Website: contoureducation.com.au | Phone: 1800 888 300
Email: hello@contoureducation.com.au

VCE Chemistry ¾
AOS 2 Revision I [2.5]
Contour Check Solutions

Contour Check

[2.1] -Introduction to Electrolysis (Checkpoints)

- [2.1.1] - Identify Differences between Galvanic & Electrolysis for Electrodes, Energy Conversions, Electron Flow Pg 3-6
- [2.1.2] - Write Equations & Calculate EMF Required for Electrolytic Reactions Pg 7-12

[2.2] -Features of Electrolytic Cells (Checkpoints)

- [2.2.1] - Find Electrolytic Reactions in Non-Standard Conditions (Molten & High Concentration) Pg 13-16
- [2.2.2] - Identify Features of Electrolytic Cells & Their Purpose Pg 17-21
- [2.2.3] - Identify Key Features, Write Reactions & Relate to Sustainability & Green Chemistry Principles Regarding Production of Green Hydrogen (PEM & Artificial Photosynthesis) Pg 22-24

[2.3] -Secondary Cells & Connected Cells (Checkpoints)

- [2.3.1] - Write Discharge & Recharge Reactions in Secondary Cells & Redox Flow Batteries Pg 25-29
- [2.3.2] -Identify Factors Which Affect Rechargeability & Compare Similarities/Differences between Secondary Cells and Other Cells Pg 30-34
- [2.3.3] -Find Reactions Occurring in Connected Cells Pg 35-37

[2.4] -Electroplating (Checkpoints)

- [2.4.1] -Identify the Electroplating Setup (Location of Object) & Find the Electroplating Reactions Pg 38-40
- [2.4.2] -Find Next Order Reactions During Electrolysis Pg 41-42
- [2.4.3] - Apply Faraday's Laws to Electroplating Calculations Pg 43-46

[2.1 - 2.4] - Overall (VCAA Qs) Pg 47-72

Section A: [2.1] - Introduction to Electrolysis (Checkpoints) (27 Marks)

Sub-Section [2.1.1]: Identify Differences between Galvanic & Electrolysis for Electrodes, Energy Conversions, Electron Flow

Question 1 (3 marks)

a. Which of the following statements about electrolytic cells is correct? (1 mark)

A. Electrolytic cells undergo exothermic reactions that result in a lower enthalpy of products than reactants.

B. Electrolytic cells are expensive to produce due to their requirement for Pt(s) catalysts.

C. Electrolytic cells have the weakest oxidant and weakest reductant react together.

D. Electrolytic cells are an example of non-spontaneous reactions that produce chemical energy.

b. Determine the **false** statement regarding the properties of galvanic and electrolytic cells. (1 mark)

A. Reduction in galvanic and electrolytic cells occurs at the cathode.

B. Electrons flow from cathode to anode in an electrolytic cell.

C. Galvanic cells convert chemical energy to electrical energy whilst electrolytic cells convert electrical energy to chemical energy.

D. The cathode is positively charged in a galvanic cell.

c. The cathode of an electrolytic cell: (1 mark)

A. Has a positive polarity and reduction occurs.

B. Has a positive polarity and oxidation occurs.

C. Has a negative polarity and reduction occurs.

D. Has a negative polarity and oxidation occurs.

Space for Personal Notes

Question 2 (2 marks)

Betty has set up a cell at school which requires electrical input to operate.

a. State the energy conversion occurring in the cell. (1 mark)

Electrical energy → Chemical energy (1).

b. Hence, determine whether electrolysis is exothermic or endothermic (1 mark)

Endothermic.

Space for Personal Notes

Question 3 (2 marks)

a. Select the option which accurately describes the electron flow of galvanic and electrolytic cells. (1 mark)

A.

Cell type	**Electron flow**
Galvanic	Positively to negatively charged electrode.
Electrolytic	Positively to negatively charged electrode.

B.

Cell type	**Electron flow**
Galvanic	Positively to negatively charged electrode.
Electrolytic	Negatively to negatively charged electrode.

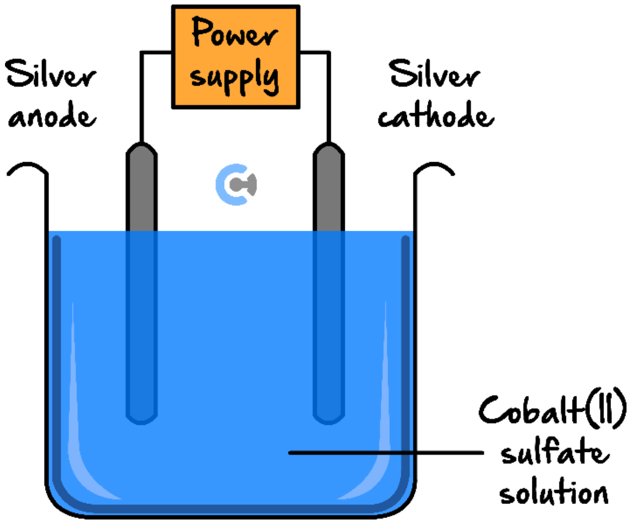
C.

Cell type	**Electron flow**
Galvanic	Negatively to positively charged electrode.
Electrolytic	Positively to negatively charged electrode.

D.

Cell type	**Electron flow**
Galvanic	Negatively to positively charged electrode.
Electrolytic	Negatively to positively charged electrode.

b. Katie is analysing the electrolytic cell shown below:



Determine the purpose of the Cobalt (II) sulphate solution. (1 mark)

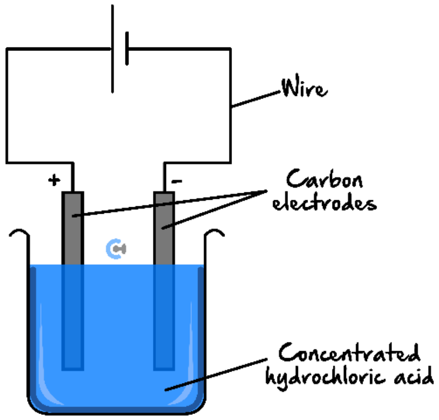
To complete the circuit due to the absence of a salt bridge (1).

Space for Personal Notes

Sub-Section [2.1.2]: Write Equations & Calculate EMF Required for Electrolytic Reactions

Question 4 (9 marks)

a. Michelle is electrolysing 3.0 *M* of HCl in the cell shown below.



i. Write the equation which occurs at the cathode. (1 mark)

$$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$$

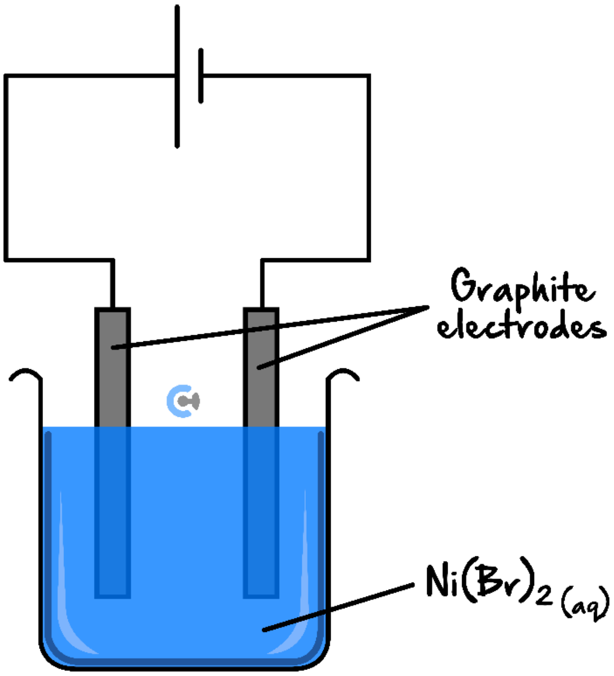
ii. Write the equation which occurs at the anode. (1 mark)

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$$

iii. Hence, write the overall equation for the cell. (1 mark)

$$2HCl_{(aq)} \longrightarrow Cl_{2(g)} + H_{2(g)}$$

b. Dai is operating a cell containing 1.0 *M* of nickel (II) bromide, shown below.



i. Write the equation which occurs at the cathode. (1 mark)

$$Ni^{2+}(aq) + 2e^{-} \rightleftharpoons Ni(s)$$

ii. Write the equation which occurs at the anode. (1 mark)

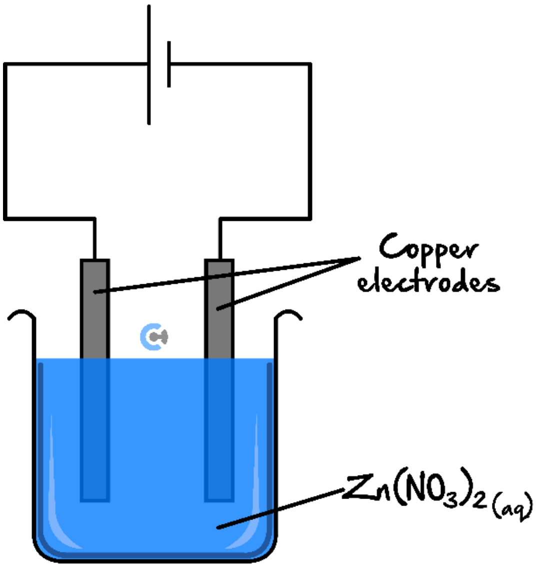
$$2Br^{-}_{(aq)} \rightarrow 2e^{-} + Br_{(l)}$$

iii. Hence, write the overall equation for the cell. (1 mark)

$$CO^{2+}(aq) + 2Br^{-}(aq) \rightarrow CO(s) + Br(I)$$

$$CoBr(aq) \rightarrow CO(s) + Br(I)$$

c. Henry is operating a cell containing 1.0 *M* of zinc nitrate. The cell is shown below.



i. Write the equation which occurs at the cathode. (1 mark)

$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$

ii. Write the equation which occurs at the anode. (1 mark)

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$

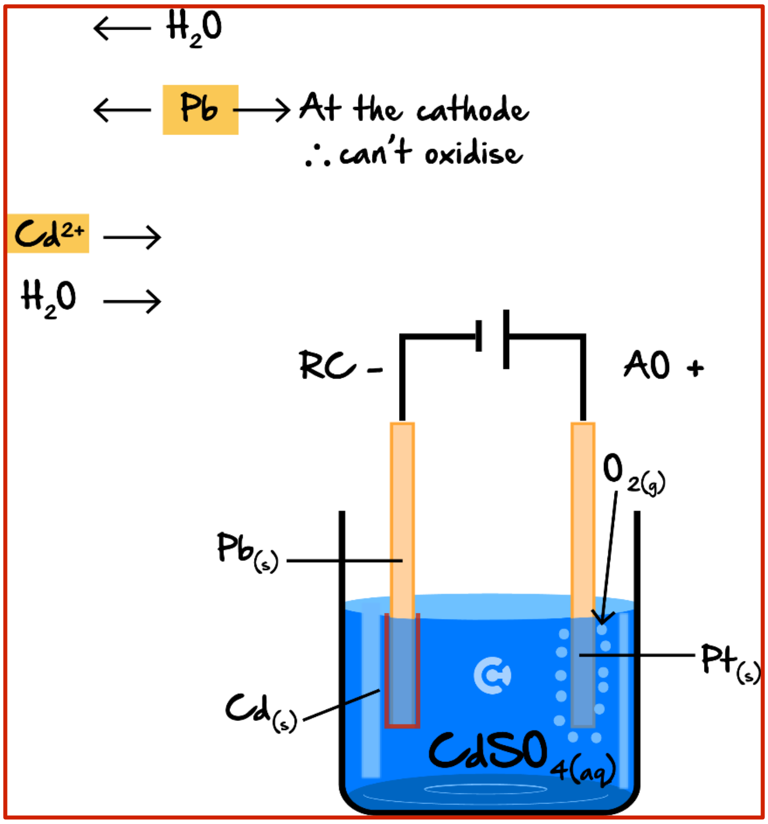
iii. Hence, write the overall equation for the cell. (1 mark)

$Cu_{(s)} + Zn^{2+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + Zn_{(s)}$

Space for Personal Notes

Question 5 (5 marks)

a. Draw a labelled cell for the electrolysis of $CdSO_4(aq)$ with a lead cathode and platinum anode. (2 marks)



b. Write the balanced half-equations occurring at the:

i. Anode. (0.5 marks)

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

ii. Cathode. (0.5 marks)

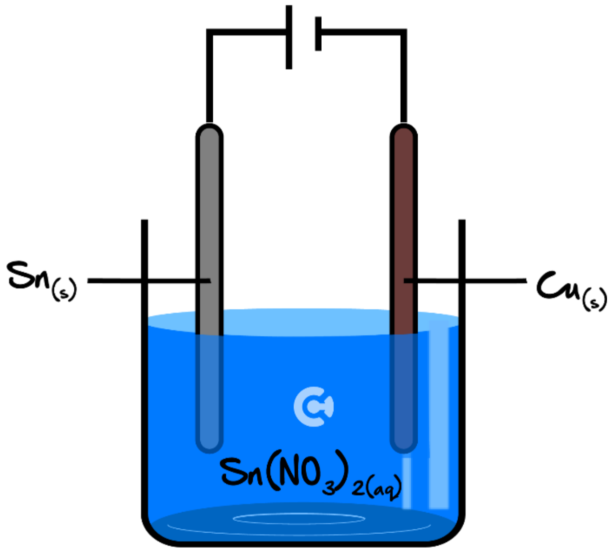
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

c. List any observations made at either electrode. (2 marks)

- Decrease in pH at the anode.
- O_2 bubbles at the anode.
- Increase in mass/size of the cathode.

Question 6 (3 marks)

A solution of Tin (II) nitrate is being electrolysed in an electrolytic cell containing a copper cathode and tin anode. A diagram of the electrolytic cell is shown in the diagram below.



a. Write the balanced half-equations occurring at the:

i. Anode. (0.5 marks)

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

ii. Cathode. (0.5 marks)

$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

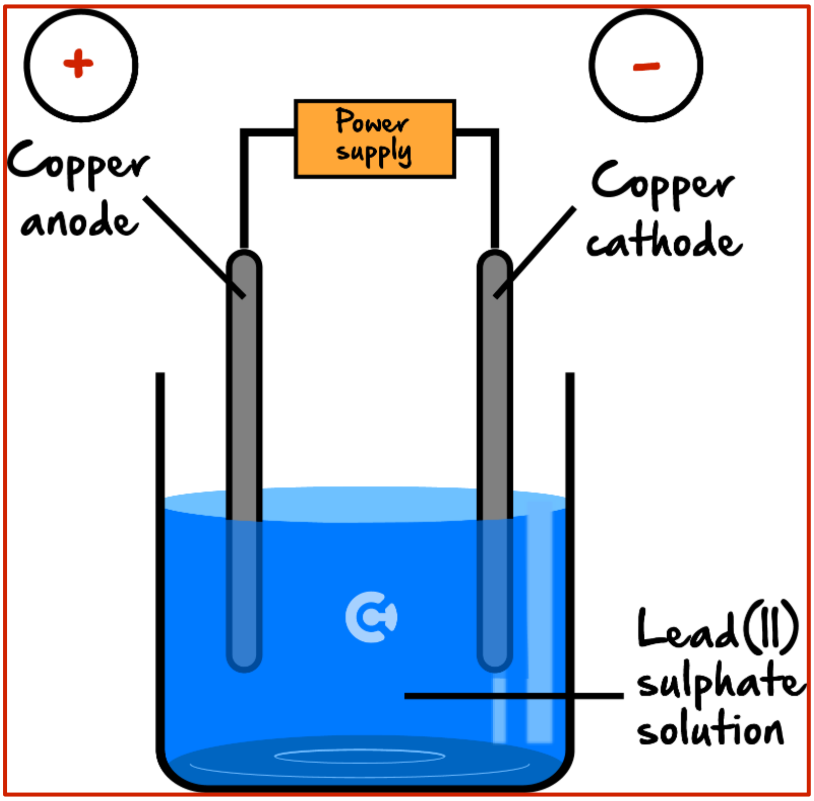
b. The cell is able to be ran for 30 minutes under standard conditions. When a separator membrane was added that prevented the flow of ions, the cell was only able to be run for 5 minutes. Explain what the separator membrane did to limit the run time of the cell. (2 marks)

The cell operates through Sn^{2+} ions being produced at the anode, that are then reduced at the cathode, effectively transferring Sn from anode to cathode. By adding the separator membrane, the Sn^{2+} produced at the anode was no longer able to move to the cathode. Thus, the cell stopped operating after it ran out of Sn^{2+} in the electrolyte and was unable to access the Sn^{2+} in the anode.

Space for Personal Notes

Question 7 (3 marks)

The electrolysis of 1 *M* solution of Lead (II) sulphate ($PbSO_4$) is undertaken with copper electrodes at 25°C. A diagram of this electrolytic process is shown in the following diagram below.



a. Label the polarity of each electrode in the diagram above. (1 mark)

b. Write the half-equations which occur at the:

i. Cathode. (0.5 marks)

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

ii. Anode. (0.5 marks)

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

c. List any observations seen at either electrode. (1 mark)

- Increase in mass/size of the cathode.
- Decrease in mass/size of the anode.
- Increase in intensity of blue colour.

