

# Chapter 12 Electrochemistry

Electrochemistry is the branch of physical chemistry which deals with the relationship between electrical energy and chemical changes taking place in redox reactions

# **Electrolytes and Electrolysis**

(1) **Definition :** "The substances whose aqueous solution undergo decomposition into ions when electric current is passed through them are known as **electrolytes** and the whole process is known as **electrolysis** or **electrolytic decomposition**."

Solutions of acids, bases, salts in water and fused salts etc. are the examples of electrolytes. Electrolytes may be **weak** or **strong**. Solutions of cane sugar, glycerine, alcohol etc., are examples of **non**electrolytes.

(2) **Electrolytic cell or Voltameter :** The device in which the process of electrolysis or electrolytic decomposition is carried out is known as *electrolytic cell* or *voltameter*.

(i) Voltameter convert *electrical energy into chemical energy*.

(ii) The electrode on which oxidation takes place is called **anode** (or +ve pole) and the electrode on which reduction takes place is called **cathode** (or -ve pole)

(iii) During electrolysis in voltameter cations are discharged on cathode and anions on anode.

(iv) In voltameter, outside the *electrolyte* electrons flow from *anode to cathode* and *current* flow from *cathode to anode*.

For voltameter,  $E_{cell} = -ve$  and  $\Delta G = +ve$ .

(v) The anions on reaching the anode give up their electrons and converted into the neutral atoms.

**At anode** :  $A^- \longrightarrow A + e^-$  (Oxidation)

(vi) On the other hand cations on reaching the cathode take up electrons supplied by battery and converted to the neutral atoms.

At cathode :  $B^+ + e^- \longrightarrow B$  (Reduction)

This overall change is known as *primary change* and products formed is known as *primary products*.

The primary products may be collected as such or they undergo further change to form molecules or compounds. These are called *secondary products* and the change is known as *secondary change*.

(3) **Preferential discharge theory :** According to this theory "If more than one type of ion is attracted towards a particular electrode, then the ion is discharged one which requires least energy or ions with lower discharge potential or which occur low in the electrochemical series".

The potential at which the ion is discharged or deposited on the appropriate electrode is termed *the* 

*discharge or deposition potential*, (D.P.). The values of discharge potential are different for different ions.

For cations:  $Li^+, K^+, Na^+, Ca^{2+}, Mg^{2+}, Al^{3+}, Zn^{2+},$  $Fe^{2+}, Ni^{2+}, H^+, Cu^{2+}, Hg^{2+}, Ag^+, Au^{3+}.$ 

The decreasing order of discharge potential or the increasing order of deposition of some of the ions is given below,

**For anions :**  $SO_4^{2-}$ ,  $NO_3^-$ ,  $OH^-$ ,  $Cl^-$ ,  $Br^-$ ,  $\Gamma$ .

Electrolyte	Electrode	Product at cathode	Product at anode
Aqueous NaOH	<i>Pt</i> or Graphite	$2H^+ + 2e^- \rightarrow H_2$	$2OH^- \rightarrow \frac{1}{2}O_2 + H_2O + 2e^-$
Fused NaOH	<i>Pt</i> or Graphite	$Na^+ + e^- \rightarrow Na$	$2OH^- \rightarrow \frac{1}{2}O_2 + H_2O + 2e^-$
Aqueous NaCl	<i>Pt</i> or Graphite	$2H^+ + 2e^- \rightarrow H_2$	$2Cl^- \rightarrow Cl_2 + 2e^-$
Fused NaCl	<i>Pt</i> or Graphite	$Na^+ + e^- \rightarrow Na$	$2Cl^- \rightarrow Cl_2 + 2e^-$
Aqueous <i>CuSO</i> <sub>4</sub>	<i>Pt</i> or Graphite	$Cu^{2+} + 2e^- \rightarrow Cu$	$2OH^- \rightarrow \frac{1}{2}O_2 + H_2O + 2e^-$
Aqueous CuSO <sub>4</sub>	<i>Cu</i> electrode	$Cu^{2+} + 2e^- \rightarrow Cu$	<i>Cu</i> oxidised to $Cu^{2+}$ ions
Dilute <i>H</i> <sub>2</sub> <i>SO</i> <sub>4</sub>	<i>Pt</i> electrode	$2H^+ + 2e^- \rightarrow H_2$	$2OH^- \rightarrow \frac{1}{2}O_2 + H_2O + 2e^-$
Conc. <i>H</i> <sub>2</sub> <i>SO</i> <sub>4</sub>	<i>Pt</i> electrode	$2H^+ + 2e^- \rightarrow H_2$	Peroxodisulphuric $acid(H_2S_2O_8)$
Aqueous <i>AgNO</i> <sub>3</sub>	<i>Pt</i> electrode	$Ag^+ + e^- \rightarrow Ag$	$2OH^- \rightarrow \frac{1}{2}O_2 + H_2O + 2e^-$
Aqueous <i>AgNO</i> <sub>3</sub>	Ag electrode	$Ag^+ + e^- \rightarrow Ag$	$Ag$ oxidised to $Ag^+$ ions

Table : 12.1 Products of electrol	ysis of some electrolytes
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(4) **Application of electrolysis :** Electrolysis has wide applications in industries. Some of the important applications are, as follows,

(i) Production of hydrogen by electrolysis of water.

(ii) Manufacture of heavy water  $(D_2 O)$ .

(iii) The metals like *Na*, *K*, *Mg*, *Al*, etc., are obtained by electrolysis of fused electrolytes.

(iv) Non-metals like hydrogen, fluorine, chlorine are obtained by electrolysis.

(v) In this method pure metal is deposited at cathode from a solution containing the metal ions Ag, Cu etc.

(vi) Compounds like *NaOH*, *KOH*,  $Na_2CO_3$ , *KClO*<sub>3</sub>, white lead, *KMnO*<sub>4</sub> etc. are synthesised by electrosynthesis method.

(vii) *Electroplating* : The process of coating an inferior metal with a superior metal by electrolysis is known as *electroplating*. The aim of electroplating is, to prevent the inferior metal from corrosion and to make it more attractive in appearance. The object to be plated is made the cathode of an electrolytic cell that contains a solution of ions of the metal to be deposited.

For	Anode	Cathode	Electrolyte

electroplati ng			
With copper	Си	Object	CuSO $_4$ +dilute $H_2SO_4$
With silver	Ag	Object	$K[Ag(CN)_2]$
With nickel	Ni	Object	Nickel ammonium sulphate
With gold	Au	Object	$K[Au(CN)_2]$
With zinc	Zn	Iron objects	ZnSO $_4$
With tin	Sn	Iron objects	SnSO <sub>4</sub>

**Thickness of coated layer :** Let the dimensions of metal sheet to be coated be  $(a cm \times b cm)$ .

Thickness of coated layer = c cm

Volume of coated layer =  $(a \times b \times c)cm^3$ 

Mass of the deposited substance = Volume × density =  $(a \times b \times c) \times dg$ 

$$\therefore \quad (a \times b \times c) \times d = \frac{I \times t \times E}{96500}$$

Using above relation we may calculate the thickness of coated layer.

# Faraday's laws of electrolysis

The laws, which govern the deposition of substances (In the form of ions) on electrodes during the process of electrolysis, is called *Faraday's laws of electrolysis*. These laws given by *Michael Faraday in* 1833.

(1) Faraday's first law : It states that,

"The mass of any substance deposited or liberated at any electrode is directly proportional to the quantity of electricity passed."i.e.,  $W \propto Q$ 

Where,

W = Mass of ions liberated in gm,

Q = Quantity of electricity passed in Coulombs

= Current in Amperes (I) × Time in second (t)  $\therefore W \propto I \times t$  or W = Z×I×t

In case current efficiency  $(\eta)$  is given, then

$$W = Z \times I \times t \times \frac{\eta}{100}$$

where, *Z* = constant, known as *electrochemical equivalent* (ECE) of the ion deposited.

When a current of 1 Ampere is passed for 1 second (*i.e.*, Q = 1 ), then, W = Z

Thus, *electrochemical equivalent* (ECE) may be defined as "the mass of the ion deposited by passing a current of one Ampere for one second (i.e., by passing Coulomb of electricity)". It's unit is gram per coulomb.

Coulomb is the unit of electrical charge.

96500 Coulombs =  $6.023 \times 10^{23}$  electrons = 1 mole electrons.

1 Coulomb = 
$$\frac{6.023 \times 10^{23}}{96500}$$
 = 6.28 × 10<sup>18</sup> electrons,

or 1 electronic charge =  $1.6 \times 10^{-19}$  Coulomb.

(2) Faraday's second law : It states that,

"When the same quantity of electricity is passed through different electrolytes, the masses of different ions liberated at the electrodes are directly proportional to their chemical equivalents (Equivalent weights)." i.e.,

$$\frac{W_1}{W_2} = \frac{E_1}{E_2} \text{ or } \frac{Z_1 I t}{Z_2 I t} = \frac{E_1}{E_2} \text{ or } \frac{Z_1}{Z_2} = \frac{E_1}{E_2} \qquad (\because W = Z I t)$$

Thus the electrochemical equivalent (Z) of an element is directly proportional to its equivalent weight (E), *i.e.*,

 $E \propto Z$  or E = FZ or  $E = 96500 \times Z$ 

where,  $F = Faraday \ constant = 96500 \ C \ mol^{-1}$ 

So, 1 Faraday = 1F =Electrical charge carried out by one mole of electrons.

1*F* = Charge on an electron × Avogadro's number.  
1*F* = 
$$e^- \times N = (1.602 \times 10^{-19} c) \times (6.023 \times 10^{23} mol^{-1}).$$

Number of Faraday = 
$$\frac{\text{Number of electrons passed}}{6.023 \times 10^{23}}$$

(3) Faraday's law for gaseous electrolytic product For the gases, we use

$$V = \frac{It V_e}{96500}$$

where, V = Volume of gas evolved at *S.T.P.* at an electrode

 $V_e$  = Equivalent volume = Volume of gas evolved at an electrode at *S.T.P.* by 1 Faraday charge

(4) **Quantitative aspects of electrolysis** : We know that, one Faraday (1*F*) of electricity is equal to the charge carried by one mole  $(6.023 \times 10^{23})$  of electrons. So, in any reaction, if one mole of electrons are involved, then that reaction would consume or produce 1*F* of electricity. Since 1*F* is equal to 96,500 Coulombs, hence 96,500 Coulombs of electricity would cause a reaction involving one mole of electrons.

If in any reaction, *n* moles of electrons are involved, then the total electricity (*Q*) involved in the reaction is given by,  $Q = nF = n \times 96,500 C$ 

Thus, the amount of electricity involved in any reaction is related to,

(i) The number of moles of electrons involved in the reaction,

(ii) The amount of any substance involved in the reaction.

Therefore, 1 Faraday or 96,500 *C* or 1 mole of electrons will reduce,

(a) 1 mole of monovalent cation,(b) 1/2mole of divalent cation,

(c) 1/3 mole of trivalent cation, (d) 1/n mole of n valent cations.

# Metallic and Electrolytic conductors

All substances do not conduct electrical current. The substances, which allow the passage of electric current, are called **conductors**. The best metal conductors are such as copper, silver, tin, etc. On the other hand, the substances, which do not allow the passage of electric current through them, are called **non-conductors or insulators**. Some common examples of insulators are rubber, wood, wax, etc.

The conductors are broadly classified into two types,

Metallic and electrolytic conductors.

Metallic conduction	Electrolytic conduction	
(i) It is due to the flow of electrons.	(i) It is due to the flow of ions.	
(ii) It is not accompanied by decomposition of the substance.(Only physical changes occurs)	(ii) It is accompanied by decomposition of the substance. (Physical as well as chemical change occur)	
(iii) It does not involve	(iii) It involves transfer of	

transfer of matter.	matter in the form of ions.		
(iv) Conductivity decreases with increase in temperature.	(iv) Conductivity increases with increases in temperature and degree of hydration due to decreases in viscosity of medium.		

