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VCE Chemistry ¾
Electroplating [2.4]
Homework Solutions

Admin Info & Homework Outline:

Student Name	
Questions You Need Help For	
Compulsory Questions	Pg 2 – Pg 19
Supplementary Questions	Pg 20 – Pg 28

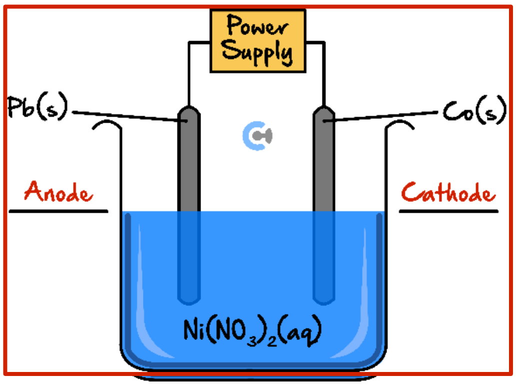
Section A: Compulsory Questions (81 Marks)

Sub-Section [2.4.1]: Identify The Electroplating Setup (Location of Object) & Find The Electroplating Reactions

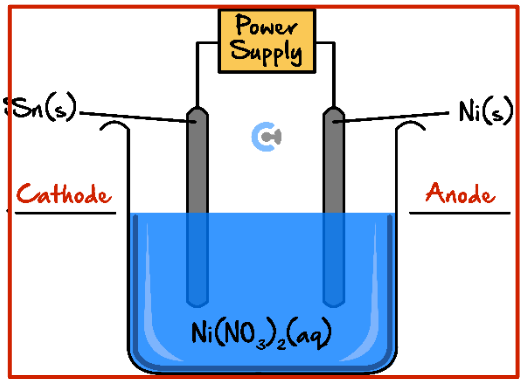
Question 1 (4 marks)

Identify the anode and the cathode in each of these cells.

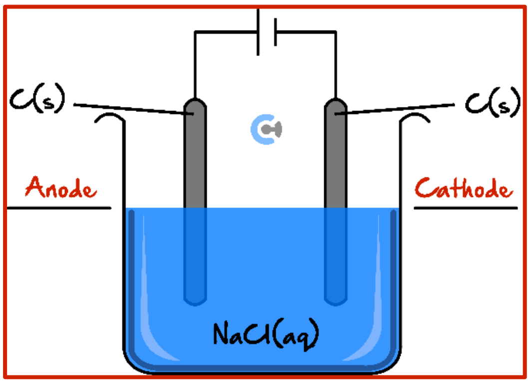
a. (1 mark)



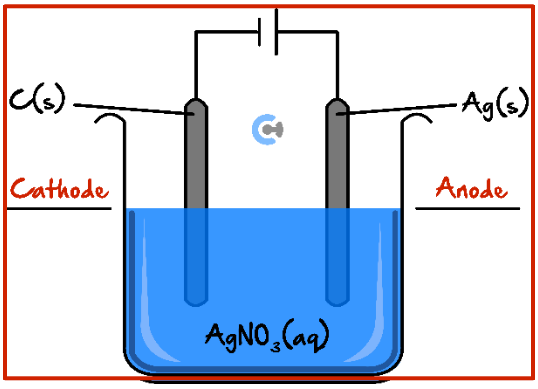
b. (1 mark)



c. (1 mark)



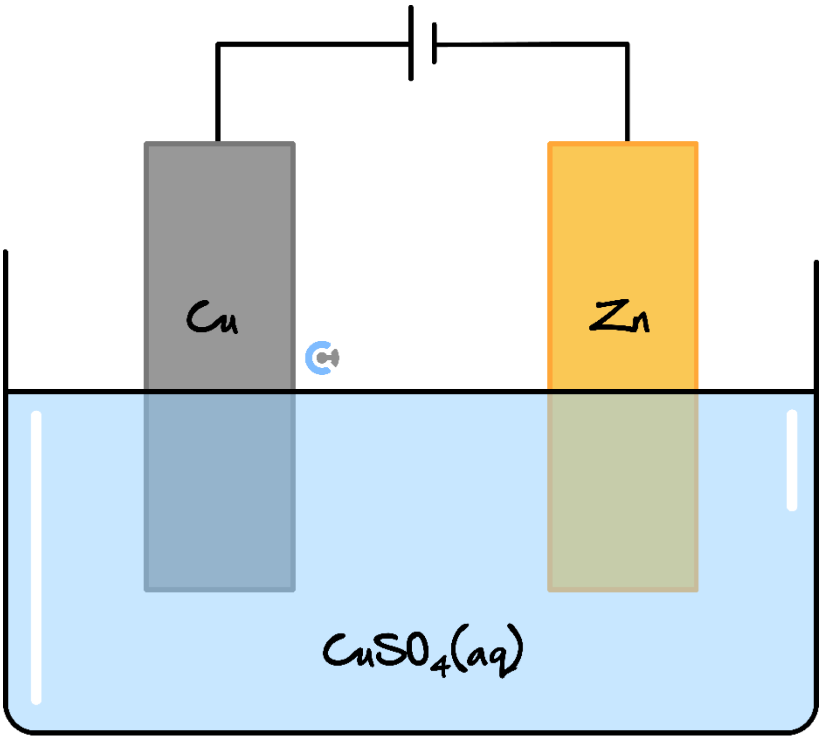
d. (1 mark)



Question 2 (6 marks)

For each of the following, write the balanced half-equation at each electrode, and the overall balanced equation.

a. (2 marks)

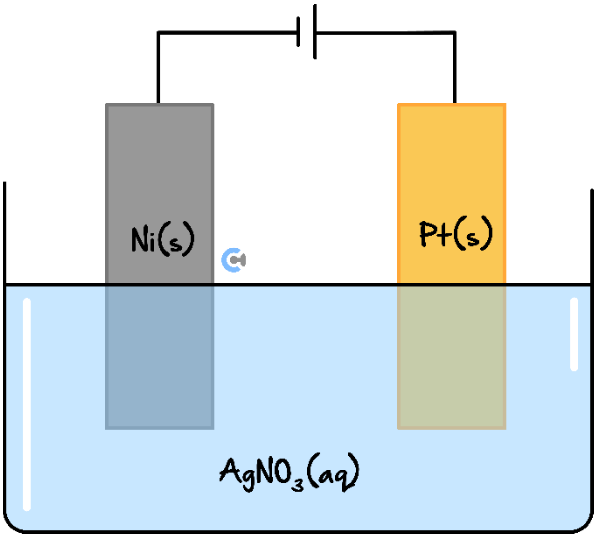


Anode: $Cu(s) \rightarrow 2e^- + Cu^{2+}(aq)$

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

Overall: $Cu(s)(Anode) \rightarrow Cu(s)(Cathode)$

b. (2 marks)