Section B: [2.2] - Features of Electrolytic Cells (Checkpoints) (47 Marks)

Sub-Section [2.2.1]: Find Electrolytic Reactions in Non-Standard Conditions (Molten & High Concentration)

Question 8 (3 marks)

A solution of magnesium bromide is being electrolysed using a copper cathode and carbon anode at molten conditions.

a. Write a balanced half-equation (including states) for the reaction occurring at the:

i. Anode. (1 mark)

$2Br^-(l) \rightarrow Br_2(l) + 2e^-$

ii. Cathode. (1 mark)

$Mg^{2+}(l) + 2e^- \rightarrow Mg(l)$

b. State the voltage required for this cell to operate. (1 mark)

Greater than 3.46 V.

Space for Personal Notes

Question 9 (4 marks)

Molten aluminium nitrate is being electrolysed using a nickel cathode and iron anode.

a. Write the balanced half-equation occurring at the:

i. Negative electrode. (1 mark)

$Al^{3+}(l) + 3e^- \rightarrow Al(l)$

ii. Positive electrode. (1 mark)

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

b. If we changed the nickel cathode into a zinc cathode, state any changes that might occur to the reaction. (2 marks)

No change in the reaction since at the cathode only reduction can occur so even though zinc in essence is a stronger reductant than iron, but it cannot react.

Space for Personal Notes

Question 10 (2 marks)

An electrolytic cell is set up using inert electrodes inside a 1.0 M solution of NaCl. A diagram of the shell is shown below.

a. Over the progression of the reaction:

A. Bubbles will be served at both electrodes.

B. Sodium metal will begin to plate the cathode and chlorine gas at the anode.

C. Sodium ions will be produced at the positive electrode and chlorine gas will be produced at the cathode.

D. Bubbles will only appear at one of the electrodes.

b. The concentration of NaCl is then increased to 5.0 M. The reaction at the negative electrode will now be:

A. $Na^+(aq) + e^- \rightarrow Na(s)$

B. $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

C. $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

D. $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Space for Personal Notes

Question 11 (6 marks)

A molten cell is set up containing a solution of potassium sulphate and is mixed with a solution of lithium fluoride. The electrolytic cell is then set up using a graphite cathode and lead anode.

a. Write the half-equation occurring at the:

i. Anode. (1 mark)

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

ii. Cathode. (1 mark)

$K^+(l) + e^- \rightarrow K(l)$

b. Explain an observation that takes over the course of the reaction. (2 marks)

The lead anode will decrease in size as it undergoes oxidation into lead ions.

c. Suggest an improvement to this cell which would allow it to be operational over a longer period of time. (2 marks)

Swap the anode and cathode, as then the lead electrode will no longer degrade over time. Or else, once the electrode is witted down to the point where it doesn't touch the electrolyte, the cell will stop working.

Space for Personal Notes

Sub-Section [2.2.2]: Identify Features of Electrolytic Cells & Their Purpose

Question 12 (2 marks)

a. Which of the following is an advantage of using a molten electrolyte in an electrolytic cell?

A. Molten electrolytes reduce the melting point of the electrodes.

B. Molten electrolytes decrease the conductivity of the cell.

C. Molten electrolytes allow for the electrolysis of compounds that are insoluble.

D. Molten electrolytes prevent any ion movement.

b. Why might an electrolytic cell use a lead anode in the electrolysis of a fluoride salt?

A. Lead reacts vigorously with fluorine gas, enhancing the reaction.

B. Lead is a poor conductor of electricity.

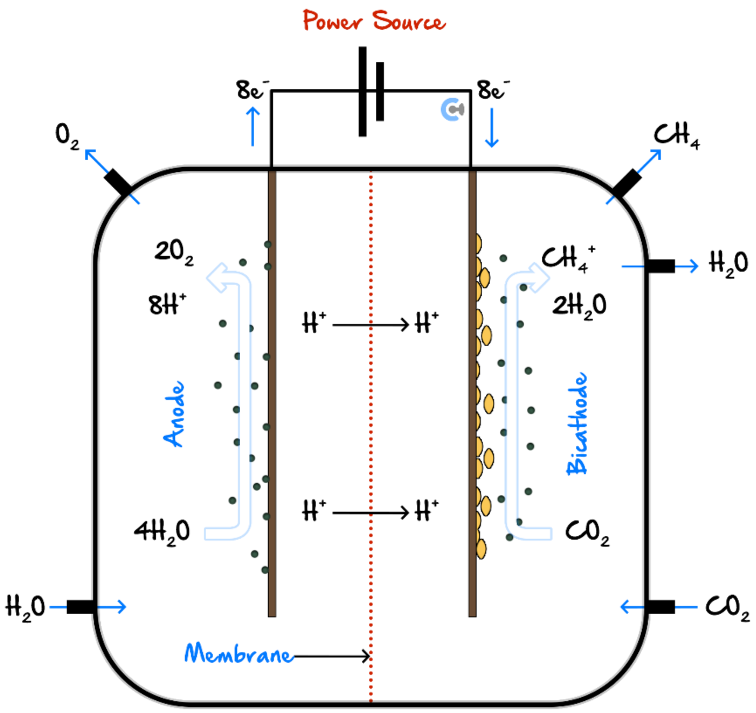
C. Lead is resistant to oxidation and minimises anode degradation.

D. Lead actively releases oxygen gas when in contact with fluoride ions.

Space for Personal Notes

Question 13 (3 marks)

Electromethanogenesis cells use bacteria and organic matter to turn CO_2 back into CH_4 (biogas) as a form of renewable energy production. A diagram is shown below.



a. What is the reaction that occurs at the anode? (1 mark)

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

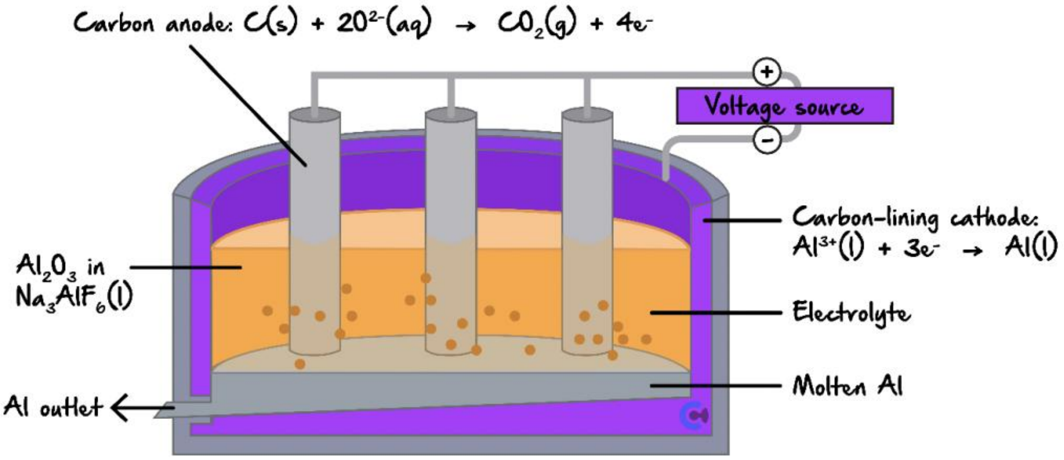
b. Biogas is the primary product, is it renewable? Explain. (2 marks)

It depends on where the power source is sourced from, if the power source comes from a traditional source, then it isn't renewable since it is still derived from fossil fuels.

Space for Personal Notes

Question 14 (5 marks)

Consider the following Hall-Heroult cell.



a. If the cell is typically enclosed, explain why it is necessary to include a gas pump that is separate from the Al outlet. (2 marks)

This is because if the cell is enclosed there would be a gas build-up, and if the gas build-up is too much the battery would burst from the internal pressure.

b. If the aluminium oxide was dissolved in a solution of water, what would occur to the aluminium being produced? (2 marks)

It would undergo water electrolysis instead.

$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$

$2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$

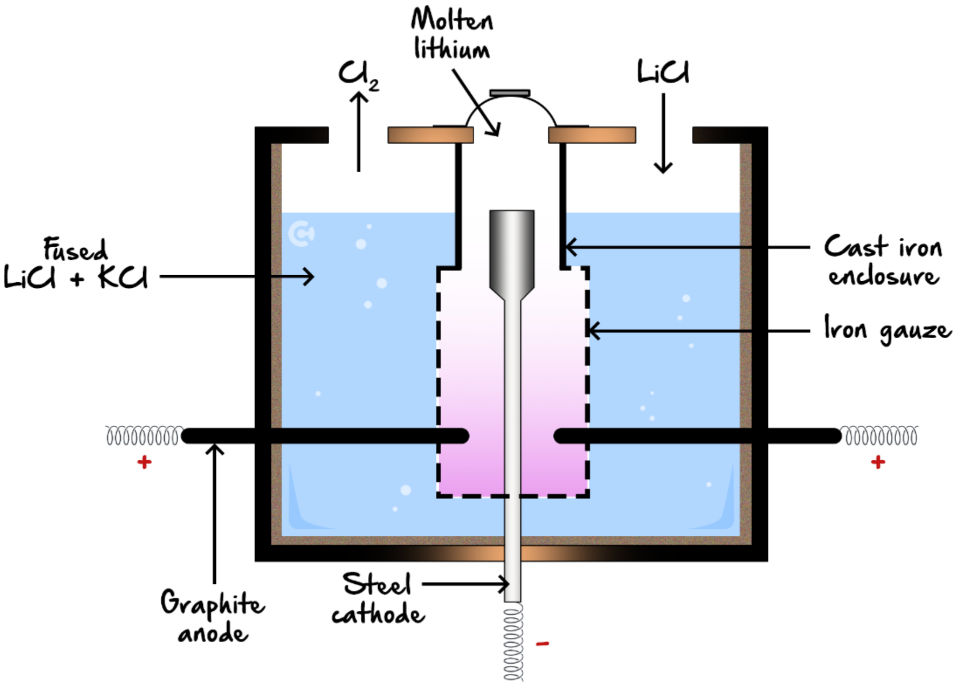
c. What is the overall reaction occurring in the cell? (1 mark)

$2Al_2O_3(l) + 3C(s) \rightarrow 4Al(l) + 3CO_2(g)$

Space for Personal Notes

Question 15 (7 marks)

Lithium, an alkali metal exhibits a diverse array of scientific applications, which is increasingly important in today's digital society. The below is an example of an industrial cell that uses lithium.



a. Write the reaction occurring at the:

i. Anode. (1 mark)

$2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$

ii. Cathode. (1 mark)

$Li^+(l) + e^- \rightarrow Li(l)$

b.

i. By using molten conditions, the cell operates much more safely than under standard conditions. Explain how this is possible. (2 marks)

Under standard conditions, water will be present. Hence, as lithium is a strong reductant and will react with water causing the production of H_2 gas which is flammable can cause explosion.

ii. Is the production of pure lithium possible under standard conditions? (1 mark)

No, it is not because water will always reduce in preference over lithium because it is a weak oxidant.

c. An iron gauze is placed around the cathode, explain its purpose. (2 marks)

Iron gauze prevents lithium metal produced to react with any other products or ions, keeping the lithium obtained pure as it is very reactive.

Space for Personal Notes

Sub-Section [2.2.3]: Identify Key Features, Write Reactions & Relate to Sustainability & Green Chemistry Principles Regarding Production of Green Hydrogen (PEM & Artificial Photosynthesis)

Question 16 (2 marks)

What is a key challenge with the practical usage of artificial photosynthesis, with regards to the product created?

As the main product is hydrogen gas, the storage of it is important because it is a flammable gas which may cause an explosion is not handled properly.

Question 17 (3 marks)

The PEM electrolyser is an innovative method of producing hydrogen gas.

a. Write its half-equations for the reactions at both the cathode and anode. (2 marks)

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

b. The setup of a PEM electrolyser is very specific. Describe the qualities of a typical electrode you might find in one of these cells. (1 mark)

The electrode needs to be PICCY (porous, inert, catalytic, conducts electricity).

Space for Personal Notes

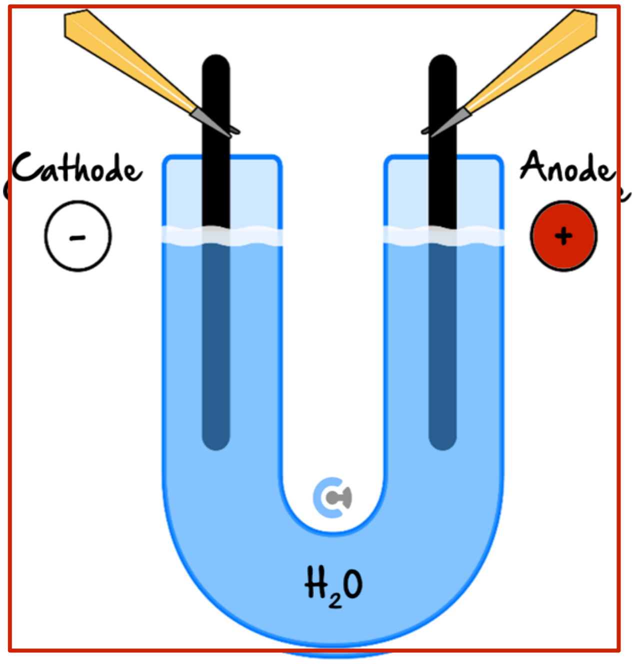
Question 18 (3 marks)

Explain the environmental advantages that you would get if you preferred to use artificial photosynthesis over a PEM electrolyser to produce hydrogen gas.

Artificial photosynthesis uses sunlight to power the electrochemical cell whereas the PEM electrolyser still requires a traditional power source that is from either solar or wind but it is then converted to electrical first before chemical. This makes artificial photosynthesis better as it is derived from the sun directly and there are no intermediate conversions.

Question 19 (7 marks)

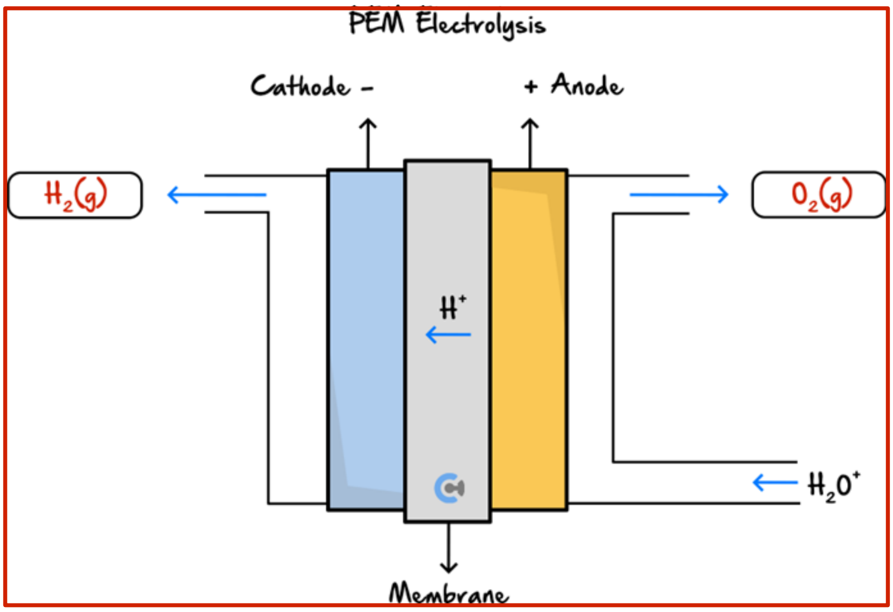
Hydrogen production can be done through two primary ways, electrolysis and steam reforming. An example of an electrolytic cell producing hydrogen gas is shown below.



a. This cell does not produce 'green' hydrogen. Explain why. (2 marks)

The energy that is used for hydrogen gas production in a simple electrolytic cell could be from any number of sources. With most of these sources being through unsustainable practice such as coal fired plants.

Polymer electrolyte membrane electrolysers (PEMs) are a method of producing green hydrogen gas.



b. On the diagram above, label the products at the anode and cathode in the boxes provided. (1 mark)

c.

i. Explain how hydrogen gas produced through PEM electrolysis can be considered 'green' whereas, that from simple electrolysis cannot. (2 marks)

A PEM electrolyser uses photovoltaic (solar) or wind energy to power the electrolytic reactions. Therefore, there are no harmful emissions in the entire process of using hydrogen gas, hence making the production of H_2 completely sustainable and green. Simple electrolysis can source energy from anywhere, and as such could be from unsustainable sources.

ii. Hence or otherwise, comment on the sustainability of the PEM electrolyser ensuring that you reference green chemistry principles and sustainable development goals. (2 marks)

Catalysis (green chem principle)
Clean water and sanitation, affordable and clean energy, climate action, sustainable cities and communities, industry, innovation and infrastructure, responsible consumption and production (Sustainable development goals).

Space for Personal Notes

Section C: [2.3] - Secondary Cells & Connected Cells (Checkpoints) (46 Marks)

Sub-Section [2.3.1]: Write Discharge & Recharge Reactions in Secondary Cells & Redox Flow Batteries

Question 20 (2 marks)

a. Zinc-bromine rechargeable batteries are commonly used within electric vehicles for efficient and quick releases of large amounts of energy.

