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VCE Chemistry ¾
Secondary Cells & Connected Cells [0.14]
Workshop Solutions

Error Logbook:

New Ideas / Concepts	Didn't Read Question
Pg / Q #: ____________ Notes:	Pg / Q #: ____________ Notes:
Algebraic / Arithmetic / Calculator Input Mistake	**Working Out Not Detailed Enough**
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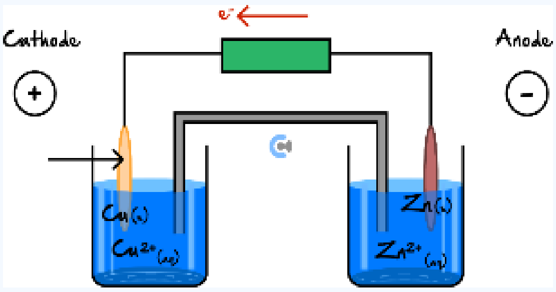
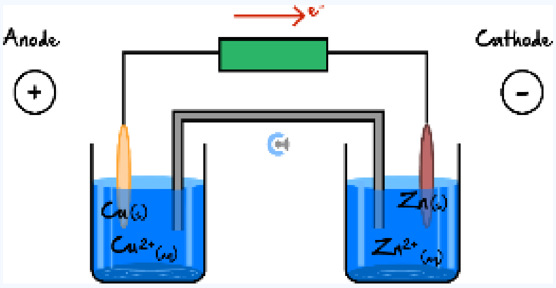
Section A: Recap (3 Marks)

Learning Objective: [2.3.1] – Write Discharge & Recharge Reactions in Secondary Cells

Definition

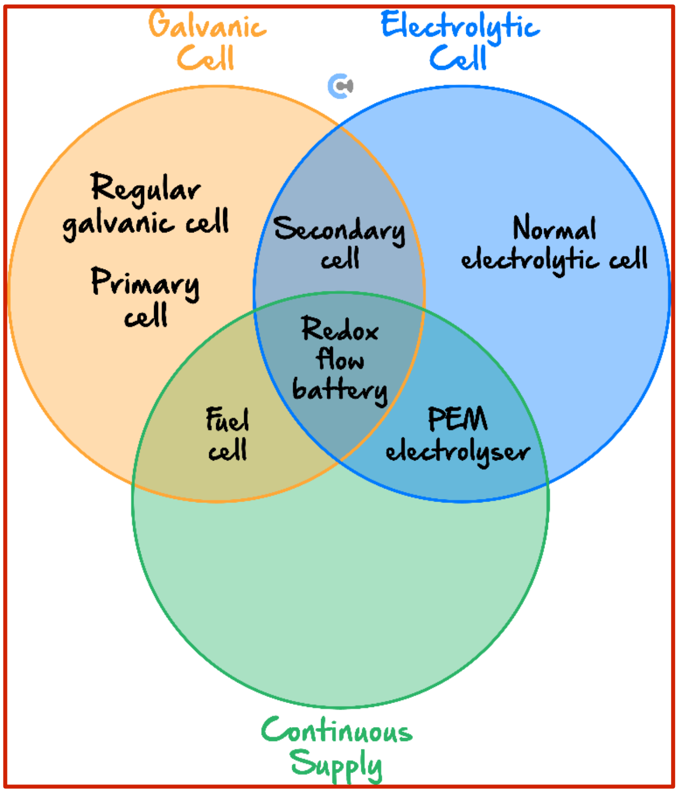
Primary Cells	Secondary Cells
[rechargeable] / [non-rechargeable]	[rechargeable] / [non-rechargeable]

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].

Definition

Learning Objective: [2.3.2] – Identify Factors which Affect Rechargeability & Compare Similarities / Differences between Secondary Cells and Other Cells

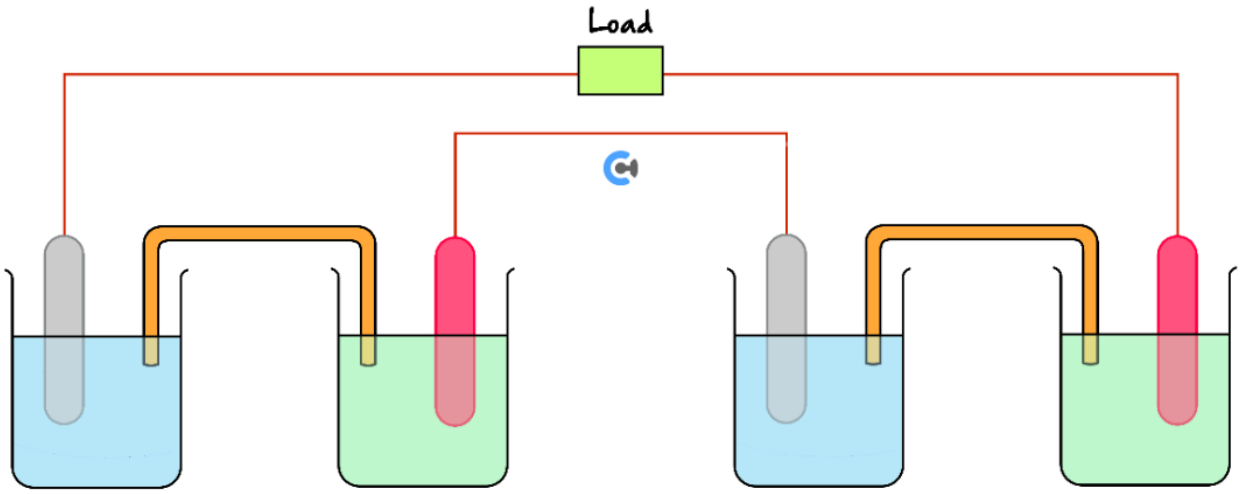


Primary Cell	Secondary Cell
[rechargeable] / [non-rechargeable]	[rechargeable] / [non-rechargeable]
Can act as [galvanic] / [electrolytic] cell.	Can act as [galvanic] / [electrolytic] cell.
[chemical to electrical] / [electrical to chemical]	[chemical to electrical] / [electrical to chemical]
[cheap] / [expensive]	[cheap] / [expensive]

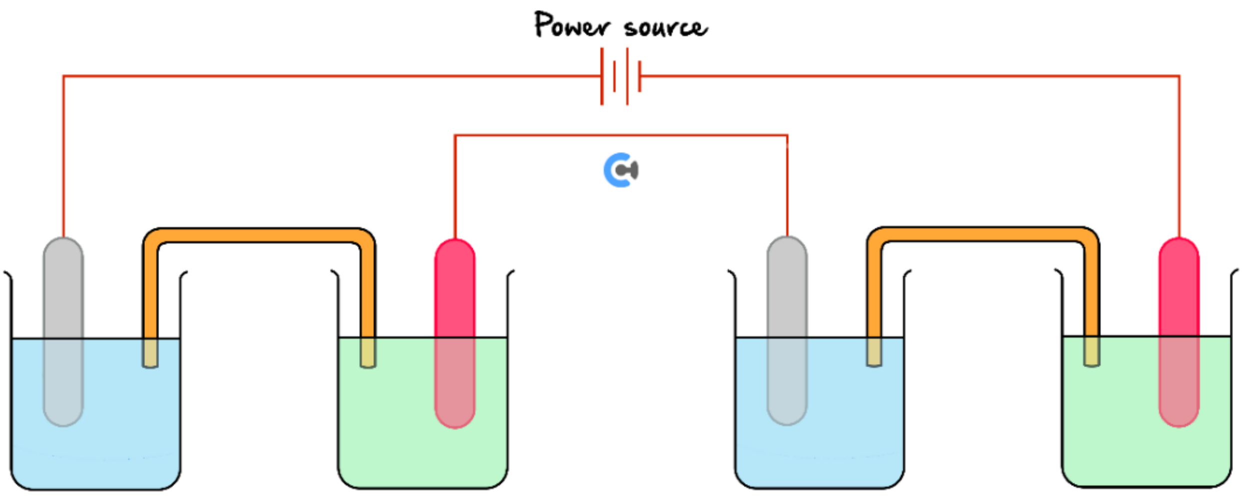
- **Reasons for Rechargeability:**
  - Products remain in contact with electrodes.
  - No side reactions.
- **Reasons for decreased battery life:**
  - Products fall off.
  - Overheating.

