## Numerical Based on Faraday's Law

**Q.1-**What mass of copper will be deposited at cathode if a current of 0.22 amp flows through the cell containing $CuSO_4$ for 1.5 hours ?

**Ans.-** Given that,

Strength of Current, I = 0.22 ampere

Time, t = 1.5 hrs

= (1.5 × 60 × 60 sec)

= 5400 sec

Mass of Copper deposited, m = ?

Here, the amount of charge passing through the cell,

Q = It

= 0.22 amp × 5400 sec

= 1200 C

$\because$ m = zIt

$= \frac{E}{96500} It$

$\therefore$ m = $\frac{31.75}{96500} \times 1200$ C

= 0.39 g

**Q.2-** A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5.0 A for 20 min. What mass of nickel will be deposited at cathode ? (Atomic mass of Ni= 58.7 g $mol^{-1}$)

**Ans.-**Given that,

Strength of Current, I = 5.0 A

Time, t = 20 min

= (20 × 60 sec)

= 1200 sec

Mass of Nickel deposited, m = ?

Here, the amount of charge passing through the cell,

Q = It

= 5.0 amp × 1200 sec

= $6.0 \times 10^3$ C

$\because$ m = zIt

$= \frac{E}{96500} It$

$\therefore$ m = $\frac{58.7}{2 \times 96500} \times 6.0 \times 10^3$ C

= 1.82 g

**Q.3-**How many moles of mercury will be produced by electrolysing 10 M $Hg(NO_3)_2$ solution with a current of 200 amp for 3 hours ? (Molar mass of $Hg(NO_3)_2$ = 200.6 $gmol^{-1}$)

**Ans.-**Given that,

Strength of Current, I = 200 ampere

Time, t = 3 hrs

= (3 × 60 × 60 sec)

= 10800 sec

Moles of Mercury deposited = ?

Here, the amount of charge passing through the cell,

$$Q = It$$
$$= 200 \text{ amp} \times 10800 \text{ sec}$$
$$= 2.16 \times 10^6 \text{ C}$$

$$\because \quad m = zIt$$
$$= \frac{E}{96500} It$$

$$\therefore \quad m = \frac{200.6}{2 \times 96500} \times 2.16 \times 10^6 \text{ C}$$
$$= 22.45 \text{ g}$$

$\therefore$ No. Of moles of $Hg(NO_3)_2$,

$$n = \frac{Mass\ of\ Hg(NO_3)_2\ (g)}{GMM\ of\ Hg(NO_3)_2}$$
$$= \frac{22.45}{200.6}$$
$$= 0.112 \text{ moles}$$

**Q.4.-** Chromium metal can be plated out from acidic solution containing $CrO_3$ as per the following equaqion-

$$CrO_3 + 6H^+ + 6e^- \rightarrow Cr\ (s) + 3H_2O$$

Calculate the mass of Chromium that will be plated out by 12000 C of charge.

**Ans.-** From given equation, it is clear that, 1 mole of chromium is deposited by 6 moles of electrons. Thus,

1 mole of Cr ≡ 6 moles of electrons

$\therefore$ 52 g of Cr ≡ 6 Faraday of charge

or 52 g of Cr ≡ 6 × 96500 C of charge

$\therefore$ x g of Cr (say) ≡ 12000 C of charge

$$\frac{52\ g}{x\ g} = \frac{6 \times 96500\ C}{12000\ C}$$

$$\therefore \quad x = \frac{52 \times 12000}{6 \times 96500}$$
$$= 1.077 \text{ g}$$

**Q.5.–** A current of 2 ampere is passed through copper (Equivalent mass = 31.5 g $mol^{-1}$) and silver voltameters connected in series. If 1.08 g of silver (Equivalent mass = 108 g $mol^{-1}$) is deposited on cathode in silver voltameter then find the amount of copper deposited at cathode in copper voltameter.

**Ans.-**Given that,

Mass of Ag deposited, $m_{Ag}$ = 1.08 g

Mass of Cu deposited, $m_{Cu}$ = ?

Equivalent mass of Ag, $E_{Ag}$ = 108 $gmol^{-1}$

Equivalent mass of Cu, $E_{Cu}$ = 31.5 $gmol^{-1}$

From Faraday's second law of electrolysis-

$$\frac{Mass\ of\ Copper}{Mass\ of\ Silver} = \frac{E_{Cu}}{E_{Ag}}$$

$$\therefore \quad \frac{m_{Cu}}{1.08\ g} = \frac{31.5\ g\ mol^{-1}}{108\ g\ mol^{-1}}$$

or $$m_{Cu} = \frac{1.08 \times 31.5}{108}$$

$$= 0.315\ g$$