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VCE Chemistry ¾
Primary Cells & Faraday's Laws [0.9]
Workshop

next week
transester
bond enthalpy

Friday

Fri 1cev CH34 415-615
Fri CH34 645-845

Error Logbook:

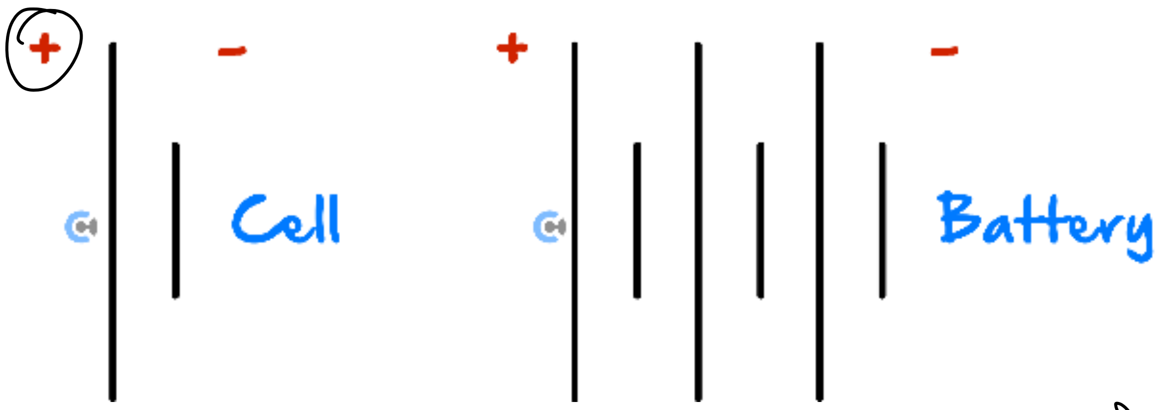
Mistake/Misconception #1		Mistake/Misconception #2	
Question #:	Page #:	Question #: 15	Page #: 19
Notes: galvanic cells, electrolytic cells 		Notes: $pH = -\log_{10}[H^+]$ $X^{2+} \rightarrow X^{3+} + e^-$ 1:1 ratio	

Mistake/Misconception #3		Mistake/Misconception #4	
Question #:	Page #:	Question #:	Page #:
Notes:		Notes:	

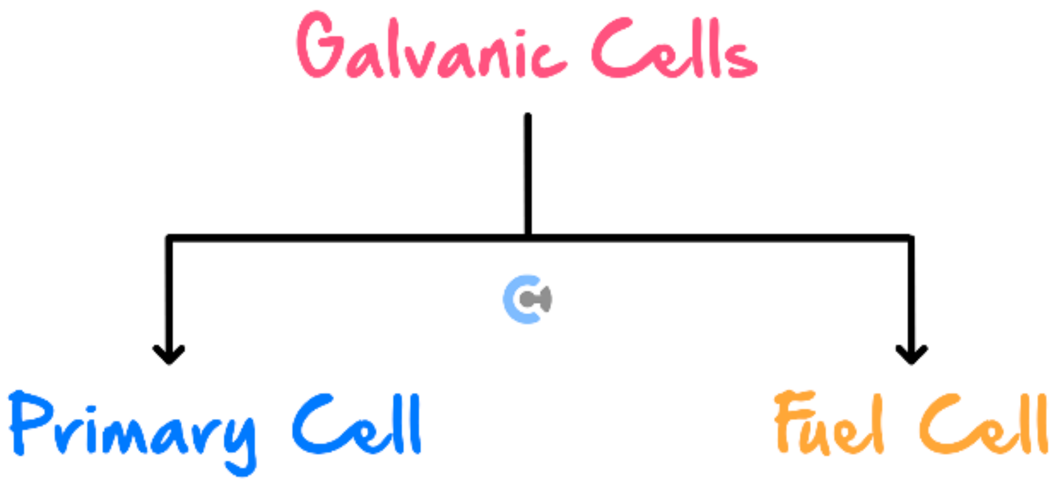
Section A: Recap (6 Marks)

Learning Objective: [1.10.1] - Identify features of primary cells & how they operate

Definition



- **Primary Cell:** Galvanic cell which produces electrical energy and cannot be recharged.
  - Electrolyte purpose: Allows for the movement of ions to balance charges around the electrodes.
  - Provides the internal circuit in the cell.
- Comparison of Primary & Fuel Cells:



Type of Cell	Primary Cells	Fuel Cells
Type of Reaction	[Spontaneous]/[Non-Spontaneous]	[Spontaneous]/[Non-Spontaneous]
Reactants	[Stored in each half-cell]/ [Continuous supply of reactants]	[Stored in each half-cell]/ [Continuous supply of reactants]
Electrodes	CY	PICCY
Reactivity of Electrodes	Electrodes [may]/[will not] participate in the chemical reaction.	Electrodes [may]/[will not] participate in the chemical reaction.
Price	[Cheap]/[Expensive]	[Cheap]/[Expensive]
Electrolyte	[Each half-cell has its electrolyte]/ [Shared electrolyte]	[Each half-cell has its electrolyte]/ [Shared electrolyte]
Energy Efficiency	[High]/[Low] Energy Efficiency	[High]/[Low] Energy Efficiency

Definition

Learning Objective: [1.10.2] - Apply Faraday's First & Second Law and $Q = It$ & $Q = n(e^-)F$ to calculations

- Electric Charge (Q): The amount of charged particles (e.g., electrons) which are present, measured in Coulombs (C).
- Electric Current (I): The rate of flow of charged particles per second, measured in amperes (A).
- Electric Charge Formula:

$$Q = It$$

- Faraday's First Law of Electrochemistry:

"The amount of chemical change being produced by a current at an electrode-electrolyte interface is proportional to the quantity of electricity used."