The electrolyte may, therefore, be defined as the substance whose aqueous solution or fused state conduct electricity accompanied by chemical decomposition. The conduction of current through electrolyte is due to the movement of ions.

On the contrary, substances, which in the form of their solutions or in their molten state do not conduct electricity, are called non-electrolytes.

# **Electrolytic conduction**

When a voltage is applied to the electrodes dipped into an electrolytic solution, ions of the electrolyte move and, therefore, electric current flows through the electrolytic solution. The power of the electrolytes to conduct electric current is termed *conductance* or *conductivity*.

(1) **Ohm's law :** This law states that the current flowing through a conductor is directly proportional to the potential difference across it, *i.e.*,  $I \propto V$ 

where I is the current strength (In Amperes) and V is the potential difference applied across the conductor (In Volts)

or 
$$I = \frac{V}{R}$$
 or  $V = IR$ 

where *R* is the constant of proportionality and is known as resistance of the conductor. It is expressed in Ohm's and is represented as  $\Omega$  The above equation is known as Ohm's law. Ohm's law may also be stated as,

"the strength of current flowing through a conductor is directly proportional to the potential difference applied across the conductor and inversely proportional to the resistance of the conductor."

(2) **Resistance** : *It measures the obstruction to the flow of current*. The resistance of any conductor is directly proportional to the length (*l*) and inversely proportional to the area of cross-section (*a*) so that

$$R \propto \frac{l}{a}$$
 or  $R = \rho \frac{l}{a}$ 

where  $\rho$  (rho) is the constant of proportionality and is called specific resistance or resistivity. The resistance depends upon the nature of the material.

**Units** : The unit of resistance is ohm ( $\Omega$ ). In terms of SI, base unit is equal to  $(kgm^2)/(s^3A^2)$ .

(3) **Resistivity or specific resistance :** We know that resistance *R* is

$$R = \rho \frac{l}{a}$$
; Now, if  $l = 1 cm, a = 1 cm^2$  then  $R = \rho$ 

Thus, resistivity is defined as the resistance of a conductor of 1 cm length and having area of cross-section equal to  $1 cm^2$ .

**Units** : The units of resistivity are  $\rho = R. \frac{a}{l} = Ohm \frac{cm^2}{cm} = Ohm. cm$ 

Its *SI* units are Ohm metre 
$$(\Omega m)$$
. But quite often Ohm centimetre  $(\Omega cm)$  is also used.

(4) **Conductance :** It is a measure of the ease with which current flows through a conductor. It is an additive property. It is expressed as *G*. It is reciprocal of the resistance, *i.e.*,

$$G = \frac{1}{R}$$

**Units** : The units of conductance are reciprocal  $Ohm(ohm^{-1})$  or mho. Ohm is also abbreviated as  $\Omega$  so that  $Ohm^{-1}$  may be written as  $\Omega^{-1}$ .

According to SI system, the units of electrical conductance is Siemens, S (*i.e.*,  $1S = 1\Omega^{-1}$ ).

(5) **Conductivity :** The inverse of resistivity is called conductivity (or specific conductance). It is represented by the symbol,  $\kappa$  (Greek kappa). The *IUPAC* has recommended the use of term conductivity over specific conductance. It may be defined as, the conductance of a solution of 1 cm length and having 1 sq. cm as the area of cross-section. In other words, conductivity is the conductance of one centimetre cube of a solution of an electrolyte.

Thus, 
$$\kappa = \frac{1}{\rho}$$

Units : The units of conductivity are

$$\kappa = \frac{1}{Ohm.\,cm} = Ohm^{-1}\,cm^{-1}$$
 or  $\Omega^{-1}\,cm^{-1}$ 

In *SI* units, *l* is expressed in *m* area of crosssection in  $m^2$  so that the units of conductivity are  $S m^{-1}$ .

(6) Molar conductivity or molar conductance : Molar conductivity is defined as the conducting power of all the ions produced by dissolving one mole of an electrolyte in solution.

It is denoted by  $\Lambda$  (lambda). Molar conductance is related to specific conductance ( $\kappa$ ) as,

$$\Lambda = \frac{\kappa}{M}$$

where, M is the molar concentration.

If *M* is in the units of molarity *i.e.*, moles per litre (mol  $L^{-1}$ ), the  $\Lambda$  may be expressed as,

$$\Lambda = \frac{\kappa \times 1000}{M}$$

For the solution containing 1 *gm* mole of electrolyte placed between two parallel electrodes of 1 *sq. cm* area of cross-section and one *cm* apart,

Conductanc  $e(G) = Conductivity = Molar conductivity(\Lambda)$ 

But if solution contains 1 *gm* mole of the electrolyte therefore, the measured conductance will be the molar conductivity. Thus,

Molar conductivity( $\Lambda$ ) = 100 × Conductivity

In other words,  $(\Lambda) = \kappa \times V$ 

where V is the volume of the solution in  $cm^3$  containing one gram mole of the electrolyte.

If M is the concentration of the solution in mole per litre, then

*M* mole of electrolyte is present in  $1000 \text{ cm}^3$ 

1 mole of electrolyte is present in 
$$=\frac{1000}{M}cm^3$$
 of

solution

Thus,  $\Lambda = \kappa \times \text{Volume in } cm^3 \text{ containing 1 mole of electrolyte.}$ 

or  $\Lambda = \frac{\kappa \times 1000}{M}$ 

**Units of Molar Conductance** : The units of molar conductance can be derived from the formula ,

$$\Lambda = \frac{\kappa \times 1000}{M}$$

The units of  $\kappa$  are  $S cm^{-1}$  and units of  $\Lambda$  are,  $\Lambda = S cm^{-1} \times \frac{cm^3}{mol} = S cm^2 mol^{-1} = S cm^2 mol^{-1}$ 

According to *SI* system, molar conductance is expressed as  $S m^2 mol^{-1}$ , if concentration is expressed as  $mol m^{-3}$ .

(7) **Equivalent conductivity :** It is defined as the conducting power of all the ions produced by dissolving one gram equivalent of an electrolyte in solution.

It is expressed as  $\Lambda_e$  and is related to specific conductance as

 $\Lambda_e = \frac{\kappa \times 1000}{C} = \kappa \times \frac{1000}{M} \qquad (M \text{ is Molarity of the}$ 

solution)

where *C* is the concentration in gram equivalent per litre (or Normality). This term has earlier been quite frequently used. Now it is replaced by molar conductance. The units of equivalent conductance are  $Ohm^{-1} cm^2 (gm equiv)^{-1}$ .

# (8) Experimental measurement of conductance

(i) The conductance of a solution is reciprocal of the resistance, therefore, the experimental

determination of the conductance of a solution involves the measurement of its resistance.

(ii) **Calculation of conductivity** : We have seen that conductivity ( $\kappa$ ) is reciprocal of resistivity ( $\rho$ ), *i.e.*,

$$\kappa = \frac{1}{\rho}$$
 and  $\rho = R \frac{a}{l}$   
 $\therefore \quad \kappa = \frac{1}{R} \left( \frac{l}{a} \right)$  or  $\kappa = G \left( \frac{l}{a} \right)$ 

where *G* is the conductance of the cell, *l* is the distance of separation of two electrodes having cross section area  $a cm^2$ .

The quantity  $\left(\frac{l}{a}\right)$  is called cell constant and is

expressed in  $cm^{-1}$ . Knowing the value of cell constant and conductance of the solution, the specific conductance can be calculated as,

 $\kappa = G \times \text{Cell constant}$ 

*i.e.*, Conductivity = Conductanc  $e \times Cell$  constant

# Factors affecting the electrolytic conductance

In general, conductance of an electrolyte depends upon the following factors,

(1) **Nature of electrolyte :** The conductance of an electrolyte depends upon the number of ions present in the solution. Therefore, the greater the number of ions in the solution the greater is the conductance. The number of ions produced by an electrolyte depends upon its nature. The strong electrolytes dissociate almost completely into ions in solutions and, therefore, their solutions have high conductance. On the other hand, weak electrolytes, dissociate to only small extents and give lesser number of ions. Therefore, the solutions of weak electrolytes have low conductance.

(2) **Concentration of the solution :** The molar conductance of electrolytic solution varies with the concentration of the electrolyte. In general, the molar conductance of an electrolyte increases with decrease in concentration or increase in dilution.

The molar conductance of strong electrolyte ( $HCl, KCl, KNO_3$ ) as well as weak electrolytes ( $CH_3COOH, NH_4OH$ ) increase with decrease in concentration or increase in dilution. The variation is however different for strong and weak electrolytes.

The variation of molar conductance with concentration can be explained on the basis of conducting ability of ions for weak and strong electrolytes.

For weak electrolytes the variation of  $\Lambda$  with dilution can be explained on the bases of number of ions in solution. The number of ions furnished by an

electrolyte in solution depends upon the degree of dissociation with dilution. With the increase in dilution, the degree of dissociation increases and as a result molar conductance increases. The limiting value of molar conductance ( $\Lambda^0$ ) corresponds to degree of dissociation equal to 1 *i.e.*, the whole of the electrolyte dissociates.

Thus, the degree of dissociation can be calculated at any concentration as,

 $\alpha = \frac{\Lambda^c}{\Lambda^0}$ 

where  $\alpha$  is the degree of dissociation,

 $\Lambda^c$  is the molar conductance at concentration  ${\it C}$  and

 $\Lambda^0\,$  is the molar conductance at infinite dilution.

For strong electrolytes, there is no increase in the number of ions with dilution because strong electrolytes are completely ionised in solution at all concentrations (Bv definition). However. in concentrated solutions of strong electrolytes there are strong forces of attraction between the ions of opposite charges called inter-ionic forces. Due to these interionic forces the conducting ability of the ions is less in concentrated solutions. With dilution, the ions become far apart from one another and inter-ionic forces decrease. As a result, molar conductivity increases with dilution. When the concentration of the solution becomes very-very low, the inter-ionic attractions become negligible and the molar conductance approaches the limiting value called molar conductance at infinite dilution. This value is characteristic of each electrolyte.

(3) **Temperature** : The conductivity of an electrolyte depends upon the temperature. With increase in temperature, the conductivity of an electrolyte increases.

# **Migration of ions**

Electricity is carried out through the solution of an electrolyte by *migration of ions*. Therefore,

(1) Ions move toward oppositely charged electrodes at different speeds.

(2) During electrolysis, ions are discharged or liberated in equivalent amounts at the two electrodes, *no matter what their relative speed is.* 

(3) Concentration of the electrolyte changes around the electrode due to difference in the speed of the ions.

(4) Loss of concentration around any electrode is proportional to the speed of the ion that moves away from the electrode, so

Loss around anode	_ Speed of cation
Loss around cathode	Speed of anion

The relation is valid only when the discharged ions do not react with atoms of the electrodes. But when the ions combine with the material of the electrode, the concentration around the electrode shows an increase.

# Transport number or Transference number

(1) **Definition :** "The fraction of the total current carried by an ion is known as transport number, transference number or Hittorf number may be denoted by sets symbols like  $t_+$  and  $t_-$  or  $t_c$  and  $t_a$  or  $n_c$  and  $n_a$ ".

From this definition,

 $t_a = \frac{\text{Current carried by an anion}}{\text{Total current passed through the solution}}$  $t_c = \frac{\text{Current carried by a cation}}{\text{Total current passed through the solution}}$ evidently,  $t_a + t_c = 1$ .

(2) **Determination of transport number :** Transport number can be determined by *Hittorf's method, moving boundary method, emf method and from ionic mobility.* 

# (3) Factors affecting transport number

A rise in temperature tends to bring the transport number of cation and anion more closer to 0.5

(4) **Transport number and Ionic mobility :** Ionic mobility or Ionic conductance is the conductivity of a solution containing 1 g ion, at infinite dilution, when two sufficiently large electrodes are placed 1 cm apart.

