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VCE Chemistry ¾
AOS 2 Revision I [2.5]
Workbook Solutions

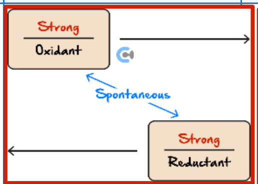
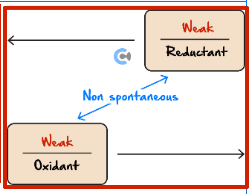
Outline:

Section A: Introduction to Electrolysis (21 Marks)

Sub-Section: Recap

Cheat Sheet

[2.1.1] – Identify differences between galvanic & electrolysis for electrodes, energy conversions, electron flow

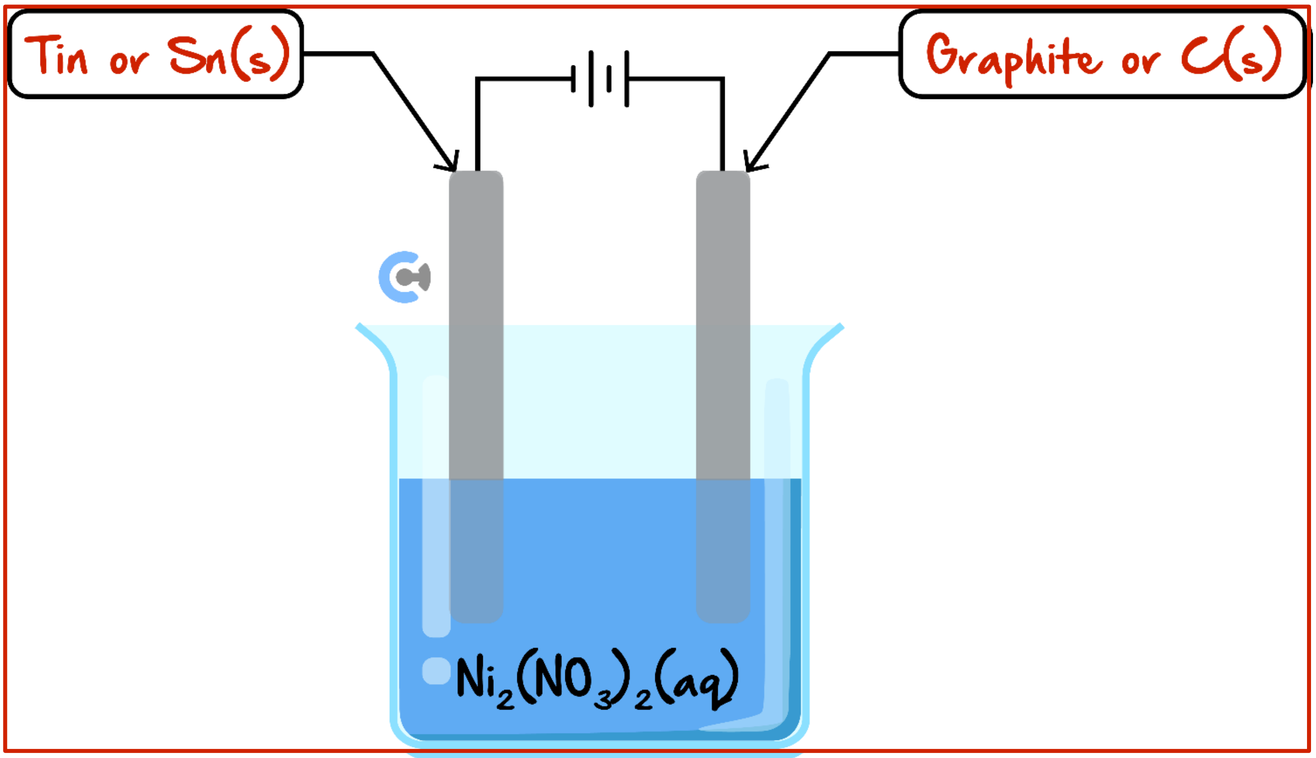
	Galvanic cells	Electrolytic cells
Spontaneous Reaction	[Yes] / [No]	[Yes] / [No]
Energy Conversion	Chemical → Electrical	Electrical → Chemical
Type of Reaction	[Exothermic] / [Endothermic]	[Exothermic] / [Endothermic]
Oxidant / Reductant Relative Strength		
Electron Flow	Anode → cathode	Anode → cathode
Anode	AO −	AO +
Cathode	RC +	RC −
Salt-Bridge / Electrolyte Ion Flow	Cations → [cathode] / [anode]	Cations → [cathode] / [anode]

[2.1.2] – Write equations & calculate EMF required for electrolytic reactions

- When predicting electrolytic reactions, do not forget to include ___water___.
- Metals at the cathode are ___unreactive___.
- Voltage required is ___greater than the difference___

Question 1 (4 marks) **Walkthrough.**

During the electrolysis of $Ni(NO_3)_2(aq)$ with a tin cathode and graphite anode, the following setup is used below.



a. In the boxes provided above, label the material used at each electrode. (1 mark) **[2.1.1]**

b. Write the half-equations which occur at the: (2 marks**) [2.1.2]**

Tin electrode: $\boxed{Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)}$

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

c. State the voltage required to be inputted for the reaction to take place. (1 mark) **[2.1.2]**

$\boxed{> 1.48\ V}$

Space for Personal Notes

Sub-Section: Question Set A

INSTRUCTION: 6 Marks. 5 Minutes Writing.

Question 2 (1 mark) **[2.1.1]**

In an electrolytic cell:

A. Reduction occurs at the positive electrode.

B. Oxidation occurs at the positive electrode.

C. Electrons flow from the negative electrode to the positive electrode.

D. The reaction at the cathode will always involve the plating of a metal.

Question 3 (3 marks)

Copper sulphate ($CuSO_4$) solution undergoes electrolysis using inert electrodes.

a. Write the balanced half-equation for the reaction occurring at the positive electrode. (1 mark) **[2.1.2]**

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

b. Write the balanced half-equation for the reaction occurring at the negative electrode. (1 mark) **[2.1.2]**

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

c. Find the voltage required to be inputted for the reaction to occur. (1 mark) **[2.1.2]**

$$0.34 - 1.23 = -0.89\ V$$
$$> 0.89\ V$$

Space for Personal Notes

Question 4 (2 marks) **[2.1.2]**

For the electrolysis of $Fe(NO_3)_2(aq)$ with copper cathode and silver anode.

Write out the balanced equation for the overall reaction taking place.

$$\boxed{3Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Fe(s)}$$

Space for Personal Notes

Sub-Section: Question Set B

INSTRUCTION: 9 Marks. 9 Minutes Writing.

Question 5 (3 marks)

An electrolytic cell involves some energy being inputted into a solution containing silver nitrate.

a. Write the overall equation. (2 marks) **[2.1.2]**

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

b. State the EMF required for the cell to operate. (1 mark) **[2.1.2]**

$$\boxed{> 0.43\ V}$$

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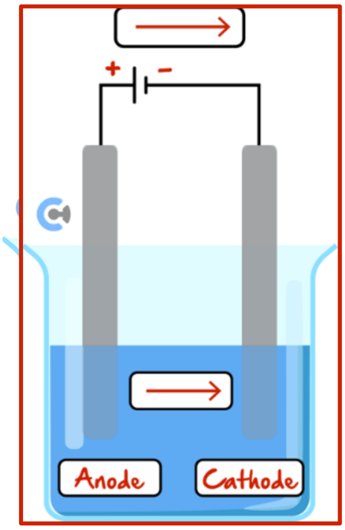
Question 6 (6 marks)

The electrolysis of a solution containing 1.0 *M* copper chloride is undertaken. Inert electrodes are used.



a. In the box provided, label the electrodes as either anode or cathode. (1 mark) **[2.1.1]**

b. Write the balanced half-equations for the:

i. Negative electrode. (1 mark) **[2.1.2]**

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

ii. Positive electrode. (1 mark) **[2.1.2]**

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

c. Label in the diagram above the direction of cations and electrons in the boxes provided. (1 mark) **[2.1.1]**

d. As the reaction proceeds, there are four observations that can be made regarding the solution.