- **Relationship:** ↑Q ∝ m↑
- **Law Simplified:** As more electrical charge passes through the cell, the amount of substances that react/are produced [increases]/[decreases]/[same].

- $F =$ 96500 C mol^{-1}

- Faraday's Constant Formula: $Q = n(e^-)F$

- **Faraday's Second Law**:

"If the same amount of electricity is passed through different electrolytes, the masses of ions deposited at the electrodes are directly proportional to their chemical equivalents."

- **Law Simplified**: Consider the stoichiometric ratio – n(metal) : n(e⁻)

- Steps for calculation:

1. Write any half-eq.
2. $Q = It$
3. $n(e^-) = \frac{Q}{F}$
4. Stoich Ratios n(metal) : n(e⁻)
5. m(metal) = n × Mr

Learning Objective: [1.10.3] - Calculate the Charge of a Metal

Definition

- Steps to find the charge of ion:

1. Find n(metal) by n(metal)= $\frac{m}{Mr}$
2. Find $n(e^-)$ by $n(e^-) =$ $\frac{Q}{F}$
3. Find the stoichiometric ratio by dividing by the smallest no. present.

Question 1 (6 marks) **Walkthrough.**

$Q = It$ → $Q = n(e^-)F$

A current of 7.50 A is passed through for 2.50 minutes in a galvanic cell, which includes a $Cu^{2+}(aq)/Cu(s)$ and a $Ni^{2+}(aq)/Ni(s)$ half-cell.

a. Find the change in mass of the anode. (4 marks)

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

$Q = It = 7.5A \times 2.5 \times 60s = 1125\ C$

$n(e^-) = \frac{Q}{F} = \frac{1125C}{96500C/mol} = 0.0117\ mol$

$n(Ni) = \frac{1}{2}n(e^-) = 0.00583\ mol$

$m(Ni) = n \times M = 0.00583 \times 58.7 = 0.342\ g$

decreases by 0.342g

b. Find the amount of electrical energy that is produced. (2 marks)

$E = VIt \quad V = 0.34 - (-0.25) = 0.59V$

$E = 0.59V \times 7.5A \times 2.5 \times 60s = 664\ J$

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Section B: Warm Up (19 Marks)

INSTRUCTION: 19 Marks. 12 Minutes Writing.

Question 2 (2 marks)

A galvanic cell produced 3.50 A of current and an electric charge of 100 C passed through the cell throughout this period.

a. Calculate how long the cell must have been operating for. (1 mark)

$t = \frac{Q}{I} = \frac{100C}{3.5A} = 28.6s$

b. What amount of electrons in mol, must have passed through the cell during this time? (1 mark)

$n(e^-) = \frac{Q}{F} = \frac{100C}{96500C/mol} = 0.00104 mol$

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Question 3 (3 marks)

If a cell produces 1.82 g of metal at the cathode during one trial of an experiment in which 812 C of charge is passed through and then produces 2.34 g of the metal during a subsequent trial:

a. Calculate the charge in C, during this subsequent trial. (1 mark)

$812\ C \times \frac{2.34}{1.82} = 1044\ C$

$= 1.04 \times 10^3\ C$

10.1 606

b. If it is known that the cell was operating for 10 minutes in the original trial, how much **longer**, in minutes, must the cell have been run for during the subsequent trial if the rate of electron flow was unchanged? (2 marks)

$10 min \times \frac{2.34}{1.82} = 12.9 min$ → $10 \times 60s = 600s$

2.9 min

$6.0 \times 10^2 s$

Question 4 (2 marks)

One of the electrodes is made up of an unknown metal, Y.

a. Given there were 0.124 mol of electrons passing through, and the amount of Y which reacted is 0.041 mol, find the charge of the ion Y^{x+}. (1 mark)

n(Y) : n(e⁻)

0.041 mol : 0.124 mol

0.041 0.041

1 : 3.02

+3

b. If there is an increase in the mass of the electrode Y, must it be the cathode or the anode? Justify your answer using a relevant half-equation. (1 mark)

Increase means $Y_{(s)}$ is being produced, so it's present at the cathode

$Y^{3+}_{(aq)} + 3e^- \rightarrow Y_{(s)}$ reduction

Question 5 (7 marks)

Batteries are commonly called primary cells.

a. What is a primary cell? (1 mark)

A galvanic cell which cannot be recharged.

b. State which metal is the most commonly used in batteries and explain why this is the case about the voltage produced. (2 marks)

Li. It is the strongest reductant on the ECS. When paired with an oxidant, it produces a larger EMF than with a weaker reductant.

c. A particular primary cell involves the reduction of nickel (II) nitrate into nickel metal. A current of 2.30 *A* is produced over 5.50 *min*. Find the mass of nickel metal that was produced at that time. (4 marks)

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

$Q = It = 2.3A \times 5.5 \times 60s = 759\ C$

$n(e^-) = \frac{Q}{F} = \frac{759C}{96500C/mol} = 0.00787mol$

$n(Ni) = \frac{1}{2}n(e^-) = \frac{1}{2} \times 0.00787 = 0.00393mol$

$m(Ni) = n \times M = 0.00393 \times 58.7 = 0.231g$

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Question 6 (5 marks)

Emina measures that after letting a galvanic cell operate for half an hour, 2.50 g of solid lead were evolved at one of the electrodes.

a. Write the appropriate balanced half-equation aligning with her observations. (1 mark)

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

b. Calculate the amount in mol, of electrons produced in this cell. (2 marks)

$n(Pb) = \frac{m}{Mr} = \frac{2.50g}{207.2} = 0.012 mol$

$n(e^-) = 2n(Pb) = 2 \times 0.012 = 0.0241 mol$

c. Calculate the current generated by this cell. (2 marks)

$Q = n(e^-)F = 0.024 mol \times 96500 = 2329 C$

$I = \frac{Q}{t} = \frac{2329C}{30 \times 60 s} = 1.29 A$

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Section C: Ramping Up (15 Marks)

INSTRUCTION: 15 Marks 11 Minutes Writing.

