

Chemical Effects of Electric Current

Introduction

The reader is well acquainted with the passage of electric current through metallic conductors, *e.g.*, copper, aluminium etc. In such conductors, current conduction is due to the movement of free electrons and there is no chemical or physical change except the rise in temperature. However, conduction of current through *certain salt solutions is quite different. Such liquids provide a large number of oppositely charged atoms (called ions) and are known as *electrolytes* *e.g.*, acids (H_2SO_4 , HCl etc.), solutions of inorganic compounds (NaCl , CuSO_4 , AgNO_3 etc.), hydroxides of metals (KOH , NaOH etc). In an electrolyte, conduction is due to the movement of ions (an electrolyte has no free electrons) and chemical changes occur so long as the conduction takes place. Thus passage of electric current through an electrolyte causes chemical changes *i.e.* electrical energy is converted into chemical energy. The converse of this is also true *i.e.*, we can produce electrical energy from chemical energy. In this chapter, we shall study about the close relationship between electrical energy and chemical energy.

10.1. Electric Behaviour of Liquids

Some liquids conduct current while others do not permit the passage of current through them. On the basis of electrical conductivity, the liquids may be divided into three classes *viz.*

(i) Those liquids which do not conduct current are called **insulators** *e.g.*, mineral oils, distilled water etc.

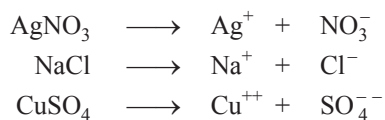
(ii) Those liquids which conduct current due to drifting of free electrons are called **conductors** *e.g.*, mercury.

(iii) Those liquids which conduct current due to drifting of †ions are known as **electrolytes** *e.g.*, solutions of CuSO_4 , AgNO_3 etc. This is the most important class of liquid conductors.

10.2. Electrolytes

A liquid which conducts electric current due to the drifting of ions is called an electrolyte.

Salts like silver nitrate (AgNO_3), sodium chloride (NaCl), copper sulphate (CuSO_4), etc. when dissolved in water dissociate into ions. Their ionic dissociation can be represented as under :



The atom or group of atoms having positive charge is called a *positive ion*. On the other hand, the atom or group of atoms having negative charge is called a *negative ion*. For example, when NaCl

* Conduction of current is possible only in those liquids which break up into oppositely charged atoms called *ions*. Such liquids are called *electrolytes*. There are, however, many substances (*e.g.*, sugar) which dissolve without splitting into ions. Solutions of these substances do not conduct current and are called non-electrolytes.

† An electrolyte has no free electrons.

is dissolved in water, it splits into positive ions (Na^+) and negative ions (Cl^-). The conduction of current through an electrolyte is due to the drifting of negative and positive ions within the liquid.

10.3. Mechanism of Ionisation

The splitting up of an ionic compound in solution into ions is known as **ionisation** or **ionic dissociation**. Let us take the example of sodium chloride (NaCl). The structure of this solid crystalline salt is made up of Na^+ and Cl^- ions. When in solid state, there is a very strong force of attraction between Na^+ and Cl^- ions which holds them together as a molecule of NaCl . However, when sodium chloride is dissolved in water, the force of attraction between the ions (Na^+ and Cl^-) of sodium chloride molecule is tremendously *reduced due to high permittivity of water ($K = 81$). In fact, the force of attraction between ions reduces 81 times. The result is that sodium ion (Na^+) and Cl^- ion get separated. This process is called ionisation. It may be noted that as soon as sodium chloride is dissolved in water, ions are formed. In other words, ions are present in an electrolytic solution even before it conducts electric current.

10.4. Electrolysis

The conduction of electric current through the solution of an electrolyte together with the resulting chemical changes is called **electrolysis**.

Fig. 10.1 shows the process of electrolysis in a copper voltameter. When copper sulphate (CuSO_4) is dissolved in water, it splits up into its components viz. the positive copper ions (Cu^{++}) and negative sulphate ions (SO_4^{--}). This process is called **ionisation. When d.c. voltage is applied across the electrodes, the negative sulphate ions (SO_4^{--}) move towards the anode (+ve electrode) and positive copper ions move towards the cathode (–ve electrode) causing the following chemical changes :