Ionic mobilities  $(\lambda_a \text{ or } \lambda_c) \propto \text{speeds of ions } (u_a \text{ or } u_c)$ 

Unit of ionic mobility is Ohm<sup>-1</sup> cm<sup>2</sup> or V<sup>-1</sup>S<sup>-1</sup>cm<sup>2</sup>

Ionic mobility and transport number are related

as,

$$\lambda_a \operatorname{or} \lambda_c = t_a \operatorname{or} t_c \times \lambda_\infty$$

Absolute ionic mobility is the mobility with which the ion moves under unit potential gradient. It's unit is  $cm \sec^{-1}$ .

Absolute ionic mobility =  $\frac{\text{Ionic mobility}}{96.500}$ 

# Kohlrausch's law

(1) Kohlrausch law states that, "At time infinite dilution, the molar conductivity of an electrolyte can be expressed as the sum of the contributions from its individual ions" i.e.,  $\Lambda_m^{\infty} = v_+ \lambda_+^{\infty} + v_- \lambda_-^{\infty}$ , where,  $v_+$  and  $v_-$  are the number of cations and anions per formula unit of electrolyte respectively and,  $\lambda_+^{\infty}$  and  $\lambda_-^{\infty}$  are the molar conductivities of the cation and anion at infinite dilution respectively. The use of above equation in

expressing the molar conductivity of an electrolyte is illustrated as,

The molar conductivity of *HCl* at infinite dilution can be expressed as,

$$\begin{split} \Lambda^{\infty}_{HCl} &= v_{H^+} \lambda^{\infty}_{H^+} + v_{Cl^-} \lambda^{\infty}_{Cl^-} \text{; For } HCl, \quad v_{H^+} = 1 \quad \text{and} \\ v_{Cl^-} &= 1. \quad \text{So,} \quad \Lambda^{\infty}_{HCl} = (1 \times \lambda^{\infty}_{H^+}) + (1 \times \lambda^{\infty}_{Cl^-}) \text{; Hence,} \\ \Lambda^{\infty}_{HCl} &= \lambda^{\infty}_{H^+} + \lambda^{\infty}_{Cl^-} \end{split}$$

(2) **Applications of Kohlrausch's law :** Some typical applications of the Kohlrausch's law are described below,

(i) **Determination of**  $\Lambda_m^{\infty}$  for weak electrolytes : The molar conductivity of a weak electrolyte at infinite dilution  $(\Lambda_m^{\infty})$  cannot be determined by extrapolation method. However,  $\Lambda_m^{\infty}$  values for weak electrolytes can be determined by using the Kohlrausch's equation.

 $\Lambda^{\infty}_{CH_{3}COOH} = \Lambda^{\infty}_{CH_{3}COONa} + \Lambda^{\infty}_{HCl} - \Lambda^{\infty}_{NaCl}$ 

(ii) **Determination of the degree of ionisation of a weak electrolyte**: The Kohlrausch's law can be used for determining the degree of ionisation of a weak electrolyte at any concentration. If  $\lambda_m^c$  is the molar conductivity of a weak electrolyte at any concentration *C* and,  $\lambda_m^\infty$  is the molar conductivity of a electrolyte at infinite dilution. Then, the degree of ionisation is given

by, 
$$\alpha_c = \frac{\Lambda_m^c}{\Lambda_m^\infty} = \frac{\Lambda_m^c}{(\nu_+ \lambda_+^\infty + \nu_- \lambda_-^\infty)}$$

Thus, knowing the value of  $\Lambda_m^c$ , and  $\Lambda_m^\infty$  (From the Kohlrausch's equation), the degree of ionisation at any concentration ( $\alpha_c$ ) can be determined.

(iii) **Determination of the ionisation constant of a weak electrolyte** : Weak electrolytes in aqueous solutions ionise to a very small extent. The extent of ionisation is described in terms of the degree of ionisation ( $\alpha$ ). In solution, the ions are in dynamic equilibrium with the unionised molecules. Such an equilibrium can be described by a constant called **ionisation constant**. For example, for a weak electrolyte *AB*, the ionisation equilibrium is, *AB* =  $A^+ + B^-$ ; If *C* is the initial concentration of the electrolyte *AB* in solution, then the equilibrium concentrations of various species in the solution are,  $[AB] = C(1 - \alpha), [A^+] = C\alpha$  and  $[B^-] = C\alpha$ 

Then, the ionisation constant of AB is given by,

$$K = \frac{[A^+][B^-]}{[AB]} = \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{(1-\alpha)}$$

We know, that at any concentration *C*, the degree of ionisation ( $\alpha$ ) is given by,  $\alpha = \Lambda_m^c / \Lambda_m^\infty$ 

Then, 
$$K = \frac{C(\Lambda_m^c / \Lambda_m^\infty)^2}{[1 - (\Lambda_m^c / \Lambda_m^\infty)]} = \frac{C(\Lambda_m^c)^2}{\Lambda_m^\infty (\Lambda_m^\infty - \Lambda_m^c)}$$
; Thus,

knowing  $\Lambda_m^{\infty}$  and  $\Lambda_m^c$  at any concentration, the ionisation constant (*K*) of the electrolyte can be determined.

(iv) **Determination of the solubility of a sparingly soluble salt**: The solubility of a sparingly soluble salt in a solvent is quite low. Even a saturated solution of such a salt is so dilute that it can be assumed to be at infinite dilution. Then, the molar conductivity of a sparingly soluble salt at infinite dilution  $(\Lambda_m^{\infty})$  can be obtained from the relationship,

The conductivity of the saturated solution of the sparingly soluble salt is measured. From this, the conductivity of the salt ( $\kappa_{salt}$ ) can be obtained by using the relationship,  $\kappa_{salt} = \kappa_{sol} - \kappa_{water}$ , where,  $\kappa_{water}$  is the conductivity of the water used in the preparation of the saturated solution of the salt.

From equation (i) and (ii);

$$C_m = \frac{1000 \ \kappa_{salt}}{(\nu_+ \ \lambda_+^{\infty} + \nu_- \ \lambda_-^{\infty})}, \quad C_m \text{ is the molar concentration}$$

of the sparingly soluble salt in its saturated solution. Thus,  $C_m$  is equal to the solubility of the sparingly soluble salt in the *mole per litre* units. The solubility of the salt in *gram per litre* units can be obtained by multiplying  $C_m$  with the molar mass of the salt.

# **Electrochemical or Galvanic cell**

"Electrochemical cell or Galvanic cell is a device in which a spontaneous redox reaction is used to convert chemical energy into electrical energy i.e. electricity can be obtained with the help of oxidation and reduction reaction".



(i) Electrochemical cell consists of two vessels, two electrodes, two electrolytic solutions and a salt bridge.

(ii) The two electrodes taken are made of different materials and usually set up in two separate vessels.

(iii) The electrolytes are taken in the two different vessels called as half - cells.

(iv) The two vessels are connected by a salt bridge/porous pot.

(v) The electrode on which oxidation takes place is called the **anode** (or – *ve* pole) and the electrode on which reduction takes place is called the **cathode** (or + *ve* pole).

(vi) In electrochemical cell, ions are discharged only on the cathode.

(vii) Like electrolytic cell, in electrochemical cell, from outside the electrolytes electrons flow from anode to cathode and current flow from cathode to anode.

(viii) For electrochemical cell,

 $E_{cell} = +ve, \ \Delta G = -ve.$ 

(ix) In a electrochemical cell, cell reaction is exothermic.

## (2) Salt bridge and its significance

(i) Salt bridge is U – shaped glass tube filled with a gelly like substance, agar – agar (plant gel) mixed with an electrolyte like *KCl*, *KNO*<sub>3</sub>, *NH*<sub>4</sub>*NO*<sub>3</sub> etc.

(ii) The electrolytes of the two half-cells should be inert and should not react chemically with each other.

(iii) The cation as well as anion of the electrolyte should have same ionic mobility and almost same transport number, viz.  $KCl, KNO_3, NH_4NO_3$  etc.

(iv) The following are the functions of the salt bridge,

(a) It connects the solutions of two half - cells and completes the cell circuit.

(b) It prevent transference or diffusion of the solutions from one half cell to the other.

(c) It keeps the solution of two half - cells electrically neutral.

(d) It prevents liquid – liquid junction potential *i.e.* the potential difference which arises between two solutions when they contact with each other.

#### (3) Representation of an electrochemical cell

The cell may be written by arranging each of the pair left – right, anode – cathode, oxidation – reduction, negative and positive in the alphabetical order as,



(4) **Reversible and irreversible cells :** A cell is said to be reversible if the following two conditions are fulfilled

(i) The chemical reaction of the cell stops when an exactly equal external *emf* is applied.

(ii) The chemical reaction of the cell is reversed and the current flows in opposite direction when the external *emf* is slightly higher than that of the cell. Any other cell, which does not obey the above two conditions, is termed as irreversible. Daniell cell is reversible but  $Zn|H_2SO_4|Ag$  cell is irreversible in nature

(5) **Types of electrochemical cells :** Two main types of electrochemical cells have been reported, these are,

(i) *Chemical cells* : The cells in which electrical energy is produced from the energy change accompanying a chemical reaction or a physical process are known as chemical cells. Chemical cells are of two types,

(a) *Chemical cells without transference* : In this type of chemical cells, the liquid junction potential is neglected or the transference number is not taken into consideration. In these cells, one electrode is reversible to cations while the other is reversible to the anions of the electrolyte.

(b) Chemical cells with transference : In this type of chemical cells, the liquid-liquid junction potential or diffusion potential is developed across the boundary between the two solutions. This potential develops due to the difference in mobilities of +ve and -ve ions of the electrolytes.

(6) **Concentration cells :** "A cell in which electrical energy is produced by the transference of a substance from a system of high concentration to one at low concentration is known as **concentration cells**". Concentration cells are of two types.

(i) *Electrode concentration cells* : In these cells, the potential difference is developed between two electrodes at different concentrations dipped in the same solution of the electrolyte. For example, two hydrogen electrodes at different gaseous pressures in the same solution of hydrogen ions constitute a cell of this type.

$$\frac{Pt, H_2(\text{pressure } p_1)}{\text{Anode}} | H^+| = \frac{H_2(\text{pressure } p_2)Pt}{\text{Cathode}};$$

$$E_{\text{cell}} = \frac{0.0591}{2} \log \frac{(p_1)}{(p_2)}$$
 at 25 °C If  $p_1 > p_2$ , oxidation occurs

at L. H. S. electrode and reduction occurs at R. H. S. electrode.

In the amalgam cells, two amalgams of the same metal at two different concentrations are immersed in the same electrolytic solution.  $M(Hg C_1) | M^{n+} | Zn(Hg C_2)$ 

P

The *emf* of the cell is given by the expression,  $E_{\text{cell}} = \frac{0.0591}{n} \log \frac{C_1}{C_2}$  at 25 ° C

(ii) *Electrolyte concentration cells* : In these cells, electrodes are identical but these are immersed in solutions of the same electrolyte of different concentrations. The source of electrical energy in the cell is the tendency of the electrolyte to diffuse from a solution of higher concentration to that of lower concentration. With the expiry of time, the two concentrations tend to become equal. Thus, at the start the *emf* of the cell is maximum and it gradually falls to zero. Such a cell is represented in the following manner ( $C_2$  is greater then  $C_1$ ).

$$M \mid M^{+n}(C_1) \mid \mid M^{n+}(C_2) \mid M$$
or
$$\frac{Zn \mid Zn^{2+}(C_1) \mid Zn^{2+}(C_2) \mid}{Zn^{2+}(C_2) \mid Zn^{2+}(C_2) \mid}$$

or  $\frac{Zn|Zn^{-1}(C_1)|}{\text{Anode}} \| \frac{Zn^{2+}(C_2)|Zn}{\text{Cathode}} \|$ 

The *emf* of the cell is given by the following expression,

$$E_{\text{cell}} = \frac{0.0591}{n} \log \frac{C_{2(R,H,S)}}{C_{1(L,H,S)}} e \text{ at } 25^{\circ} \text{ C}$$

The concentration cells are used to determine the solubility of sparingly soluble salts, valency of the cation of the electrolyte and transition point of the two allotropic forms of metal used as electrodes, etc.