Question 7 (1 mark)

Which one of the following statements is correct?

A. Fuel cells and galvanic cells produce heat.

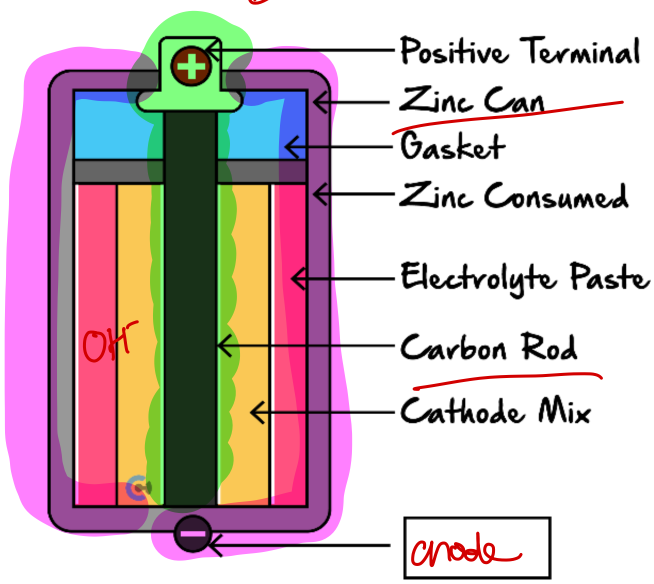
B. Fuel cells have inert electrodes, whereas galvanic cells do not.

C. Fuel cells have porous electrodes, whereas galvanic cells do not.

D. Fuel cells and galvanic cells operate when electrons flow through the electrolyte.

Question 8 (8 marks)

Use the diagram below to answer the following questions:



a. Label the bottom section of the cell as the anode or the cathode in the box provided above. (1 mark)

b. This particular primary cell is often referred to as an "alkaline" cell. Suggest why. (1 mark)

The electrolyte is made of hydroxide (OH^-), which is alkaline.

c. Given the overall equation occurring is: $Zn(s) + MnO_2(s) + 2H_2O(l) \rightarrow Mn(OH)_2(s) + Zn(OH)_2(s)$:

2H2O

i. Write the balanced half-equation involving manganese (IV) dioxide. (1 mark)

$MnO_2 + 2H^+ + 2e^- \rightarrow Mn(OH)_2$

$+2OH^-$ $+2OH^-$

$MnO_{2(s)} + 2H_2O_{(l)} + 2e^- \rightarrow Mn(OH)_{2(s)} + 2OH^-_{(aq)}$

ii. Circle the following: (1 mark)

Site of MnO_2 reaction: [Anode]/[Cathode]

iii. Fill out the following table. You should refer to the diagram. (2 marks)

Anode material	Cathode material
Zinc	Carbon

d. Explain the electrolyte paste's role - in terms of how it migrates throughout this particular cell. (2 marks)

It acts as the internal circuit to complete the circuit. (1)
Hydroxide is produced at the cathode, & it moves through the electrolyte paste & arrives at the anode. (1)
anions → anode

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Question 9 (6 marks)

If a galvanic cell were set up between Co^{2+}/Co and Fe^{2+}/Fe:

a. Calculate the voltage this cell would produce. (1 mark)

EMF = -0.28V − (-0.44V) = 0.16V

b.

i. If the cell was left running for two and a half minutes and a current of 4.00 A was generated, calculate the charge produced in C. (1 mark)

Q = It = 4A × 2.5 × 60s = 600 C

ii. What amount of electrons in mol, must have flown towards the cathode in this cell? (1 mark)

n(e⁻) = Q/F = 600/96500 = 0.00622 mol

c.

i. Write the half-equation occurring at the cathode. (1 mark)

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

ii. Thus, what mass of metal would have been deposited on the cathode? (2 marks)

n(Co) = ½ n(e⁻) = ½ × 0.00622 = 0.00311 mol

m(Co) = n × M = 0.00311 × 58.9 = 0.183 g

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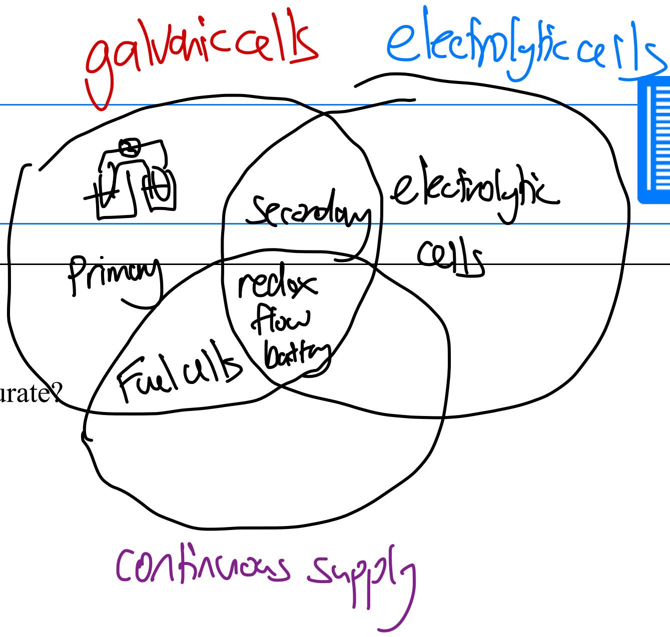
Section D: Getting Trickier I (12 Marks)



INSTRUCTION: 12 Marks. 10 Minutes Writing.

