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VCE Chemistry ¾
Electroplating [0.15]
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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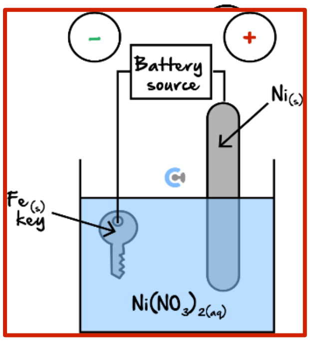
Section A: Recap

Learning Objective: [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: $> 0\ V$

Space for Personal Notes

Definition

Learning Objective: [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.

Definition

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

- Equations:

$$Q = It$$

$$Q = n(e^-)F$$

- Typical Steps:

1. $Q = It$
2. $n(e^-) = \frac{Q}{F}$
3. Stoich 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: Stoich ratio $n(\text{metal}) : n(e^-)$
- **Molar Mass: Charge Ratio** $(\boldsymbol{M_r/z})$
  - **Use:** compare mass deposited for different metals.
  - **Formula:** $\frac{M_r}{\text{charge}}$

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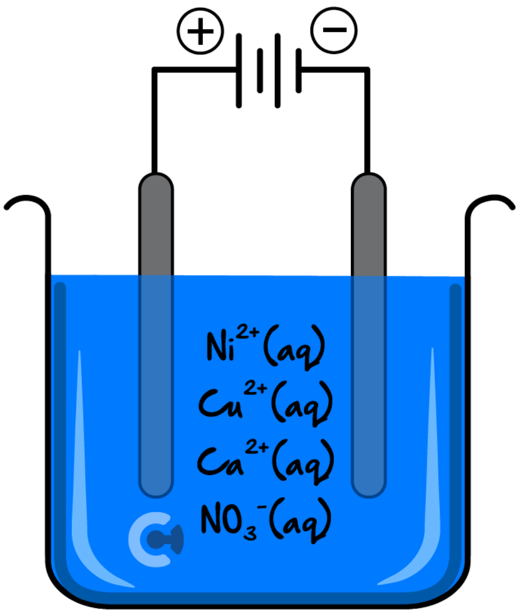
Question 1 (5 marks) **Walkthrough.**

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:

i. Positive electrode. (1 mark)

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

ii. Negative electrode. (1 mark)

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

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 which takes place at this electrode. (1 mark)

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

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

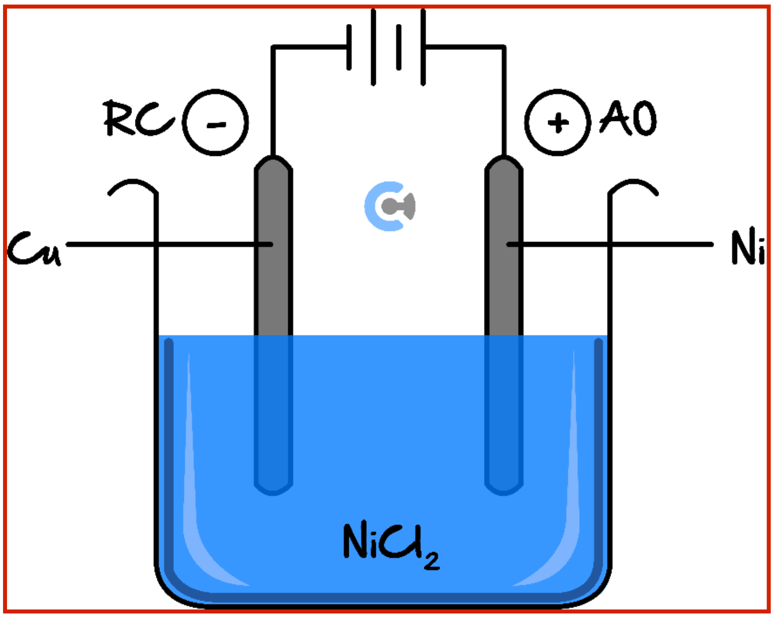
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$

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

Question 2 (5 marks) **Walkthrough.**

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

a. Draw the electroplating cell. (1 mark)



b. Find the change in mass of the key chain, if 4.20 *A* of current is passed through for 15.0 minutes. (4 marks)

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

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

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

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

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

$m(Ni) = n \times M = 0.0196 \times 58.7 = 1.150\,g$

increase by 1.150g

Space for Personal Notes

Section B: Warm Up (16 Marks)

INSTRUCTION: 16 Marks. 10 Minutes Writing.

Question 3 (1 mark)

During electroplating, indicate which electrode each material should be placed at.

Object to be plated (e.g. key)	Metal used for electroplating
[cathode] / [anode]	[cathode] / [anode]
[positive] / [negative] terminal	[positive] / [negative] terminal

Question 4 (2 marks)

a. If a current of 3.30 mA is run for 2.00 days, determine the amount of charge which passes through the wire. (1 mark)

$$Q = It$$
$$Q = 3.3 \times 10^{-3} \times 2 \times 24 \times 3600 = 5.7 \times 10^{2}\ C$$

b. 3600 mC of charge is inputted, determine the number of electrons involved. (1 mark)

$$\frac{3600 \times 10^{-3}}{96500} = 3.731 \times 10^{-5}\ mol\ (4\ sf)$$

Space for Personal Notes

Question 5 (3 marks)

A solution of Vanadium ions is reduced at a cathode, whereby it is found that a current of 9.00 A is produced over 1 hour. It was found that 2.87 g of Vanadium metal is deposited at the negative electrode.