List all **four** observations and state the substances used or produced which are responsible for the observation. (2 marks) **[2.1.2]**

Bubbles formed - O_2 gas.
pH decreased - H^+ produced.
Colour changes from blue to colourless (Cu^{2+} to Cu).
Solid coating at cathode (Cu formed).

Sub-Section: Additional Questions

Question 7 (1 mark) **[2.1.2]**

An electrolytic cell that contains a solution of magnesium nitrate and sodium chloride is electrolysed. The positive terminal of the power source is attached to a gold electrode and the negative terminal is attached to a copper electrode.

Which of the following is true?

A. Both electrodes will have no change in mass/size.

B. No bubbles will be observed.

C. The overall pH increases.

D. The overall pH decreases.

Question 8 (1 mark) **[2.1.2]**

Platinum electrodes are placed in a solution of 1.0 *M* concentration that contains zinc nitrate, $Zn(NO_3)_2$ and sodium chloride, NaCl and is connected to a power source.

Which one of the following statements about the reaction is correct?

A. Zinc metal is produced at the anode and chlorine gas is produced at the cathode.

B. Zinc metal is produced at the cathode and oxygen gas is produced at the anode.

C. Hydrogen gas is produced at the anode and chlorine gas is produced at the cathode.

D. Hydrogen gas is produced at the cathode and oxygen gas is produced at the anode.

Space for Personal Notes

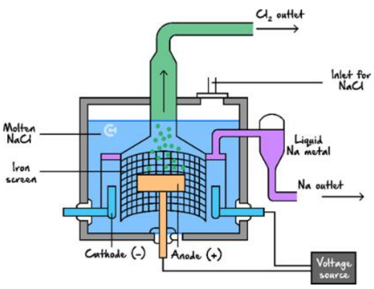
Section B: Features of Electrolytic Cells (29 Marks)

Cheat Sheet

[2.2.1] - Find electrolytic reactions in non-standard conditions (molten & high concentration)

- **High Concentration:**
  - Chloride ions at concentrations greater than $4.0\ M$ concentration become a [stronger] / [weaker] reductant and [react] / [do not] react in preference to water.
  - Sodium ions at concentrations greater than $4.0\ M$ concentration [react] / [do not] react in preference to water.
- **Molten Concentration:** water is not present, and the state of ions is (l) not (aq).

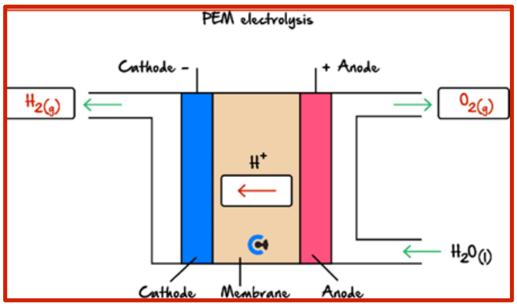
[2.2.2] - Identify features of electrolytic cells & their purpose



- Molten Electrolyte: React species weaker than water.
- Iron at the cathode: Cathode unreactive - cheaper.
- Other Electrolytes (e.g., $CaCl_2$) added: Lower melting point of electrolyte.
- Barrier within the cell: Prevent products from spontaneously re-reacting.
  - Still allow flow of ions
- Products constantly removed:
  - Don't re-react back.
  - No interfere reaction.
- Enclosed container: Prevent O_2 from outside reacting spontaneously.

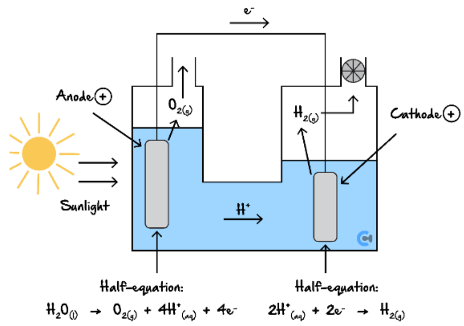
[2.2.3] - Identify key features, write reactions & relate to sustainability & green chemistry principles regarding production of green hydrogen (PEM & artificial photosynthesis)

- Both PEM electrolyser & artificial photosynthesis involve electrolysis of acidic water.
- PEM Electrolyser:



Cathode	Anode
$2H^+(s) + 2e^- \rightarrow H_2(g)$	$2H_2O(l) \rightarrow O_2(g) + 4H^+(s) + 4e^-$

  - Energy Used: Solar/wind energy
  - Green Chemistry Principle: Catalysis.
- **Artificial Photosynthesis:**



  - Energy Conversion: Solar → Chemical.
  - Green Chemistry Principle: Catalysis, design for energy efficient

Question 9 Walkthrough.

A mixture of calcium chloride is electrolysed at high concentrations and is compared to when molten calcium chloride is electrolysed.

a. Write the half-equations that occur when it is electrolysed at **higher concentrations** at the:

i. Cathode. **[2.2.1]**

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

ii. Anode. **[2.2.1]**

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

b. Write the half-equations that occur when it is electrolysed at **molten conditions** at the:

i. Negative electrode. **[2.2.1]**

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

ii. Positive electrode. **[2.2.1]**

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

Space for Personal Notes

Sub-Section: Question Set A

INSTRUCTION: 9 Marks. 9 Minutes Writing.

Question 10 (2 marks)

A molten mixture of lithium bromide is electrolysed using an iron cathode and a graphite anode.

Write the half-equations which occur at the:

a. Positive electrode. (1 mark) **[2.2.1]**

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

b. Negative Electrode. (1 mark) **[2.2.1]**

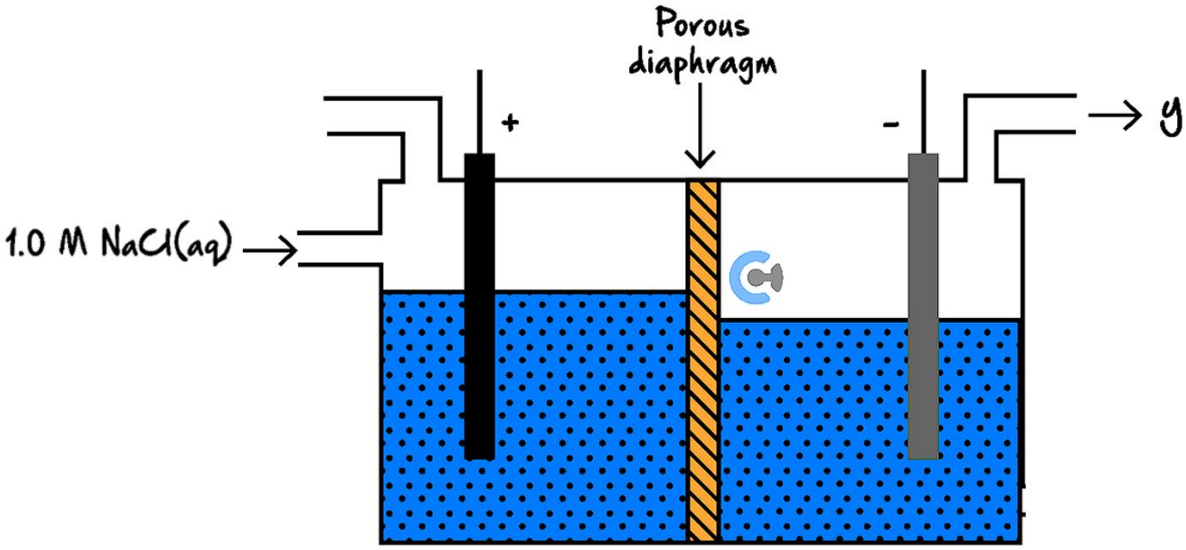
$$\boxed{Li^{+}(l) + e^{-} \rightarrow Li(l)}$$

Space for Personal Notes

Question 11 (7 marks)

Inspired from VCAA Chemistry Exam 2 2004
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem22004.pdf#page=11

A student carries out the electrolysis of a 1.0 *M* solution of sodium chloride using graphite electrodes. The setup is shown below, whereby it is known that gas *y* is formed at the negative electrode as shown.



a. Write an equation for the half-reaction that occurs at the electrode which produces gas *y*. (1 mark) **[2.2.1]**

Marks	0	1	Average
%	52	48	0.5

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Too many students chose to discharge sodium via $Na^+(aq) + e^- \rightarrow Na(s)$.

b. Two different gases are produced at the anode. Write equations for the half-reactions that result in the formation of these two gases.

i. The equation for the half-reaction that produces gas 1. (1 mark) **[2.2.1]**

Marks	0	1	Average
%	33	67	0.7

$2Cl^- \rightarrow Cl_2 + 2e^-$

ii. Equation for the half-reaction that produces gas 2. (1 mark) **[2.2.1]**

Marks	0	1	Average
%	40	60	0.7

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

c. Using the same current and electrodes, the student carries out a second electrolysis, this time of a saturated solution (approximately 6 *M*) of sodium chloride instead of a 1.0 *M* solution. What difference, if any, would you expect in the product or products formed at the following electrodes? (2 marks) **[2.2.1]**

Cathode:

Marks	0	1	Average
%	67	33	0.4

No change

Anode:

ii.