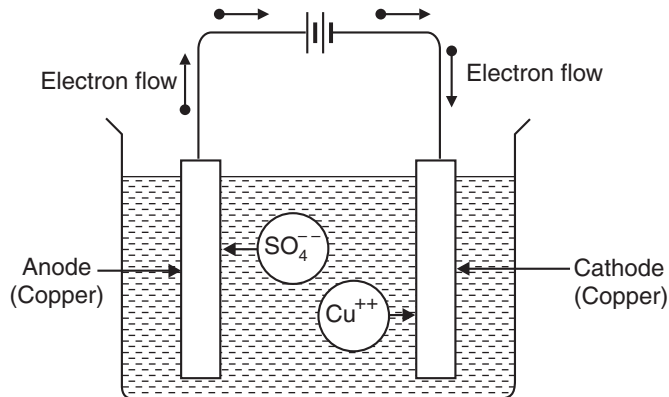
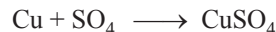


Fig. 10.1

At anode. A sulphate ion (SO_4^{--}) on reaching the anode gives its two extra electrons to it and becomes sulphate radical. These given up electrons continue their journey towards the cathode *via* the external circuit. Now the sulphate radical cannot exist and, therefore, it acts chemically on the anode material to form copper sulphate according to the following reaction :



Thus copper from anode continuously dissolves into the solution so long as this action takes place.

At cathode. At the same time, a copper ion (Cu^{++}) on reaching the cathode takes two electrons from it (these are the same electrons given by the sulphate ion at the anode and have come to cathode

* $F = \frac{1}{4\pi\epsilon_0 K} \frac{q_1 q_2}{r^2}$. For air, $K = 1$ and for water at room temperature, $K = 81$.

** The reader may recall that copper sulphate is an ionic compound *i.e.*, each molecule of CuSO_4 is formed due to the attraction between oppositely charged atoms viz Cu^{++} and SO_4^{--} . When dissolved in water, the force of attraction between them is tremendously reduced due to high relative permittivity of water. The result is that Cu^{++} and SO_4^{--} get separated. These charged atoms (Cu^{++} and SO_4^{--}) are called *ions*.

via the external circuit). The copper ion (Cu^{++}) combines with these two electrons to become copper atom and gets deposited on the cathode.



Thus copper from the solution (CuSO_4) gets deposited on the cathode.

The following points may be carefully noted :

(i) *Electrolysis is possible only if d.c. potential is applied to the electrodes.* It is because we are to attract ions of only one kind to each electrode.

(ii) *During electrolysis, either anode material gets deposited over the cathode or gases are liberated at the two electrodes.*

(iii) The resulting chemical changes during electrolysis take place so long as the current flows through the electrolyte. When the current through the electrolyte ceases, chemical action also ceases.

10.5. Back e.m.f. or Polarisation Potential

The process of electrolysis is carried out in an apparatus called *voltmeter* or *electrolytic cell*. When external d.c. voltage is applied across the electrodes, an e.m.f. is set up between each electrode and the electrolyte which opposes the external d.c. voltage. This opposing e.m.f. is called *back e.m.f.* (E_b) of the electrolyte and is produced due to the coating of electrodes by the products of electrolysis. This effect is called polarisation and for this reason, back e.m.f. is also called *polarisation potential*.

The e.m.f. set up in the voltmeter which opposes the external d.c. voltage is called back e.m.f. of the electrolyte.

The value of back e.m.f. is different for different electrolytes. For acids and alkalies which evolve hydrogen and oxygen, its value is about 1.7 V. For other electrolytes, the value of back e.m.f. depends on the particular salt and generally lies between 0.5 V and 2 V for normal solutions.

Voltage equation for electrolysis. For electrolysis, the applied external d.c. voltage V must overcome the back e.m.f. (E_b) and voltage drop (IR_e) in the electrolyte *i.e.*

$$V = E_b + IR_e \quad \dots (i)$$

where V = External d.c. voltage

E_b = Back e.m.f. of electrolyte

I = Circuit current

R_e = Resistance of electrolyte

Therefore, in order to carry out electrolysis at an appreciable rate, the external d.c. voltage V must be atleast equal to $E_b + IR_e$. If the external d.c. voltage is less than this value, electrolysis will not take place.

10.6. Faraday's Laws of Electrolysis

Faraday performed a series of experiments to determine the factors which govern the mass of an element deposited or liberated during electrolysis. He summed up his conclusions into two laws, known as Faraday's laws of electrolysis.