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

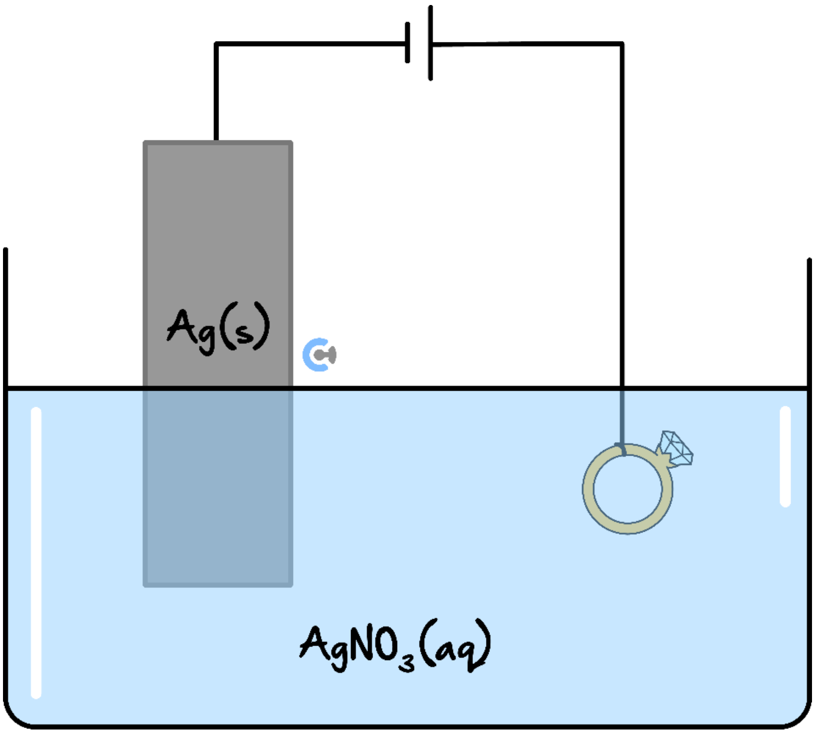
Cathode: $Ag^+(aq) + e^- \rightarrow Ag(s)$

Overall: $2H_2O(l) + 4Ag^+(aq) \rightarrow 4H^+(aq) + O_2(g) + 4Ag(s)$
Or alternatively: $2H_2O(l) + 4AgNO_3(aq) \rightarrow 4HNO_3(aq) + O_2(g) + 4Ag(s)$

Space for Personal Notes

Question 3 (7 marks)

Robert is a novice jeweller who wishes to electroplate a copper ring with silver. He sets up an electroplating apparatus as shown.



a. A fellow jeweller, Simon, points out that there is a mistake in Robert's setup. Identify the mistake Robert has made and propose a solution. (2 marks)

The ring has been attached to the positive terminal of the power supply and the silver electrode has been attached to the negative terminal (1). The silver in the silver nitrate solution will be reduced onto the silver electrode (cathode) rather than the ring (2). A solution to Robert's problem would be to swap the positions of the ring and silver electrode.

b. Robert fixes his mistake and turns the power supply on. A thin layer of silver is electroplated onto the ring. Write the balanced half-equations for the reactions occurring at the anode and the cathode. (2 marks)

Anode: $Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$

Cathode: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

c. Robert turns the power supply in his apparatus off and removes the ring. He then replaces the electrolyte for the next experiment. Simon argues that Robert should not have done this because the concentration of the electrolyte has decreased.

Evaluate Simon's statement and determine whether he is correct. (3 marks)

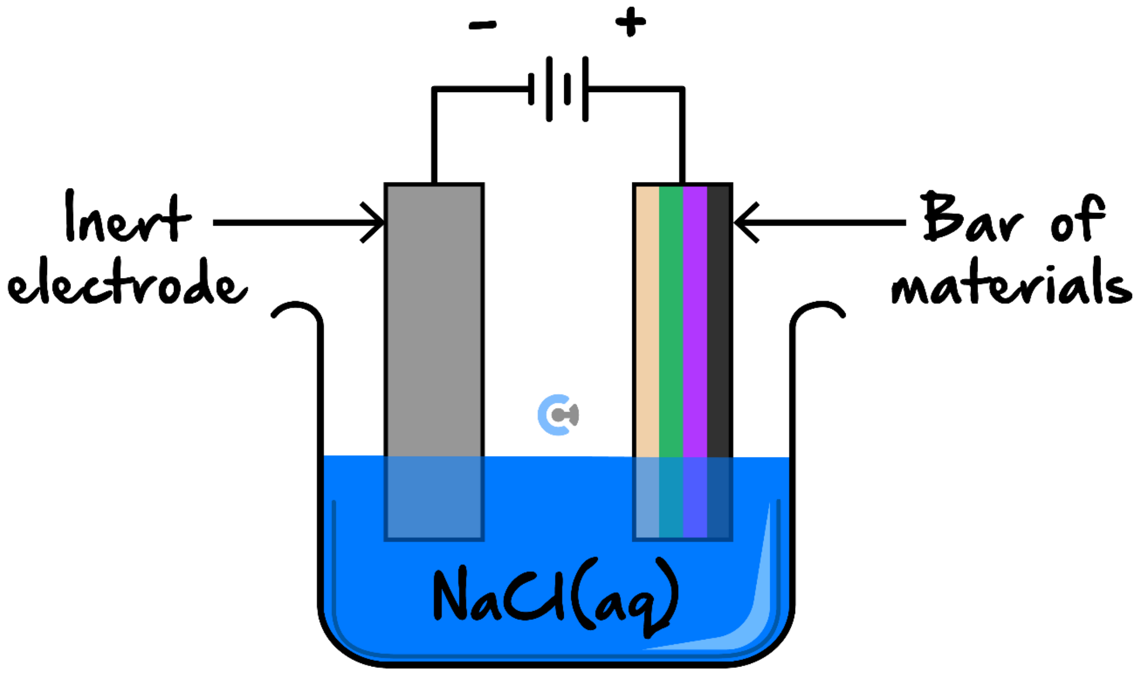
Simon is NOT correct (1). The concentration of the $AgNO_3$ solution remains the same (2). The Ag^+ that was used by the reaction at the cathode was replaced by the Ag^+ produced at the anode (3), and so the net change in Ag^+ concentration is zero.

