FARADAY'S FIRST LAW

States that the mass of a substance formed during electrolysis is proportional to the quantity of electricity supplied, Q.

$$Q = It$$

where I = *current (A)*
t = *time (s)*

Charge of 1 electron, e^- = 1.6×10^{-19} C

Charge of 1 mol electrons

$= 6.02 \times 10^{23} \times 1.6 \times 10^{-19}$ C

$\approx$ **96500 C**

$=$ **1 F**

Faraday constant

F = **96500 C / mol**

- **To obtain 1 mol of product at electrode, quantity of electrical charge required :**

Example :

$$Ag^{+}(aq) \quad + \quad e^{-} \quad \rightarrow \quad Ag(s)$$

1 mol e^- — 1 mol Ag

1 F

96500 C

$$Cu^{2+}(aq) \quad + \quad 2e^{-} \quad \rightarrow \quad Cu(s)$$

2 mol e^- — 1 mol Cu

2 F

2×96500 C

Example 1:

A current of 0.50 A was applied to an electrolytic cell containing a $Cu(NO_3)_2$ solution for 5 hours at 25 °C and 1 atm.

(a) Calculate the mass of Cu deposited. (A_r Cu = 63.5)

(b) Calculate the volume of gas collected at anode.

Solution :

cathode (red) : $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

2F   1 mol

$Q = It$

$= 0.5\ A \times (5 \times 60 \times 60)\ s$

$= 9000\ C$

$= 9000\ C \times \dfrac{1\ F}{96500\ C}$

$= 0.09326\ F$

$2\ F \equiv 1\ mol\ Cu$

$0.09326\ F \equiv \dfrac{2}{0.09326} \times 1\ mol$

$= 0.04663\ mol\ Cu$

$\therefore$ Mass of Cu deposited

$= 0.04663\ mol \times 63.5\ g\ mol^{-1}$

$= 2.96\ g$

(b) Calculate the volume of gas collected at anode.

Anode (ox) : $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$

1 mol 4 F

$$4\text{ F} \equiv 1\text{ mol } O_2$$

$$0.09326\text{ F} \equiv \frac{0.09326}{4} \times 1\text{ mol}$$

$$= 0.02332\text{ mol } O_2$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$= \frac{0.02332\text{ mol} \times 0.08206\text{ L atm mol}^{-1}\text{ K}^{-1} \times 298\text{ K}}{1\text{ atm}}$$

$$= 0.5703\text{ L}$$

Example 2 :

A $CrCl_3$ solution is electrolyzed using inert electrodes. How many minutes needed to plate out 25 g Cr from the solution using a current of 2.75 A ? (A$_r$ Cr = 52)

Solution :

$$\text{Mol of Cr} = \frac{25}{52} = 0.4808 \text{ mol}$$

cathode (red) : $Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$

3 F 1 mol

$$1 \text{ mol Cr} \equiv 3 \text{ F}$$

$$0.4808 \text{ mol Cr} \equiv 0.4808 \times 3 \text{ F}$$

$$= 1.442 \text{ F}$$

$$Q = It$$

$$1.442 \times 96500 \text{ C} = 2.75 \text{ A} \times t$$

$$t = 50612 \text{ s}$$

$$= 843.5 \text{ min}$$

Example 3:

When an aqueous solution containing the gold ion, Au^{n+} is electrolysed by a current of 2.0 A for 4.0 hour, 19.7 g of of gold was deposited on the cathode. Calculate the charge on the gold ion.

Solution :

$$Au^{n+}(aq) + ne^- \rightarrow Au(s)$$

$$\text{Mol of Au formed} = \frac{19.7}{197}$$
$$= 0.10 \text{ mol}$$

$$Q = It$$
$$= 2.0 \text{ A} \times (4.0 \times 60 \times 60) \text{ s}$$
$$= 2.88 \times 10^4 \text{ C} = \frac{2.88 \times 10^4}{96500}$$
$$= 0.2984 \text{ F}$$

$$0.10 \text{ mol Au} \equiv 0.2984 \text{ F}$$

$$1 \text{ mol Au} \equiv \frac{1}{0.10} \times 0.2984$$
$$= 2.984$$
$$\approx 3 \text{ F}$$

n = 3

$\therefore$ *Charge of gold ion = +3*