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VCE Chemistry ¾
AOS 1 Revision II [1.12]
Contour Check Solutions

Contour Checklist

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Section A: [1.6] - Introduction to Redox (Checkpoints) (43 Marks)

Sub-Section [1.6.1]: Apply Oxidation Numbers to Find Oxidant & Reductant

Question 1 (4 marks)

State the oxidation number for the element specified in the molecule/ion provided.

a. Oxidation number of chromium in CrO_4^{2-}. (1 mark)

$\boxed{Cr = +6, O = -2}$

b. Oxidation number of sulphur in SO_3^{2-}. (1 mark)

$\boxed{S = +4, O = -2}$

c. Oxidation number of phosphorus in $H_2PO_4^{-}$. (1 mark)

$\boxed{P = +5, H = +1, O = -2}$

d. Oxidation number of nitrogen in NO_3^{-}. (1 mark)

$\boxed{N = +5, O = -2}$

Space for Personal Notes

Question 2 (4 marks)

Natalie is investigating the following chemical reaction:

$$ZnO(s) + C(s) \rightarrow Zn(s) + CO_2(g)$$

a. Find the oxidation numbers for all atoms in the following molecules:

i. C. (1 mark)

$\boxed{C = 0}$

ii. CO_2. (1 mark)

$\boxed{C = +4, O = -2}$

b. Hence, determine and justify whether C (carbon) is a reductant or an oxidant. (2 marks)

C is a reductant (1). This is because the oxidation number of Carbon in C to CO_2 increases from 0 to +4, meaning C has undergone oxidation (2).

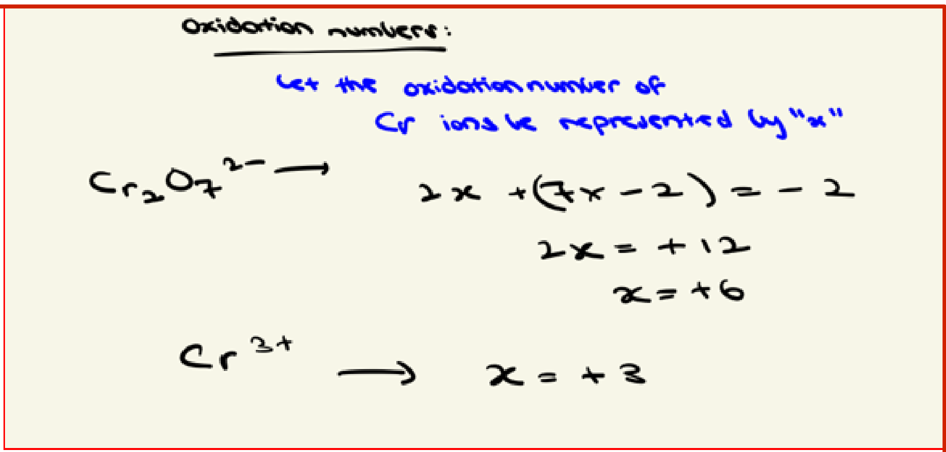
Space for Personal Notes

Question 3 (4 marks)

Kanta is observing the following reaction occurring at school:

$$2Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6I^-(aq) \rightarrow 2Cr^{3+}(aq) + 3I_2(g) + 7H_2O(l)$$

His friend explains that the reducing agent in this reaction is $Cr_2O_7^{2-}$. Evaluate Kanta's friend's statement, using calculations as justification.



Oxidation numbers:

Let the oxidation number of Cr ions be represented by "x"

$Cr_2O_7^{2-} \rightarrow$ $2x + (7\times-2) = -2$

$2x = +12$

$x = +6$

$Cr^{3+} \rightarrow$ $x = +3$

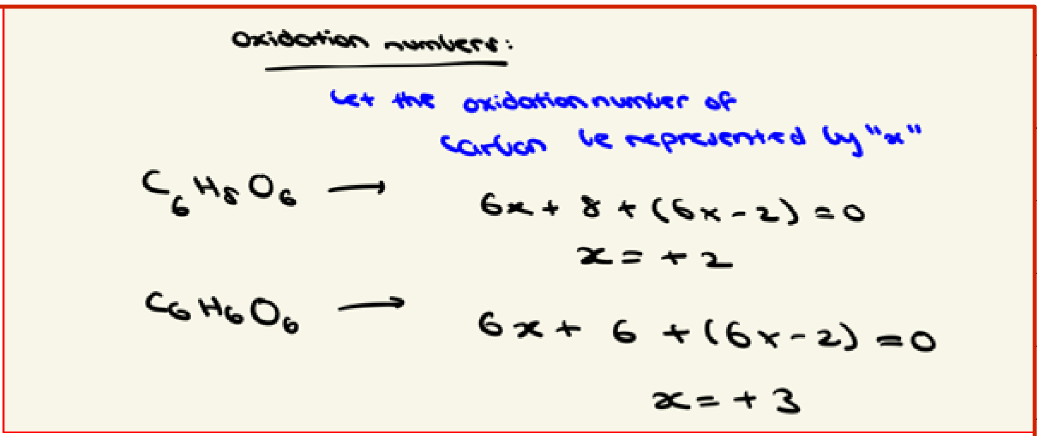
Micah's friend is incorrect (1). The chromium ions from $Cr_2O_7^{2-}$ to Cr^{3+} have an oxidation number decrease from +6 to +3 (2). This means that the $Cr_2O_7^{2-}$ has reduced and is the oxidising agent (3).

Question 4

In an acidic solution, ascorbic acid ($C_6H_8O_6$) reacts with dichromate ions ($Cr_2O_7^{2-}$), resulting in the formation of chromium (III) ions (Cr^{3+}) and dehydroascorbic acid ($C_6H_6O_6$).

$$C_6H_8O_6(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow C_6H_6O_6(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$$

A friend claims that $\mathbf{C_6H_8O_6}$ is the reducing agent in this reaction. Evaluate this claim and justify your response with the relevant calculations.



Oxidation numbers:

Let the oxidation number of carbon be represented by "x"

$C_6H_8O_6 \rightarrow$ $6x + 8 + (6\times-2) = 0$

$x = +2$

$C_6H_6O_6 \rightarrow$ $6x + 6 + (6\times-2) = 0$

$x = +3$

The friend is correct (1). The carbon oxidation number in $C_6H_8O_6$ to $C_6H_6O_6$ has an oxidation number increase from +2 to +3 (2). This means that the $C_6H_8O_6$ has oxidised and is the reducing agent (3).

Sub-Section [1.6.2]: Apply KOHES to Write Balanced Half-Equations in Acidic & Basic Conditions

Question 5 (2 marks)

Perchlorate ions (ClO_4^-) turn into chlorine gas (Cl_2) in a laboratory.

a. Write the half-equation in acidic conditions. (1 mark)

$2ClO_4^{2-}(aq) + 16H^+ + 14e^- \rightarrow Cl_2(g) + 8H_2O(l)$

b. Write the half-equation in alkaline conditions. (1 mark)

$2ClO_4^{2-}(aq) + 8H_2O(l) + 14e^- \rightarrow Cl_2(g) + 16OH^-(aq)$

Question 6 (4 marks)

Complete the balanced half-equation for each of the following, and state whether it is a reduction or oxidation reaction.

a. Copper (II) ions turning into copper solid. (1 mark)

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (Reduction, Cu goes from +2 to 0.)

b. Silver oxide (Ag_2O) turning into silver solid. (1 mark)

$Ag_2O(s) + 2H^+(aq) + 2e^- \rightarrow 2Ag(s) + H_2O(l)$

Reduction

c. Butanol (C_4H_9OH) turning into butanoic acid ($C_4H_8O_2$). (1 mark)

$C_4H_9OH\ (l) + H_2O(l) \rightarrow C_4H_8O_2\ (l) + 4H^+(aq) + 4e^-$

Oxidation

d. Nitrous oxide (N_2O) turning into nitrogen gas (N_2). (1 mark)

$N_2O\ (g) + 2H^+(aq) + 2e^- \rightarrow N_2(g) + H_2O(l)$

Reduction

Question 7 (4 marks)

In acidic conditions, potassium permanganate ($KMnO_4$) can react and turn into manganese ions (Mn^{2+}).

a. Write a balanced half-equation for this process. (1 mark)

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

b. State whether this is an oxidation or reduction reaction and justify why. (2 marks)

- This is a **reduction reaction** because the oxidation number of manganese decreases.
- In **MnO_4^-**, the oxidation number of manganese is **+7**, but in **Mn^{2+}**, it is **+2**. Since the oxidation number of manganese decreases, it means that **MnO_4^-** is gaining electrons (reduction).

c. Hence or otherwise, is Mn^{2+} an oxidant or reductant? (1 mark)

Mn^{2+} is a **reductant (reducing agent)** because it has a low oxidation state of **+2** and can easily donate electrons to other species (undergoing oxidation itself).

Question 8 (4 marks)

In acidic conditions, potassium dichromate ($K_2Cr_2O_7$) reacts and turns into chromium (III) ions (Cr^{3+}).

a. Write a balanced half-equation for the reduction of potassium dichromate to chromium (III) ions in acidic conditions. (1 mark)

$$\mathrm{Cr_2O_7^{2-}}(aq) + 14\mathrm{H}^+(aq) + 6e^- \rightarrow 2\mathrm{Cr}^{3+}(aq) + 7\mathrm{H_2O}(l)$$

b. State whether this is an oxidation or reduction reaction and justify why. (2 marks)

The oxidation number of chromium in **dichromate** ($\mathrm{Cr_2O_7^{2-}}$) is **+6**, while in **chromium(III)** ions ($\mathrm{Cr^{3+}}$), the oxidation number is **+3**.

Since the oxidation number of chromium decreases, it indicates that the **chromium has gained electrons** in the reaction (reduction).

This is further evidenced by the presence of **electrons** on the left-hand side of the half-equation, indicating a gain of electrons.

c. Hence or otherwise, is Cr^{3+} an oxidant or reductant? (1 mark)

Chromium(III) ions (Cr^{3+}) have an oxidation state of **+3**, which is relatively low. Therefore, they can easily **lose electrons** to be oxidized to a higher oxidation state (e.g., $\mathrm{Cr^{6+}}$ or $\mathrm{Cr_2O_7^{2-}}$).

A species with a low oxidation state (like Cr^{3+}) is a good **reducing agent** because it can donate electrons to other species and undergo **oxidation** itself.

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Sub-Section [1.6.3]: Apply KOHES to Write Balanced Half-Equations and Overall Equations in Acidic & Basic Conditions

Question 9 (4 marks)

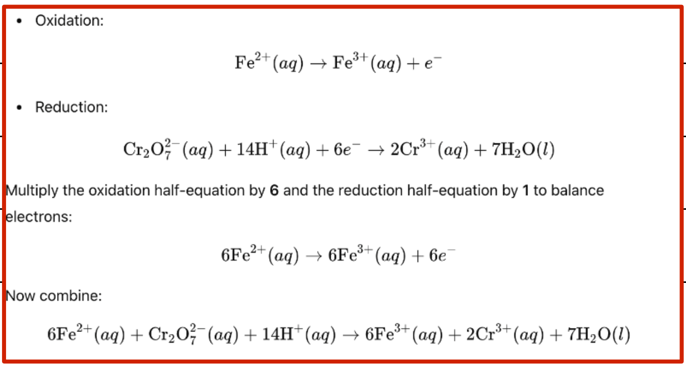
Express the overall equation using the half-equations provided.

a. **Oxidation half-equation:** (2 marks)

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$$

Reduction half-equation:

$$Cr_2O_7{}^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$



- Oxidation:

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$$

- Reduction:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

Multiply the oxidation half-equation by **6** and the reduction half-equation by **1** to balance electrons:

$$6Fe^{2+}(aq) \rightarrow 6Fe^{3+}(aq) + 6e^-$$

Now combine:

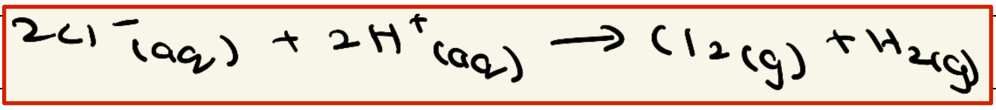
$$6Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 6Fe^{3+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$$

b. **Oxidation half-equation:** (2 marks)

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

Reduction half-equation:

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



$$2Cl^-(aq) + 2H^+(aq) \rightarrow Cl_2(g) + H_2(g)$$

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Question 10 (4 marks)

Zinc (II) Zn^{2+} can be formed from zinc solid, Zn, when reacted with dichromate ions ($Cr_2O_7^{2-}$). Cr^{3+} ions are formed in the process.

Write the balanced equation for:

a. The oxidation reaction. (1 mark)

$$\boxed{\mathrm{Zn}(s) \rightarrow \mathrm{Zn}^{2+}(aq) + 2e^-}$$

b. The reduction reaction. (1 mark)

$$\boxed{\mathrm{Cr_2O_7}^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2\mathrm{Cr}^{3+}(aq) + 7H_2O(l)}$$

c. The overall reaction. (2 marks)

Oxidation reaction (multiplied by 3):

$$3\mathrm{Zn}(s) \rightarrow 3\mathrm{Zn}^{2+}(aq) + 6e^-$$

Reduction reaction (no change):

$$\mathrm{Cr_2O_7}^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2\mathrm{Cr}^{3+}(aq) + 7H_2O(l)$$

Now, combine the two reactions:

$$3\mathrm{Zn}(s) + \mathrm{Cr_2O_7}^{2-}(aq) + 14H^+(aq) \rightarrow 3\mathrm{Zn}^{2+}(aq) + 2\mathrm{Cr}^{3+}(aq) + 7H_2O(l)$$

Space for Personal Notes

Question 11 (4 marks)

Sodium (Na) reacts with water in an alkaline environment to form sodium hydroxide and hydrogen gas.

Write the balanced equation for:

a. The oxidation reaction. (1 mark)

$$\boxed{Na(s) \rightarrow Na^{+}(aq) + e^{-}}$$

b. The reduction reaction. (1 mark)

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

c. The overall reaction. (2 marks)

- **Oxidation reaction:**

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

- **Reduction reaction:**

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

Now, combine them:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

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Question 12 (5 marks)

In the paper industry, bleaching is a crucial process to eliminate colour from pulp, ensuring the production of high-quality paper. Chlorine or chlorine compounds are commonly used in redox reactions to oxidise and remove impurities, enhancing the paper's brightness and quality.

a. Chlorine gas is used in a reaction with water. This purifies pulp and produces Hydrochloric acid(HCl) and hypochlorous acid (HOCl).

Write a balanced equation for the:

i. Oxidation reaction. (1 mark)

$Cl_2(g) + 2H_2O(l) \rightarrow 2HOCl(aq) + 2H^+(aq) + 2e^-$

ii. Reduction reaction. (1 mark)

$Cl_2(g) + H^+(aq) + e^- \rightarrow HCl(aq)$

iii. Overall reaction. (1 mark)

verall reaction. (1 mark)

$\overset{0}{Cl_2}(g) + H_2O(l) \rightarrow \overset{+1\ -2\ +1}{HOCl}(aq) + \overset{+1\ -1}{HCl}(aq)$

b. Hydrochloric acid (HCl) is also often used as an alternative to using chlorine gas. HCl can undergo the following reaction:

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

State and explain whether the above reaction is a redox reaction. (2 marks)

No, there is no change of oxidation numbers across any of the species. This reaction rather is an acid-base reaction.

Space for Personal Notes

Section B: [1.7] - Spontaneous Redox Reactions (Checkpoints) (34 Marks)

Sub-Section [1.7.1]: Apply the ECS to Predict Spontaneous Reactions

Question 13 (3 marks)

Write the (i) oxidation and (ii) half equations for each of the following spontaneous reactions.

a. Sarah reacts magnesium metal (Mg) with hydrochloric acid (HCl) in an aqueous solution.

i. Oxidation half equation. (0.5 marks)

i) **Oxidation Half Equation:**

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

ii. Reduction half equation. (0.

ii) **Reduction Half Equation:**

$2H^+ + 2e^- \rightarrow H_2$

b. Liam reacts aluminium metal (Al) with iron (III) chloride ($FeCl_3$) in aqueous solution.

i. Oxidation half equation. (0.5 marks)

i) **Oxidation Half Equation:**

$Al \rightarrow Al^{3+} + 3e^-$

ii. Reduction half equation. (0

ii) **Reduction Half Equation:**

$Fe^{3+} + 3e^- \rightarrow Fe$

c. Emma reacts lead (II) nitrate ($Pb(NO_3)_2$) with solid zinc (Zn).

i. Oxidation half equation. (

i) **Oxidation Half Equation:**

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

ii. Reduction half equation. (

ii) **Reduction Half Equation:**

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

Question 14 (3 marks)

Write the (i) oxidation and (ii) half equations for each of the following spontaneous reactions.

a. Rebecca reacts tin (II) chloride ($SnCl_2$) with iron (III) nitrate ($Fe(NO_3)_3$) in an aqueous solution.

i. Oxidation half equation. (0.5 marks)

i) **Oxidation Half Equation:**

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$$

ii. Reductio

ii) **Reduction Half Equation:**

$$2Fe^{3+}(aq) + 2e^- \rightarrow 2Fe^{2+}(aq)$$

b. Kevin reacts hydrogen peroxide (H_2O_2) with permanganate ions (MnO_4^-) in an aqueous solution.

i. Oxidation half equation. (0.5 marks)

i) **Oxidation Half Equation:**

$$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$$

ii. Red

ii) **Reduction Half Equation:**

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

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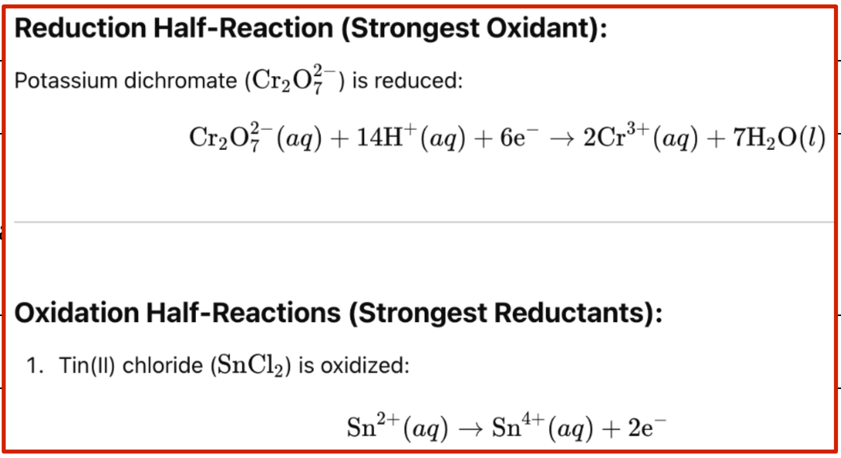
Question 15 (4 marks)

Write the (i) oxidation and (ii) half equations for each of the following spontaneous reactions.

a. Sophia adds tin (II) chloride ($SnCl_2$) to a solution containing potassium dichromate ($K_2Cr_2O_7$) and hydrogen peroxide (H_2O_2) in aqueous conditions.

i. Oxidation half equation. (1 mark)



Reduction Half-Reaction (Strongest Oxidant):

Potassium dichromate ($Cr_2O_7^{2-}$) is reduced:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

Oxidation Half-Reactions (Strongest Reductants):

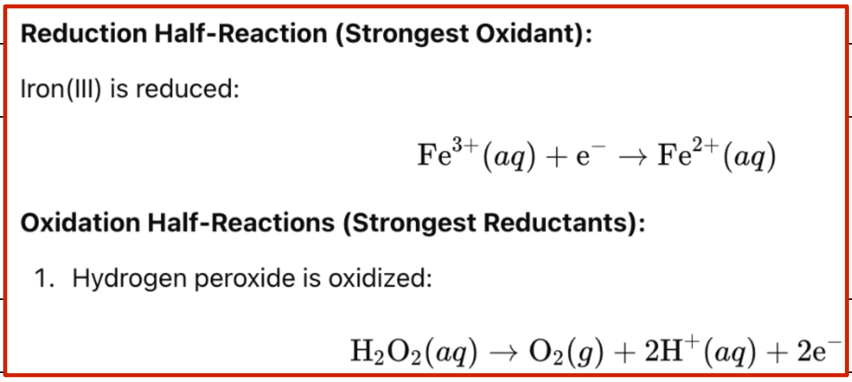
1. Tin(II) chloride ($SnCl_2$) is oxidized:

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$$

ii. Reduction h

b. Angela adds iron (III) chloride ($FeCl_3$), hydrogen peroxide (H_2O_2), and sulphuric acid (H_2SO_4) in aqueous medium.

i. Oxidation half equation. (1 mark)



Reduction Half-Reaction (Strongest Oxidant):

Iron(III) is reduced:

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$$

Oxidation Half-Reactions (Strongest Reductants):

1. Hydrogen peroxide is oxidized:

$$H_2O_2(aq) \rightarrow O_2(g) + 2H^+(aq) + 2e^-$$

ii. Reduction half

Space for Personal Notes

Question 16 (2 marks)

Vivek is given the following electrochemical series for some reactions that are not occurring at SLC.

$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	$+0.1$
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	$+0.25$
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	$+0.69$
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.90

Vivek adds copper (II), iodide, and tin (IV) ions to a beaker. Following this, he adds a strip of zinc metal.

Predict the half equations of the reaction occurring.

a. Oxidation reaction. (1 mark)

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

b. Reduction reaction. (1 mark)

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

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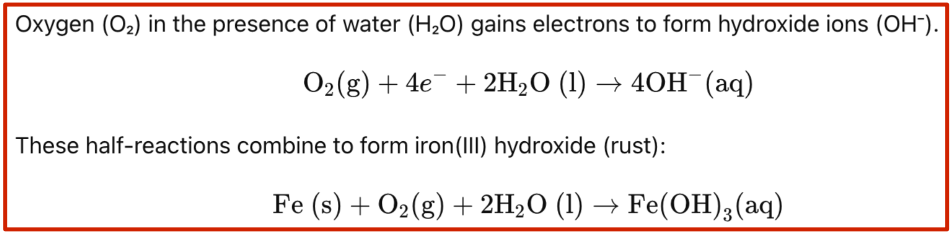
Sub-Section [1.7.2]: Identify Differences Between Direct & Indirect Redox Reactions, & Features of ECS

Question 17 (4 marks)

Betty is analysing a reaction between oxygen gas (O_2) and iron metal (Fe).

a. State and justify using chemical equations whether the reaction is spontaneous. (2 marks)

Oxygen gas and iron metal can react spontaneous due to a transfer of electrons from the oxidation of iron, producing electrons, which are consumed in the reduction of oxidation gas (1). This is shown in these half equations; (2)



Oxygen (O_2) in the presence of water (H_2O) gains electrons to form hydroxide ions (OH^-).