Question 10 (1 mark)

Which one of the following statements is the **most** accurate?

A. All fuel cells are galvanic cells.

B. All galvanic cells are primary cells.

C. All secondary cells have porous electrodes.

D. All fuel cells are more efficient than all secondary cells.

Question 11 (4 marks)

a. If a cell runs for 10 minutes and a current of 10.0 A is generated, calculate the charge passing through the cell. (1 mark)

$Q = It = 10A \times 10 \times 60s = 6000C = 6.0 \times 10^3 C$

b.

i. Had the current been turned up to 15.0 A and the charge produced were unchanged, what would the new time have been, in minutes? (1 mark)

$t = \frac{Q}{I} = \frac{6000C}{15A} = 400s = 6.7 min$

ii. Hence, how many moles of electrons would have flown through the external circuit? (1 mark)

$n(e^-) = \frac{Q}{F} = \frac{6000C}{96500 C/mol} = 0.062 mol$

120s + 43s = 163s

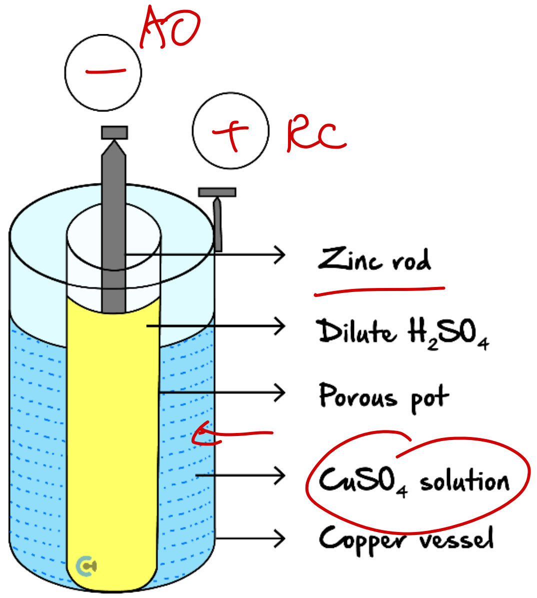
2×60 + 43

c. Now, suppose the charge flowing through a separate galvanic cell was known to be $3.8\ C$, and a stopwatch showed that this cell was operating for 2 minutes and 43 seconds, what would one expect the ammeter to have read in mA? (1 mark)

$I = \frac{Q}{t} = \frac{3.8C}{(2\times60+43)} = 0.0233A = 23mA$

Question 12 (7 marks)

Another primary cell, which contains sulphuric acid, H_2SO_4, is shown below:



a. Given the following half-equation:

$$CuSO_4(aq) + 2H^+(aq) + 2e^- \rightarrow Cu(s) + H_2SO_4(aq)$$

Label the polarity of the electrodes in the circles provided in the diagram. (1 mark)

b. Explain why the equation above can also be written as: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. (1 mark)

$SO_4^{2-}(aq)$ stays the same before & after reaction, & acts as a spectator ion.

Additionally, $CuSO_4(aq)$ is aqueous, & can dissociate into cations & anions.

either

c.

i. If the other half-equation is:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + 2H^+(aq) + 2e^-$$

Write the overall **full** equation. (1 mark)

~~$CuSO_4 + 2H^+ + Zn + H_2SO_4 \rightarrow Cu + H_2SO_4 + Zn + 2H^+$~~

$CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$

ii. Write the overall **ionic** equation. (1 mark)

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

d. Would this cell be able to operate without the sulphuric acid present? Justify your reasoning, ensuring to refer to the role of the porous pot. (3 marks)

(1) No. While it does not react, it acts as the electrolyte.

(1) It helps to balance the build-up of charge & complete the circuit as the SO_4^{2-} moves to the anode & H^+ moves to the cathode.

(1) The porous pot allows these ions to pass through & arrive at each electrode while separating half-cells.

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Section E: Getting Trickier II (13 Marks)

INSTRUCTION: 13 Marks. 12 Minutes Writing.

Question 13 (6 marks)

A hydrogen fuel cell is run with an acidic electrolyte.

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

i. Positive electrode. (1 mark)

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

ii. Negative electrode. (1 mark)

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

1.23V

b. Find the amount of electrical energy produced, given that 3.00 *L* of oxygen gas has reacted at SLC. (4 marks)