The reaction occurring in a zinc-bromine cell during discharge is shown below.

$$Zn(s) + Br_2(l) \rightarrow ZnBr_2(s)$$

An observation noted during recharge will be: (1 mark)

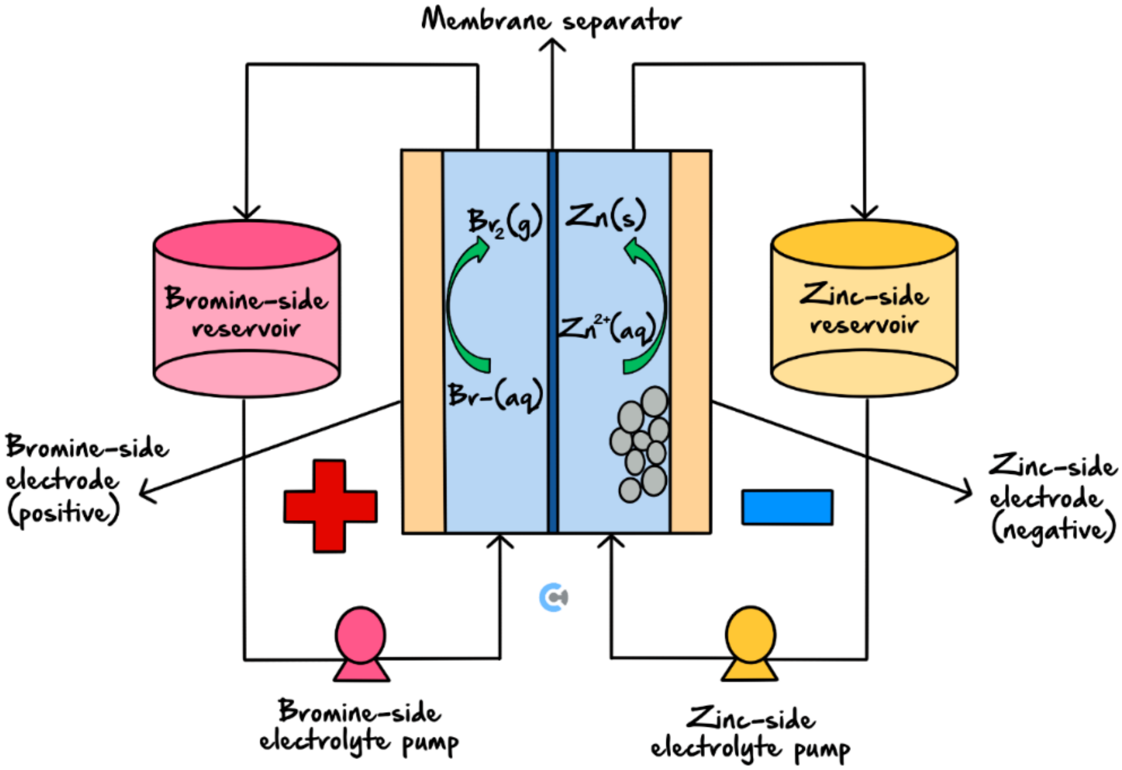
A. Decrease in mass of the cathode.

B. Increase in intensity of brown in the electrolyte.

C. Release of thermal energy.

D. A voltage produced of $+1.85\ V$.

b. An image of a zinc-bromine redox flow battery is shown below.



It can both discharge and recharge. The reactions that occur within the cell are shown below:

$$Br_2(l) + 2e^- \leftrightharpoons 2Br^-(aq) \qquad E^0 = +1.09\ V$$

$$Zn^{2+}(aq) + 2e^- \leftrightharpoons Zn(s) \qquad E^0 = -0.76\ V$$

An observation that would be made during the recharge of the cell would be: (1 mark)

A. Increase in mass at the bromine-side reservoir.

B. Decrease in mass at the zinc-side reservoir.

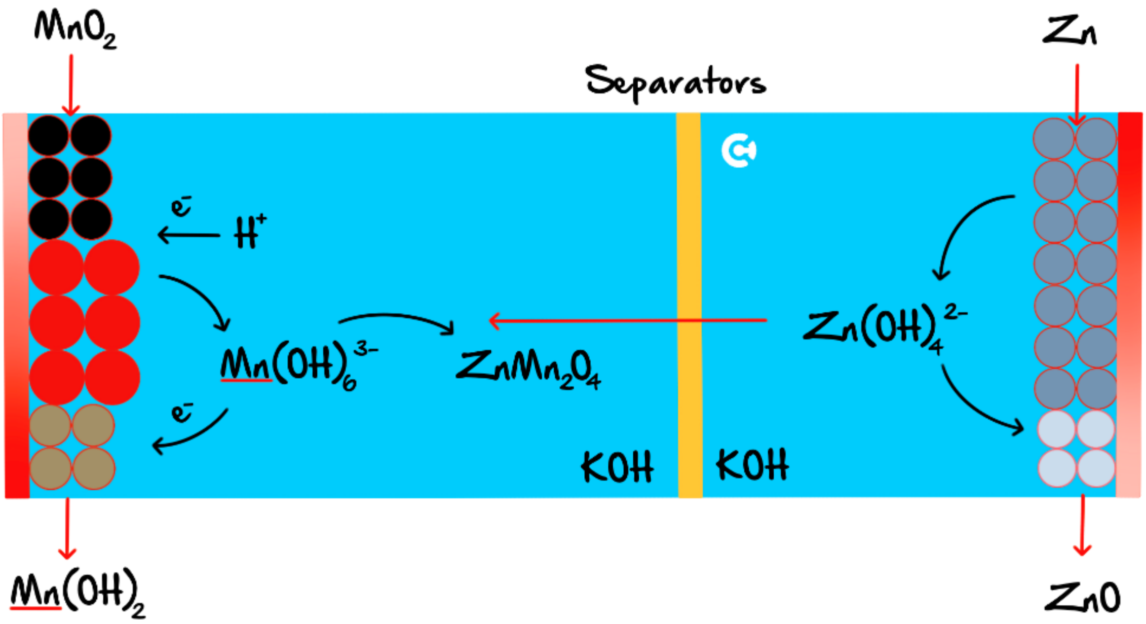
C. Brown bromine liquid being produced at the anode.

D. Zinc solid being produced at the anode.

Space for Personal Notes

Question 21 (4 marks)

A manganese dioxide zinc half-cell is shown below. It uses multi-step reactions to both recharge and discharge.



The overall reaction occurring during discharge is shown below:

$$MnO_2(s) + Zn(s) + H_2O(l) \rightarrow ZnO(s) + Mn(OH)_2$$

a. Write the balanced equation for the reaction that occurs at the anode during discharge. (1 mark)

$Zn(s) + 4OH^-(aq) \rightarrow ZnO(s) + 2H_2O(l) + 4e^-$

b. Write the balanced equation for the oxidation reaction occurring during recharge. (1 mark)

$Mn(OH)_2(s) + 2OH^-(aq) \rightarrow MnO_2(s) + 2H_2O(l) + 2e^-$

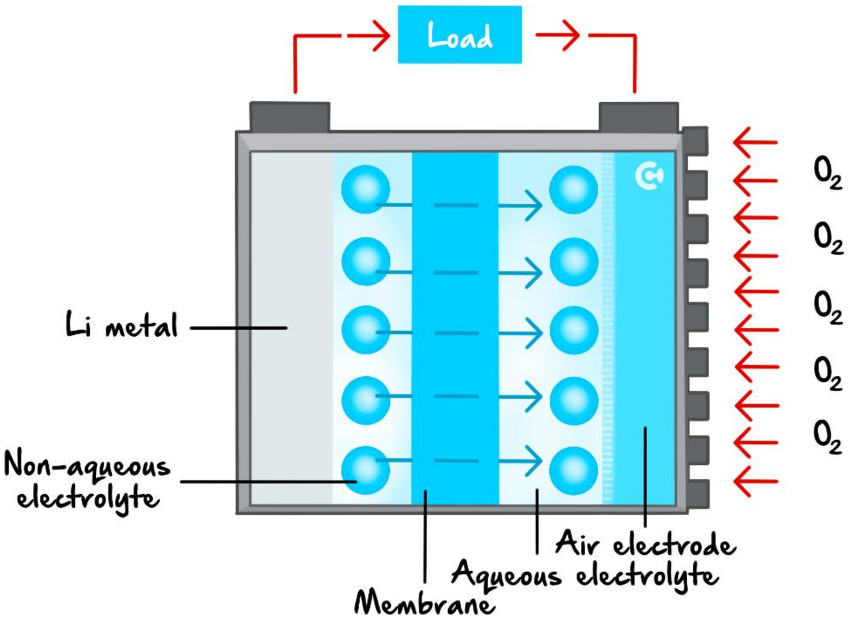
c. Explain the purpose of the separator in this secondary cell. (2 marks)

To allow for the flow of $Zn(OH)_4{}^{2-}$ ions.
Allow for the balance of the build up of charge.

Question 22 (6 marks)

Lithium-air batteries, celebrated for their promising high energy density, hold potential as next-generation power sources in various applications.

A diagram of a lithium-air battery is shown below.



The following reaction occurs in an acidic electrolyte whilst the cell is undergoing discharge:

$$4Li(s) + O_2(g) + 4H^+(aq) \rightarrow 4Li^+(aq) + 2H_2O(l)$$

a.

i. Write the reaction occurring at the cathode as the cell is producing energy. (1 mark)

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

ii. Hence or otherwise, identify the conjugate oxidant of the discharge reaction. (1 mark)

$Li^+(aq)$

b. Determine the voltage that must be inputted in order for the cell to be recharged. (1 mark)

$> 4.27\ V$

c. A non-aqueous electrolyte is used within some part of the cell. Explain a potential hazard in using an aqueous electrolyte throughout the entire cell. (3 marks)

A non-aqueous electrolyte contains no water. The lithium metal is a very strong reductant, and will spontaneously react with water under the reaction $H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$. In this, it produces $H_2(g)$ which is highly flammable, which can lead to an explosion.

Space for Personal Notes

Sub-Section [2.3.2]: Identify Factors Which Affect Rechargeability & Compare Similarities/Differences between Secondary Cells and Other Cells

Question 23 (1 mark)

Secondary cells are used in phone batteries due to their rechargability. Which of the following is a unique feature about the reactions in secondary cells?

A. The reactions both require the input of electrical energy in order to operate.

B. The products of discharge remain in contact with the electrodes.

C. The electrodes are inert and allow for theoretically indefinite rechargeability.

D. The discharge reactions of the cathode will occur at the other electrode during recharge.

Question 24 (2 marks)

a. Which of the following statements about secondary cells is correct? (1 mark)

A. Secondary cells can force electrons to travel against natural electrostatic forces to make the cathode negative.

B. Secondary cells minimise side reactions by always having a membrane between electrodes.

C. Secondary cells typically operate using porous electrodes in order to increase the surface area and increase the rate of reaction.

D. Secondary cells allow electrons to always follow natural electrostatic forces.

b. Which of the following statements accurately describes a feature of secondary cells? (1 mark)

A. Exhibit minimal self-discharge rates compared to primary cells.

B. Predominantly used in low-drain devices due to their limited capacity.

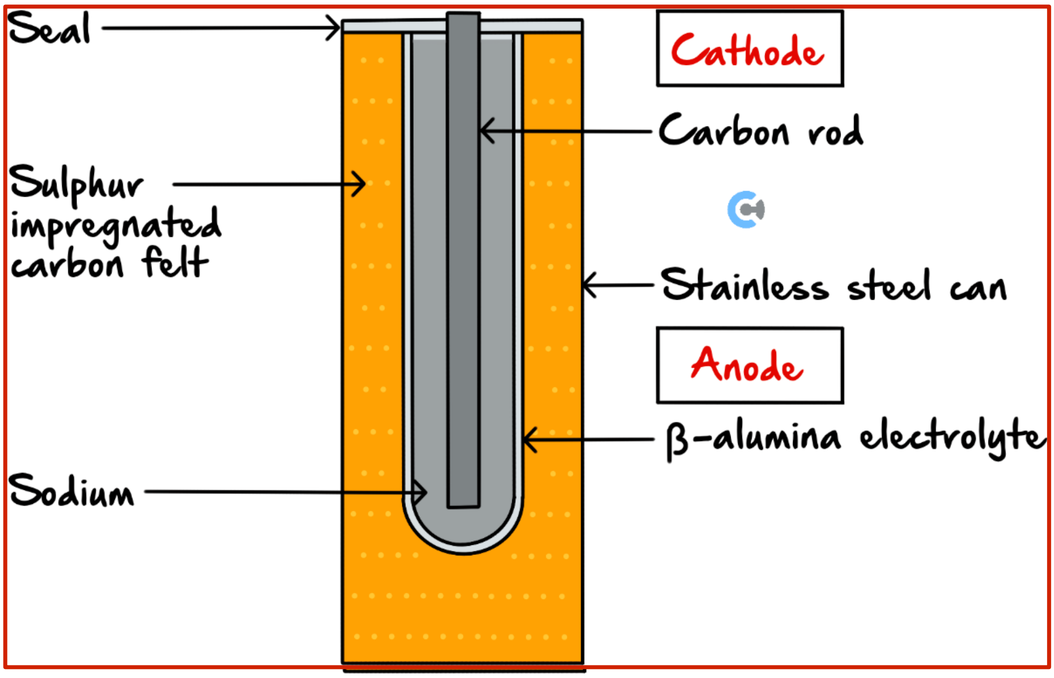
C. Rely on non-reversible chemical reactions for energy conversion.

D. Typically operate optimally in extreme temperature conditions due to their robust design.

Space for Personal Notes

Question 25 (6 marks)

Sodium-sulphur batteries are used for large-scale energy production and storage due to their high energy density and long life cycle. This allows them to have applications in storing renewably generated energy from sources such as wind and solar. It utilised molten sodium and sulphur, with a solid β-alumina electrolyte. A diagram of a sodium-sulphur cell is shown below.



The overall reaction occurring within the cell during discharge is:

$$Na(l) + S(l) \rightarrow Na_2S(l)$$

a. In the boxes provided above, label the location of the anode and cathode during discharge. (1 mark)

b. Write the reaction for the reaction occurring at the positive electrode as energy is being produced. (1 mark)

$$S(l) + 2e^- \rightarrow S^{2-}(l)$$

c.

i. β-alumina is added to the electrolyte for proper operation of the cell, it can be collected and reused after the reaction has been completed. Explain the role of the β-alumina within the cell. (2 marks)

Alumina is reducing the melting point of the substances within the cell through forming temporary partial ionic bonds within the metallic/ionic lattices of the products/reactants. This then allows for a lower temperature to be maintained and as such less energy has to be inputted.

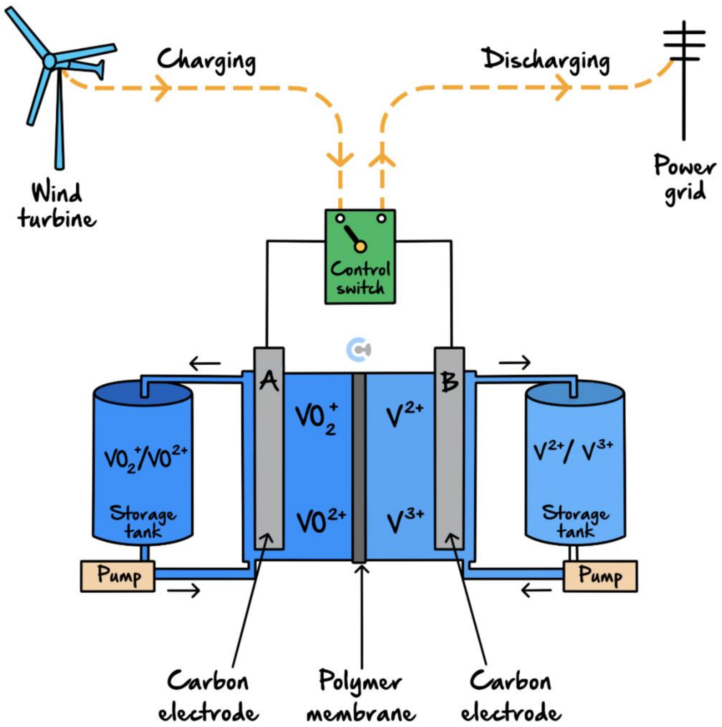
ii. There is a separator between either of the electrodes that allows for the flow of $Na^+(aq)$ ions. Explain two purposes for Na^+ions within the cell. (2 marks)

Na^+ acts as the electrolyte within the cell. It balances the build up of charge to ensure that the cell remains electircally neutral, allowing intended reactions to occur.

