

# ***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,  $e^- = 1.6 \times 10^{-19} \text{ C}$**

**Charge of 1 mol electrons**

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

$$**\approx 96500 \text{ C}**$$

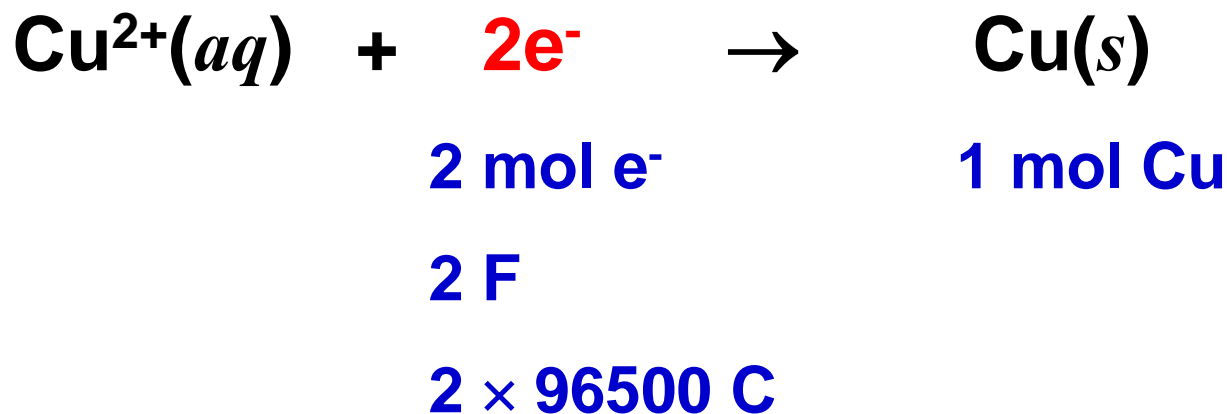
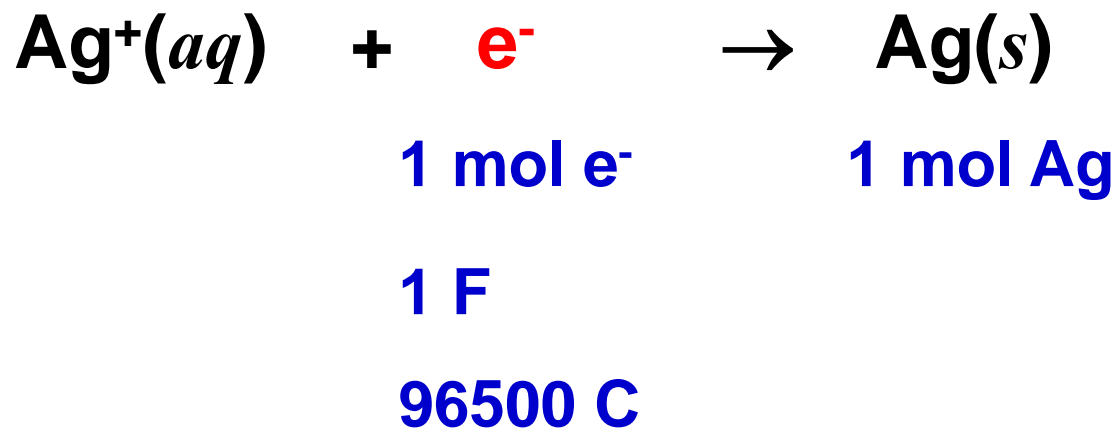
$$**= 1 \text{ F}**$$

***Faraday constant***

$$**F = 96500 \text{ C / mol}**$$

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

**Example :**



## Example 1:

A current of 0.50 A was applied to an electrolytic cell containing a  $\text{Cu}(\text{NO}_3)_2$  solution for 5 hours at 25 °C and 1 atm.