Space for Personal Notes

Sub-Section [2.4.2]: Find Next Order Reactions During Electrolysis

Question 4 (3 marks)

A bar consisting of four materials is attached to the positive terminal of a power supply. An inert electrode is attached to the negative terminal. Both the electrode and the bar are submerged in a saltwater solution as shown below.



In each scenario, the materials in the bar are given. Circle the material that is first to react.

a. Lead, cobalt, nickel, tin. (0.5 marks)

b. Nickel, copper, zinc, manganese. (0.5 marks)

c. Silver, calcium, lead, iodine. (0.5 marks)

d. Copper, potassium, sodium, silver. (0.5 marks)

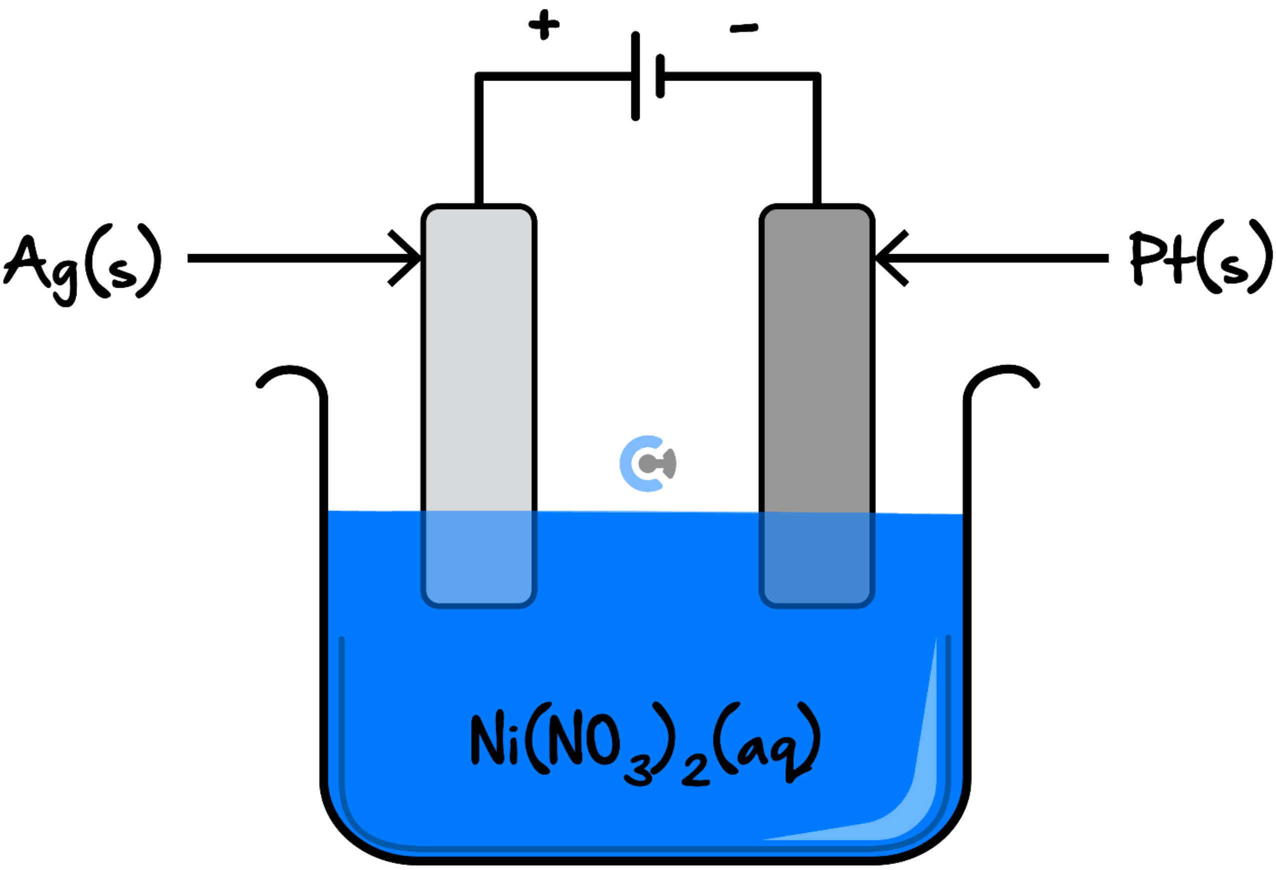
e. Copper, aluminium, magnesium, tin. (0.5 marks)

f. Iron, lead, silver, copper. (0.5 marks)

Question 5 (4 marks)

The electroplating setups below will operate continuously. For each setup, write the balanced half-equations for **all** reactions that will occur at the anode and cathode after some time elapsed, in the order they take place.

a.