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

These half-reactions combine to form iron(III) hydroxide (rust):

$$Fe\ (s) + O_2(g) + 2H_2O\ (l) \rightarrow Fe(OH)_3(aq)$$

b. When Betty sets up this reaction in the laboratory, she notices no reaction occurring. Provide a justification for this observation. (1 mark)

The rate of reaction is extremely slow, meaning no physical evidence of a reaction occurring is present (1).

c. Provide a physical application of this chemical reaction. (1 mark)

Iron metal rusting (1)

Space for Personal Notes

Question 18 (2 marks)

Precious adds a strip of magnesium (Mg) to a beaker containing a 1.0 *M* aqueous solution of aluminium phosphate ($AlPO_4$).

a. Predict the half-reactions occurring in the beaker.

i. Oxidation reaction. (0.5 marks)

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

ii. Reduction reaction. (0.5 marks)

$$2H_2O(l) + 2e^{-} \rightleftharpoons H_2(g) + 2OH^{-}(aq)$$

b. State the type of energy conversion occurring in this chemical reaction. (1 mark)

Chemical energy → Heat/thermal energy

Question 19 (2 marks)

State the role of the Standard Hydrogen Electrode (SHE) in the electrochemical series.

The Standard Hydrogen Electrode (SHE) serves as the reference electrode in the electrochemical series. It is assigned a standard electrode potential of 0.00 *V* under standard conditions (1 *M* concentration, 1 *atm* pressure, and 25°C). (1) The SHE provides a consistent baseline for comparing the electrode potentials of other half-reactions, allowing the determination of the relative tendency of substances to gain or lose electrons, allowing other the construction of an ECS based on whether other reagents can spontaneously react with it (2).

Space for Personal Notes

Question 20 (3 marks)

Discuss the limitations of the electrochemical series when applied to real-world chemical reactions.

The electrochemical series is a useful tool for predicting the relative reactivity of different substances, but it has several limitations when applied to real-world reactions:

1. **Conditions Dependence:** The electrochemical series is based on standard conditions (1 M concentration, 1 atm pressure, and 25°C). In real-world reactions, conditions may differ (e.g., varying concentrations, temperatures, or pressures), which can affect the actual electrode potentials and the feasibility of a reaction.
2. **Reaction Kinetics:** The series does not account for the rate of reaction. Even if a reaction is thermodynamically favourable (based on electrode potentials), it may occur very slowly or not at all due to high activation energy or other kinetic barriers.
3. **Complexity of Real-World Systems:** In real systems, multiple factors such as the presence of catalysts, the medium (e.g., acidic or basic solutions), and side reactions can influence the outcome. The electrochemical series only provides a simplified view and may not fully predict all the variables that impact the reaction in practice.

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Sub-Section [1.7.3]: Find the Strongest Oxidants/Reductants by Constructing Your Own ECS

Question 21 (4 marks)

a. There are three unknown substances J, K, and L. The following half-equations are given, but their *E*° values are not given.

Reaction
$J^{2+}(aq) + 2e^- \rightarrow J(s)$
$K^{2+}(aq) + 2e^- \rightarrow K(s)$
$L^{2+}(aq) + 2e^- \rightarrow L(s)$

It is known that when J is mixed into a solution of L^{2+}, a reaction begins to occur.

It is also known that when K is mixed into a solution of L^{2+}, no reaction occurs.

Rank the three metals in terms of their decreasing oxidant strength. (2 marks)

$\boxed{K^{2+}, L^{2+}, J^{2+}}$

b. There are three unknown substances A, B, and C. The following half-equations are given, but their $E°$ values are not given.

Reaction
$A^{2+}(aq) + 2e^- \rightarrow A(s)$
$B^{2+}(aq) + 2e^- \rightarrow B(s)$
$C^{2+}(aq) + 2e^- \rightarrow C(s)$

i. A plastic rod coated in metal A reacts vigorously with a solution of $B(NO_3)_2$ and CSO_4.

ii. Metal C is able to react with B^{2+} but not with A^{2+}.

Rank the three metals in terms of their decreasing oxidant strength. (2 marks)

$\boxed{\text{B, C, A}}$

Space for Personal Notes

Question 22 (3 marks)

Four unknown substances P, Q, R, and S are present. The half-equations are provided without $E°$ values.

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

$$Q^{+}(aq) + e^- \rightarrow Q(g)$$

$$R^{3+}(aq) + 3e^- \rightarrow R(s)$$

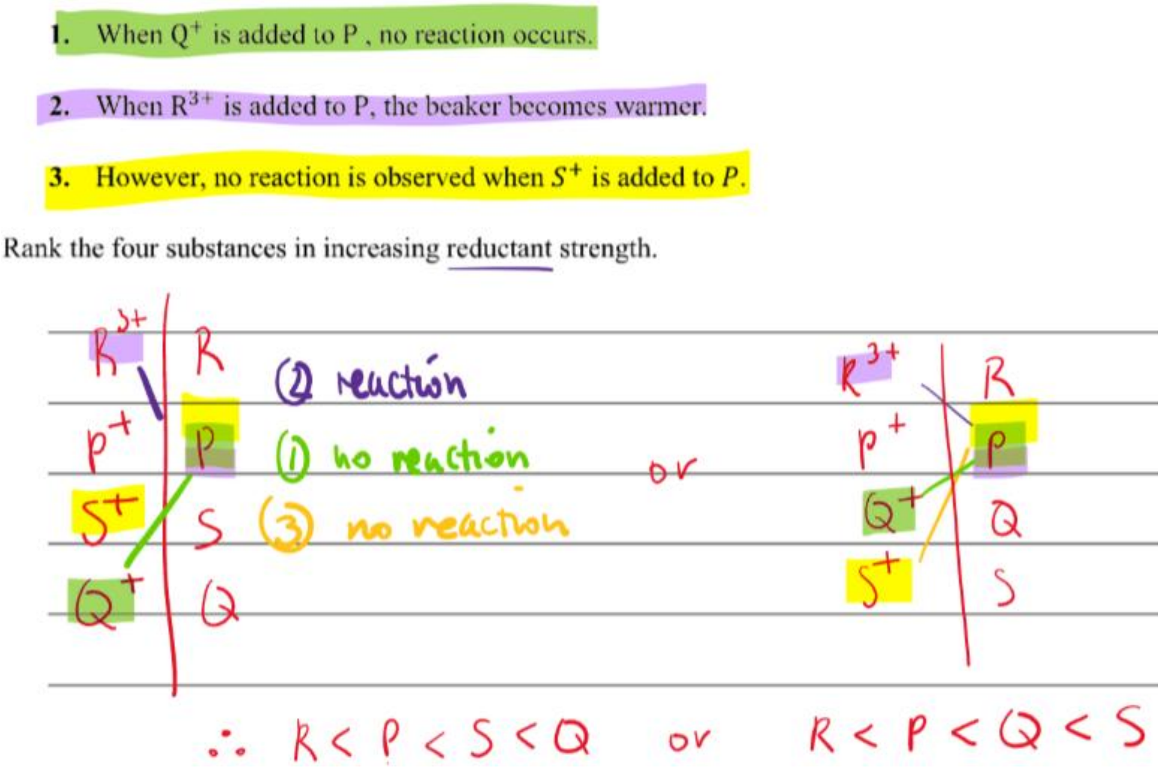
$$S^{+}(aq) + e^- \rightarrow S(l)$$

The following observations are recorded:

1. When Q^+ is added to P , no reaction occurs.
2. When R^{3+} is added to P, the beaker becomes warmer.
3. However, no reaction is observed when S^+ is added to P.

Rank the four substances in increasing reductant strength.

$R < P < S < Q$



1. When Q^+ is added to P , no reaction occurs.
2. When R^{3+} is added to P, the beaker becomes warmer.
3. However, no reaction is observed when S^+ is added to P.

Rank the four substances in increasing reductant strength.

∴ R < P < S < Q or R < P < Q < S

Question 23 (4 marks)

A scientist is given five metals and 1 *M* solutions of their nitrates.

The metals are labelled A, B, C, D, and E and the solutions are labelled A^{2+}, B^{2+}, C^{2+}, D^{2+}, and E^{2+}.

The scientist carries out several experiments, and the results obtained are as follows:

1. When metal A is dipped into all solutions, it reacts vigorously with D^{2+} and E^{2+}, but no reaction occurs with B^{2+} or C^{2+}.
2. When metal D is dipped into all solutions, it reacts only with E^{2+}.
3. Metal C, when dipped into solutions, reacts with A^{2+} but not with B^{2+}.

a. State the strongest oxidant and reductant. (2 marks)

B → Strongest reducatant
E^{2+} → Strongest oxidant

b. Rank the metals in terms of decreasing $E°$ values. (2 marks)

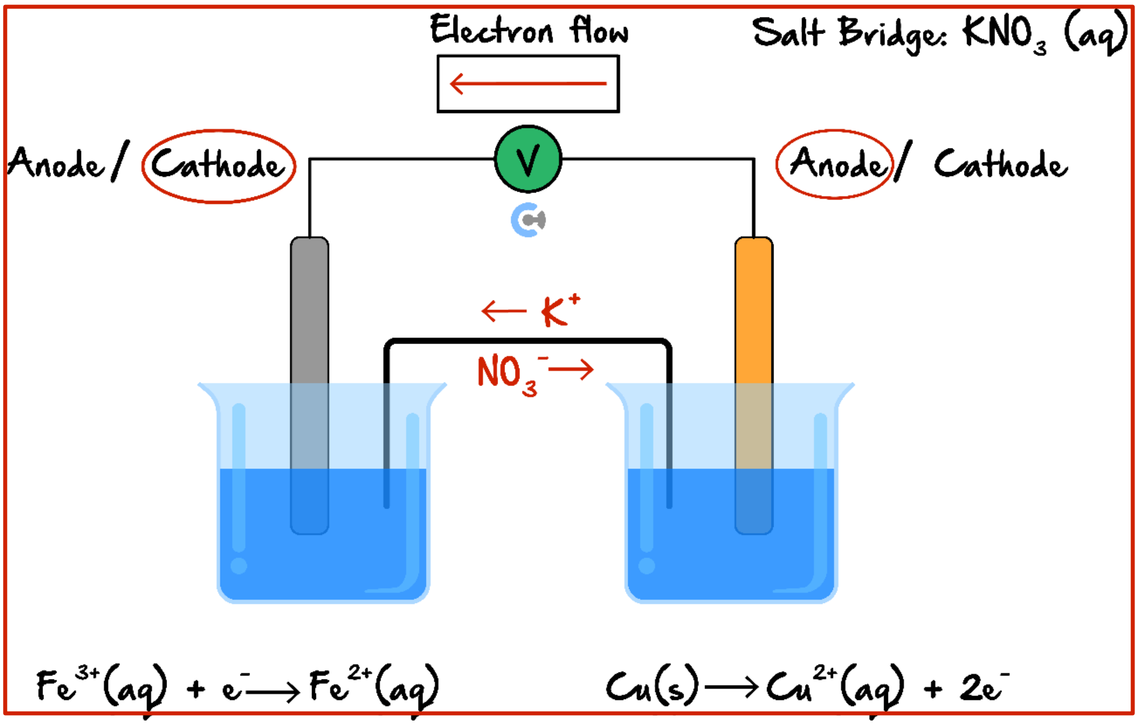
$E > D > A > C > B$

Space for Personal Notes

Section C: [1.8] - Galvanic Cells (Checkpoints) (26 Marks)

Sub-Section [1.8.1]: Identify Electrodes, Salt Bridge/Electron Movement during Galvanic

Question 24 (3 marks)



Use the provided diagram of a galvanic cell, and label the cathode, anode, electron flow and ion movement in the salt bridge.

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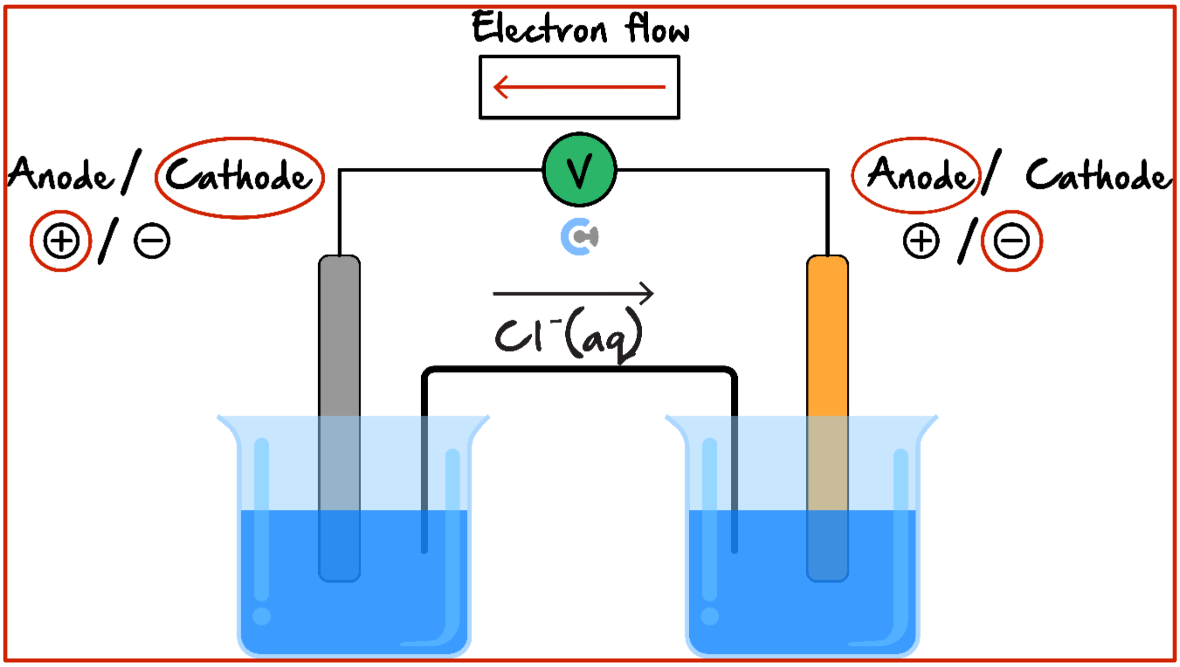
Question 25 (3 marks)

Use the provided galvanic cell diagram to label the cathode, anode, electrode polarities and electron flow.



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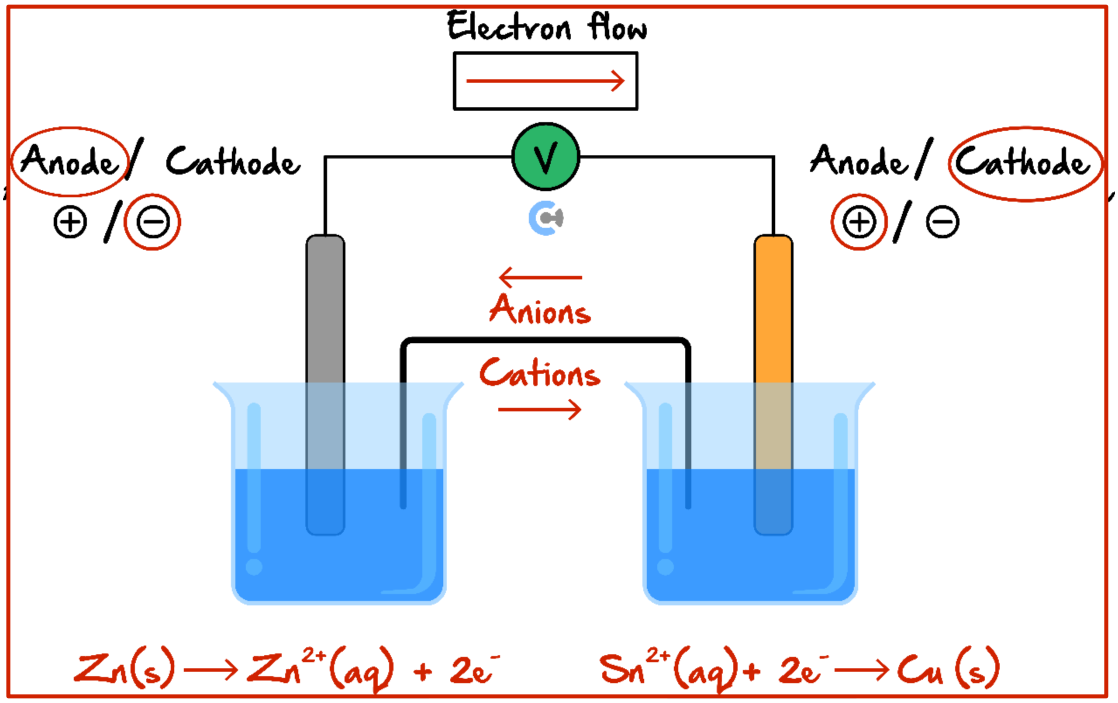
Question 26 (3 marks)

Use the provided half-equations:

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

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

Draw the galvanic cell, labelling the anode, cathode, electron flow, ion movement in the salt bridge and electrode polarities.



$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$ $Sn^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

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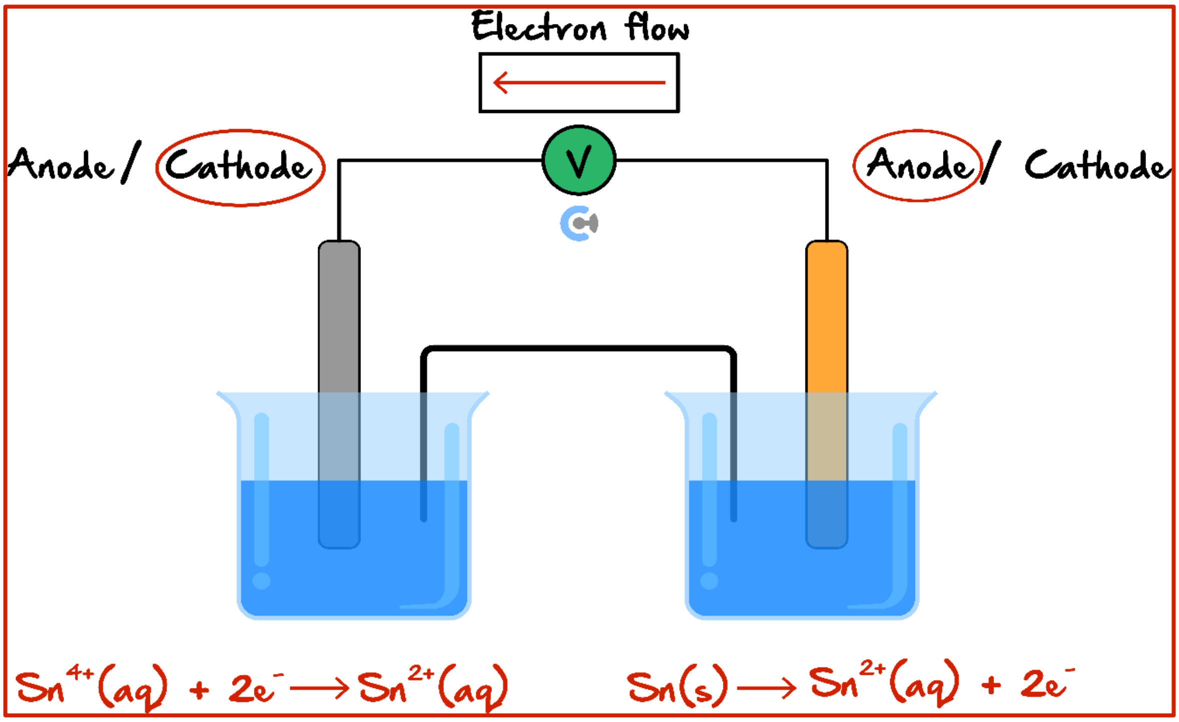
Sub-Section [1.8.2]: Write Reactions in Galvanic Cells & Calculate the Maximum EMF Produced

Question 27 (5 marks)

Matthias is interested in the Sn-Sn^{4+}(aq) galvanic cell.



a. Write the half-equations for this cell. (2 marks)

1. **Anode (oxidation):**

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

2. **Cathode (reduction):**

$$\mathrm{Sn}^{4+}(aq) + 2e^- \rightarrow \mathrm{Sn}^{2+}(aq)$$

b. Label the cathode, anode, their polarities and electron flow in the diagram above. (3 marks)

Space for Personal Notes

Question 28 (7 marks)

Angel is curious about the Cu^{2+}/Cu^{+} and Cr^{3+}/Cr^{2+} half-cell. The half-equation for the cells and their relative E° values are provided.

$$Cr^{3+}(aq) + e^{-} \rightarrow Cr^{2+}(aq) \quad E^\circ = -0.41\ V$$

$$Cu^{2+}(aq) + e^{-} \rightarrow Cu^{+}(aq) \quad E^\circ = +0.15\ V$$

a. Write the half-equations for the galvanic cell. (2 marks)

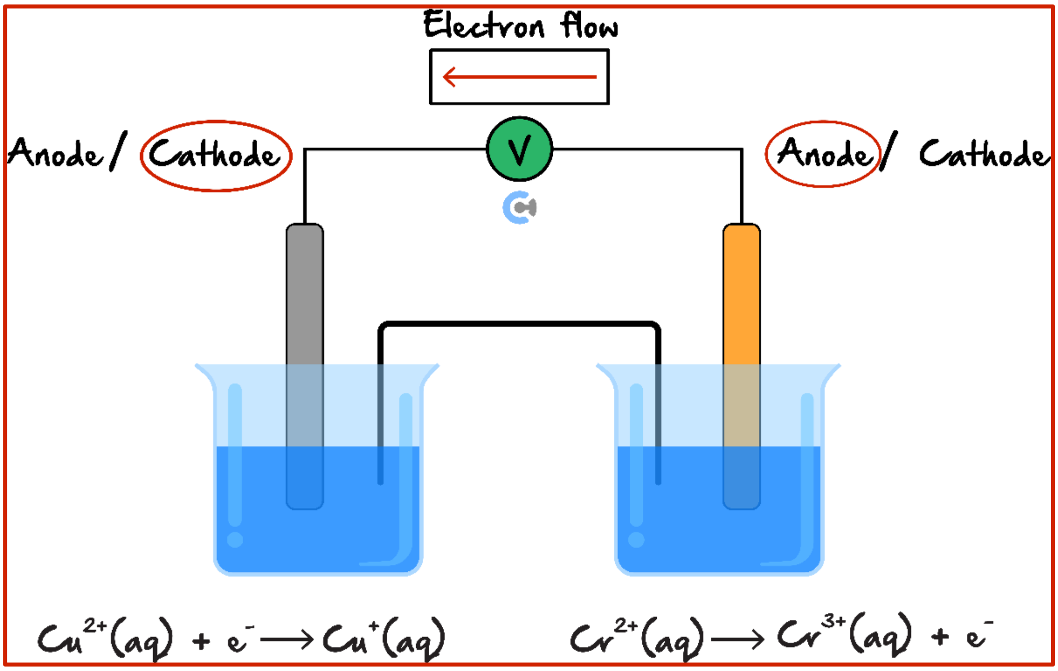
Oxidation: $Cr^{3+}(aq) \rightarrow Cr^{2+}(aq) + e^{-}$
Reduction: $Cu^{2+}(aq) + e^{-} \rightarrow Cu^{+}(aq)$

b. Calculate the EMF of the cell. (2 marks)

$$E^\circ_{cell} = 0.15 - (-0.41)$$

$$E^\circ_{cell} = 0.15 + 0.41 = 0.56\ V$$

c. Label the galvanic cell in Angel's experiment. (3 marks)

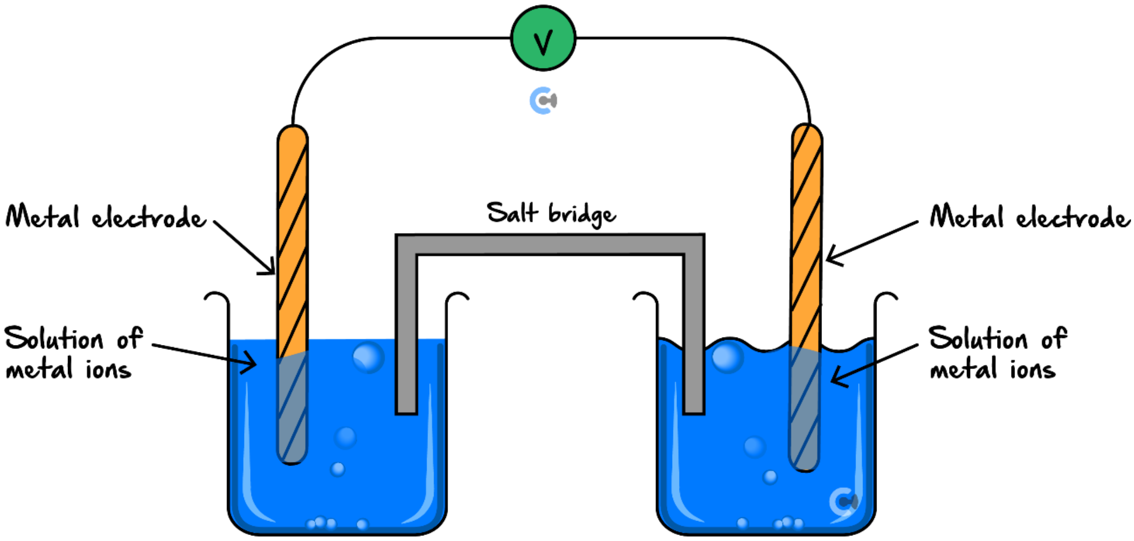


Question 29 (1 mark)

Four half-cells are constructed as follows.

