{2+}(aq) was 0.050 mol dm^{-3}.

The Nernst equation can be used to calculate the value of an electrode potential at different concentrations.

$$E = E^\ominus + (0.059/z)\log[Co^{2+}(aq)]$$ Nernst equation

(i) Use the *Data Booklet* and the Nernst equation to calculate the value of *E* for the Co^{2+}/Co half-cell in this experiment.

E for Co^{2+}/Co = ... V [1]

(ii) Suggest how this change will affect the overall cell potential, E_{cell}, compared to $E^\ominus_{cell}$ in **(d)(iii)**.
Circle your answer.

less positive   no change   more positive

[1]

(f) Iron(III) ions can oxidise vanadium metal.

Construct an equation for the reaction of an excess of iron(III) ions with vanadium metal.
Use of the *Data Booklet* will be helpful.

.... [2]

2. 9701/41/M/J/16 Q4

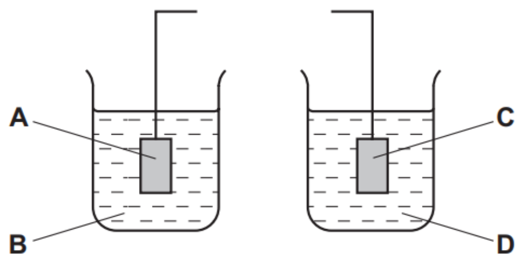
(a) **(i)** Define the term *standard cell potential*, $E^{\ominus}_{cell}$.

....

....

.... [1]

The following incomplete diagram shows the apparatus that can be used to measure the $E^{\ominus}_{cell}$ for a cell composed of the Fe^{3+}/Fe^{2+} and Ag^{+}/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 mol dm^{-3} Ag^{+}(aq) instead of the standard Ag solution.

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

E_{cell} = ... V [2]

3. 9701/42/M/J/16 Q2c

(c) (i) Draw a fully labelled diagram of the equipment needed to measure the voltage of an electrochemical cell consisting of the standard hydrogen electrode and the standard Cu/Cu^{2+} electrode.

[4]

(ii) For the cell drawn in **(i)**, calculate the $E^{\ominus}_{cell}$ and state which electrode is positive.

$E^{\ominus}_{cell}$ = ... identity of the positive electrode ....
[1]

(c) (i) Draw a fully labelled diagram of the equipment needed to measure the voltage of an electrochemical cell consisting of the standard hydrogen electrode and the standard Cu/Cu^{2+} electrode.

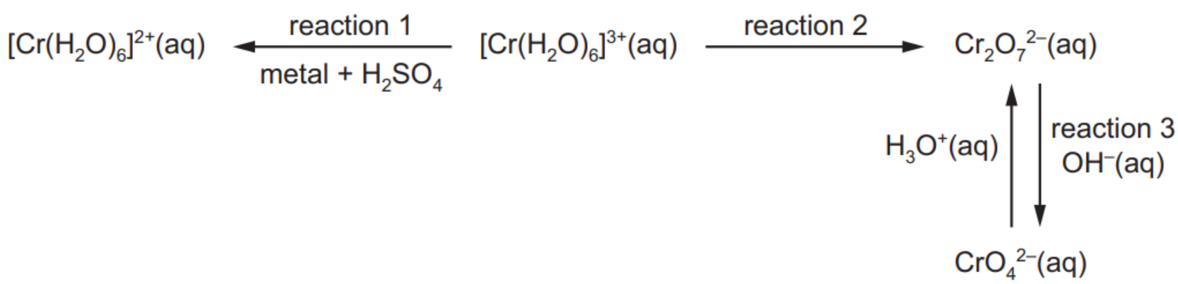
[4]

(ii) For the cell drawn in **(i)**, calculate the $E^{\ominus}_{cell}$ and state which electrode is positive.

$E^{\ominus}_{cell}$ = ... identity of the positive electrode ....
[1]

4. 9701/42/F/M/16 Q5

Some reactions of chromium ions are shown below.



(a) **(i)** Use the *Data Booklet* to suggest a suitable metal to carry out reaction 1.

.... [1]

(ii) Use $E^{\ominus}$ values to explain your answer to **(i)** by calculating the $E^{\ominus}_{cell}$.

....

....

.... [2]

(b) A student suggested that reaction 2 could be carried out using acidified hydrogen peroxide solution.

Use the *Data Booklet* to show whether or not this reaction is feasible.

....

....

.... [2]

(c) Explain using oxidation numbers whether or not reaction 3 is a redox reaction.

....

....

.... [2]

(d) The student used an acidified solution of $Cr_2O_7^{2-}$(aq) to electroplate a steel box with chromium metal.