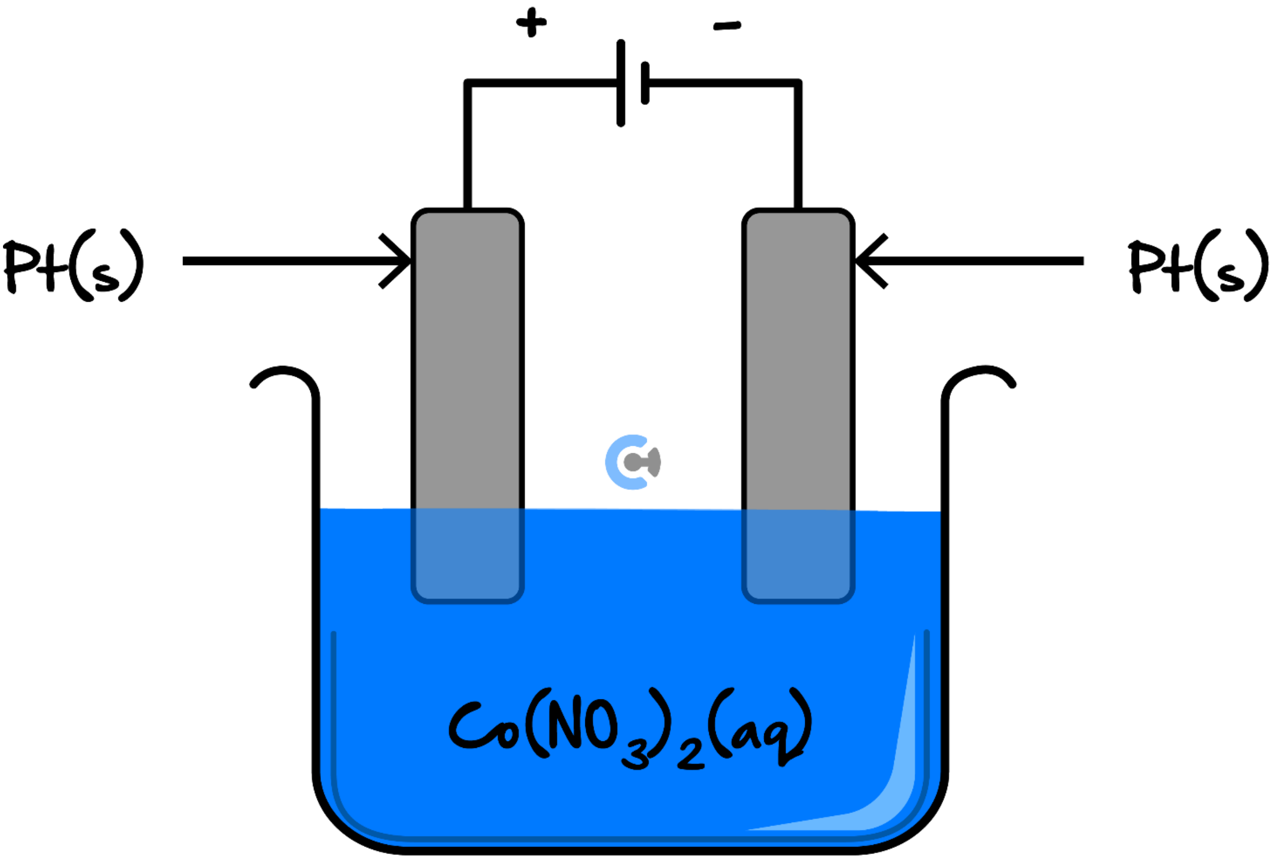
i. Anode:

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

ii. Cathode:

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

b.



i. Anode:

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

ii. Cathode:

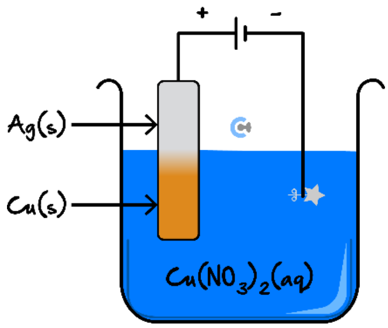
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

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

Space for Personal Notes

Question 6 (8 marks)

Robert the jeweller is working on a pair of earrings and needs to electroplate the core platinum structures with a layer of silver. He sets up his electroplating apparatus once more as shown below. This setup involves a special electrode that allows Robert to first electroplate the earring with copper to further strengthen the platinum structure, then add a layer of silver for cosmetics.



a. Explain why the copper is electroplated onto the earring before the silver. (2 marks)

Copper is a stronger reductant than silver (1) and therefore is preferentially oxidised (2). There are already copper ions in the electroplating solution which get reduced onto the earring (+1 to a maximum of two marks.).

b. Write the half-equation for the reaction occurring at the anode while the silver is being electroplated onto the earring. (1 mark)

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

Robert finishes plating one of the earrings. When he is setting up to plate the second earring, Robert accidentally inserts the special electrode upside-down; that is the silver portion is on the bottom.

c. Write the overall equation for the reaction that occurs initially for this earring. (1 mark)

$Cu(s)(Anode) \rightarrow Cu(s)\left(\frac{earring}{cathode}\right)$

d. Write the overall equation for the reaction that occurs once the copper in the special electrode has all been oxidised. (1 mark)

There is no reaction (1).

e. Describe and explain what occurs when all of the copper in the special electrode has been oxidised. (3 marks)

Since, the copper is above the silver, once all of the copper has been oxidised, there will be nothing to connect the silver to the electrode (1). The silver will have no connection to the power supply and therefore will not undergo an electrolytic reaction (2). The silver portion of the special electrode will fall to the bottom of the electroplating apparatus (3).
(Note: No spontaneous reaction will occur.)

Space for Personal Notes

Sub-Section [2.4.3]: Apply Faraday's Laws To Electroplating Calculations

Question 7 (10 marks)

a. Find the moles of electrons required to:

i. Electroplate a ring with 2.0 mol of silver. (1 mark)

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

ii. Electroplate a circuit board with 4.0 mol of copper. (1 mark)

$$n(copper) \times \frac{2}{1} = n(e^-)$$
$$4 \times \frac{2}{1} = 8\ mol$$

iii. Electroplate a necklace with 10.00 g of nickel. (2 marks)

$$n(nickel) = \frac{10.00}{58.69} = 0.1703\ mol$$
$$n(e^-) = n(nickel) \times \frac{2}{1}$$
$$n(e^-) = 0.1703 \times \frac{2}{1} = 0.3408\ mol$$

iv. Electroplate a nail with 5.00 g of zinc. (2 marks)

$$n(zinc) = \frac{5.00}{65.38} = 0.076\ mol$$
$$n(e^-) = n(zinc) \times \frac{2}{1}$$
$$n(e^-) = 0.076 \times \frac{2}{1} = 0.153\ mol$$

b. 2.0 mol of electrons are passed through some electrolytic setups:

i. Find the current which passes through an electroplating cell with silver attached to the positive terminal, and platinum to the negative terminal if the cell runs for 3.00 h. (2 marks)

