

Faraday's law of electrolysis

76.

- Charge $Q = I \times t = 0.5 \times 600 = 300 \text{ C}$.
- $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$, so 1 mole e^- deposits 1 equivalent (eq. wt = 108 g).
- Mass $m = (\text{equivalent weight}) \times \frac{Q}{F} = 108 \times \frac{300}{96500}$.
- $m \approx 108 \times 0.003109 = 0.33598 \text{ g} \approx 0.336 \text{ g}$.

Answer: (b) 0.336 g

77. Answer: (c) Ampere

78.

- Moles $\text{H}_2 = 0.112/22.4 = 0.005 \text{ mol}$.
- Electrons needed = $0.005 \times 2 = 0.01 \text{ mol e}^-$.
- Charge $Q = 0.01 \times 96500 = 965 \text{ C}$.

Answer: (c) 965 Coulomb

79.

- Resistance $\propto 1/\text{conductance}$ and conductance increases with ion concentration. The lowest concentration gives the highest resistance.

Answer: (a) 0.05 N NaCl

80. (b) In 5 gm CuO , 4 gm Cu and 1 gm O be present.

Element	Wt.	At Wt.	$Wt. / At. Wt. = x$	Ratio
Cu	4 gm	63.5	$4/63.5 = .0625$	$\frac{.0625}{.0625} = 1$



O	1 gm	16	$1/16 = .0625$	$\frac{.0625}{.0625} = 1$
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Emperical formula = CuO of oxide

In this oxide, oxidation no. of $Cu = +2$

Equivalent weight = $\frac{\text{Molecular weight}}{\text{Oxidation no.}} = \frac{63.5}{2} \approx 31.75$ but Equivalent weight should be an

integral no. = 32

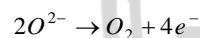
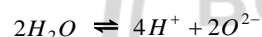
81. (c) Given, Current = 241.25 *columb*

1 *coulomb* current will deposit = $1.118 \times 10^{-3} \text{ gm Ag}$.

\therefore 241.25 current will deposit = $1.118 \times 10^{-3} \times 241.25$

= 0.27 gm silver.

82. (b) Reaction for electrolysis of water is



$\therefore n = 4$ so 4 Faraday charge will liberate

1 mole = 22.4 dm^3 oxygen

\therefore 1 Faraday charge will liberate $\frac{22.4}{4} = 5.6 \text{ dm}^3 O_2$.

83. (a) $Na^+ + e^- \rightarrow Na$

Charge (in F) = moles of e^- used = moles of Na deposited

$$= \frac{11.5}{23} \text{ gm} = 0.5 \text{ Faraday.}$$

84. (c) Hydrolysis of water : $2H_2O \rightleftharpoons 4H^+ + 4e^- + O_2$

4 F charge will produce = 1 mole $O_2 = 32 \text{ gm } O_2$

1 F charge will produce = $\frac{32}{4} = 8 \text{ gm } O_2$.

85. (c) In a galvanic cell, the electrons flow from anode to cathode through the external circuit.

At anode (–ve pole) oxidation and at cathode (+ pole) reduction takes place.



86. (e) Number of equivalents of silver formed = Number of equivalents of copper formed.

In $AgNO_3$, Ag is in +1 oxidation state.

In $CuSO_4$, Cu is in +2 oxidation state.

$$\text{Equivalent weight of } Ag = \frac{108}{1} = 108$$

$$\text{Equivalent weight of } Cu = \frac{63.6}{2} = 31.8$$

$$\frac{M_1}{M_2} = \frac{E_1}{E_2} ; \frac{10.79}{M_{Cu}} = \frac{108}{31.8}$$

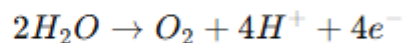
$$M_{Cu} = \frac{10.79 \times 31.8}{108} = 3.2 \text{ gm.}$$

82.

Answer: (a) 11.2 dm^3

Explanation:

- Electrolysis reaction at anode:



- 4 Faradays produce 1 mole of O_2 .
- 1 mole O_2 at NTP = 22.4 L.
- 1 F produces: $22.4/2 = 11.2 \text{ dm}^3 O_2$

83. Answer: (a) 0.5 F

Explanation:

- Molar mass of Na = 23 g/mol \rightarrow 1 g eq. = 23 g (since Na is monovalent)
- Fraction of 1 F required: $11.5/23 = 0.5 \text{ F}$

84.

Answer: (c) 8 g of oxygen



Explanation:

- 4 F produce 1 mole $O_2 = 32$ g
- 1 F produces $32 / 4 = 8$ g of O_2

85. Answer: (c) Anode \rightarrow Cathode through external circuit

Explanation:

- Electrons are released at the **anode (oxidation)** and flow through the external circuit to the **cathode (reduction)**.