Calculate how long it would take for a current of 0.125A to deposit 0.0312g of chromium metal.

time = ... [3]

5. 9701/41/O/N/17 Q3

(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 Pb^{2+}(aq)/Pb(s) electrode. Include the necessary chemicals.

[4]

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

Suggest how the E for this electrode would differ from its $E^\ominus$ value if the concentration of $Pb^{2+}(aq)$ 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_2. The electrolyte is $H_2SO_4(aq)$.

When a lead-acid cell is in use, Pb^{2+} ions are precipitated out as $PbSO_4(s)$ at the negative electrode.

$$Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$$

(i) Calculate the mass of Pb that is converted to $PbSO_4$ when a current of 0.40 A is delivered by the cell for 80 minutes.

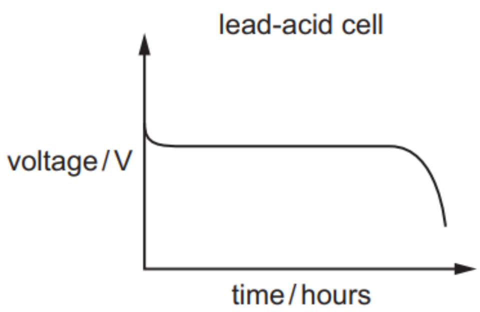
mass of Pb = ... g [2]

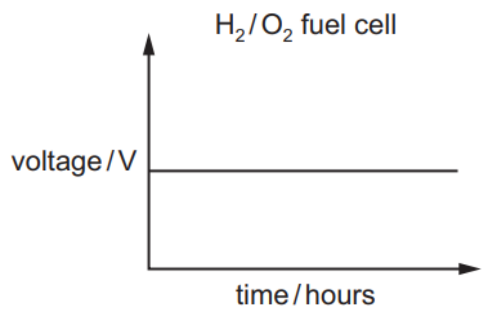
(ii) Complete the half-equation for the reaction taking place at the positive electrode.

$PbO_2(s)$ + $SO_4^{2-}(aq)$ + ... + ... → $PbSO_4(s)$ + ...

[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]

6. 9701/42/O/N/17 Q4

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 $E^\ominus_{cell}$.

$E^\ominus_{cell}$ = ... V [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]

(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 $E^\ominus_{cell}$.

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

Does a reaction occur? ....

equation ....

$E^\ominus_{cell}$ ....

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

Does a reaction occur? ....

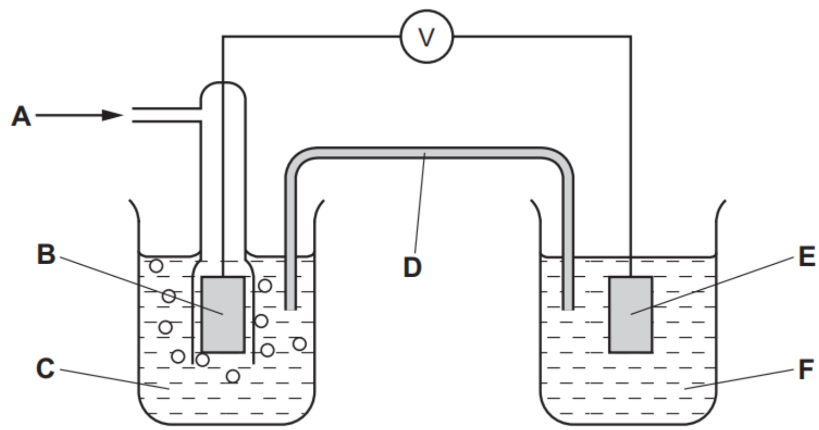
equation ....

$E^\ominus_{cell}$ ....

[3]

7. 9701/42/F/M/17 Q3

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



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

A .... **D** ....

B .... **E** ....

C .... **F** ....

[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 $Fe^{3+}(aq)/Fe^{2+}(aq)$ half-cell was connected to a $Cu^{2+}(aq)/Cu(s)$ half-cell.

Determine the standard cell potential, $E^\ominus_{cell}$, when these two half-cells are connected by a wire and the circuit is completed.
Use the *Data Booklet* to help you.

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

(c) **(i)** The $E^\ominus$ of $Ni^{2+}(aq)/Ni(s)$ is –0.25 V.

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

....

....

.... [1]

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

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

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

E = ... V [2]

8. 9701/41/O/N/18 Q8e

(e) In a methanol-oxygen fuel cell, $CH_3OH(l)$ and $O_2(g)$ are in contact with two inert electrodes immersed in an acidic solution.

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

$$CH_3OH + H_2O \rightleftharpoons CO_2 + 6H^+ + 6e^- \qquad E^\ominus = -0.02\,V$$

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

....