$It = Fn(e^-) \Rightarrow I = \frac{Fn(e^-)}{t}$

$I = \frac{96500 \times 2}{3 \times 60 \times 60} = 17.87$ amperes

ii. An electroplating cell is set up with inert electrodes and a potassium chloride electrolyte at 1.0 M concentration. Find the volume of O_2 evolved at the anode at SLC. (2 marks)

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

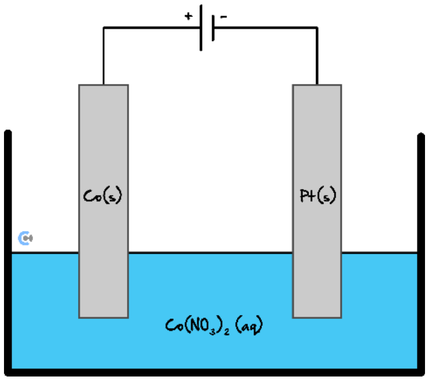
$2 \times \frac{1}{4} = 0.5$ moles

$V(O_2) = 0.5 \times 24.8 = 12.4\ L$

Space for Personal Notes

Question 8 (14 marks)

a. The following electroplating apparatus has been running for 1.00 h. The current through the apparatus is 5.00 A.



i. Find the mass of cobalt, in grams, which has been oxidised at the anode. (3 marks)

$It = Fn(e^-) \Rightarrow n(e^-) = \frac{It}{F}$

$n(e^-) = \frac{It}{F} = \frac{5\times(1\times60\times60)}{96500} = 0.1865\ mol$

$n(\text{Co}) = n(e^-) \times \frac{1}{2}$

$n(\text{Co}) = 0.1865 \times \frac{1}{2} = 0.0933$

$m(\text{Co}) = n(\text{Co}) \times 58.93 = 0.0933 \times 58.93 = 5.496\ g$

ii. Find the mass of cobalt has been reduced at the cathode. (2 marks)

$n(e^-) = 0.1865\ mol$ (from **part a.i.**)

$n(\text{Co}) = n(e^-) \times \frac{1}{2}$

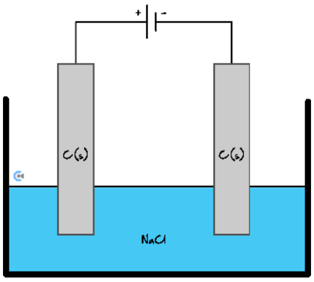
$n(\text{Co}) = 0.1865 \times \frac{1}{2} = 0.0933$

$m(\text{Co}) = n(\text{Co}) \times 58.93 = 0.0933 \times 58.93 = 5.496\ g$

iii. Has the concentration of $Co(NO_3)_2$ changed? Explain your answer. (2 marks)

No (1), the Co^{2+} ions taken out of the $Co(NO_3)_2$ by the reduction reaction, were replaced by the same amount of Co^{2+} ions produced at the anode by the oxidation reaction (2).

b. Another electrolytic cell has been left running for 9.00 days at SLC. The current through this cell is a constant 10.00 A.



i. Find the volume of H_2 produced by this cell. (3 marks)

$It = Fn(e^-) \Rightarrow n(e^-) = \frac{It}{F}$

$n(e^-) = \frac{It}{F} = \frac{10 \times (9 \times 24 \times 60 \times 60)}{96500} = 80.58\ mol$

$n(H_2) = n(e^-) \times \frac{1}{2}$

$n(H_2) = 80.58 \times \frac{1}{2} = 40.29\ mol$

$V(H_2) = n(H_2) \times 24.8 = 40.29 \times 24.8 = 999.2\ L$

ii. Find the volume of O_2 produced by this cell. (2 marks)

$It = Fn(e^-) \Rightarrow n(e^-) = \frac{It}{F}$

$n(e^-) = \frac{It}{F} = \frac{10 \times (9 \times 24 \times 60 \times 60)}{96500} = 80.58$ moles

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

$n(O_2) = 80.58 \times \frac{1}{2} = 20.15$

$V(O_2) = n(H_2) \times 24.8 = 20.15 \times 24.8 = 499.6\ L$

iii. Explain the relationship between the volume of H_2 produced and O_2 produced. Use any relevant equation to justify your answer. (2 marks)

Overall Equation: $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ (1)
Hydrogen gas and oxygen are produced in the ratio 2: 1. This means for every litre of oxygen produced, two liters of hydrogen will be produced (2).

Question 9 (10 marks)

a. An electroplating apparatus is set up using a platinum electrode connected to the positive terminal, a copper disc connected to the negative terminal to be electroplated and a 250 mL solution of 1.0 M $AgNO_3$. The current through the setup is 5.00 A.

i. State the amount of Ag^+ ions, in moles, present in the solution. (1 mark)

$n(Ag^+) = [Ag^+] \times V$
$n(Ag^+) = 1.0 \times 0.250 = 0.25\ mol$

ii. How long, to the nearest minute, would it take to reduce the concentration of the $AgNO_3$ solution to 0 M? (2 marks)

$n(e^-) = n(Ag^+) \times \frac{1}{1}$
$n(e^-) = 0.250 \times \frac{1}{1} = 0.250\ mol$ (1)
$It = Fn(e^-) \Rightarrow t = \frac{Fn(e^-)}{I}$
$t = \frac{96500 \times 0.250}{5} = 4825$ seconds $= 80.42$ minutes (2)
$t = 81$ minutes (accept $t = 80$ minutes)

b. Another electroplating apparatus is set up using a silver electrode, an iron nail to be galvanised and 200 mL of a 2.0 M solution of $ZnSO_4$. 7.00 A of current is passed through the setup.

i. Write the balanced equation for the overall reaction. (2 marks)

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

ii. Find the time, in minutes, required to oxidise 3.00 g of silver. (2 marks)

$n(\text{Ag}) = \frac{3}{107.87} = 0.0278$

$n(e^-) = n(\text{Ag}) \times \frac{1}{1}$

$n(e^-) = 0.0278 \times \frac{1}{1} = 0.0278$ moles (1)

