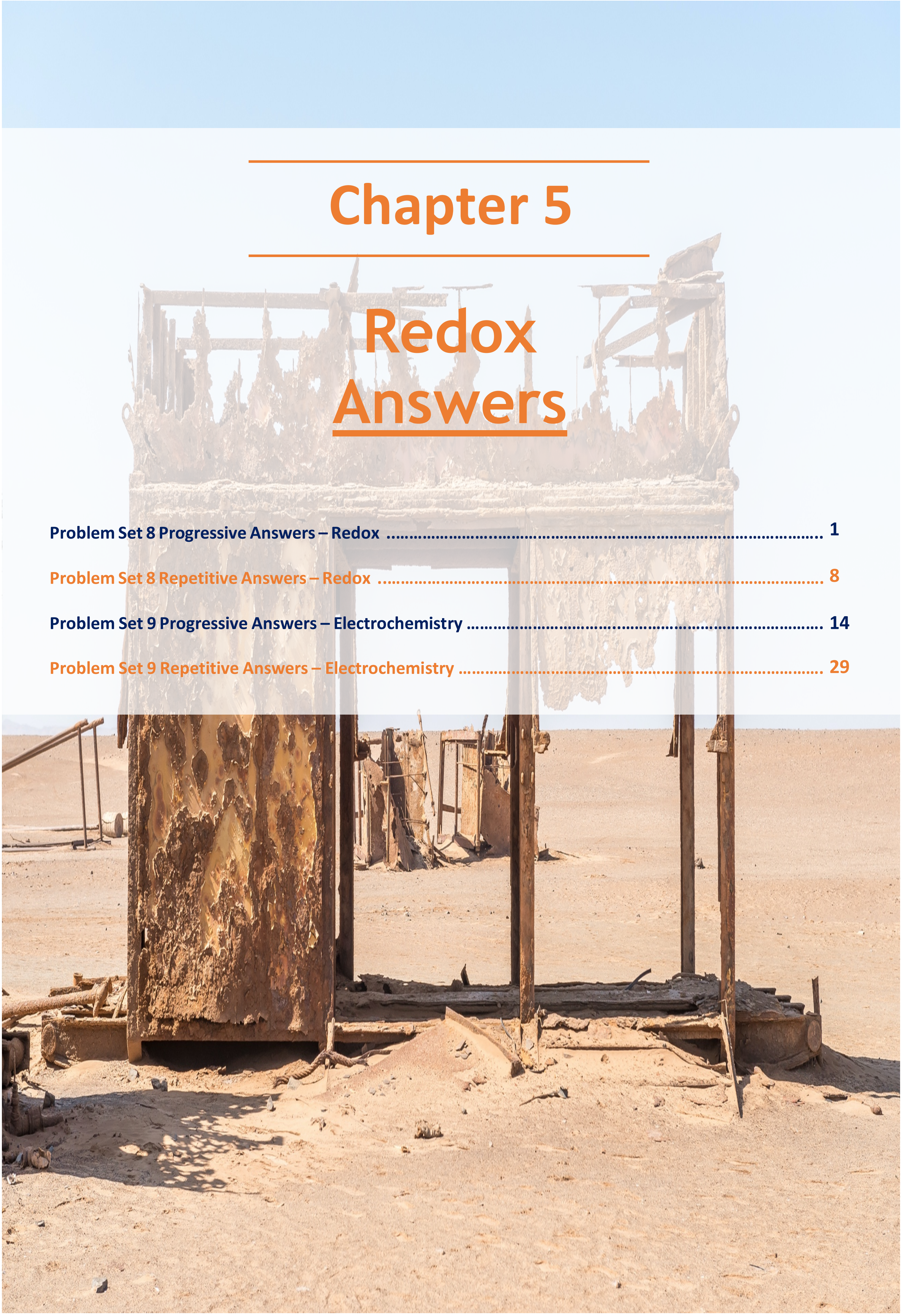


Chapter 5

Redox Answers

Problem Set 8 – Redox

Progressive Questions

Concept 1

Oxidation and Reduction – Progressive Questions Answers

Oxidation Numbers: Q1, Q2, Q3, Q4, Q5

1. **[9 marks]**

Student	Substance	Student's Oxidation Number Guess	Actual Oxidation Number	Progresses to next stage (✔ or X)
Luke	$Al_{(s)}$	0	0	✔ (1)
Peter	$O_{2\,(g)}$	0	0	✔ (1)
Janet	$Cd^{2+}_{(aq)}$	**Unsure**	$+2$	✔ (1)
Sarah	$F^{-}_{(aq)}$	-1	-1	✔ (1)
Alexa	$MnO^{-}_{4\,(aq)}$	**Unsure**	$+7$	✔ (1)
Dylan	$HClO_{(aq)}$	**Unsure**	$+1$	✔ (1)
Jamie	$H_2O_{2\,(aq)}$	-1	-1	✔ (1)
Rupert	$H_2C_2O_{4\,(aq)}$	$+3$	$+3$	✔ (1)
Tyler	$K_2Cr_2O_{7\,(aq)}$	$+6$	$+6$	✔ (1)
Rabea	$S_2O^{2-}_{3\,(aq)}$	$+2$	$+2$	✔ (1)
Tom	$MgH_{2\,(s)}$	**Unsure**	-1	✔ (1)

2. **[8 marks]**

$NH_4Cl_{(aq)}$ $(+1)\,(-1)$ — $N = -3$ **(1)**

$OF_{2\,(g)}$ (-1) — $O = +2$ **(1)**

$Cr(OH)_{3\,(s)}$ $(-2)(+1)$ — $Cr = +3$ **(1)**

$BaO_{2\,(aq)}$ $(+4)$ — $O = -2$ **(1)**

$CrO^{2-}_{4\,(aq)}$ (-2) — $Cr = +6$ **(1)**

$ClO^{-}_{4\,(aq)}$ (-2) — $Cl = +7$ **(1)**

$Fe(NO_3)_{3\,(aq)}$ (-2) — $N = +5$ **(1)**

$H_3PO_{4\,(aq)}$ $(+1)\,(-2)$ — $P = +5$ **(1)**

Points to note: For OF_2, the fluorine is more electronegative than oxygen which is why oxygen has an oxidation number of $+2$ not -2. Additionally, when you have known ions such as OH^- or NO_3^- you can assume they have these overall charges within the molecule.

3. **[6 marks]**

Rupert: $$\underset{(+2)(-2)}{CuO_{(s)}} + \underset{(0)}{C_{(s)}} \rightarrow \underset{(0)}{Cu_{(s)}} + \underset{(+2)(-2)}{CO_{(g)}}$$ **(1)**

Oxidised	Reduced
C	CuO

∴ Rupert is **correct** because C is **oxidised** as its oxidation number has **increased**, and Cu is **reduced** as its oxidation number has **decreased**. ∴ Rupert will progress to the next stage. **(1)**

Tyler: $$\underset{(+3)}{2Cr^{3+}_{(aq)}} + \underset{(0)}{3H_{2\,(g)}} + \underset{(-2)(+1)}{6OH^{-}_{(aq)}} \rightarrow \underset{(0)}{2Cr_{(s)}} + \underset{(+1)(-2)}{6H_2O_{(l)}}$$ **(1)**

Oxidised	Reduced
H_2	Cr^{3+}

∴ Tyler is **correct** because H_2 is **oxidised** as its oxidation number has **increased**, and Cr^{3+} is **reduced** as its oxidation number has **decreased**. ∴ Tyler will progress to the next stage. **(1)**

Sarah: $$\underset{(+6)\ (-2)}{Cr_2O^{2-}_{7\,(aq)}} + \underset{(+1)\ (-1)}{3H_2O_{2\,(aq)}} + \underset{(+1)}{8H^{+}_{(aq)}} \rightarrow \underset{(+3)}{2Cr^{3+}_{(aq)}} + \underset{(+1)(-2)}{7H_2O_{(l)}} + \underset{(0)}{3O_{2\,(g)}}$$ **(1)**

Oxidised	Reduced
$Cr_2O_7^{2-}$	H_2O_2

∴ Sarah is **incorrect** because H_2O_2 is actually **oxidised** as its oxidation number has **increased**, and $Cr_2O_7^{2-}$ is **reduced** as its oxidation number has **decreased**. ∴ Sarah will be eliminated. **(1)**

Point to note: For Sarah, it is important to remember that oxygen has an oxidation number of $+1$ in peroxides.

Marking Criteria	Marks Allocated
• Correctly determines oxidation numbers for each reaction	$1-3$
• States whether student is right or wrong	$1-3$
Total	**6**

4. **[6 marks]**

Luke: $$\underset{(+2)\ (-2)(+1)}{4Fe(OH)_{2\,(s)}} + \underset{(0)}{O_{2\,(aq)}} + \underset{(+1)(-2)}{2H_2O_{(l)}} \rightarrow \underset{(+3)\ (-2)(+1)}{4Fe(OH)_{3\,(s)}}$$ **(1)**

Reductant: $Fe(OH)_2$

∴ Luke is **correct** because $Fe(OH)_2$ is oxidised and is therefore the **reductant**. **(1)**

Dylan: $$\underset{(+1)\ (+3)(-2)}{H_2C_2O_{4\,(aq)}} + \underset{(+1)(+1)(-2)}{HClO_{(aq)}} \rightarrow \underset{(+4)(-2)}{2CO_{2\,(g)}} + \underset{(+1)}{H^{+}_{(aq)}} + \underset{(-1)}{Cl^{-}_{(aq)}} + \underset{(+1)\ (-2)}{H_2O_{(l)}}$$ **(1)**

Reductant: $HClO$

∴ Dylan is **incorrect** because $H_2C_2O_4$ is oxidised and is therefore the **reductant**, not $HClO$. **(1)**

Peter: $$\underset{(+7)\ (-2)}{2MnO^{-}_{4\,(aq)}} + \underset{(+1)}{16H^{+}_{(aq)}} + \underset{(0)}{5Pb_{(s)}} + \underset{(+6)(-2)}{5SO^{2-}_{4\,(aq)}} \rightarrow \underset{(+2)}{2Mn^{2+}_{(aq)}} + \underset{(+1)\ (-2)}{8H_2O_{(l)}} + \underset{(+2)(+6)(-2)}{5PbSO_{4\,(s)}}$$

Reductant: Pb

∴ Luke is **correct** because Pb is oxidised and is therefore the **reductant**. **(1)**

Marking Criteria	Marks Allocated
• Correctly determines oxidation numbers for each reaction	$1-3$
• States whether student is right or wrong	$1-3$
Total	**6**

5. **[5 marks]**

$$\underset{(-8/3)(+1)}{C_3H_{8\,(g)}} + \underset{(0)}{O_{2\,(g)}} \rightarrow \underset{(+4)(-2)}{CO_{2\,(g)}} + \underset{(+1)(-2)}{H_2O_{(l)}}$$

Yes or No

$NaHCO_{3\,(s)} \rightarrow Na_2CO_{3\,(s)} + H_2O_{(l)} + CO_{2\,(g)}$ — Yes or No
$(+1)(+1)(+4)(-2)$ $\quad (+1)\ (+4)(-2)$ $\quad (+1)(-2)$ $\quad (+4)(-2)$

$Cu_{(s)} + H_2SO_{4\,(aq)} \rightarrow CuSO_{4\,(aq)} + H_{2\,(g)}$ — Yes or No
(0) $\quad (+1)(+6)(-2)$ $\quad (+2)(+6)(-2)$ $\quad (0)$

$H_2S_2O_{7\,(l)} + H_2O_{(l)} \rightarrow 2H_2SO_{4\,(l)}$ — Yes or No
$(+1)\ (+6)\ (-2)$ $\quad (+1)(-2)$ $\quad (+1)(+6)(-2)$

$C_6H_{12}O_{6\,(s)} + 6O_{2\,(g)} \rightarrow 6CO_{2\,(g)} + 6H_2O_{(l)}$ — Yes or No
$(0)\ (+1)\ \ (-2)$ $\quad (0)$ $\quad (+4)(-2)$ $\quad (+1)(-2)$

Points to Note: A reaction is only a redox reaction if there is a change in oxidation numbers. For part (a), C in C_3H_8 has an oxidation number of $-8/3$. This is not theoretically possible, but this just represents the average oxidation number amongst the C atoms.