First law. *The mass of an element deposited or liberated at an electrode is directly proportional to the quantity of electricity that passes through the electrolyte.*

If m is the mass of an element deposited or liberated due to the passage of I amperes for t seconds, then according to first law,

$$m \propto Q$$

$$\text{or } m \propto It \quad (\because Q = It)$$

$$\text{or } m = ZIt \text{ or } m = ZQ$$

where Z is a constant known as *electro-chemical equivalent* (E.C.E.) of the element. It has the same value for one element but different for other elements.

If $Q = 1$ coulomb, then, $m = Z$.

Hence **electro-chemical equivalent** (E.C.E.) of an element is equal to the mass of element deposited or liberated by the passage of 1 coulomb of electricity through the electrolyte. Its unit is gm/C or kg/C.

For example, E.C.E. of copper is 0.000304 gm/C. It means that if 1 coulomb of electricity is passed through a solution of CuSO_4 , then mass of copper deposited on the cathode will be 0.000304 gm.

The validity of first law is explained by the fact that current inside the electrolyte is carried by the ions themselves. Hence the masses of the chemical substances reaching the anode and cathode are proportional to the quantity of electricity carried by the ions *i.e.*, mass of an ion liberated at any electrode is proportional to the quantity of electricity passed through the electrolyte.

Second law. The mass of an element deposited or liberated during electrolysis is directly proportional to the chemical equivalent weight of that element *i.e.*

$$m \propto \text{Chemical equivalent weight of the element (E)}$$

Faraday's second law is illustrated in Fig. 10.2 where silver and copper voltameters are connected in series. When the same current is passed for the same time through the two voltameters, it will be seen that the masses of silver (Ag) and copper (Cu) deposited on the respective cathodes are in the ratio of 108 : 32. These values of 108 and 32 are respectively the equivalent weights of silver and copper.

$$\frac{\text{Mass of silver deposited}}{\text{Mass of copper deposited}} = \frac{\text{Eq. wt. of Ag}}{\text{Eq. wt. of Cu}} = \frac{108}{32}$$

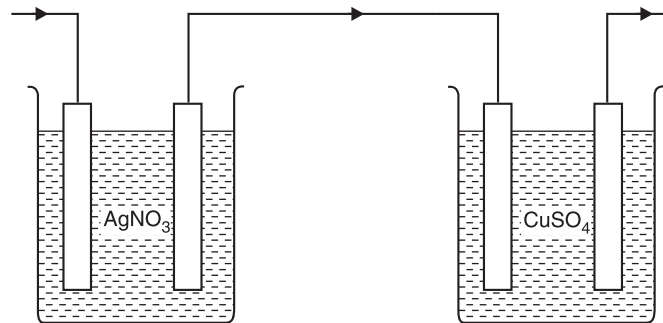


Fig. 10.2

Faraday's second law can be explained as follows. The negative ions (*i.e.* NO_3^- and SO_4^{2-}) from the solutions give up their respective extra electrons to the anodes. These electrons come to cathodes *via* the external circuit and are taken up by the positive ions (Ag^+ and Cu^{++}) to become metallic atoms and get deposited on the respective cathodes. Suppose 10 electrons are flowing in the external circuit. Since silver is monovalent (*i.e.*, its valency is 1), 10 silver ions must be liberated at the cathode of silver voltameter. Again copper is bivalent (*i.e.*, its valency is 2) and hence 5 copper ions must be liberated at the cathode of copper voltameter. This means that mass of an element (silver or copper) liberated is directly proportional to the atomic weight and inversely proportional to the valency of that element *i.e.*

$$\text{Mass liberated, } m \propto \frac{\text{Atomic weight}}{\text{Valency}}$$

$$\propto \text{Chemical equivalent wt. of the element}$$

i.e.

$$m \propto E$$

10.7. Relation Between E and Z

Suppose the same amount of charge (Q) is passed through the solutions of two electrolytes. If m_1 and m_2 are the masses of the substances liberated/deposited and Z_1 and Z_2 are their electro-chemical equivalents, then,

$$m_1 = Z_1 Q ; m_2 = Z_2 Q$$

$$\therefore \frac{m_1}{m_2} = \frac{Z_1}{Z_2} \quad \text{But} \quad \frac{m_1}{m_2} = \frac{E_1}{E_2} \quad \dots \text{Faraday's second law}$$

$$\therefore \frac{E_1}{E_2} = \frac{Z_1}{Z_2}$$

$$\text{or} \quad \frac{E}{Z} = \text{Constant}$$

Thus the ratio E/Z is the same for all substances. This constant is called Faraday constant $F (=E/Z)$.