Marks	0	1	Average
%	55	45	0.5

A greater proportion of chlorine evolved (or only chlorine evolved).

d. List **two** functions of the diaphra

Prevent products from coming in contact and re-reacting.
Allow flow of ions.

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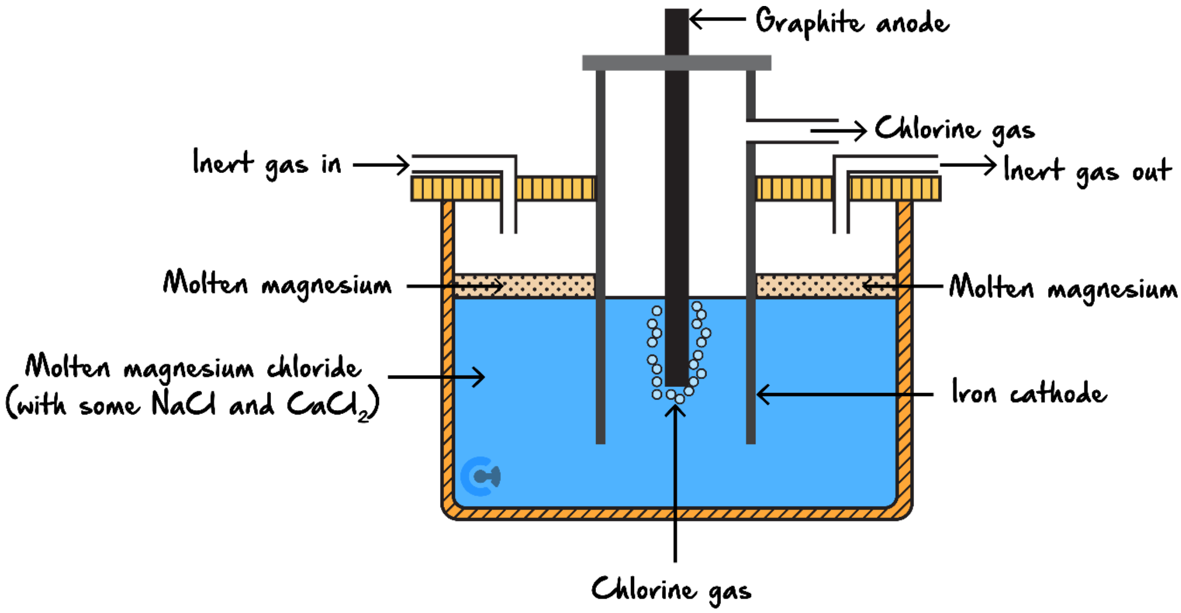
Sub-Section: Question Set B

INSTRUCTION: 15 Marks. 15 Minutes Writing.

Question 12 (8 marks)

Magnesium is one of the most abundant elements on Earth. It is used extensively in the production of magnesium-aluminium alloys. It is produced by the electrolysis of molten magnesium chloride. A schematic diagram of the electrolytic cell is shown below.



The design of this cell takes into account the following properties of both magnesium metal and magnesium chloride:

- Molten magnesium reacts vigorously with oxygen.
- At the temperature of molten magnesium chloride, magnesium is a liquid.
- Molten magnesium has a lower density than molten magnesium chloride and forms a separate layer on the surface.

a. Write a balanced half-equation for the reaction occurring at each of: (2 marks) **[2.2.1]**

Anode: ______________________

Cathode: ______________________

Question 9a.

Marks	0	1	2	Average
%	59	26	15	0.6

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

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

Most students missed out on marks for the use of incorrect states and/or use of equilibrium arrows.

b. Explain why an inert gas is constantly blown through the cathode compartment. (1 mark) **[2.2.2]**

Question 9b.

Marks	0	1	Average
%	78	22	0.2

To prevent:
- molten Mg reacting with oxygen in the air
- contact between Mg and air/oxygen.

c. In this cell, NaCl and $CaCl_2$ are added to the mixture. Propose **one** purpose for the addition of this and how it helps with the operation of the cell. (2 marks) **[2.2.2]**

These ions help weaken the bonds in the magnesium chloride, lowering the melting point. As a result, less energy is required to main molten state.

d. What difference would it make to the half-cell reactions if the graphite anode were replaced with an iron anode? Write the half-equation for any different half-cell reaction. Justify your answer. (3 marks) **[2.2.1]**

Question 9d.

Marks	0	1	2	3	Average
%	56	20	21	3	0.7

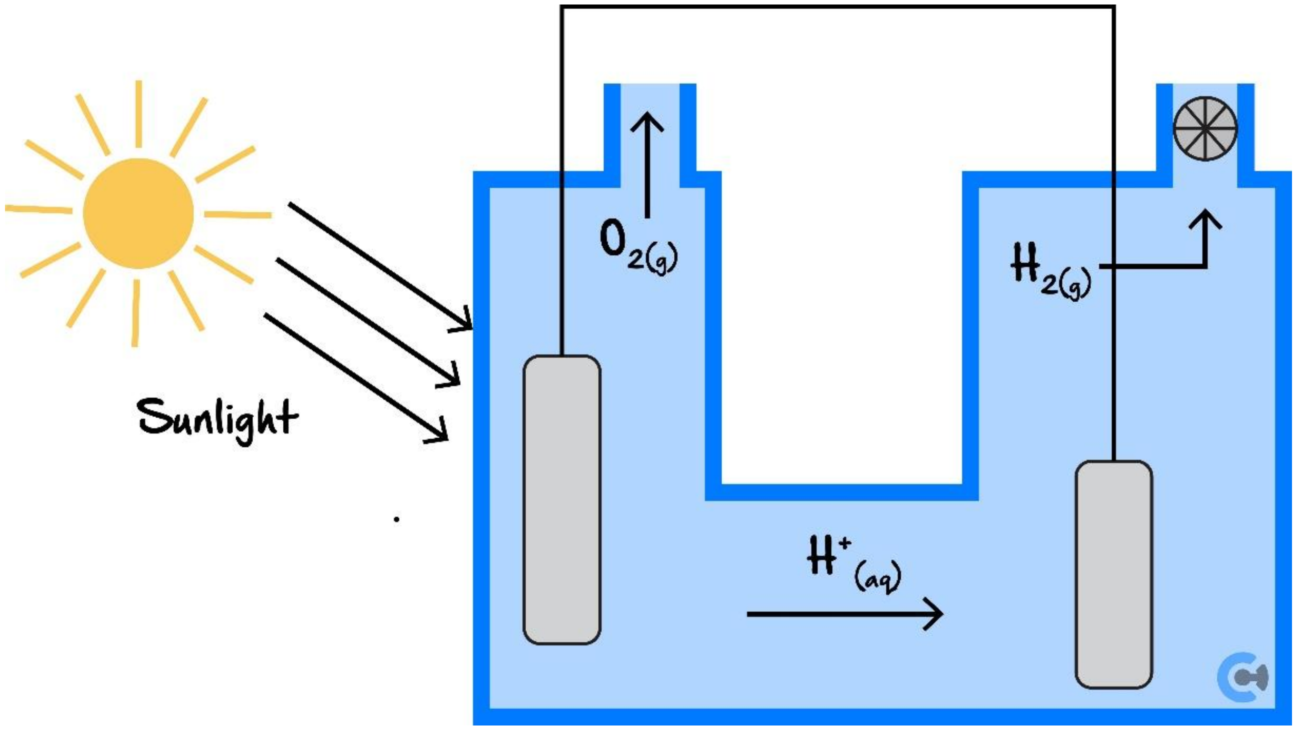
According to the electrochemical series Fe is a stronger reductant than Cl^-.
At the anode, Fe would be oxidised instead of Cl^-/Fe^{2+} would be produced rather than Cl_2.
Half-equation: $Fe(s) \rightarrow Fe^{2+}(l) + 2e^-$
The cations $Fe^{2+}(l)$ would migrate to the cathode/Fe^{2+} is a stronger oxidant than Mg^{2+} hence Fe could be produced/cathode half-equation would be $Fe^{2+}(l) + 2e^- \rightarrow Fe(s)$.