Find the charge of the Vanadium ions.

$n(V) = \frac{2.87}{50.9} = 5.64 \times 10^{-2}$ mol

$Q = 9 \times 3600 = 3.24 \times 10^{4}$ C

$n(e^-) = \frac{3.24 \times 10^{4}}{96500} = 3.36 \times 10^{-1}$ mol

$\frac{n(e^-)}{n(V)} = \frac{5.64 \times 10^{-2}}{3.36 \times 10^{-1}} \approx 6$

$\therefore V^{6+}$

Question 6 (4 marks)

A solution of iron (II) nitrate is electrolysed using inert electrodes.

a. Write the half equation which occurs at the negative electrode. (1 mark)

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

b. The negative electrode is seen to increase in mass by 700 mg.

i. Find the amount of electrons which must have passed through the cell. (2 marks)

$n(Fe) = \frac{700 \times 10^{-3}}{55.8}$

$= 1.25 \times 10^{-2}$ mol

$n(e^-) = 2 \times 1.25 \times 10^{-2} = 2.51 \times 10^{-2}$ mol

$2.51 \times 10^{-2} \times 6.02 \times 10^{23} = 1.51 \times 10^{22}\ e^-$

ii. Find the electric charge supplied by the power source. (1 mark)

$Q = 2.51 \times 10^{-2} \times 96500$

$= 2.42 \times 10^{3}$ C 3sf

Question 7 (6 marks)

The following solution of zinc (II) fluoride is electrolysed.

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

i. Positive electrode. (1 mark)

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

ii. Negative electrode. (1 mark)

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

b. The positive electrode is found to produce oxygen gas, while the negative electrode deposits a solid at the electrode. Compare the mass of oxygen gas produced to the mass of the solid deposited at the negative electrode. (3 marks)

Assume 1mol of e^-

$n(Zn) = \frac{1}{2}n(e^-) = 0.5mol$

$m(Zn) = n \times M = 0.5 \times 65.4 = 32.7g$

$n(O_2) = \frac{1}{4}n(e^-) = 0.25mol$

$m(O_2) = n \times M = 0.25 \times 32 = 8g$

mass deposited at negative electrode is greater

c. After some time elapses, the reaction which occurs at one of the electrodes is seen to change. Write the new half-equation which occurs at this electrode. (1 mark)

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

Space for Personal Notes

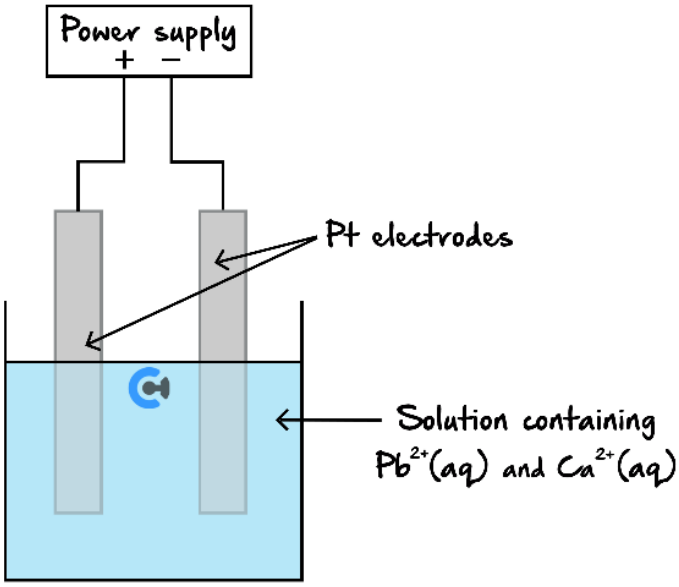
Section C: Ramping Up (14 Marks)

INSTRUCTION: 14 Marks. 11 Minutes Writing.

Question 8 (5 marks)

Inspired from VCAA Chemistry Exam 2005
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2005chem2.pdf#page=17

A mineral ore contains a mixture of compounds of Lead and Calcium, in approximately equal proportions. A chemist extracts the metal ions by roasting the ore in the air. The solution that contains the $Pb^{2+}(aq)$ and $Ca^{2+}(aq)$ is then placed in an electrolytic cell as shown in the diagram below.



a. Label the anode and cathode of the cell. (1 mark)

b. When the current begins to flow in the cell, write equations for the half-reaction that is likely to occur at the:

i. Positive electrode. (1 mark)

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

ii. Negative electrode. (1 mark)

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

c. After some time has elapsed, a new half-reaction occurs at one of the electrodes. Write the equation for this half-reaction. (1 mark)