Space for Personal Notes

Question 26 (9 marks)

Vanadium redox flow batteries can be used as a means of adding energy generated by solar or wind energy to the power grid. An example of a typical vanadium redox flow battery is shown below.



a.

i. Given that electrode *B* is negative during discharge, state the reaction occurring at electrode *A* during recharge. (2 marks)

@ negative electrode *B* during discharge: $V^{2+} \rightarrow V^{3+} + e^-$
Positive *A* recharge is an oxidation, therefore: $VO^{2+} + H_2O \rightarrow VO_2^+ + 2H^+ + e^-$

ii. Hence or otherwise, explain why a polymer membrane must be added to the cell using relevant half-equations to support your response. (2 marks)

A polymer membrane prevents the movement of the products electrolysis from spontaneously reacting with one another. The reaction $V^{2+} + VO_2^+ + 2H^+ \rightarrow CO^{2+} + V^{3+} + H_2O$, which would cause the production of thermal energy as opposed to electrical energy.

b. Explain how the control switch will allow for non-spontaneous reactions to occur when taking energy from the wind turbine. (3 marks)

When the control switch is set to discharging, electrons are travelling in their natural directions towards the positive cathode, where they are gained by oxidants to undergo reduction. When the wind turbine produces electrical energy that is inputted to the redox flow battery that forces electrons to move in the against their natural electrostatic attraction. Therefore, the electrons travel towards the negative electrode, whereby they must be gained, causing non-spontaneous reduction to occur.

c. It can be claimed that the cell could be considered a fuel cell. Comment on the accuracy of this claim. (2 marks)

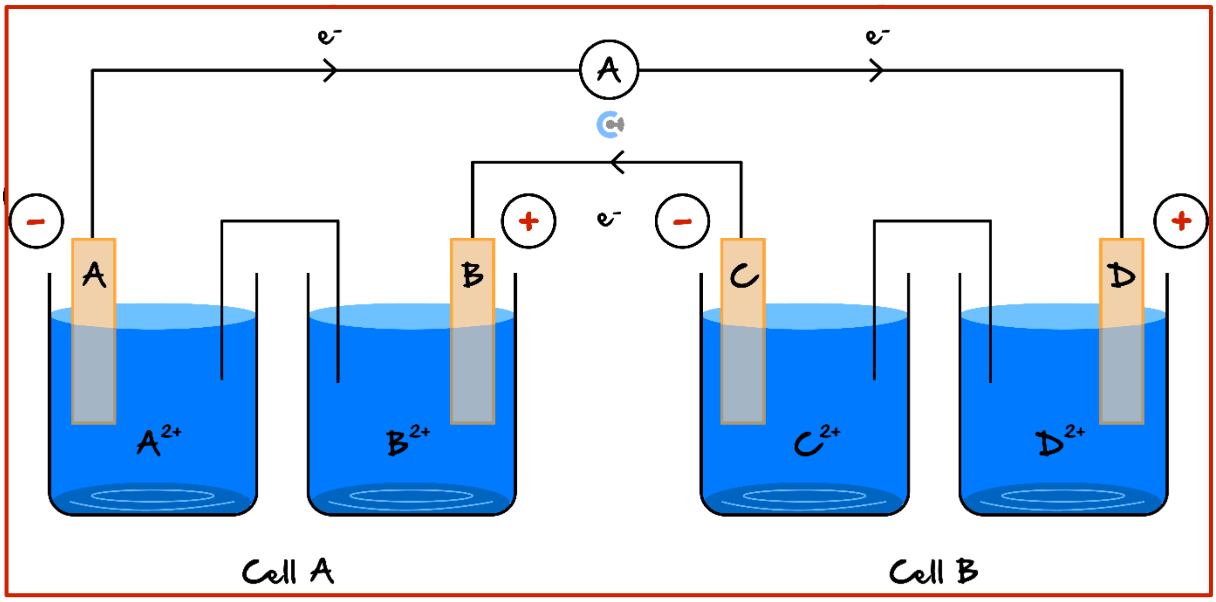
A fuel cell is one that requires a constant supply of reactants in order to operate. In this case, the cell could be considered to have a constant supply of reactants as they are continuously pumped in from the storage tanks. However, fuel cells only convert chemical energy into electrical energy, whereas this redox flow battery is able to both convert chemical into electrical and electrical into chemical.

Space for Personal Notes

Sub-Section [2.3.3]: Find Reactions Occurring in Connected Cells

Question 27 (3 marks)

Two galvanic cells are connected together, as shown in the diagram.



a. Label the polarities of each electrode in the box provided above. (1 mark)

b. Write the half-equation which occurs at the following electrodes. States are not required.

i. Electrode *A*. (1 mark)

$A \rightarrow A^{2+} + 2e^-$

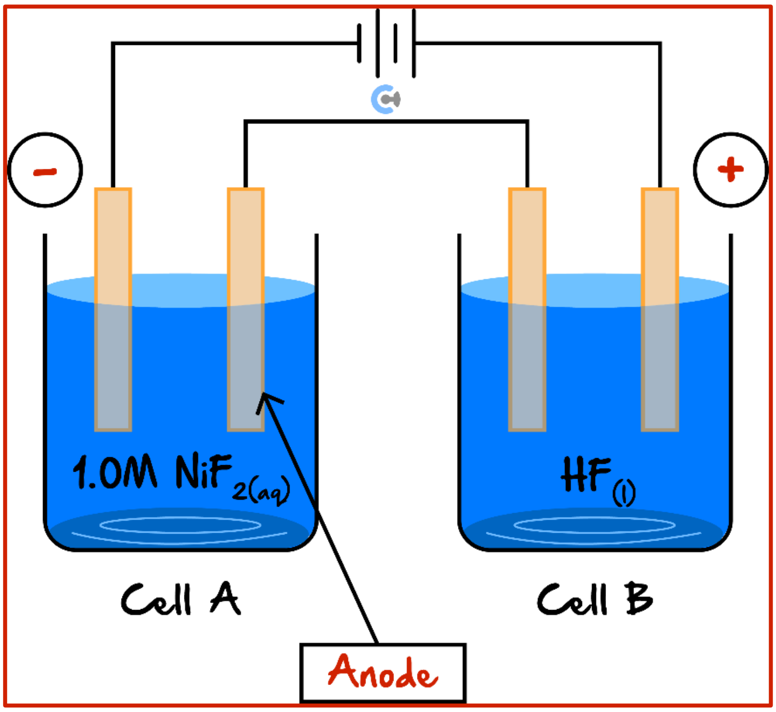
ii. Electrode *D*. (1 mark)

$D^{2+} + 2e^- \rightarrow D$

Space for Personal Notes

Question 28 (6 marks)

The following connected cell is constructed and is connected to a power source.



a. Write the polarities of the two electrodes in the circles provided above. (1 mark)

b. Label the electrode labelled in the diagram as the cathode or the anode. (1 mark)

c. Write the half-equations occurring at the anode of:

i. Cell A. (1 mark)

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

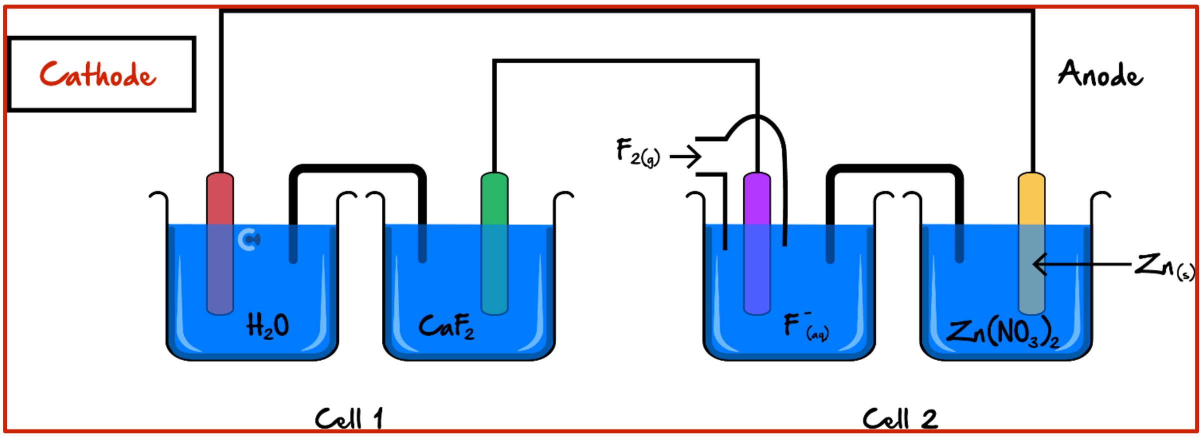
ii. Cell B. (1 mark)

$2F^-(l) \rightarrow F_2(g) + 2e^-$

d. Find the total voltage that the power source must supply for the cell to operate. (2 marks)

Cell 1: Requires $> 1.28\ V$
Cell 2: Requires $> 2.87\ V$
Total EMF required $> 4.15\ V$

Question 29 (7 marks)



a. In the box provided above, label whether the electrode indicated is the anode or cathode. (1 mark)

b.

i. Using relevant half-equations, justify whether cell 2 is a galvanic or electrolytic cell. (2 marks)

Cell 2 has a zinc anode. Therefore, the fluorine half cell is the cathode. The spontaneous reaction between zinc and fluorine will have fluorine be the cathode since it is a stronger oxidant than Zn^{2+}. Therefore, cell 2 must be a galvanic cell as indirect spontaneous reactions are taking place.

ii. Hence or otherwise, determine the overall reaction occurring in cell 1. (1 mark)

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

iii. If the $F^-(aq)/F_2(g)$ half cell is swapped out for a $Cu^{2+}(aq)/Cu(s)$ half cell, explain whether the entire cell will operate. (3 marks)

If swapped with a Cu^{2+}/Cu half-cell, Cell 2 will still operate as a galvanic cell. It will produce a voltage of $+1.10\ V$. However, Cell 1 which is the reduction and oxidation of water, requires a voltage input of $> 2.09\ V$. Therefore, the electrolytic cell will not operate any longer.

Section D: [2.4] - Electroplating (Checkpoints) (37 Marks)

Sub-Section [2.4.1]: Identify The Electroplating Setup (Location of Object) & Find The Electroplating Reactions

Question 30 (1 mark)

When electroplating a metallic key with Cu metal:

A. The key must be connected to the negative terminal of the power supply.

B. The electrolyte can be a solution of $CuSO_4$.

C. The anode can be made from Cu metal.

D. All of the above.

Question 31 (1 mark)

A student wanted to cover an iron key with copper metal. Which of the following experimental setups is **incorrect**?

A. The student connected the key to the negative terminal of a power supply.

B. The student used copper (II) sulphate solution as the electrolyte.

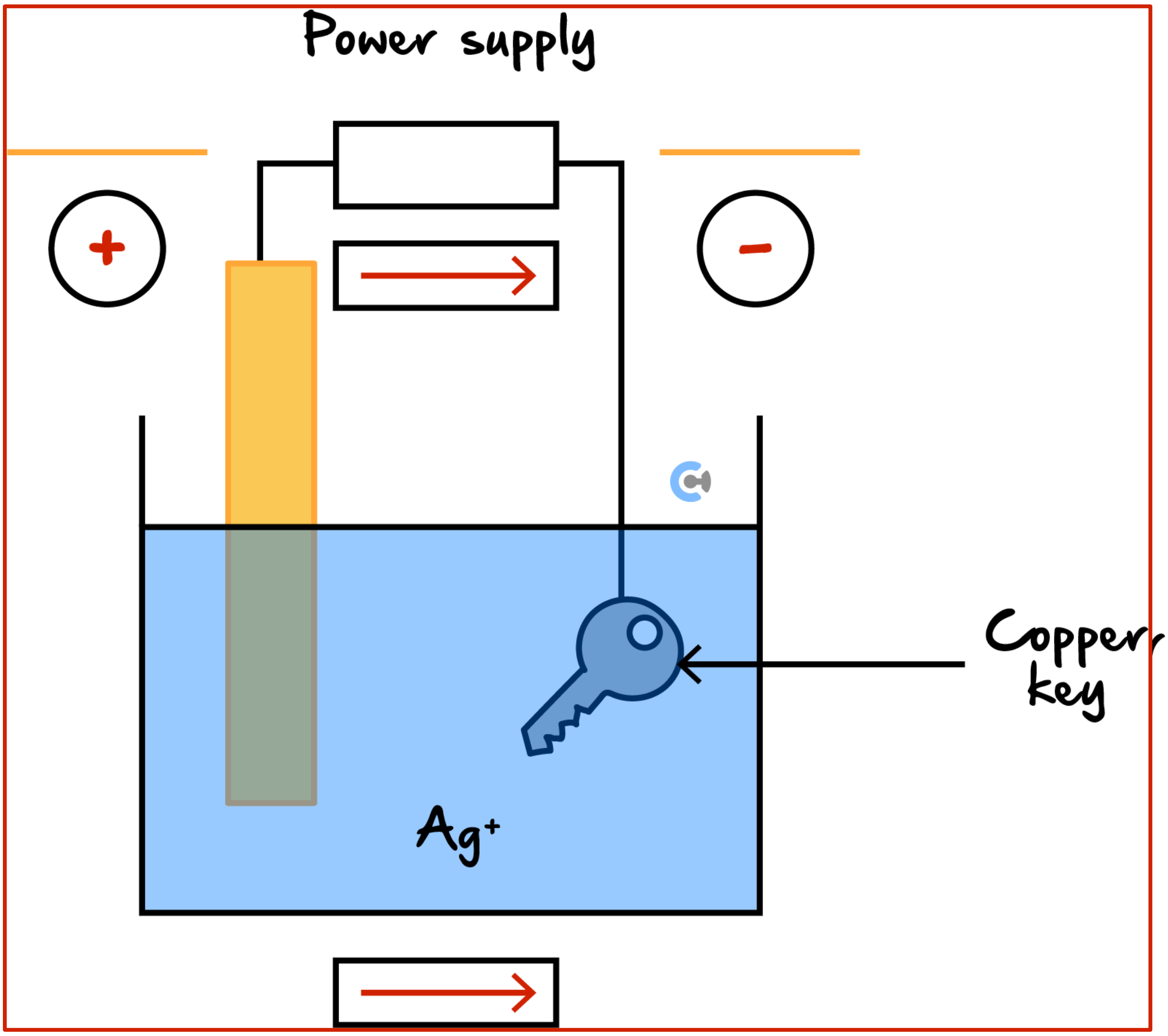
C. The student used an iron rod as the cathode.

D. The student connected an iron rod to the positive terminal of a power supply.

Space for Personal Notes

Question 32 (10 marks)

The following electroplating apparatus has been set up by a student A. 1.2 M solution of $AgNO_3$ was used as an electrolyte.



a. Label the following on the diagram.

i. The cathode and anode on the **orange** lines. (1 mark)

ii. The charges of the cathode and anode in the circles. (1 mark)

iii. An arrow showing the direction of the current in the **purple** box. (1 mark)

iv. An arrow showing the direction of Ag^+ ions in the **blue** box. (1 mark)

b. Determine the oxidation and reduction half-reactions for this cell. (1 mark)

Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$
Reduction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

c. The student runs the electroplating apparatus for 30 minutes and makes several observations.

- The key increases in mass.
- There are bubbles produced at the silver electrode once 15 minutes have passed.
- The silver electrode increases in mass.

Evaluate these observations. If the observation is impossible, explain why. (5 marks)

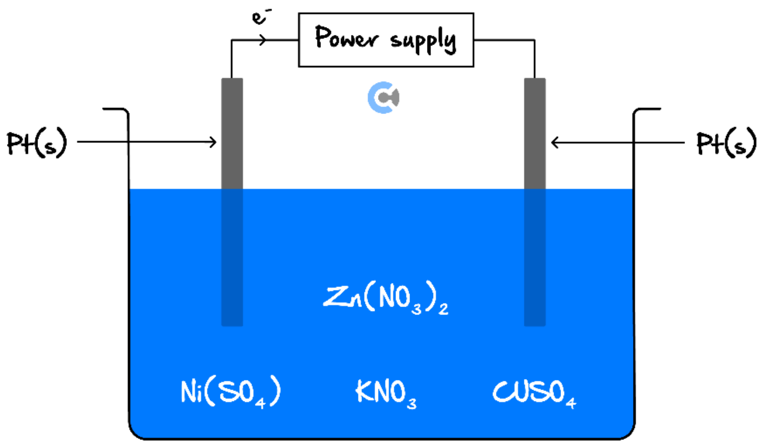
The key does increase in mass (1). Bubbles should not appear (2) at the silver electrode as there is no gas being produced at the anode (3). The silver electrode cannot increase (4) in mass as Ag is being oxidised into the solution (5).

Space for Personal Notes

Sub-Section [2.4.2]: Find Next Order Reactions During Electrolysis

Question 33 (8 marks)

Ziggy sets up an electrolytic cell containing two platinum electrodes with 0.1 M of $CuSO_4(aq)$, $NiSO_4(aq)$, $Zn(NO_3)_2(aq)$ and $KNO_3(aq)$ as shown in the diagram below:



a. State and explain what metal will accumulate on the cathode immediately when the power supply is turned on. (2 marks)

Copper (1). Copper cations are the strongest oxidants present (2).

b. Ziggy notes that after 5 minutes, the solution, which was initially blue, has turned transparent.

i. Explain this observation. (1 mark)

$CuSO_4$ (blue in colour) has all been reduced onto the cathode.

ii. At this point, explain which metal is being electroplated at the cathode. (2 marks)

Nickel (1). Nickel cations are the next strongest oxidants after copper cations (2).

c. After 30 minutes of the cell running, Ziggy notices bubbling occurring at the cathode. His friend, Gain, claims that "the bubbling is due to impurity metal ions in the solution reacting with water." Evaluate Gain's statement and explain the reason for the bubbling. (3 marks)

Gain is incorrect (1). The bubbling is caused by the electrolysis of water (2). This occurs because all of the copper, nickel and zinc cations have reacted and H_2O is the strongest oxidant after zinc cations (3).