One mark each was awarded for:
- explaining why Fe^{2+} is produced at the anode
- the correct anode half-equation
- explanation of, or half-equation for, production of Fe at the cathode or other valid consequence of the production of Fe at the anode.

Space for Personal Notes

Question 13 (7 marks)

For the artificial photosynthesis cell below:



a. Determine the reaction occurring at the: (2 marks) **[2.2.3]**

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

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

b. The cell produces hydrogen gas which has some safety concerns.

i. State **one** safety precaution which should be taken to mitigate the safety concerns. (1 mark) **[2.2.3]**

Keep away from ignition sources, keep in.

ii. Hydrogen gas is often stored in a pressurised vessel as a liquid. Propose **one** reason why. (1 mark) **[2.2.3]**

Easier to transport / can store more in the same space.

c. By referring to **one** green chemistry principle, explain how the cell produces hydrogen gas to meet society's demands. (2 marks)

Solar energy used → design for energy efficiency → has minimal environmental and economic impacts.

d. State **one** sustainability advantage the artificial photosynthesis cell has, with reference to the United Nations Sustainable Development Goal 12. (1 mark)

It is considered 'responsible' consumption of energy (solar energy) which is renewable source of energy.
or
Considered 'responsible' consumption of reactants (water) which is non-polluting to obtain.
or
Considered 'responsible' production of products, as hydrogen gas is desired, and oxygen gas is non-polluting and has minimal downsides.

Space for Personal Notes

Sub-Section: Additional Questions

Question 14 (5 marks)

An aqueous solution of 1.0 *M* calcium iodide is electrolysed using graphite electrodes.

a. Write the equations for the reaction:

i. At the cathode. (1 mark) **[2.2.1]**

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

ii. At the anode. (1 mark) **[2.2.1]**

$2I^-(aq) \rightarrow 2e^- + I_2(s)$

iii. The electrolysis of this solution is dangerous. Explain why it is dangerous, and identify **two** safety precautions that can be implemented to help mitigate this risk. (2 marks) **[2.2.3]**

Hydrogen gas is produced which is highly flammable. This danger can be mitigated by keeping the cell away from ignition sources and in a well-ventilated area.

b. Write the overall reaction that occurs if molten conditions are used instead. (1 mark) **[2.2.1]**

$2I^-(l) + Ca^{2+}(l) \rightarrow Ca(l) + I_2(l)$

Space for Personal Notes

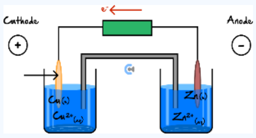
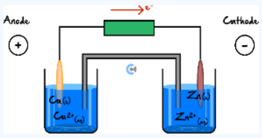
Section C: Secondary Cells & Connected Cells (24 Marks)

Sub-Section: Recap

Cheat Sheet

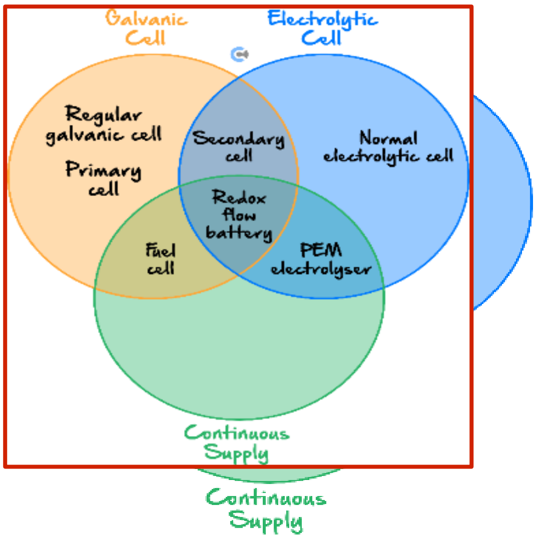
[2.3.1] – Write discharge & recharge reactions in secondary cells

Primary Cells	Secondary Cells
[Rechargeable] / [Non-rechargeable]	[Rechargeable] / [Non-rechargeable]

Discharge (Galvanic)	Recharge (Electrolytic)
	
Electron flow: [Left] / [Right]	Electron flow: [Left] / [Right]
Cathode: RC + Anode: AO –	Cathode: RC – Anode: AO +
Left Electrode Polarity: [+] / [–]	Left Electrode Polarity: [+] / [–]
Left Electrode Type: [Cathode] / [Anode]	Left Electrode Type: [Cathode] / [Anode]

- During discharge/recharge:
  - Polarities [stays same] / [swap].
  - Type of electrode (cathode/anode) [stays same] / [swap].

2.3.2] – Identify factors which affect rechargeability & compare similarities/differences between secondary cells and other cells



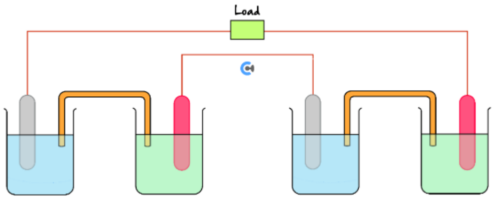
Primary Cell	Secondary Cell
[Rechargeable] / [Non-rechargeable]	[Rechargeable] / [Non-rechargeable]
Can act as [galvanic] / [electrolytic] cell.	Can act as [galvanic] / [electrolytic] cell.
[Chemical to electrical] / [Electrical to chemical].	[Chemical to electrical] / [Electrical to chemical].
[Cheap] / [Expensive]	[Cheap] / [Expensive]

- Reasons for Rechargeability:
  - Products remain in contact with electrodes.
  - No side reactions.
- Reasons for decreased battery life:
  - Products fall off.
  - Overheating.

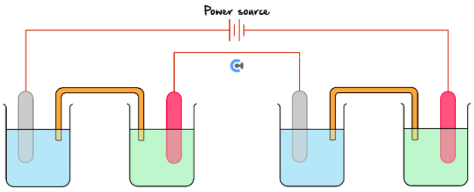
Cheat Sheet

[2.3.3] – Find reactions occurring in connected cells

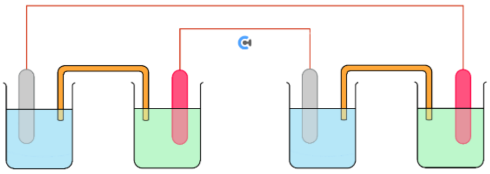
- Connected Galvanic Cells:



- Connected Electrolytic Cells:



- Connected Galvanic - Electrolytic Cells:



- TIPS:
  - First find: Direction of e^- flow.
  - Treat each cell as: Separate.

Question 15 Walkthrough.