Faraday constant (F). The value of Faraday constant is found to be 96500 C i.e. $F = 96500$ C.

Hence **Faraday constant** is the quantity of charge (i.e., 96500 C) required to liberate/deposit one gram equivalent (chemical equivalent in gram) of the substance during electrolysis.

For example, chemical equivalent of silver is 108. When a charge of 96500 C is passed through a silver voltameter, then mass of silver deposited on the cathode will be 108g. Again chemical equivalent of copper is 31.75. If a charge of 96500 C is passed through a copper voltameter, then mass of copper deposited on the cathode will be 31.75 g.

Finding the value of F . According to Faraday's first law of electrolysis,

$$m = ZQ$$

Suppose M is the mass of one mole of the substance. If, during electrolysis, the mass of the substance to be deposited is M and p is the valency of the depositing atom, then $N_A (= 6.023 \times 10^{23})$ atoms will deposit on the electrode.

$$\text{Now } m = M \text{ and } Q = N_A p e$$

where e = Charge on electron

$$\therefore M = Z N_A p e \quad \dots \text{Faraday's first Law}$$

$$\text{or} \quad Z = \frac{1}{N_A e} \cdot \frac{M}{p}$$

$$\text{But } \frac{M}{p} = \frac{\text{Mass of one mole}}{\text{Valency}} = \text{Chemical Equivalent of the substance}$$

$$\therefore Z = \frac{1}{N_A e} E$$

Now, $N_A e$ is a constant, called Faraday constant F .

$$\therefore Z = \frac{E}{F}$$

$$\text{Now, } 1F = N_A e = (6.023 \times 10^{23}) \times 1.602 \times 10^{-19} = 96485 \text{ C} \approx 96500 \text{ C}$$

10.8. Deduction of Faraday's Laws of Electrolysis

Suppose a charge Q is passed through an electrolyte during electrolysis and the mass liberated/deposited on the cathode is m .

$$\therefore m = ZQ$$

$$\text{or} \quad m = \frac{E}{F} Q \quad \left(\because Z = \frac{E}{F} \right) \quad \dots (i)$$

Equation (i) is the **fundamental equation of electrolysis** and contains Faraday's two laws of electrolysis.

(i) It is clear from equation (i) that :

$$m \propto Q \quad \dots \text{Faraday's first law}$$

(ii) If the same charge is passed (i.e. Q is constant) through different electrolytes during electrolysis, then,

$$m \propto E \quad \dots \text{Faraday's second law}$$

Example 10.1. A current passes through two voltmeters in series, one having silver plates and a solution of AgNO_3 , and the other copper plates and a solution of CuSO_4 . After the current has ceased to flow, 3.6 gm of silver have been deposited. How much copper will have deposited in the other voltmeter? Take E.C.E. of silver as 0.001118 gm/C and that of copper as 328.86×10^{-6} gm/C.

Solution. For silver voltmeter, we have,

$$m_1 = Z_1 It$$

$$\text{or} \quad It = \frac{m_1}{Z_1} = \frac{3.6}{0.001118}$$

$$\text{For copper voltmeter, } m_2 = Z_2 It = (328.86 \times 10^{-6}) \times \frac{3.6}{0.001118} = \mathbf{1.06 \text{ gm}}$$

Example 10.2. If 16 amperes deposit 12 gm of silver in 9 minutes, how much copper would 10 amperes deposit in 15 minutes? At. wt. of silver = 108 and At. wt. of copper = 63.5.

Solution. 16A in 9 minutes deposit silver = 12 gm

$$10\text{A in 15 minutes deposit silver} = 12 \times \frac{10}{16} \times \frac{15}{9} = 12.5 \text{ gm}$$

$$\text{Eq. wt. of silver} = \text{At. wt./Valency} = 108/1 = 108 ; \text{Eq. wt. of copper} = 63.5/2 = 31.75$$

Let m gm be the mass of copper deposited by 10 A in 15 minutes. Then by Faraday's second law of electrolysis,

$$\frac{\text{mass of Cu deposited}}{\text{mass of Ag deposited}} = \frac{\text{Eq. wt. of Cu}}{\text{Eq. wt. of Ag}}$$

$$\text{or} \quad \frac{m}{12.5} = \frac{31.75}{108} \quad \therefore m = \frac{31.75}{108} \times 12.5 = \mathbf{3.67 \text{ gm}}$$