$n(O_2) = \frac{V}{V_m} = \frac{3}{24.8} = 0.121 \text{ mol}$

$n(e^-) = 4n(O_2) = 0.484 \text{ mol}$

$Q = n(e^-)F = 0.484 \times 96500 = 46694 \text{ C}$

$E = VIt = VQ = 46694 \times 1.23 = 57433 \text{ J}$

$= 57.4 \text{ kJ}$

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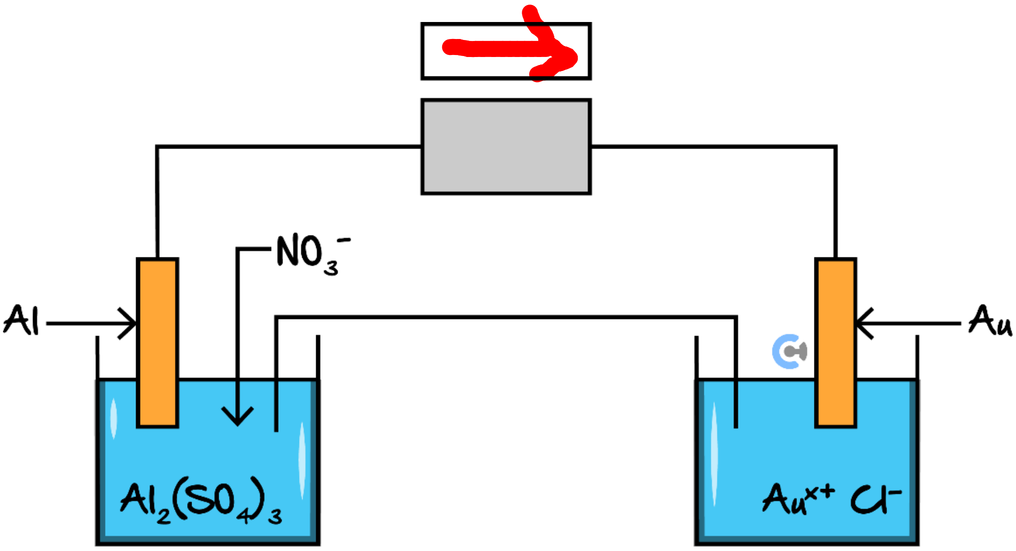
Question 14 (7 marks)

Matthew, a lucky lad, finds a gold bar in his school lab. He's a bit greedy, however, and wishes to produce more gold. So, he finds some useful equipment and sets up the following galvanic cell:



a. Label the direction of electron flow in the box provided. (1 mark)

b. The aluminium electrode is seen to decrease in mass by 2.93 g.

i. Find the amount of electrons in mol, which would have flown through the wire. (2 marks)

n(Al) = m/M = 2.93/27.0 = 0.11 mol
n(e^-) = 3 × n(Al) = 0.326 mol

ii. It is found that 21.38 g of gold were deposited at the other electrode. Calculate the **charge**, x, on the gold ions. (2 marks)

n(Au) = m/M = 21.38/197.0 = 0.109 mol
n(Au) ∶ n(e^-) = 0.109 ∶ 0.326 = 1 ∶ 3
Therefore,Au^(3+) (aq) and the solution is gold (III) chloride.

iii. If the gold ions had a greater charge, would this have benefitted Matthew or not? That is, would more or less solid gold have been evolved than 21.38 g? Justify with reference to Faraday's laws of electrochemistry. (2 marks)

According to Faraday's second law, for every mole of metal, x mol of electrons will be required, where x is the charge on the metal ion. Therefore, if the charge were greater, more moles of electrons would be needed to get the same amount of gold. Since $n(e^-)$ is fixed, less gold would have been made → bad for Matthew.

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Let's take a BREAK!

Section F: VCAA-Level Questions I (11 Marks)

INSTRUCTION: 11 Marks. 0.5 Minutes Reading. 10 Minutes Writing.

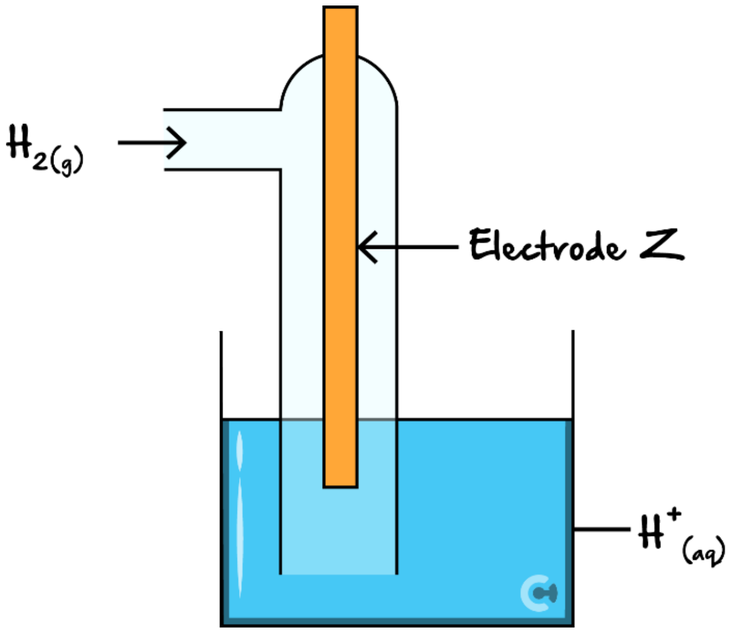
7:15

Question 15 (11 marks)

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

The following diagram represents a half-cell for the reaction:

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



a.

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

Pt(s), graphite, Au(s)

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

- The temperature at which it must operate 25°C.
- The required pH of the solution of H^{+}(aq) ions 0.