Marking Criteria	Marks Allocated
• Correctly circles 'Yes' or 'No'	$1-5$
Total	**5**

Half Equations and Overall Equations: Q6, Q7, Q8, Q9

6. **[7 marks]**

(a) $Ag_{(s)} \rightarrow Ag^{2+}_{(aq)} + 2e^-$ **(1)**

(b) $Cl_{2\,(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$ **(1)**

(c) $MnO^-_{4\,(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$ **(1)**

(d) $Cl_{2\,(g)} + 2H_2O_{(l)} \rightarrow 2HClO_{(aq)} + 2H^+_{(aq)} + 2e^-$ **(1)**

(e) $H_2O_{2\,(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow 2H_2O_{(l)}$ **(1)**

(f) $2Cr^{3+}_{(aq)} + 7H_2O_{(l)} \rightarrow Cr_2O^{2-}_{7\,(aq)} + 14H^+_{(aq)} + 6e^-$ **(1)**

(g) $N_2O_{4\,(aq)} + 2H_2O_{(l)} \rightarrow 2NO^-_{3\,(aq)} + 4H^+_{(aq)} + 2e^-$ **(1)**

Marking Criteria	Marks Allocated
• Writes the correct half equation	$1-7$
Total	**7**

7. **[14 marks]**

(a)

Oxidation Half-Equation: $H_{2\,(g)} \rightarrow 2H^+_{(aq)} + 2e^-$

Reduction Half-Equation: $Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$ **(1)**

$H_{2\,(g)} + 2Fe^{3+}_{(aq)} \rightarrow 2H^+_{(aq)} + 2Fe^{2+}_{(aq)}$ **(1)**

(b)

Oxidation Half-Equation: $Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$

Reduction Half-Equation: $H_2O_{2\,(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow 2H_2O_{(l)}$ **(1)**

$2Cr_{(s)} + 3H_2O_{2\,(aq)} + 6H^+_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 6H_2O_{(l)}$ **(1)**

(c)

Oxidation Half-Equation: $2S_2O^{2-}_{3\,(aq)} \rightarrow S_4O^{2-}_{6\,(aq)} + 2e^-$

Reduction Half-Equation: $I_{2\,(s)} + 2e^- \rightarrow 2I^-_{(aq)}$ **(1)**

$2S_2O^{2-}_{3\,(aq)} + I_{2\,(s)} \rightarrow S_4O^{2-}_{6\,(aq)} + 2I^-_{(aq)}$ **(1)**

(d)

Oxidation Half-Equation

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

Reduction Half-Equation

$PbO_{2\,(s)} + 4H^+_{(aq)} + 2e^- \rightarrow Pb^{2+}_{(aq)} + 2H_2O_{(l)}$ **(1)**

$Ni_{(s)} + PbO_{2\,(s)} + 4H^+_{(aq)} \rightarrow Ni^{2+}_{(aq)} + Pb^{2+}_{(aq)} + 2H_2O_{(l)}$ **(1)**

(e)

Oxidation Half-Equation

$Mn^{2+}_{(aq)} + 4H_2O_{(l)} \rightarrow MnO^-_{4\,(aq)} + 8H^+_{(aq)} + 5e^-$

Reduction Half-Equation

$N_{2\,(g)} + 8H^+_{(aq)} + 6e^- \rightarrow 2NH^+_{4\,(aq)}$ **(1)**

$6Mn^{2+}_{(aq)} + 24H_2O_{(l)} + 5N_{2\,(g)} \rightarrow 6MnO^-_{4\,(aq)} + 8H^+_{(aq)} + 10NH^+_{4\,(aq)}$ **(1)**

Marking Criteria	Marks Allocated
• Correctly writes both half-equations	$1-5$
• Balances half-equations to get overall equation	$1-5$
Total	**10**
Note: Give follow through marks for balancing if half-equations are incorrect	

8. **[10 marks]**

Scenario	Jamie and Alexa's Equation	Correct or Incorrect?
Chlorine gas is bubbled through a solution of **potassium iodide**	$Cl_{2\,(g)} + 2I^-_{(aq)} \rightarrow I_{2\,(s)} + 2Cl^-_{(aq)}$	Correct (2)
Zinc metal is dissolved in a solution of **hydrochloric acid**	$2Cl^-_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + Cl_{2\,(g)}$	Incorrect (2)
Acidified **potassium permanganate** solution is added to **solid iron**	$MnO^-_{4\,(aq)} + 8H^+_{(aq)} + 2Fe_{(s)} \rightarrow Mn^{2+}_{(aq)} + 8H_2O_{(l)} + 2Fe^{2+}_{(aq)}$	Incorrect (2)
Acidified **perchloric acid ($HClO$)** is added to a **sodium fluoride** solution	$2F^-_{(aq)} + 2HClO_{(aq)} + 2H^+_{(aq)} \rightarrow F_{2\,(g)} + Cl_{2\,(g)} + 2H_2O_{(l)}$	Incorrect (2)
Sulfur dioxide is bubbled through an **acidified dichromate ($Cr_2O_7^{2-}$) solution**, producing sulfate ions	$2SO_{2\,(g)} + Cr_2O_7^{2-} + 6H^+_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 3H_2O_{(l)} + 2SO^{2-}_{4\,(aq)}$	Incorrect (2)

Point to note: For each of these scenarios, write out the half-equations and then the overall balanced equation to determine if they are correct. The workings below are as follows:

(a) Half-equations: $2I^-_{(aq)} \rightarrow I_{2\,(s)} + 2e^-$ and $Cl_{2\,(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$
Overall equation: $2I^-_{(aq)} + Cl_{2\,(g)} \rightarrow I_{2\,(s)} + 2Cl^-_{(aq)}$ $\therefore$ the equation is **correct**.

(b) In this scenario, hydrogen ions will actually oxidise to form hydrogen gas ($E^0 = 0V$) in preference to the chlorine ions ($-1.36V$). Therefore, the equation is incorrect.

(c) Half-equations: $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ and $MnO^-_{4\,(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$
Overall equation: $5Fe_{(s)} + 2MnO^-_{4\,(aq)} + 16H^+_{(aq)} \rightarrow 5Fe^{2+}_{(aq)} + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$
$\therefore$ the equation is incorrect

(d) In this scenario, no reaction will occur because it is not spontaneous. The overall E^0 is $1.49 - 2.89 = -1.4V$

(e) Half-equations: $SO_{2\,(g)} + 2H_2O_{(l)} \rightarrow SO^{2-}_{4\,(aq)} + 4H^+_{(aq)} + 2e^-$ and $Cr_2O^{2-}_{7\,(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$
Overall equation: $3SO_{2\,(g)} + Cr_2O^{2-}_{7\,(aq)} + 2H^+_{(aq)} \rightarrow 3SO^{2-}_{4\,(aq)} + 2Cr^{3+}_{(aq)} + H_2O_{(l)}$
$\therefore$ the equation is incorrect

9. **[8 marks]**

Reaction and Observation	True	False
Reactants: A solid piece of **chromium** and a solution of **copper sulfate** Observation: A **silver metal** is added to a **colourless solution**, to produce a **green solution** and a **blue metal**.		✔ **(1)**
Explanation (For the box you ticked): The **copper sulfate** solution is **blue**, not colourless, and **copper** is a **salmon pink** coloured metal, not blue. **(1)**		
Reactants: **Chlorine gas** is bubbled through a solution of **hydrogen peroxide** Observation: A **greenish-yellow gas** is bubbled through a **colourless solution**, to produce a **colourless solution**.		✔ **(1)**
Explanation (For the box you ticked): **Hydrogen peroxide** will be **oxidised**, and will therefore **produce oxygen gas** (so a colourless, odourless gas will be produced). **(1)**		
Reactants: A **strip of silver** added to a solution of **bromine water** Observation: A **solid silver strip** is added to a **red solution**, to produce a **colourless solution**.		✔ **(1)**
Explanation (For the box you ticked): **Silver bromide** will form as a precipitate, so a **white precipitate** will be **formed**. **(1)**		
Reactants: A solution of **oxalic acid** ($H_2C_2O_4$) is added to a solution of **iron (III) nitrate** Observation: A **colourless solution** is added to a **colourless solution**, to produce a **colourless solution** and a **colourless**, **pungent gas**.		✔ **(1)**
Explanation (For the box you ticked): **Iron (III) nitrate** is a **pale brown colour**, and **carbon dioxide** will be produced which is **odourless**, not pungent. **(1)**		

Marking Criteria	Marks Allocated
• Selects correct true/false box	$1-4$
• Provides adequate explanation of **all** the issues with the observation	$1-4$
Total	**8**

Reaction Tendency: Q10, Q11, Q12

10. **[8 marks]**

Student	Reactants	Student's Guess (Yes or No)	Progresses to next stage (✔ or X)
Rupert	**Copper nitrate** and **gold**	No	✔ (1)
	Potassium permanganate and **sodium fluoride**	No	X (1)
Luke	**Chlorine gas** and **sodium fluoride solution**	Yes	✔ (1)
	Fluorine gas and **hydrogen peroxide solution**	Yes	✔ (1)
Tyler	**Copper metal** and **tin (II) carbonate solution**	No	✔ (1)
	Solid manganese and **aluminium sulfate solution**	Yes	✔ (1)
Peter	**Oxalic acid** and **zinc nitrate solution**	Yes	X (1)
	Sodium dichromate solution and **magnesium chloride**	No	X (1)

The E_0 values for each of the reactions are as follows:

(a) $E^0 = -0.34 - 1.50 = \mathbf{-1.84V}$ **(1)**

(b) $E^0 = -1.23 + 1.51 = \mathbf{+0.28V}$ **(1)**

Note: Water will be oxidised ($\mathbf{2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^-} \quad \mathbf{E_0 = -1.23V}$)

(c) $E^0 = -1.23 + 1.36 = \mathbf{+0.13V}$ **(1)**

Note: Water will be oxidised ($\mathbf{2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^-} \quad \mathbf{E_0 = -1.23V}$)

(d) $E^0 = 2.89 - 0.70 = \mathbf{+2.19V}$ **(1)**

(e) $E^0 = -0.34 - 0.14 = \mathbf{-0.48V}$ **(1)**

(f) $E^0 = 1.18 - 0.83 = \mathbf{+0.35V}$ **(1)**

Note: Water will be reduced ($\mathbf{2H_2O + 2e^- \rightarrow H_{2\,(g)} + 2OH^-_{(aq)}} \quad \mathbf{E_0 = -0.83V}$)

(g) $E^0 = 0.43 - 0.76 = \mathbf{-0.33V}$ **(1)**

(h) $E^0 = -1.23 + 1.36 = \mathbf{0.13V}$ **(1)**

Note: Water will be oxidised ($\mathbf{2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^-} \quad \mathbf{E_0 = -1.23V}$)

∴ We can conclude that **Peter** will be eliminated!

Marking Criteria	Marks Allocated
• Determines correct E_0 value and if the student's guess is correct	$1 - 8$
Total	**8**

11. **[4 marks]**

Increasing oxidant strength: $F^-_{(aq)}$ $Mg_{(s)}$ $Zn_{(s)}$ $H_2O_{(l)}$ $Fe^{3+}_{(aq)}$ $Cr_2O_7^{2-}{}_{(aq)}$ **(2)**

Increasing reductant strength: $MnO_4^-{}_{(aq)}$ $Cl_{2\,(g)}$ $H_2O_{(l)}$ $Br_{2\,(l)}$ $Cu_{(s)}$ $H^+_{(aq)}$ **(2)**

Marking Criteria	Marks Allocated
• Corrected order only has one mistake	$1-2$
• Corrected order has no mistakes	$3-4$
Total	**4**

12. **[8 marks]**

Reaction	**Rupert's Prediction (Yes or No)**	**Tyler's Prediction (Yes or No)**	**Correct Person (Rupert or Tyler)**
A piece of **steel wool (iron)** is added to a solution of **bromine water**	**Yes**	**No**	**Rupert** **(1)**
A **gold plated earing** is added to a solution of **perchloric acid ($HClO$)**	**No**	**Yes**	**Tyler** **(1)**
A **tin can** is added to a solution of **lead nitrate**	**No**	**Yes**	**Tyler** **(1)**
A **lead dioxide** and **sulfate ion solution** from a car battery is added to some **solid magnesium**	**Yes**	**No**	**Rupert** **(1)**
Some **solid aluminium** from the fridge door handle is added to a **manganese nitrate solution**	**Yes**	**No**	**Rupert** **(1)**
A **'chromium plated' fork** is added to a solution of **potassium permanganate**	**No**	**Yes**	**Tyler** **(1)**
Solid copper from a copper pipe is added to some **iron(III) sulfate solution**	**No**	**Yes**	**Tyler** **(1)**

∴ Tyler will win!

The E_0 values for each reaction are as follows:

(a) $E^0 = 0.44 + 1.08 = \mathbf{+1.52V}$
(b) $E^0 = -1.29 + 1.49 = \mathbf{+0.20V}$
Note: Water will be oxidised ($\mathbf{2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^-}$ $\quad \mathbf{E_0 = -1.23V}$)
(c) $E^0 = 0.14 - 0.13 = \mathbf{+0.01V}$
(d) $E^0 = 2.36 + 1.69 = \mathbf{+4.05V}$
(e) $E^0 = 1.68 - 1.18 = \mathbf{+0.50V}$
(f) $E^0 = -0.74 + 1.51 = \mathbf{+0.77V}$
(g) $E^0 = 1.68 - 1.18 = \mathbf{+0.50V}$
(h) $E^0 = -0.34 + 0.77 = \mathbf{+0.43}$

Marking Criteria	Marks Allocated
• Determines correct E^0 value and correct guess	$1-8$
Total	**8**

Problem Set 8 – Redox
Repetitive Questions

Concept 1
Oxidation and Reduction – Progressive Questions Answers

Oxidation Numbers: 1.1, 1.2, 1.3, 1.5

1.1 **[12 marks]**

(a) $Mn_{(s)}$ $(+1)\ (-1)$
$Mn = 0$ **(1)**

(e) $NO_{2\,(g)}$ $(+4)(-2)$
$O = +4$ **(1)**

(i) $CaBr_{2\,(s)}$ $(+2)(-1)$
$Br = -1$ **(1)**

(b) $H_2O_{(l)}$ $(+1)(-2)$
$O = -2$ **(1)**

(f) $S_2O^{2-}_{3\,(aq)}$ $(+2)(-2)$
$S = +2$ **(1)**

(j) $NH^{+}_{4\,(aq)}$ $(-3)(+1)$
$N = -3$ **(1)**

(c) $Ca^{2+}_{(aq)}$ $(+2)$
$Ca = +2$ **(1)**

(g) $F_2O_{(g)}$ $(-1)(+2)$
$O = +2$ **(1)**

(k) $S_4O^{2-}_{6\,(aq)}$ $(+2.5)(-2)$
$S = +2.5$ **(1)**

(d) $Br_{2\,(l)}$ (0)
$Br = 0$ **(1)**

(h) $ClO^{-}_{4\,(aq)}$ $(+7)(-2)$
$Cl = +7$ **(1)**

(l) $Al(OH)_{3\,(s)}$ $(+3)\ (-2)(+1)$
$Al = +3$ **(1)**

Marking Criteria	Marks Allocated
• Determines correct oxidation number	$1 - 12$
Total	**12**

1.2 **[8 marks]**

(a) $Na_2Cr_2O_{7\,(aq)}$ $(+1)\ \ (+6)\ \ (-2)$
$Cr = +6$ **(1)**

(e) $Ni(NO_3)_{2\,(aq)}$ $(+2)\ (+5)(-2)$
$N = +5$ **(1)**

(b) $H_2O_{2\,(aq)}$ $(+1)(-1)$
$O = -1$ **(1)**

(f) $CdCO_{3\,(s)}$ $(+2)(+4)(-2)$
$C = +4$ **(1)**

(c) $NaH_{(aq)}$ $(+1)(-1)$
$H = -1$ **(1)**

(g) $Mn_2O_{3\,(s)}$ $(+3)\ \ (-2)$
$Mn = +3$ **(1)**

(d) $PbSO_{4\,(s)}$ $(+2)(+6)(-2)$
$O = -1$ **(1)**

(h) $(NH_4)_2Cr_2O_{7\,(s)}$ $(-3)(+1)\ \ (+6)\ (-2)$
$Cr = +6$ **(1)**

Points to note: For ions such as NO_3^-, NH_4^+, $Cr_2O_7^{2-}$ etc. you can assume they have these overall charges within the molecule, and you can therefore determine the oxidation numbers of elements such as N or Cr by just working within the ion rather than the whole molecule. For part (c), NaH is a metal hydride so hydrogen has an oxidation number of -1.