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

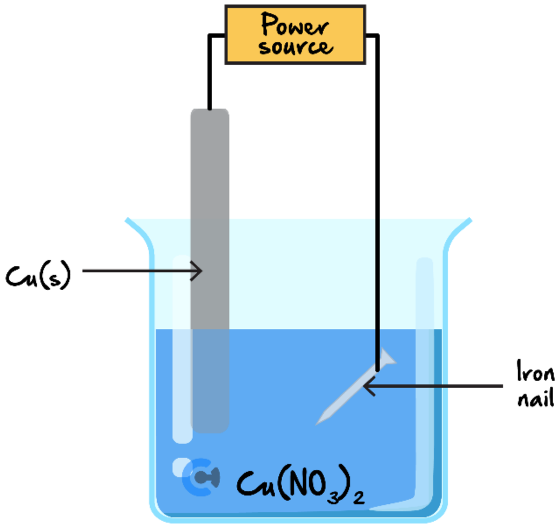
d. If the chemist had used Copper electrodes instead of Platinum electrodes, how would this have affected the half-reaction at the anode? (1 mark)

The reaction $Cu \rightarrow Cu^{2+} + 2e^-$ would have occurred (or Copper would be oxidised/a stronger reductant).

Space for Personal Notes

Question 9 (9 marks)

The following setup is used to electroplate an iron nail:



a. Explain which terminal of the power source the nail should be connected to and why. (2 marks)

Negative terminal. Object to be plated must always be the cathode (1) as this prevents it from oxidising, and allows ions to reduce and coat it (2).

b.

i. If 2.00 *A* are supplied for 20 minutes, calculate the change in mass of the anode. (3 marks)

$n(e^-) = \frac{q}{F} = \frac{It}{F} = \frac{2 \times 20 \times 60}{96500} = 2.5 \times 10^{-2}$ mol

$n(Cu) = \frac{1}{2} \times n(e^-) = 1.24 \times 10^{-2}$ mol

$m(Cu) = n \times M = 1.24 \times 10^{-2} \times 63.5 = 0.79$ g

$\Delta m = -0.79$ g

ii. Hence or otherwise, state the change in mass of the cathode. Justify your reasoning. No calculations are necessary. (1 mark)

It would be the same since the cathode half-equation is just the reverse of the anode, and thus it's the same M/z ratio, so change in mass $= g$.

c.

i. Had the cell been run for longer, but with the same current, use Faraday's laws to explain whether the change in mass observed at each electrode would be greater, lesser, or the same. (1 mark)

Q increases, and therefore according to Faraday's law, mass change observed will also increase as they are directly proportional (1).

ii. State the effect this would have on the change in mass of the entire cell. (1 mark)

Nothing, as everything remains within the cell. (This answer would hold even if this were not electroplating, as long as no gases are produced).

d. Suggest a purpose for plating the nail using copper. (1 mark)

Improving its appearance; giving it some of copper's properties such as electrical conductivity. Corrosion resistance is not the best answer as while Cu blocks water and Oxygen from the nail, as soon as there's a scratch, the nail will rust as Cu is a weaker reductant than iron.

Space for Personal Notes

Section D: Getting Trickier I (12 Marks)

INSTRUCTION: 12 Marks. 10 Minutes Writing.

The following information applies to the two questions that follow

Question 10 (1 mark)

An aqueous solution was prepared by dissolving 0.1 *mol* of each of $Sn(NO_3)_2$, $AgNO_3$, $Mn(NO_3)_2$ and $Co(NO_3)_2$ in 0.1 *L* of deionised water. The electrolysis of this aqueous solution was then carried out using graphite electrodes.

At the end of the electrolysis, the metal coatings deposited on the cathode in order from inside to outside are:

A. Sn, Mn, Ag, Co

B. Co, Sn, Ag

C. Ag, Sn, Co, Mn

D. Ag, Sn, Co

Question 11 (1 mark)

The order of increasing mass deposited at the cathode is:

A. Mn, Ag, Sn, Co

B. Co, Mn, Ag, Sn

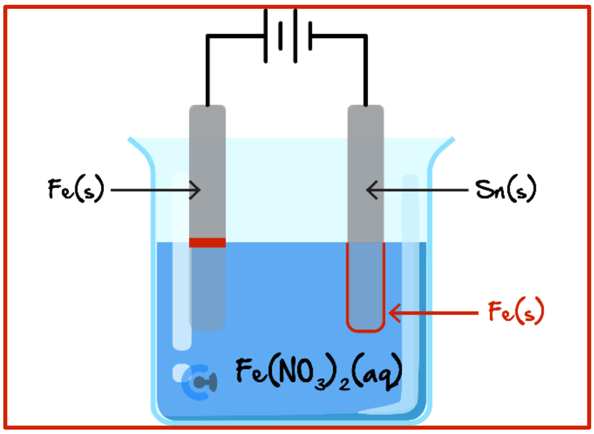
C. Mn, Co, Sn, Ag

D. Ag, Sn, Co, Mn

Space for Personal Notes

Question 12 (6 marks)

Consider an aqueous solution which contains both Fe(II) nitrate. The solution is electrolysed using a large iron anode and large tin cathode, as shown below:



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

i. Cathode. (1 mark)

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

ii. Anode. (1 mark)

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

b. A fellow chemistry pirate Usopp stated, "After 2 minutes, the colour of the solution faded."