....

.... [1]

(ii) Use $E^\ominus$ values to calculate the $E^\ominus_{cell}$ for this reaction.

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

9. 9701/42/O/N/18 Q5c

Sodium nitrite, $NaNO_2$, is a decomposition product from heating sodium nitrate, $NaNO_3$.

A student analysed a sample of sodium nitrite by titration with aqueous cerium(IV) ions, $Ce^{4+}(aq)$. The equation for the titration reaction is shown.

$$NO_2^-(aq) + 2Ce^{4+}(aq) + H_2O(l) \rightarrow 2Ce^{3+}(aq) + NO_3^-(aq) + 2H^+(aq)$$

(c) Acidified manganate(VII) ions, MnO_4^-, can also be used to analyse solutions containing nitrite ions, NO_2^-, by titration. In acidic solution, NO_2^- ions exist as HNO_2.

(i) Use the *Data Booklet* to construct an ionic equation for this reaction.

....

....

.... [2]

(ii) Use $E^\ominus$ values to calculate the $E^\ominus_{cell}$ for this reaction.

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

10. 9701/41/M/J/18 Q3

(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
$AgNO_3$(aq)		
concentrated NaCl(aq)		
$CuSO_4$(aq)		

[3]

(b) Molten calcium iodide, 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 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 1.11g.

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

oxidation number of Fe = ... [3]

11. 9701/42/M/J/18 Q3

(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 $CuCl_2$(aq)		
concentrated $MgCl_2$(aq)		

[3]

(b) **(i)** The electrolysis of molten $ZnBr_2$ 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 = ... cm^3 [3]

12. 9701/42/M/J/18 Q5d

(d) $Co^{2+}(aq)$ can be oxidised to $Co^{3+}(aq)$.

(i) Use the *Data Booklet* to suggest a suitable oxidising agent for this reaction.

.... [1]

(ii) Calculate the $E^{\ominus}_{cell}$ of this reaction.

$E^{\ominus}_{cell}$ = ... V [1]

(iii) Write an equation for the reaction between Co^{2+} and the oxidising agent you chose in **(d)(i)**.

.... [1]

13. 9701/42/F/M/18 Q2

(a) Describe the trend in the reactivity of the halogens Cl_2, Br_2 and I_2 as oxidising agents. Explain this trend using values of $E^{\ominus}$ (X_2/X^-) from the *Data Booklet*.

....

....

....

....

.... [2]

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

.... [1]

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

$$Cl_2 + 2OH^- \rightleftharpoons Cl^- + ClO^- + H_2O$$

$E^{\ominus}_{cell}$ = ... V [2]

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

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

E_{cell} becomes less positive than $E^{\ominus}_{cell}$.	
E_{cell} stays the same as $E^{\ominus}_{cell}$.	
E_{cell} becomes more positive than $E^{\ominus}_{cell}$.	

Explain your answer.

....

....

....
[2]

(c) A half-equation involving bromate(V) ions, BrO_3^-, and bromide ions is shown.

$$BrO_3^-(aq) + 3H_2O(l) + 6e^- \rightleftharpoons Br^-(aq) + 6OH^-(aq) \qquad E^{\ominus} = +0.58\,V$$

(i) An alkaline solution of chlorate(I), 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 $E^{\ominus}_{cell}$ for the reaction in **(i)**.

$E^{\ominus}_{cell}$ = ... V [1]

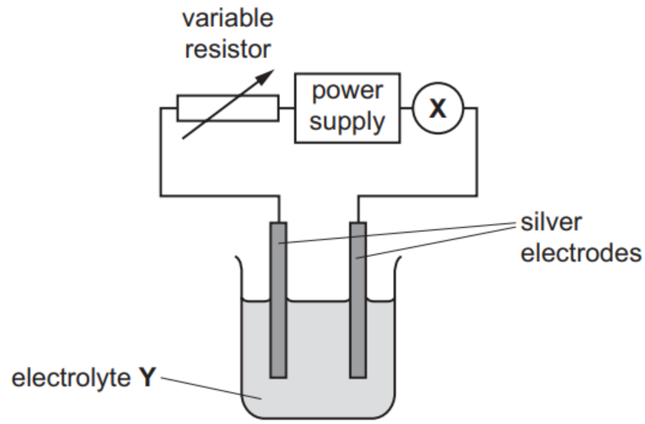
(iii) When a concentrated solution of bromic(V) acid, $HBrO_3$, is warmed, it decomposes to form bromine, oxygen and water only.

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

.... [1]

14. 9701/42/F/M/18 Q6

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 0.200A was passed through the cell for 40.0 minutes. The mass of the silver cathode increased by 0.500g.