(7) **Heat of reaction in an electrochemical cell** : Let *n* Faraday charge flows out of a cell of *emf E*, then

$$-\Delta G = nFE \qquad \dots \dots (i)$$

Gibbs – Helmholtz equation from thermodynamics may be given as

From equation (i) and (ii) we get,

$$-nFE = \Delta H + T \left[ \frac{\partial (-nFE)}{\partial T} \right]_{p} = \Delta H - nFT \left( \frac{\partial E}{\partial T} \right)_{p}$$
$$\Delta H = -nFE + nFT \left( \frac{\partial E}{\partial T} \right)_{p}$$

where  $\left(\frac{\partial E}{\partial T}\right)_p$  = Temperature coefficient of cell

**Case I:** When  $\left(\frac{\partial E}{\partial T}\right)_p = 0$ , then  $\Delta H = -nFE$ 

**Case II:** When  $\left(\frac{\partial E}{\partial T}\right)$  > 0, then  $nFE > \Delta H$ , *i.e.* 

process inside the cell is endothermic.

**Case III:** When 
$$\left(\frac{\partial E}{\partial T}\right) < 0$$
, then  $nFE < \Delta H$ , *i.e.*,

process inside the cell is exothermic.

# Some Commercial cell (Batteries)

One of the main use of galvanic cells is the generation of portable electrical energy. These cells are also popularly known as **batteries**. The term battery is generally used for two or more Galvanic cells connected in series. Thus, a **battery** is an arrangement of electrochemical cells used as an energy source. The basis of an electrochemical cell is an oxidation – reduction reaction.

**Types of commercial cells :** There are mainly two types of commercial cells,

(1) **Primary cells :** In these cells, the electrode reactions cannot be reversed by an external electric energy source. In these cells, reactions occur only once and after use they become dead. Therefore, they are **not chargeable.** Some common example are, dry cell, mercury cell, Daniell cell and alkaline dry cell

(i) Voltaic cellCu rod  $\mathbf{P} \ominus \mathbf{P} \mathbf{Z}n$  rod



Cathode : Cu rodAnode : Zn rodElectrolyte : dil.  $H_2SO_4$ Emf : 1.08 V

**At cathode :**  $Cu^{2+} + 2e^- \rightarrow Cu$ 

**At Anode** :  $Zn \rightarrow Zn^{2+} + 2e^{-}$ 

**Over all reaction** : 
$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$



Cathode : Cu rodAnode : Zn rodElectrolyte : dil.  $H_2SO_4$ Emf : 1.1 V

At cathode :  $Cu^{2+} + 2e^- \rightarrow Cu$ 

**At Anode** :  $Zn \rightarrow Zn^{2+} + 2e^{-}$ 

**Over all reaction** :  $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ (iii) **Lechlanche cell (Dry cell)** 



**Cathode** : Graphite rod **Anode** : Zn pot **Electrolyte** : Paste of  $NH_4Cl + ZnCl_2$  in starch Emf : 1.2 V to 1.5 V At cathode :  $NH_4^+ + MnO_2 + 2e^- \rightarrow MnO(OH)^- + NH_3$ At Anode :  $Zn \rightarrow Zn^{2+} + 2e^-$ Over all reaction :  $Zn + NH_4^+ + MnO_2 \rightarrow Zn^{2+} + MnO(OH)^- + NH_3$ (iv) Mercury cell

Cathode : Mercury (II) oxide	Anode :	
Zn rod		
<b>Electrolyte</b> : Paste of <i>KOH</i> + <i>ZnO</i>	Emf :	
1.35 V		
<b>At cathode</b> : $HgO_{(s)} + H_2O_{(l)} + 2e^- \rightarrow Hg_{(l)}$	$+2OH^{-}_{(aq)}$	
At Anodo: $7n \pm 20 H^{-}$ > $7n0 \pm H$	$0 + 2a^{-}$	

(amalgam) (arg) 
$$\rightarrow ZhO_{(s)} + H_2O_{(l)} + 2e$$

**Over all reaction :**  $Zn_{(s)} + HgO_{(s)} \rightarrow ZnO_{(s)} + Hg_{(l)}$ 

(2) **Secondary cells :** In the secondary cells, the reactions can be reversed by an external electrical energy source. Therefore, these cells can be *recharged* by passing electric current and used again and again. These are also celled *storage cells*. Examples of secondary cells are, lead storage battery and nickel – cadmium storage cell.

In charged	Lead storage cell	Alkali cell	
	$\begin{array}{c} + & & & \\ - & & & \\$	$Ni(OH)_2$ + $Fe(OH)_2$ Perforated steel grid + $Li(OH), 1\%$	
Positive electrode	Perforated lead plates coated with <i>PbO</i> <sub>2</sub>	Perforated steel plate coated with $Ni(OH)_4$	
Negative electrode	Perforated lead plates coated with pure lead	Perforated steel plate coated with Fe	
Electrolyte	dil. H <sub>2</sub> SO <sub>4</sub>	20% solution of <i>KOH</i> + 1% <i>LiOH</i>	
During charging	Chemical reaction	Chemical reaction	
Province	At anode : $PbSO_4 + 2H^+ + 2e^- \rightarrow Pb + H_2SO_4$ At cathode : $PbSO_4 + SO_4^- + 2H_2O - 2e^- \rightarrow PbO_2$ + $2H_2SO_4$ Specific gravity of $H_2SO_4$ increases and when specific gravity becomes 1.25 the cell is fully charged. Emf of cell: When cell is fully charged then $E = 2.2$ volt	At anode : $Ni (OH)_2 + 2OH^+ - 2e^- \rightarrow Ni(OH)_4$ At cathode : $Fe(OH)_2 + 2K^+ + 2e^- \rightarrow Fe + 2KOH$ Emf of cell : When cell is fully charged then $E = 1.36$ volt	
During	Chemical reaction $At anoda = Db + SO = -2a^2 + DbSO$	<b>Chemical reaction</b>	
uischargilig	At cathode : $Pb + SO_4 - 2e \rightarrow PDSO_4$ At cathode : $PbO_2 + 2H^+ + 2e^- + H_2SO_4 \rightarrow PbSO_4$ + $2H_2O$ Specific gravity of $H_2SO_4$ decreases and when specific gravity falls below 1.18 the cell	At above : $Fe + 2OH^{-} - 2e^{-} \rightarrow Fe(OH)_{2}$ At cathode : $Ni(OH)_{4} + 2K^{+} + 2e^{-} \rightarrow Ni(OH)_{2} + 2KOH$ Emf of cell : When emf of cell falls below 1.1 V it requires charging.	

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	<i>Emf of cell</i> : When emf of cell falls below 1.9 <i>volt</i> the cell requires recharging.	
Efficiency	80%	60%

# Fuel cells

These are Voltaic cells in which the reactants are continuously supplied to the electrodes. These are designed to convert the energy from the combustion of fuels such as  $H_2$ , CO,  $CH_4$ , etc. directly into electrical energy. The common example is **hydrogen-oxygen fuel cell** as described below,

In this cell, hydrogen and oxygen are bubbled through a porous carbon electrode into concentrated aqueous sodium hydroxide or potassium hydroxide. Hydrogen (the fuel) is fed into the anode compartment where it is oxidised. The oxygen is fed into cathode compartment where it is reduced. The diffusion rates of the gases into the cell are carefully regulated to get maximum efficiency. The net reaction is the same as burning of hydrogen and oxygen to form water. The reactions are

At anode :  $2[H_2(g) + 2OH^-](aq) \longrightarrow 2H_2O(l) + 2e^-$ At cathode :  $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ Overall reaction :  $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$ 

Each electrode is made of porous compressed carbon containing a small amount of catalyst (*Pt*, *Ag* or *CoO*). This cell runs continuously as long as the reactants are fed. Fuel cells convert the energy of the fuel directly into electricity EMF of fuel cell is 1.23 *V*. This cell has been used for electric power in the *Apollo space programmite*. The important advantages of fuel cells are  $Anode \bigcirc \uparrow_1 \oplus Cathode$ 



(1) **High efficiency :** The fuel cells convert the energy of a fuel directly into electricity and therefore, they are more efficient than the conventional methods of generating electricity on a large scale by burning hydrogen, carbon fuels. Though we expect 100 % efficiency in fuel cells, so far **60** – **70% efficiency** has been attained. The conventional methods of production of electrical energy involve combustion of a fuel to liberate heat which is then used to produce electricity. The efficiency of these methods is only about 40%.

(2) **Continuous source of energy :** There is no electrode material to be replaced as in ordinary

battery. The fuel can be fed continuously to produce power. For this reason,  $H_2 - O_2$  fuel cells have been used in space crafts.

(3) **Pollution free working :** There are no objectionable byproducts and, therefore, they do not cause pollution problems. Since fuel cells are efficient and free from pollution, attempts are being made to get better commercially practical fuel cells.

## **Electrode Potential**

(1) When a metal (M) is placed in a solution of its ions  $(M^{++})$ , either of the following three possibilities can occurs, according to the *electrode potential solution pressure theory of Nernst.* 

(i) A metal ion  $M^{n+}$  collides with the electrode, and undergoes no change.

(ii) A metal ion  $M^{n+}$  collides with the electrode, gains *n* electrons and gets converted into a metal atom *M*, (*i.e.* the metal ion is reduced).

# $M^{n+}(aq) + ne^{-} \longrightarrow M(s)$

(iii) A metal atom on the electrode *M* may lose an electrons to the electrode, and enter to the solution as  $M^{n+}$ , (*i.e.* the metal atom is oxidised).  $M(s) \longrightarrow M^{n+}(aq) + ne^{-}$ .

Thus, "the electrode potential is the tendency of an electrode to lose or gain electrons when it is in contact with solution of its own ions."

(2) The magnitude of electrode potential depends on the following factors,

(i) Nature of the electrode, (ii) Concentration of the ions in solution, (iii) Temperature.

(3) **Types of electrode potential :** Depending on the nature of the metal electrode to lose or gain electrons, the electrode potential may be of two types,

(i) **Oxidation potential** : When electrode is negatively charged with respect to solution, *i.e.*, it acts as anode. Oxidation occurs.  $M \longrightarrow M^{n+} + ne^{-}$ 

(ii) **Reduction potential** : When electrode is positively charged with respect to solution, *i.e.* it acts as cathode. Reduction occurs.  $M^{n+} + ne^- \longrightarrow M$ 

(4) **Standard electrode potential :** "If in the half cell, the metal rod (M) is suspended in a solution of one molar concentration, and the temperature is kept at 298 K, the electrode potential is called **standard electrode potential**, represented usually by  $E^{\circ}$ ". 'or'

The standard electrode potential of a metal may be defined as "the potential difference in volts developed in a cell consisting of two electrodes, the pure metal in

contact with a molar solution of one of its ions and the normal hydrogen electrode (NHE)".