Evaluate Usopp's statement. Justify your answer. (2 marks)

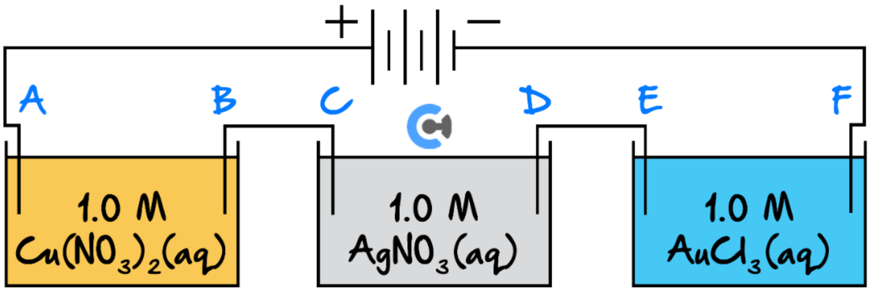
Usopp is incorrect.
2 minutes is insufficient time for the large anode to be completely oxidised. As per the databook, the colour of Fe^{2+} ion is pale green, due to the cathode reaction $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ the concentration of $Fe^{2+}(aq)$ decreases at the cathode. Simultaneously, the reaction at the anode $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$, is producing Fe^{2+} thereby there is no net change in Fe^{2+} concentration thereby the colour of the solution does not change and remains pale green.

c. After some time, the electroplating stops working. Explain this observation and draw the appearance of the cathode and anode in the diagram. (2 marks)

After some time, Fe(s) from the anode is completely oxidised and transferred to the Sn(s) cathode. Hence, the anode is no longer present in the solution thereby resulting in an open circuit. Thus, current cannot flow between the electrodes for non-spontaneous reaction to occur thereby preventing electroplating from occurring.

Question 13 (4 marks)

The connected cell given below is used, where it is found that in all three half-cells, water does not reduce at the cathode.



a. In the middle cell, state whether electrodes C and D are cathodes or anodes. (1 mark)

Electrode C	**Electrode D**
Anode	Cathode

b. In the left cell, write the overall balanced equation for the reaction which occurs. (1 mark)

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

c. For 15.0 minutes, 2.50 A of current is passed through.

i. Determine which electrode (A-F) will have the greatest an ... ted. (1 mark)

assuming 1.0 mol for e^- pass through

$n(Cu) = \frac{1}{2} n(e^-) = 0.5$ mol

$n(Ag) = n(e^-) = 1.0$ mol

$n(Au) = \frac{1}{3} n(e^-) = 0.33$ mol

Ag

ii. Determine which electrode (A-F) will have the greatest mass, in grams of metal deposited. (1 mark)

$m(Cu) = n \times M = 0.5 \times 63.5 = 31.75$ g

$m(Ag) = 1 \times 107.9 = 107.9$ g

$m(Au) = 0.33 \times 197 = 65.0$ g

Ag

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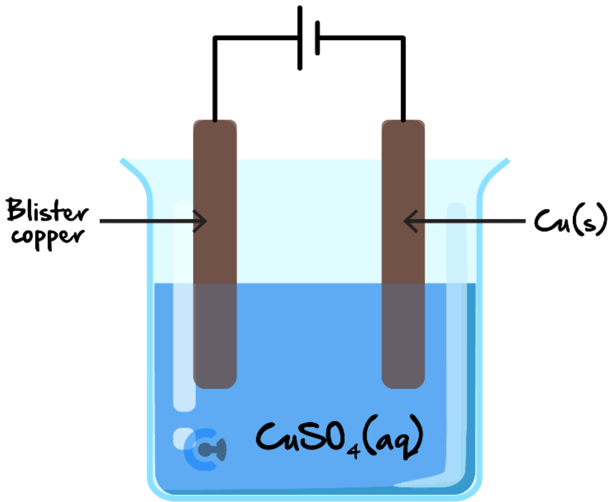
Section E: Getting Trickier II (9 Marks)

INSTRUCTION: 9 Marks. 8 Minutes Writing.

Question 14 (9 marks)

Blister copper is an impure form of copper which contains other materials such as nickel, iron, silver, cobalt, and zinc scattered throughout the blister copper.

Electro-refining copper is a method by which the impurities can be removed, which involves the following setup:



a. State whether the blister copper functions as the anode or cathode in this cell. (1 mark)

Anode

b. Electro-refining is another version of electroplating where the negative electrode has a copper coating formed on its surface.