The charge on the electron is -1.60×10^{-19} 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]

15. 9701/41/O/N/19 Q1

An electrochemical cell is constructed using two half-cells.

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

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

- Sn^{4+}/Sn^{2+} half-cell ....

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

(b) The cell is operated at 298 K.

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

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

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

E = ... V [2]

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

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

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

.... [2]

(c) Aluminium is produced industrially by electrolysis of a melt containing large amounts of 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]

16. 9701/42/O/N/19 Q1

An electrochemical cell is constructed using two half-cells.

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

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

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

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

(b) The cell is operated at 298 K.

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

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

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

E = ... V [2]

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

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

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

.... [2]

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

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

The relative atomic mass of copper is 63.5.

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

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

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

....

....

....

.... [2]

17. 9701/41/M/J/19 Q3c

(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 $E^{\ominus}_{cell}$ for this cell.

$E^{\ominus}_{cell}$ = ... V [1]

(iii) A current of 2.5×10^{-5} 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]

18. 9701/42/M/J/19 Q1d

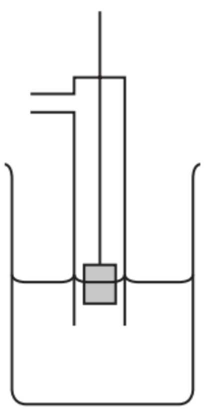
(d) Define the term *standard electrode potential, E*$^{\ominus}$.

....

....

.... [1]

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



[4]

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

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

[1]

19. 9701/42/F/M/19 Q4c

(c) (i) Use standard electrode potential data from the *Data Booklet* to calculate $E^{\ominus}_{cell}$ for the reaction.

$$2Cu^{2+}(aq) + 4I^{-}(aq) \rightarrow 2CuI(s) + I_2(aq)$$

$E^{\ominus}_{cell}$ = ... V [1]

(ii) Explain how the value of $E^{\ominus}_{cell}$ calculated in **(i)** predicts that the reaction is **not** likely to occur.

....

.... [1]

In an experiment, a solution of $I^{-}(aq)$ is added to a solution of $Cu^{2+}(aq)$. A reaction **does** occur and a precipitate of sparingly soluble CuI(s) is formed.

The concentration of $Cu^{2+}(aq)$ remaining in the solution is 1.00 mol dm^{-3}.
The concentration of $Cu^{+}(aq)$ in a saturated solution of CuI is 1.3×10^{-6} mol dm^{-3}.

(iii) Use the Nernst equation to calculate the electrode potential, *E*, for the Cu^{2+}/Cu^{+} half cell in this experiment.

$E(Cu^{2+}/Cu^{+})$ = ... V [2]

(iv) Copper(I) chloride is also sparingly soluble in water.

Suggest why the following reaction does **not** occur.

$$2Cu^{2+}(aq) + 4Cl^{-}(aq) \not\longrightarrow 2CuCl(s) + Cl_2(aq)$$

....

.... [1]

20. 9701/41/O/N/20 Q3

(a) Identify the substances liberated at the anode and at the cathode during the electrolysis of aqueous sodium sulfate, $Na_2SO_4(aq)$.

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 = ... dm^3

mass of sodium = ... g
[4]

(c) The equation representing the standard electrode potential, $E^⦵$, for the reduction of $MnO_4^-(aq)$ to $Mn^{2+}(aq)$ in acid solution is given.

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l) \qquad E^⦵ = +1.52\,V$$

(i) Draw a diagram of the apparatus that would be used to measure the $E^⦵$ 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 Mn^{2+} ions to MnO_4^- ions under standard conditions.

Write an equation for the reaction.

....

....

.... [2]

21. 9701/42/O/N/20 Q4

(a) Identify the substances liberated at the anode and at the cathode during the electrolysis of saturated $KCl(aq)$.

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 0.750A.

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

volume of oxygen = ... cm^3 [3]

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

(i) Give the $E^{\ominus}$ 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 $Sn^{2+}(aq)$ 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 $Mn^{2+}(aq)$ under standard conditions.

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

.... [1]

(d) An electrochemical cell can be made by connecting an Fe^{3+}/Fe^{2+} half-cell to an $S_2O_8^{2-}/SO_4^{2-}$ half-cell under standard conditions.

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

$E^{\ominus}_{cell}$ = ... V [1]

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

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

in the $S_2O_8^{2-}/SO_4^{2-}$ 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.

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

....

$S_2O_8^{2-}/SO_4^{2-}$ half-cell ....