- **Half-cell I:** An electrode of metal X in a 1.0 *M* solution of X^+ ions.
- **Half-cell II:** An electrode of metal Y in a 1.0 *M* solution of Y^+ ions.
- **Half-cell III:** An electrode of metal Z in a 1.0 *M* solution of Z^+ ions.
- **Half-cell IV:** An electrode of W metal in a 1.0 *M* solution of W^{2+} ions.

The half-cells are connected in pairs as shown below to form a series of galvanic cells.



Half-Cells Used	Positive Electrode	Negative Electrode	Voltage Produced (*V*)
I and IV	X	W	0.35
II and IV	W	Y	0.45
III and IV	W	Z	1.20
II and III	Y	Z	0.65

Which of the following alternatives lists the metals in order of increasing strength as reductants?

A. Z, Y, W, X

B. X, W, Y, Z

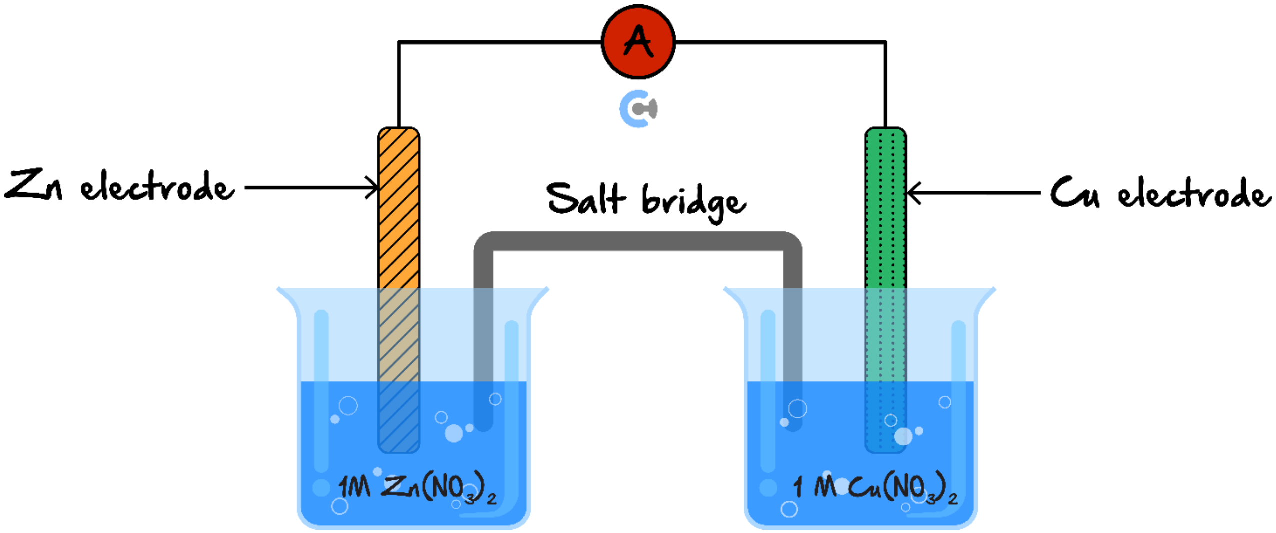
C. W, X, Z, Y

D. Z, W, X, Y

Sub-Section [1.8.3]: Identify & Explain Observations during Operation of Galvanic Cells

Question 30 (1 mark)

A galvanic cell is set up as shown in the diagram below with Zn/Zn^{2+} and Cu/Cu^{2+} half-cells.



When this cell is operating:

A. A gas forms at the Zn electrode.

B. The mass of the Cu electrode increases.

C. Zn^{2+} ions move towards the salt bridge.

D. Electrons move from the Cu electrode to the Zn electrode.

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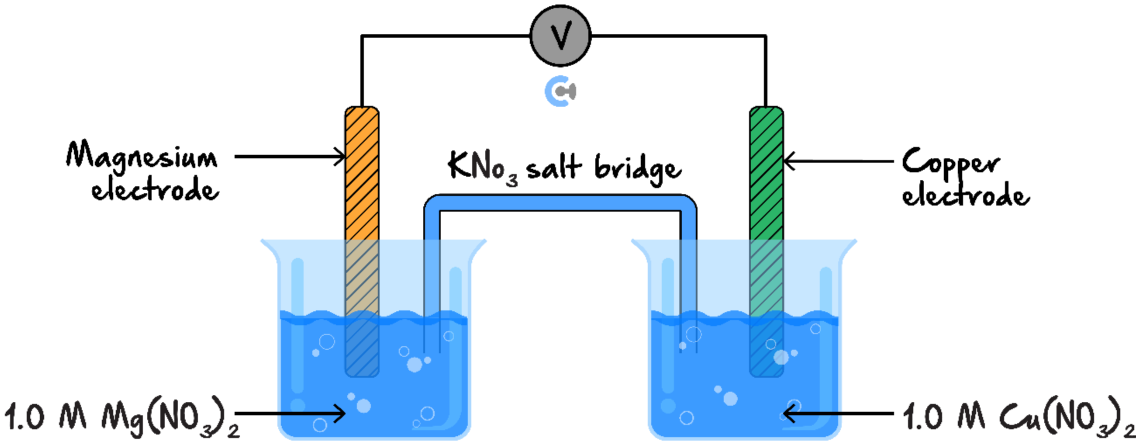
Question 31 (2 marks)

The following diagram is relevant to the two questions below.



a. Which one of the following statements about the cell above is true as the cell discharges? (1 mark)

A. The copper electrode is the anode.

B. The concentration of Mg^{2+} ions will increase.

C. The maximum voltage delivered by this cell will be 2.71 V.

D. Electrons in the external circuit will flow from the magnesium electrode to the copper electrode.

b. What should be observed at the magnesium electrode as the cell discharges? (1 mark)

A. No change will be observed at this electrode.

B. The electrode will become thinner as magnesium dissolves into the solution.

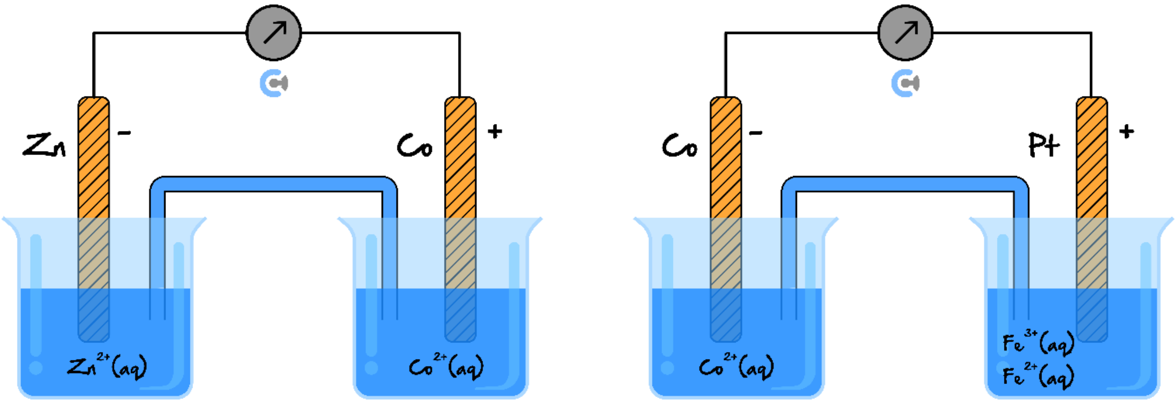
C. Crystals will form over the surface of the electrode.

D. Bubbles of gas will form over the surface of the electrode.

Space for Personal Notes

Question 32 (1 mark)

Two standard galvanic cells are shown below.



On the basis of the polarity of the electrodes shown above, which one of the following reactions would **not** be expected to occur spontaneously? (1 mark)

A. $Co^{2+}(aq) + Zn(s) \rightarrow Co(s) + Zn^{2+}(aq)$

B. $2Fe^{3+}(aq) + Co(s) \rightarrow 2Fe^{2+}(aq) + Co^{2+}(aq)$

C. $2Fe^{3+}(aq) + Zn(s) \rightarrow 2Fe^{2+}(aq) + Zn^{2+}(aq)$

D. $2Fe^{2+}(aq) + Co^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Co(s)$

Space for Personal Notes

Section D: [1.9] – Fuel Cells (Checkpoints) (62 Marks)

Sub-Section [1.9.1]: Write Fuel Cell Half & Overall Reactions in Acidic Conditions

Question 33 (2 marks)

Consider a fuel cell between hydrogen gas, as the fuel, and oxygen gas. Remember to include states in your answer.

a. Hydrogen gas. (1 mark)

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

b. Oxygen gas. (1 mark)

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

Question 34 (3 marks)

For each of the following, write the balanced half-equation for the reaction occurring at the anode in acidic conditions. Assume that carbon dioxide is produced.

a. A fuel cell involving ethanol as a reactant. (1 mark)

$$C_2H_5OH(l) + 3H_2O(l) \rightarrow 2CO_2(g) + 12H^+(aq) + 12e^-$$

b. A fuel cell involving ethane as a reactant. (1 mark)

$$C_2H_6(g) + 4H_2O(l) \rightarrow 2CO_2(g) + 14H^+(aq) + 14e^-$$

c. A fuel cell involving propanol as a reactant. (1 mark)

$$C_3H_7OH(l) + 5H_2O(l) \rightarrow 3CO_2(g) + 18H^+(aq) + 18e^-$$

Question 35 (4 marks)

Write the half-equations for the fuel cell reaction involving butanol and oxygen gas.

a. Reduction half-equation. (1 mark)

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

b. Oxidation half-equation. (2 marks)

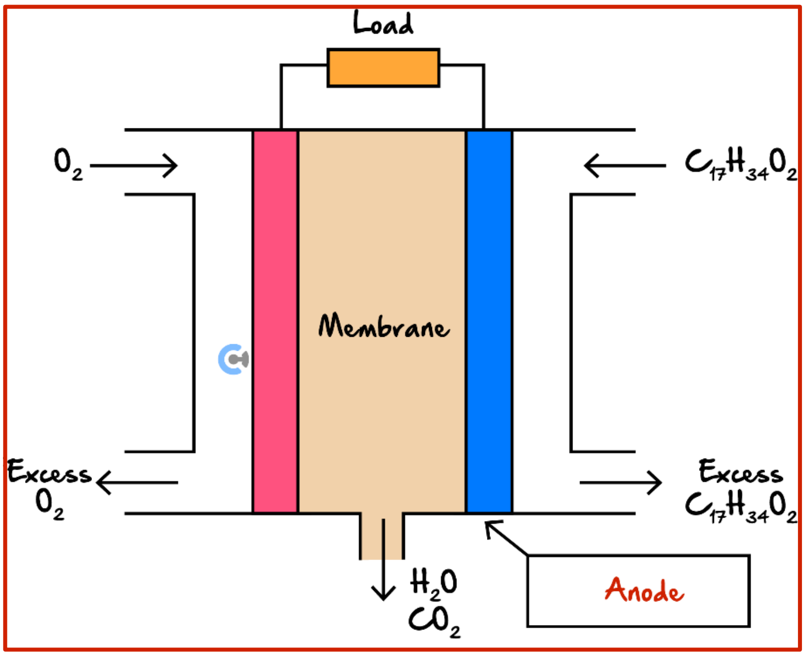
$$\boxed{C_4H_9OH(l) + 7H_2O(l) \rightarrow 4CO_2(g) + 24H^+(aq) + 24e^-}$$

c. Overall reaction. (1 mark)

$$\boxed{2C_4H_9OH(l) + 12O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)}$$

Question 36 (7 marks)

Biodiesel is an example of a renewable fuel. Some farming equipment manufacturers have tried to make a fuel cell in order to more efficiently use leftover livestock feed. A diagram of an acidic biodiesel fuel cell is shown below.



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

b. Write the half-equation for the reaction occurring at the anode. (2 marks)

$$C_{17}H_{34}O_2 + 32H_2O(l) \rightarrow 17CO_2(g) + 98H^+(aq) + 98e^-$$

c. Write a balanced equation for the overall reaction which takes place at SLC. (2 marks)

$$C_{17}H_{34}O_2(l) + \frac{49}{2}O_2(g) \rightarrow 17CO_2(g) + 17H_2O(l)$$

d. Explain whether this cell would be considered renewable or not. (2 marks)

The cell is renewable. The biodiesel used in the fuel cell is produced through transesterification of fats and oils. Therefore, it can be produced in a relatively short time period by natural processes and hence, is renewable.

Space for Personal Notes

Sub-Section [1.9.2]: Identify Key Features of Fuel Cells Including Continuous Supply, Electrolyte Movement, and Properties of Electrodes

Question 37 (2 marks)

Explain the key characteristics of a fuel cell.

The continuous supply of reactants, this is due to it not starting with a set amount of reactants at either electrode.

Question 38 (3 marks)

For the following table, mark each statement as either True or False, with regards to a fuel cell involving methane and oxygen gas in acidic conditions.

Statement	**True**	**False**
a. The overall reaction is the same as combustion.	✓	
b. Electrons flow from the anode to the cathode.	✓	
c. Oxygen reacts at the anode.		✓
d. Electrodes must always be inert.	✓	
e. The energy conversion is direct.	✓	

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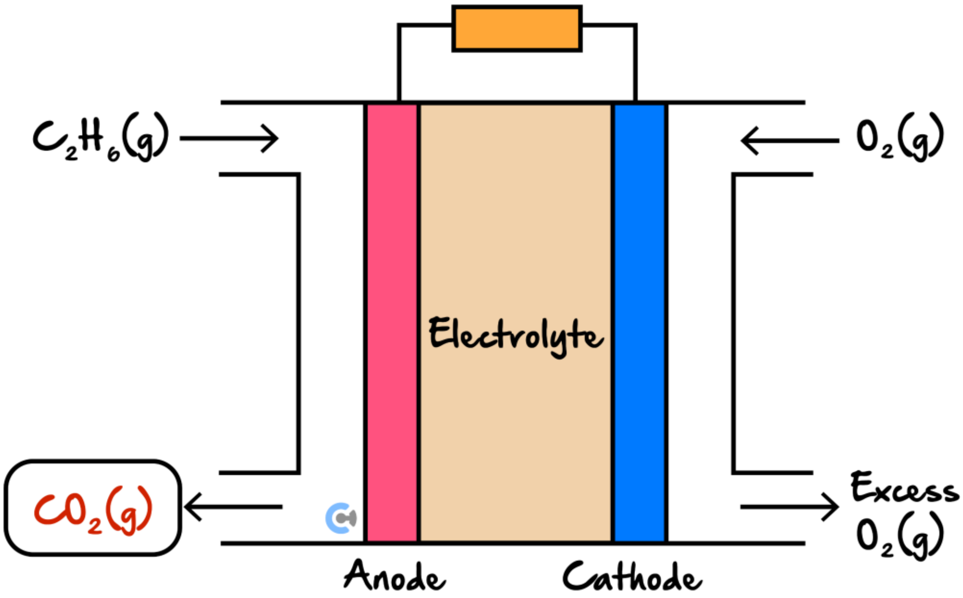
Question 39 (5 marks)

Ethane is a common fuel used in fuel cells.



a. Write the balanced redox reaction for:

i. The half-reaction occurring at the anode. (1 mark)

$C_2H_6(g) + 4H_2O(l) \rightarrow 2CO_2(g) + 14H^+(aq) + 14e^-$

ii. Label the main product in the blank box provided above. (1 mark)

b. State three different qualities electrodes in fuel cells must possess. (3 marks)

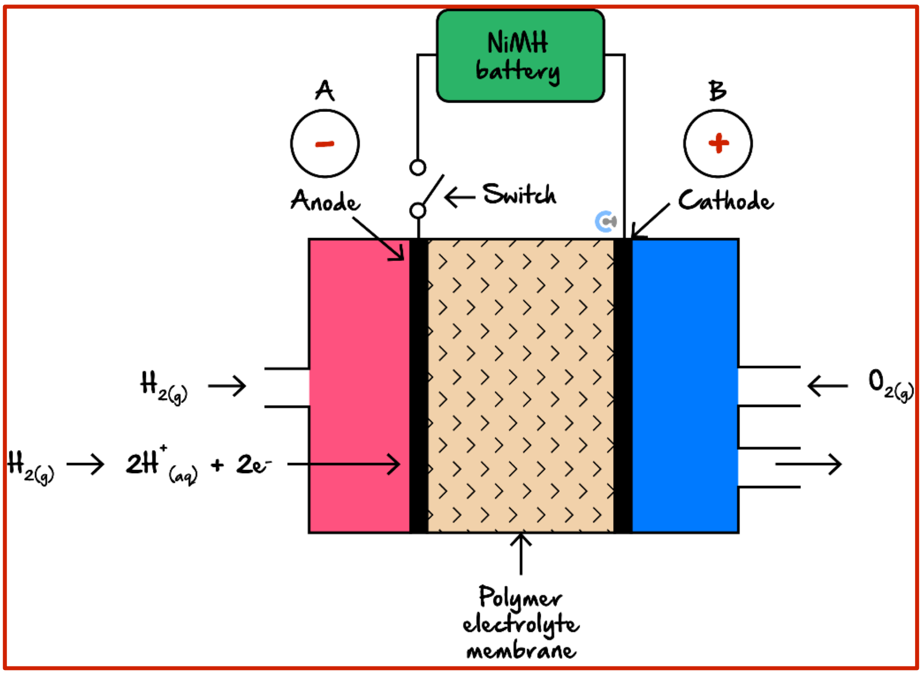
Any three of PICCY:

- ➤ Porous, Inert, Catalyst, Conduct Electricity.

Space for Personal Notes

Question 40 (6 marks)

An example of the real-life design of a hydrogen fuel cell is shown below.



a. On the diagram above, indicate the polarity of the anode and cathode in circles A and B. (1 mark)

b. Write the overall reaction occurring in the cell. (1 mark)

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

c. Explain the function of the polymer electrolyte membrane in the operation of the cell. (2 marks)

The electrolyte allows for the flow of ions that are produced at one electrode to the other electrode. Additionally, it acts as a catalyst to increase the rate of reaction within the fuel cell.

d. State and explain whether this cell is more efficient than a typical combustion engine. (2 marks)

This cell is more efficient than a typical combustion engine as it converts chemical energy into electrical energy.

Sub-Section [1.9.3]: Explain Advantages & Disadvantages of Fuel Cells with Reference to Green Chemistry Principles

Question 41 (2 marks)

State and explain one reason why we would use a fuel cell over a galvanic cell.

Sample Response: Fuel cells are more efficient than galvanic cells while still achieving the same energy conversion, meaning for the same amount of energy less fuel is used.

Question 42 (2 marks)

One environmental advantage of using hydrogen fuel cells instead of gas-powered engines is the reduction of greenhouse gas emissions. State one other environmental advantage, referencing green chemistry principles.

Hydrogen fuel cells produce only water as a byproduct, eliminating harmful pollutants such as carbon monoxide and nitrogen oxides, which aligns with the green chemistry principle of design for degradation.

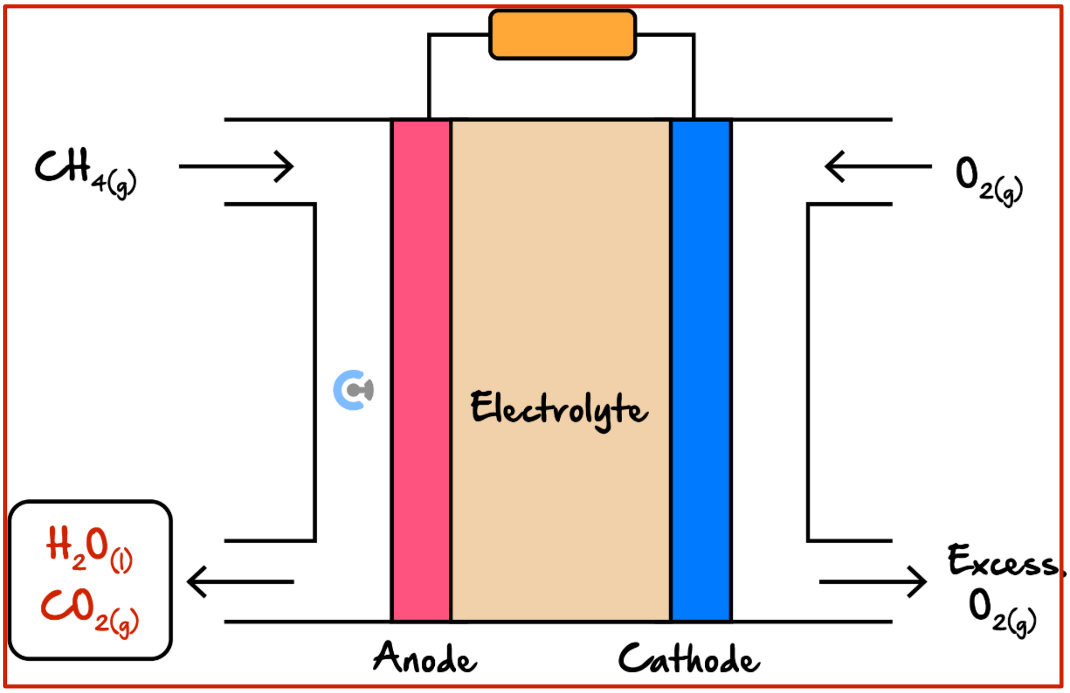
Question 43 (3 marks)

State two limitations of hydrogen fuel cells in practical applications.