$It = Fn(e^-) \Rightarrow t = \frac{Fn(e^-)}{I}$

$t = \frac{96500 \times 0.0278}{7} = 383.398$ seconds $= 6.39$ minutes (2)

iii. Find the mass of zinc plated on the iron nail after 3.00 g of silver has been oxidised. (3 marks)

From **part III.**, $t = 383.398$ seconds

$It = Fn(e^-) \Rightarrow \frac{It}{F} = n(e^-)$

$n(e^-) = \frac{7.00 \times 383.398}{96500} = 0.0278$ moles (1)

$n(\text{Zn}) = n(e^-) \times \frac{1}{2}$

$n(\text{Zn}) = 0.0278 \times \frac{1}{2} = 0.0139$ moles

$m(\text{Zn}) = 65.38 \times 0.0139 = 0.909$ grams (2)

OR

$n(\text{Ag}) \times \frac{1}{2} = n(\text{Zn}) \Rightarrow m(\text{Ag}) \times \frac{1}{2} \times \frac{1}{107.87} \times 65.38 = m(\text{Zn})$ (1)

$m(\text{Zn}) = 3.00 \times \frac{1}{2} \times \frac{1}{107.87} \times 65.38 = 0.9088 = 0.909$ grams (2)

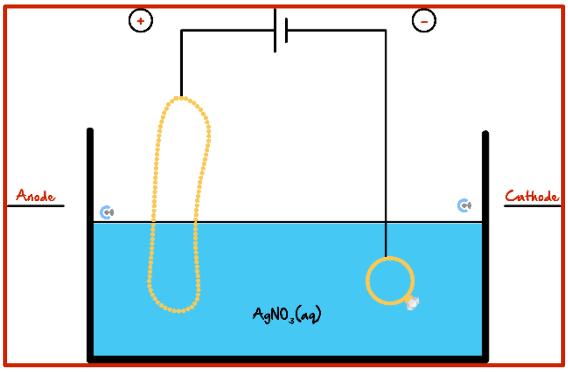
Space for Personal Notes

Sub-Section: Final Boss

Question 10 (15 marks)

Max the jeweller needs to electroplate a platinum ring with a protective layer of silver. Max uses 500 mL of a 2.00 M $AgNO_3$ electroplating solution.

A platinum electrode is used and attached to the positive terminal, whereby 7.00 A of current will run through the setup when the power supply is turned on. The electroplating apparatus Max uses is represented below.



a. Label each electrode as positive or negative in the circles provided on the diagram. (1 mark)

b. Label the anode and the cathode on the diagram above. (1 mark)

c. Write the half-equations for the reactions that will initially occur at the anode and the cathode. (2 marks)

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

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

Max turns on his electroplating apparatus and takes a nap.

d. Find the time, in hours, that the silver be reduced onto the ring. (4 marks)

$n(Ag^+) = [Ag^+] \times V$

$n(Ag^+) = 2.00 \times 0.500 = 1.00\ mol$ (1)

$n(e^-) = n(Ag^+) \times \frac{1}{1}$

$n(e^-) = 1.00 \times \frac{1}{1} = 1.00\ mol$ (2)

$It = Fn(e^-) \Rightarrow t = \frac{Fn(e^-)}{I}$

$t_1 = \frac{96500 \times 1.00}{7.00} = 13785.71\ s = 229.76\ min = 3.83\ h$ (3)

e. Write the half-equations for the reactions that occur at the anode and the cathode once the Ag^+ ions in the solution have all been reduced onto the ring. (2 marks)

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

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

f. Identify and explain one safety concern in this setup. (1 mark)

Hydrogen gas is produced (1). Hydrogen gas is very flammable (needed for mark).

The volume of hydrogen gas in the room should not exceed 10 L.

g. Find the maximum length of time, in hours, for which Max can sleep before he should turn his electroplating apparatus off. (4 marks)

$n(H_2)_{max} = \frac{10}{24.8} = 0.4032\ mol$ (1)

$n(e^-)_{max} = \frac{2}{1} \times n(H_2)_{max}$

$n(e^-)_{max} = \frac{2}{1} \times 0.4032 = 0.8065\ mol$ (2)

$It = Fn(e^-) \Rightarrow t = \frac{Fn(e^-)}{I}$

$t_2 = \frac{96500 \times 0.8065}{7.00} = 11117.51\ s = 185.29\ min = 3.088\ h$ (3)

Adding the Silver electroplating time from **part d.**

$t_1 + t_2 = 3.83 + 3.088 = 6.918\ h = 6.92\ h$ (4)

(Accept 6 hours, 55 min)

Space for Personal Notes

Section B: Supplementary Questions (37 Marks)

Sub-Section [2.4.1]: Identify The Electroplating Setup (Location of Object) & Find The Electroplating Reactions

Question 11 (1 mark)

When electroplating a metallic key with Cu metal:

A. The key must be connected to the negative terminal of the power supply.

B. The electrolyte can be a solution of $CuSO_4$.

C. The anode can be made from Cu metal.

D. All of the above.

Question 12 (1 mark)

A student wanted to cover an iron key with copper metal. Which of the following experimental set-ups is **incorrect**?