Marking Criteria	Marks Allocated
• Determines correct oxidation number	$1-8$
Total	**8**

1.3 **[12 marks]**

(a) $3\underset{(+2)}{Fe^{2+}_{(aq)}} + \underset{(+3)}{Cr^{3+}_{(aq)}} \rightarrow 3\underset{(+3)}{Fe^{3+}_{(aq)}} + \underset{(0)}{Cr_{(s)}}$

Oxidised	Reduced	Oxidant	Reductant	
Fe^{2+}	Cr^{3+}	Cr^{3+}	Fe^{2+}	**(2)**

(b) $\underset{(0)}{F_{2\,(g)}} + \underset{(+1)\,(-1)}{H_2O_{2\,(aq)}} \rightarrow 2\underset{(-1)}{F^-_{(aq)}} + \underset{(0)}{O_{2\,(g)}} + 2\underset{(+1)}{H^+_{(aq)}}$

Oxidised	Reduced	Oxidant	Reductant	
H_2O_2	F_2	F_2	H_2O_2	**(2)**

(c) $\underset{(+1)(+3)(-2)}{H_2C_2O_{4\,(aq)}} + \underset{(0)}{S_{(s)}} \rightarrow 2\underset{(+4)(-2)}{CO_{2\,(g)}} + \underset{(+1)(-2)}{H_2S_{(aq)}}$

Oxidised	Reduced	Oxidant	Reductant	
$H_2C_2O_4$	S	S	$H_2C_2O_4$	**(2)**

(d) $6\underset{(+1)(+1)(-2)}{HClO_{(aq)}} + 2\underset{(+3)}{Cr^{3+}_{(aq)}} + \underset{(+1)(-2)}{H_2O_{(l)}} \rightarrow 3\underset{(0)}{Cl_{2\,(g)}} + \underset{(+6)\,(-2)}{Cr_2O^{2-}_{7\,(aq)}} + 8\underset{(+1)}{H^+_{(aq)}}$

Oxidised	Reduced	Oxidant	Reductant	
Cr^{3+}	$HClO$	$HClO$	Cr^{3+}	**(2)**

(e) $2\underset{(+2)(-2)}{S_2O^{2-}_{3\,(aq)}} + \underset{(+4)\,(-2)}{PbO_{2\,(s)}} + \underset{(+6)(-2)}{SO^{2-}_{4\,(aq)}} + 4\underset{(+1)}{H^+_{(aq)}} \rightarrow \underset{(+2.5)(-2)}{S_4O^{2-}_{6\,(aq)}} + \underset{(+2)(+6)(-2)}{PbSO_{4\,(s)}} + 2\underset{(+1)\,(-2)}{H_2O_{(l)}}$

Oxidised	Reduced	Oxidant	Reductant	
$S_2O_3^{2-}$	PbO_2	PbO_2	$S_2O_3^{2-}$	**(2)**

(f) $2\underset{(+7)\,\,(-2)}{MnO^-_{4\,(aq)}} + 10\underset{(-1)}{Br^-_{(aq)}} + 16\underset{(+1)}{H^+_{(aq)}} \rightarrow 2\underset{(+2)}{Mn^{2+}_{(aq)}} + 5\underset{(0)}{Br_{2\,(l)}} + 8\underset{(+1)\,\,(-2)}{H_2O_{(l)}}$

Oxidised	Reduced	Oxidant	Reductant	
Br^-	MnO_4^-	MnO_4^-	Br^-	**(2)**

Points to note: For part (e), S in $S_4O_6^{2-}$ has an oxidation number of $+2.5$. This is not theoretically possible, but this just represents the average oxidation number amongst the S atoms.

Marking Criteria	Marks Allocated
• Writes correct oxidation numbers	$1-6$
• Determines correct oxidised and reduced species, and correct oxidant/reductant	$1-6$
Total	**12**

1.5 **[8 marks]**

$\underset{(+3)\ (-1)}{FeCl_{3(s)}} \rightarrow \underset{(+3)}{Fe^{3+}_{(aq)}} + \underset{(-1)}{3Cl^{-}_{(aq)}}$	Yes or No
$\underset{(0)}{Mg_{(s)}} + \underset{(+1)(-1)}{2HCl_{(aq)}} \rightarrow \underset{(+2)\ (-1)}{MgCl_{2\,(aq)}} + \underset{(0)}{H_{2\,(g)}}$	Yes or No
$\underset{(-4)(+1)}{CH_{4\,(g)}} + \underset{(0)}{O_{2\,(g)}} \rightarrow \underset{(+4)(-2)}{CO_{2\,(g)}} + \underset{(+1)(-2)}{2H_2O_{(l)}}$	Yes or No
$\underset{(-8/3)(+1)}{2C_3H_{8\,(g)}} + \underset{(0)}{7O_{2\,(g)}} \rightarrow \underset{(+4)(-2)}{2CO_{2\,(g)}} + \underset{(+1)(-2)}{8H_2O_{(l)}} + \underset{(+2)(-2)}{2CO_{(g)}} + \underset{(0)}{2C}$	Yes or No
$\underset{(+1)(-2)(+1)}{NaOH_{(aq)}} + \underset{(+1)(-1)}{HCl_{(aq)}} \rightarrow \underset{(+1)(-1)}{NaCl_{(aq)}} + \underset{(+1)(-2)}{H_2O_{(l)}}$	Yes or No
$\underset{(+1)\ (-1)}{AgBr_{(s)}} \rightarrow \underset{(+1)}{Ag^{+}_{(aq)}} + \underset{(-1)}{Br^{-}_{(aq)}}$	Yes or No
$\underset{(+7)\ (-2)}{2MnO^{-}_{4\,(aq)}} + \underset{(+1)}{16H^{+}_{(aq)}} + \underset{(-1)}{10I^{-}_{(aq)}} \rightarrow \underset{(+2)}{2Mn^{2+}_{(aq)}} + \underset{(+1)(-2)}{8H_2O_{(l)}} + \underset{(0)}{5I_{2\,(s)}}$	Yes or No
$\underset{(+1)(+4)(-2)}{K_2CO_{3\,(s)}} + \underset{(+1)(-1)}{2HCl_{(aq)}} \rightarrow \underset{(+4)(-2)}{CO_{2\,(g)}} + \underset{(+1)(-1)}{2KCl_{(aq)}} + \underset{(+1)(-2)}{H_2O_{(l)}}$	Yes or No

Points to note: For part (d), C in C_3H_8 has an oxidation number of $-8/3$. This is not theoretically possible, but this just represents the average oxidation number amongst the C atoms.

Marking Criteria	Marks Allocated
• Correctly circles 'Yes' or 'No'	$1-8$
Total	**8**

Half Equations and Overall Equations: 1.6, 1.7, 1.9

1.6 **[7 marks]**

(a) $2CO_{2\,(g)} + 2H^{+}_{(aq)} + 2e^{-} \rightarrow H_2C_2O_{4\,(aq)}$ **(1)**

(b) $2S_2O^{2-}_{3\,(aq)} \rightarrow S_4O^{2-}_{6\,(aq)} + 2e^{-}$ **(1)**

(c) $VO^{+}_{2\,(aq)} + 2H^{+}_{(aq)} + e^{-} \rightarrow VO^{2+}_{(aq)} + H_2O_{(l)}$ **(1)**

(d) $CrO^{2-}_{4\,(aq)} + 8H^{+}_{(aq)} + 3e^{-} \rightarrow Cr^{3+}_{(aq)} + 4H_2O_{(l)}$ **(1)**

(e) $H_2C_2O_{4\,(aq)} \rightarrow 2CO_{2\,(g)} + 2H^{+}_{(aq)} + 2e^{-}$ **(1)**

(f) $Pb^{2+}_{(aq)} + 2H_2O_{(l)} \rightarrow PbO_{2\,(s)} + 4H^{+}_{(aq)} + 2e^{-}$ **(1)**

(g) $HClO_{(aq)} + H^{+}_{(aq)} + 2e^{-} \rightarrow Cl^{-}_{(aq)} + H_2O_{(l)}$ **(1)**

Marking Criteria	Marks Allocated
• Correctly balanced equation	$1-7$
Total	**7**

1.7 **[12 marks]**

(a)

Oxidation Half-Equation

$$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$$

Reduction Half-Equation

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)} \quad \textbf{(1)}$$

$$3Zn^{2+}_{(aq)} + 2Al_{(s)} \rightarrow 3Zn_{(s)} + 2Al^{3+}_{(aq)} \quad \textbf{(1)}$$

(b)

Oxidation Half-Equation

$$2H_2O_{(l)} \rightarrow H_2O_{2\,(aq)} + 2H^+_{(aq)} + 2e^-$$

Reduction Half-Equation

$$VO^{2+}_{(aq)} + 2H^+_{(aq)} + e^- \rightarrow V^{3+}_{(aq)} + H_2O_{(l)} \quad \textbf{(1)}$$

$$VO^{2+}_{(aq)} + 2H^+_{(aq)} \rightarrow H_2O_{2\,(aq)} + V^{3+}_{(aq)} \quad \textbf{(1)}$$

Point to note: Part (b) has an error, it should be water that is the specified species, not hydrogen ions, for the oxidation into hydrogen peroxide.

(c)

Oxidation Half-Equation

$$2NH^+_{4\,(aq)} \rightarrow N_{2\,(g)} + 8H^+_{(aq)} + 6e^-$$

Reduction Half-Equation

$$Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)} \quad \textbf{(1)}$$

$$4NH^+_{4\,(aq)} + 2Au^{3+}_{(aq)} \rightarrow N_{2\,(g)} + 8H^+_{(aq)} + 2Au_{(s)} \quad \textbf{(1)}$$

(d)

Oxidation Half-Equation

$$H_2O_{2\,(aq)} \rightarrow O_{2\,(g)} + 2H^+_{(aq)} + 2e^-$$

Reduction Half-Equation

$$2ClO^-_{3\,(aq)} + 12H^+_{(aq)} + 10e^- \rightarrow Cl_{2\,(g)} + 6H_2O_{(l)} \quad \textbf{(1)}$$

$$2ClO^-_{3\,(aq)} + 5H_2O_{2\,(aq)} + 2H^+_{(aq)} \rightarrow 5O_{2\,(g)} + Cl_{2\,(g)} + 6H_2O_{(l)} \quad \textbf{(1)}$$

(e)

Oxidation Half-Equation

$$PbSO_{4\,(aq)} + 2H_2O_{(l)} \rightarrow PbO_{2\,(s)} + SO^{2-}_{4\,(aq)} + 4H^+_{(aq)} + 2e^-$$

Reduction Half-Equation

$$2CO_{2\,(g)} + 2H^+_{(aq)} + 2e^- \rightarrow H_2C_2O_{4\,(aq)} \quad \textbf{(1)}$$

$$PbSO_{4\,(aq)} + 2CO_{2\,(g)} + 2H_2O_{(l)} \rightarrow PbO_{2\,(s)} + SO^{2-}_{4\,(aq)} + H_2C_2O_{4\,(aq)} + 2H^+_{(aq)} \quad \textbf{(1)}$$

(f)

Oxidation Half-Equation

$$HSO^-_{3\,(aq)} + H_2O_{(l)} \rightarrow HSO^-_{4\,(aq)} + 2H^+_{(aq)} + 2e^-$$

Reduction Half-Equation

$$Cr_2O^{2-}_{7\,(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)} \quad \textbf{(1)}$$

$$3HSO^-_{3\,(aq)} + Cr_2O^{2-}_{7\,(aq)} + 8H^+_{(aq)} \rightarrow 3HSO^-_{4\,(aq)} + 2Cr^{3+}_{(aq)} + 4H_2O_{(l)} \quad \textbf{(1)}$$

Marking Criteria	Marks Allocated
• Correctly writes both half-equations	$1-6$
• Balances half-equations to get overall equation	$1-6$
Total	**12**

1.9 **[18 marks]**

(a) Overall Reaction: $Sn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Sn^{2+}_{(aq)} + Cu_{(s)}$ **(1)**

Observation: A **silver metal** is added to a **blue solution**, to produce a **salmon pink solid** and the **blue colour** of the solution **fading** **(1)**.

(b) Overall Reaction: $Cl_{2\,(g)} + H_2S_{(aq)} \rightarrow 2Cl^-_{(aq)} + S_{(s)} + 2H^+_{(aq)}$ **(1)**

Observation: A **silver metal** is added to a **colourless solution**, to produce a **green solution** and a **blue metal** **(1)**.

Point to note: If you are struggling to find the hydrogen disulfide (H_2S) oxidation equation on the formula sheet, it is just above the oxidation of hydrogen gas at $0V$.

(c) Overall Reaction: $Ni_{(s)} + HClO_{(aq)} + H^+_{(aq)} \rightarrow Ni^{2+}_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$ **(1)**

Observation: A **silver metal** is added to a **colourless solution**, to produce a **green solution** **(1)**.

Point to note: The other equation that could have been used for perchloric acid is $2HClO_{(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow Cl_{2\,(g)} + 2H_2O_{(l)}$, we should have specified in the question whether the outcome was chlorine ions or chlorine gas (our mistake!).

(d) Overall Reaction: $Br_{2\,(l)} + H_2C_2O_{4\,(aq)} \rightarrow 2Br^-_{(aq)} + 2CO_{2\,(g)} + 2H^+_{(aq)}$ **(1)**

Observation: A **colourless solution** is added to a **red solution**, to produce a **colourless solution** and a **blue metal** **(1)**.

(e) Overall Reaction: $2Au^{3+}_{(aq)} + 6I^-_{(aq)} \rightarrow 2Au_{(s)} + 3I_{2\,(s)}$ **(1)**

Observation: A **colourless solution** is added to another **colourless solution**, to produce a **purple solid** and a **yellow metal** mixture **(1)**.

(f) Overall Reaction: $5Fe^{2+}_{(aq)} + MnO^-_{4\,(aq)} + 8H^+_{(aq)} \rightarrow 5Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)} + 4H_2O_{(l)}$ **(1)**

Observation: A **pale green solution** is added to a **purple solution**, to produce a **pale pink/pale brown** solution **(1)**.