- Storage challenge: Hydrogen gas is hard to store because of its danger (very flammable).
- Limited infrastructure: Hydrogen gas stations are hard to come by and so has limited area usage compared to a traditional petrol combustion engine.

Question 44 (7 marks)

Methane can be a source of energy in combustion engines. A fuel cell involving methane is shown below.



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

i. Anode. (1 mark)

$CH_4(g) + 2H_2O(l) \rightarrow CO_2(g) + 8H^+(aq) + 8e^-$

ii. Cathode. (1 mark)

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

b. In the diagram above, label the expected product(s) at the anode. (1 mark)

c. Compare the efficiency and viability of a combustion engine as opposed to a fuel cell, referencing green chemistry principles. (2 marks)

Fuel cells are more efficient as compared to combustion engines. This aligns with the green chemistry principle of efficiency as for the same amount of energy, less environmental impact is made.

d. A scientist is aiming to produce an engine that can produce mass amounts of energy at a single time. Explain whether he should use a fuel cell or an internal combustion engine. (2 marks)

The combustion engine can pack fuel densely and the fuel can be combusted instantly. The fuel cell requires a constant supply of reactants and must react on the electrode itself, which means not a lot of the fuel can be reacted at once, compared to a combustion engine.

Space for Personal Notes

Sub-Section [1.9.4]: Write Fuel Cell Equations in Non-Acidic Conditions

Question 45 (1 mark)

Write the balanced oxidation half-equation in basic conditions of methane oxidising into carbon dioxide.

$$\boxed{CH_4(g) + 8OH^-(aq) \rightarrow CO_2(g) + 6H_2O(l) + 8e^-}$$

Question 46 (3 marks)

For each of the following, write the balanced oxidation half-equation in basic conditions. Assume CO_2 is the only carbon product formed.

a. A fuel cell involving ethane as a reactant. (1 mark)

$$\boxed{\begin{array}{c} C_2H_6(g) + 14OH^-(aq) \rightarrow 2CO_2(g) + 10H_2O(L) + 14e^- \\ C_2H_6(g) + 14OH^-(aq) \rightarrow 2CO_2(g) + 10H_2O(l) + 14e^- \end{array}}$$

b. A fuel cell involving propanol as a reactant. (1 mark)

$$\boxed{\begin{array}{c} C_3H_7OH(l) + 5H_2O(l) \rightarrow 3CO_2(g) + 18H^+(aq) + 18e^- \\ C_3H_7OH(l) + 18OH^-(aq) \rightarrow 3CO_2(g) + 13H_2O(l) + 18e^- \end{array}}$$

c. A fuel cell involving methanol as a reactant. (1 mark)

$$\boxed{\begin{array}{c} CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e^- \\ CH_3OH(l) + 6OH^-(aq) \rightarrow CO_2(g) + 5H_2O(l) + 6e^- \end{array}}$$

Space for Personal Notes

Question 47 (5 marks)

The overall equation for an unknown fuel cell is shown below, in alkaline conditions.

$$X(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

a. What is the equation for the reaction that occurs at the cathode? (1 mark)

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

b. If the fuel itself was an alkane, what could its identity be? (2 marks)

Propane

c. Hence, write the half-equation for the reaction that would occur at the anode. (2 marks)

$$\boxed{C_3H_8(g) + 20OH^-(aq) \rightarrow 3CO_2(g) + 14H_2O(l) + 20e^-}$$

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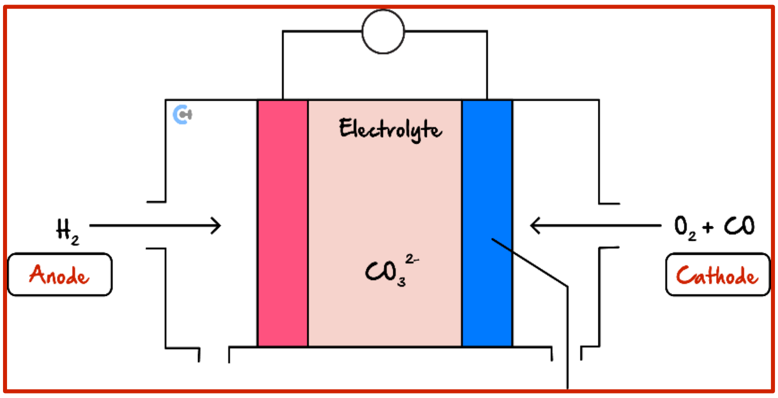
Question 48 (7 marks)

A simplified diagram of a molten carbonate fuel cell (MCFC) is shown below.



In this MCFC, hydrogen is being used as a reactant alongside $O_2(g)$ in order to produce energy. The main carbon-containing product of the cell is carbon monoxide.

a. On the diagram above, label the:

i. Anode and cathode along with the polarities of either electrode. (1 mark)

ii. The direction of electron flow in the external circuit. (1 mark)

Solution Pending

b. Write a balanced half-equation for the cathode. (1 mark)

$$O_2(g) + CO(g) + 2e^- \rightarrow CO_3^{2-}(l)$$

c. Write the balanced half-equation for the anode. (2 marks)

$$2H_2(g) + CO_3^{2-}(l) \rightarrow 2H_2O(l) + CO(g) + 2e^-$$

d. Explain one property of the electrodes in this fuel cell. (2 marks)

Any of PICCY – Porous, inert, catalysts, conducts electricity.

Space for Personal Notes

Section E: [1.10] – Primary Cells & Faraday's Laws (Checkpoints) (25 Marks)

Sub-Section [1.10.1]: Identify Features of Primary Cells & How They Operate

Question 49 (3 marks)

Sodium metal is commonly used in batteries, especially in alkaline and sodium-carbon cells for daily consumer use.

a. By referring to information provided in the Data Book, give one reason why sodium is used as a reactant in these galvanic cells. (1 mark)

Sodium is a relatively strong reductant (-2.71V). This means that many reactants will have larger EMF values and act as a suitable oxidant, allowing for the galvanic cell to operate. (1)

b. When sodium comes into contact with water, an explosion is observed. Using the half equations, explain this observation. (2 marks)

Sodium metal and water will spontaneously react releasing heat. As hydrogen gas produced accumulates, the heat produced can cause ignition of this gas causing an explosion. (1) (2) – half eqns

$$\mathrm{Na(s)} \rightarrow \mathrm{Na^+(aq)} + e^-$$

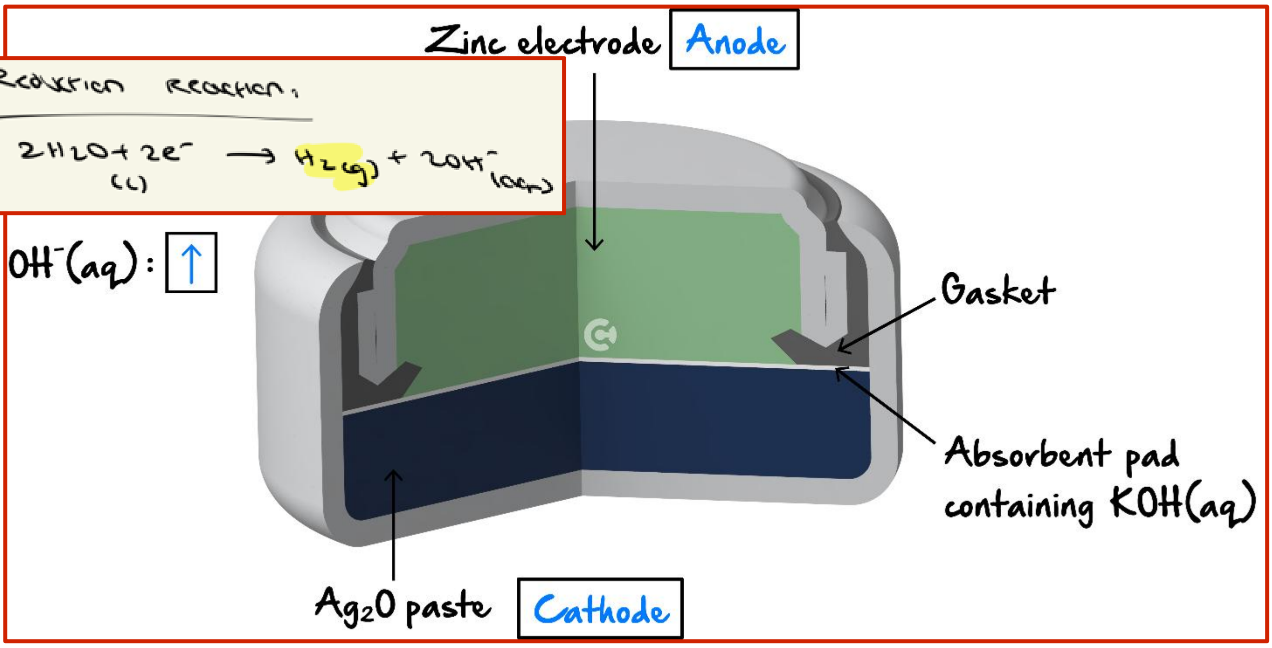
Space for Personal Notes

Question 50 (2 marks)

Zinc-mercury oxide batteries are commonly used commercially due to their long battery life and durability. A diagram of the cell has been provided below.

Reduction reaction:

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



The half equations for the cell have also been provided below:

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

$$HgO(s) + H_2O(l) + 2e^- \rightarrow Hg(l) + 2OH^-(aq)$$

a. Label the anode and the cathode in the boxes provided above. (1 mark)

b. Label the direction of movement of the electrolyte in the cell in the box provided above. (1 mark)

Space for Personal Notes

Question 51 (2 marks)

a. What is a key feature of primary cells that allows for their commercial usage over fuel cells? (1 mark)

A. Greater efficiency in the production of energy.

B. Cheaper electrodes that reduce the overall cost of the battery.

C. Separation of reactants into two half-cells is cheaper than a constant supply of reactants.

D. Less specific electrodes that are easier to source.

b. Which of the following outlines the properties required for electrodes in primary cells? (1 mark)

A. Porous, inert, catalytic, conduct electricity.

B. Reactive, catalytic, porous.

C. Conductive of electrons.

D. Porous, inert, conductive.

Space for Personal Notes

Sub-Section [1.10.2]: Apply Faraday's First & Second Law and $Q = It$ & $Q = n(e)F$ to Calculations

Question 52 (4 marks)

Answer the following questions regarding three separate galvanic cells.

a. In a galvanic cell, 250 C of electric charge passes through the circuit in 20 minutes. Calculate the current, in A, running through the cell. (1 mark)

$$I = \frac{Q}{t}$$

$$= \frac{250}{20 \times 60}$$

$$= 0.208\ A$$

b. In another galvanic cell, 1.46 A of current runs through the cell during a 10.0-minute period. Calculate the moles of electrons produced in the cell. (1 mark)

$$Q = 1.46 \times 10 \times 60$$

$$= 876\ C$$

$$\therefore \boxed{n(e^-)} \ \frac{876}{96500}$$

$$= \boxed{0.00908\ mol}$$

$$\approx 9.08 \times 10^{-3}\ mol$$

c. Calculate the moles of zinc (Zn) produced in a cell, when 5.42 A of current is running through the circuit for 35 minutes. The half equation for zinc has been provided below: (2 marks)

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

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

$$= \frac{5.42 \times 35 \times 60}{96500}$$

$$= \boxed{0.117948\ mol}$$

$$\therefore n(Zn) = \frac{0.117948}{2}$$

$$= \boxed{0.05897\ mol}$$

$$\approx 5.90 \times 10^{-2}\ mol$$

Space for Personal Notes

Question 53 (3 marks)

Joanne has set up a copper-nickel galvanic cell for a school experiment.

a. Write the half equations for the galvanic cell. (1 mark)

Oxidation Half reaction:

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

Reduction Half reaction:

$$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^{-}$$

b. Joanne ran the cell for 19.5 minutes and found that 8.75 *A* of current passed through the cell. Calculate the mass of metal deposited on the electrode. (2 marks)

$$\boxed{n(e^{-})} = \frac{19.5 \times 60 \times 8.75}{96500}$$

$$= \boxed{0.106088\ mol}$$

$$\therefore m(Cu) = \frac{0.106088}{2} \times 63.5$$

$$= 3.37g$$

c. The setup is reset and the experiment is run again and Joanne notices 4.55 *g* of copper produced. Given that 2.45 *A* of current was passed through the cell, calculate the time (in seconds) for which the cell was running.

$$n(Cu) = \frac{4.55}{63.5}$$

$$= \boxed{0.07165\ mol}$$

$$\boxed{n(e^{-})} = \frac{It}{F} = 0.07165 \times 2$$

$$\therefore t = \frac{0.07165 \times 2 \times 96500}{2.45}$$

$$= \boxed{2537.33\ s}$$

$$\approx 2.54 \times 10^{3} \text{ seconds}$$

Space for Personal Notes

Question 54 (2 marks)

Scott sets up a galvanic cell where chromium metal oxidises to form chromium (III) ions. Given that the change in mass of chromium was 2.55 g and the cell ran for 350 seconds, calculate the current running through the cell.

$$n(Cr) = \frac{2.55}{52}$$

$$= \boxed{0.049\ mol}$$

As $Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)} + 3e^-$,

$$\therefore \boxed{n(e^-)} = 3 \times 0.049$$

$$= \boxed{0.147\ mol}$$

$$\therefore \frac{It}{F} = 0.147$$

$$I = \frac{0.147 \times 96500}{350}$$

$$= 40.5\ A$$

Space for Personal No

Sub-Section [1.10.3]: Calculate the Charge of a Metal

Question 55 (1 mark)

Kevin runs a galvanic cell where 9.54 mol of iron is formed on the cathode. Given that the moles of electrons running through the cell is 3.18 mol, calculate the charge of the iron ions in the cell.

$$\frac{9.54}{3.18} = 3$$

Fe^{3+}

Question 56 (2 marks)

Shiven sets up a galvanic cell in the school laboratory and notes that 27.6 g of manganese metal deposits on the electrode. Given that the amount of electrons running through the cell is 1.01 mol, calculate the charge of the manganese ions in the cell.

$$n(Mn) = \frac{27.6}{54.9}$$

$$= \boxed{0.5027\ mol}$$

$$\therefore \frac{\boxed{n(e^-)}}{n(Mn)} = \frac{1.01}{0.5027}$$

$$\approx 2$$

$$\therefore Mn^{2+}$$

Space for Personal Notes

Question 57 (3 marks)

A solution of titanium ions is reduced at a cathode, whereby it is found that a current of 2.75 A is produced over 9.75 hours. It is found that 15.9 g of titanium metal is deposited at the cathode. Find the charge of the titanium ions.

$Q = It = 2.75 \times 9.75 \times 60 \times 60$

$= 96525C$

$n(e^-) = \frac{Q}{F} = \frac{96525C}{96500} = 1.00mol$

$n(Ti) = \frac{m}{Mr} = \frac{15.9g}{47.9} = 0.332mol$

$n(Ti) : n(e^-)$

$0.332 : 1.00$

$1 : 3$

Ti^{3+}

Question 58 (3 marks)

A solution of iodine ions is reduced at a cathode, whereby it is found that a current of 5.80 A is produced over 4.04 hours. It is found that 22.2 g of iodine metal is deposited at the cathode. Find the charge of the iodine ions.

$Q = It = 5.80A \times 4.04 \times 60 \times 60$

$= 84355C$

$n(e^-) = \frac{Q}{F} = \frac{84355C}{96500} = 0.874mol$

$n(I) = \frac{m}{Mr} = \frac{22.2g}{126.9} = 0.175mol$

$n(I) : n(e^-)$

$\frac{0.175}{0.175} : \frac{0.874}{0.175}$

$1 : 4.99$

$\therefore I^{5+}$

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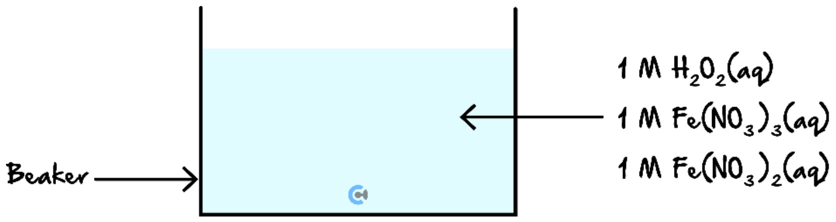
Section F: [1.6 - 1.10] - Overall (VCAA Qs) (75 Marks)

Question 59 (1 mark)

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

At standard conditions, solutions of hydrogen peroxide, H_2O_2, iron(III) nitrate, $Fe(NO_3)_3$, and iron(II) nitrate, $Fe(NO_3)_2$, were added to a beaker. The initial concentrations of H_2O_2, $Fe(NO_3)_3$ and $Fe(NO_3)_2$ in the beaker were all 1 M.



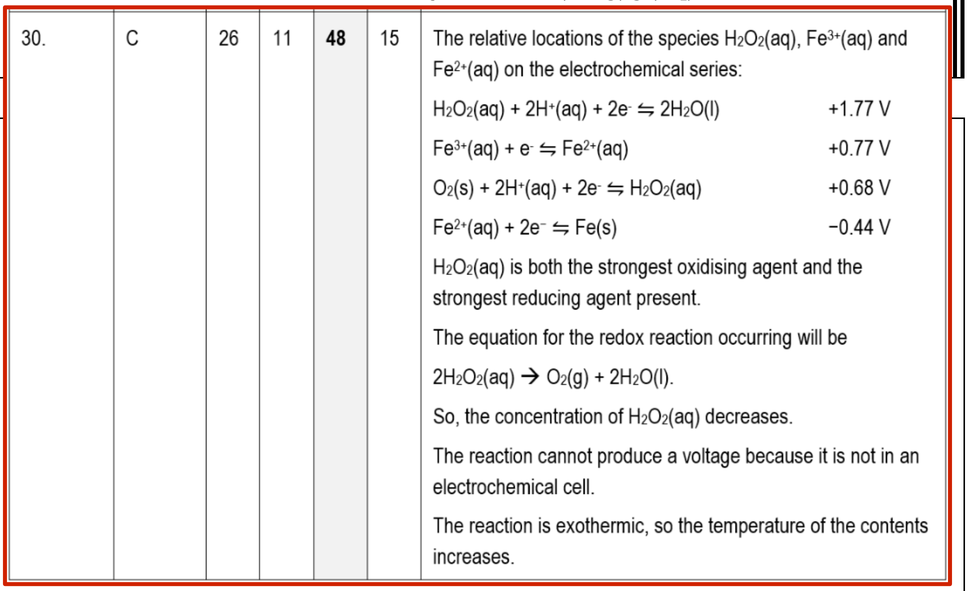
Which one of the following statements is correct?

A. Iron, Fe, is deposited at the bottom of the beaker.

B. The two half-reactions in the beaker immediately produce 1.09 V.

C. The concentration of $H_2O_2(aq)$ decreases immediately since it is the strongest reducing agent.

D. The temperature of the contents in the beaker decreases immediately when $Fe(NO_3)_3(aq)$ reacts with $H_2O_2(aq)$.



30.	C	26	11	**48**	15	The relative locations of the species $H_2O_2(aq)$, $Fe^{3+}(aq)$ and $Fe^{2+}(aq)$ on the electrochemical series: $H_2O_2(aq) + 2H^+(aq) + 2e^- \leftrightharpoons 2H_2O(l)$ +1.77 V $Fe^{3+}(aq) + e^- \leftrightharpoons Fe^{2+}(aq)$ +0.77 V $O_2(s) + 2H^+(aq) + 2e^- \leftrightharpoons H_2O_2(aq)$ +0.68 V $Fe^{2+}(aq) + 2e^- \leftrightharpoons Fe(s)$ −0.44 V $H_2O_2(aq)$ is both the strongest oxidising agent and the strongest reducing agent present. The equation for the redox reaction occurring will be $2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$. So, the concentration of $H_2O_2(aq)$ decreases. The reaction cannot produce a voltage because it is not in an electrochemical cell. The reaction is exothermic, so the temperature of the contents increases.

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Question 60 (1 mark)

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

Consider the following half-equation.

$$ClO_2(g) + e^- \rightleftharpoons CIO_2^-(aq)$$

It is also known that:

- $ClO_2(g)$ will oxidise $HI(aq)$, but not $HCl(aq)$.
- $Fe^{3+}(aq)$ will oxidise $HI(aq)$, but not $NaClO_2(aq)$.

Based on this information, $Fe^{2+}(aq)$ can be oxidised by:

A. $Cl_2(g)$ and $I_2(aq)$.

B. $Cl_2(g)$, but not $ClO_2(g)$

Solution Pending

C. $ClO_2(g)$and $Cl_2(g)$, but not $I_2(aq)$.

D. $Cl_2(g)$, $ClO_2(g)$and $I_2(aq)$.

Question 61 (1 mark)

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

The reaction between hydrogen peroxide and ammonium ions is represented by the following equation.

$$3H_2O_2(aq) + 2NH_4^+(aq) \rightarrow N_2(g) + 2H^+(aq) + 6H_2O(l)$$

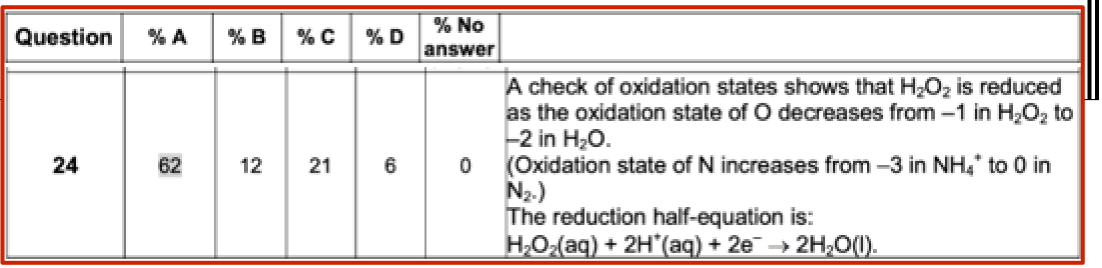
Which one of the following is the correct half-equation for the reduction reaction?

A. $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$

B. $2NH_4^+(aq) \rightarrow N_2(g) + 8H^+(aq) + 6e^-$

C. $2NH_4^+(aq) + 2e^- \rightarrow N_2(g) + 4H_2(g)$

D. $H_2O_2(aq) + 2H_2O$



Question	% A	% B	% C	% D	% No answer	
24	62	12	21	6	0	A check of oxidation states shows that H_2O_2 is reduced as the oxidation state of O decreases from –1 in H_2O_2 to –2 in H_2O. (Oxidation state of N increases from –3 in NH_4^+ to 0 in N_2.) The reduction half-equation is: $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$.