1.0 M pH = −log₁₀[H⁺] pH = −log₁₀(1)

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

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

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

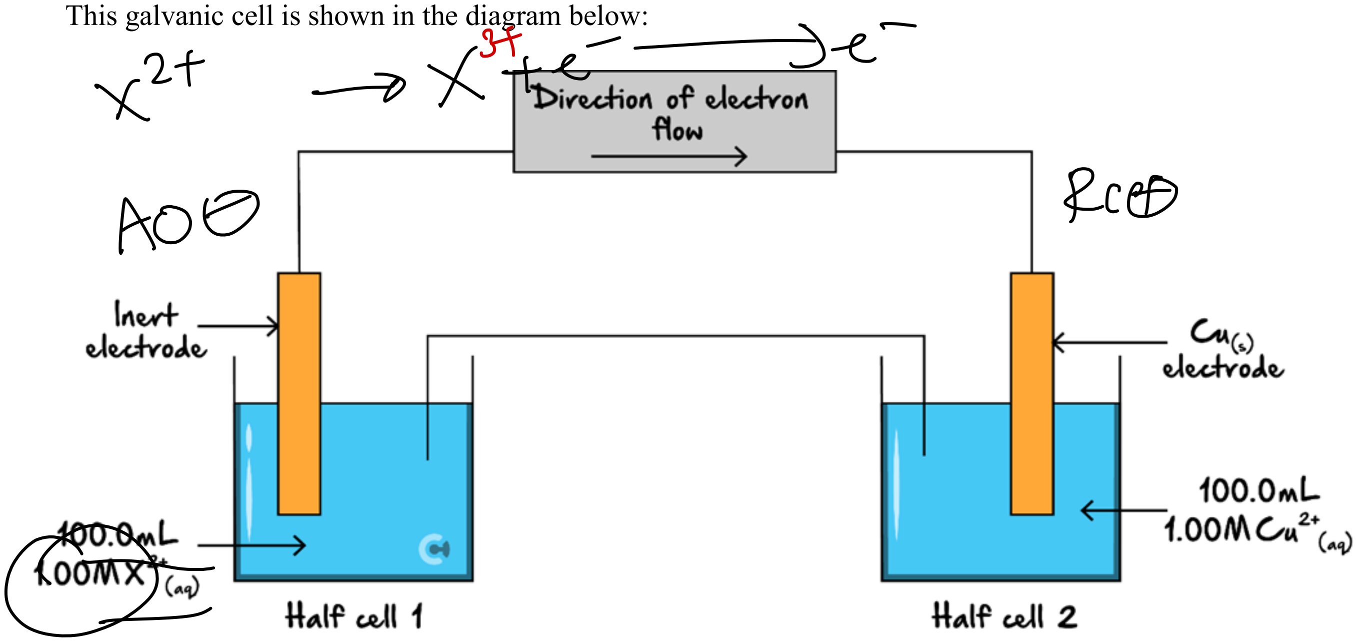
The stronger reductant is Ni(s) ①.

Explanation if pH of half-cell 1 increased, it must be becoming more basic. This means H^+ ions are being used up, according to: $2H^+(aq) + 2e^- \rightarrow H_2(g)$ ①
This means Ni(s) must be the stronger reductant.

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

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

This galvanic cell is shown in the diagram below:



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

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

$\Delta c = 1.00M - 0.725M = 0.275M$

$\Delta n = \Delta c V = 0.275M \times 0.1L = 0.0275 mol$

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

$n(e^{-}) = \frac{Q}{F} = \frac{2654C}{96500C/mol} = 0.0274 mol$ ① $n(e^{-})$

$n(X^{2+})_{reacted} : n(e^{-})$

$0.0275 : 0.0274$

ratio is 1:1 ①

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

+3

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

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

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Section G: Multiple Choice Questions (7 Marks)

INSTRUCTION: 7 Marks. 7 Minutes Writing.

Question 16 (1 mark)

The following reactions occur in a primary cell battery:

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

$$2MnO_2 + 2e^- + H_2O \rightarrow Mn_2O_3 + 2OH^-$$

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

A. The reaction produces heat and Zn reacts directly with MnO_2.

B. The reaction produces heat and Zn does not react directly with MnO_2.

C. The reaction does not produce heat and Zn reacts directly with MnO_2.

D. The reaction does not produce heat and Zn does not react directly with MnO_2.

Question 17 (1 mark)

A hydrogen-oxygen fuel cell uses 1.00×10^{-5} *mol* of hydrogen gas per second of operation. The current produced by this cell is:

A. 0.483 *A*

B. 0.965 *A*

C. 1.93 *A*

D. 3.86 *A*

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

$n(e^-) = 2(H_2) = 2 \times 10^{-5}$ mol/s

$Q = n(e^-)F = 2 \times 10^{-5}$ mol/s $\times 96500$ C/mol

$= 1.93$ C/s $= 1.93$ A

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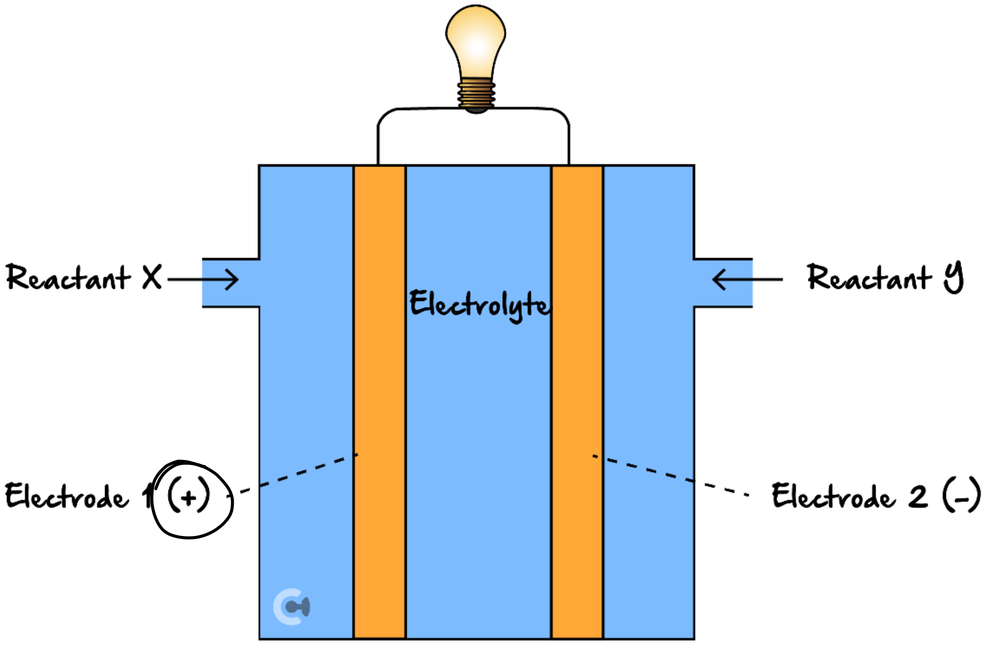
$I = \frac{Q}{t}$