Space for Personal Notes

Sub-Section [2.4.3]: Apply Faraday's Laws To Electroplating Calculations

Question 34 (4 marks)

Aanya sets up an electrolytic cell and leaves it running for a period of time.

a. The cell is running for 30 minutes with 4.00 A of current running through the cell. Calculate the electrical charge running through the cell. (2 marks)

$$Q = I\,t$$
$$Q = 4.00 \times 30 \times 60 = 7200\ Coulombs$$

b. Aanya decides to decrease the time that the cell runs to 10 minutes and finds that 900 C of electrical charge runs through the cell. Calculate the current running through the cell. (2 marks)

$$Q = I\,t \Rightarrow I = \frac{Q}{t}$$
$$I = \frac{900}{10 \times 60} = 1.5\ A$$

Space for Personal Notes

Question 35 (5 marks)

a. Daniel passes 2.450 mol electrons through a cell. Calculate the amount of electrical charge that passes through. (2 marks)

$Q = F\,n(e^-)$

$Q = 96500 \times 2.450 = 2.364 \times 10^5\ Coulombs$

b. In another experiment, Daniel runs the cell for 45 minutes, at 4.20 A. Calculate the number of electron moles running through the cell. (3 marks)

$I\,t = F\,n(e^-) \Rightarrow n(e^-) = \frac{I\,t}{F}$

$n(e^-) = \frac{4.20 \times 45 \times 60}{96500} = 0.1175\ moles$

Space for Personal Notes

Question 36 (5 marks)

An electrolytic cell was set up using an unknown, molten metal salt, MBr_3. A current of $1.25\ A$ was passed through the molten compound for 50.0 minutes to deposit $0.675\ g$ of the metal.

a. Write a balanced half-equation for the anode and cathode reactions in this electrolytic cell. (2 marks)

$2Br^-_{(l)} \rightarrow Br_{2\,(l)} + 2e^-$

$M^{3+}_{(l)} + 3e^- \rightarrow M_{(l)}$

b. Calculate the charge passed through the cell. (1 mark)

$Q = It \qquad Q = 50 \times 60 \times 1.25 = 3750\ Coulombs$

c. Calculate the moles of metal deposited. (2 marks)

$It = Fn(e^-) \Rightarrow n(e^-) = \frac{It}{F}$

$n(e^-) = \frac{50\times60\times1.25}{96500} = 0.03886\ moles$

$n(M) = \frac{1}{3} \times n(e^-)$

$n(M) = \frac{1}{3} \times 0.03886 = 0.01295\ moles$

d. Identify the metal deposited. (2 marks)

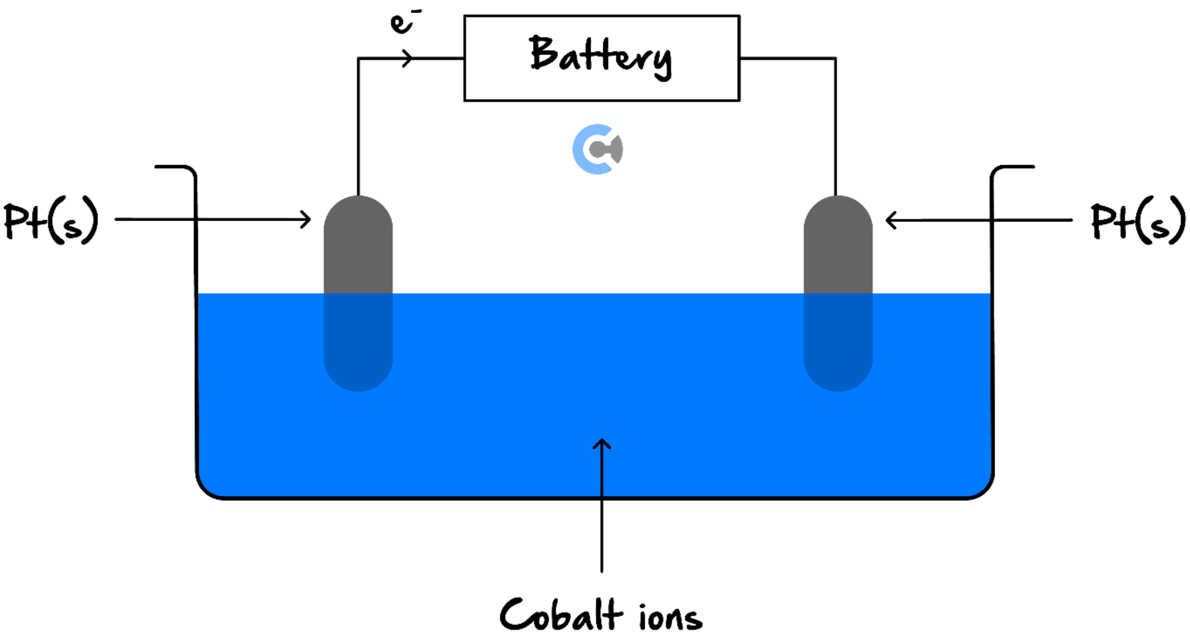
$M_R = \frac{m(M)}{n(M)} = \frac{0.675}{0.01295} = 52.11\ g/mol$

corresponds most closely with Chromium

Space for Personal Notes

Question 37 (3 marks)

Kynan sets up a cell shown below:



Kynan runs the cell for 20.0 minutes with 5.50 A running through the cell and finds that 2.0 g of metal has been deposited on the cathode.

Find the charge of the cobalt cation.

moles of cobalt:

$n(Co) = \frac{2.0}{58.93} = 0.0339\ moles$

moles of electrons:

$I\,t = F\,n(e^-) \Rightarrow n(e^-) = \frac{I\,t}{F}$

$n(e^-) = \frac{20 \times 60 \times 5.50}{96500} = 0.0684\ moles$

ratio:

$\frac{n(Co)}{n(e^-)} = \frac{0.0339}{0.0684} \approx \frac{1}{2}$

$\Rightarrow Co^{a+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$

charge on each side must be equal $\Rightarrow$ *charge of Co cation* $= 2+$

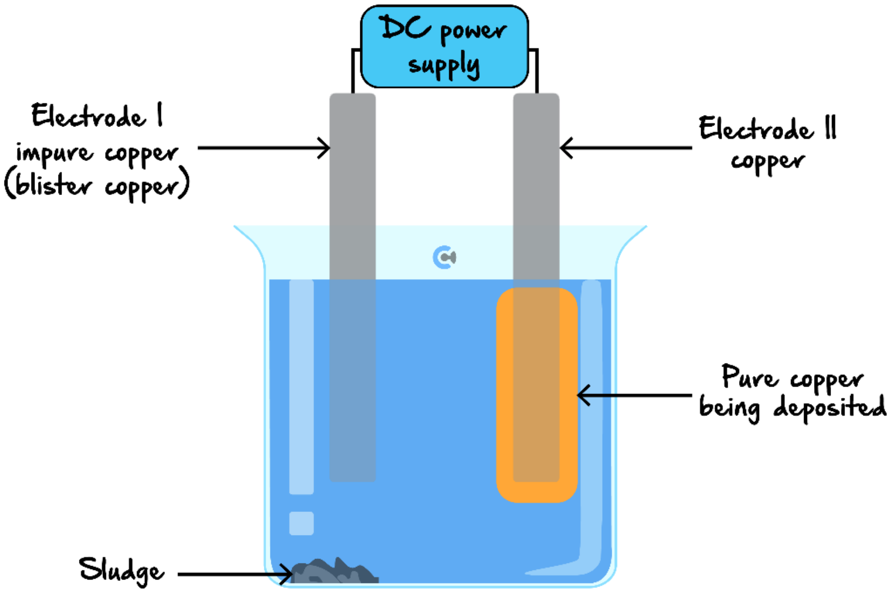
Space for Personal Notes

Section E: [2.1 - 2.4] - Overall (VCAA Qs) (89 Marks)

The following information applies to the two questions that follow.

Inspired from VCAA Chemistry Exam 2015
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2015/2015chem-w.pdf

An electrolytic cell is set up to obtain pure copper from an impure piece of copper called 'blister copper'. The electrolyte solution contains both copper (II) sulphate and sulphuric acid. The blister copper, Electrode I, contains impurities such as zinc, cobalt, silver, nickel and iron. The cell voltage is adjusted so that only copper is deposited on Electrode II. Sludge, which contains some of the solid metal impurities present in the blister copper, forms beneath Electrode I. The other impurities remain in the solution as ions.
The diagram below represents the cell:



Question 38 (1 mark)

The solid metal impurities that are found in the sludge are:

A. Nickel and cobalt.

B. Cobalt, nickel and iron.

C. Nickel and iron.

D. Silver

The relative positions in the electrochemical series of the metals present in the impure Cu was a key reference in this question:
$Au^{+}(aq) + e^{-} \rightleftharpoons Au(s)$ $E^0 = 1.68$ V
$Ag^{+}(aq) + e^{-} \rightleftharpoons Ag(s)$ $E^0 = 0.80$ V
$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s)$ $E^0 = 0.34$ V
$Ni^{2+}(aq) + 2e^{-} \rightleftharpoons Ni(s)$ $E^0 = -0.23$ V
$Co^{2+}(aq) + 2e^{-} \rightleftharpoons Co(s)$ $E^0 = -0.28$ V
$Fe^{2+}(aq) + 2e^{-} \rightleftharpoons Fe(s)$ $E^0 = -0.44$ V
$Zn^{2+}(aq) + 2e^{-} \rightleftharpoons Zn(s)$ $E^0 = -0.76$ V

28	11	30	10	48	0	Since pure copper is deposited at Electrode II it must be the site of reduction, i.e. $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$ At Electrode I, Cu, and any metals that are stronger reductants than Cu (Zn, Co, Ni and Fe) will be oxidised; half-equation $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$ Metals that are weaker reductants than Cu (Ag and Au) will not be oxidised and will collect under Electrode I in the sludge. It is essential that Ag and Au are not oxidised at Electrode I, because their ions would be reduced in preference to $Cu^{2+}(aq)$ and would impact on the purity of the Cu collected at Electrode II. The selection of alternative B ignores the fact that Co, N and Fe cannot be present as solids if Cu(s) has been oxidised because they are stronger reductants than Cu and, under a voltage high enough to oxidise Cu, would also be oxidised

Question 39 (1 mark)

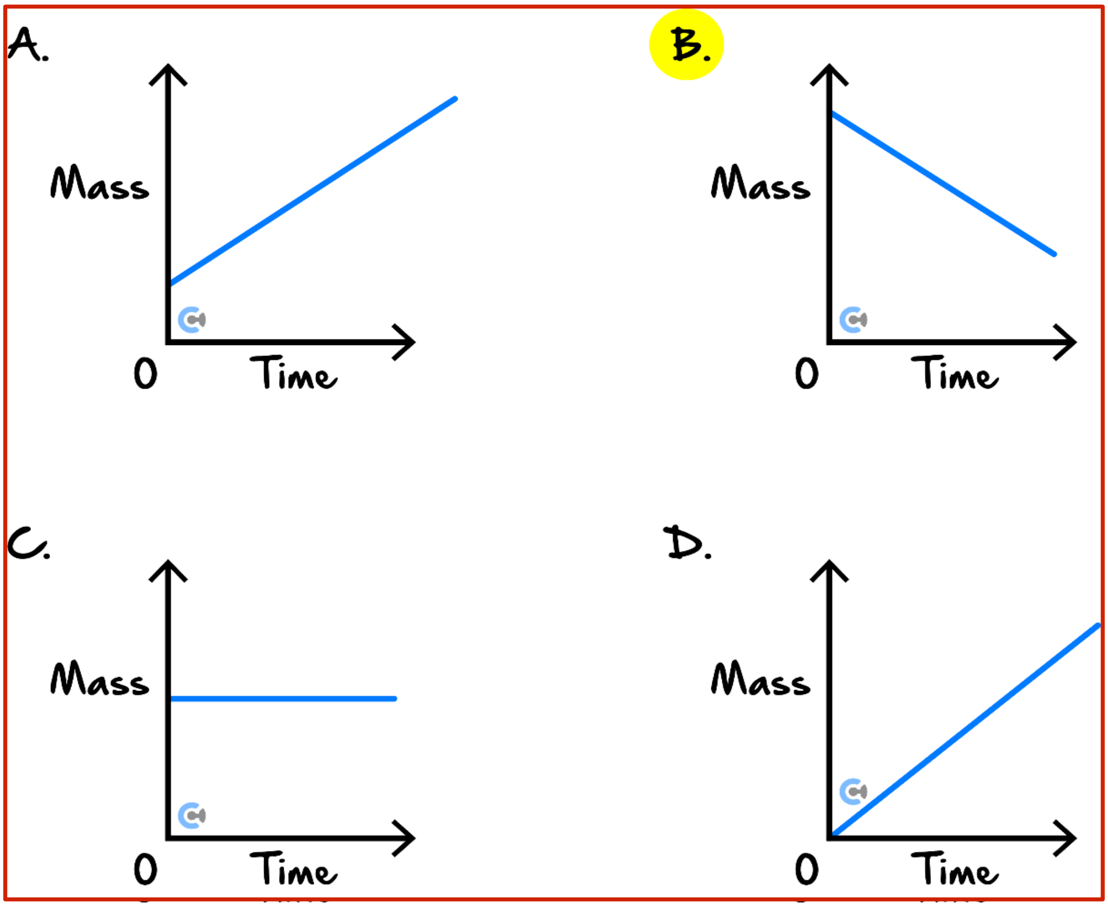
Which of the following correctly shows both the equation for the reaction occurring at the cathode and the polarity of Electrode I?

	Cathode reaction	Polarity of Electrode I
A.	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	Positive
B.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$	
C.	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	
D.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$	

29	31	12	38	18	0	Electrorefining is an electrolytic process, so electrons move from the (+) electrode to the (–) electrode. Since the electrons are moving from Electrode I, the site of oxidation, to Electrode II, the site of reduction, the electrode signs are: Electrode I – positive, Electrode II – negative. The reaction at the cathode, the site of reduction, must be $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ because copper is deposited there. Selection of alternative C suggested that students considered the question to be asking for the sign of the electrode at which the reduction reaction occurred. Electrode I was either ignored or assumed to be the cathode. This emphasises the importance of effective and accurate reading of questions.

Question 40 (1 mark)

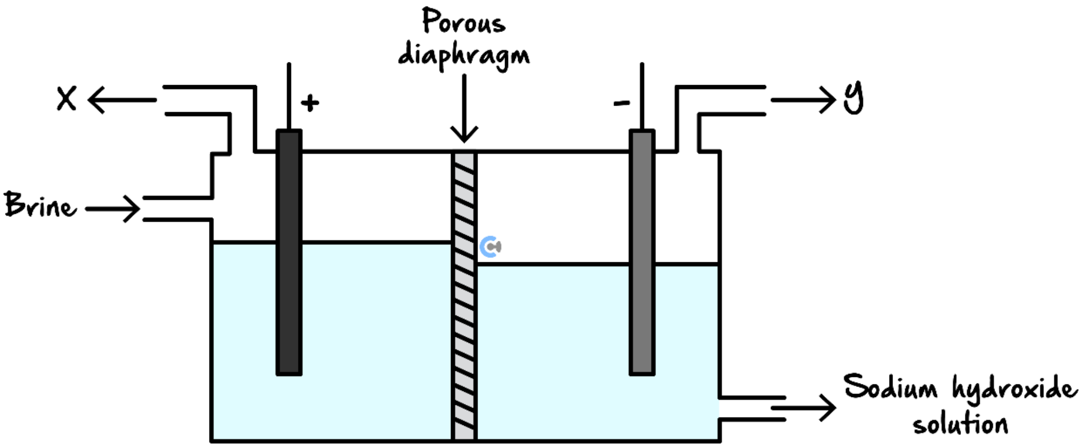
Which one of the following graphs best shows the change in mass of Electrode I over a period of time, starting from the moment the power supply is connected?



30	9	83	4	4	0	Since the metals in the electrode are oxidised, its mass must decrease with time.

The following information applies to the three questions that follow.

The diagram below represents a diaphragm cell used for the commercial production of chlorine gas:



Question 41 (1 mark)

Inspired from VCAA Chemistry Exam 2007
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2007chem2.pdf

The gases labelled *X* and *Y* are:

12	10	8	66	16

	X	*Y*
A.	Chlorine	Oxygen
B.	Oxygen	Chlorine
C.	Chlorine	Hydrogen
D.	Hydrogen	Chlorine

Question 42 (1 mark)

Inspired from VCAA Chemistry Exam 2007
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2007chem2.pdf

One function of the porous diaphragm in the cell is to:

13	4	56	26	14

A. Act as a catalyst to increase the rate of the reaction.

B. Allow movement of ions between the cell compartments.

C. Prevent sodium ions from entering the solution near the anode.

D. Prevent the electrolyte from making contact with the gases produced.

Question 43 (1 mark)

14	14	64	12	10

Inspired from VCAA Chemistry Exam 2007
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2007chem2.pdf

A highly concentrated salt solution, brine, is used as the electrolyte in this cell.
The main reason that a highly concentrated, rather than a dilute, solution is used is to:

A. Allow an electric current to pass through the cell.

B. Produce chlorine gas, in preference to oxygen gas.

C. Allow sodium hydroxide to be separated from the salt by crystallisation.

D. Create non-standard conditions that ensure hydrogen gas production.

Question 44 (1 mark)

Inspired from VCAA Chemistry Exam 2009
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2009chem2-w.pdf

Lithium metal is manufactured by electrolysis of lithium salts.
Which of the following would be the best choice for the electrolyte and the anode in a commercial cell?