A. The student connected the key to the negative terminal of a power supply.

B. The student used copper (II) sulphate solution as the electrolyte.

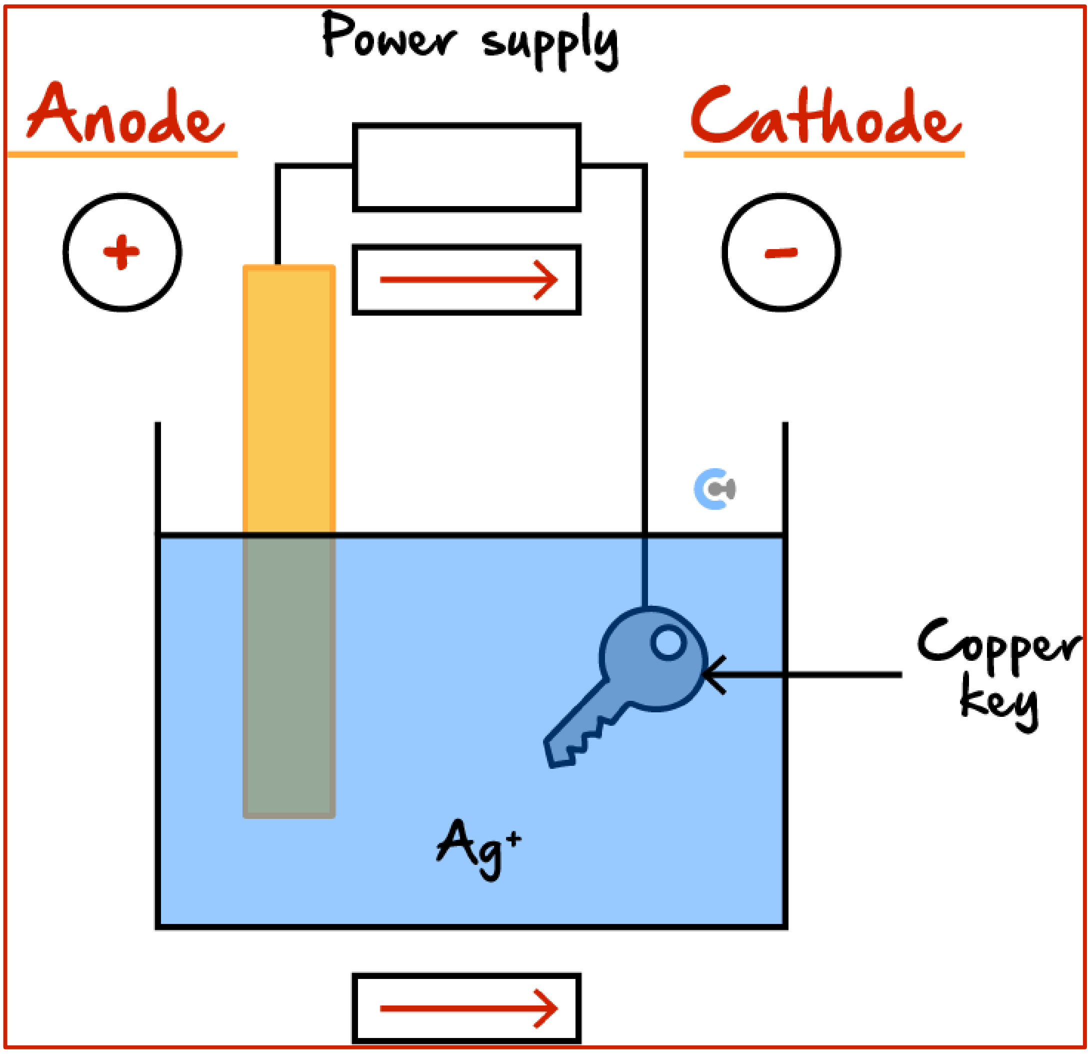
C. The student used an iron rod as the cathode.

D. The student connected an iron rod to the positive terminal of a power supply.

Space for Personal Notes

Question 13 (10 marks)

The following electroplating apparatus has been set up by a student. A 1.2 *M* solution of $AgNO_3$ was used as an electrolyte.



a. Label the following on the diagram.

i. The cathode and anode on the **orange** lines. (1 mark)

ii. The charges of the cathode and anode in the circles. (1 mark)

iii. An arrow showing the direction of the flow of electrons in the **purple** box. (1 mark)

iv. An arrow showing the direction of Ag^+ ions in the **blue** box. (1 mark)

b. Determine the oxidation and reduction half-reactions for this cell. (1 mark)

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

c. The student runs the electroplating apparatus for 30 minutes and makes several observations.

- The key increases in mass.
- There are bubbles produced at the silver electrode once 15 minutes have passed.
- The silver electrode increases in mass.

Evaluate these observations. If the observation is impossible, explain why. (5 marks)

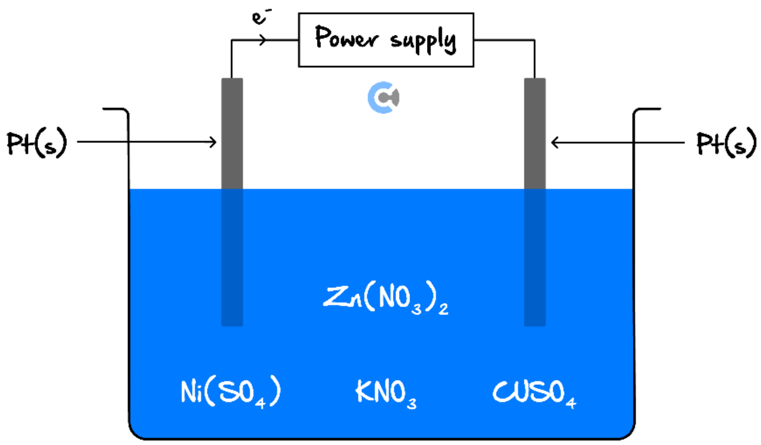
The key does increase in mass (1). Bubbles should not appear (2) at the silver electrode as there is no gas being produced at the anode (3). The silver electrode cannot increase (4) in mass as Ag is being oxidised into the solution (5).

Space for Personal Notes

Sub-Section [2.4.2]: Find Next Order Reactions During Electrolysis

Question 14 (8 marks)

Ziggy sets up an electrolytic cell containing two platinum electrodes with 0.1 M of $CuSO_4(aq)$, $NiSO_4(aq)$, $Zn(NO_3)_2(aq)$ and $KNO_3(aq)$ as shown in the diagram below.



a. State and explain what metal will accumulate on the cathode immediately when the power supply is turned on. (2 marks)

Copper (1). Copper cations are the strongest oxidants present (2).

b. Ziggy notes that after 5 minutes, the solution which was initially blue has turned transparent.

i. Explain this observation. (1 mark)

$CuSO_4$ (blue in colour) has all been reduced onto the cathode.

ii. At this point, explain which metal is being electroplated at the cathode. (2 marks)

Nickel (1). Nickel cations are the next strongest oxidants after copper cations (2).

c. After 30 minutes of the cell running, Ziggy notices bubbling occurring at the cathode. His friend, Gain, claims that "the bubbling is due to impurity metal ions in the solution reacting with water." Evaluate Gain's statement and explain the reason for the bubbling. (3 marks)

Gain is incorrect (1). The bubbling is caused by the electrolysis of water (2). This occurs because all of the copper, nickel and zinc cations have reacted and H_2O is the strongest oxidant after zinc cations (3).

