

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,

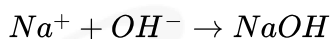
$$Z = \frac{E}{96500}$$

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

where,

E is Equivalent mass

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



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

n-factor 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. $19.3 \times 10^{-17} \text{ s}$

☐ B. $4.82 \times 10^{20} \text{ s}$

☒ C. $9.65 \times 10^{20} \text{ s}$

☐ D. $3.4 \times 10^{11} \text{ s}$

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,

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

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

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,

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

∴

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

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

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

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

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:

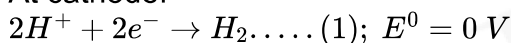
- ☒ A. 1 : 8
☐ B. 8 : 1
☐ C. 16 : 1
☐ D. 1 : 16

Electrolysis of dilute sulphuric acid:

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

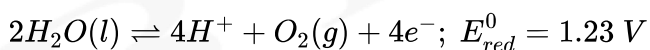
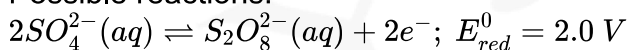
H^+ will migrate to the cathode and OH^- , SO_4^{2-} migrates to the anode.

At cathode:



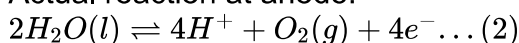
At Anode:

Possible reactions:



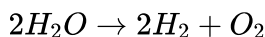
Hence, H_2O has lower reduction, it will undergo oxidation at anode and liberates O_2 .

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

\therefore
 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

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

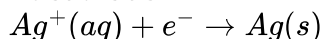
5. How many grams of silver would be deposited at cathode if 0.05 dm^3 of oxygen measured at NTP is liberated at the anode when silver nitrate solution is electrolysed between platinum electrodes?

(Given: Atomic mass of $Ag = 108$, $O = 16$)

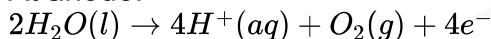
- ☒ A. 1.22 g
- ☒ B. 0.44 g
- ☒ C. 0.95 g
- ☒ D. 0.33 g

Electrolysis of $AgNO_3$

At cathode:



At anode:



For 1 mole of O_2 , 4 Faraday charge is released

At NTP,

$$1 \text{ mol of } O_2 = 22.4 \text{ L}$$

$$1 \text{ L} = \frac{1}{22.4} \text{ mole of } O_2$$

$$0.05 \text{ L} = \frac{0.05}{22.4} = 0.0022 \text{ mole of } O_2$$

For 1 mole of O_2 , 4 Faraday charge is released

For 0.0022 mole,

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

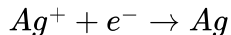
$$= 0.0088 \text{ F}$$

$$= 0.0088 \times 96500$$

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

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

For cathode reaction,



On passing 1 F or 96500 C of charge, one mole of Ag or 108 g of Ag is produced.

\therefore

850 C of charge gives,

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

$$= 0.95 \text{ g}$$

Thus, 0.95 g of silver is deposited at cathode.