....
[1]

22. 9701/41/M/J/20 Q3

The overall reaction for photosynthesis is shown.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

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

$$2H_2O \rightarrow O_2 + 4H^+ + 4e^- \qquad \text{half-equation 1}$$

(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 O_2(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 $E^\ominus_{cell}$ and deduce which electrode is positive.

$E^\ominus_{cell}$ = ... V

identity of the positive electrode = ...

[1]

23. 9701/41/M/J/20 Q7a

(iv) An electrochemical cell is set up to measure the electrode potential, *E*, for the Ag^+/Ag half-cell using the saturated $Ag_2CO_3(aq)$ 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 Ag^+/Ag half-cell.

E for Ag^+/Ag half-cell = ... V [2]

24. 9701/42/M/J/20 Q2f

$$IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$$

(f) The half-equation for the reduction of iodate(V) ions is shown.

$$IO_3^- + 6H^+ + 5e^- \rightarrow \tfrac{1}{2}I_2 + 3H_2O \qquad E^\ominus = +1.19\,V$$

Use data from the *Data Booklet* to predict whether a reaction is feasible when aqueous solutions of acidified iodate(V) ions and bromide ions are mixed. Explain your answer.

....

.... [1]

25. 9701/42/M/J/20 Q6

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

....

....

.... [2]

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

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

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

[1]

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

$E^{\ominus}_{cell}$ = ... V [1]

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

.... [1]

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

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

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

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

26. 9701/42/F/M/20 Q2b

(b) Aluminium is extracted from Al_2O_3 by electrolysis. Al_2O_3 is dissolved in cryolite in this process.

(i) The half-equation for the reaction at the anode is shown.

$$O^{2-} + C \rightarrow CO + 2e^-$$

Use this half-equation to write the ionic equation for the electrolysis of Al_2O_3.

.... [1]

(ii) Aluminium oxide is electrolysed for 3.0 hours using carbon electrodes and a current of 3.5×10^5 A.

Calculate the mass of aluminium that is formed.

mass of aluminium = ... g [3]

(iii) Cryolite can be made from SiF_4.

The first step in this conversion is the reaction of SiF_4 with H_2O, forming H_2SiF_6 and SiO_2.

Write an equation for this reaction.

.... [1]

27. 9701/42/F/M/20 Q3

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

(a) The standard electrode potential, $E^\ominus$, of Au^{3+}(aq)/Au(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, $E^\ominus_{cell}$, of Au^{3+}(aq)/Au(s) and HNO_3(aq)/NO(g).

Include all necessary chemicals.

[4]

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

	half-equation	$E^\ominus$/V
1	$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	+1.50
2	$[AuCl_4]^-(aq) + 3e^- \rightleftharpoons Au(s) + 4Cl^-(aq)$	+1.00
3	$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightleftharpoons NO(g) + 2H_2O(l)$	+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 $E^\ominus_{cell}$ of the reaction in **(a)(iii)**.

$E^\ominus_{cell}$ = ... 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 mol dm^{-3}. Concentrated nitric acid is 16 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]

28. 9701/41/O/N/21 Q1

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

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

A current of **x** A is passed through the solution for 14.0 minutes. 462 cm^3 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]

29. 9701/42/O/N/21 Q3

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.

$$HCl(g) + aq \rightarrow H^+(aq) + Cl^-(aq)$$

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

$$2HCl(g) \rightarrow H_2(g) + Cl_2(g)$$

When a current of 3.10A is passed through the solution for *Y* minutes, 351 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]

30. 9701/41/M/J/21 Q3b

(b) **(i)** Define the term *standard cell potential*, $E^{\ominus}_{cell}$.

....

....

.... [2]

(ii) Draw a fully labelled diagram of the apparatus that can be used to measure the cell potential of a cell composed of a Cu(II)/Cu electrode and an Fe(III)/Fe(II) electrode. Include all necessary reactants.

[3]

(c) The reaction between $S_2O_8^{2-}$(aq) and I^-(aq) is catalysed by adding a few drops of Fe^{3+}(aq).

(i) Use equations to show the catalytic role of Fe^{3+} in this reaction.

[2]

(ii) Fe^{3+}(aq) can oxidise I^-(aq), whereas $[Fe(CN)_6]^{3-}$(aq) cannot oxidise I^-(aq).

Use E° values to explain these observations.

....

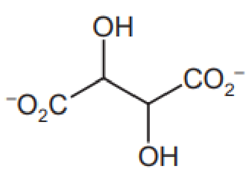
....

....

.... [2]

(d) When aqueous solutions of $S_2O_8^{2-}$ and tartrate ions are mixed the reaction proceeds very slowly. However, this reaction proceeds quickly in the presence of an Fe^{3+}(aq) catalyst. The overall equation for this reaction is shown.

tartrate ions



$$+ \ 3S_2O_8^{2-} + 2H_2O \rightarrow 2CO_2 + 2HCO_2^- + 6H^+ + 6SO_4^{2-}$$

(i) Suggest why this reaction is slow without the Fe^{3+} catalyst.