	Electrolyte	Anode
A.	LiCl solution	Iron rod
B.	Molten LiCl	Iron rod
C.	LiCl solution	Carbon rod
D.	Molten LiCl	Carbon rod

20	7	12	19	61	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $O_2(g) + 4H^+(aq) + 4e^- \rightarrow \mathbf{2H_2O}(l)$ $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\mathbf{2H_2O}(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $\mathbf{Li^+}(aq) + e^- \rightarrow Li(s)$ The electrochemical series indicates that an aqueous solution cannot be used because H_2O is a stronger oxidant than Li^+ and would be preferentially reduced at the cathode. An Fe rod cannot be used as the anode because Fe is a stronger reductant than Cl^- and would be preferentially oxidised.

Space for Personal Notes

Question 45 (1 mark)

Inspired from VCAA Chemistry Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/NHT/2023chem-nht-w.pdf

In the electrolysis of 6 M sodium chloride solution, $NaCl(aq)$, the amount of charge required to form one mole of $NaOH(aq)$ is:

A. $4.8 \times 10^4\ C$

B. $9.7 \times 10^4\ C$

C. $1.9 \times 10^5 C$

D. $3.9 \times 10^5 C$

17	B	The half-equations occurring during electrolysis of concentrated NaCl(aq) are: Anode (+): $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ Cathode: (–) $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $n(e^-) = n(NaOH) = 1$ mol

Question 46 (1 mark)

Inspired from VCAA Chemistry Exam 2013
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2013/2013chem-w.pdf

The lead-acid battery used in cars consists of secondary galvanic cells.

The following equations relate to the lead-acid battery:

$$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq) \qquad E° = -0.36\ V$$

$$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l) \qquad E° = 1.69\ V$$

When an external power source is used to recharge a flat lead-acid battery:

A. The concentration of sulphuric acid decreases.

B. $PbSO_4$ is both oxidised and reduced.

C. The mass of metallic lead decreases.

D. PbO_2 is oxidised to Pb.

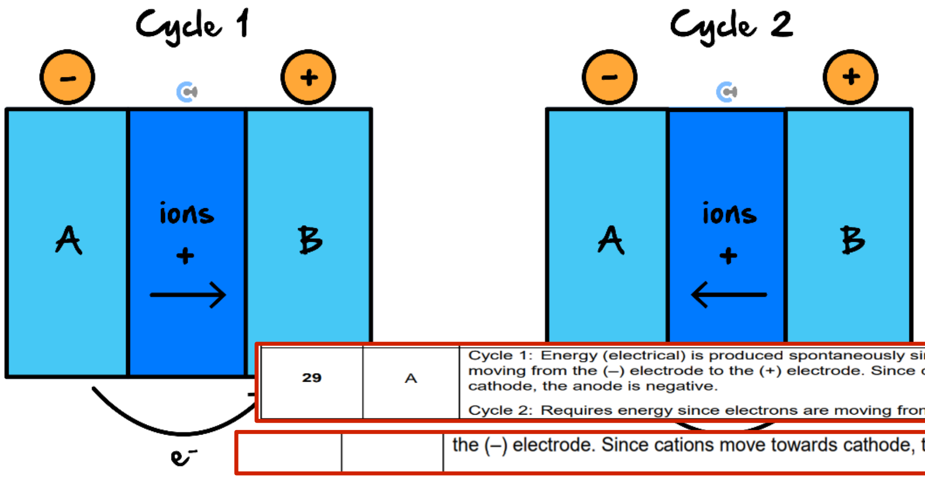
29	14	56	18	11	It was important that students made decisions based on the half-equations in order of decreasing $E°$ values, i.e. $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l)$ 1.69 V $PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$ –0.36 V Half-equations for the reaction occurring during recharging – an electrolytic process in which an external power source is used – are Anode (+) $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$ Cathode (–) $PbSO_4(s) + 2e^- \rightarrow Pb(s) + SO_4^{2-}(aq)$ Only option B was consistent with these half-equations: $PbSO_4$ is oxidised to PbO_2 (oxidation number of Pb increases from +2 to +4) and reduced to Pb (oxidation number of Pb decreases from + 2 to 0).

Space for Personal Notes

Question 47 (1 mark)

Inspired from VCAA Chemistry Exam 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/nht/2018chem-nht-w.pdf

The following diagrams represent the operation of a secondary cell during recharge and discharge, in no particular order. The diagrams of the circuits are not complete.



29	A	Cycle 1: Energy (electrical) is produced spontaneously since electrons are moving from the (–) electrode to the (+) electrode. Since cations move towards cathode, the anode is negative. Cycle 2: Requires energy since electrons are moving from the (+) electrode to the (–) electrode. Since cations move towards cathode, the anode is positive.

Which of the options below correctly describes the cell and its operation?

	Cycle 1	Cycle 2
A.	Energy produced	Anode is positive
B.	Spontaneous reaction	Energy produced
C.	Anode is positive	Energy required
D.	Spontaneous reaction	Cathode is positive

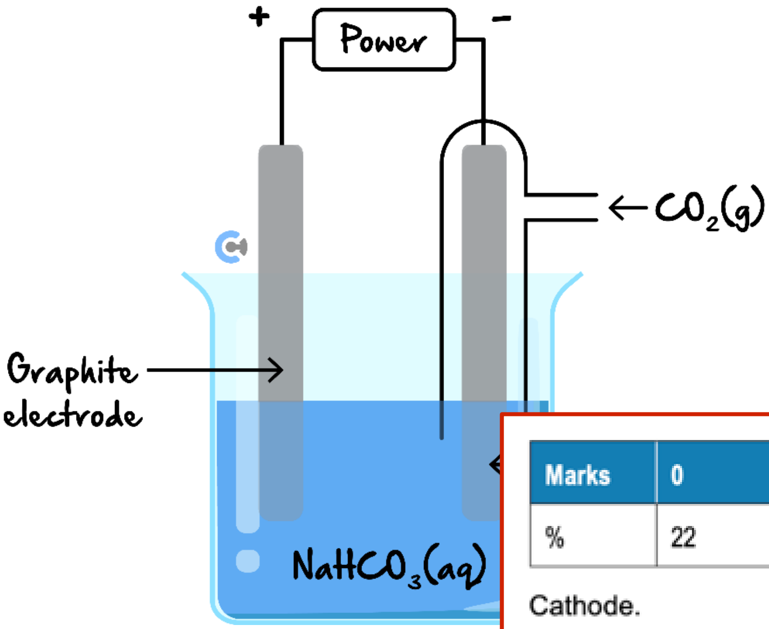
Space for Personal Notes

Question 48 (8 marks)

Inspired from VCAA Chemistry Exam 2020
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2020/2020chem-w.pdf

The electrolysis of carbon dioxide gas, CO_2, in water is one way of making ethanol, C_2H_5OH.

The diagram below shows a CO_2-H_2O electrolysis cell. The electrolyte used in the electrolysis cell is sodium bicarbonate solution, $NaHCO_3(aq)$.



Marks	0	1	Average
%	22	78	0.8

Cathode.

The following half-cell reactions occur in the CO_2-H_2O electrolysis cell.

$$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq) \qquad E° = +0.40\ V$$

$$2CO_2(g) + 9H_2O(I) + 12e^- = C_2H_3OH(l) + 12OH^-(aq)\ ) \qquad E° = -0.33\ V$$

a. Identify the Cu-Zn electrode as either the anode or the cathode in the box provided in the diagram above. (1 mark)

b. Determine the applied voltage required for the electrolysis cell to operate. (1 mark)

Marks	0	1	Average
%	91	9	0.1

>0.73 V

The vast majority of students did not recognise that for electrolysis to occur, the calculated value has to be greater than that shown by the difference in voltages on the SEP table. Use of exactly 0.73 V will not generate any reaction due to the overpotential needed.

c. Write the balanced equation for the overall electrolysis reaction. (1 mark)

Marks	0	1	Average
%	52	48	0.5

$2CO_2 + 3H_2O \rightarrow C_2H_5OH + 3O_2$

Many students wrote the reverse equation, suggesting confusion in establishing the correct direction for half-equations required for this electrolysis reaction.

d. Identify the oxidising agent in the electrolysis reaction. Give your reasoning using oxidation numbers. (2 marks)

Marks	0	1	2	Average
%	50	32	18	0.7

CO_2 *

Either of the following responses was required for the second mark. *

- CO_2 is the oxidising agent since CO_2 is reduced; the oxidation number of C decreases from +4 in CO_2 to an average of -2 in C_2H_5OH.
- CO_2 is the oxidising agent since it oxidises H_2O and the oxidation number of O in H_2O increases from -2 in H_2O to 0 in O_2.

A common mistake was referring to the element carbon as the oxidising agent rather than the required carbon dioxide. Carbon was not a reactant in this equation and could not therefore be accepted as a viable answer. Performance on this question was also impacted by errors in Question 2a. and lack of, or superficial, reference to the role of oxidation numbers. Students need to be aware that the correct convention for oxidation numbers is to always show the sign prior to the digit, i.e. +2 not 2+.

e. A current of 2.70 A is passed through the CO_2-H_2O electrolysis cell. The cell has an efficiency of 58%. Calculate the time taken, in minutes, for this cell to consume 6.05×10^{-3} mol of $CO_2(g)$. (3 marks)

Marks	0	1	2	3	Average
%	35	17	25	22	1.4

$n(CO_2)$ consumed = 6.05×10^{-3}

$n(e^-) = 6 \times n(CO_2) = 6 \times 6.05 \times 10^{-3}$ mol = 0.0363 mol *

$Q = 0.0363 \times 96500 = 3.5 \times 10^3$ C

$t = Q / I = 3.5 \times 10^3 / 2.70 = 1.3 \times 10^3$ s *

Since cell efficiency = 58 %

1297.4 s = 0.58 x time of cell operation

Time = 1297.4 / 0.58 = 2.24×10^3 s = 37 or 37.3 (minutes) *

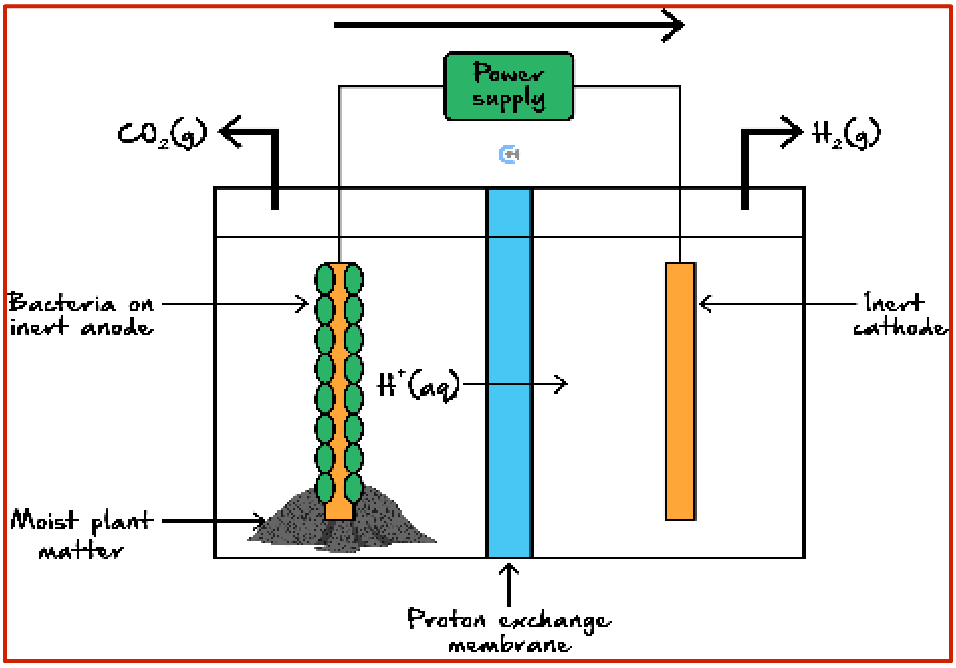
Common mistakes were failure to convert $n(CO_2)$ to $n(e^-)$ and, more significantly, failure to understand how to apply the percentage efficiency. Many students also lacked the ability to reliably manipulate mathematical relationships.

Space for Personal Notes

Question 49 (4 marks)

Inspired from VCAA Chemistry Exam 2012
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2012/2012chem2-w.pdf

Hydrogen gas is an energy source. Researchers are investigating the production of hydrogen gas in a microbial electrolysis cell.
The cell is made up of an anode half-cell and a cathode half-cell. A proton exchange separates the half-cells membrane, as shown in the diagram below:



A number of reactions take place in the cell, resulting in the production of hydrogen. These reactions are summarised below.

Anode half-cell:

- The anode half-cell contains moist plant matter and electrochemically active bacteria that live on an inert anode.
- The gaseous mixture that is present in the half-cell does not contain oxygen.
- The moist plant matter ferments to produce ethanoic acid (CH_3COOH). Bacteria on the anode consume the ethanoic acid and release hydrogen ions, electrons and carbon dioxide gas. A small voltage is then applied to reduce the H^+ ions.

Cathode half-cell:

- The cathode half-cell contains an inert cathode.
- The gaseous mixture that is present in the half-cell does not contain oxygen.
- The released hydrogen ions and electrons react to form hydrogen gas, as shown in the equation below:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

a. Ethanoic acid is converted to carbon dioxide gas and H^+ ions at the anode.

Write an equation for this reaction. (1 mark)

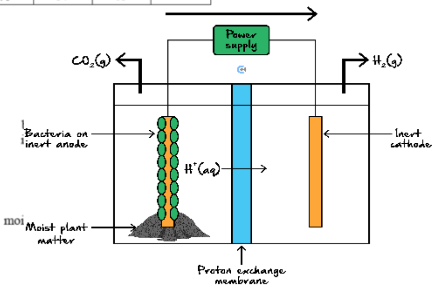
Marks	0	1	Average
%	67	33	0.3

$CH_3COOH(aq) + 2H_2O(l) \rightarrow 2CO_2(g) + 8H^+(aq) + 8e^-$

There was significant diversity in the responses to this question, mostly associated with balancing. The initial step in developing the half-equation, balancing the C atoms, was commonly missed.

b. In the diagram above, use one arrow to indicate the (
voltage is supplied to the cell by the power supply. (

Marks	0	1	Average
%	37	63	0.7



Students were expected to show that the electrons move from left to right (from the site of oxidation to the site of reduction) and in the external circuit. The most common errors showed electrons moving through the proton exchange membrane or in the wrong direction.

c. Hydrogen gas is not produced at the cathode if oxyg
Write a balanced half-equation to show the product
present in the half-cell. (1 mark)

Marks	0	1	Average
%	52	48	0.5

$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

d. Describe one difference between an electrolysis cell and a traditional fuel cell. (1 mark)

Question 7d.

Marks	0	1	Average
%	61	39	0.4

Acceptable responses included

- in a fuel cell, chemical energy is converted into electrical energy; in an electrolysis cell, electrical energy is converted to chemical energy
- in a fuel cell, the reaction produces electrical energy; in an electrolysis cell, electrical energy drives the reaction
- a fuel cell is not connected to an external power supply
- cathode/anode polarity is different in the two cells.

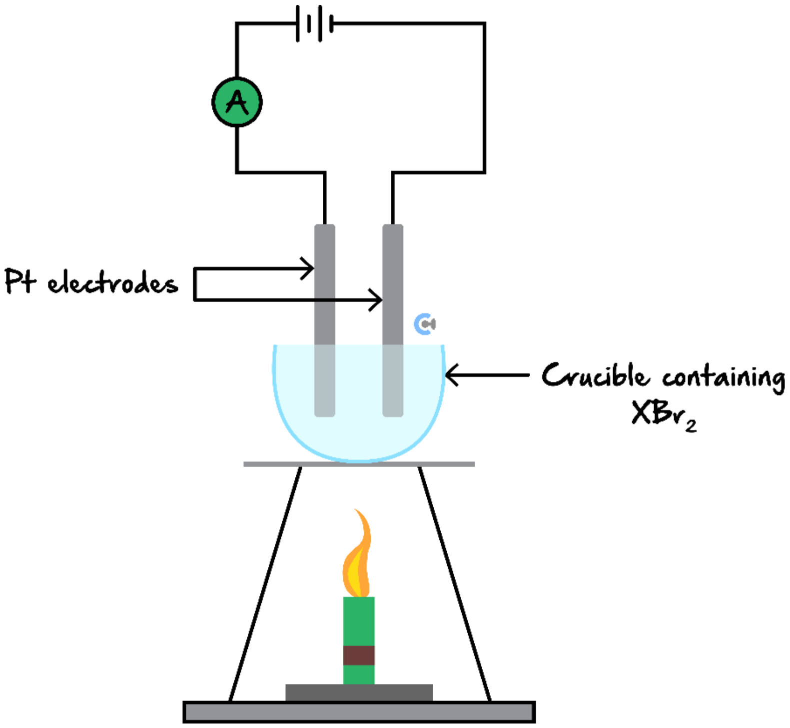
Most students clearly wanted to use the well-learned distinction between a fuel cell and a galvanic cell and referred to fuel cells having a continuous supply of reactants. This was not accepted as a distinction because some **electrolysis** cells, such as the production of NaOH by electrolysis of NaCl(aq), also have a continuous supply of reactants.