Example 10.3. A coating of nickel 1 mm thick is to be deposited on a cylinder 2 cm in diameter and 30 cm in length. Calculate the time taken if the current used is 100 A. The following data may be taken. Specific gravity of nickel = 8.9, At. wt. of nickel = 58.7 (divalent), E.C.E. of silver = 1.12 mg/C, At.wt. of silver = 108.

$$\text{Solution. Area of curved surface of cylinder} = \pi D \times l = \pi \times 2 \times 30 = 188.5 \text{ cm}^2$$

$$\text{Volume of Ni to be deposited} = \text{Area of curved surface} \times \text{thickness of Ni}$$

$$= 188.5 \times 0.1 = 18.85 \text{ cm}^3$$

$$\text{Mass of Ni to be deposited, } m = 18.85 \times 8.9 = 167.7 \text{ gm}$$

$$\text{Eq. wt. of Ni} = 58.7/2 = 29.35 ; \text{Eq. wt. of Ag} = 108/1 = 108$$

$$\frac{\text{E.C.E. of Ni}}{\text{E.C.E. of Ag}} = \frac{\text{Eq. wt. of Ni}}{\text{Eq. wt. of Ag}} \quad \text{or} \quad \frac{\text{E.C.E. of Ni}}{1.12} = \frac{29.35}{108}$$

$$\therefore \text{E.C.E. of Ni} = \frac{29.35}{108} \times 1.12 = 0.304 \text{ mg/C}$$

$$\text{Now,} \quad m = ZIt$$

$$\therefore t = \frac{m}{ZI} = \frac{167.7}{0.304 \times 10^{-3} \times 100} = 5516 \text{ seconds} = \mathbf{91.93 \text{ minutes}}$$

Example 10.4. Find the thickness of copper deposited on a plate area of 0.00025 m^2 during electrolysis if a current of 1 A is passed for 100 minutes. Density of copper = 8900 kg/m^3 and E.C.E. of copper = $32.95 \times 10^{-8} \text{ kg/C}$.

Solution. According to Faraday's law of electrolysis, mass (m) of copper deposited is

$$m = ZIt$$

$$\text{Here, } Z = 32.95 \times 10^{-8} \text{ kg/C ; } I = 1 \text{ A ; } t = 100 \text{ min} = 100 \times 60 \text{ sec.}$$

$$\therefore m = 32.95 \times 10^{-8} \times 1 \times 100 \times 60 = 0.001977 \text{ kg}$$

$$\text{Volume of Cu deposited, } v = \frac{\text{Mass}}{\text{Density}} = \frac{0.001977}{8900} = 0.222 \times 10^{-6} \text{ m}^3$$

$$\therefore \text{Thickness of Cu deposited} = \frac{v}{\text{Plate area}} = \frac{0.222 \times 10^{-6}}{0.00025} = 0.888 \times 10^{-3} \text{ m} = \mathbf{0.888 \text{ mm}}$$

Example 10.5. An ammeter is being calibrated with the aid of copper voltameter. The ammeter continually reads 2 A when a current is passed through the voltameter for 1 hour. During this time, 2.34 gm of copper was liberated. Taking the electro-chemical equivalent of copper to be $330 \times 10^{-9} \text{ kg/C}$, determine the magnitude of error of the ammeter.

Solution. Let I amperes be the actual current.

$$\text{Now, } m = ZIt$$

$$\text{Here, } m = 2.34 \times 10^{-3} \text{ kg ; } Z = 330 \times 10^{-9} \text{ kg/C ; } t = 1 \text{ hr} = 3600 \text{ s}$$

$$\therefore I = \frac{m}{Zt} = \frac{2.34 \times 10^{-3}}{330 \times 10^{-9} \times 3600} = 1.97 \text{ A}$$

$$\therefore \text{Error} = 2 - 1.97 = \mathbf{0.03 \text{ A ; ammeter reads more}}$$

This example shows that the phenomenon of electrolysis can also be used to measure the magnitude of current.

Example 10.6. Find the mass of zinc which has been dissolved in a simple zinc-copper voltaic cell when 2200 J of energy has been supplied. Assume that electromotive force (e.m.f.) is constant at 1.1 V and that electro-chemical equivalent of zinc is $0.34 \times 10^{-6} \text{ kg/C}$.