 $\label{eq:standard} \begin{array}{l} Standard & oxidation potential & for any half-cell=-(Standard reduction potential) \\ Standard & reduction potential & for any half-cell=-(Standard reduction potential) \end{array}$ 

(5) Reference electrode or reference half - cells

It is not possible to measure the absolute value of the single electrode potential directly. Only the difference in potential between two electrodes can be measured experimentally. It is, therefore, necessary to couple the electrode with another electrode whose potential is known. This electrode is termed as *reference electrode or reference half - cells*. Various types of half – cells have been used to make complete cell with spontaneous reaction in forward direction. These half – cells have been summarised in following table,

Туре	Example	Half – cell reaction	<i>Q</i> =	Reversibl e to	Electrode Potential (oxidn), E =
Gas ion half - cell	$Pt(H_2)  H^+(aq)$ $Pt(Cl_2)  Cl^-(aq)$	$\frac{1}{2}H_2(g) \rightarrow H^+(aq) + e^-$ $Cl^-(aq) \rightarrow \frac{1}{2}Cl_2(g) + e^-$	$[H^+] = \frac{1}{[Cl^-]}$	$H^+$ $Cl^-$	$E^0 - 0.0591\log[H^+]$ $E^0 + 0.0591\log[Cl^-]$
Metal – metal ion half – cell	$Ag \mid Ag^+(aq)$	$Ag(s) \rightarrow Ag^+(aq) + e^-$	$[Ag^+]$	$Ag^+$	$E^0 - 0.0591 \log[Ag^+]$
Metal insoluble salt anion half – cell	$Ag, AgCl \mid Cl^{-}(aq)$	$Ag(s) + Cl^{-}(aq) \rightarrow AgCl(s) + e^{-}$	$\frac{1}{[Cl^-]}$	Cl⁻	$E^0 + 0.0591\log[Cl^-]$
Calomel electrode	$Hg, Hg_2Cl_2 \mid Cl^-(aq)$	$2Hg(l) + 2Cl^{-}(aq) \rightarrow$ $Hg_{2}Cl_{2}(s) + 2e^{-}$	$\frac{1}{\left[Cl^{-}\right]^{2}}$	Cl⁻	$E^0 + 0.0591 \log[Cl^-]$
Metal – metal oxide hydroxide half - cell	$Hg, HgO  OH^{-}(aq)$	$Hg(l) + 2OH^{-}(aq) \rightarrow$ $HgO(s) + H_2O(l) + 2e^{-}$	$\frac{1}{\left[OH^{-}\right]^{2}}$	ОН⁻	$E^0 + 0.0591 \log[OH^-]$
Oxidation – reduction half – dell	$Pt   Fe^{2+}_{(aq)}, Fe^{3+}_{(aq)}$	$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^{-}$	$\frac{[Fe^{3+}]}{[Fe^{2+}]}$	$Fe^{2+}, Fe^{3+}$	$E^0 - 0.0591 \log \frac{[Fe^{3+}]}{[Fe^{2+}]}$

# Table : 12.2 Various Types of Half - cells

# Cell potential or EMF of the cell

(1) "The difference in potentials of the two half – cells of a cell known as **electromotive force** (emf) of the cell or cell potential."

The difference in potentials of the two half – cells of a cell arises due to the flow of electrons from anode to cathode and flow of current from cathode to anode.

Anode 
$$\leftarrow$$
 Flow of current Cathode

(2) The *emf* of the cell or cell potential can be calculated from the values of electrode potentials of two the half – cells constituting the cell. The following three methods are in use :

(i) When oxidation potential of anode and reduction potential of cathode are taken into account

 $E_{\text{cell}}^0$  = Oxidation potential of anode + Reduction

potential of cathode =  $E_{ox}^0$  (anode ) +  $E_{red}^0$  (cathode )

(ii) When reduction potentials of both electrodes are taken into account

 $E_{cell}^0$  = Reduction potential of cathode – Reduction potential of anode

$$= E_{\text{Cathode}}^{0} - E_{\text{Anode}}^{0} = E_{\text{right}}^{0} - E_{\text{left}}^{o}$$

(iii) When oxidation potentials of both electrodes are taken into account

 $E_{cell}^o$  = Oxidation potential of anode – Oxidation potential of cathode =  $E_{ox}^0$  (anode ) –  $E_{ox}^0$  (cathode )

(3) Difference between emf and potential difference

Emf	Potential difference
It is the potential difference between two electrodes when no current is flowing in the circuit.	It is the difference of the electrode potentials of the two electrodes when the cell is under operation.
It is the maximum voltage that the cell can deliver.	It is always less then the maximum value of voltage which the cell can deliver.

It is responsible for the<br/>steady flow of current in<br/>the cell.It is not responsible for<br/>the steady flow of<br/>current in the cell.

(4) Cell EMF and the spontaneity of the reaction : We know,  $\Delta G = -nFE_{cell}$ 

Nature of reaction	$\Delta \mathbf{G}(\mathbf{or} \ \Delta \mathbf{G}^{\mathbf{o}})$	E <sub>cell</sub> (or E <sup>o</sup> <sub>cell</sub> )
Spontaneous	-	+
Equilibrium	0	О
Non – spontaneous	+	-

# Nernst's equation

# (1) Nernst's equation for electrode potential

The potential of the electrode at which the reaction,

 $M^{n+}(aq) + ne^{-} \rightarrow M(s)$ 

takes place is described by the equation,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^{0} - \frac{RT}{nF} \ln \frac{[M(s)]}{[M^{n+}(aq.)]}$$
  
or  $E_{M^{n+}/M} = E_{M^{n+}/M}^{0} - \frac{2.303 \ RT}{nF} \log \frac{[M(s)]}{[M^{n+}(aq)]}$ 

above *eq*. is called the *Nernst equation*.

Where,

 $E_{M^{n+}/M} = {\rm the \ potential \ of \ the \ electrode \ at \ a \ given}$  concentration,

 $E_{M^{n+}/M}^{0}$  = the standard electrode potential

R = the universal gas constant, 8.31  $J K^{-1} mol^{-1}$ 

*T*= the temperature on the absolute scale,

n = the number of electrons involved in the electrode reaction,

F = the Faraday constant : (96500 C),

[M(s)] = the concentration of the deposited metal,

 $[M^{n+}(aq)]$  = the molar concentration of the metal ion in the solution,

The concentration of pure metal M(s) is taken as unity. So, the Nernst equation for the  $M^{n+}/M$ electrode is written as,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^{0} - \frac{2.303 \ RT}{nF} \log \frac{1}{[M^{n+}(aq)]}$$

At 298 K, the Nernst equation for the  $M^{n+}/M$  electrode can be written as,

$$E_{M^{n+}/M} = E_{M^{n+}/M}^0 - \frac{0.0591}{n} \log \frac{1}{[M^{n+}(aq)]}$$

For an electrode (half - cell) corresponding to the electrode reaction,

Oxidised form  $+ne^- \rightarrow$  Reduced form

The Nernst equation for the electrode is written as,

$$E_{half-cell} = E_{half-cell}^0 - \frac{2.303 \ RT}{nF} \log \frac{[\text{Reduced form}]}{[\text{Oxidised form}]}$$

At 298 K, the Nernst equation can be written as,

 $E_{half-cell} = E_{half-cell}^0 - \frac{0.0591}{n} \log \frac{[\text{Reduced form }]}{[\text{Oxidised form]}}$ 

# (2) Nernst's equation for cell EMF

For a cell in which the net cell reaction involving n electrons is,  $aA + bB \rightarrow cC + dD$ 

The Nernst equation is written as,

$$E_{cell} = E_{cell}^{0} - \frac{RT}{nF} \ln \frac{[\mathbf{C}]^{c}[D]^{a}}{[A]^{a}[B]^{b}}$$
  
Where,  $E_{cell}^{0} = E_{cathode}^{0} - E_{anode}^{0}$ .

The  $E_{cell}^{o}$  is called the standard cell potential.

or 
$$E_{\text{cell}} = E_{cell}^o - \frac{2.303 \, RT}{nF} \log \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

At 298 K, above eq. can be written as,

or 
$$E_{\text{cell}} = E_{cell}^o - \frac{0.0592}{n} \log \frac{[C]^c [D]^a}{[A]^a [B]^b}$$

It may be noted here, that the concentrations of *A*, *B*, *C* and *D* referred in the eqs. are the concentrations at the time the cell *emf* is measured.

(3) **Nernst's equation for Daniells cell :** Daniell's cell consists of zinc and copper electrodes. The electrode reactions in Daniell's cell are,

At anode : 
$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$$
  
At cathode :  $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$   
Net cell reaction :  
 $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$ 

Therefore, the Nernst equation for the Daniell's cell is,

$$E_{cdll} = E_{cell}^0 - \frac{2.303 RT}{2F} \log \frac{[Cu(s)][Zn^{2+}(aq)]}{[Zn(s)][Cu^{2+}(aq)]}$$

Since, the activities of pure copper and zinc metals are taken as unity, hence the Nernst equation for the Daniell's cell is,

$$E_{cdll} = E_{cell}^0 - \frac{2.303 \ RT}{2F} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]}$$

The above *eq*. at 298 *K* is,

$$E_{cdll} = E_{cell}^{o} - \frac{0.0591}{2} \log \frac{[Zn^{2+}(aq)]}{[Cu^{2+}(aq)]} V$$

For Daniells cell,  $E_{cell}^0 = 1.1 V$ 

### (4) Nernst's equation and equilibrium constant

For a cell, in which the net cell reaction involving *n* electrons is,  $aA + bB \rightarrow cC + dD$ 

The Nernst equation is

$$E_{Cell} = E_{cell}^{0} - \frac{RT}{nF} \ln \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}} \qquad \dots \dots (i)$$

At equilibrium, the cell cannot perform any useful work. So at equilibrium,  $E_{Cell}$  is zero. Also at equilibrium, the ratio

$$\frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}} = \left\lfloor \frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}} \right\rfloor_{eauilibrium} = K_{c}$$

# Relationship between potential, Gibbs energy and equilibrium constant

The electrical work (electrical energy) is equal to the product of the *EMF* of the cell and electrical charge that flows through the external circuit *i.e.*,

 $W_{\max} = nFE_{cell} \qquad \dots \dots (i)$ 

According to thermodynamics the free energy change ( $\Delta G$ ) is equal to the maximum work. In the cell work is done on the surroundings by which electrical energy flows through the external circuit, So

 $-W_{\rm max,} = \Delta G$ 

.....(ii)

from eq. (i) and (ii)  $\Delta G = -nFE_{cell}$ 

In standard conditions  $\Delta G^0 = -nFE_{cell}^0$ 

Where  $\Delta G^0$  = standard free energy change

But 
$$E_{cell}^0 = \frac{2.303}{nF} RT \log K_c$$
  
 $\therefore \Delta G^0 = -nF \times \frac{2.303}{nF} RT \log K_c$   
 $\Delta G^0 = -2.303 \text{ RT} \log K_c \text{ or } \Delta G = \Delta G^\circ + 2.303 RT \log Q$   
 $\Delta G^0 = -RT \ln K_c$  (2.303 log  $X = \ln X$ )

**Electrochemical series** 

(1) The standard reduction potentials of a large number of electrodes have been measured using standard hydrogen electrode as the reference electrode. These various electrodes can be arranged in increasing or decreasing order of their reduction potentials. The arrangement of elements in order of increasing reduction potential values is called *electrochemical series*. It is also called *activity series*, of some typical electrodes.

# (2) Characteristics of Electrochemical series

(i) The negative sign of standard reduction potential indicates that an electrode when joined with *SHE* acts as anode and oxidation occurs on this electrode. For example, standard reduction potential of zinc is -0.76 volt, When zinc electrode is joined with SHE, it acts as anode (*-ve* electrode) *i.e.*, oxidation occurs on this electrode. Similarly, the *+ve* sign of standard reduction potential indicates that the electrode when joined with *SHE* acts as cathode and reduction occurs on this electrode.

(ii) The substances, which are stronger reducing agents than hydrogen are placed above hydrogen in the series and have negative values of standard reduction potentials. All those substances which have positive values of reduction potentials and placed below hydrogen in the series are weaker reducing agents than hydrogen.

(iii) The substances, which are stronger oxidising agents than  $H^+$  ion are placed below hydrogen in the series.

(iv) The metals on the top (having high negative value of standard reduction potentials) have the tendency to lose electrons readily. These are active metals. The activity of metals decreases from top to bottom. The non-metals on the bottom (having high positive values of standard reduction potentials) have the tendency to accept electrons readily. These are active non-metals. The activity of non-metals increases from top to bottom.