Space for Pers

Question 50 (6 marks)

Inspired from VCAA Chemistry Exam 2012
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2012/2012chem2-w.pdf

A teacher demonstrated the process of electrolysis of a molten salt using an unknown metal salt, XBr_2. The apparatus was set up as shown below:



After the demonstration, the students were provided with the following information:

- A current of 1.50 amperes was applied for 30.0 minutes.
- 2.90 *g* of metal *X* was produced.

a. Write a balanced half-equation for the anode reaction in this electrolytic cell. (1 mark)

Marks	0	1	Average
%	91	9	0.1

$2Br^-(l) \rightarrow Br_2(l + g) + 2e^-$

The majority of students incorrectly included (aq) as the state for Br^- in their equations. The 'electrolysis of a molten salt' should have indicated that an aqueous solution was not present. Students should be aware that molten ionic compounds are liquids, and that oxidation occurs at the anode.

Equilibrium arrows in the half-equations on the electrochemical series reflect the fact that, in an electrochemical cell, whether a particular half-reaction proceeds to the left or the right depends on the other oxidant/reductant pair present in the cell. Once a decision has been made about the direction in which the half-reaction proceeds, equilibrium arrows are not included in the half-equation.

b.

i. Determine the amount, in *mol*, of metal *X* that was deposited on the cathode. (3 marks)

Marks	0	1	2	3	Average
%	19	5	22	54	2.1

$Q = It$

$= 1.50 \times 30.0 \times 60$

$= 2.70 \times 10^3$ C

$n(e^-) = Q \div F$

$= 2.7 \times 10^3 \div 96500$

$= 2.80 \times 10^{-2}$ mol

Reduction of metal $X^{2+} + 2e^- \rightarrow X$

$n(X) = n(e^-) \div 2$

$= 2.8 \times 10^{-2} \div 2$

$= 1.4 \times 10^{-2}$ mol

The most common error on this part of the question was the omission or incorrect use of the relationship between the n(X) and the $n(e^-)$.

ii. Identify metal *X*. (2 marks)

Marks	0	1	2	Average
%	31	5	64	1.3

$M(X) = m \div n$

$= 2.90 \div 1.4 \times 10^{-2}$

$= 207.2$ g mol^{-1}

Lead (Pb)

Since the answer to this question was dependent on the answer to part i., it may seem strange that significantly more students obtained full marks for part ii. This shows how 'consequential' marks are applied. If the element identified in part ii. was consistent with the n(X) calculated in part i., and forms ions with a +2 charge, full marks were awarded for part ii. Therefore, if

- n(X) = 0.0280 mol in part i. due to not dividing n(e), then M(X) = 103.6 g mol^{-1} → Rhodium (Rh)
- n(X) = 0.0560 mol in part ii. due to multiplying $n(e^-)$ by 2, then M(X) = 51.8 g mol^{-1} → Chromium (Cr).

This consequential mark was not awarded if a clearly monovalent metal such as K was given, since this was not consistent with the data provided as the metal ion had a charge of +2.

Question 51 (8 marks)

Inspired from VCAA Chemistry Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/NHT/2023chem-nht-w.pdf

Copper metal, Cu, is used to conduct electricity through a printed circuit board (PCB). One use of PCBs is in computers. Electrolysis can be used to deposit Cu on PCBs.

a. State the structural feature that is present in an electrolytic cell but absent in a galvanic cell. (1 mark)

Power supply / battery

b. State an advantage of using Cu over platinum, Pt, as the anode in an electrolysis cell used to deposit Cu on PCBs. (1 mark)

Provides a source of Cu^{2+} to the solution/electrolyte, which allows the cell to operate longer.

An alternative response: allows the $c(Cu^{2+})$ in the electrolyte to remain constant.

c. In a particular electrolysis cell, a number of PCBs have copper deposited on them. The electrolyte used in the cell is $Cu(NO_3)_2(aq)$. The electrolysis cell is operating at 4.0 amps.

Assuming 100% cell efficiency, for how many seconds does the cell need to operate to deposit 5.08 g of copper on the PCBs? (4 marks)

n(Cu) = 5.08 g /63.5 g mol^{-1}

= 0.0800 mol

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

$n(e^-)$ = 0.0800 × 2

= 0.160 mol

$Q = n(e^-) \times F$ = 0.160 × 96500

= 15440 C

$t = Q / I$ = 15440 / 4.0

= **3.9 × 10^3** sec

d. Another electrolysis cell is set up, at standard conditions, with two Pt electrodes and 1M sodium hydroxide, NaOH as the electrolyte.

Write the equations for the two half-reactions that occur. (2 marks)

One mark was awarded for each fully correct half-equation.

$4OH^-(aq) \rightarrow O_2(g) + 2H_2O(l) + 4e^-$

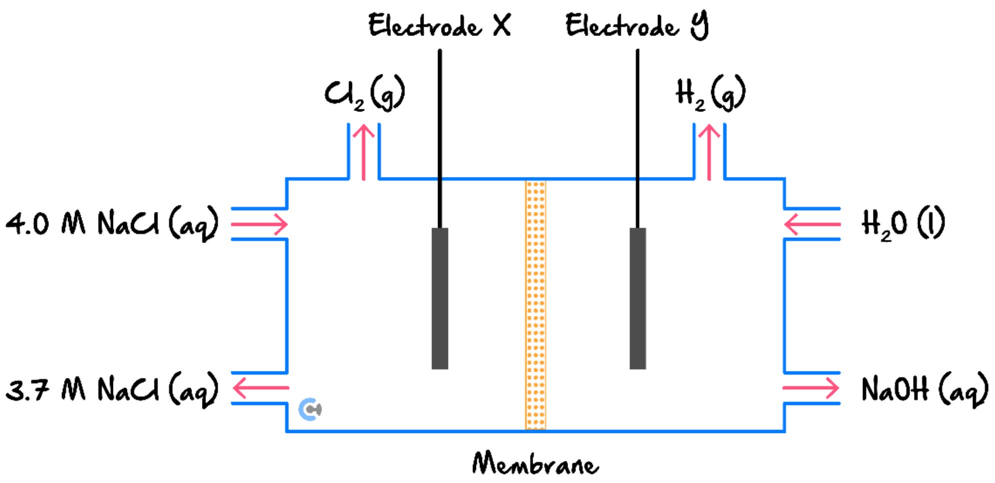
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

Space for Personal Notes

Question 52 (8 marks)

Inspired from VCAA Chemistry Exam 2022
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2022/NHT/2022ChemistryNHT-w.pdf

Chlorine gas, Cl_2, can be produced from the electrolysis of 4.0 *M* sodium chloride, NaCl(aq). The diagram below shows a simplified model of the electrolysis cell for this process:



a.

i. Which electrode- *X* or *Y*- is the anode when the electrolysis cell is operating? (1 mark)

Electrode X

ii. Name or write the formula for the ion that passes through the membrane between the half-cells to allow the electrolysis cell to operate. (1 mark)

Sodium ion / Na^+

b.

i. Write the half-equation for the production of $Cl_2(g)$. (1 mark)

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

ii. With reference to the electrochemical series, explain why $Cl_2(g)$ is produced instead of $O_2(g)$. (2 marks)

The electrochemical series predicts that (under standard conditions / 1 M NaCl) H_2O should be preferentially oxidised to produce O_2 and H+(aq).

High concentration of NaCl means that the Cl− will be oxidised to Cl_2 instead of the H_2O being oxidised.

iii. Calculate the mass of $Cl_2(g)$ that would be produced from 1.80×10^6 coulombs of charge. Assume that $Cl_2(g)$ is the only product at Electrode X. (3 marks)

$n(e^-)$ $= 1.80 \times 10^6 / 96500$ *

$= 18.7$ (mol)

$n(Cl_2)$ $= n(e^-) / 2$ *

$= 18.7 / 2$

= **9.32** (mol)

$m(Cl_2)$ $= n \times M$

$= 9.32 \times 71.0$

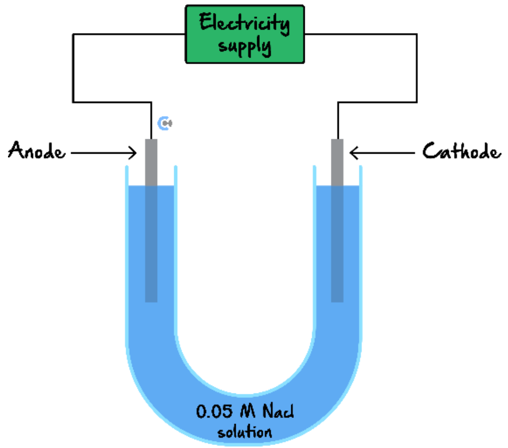
= **662 g** *

Space for Personal Notes

Question 53 (9 marks)

Inspired from VCAA Chemistry Exam 2019
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/NHT/2019chem-nht-w.pdf

A student electrolysed a 0.05 *M* sodium chloride, NaCl, solution using graphite electrodes, as shown in the setup below:



Several drops of phenol red were added to the solution next to each electrode.

The following observations were made as the reactions proceeded:

Polarity of electrode	Observation	Colour of phenol red
Positive	Bubbles formed at the electrode	Yellow
Negative	Bubbles formed at the electrode	Red

a. What colour was observed at the cathode as the electrolysis proceeded? (1 mark)

red

b. Use the electrochemical series to predict the gas expected to be formed at each electrode. (2 marks)

- At the anode: oxygen/O_2
- At the cathode: hydrogen/H_2

Appropriate half-equations were also acceptable.

c. Molten NaCI can be electrolysed commercially. The melting point of NaCl is 801°C.

i. Write the overall equation for this electrolysis. (1 mark)

$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

ii. Give two reasons why it would be difficult to carry out this electrolysis in a school. (2 marks)

Any two of:

- difficulty maintaining the required temperature or difficulty/danger in keeping temperature at or above 801 °C
- lack of correct equipment to safely collect toxic Cl_2 or difficulty in safely removing Cl_2
- difficulty keeping Na produced away from air
- difficulty of keeping Na and Cl_2 from coming into contact (violent spontaneous reaction)
- issues in sourcing suitable electrodes
- issues with maintaining electric circuit at the high temperature.

Other reasonable responses were accepted.

Students needed to give two specific responses that indicated a clear understanding of the practical issues.

d. The student constructs the summary table below to compare the processes that take place in galvanic cells and electrolytic cells.

Complete the table by writing 'true' or 'false' in each space provided. (3 marks)

Process	Galvanic cells	Electrolytic cells

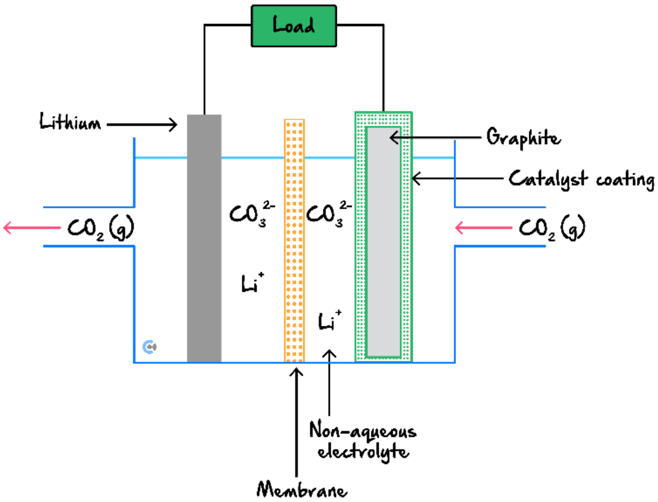
Process	Galvanic cells	Electrolytic cells
Oxidation occurs at the cathode.	false	false
Chemical energy is converted to electrical energy.	true	false
Spontaneous reactions take place.	true	false

Space for Personal Notes

Question 54 (7 marks)

Inspired from VCAA Chemistry Exam 2020
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2020/2020chem-w.pdf

Research scientists are developing a rechargeable lithium-carbon dioxide, Li-CO_2, battery. The rechargeable Li-CO_2 battery is made of lithium metal, carbon in the form of graphite (coated with a catalyst) and a non-aqueous electrolyte that absorbs CO_2.
A diagram of the rechargeable Li-CO_2 cell is shown below. One Li-CO_2 cell generates 4.5 V.



a. When the Li-CO_2 cell generates electricity, the two half-cell reactions are:

$$4Li^+ + 3CO_2 + 4e^- \rightarrow 2Li_2CO_3 + C$$

$$Li \rightarrow Li^+ + e^-$$

Write the equation for the overall recharge reaction. (1 mark)

Marks	0	1	Average
%	59	41	0.4

$2Li_2CO_3 + C \rightarrow 4Li + 3CO_2$

Here too there was an incorrect understanding of which direction each half-cell reaction needed to be written in during the recharge process. The most common mistake was that the overall reaction was written in reverse.

b. During discharge, lithium carbonate, Li_2CO_3, deposits break away from the electrode.

Describe how this might affect the performance of the battery. (2 marks)

Marks	0	1	2	Average
%	29	39	32	1.0

Reduces battery life / Performance reduced / Limits the extent of recharging / number of recharges / reduced ability to hold full charge.

As lithium carbonate breaks away from the cathode, this reduces the amount of Li_2CO_3 available for recharging.

OR

Products of electrolysis need to stay in contact with electrodes for effective recharge.

One mark was awarded for describing the effect on the performance of the battery. This was frequently not done.

One mark was awarded for describing how this affected the performance of the battery.

c. Explain why it is unsafe to use an aqueous electrolyte in the design of the Li-CO_2 battery. Include appropriate equations in your answer. (3 marks)

Marks	0	1	2	3	Average
%	58	17	18	8	0.7

Students used two approaches to respond to this question.

Galvanic cell approach

Li(s) is a very strong reducing agent (reductant) and will reduce water to produce $H_2(g)$.

$Li(s) \rightarrow Li^+(aq) + e^-$ / $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

OR

$2Li(s) + 2H_2O(l) \rightarrow H_2(g) + 2Li^+(aq) + 2OH^-(aq)$

$H_2(g)$ is explosive / flammable / builds up pressure in battery

OR

The reaction is highly exothermic and may cause battery to catch fire.

Electrolytic cell approach

Water will be reduced in preference to Li^+ ions producing $H_{2(g)}$.

At the cathode(-) $2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

At the anode(+) $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

Overall $2H_2O_{(l)} \rightarrow O_{2(g)} + 2H_{2(g)}$

Build-up of H_2/O_2 leading to a potential explosion due to pressure or a potential explosion due to spontaneous combustion.

One mark was awarded for identifying the products that could cause a safety concern.

One mark was awarded for the equations (either both half-equations or the full equation) to back up the formation of this product.

One mark was awarded for the reason this chemical poses a safety risk.

Students needed to show these three key sections in their response. Not many included the last point.

d. Could the Li-CO_2 battery be used to reduce the amount of $CO_2(g)$ in the atmosphere? Give your reasoning. (1 mark)

Marks	0	1	Average
%	51	49	0.5

No. The $CO_2(g)$ absorbed during discharge will be released during recharge.

A reason had to be given by the student in order to obtain the mark.

Question 55 (7 marks)

Inspired from VCAA Chemistry Exam 2016
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2016/2016chem-amd-w.pdf

The lithium-ion battery is a secondary cell that is now widely used in portable electronic devices. In these batteries, lithium ions, Li^+, move through a special non-aqueous electrolyte between the two electrodes. The batteries are housed in sealed containers to ensure that no moisture can enter them.

Both electrodes are made up of materials that allow the lithium ions to move into and out of their structures. The anode consists of LiC_6, where lithium is embedded in the graphite structure. Lithium cobalt oxide, $LiCOO_2$, is commonly used as the material in the cathode. The reaction at the cathode is quite complex. When the cell discharges, Li^+ ions move out of the anode and enter the cathode.

During discharge, the half-cell reaction at the anode is:

$$LiC_6 \rightarrow Li^+ + e^- + C_6$$

a. During discharge, what is the polarity of the graphite electrode? (1 mark)

Marks	0	1	Average
%	33	67	0.7

Negative (–)

Many students did not correctly apply two basic principles of electrochemical cells:

- electrons move spontaneously from the negative electrode to the positive electrode during discharge
- electrons move from the site of oxidation (anode) to the site of reduction (cathode).

b. Write the half-equation for the reaction that occurs at the cathode of a lithium-ion battery when it is recharged. (1 mark)

Marks	0	1	Average
%	66	34	0.4

$Li^{+} + e^{-} + C_6 \rightarrow LiC_6$

Responses to this question suggested that most students did not refer to the principle that the electrode that acts as the anode (site of oxidation) during discharge acts as the cathode (site of reduction) during recharging. This means that the oxidation half-equation ($LiC_6 \rightarrow Li^{+} + e^{-} + C_6$) during discharge is reversed to become the reduction half-equation ($Li^{+} + e^{-} + C_6 \rightarrow LiC_6$) during recharge.

Several students who applied this principle undermined the response by including (aq) as the state of Li^{+}, despite the inclusion of 'non-aqueous electrolyte' in the stem of the question.

c. In a lithium-ion battery, lithium metal must not be in contact with water.

Explain why and justify your answer with the use of appropriate equations. (3 marks)

Marks	0	1	2	3	Average
%	45	29	18	8	0.9

Li is a strong reductant that reacts readily with water.

