## Faraday's law of electrolysis

**76.**

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

**Answer:** (b) **0.336 g**

**77.** **Answer:** (c) **Ampere**

**78.**

- Moles $H_2$ = $0.112/22.4 = 0.005\ \mathrm{mol}$.
- Electrons needed = $0.005 \times 2 = 0.01\ \mathrm{mol\ e^-}$.
- Charge $Q = 0.01 \times 96500 = 965\ \mathrm{C}$.

**Answer:** (c) **965 Coulomb**

**79.**

- Resistance ∝ 1/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. | $Wt./At.Wt. \neq x$ | Ratio |
|---|---|---|---|---|
| *Cu* | 4 *gm* | 63.5 | 4/63.5=.0625 | $\frac{.0625}{.0625} = 1$ |

| | | | | |
|---|---|---|---|---|
| *O* | 1 *gm* | 16 | 1/16 =.0625 | $\frac{.0625}{.0625}=1$ |

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 integeral no. = 32

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

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

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

$=0.27\ gm$ silver.

**82.** (b) Reaction for electrolysis of water is

$2H_2O \rightleftharpoons 4H^+ + 2O^{2-}$

$2O^{2-}\rightarrow O_2+4e^-$

$4e^-+4H^+\rightarrow 2H_2$

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

1 mole $=22.4\ dm^3$ oxygen

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

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

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

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

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

4 *F* charge will produce = 1 mole $O_2=32\ gm\ O_2$

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

**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.

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

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

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

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

**82.**

**Answer:** (a) 11.2 $dm^3$

**Explanation:**

- Electrolysis reaction at anode:

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

- 4 Faradays produce 1 mole of $O_2$.
- 1 mole $O_2$ at NTP = 22.4 L.
- 1 F produces: $22.4/2 = 11.2 \text{ dm}^3 \text{ O}_2$

**83.** **Answer:** (a) 0.5 F

**Explanation:**

- Molar mass of Na = 23 g/mol → 1 g eq. = 23 g (since Na is monovalent)
- Fraction of 1 F required: 11.5/23=0.5 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 → 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¥ ¥text{eq.} 10.79/108=0.1 eq.
2. Copper deposition: $Cu^{2+} + 2e^- \rightarrow Cu$
   - Eq. wt of Cu = 63.5 / 2 ≈ 31.75 g/equiv
   - Mass deposited: 0.1 × 31.75 ≈ 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*

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

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

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

$2e^-$ are required to deposite one *Ca* atom

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

**89.** (d) $C_6H_5NO_2 + 6H^+ + 6e^- \rightarrow C_6H_5NH_2 + 2H_2O$

1 mole = 123 *gm* nitrogen requires 6 mole electron *e*

$= 6 \times 96500\ coulomb$ *charge*

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

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

**92.** **Answer:** (b) 18 g

**Explanation:**

1. Moles of $Cl_2$ = 11.2 / 22.4 = 0.5 mol
2. Electrons: 2 $e^-$ per mole $Cl_2$ → 0.5 × 2 = 1 mol $e^-$
3. $Al^{3+}$ + 3 $e^-$ → Al → Mass Al deposited = 27 × (1 / 3) = 9 g? Wait carefully:

Step:

- $Cl_2$: 0.5 mol → 1 mol $e^-$ required?

Actually: 2 $Cl^-$ → $Cl_2$ + 2 $e^-$ → 0.5 mol $Cl_2$ requires 1 mol $e^-$

- $Al^{3+}$ + 3 $e^-$ → Al → Mass = 27 × (1 mol $e^-$ / 3) = 9 g

So correct **mass of Al = 9 g**

**93. Answer:** (d) 11.2 $cm^3$

**Explanation:**

1. Silver deposition: $Ag^+$ + $e^-$ → Ag, mass = 0.108 g
2. Equivalent weight of Ag = 108 g/equiv → moles of electrons = 0.108 / 108 = 0.001 mol $e^-$
3. Water electrolysis at anode: 2 $H_2O$ → $O_2$ + 4 $H^+$ + 4 $e^-$
4. Moles $O_2$ = 0.001 / 4 = 0.00025 mol → Volume = 0.00025 × 22.4 L = 0.0056 L = 5.6 mL? Wait options in $cm^3$.

0.0056 L = 5.6 $cm^3$

**94. Answer:** (b) 5.6 L

**Explanation:**

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