Solution. A cell is a device which converts chemical energy into electrical energy. Faraday's law $m = ZQ$ is applicable to cells also but in this case, the mass m refers to the mass dissolved instead of mass liberated.

$$\text{Energy} = \text{Volts} \times \text{Coulombs}$$

$$\text{or } 2200 = 1.1 Q \therefore Q = 2200/1.1 = 2000 \text{ C}$$

$$\text{Now, } m = ZQ = (0.34 \times 10^{-6}) \times (2000) = 0.68 \times 10^{-3} \text{ kg} = 0.68 \text{ g}$$

$$\therefore \text{Mass of zinc dissolved} = \mathbf{0.68 \text{ g}}$$

Example 10.7. A steady direct current of 100 A flows for 5 minutes through fused sodium chloride. How much sodium will be drawn off and how much chlorine will be evolved? The atomic masses of sodium and chlorine are 23 and 35.5 respectively.

$$\text{Solution. } m = E \frac{Q}{F}$$

$$\text{Sodium } E_{\text{Na}} = \frac{23}{1} = 23 ; \frac{Q}{F} = \frac{100 \times 5 \times 60}{96500} = 0.311$$

$$\therefore m = 23 \times 0.311 = \mathbf{7.15 \text{ g}}$$

$$\text{Chlorine } E_{\text{Cl}} = 35.5/1 = 35.5 ; Q/F = 0.311 \quad (\text{same as before})$$

$$\therefore m = 35.5 \times 0.311 = \mathbf{11.04 \text{ g}}$$

Example 10.8. A steady current of 10.0 A is passed through a water voltameter for 300 s. Estimate the volume of hydrogen evolved at standard temperature and pressure. Use the known value of Faraday constant. Relative molecular mass of H_2 is 2.016 and molar volume = 22.4 litres (volume of 1 mole of an ideal gas at S.T.P.).

Solution.

$$m = Z I t$$

$$\text{Now, } Z = \frac{E}{F} = \frac{M}{pF} \quad \text{or} \quad m = \frac{M I t}{pF}$$

Here, $M = 2.016$; $p = 2$; $F = 96500 \text{ C}$; $I = 10.0 \text{ A}$; $t = 300 \text{ s}$

$$\therefore m = \frac{2.016 \times 10 \times 300}{2 \times 96500} = 0.0313 \text{ g}$$

$$\therefore \text{Volume of } 0.0313 \text{ g of } H_2 \text{ at STP} = \frac{22.4 \times 0.0313}{2.016} = \mathbf{0.35 \text{ litres}}$$

Example 10.9. The potential difference across the terminals of a battery of e.m.f. 12 V and internal resistance 2Ω drops to 10 V when it is connected to a silver voltameter. Calculate the silver deposited at the cathode in half an hour. Atomic weight of silver is 107.9 g mol^{-1} .

Solution. E.M.F. of battery, $E = 12 \text{ volts}$; Terminal p.d. of battery, $V = 10 \text{ volts}$; Internal resistance of battery, $r = 2 \Omega$; Resistance of voltameter = R .

$$\therefore r = \frac{E - V}{V} \times R \quad \text{or} \quad 2 = \frac{12 - 10}{10} \times R \quad \therefore R = 10 \Omega$$

$$\therefore \text{Circuit current, } I = \frac{E}{R + r} = \frac{12}{10 + 2} = 1 \text{ A}$$

$$\text{Now, } E_{Ag} = \frac{\text{Atomic weight}}{\text{Valency}} = \frac{107.9}{1} = 107.9 \text{ g mole}^{-1}$$

$$\text{Electrochemical equivalent, } Z = \frac{E_{Ag}}{F} = \frac{107.9}{96500} \text{ g C}^{-1}$$

\therefore Mass of silver deposited in half hour ($t = 30 \times 60 \text{ s}$) is

$$m = Z I t = \frac{107.9}{96500} \times 1 \times 30 \times 60 = \mathbf{2.01 \text{ g}}$$

Example 10.10. A silver and copper voltameters are connected across a 6 V battery of negligible resistance. In half hour, 1 g of copper and 2 g of silver are deposited. Calculate the rate at which energy is supplied by the battery. Given that E.C.E. of Cu is $3294 \times 10^{-7} \text{ g/C}$ and that of silver is $1118 \times 10^{-6} \text{ g/C}$.