Question 18 (1 mark)

$$C_2H_5OH(aq) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$



Choose the alternative that correctly labels electrodes 1 and 2 and identifies the reactants:

	Electrode 1	Electrode 2	Reactant X	Reactant Y
A.	Anode	Cathode	$O_2(g)$	$C_2H_5OH(aq)$
B.	Anode	Cathode	$C_2H_5OH(aq)$	$O_2(g)$
C.	Cathode	Anode	$O_2(g)$	$C_2H_5OH(aq)$
D.	Cathode	Anode	$C_2H_5OH(aq)$	$O_2(g)$

Space for Personal Notes

Question 19 (1 mark)

A particular galvanic cell includes metal, M, which involves the metal ion M^{x+}. During cell operation, a current of 1.50 amperes was produced for 180 seconds. The galvanic cell resulted in 0.0014 mol of metal.

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

A. 1

B. 2

C. 3

D. 4

Question 20 (1 mark)

In delivering an electric current, a particular cell uses 0.750 g of ethanol in 30.0 minutes. The half-equation for the reaction of ethanol is:

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

The electric current flowing, in amps, is:

A. 0.175

B. 0.874

C. 10.5

D. 629

Space for Personal Notes

Question 21 (1 mark)

1 F is equivalent to the charge on 1 mol of electrons. The mass of nickel, Ni, that can be electroplated onto the platinum, Pt, electrode with 320 F of charge is:

A. $9.73 \times 10^{-2}\ g$

B. $1.95 \times 10^{-1}\ g$

C. $9.39 \times 10^{3}\ g$

D. $1.88 \times 10^{4}\ g$

Question 22 (1 mark)

A fuel cell is set up based on the oxidation of methane. Again, the equation for the anode half-reaction is:

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

Assuming that all the energy of the oxidation reaction is converted to electricity, the amount of electric charge, in coulomb, obtained from the oxidation of one mole of methane is closest to:

A. 8×10^{2}

B. 1×10^{3}

C. 8×10^{5}

D. 1×10^{6}

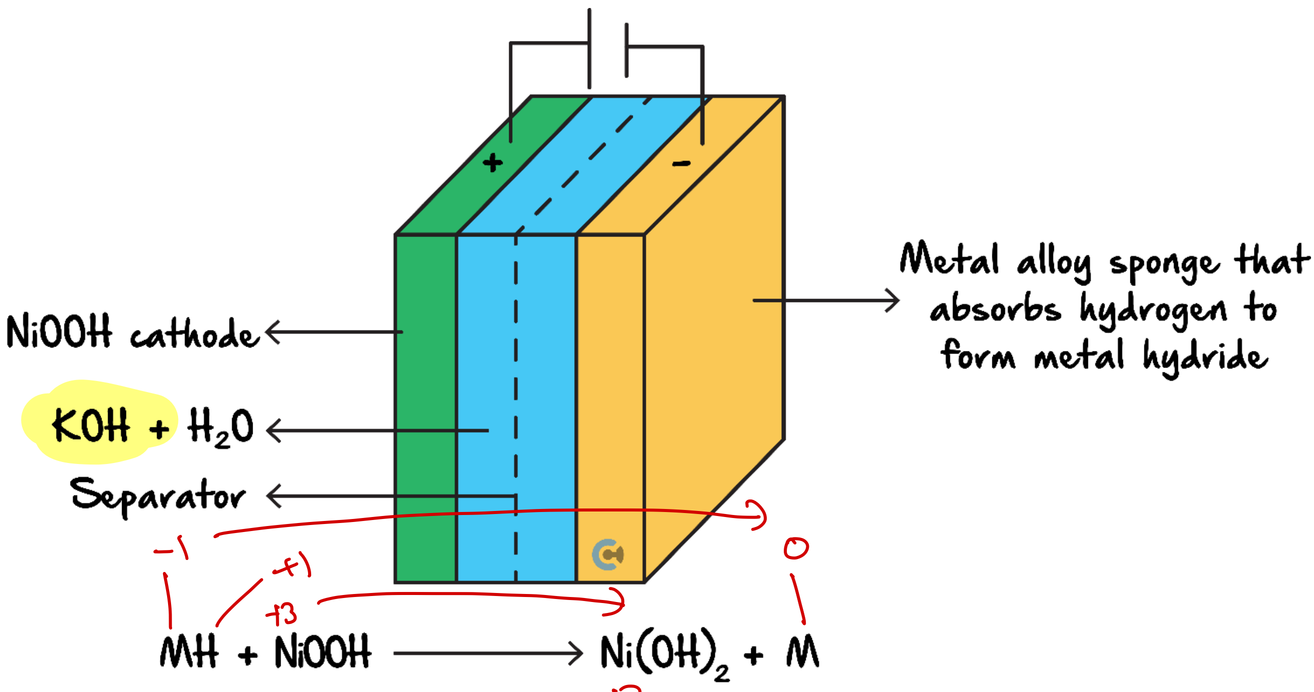
Space for Personal Notes

Section H: VCAA-Level Questions II (10 Marks)

INSTRUCTION: 10 Marks. 0.5 Minutes Reading. 9 Minutes Writing.

