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 :
$\text{Current, I } = 9.65\ A$
$\text{Time, t } = 1000\ s$

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

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

$Na^{+} + OH^{-} \rightarrow NaOH$
Molecular mass of $NaOH$ is $40\ g\,mol^{-1}$
n-factor the reaction is 1

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

$\text{Charge, Q } = I \times t$
$Q = 9.65 \times 1000$
Hence,
$W = Z \times i \times t$
$W = \frac{E}{96500} \times 9.65 \times 1000$

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

$W = 4\ g$

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 :
$\text{Current } = 10^{-16}\ A$
$\text{Charge } = 96500\ C$
(Charge on 1 mole of electrons)

Using the formula,
$\text{Charge (Q) } = \text{Current (A)} \times \text{Time (T)}$

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:

$$\text{Molar mass of gold } = 197\ g/mol$$
$$\text{Molar mass of copper } = 63.5\ g/mol$$

- **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,

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

$\therefore$

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

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

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

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

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,

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

$$t=5\ hour=5\times 3600\ s$$

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

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

$$I=0.804\ A$$

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:
$2H^+ + 2e^- \rightarrow H_2.....(1);\ E^0 = 0\ V$

At Anode:
Possible reactions:
$2SO_4^{2-}(aq) \rightleftharpoons S_2O_8^{2-}(aq) + 2e^-;\ E^0_{red} = 2.0\ V$

$2H_2O(l) \rightleftharpoons 4H^+ + O_2(g) + 4e^-;\ E^0_{red} = 1.23\ V$

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

Actual reaction at anode:
$2H_2O(l) \rightleftharpoons 4H^+ + O_2(g) + 4e^-...(2)$

Multiplying equation (1) by 2 and adding to equation (2) gives overall reaction.
Overall reaction:
$2H_2O \rightarrow 2H_2 + O_2$

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:
$Ag^+(aq) + e^- \rightarrow Ag(s)$
At anode:
$2H_2O(l) \rightarrow 4H^+(aq) + O_2(g) + 4e^-$
For 1 mole of O_2, 4 Faraday charge is released
At NTP,
$1\ mol$ of $O_2 = 22.4\ L$

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

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

For 1 mole of O_2, 4 Faraday charge is released
For 0.0022 mole,
$= 4 \times 0.0022\ F$
$= 0.0088\ F$
$= 0.0088 \times 96500$
$\approx 850\ C$
Charge released, Q $= 850\ C$
For cathode reaction,
$Ag^+ + e^- \rightarrow Ag$
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,

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

$= 0.95\ g$

Thus, $0.95\ g$ of silver is deposited at cathode.