Space for Personal Notes

Sub-Section [2.4.3]: Apply Faraday's Laws To Electroplating Calculations

Question 15 (4 marks)

Aanya sets up an electrolytic cell and leaves it running for a period of time.

a. The cell is running for 30 minutes with 4.00 A of current running through the cell. Calculate the electrical charge running through the cell. (2 marks)

$Q = I\,t$

$Q = 4.00 \times 30 \times 60 = 7200\ Coulombs$

b. Aanya decides to decrease the time that the cell runs to 10 minutes and finds that 900 C of electrical charge runs through the cell. Calculate the current running through the cell. (2 marks)

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

$I = \frac{900}{10 \times 60} = 1.5\ A$

Space for Personal Notes

Question 16 (5 marks)

a. Daniel passes 2.450 mol electrons through a cell. Calculate the amount of electrical charge which passes through. (2 marks)

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

$Q = 96500 \times 2.450 = 2.364 \times 10^5$ *Coulombs*

b. In another experiment, Daniel runs the cell for 45 minutes, at 4.20 A. Calculate the number of electron moles running through the cell. (3 marks)

$I\,t = F\,n(e^-) \Rightarrow n(e^-) = \frac{I\,t}{F}$

$n(e^-) = \frac{4.20 \times 45 \times 60}{96500} = 0.1175$ *moles*

Space for Personal Notes

Question 17 (5 marks)

An electrolytic cell was set up using an unknown, molten metal salt, MBr_3. A current of 1.25 A was passed through the molten compound for 50.0 minutes to deposit 0.675 g of the metal.

a. Write a balanced half-equation for the anode and cathode reactions in this electrolytic cell. (2 marks)

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

$M^{3+}_{(l)} + 3e^- \rightarrow M_{(l)}$

b. Calculate the charge passed through the cell. (1 mark)

$Q = It \qquad Q = 50 \times 60 \times 1.25 = 3750\ Coulombs$

c. Calculate the moles of metal deposited. (2 marks)

$It = Fn(e^-) \Rightarrow n(e^-) = \frac{It}{F}$

$n(e^-) = \frac{50 \times 60 \times 1.25}{96500} = 0.03886\ moles$

$n(M) = \frac{1}{3} \times n(e^-)$

$n(M) = \frac{1}{3} \times 0.03886 = 0.01295\ moles$

d. Identify the metal deposited. (2 marks)

$M_R = \frac{m(M)}{n(M)} = \frac{0.675}{0.01295} = 52.11\ g/mol$

corresponds most closely with Chromium

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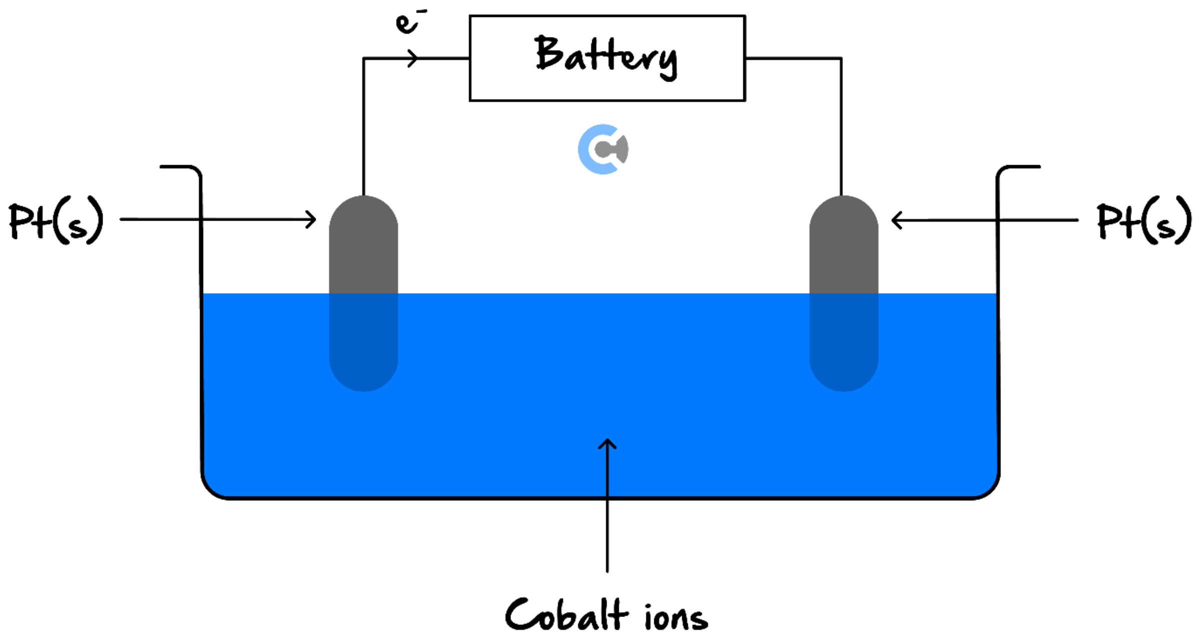
Question 18 (3 marks)

Kynan sets up a cell shown below.



Kynan runs the cell for 20.0 minutes with 5.50 A running through the cell and finds that 2.0 g of metal has been deposited on the cathode.

Find the charge of the cobalt cation.

moles of cobalt:

$n(Co) = \frac{2.0}{58.93} = 0.0339\ moles$

moles of electrons:

$I\,t = F\,n(e^-) \Rightarrow n(e^-) = \frac{I\,t}{F}$

$n(e^-) = \frac{20\times60\times5.50}{96500} = 0.0684\ moles$

ratio:

$\frac{n(Co)}{n(e^-)} = \frac{0.0339}{0.0684} \approx \frac{1}{2}$

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

charge on each side must be equal $\Rightarrow$ *charge of Co cation* $= 2+$

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