Discharge $\rightarrow$ Galvanic | RC $\oplus$ $NiO_2(s) + 2H_2O(l) + 2e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$ AO $\oplus$ | Recharge $\leftarrow$ Electrolytic

AO $\ominus$ $Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2 + 2e^-$ RC $\ominus$

The electrode reactions that occur when the nickel-cadmium battery is producing electrical energy are shown below.

$$NiO_2(s) + 2H_2O(l) + 2e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$$

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

a. Write the balanced half-equation for the reaction which takes place at the negative electrode during discharge. **[2.3.1]**

$$\boxed{Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2(s) + 2e^-}$$

b. Write the balanced half-equation for the reaction which takes place at the positive electrode during recharge. **[2.3.1]**

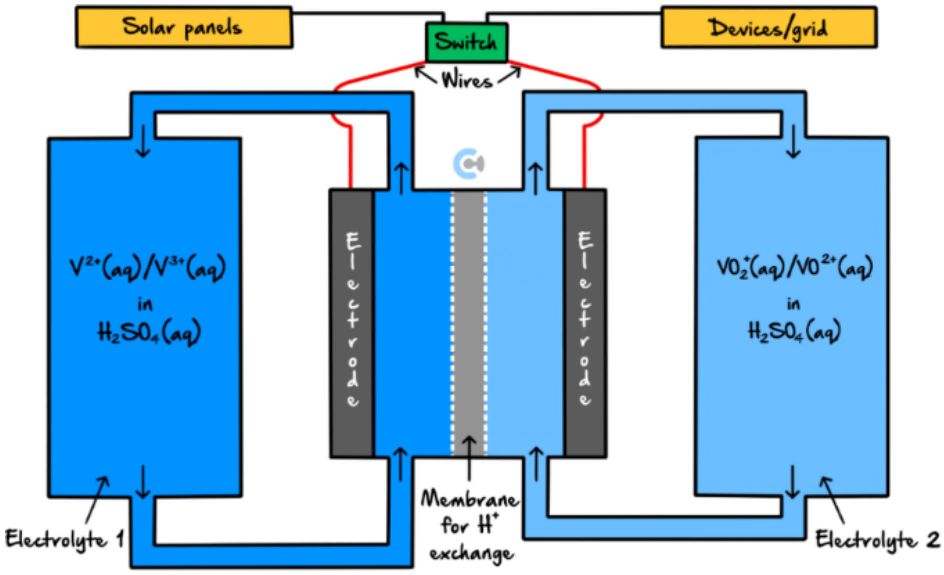
$$\boxed{Ni(OH)_2(s) + 2OH^-(aq) \rightarrow NiO_2(s) + 2H_2O(l) + 2e^-}$$

Space for Personal Notes

Question 16 Walkthrough. [2.3.1]

An increasingly popular battery for storing energy from solar panels is the vanadium redox battery. The battery takes advantage of the four oxidation states of vanadium that are stable in aqueous acidic solutions.

A schematic diagram of a vanadium redox battery is shown below.



The two relevant half-equations for the vanadium redox battery are:

$$VO_2^+(aq) + 2H^+(aq) + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l) \qquad E^\circ = +1.00\ V$$

$$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq) \qquad E^\circ = -0.26\ V$$

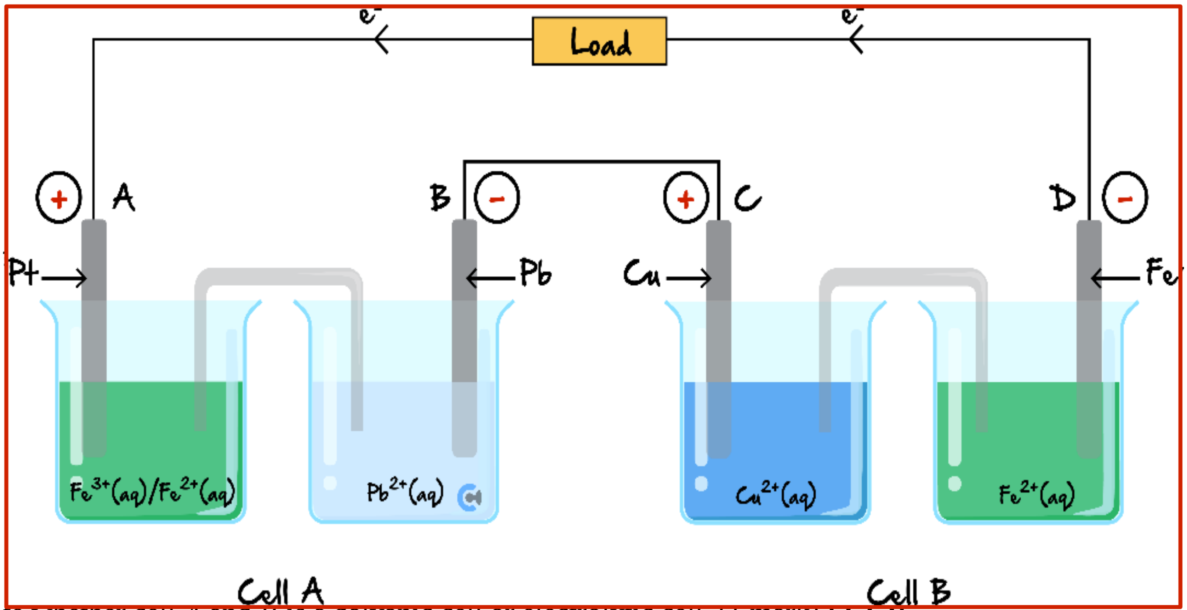
Write the balanced equation for the overall reaction that takes place when the cell is recharged.

$$\boxed{V^{3+}(aq) + VO^{2+}(aq) + H_2O(l) \rightarrow V^{2+}(aq) + VO_2^+(aq) + 2H^+(aq)}$$

Space for Personal Notes

Question 17 (7 marks) **Walkthrough.**

Consider the following:



a. State whether cell *A* and *B* is a galvanic cell or electrolytic cell. (1 mark) **[2.3.3]**

Cell *A*	**Cell *B***
Galvanic	Galvanic

b. Label the polarities of the electrodes in the circles provided above. (1 mark) **[2.3.3]**

c. Write the balanced half-equation for the reaction which takes place at:

i. Electrode A. (1 mark) **[2.3.3]**

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

ii. Electrode B. (1 mark) **[2.3.3]**

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

iii. Electrode C. (1 mark) **[2.3.3]**

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

iv. Electrode D. (1 mark) **[2.3.3]**

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

d. Find the overall EMF produced by the cell. (1 mark) **[2.3.3]**

$$\boxed{0.90 + 0.78 = 1.68\,V}$$

Space for Personal Notes

Sub-Section: Question Set A

INSTRUCTION: 9 Marks. 7 Minutes Writing.

Question 18 (7 marks)

The lead-acid battery is made up of a series of secondary cells in which the following half-reactions are utilised.

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

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

a. During discharge, write the half-equation which occurs at each electrode. (2 marks) **[2.3.1]**

Cathode: $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

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

b. When the battery is discharging state, how the pH changes? (1 mark) **[2.3.1]**

H^+ concentration decreases resulting in a higher pH.

c. State **one** feature of the lead-acid battery which makes it rechargeable. (1 mark) **[2.3.2]**

Products remain in contact with electrode.

d. The reaction which occurs at the anode when the battery is recharging is: (1 mark) **[2.3.1]**

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

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

C. $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2\ (s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

D. $PbO_2\ (s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$

During recharging, the direction of electron flow is reversed (compared to discharging) and electrons are 'forced' to flow from the (+) electrode to the (−) electrode. Since electrons always leave the anode (site of oxidation), then the (+) electrode is the anode during recharging.

The half-reaction occurring at the (+) electrode during discharging was $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$.
So the half-reaction occurring at the (+) electrode during recharging is $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^-$.

e. When the lead-acid battery is recharging the energy transformation occurring is: (1 mark) **[2.3.2]**

A. Chemical → Electrical + Heat.

B. Kinetic → Chemical + Electrical.

C. Electrical → Chemical + Heat.

D. Electrical → Light + Kinetic + Heat.

Discharging is a spontaneous reaction where the oxidant is in the half-equation with the higher $E°$ value. When the half-equations are arranged in order of decreasing $E°$ values, i.e.