Question 62 (1 mark)

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

Hydrogen, H_2, fuel cells and H_2-powered combustion engines can both be used to power cars. Three statements about H_2 fuel cells and H_2-powered combustion engines are given below:

I. Neither H_2 fuel cells nor H_2-powered combustion engines produce greenhouse gases.
II. Less H_2 is required per kilometre travelled when using an H_2-powered combustion engine than when using H_2 fuel cells.
III. More heat per kilogram of H_2 is generated in an H_2-powered combustion engine than in H_2 fuel cells.

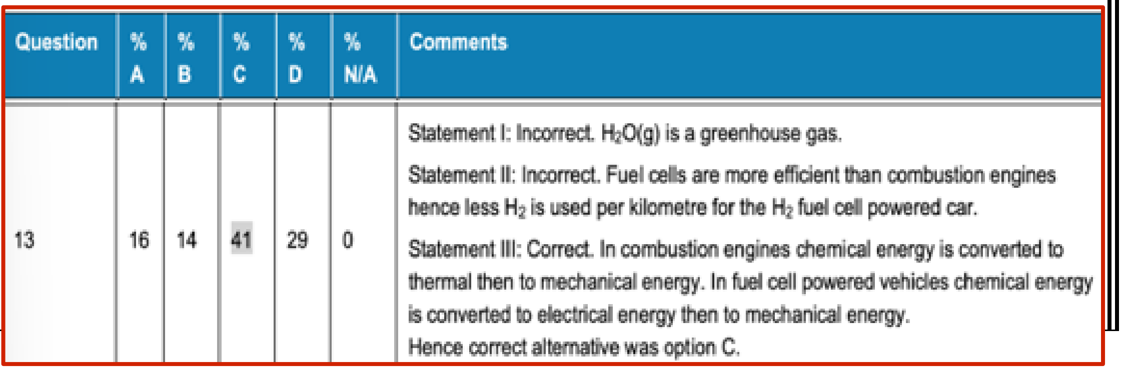
Which of the statements above is correct?

A. II only

B. I and II only

C. III only

D. I and III only



Question	% A	% B	% C	% D	% N/A	Comments
13	16	14	41	29	0	Statement I: Incorrect. $H_2O(g)$ is a greenhouse gas. Statement II: Incorrect. Fuel cells are more efficient than combustion engines hence less H_2 is used per kilometre for the H_2 fuel cell powered car. Statement III: Correct. In combustion engines chemical energy is converted to thermal then to mechanical energy. In fuel cell powered vehicles chemical energy is converted to electrical energy then to mechanical energy. Hence correct alternative was option C.

Question 63 (1 mark)

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

Hydrogen peroxide solutions are commercially available and have a range of uses. The active ingredient, hydrogen peroxide, H_2O_2, undergoes decomposition in the presence of a suitable catalyst according to the reaction:

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

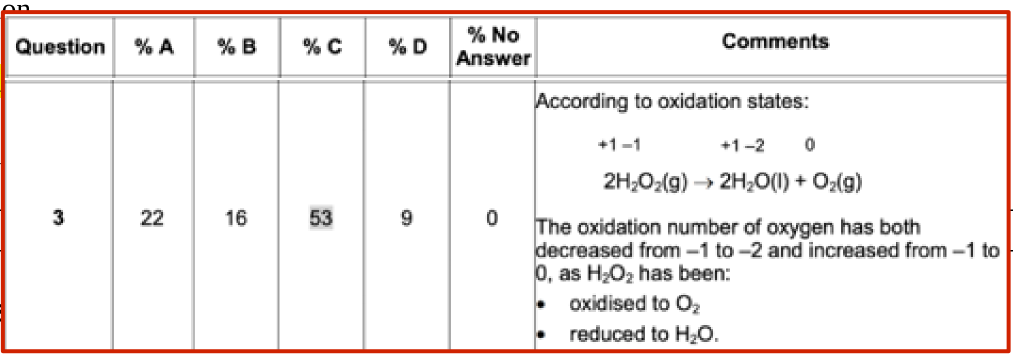
In this reaction, oxygen:

A. Only undergoes oxidation.

B. Only undergoes reductio

C. Undergoes both oxidati

D. Undergoes neither oxid



Question	% A	% B	% C	% D	% No Answer	Comments
3	22	16	53	9	0	According to oxidation states: $2\overset{+1\,-1}{H_2O_2}(g) \rightarrow 2\overset{+1\,-2}{H_2O}(l) + \overset{0}{O_2}(g)$ The oxidation number of oxygen has both decreased from –1 to –2 and increased from –1 to 0, as H_2O_2 has been: • oxidised to O_2 • reduced to H_2O.

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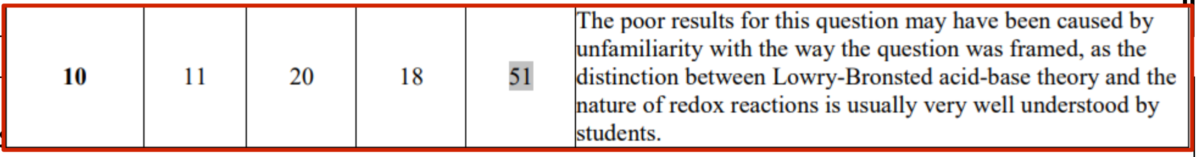
Question 64 (1 mark)

The following two unbalanced equations represent processes which are part of the nitrogen cycle.

I. $NH_3(aq) \rightarrow NH_4^+(aq)$
II. $NH_4^+(aq) \rightarrow NO_3^-(aq)$

Which one of the following alternatives correctly describes the reactants in each of these processes?

	In process I, $NH_3(aq)$ is	**In process II, the $NH_4^+(aq)$ ion is**
A.	an acid	reduced
B.	a base	reduced
C.	an acid	oxidised
D.	a base	oxidised



10	11	20	18	51	The poor results for this question may have been caused by unfamiliarity with the way the question was framed, as the distinction between Lowry-Bronsted acid-base theory and the nature of redox reactions is usually very well understood by students.

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Question 65 (1 mark)

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

Consider the half-cell equations and their half-cell potentials in the table below.

Half-Cell Equations	**Standard Electrode Potential ($E°$) in volts at 25°C**
$Mn^{3+}(aq) + e^- \rightleftharpoons Mn^{2+}(aq)$	$+1.56$
$SO_3^{2-}(aq) + 3H_2O(l) + 4e^- \rightleftharpoons S_2O_3^{2-}(aq) + 6OH^-(aq)$	-0.57
$SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$	-0.94
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	-1.18

A galvanic cell contains $S_2O_3^{2-}(aq)$, $SO_3^{2-}(aq)$ and $OH^-(aq)$ with a platinum electrode in the first half-cell and $Mn^{3+}(aq)$ and $Mn^{2+}(aq)$ in the second half-cell. All solutions are IM solutions and the cell is at SLC.

The galvanic cell will deliver:

A. 2.50 V if the electrode in the second half-cell is Mn(s).

B. 2.50 V if the electrode in the second half-cell is Pt(s).

C. 1.75 V if the electrode in the second half-cell is Mn(s).

D. 1.75 V if the electrode in the second half-cell is Pt(s).

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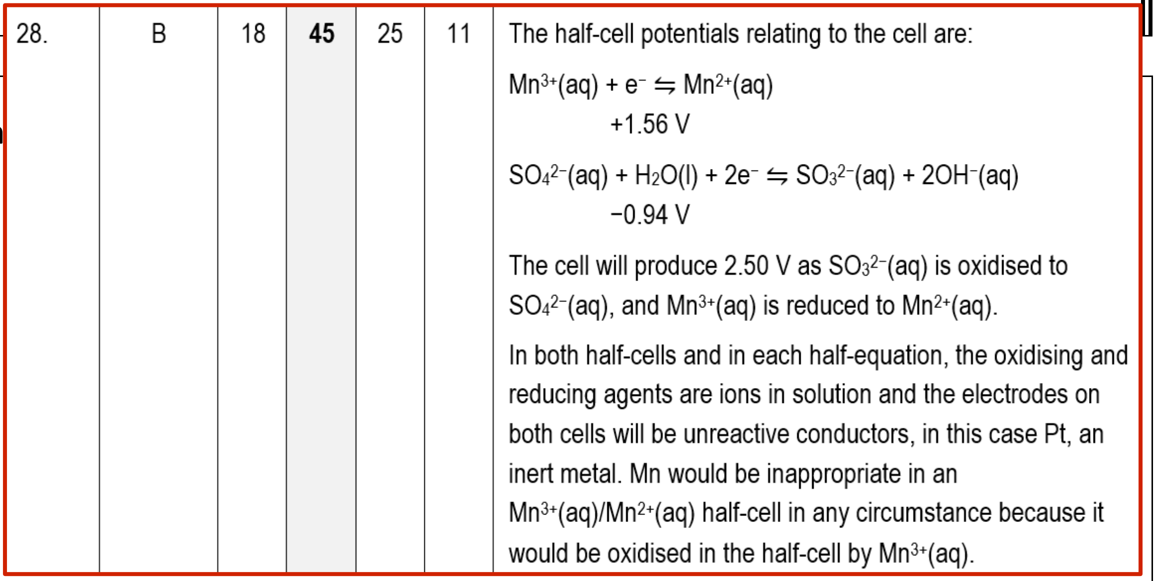


28.	B	18	**45**	25	11	The half-cell potentials relating to the cell are: $Mn^{3+}(aq) + e^- \rightleftharpoons Mn^{2+}(aq)$ +1.56 V $SO_4^{2-}(aq) + H_2O(l) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + 2OH^-(aq)$ −0.94 V The cell will produce 2.50 V as $SO_3^{2-}(aq)$ is oxidised to $SO_4^{2-}(aq)$, and $Mn^{3+}(aq)$ is reduced to $Mn^{2+}(aq)$. In both half-cells and in each half-equation, the oxidising and reducing agents are ions in solution and the electrodes on both cells will be unreactive conductors, in this case Pt, an inert metal. Mn would be inappropriate in an $Mn^{3+}(aq)/Mn^{2+}(aq)$ half-cell in any circumstance because it would be oxidised in the half-cell by $Mn^{3+}(aq)$.

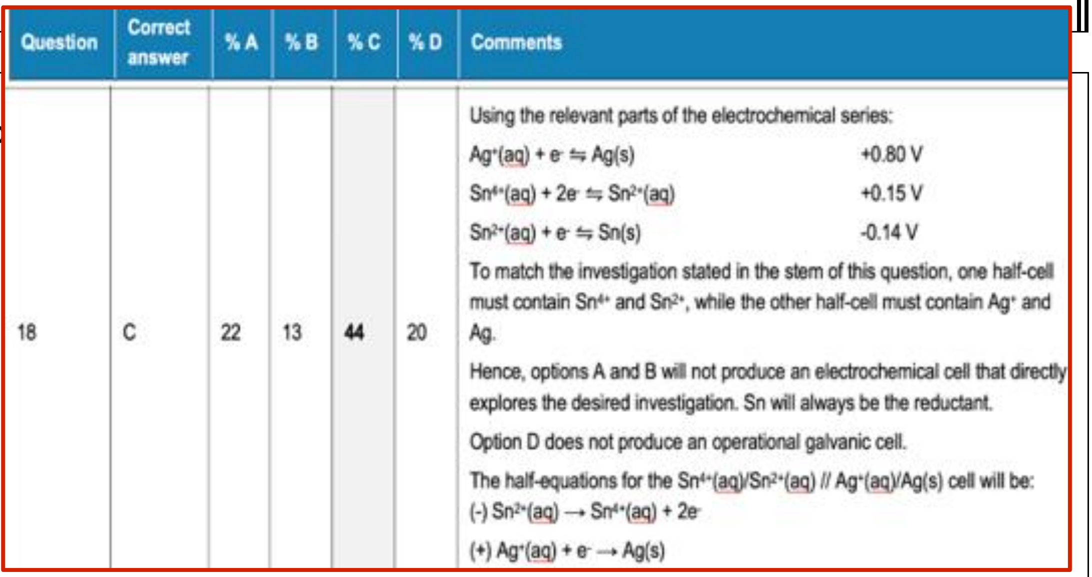
Question 66 (1 mark)

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

A student wants to investigate a galvanic cell consisting of Sn^{4+}/Sn^{2+} and Ag^{+}/Ag half-cells.

Which one of the following combinations of electrodes and solutions will produce an operational galvanic cell?

	Sn^{4+}/Sn^{2+} Half-Cell		**Ag^{+}/Ag Half-Cell**	
	Electrode	**Solution(s)**	**Electrode**	**Solution**
A.	Sn	1 M $Sn(NO_3)_2$	Graphite	1 M $AgNO_3$
B.	Sn	1 M $Sn(NO_3)_4$, 1 M $Sn(NO_3)_2$	Graphite	1 M $AgNO_3$
C.	Graphite	1 M $Sn(NO_3)_4$, 1 M $Sn(NO_3)_2$	Ag	1 M $AgNO_3$
D.	Graphite	1 M $Sn(NO_3)_4$	Ag	1 M $AgNO_3$



Question	Correct answer	% A	% B	% C	% D	Comments
18	C	22	13	**44**	20	Using the relevant parts of the electrochemical series: $Ag^{+}(aq) + e^{-} \leftrightharpoons Ag(s)$ +0.80 V $Sn^{4+}(aq) + 2e^{-} \leftrightharpoons Sn^{2+}(aq)$ +0.15 V $Sn^{2+}(aq) + e^{-} \leftrightharpoons Sn(s)$ -0.14 V To match the investigation stated in the stem of this question, one half-cell must contain Sn^{4+} and Sn^{2+}, while the other half-cell must contain Ag^{+} and Ag. Hence, options A and B will not produce an electrochemical cell that directly explores the desired investigation. Sn will always be the reductant. Option D does not produce an operational galvanic cell. The half-equations for the $Sn^{4+}(aq)/Sn^{2+}(aq)$ // $Ag^{+}(aq)/Ag(s)$ cell will be: (-) $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^{-}$ (+) $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

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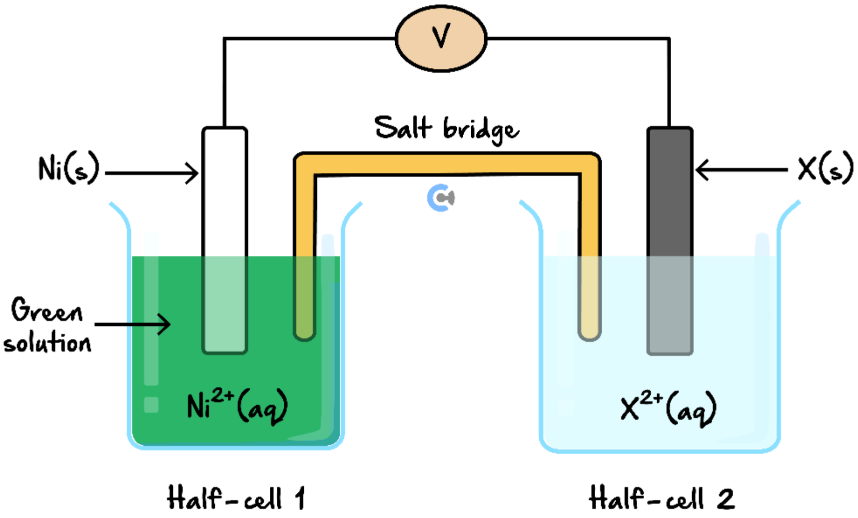
Question 67 (1 mark)

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

At the start of the day, a student set up a galvanic cell using two electrodes: nickel, Ni, and metal *X*. This setup is shown in the diagram below.



Consider the following alternative metals that could be used to replace metal *X*:

1. Zinc, Zn
2. Lead, Pb
3. Cadmium, Cd
4. Copper, Cu

At the end of the day, the student checked the colour of the solution in half-cell 1 and observed that the solution was a darker green colour.

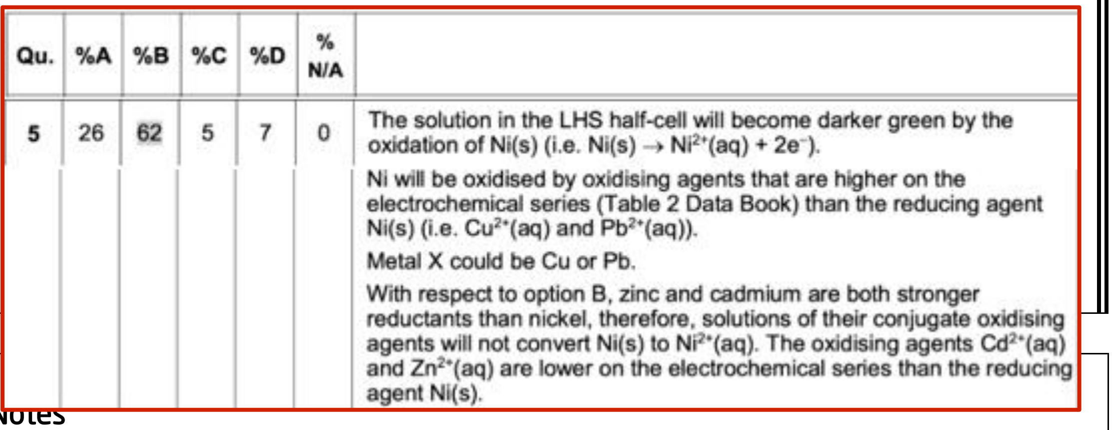
Which of the alternative metals could cause the colour of half-cell 1 to become a darker green?

A. Metals 1 and 3

B. Metals 2 and 4

C. Metals 1, 2 and 3

D. Metals 3 and 4



Qu.	%A	%B	%C	%D	% N/A	
5	26	62	5	7	0	The solution in the LHS half-cell will become darker green by the oxidation of Ni(s) (i.e. Ni(s) → $Ni^{2+}(aq) + 2e^-$). Ni will be oxidised by oxidising agents that are higher on the electrochemical series (Table 2 Data Book) than the reducing agent Ni(s) (i.e. $Cu^{2+}(aq)$ and $Pb^{2+}(aq)$). Metal X could be Cu or Pb. With respect to option B, zinc and cadmium are both stronger reductants than nickel, therefore, solutions of their conjugate oxidising agents will not convert Ni(s) to $Ni^{2+}(aq)$. The oxidising agents $Cd^{2+}(aq)$ and $Zn^{2+}(aq)$ are lower on the electrochemical series than the reducing agent Ni(s).

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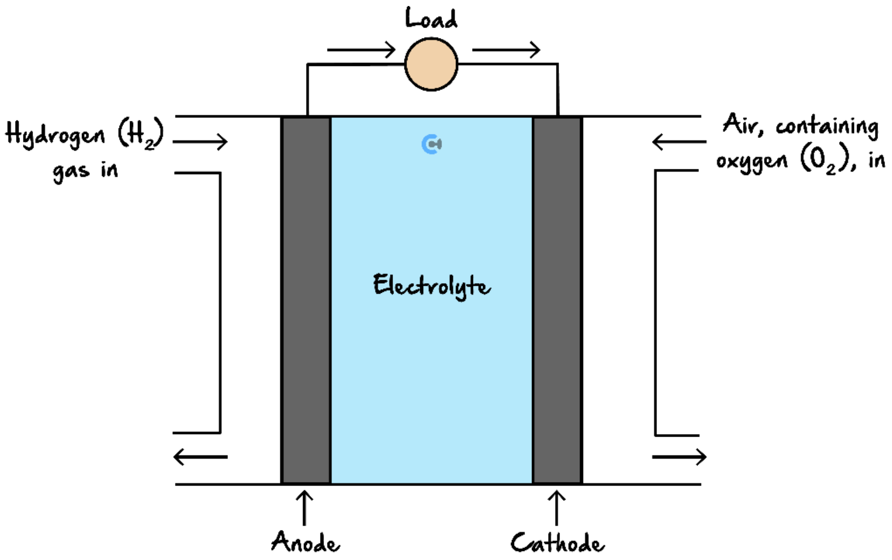
Question 68 (8 marks)

Inspired from VCAA Chemistry Exam 2017
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2017/2017chem-w.pdf#page=26

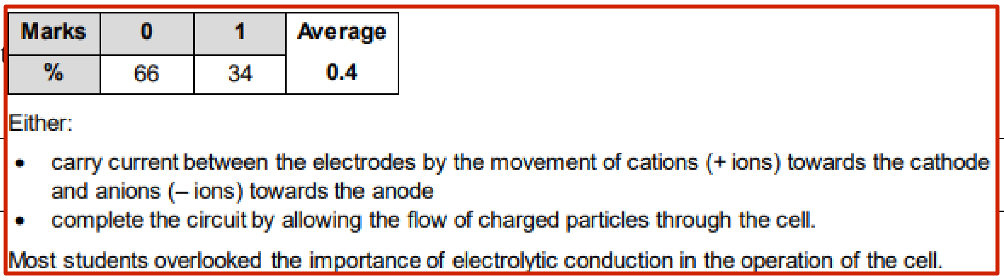
Submarines operate both on the surface and underwater. When operating underwater, the submarine acts as a closed system, where there is no interaction with the atmosphere. Most types of submarines use both batteries and diesel engines to provide their energy requirements. A new type of submarine uses proton exchange membrane (PEM) fuel cells and diesel engines.

Below is a diagram of a PEM fuel cell.



a.

i. State the funct



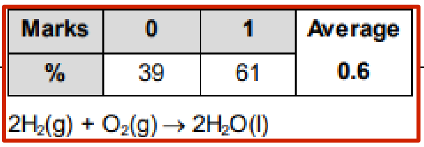
Marks	0	1	Average
%	66	34	0.4

Either:

- carry current between the electrodes by the movement of cations (+ ions) towards the cathode and anions (– ions) towards the anode
- complete the circuit by allowing the flow of charged particles through the cell.

Most students overlooked the importance of electrolytic conduction in the operation of the cell.

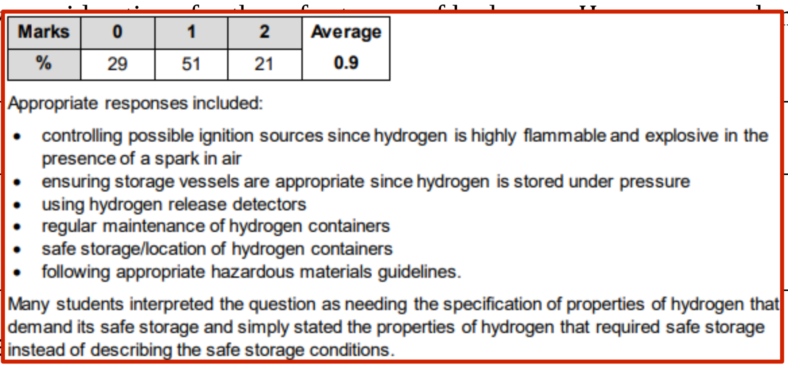
ii. Write the balanced overall redox reaction that occurs in this PEM fuel cell. (1 mark)



Marks	0	1	Average
%	39	61	0.6

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

iii. Give **two** safety arine. (2 marks)



Marks	0	1	2	Average
%	29	51	21	0.9

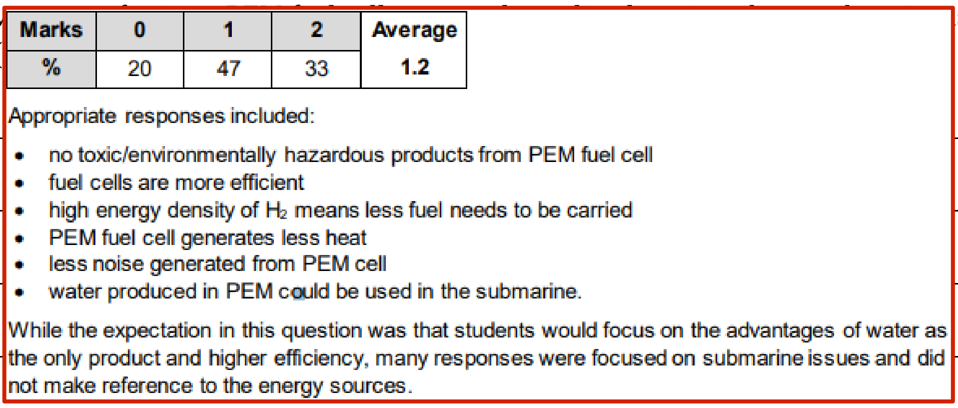
Appropriate responses included:

- controlling possible ignition sources since hydrogen is highly flammable and explosive in the presence of a spark in air
- ensuring storage vessels are appropriate since hydrogen is stored under pressure
- using hydrogen release detectors
- regular maintenance of hydrogen containers
- safe storage/location of hydrogen containers
- following appropriate hazardous materials guidelines.