86. Answer: (e) 3.2 g

Step-by-Step:

1. Silver deposition: $Ag^+ + e^- \rightarrow Ag$
 - Eq. wt of Ag = 108 g
 - Moles of electrons used: $10.79/108=0.1$ eq. $10.79 / 108 = 0.1$ eq.
2. Copper deposition: $Cu^{2+} + 2e^- \rightarrow Cu$
 - Eq. wt of Cu = $63.5 / 2 \approx 31.75$ g/equiv
 - Mass deposited: $0.1 \times 31.75 \approx 3.2$ g

87. (b) Laws of electrolysis were proposed by Faraday.

88. (a) Given, Current (i) = 25 mA = 0.025 A

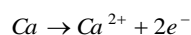
Time (t) = 60 sec

$Q = i t = 60 \times 0.025 = 1.5$ coulombs



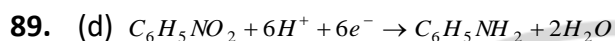
$$\text{No. of electrons} = \frac{1.5 \times 6.023 \times 10^{23}}{96500}$$

$$e^- = 9.36 \times 10^{18}$$



$2e^-$ are required to deposit one Ca atom

$$9.36 \times 10^{18} e^- \text{ will be used to deposit} = \frac{9.36 \times 10^{18}}{2} = 4.68 \times 10^{18}.$$



1 mole = 123 gm nitrogen requires 6 mole electron e^-

= 6×96500 coulomb charge

$$\therefore 12.3 \text{ gm nitrobenzene will require} = \frac{6 \times 96500 \times 12.3}{123} = 6 \times 9650 = 57900 \text{ C.}$$

90. (c) Au and Ag settle down below the anode as anode mud during the process of electrolytic refining of copper.

91. Answer: (b) 41 hrs

Explanation:

1. Zinc oxidation: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$
2. Moles of Zn: $50 / 65 \approx 0.769 \text{ mol}$
3. Each mole of Zn releases 2 moles of electrons \rightarrow total moles of e^- : $0.769 \times 2 \approx 1.538 \text{ mol}$
4. Total charge: $Q = 1.538 \times 96500 \approx 148,400 \text{ C}$
5. Current $I = 1 \text{ A} \rightarrow$ time $t = Q / I = 148,400 \text{ s} \approx 41.1 \text{ hrs}$

92. Answer: (b) 18 g

Explanation:



1. Moles of $\text{Cl}_2 = 11.2 / 22.4 = 0.5 \text{ mol}$
2. Electrons: 2 e^- per mole $\text{Cl}_2 \rightarrow 0.5 \times 2 = 1 \text{ mol e}^-$
3. $\text{Al}^{3+} + 3 \text{ e}^- \rightarrow \text{Al} \rightarrow \text{Mass Al deposited} = 27 \times (1 / 3) = 9 \text{ g}$? Wait carefully:

Step:

- Cl_2 : $0.5 \text{ mol} \rightarrow 1 \text{ mol e}^-$ required?

Actually: $2 \text{ Cl}^- \rightarrow \text{Cl}_2 + 2 \text{ e}^- \rightarrow 0.5 \text{ mol Cl}_2$ requires 1 mol e^-

- $\text{Al}^{3+} + 3 \text{ e}^- \rightarrow \text{Al} \rightarrow \text{Mass} = 27 \times (1 \text{ mol e}^- / 3) = 9 \text{ g}$

So correct mass of Al = 9 g

93. Answer: (d) 11.2 cm^3

Explanation:

1. Silver deposition: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$, mass = 0.108 g
2. Equivalent weight of Ag = $108 \text{ g/equiv} \rightarrow$ moles of electrons = $0.108 / 108 = 0.001 \text{ mol e}^-$
3. Water electrolysis at anode: $2 \text{ H}_2\text{O} \rightarrow \text{O}_2 + 4 \text{ H}^+ + 4 \text{ e}^-$
4. Moles $\text{O}_2 = 0.001 / 4 = 0.00025 \text{ mol} \rightarrow \text{Volume} = 0.00025 \times 22.4 \text{ L} = 0.0056 \text{ L} = 5.6 \text{ mL}$? Wait options in cm^3 .

$$0.0056 \text{ L} = 5.6 \text{ cm}^3$$

94. Answer: (b) 5.6 L

Explanation:





1. Moles of O_2 : $4 / 32 = 0.125 \text{ mol}$
2. H_2 is produced at cathode: $2 H^+ + 2 e^- \rightarrow H_2$
3. $4 e^-$ produce 1 $O_2 \rightarrow 2 e^-$ produce 1 $H_2 \rightarrow$ Moles $H_2 = 0.125 \times 2 = 0.25 \text{ mol}$
4. Volume H_2 at NTP = $0.25 \times 22.4 = 5.6 \text{ L}$

