Faraday's laws of electrolysis

The first law of Faraday states that the mass of a substance deposited or liberated at a particular electrode is proportional to the amount of charge passing through the electrodes. W or the mass of the substance directly proportional to Q where Q is( It), if( I) is the current passing is one ampere and t the time taken as one second then( W) will be equal to Z. (Z) is electrochemical equivalent. W=ZQ or W=Zit

W=Z when I is one ampere and t is one second.

Faraday's second law it states that the amount of different substances liberated by the same quantity of electricity passing through the electrolytic solution are proportional to their chemical Equivalent weight.( atomic mass of the metal ÷ number of electron required to reduce the cation).

Equivalent charge is equal to 1 Faraday equal to 96500 coulomb.

W = Equivalent weight ÷1 Faraday (96500 )X( it)

Charge of one electron is = 1.6 into 10 ^ - 19 .

One mole = 6.022 into 10 ^ 23 electrons therefore 1 mole of electron charge will carry 1.6 ×10 ^ - 19 ×6.022 × 10 ^ 23 = 96500 C(1F).

There were no constant current sources available during Faraday's times. The general practice was to put a coulometer (a standard electrolytic cell) for determining the quantity of electricity passed from the amount of metal (generally silver or copper) deposited or consumed. However, coulometers are now obsolete and we now have constant current ($I$) sources available and the quantity of electricity $Q$, passed is given by

$$Q = It$$

$Q$ is in coloumbs when $I$ is in ampere and $t$ is in second.

The amount of electricity (or charge) required for oxidation or reduction depends on the stoichiometry of the electrode reaction. For example, in the reaction:

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s) \qquad (3.30)$$

One mole of the electron is required for the reduction of one mole of silver ions.

We know that charge on one electron is equal to $1.6021\times 10^{-19}$C.

Therefore, the charge on one mole of electrons is equal to:

$$N_A \times 1.6021 \times 10^{-19}\ C = 6.02 \times 10^{23}\,mol^{-1} \times 1.6021 \times 10^{-19}$$
$$C = 96487\ C\ mol^{-1}$$

This quantity of electricity is called **Faraday** and is represented by the symbol **F**.

For approximate calculations we use 1F Y 96500 C $mol^{-1}$.

For the electrode reactions:

$$Mg^{2+}(l) + 2e^{-} \longrightarrow Mg(s) \qquad (3.31)$$

$$Al^{3+}(l) + 3e^{-} \longrightarrow Al(s) \qquad (3.32)$$

It is obvious that one mole of $Mg^{2+}$ and $Al^{3+}$ require 2 mol of electrons (2F) and 3 mol of electrons (3F) respectively. The charge passed through the electrolytic cell during electrolysis is equal to the product of current in amperes and time in seconds. In commercial production of metals, current as high as 50,000 amperes are used that amounts to about 0.518 F per second.

3.10 A solution of $CuSO_4$ is electrolysed for 10 minutes with a current of 1.5 amperes. What is the mass of copper deposited at the cathode?

ion $t$ = 600 s charge = current × time = 1.5 A × 600 s = 900 C

According to the reaction:

$Cu^{2+}(aq) + 2e^{-} = Cu(s)$

We require 2F or 2 × 96487 C to deposit 1 mol or 63 g of Cu.

For 900 C, the mass of Cu deposited

$= (63\ g\ mol^{-1} \times 900\ C)/(2 \times 96487\ C\ mol^{-1}) = 0.2938\ g.$

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Products of electrolysis depend on the nature of material being electrolysed and the type of electrodes being used. If the electrode is inert (e.g., platinum or gold), it does not participate in the chemical reaction and acts only as source or sink for electrons. On the other hand, if the electrode is reactive, it participates in the electrode reaction. Thus, the products of electrolysis may be different for reactive and inert

**Batteries it is the source of energy in which chemical energy converted into electrical energy with the help of redox reaction. A battery is made up of number of cells compact together to give the desired voltage ,there are two types of batteries primary battery and secondary battery.**

Primary battery which is used only once and after use over a period of time the battery becomes dead and cannot be reused again it is also known as **Leclanche cell**.

The zinc container act as the anode and the cathode is made of carbon (graphite)rod surrounded by powdered manganese dioxide and carbon

The space between the electrode is filled by a paste of Ammonium Chloride and zinc chloride, the electrode reactions are complex .