Table : 12.3 Standard	reduction	electrode po	otentials at
298K			

Elen	nent	Electrode Reac (Reduction)	tion )	Standard Electrode Reduction potential E <sup>0</sup> , volt
Li	N /	$Li^+ + e^- = Li$	Δ	-3.05
Κ	$ \backslash / $	$K^+ + e^- = K$		-2.925
Ва		$Ba^{++}+2e = Ba$		-2.90
Sr		$Sr^{++}+2e=Sr$		-2.89
Са	50	$Ca^{2+} + 2e^- = Ca$	4	-2.87
Na	asin	$Na^+ + e^- = Na$	l B	-2.714
Mg	incre	$Mg^{2+} + 2e^{-} = Mg$	easir	-2.37
Al	ons l ent	$Al^{3+} + 3e^{-} = Al$	Incr	-1.66
	tendency to accept electr strength as oxidising age		rength as reducing agent	ndency to lose electrons

Mn	$Mn^{++}+2e = Mn$	-1.18
Zn	$Zn^{2+}+2e^{-}=Zn$	-0.7628
Cr	$Cr^{3+}+3 e^{-} = Cr$	-0.74
Fe	$Fe^{2+}+2e^{-}=Fe$	-0.44
Cd	$Cd^{2+}+2e^{-}=Cd$	-0.403
Со	$Co^{++} + 2e = Co$	-0.27
Ni	$Ni^{2+}+2e^- = Ni$	-0.25
Sn	$Sn^{2+}+2e^{-}=Sn$	-0.14
Pb	$Pb^{++}+2e = Pb$	-0.12
$H_2$	$2H^++2e^- = H_2$	0.00
Си	$Cu^{2+} + 2e^- = Cu$	+0.337
$I_2$	$I_2 + 2e^- = 2I^-$	+0.535
Hg	$Hg^{2+}+2e^{-}=Hg$	+0.885
Ag	$Ag^+ + e^- = Ag$	+0.799
$Br_2$	$Br_2+2e^-=2Br^-$	+1.08
Pt	$Pt^{++}+2e = Pt$	+1.20
$Cl_2$	$Cl_2+2e^- = 2Cl^-$	+1.36
Au	$Au^{3+}+3e^{-}=Au$	+1.50
$F_{2}$	$F_2 + 2e^- = 2F^-$	+2.87

# (3) Application of Electrochemical series

(i) **Reactivity of metals**: The activity of the metal depends on its tendency to lose electron or electrons, *i.e.*, tendency to form cation  $(M^{n+})$ . This tendency depends on the magnitude of standard reduction potential. The metal which has high negative value (or smaller positive value) of standard reduction potential readily loses the electron or electrons and is converted into cation. Such a metal is said to be chemically active. The *chemical reactivity of metals decreases from top to bottom in the series*. The metal higher in the series is more active than the metal lower in the series. For example,

(a) Alkali metals and alkaline earth metals having high negative values of standard reduction potentials are chemically active. These react with cold water and evolve hydrogen. These readily dissolve in acids forming corresponding salts and combine with those substances which accept electrons.

(b) Metals like *Fe*, *Pb*, *Sn*, *Ni*, Co, etc., which lie a little down in the series do not react with cold water but react with steam to evolve hydrogen.

(c) Metals like *Cu*, *Ag* and *Au* which lie below hydrogen are less reactive and do not evolve hydrogen from water.

(ii) *Electropositive character of metals* : The electropositive character also depends on the tendency to lose electron or electrons. Like reactivity, the electropositive character of metals decreases from top to bottom in the electrochemical series. On the basis of standard reduction potential values, metals are divided into three groups

(a) *Strongly electropositive metals* : Metals having standard reduction potential near about – 2.0 *volt* or more

negative like alkali metals, alkaline earth metals are strongly electropositive in nature.

(b) *Moderately electropositive metals* : Metals having values of reduction potentials between 0.0 and about – 2.0 *volt* are moderately electropositive *Al*, *Zn*, *Fe*, *Ni*, *Co*, etc., belong to this group.

(c) *Weakly electropositive* : The metals which are below hydrogen and possess positive values of reduction potentials are weakly electropositive metals. *Cu, Hg, Ag,* etc., belong to this group.

# (iii) Displacement reactions

(a) To predict whether a given metal will displace another, from its salt solution: A metal higher in the series will displace the metal from its solution which is lower in the series, i.e., The metal having low standard reduction potential will displace the metal from its salt's solution which has higher value of standard reduction potential. A metal higher in the series has greater tendency to provide electrons to the cations of the metal to be precipitated.

(b) Displacement of one nonmetal from its salt solution by another nonmetal: A non-metal higher in the series (towards bottom side), *i.e.*, having high value of reduction potential will displace another non-metal with lower reduction potential, *i.e.*, occupying position above in the series. The non-metal's which possess high positive reduction potentials have the tendency to accept electrons readily. These electrons are provided by the ions of the nonmetal having low value of reduction potential,. Thus,  $Cl_2$  can displace bromine and iodine from bromides and iodides.

$Cl_2 + 2KI \rightarrow 2KCl + I_2$	
$2I^- \rightarrow I_2 + 2e^-$	(Oxidation)
$Cl_2 + 2e^- \rightarrow 2Cl^-$	(Reduction)

[The activity or electronegative character or oxidising nature of the nonmetal increases as the value of reduction potential increases.]

(c) Displacement of hydrogen from dilute acids by metals : The metal which can provide electrons to  $H^+$  ions present in dilute acids for reduction, evolve hydrogen from dilute acids.

$$Mn \rightarrow Mn^{n+} + ne^{-} \qquad \dots (Oxidation)$$
$$2H^{+} + 2e^{-} \rightarrow H_{2}$$
$$\dots (Reduction)$$

The metal having negative values of reduction potential possess the property of losing electron or electrons.

Thus, the metals occupying top positions in the electrochemical series readily liberate hydrogen from dilute acids and on descending in the series tendency to liberate hydrogen gas from dilute acids decreases.

The metals which are below hydrogen in electrochemical series like *Cu, Hg, Au, Pt,* etc., do not evolve hydrogen from dilute acids.

(d) *Displacement of hydrogen from water* : Iron and the metals above iron are capable of liberating hydrogen from water. The tendency decreases from top to bottom in electrochemical series. Alkali and alkaline earth metals liberate hydrogen from cold water but *Mg*, *Zn* and *Fe* liberate hydrogen from hot water or steam.

(iv) **Reducing power of metals**: Reducing nature depends on the tendency of losing electron or electrons. More the negative reduction potential, more is the tendency to lose electron or electrons. Thus reducing nature decreases from top to bottom in the electrochemical series. The power of the reducing agent increases, as the standard reduction potential becomes more and more negative. Sodium is a stronger reducing agent than zinc and zinc is a stronger reducing agent than iron. (decreasing order of reducing nature)

Element :			Na	>	Zn	>	Fe
Reduction	potential	:	- 2.71		-0.76		-0.44

Alkali and alkaline earth metals are strong reducing agents.

(v) **Oxidising nature of non-metals** : Oxidising nature depends on the tendency to accept electron or electrons. More the value of reduction potential, higher is the tendency to accept electron or electrons. Thus, oxidising nature increases from top to bottom in the electrochemical series. The strength of an oxidising agent increases as the value of reduction potential becomes more and more positive.

 $F_2$  (Fluorine) is a stronger oxidant than  $Cl_2, Br_2$ and  $I_2, Cl_2$  (Chlorine) is a stronger oxidant than  $Br_2$  and  $I_2$ Element:

#### **Reduction potential :**

 $\begin{array}{cccc} I_2 & Br_2 & Cl_2 & F_2 \\ + 0.53 + 1.06 + 1.36 + 2.85 \end{array}$ 

Oxidising nature increases

# Thus, in electrochemical series



Botto Highest positive value of reduction potential

(vi) *Thermal stability of metallic oxides* : The thermal stability of the metal oxide depends on its electropositive nature. As the electropositivity

decreases from top to bottom, the thermal stability of the oxide also decreases from top to bottom. The oxides of metals having high positive reduction potentials are not stable towards heat. The metals which come below copper form unstable oxides, i.e., these are decomposed on heating.

$$Ag_{2}O \xrightarrow{\Delta} 2Ag + \frac{1}{2}O_{2}$$

$$2HgO \xrightarrow{\Delta} 2Hg + O_{2} ; \begin{bmatrix} BaO \\ Na_{2}O \\ Al_{2}O_{3} \end{bmatrix} \xrightarrow{\Delta} \text{No decomposit ion}$$

(vii) **Extraction of metals** : A more electropositive metal can displace a less electropositive metal from its salt's solution. This principle is applied for the extraction of Ag and Au by cyanide process. silver from the solution containing sodium argento cyanide,  $NaAg(CN)_2$ , can be obtained by the addition of zinc as it is more electro-positive than Ag.

$$2NaAg(CN)_2 + Zn \rightarrow Na_2Zn(CN)_4 + 2Ag$$

#### Corrosion

(1) When metals are exposed to atmospheric conditions, they react with air or water in the environment to form undesirable compounds (usually oxides). This process is called **corrosion**. Almost all metals except the least active metals such as gold, platinum and palladium are attacked by environment *i.e.*, undergo corrosion. For example, silver tarnishes, copper develops a green coating, lead or stainless steel lose their lusture due to corrosion. Corrosion causes enormous damage to building, bridges, ships and many other articles made of iron.

Thus **corrosion** is a process of deterioration of a metal as a result of its reaction with air or water (environment) surrounding it.

In case of iron, corrosion is called rusting. Chemically, rust is hydrated form of ferric oxide,  $Fe_2O_3$ .  $xH_2O$ . Rusting of iron is generally caused by moisture, carbon dioxide and oxygen present in air. It has been observed that rusting takes place only when iron is in contact with moist air. Iron does not rust in dry air and in vacuum.

(2) **Factors which affect corrosion** : The main factors which affect corrosion are

More the reactivity of metal, the more will be the possibility of the metal getting corroded.

The impurities help in setting up voltaic cells, which increase the speed of corrosion

Presence of electrolytes in water also increases the rate of corrosion

Presence of  $CO_2$  in natural water increase rusting of iron.

(v) When the iron surface is coated with layers of metals more active than iron, then the rate of corrosion is retarded.

A rise in temperature (with in a reasonable limit) increases the rate of corrosion.

(3) **Classification of corrosion process** : Depending upon the nature of corrosion, and the factors affecting it, the corrosion may be classified as follows.

(i) *Chemical corrosion* : Such corrosion, generally takes place when

(a) Reactive gases come in contact with metals at high temperatures *e.g.*, corrosion in chemical industry.

(b) Slow dissolution of metal takes place when kept in contact with non conducting media containing organic acids.

(ii) **Bio-chemical corrosion or Bio-corrosion**: This is caused by the action of microorganisms. Soils of definite composition, stagnant water and certain organic products greatly favour the bio-corrosion.

(iii) *Electrochemical corrosion* : It occurs in a gaseous atmosphere in the presence of moisture, in soils and in solutions.

(4) Mechanism of rusting of iron : *Electrochemical theory of rusting*.

The overall rusting involves the following steps,

(i) Oxidation occurs at the anodes of each electrochemical cell. Therefore, at each anode neutral iron atoms are oxidised to ferrous ions.

At anode :  $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^{-}$ .

Thus, the metal atoms in the lattice pass into the solution as ions, leaving electrons on the metal itself. These electrons move towards the cathode region through the metal.



(ii) At the cathodes of each cell, the electrons are taken up by hydrogen ions (reduction takes place). The  $H^+$  ions are obtained either from water or from acidic substances (e.g.  $CO_2$ ) in water

$$H_2O \longrightarrow H^+ + OH^-$$
 or  $CO_2 + H_2O \longrightarrow H^+ + HCO_3^-$ 

At cathode :  $H^+ + e^- \longrightarrow H$ 

The hydrogen atoms on the iron surface reduce dissolved oxygen.  $4H + O_2 \longrightarrow 2H_2O$ 

Therefore, the overall reaction at cathode of different electrochemical cells may be written as,

 $4H^+ + O_2 + 4e^- \longrightarrow 2H_2O$ 

(iii) The overall redox reaction may be written by multiplying reaction at anode by 2 and adding reaction at cathode to equalise number of electrons lost and gained i.e.