Many students interpreted the question as needing the specification of properties of hydrogen that demand its safe storage and simply stated the properties of hydrogen that required safe storage instead of describing the safe storage conditions.

b.

i. State **two** adv s underwater. (



Marks	0	1	2	Average
%	20	47	33	1.2

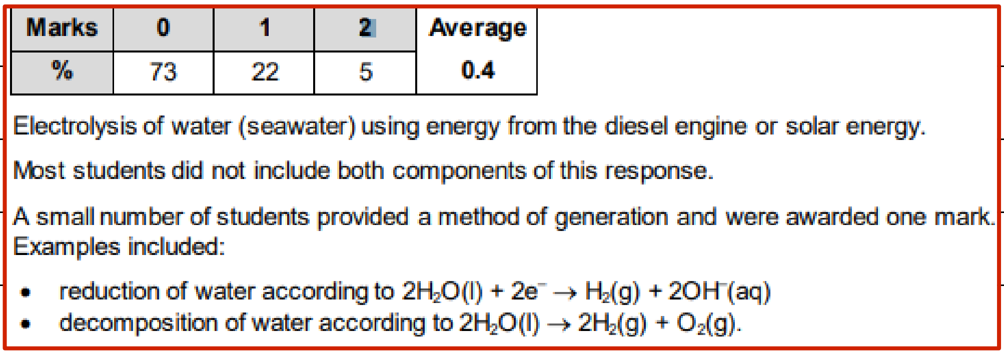
Appropriate responses included:

- no toxic/environmentally hazardous products from PEM fuel cell
- fuel cells are more efficient
- high energy density of H_2 means less fuel needs to be carried
- PEM fuel cell generates less heat
- less noise generated from PEM cell
- water produced in PEM could be used in the submarine.

While the expectation in this question was that students would focus on the advantages of water as the only product and higher efficiency, many responses were focused on submarine issues and did not make reference to the energy sources.

ii. Most submarines generate more H_2 gas for their fuel cells when travelling on the surface.

Explain how the H_2 gas could be generated. (2 marks)



Marks	0	1	2	Average
%	73	22	5	0.4

Electrolysis of water (seawater) using energy from the diesel engine or solar energy.

Most students did not include both components of this response.

A small number of students provided a method of generation and were awarded one mark. Examples included:

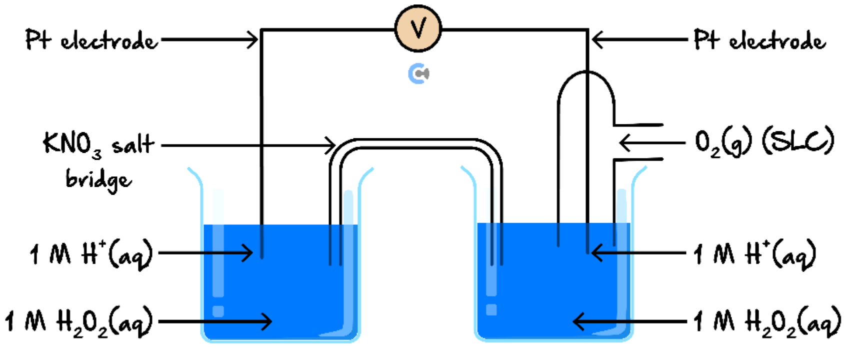
- reduction of water according to $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
- decomposition of water according to $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$.

Space for Personal Notes

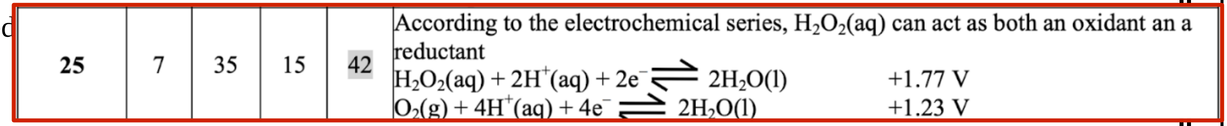
Question 69 (1 mark)

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

A student constructs the following galvanic cell.



The student pred



25	7	35	15	42	According to the electrochemical series, H_2O_2(aq) can act as both an oxidant an a reductant $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$ +1.77 V $O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$ +1.23 V

However, no reaction is observed.

This is most likely because:

A. The difference between the $E°$ values is too small for a reaction to occur.

B. Hydrogen peroxide will oxidise water in preference to itself.

C. The student did not construct standard half-cells.

D. The rate of the reaction is extremely slow.

Question 70 (1 mark)

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

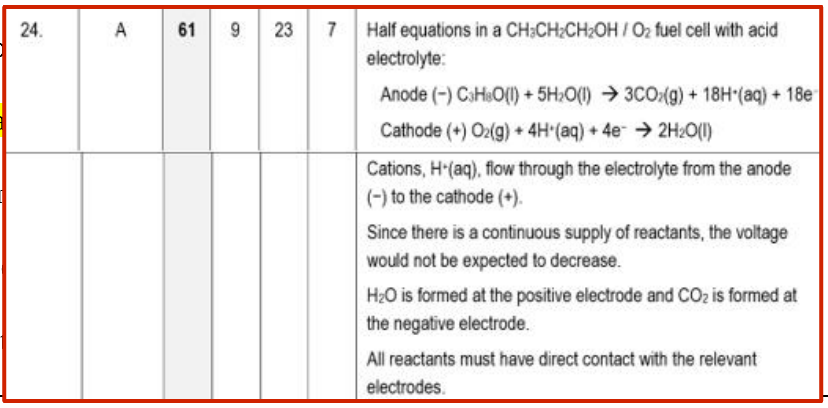
Which one of the following statements abo

A. Positive ions flow through the membra

B. The voltage decreases since the reactar

C. H_2O is formed at the negative electrod

D. Since it is a gas, O_2 is the only reactan



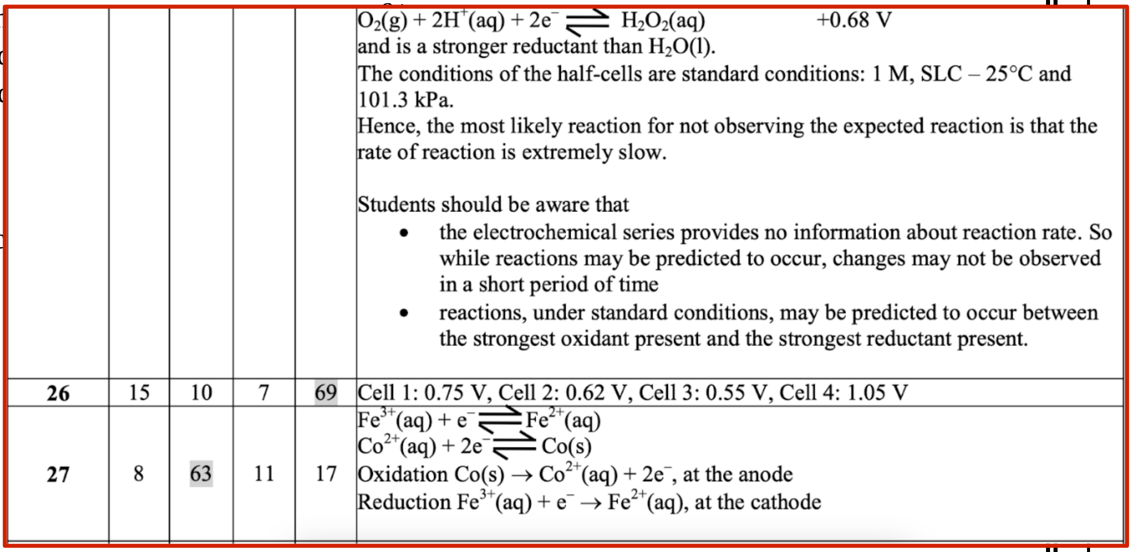
24.	A	**61**	9	23	7	Half equations in a $CH_3CH_2CH_2OH$ / O_2 fuel cell with acid electrolyte: Anode (−) $C_3H_8O(l) + 5H_2O(l) \rightarrow 3CO_2(g) + 18H^+(aq) + 18e^-$ Cathode (+) $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$
						Cations, H^+(aq), flow through the electrolyte from the anode (−) to the cathode (+). Since there is a continuous supply of reactants, the voltage would not be expected to decrease. H_2O is formed at the positive electrode and CO_2 is formed at the negative electrode. All reactants must have direct contact with the relevant electrodes.

Use the following information to answer the two questions that follow.

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

Four standard galvanic cells are set up as indicated below.

Cell I: A Br_2/Br^- standard half-cell connected to a Cu^{2+}/Cu standard half-cell
Cell II: An Sn^{2+}/Sn star
Cell III: A Br_2/Br^- stand
Cell IV: A Co^{2+}/Co stand



					$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$ +0.68 V and is a stronger reductant than $H_2O(l)$. The conditions of the half-cells are standard conditions: 1 M, SLC – 25°C and 101.3 kPa. Hence, the most likely reaction for not observing the expected reaction is that the rate of reaction is extremely slow. Students should be aware that • the electrochemical series provides no information about reaction rate. So while reactions may be predicted to occur, changes may not be observed in a short period of time • reactions, under standard conditions, may be predicted to occur between the strongest oxidant present and the strongest reductant present.
26	15	10	7	69	Cell 1: 0.75 V, Cell 2: 0.62 V, Cell 3: 0.55 V, Cell 4: 1.05 V
27	8	63	11	17	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$ $Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ Oxidation $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$, at the anode Reduction $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$, at the cathode

Question 71 (1 mark)

Which cell would be expec

A. I

B. II

C. III

D. IV

Question 72 (1 mark)

The reaction occurring at the cathode as cell IV is discharged is:

A. $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

B. $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

C. $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$

D. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

Space for Personal Notes

Question 73 (1 mark)

Inspired by VCAA Chemistry Exam 2011
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2011chem2-w.pdf

An ornament was coated with a metal, M, by electrolysis of a solution of the metal ion, M^{x+}. During the electrolysis, a current of 1.50 amperes was applied for 180 seconds. The ornament was coated in 0.0014 mol of metal.

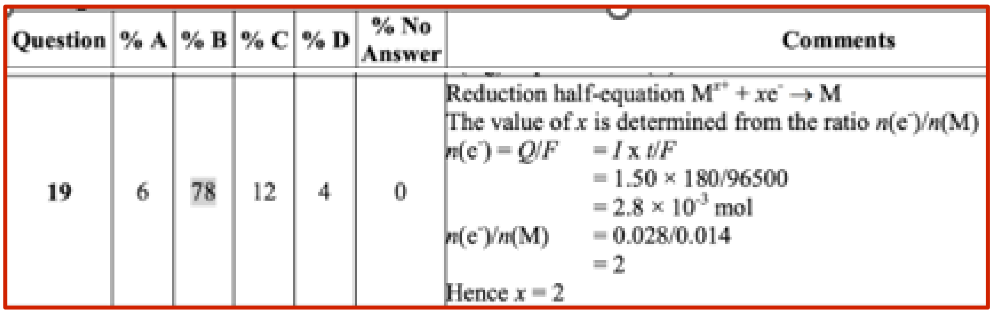
The value of x in M^{x+} is:

A. 1

B. 2

C. 3

D. 4



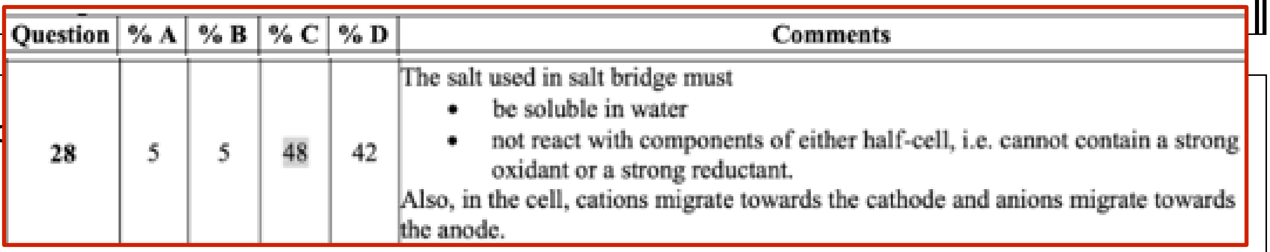
Question	% A	% B	% C	% D	% No Answer	Comments
19	6	78	12	4	0	Reduction half-equation $M^{x+} + xe^- \rightarrow M$ The value of x is determined from the ratio $n(e^-)/n(M)$ $n(e^-) = Q/F \quad = I \text{ x } t/F$ $= 1.50 \times 180/96500$ $= 2.8 \times 10^{-3}$ mol $n(e^-)/n(M) \quad = 0.028/0.014$ $= 2$ Hence $x = 2$

Question 74 (1 mark)

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

The main reason an aqueous solution of potassium nitrate, KNO_3, is used in salt bridges is:

A.	$K^+(aq)$ is a strong oxidant.	$NO_3^-(aq)$ is a weak reductant.
B.	$K^+(aq)$ is a weak reductant.	$NO_3^-(aq)$ is a strong oxidant.
C.	$K^+(aq)$ salts are soluble in water.	$NO_3^-(aq)$ salts are soluble in water.
D.	$K^+(aq)$ ions will migrate to the anode half-cell.	$NO_3^-(aq)$ ions will migrate to the cathode half-cell.



Question	% A	% B	% C	% D	Comments
28	5	5	48	42	The salt used in salt bridge must • be soluble in water • not react with components of either half-cell, i.e. cannot contain a strong oxidant or a strong reductant. Also, in the cell, cations migrate towards the cathode and anions migrate towards the anode.

Space fo

Question 75 (1 mark)

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

Consider two types of drones: petrol-powered and hydrogen-fuel-cell-powered.

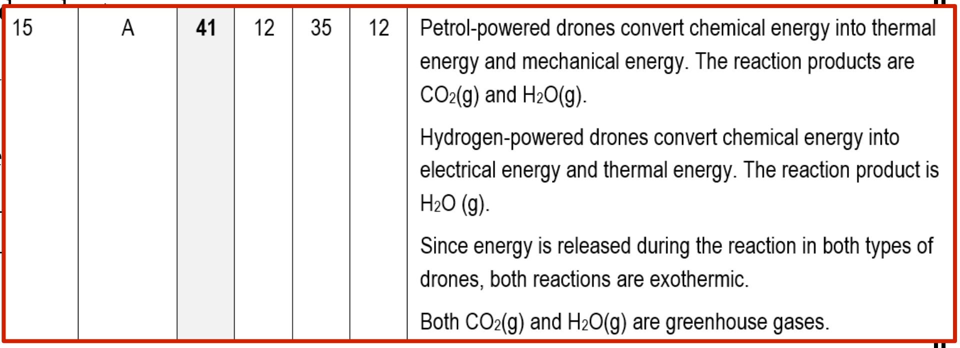
Which one of the following statements is correct about petrol-powered drones and hydrogen-fuel-cell-powered drones?

A. They both produce greenhouse gases.

B. Only petrol-powered drones prod

C. They both have the same energy

D. The overall reactions in hydroge



15	A	**41**	12	35	12	Petrol-powered drones convert chemical energy into thermal energy and mechanical energy. The reaction products are $CO_2(g)$ and $H_2O(g)$. Hydrogen-powered drones convert chemical energy into electrical energy and thermal energy. The reaction product is H_2O (g). Since energy is released during the reaction in both types of drones, both reactions are exothermic. Both $CO_2(g)$ and $H_2O(g)$ are greenhouse gases.

Question 76 (1 mark)

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

Corrosion of an iron pipe can be prevented by connecting it to a magnesium bar buried in the ground. The magnesium corrodes in preference to the iron.

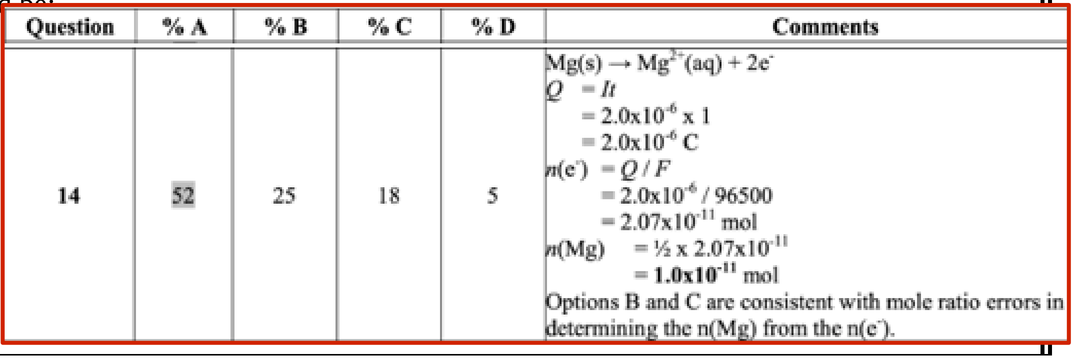
If the average current flowing between the two metals is $2.0 \times 10^{-6}\ A$, the amount of magnesium metal, in *mol*, reacting each second, would

A. 1.0×10^{-11}

B. 2.1×10^{-11}

C. 4.1×10^{-11}

D. 0.19



Question	% A	% B	% C	% D	Comments
14	52	25	18	5	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$ $Q = It$ $= 2.0 \times 10^{-6} \times 1$ $= 2.0 \times 10^{-6}$ C $n(e^-) = Q / F$ $= 2.0 \times 10^{-6} / 96500$ $= 2.07 \times 10^{-11}$ mol $n(Mg) = ½ \times 2.07 \times 10^{-11}$ $= \mathbf{1.0 \times 10^{-11}}$ mol Options B and C are consistent with mole ratio errors in determining the n(Mg) from the n(e⁻).

Question 77 (1 mark)

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

Methane, CH_4, and methan

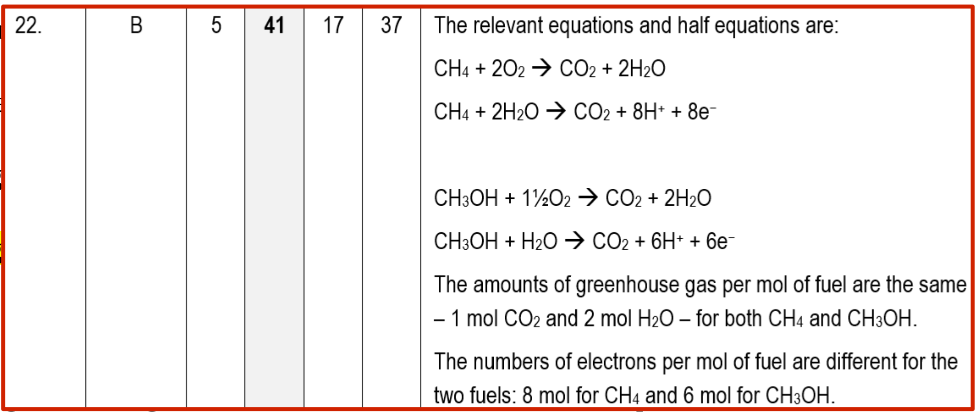
Methane and methanol fue

A. The same amount of g

B. The same amount of g

C. A different amount of

D. A different amount of



22.	B	5	**41**	17	37	The relevant equations and half equations are: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ $CH_4 + 2H_2O \rightarrow CO_2 + 8H^+ + 8e^-$ $CH_3OH + 1\frac{1}{2}O_2 \rightarrow CO_2 + 2H_2O$ $CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$ The amounts of greenhouse gas per mol of fuel are the same – 1 mol CO_2 and 2 mol H_2O – for both CH_4 and CH_3OH. The numbers of electrons per mol of fuel are different for the two fuels: 8 mol for CH_4 and 6 mol for CH_3OH.

Question 78 (1 mark)

Inspired by VCAA Chemistry Exam 2004
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem22004.pdf

A student is planning to set up a demonstration of a galvanic cell using half-cells constructed as follows.

- Half cell 1: a calcium electrode in a beaker containing an aqueous solution of Ca^{2+} ions
- Half cell 2: a platinum electrode in a beaker containing an aqueous solution of a mixture of Sn^{4+} and Sn^{2+} ions

A salt bridge would connect the two beakers. The electrodes would be attached to a voltmeter.

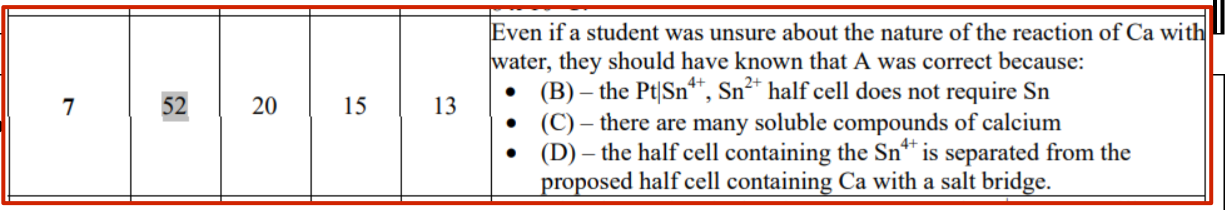
This particular cell is impractical because:

A. Solid calcium (Ca) will react directly to reduce water to hydrogen gas.

B. There is no solid tin (Sn) in the half-cell containing $Sn^{4+}(aq)$ and $Sn^{2+}(aq)$.

C. There are no known ionic compounds of calcium that are soluble in water.

D. $Sn^{4+}(aq)$ will be in contact with Ca and will oxidise it to $Ca^{2+}(aq)$.



7	52	20	15	13	Even if a student was unsure about the nature of the reaction of Ca with water, they should have known that A was correct because: • (B) – the Pt|Sn^{4+}, Sn^{2+} half cell does not require Sn • (C) – there are many soluble compounds of calcium • (D) – the half cell containing the Sn^{4+} is separated from the proposed half cell containing Ca with a salt bridge.