Learning Objective: [2.3.3] – Find Reactions Occurring in Connected Cells

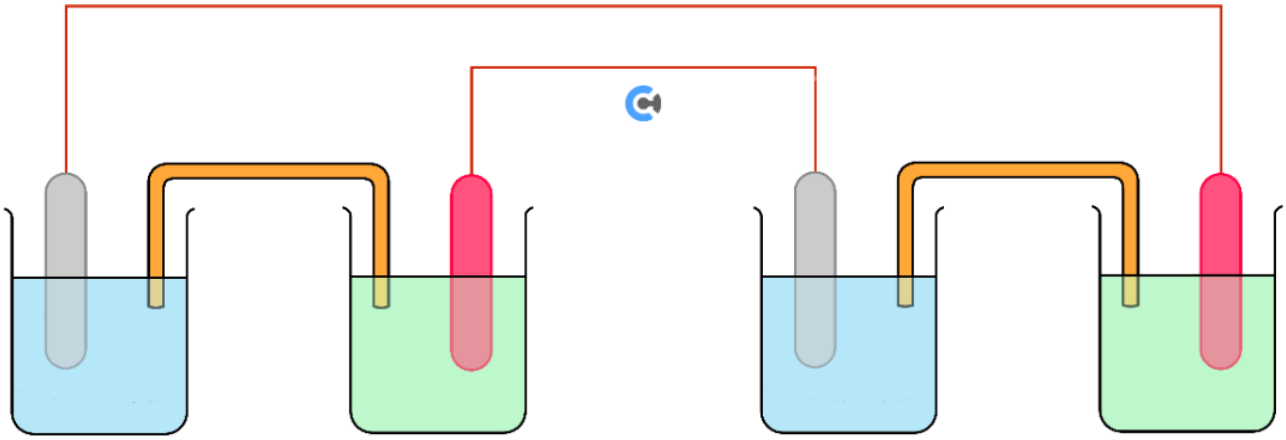
- **Connected** Galvanic **Cells:**



- **Connected** Electrolytic **Cells:**



Connected Galvanic - Electrolytic **Cells:**



TIPS:

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

Question 1 Walkthrough.

Discharge → Galvanic	RC ⊕	$NiO_2(s) + 2H_2O(l) + 2e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$	AO ⊕	Recharge ← Electrolytic
	AO ⊖	$Cd(s) + 2OH^-(aq) \rightarrow Cd(OH)_2 + 2e^-$	RC ⊖	

The electrode reactions which 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.

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

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

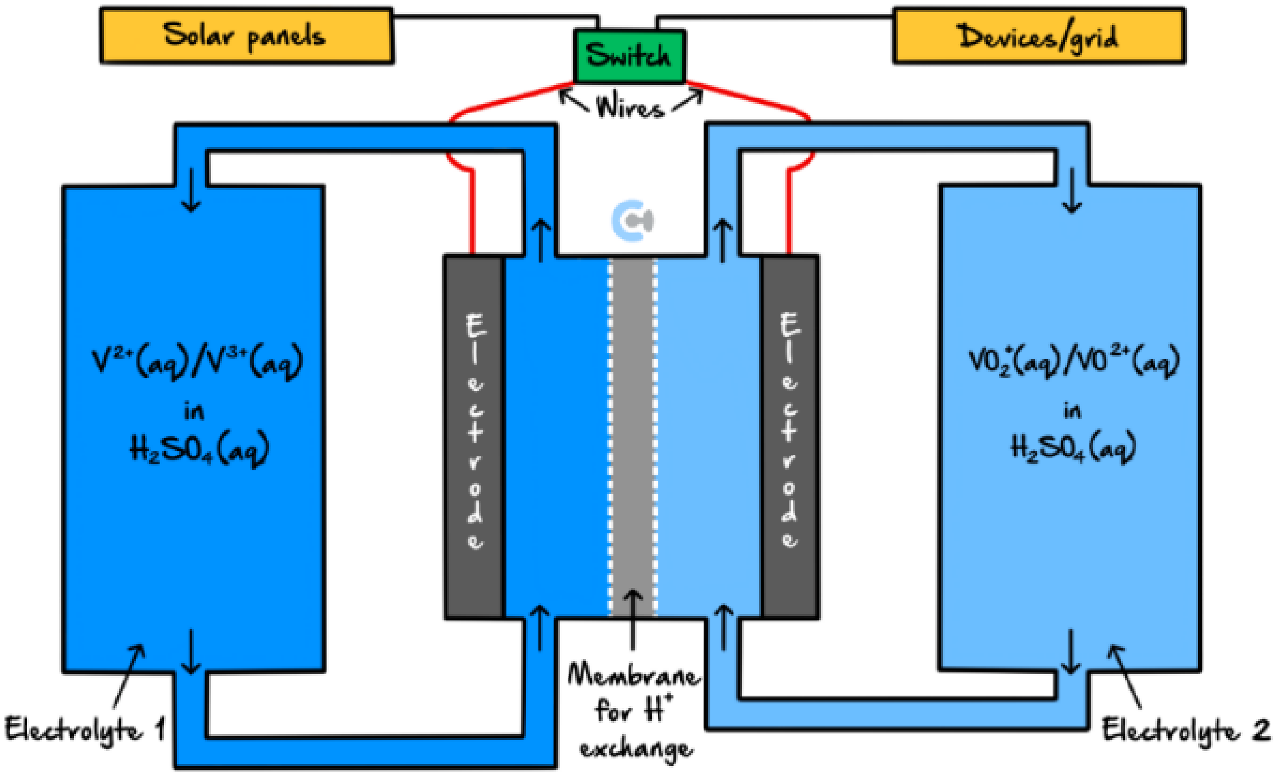
$$Ni(OH)_2(s) + 2OH^-(aq) \rightarrow NiO_2(s) + 2H_2O(l) + 2e^-$$

Space for Personal Notes

Question 2 Walkthrough.