(g) Overall Reaction: $2Cr_{(s)} + Cr_2O^{2-}_{7\,(aq)} + 14H^+_{(aq)} \rightarrow 3Cr^{3+}_{(aq)} + 7H_2O_{(l)}$ **(1)**

Observation: A **silver/grey metal** is added to a **orange solution**, to produce a **green solution** **(1)**.

(h) Overall Reaction: $2Cl_{2\,(g)} + 2H_2O_{(l)} \rightarrow 4Cl^-_{(aq)} + O_{2\,(g)} + 4H^+_{(aq)}$ **(1)**

Observation: A **greenish-yellow gas** is bubbled through a **colourless solution**, to produce a **colourless, odourless gas** and a **colourless solution** **(1)**.

Point to note: Water will be oxidised in this scenario ($2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^- \quad E_0 = -1.23V$)

(i) Overall Reaction: **No reaction ($E^0 < 0$)** **(1)**

Observation: A **yellow solid** is added to a **colourless solution**, and no reaction will occur from this **(1)**.

Marking Criteria	Marks Allocated
• Writes correct overall equation	$1-9$
• Writes an observation with all key highlighted components	$1-9$
Total	**18**

Reaction Tendency: 1.101, 1.111

1.101 **[16 marks]**

(a) $E^0 = 0.24 - 0.28 = \mathbf{-0.04V}$, ∴ the reaction will **not** be spontaneous **(1)**

(b) $E^0 = -1.23 + 1.08 = \mathbf{-0.15V}$, ∴ the reaction will **not** be spontaneous **(1)**

Note: Water would be oxidised in this scenario ($\mathbf{2H_2O \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^-}$ $\mathbf{E_0 = -1.23V}$) in preference to the chromium ions, if it were spontaneous.

(c) $E^0 = -0.4 + 0.54 = \mathbf{+0.14V}$, ∴ the reaction **will be** spontaneous **(1)**

Note: The OH^- ions are oxidised in this scenario ($\mathbf{4OH^-_{(aq)} \rightarrow O_{2\,(g)} + 2H_2O_{(l)} + 4e^-}$ $\mathbf{E_0 = -0.40V}$)

(d) $E^0 = 0.4 + 0 = \mathbf{+0.4V}$, ∴ the reaction **will be** spontaneous **(1)**

Note: The H^+ ions are reduced in this scenario ($\mathbf{2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}}$ $\mathbf{E_0 = 0V}$)

(e) $E^0 = -1.23 + 1.63 = \mathbf{+0.4V}$, ∴ the reaction **will be** spontaneous **(1)**

Note: Water is oxidised in this scenario. You could also choose to use the E_0 of the gold ions, depending on the perchloric acid equation you selected.

(f) $E^0 = -0.54 + 1.76 = \mathbf{+1.22V}$, ∴ the reaction **will be** spontaneous **(1)**

Note: The hydrogen peroxide will be reduced in preference to the iron (III) ions.

(g) $E^0 = 1.18 - 1.08 = \mathbf{+0.1V}$, ∴ the reaction **will be** spontaneous **(1)**

(h) $E^0 = -0.34 - 0.13 = \mathbf{-0.47V}$, ∴ the reaction will **not** be spontaneous **(1)**

Note: The hydrogen peroxide will be reduced in preference to the iron (III) ions.

Marking Criteria	Marks Allocated
• Correct E^0 value	$1-8$
• Correctly determines whether spontaneous or not	$1-8$
Total	**16**

1.111 **[8 marks]**

Point to note: Multiple errors were made with this question, and it is not possible to solve so please ignore it.

Problem Set 9 – Electrochemistry
Progressive Questions

Concept 1
Oxidation and Reduction – Progressive Questions Answers

Galvanic Cells: Q1, Q2, Q3, Q4, Q5

1. **[11 marks]**

(a) In order to **determine** the **voltage** produced by each half reaction, one of the half reactions needed to be assigned a potential energy of **zero** **(1)**. To do this the hydrogen half-cell was set as the 'standard' **(1)**, with a **voltage of zero**. This half-cell could then be used to compare the relative voltages of all other half-reactions **(1)**, and from doing this they were able to create the 'Standard Reduction Potentials Table' **(1)**.

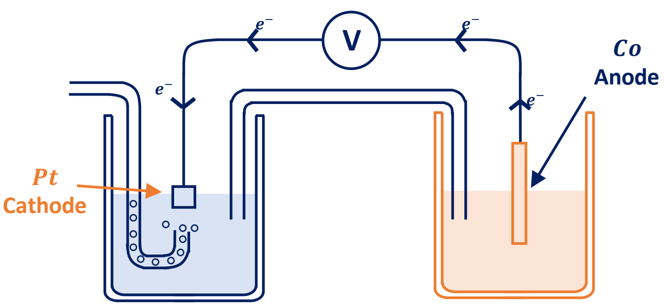
Marking Criteria	Marks Allocated
• To compare half-reactions, one reaction need to be set as a standard. The hydrogen half-cell was selected • The hydrogen half-cell was allocated a potential energy of zero. • This half-cell was used to compare the **relative voltages** of all other half-reactions • Repeating this process across all half-reactions allowed for the creation of the Standard Reduction Potentials Table	$1-4$
Total	**4**

(b) Limitations of the Standard reductions potential table include:

- It only applies to aqueous solutions. **(1)**
- The values of E^0 will change depending on the temperature, pressure, acidity and solution concentration (i.e. the conditions must be at STP). **(1)**
- The values of E^0 do not indicate anything about reaction rate. **(1)**

Marking Criteria	Marks Allocated
• States two limitations	$1-2$
Total	**4**

(c)



Oxidation half-cell: $Co_{(s)} \rightarrow Co^{2+}_{(aq)} + 2e^-$

Reduction half-cell: $2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$

$$Co_{(s)} + 2H^+_{(aq)} \rightarrow Co^{2+}_{(aq)} + H_{2\,(g)}$$

$$E^0 = 0.28 + 0 = +0.28V$$

Marking Criteria	Marks Allocated
• Labels anode, cathode and electron flow correctly	$1-3$
• Correct overall equation and E^0 calculation	$1-2$
Total	**5**

2. **[6 marks]**

Reaction and Observation	True	False
Cell 1 Observation: In the **left half-cell** the **silver electrode** will **decrease in size** over time, and in the **right half-cell** a **colourless, odourless gas** will be produced.		✔ **(1)**
Explanation (For the box you ticked): In the right cell, no colourless gas will be produced. The reduction of H_2O_2 produces $H_2O_{(l)}$ so **no observations** will occur **(1)**.		
Cell 2 Observation: In the **left half-cell** the **solution** will become a **lighter red colour** over time, and in the **right half-cell** the **silver electrode** will **increase in size** over time.		✔ **(1)**
Explanation (For the box you ticked): In the right cell, the iron anode will **decrease in size** over time and will oxidise to form pale green $Fe^{2+}_{(aq)}$ ions.		
Cell 3 Observation: In the **left half-cell** the **solution** will become a **lighter purple and more pale pink colour** over time, and in the **right half-cell** a **colourless, odourless gas** will be produced.		✔ **(1)**
Explanation (For the box you ticked): In the right hand cell, the zinc electrode will **decrease in size** over time. No colourless, odourless gas will be produced **(1)**.		

Marking Criteria	Marks Allocated
• Selects correct True/False box	$1-4$
• Provides adequate explanation of **all** the issues with the observation	$1-4$
Total	**8**

3. **[10 marks]**

(a) **Anode**: $Co_{(s)}$ **Cathode**: $C_{(s)}$ **Voltage**: $\mathbf{0.28 + 1.36 = +1.64V}$ **(2)**

(b) **Anode**: $Cr_{(s)}$ **Cathode**: $Au_{(s)}$ **Voltage**: $\mathbf{0.74 + 1.50 = +2.24V}$ **(2)**

(c) **Anode**: $Cu_{(s)}$ **Cathode**: $Pt_{(s)}$ **Voltage**: $\mathbf{-0.34 + 1.36 = +1.02V}$ **(2)**

(d) **Anode**: $Mn_{(s)}$ **Cathode**: $Pt_{(s)}$ **Voltage**: $\mathbf{1.18 + 1.76 = +2.94V}$ **(2)**

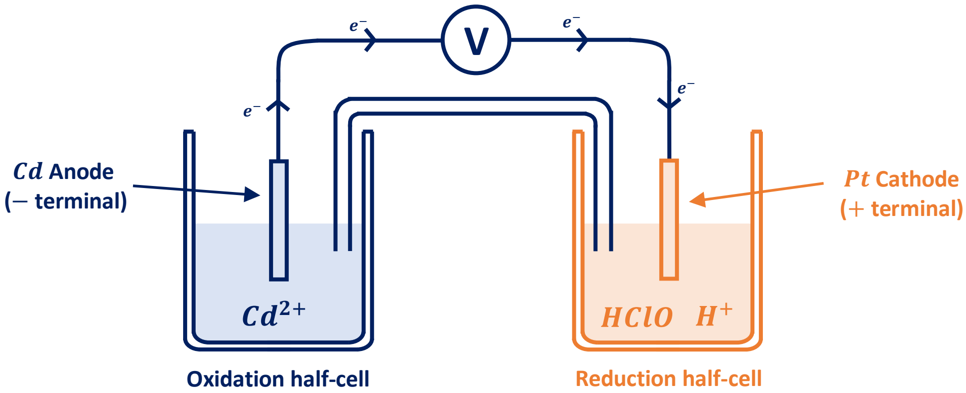
(e) **Anode**: $Al_{(s)}$ **Cathode**: $C_{(s)}$ **Voltage**: $\mathbf{1.68 + 0.77 = +2.45V}$ **(2)**

Marking Criteria	Marks Allocated
• Determines correct anode and cathode	$1-5$
• Determines correct voltage	$1-5$
Total	**10**

4. **[13 marks]**

(a)

$$E^0 = 0.40 + 1.63 = +2.03V$$



$$Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^- \quad 2HClO_{(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow Cl_{2\,(g)} + 2H_2O_{(l)}$$

$$Cd_{(s)} + 2HClO_{(aq)} + 2H^+_{(aq)} \rightarrow Cd^{2+}_{(aq)} + Cl_{2\,(g)} + 2H_2O_{(l)}$$

Point to note: The other equation and voltage for hypochlorous acid could also be used: $HClO_{(aq)} + H^+_{(aq)} + e^- \rightarrow Cl^-_{(aq)} + H_2O_{(l)} \quad E^0 = 1.49V$, as the products of the reaction was not specified.

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Determines correct cell E^0 • Labels correct ions • Determines correct overall cell equation	$1-4$
Total	**4**

(b)

(i) The salt bridge allows for the **controlled flow of ions** across the half-cells but forces the electrons to still flow through an external circuit **(1)**. This **maintains electrical neutrality** and **completes the circuit** **(1)**.

Marking Criteria	Marks Allocated
• Allows for the controlled flow of ions • Maintains electrical neutrality/completes the circuit	$1-2$
Total	**2**

(ii) Overtime the cell **emf** will start to **slowly decrease** over time **(1)**, because the rate of oxidation and reduction will slowly decrease over time **(1)**.

Marking Criteria	Marks Allocated
• Emf will slowly decrease over time, as oxidation/reduction rates decrease	$1-2$
Total	**2**

(iii) If the concentration of $HClO$ was increased, the rate of reduction would increase **(1)** and the cell emf would increase **(2)**.

Marking Criteria	Marks Allocated
• The rate of reduction would increase • The cell emf would therefore increase	$1-2$
Total	**2**

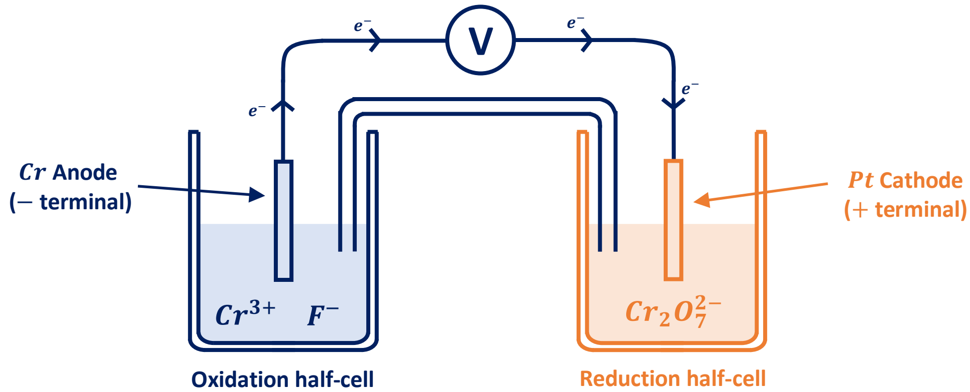
(iv) In galvanic cells there is not a significant enough charge to pull the ions towards the opposite cell **(1)**. Therefore the Cd^{2+} ions will move towards the anode **(1)**, and the $HClO$ and H^+ ions will move towards the cathode **(1)**.

Marking Criteria	Marks Allocated
• Galvanic cells do not produce a significant enough charge to pull ions to the opposite electrodes • Cd^{2+} ions move towards the anode • $HClO$ and H^+ ions move towards the cathode	$1-3$
Total	**3**

5. **[12 marks]**

(a)

$$E^0 = 0.74 + 1.36 = +2.10V$$



$$Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^- \qquad Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 6e^- \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$$

$$2Cr_{(s)} + Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} \rightarrow 4Cr^{3+}_{(aq)} + 7H_2O_{(l)}$$

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Determines correct cell E^0 • Labels correct ions • Determines correct overall cell equation	$1-4$
Total	**4**

(b) Peter is very wrong, he has the cells on the wrong sides and has everything mixed up **(1)**. It should be closer to: "In the **left half-cell** the **silver electrode** will **decrease in size** over time and the surrounding solution will become a **darker green** **(1)**, and in the **right half-cell** the **orange solution colour** will fade and become a **green colour** **(1)**."