Question 23 (10 marks)

A diagram of a nickel-metal hydride cell is shown below:



a. Write the reaction occurring at the cathode. (1 mark)

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

b. State **two** functions of the electrolyte in this particular cell. (2 marks)

$NiOOH_{(s)} + H_2O_{(l)} + e^- \rightarrow Ni(OH)_{2(s)} + OH^-_{(aq)}$

c. These types of nickel-metal hydride cells are efficient and portable.

i. An average battery contains 2.00 *g* of NiOOH(s), calculate the battery life of the average battery, given it produces a 2.00 *A* current. (3 marks)

$n(NiOOH) = \frac{m}{M} = \frac{2}{58.7+32+1} = 0.0218\ mol$

$n(e^-) = n(NiOOH) = 0.0218\ mol$

$Q = n(e^-)F = 0.0218 \times 96500 = 2105\ C$

$t = \frac{Q}{I} = \frac{2105\ C}{2\ A} = 1052\ s$

$= 1.05 \times 10^3\ s$

ii. Describe the effect the electrolyte has on the overall battery life of the cell. (1 mark)

The electrolyte has no impact on the overall battery life of the cell as it does not participate in overall reactions, and thus never has a net gain or loss of ions.

d. Lithium is a more common reactant used in batteries.

i. State why lithium is a more common reactant used in batteries. (1 mark)

Lithium is a strong reductant and wil produce a large EMF.

ii. Explain whether lithium can viably replace the reductant in this cell. (2 marks)

No. Lithium is such a strong reductant that it will react with water spontaneously. Since water is used within the electrolyte, it will spontaneously react with the lithum and degrade the anode thus rendering the cell inactive. Additionally it will produce hydrogen gas that can explode.

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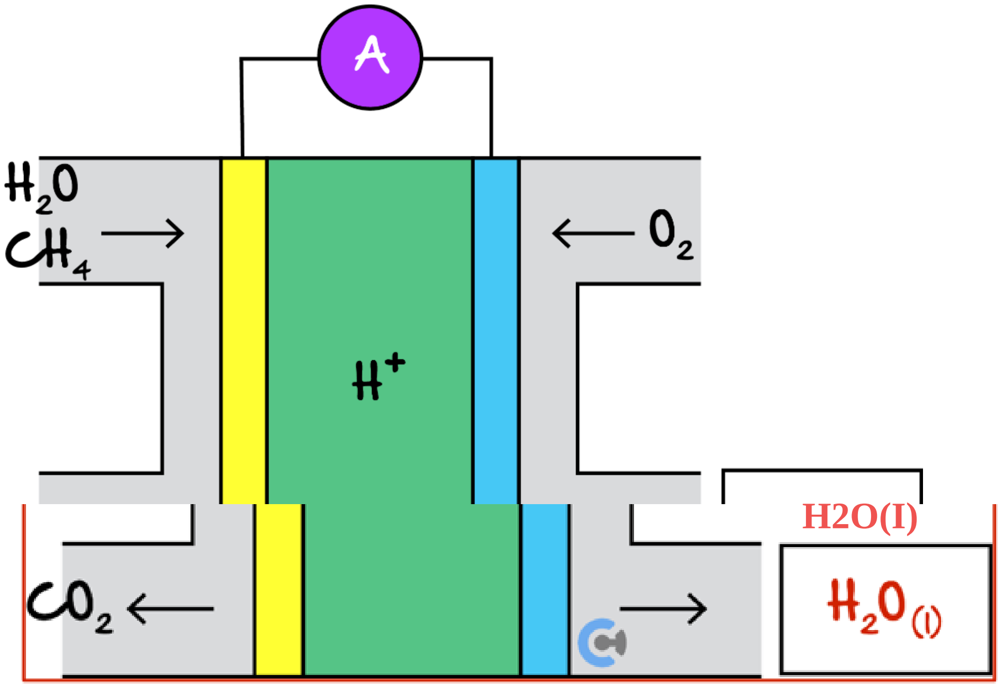
Section I: Extension Questions (10 Marks)

Question 24 (10 marks)

An acidic methane fuel cell is depicted below:



a. Write the balanced half-equation occurring at the cathode. (1 mark)

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

b. Write the other half-equation. (1 mark)

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

c. Hence, fill in the box on the diagram with the appropriate product(s). (1 mark)

d. Outline two design features which differentiate fuel cells from primary cells. (2 marks)

Continuous supply of reactants whereas primary has stored reactants.
PICCY electrodes versus merely electrically conductive electrodes.

e. If 22.0 L of $CO_2(g)$ were evolved at SLC, calculate the amount of electrons in mol, which must have been released by the anode. (2 marks)

$$n(CO_2) = V/Vm = 22.0/24.8 = 0.887\ mol$$
$$n(e^-) = 8 \times n(CO_2) = 7.097\ mol$$

f. If the ammeter in the external circuit reads 3.50 A, calculate how long this cell must have been operating, in **days**. (3 marks)

$$Q = n(e^-) \times F = 7.097 \times 96500 = 684{,}839\ C$$
$$t = Q/I = 684839/3.50 = 195{,}669\ s$$
$$t = 195669/(60 \times 60 \times 24) = \mathbf{2.26\ days}$$

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