

1. A current of 9.65 ampere is passed through the aqueous solution  $NaCl$  using suitable electrodes for 1000 s . The amount of  $NaOH$  formed during electrolysis is :

- A. 2.0 g
- B. 4.0 g
- C. 6.0 g
- D. 8.0 g

Given :

Current,  $I = 9.65 \text{ A}$

Time,  $t = 1000 \text{ s}$

By Faraday's First law,

The mass deposited/released of any substance during electrolysis is proportional to the amount of charge passed into the electrolyte.

$$W \propto Q$$

$$W = ZQ$$

where,

W: Mass deposited or liberated

Q: Amount of charge passed

Z: Electrochemical equivalent of the substance

Since,

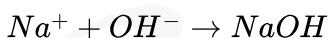
$$Z = \frac{E}{96500}$$

$$W = \frac{E}{96500} \times Q$$

where,

E is Equivalent mass

$$E = \frac{\text{Molecular Mass}}{n - \text{factor}}$$



Molecular mass of  $NaOH$  is  $40 \text{ g mol}^{-1}$

n-factor of the reaction is 1

$$E = \frac{40}{1}$$

Charge,  $Q = I \times t$

$$Q = 9.65 \times 1000$$

Hence,

$$W = Z \times i \times t$$

$$W = \frac{E}{96500} \times 9.65 \times 1000$$

$$W = \frac{\left(\frac{40}{1}\right)}{96500} \times 9.65 \times 1000$$

$$W = 4 \text{ g}$$

2. The time taken by the galvanic cell which operates almost ideally under reversible conditions at a current of  $10^{-16} A$  to deliver 1 mole of electron is :

- A.  $19.3 \times 10^{-17} s$
- B.  $4.82 \times 10^{20} s$
- C.  $9.65 \times 10^{20} s$
- D.  $3.4 \times 10^{11} s$

Given :

Current =  $10^{-16} A$

Charge =  $96500 C$

(Charge on 1 mole of electrons)

Using the formula,

Charge (Q) = Current (A)  $\times$  Time (T)

Therefore,

$$T = \frac{96500}{10^{-16}}$$

$$T = 9.65 \times 10^{20} s$$

3. In an electrolysis experiment, a current was passed for 5 hours through two cells connected in series. The first cell contains a solution of gold salt and the second cell contains copper sulphate solution. 9.85 g of gold was deposited in the first cell. If the oxidation number of gold is +3, find the amount of copper deposited on the cathode in the second cell and the magnitude of the current in ampere respectively.

Given:

Molar mass of gold = 197 g/mol

Molar mass of copper = 63.5 g/mol

- A. 4.67 g and 0.804 A
- B. 2.65 g and 2 A
- C. 70.5 g and 0.35 A
- D. 70.5 g and 0.26 A

Faradays second law of electrolysis:

If equal amount of charge (Q) is passed through two different solutions, the amount of substance deposited/liberated (W) is proportional to their chemical equivalent weights (E).

Thus,

$$\frac{W_1}{W_2} = \frac{E_1}{E_2} = \frac{Z_1}{Z_2}$$

$$\therefore \frac{\text{Mass of Au deposited}}{\text{Mass of Cu deposited}} = \frac{\text{Equivalent mass of Au}}{\text{Equivalent mass of Cu}}$$

$$\text{Equivalent mass of Au} = \frac{197}{3};$$

$$\text{Equivalent mass of Cu} = \frac{63.5}{2}$$

$$\text{Mass of copper deposited} = 9.85 \times \frac{63.5}{2} \times \frac{3}{197} = 4.76 \text{ g}$$

Let 'Z' be the electrochemical equivalent of Cu.