Marking Criteria	Marks Allocated
• States that Peter is incorrect • Gives correct anode observations • Gives correct cathode observations	$1-3$
Total	**3**

(c)

Dr Lachlan's Statement	Peter's Prediction (True or False)	Correct Answer (Yes or No)
"If the **concentration** of **potassium dichromate** was **increased**, the **voltage produced** would be **larger**"	**True**	**Yes**
"The role of the **electrolyte** is to **maintain electrical neutrality** and **complete the circuit**"	**True**	**Yes**
"If **bromine water** was used instead of **potassium dichromate**, the cell would **no longer operate**"	**False**	**No**
"The temperature of the cells must be at $25°C$ to **accurately predict** the overall voltage using the **Standard Reduction Potential Table**"	**True**	**Yes**
"As time progresses, the overall cell **voltage** will **slowly increase**"	**True**	**Yes**

∴ Peter **will not** be allowed to stay at the factory

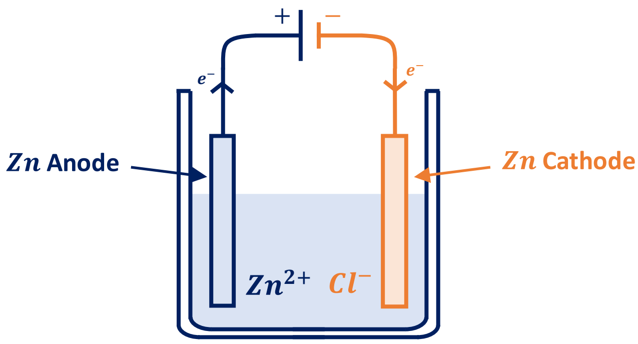
Point to note: The third statement is 'true', the cell would have an overall voltage of $+1.92V$

Marking Criteria	Marks Allocated
• Correctly determines if Peter is correct for each statement	$1-5$
Total	**5**

Electrolytic Cells: Q6, Q7, Q8, Q9

6. **[9 marks]**

(a)



Oxidation half-reaction

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

Reduction half-reaction

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

$$Zn_{(s)} + Zn^{2+}_{(aq)} \rightarrow Zn_{(s)} + Zn^{2+}_{(aq)}$$ or **None**

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Determines correct cell E^0 • Labels correct ions • Labels positive and negative electrodes • Determines correct overall cell equation	$1-5$
Total	**5**

(b) On the left side the **grey electrode** will **decrease in size** over time **(1)**, and on the **right side** the **grey electrode** will **increase in size** over time **(1)**.

Marking Criteria	Marks Allocated
• Gives all correct anode observations • Gives all correct cathode observations	$1-2$
Total	**2**

(c) The electrodes in an electrolytic cell are dipped into the **same electrolyte solution (1)**, meaning they **don't need a salt bridge** to separate them into half-cells **(1)**.

Marking Criteria	Marks Allocated
• Operate in the same electrolyte solution/same cell • Therefore don't need a salt bridge to be separated	$1-2$
Total	**2**

7. **[9 marks]**

(a) **Towards anode: Br^-** **Towards Cathode: Cr^{3+}** **(1)**

Observations: On the left side the **surrounding solution** will become a **darker red colour (1)**, and on the **right side** the size of the **silver electrode** will **increase** and the **surrounding solution** will become a **lighter green colour (1)**.

(b) **Towards anode: F^-** **Towards Cathode: Na^+** **(1)**

Observations: On the **left side** a **colourless, odourless gas** will be produced **(1)**, and on the **right side** a **colourless, odourless gas** will also be produced **(1)**.

Point to note: Water will be both oxidised and reduced in this case, in preference to the sodium and fluorine ions.

(c) **Towards anode: I^-** **Towards Cathode: Pb^{2+}** **(1)**

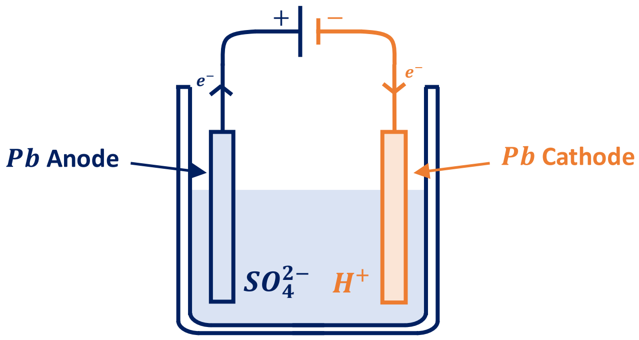
Observations: On the left side the size of the **silver electrode** will **decrease over time (1)**, and on the **right side** the size of the **silver electrode** will **increase** with a **grey coating (1)**.

Point to note: The lead electrode will be oxidised and the lead ions will be reduced

Marking Criteria	Marks Allocated
• Determines correct directions for flow of ions	$1-3$
• Gives all correct anode observations	$1-3$
• Gives all correct cathode observations	$1-3$
Total	**9**

8. **[7 marks]**

(a)



Oxidation half-reaction

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

Reduction half-reaction

$$2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$$

$$Pb_{(s)} + 2H^+_{(aq)} \rightarrow Pb^{2+}_{(aq)} + H_{2\,(g)}$$

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Labels correct ions • Labels positive and negative electrodes • Correct half equations • Determines correct overall cell equation	$1 - 5$
Total	**5**

(b) Tyler should say that the **no voltage** needs to be applied because the reaction is spontaneous **(1)**, as seen from the calculation $E^0 = \mathbf{0.13} + \mathbf{0} = \mathbf{+0.13V}$ **(1)**.

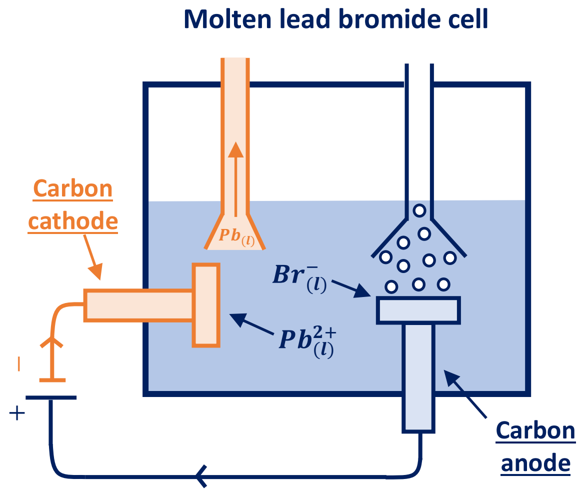
Marking Criteria	Marks Allocated
• No voltage should be applied, reaction is spontaneous • Shows calculation	$1 - 2$
Total	**2**

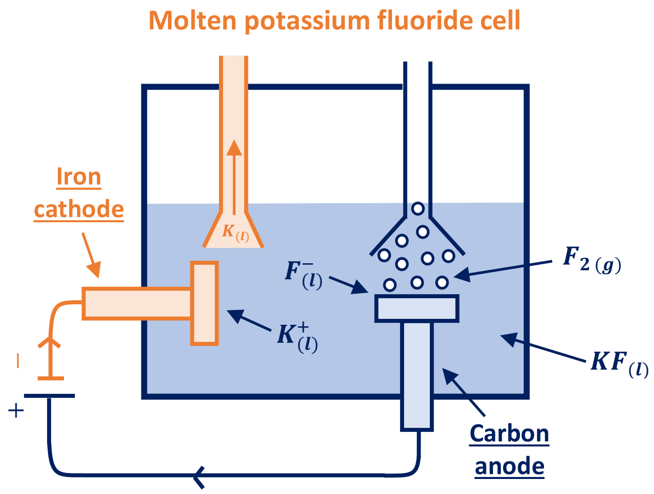
9. **[12 marks]**

(a) These are molten cells, meaning that the lead bromide and potassium fluoride are in a **liquid form** **(1)**, with **no water present**. **(1)**

Marking Criteria	Marks Allocated
• The molten cells are in a liquid state, not aqueous • Water is not present and therefore cannot be oxidised/reduced	$1 - 2$
Total	**2**

(b)





Marking Criteria	Marks Allocated
• Correctly labelled anode and cathode	$1-2$
• Correct direction of electron flow and correct ion flow	$1-2$
Total	4

(c)

Molten lead bromide cell

Oxidation Half-Equation

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

Reduction Half-Equation

$Pb^{2+}_{(l)} + 2e^- \rightarrow Pb_{(l)}$ **(1)**

$2Br^-_{(l)} + Pb^{2+}_{(l)} \rightarrow Br_{2\,(g)} + Pb_{(l)}$ **(1)**

$E^0 = -1.08 - 0.13 = -1.21V$ **(1)**

Molten potassium fluoride cell

Oxidation Half-Equation

$2F^-_{(l)} \rightarrow F_{2\,(g)} + 2e^-$

Reduction Half-Equation

$K^+_{(l)} + e^- \rightarrow K_{(l)}$ **(1)**

$2F^-_{(l)} + 2K^+_{(l)} \rightarrow F_{2\,(g)} + 2K_{(l)}$ **(1)**

$E^0 = -2.89 - 2.94 = -5.83V$ **(1)**

Marking Criteria	Marks Allocated
• Correct half equations	$1-2$
• Correct overall equation	$1-2$
• Correct E^0 of cell	$1-2$
Total	6

Concept 2

Real World Electrochemistry – Progressive Questions Answers

Primary, Secondary and Fuel Cells: Q1, Q2, Q3, Q4

1. **[12 marks]**

(a) **Primary cells** such as the **Leclanché cell** are those cells with a **fixed amount** of each **reactant**, and once used up **cannot be recharged** **(1)**. Like primary cells, **secondary cells** such as the **lead-acid accumulator** also have a **fixed amount** of each reactant, but once used up they can be **recharged** by applying an external voltage, unlike in primary cells **(1)**. **Fuel cells** such as the **hydrogen fuel cell**, are different because they have their **reactants continually fed into them** and will therefore continually produce electricity for as long as they have fuel fed into them **(1)**.

Marking Criteria	Marks Allocated
• Primary cells have fixed amounts of reactants and cannot be recharged • Secondary cells have fixed amount of reactants and can be recharged • Fuel cells have reactants continually fed into them	$1-3$
• One correct example of each cell	1
Total	**4**

(b)

(i) The Leclanché cell is classified as a primary cell because it has a **fixed amount of reactants** and **cannot be recharged** **(1)**.

Marking Criteria	Marks Allocated
• The Leclanché cell has fixed amounts of reactants and cannot be recharged	1
Total	**1**

(ii)

Oxidation Half-Equation

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \quad \textbf{(1)}$$

Reduction Half-Equation

$$2MnO_{2\,(s)} + 2NH^+_{4\,(aq)} + 2e^- \rightarrow Mn_2O_{3\,(s)} + 2NH_{3\,(aq)} + H_2O_{(l)} \quad \textbf{(1)}$$

Marking Criteria	Marks Allocated
• Correct half equations	$1-2$
Total	**2**

(iii)

$$\underset{(0)}{Zn_{(s)}} \rightarrow \underset{(+2)}{Zn^{2+}_{(aq)}} + 2e^-$$

$$2\underset{(+4)}{Mn}\underset{(-2)}{O_{2\,(s)}} + 2\underset{(-3)}{N}\underset{(+1)}{H^+_{4\,(aq)}} + 2e^- \rightarrow \underset{(+3)}{Mn_2}\underset{(-2)}{O_{3\,(s)}} + 2\underset{(-3)}{N}\underset{(+1)}{H_{3\,(aq)}} + \underset{(+1)}{H_2}\underset{(-2)}{O_{(l)}}$$

Based on the oxidation numbers we can deduce that $MnO_{2\,(s)}$ is the **oxidant** **(1)**, and $Zn_{(s)}$ is the **reductant** **(1)**.

(iv) **Advantages** of the Leclanché cell include but are not limited to being: **cheap**, **portable** and **suitable** for powering small devices with a **low required voltage** **(1)**. **Disadvantages** of the Leclanché cell include but are not limited to: **cannot be recharged**, **need constant replacement**, the zinc casing **slowly dissolves** and can **cause a leak**, and the cell **contains toxic contents** **(1)**.

Marking Criteria	Marks Allocated
• Gives both advantages and disadvantages • States three correct advantages • States three correct disadvantages	$1-3$
Total	**3**

2. **[12 marks]**

(a) The lead-acid accumulator is classified as a secondary cell because it has a **fixed amount of reactants** and **can be recharged** **(1)**.

Marking Criteria	Marks Allocated
• The lead-acid accumulator has fixed amounts of reactants and can be recharged	1
Total	**1**

(b)

Oxidation Half-Equation

$$Pb_{(s)} + SO_{4\,(aq)}^{2-} \rightarrow PbSO_{4\,(s)} + 2e^{-} \qquad \textbf{(1)}$$

Reduction Half-Equation

$$PbO_{2\,(s)} + SO_{4\,(aq)}^{2-} + 4H_{(aq)}^{+} + 2e^{-} \rightarrow PbSO_{4\,(s)} + 2H_2O_{(l)} \qquad \textbf{(1)}$$

Marking Criteria	Marks Allocated
• Correct half equations	$1-2$
Total	**2**

(c) $E^{0} = 0.36 + 1.69 = +2.05V$ **(1)**

(d) If this was a car battery there would be **six** of these lead-acid accumulators placed in **series** in order to give approximately a **12V battery** **(1)**.

Marking Criteria	Marks Allocated
• Six lead-acid accumulators placed in series	1
Total	**1**

(e) Increasing the H_2SO_4 concentration will **increase** the **rates of oxidation** and **reduction** **(1)**, and therefore **increase** the **emf** produced by the lead-acid accumulator **(1)**.

Marking Criteria	Marks Allocated
• Increase the rates of oxidation and reduction • Increase the cell emf	$1-2$
Total	**2**

(f) The oxidation and reduction of $PbSO_4$ is not a spontaneous process. Therefore the $PbSO_4$ must be in contact with the electrodes so that the electrons can be pulled from the $PbSO_4$ or provided for the $PbSO_4$ in order to allow this process to occur **(1)**.

Marking Criteria	Marks Allocated
• Not a spontaneous process, so it needs to be in contact for the electrons to be pulled/provided and allow the process to occur.	1
Total	**1**

3. **[6 marks]**

Question	Jamie, Dylan and Alexa's answer	Correct Answer (Yes or No)
What **product(s)** is produced from this fuel cell?	H_2O	**Yes**
What is the **half-reaction** occurring at the **anode**?	$H_{2\,(g)} + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^-$	**No**
What is the **half-reaction** occurring at the **cathode**?	$O_{2\,(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}$	**Yes**
What is the **overall** E^0 of this fuel cell?	$E^0 = 0.4 + 1.23 = 1.63V$	**No**
If a OH^- **electrolyte** was used instead of a H^+ **electrolyte**, would the **voltage** be **larger**?	**No**	**Yes**

∴ the trio will **have to come back tomorrow** to try again **(1)**.