Solution. We know that : $I = \frac{m}{Zt}$

$$\text{For copper voltameter, } I_1 = \frac{m_1}{Z_1 t} = \frac{1}{3294 \times 10^{-7} \times 1800} = 1.687 \text{ A}$$

$$\text{For silver voltameter, } I_2 = \frac{m_2}{Z_2 t} = \frac{2}{1118 \times 10^{-6} \times 1800} = 0.994 \text{ A}$$

Total current I drawn from the battery is

$$I = I_1 + I_2 = 1.687 + 0.994 = 2.681 \text{ A}$$

Rate at which energy is supplied by the battery is

$$P = VI = 6 \times 2.681 = \mathbf{16.1 \text{ W}}$$

Example 10.11. A refining plant employs 1000 electrolytic cells for copper refining. A current of 5000 A is used and the voltage per cell is 0.25 V. If the plant works for 100 hours/week, determine the annual output of refined copper and the energy consumed in kWh/tonne. The E.C.E. of copper = 1.1844 kg/1000 Ah.

Solution. Since the voltage drop across each electrolytic cell is less than 1 V, a number of cells are connected in series so that the generator can supply current at reasonable voltage.

$$\text{Supply voltage, } V = 0.25 \times 1000 = 250 \text{ volts}$$

$$\text{Circuit current, } I = 5000 \text{ A}$$

$$\text{Plant working time/year, } t = 100 \times 52 = 5200 \text{ hours/year}$$

$$\text{E.C.E. of copper, } Z = 1.1844 \text{ kg/1000 Ah}$$

$$= \frac{1.1844}{1000 \times 3600 \text{ As}} = 0.329 \times 10^{-6} \text{ kg/C} \quad (\because \text{As} = \text{C})$$

The amount (m) of refined copper per year is

$$m = Zit = 0.329 \times 10^{-6} \times 5000 \times 5200 \times 3600 = 30794 \text{ kg}$$

$$= \frac{30794}{1000} \text{ tonne} = 30.794 \text{ tonne}$$

$$\text{Energy consumption/year} = VIt = 250 \times 5000 \times 5200 \text{ Wh} = 6500 \times 10^6 \text{ Wh}$$

$$= \frac{6500 \times 10^6}{1000} \text{ kWh} = 6500 \times 10^3 \text{ kWh}$$

Since this energy consumption is for refining 30.794 tonne of copper,

$$\therefore \text{Energy consumption/tonne} = \frac{6500 \times 10^3}{30.794} = 211.08 \times 10^3 \text{ kWh/tonne}$$

Tutorial Problems

1. A current of 5 A flows for 40 minutes through an electrolyte which is a solution of a salt of chromium in water. Calculate the mass of chromium liberated. The electro-chemical equivalent of chromium is $90 \times 10^{-9} \text{ kg/C}$. **[1.08 gm]**
2. How long will it take to deposit, from a copper sulphate solution, a coating of copper 0.05 mm thick on an area of 118 cm^2 if the supply p.d. is 4.5 volts and the total resistance of the circuit is 2.3Ω . Specific gravity of copper is 8.93 and E.C.E. of copper = 0.329 mg/C . **[2.269 hr]**
3. A metal plate having a surface area of 115 cm^2 is to be silver plated. If a current of 1.5 A is passed for 1 hour and 30 minutes, what thickness of copper will be deposited? Specific gravity of silver = 10.5 and E.C.E. of silver = 1.118 mg/C . **[0.075 mm]**
4. A worn shaft is to be reconditioned by depositing chromium on its curved surface to a thickness of 0.1 mm. The shaft has a diameter of 3.5 cm and a length of 80 cm. If a current of 4.4 A is passed, calculate how long the plate will take. Density of chromium = 6600 kg/m^3 and E.C.E. of chromium = $90 \times 10^{-9} \text{ kg/C}$. **[41 hours, 44 minutes]**
5. Due to an error, a car battery is overcharged with a current of 5 A for 10 hours. Given that the electro-chemical equivalents of hydrogen and oxygen are $10.4 \times 10^{-9} \text{ kg/C}$ and $83.2 \times 10^{-9} \text{ kg/C}$ respectively, calculate the volume of distilled water which must be added to compensate for the loss. **[168 c.c.]**

10.9. Practical Applications of Electrolysis

The phenomenon of electrolysis has many industrial and commercial applications. A few of them are discussed below by way of illustration.

(i) **Electroplating.** The process of depositing a thin layer of superior metal (e.g., gold, silver, nickel, etc.) over an inferior metal (e.g., iron) is known as *electroplating*. The aim of electroplating