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

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

it can be deduced that the reduction of strongest oxidant, $PbO_2(s)$, is accompanied by a decrease in $[H^+]$ and so the pH increases.

f. When recharging the lead-acid battery the positive terminal of the power supply should be connected to the: (1 mark) **[2.3.1]**

A. Positive terminal of the battery where oxidation will occur.

B. Positive terminal of the battery where reduction will occur.

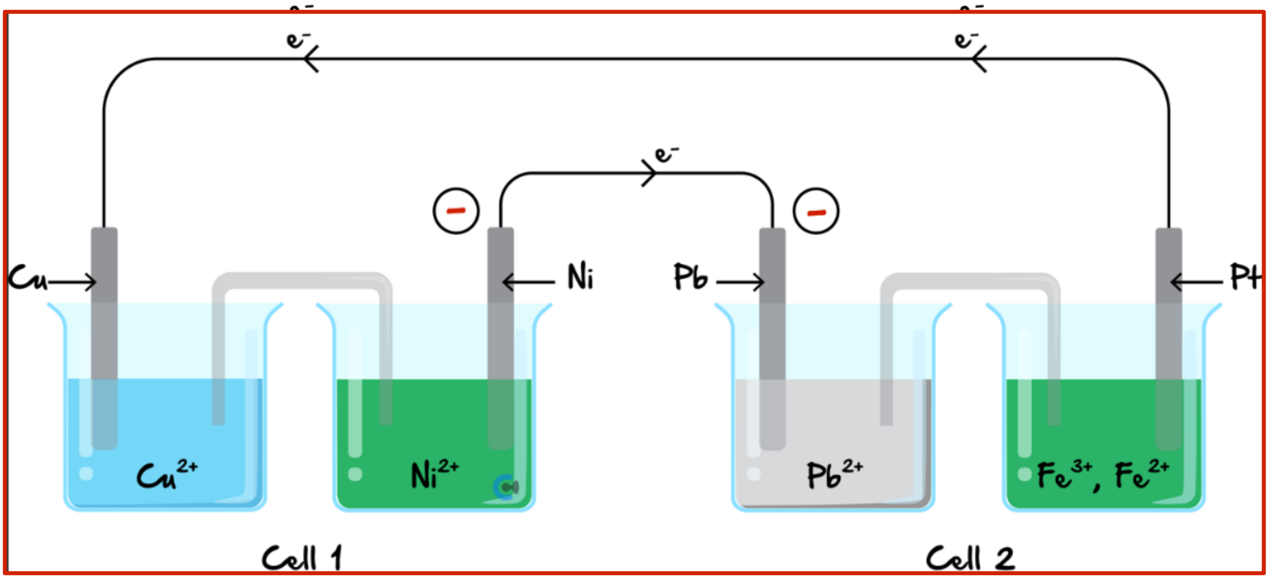
C. Negative terminal of the battery where oxidation will occur.

D. Negative terminal of the battery where reduction will occur.

Space for Personal Notes

Question 19 (2 marks)

The following connected-cell is to be investigated.



a. Label the polarities of the nickel and lead electrode in the circles provided above. (1 mark) **[2.3.3]**

b. The energy transformation occurring in each cell is: (1 mark) **[2.3.3]**

	Cell 1	**Cell 2**
A.	Chemical → Electrical	Chemical → Electrical
B.	Chemical → Electrical	Electrical → Chemical
C.	Electrical → Chemical	Chemical → Electrical
D.	Electrical → Chemical	Electrical → Chemical

Space for Personal Notes

Sub-Section: Question Set B

INSTRUCTION: 8 Marks. 5 Minutes Writing.

Question 20 (7 marks)

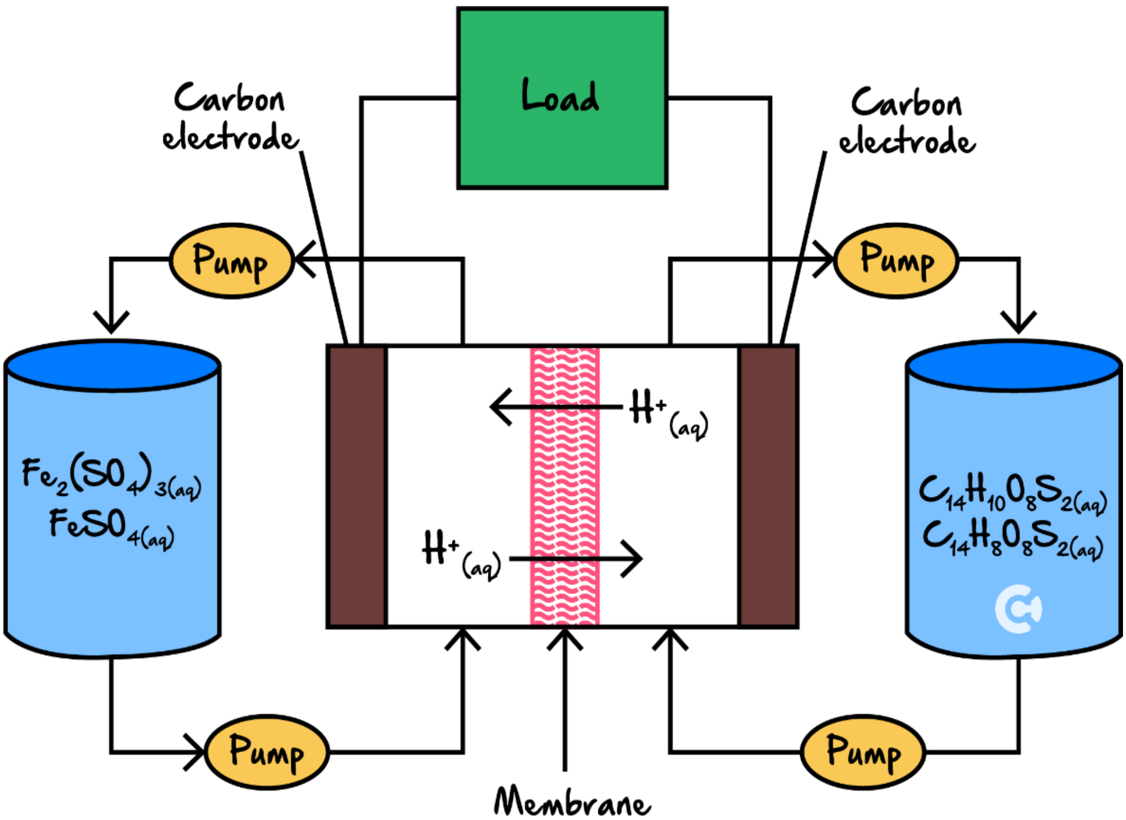
Inspired from VCAA Chemistry Exam 2023
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/2023chemistry-w.pdf#page=24

Scientists are currently researching an experimental secondary cell.

The following reaction takes place in the experimental cell during discharge.

$$Fe_2(SO_4)_3(aq) + C_{14}H_{10}O_8S_2(aq) \rightarrow 2FeSO_4(aq) + C_{14}H_8O_8S_2(aq) + HSO_4^-(aq) + H^+(aq)$$

A diagram of the experimental cell is shown below.



a. State the energy transformations that occur in the experimental cell during discharge. (1 mark) **[2.3.2]**

Chemical energy → Electrical energy.

b. Which reactant is the oxidising agent in the experimental cell during discharge? Use oxidation numbers to justify your answer. (2 marks) **[2.3.1]**

The first mark was awarded for either $Fe_2(SO_4)_3$ or Fe^{3+} as the oxidising agent.

The second mark was awarded for indicating that the oxidation number for iron changed from +3 to +2 during this process and hence why Fe^{3+} was the oxidising agent.

c.

i. Write the half-equation for the reaction that occurs in the $C_{14}H_8O_8S_2/C_{14}H_{10}O_8S_2$ half-cell during recharge. (1 mark) **[2.3.1]**

$$C_{14}H_8O_8S_2 + 2H^+ + 2e^- \rightarrow C_{14}H_{10}O_8S_2$$

ii. State the polarity of the $C_{14}H_8O_8S_2/C_{14}H_{10}O_8S_2$ half-cell electrode during recharge. (1 mark) **[2.3.1]**

Negative

iii. Explain how the polarity of the electrodes is established during recharge to allow the recharge to occur. (2 marks) **[2.3.2]**

The first mark was awarded for the recognition that the electrode polarity does not change when the system is changed from discharge to recharge. The relative reductive and oxidative strength of the chemicals present in the galvanic cell is what determines the polarity, and this is fixed for both the discharge and recharge processes.
The second mark was awarded for the recognition that the external power source causes the process to change at each electrode. This means that during the recharge the positive electrode becomes the anode and oxidation occurs here. This reversal of processes means that the original chemicals can be re-formed and hence the cell is recharged.
Students frequently referred to the electrode polarity 'swapping' during recharge and this is an incorrect statement.