Point to note: For statement 2, in an **acidic hydrogen fuel cell** the anode half-reaction is: $H_{2\,(g)} \rightarrow 2H^+_{(aq)} + 2e^-$. For statement 4, the actual E^0 of the cell is: $E^0 = 0 + 1.23 = +1.23V$. For statement 5, they are correct because the voltage produced in both the acidic and basic cell is always $+1.23V$.

Marking Criteria	Marks Allocated
• Correctly determines if Jamie, Dylan and Alexa are correct for each statement	$1-5$
• States they will have to come back tomorrow	1
Total	**6**

4. **[10 marks]**

(a) **Advantages** of the hydrogen fuel cell include but are not limited to: the **only by-product is water**, generates electricity **as long as fuel is supplied**, **different fuels** can be used, converts **chemical energy directly to electrical energy**, and **uses O_2 and H_2 gas** as opposed to non-renewable fossil fuels **(1)**.

Disadvantages of the hydrogen fuel cell include but are not limited to: H_2 gas is highly **explosive** and O_2 gas is highly **flammable**, fuel cells are **expensive** to produce, H_2 gas is currently sourced from **steam reforming** which uses **fossil fuels** and produces **greenhouse gases**, fuel cells require a **constant supply** of reactants to keep operating, fuel cells need **more development** and **infrastructure** for wide-spread implementation **(1)**.

Marking Criteria	Marks Allocated
• States two advantages and two disadvantages	$1-2$
Total	**2**

(b)

Anode half-reaction:

$$H_2 + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^- \quad \textbf{(1)}$$

Cathode half-reaction:

$$O_{2\,(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)} \quad \textbf{(1)}$$

$$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(l)} \quad \textbf{(1)}$$

Marking Criteria	Marks Allocated
• Correct half equations	$1-2$
• Correct overall equation	1
Total	**2**

(c) $E^0 = 0.40 + 0.83 = +1.23V$ **(1)**

(d) Hydrogen gas is **highly explosive** and oxygen gas is **highly flammable** **(1)**, meaning there could be catastrophic damage if it were to crack open and be exposed to an open flame **(1)**.

Marking Criteria	Marks Allocated
• H_2 gas is highly explosive and O_2 gas is highly flammable • If exposed to an open flame there would be a large explosion/fire	$1-2$
Total	**2**

(e) Increasing the OH^- concentration will **increase** the **rates of oxidation** and **reduction** **(1)**, and therefore **increase** the **emf** produced by the hydrogen fuel cell **(1)**.

Marking Criteria	Marks Allocated
• Increase the rates of oxidation and reduction • Increase the emf	$1-2$
Total	**2**

Electrorefining and Electroplating: Q5, Q6

5. **[7 marks]**

(a) Based on the standard reduction potentials table, the species that would be oxidised are **copper**, **zinc**, **nickel**, **magnesium**, **tin** and **lead** **(2)**. However the only species that would be reduced is the **copper ions** **(2)**.

Marking Criteria	Marks Allocated
• Copper, zinc, nickel, magnesium, tin and lead will be oxidised	$1-2$
• Copper ions will be reduced	2
Total	**4**

(b) The anode mud is collected as it contains extremely valuable metals such as gold and silver **(1)**.

Marking Criteria	Marks Allocated
• The anode mud contains metals such as gold and silver, and is therefore extremely valuable	1
Total	**1**

(c) As the **copper ions** have the highest reduction potential of $+0.34V$ **(1)**, they will be reduced in preference to the Ni^{2+} and Pb^{2+} ions which have E^0 values of $-0.13V$ and of $-0.24V$ respectively **(1)**.

Marking Criteria	Marks Allocated
• Copper ions have the highest reduction potential of $+0.34V$ • They will therefore be reduced in preference to the Ni^{2+} and Pb^{2+} ions	$1-2$
Total	**2**

6. **[8 marks]**

(a) In this electroplating an external voltage is used to allow this non-spontaneous process to occur. In this process, the **silver anode** will **oxidise** to produce silver ions in the solution: $Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-$ **(1)**. The silver ions in the electrolyte will then reduce onto the key cathode: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ **(1)**. This process will form a thin coating of silver around the key, giving it a 'silver plated' finish **(1)**.

Marking Criteria	Marks Allocated
• An external voltage is applied to allow this process to occur • Silver will oxidise at the anode: $Ag_{(s)} \rightarrow Ag^+_{(aq)} + e^-$ • Silver ions will reduce at the cathode: $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$ • This will form a thin silver coating around the key	$1-4$
Total	**4**

(b) On the **left-side** the **silver electrode** will **decrease in size** over time **(1)**, and on the **right-side** the key will **increase in size** over time from a **silver coating** **(1)**.

Marking Criteria	Marks Allocated
• Gives all correct anode observations • Gives all correct cathode observations	$1-2$
Total	**2**

(c) This process would be **reversed** and over-time the **iron key** would **decrease in size** **(1)**, and the **silver electrode** would become **coated** with more **silver** (as the silver ions will reduce in preference to the Fe^{2+} ions) **(1)**.

Marking Criteria	Marks Allocated
• Process would be reversed • Key would decrease in size and silver electrode would increase in size from silver coating	$1-2$
Total	**2**

Corrosion: Q7, Q8

7. **[7 marks]**

(a) Adjustments that Tyler could have made include but are not limited to having: **oxygen and water present**, **high ion concentrations** in the water, **acidic conditions**, a **higher temperature**, and a **contact with a metal** with a **higher reduction potential** **(3)**.

Marking Criteria	Marks Allocated
• Gives three examples	$1-3$
Total	**3**

(b)

Stage	Description	Equation
1.	Fe is oxidised to form Fe^{2+} ions, and O_2 gas is reduced to form OH^- ions	**Anode:** $\boldsymbol{Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-}$ **Cathode:** $\boldsymbol{O_{2\,(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}}$ **Overall:** $\boldsymbol{2Fe_{(s)} + O_{2\,(g)} + 2H_2O_{(l)} \rightarrow 2Fe^{2+}_{(aq)} + 4OH^-_{(aq)}}$ **(1)**
2.	The Fe^{2+} and OH^- ions will precipitate to form $Fe(OH)_2$	$\boldsymbol{Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2\,(s)}}$ **(1)**
3.	The $\boldsymbol{Fe(OH)_2}$ will further oxidise to form **iron (III) hydroxide** **(1)**	$4Fe(OH)_{2\,(s)} + O_{2\,(aq)} + 2H_2O_{(l)} \rightarrow 4Fe(OH)_{3\,(s)}$
4.	The $\boldsymbol{Fe(OH)_3}$ will **lose its water** to form **hydrated iron (III) oxide** **(1)**.	$Fe(OH)_{3\,(s)}$ to $Fe_2O_3 \cdot xH_2O_{(s)}$

Marking Criteria	Marks Allocated
• Provides correct equations for stages 1 and 2	$1-2$
• Provides similar explanations to the ones provided for stages 3 and 4	$1-2$
Total	**4**

8. **[14 marks]**

(a) The **first method** they could use would be to **exclude** oxygen and water by using different materials to **cover the iron surfaces** **(1)**. This will **prevent oxygen and water** from coming in contact with the iron and therefore prevent corrosion from occurring **(1)**.

The **second method** they could use would be a **sacrificial anode**, where a metal with a **higher oxidation potential** than iron (e.g. Zn) is **connected** via an external circuit to the iron object **(1)**. In this 'galvanic cell', the zinc will act as the sacrificial anode and be oxidised in preference to the iron and the electrons will build-up at the iron cathode, preventing it from being oxidised **(1)**.

The **third method** they could use is **cathodic protection**, where an **external voltage** is applied to **build-up electrons** at the iron **(1)**. This will cause the iron to act as a **cathode** and therefore **prevent it** from being oxidised **(1)**.

Marking Criteria	Marks Allocated
• Briefly explains the implementation of each method	$1-3$
• Briefly explains how each method will prevent the oxidation of the iron	$1-3$
Total	**6**

(b) By coating the structures with paint, they are preventing **oxygen** and **water** from **coming in contact** with the iron structures **(1)**, which will **prevent the oxidation** of iron from being able to occur **(1)**.

Marking Criteria	Marks Allocated
• The paint prevents water and oxygen from coming in contact with the iron • This prevents the oxidation from occurring	$1-2$
Total	**2**

(c) The zinc pieces will be **oxidised in preference** to the iron, and therefore the zinc pieces will act as 'sacrificial anodes' **(1)**. This can be seen in the equations: $\boldsymbol{Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \quad E^0 = 0.76V}$ and $\boldsymbol{Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^- \quad E^0 = 0.44V}$. The **electrons** will then **build-up** at the iron cathode and **prevent** it from being oxidised **(1)**.

Marking Criteria	Marks Allocated
• The zinc pieces will be oxidised in preference to the iron • Relevant equations used • The build-up of electrons will prevent the iron from being oxidised	$1-3$
Total	**3**

(d) By applying an external voltage to the iron structures, this will cause **electrons to build-up** at the iron cathode **(1)**, which will **prevent** the iron structures from being oxidised **(1)**. Additionally by having an inert electrode as the anode, it will **not be oxidised** and will therefore not need to be replaced **(1)**.

Marking Criteria	Marks Allocated
• The external voltage will build-up electrons in the iron structures • This will prevent the iron structures from being oxidised • The inert electrode will also not be oxidised, so it doesn't need replacing	$1-3$
Total	**3**

Problem Set 9 – Electrochemistry
Repetitive Questions

Concept 1
Galvanic and Electrolytic Cells – Progressive Questions Answers

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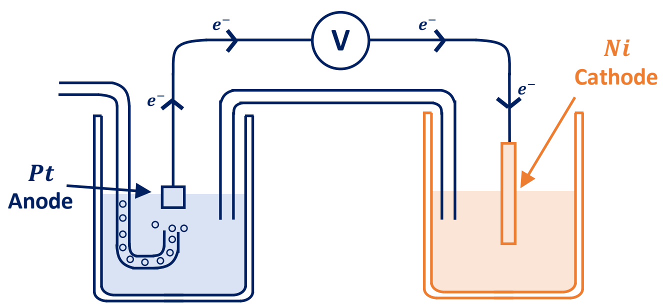
Galvanic Cells: 1.1, 1.2, 1.3, 1.4

1.1 **[10 marks]**

(a) In order to **determine** the **voltage** produced by each half reaction, one of the half reactions needed to be assigned a potential energy of **zero** **(1)**. The hydrogen half-cell was given this potential energy of zero, and could therefore be used to **compare the relative voltages** of all other half-reactions **(1)**. In terms of the **nickel/nickel nitrate cell**, by pairing it with the hydrogen half-cell they will be able to measure the voltage and determine its E^0 value **(1)**.

Marking Criteria	Marks Allocated
• The hydrogen half-cell was assigned a potential energy of zero. • This half-cell was used to compare the relative voltages of all other half-reactions • When paired with the nickel/nickel nitrate cell, they will be able to measure the voltage and determine its E^0 value	$1 - 4$
Total	**4**

(b)



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

Reduction half-cell: $2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$

$$Ni_{(s)} + 2H^+_{(aq)} \rightarrow Ni^{2+}_{(aq)} + H_{2\,(g)}$$

$$E^0 = 0.24 + 0 = +0.24V$$

Marking Criteria	Marks Allocated
• Labels anode, cathode and electron flow correctly	$1 - 3$
• Correct overall equation and E^0 calculation	$1 - 2$
Total	**5**

(c) **Solid Iodine**: $E^0 = 0.54 + 0 = +0.54V$ **(1)**

Gold: $E^0 = 0V$ (this redox reaction is non-spontaneous and would not occur) **(1)**

Lead Sulfate: $E^0 = 0V$ (the oxidation/reduction of lead sulfate is non-spontaneous, so it would not occur) **(1)**

Marking Criteria	Marks Allocated
• Determines correct E^0 values	$1 - 3$
Total	**5**

1.2 **[12 marks]**

(a) In the **left half-cell** a **colourless, odourless gas** will be produced **(1)**, and in the **right half-cell** the **blue solution colour** will **fade** and the size of the **salmon pink electrode** will **increase over time** **(1)**.

(b) In the **left half-cell** the size of the **silver electrode** will **increase** over time **(1)**, and in the **right half-cell** the **deep green solution colour** will **become darker** and the size of the **silver electrode** will **decrease over time** **(1)**.

(c) In the **left half-cell** the **greenish-yellow gas** will dissolve into the **colourless solution** **(1)**, and in the same half-cell a **colourless, odourless gas** will be produced. **No observations** will be made in the **right half-cell** **(1)**.

Point to note: This is a tricky question. As chlorine gas is reduced in the left half-cell and there are no species that can be oxidised in the right half-cell, the water in the solution of the left half-cell will be oxidised in that half-cell. This is because the electrons would rather directly transfer to the water in the left half-cell, then travel round to the right half-cell and oxidise it there.

(d) In the **left half-cell** a **greenish-yellow gas** will be produced **(1)**, and in the **right half-cell** the **pale brown/green solution colour** will **become** more of a **pale brown colour** **(1)**.

(e) In the **left half-cell** a **yellow gas** will dissolve into the **colourless solution** **(1)**, and in the **right half-cell** the size of the **gold electrode** will **increase in size** over time from being coated in a **purple solid** **(1)**.

(f) In the **left half-cell** the **size** of the **silver electrode** will **increase** over time **(1)**, and in the **right half-cell** the **size** of the **silver electrode** will **decrease** over time **(1)**.

Marking Criteria	Marks Allocated
• Gives all correct anode observations	$1-6$
• Gives all correct cathode observations	$1-6$
Total	**12**

1.3 **[10 marks]**

(a) **Anode**: $Pt_{(s)}$ **Cathode**: $C_{(s)}$ **Voltage**: $\mathbf{0+1.08=+1.08V}$ **(2)**

(b) **Anode**: $Zn_{(s)}$ **Cathode**: $Cu_{(s)}$ **Voltage**: $\mathbf{0.76+0.34=+1.10V}$ **(2)**

(c) **Anode**: $Mg_{(s)}$ **Cathode**: $Mg_{(s)}$ **Voltage**: $\mathbf{2.36-0.83=+1.53V}$ **(2)**

Point to note: In this question water will be reduced, and it will be reduced in the same cell as the magnesium. This is because the electrons would prefer to go directly to the water in the magnesium cell rather than travelling through the external circuit.