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:

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

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

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

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

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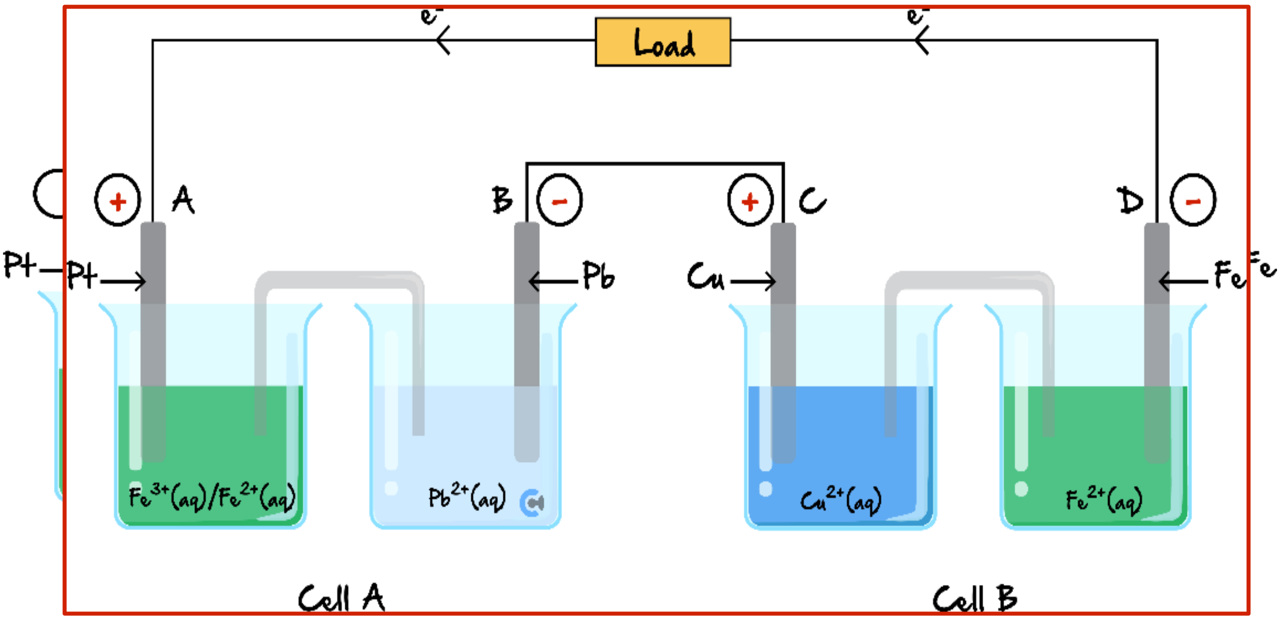
Question 3 (3 marks) **Walkthrough.**

The following connected cell is set up.



a. State whether cell *A* is a galvanic cell or electrolytic cell. (1 mark)

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

b. Label the polarities of the electrodes in the circles provided above. (1 mark)

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

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

Electrode *B*: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$

Electrode *C*: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Electrode *D*: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

d. Find the overall EMF produced by the cell. (1 mark)

$0.9\ V + 1.1\ V = 2.00\ V$

Section B: Warm Up (13 Marks)

INSTRUCTION: 13 Marks. 1 Minute Reading. 8 Minutes Writing.

Question 4 (1 mark)

Rechargeable batteries:

A. Use reversible reactions.

B. Operate as galvanic cells during recharge.

C. Require a continuous flow of reactants to operate.

D. Have fewer side reactions as temperature increases.

Question 5 (2 marks)

Circle what type of reaction takes place at each electrode, and whether it acts as the cathode / anode.

a. During discharge reactions. (1 mark)

Positive electrode	Negative electrode half-equation
[reduction] / [oxidation] half-equation.	[reduction] / [oxidation] half-equation.
[cathode] / [anode]	[cathode] / [anode]

b. During recharge reactions. (1 mark)

Positive electrode	Negative electrode half-equation
[reduction] / [oxidation] half-equation.	[reduction] / [oxidation] half-equation.
[cathode] / [anode]	[cathode] / [anode]

Space for Personal Notes

Question 6 (4 marks)

The rechargeable lithium-ion cell can be used. When the cell is being used, the electrode reactions are represented by the following equations during **discharge**:

$$CoO_2(s) + Li^+(aq) + e^- \rightarrow LiCoO_2(s)$$

$$LiC_6(s) \rightarrow Li^+(aq) + C_6(s) + e^-$$

a. Write the molecular formula for the oxidant as the cell produces electrical energy. (1 mark)

$CoO_2(s)$

b. Write the overall reaction which takes place during discharge. (1 mark)

$LiC_6(s) + CoO_2(s) \rightarrow LiCoO_2(g) + C_6(s)$

c. Suggest a suitable electrolyte for this cell. (1 mark)

$LiNO_3$, LiCl etc

d. Write the balanced half-equation for the reaction which takes place at the negative electrode when the cell is recharging. (1 mark)

$Li^+(aq) + C_6(s) + e^- \rightarrow LiC_6(s)$

Space for Personal Notes

Question 7 (6 marks)

In a lead-acid accumulator, it is known that $Pb(s)$ reacts into $PbSO_4(s)$, and $PbO_2(s)$ reacts into $PbSO_4(s)$ when the battery is discharging.

a. Using oxidation numbers, explain whether the reaction of $PbO_2(s)$ into $PbSO_4(s)$ occurs at the cathode or anode. (1 mark)

Cathode – oxidation number decreases from $+4 \rightarrow +2$, meaning that it is reduction.

b. Given that it uses sulfuric acid as the electrolyte, write the balanced chemical equation for the discharge reaction which takes place at the:

i. Anode. (1 mark)

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

ii. Cathode. (1 mark)

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

c. Write the balanced equation for the overall recharge reaction which takes place. (1 mark)

$$2PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + Pb(s) + H_2SO_4(aq)$$

d. State **two** properties of the lead-acid accumulator that make it rechargeable. (2 marks)

The reactants and products are solid in state, meaning the substances remain in contact with the electrode do not participate in side reactions.

Space for Personal Notes

Section C: Ramping Up (9 Marks)

INSTRUCTION: 9 Marks. 1 Minute Reading. 7 Minutes Writing.

Question 8 (1 mark)

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

The following equations relate to the lead acid battery.

$$PbSO_4(s) + 2e^- \rightleftharpoons Pb(s) + SO_4^{2-}(aq) \qquad E^0 = -0.36\ V$$

$$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightleftharpoons PbSO_4(s) + 2H_2O(l) \quad E^0 = +1.69\ V$$

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

A. The concentration of sulfuric acid decreases.

B. $PbSO_4$ is both oxidised and reduced.

C. The mass of metallic lead decreases.

D. PbO_2 is oxidised to Pb.

Question 9 (1 mark)

The discharge reaction in a vanadium redox battery is represented by the following equation:

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

When the vanadium redox battery is recharging:

A. H^+ is the reducing agent.

B. H_2O is the oxidising agent.

C. VO^{2+} is the reducing agent.

D. VO_2^+ is the oxidising agent.

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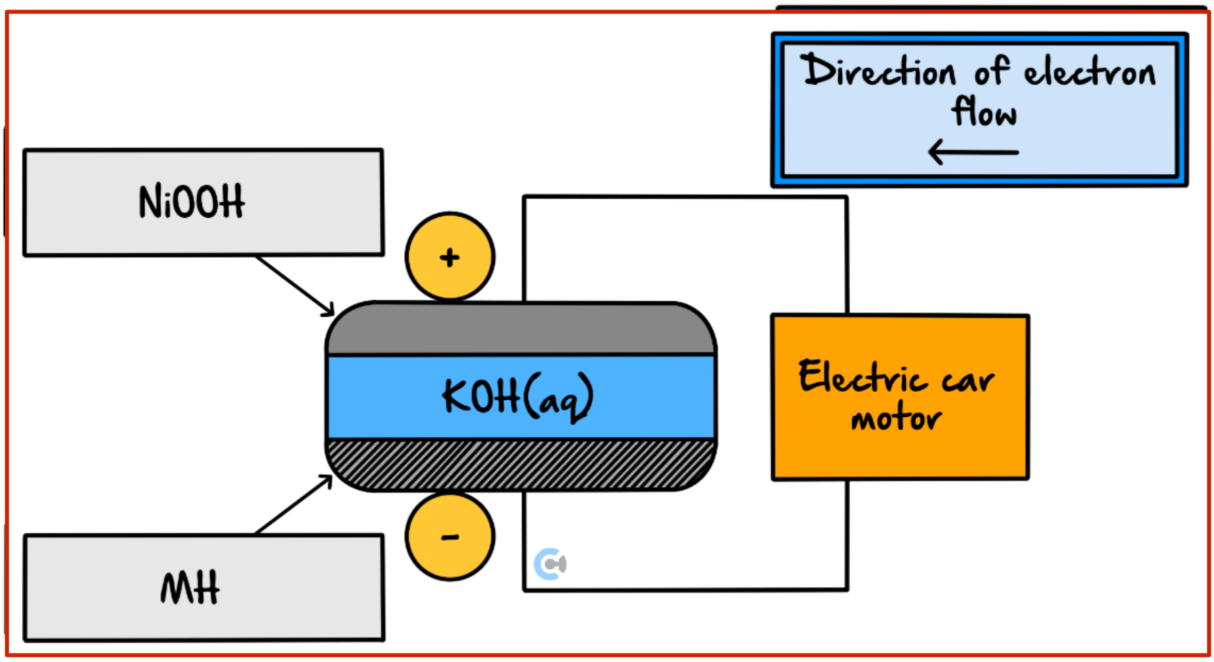
Question 10 (3 marks)

Inspired from VCAA Chemistry Exam 2015
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2015/2015chem-w.pdf#page=37

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



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

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

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

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

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

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

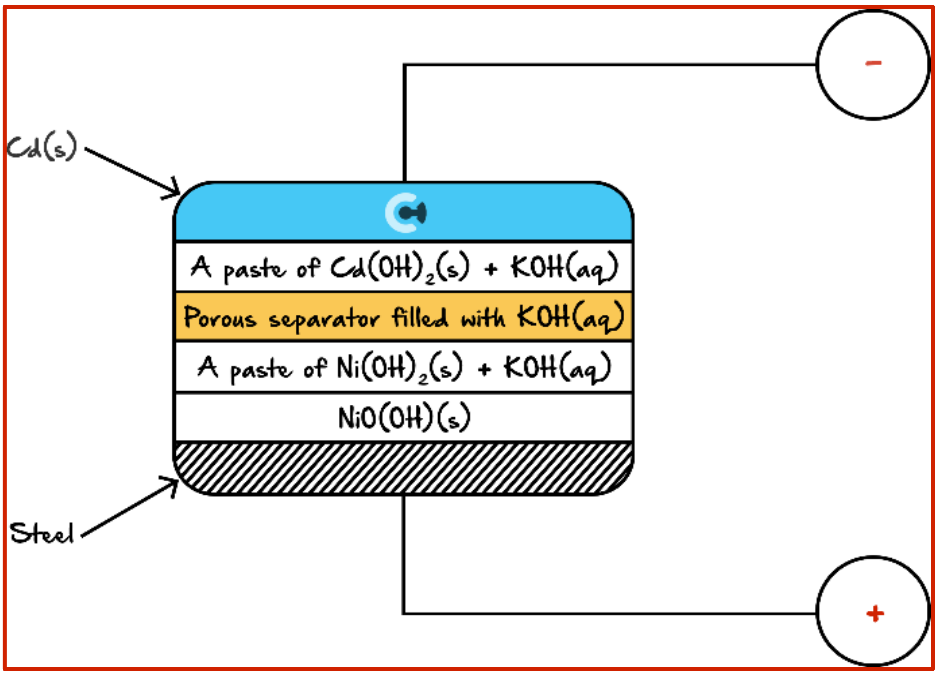
b. In the boxes on the diagram above, indicate which is the MH electrode and which is the NiOOH electrode. (1 mark)

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

Question 11 (4 marks)

Inspired from VCAA Chemistry Exam 2008
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2008chem2-w.pdf#page=21

A rechargeable galvanic cell, also based on nickel and cadmium (NiCd cell), has been commercially available for a number of years and has been used to power small appliances such as mobile phones.
A simplified diagram of a NiCd cell is given below.



The overall cell reaction for the cell when discharging is:

$$Cd(s) + 2NiO(OH)(s) + 2H_2O(l) \rightarrow Cd(OH)_2(s) + 2Ni(OH)_2(s)$$

a. Identify the positive and the negative electrodes by writing '+' or '−' in the circles provided in the diagram. (1 mark)

b. What feature of this secondary cell enables it to be recharged? (1 mark)

The products of the discharge reaction, $Cd(OH)_2$ and $Ni(OH)_2$, stay in contact with the electrodes

c. Give the equation for the half-reaction that takes place at the negative electrode when the cell is discharging. (1 mark)

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

d. Give the equation for the half-reaction that takes place at the electrode connected to the negative terminal of the power supply when the cell is recharging. (1 mark)

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

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Section D: Getting Trickier I (11 Marks)

INSTRUCTION: 11 Marks. 1 Minute Reading. 9 Minutes Writing.

Question 12 (1 mark)

Two types of electrochemical cells are - the primary cell and the secondary cell. Which one of these features is exhibited by only one of these cells?

A. An energy transformation is from chemical energy to electrical energy.

B. The efficiency of the cell is close to 100%.

C. Products of the cell reaction remain in contact with the electrodes.

D. A spontaneous redox reaction is the overall reaction of the cell.

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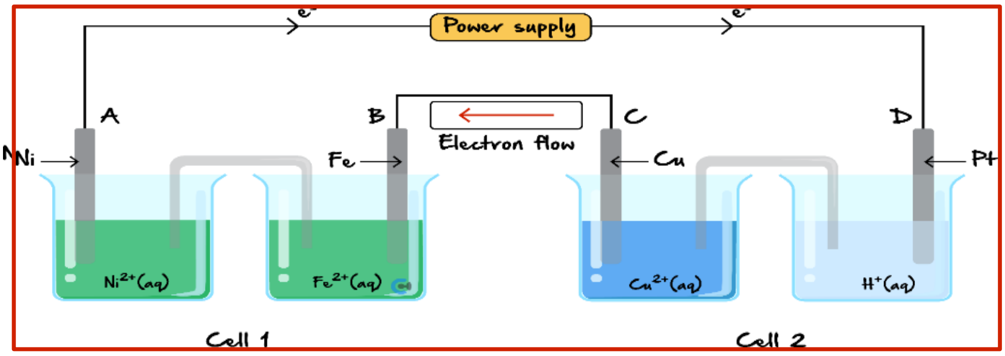
Question 13 (10 marks)

Cells 1 and 2 are connected together to a power supply as depicted in the diagram below.