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Question 21 (1 mark) **Additional Question.**

The lead-acid accumulator, used as a common car battery, converts chemical energy into electrical energy via the electrode reactions.

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

$$PbO_2(s) + SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$$

When the lead-acid accumulator is recharge

A. Pb is produced at the negative electrod

B. The pH increases.

C. $PbSO_4$ is produced at the positive elect

D. The changes in the oxidation numbers c

The discharging reactions with electrode signs are

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

(+) $PbO_2(s) + SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$

Therefore the recharging reactions are

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

(+) $PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + SO_4{}^{2-}(aq) + 4H^+(aq) + 2e^-$

Checking the various alternatives

A. Pb is produced at the (-) electrode

B. pH will decrease as H^+ ions are produced at (+) electrode

C. $PbSO_4$ is consumed

D. Oxidation number changes of Pb are from +2 to 0 and +2 to +4

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Section D: Electroplating (21 Marks)

Sub-Section: Recap

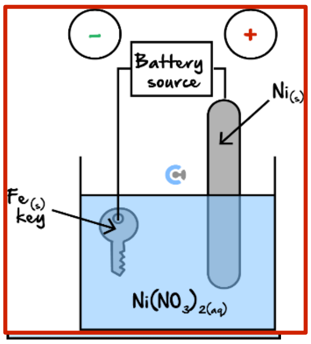
Cheat Sheet

[2.4.1] - Identify the electroplating setup (location of object) & find the electroplating reactions

- Definition: Coat a metal.

Object	Metal Used
[Cathode] / [Anode]	[Cathode] / [Anode]
[Positive] / [Negative]	[Positive] / [Negative]

- Setup:



Anode Reaction	Cathode Reaction
$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$	$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

- Concentration of Electrolyte: Remains constant.
- EMF: $> 0\ V$

[2.4.2] - Find next-order reactions during electrolysis

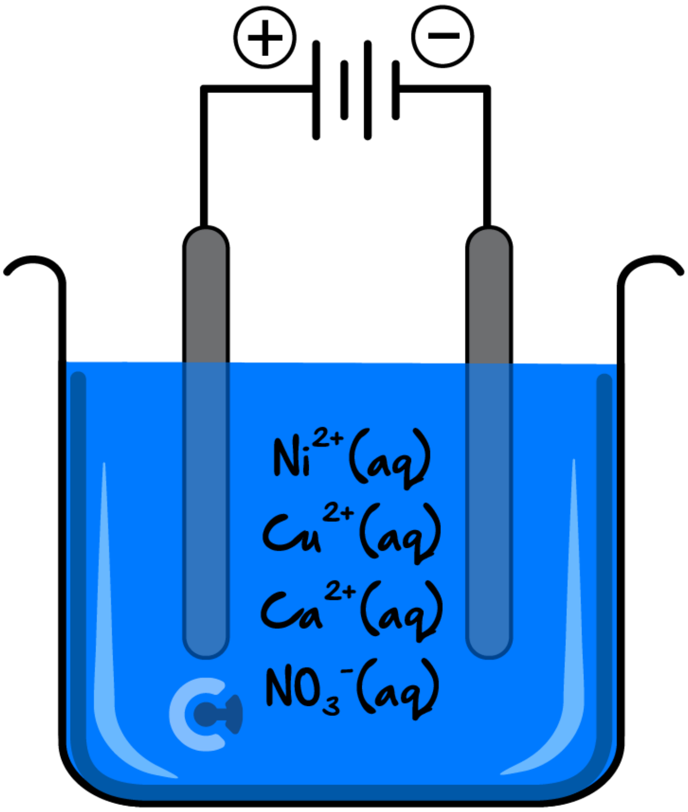
- Assume the current strongest oxidant runs out.
- Move to the next strongest oxidant.
- End game scenarios: Electrolysis of water. Metal oxidises at anode.

[2.4.3] - Apply Faraday's laws to electroplating calculations

- Equations: $Q = It$ $\quad Q = n(e^-)F$
- Typical Steps:
  1. $Q = It$
  2. $n(e^-) = \frac{Q}{F}$
  3. Stoichiometry ratios $n(\text{Zn}) = \frac{1}{2}n(e^-)$
  4. $m = n \times Mr$
- Faraday's First Law: $Q \propto m$
- Faraday's Second Law: Stoichiometry ratio $n(\text{metal}) : n(e^-)$
- **Molar Mass: Charge Ratio (M_r/z)**
  - Use: Compare mass deposited for different metals.
  - **Formula:** $\frac{M_r}{\text{Charge}}$

Question 22 (5 marks) **Walkthrough.**

Consider the electrolysis of the following electrolytic cell, which contains aqueous solutions of 0.1 *M* concentrations of $Ni(NO_3)_2(aq), Cu(NO_3)_2(aq), Ca(NO_3)_2(aq)$ and inert electrodes.



a. Write the balanced half-equation for the reactions which occur at the: (2 marks) **[2.4.2]**

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

Negative electrode: $\boxed{Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)}$

b. After some time has elapsed, the reaction which takes place at one of the electrodes is observed to change.

i. Write the balanced half-equation for the next reaction that takes place at this electrode. (1 mark) **[2.4.2]**

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

ii. Write the balanced half-equation for the next reaction which takes place at this same electrode afterwards. (1 mark) **[2.4.2]**

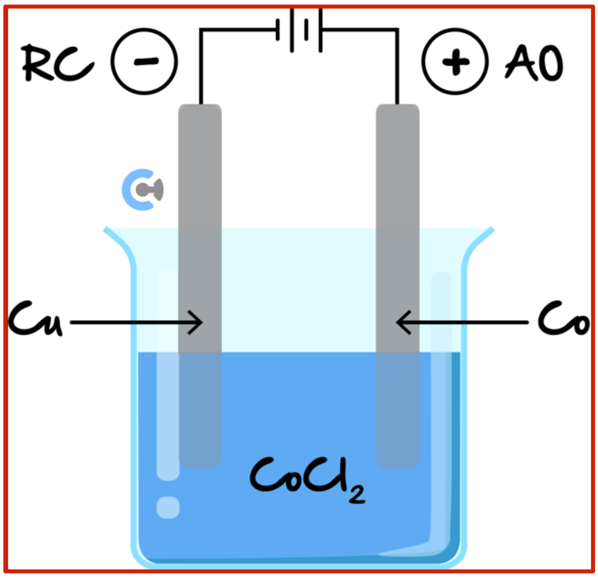
$\boxed{2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)}$

iii. Hence or otherwise, draw the products that form at this electrode. (1 mark) **[2.4.2]**

Question 23 Walkthrough.

