

Faraday's law of electrolysis

1. (c) $Ag^+ + e^- \rightarrow Ag$; $E_{Ag} = \frac{\text{Atomic Mass}}{1} = 108$

$$\text{Number of faraday} = \frac{W_{Ag}}{E_{Ag}} = \frac{108}{108} = 1.$$

2. (a) $W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{108 \times 9.65}{96500} = 1.08 \times 10^{-2} \text{ gm} = 10.8 \text{ mg}$

3. (b) $Fe^{2+} + 2e^- \rightarrow Fe$; $E_{Fe} = \frac{56}{2} = 28$

$$W_{Fe} = E_{Fe} \times \text{Number of faraday} = 28 \times 3 = 84 \text{ gm}.$$

4. (c) $W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{107.87 \times 965}{96500} = 1.0787 \text{ gm}.$

5. (c) $Al^{3+} + 3e^- \rightarrow Al$

$$E_{Al} = \frac{27}{3} = 9$$

$$W_{Al} = E_{Al} \times \text{No. of faradays} = 9 \times 5 = 45 \text{ gm}.$$

6. (c) *Cu* voltameter or *Cu* or *Ag* coulometer are used to detect the amount deposited on an electrode during passage of know charge through solution.

7. **Answer:** (c) Gram/coulomb

Explanation:

Electrochemical equivalent (*Z*) is defined as the mass of a substance deposited or liberated per unit charge passed through the electrolyte. Its formula is:

$$Z = \frac{m}{Q} = \frac{m}{It}$$



where m = mass deposited, I = current, t = time, and $Q=It$ total charge. Hence, the unit of Z is **gram per coulomb (g/C)**.

8. (b)
$$\frac{\text{Weight of Cu}}{\text{Weight of H}_2} = \frac{\text{Eq. weight of Cu}}{\text{Eq. weight. of H}}$$

$$\frac{\text{Weight of Cu}}{0.50} = \frac{63.6 / 2}{1}$$

$$\text{Weight of Cu} = 15.9 \text{ gm.}$$

9. (c) $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$
2 Faradays will deposit = 1 g atom of Cu = 63.5 g .

10. **Answer:** (a) Anions move towards anode, cations towards cathode

Explanation:

During electrolysis:

- **Cations** (positive ions) move toward the **cathode** (negative electrode) to gain electrons (reduction).
- **Anions** (negative ions) move toward the **anode** (positive electrode) to lose electrons (oxidation).

11. **Answer:** (c) Coulomb mole⁻¹

Explanation:

One Faraday is the charge of one mole of electrons. Its value is approximately 96500 C/mol. Hence, the unit of Faraday is **Coulomb per mole (C/mol)**.



12. (a) At cathode; $Al^{3+} + 3e^- \rightarrow Al$

$$E_{Al} = \frac{27}{3} = 9$$

$$W_{Al} = E_{Al} \times \text{No. of faradays} = 9 \times 0.1 = 0.9 \text{ gm}.$$

13. **Answer:** (c) $m = Z \times Q = Z \times I \times t$

Explanation:

The **first law of electrolysis** (Faraday's First Law) states:

The mass (m) of a substance deposited or liberated at an electrode is directly proportional to the total charge (Q) passed through the electrolyte.

Mathematically:

$$m = Z \cdot Q = Z \cdot I \cdot t$$

Where:

- m = mass of substance deposited
- Z = electrochemical equivalent
- $Q = \text{total charge} = \text{current} \times \text{time}$

14. (b) $W = zit$; $W = \frac{32.69 \times 5 \times 60 \times 40}{96500} = 4.065 \text{ gm}.$

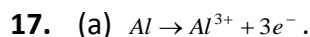
15. (b) $m = Z \times 4 \times 120$; $M = Z \times 6 \times 40$

$$\frac{M}{m} = \frac{6 \times 40}{4 \times 120} = \frac{1}{2}; \quad M = m / 2.$$

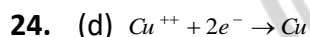
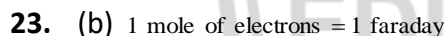
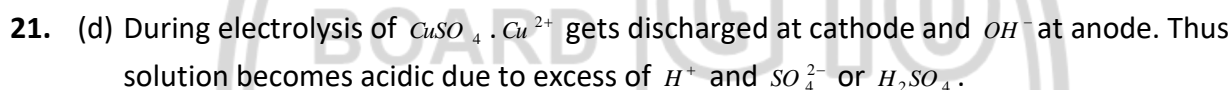
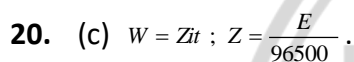
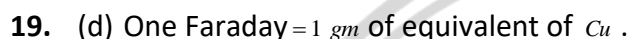
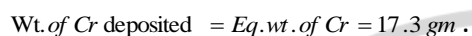
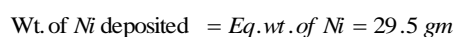
16. (c) $W_{\text{metal}} = \frac{E \times I \times t}{96500} = \frac{E \times 3 \times 50 \times 60}{96500}$



$$E = \frac{96500 \times w}{3 \times 50 \times 60} = \frac{96500 \times 1.8}{3 \times 50 \times 60} = 19.3 .$$



The charged obtained is $3 \times 96500 \text{ C}$.



$$E_{Cu} = \frac{63.54}{2} = 31.77$$

Amount of electricity required to deposit .6354 gm of Cu = $\frac{96500 \times 0.6354}{31.77} = 1930 \text{ Coulombs}$.