a. Draw an arrow in the box provided above to indicate the direction of electron flow between electrodes B and C. (1 mark)

b. Cell 2 is to be investigated first.

i. State whether electrodes C and D in cell 2 are a cathode / anode, including the polarity of the electrodes. (1 mark)

Electrode *C*	**Electrode *D***
Anode, Positive	Cathode, Negative

ii. Write the balanced half-equation for the reaction which takes place at electrode C. (1 mark)

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

iii. As the cell reacts, state two observations which can be made at electrode D, and explain why they come about. (2 marks)

Bubbles forming, ($H_2(g)$) forming,
pH increase as it becomes less acidic
$2H^+(aq) + 2e^- \rightarrow H_2(g)$

iv. Find the EMF across cell 2. (1 mark)

$-0.34\ V$

For cell 1, both electrolytes are equally green at the beginning.

c. Write the balanced half-equations for the reaction which occurs at:

i. Electrode A. (1 mark)

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

ii. Electrode B. (1 mark)

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

d. Hence or otherwise, explain whether the electrolyte in electrode A or B will have a greater green intensity after the cell has operated for 2.00 min. (1 mark)

Electrode B – forms $Ni^{2+}(aq)$ which is green, increasing green intensity.

e. In order for the whole cell to operate, find the voltage that needs to be inputted into the cell. (1 mark)

$-0.34\ V + (-0.19\ V) = -0.53\ V$

voltage inputted is $> 0.53\ V$.

Space for Personal Notes

Section E: Getting Trickier II (8 Marks)

INSTRUCTION: 8 Marks. 1 Minute Reading. 7 Minutes Writing.

Question 14 (1 mark)

The silver oxide-zinc battery is rechargeable and utilises sodium hydroxide, NaOH, solution as the electrolyte. The battery is used as a backup in spacecraft if the primary energy supply fails. The overall reaction during discharge is:

$$Zn + Ag_2O \rightarrow ZnO + 2Ag$$

When the silver oxide-zinc battery is being recharged, the reaction at the anode is:

A. $2Ag + 2OH^- \rightarrow Ag_2O + H_2O + 2e^-$

B. $Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$

C. $ZnO + H_2O + 2e^- \rightarrow Zn + 2OH^-$

D. $Zn + 2OH^- \rightarrow ZnO + H_2O + 2e^-$

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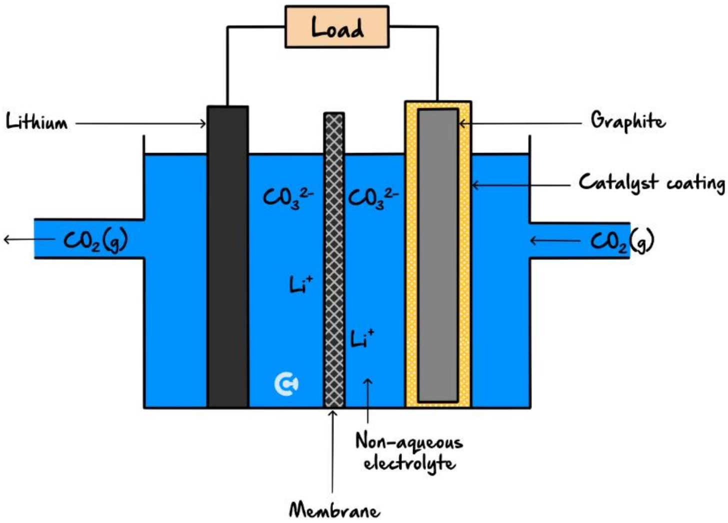
Question 15 (7 marks)

Inspired by VCAA Chemistry 2020 Exam
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2020/2020chem-w.pdf#page=22

Research scientists are developing a rechargeable lithium–carbon dioxide, $Li-CO_2$, battery. The rechargeable $Li-CO_2$ battery is made of lithium metal, carbon in the form of graphite (coated with a catalyst) and a non-aqueous electrolyte that absorbs CO_2.

A diagram of the rechargeable $Li-CO_2$ cell is shown below. One $Li-CO_2$ cell generates 4.5 V.



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

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

Marks	0	1	Average
%	59	41	0.4

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

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

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

Marks	0	1	2	Average
%	29	39	32	1.0

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

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

OR

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

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

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

c. ...priate

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

Students used two approaches to respond to this question.
Galvanic cell approach:
$Li(s)$ is a very strong reducing agent (reductant) and will reduce water to produce $H_2(g)$.

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

OR

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

$H_2(g)$ is explosive / flammable / builds up pressure in battery.
OR
The reaction is highly exothermic and may cause battery to catch fire.
Electrolytic cell approach:
Water will be reduced in preference to Li^+ ions producing $H_2(g)$.
At the cathode (−) $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
At the anode (+) $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e$ Overall $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$
Build-up of H_2 / O_2 leading to a potential explosion due to pressure or a potential explosion due to spontaneous combustion.

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

Marks	0	1	Average
%	51	49	0.5

No. The $CO_2(g)$ absorbed during discharge will be released during recharge.
A reason had to be given by the student in order to obtain the mark.

Let's take a BREAK!

Space for Personal Notes

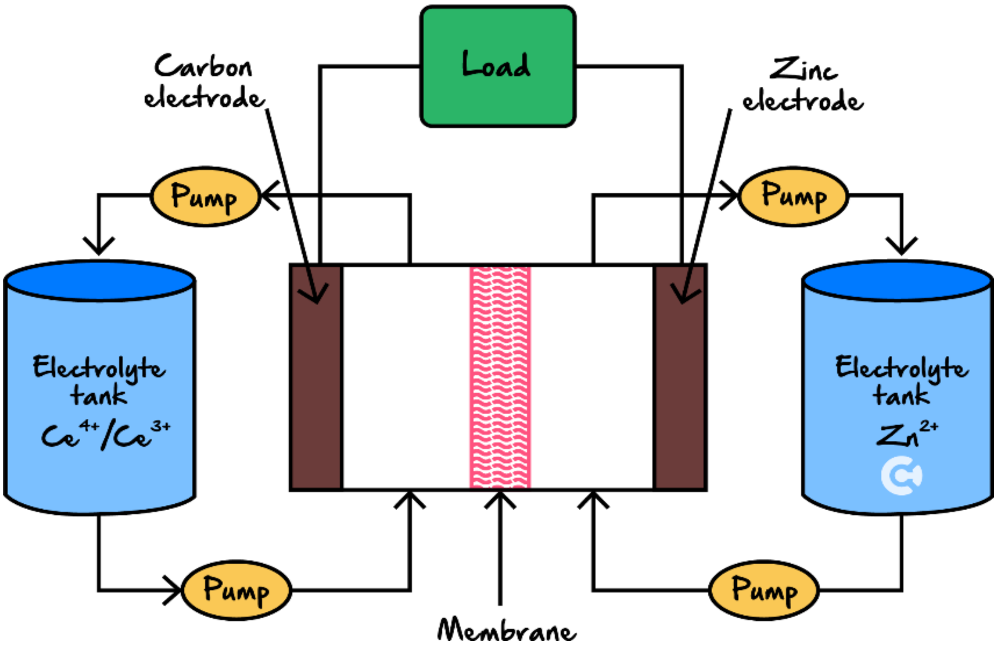
Section F: VCAA-Level Questions I (8 Marks)

INSTRUCTION: 8 Marks. 30 Seconds Reading. 8 Minutes Writing.