Michael wants to electroplate cobalt metal onto his copper key chain. To do so, he attaches the positive terminal of the power source a sheet of cobalt metal, and he attaches the negative terminal to the copper key chain. The electrolyte is comprised of cobalt (II) chloride.

a. Draw the electroplating cell. **[2.4.1]**



b. Find the change in mass of the key chain, if 4.20 *A* of current is passed through for 15.0 minutes. **[2.4.3]**

RC⊖: $Co^{2+} + 2e^- \rightarrow Co$ AO⊕: $Co \rightarrow Co^{2+} + 2e^-$

$Q = It = 4.2 \times 15 \times 60 = 3780\ C$

$n(e^-) = \frac{Q}{F} = \frac{3780}{96500} = 0.0392\ mol$

$n(Co) = \frac{1}{2} n(e^-) = 0.0196\ mol$

$m(Co) = n \times M_r = 0.0196 \times 58.9 = 1.15\ g$

increase by 1.15g

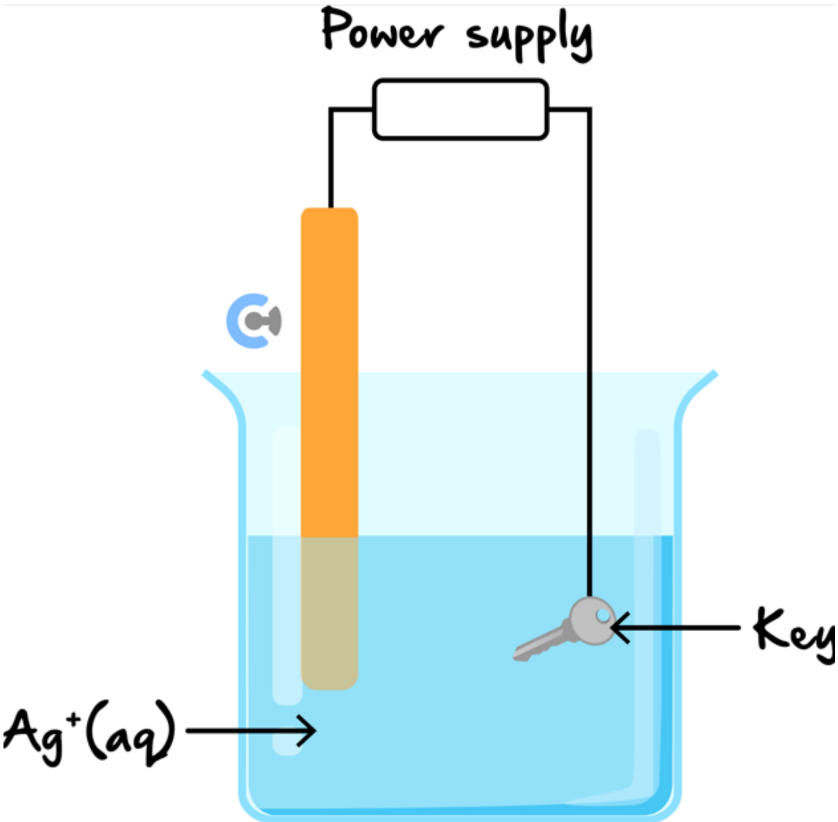
Sub-Section: Questions

INSTRUCTION: 10 Marks. 13 Minutes Writing.

Question 24 (1 mark) **[2.4.1]**

Inspired from VCAA Chemistry Exam 2002
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/chem22002.pdf#page=4

A student decided to silver-plate a locker key with a silver electrode using the apparatus shown:



In this cell, the silver electrode is the:

A. Anode, and is connected to the positive terminal of the power supply.

B. Anode, and is connected to the negative terminal of the power supply.

C. Cathode, and is connected to the positive terminal of the power supply.

D. Cathode, and is connected to the negative terminal of the power supply.

Question 25 (6 marks)

A nickel rod is placed as the anode and a metal key is used as the cathode. A solution of 1.0 M $Ni(NO_3)_2$ is placed in the cell. $Ni^{2+}(aq)$ ions are green in colour.

a. Determine the reaction at the cathode. (1 mark) **[2.4.1]**

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

b. Determine the reaction at the anode. (1 mark) **[2.4.1]**

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

c. After 10 minutes, with a low current of 2.00 A, the beaker is electrolysed.

i. Explain how the colour of the solution changes. (1 mark) **[2.4.1]**

Nickel key from anode is oxidised producing Ni^{2+} whilst simultaneously Ni^{2+} in solution is reduced at cathode forming coating. Hence, Ni^{2+} concentration stays constant and remains same intensity of green.

ii. Find the expected change in mass of the metal key. (3 marks) **[2.4.3]**

$Q = It = 2A \times 10 \times 60 = 1200\ C$

$n(e^-) = \frac{Q}{F} = \frac{1200}{96500} = 0.0124\ mol$

$n(Ni) = \frac{1}{2} n(e^-) = 0.00622\ mol$

$m(Ni) = n \times M_r =$ ...5g

Increases by 0.365 g.

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Question 26 (1 mark) **[2.4.2]**

A solution containing 0.1 mole each of Ag^{+}, Ni^{2+}, Co^{2+} and Mg^{2+} ions was prepared by dissolving their respective nitrates in $1.0\ L$ of deionised water. This solution was then subjected to electrolysis using platinum electrodes. As electrolysis proceeded, the metal ions were sequentially reduced and deposited onto the cathode.

In which order did the metals deposit on the cathode, from the first to the last?

Ag, Co, Ni (Mg does not react as it is lower than water on the ECS.)

Question 27 (2 marks) **[2.4.3]**

The passage of $2960\ C$ of electric charge through a molten vanadium compound yields $0.39\ g$ of vanadium metal. Find the oxidation number of vanadium in the compound.

$2960\ C/96500\ C\ mol^{-1} = 0.0307\ mol$

$0.39\ g$ of V equals $0.39/50.9 = 0.0077\ mol$. Thus, each mole of V must have provided $4\ mol$ of electrons so that the oxidation number was $+4$.

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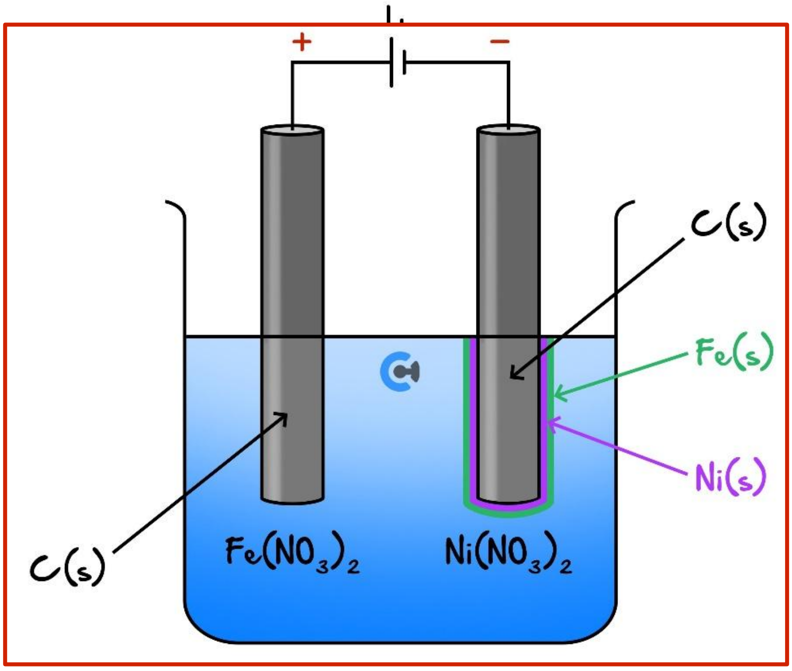
Sub-Section: Additional Questions

Question 28 (5 marks)

Consider an aqueous solution of both iron (II) nitrate and nickel nitrate. The solution is electrolysed using inert electrodes, as shown below:



a. Write the balanced half-equations for the reaction which takes place at the:

i. Cathode. (1 mark) **[2.4.2]**

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

ii. Anode. (1 mark) **[2.4.2]**

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

b. After some time has elapsed, a new equation occurs at one of the electrodes. Write the half-equation for the new reaction which takes place. (1 mark) **[2.4.2]**

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

c. As the cell operates, a coating is seen to form over one of the electrodes. Draw the coating(s) that form at the electrode on the diagram above. (2 marks) **[2.4.2]**

Question 29 (1 mark) **[2.4.3]**

Inspired from VCAA Chemistry NHT 2018
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2018/nht/2018chem-nht-w.pdf#page=11

An electroplating cell containing two platinum electrodes and an electroplating solution is operated at 5.0 A for 600 s. After the cell is turned off, 0.54 g of metal is found to have been deposited on the cathode. Which electroplating solution was used in this process?

A. 1 M $AgNO_3$

B. 1 M $Ni(NO_3)_2$

C. 1 M $Pb(NO_3)_2$

D. 1 M $Cr(NO_3)_3$

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