Write the half-equation occurring at the negative electrode. (1 mark)

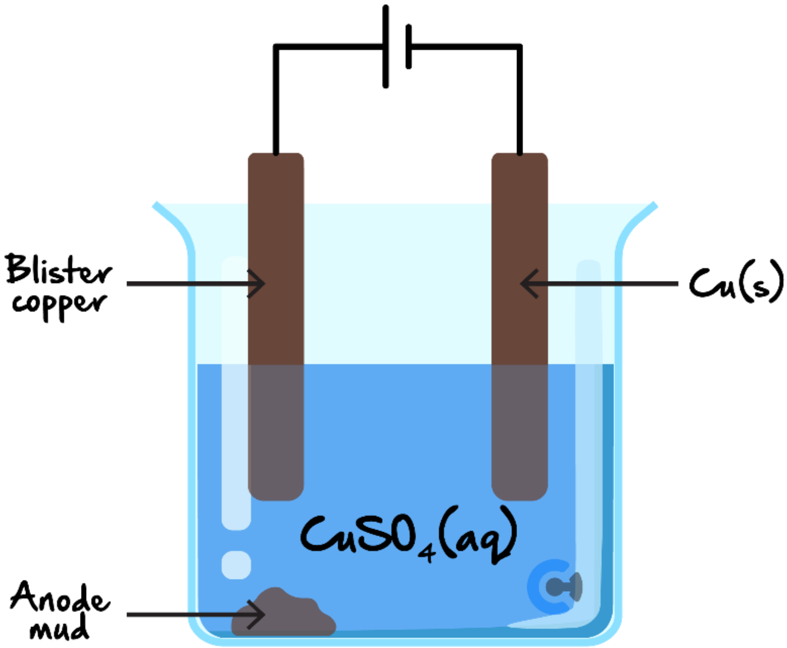
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

c. As this cell operates, the anode is seen to decrease in mass.

State why this is the case. Write an appropriate half-equation to justify your answer. (1 mark)

Because the solid blister copper oxidises into Cu^{2+}: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$, causing the electrode to become smaller.

d. Over time, a sediment is seen to form at the bottom of the beaker, known as anode mud, as shown below:



i. Explain what this would be made out of and why it forms. (2 marks)

Made out of the impurities in the blister copper, such as silver. (1)
It forms because these are weaker reductants than copper and therefore are unable to oxidise, so as the copper oxidises, the impure metals fall off the electrode and sink to the bottom as they are insoluble in water. (2)

ii. Given that this cell is powered with 5.5 A for 15 minutes, calculate the expected increase in mass at the cathode. (3 marks)

$Q = It = 5.50 \times 15 \times 60 = 4.95 \times 10^{3}\ C$

$n(e^-) = \frac{Q}{F} = \frac{4.95 \times 10^{3}}{96500} = 5.13 \times 10^{-2}\ mol$

$n(Cu) = \frac{1}{2} n(e^-) = 0.5 \times 5.13 \times 10^{-2} = 2.56 \times 10^{-2}\ mol$

$m(Cu) = n \times M = 2.56 \times 10^{-2} \times 63.5 = 1.63\ g$

After some time elapses, some ions are found to be present in the electrolyte of the electrolytic cell.

e. Apart from copper (II) ions and sulphate, list **two** other ions which will be present in the electrolyte. (1 mark)

$\boxed{2 \text{ of } Ni^{2+}, Fe^{2+}, Zn^{2+} Co^{2+}}$

Let's take a BREAK!

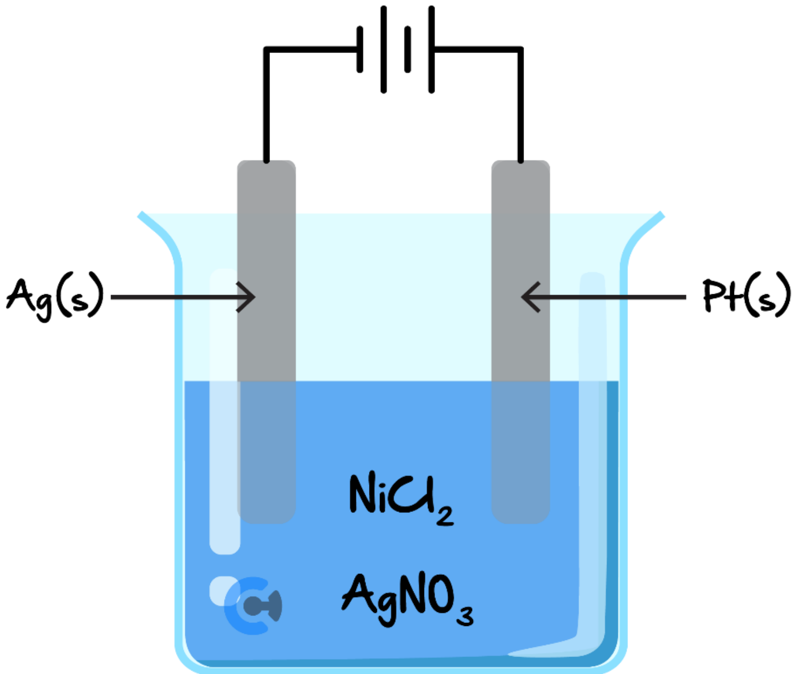
Space for Personal Notes

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

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

Question 15 (8 marks)

The following cell is constructed by passing charge through it, using 0.100 M solutions in a 50.0 mL beaker.