**Oxi.** half reaction :  $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}] \times 2$ 

 $(E = -0.44\,V)$ 

**Red. half reaction :**  $4H^+ + O_2 + 4e^- \rightarrow 2H_2O$ (E = 1.23V)

**Overall cell reaction :**  $2Fe(s) + 4H^+ + O_2 \rightarrow 2Fe^{2+}(aq) + 2H_2O$ 

$$(E_{\text{Cell}} = 1.67 V)$$

The ferrous ions are oxidised further by atmospheric oxygen to form rust.

$$4Fe^{2+}(aq) + O_2(g) + 4H_2O \longrightarrow 2Fe_2O_3 + 8H^+ \text{ and}$$

$$Fe_2O_3 + xH_2O \longrightarrow Fe_2O_3 \cdot xH_2O$$
Rust
Rust

It may be noted that salt water accelerates corrosion. This is mainly due to the fact that salt water increases the electrical conduction of electrolyte solution formed on the metal surface. Therefore, rusting becomes more serious problem where salt water is present.

(5) **Corrosion protection** : Corrosion of metals can be prevented in many ways. Some commonly used methods are

# (i) By surface coating

(a) By applying, oil, grease, paint or varnish on the surface.

(b) By coating/depositing a thin layer of any other metal which does not corrode. For example, iron surface can be protected from corrosion by depositing a thin layer of zinc, nickel or chromium on it. Copper/brass can be protected by coating it with a thin layer of tin. Tinning of brass utensils is a very common practice in our country.

(c) *By Galvanization* : Prevention of corrosion of iron by *Zn* coating.

(ii) **By connecting metal to a more** *electropositive metal* : As long as the more electropositive metal is there, the given metal does not get corroded. For example, iron can be protected from corrosion by connecting it to a block/plate of zinc or

magnesium. This method of corrosion protection is called *cathodic protection*.

(iii) **By forming insoluble phosphate or chromate coating**: Metal surfaces are treated with phosphoric acid to form an insoluble phosphate. Formation of a thin chromate layer also prevents the corrosion of metals.

(iv) **Using anti – rust solutions** : Solutions of alkaline phosphates and alkaline chromates are generally used as anti – rust solutions. For example, iron articles are dipped in boiling alkaline sodium phosphate solutions, when a protective insoluble sticking film of iron phosphate is formed.



- When two or more ions compete at the electrodes then the ion with higher reduction potential gets liberated at the cathode while the one with lower reduction potential at the anode.
- Cell constant is determined with the help of conductivity bridge, where a standard solution of KCl is used.
- ✓ If the external *EMF* is slightly more than the actual *EMF*, the current will flow into the cell and reverse reaction takes place.
- Identification of cathode and anode is done by the use of galvanometer.
- *KCl* / *NaCl* / *NH*<sub>4</sub>*Cl* etc., can not be used in the salt bridge of a cell containing silver salt as one of the electrodes as  $Cl^-$  ions form a ppt. of *AgCl* with silver ion.
- Weston cell is a common example of standard cell. The *emf* of a standard cell does not change with temperature.
- **\not** In Appolo moon flights,  $H_2 O_2$  fuel cell was the source of energy and drinking water.
- Conductivity water is the highly purified water whose on conductance is very small. It is prepared by the demineralisation of ordinary water by passing through cation and anion exchange resins.



- conduct an electric current quite well[MP PMT 1987](a) Glycerol(b) HCl
  - (c) Sugar (d) Pure water

- **9.** On the electrolysis of aqueous solution of sodium sulphate, on cathode we get
  - (a) Na (b)  $H_2$
  - (c)  $SO_2$  (d)  $SO_3$
- **10.** Electrolysis involves oxidation and reduction respectively at
  - [CPMT 1973; AMU 1983; NCERT 1983, 84; MH CET 2001] (a) Anode and cathode (b) Cathode and anode
  - (c) At both the electrodes (d) None of the above
- **11.** Which of the following compounds will not undergo decomposition on passing electricity through aqueous solution
  - (a) Sugar (b) Sodium Chloride
  - (c) Sodium Bromide (d) Sodium Acetate
- During the electrolysis of an electrolyte, the number of ions produced, is directly proportional to the [AFMC 2002]
  - (a) Time consumed
  - (b) Electro chemical equivalent of electrolysis
  - (c) Quantity of electricity passed
  - (d) Mass of electrons
- **13.** When the sample of copper with zinc impurity is to be purified by electrolysis, the appropriate electrodes are

# [AIEEE 2002]

Cathode	Anode
(a) Pure zinc	Pure copper
(b) Impure sample	Pure copper
(c) Impure zinc	Impure sample
(d) Pure copper	Impure sample

**14.** In the electrolytic cell, flow of electrons is from

# [IIT Screening 2003]

- (a) Cathode to anode in solution
- (b) Cathode to anode through external supply
- (c) Cathode to anode through internal supply
- (d) Anode to cathode through internal supply
- An electric current is passed through an aqueous solution of the following. Which one shall decompose [NCERT 1972]
  - (a) Urea (b) Glucose
  - (c)  $AgNO_3$  (d) Ethyl alcohol
- **16.** The electric conduction of a salt solution in water depends on the
  - (a) Shape of its molecules
  - (b) Size of its molecules
  - (c) Size of solvent molecules
  - (d) Extent of its ionization
- **17.** A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively[**IIT 1987,96**]
  - (a)  $H_2$ ,  $O_2$  (b)  $O_2$ ,  $H_2$
  - (c)  $O_2$ , Na (d)  $O_2$ , SO 2

18.	On electrolysing a solution of dilute $H_2SO_4$		(d) Is readily decomposed		
	between platinum electrodes, the gas evolved at the anode is	27.	Which is responsible for electrical conduction of molten sodium chloride[MADT Bihar 1995]		
	[NCERT 1977, 79; MNR 1980; CBSE PMT 1992]		(a) Free electrons		
	(a) $SO_2$ (b) $SO_3$		(b) Free ions		
	(c) $O_2$ (d) $H_2$		(c) Free molecules		
19.	The addition of a polar solvent to a solid		(d) Atoms of sodium and chlorine		
-91	electrolyte results in	28 <sub>-[1</sub>	NCERT 1973] ysis of aqueous copper sulphate, the gas at anode and cathode is [AFMC 1995]		
	(a) Polarization (b) Association		(a) $O_2$ and $H_2$ (b) $SO_2$ and $H_2$		
heat	(c) Ionization (d) Non-Ilberation of		(c) $H_2$ and $O_2$ (d) $SO_3$ and $O_2$		
20.	During the electrolysis of fused <i>NaCl</i> which	29.	Use of electrolysis is [AFMC 1995]		
-0.	reaction occurs at anode[NCERT 1973: AFMC 1992: MP	-5. PMT 20	<b>o6a)</b> Electroplating (b) Electrorefining		
	(a) Chloride ions are oxidized		(c) (a) and (b) both (d) None of these		
	(b) Chloride ions are reduced	30.	Sodium is made by the electrolysis of a molten		
	(c) Sodium ions are oxidised	-	mixture of about 40% NaCl and 60% CaCl <sub>2</sub>		
	(d) Sodium ions are reduced		because		
21.	The amount of ion discharged during electrolysis		[CBSE PMT 1995]		
	is not directly proportional to [NCERT 1973]		(a) $CaCl_2$ helps in conduction of electricity		
	(a) Resistance		(b) This mixture has a lower melting point than		
	(b) Time	NaC	71		
	(c) Current		(c) $Ca^{++}$ can displace Na from NaCl		
	(d) Chemical equivalent of the ion		(d) $Ca^{++}$ can reduce NaCl to Na		
22.	Electrolysis of aqueous <i>HCl</i> solution produces		Electrolysis is a process in which the cations and		
	[CPMT 1987]	51.	anions of the electrolyte are		
	(a) $H_2$ gas at the anode		(a) Hydrated (b) Hydrolysed		
	(b) $H_2$ gas at the cathode		(c) Charged (d) Discharged		
	(c) $Cl_2$ gas at the cathode	32.	Degree of ionisation of a solution depends upon		
	(d) $Cl_2$ and $O_2$ gases both at the anode		[BHU 1998]		
23.	During electrolysis of <i>NaCl</i> solution, part of the		(a) Temperature (b) Nature of the		
-0.	reaction is $Na^+ + e^- \rightarrow Na$ . This is termed as[NCERT 19	elect	trolyte		
	(a) Oxidation (b) Reduction		(c) Nature of the solvent(d) None of these		
	(c) Deposition (d) Cathode reaction	33.	Which of the following is non-electrolytes		
24.	When a solution of an electrolyte is heated the		[KCET (Med.) 1999]		
	conductance of the solution	[]	ксет 1991] (b) CaCl <sub>2</sub>		
	(a) Increases because of the electrolyte conducts		(c) $C_{12}H_{22}O_{11}$ (d) $CH_3COOH$		
bette	er	34.	When a molten ionic hydride is electrolysed		
	(b) Decreases because of the increased heat		[JIPMER 1999]		
	(c) Decreases because of the dissociation of the		(a) Hydrogen is liberated at the cathode		
	(d) Increases because the electrolyte is		(b) Hydrogen is liberated at the anode		
disso	ciated more		(c) There is no reaction		
25.	The passage of current liberates $H_2$ at cathode		(d) $H^-$ ions produced migrate to the cathode		
	and $Cl_2$ at anode. The solution is [EAMCET 1979,87]	35.	During electrolysis, the species discharged at cathode are		
	(a) Copper chloride in water (b) <i>NaCl</i> in water		[AFMC 2000]		
	(c) $H_2SO_4$ (d) Water		(a) Ions (b) Cation		
26.	Pure water does not conduct electricity because it		(c) Anion (d) All of these		
	[Manipal MEE 1995]	36.	Electrolysis of molten anhydrous calcium chloride		
	(a) Has a low boiling point	•••	produces		
	(b) Is almost totally unionized		[AIIMS 2000]		

(c) Is neutral

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2 <b>T 1980</b> gen be the <b>1973, 7</b> 7
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(a) 40.65 *gm* 