Question 16 (8 marks)

Inspired by VCAA Chemistry 2019 Exam
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2019/2019chem-w.pdf#page=30

The zinc–cerium battery is a commercial rechargeable battery that comprises a series of cells. During recharging, the cells use energy from wind farms or solar cell panels. During discharging, energy is supplied to electric grids to power local factories and homes. The electrolytes are stored in separate storage tanks and are pumped into and out of each cell when in use. A membrane separates the two electrodes that are immersed in 1 M methanesulfonic acid, CH_3SO_3H. A diagram representing a zinc–cerium cell is shown below.



The following half-cell reactions occur in the zinc–cerium cell.

$$Zn(CH_3SO_3)_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons Zn(s) + 2CH_3SO_3H(aq) \quad E^0 = -0.76\ V$$

$$Ce(CH_3SO_3)_4(aq) + H^+(aq) + e^- \rightleftharpoons Ce(CH_3SO_3)_3(aq) + CH_3SO_3H(aq) \quad E^0 = 1.64\ V$$

a. Write the equation for the overall discharge reaction. (1 mark)

$Zn(s) + 2Ce(CH_3SO_3)_4(aq) \rightarrow 2Ce(CH_3SO_3)_3(aq) + Zn(CH_3SO_3)_2(aq)$

$Zn(s) + 2Ce^{4+}(aq) \rightarrow Zn^{2+}(aq) + 2Ce^{3+}(aq)$

b. Identify the oxidising agent during discharging and justify your answer using oxidation numbers. (2 marks)

Oxidising agent is $Ce(CH_3SO_3)_4$ / Ce^{4+}

Oxidation number of Ce decreases from +4 [in $Ce(CH_3SO_3)_4$] to +3 [in $Ce(CH_3SO_3)_3$] or $Ce(CH_3SO_3)_4$ is reduced according to

$Ce(CH_3SO_3)_4(aq) + H^+(aq) + e^- \rightarrow Ce(CH_3SO_3)_3(aq) + CH_3SO_3H(aq)$

+4 +3

One mark was awarded for the correct oxidising agent.

One mark was awarded for accurate justification using oxidation numbers.

c. Determine the theoretical voltage produced by a single cell as it discharges. (1 mark)

E_{cell} = E°(oxidant) - E°(reductant)

= 1.64 – (-0.76)

= 2.40 V

Cell voltages are not negative.

d. Write the ionic equation for the reaction occurring at the positive electrode during recharging. (1 mark)

$Ce(CH_3SO_3)_3(aq) + CH_3SO_3H(aq) \rightarrow Ce(CH_3SO_3)_4(aq) + H^+(aq) + e^-$

$Ce^{3+}(aq) \rightarrow Ce^{4+}(aq) + e^-$

Errors in previous parts of Question 7 clearly influenced responses to this part. During recharging the positive electrode is the anode, hence the site of oxidation.

e. Other than transporting ions between the electrodes, describe one function of the membrane in the zinc–cerium cell. (1 mark)

Acceptable responses included:

- to prevent the oxidant and reductant from coming into direct contact
- to prevent a spontaneous redox reaction occurring when the reductant and oxidant come into contact with each other
- to separate the two half-cells
- to prevent the excessive release of thermal energy in the cell.

One mark was awarded for a valid description of the function of the membrane.

f. Specify one factor that would limit the life of

One of:

- loss/breakdown/oxidation/corrosion of the Zn electrode
- side reactions at the electrodes
- electrolysis of water during recharging
- (-) $2H^+(aq) + 2e^- \rightarrow H_2(g)$; (+) $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
- significant temperature change
- build-up of gases around electrode.

Many responses to this question were not pertinent. Several simply stated 'temperature', but such a cell is designed to operate at a relatively low temperature (aqueous solutions are used) although it may be affected by a major change in temperature (e.g. electrolyte evaporation at high temperatures).

g. Experts have regarded the zinc–cerium cell a ould this be the case? (1 mark)

Fuel cell: supply of reactants ($Zn^{2+}/Ce^{4+}/Ce^{3+}$) from outside the cell.

Secondary cell: rechargeable/discharge reaction can be reversed.

One mark was awarded for appropriate reference to both fuel cell property **and** secondary cell property.

The secondary cell property was much more evident than the fuel cell property.

Space for Personal Notes

Section G: Multiple Choice Questions (6 Marks)

INSTRUCTION: 6 Marks. 1 Minute Reading. 6 Minutes Writing.

Question 17 (2 marks)

The overall discharge reaction for a lead-acid battery is:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

a. During recharge, the reaction at the cathode is: (1 mark)

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

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

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

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

b. When the lead-acid battery is discharging, the oxidising agent is: (1 mark)

A. Pb

B. PbO_2

C. $PbSO_4$

D. H_2SO_4

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Question 18 (1 mark)

Inspired by VCAA Chemistry 2023 Exam
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2023/NHT/2023chem-nht-w.pdf#page=6

Consider the following statements:

I. The half-cells must be separated for this cell to operate.
II. This cell requires reactants.
III. The anode in this cell is always negative.

Which one of the following combinations of statements applies to secondary cells and fuel cells?

	Secondary cells	Fuel cells
A.	I and II only	I, II and III
B.	I and II only	II and III only
C.	I and III only	I, II and III only
D.	I and III only	II and III only

Question 19 (1 mark)

Below are the half-equations for the lead-acid accumulator cell. The lead-acid accumulator cell can operate as an electrolytic cell.

$$PbSO_4(s) + H^+(aq) + 2e^- \rightarrow Pb(s) + HSO_4^-(aq) \qquad E^0 = -0.36\ V$$

$$PbO_2(s) + HSO_4^-(aq) + 3H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l) \qquad E^0 = 1.69\ V$$

Which one of the following correctly describes the lead-acid accumulator cell when it is operating as a galvanic cell and when it is operating as an electrolytic cell?

	Galvanic cell	Electrolytic cell
A.	H^+ ions react at the cathode.	HSO_4^- ions react at both electrodes.
B.	H^+ ions react at the anode.	HSO_4^- ions react at both electrodes.
C.	H^+ ions react at the cathode.	HSO_4^- ions are produced at both electrodes.
D.	H^+ ions react at the anode.	HSO_4^- ions are produced at both electrodes.

Question 20 (1 mark)

Lithium-ion rechargeable batteries are used in mobile phones. Environmental conditions can affect the number of charge cycles for a lithium-ion battery until the end of its useful life.

Which of the following environmental conditions would result in the largest number of charge cycles for a lithium-ion battery?

	Minimum temperature (°C)	**Maximum temperature** (°C)
A.	−8	11
B.	9	21
C.	18	37
D.	28	40

Question 21 (1 mark)

Inspired by VCAA Chemistry 2017 Exam
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2017/nht/2017chem-nht-w.pdf#page=14

Primary cells, secondary cells, and fuel cells can all be used as sources of energy.
The features that can be associated with one or more of the three types of cells include the following:

1. The cell is rechargeable.

2. The cell requires a continuous supply of reactants.

3. Reduction takes place at the cathode when discharging.

4. Oxidation takes place at the positive electrode when recharging.

Which of the above features are displayed by secondary cells?