a. The cell is then switched on, whereby sufficient voltage is inputted into the cell.

i. In the beginning, silver is deposited onto the silver electrode. Write the balanced half-equation for this reaction. (1 mark)

$Ag^+(aq) + e^- \rightarrow Ag(s)$

ii. Explain what will happen to the concentration of silver ions in the electrolyte as the cell operates. (1 mark)

Decreases, as Ag^+ is being used yet it is not being replenished from the anode.

b. After the cell completely stops operating, it is seen that the cathode has increased in mass due to the formation of 2 layers on top of it.

i. Write the balanced half-equation for the second-order reduction reaction which occurs at the cathode. (1 mark)

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

ii. Identify what these coatings are comprised of, stating them in the order in which they form. (1 mark)

silver, nickel

c. Calculations are then done to find the point at which silver ions completely run out in the cell. Throughout the operation of the cell, 4.80 A of current is constantly passing through the cell.

i. Find the amount, in moles, of silver ions present at the beginning of the reaction in the 0.100 M of electrolyte, given that the electrolyte volume of electrolyte remains at 50.0 mL throughout the reaction. (1 mark)

$$n(\text{Ag}^+) = c \times V = 0.1\ M \times 0.050\ L = 0.0050\ mol$$

ii. Calculate the time taken for the silver ions to completely react. (3 marks)

$$n(e^-) = n(\text{Ag}^+) = 0.0050\ mol$$
$$Q = n(e^-)F = 0.0050\ mol \times 96500\ C\ mol^{-1} = 482.5\ C$$
$$t = \frac{Q}{I} = \frac{482.5}{4.8} = 100.5\ s = 101\ s$$

iii. If the cell runs for a total time of 200 s, find the total mass of metal deposited at the cathode.

$$Q = It = 4.80\ A \times (200 - 101)s = 478\ C$$
$$n(e^-) = \frac{Q}{F} = 0.00495\ mol$$
$$n(\text{Ni}) = \frac{1}{2}n(e^-) = 0.00247\ mol$$
$$m(\text{Ni}) = n \times Mr = 0.00247 \times 58.7 = 0.145\ g$$
$$m(\text{Ag}) = n \times Mr = 0.0050\ mol \times 107.9\frac{g}{mol} = 0.5375\ g$$
$$m(\text{total}) = m(\text{Ni}) + m(\text{Ag}) = 0.1452\ g + 0.5375\ g = 0.683\ g$$

Space for Personal Notes

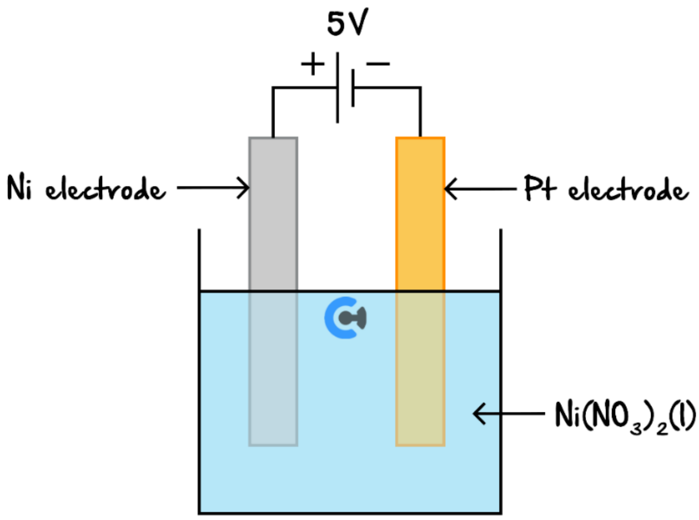
Section G: Multiple Choice Questions (5 Marks)

INSTRUCTION: 5 Marks. 5 Minutes Writing.

Inspired from VCAA Chemistry Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/2021chem-w.pdf#page=10

The following information applies to the two questions that follow

An electrolysis cell with a 5 V power supply is shown below:



Question 16 (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. $1.88 \times 10^{4}\ g$

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

$n(e^-)$ = 320 mol

$n(Ni)$ = $n(e^-)$ / 2

= 160 mol

$m(Ni)$ = 160 mol x 58.7 g mol^{-1}

= 9.39 x 10^3 g

Students who selected Option D did not make use of the reduction half-equation.

Question 17 (1 mark)

Using the electrochemical series, which one of the following changes to the electrolysis cell may reduce the amount of Ni electroplated onto the Pt electrode?