(d) **Anode**: $Ag_{(s)}$ **Cathode**: $C_{(s)}$ **Voltage**: $\mathbf{-0.80+1.51=+0.71V}$ **(2)**

(e) **Anode**: $Al_{(s)}$ **Cathode**: $Pb_{(s)}$ **Voltage**: $\mathbf{1.68-0.13=+1.55V}$ **(2)**

Marking Criteria	Marks Allocated
• Determines correct anode and cathode	$1-5$
• Determines correct voltage	$1-5$
Total	**10**

1.4 **[14 marks]**

(a) **Similarities** between galvanic and electrolytic cells include but are not limited to:

- Both have an **anode** and **cathode**
- Both have an **oxidised** and **reduced** species
- **Oxidation** occurs at the **anode**, **reduction** occurs at the **cathode**
- **Electrons flow** from the **anode** to the **cathode**
- Both have an **electrolyte**
- Both have an **external circuit**
- Both have **electrodes**

Marking Criteria	Marks Allocated
• States four correct similarities	$1-2$
Total	**2**

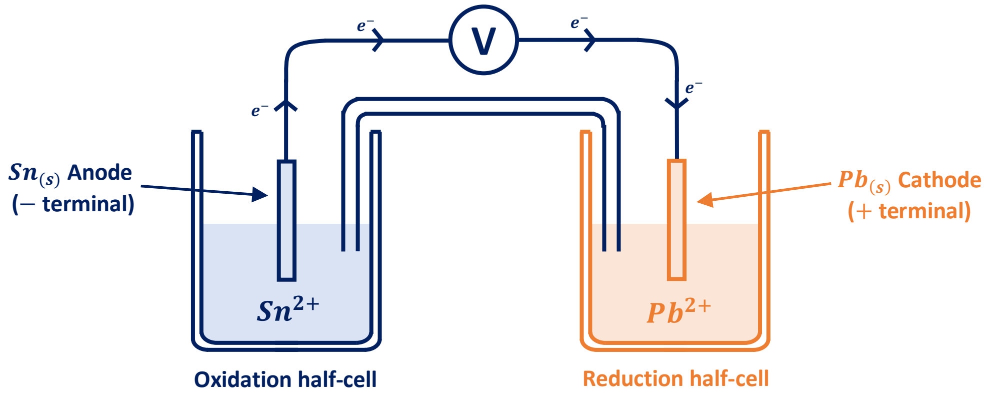
(b) **Differences** between galvanic and electrolytic cells include but are not limited to:

- **Galvanic cells** are **spontaneous**, **electrolytic cells** are **non-spontaneous**.
- **Electrolytic cells** are connected to an **external voltage source**.
- **Galvanic cells** have a **negative anode** and a **positive cathode**, **electrolytic cells** have a **positive anode** and a **negative cathode**.
- **Galvanic cells** have **two cells**, **electrolytic cells** only have **one cell**.
- **Galvanic cells** convert **chemical energy** to **electrical energy**, **electrolytic cells** convert **electrical energy** to **chemical energy**.

Marking Criteria	Marks Allocated
• States four correct differences	$1-2$
Total	**2**

(c)

$$E^0 = 0.14 - 0.13 = +0.01V$$



$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^-$ $\qquad$ $Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

$$\underline{\overline{Sn_{(s)} + Pb^{2+}_{(aq)} \rightarrow Sn^{2+}_{(aq)} + Pb_{(s)}}}$$

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Determines correct cell E^0 • Labels correct ions • Determines correct overall cell equation	$1-4$
Total	**4**

(d) The salt bridge allows for the **controlled flow of ions** across the half-cells but forces the electrons to still flow through an external circuit **(1)**. This **maintains electrical neutrality** and **completes the circuit** **(1)**.

Marking Criteria	Marks Allocated
• Allows for the controlled flow of ions • Maintains electrical neutrality/completes the circuit	$1-2$
Total	**2**

(e) If the concentration of the lead ions was increased, the **rate of reduction** would **increase** **(1)**, so the **emf** of the cell would **increase** **(1)**.

Marking Criteria	Marks Allocated
• Rate of reduction will increase • Cell emf will increase	$1-2$
Total	**2**

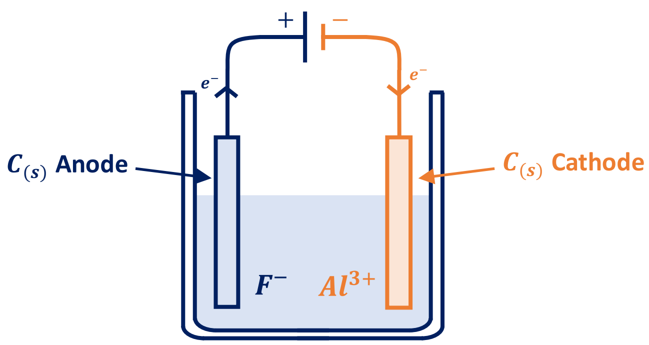
(f) If a gold electrode was used instead of tin, the **lead electrode** would now become the **anode** **(1)** and the **gold electrode** would become the **cathode** **(1)**.

Marking Criteria	Marks Allocated
• Lead electrode becomes the anode • Gold electrode becomes the cathode	$1-2$
Total	**2**

Electrolytic Cells: 1.6, 1.7, 1.9

1.6 **[10 marks]**

(a)



Oxidation half-reaction **Reduction half-reaction**

$$2H_2O_{(l)} \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 4e^- \quad 2H_2O_{(l)} + 2e^- \rightarrow H_{2\,(g)} + 2OH^-_{(aq)}$$

$$6H_2O_{(l)} \rightarrow O_{2\,(g)} + 4H^+_{(aq)} + 2H_{2\,(g)} + 4OH^-_{(aq)}$$

Marking Criteria	Marks Allocated
• Labels correct anode and cathode • Labels correct ions • Labels positive and negative electrodes • Correct half equations • Determines correct overall cell equation	$1-5$
Total	**5**

(b) On the **left side** a **colourless, odourless gas** will be produced **(1)**, and on the **right side** a **colourless, odourless gas** will also be produced **(1)**.

Point to note: Water will be both oxidised and reduced in this case, in preference to the aluminium and fluorine ions.

Marking Criteria	Marks Allocated
• Correct anode observations • Correct cathode observations	$1-2$
Total	**2**

(c) The electrodes in an electrolytic cell are dipped into the **same electrolyte solution** **(1)**, meaning they **don't need** to be separated into half-cells **(1)**.

Marking Criteria	Marks Allocated
• Dipped in same electrolyte solution, therefore don't need separate half-cells	$1-2$
Total	**2**

1.7 **[12 marks]**

(a) On the **left side** the **salmon pink electrode** will **decrease in size** over time **(1)**, and on the **right side** the size of the **salmon pink electrode** will **increase** becoming coated with a **silver coating** **(1)**.

(b) On the **left side** a **colourless, odourless gas** will be produced **(1)**, and on the **right side** a **colourless, odourless gas** will also be produced **(1)**.

Point to note: In this scenario, water will be both oxidised and reduced.

(c) On the **left side** the **grey electrode** will **decrease in size** over time **(1)**, and on the **right side** the size of the **black electrode** will **increase** becoming coated with a **silver coating** **(1)**.

(d) On the **left side** a **colourless, odourless gas** will be produced **(1)**, and on the **right side** a **colourless, odourless gas** will also be produced **(1)**.

Point to note: In this scenario, water will be both oxidised and reduced.

(e) On the **left side** the **silver electrode** will **decrease in size** over time **(1)**, and on the **right side** the size of the **silver electrode** will **increase** becoming coated with a **silver coating** **(1)**.

(f) On the **left side** the **silver electrode** will **decrease in size** over time **(1)**, and on the **right side** the size of the **silver electrode** will **increase** becoming coated with a **purple solid** **(1)**.

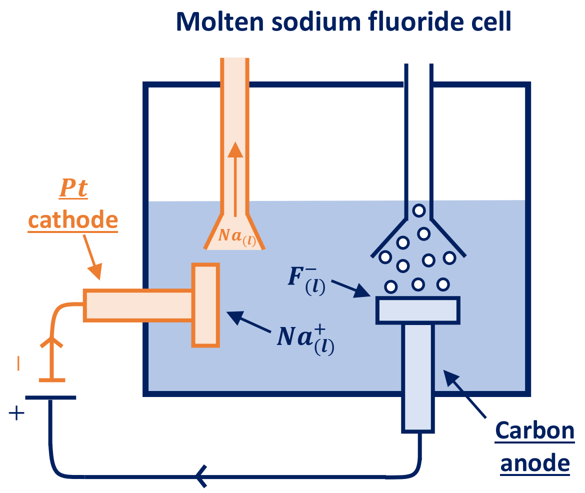
Marking Criteria	Marks Allocated
• Gives all correct anode observations	$1-6$
• Gives all correct cathode observations	$1-6$
Total	**12**

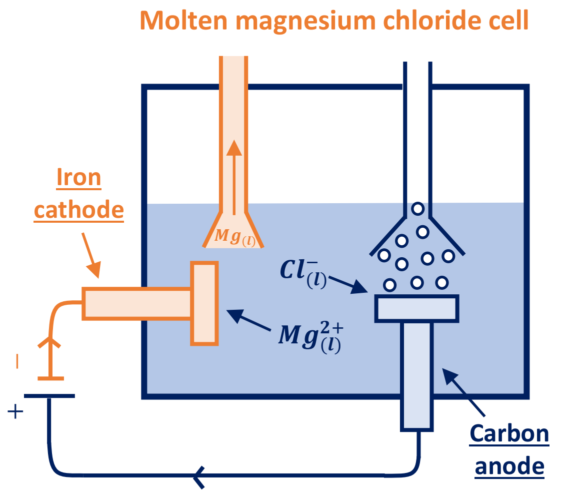
1.9 **[12 marks]**

(a) These are molten cells, meaning that the sodium fluoride and magnesium chloride are in a **liquid form** **(1)**, with **no water present**. **(1)**

Marking Criteria	Marks Allocated
• The molten cells are in a liquid state, not aqueous • Water is not present and therefore cannot be oxidised/reduced	$1-2$
Total	**2**

(b)





Marking Criteria	Marks Allocated
• Correctly labelled anode and cathode	$1-2$
• Correct direction of electron flow and direction of ions	$1-2$
Total	**4**

(c)

Molten sodium fluoride cell

Oxidation Half-Equation **Reduction Half-Equation**

$2F^-_{(l)} \rightarrow F_{2\,(g)} + 2e^-$ $Na^+_{(l)} + e^- \rightarrow Na_{(l)}$ **(1)**

$$2F^-_{(l)} + 2Na^+_{(l)} \rightarrow F_{2\,(g)} + 2Na_{(l)}$$ **(1)**

$$E^0 = -2.89 - 2.71 = -5.60V$$ **(1)**

Molten magnesium chloride cell

Oxidation Half-Equation **Reduction Half-Equation**

$2Cl^-_{(l)} \rightarrow Cl_{2\,(g)} + 2e^-$ $Mg^{2+}_{(l)} + 2e^- \rightarrow Mg_{(l)}$ **(1)**

$$2Cl^-_{(l)} + Mg^{2+}_{(l)} \rightarrow Cl_{2\,(g)} + Mg_{(l)}$$ **(1)**

$$E^0 = -1.36 - 2.36 = -3.72V$$ **(1)**

Marking Criteria	Marks Allocated
• Correct half equations	$1-2$
• Correct overall equation	$1-2$
• Correct E^0 of cell	$1-2$
Total	**6**

Concept 2

Real World Electrochemistry – Progressive Questions Answers

Primary, secondary and fuel cells: 2.1, 2.3

2.1 **[16 marks]**

(a) **Primary cells** such as the **Leclanché cell** **(1)** are those cells with a **fixed amount** of each **reactant**, and once used up **cannot be recharged** **(1)**. Like primary cells, **secondary cells** such as the **lead-acid accumulator** **(1)** also have a **fixed amount** of each reactant, but once used up they can be **recharged** by applying an external voltage, unlike in primary cells **(1)**. **Fuel cells** such as the **hydrogen fuel cell** **(1)**, are different because they have their **reactants continually fed into them** and will therefore continually produce electricity for as long as they have fuel fed into them **(1)**.

Marking Criteria	Marks Allocated
• Primary cells have fixed amounts of reactants and cannot be recharged • Secondary cells have fixed amount of reactants and can be recharged • Fuel cells have reactants continually fed into them	$1-3$
• One correct example of each cell	$1-3$
Total	**6**

(b) **Advantages** of the Leclanché cell include but are not limited to being: **cheap**, **portable** and **suitable** for powering small devices with a **low required voltage** **(1)**. **Disadvantages** of the Leclanché cell include but are not limited to: **cannot be recharged**, **need constant replacement**, the zinc casing **slowly dissolves** and can **cause a leak**, and the cell **contains toxic contents** **(1)**.

Advantages of the lead-acid accumulator include but are not limited to being: **rechargeable**, last for a **long time**, **relatively cheap**, and it **produce large currents** **(1)**. **Disadvantages** of the lead-acid accumulator include but are not limited to being: full of **toxic lead** and **corrosive sulfuric acid**, having a **low energy density**, and having a **limited number** of **discharge-recharge cycles** **(1)**.

Advantages of the hydrogen fuel cell include but are not limited to: the **only by-product is water**, generates electricity **as long as fuel is supplied**, **different fuels** can be used, converts **chemical energy directly to electrical energy**, and **uses O_2 and H_2 gas** as opposed to non-renewable fossil fuels **(1)**.

Disadvantages of the hydrogen fuel cell include but are not limited to: H_2 gas is highly **explosive** and O_2 gas is highly **flammable**, fuel cells are **expensive** to produce, H_2 gas is currently sourced from **steam reforming** which uses **fossil fuels** and produces **greenhouse gases**, fuel cells require a **constant supply** of reactants to keep operating, fuel cells need **more development** and **infrastructure** for wide-spread implementation **(1)**.