A. 1,2 and 3

B. 2, 3 and 4

C. 1, 3 and 4

D. 1, 2, 3 and 4

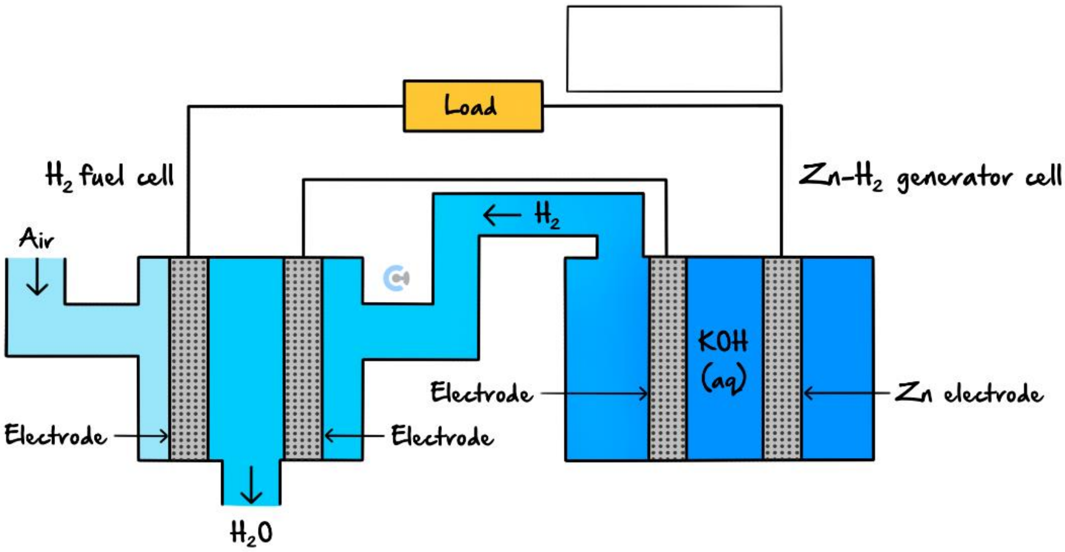
Section H: VCAA-Level Questions II (9 Marks)

INSTRUCTION: 9 Marks. 30 Seconds Reading. 9 Minutes Writing.

Question 22 (9 marks)

Inspired from VCAA Chemistry NHT Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/NHT/2021chem-nht-w.pdf#page=26

Researchers have investigated generating hydrogen, H_2, gas for hydrogen fuel cells by reacting zinc, Zn, and water, H_2O, in an electrochemical cell in series with an H_2 fuel cell. The diagram below represents an alkaline H_2 fuel cell in series with a Zn– H_2 generator cell.



The reactions that occur at each electrode in the Zn– H_2 generator cells are given below.

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

a. Write the overall reaction to the production of H_2 gas in the Zn– H_2 generator cell. (1 mark)

$Zn(s) + H_2O(l) \rightarrow ZnO(s) + H_2(g)$

b. Write the half-equation that occurs at the anode of the alkaline H_2 fuel cell. (1 mark)

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

c. In the box provided in the diagram above, draw an arrow to show the direction of the flow of electrons. (1 mark)

Correct answer: Arrow pointing to the left.

d. In terms of H_2 gas flow and electron flow, explain why it is theoretically possible to connect the H_2 fuel cell in series with the $Zn-H_2$ generator cell. (2 marks)

The amount of H_2 produced in the generator cell is equal to the amount of H_2 consumed in the fuel cell.

The half-equations show that n(e^-) released (2 per mol H_2 produced) in the generator cell is equal to the
n(e^-) used (2 per mol H_2 consumed) in the fuel cell.

Two electrons are transferred in the production of each mole of hydrogen in the Zn-H_2 generator cell and two electrons are transferred in the reaction of each mole of hydrogen consumed in the hydrogen fuel cell.

Students were awarded:

- 1 mark for the correct explanation of the mass (mole) balance
- 1 mark for the correct explanation for electron balance

e. Explain why the $Zn-H_2$ generator cell must be well-sealed to prevent contact with the atmosphere in order to produce H_2. Include any relevant equations in your answer. (2 marks)

Oxygen is a stronger oxidant than H_2O, so would be reduced at the cathode.

$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$ would occur at the cathode and no H_2 would be produced.

The reduction of O_2 is higher on the electrochemical series, so water will not be reduced, and hydrogen will not be produced.

Students were awarded:

- 1 mark for the half-equation for the reduction of O_2
- 1 mark for the correct explanation.

f. Assuming the system is 100% efficient, describe all of the energy conversions that occur in a combined $Zn-H_2$ generator cell and H_2 fuel cell system. (2 marks)

Zn-H_2 generator cell converts chemical energy of zinc to electrical energy **and** chemical energy in the form of hydrogen.

H_2 fuel cell converts chemical energy of hydrogen to electrical energy

Students were awarded:

- 1 mark for the correct conversion for the Zn-H_2 generator cell
- 1 mark for the correct conversion for the H_2 fuel cell

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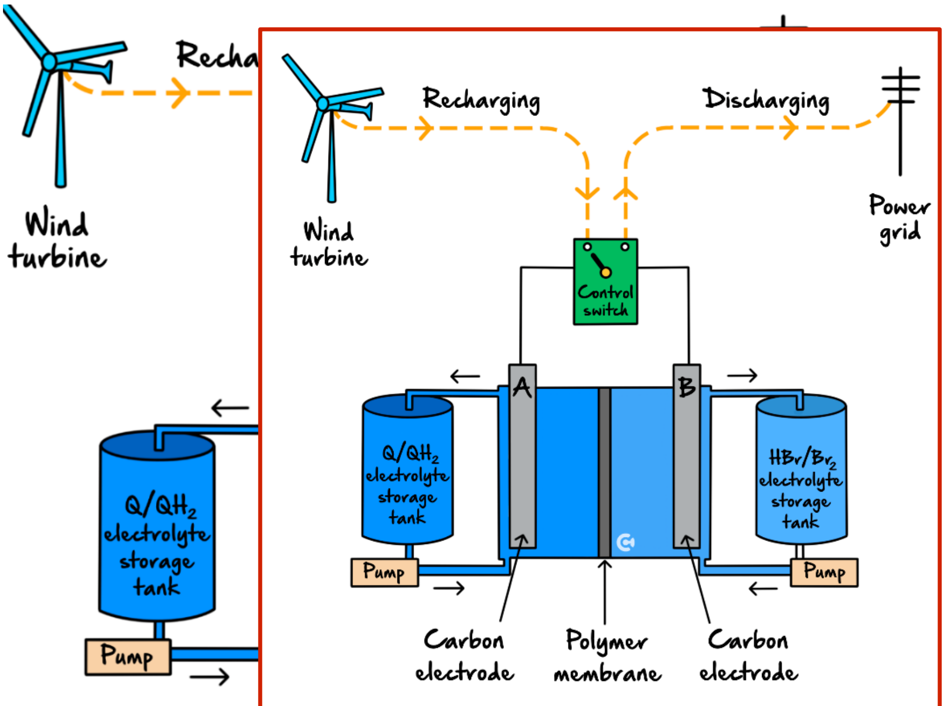
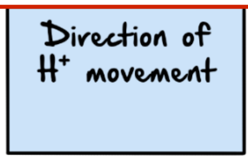
Section I: Extension Questions (14 Marks)

Question 23 (7 marks)

Redox flow batteries are used to store the excess electrical energy generated by commercial wind and solar farms. The batteries are recharged using electricity generated by wind turbines or solar cells. A scientific report, published in January 2014, described a redox flow battery that used a family of chemicals commonly occurring in plants such as rhubarb. These are organic and are known as quinones and hydroquinones. A diagram showing how such a redox flow battery might operate is provided below. In the diagram, Q represents the quinone and QH_2 represents the corresponding hydroquinone. The researchers made a model of the redox flow battery using aqueous solutions of the redox pairs, Q / QH_2 and Br_2 / Br^-. Refer to the diagram below.