A. Replacing the Ni electrode with a Cu electrode.

B. Replacing $Ni(NO_3)_2(l)$ with 1 M $Ni(NO_3)_2(aq)$.

C. Replacing the Pt electrode with $Pb(s)$.

D. Replacing $Ni(NO_3)_2(l)$ with $NiCl_2(l)$.

Question 18 (1 mark)

Inspired from VCAA Chemistry Exam 2021
https://www.vcaa.vic.edu.au/Documents/exams/chemistry/2021/2021chem-w.pdf#page=5

An electrolysis cell consumed a charge of 4.00 C in 5.00 minutes. This represents a consumption of:

A. 4.15×10^{-5} mol of electrons.

B. 2.07×10^{-4} mol of electrons.

C. 1.93×10^{4} mol of electrons.

D. 2.41×10^{4} mol of electrons.

Since $Q = n(e^-) \times F$

$n(e^-) = Q / F$

$= 4.00\ C / 96500\ C\ mol^{-1}$

$= 4.15 \times 10^{-5}$ mol

Selection of Option B is consistent with confusing current with charge and using, inappropriately, $Q = It$

Question 19 (1 mark)

A solution of Copper ions is undergoing electrolysis using Copper electrodes. A student would observe:

A. The anode increasing in mass and the cathode decreasing in mass.

B. The concentration of the Copper ions staying constant.

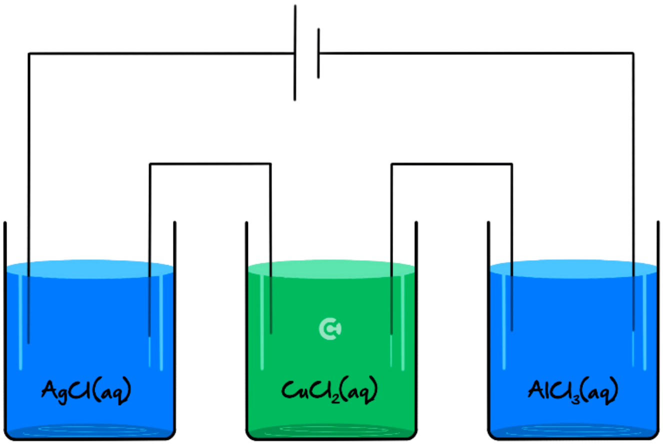
C. Electrons flowing from the negative to the positive electrode.

D. The pH of the solution decreasing.

Space for Personal Notes

Question 20 (1 mark)

Aqueous solutions of AgCl, $CuCl_2$, and $AlCl_3$ are connected in series to a power supply.



After 0.60 moles of electrons have passed through this circuit, the amounts of metal deposited will be, in moles,

A. 0.20 Ag, 0.20 Cu, 0.20 Al

B. 0.60 Ag, 0.30 Cu, 0.20 Al

C. 0.60 Ag, 0.30 Cu, 0 Al

D. 0.60 Ag, 1.20 Cu, 1.80 Al

Question 4

Answer: C

Explanation:

Aluminium will not deposit from an aqueous solution hence C is correct.
Ag^+ will have the same number of mole as the electrons.
Cu^{2+} will have half the number of mole of the electrons.

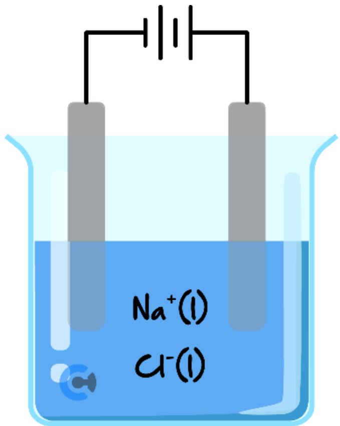
Space for Personal Notes

Section H: VCAA-Level Questions II (8 Marks)

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

Question 21 (8 marks)

The following electrochemical cell, with inert electrodes, has been constructed at **1000**°C by Bobby, who wishes to verify Faraday's findings. The density of molten sodium is 0.75 g/mL.



a. Write the overall equation occurring in this particular cell. (1 mark)

$2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

b. The products of electrolysis are immediately removed so as to avoid them spontaneously reacting.

Calculate the amount of electrons, in mol, which must have passed through the cell if 200 mL of molten sodium was collected. (2 marks)

$m(Na) = d \cdot V = 0.75 \times 200 = 150g$

$n(Na) = \frac{m}{M} = \frac{150}{23} = 6.52\ mol$

$Na^+ + e^- \rightarrow Na$

$n(e^-) = n(Na) = 6.52\ mol$

c.

i. Given that Bobby supplied a current of 30.0 A for 5.7 hours, calculate an experimental value of Faraday's constant, in $C\ mol^{-1}$. (2 marks)

$Q = It = 30 \times 5.7 \times 60 \times 60 = 6.156 \times 10^5\ C$

$F = \frac{Q}{n(e)} = \frac{6.156 \times 10^5}{6.52} = 100\ 075\ C/mol$

ii. Calculate the elementary charge, e, of each individual electron, using the value for Faraday's constant derived in **part c. i.** (1 mark)

$e = \frac{F}{N_A} = \frac{100\ 075}{6.02 \times 10^{23}} = -1.66 \times 10^{-19}\ C$

d. Propose **two** reasons as to why this cell is not used in industry to electroplate sodium metal onto objects. Justify your answer. (2 marks)

Unsafe/expensive to maintain such high temperatures (1).

There's no point plating Sodium onto anything as the moment it comes into contact with Oxygen it will re-react into Na^+ (and cause an explosion) (2).