$$E = Z \times 96500$$

$$Z = \frac{E}{96500} = \frac{63.5}{2 \times 96500}$$

By Faradays first law,

$$W = Z \times I \times t$$

$$t = 5 \text{ hour} = 5 \times 3600 \text{ s}$$

$$4.76 = \frac{63.5}{2 \times 96500} \times I \times 5 \times 3600$$

$$I = \frac{4.76 \times 2 \times 96500}{63.5 \times 5 \times 3600}$$

$$I = 0.804 \text{ A}$$

4. On electrolysis of dil.  $H_2SO_4$  with platinum electrodes, the amount of substance liberated at the cathode and anode are in the ratio:

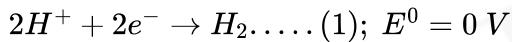
- A. 1 : 8
- B. 8 : 1
- C. 16 : 1
- D. 1 : 16

Electrolysis of dilute sulphuric acid:

In dilute sulphuric acid, we have  $H^+$ ,  $OH^-$  and  $SO_4^{2-}$  ions.

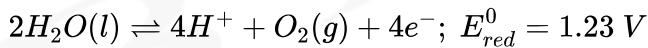
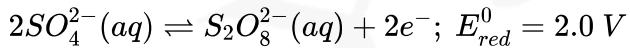
$H^+$  will migrate to the cathode and  $OH^-$ ,  $SO_4^{2-}$  migrates to the anode.

At cathode:



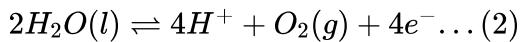
At Anode:

Possible reactions:



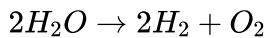
Hence,  $H_2O$  has lower reduction, it will undergo oxidation at anode and liberates  $O_2$ .

Actual reaction at anode:



Multiplying equation (1) by 2 and adding to equation (2) gives overall reaction.

Overall reaction:



2 moles of  $H_2O$  electrolyse to give 2 moles of  $H_2$  and 1 mole of  $O_2$

∴

No. of moles of  $H_2$  liberated at cathode = 2

Amount of  $H_2$  liberated at cathode = 4 g

No. of moles of  $O_2$  liberated at anode = 1

Amount of  $O_2$  liberated at anode = 32 g

∴

Ratio of the amount of substance liberated at the cathode and anode = 4 : 32  $\Rightarrow$  1 : 8

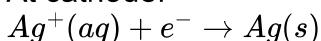
5. How many grams of silver would be deposited at cathode if  $0.05 \text{ dm}^3$  of oxygen measured at NTP is liberated at the anode when silver nitrate solution is electrolysed between platinum electrodes?

(Given: Atomic mass of  $\text{Ag} = 108$ ,  $O = 16$ )

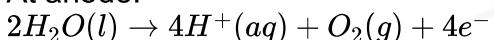
- A.  $1.22 \text{ g}$
- B.  $0.44 \text{ g}$
- C.  $0.95 \text{ g}$
- D.  $0.33 \text{ g}$

Electrolysis of  $\text{AgNO}_3$

At cathode:



At anode:



For 1 mole of  $\text{O}_2$ , 4 Faraday charge is released

At NTP,

1 mol of  $\text{O}_2 = 22.4 \text{ L}$

$$1 \text{ L} = \frac{1}{22.4} \text{ mole of } \text{O}_2$$

$$0.05 \text{ L} = \frac{0.05}{22.4} = 0.0022 \text{ mole of } \text{O}_2$$

For 1 mole of  $\text{O}_2$ , 4 Faraday charge is released

For 0.0022 mole,

$$= 4 \times 0.0022 \text{ F}$$

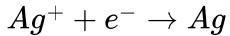
$$= 0.0088 \text{ F}$$

$$= 0.0088 \times 96500$$

$$\approx 850 \text{ C}$$

Charge released,  $Q = 850 \text{ C}$

For cathode reaction,



On passing 1 F or 96500 C of charge, one mole of  $\text{Ag}$  or 108 g of  $\text{Ag}$  is produced.

∴

850 C of charge gives,

$$= \frac{850 \times 108}{96500}$$

$$= 0.95 \text{ g}$$

Thus, 0.95 g of silver is deposited at cathode.