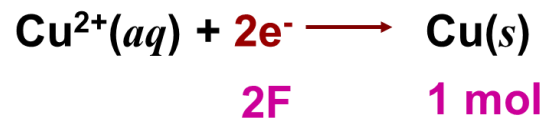
(a) Calculate the mass of Cu deposited.

( $A_r \text{ Cu} = 63.5$ )

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

**Solution :**

cathode  
(red) :

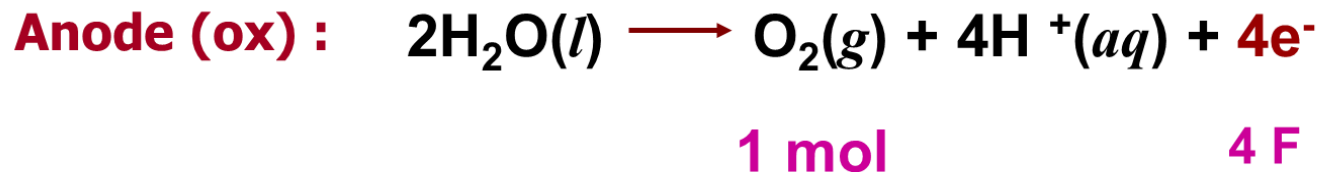


$$\begin{aligned} Q &= It \\ &= 0.5 \text{ A} \times (5 \times 60 \times 60) \text{ s} \\ &= 9000 \text{ C} \\ &= 9000 \text{ C} \times \frac{1 \text{ F}}{96500 \text{ C}} \\ &= 0.09326 \text{ F} \end{aligned}$$

$$\begin{aligned} 2 \text{ F} &\equiv 1 \text{ mol Cu} \\ 0.09326 \text{ F} &\equiv \frac{2}{0.09326} \times 1 \text{ mol} \\ &= 0.04663 \text{ mol Cu} \end{aligned}$$

$$\begin{aligned} \therefore \text{Mass of Cu deposited} &= 0.04663 \text{ mol} \times 63.5 \text{ g mol}^{-1} \\ &= 2.96 \text{ g} \end{aligned}$$

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



$$\begin{aligned} 4 \text{ F} &\equiv 1 \text{ mol O}_2 \\ 0.09326 \text{ F} &\equiv \frac{0.09326}{4} \times 1 \text{ mol} \\ &= 0.02332 \text{ mol O}_2 \end{aligned}$$

$$PV = nRT$$

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

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

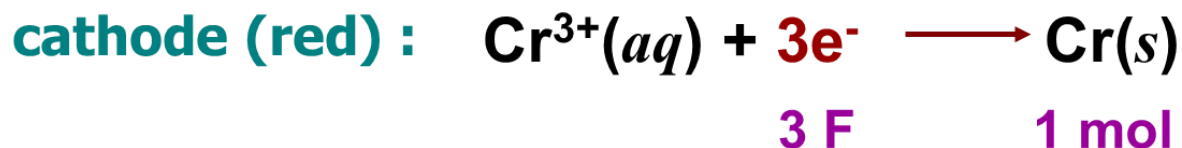
$$= 0.5703 \text{ L} //$$

## Example 2 :

A  $\text{CrCl}_3$  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 ? ( $A_r \text{ Cr} = 52$ )

### Solution :

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



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

$$Q = It$$

$$\begin{aligned} 0.4808 \text{ mol Cr} &\equiv 0.4808 \times 3 \text{ F} \\ &= 1.442 \text{ F} \end{aligned}$$

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

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

$$= 843.5 \text{ min} //$$

### Example 3:

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

### Solution :

$$\begin{aligned}\text{Mol of Au formed} &= \frac{19.7}{197} \\ &= 0.10 \text{ mol}\end{aligned}$$

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



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

$$1 \text{ mol Au} \equiv \frac{1}{0.10} \times 0.2984$$

$$= 2.984$$

$$\approx 3 \text{ F}$$

$$n = 3$$

$\therefore$  Charge of gold ion = +3