Space for P

Question 79 (10 marks)

Inspired from VCAA Chemistry Exam 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/2018chem-w.pdf#page=28

Redox reactions occur in the human body as well as in electrochemical cells.

a. Nicotinamide adenine dinucleotide (NAD) is a vital coenzyme for energy production in the human body. It exists in two forms: an oxidised form, NAD^+, and a reduced form, NADH.

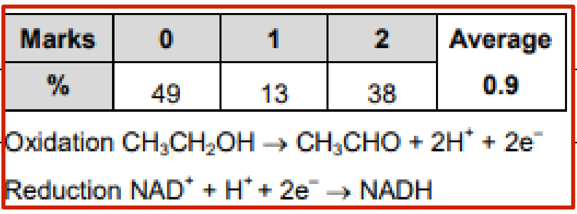
NAD is involved in the conversion of ethanol, CH_3CH_2OH, to ethanal, CH_3CHO, in the human body. The overall equation for this redox reaction is

$$CH_3CH_2OH + NAD^+ \rightarrow CH_3CHO + NADH + H^+$$

i. Write the two half-equations for this redox reaction. States are not required. (2 marks)

Oxidation half-equation ____________________

Reduction half-equation ____________________

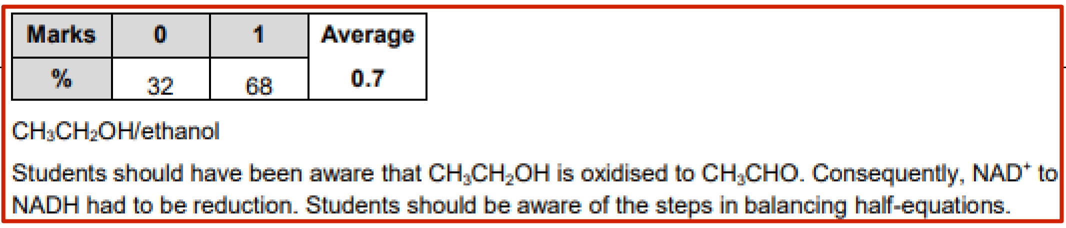


Marks	0	1	2	Average
%	49	13	38	0.9

Oxidation $CH_3CH_2OH \rightarrow CH_3CHO + 2H^+ + 2e^-$

Reduction $NAD^+ + H^+ + 2e^- \rightarrow NADH$

ii. Identify the reducing agent in this redox reaction. (1 mark)

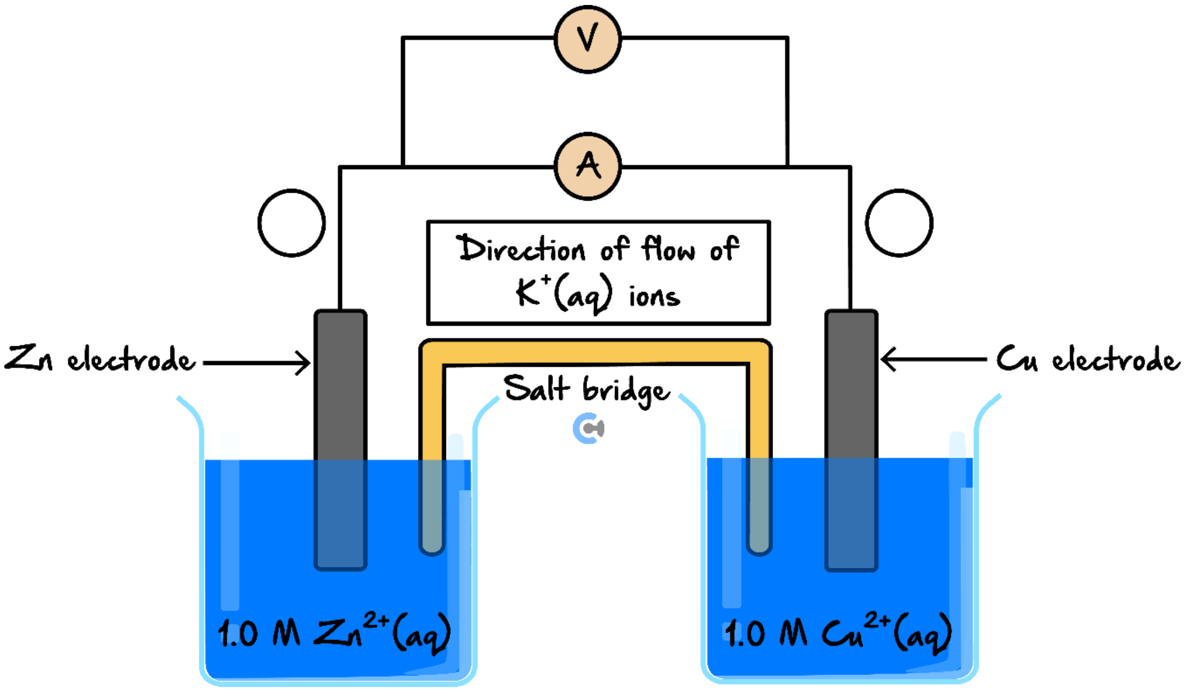


Marks	0	1	Average
%	32	68	0.7

CH_3CH_2OH/ethanol

Students should have been aware that CH_3CH_2OH is oxidised to CH_3CHO. Consequently, NAD^+ to NADH had to be reduction. Students should be aware of the steps in balancing half-equations.

b. The Daniell cell, a type of galvanic cell, was first constructed in the mid-1800s and this type of cell is still in use today. A diagram of the Daniell cell is shown below.



i. Label the polarity of the electrodes by placing a positive (+) o cles next to the electrodes on the diagram above. (1 mark)

Marks	0	1	Average
%	13	87	0.9

(+) Cu electrode, (–) Zn electrode

ii. Use the electrochemical series to determine the theoretical voltage of this cell. (1 mark)

Marks	0	1	Average
%	19	81	0.8

1.10 V

iii. The electrolyte in the salt bridge is (aq).

In the box above the salt bridge, u of flow of K^+(aq) ions. (1 mark)

Marks	0	1	Average
%	27	73	0.8

Left to right

iv. List **two** visible changes that are likely to be observed when the Daniell cell has been operating for some time. (2 marks)

Marks	0	1	2	Average
%	24	27	49	1.3

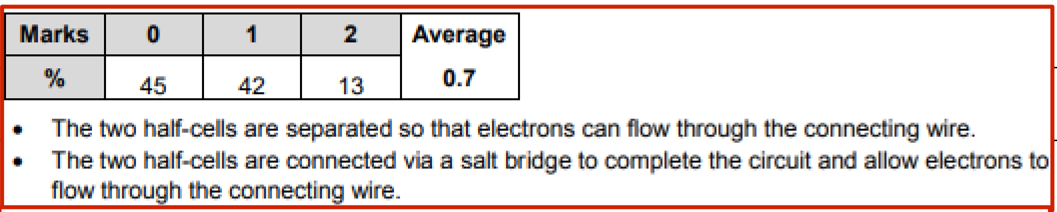
Acceptable responses included two of:

- Zn electrode decreases in size
- Cu electrode increases in size (is coated)
- Cu^{2+} solution becomes less blue (paler blue)/fades
- voltage or current drops.

Mass changes were a relatively common response but were not considered to be 'visible'.

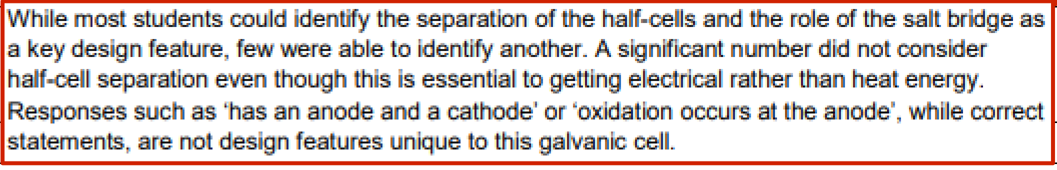
Some responses indicated issues with effective use of the electrochemical series in identifying what was happening in the half-cells.

c. What design features of the Daniell cell structure would allow it to produce electrical energy? (2 marks)



Marks	0	1	2	Average
%	45	42	13	0.7

- The two half-cells are separated so that electrons can flow through the connecting wire.
- The two half-cells are connected via a salt bridge to complete the circuit and allow electrons to flow through the connecting wire.



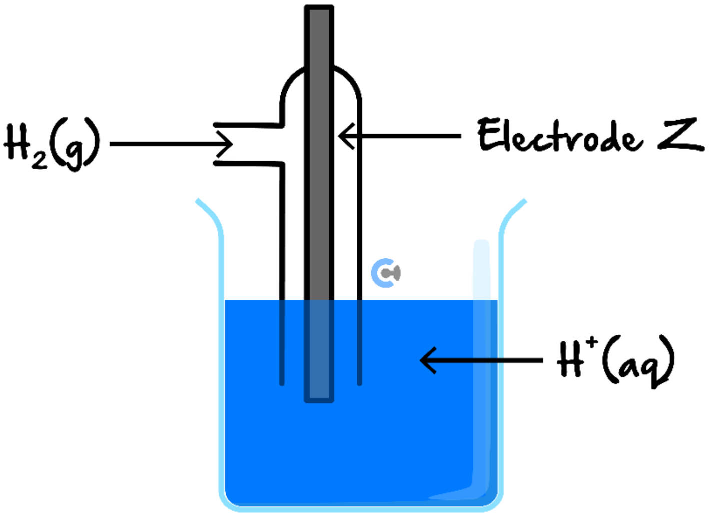
While most students could identify the separation of the half-cells and the role of the salt bridge as a key design feature, few were able to identify another. A significant number did not consider half-cell separation even though this is essential to getting electrical rather than heat energy. Responses such as 'has an anode and a cathode' or 'oxidation occurs at the anode', while correct statements, are not design features unique to this galvanic cell.

Question 80 (11 marks)

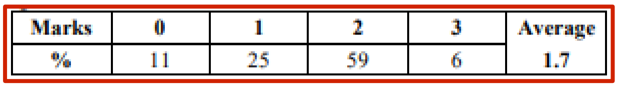
Inspired from VCAA Chemistry Exam 2007
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2007chem2.pdf#page=18

The following diagram represents a $H^+(aq)/H_2(g)$ half cell for the reaction:

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



a.



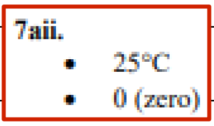
Marks	0	1	2	3	Average
%	11	25	59	6	1.7

i. For this half cell, identify an appropriate material for electrode Z. (1 mark)

7ai.
Pt, graphite, Pd or Au

ii. For this half cell to be a **standard** half cell, state: (2 marks)

- The temperature at which it must operate ____
- The required **pH** of the solution of $H^+(aq)$ ions ____



7aii.
- 25°C
- 0 (zero)

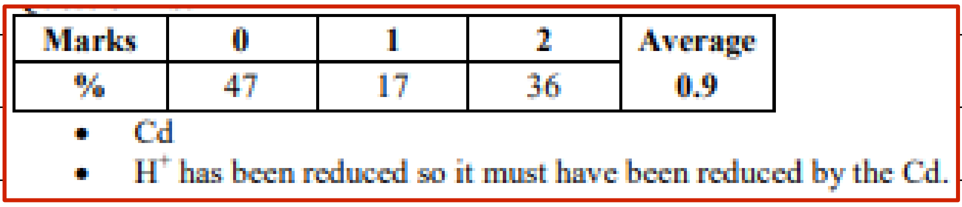
b. A galvanic cell consists of the following half cells which have been set up under standard conditions.

- Half cell 1: the $H^+(aq)/H_2(g)$ half cell described in **part a.**
- Half cell 2: an iron (Fe) electrode in a solution containing $Fe^{2+}(aq)$.

After some time, the pH in half cell 1 has increased. Use this information to identify the species in this galvanic cell which is the stronger reductant and explain how you reached this conclusion. (2 marks)

The stronger reductant is ________________________________

Explanation ________________________________



Marks	0	1	2	Average
%	47	17	36	0.9

- Cd
- H^+ has been reduced so it must have been reduced by the Cd.

The stronger reductant is Fe(s)

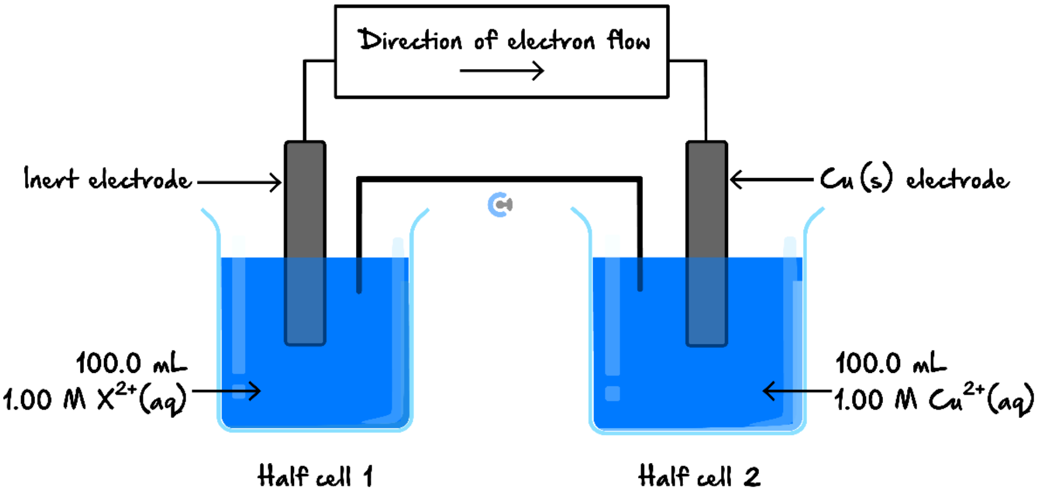
Explanation As cell #1 is increasing in pH, this indicates $H^+(aq)$ is being used up. Thus, $H^+(aq)$ is reducing according to $2H^+(aq) + 2e^- \rightarrow H_2(g)$. Therefore, Fe(s) is the stronger reductant since it causes $H^+(aq)$ to be reduced.

① : $H^+(aq)$ has been reduced so it must have been reduced by Cd(s)

c. A second galvanic cell consists of the following half-cells.

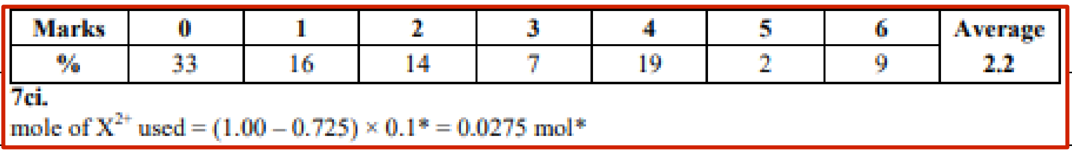
- Half cell 1: an inert electrode in 100.0 mL solution of 1.00 M X^{2+}(aq).
- Half cell 2: an electrode of Cu(s) in 100.0 mL solution of 1.00 M Cu^{2+}(aq).

This galvanic cell is shown in the diagram below.



After discharging 2654 C of electricity, the concentration of the X^{2+}(aq) in solution in half cell 1 was found to be 0.725 M. The volume of the solutions in the two half cells had not changed.

i. Calculate the amount, in mol, of X^{2+}(aq) that reacted in half cell 1. (2 marks)

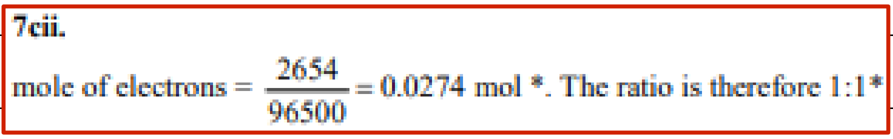


Marks	0	1	2	3	4	5	6	Average
%	33	16	14	7	19	2	9	2.2

7ci.

mole of X^{2+} used = (1.00 – 0.725) × 0.1* = 0.0275 mol*

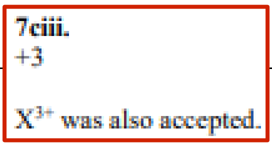
ii. Calculate the ratio of $n(X^{2+})$ reacted to $n(e^-)$ that passed through the cell. That is, calculate: $n(X^{2+})_{reacted}: n(e^-)$. (2 marks)



7cii.

mole of electrons = $\frac{2654}{96500}$ = 0.0274 mol *. The ratio is therefore 1:1*

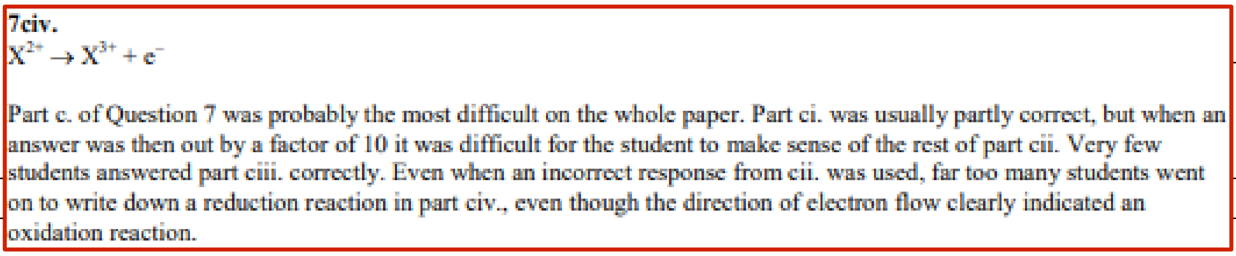
iii. State the oxidation state of the product of the half-reaction in half cell 1. (1 mark)



7ciii.
+3

X^{3+} was also accepted.

iv. Write an equation for the half-reaction that occurred at the electrode of half cell 1. (1 mark)



7civ.
$X^{2+} \rightarrow X^{3+} + e^-$

Part c. of Question 7 was probably the most difficult on the whole paper. Part ci. was usually partly correct, but when an answer was then out by a factor of 10 it was difficult for the student to make sense of the rest of part cii. Very few students answered part ciii. correctly. Even when an incorrect response from cii. was used, far too many students went on to write down a reduction reaction in part civ., even though the direction of electron flow clearly indicated an oxidation reaction.

Space for Personal Notes

Question 81 (4 marks)

Inspired from VCAA Chemistry Exam 2010
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2010chem2-w.pdf

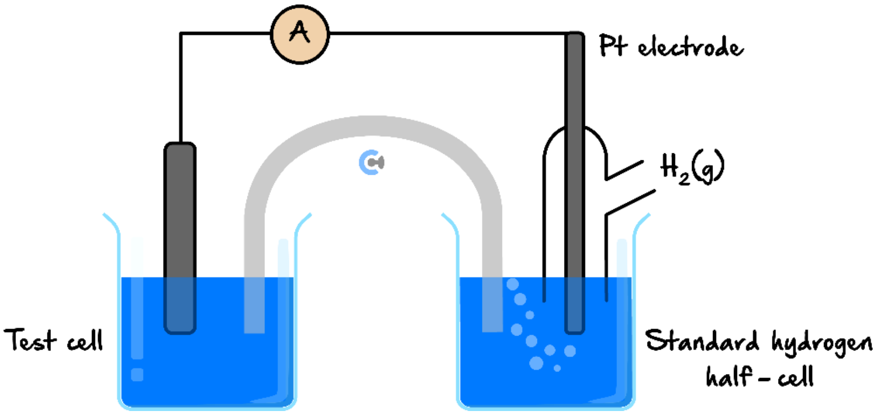
In a problem-solving activity, a student is given the following information regarding three half-equations.

However, although the three numerical values of $E°$ are correct, they have been incorrectly assigned to the three half-equations.

Half-Equation	$E°$
$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	$-0.40\ V$
$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	$-0.36\ V$
$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	$+0.22\ V$

The objective of this task is to correctly assign the $E°$ values to the corresponding half-equation.

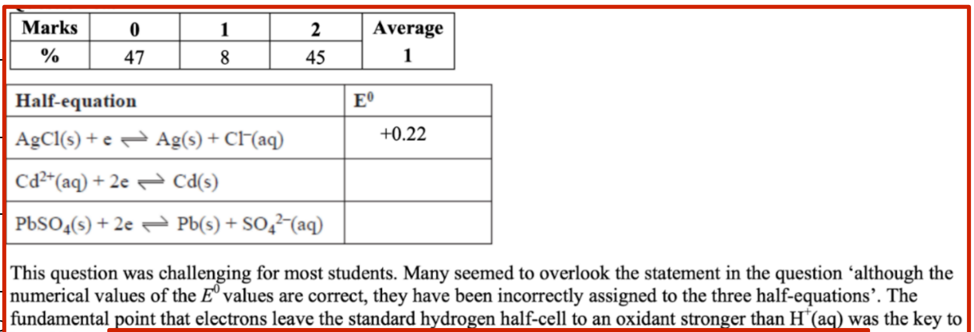
To do this, the student constructs standard half-cells for each of the above half-reactions. These half-cells are connected, one at a time, to a standard hydrogen half-cell as indicated in the diagram below.



The following observations were made either during or after the electrochemical cell discharged electricity for several minutes.

Experiment	**Half-Cell Reaction Being Investigated**	**Experimental Notes**
1	$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	Electron flow was detected passing from the standard hydrogen half-cell to the half-cell containing the silver electrode.
2	$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	The mass of the cadmium electrode decreased.
3	$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	The pH of the solution in the standard hydrogen half-cell increased.

a. The above information can only be used to assign **one** of the *E*° values to its corresponding half-equation. Identify this half-equation by placing the correct *E*° value next to its corresponding half-equation in the table below. (2 marks)



Marks	0	1	2	Average
%	47	8	45	1

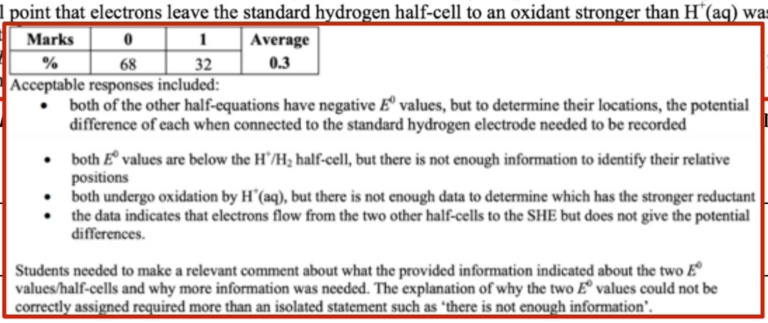
Half-equation	E^0
$AgCl(s) + e \rightleftharpoons Ag(s) + Cl^-(aq)$	+0.22
$Cd^{2+}(aq) + 2e \rightleftharpoons Cd(s)$	
$PbSO_4(s) + 2e \rightleftharpoons Pb(s) + SO_4^{2-}(aq)$	

This question was challenging for most students. Many seemed to overlook the statement in the question 'although the numerical values of the E^0 values are correct, they have been incorrectly assigned to the three half-equations'. The fundamental point that electrons leave the standard hydrogen half-cell to an oxidant stronger than $H^+(aq)$ was the key to identifying t

Only 'one' *E* positive E^0

value, but in

F



Marks	0	1	Average
%	68	32	0.3

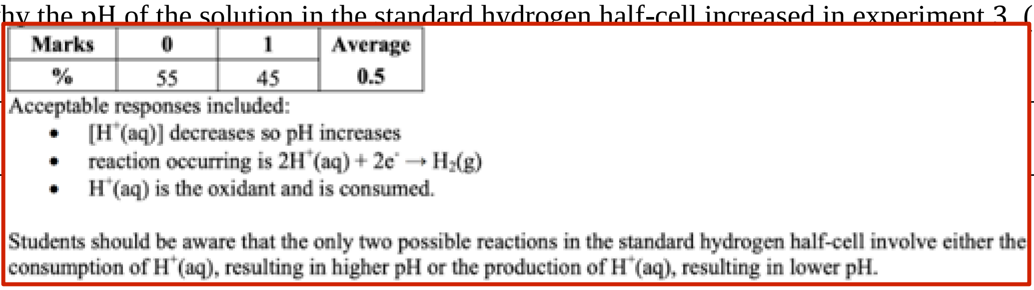
Acceptable responses included:

- both of the other half-equations have negative E^0 values, but to determine their locations, the potential difference of each when connected to the standard hydrogen electrode needed to be recorded
- both E^0 values are below the H^+/H_2 half-cell, but there is not enough information to identify their relative positions
- both undergo oxidation by $H^+(aq)$, but there is not enough data to determine which has the stronger reductant
- the data indicates that electrons flow from the two other half-cells to the SHE but does not give the potential differences.