....

.... [1]

(ii) Use the overall equation to deduce the half-equation for the oxidation of tartrate ions, $C_4H_4O_6^{2-}$, to carbon dioxide, CO_2, and methanoate ions, HCO_2^-.

$C_4H_4O_6^{2-}$ + ... $\rightleftharpoons$ .... [1]

31. 9701/42/M/J/21 Q1e

(e) The $[Cr_2(O_2CCH_3)_4(H_2O)_2]$ complex reacts with aqueous acid to form Cr^{2+}(aq) ions.

Cr^{2+}(aq) ions react with O_2(aq) under acidic conditions. Cr^{3+}(aq) ions are formed.

Use the *Data Booklet* to answer the following questions.

(i) Construct an ionic equation for the reaction of Cr^{2+}(aq) with O_2(aq) under acidic conditions.

.... [2]

(ii) Calculate $E^{\ominus}_{cell}$ for the reaction in **(e)(i)**.

$E^{\ominus}_{cell}$ = ... V [1]

32. 9701/42/M/J/21 Q3a,b

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

....

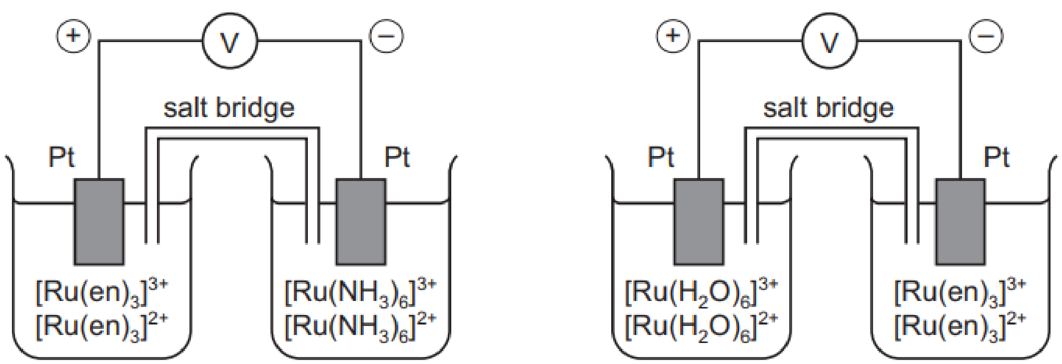
....

.... [2]

Three redox systems, **A**, **B** and **C**, are shown. The ligand 1,2-diaminoethane, $H_2NCH_2CH_2NH_2$, is represented by en.

A	$[Ru(H_2O)_6]^{3+}$ + e^- $\rightleftharpoons$ $[Ru(H_2O)_6]^{2+}$
B	$[Ru(NH_3)_6]^{3+}$ + e^- $\rightleftharpoons$ $[Ru(NH_3)_6]^{2+}$
C	$[Ru(en)_3]^{3+}$ + e^- $\rightleftharpoons$ $[Ru(en)_3]^{2+}$

Two electrochemical cells are set up to compare the standard electrode potentials, E°, 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° for the three half-cells.

E°	redox system
most negative	
↑	
least negative	

[1]

(b) **(i)** An electrochemical cell consists of a Br_2/Br^- half-cell and a Ag^+/Ag half-cell, under standard conditions.

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

$E^\ominus_{cell}$ = ... V

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

(ii) Water is added to the Ag^+/Ag half-cell in **(b)(i)**.

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

less positive	no change	more positive

Explain your answer.

....

....

.... [2]

33. 9701/42/M/J/21 Q4b

(b) **(i)** Identify the substances formed at the anode and at the cathode during the electrolysis of saturated $CaCl_2(aq)$.

at the anode ....

at the cathode ....
[1]

(ii) Calcium can be produced by the electrolysis of molten calcium chloride, $CaCl_2(l)$.

Calculate the mass, in g, of Ca formed when a current of 0.75A passes through $CaCl_2(l)$ for 60 minutes.
[A_r: Ca, 40.1]

mass of Ca = ... g [2]

34. 9701/42/F/M/21 Q1b

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

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

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

aqueous solution of Co^{3+}

....

....

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

....

....
[3]

35. 9701/42/F/M/21 Q2b

Fe_3O_4(l) can be electrolysed using inert electrodes to form Fe.

(ii) Write the half-equation for the reaction that occurs at the anode during the electrolysis of Fe_3O_4(l).