Students need to recognise that oxidation occurred at the (−) electrode and reduction occurred at the (+) electrode. The question stated that QH_2 was converted to *Q* and Br_2 was converted to HBr, from which it could be deduced that Br_2 was reduced (the oxidation number of Br decreases from 0 to −1) and so QH_2 was oxidised. Given that the diagram suggested the movement of H^+ ions, the half-equations should have included H^+. As the question referred to the Br_2 / Br^- redox pair, $Br_2 + 2e^- \rightarrow 2Br^-$ was also accepted for the half-equation at the (+) electrode.

During discharge, QH_2 is converted to Q and Br_2 is converted to HBr.

a. Write balanced half-equations for the reactions occurring at the positive and negative electrodes as the cell is **discharged**. Assume the electrolytes are acidic. (2 marks)

Positive electrode: ______________________________

Negative electrode: ______________________________

Marks	0	1	2	Average
%	50	25	25	0.8

(+) electrode $Br_2 + 2H^+ + 2e^- \rightarrow 2HBr$/$Br_2 + 2e^- \rightarrow 2Br^-$
(−) electrode $QH_2 \rightarrow Q + 2H^+ + 2e^-$

b. Write an overall equation for the reaction that occurs when the cell is **recharged**. (1 mark)

Marks	0	1	Average
%	57	43	0.5

$2HBr + Q \rightarrow Br_2 + QH_2$ / $2Br^- + 2H^+(aq) + Q \rightarrow Br_2 + QH_2$

One mark is to be awarded for a 'recharging' equation consistent with the 'discharging' half-equations in **part a**. – that is, if the equation was balanced and consistent with the reverse of the overall equation implied in **part a**. While students realised that they needed to add together the reverse of their half-equations from **part a.**, a significant proportion provided an unbalanced overall half-equation.

c. The researchers reported that their tests indicated that only hydrogen ions were able to move through the polymer membrane separating the cells.

i. In the box provided on the diagram, use an arrow, → or ←, to indicate the direction of movement of hydrogen ions as the cell is **recharged**. (1 mark)

ii. Why is it important that the other reactants in the half-cells are not able to pass through the polymer? (1 mark)

- If the oxidant and reductant are allowed to come into contact they can react together in the solution rather than individually at the electrodes.
- Chemical energy would be converted to thermal energy rather than electrical energy.
- To prevent the oxidant and reductant coming into contact with each other.
- To prevent spontaneous direct reaction between the oxidant and reductant.
- To ensure that the cell can be recharged.

d. The researchers also reported that the voltage applied to the cell during recharging was kept below 1.5 *V* to avoid the electrolysis of water.

Write an equation for the overall reaction that occurs when water is electrolysed. (1 mark)

Marks	0	1	Average
%	86	14	0.2

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

Few students recalled that

It is possible to deduce the overall equation from the half-equations for the oxidation and reduction of water – that is $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ and $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ from which the overall equation $6H_2O(l) \rightarrow O_2(g) + 2H_2(g) + 4H^+(aq) + 4OH^-(aq)$ can be deduced and simplified down to $2H_2O(l) \rightarrow O_2(g) + 2H_2(g)$

overall equation for the decomposition of water.

e. Q… a… …lso e…

What is meant by the term 'renewable' in this context? (1 mark)

Marks	0	1	Average
%	43	57	0.6

Rhubarb (the raw material for the quinones) can readily be replenished. Rhubarb is a plant-based source of quinones and can be easily grown. Many students answered this question well. Statements such as 'can be produced as fast as it is needed' were also acceptable. Some students made incorrect statements such as 'recycled', 'used again', 'environmentally friendly' and 'carbon neutral', suggesting that further improvement is needed in achieving real understanding of the concept of renewability.

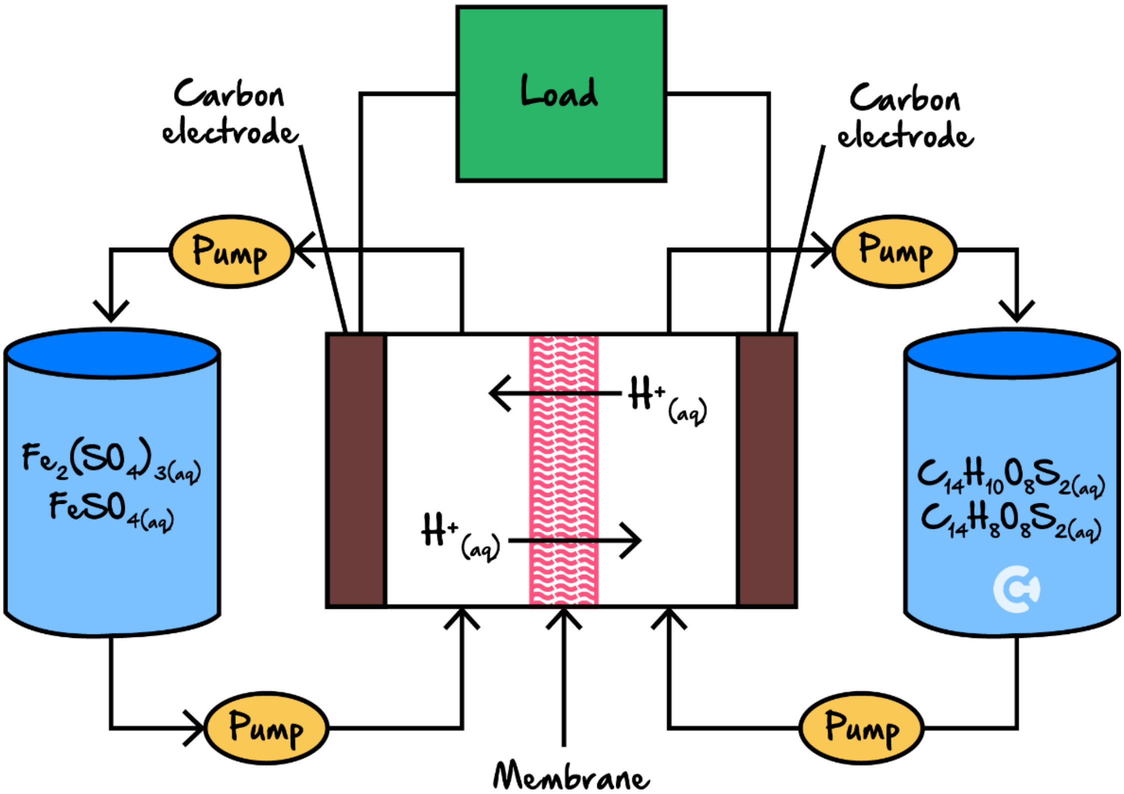
Question 24 (7 marks)

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)

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)

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)

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

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)

Negative

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

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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