Faraday's law of electrolysis

51. **Answer:** (c) Coulomb per equivalent

Explanation:

- Faraday is defined as the **charge required to deposit one gram equivalent of a substance**.
- Its value is approximately **96500 C per equivalent**.
- Hence, the **dimension of Faraday** is **Coulomb per equivalent**.

52. (b) For deposition of one equivalent silver required charged is 96500 *C*.

53. (b) $Cu^{++} + 2e^- \rightarrow Cu$; $E_{Cu} = \frac{63.55}{2} = 31.75\ gm\ Cu$.

54. (a) $Q = 2.5 \times 386 = 96500\ C$

$2F(2 \times 96500\ C)$ deposited $Cu = 63.5\,g$

$\therefore$ Hence 965 *C* will deposited; $Cu = 0.3175\ gm$.

55. (c) $\frac{\text{Wt. of } Cu}{\text{Wt. of } Ag} = \frac{\text{Eq. wt. of } Cu}{\text{Eq. wt. of } Ag}$; $\frac{\text{Wt. of } Cu}{1.08} = \frac{63.5/2}{108}$

Wt. of $Cu = 0.3177\ gm$.

56. (c) 1 *g* atom of *Al* = 3 equivalent of $Al = 3$ faraday charge

3 *mole* electrons = 3 *N* electron.

57. (c) At cathode : $Al^{3+} + 3e^- \rightarrow Al$

$E_{Al} = \frac{\text{Atomic mass}}{3}$

At cathode : $Cu^{2+} + 2e^- \rightarrow Cu$

$E_{Cu} = \frac{\text{Atomic mass}}{2}$

At cathode : $Na^{+} + e^{-} \rightarrow Na$

$E_{Na} = \frac{\text{Atomic mass}}{1}$

For the passage of 3 faraday;

mole atoms of Al deposited = 1

mole atoms of *Cu* deposited $= \frac{1 \times 3}{2} = 1.5$

mole atoms of *Na* deposited $= 1 \times 3 = 3$.

58. (d) At cathode: $Ag^{+} + e^{-} \rightarrow Ag$

At Anode: $2OH^{-} \rightarrow H_2O + \frac{1}{2}O_2 + 2e^{-}$

$$E_{Ag} = \frac{108}{1} = 108; \; E_{O_2} = \frac{\frac{1}{2} \times 32}{2} = 8$$

$$\frac{W_{Ag}}{E_{Ag}} = \frac{W_{O_2}}{E_{O_2}}; \; W_{Ag} = \frac{1.6 \times 108}{8} = 21.6 \; gm .$$

59. (d) *KI* is an electrolyte.

60. (d) Number of *gm* equivalent = Number of faraday pass

4 *gm* = 4 faraday.

61. (c) Eq. of *Al* = $\frac{13.5}{27/3} = 1.5$.

Thus 1.5 Faraday is needed.

62. **Answer:** (c) 96500 C mol^{-1}

Explanation:

- One Faraday (F) is the total electric charge carried by **1 mole of electrons**.

- Its numerical value is approximately **96500 coulombs per mole** of electrons.
- This is used in **Faraday's laws of electrolysis** to relate charge passed to the mass of substance deposited or liberated.

63. (b) Electricity required

= No. of *gm* equivalent $\times$ 96500 *coulombs*

= 0.5 $\times$ 96500 = 48250 *C.*

64. (a) Equivalent weight of silver = 107.870 *g*.

1 Faraday = 96500 *coulomb.*

65. **Answer:** (b) Coulomb

Explanation:

- Electric charge Q is calculated as: $Q = I \cdot t$
- If I=1 A
- and t=1 st = 1¥,st=1s,then Q=1 · 1=1 C
- Hence, **1 coulomb** is the charge transported by a current of 1 ampere in 1 second.

66. **Answer:** (b) Equivalent weight

Explanation:

- According to **Faraday's first law of electrolysis**, the mass of substance deposited or liberated at an electrode is directly proportional to its **equivalent weight** and the total charge passed:

$$m = Z \cdot Q$$

Where Z = electrochemical equivalent

$= \frac{\text{Equivalent weight}}{\text{Faraday}}.$

67. (a) Equivalent weight and atomic weight of *Na* metal are the same, so 1*g* atom of *Na* is deposited by one Faraday of current.

68. (a) $Al \rightarrow Al^{3+} + 3e^-$.

69.

Answer: (b) 11.2 g

Step-by-Step Solution:

1. **Electrolysis reaction:**

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

2. **Equivalent weight of Fe:**

$$\text{Equivalent weight} = \frac{\text{Atomic weight}}{\text{Valency}} = \frac{56}{3} \approx 18.67\,\text{g/equiv}$$

3. **Mass deposited using Faraday's law:**

$$m = (\text{Equivalent weight}) \times (\text{No. of Faradays}) = 18.67 \times 0.6 \approx 11.2\,\text{g}$$

70. (c) $\because$ 1*F* obtained from 1 *g* equivalent

$\therefore$ 2.5 *F* obtained from 2.5 *g* equivalent.

75. (c) Faraday constant depends upon the current passed.