

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

54. (a)  $Q = 2.5 \times 386 = 96500 \text{ C}$   
 $2F(2 \times 96500 \text{ C})$  deposited  $Cu = 63.5 \text{ g}$   
 $\therefore$  Hence 965 C will deposited;  $Cu = 0.3175 \text{ gm}$ .

55. (c)  $\frac{\text{Wt. of Cu}}{\text{Wt. of Ag}} = \frac{\text{Eq. wt. of Cu}}{\text{Eq. wt. of Ag}}$ ;  $\frac{\text{Wt. of Cu}}{1.08} = \frac{63.5/2}{108}$   
 $\text{Wt. of Cu} = 0.3177 \text{ gm}$ .

56. (c) 1 g atom of  $Al = 3$  equivalent of  $Al = 3$  faraday charge  
 $3 \text{ mole electrons} = 3 N$  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:  $2OH^- \rightarrow H_2O + \frac{1}{2}O_2 + 2e^-$

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

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

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)  $96500 \text{ C mol}^{-1}$

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 $\cancel{\text{m}}$ , st=1s, then  $Q=1 \cdot 1=1$  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  $1g$  atom of  $Na$  is deposited by one Faraday of current.

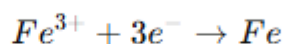
68. (a)  $Al \rightarrow Al^{3+} + 3e^-$ .

69.

Answer: (b) 11.2 g

Step-by-Step Solution:

1. Electrolysis reaction:



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)  $\therefore 1F$  obtained from  $1g$  equivalent  
 $\therefore 2.5 F$  obtained from  $2.5 g$  equivalent.

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