.... [1]

(iii) Calculate the maximum mass of iron metal formed when Fe_3O_4(l) is electrolysed for six hours using a current of 50 A.

Assume the one Fe^{2+} and two Fe^{3+} ions are discharged at the same rate.

mass of iron = ... g [3]

(c) $LiFePO_4$ can be used in lithium-ion rechargeable batteries.

When the cell is charging, lithium reacts with a graphite electrode to form LiC_6.

When the cell is discharging, the half-equations for the two processes that occur are as follows.

anode half-equation $LiC_6 \rightarrow 6C + Li^+ + e^-$

cathode half-equation $Li^+ + FePO_4 + e^- \rightarrow LiFePO_4$

(i) State one possible advantage of developing cells such as lithium-ion rechargeable batteries.

.... [1]

(ii) Use the cathode half-equation to determine the change, if any, in oxidation states of lithium and iron at the **cathode** during discharging.

metal	change in oxidation state during discharging	
	from	to
lithium		
iron		

[1]

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

.... [1]

36. 9701/42/F/M/21 Q3c

(c) The decomposition of hydrogen peroxide, H_2O_2, is catalysed by acidified IO_3^-.

H_2O_2 reduces acidified IO_3^- as shown.

$$5H_2O_2 + 2H^+ + 2IO_3^- \rightarrow I_2 + 5O_2 + 6H_2O$$

This reaction is followed by the oxidation of I_2 by H_2O_2.

half-equation	$E^\ominus$/V
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.77
$IO_3^- + 6H^+ + 5e^- \rightleftharpoons \frac{1}{2}I_2 + 3H_2O$	+1.19
$O_2 + 2H^+ + 2e^- \rightleftharpoons H_2O_2$	+0.68

(i) Use the data to show that the separate reactions of H_2O_2 with IO_3^- and with I_2 are both feasible under standard conditions.

In your answer, give the equation for the reaction of H_2O_2 with I_2.

....

....

....

....

....

.... [3]

(ii) Write the overall equation for the decomposition of H_2O_2 catalysed by acidified IO_3^-.

.... [1]

37. 9701/41/O/N/22 Q3

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

Table 3.1

electrode reaction	$E^\ominus$ / V
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$2HOCl + 2H^+ + 2e^- \rightleftharpoons Cl_2 + 2H_2O$	+1.64
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89
$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$	+0.15
$Sn^{2+} + 2e^- \rightleftharpoons Sn$	–0.14
$V^{2+} + 2e^- \rightleftharpoons V$	–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

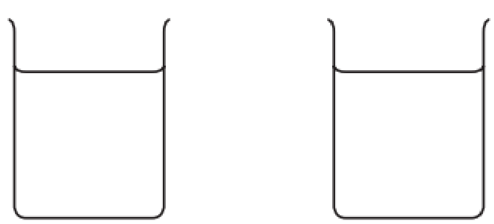
(a) Standard electrode potentials are measured under standard conditions.

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

....

.... [1]

(ii) Complete the diagram below to show how $E^\ominus$ (Sn^{4+}/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 mol dm^{-3} Sn^{2+}(aq) and 1.0 mol dm^{-3} Cl^{-}(aq) are mixed.

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

....

....

....

.... [2]

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

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

.... [2]

(b) A solution of $SnCl_2$(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.

Al_2O_3(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]

38. 9701/42/O/N/22 Q3

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

Table 3.1

electrode reaction	$E^⦵$/V
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	–2.38
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	–1.18
$Mn^{3+} + e^- \rightleftharpoons Mn^{2+}$	+1.49
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + e^- \rightleftharpoons MnO_4^{2-}$	+0.56
$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.67
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.52

(a) An electrochemical cell can be constructed from a Mg^{2+}/Mg half-cell and a MnO_4^-/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.

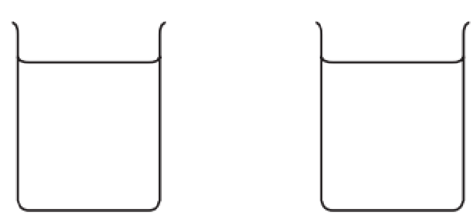
....

....

....

.... [2]

(ii) Complete the diagram below to show an electrochemical cell constructed from a Mg^{2+}/Mg half-cell and a MnO_4^-/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, $E^{\ominus}_{cell}$, of this cell.

$E^{\ominus}_{cell}$ = ... 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 Mg^{2+} / 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.

4.75×10^{22} magnesium atoms are produced at the cathode.

Calculate the value of the current used.

current = ... A [2]

39. 9701/41/M/J/22 Q2e

(iii) Table 2.1 lists relevant electrode potentials for some electrode reactions.

