Faraday's law of electrolysis

26. **Answer:** (c) A mixture of electrolytes is used

Explanation:

Faraday's laws assume that a **single electrolyte** is undergoing electrolysis. When a mixture of electrolytes is used, preferential deposition or discharge of ions can occur depending on factors like electrode potential, ion concentration, and mobility. This causes deviations from the predicted masses, and Faraday's laws no longer strictly hold.

27. (a) $W = ZQ$; $W = Zit$.

28. (d) $Ca^{++} + 2e^{-} \rightarrow Ca$

$E_{Ca} = \frac{40}{2} = 20$

$W_{Ca} = E_{Ca}$ × No. of faradays $= 20 \times 0.04 = 0.8\, gm$.

29. (c) $E_{\text{metal}} = \frac{\text{Weight of metal} \times 96500}{\text{Number of coulombs}}$

$= \frac{22.2 \times 96500}{2 \times 5 \times 60 \times 60} = 59.5$

Oxidation number of the metal $= \frac{177}{59.5} = +3$

30. (a) Quantity of electricity passed = $\frac{25}{1000} \times 60 = 1.5$

$2F = 2 \times 96500\ C$ deposit $Ca = 1 mole$

$\therefore$ 1.5 C will deposit $Ca = \frac{1}{2 \times 96500} \times 1.5\, mole$

$= \frac{1}{2 \times 96500} \times 1.5 \times 6.023 \times 10^{23}$ *atom* $= 4.68 \times 10^{18}$.

31. (b) Equivalent of *Cl* deposited

= No. of *Faraday* passed = 0.5

Wt. of $Cl = 0.5 \times Eq.wt. = 0.5 \times 35.5 = 17.75\ gm$.

32. (b) At Andoe $At\ Anode\ \ Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$

$$E_{Cl_2} = \frac{35.5 \times 2}{2} = 35.5$$

$$W_{Cl_2} = \frac{E_{Cl_2} \times I \times t}{96500} = \frac{35.5 \times 2 \times 30 \times 60}{96500} = 1.32\ gm\ .$$

33. **Answer:** (b) 9.0 g

Explanation:

Using **Faraday's laws of electrolysis**:

$$m = \frac{M \cdot Q}{nF}$$

Where:

- mmm = mass of the element deposited
- MMM = molar mass of the element
- QQQ = total charge passed
- nnn = number of electrons involved per atom in the reaction
- FFF = Faraday constant

For K:

$K^+ + e^- \rightarrow K$, n=1

Using the ratio formula:

$$\frac{m_{Al}}{M_{Al}/n_{Al}} = \frac{m_K}{M_K/n_K}$$

$$m_{Al} = \frac{M_{Al}}{n_{Al}} \cdot \frac{m_K \cdot n_K}{M_K} = \frac{27}{3} \cdot \frac{19.5 \cdot 1}{39} = 9\,\text{g}$$

34. (b) It is Faraday's law.

35. (b) Equivalent wt. of O_2 = Equivalent wt. of *Cu*

36. (d) $O_2\% = 20\%$

Metal % = 80% $= \frac{80}{20} \times 8 = 32\ g$ of metal .

37. (b) $V = \frac{827 \times 10^3}{4 \times 96500} = 2.14\ V$.

38. (b) $Ag^+ \xrightarrow{+e^-} Ag$, 96500 *C* will liberate silver =108 *gm*. 9650*C* will liberate silver = 10.8 *gm.*

39. (a) One mole of monovalent metal ion means charge of N electron i.e. 96500 *C* or 1 *Faraday.*

40.

Answer: (c) 1.0 A

Step-by-Step Solution:

1. Volume of H_2 collected: 112 mL = 0.112 L
2. Moles of H_2 at N.T.P.:

$$\text{Moles of } H_2 = \frac{\text{Volume}}{\text{Molar volume}} = \frac{0.112}{22.4} = 0.005\,\text{mol}$$

3. Electrons required per mole of H_2: 2 electrons (because $2\,H^+ + 2\,e^- \rightarrow H_2$)
4. Total moles of electrons needed:
   $0.005 \times 2 = 0.01$ mol e^-
5. Charge passed (Q):
   Q = 0.01 × 96500 C = 965 C
6. Current (I):
   I = Q / t = 965 / 965 = 1 A

41.

Answer: (a) 0.0093 mol

Explanation:

1. Time: $t = 30\,\text{min} = 1800\,\text{s}$
2. Current: $I = 1\,\text{A}$ → Total charge:

$$Q = I \cdot t = 1 \times 1800 = 1800\,\text{C}$$

3. Reaction at anode:

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

4. Moles of electrons:

$$n_{e^-} = \frac{Q}{F} = \frac{1800}{96500} \approx 0.01865\,\text{mol}$$

5. Moles of Cl_2 liberated: 2 electrons produce 1 mole of Cl_2 →

$$n_{Cl_2} = \frac{0.01865}{2} \approx 0.00933\,\text{mol}$$

So, 0.0093 mol of Cl_2 is liberated.

42. (a) 1 Faraday involves charge of 1 *mole* electrons.

43. (a) *Coulomb* = *ampere* (A) × *second* (S).

44. (b) $E = -\dfrac{13.6}{n^2}$ for He^+ $n = 1$

$$E = -\frac{13.6}{1^2} = -13.6\,eV.$$

45. (c) $w \propto E$ if *i* and *t* are constant.

46. **Answer:** (d) Charge on one mole of electrons

Explanation:

- The charge required to deposit or liberate **1 gram equivalent** of a substance is **1 Faraday**, which is equal to the **charge on one mole of electrons**.
- Numerically, 1 Faraday = **96500 C**.

- This is the fundamental basis of **Faraday's laws of electrolysis**.

47. (d) Charge (Coulombs) pass per second $= 10^{-6}$

number of electrons passed per second

$$= \frac{10^{-6}}{1.602 \times 10^{-19}} = 6.24 \times 10^{12}.$$

48. (d) At cathode;

$$Fe^{2+} + 2e^- \rightarrow Fe \;;\; Fe^{3+} + 3e^- \rightarrow Fe$$

$$(E_{Fe})_1 = \frac{\text{Atomic.weight}}{2}; (E_{Fe})_2 = \frac{\text{Atomic.weight}}{3}$$

Ratio of weight of *Fe* liberated

$$= \frac{\text{Atomic weight}}{3} : \frac{\text{Atomic weight}}{2} = 3:2.$$

49. (b) 31.75 *g* copper gets deposited at cathode on passing 96500 *coulomb* charge. We know that 31.75 *gm* of *Cu* is equal to 0.5 mole of *Cu* deposited at cathode on passing 1*F* of current.

50.

Answer: (b) 11.2 litres

Step-by-Step Solution:

1. Reaction at cathode:

$$Al^{3+} + 3e^- \rightarrow Al$$

2. Reaction at anode:

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

3. Moles of Al deposited:

$$\text{Moles of Al} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{0.9}{27} = 0.03333\,\text{mol}$$

4. Electrons required per mole of Al: 3

$$\text{Total moles of electrons} = 0.03333 \times 3 = 0.1\,\text{mol e}^-$$

5. Moles of Cl_2 liberated: 2 e^- produce 1 mole of Cl_2 →

$$n_{Cl_2} = \frac{0.1}{2} = 0.05\,\text{mol}$$

6. Volume of Cl_2 at N.T.P.: 1 mole = 22.4 L

$$V = n \times 22.4 = 0.5 \times 22.4 = 11.2\,\text{L}$$