- $Li(s) \rightarrow Li^{+}(aq) + e^{-}/2H_2O(l) + 2e^{-} \rightarrow H_2(g) + 2OH^{-}(aq)$
- $2Li(s) + 2H_2O(l) \rightarrow H_2(g) + 2LiOH(aq)$

Hydrogen gas is explosive/Heat is generated/LiOH is a strong base/current does not flow.

One mark each was awarded for:

- appropriate reference to lithium's ready reaction with water
- both half-equations correctly balanced or a correctly balanced overall equation
- a significant consequence of lithium in contact with water

Responses to this question indicated that many students struggled to apply their understandings in this context. Many students focused on the battery itself rather than stating what happens when lithium is in contact with water and why is this an issue. Consequently, many students tried to frame responses around recharging problems that might arise if water was present in the battery.

Many students seemed to consider Li and Li+ to be interchangeable. Comments such as 'water is a stronger oxidant than lithium' or 'Li+ will be oxidised by water' were common.

It was evident that many students did not read the question carefully and their answers were based on recharging problems that might arise if water were present. Other students thought that water would reduce in preference to Li+ during discharge.

d. Identify **one** design feature of the lithium-ion battery that enables it to be recharged (1 mark)

Marks	0	1	Average
%	86	14	0.2

The movement of lithium ions into and out of the electrodes enables the reactions at the electrodes to be reversed.

Most students simply stated the general characteristic of secondary cells that products of discharge remain at the electrodes. The question was about the lithium-ion battery and the lithium ions are integral to the operation of cell in both discharging and recharging, so their movement into and out of the structure of the electrodes was the required design feature.

e. What is **one** advantage of using a secondary cell compared to using a fuel cell? (1 mark)

Marks	0	1	Average
%	25	75	0.8

Possible responses included:

- A secondary cell is more convenient for on–off usage since it does not need a continuous external supply of reactants.
- A secondary cell can be recharged (electrically), whereas a fuel cell continuously needs fresh reactants.
- Secondary cells are usually cheaper than fuel cells.
- Secondary cells are more suitable for most of today's electronic devices.
- Secondary cells are storage devices.

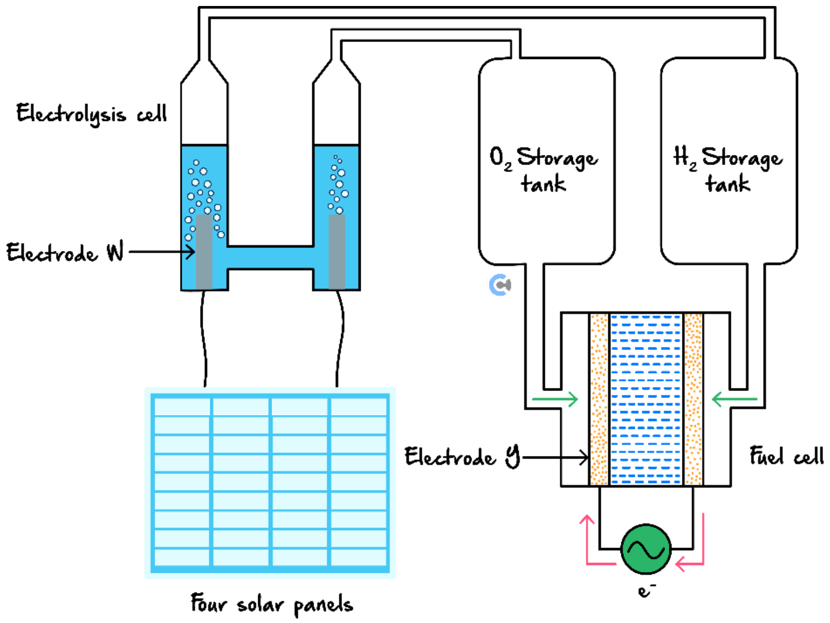
Space for Personal Notes

Question 56 (11 marks)

Inspired from VCAA Chemistry Exam 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/2018chem-w.pdf

An energy company investigates the feasibility of supplying energy while reducing greenhouse gas emissions. Solar panels collect energy from the sun during daylight hours, and this energy is used to electrolyse water, H_2O to produce oxygen gas, O_2, and hydrogen gas, H_2. These gases are stored separately and then used in a fuel cell to produce energy when required.

The diagram below shows a simplified representation of the setup used:



a.

i. State the polarity of Electrode W in the electrolysis cell. (1 mark)

Marks	0	1	Average
%	41	59	0.6

Negative/(–)

Many students did not make the link between H_2 in the fuel cell where it is oxidised at the negative electrode and the 'like connects to like' connection to the electrolysis cell. Also, in the electrolysis of water, reduction (to produce H_2) occurs at the negative electrode.

ii. The fuel cell operates in an alkaline environment.

Write the half-equation for the reaction that takes place at Electrode Y. (1 mark)

Marks	0	1	Average
%	66	34	0.4

$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$

The frequency of the half-equation for an acidic environment suggested that many students copied the first O_2 reduction half-equation on the electrochemical series rather than considering the significance of the alkaline electrolyte.

Even though equilibrium arrows were shown on the electrochemical series – to indicate that, depending on the combination of oxidant and reductant, both directions are possible – they are not appropriate for cell-specific half-equations.

b. Each of the four solar panels produces an average current of 5.20 A and operates over an eight-hour period. The electrical energy generated is used by the electrolysis cell to produce O_2 and H_2.

i. Calculate the amount, in moles, of H_2 produced by the electrolysis cell. (3 marks)

Marks	0	1	2	3	Average
%	19	21	49	12	1.6

$Q = It = 5.20 \times (8 \times 60 \times 60) \times 4$

$= 5.99 \times 10^5$ C

$n(e^-) = Q / F = 5.99 \times 10^5/96\ 500$

$= 6.21$ mol

$H_2(g) + 2OH^-(aq) \rightarrow 2H_2O(l) + 2e^-$

$n(H_2) = 6.21/2$

$= 3.10$ mol

Marks were awarded for the accurate calculations of Q, $n(e^-)$ and $n(H_2^-)$.

Common errors associated with solutions to this question included:

- ignoring the fact that there were 4 solar cells
- not converting time into seconds in the calculation of Q
- assuming, incorrectly, that the $n(H_2)$ was the same as the $n(e^-)$.

The eight-hour period indicates that this quantity was not a factor in determination of the appropriate number of significant figures. The current of 5.20 A was the determining factor in the answer having three significant figures.

ii. Determine the pressure this amount of H_2 gas would exert at SLC in a 10.0 L H_2 tank. (1 mark)

Marks	0	1	Average
%	45	55	0.6

$p = nRT / V$

$= 3.10 \times 8.31 \times 298/10.0$

$= 769$ kPa

The most logical approach to this question was to calculate the pressure exerted at 25 °C in a 10.0 L container by 3.10 L of H_2 using the general gas equation.

c. The fuel cell produces 3553 kJ when 20 mol of H_2 is consumed. Another possible energy source is a generator using petrodiesel as a fuel. The generator operates with an efficiency of 35%. A particular petrodiesel containing a range of hydrocarbons has been found to have a heat content of 45 $kJ\ g^{-1}$. The formula for this petrodiesel can be represented by $C_{12}H_{24}$ ($M = 168\ g\ mol^{-1}$).

i. Calculate the mass of petrodiesel required to produce 3553 kJ. (2 marks)

Marks	0	1	2	Average
%	38	34	28	0.9

m(petrodiesel) at 100 per cent efficiency

= 3553 kJ/45 kJ g^{-1}

= 79 g

At 35% efficiency:

m(petrodiesel) = 79 g/0.35

= 2.3×10^2 g

Alternatively:

Energy = 3553/0.35

= 1.02×10^4 kJ

m(petrodiesel) = $1.02 \times 10^4/45.0$

= 2.3×10^2 g

Students needed to accurately divide the energy by 0.35 (or 35/100) and 45. While most students were able to execute one of these requirements, a significant number did not realise that kJ/kJ g^{-1} gives the mass in grams.

ii. Calculate the mass of CO_2(g) released when 3553 kJ of energy is produced from petrodiesel. (2 marks)

Marks	0	1	2	Average
%	38	14	48	1.1

$n(C_{12}H_{24}) = 2.3 \times 10^2/168$

= 1.34 mol

$n(CO_2)$ = 12 × $n(C_{12}H_{24})$ = 12 × 1.34

= 16.1 mol

$m(CO_2)$ = 16.1 × 44.0

= 709 g (640 g)

Students were required to calculate the $n(CO_2)$ from the m(petrodiesel) determined in part ci. A significant number did not recognise that 1 mol $C_{12}H_{24}$ will produce 12 mol CO_2.

iii. How would the mass of CO_2 produced from the combustion of this petrodiesel compared with the mass of CO_2 produced by the fuel cell? (1 mark)

Marks	0	1	Average
%	34	66	0.7

The amount of CO_2 from the combustion of petrodiesel is greater since the H_2-O_2 fuel cell does not produce CO_2.

Space for Personal Notes

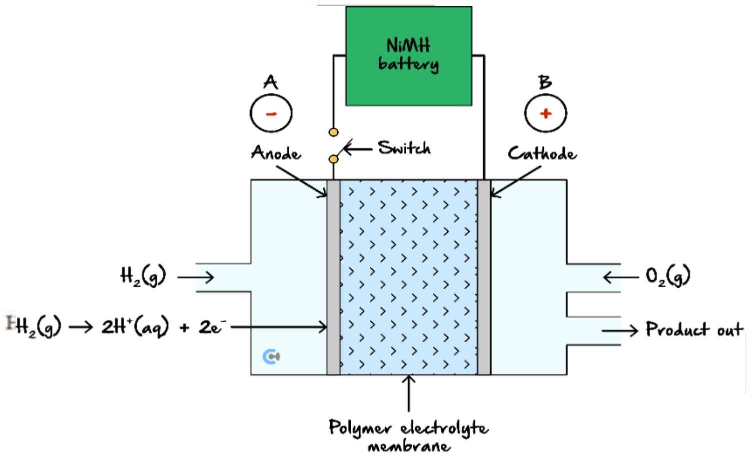
Question 57 (11 marks)

Inspired from VCAA Chemistry Exam 2016
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2016/2016chem-amd-w.pdf

A car manufacturer is planning to sell hybrid cars powered by a type of hydrogen fuel cell connected to a nickel metal hydride, NiMH, battery.

A representation of the hydrogen fuel cell is given below:

Marks	0	1	2	Average
%	12	32	56	1.5



The overall cell reaction is:

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

a.

i. On the diagram above, indicate the polarity of t identify the product of the reaction in the box *C*

ii. Write an equation for the reaction that occurs at

Marks	0	1	Average
%	46	54	0.6

Cathode reaction: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(g)$

Students should be aware that in a fuel cell the fuel is oxidised at the anode and oxygen is reduced at the cathode. The half-equation for the reduction of O_2 in an acidic fuel cell can be deduced using the electrochemical series. However, when writing half-equations for specific cells, equilibrium arrows are not included. The electrochemical series shows the half-equations as reversible because the direction is determined by the other reacting species.

Whilst the half-equation provided in the electrochemical series in the data book shows $H_2O(aq)$, that was not appropriate in this question because the cell reaction given showed $H_2O(g)$.

Cathode reaction ____

iii. Identify one advantage and one disadvantage of using this fuel cell instead of a petrol engine to power the car. (2 marks)

Advantage____

Disadvantage ____

Marks	0	1	2	Average
%	30	42	28	1

Advantages

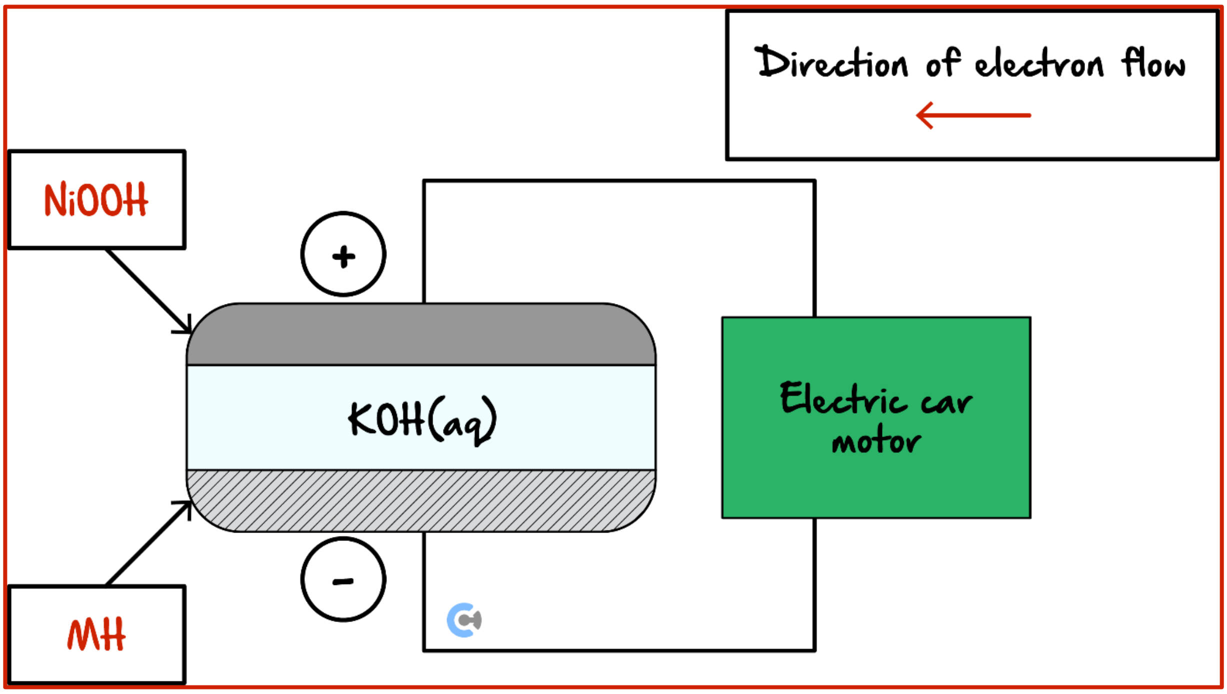
- highly efficient compared to petrol engine
- no CO_2 or other pollutants such as CO or unburnt hydrocarbons produced

Disadvantages

- relative cost
- difficulties in accessing and storing hydrogen
- risk of leaking hydrogen reacting explosively

Many students gave responses such as 'no greenhouse gases' as an advantage, and 'needs a constant supply of fuel' as a disadvantage. Such responses suggested students had a lack of awareness of the greenhouse gas nature of water vapour and a limited understanding of the role of the petrol tank in a common vehicle.

b. The storage battery to be used in the hybrid cars is comprised of a series of nickel metal-hydride cells, NiMH cells. MH represents a metal hydride alloy that is used as one electrode. The other electrode contains nickel oxide hydroxide, NiOOH. The electrolyte is aqueous KOH.



The simplified equation for the reaction at the anode while **recharging** is:

$$Ni(OH)_2(s) + OH^-(aq) \rightarrow NiOOH(s) + H_2O(l) + e^-$$

The simplified equation for the reaction at the cathode while **recharging** is:

$$M(s) + H_2O(l) + e^- \rightarrow MH(s) + OH^-(aq)$$

i. What is the overall equation for the discharging reaction? (1 mark)

Marks	0	1	Average
%	39	61	0.6

$MH(s) + NiOOH(s) \rightarrow M(s) + Ni(OH)_2(s)$

The equation for the discharging reaction was the reverse of the sum of the half-equations provided for the recharging reaction.

ii. In the boxes on the diagram above, indicate which is the MH electrode and which is the NiOOH electrode. (1 mark)

Marks	0	1	Average
%	42	58	0.6

iii. In the bold box provided above the cell diagram, use an arrow, → or ← to indicate the direction of the electron flow as the cell is discharging. (1 mark)

iv. The battery discharged for 60 minutes, producing a current of 1.15 A.

What mass, in grams, of NiOOH would be used during this period? (3 marks)

Marks	0	1	2	3	Average
%	22	6	13	59	2.1

$Q = It$
$= 1.15 \times 60 \times 60)$
$= 4.14 \times 10^3$ C
$n(e^-) = Q/F$
$= 4.14 \times 10^3/96\ 500$
$= 0.0429$ mol
n(NiOOH) = 0.0429 mol
m(NiOOH) $= n$(NiOOH) x M(NiOOH)
$= 0.0429 \times 91.7$
$= 3.9$ g

One mark each was awarded for:

- accurate calculation of charge
- accurate calculation of n(NiOOH)
- accurate calculation of m(NiOOH).

The most common error was using the molar mass of Ni rather than the required molar mass of NiOOH.

Space for Personal Notes

Website: contoureducation.com.au | **Phone:** 1800 888 300 | **Email:** hello@contoureducation.com.au

VCE Chemistry ¾

Free 1-on-1 Support

Be Sure to Make the Most of These (Free) Services!

- Experienced Contour tutors ($45+$ raw scores, $99+$ ATARs).
- For fully enrolled Contour students with up-to-date fees.
- After school weekdays and all-day weekends.

1-on-1 Video Consults	Text-Based Support
• Book via bit.ly/contour-chemistry-consult-2025 (or QR code below). • One active booking at a time (must attend before booking the next).	• Message +61 440 137 304 with questions. • Save the contact as “Contour Chemistry”.

Booking Link for Consults

bit.ly/contour-chemistry-consult-2025

Number for Text-Based Support

+61 440 137 304