Students needed to make a relevant comment about what the provided information indicated about the two E^0 values/half-cells and why more information was needed. The explanation of why the two E^0 values could not be correctly assigned required more than an isolated statement such as 'there is not enough information'.

b. Explain why the other two *E* mark)

c. Explain why the pH of the solution in the standard hydrogen half-cell increased in experiment 3. (1 mark)



Marks	0	1	Average
%	55	45	0.5

Acceptable responses included:

- $[H^+(aq)]$ decreases so pH increases
- reaction occurring is $2H^+(aq) + 2e^- \rightarrow H_2(g)$
- $H^+(aq)$ is the oxidant and is consumed.

Students should be aware that the only two possible reactions in the standard hydrogen half-cell involve either the consumption of $H^+(aq)$, resulting in higher pH or the production of $H^+(aq)$, resulting in lower pH.

Space for Personal Notes

Question 82 (10 marks)

Inspired from VCAA Chemistry Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/NHT/2023chem-nht-w.pdf#page=20

A galvanic cell can be constructed from a $Ni^{2+}(aq)/ Ni(s)$ half-cell and an $Ag^{+}(aq)/Ag(s)$ half-cell.

a. State the maximum cell voltage, under standard conditions, of the $Ag^{+}(aq)/Ag(s)//Ni^{2+}(aq)/ Ni(s)$ galvanic cell. (1 mark)

1.05 V [0.80 – (-0.25)]

b. Write the overall equation for the $Ag^{+}(aq)/Ag(s)//Ni^{2+}(aq)/ Ni(s)$ galvanic cell. (1 mark)

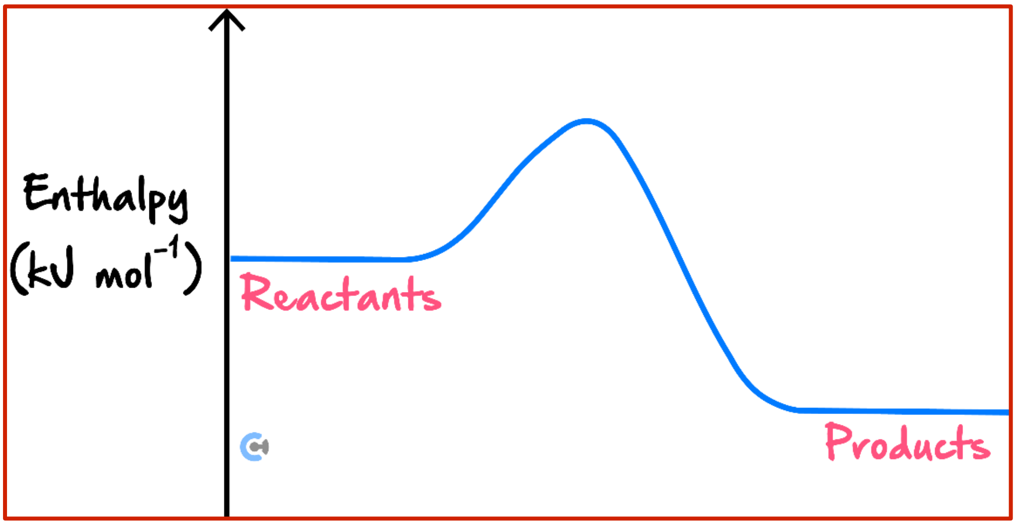
$2Ag^{+}(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$

c. Microbial fuel cells, MFCs, use bacteria to metabolise biomass. In an MFC, the bacteria decompose biomass, in the absence of oxygen, through a process called anaerobic cellular respiration.

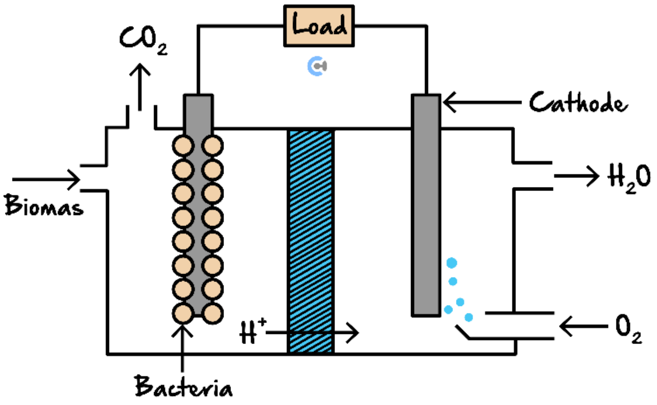
i. Write the equation for respiration that occurs in the cells of the human body. (1 mark)

$C_6H_{12}O_6(aq) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

ii. On the axes below, sketch the energy profile diagram for respiration. (1 mark)



The MFC anode has bacteria growing on it. The bacteria decompose organic matter and produce H^+ ions and electrons. The bacteria transfer electrons directly to the anode. One half-cell has oxygen bubbled into it. In the other half-cell, oxygen is excluded. A diagram of an MFC is given below.



d. Write the balanced half-equation for the reaction occurring at the cathode. (1 mark)

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

e. Explain why the two half-cells are separated. (2 marks)

One mark was awarded for indication of the need to produce/harness electrical energy.

One mark was awarded for explanation of how the electrical energy is harnessed *or* the need to keep O_2 away from the oxidation half-cell / anaerobic bacteria.

Acceptable responses included the following:

The half-cells need to be separated so that electrons can travel between the two half-cells in connecting wires (through the load) and the electrical energy harnessed.

The bacteria are anaerobic, and the oxygen needs to be excluded from the oxidation half-cell; if the half-cells are not separated, electricity will not be produced.

If there was no separation of the two half-cells, oxidation and reduction would occur together and electrons would not move through the external circuit.

f. For the cell shown in the diagram above, state the feature that causes the overall movement of H^+ ions across the membrane. Explain your answer. (2 marks)

One mark was awarded for indicating the feature as an 'electrode – anode / cathode, or electrodes'.

One mark was awarded for linking the feature to movement of H^+ ions.

For example:

As the electrons enter the cathode, the O_2 half-cell becomes negatively charged. The mobile H^+ ions are attracted to the cathode from the biomass half-cell and pass through the membrane.

As electrons leave the anode, the biomass half-cell becomes positively charged. The mobile H^+ ions are repelled from the anode and biomass half-cell and pass through the membrane.

At the cathode the $[H^+]$ decreases and at the anode the $[H^+]$ increases. The H^+ ions move from the higher concentration through the membrane to the lower concentration.

g. State an advantage of an MFC over the $Ag^+(aq)/Ag(s)//Ni^{2+}(aq)/Ni(s)$ galvanic cell. (1 mark)

Possible responses included the following:

As long as oxygen and biomass are supplied, the MFC will keep working, whereas the galvanic cell will run out of reactants.

The MFC produces energy from a renewable resource, whereas the galvanic cell does not.

The reagents (biomass, oxygen) for a MFC are cheaper and more sustainable than the reagents (silver in the form of Ag^+, nickel) for an $Ag^+(aq)/Ag(s)/ /Ni^{2+}(aq)/ Ni(s)$ galvanic cell.

The operation of a MFC decomposes organic matter, which is a way of processing/recycling waste, and is more sustainable than the operation of an $Ag^+(aq)/Ag(s)/ /Ni^{2+}(aq)/ Ni(s)$ galvanic cell.

Question 83 (7 marks)

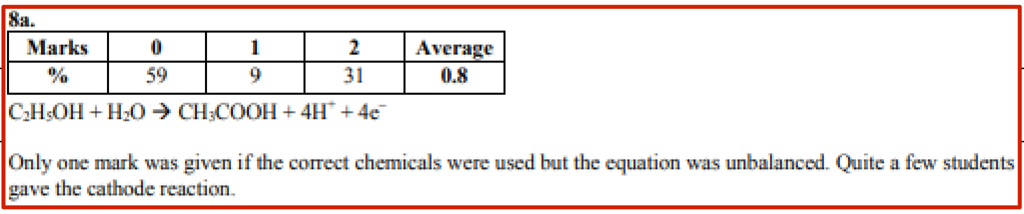
Inspired from VCAA Chemistry Exam 2005
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2005chem2.pdf#page=18

One type of 'breathalyser' instrument used by police for the measurement of the concentration of alcohol in a driver's breath is a fuel cell. An acidic electrolyte is used. Ethanol is oxidised to ethanoic acid at one electrode and oxygen from the air is converted to water at the other.

The overall equation for this reaction is:

$$\mathrm{C_2H_2OH(aq) + O_2(g) \rightarrow CH_3COOH(aq) + H_2O(l)}$$

a. Write the equation for the half-reaction at the anode. (2 marks)



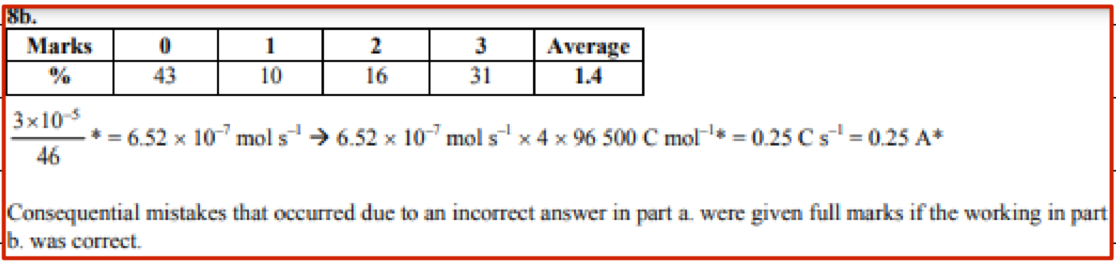
8a.

Marks	0	1	2	Average
%	59	9	31	0.8

$C_2H_5OH + H_2O \rightarrow CH_3COOH + 4H^+ + 4e^-$

Only one mark was given if the correct chemicals were used but the equation was unbalanced. Quite a few students gave the cathode reaction.

b. A motorist who has consumed alcohol blows into the fuel cell. If the breath entering the cell provides alcohol at the rate of 3.0×10^{-5} g per second, calculate the maximum current, in amps, that the cell would produce. (3 marks)



8b.

Marks	0	1	2	3	Average
%	43	10	16	31	1.4

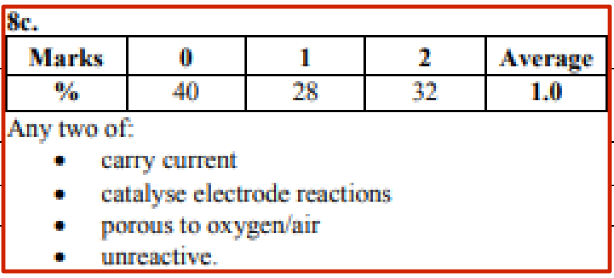
$\frac{3\times10^{-5}}{46}$* $= 6.52 \times 10^{-7}$ mol s^{-1} → 6.52×10^{-7} mol s^{-1} × 4 × 96 500 C mol^{-1}* = 0.25 C s^{-1} = 0.25 A*

Consequential mistakes that occurred due to an incorrect answer in part a. were given full marks if the working in part b. was correct.

c. The nature of the electrodes in the cell is essential to the effective operation of the breathalyser. State **two** important functions that the electrodes must perform. (2 marks)

Function 1

Function 2



8c.

Marks	0	1	2	Average
%	40	28	32	1.0

Any two of:

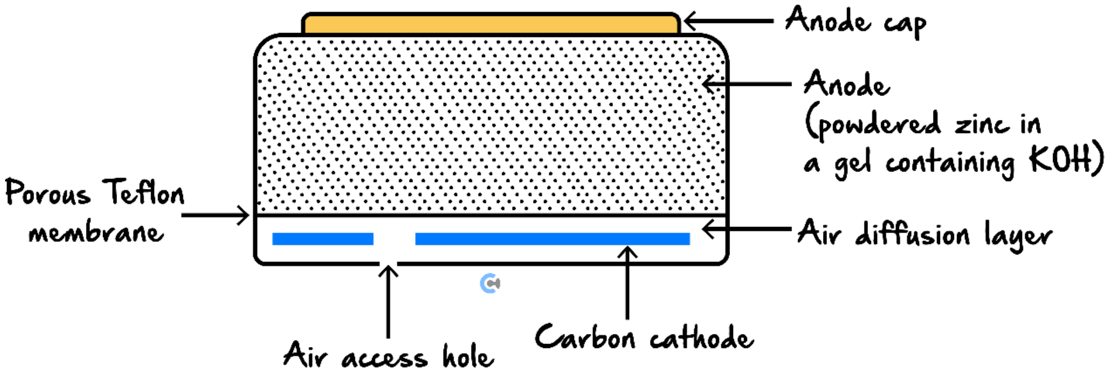
- carry current
- catalyse electrode reactions
- porous to oxygen/air
- unreactive.

Space for Personal Notes

Question 84 (6 marks)

Inspired from VCAA Chemistry Exam 2014
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2014/2014chem-amd-w.pdf#page=38

The following diagram shows a cross-section of a small zinc-air button cell, a button cell that is used in hearing aids.

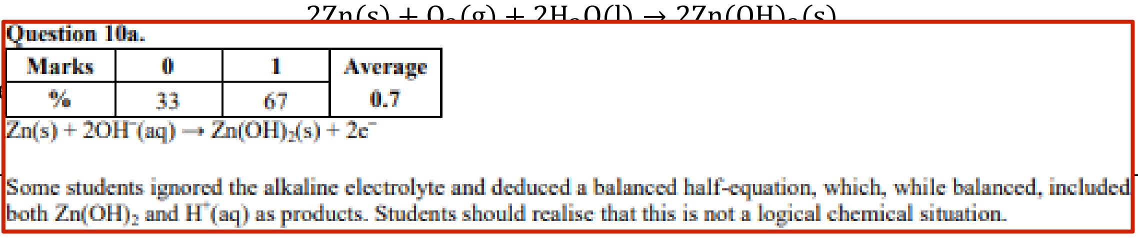


The zinc acts as the anode. It is in the form of a powder dispersed in a gel (a jelly-like substance) that also contains potassium hydroxide. The cathode consists of a carbon disc. Oxygen enters the cell via a porous Teflon membrane. This membrane also prevents any chemicals from leaking out.

The following reaction takes place as the cell discharges.

$$2Zn(s) + O_2(g) + 2H_2O(l) \rightarrow 2Zn(OH)_2(s)$$

a. Writ



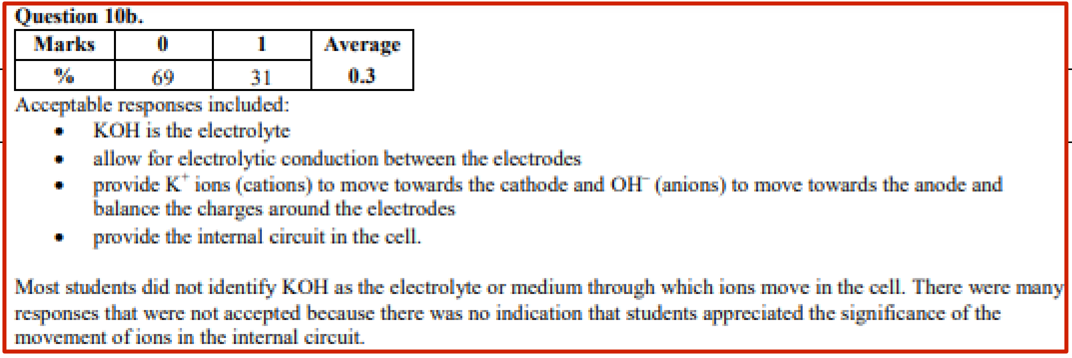
Question 10a.

Marks	0	1	Average
%	33	67	0.7

$Zn(s) + 2OH^-(aq) \rightarrow Zn(OH)_2(s) + 2e^-$

Some students ignored the alkaline electrolyte and deduced a balanced half-equation, which, while balanced, included both $Zn(OH)_2$ and $H^+(aq)$ as products. Students should realise that this is not a logical chemical situation.

b. Suggest **one** role of potassium hydroxide in this cell. (1 mark)



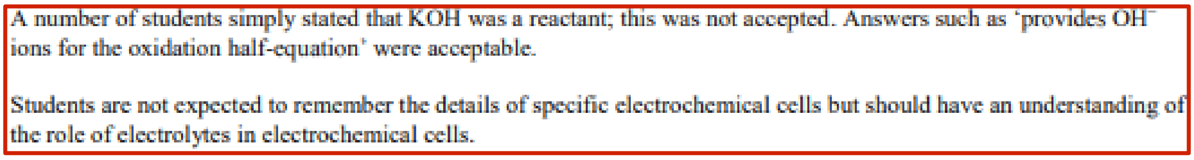
Question 10b.

Marks	0	1	Average
%	69	31	0.3

Acceptable responses included:

- KOH is the electrolyte
- allow for electrolytic conduction between the electrodes
- provide K^+ ions (cations) to move towards the cathode and OH^- (anions) to move towards the anode and balance the charges around the electrodes
- provide the internal circuit in the cell.

Most students did not identify KOH as the electrolyte or medium through which ions move in the cell. There were many responses that were not accepted because there was no indication that students appreciated the significance of the movement of ions in the internal circuit.

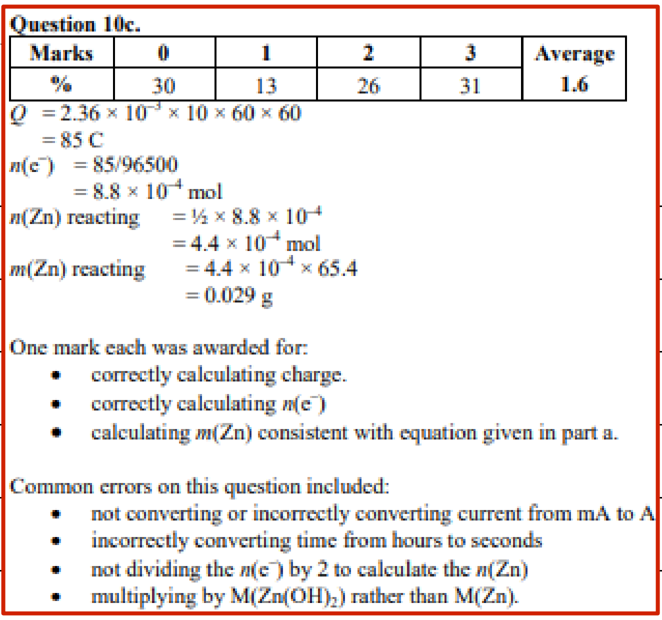


A number of students simply stated that KOH was a reactant; this was not accepted. Answers such as 'provides OH^- ions for the oxidation half-equation' were acceptable.

Students are not expected to remember the details of specific electrochemical cells but should have an understanding of the role of electrolytes in electrochemical cells.

c. A zinc-air button cell is r

What mass of zinc metal



Question 10c.

Marks	0	1	2	3	Average
%	30	13	26	31	1.6

$Q = 2.36 \times 10^{-3} \times 10 \times 60 \times 60$

$= 85$ C

$n(e^-) = 85/96500$

$= 8.8 \times 10^{-4}$ mol

n(Zn) reacting $= ½ \times 8.8 \times 10^{-4}$

$= 4.4 \times 10^{-4}$ mol

m(Zn) reacting $= 4.4 \times 10^{-4} \times 65.4$

$= 0.029$ g

One mark each was awarded for:

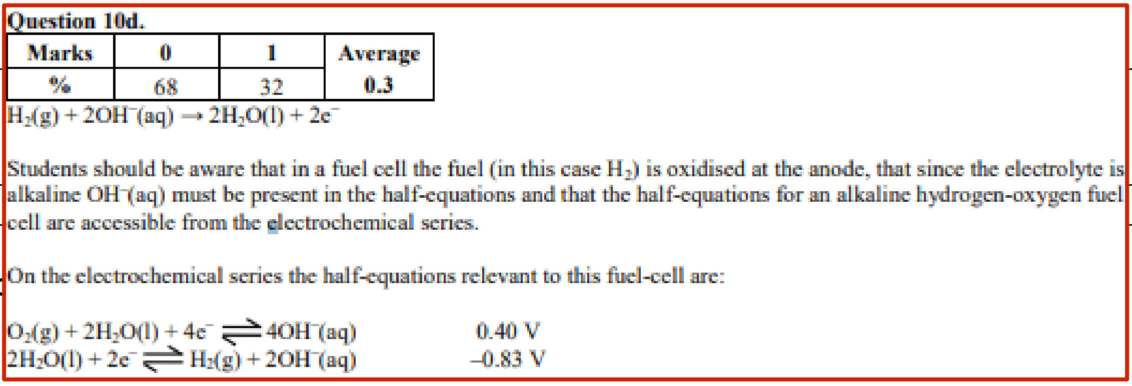
- correctly calculating charge.
- correctly calculating $n(e^-)$
- calculating m(Zn) consistent with equation given in part a.

Common errors on this question included:

- not converting or incorrectly converting current from mA to A
- incorrectly converting time from hours to seconds
- not dividing the $n(e^-)$ by 2 to calculate the n(Zn)
- multiplying by M(Zn(OH)$_2$) rather than M(Zn).

d. A hydrogen-oxygen fuel cell can operate with an alkaline electrolyte such as potassium hydroxide. In this cell, the reaction at the cathode is the same as that in the zinc-air cell. A porous carbon cathode is used.

Write the half-equation for the reaction that occurs at the anode in a hydrogen-oxygen cell with an alkaline electrolyte. (1 mark)



Question 10d.

Marks	0	1	Average
%	68	32	0.3

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

Students should be aware that in a fuel cell the fuel (in this case H_2) is oxidised at the anode, that since the electrolyte is alkaline OH^-(aq) must be present in the half-equations and that the half-equations for an alkaline hydrogen-oxygen fuel cell are accessible from the electrochemical series.

On the electrochemical series the half-equations relevant to this fuel-cell are:

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

$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$ –0.83 V

Space fo