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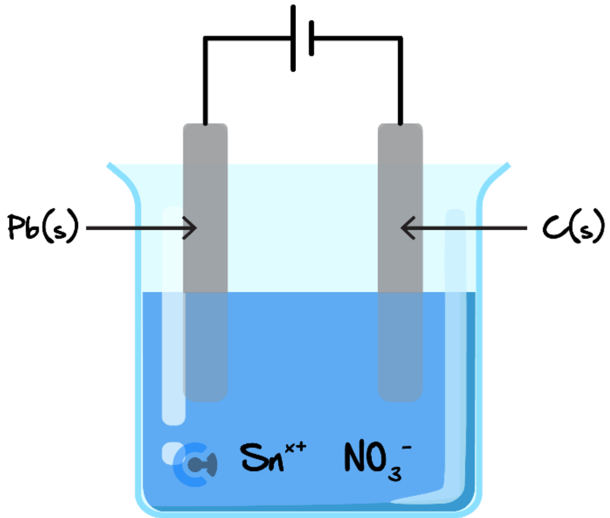
Section I: Extension Questions (16 Marks)

Question 22 (6 marks)

Students are trying to determine the oxidation state of Tin in a solution of Tin nitrate:



a. The graphite electrode was coated in 25 g of tin when 23.0 A were supplied for 36 minutes. Calculate the amount of electrons, in mol, flowing through the circuit. (1 mark)

$$n(e^-) = \frac{Q}{F} = \frac{It}{F} = \frac{23 \times 36 \times 60}{96500}$$

$$= 0.515\ mol$$

b. Given that the cell is only 80% efficient, calculate the theoretical amount of tin which **should** have deposited on the cathode, in mol, had the cell been 100% efficient. (2 marks)

$$n(Sn)_{actual} = \frac{m}{M} = \frac{25}{118.7} = 0.211\ mol$$

$$n(Sn)_{Theoretical} = \frac{0.211}{0.8} = 0.263\ mol$$

c. Hence or otherwise, calculate the charge on the Tin ion, and write the half-equation occurring at this electrode. (2 marks)

$n(Sn) : n(e^-) = \frac{0.263}{0.263} : \frac{0.515}{0.263} = 1 : 2$

$\therefore Sn^{2+}$ // $Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn(s)$ //

d. Calculate the voltage supplied by the battery in order for this cell to operate. (1 mark)

$EMF = -0.14 - (-0.13) = -0.01$

$\therefore E_{supplied} : > 0.01\ V$ //

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Question 23 (10 marks)

a. In the electrolysis of water with just enough voltage supplied (surplus of $0.01\ V$) a current of $5.2\ A$ is passed through the system for one hour.

Calculate the mass of oxygen gas, in grams, evolved at the anode at SLC, assume 48% efficiency and $1.5\ mg/L$ density. (4 marks)

$Q = It$
$= 5.2 \times 3600 = 1.87 \times 10^{4}\ C$
$n(e^-) = \frac{1.87 \times 10^{4}}{96500} = 1.94 \times 10^{-1}\ mol$

$2H_2O(l) \rightarrow 4e^- + 4H^+(aq) + O_2(g)$
$n(O_2) = \frac{1.99 \times 10^{-1}}{4} = 4.85 \times 10^{-2}\ mol$

Efficiency
$n_T(O_2) = 4.85 \times 10^{-2} \times 0.48 = 2.33 \times 10^{-2}\ mol$
$V(O_2) = 2.33 \times 10^{-2} \times 24.8 = 5.77 \times 10^{-1}\ L$
$d = \frac{m}{V}$
$m = Vd = 5.77 \times 10^{-1} \times 1.5 = 8.66 \times 10^{-1}\ mg$
$\therefore 8.7 \times 10^{-4}\ g$ 2sf

b. Identify and explain which of Faraday's two laws of electrochemistry are involved in **part a.** (3 marks)

Faraday's first law of electrochemistry is involved in **part a.** As per Faraday's first law, the mass of a substance deposited or produced at the electrode (which in this case is Oxygen gas) is directly proportional to the charge (Q) supplied. Hence, the greater the charge, the greater the mass of Oxygen gas produced.

c. It is known that another galvanic cell with HCl(aq) solution with platinum cathode and nickel anode with $Ni(NO_3)_2$(aq) solution is connected in series with the cell supplying 5.2 A in **part a.**

Calculate the new energy required for the same volume of Oxygen gas produced at the same conditions (same efficiency). (3 marks)

H^+

Ni

$EMF_{H_2O} = -0.83 - 1.23$
$= -2.06V$
∴ > 2.06V supplied
Hence 2.07V used
∴ 2.07V supplied

$EMF = 0 - (-0.25) = +0.25V$

$EMF_{Total} = 0.25 + 2.07 = +2.32V$

$n_{required}(e^-) = 1.94 \times 10^{-1} \times \frac{100}{48} = 4.04 \times 10^{-1}$ mol

$Q = n(e^-)F = 4.04 \times 10^{-1} \times 96500 = 3.9 \times 10^{4}$ C

$V = \frac{E}{Q}$

$2.32 = \frac{E}{3.9 \times 10^4}$

$E = 9.0 \times 10^4$ J 2sf

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