Table 2.1

electrode reaction	$E^\ominus$/V
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	–1.18
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$2HOCl + 2H^+ + 2e^- \rightleftharpoons Cl_2 + 2H_2O$	+1.64
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.67

Suggest the formula of the manganese species formed when Mn^{2+}(aq) reacts with Cl_2.

State the type of reaction.

formula of manganese species formed ....

type of reaction ....

[1]

40. 9701/41/M/J/22 Q3c

(c) **(i)** Define standard electrode potential, $E^\ominus$.

....

....

.... [1]

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

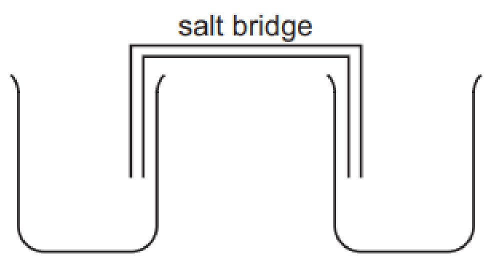
State the function of the salt bridge. Explain your answer.

....

.... [1]

(iii) Complete the diagram of the apparatus that can be used to measure the $E^\ominus$ of the $Cr_2O_7^{2-}(aq), H^+(aq)/Cr^{3+}(aq)$ electrode against the standard hydrogen electrode.

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



[3]

(iv) The $E^\ominus$ of the $Cr_2O_7^{2-}(aq), H^+(aq)/Cr^{3+}(aq)$ 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	$E^⦵$/V
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$CH_3CHO + 2H^+ + 2e^- \rightleftharpoons CH_3CH_2OH$	–0.61
$CH_3COOH + 2H^+ + 2e^- \rightleftharpoons CH_3CHO + H_2O$	–0.94
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23

(i) Ethanal is oxidised to ethanoic acid in the presence of $Cr_2O_7^{2-}$ ions.

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

ionic equation ....

$E^⦵_{cell}$ = ... V
[2]

(ii) In an ethanol-oxygen fuel cell, $CH_3CH_2OH(l)$ and $O_2(g)$ are in contact with two inert electrodes immersed in an acidic solution.

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

$$2CH_3CH_2OH + O_2 \rightarrow 2CH_3COOH + 2H_2O \qquad E^⦵_{cell} = +2.01\,V$$

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

$\Delta G^⦵$ = ... $kJ\,mol^{-1}$ [2]

41. 9701/41/M/J/22 Q4c

(c) Standard electrode potentials can be used to compare the stability of different complex ions for a given transition element.

Table 4.1 lists electrode potentials for some electrode reactions for Fe^{3+}/Fe^{2+} systems.

Table 4.1

electrode reaction	$E^{\ominus}$/V
$[Fe(H_2O)_6]^{3+} + e^- \rightleftharpoons [Fe(H_2O)_6]^{2+}$	+0.77
$[Fe(CN)_6]^{3-} + e^- \rightleftharpoons [Fe(CN)_6]^{4-}$	+0.36
$[Fe(bipy)_3]^{3+} + e^- \rightleftharpoons [Fe(bipy)_3]^{2+}$	+0.96

Use relevant data from Table 4.1 to state which iron(III) complex is hardest to reduce. Explain your choice.

iron(III) complex ....

explanation ....

....

[1]

42. 9701/42/M/J/22 Q5

(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
$PbBr_2(l)$		
concentrated $NaCl(aq)$		
$Cu(NO_3)_2(aq)$		

[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 0.600 A is passed through the acid for 30.0 minutes, the anode decreases in mass by 0.350 g.

(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]

43. 9701/42/F/M/22 Q3c

(c) Acidified Ti^{3+}(aq) reacts with oxygen dissolved in water as shown.

$$4Ti^{3+} + O_2 + 2H_2O \rightarrow 4TiO^{2+} + 4H^+ \qquad \Delta G^\ominus = -436.1\,kJ\,mol^{-1}$$

The standard reduction potential, $E^\ominus$, of $O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$ is +1.23 V.

(i) Calculate the standard reduction potential, $E^\ominus$, in V, of the TiO^{2+}(aq) / Ti^{3+}(aq) half-cell. Show your working.

$E^\ominus$ = ... V [3]

(iii) Acidified $[Ti(C_6H_5O_7)_2]^{3-}$(aq) does not react with oxygen dissolved in water, unlike acidified Ti^{3+}(aq).

Suggest what this means for the value of the standard reduction potential, $E^\ominus$, of the following half-cell.

$$[Ti(C_6H_5O_7)_2]^{2-}(aq) + e^- \rightleftharpoons [Ti(C_6H_5O_7)_2]^{3-}(aq)$$

Explain your answer.

....

.... [1]