Marking Criteria	Marks Allocated
• States two correct advantages of each cell	$1-3$
• States two correct disadvantages of each cell	$1-3$
Total	**6**

(c) In primary cells, they have a **limited amount of reactants** and cannot be recharged **(1)**. Therefore as the reactants are **used up over time**, the **rates of oxidation** and **reduction** will slowly decrease and therefore decrease the emf of the cell over time **(1)**. In secondary cells, they also have a limited amount of reactants, but they can do this at a

slower rate because they **can be recharged** **(1)**. Whilst they can be recharged, the effectiveness of these recharge cycles **slowly starts to decrease** and therefore the maximum emf output of the cell will decrease until the recharging process stops working **(1)**.

Marking Criteria	Marks Allocated
• Primary cells have a limited amount of reactants, which will decrease over time. • This will decrease the rate of oxidation/reduction and therefore decrease the emf of the cell. • Secondary cells do this at a slower rate because they can be recharged. • However, the recharge process slowly loses its effectiveness so the maximum emf output keeps decreasing.	$1-4$
Total	**4**

2.2 **[11 marks]**

(a) In an acidic electrolyte the anode equation is: $\boldsymbol{H_{2\,(g)} \rightarrow 2H^{+}_{(aq)} + 2e^{-}}$ and the cathode equation is: $\boldsymbol{O_{2\,(g)} + 4H^{+}_{(aq)} + 4e^{-} \rightarrow 2H_2O_{(l)}}$ **(1)**, whereas in a basic electrolyte the anode equation is: $\boldsymbol{H_2 + 2OH^{-}_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^{-}}$ and the cathode equation is: $\boldsymbol{O_{2\,(g)} + 2H_2O_{(l)} + 4e^{-} \rightarrow 4OH^{-}_{(aq)}}$ **(1)**. The key difference is that in the **acidic electrolyte** the half-reactions involve the $\boldsymbol{H^{+}}$ **ion** **(1)**, whereas in the **basic electrolyte** the half-reactions involve the $\boldsymbol{OH^{-}}$ **ion** **(1)**.

Marking Criteria	Marks Allocated
• States equations in acidic electrolyte • States equations in basic electrolyte • Explains the acidic equations involve the H^{+} ion • Explains the basic equations involve the OH^{-} ion	$1-4$
Total	**4**

(b) In both the acidic and basic electrolytes the overall reaction equation is: $\boldsymbol{2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(l)}}$ **(1)** and the overall voltage produced by both cells is: $\boldsymbol{E^{0} = 1.23V}$ **(1)**. The reason they both have the same equations is because the $\boldsymbol{H^{+}}$ **and** $\boldsymbol{OH^{-}}$ **ions** end up **cancelling each other out** in the half equations **(1)**, and they both have the **same voltage output** because they are the same overall reaction **(1)**.

Marking Criteria	Marks Allocated
• Shows same overall equation • Shows E^{0} value of $1.23V$ • Both have the same equations because H^{+} and OH^{-} ions cancel each other out • Both have same voltage output because they are the same overall reaction	$1-4$
Total	**4**

(c) The rates of oxidation and reduction would increase, so the emf of the cell would increase **(1)**.

Marking Criteria	Marks Allocated
• Emf of the cell would increase	1
Total	**1**

(d) The International Partnership for Hydrogen and Fuel Cells in the Economy (IPHE) was established to:

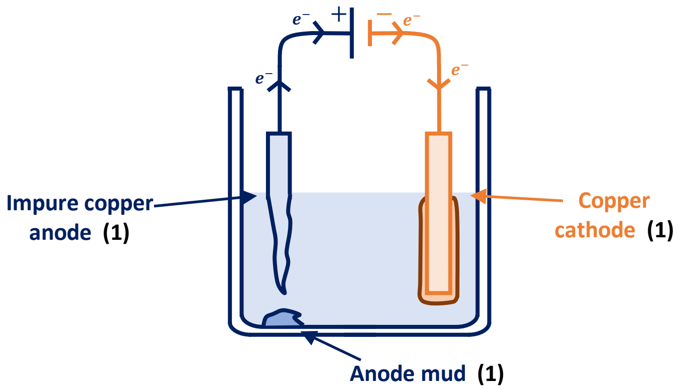
- Encourage the research and development of the hydrogen fuel-cell on a global scale; and
- To develop codes and standards around their development and use.

Marking Criteria	Marks Allocated
• Encourage the research and development of the hydrogen fuel-cell on a global scale; and • To develop codes and standards around their development and use.	$1-2$
Total	**2**

Electrorefining and electroplating: 2.4, 2.5

2.4 **[12 marks]**

(a)



Marking Criteria	Marks Allocated
• Labels anode, cathode and anode mud correctly	$1-3$
Total	**3**

(b) With an E^0 value of $-0.34V$, **copper** will be **oxidised at the anode** **(1)** along with **tin**, **lead**, **chromium**, **magnesium**, **manganese** and **cadmium**, as they have E^0 values of: $\mathbf{0.13V}, \mathbf{0.74V}, \mathbf{2.36V}, \mathbf{1.18V}$ and $\mathbf{0.40V}$ respectively **(2)**. However, at the cathode only the copper will be oxidised spontaneously, with an E^0 value of: $\mathbf{0.34V}$ **(1)**.

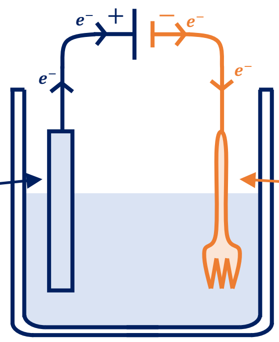
Marking Criteria	Marks Allocated
• Explains copper will be oxidised at the anode • Explains tin, lead, chromium, magnesium, manganese and cadmium will also be oxidised at the anode • Explains only copper will be reduced at the cathode • Uses E^0 values to support explanations	$1-4$
Total	**4**

(c) With **gold** and **silver** having E^0 values of $-1.50V$ and $-0.80V$ respectively, neither of these valuable metals will oxidise into the solution **(2)**. Therefore Tom should collect this anode mud as it will contain valuable metals such as gold and silver **(1)**.

Marking Criteria	Marks Allocated
• Gold and silver will not oxidise in the solution • The anode mud contains valuable metals and should therefore be collected • Uses E^0 values to support explanations	$1-3$
Total	**3**

2.5 **[12 marks]**

(a)



Marking Criteria	Marks Allocated
• Correct drawing/labelling of copper anode and fork cathode	$1-2$
Total	**2**

(b) Any soluble copper solution such as $CuNO_{3\,(aq)}$ or $CuSO_{4\,(aq)}$ **(1)**, with no species that will be oxidised/reduced in preference to the copper

Marking Criteria	Marks Allocated
• A soluble copper solution with no ions that will be oxidised/reduced in preference to copper	1
Total	**1**

(c) The anode is made of the metal that will be used to plate the fork **(1)**, the role of the copper anode is to **replace** the ions in the electrolyte solution so that the electrolysis process can continue **(1)**.

Marking Criteria	Marks Allocated
• Anode is the metal that will coat the fork • Will replace ions in the electrolyte solution as they are plated onto the fork	$1-2$
Total	**2**

(d) On the **left-side** the **salmon pink electrode** will **decrease in size** over time **(1)**, and on the **right-side** the **fork** will **increase in size** over time from a **salmon pink coating** **(1)**.

Marking Criteria	Marks Allocated
• Gives all correct anode observations • Gives all correct cathode observations	$1-2$
Total	**2**

(d) This process would be **reversed** and over-time the **silver fork** would **decrease in size** **(1)**, and the **salmon pink electrode** would become **coated** with a **silver metal** **(1)**.

Marking Criteria	Marks Allocated
• Process would be reversed • Key would decrease in size and silver electrode would increase in size from silver coating	$1-2$
Total	**2**

Corrosion: 2.7, 2.8

2.7 **[12 marks]**

(a) The process of corrosion occurs in four steps.

In step 1, an exposed part of the **iron** is **oxidised** into **Fe^{2+} ions**, and at the same time, but on another part of the surface, the electrons are used to **reduce oxygen** in the presence of **water** to form **OH^- ions**, with the overall equation being: $2Fe_{(s)} + O_{2\,(aq)} + 2H_2O_{(l)} \rightarrow 2Fe^{2+}_{(aq)} + 4OH^-_{(aq)}$ **(1)**.

In step 2, the Fe^{2+} and OH^- ions will react with in the water to form an **iron (II) hydroxide precipitate**: $Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2\,(s)}$ **(1)**. In step 3, if in the presence of oxygen, the $Fe(OH)_2$ will further oxidise to form **iron (III) hydroxide**: $4Fe(OH)_{2\,(s)} + O_{2\,(aq)} + 2H_2O_{(l)} \rightarrow 4Fe(OH)_{3\,(s)}$ **(1)**.

In step 4, if surrounded by **air**, overtime the $Fe(OH)_3$ will **lose its water** to form **hydrated iron (III) oxide ($Fe_2O_3 \cdot xH_2O$)**, which is **rust**. Once formed as **rust**, this **brown** and **flaky material** will easily **peel away** from the surface, **exposing new iron** to undergo this same process **(1)**.

Marking Criteria	Marks Allocated
• Briefly explains each step in general terms and uses an appropriate equation	$1-4$
Total	**4**
Note: Can be lenient with wording, so long as equations are correct	

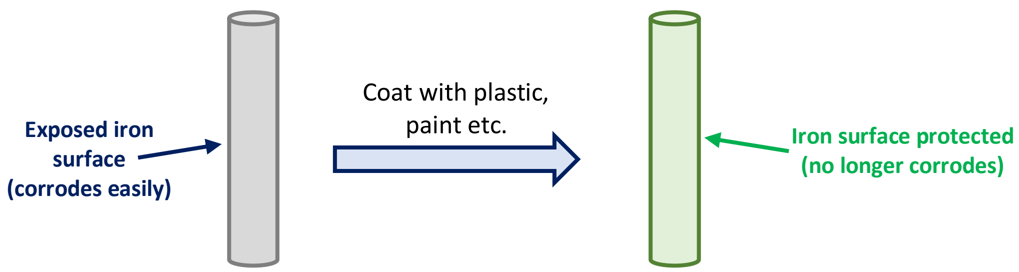
(b) Methods to increase the rate of corrosion include but are not limited to having: **oxygen and water present**, **high ion concentrations** in the water, **acidic conditions**, a **higher temperature**, and a **contact with a metal** with a **higher reduction potential** **(3)**.

Marking Criteria	Marks Allocated
• Briefly discusses three examples	$1-3$
Total	**3**

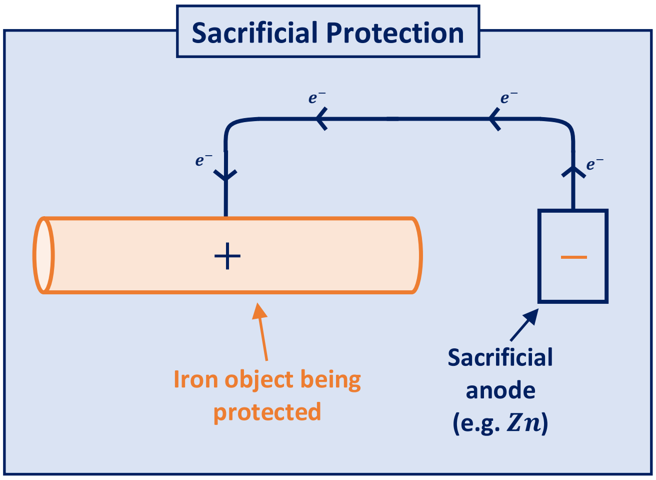
2.8 **[12 marks]**

Can choose two out of the three following methods.

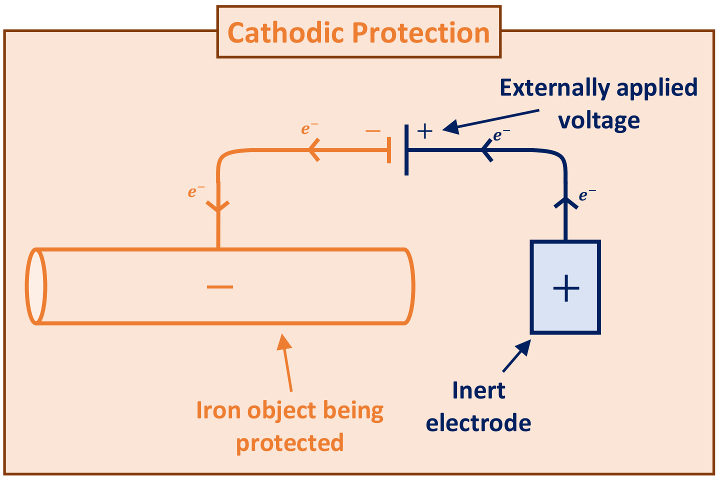
The **first method** that could be used would be to **exclude oxygen and water** by using different materials to **cover the iron surfaces** **(1)**. This will **prevent oxygen and water** from coming in contact with the iron and therefore prevent corrosion from occurring **(1)**. Examples of materials that could be used to coat the iron object include paint, plastic, grease or non-corrosive metals **(1)**, a sample diagram is shown below **(1)**.



The **second method** that could be used would be a **sacrificial anode**, where a metal with a **higher oxidation potential** than iron (e.g. Zn) is **connected** via an external circuit to the iron object **(1)**. In this 'galvanic cell', the zinc will act as the sacrificial anode and be oxidised in preference to the iron and the electrons will build-up at the iron cathode, preventing it from being oxidised **(1)**. Examples of materials that could be used are any metals with a higher oxidation potential than iron and that aren't too reactive, such as zinc or chromium **(1)**, a sample diagram is shown below **(1)**.



The **third method** they could use is **cathodic protection**, where an **external voltage** is applied to **build-up electrons** at the iron **(1)**. This will cause the iron to act as a **cathode** and therefore **prevent it** from being oxidised **(1)**. Examples of inert electrodes that could be used in this process include carbon and platinum **(1)**, a sample diagram is shown below **(1)**.



Marking Criteria	Marks Allocated
• Provides a general overview of the method • Discusses how it prevents the corrosion of iron • Provides examples of materials that can be used for each method • Provides adequate diagram	$(1-4)\times 2$
Total	**8**