(c) 0.4065 gm

(b) 4.065 gm

(d) 65.04 gm

**4.** A silver cup is plated with silver by passing 965 coulombs of electricity, the amount of silver deposited is

(d)  $E = \frac{C \times t \times 96500}{C \times t \times 96500}$ In an electroplating experiment m g of silver is 15. 96500 × m (c) *E* =  $C \times t$ deposited, when 4 amperes of current flows for 2 m minutes. The amount (in gms) of silver deposited 23. How many Faradays are required to generate one gram atom of magnesium from *MgCl*<sub>2</sub> [MADT Bihar 1982] by 6 amperes of current flowing for 40 seconds will be [MNR 1991] (a) 1 (b) 2 (a) 4m (b) *m*/2 (c) 3 (d) 4 (d) 2m (c) m/4To deposit 0.6354 qm of copper by electrolysis of 24. On passing 3 *ampere* of electricity for 50 minutes, aqueous cupric sulphate solution, the amount of 16. 1.8 gram metal deposits. The equivalent mass of electricity required (in coulombs) is metal is (a) 9650 (b) 4825 [MP PMT 1992] (c) 3860 (d) 1930 (a) 20.5 (b) 25.8 In electrolysis of a fused salt, the weight of the 25. (c) 19.3 (d) 30.7 deposit on an electrode will not depend on[CPMT 1973] The desired amount of charge for obtaining one 17. (a) Temperature of the bath mole of Al from  $Al^{3+}$ (b) Current intensity (a)  $3 \times 96500 C$ (b) 96500 C (c) Electrochemical equivalent of ions (d) Time for electrolysis (c)  $\frac{96500}{C}$ (d)  $\frac{96500}{2}C$ 26. Faraday's laws of electrolysis will fail when[NCERT 1971] (a) Temperature is increased On passing one faraday of electricity through the 18. electrolytic cells containing  $Ag^+$ ,  $Ni^{+2}$  and  $Cr^{+3}$  ions (b) Inert electrodes are used (c) A mixture of electrolytes is used solution. the deposited Ag(At. wt. = 108), Ni(At. wt. = 59) and Cr(At. wt. = 52) is[AIIMS 1982] In none of these cases According to the first law of Faraday, the weight 27. Ag Ni Crof a substance discharge at the electrode is (a) 108 gm 29.5 qm 17.3 gm (a) W = ZQ(b) W = eF(b) 108 *qm* 59.0 qm 52.0 qm (c)  $W = \frac{Z}{E} It$ (d) W = ZI(c) 108.0 *qm* 108.0 *qm* 108.0 gm (d) 108 gm 117.5 qm 166.0 gm When 0.04 faraday of electricity is passed 28. One Faraday of electricity when passed through a through a solution of  $CaSO_4$ , then the weight of solution of copper sulphate deposits [CPMT 1978]  $Ca^{2+}$  metal deposited at the cathode is [BHU 1996] (a) 1 mole of Cu (b) 1 qm atom of Cu (a) 0.2 *qm* (b) 0.4 qm (c) 1 molecule of Cu (d) 1 qm equivalent of (c) 0.6 qm (d) 0.8 qm A current 2.0 A is passed for 5 hours through a 29. When 1 coulomb of charge is passed through molten metal salt deposits 22 g of metal (At. wt. electrolyte solution, then the mass deposited is =177). The oxidation state of the metal in the equal to metal salt is [KCET 1996] (a) Equivalent weight (a) + 1(b) + 2(b) Atomic weight (c) + 3(d) + 4(c) Electrochemical equivalent 30. How many atoms of calcium will be deposited from a solution of  $CaCl_2$  by a current of 25 (d) Chemical equivalent milliamperes flowing for 60 seconds The platinum electrodes were immersed in a 21. solution of cupric sulphate and electric current (a)  $4.68 \times 10^{18}$ (b)  $4.68 \times 10^{15}$ passed through the solution. After some time it

> 31. On passing 0.5 faraday of electricity through *NaCl* , the amount of *Cl* deposited on cathode is

# [BHU 1997; RPET 1999]

(a) 35.5 <i>gm</i>	(b) 17.75 gm
(c) 71 gm	(d) 142 gm

32. What is the amount of chlorine evolved when 2 amperes of current is passed for 30 minutes in an aqueous solution of *NaCl* [BHU 1998; AIIMS 1999] (a) 66 a(h) 122a

(u) 00 g	(0) 1.32 9
(c) 33 g	(d) 99 g

disappeared with evolution of gas at the electrode. The colourless solution contains [NCERT 1984] (b) Copper hydroxide

(c) Copper sulphate (d) Sulphuric acid

(a) Platinum sulphate

was found that colour of copper sulphate

On passing *C* ampere of electricity through a 22. electrolyte solution for t second, m gram metal deposits on cathode. The equivalent weight E of the metal is [MP PMT 1990]

(a) 
$$E = \frac{C \times t}{m \times 96500}$$
 (b)  $E = \frac{C \times m}{t \times 96500}$ 

- Си
- 20.

- 19.

- (c)  $4.68 \times 10^{12}$

- (d)  $4.68 \times 10^9$

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33.	On passing a current th	rough KCl solution, $19.5 g$		(a) 0.66 <i>mole</i>	(b) 0.33 <i>mole</i>	
	of potassium is deposit	ed. If the same quantity of		(c) 0.66 gm	(d) 0.33 gm	
	electricity is passed	through a solution of	42.	The number of elec	ctrons involved in redox	
	aluminium chloride, t	he amount of aluminium		reactions when a Fara	day of electricity is passed	
	deposited is	[EAMCET 1997]		through an electrolyte	in solution is[NCERT 1982; Pb. PM	Ĩ
	(a) 4.5 g	(b) 9.0 <i>g</i>		(a) $6 \times 10^{23}$	(b) $6 \times 10^{-23}$	
	(c) 13.5 g	(d) 27 g		(c) 96500	(d) $8 \times 10^{19}$	
	(e) None is correct		43.	Coulomb is equal to	[Orissa JEE 2002]	
34.	Electrolysis rules of F	araday's states that mass	_	(a) ampere × second	(b) ampere × minute	
	deposited on electrode	is proportional to [CBSE PMT	2000]	(c) watt $\times$ second	(d) volt $\times$ second	
	(a) $m \propto I^2$	(b) $m \propto Q$	44.	The energy required	to release 1 electron from	
	(c) $m \propto Q^2$	(d) None of these		<i>He</i> <sup>+</sup> <b>is</b>		
35.	A current being passe	d for two hour through a			[Orissa JEE 2002]	
00	solution of an acid libe	erating 11.2 <i>litre</i> of oxygen		(a) + 54.4 <i>eV</i>	(b) – 13.6 <i>eV</i>	
	at NTP at anode. Wh	at will be the amount of		(c) + 27.2 <i>eV</i>	(d) Cannot be predicted	
	copper deposited at	the cathode by the same	45.	Faraday's laws of elect	rolysis are related to the[IIT 198;	3]
	current when passed th	rough a solution of copper		(a) Atomic number of a	cation	
	sulphate for the same t	Ime [BVP 2003]		(b) Atomic number of a	anion	
	(a) 16 g	(b) 63 g		(c) Equivalent weight	of the electrolyte	
- 6	(c) 31.5 g	(d) 8 g		(d) Speed of the cation		
36.	In a metal oxide, there	is 20% oxygen by weight.	46.	The electric charge for	electrode decomposition of	
	(a) 40	(b) $64$		one gram equivalent of	a substance is[IIT 1984; KCET 19	<b>}</b> 2
	(a) $40$	(d) 22		(a) One ampere per sec	cond	
27	(c) / 2 On the basis of the info	(u) 52		(b) 96500 coulombs pe	er second	
3/•				(c) One ampere for one	e nour	
	reaction $\frac{-Al}{3} + O_2 \rightarrow \frac{-Al}{3}$	$_{3}^{-}Al_{2}O_{3},\Delta G = -827  kJmol^{-1}$ of	4-	(d) Charge on one mole	e of electrons	
	$O_2$ , the minimum <i>emf</i>	required to carry out an	47.	through a cross-section	on of copper wire carrying	
	electrolysis of $Al_{2}O_{2}$ is	$(F = 96500C \ mol^{-1})$ [CBSE PM]	T 2003	1 10 <sup>-6</sup> amperes of curren	t per second is found to be <b>EAM</b>	E
	(a) 8 = 6 V	(b) 2 14 V		(a) $1.6 \times 10^{-19}$	(b) $6 \times 10^{-35}$	-
	(a) $0.50 V$	(d) $6.42 V$		(a) $6 \times 10^{-16}$	(d) $6 \times 10^{12}$	
28	Then during electrolys	is of a solution of $AaNO_{2}$	40	(c) $0 \times 10$	$(0) 0 \times 10$	
50.	9650 coulombs of c	harge pass through the	40.	ferrous chloride and	another acidified ferric	
	electroplating bath, th	e mass of silver deposited		chloride are connected	in series. The ratio of iron	
	in the cathode will be	_		deposited at cathode	s in the two cells when	
		[AIEEE 2003]		electricity is passed the	rough the cells will be	
	(a) 1.08 g	(b) 10.8 <i>g</i>			[CPMT 1989]	
	(c) 21.6 <i>g</i>	(d) 108 g		(a) 3 : 1	(b) 2 : 1	
39.	Total charge on 1 mole of	of a monovalent metal ion is		(c) 1:1	(d) 3 : 2	
	equal to		<b>49</b> .	When 96500 coulom	b of electricity is passed	
		[DPMT 2001]		through a copper sulph	nate solution, the amount of	
	(a) $9.65 \times 10^4$ <i>Coulomb</i>	(b) $6.28 \times 10^{18}$ Coulomb		copper deposited will t		
	(c) $1.6 \times 10^{-19}$ Coulomb	(d) None of these		(a) 0.25 mol	(b) 0.50 <i>mol</i>	
40	When an electric of	wront is passed through		(c) 1.00 mol	(d) 2.00 mol	
40.	acidulated water 112 m	d of hydrogen gas at N T P	50.	During electrolysis of	fused aluminium chloride	
	collect at the cathode i	n 965 <i>seconds</i> . The current			i was deposited on the	
	passed, in amperes is			catnode. The volume (	or chlorine liberated at the	
	-	[MNR 1991; UPSEAT 2001]		(a) a a k itrac	(b) 11 2 litrae	
	(a) 1.0	(b) 0.5		(a) $2.24$ infes	(b) 11.2 $uures$	
	(c) 0.1	(d) 2.0		(c) 1.12 <i>litres</i>	(a) 5.6 litres	
41.	How much chlorine w	ill be liberated on passing	51.	raraday has the dimen	SIGHS OF [MP PET 1995]	
	one ampere current for	r 30 minutes through NaCl		(a) Coulombs		
	solution	[BVP 2003]		(b) Coulomb equivalen	t	

	<ul><li>(c) Coulomb per equival</li><li>(d) Coulomb per degree</li></ul>	ent Kelvin
52.	The required charge for silver deposite on cathoo	one equivalent weight of de is
	(a) $9.65 \times 10^7 C$	(b) $9.65 \times 10^4 C$
	(c) $9.65 \times 10^3 C$	(d) $9.65 \times 10^5 C$
53.	96500 <i>coulombs</i> of elect $CuSO_4$ solution	tric current liberates from [MP PMT 1995]
	(a) 63.5 gm Cu	(b) 31.75 gm Cu
	(c) 96500 gm Cu	(d) 100 gm Cu
54.	A current of strength 2.5 $CuSO_4$ solution for 6	5 <i>amp</i> was passed through minutes 26 seconds. The
	amount of copper depos	ited is
	(Atomic weight of $Cu = 6$	53.5)
	(1  faraday = 96500  coul)	ombs)
	(a) $0.2175 a$	(b) $2.175 a$
	(a) $0.31/5 g$ (c) $0.635 a$	(d) $6.35 a$
55.	A certain quantity of ele	ectricity is passed through
55	an aqueous solution of	f $A_{gNO_3}$ and cupric salt
	solution connected in s	eries. The amount of $A_g$
	deposited is 1.08 gm,	the amount of copper
	deposited is (atomic w 108)	eight of <i>Cu</i> = 63.5; <i>Ag</i> = [EAMCET 1986]
	(a) 0.6454 <i>g</i>	(b) 6.354 g
56.	<ul> <li>(c) 0.3177 g</li> <li>The number of electron atom of aluminium (at. of aluminium chloride Avogadro's number)</li> </ul>	(d) 3.177 g s required to deposit 1gm wt. = 27) from a solution e will be (where N is
		[AIIMS 1992]
	(a) 1 <i>N</i>	(b) 2 <i>N</i>
	(c) 3 <i>N</i>	(d) 4 <i>N</i>
57.	Three faradays of elect molten $Al_2O_3$ , aqueous	ricity are passed through solution of $CuSO_4$ and
	The amount of Al.Cu	and $Na$ deposited at the
	cathodes will be in the r	atio of
		[BHU 1990]
	(a) 1 mole : 2 mole : 3 m	iole
	(b) 3 mole : 2 mole : 1m	ole
	(c) 1 mole : $1.5$ mole : $3^{\circ}$	mole
58	An electrolytic cell cont	ains a solution of $A_{\alpha}$ SO
J <b>C</b> .	and have platinum elect	rodes A current is passed
	until 1.6 $gm$ of $O_2$ has be	een liberated at anode. The
	amount of silver deposite	ed at cathode would be
	(a) 107.88 <i>gm</i>	(b) 1.6 <i>gm</i>
	(c) 0.8 <i>gm</i>	(d) 21.60 gm
59.	The aqueous solution of	of which of the following

decomposes on passing electric current[EAMCET 1973] (a) Canesugar (b) Urea

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	(c) Methanol	(d) Potassium iodide
60.	The number of Faraday	vs needed to reduce 4 gram
	equivalents of $Cu^{++}$ to	Cu metal will be[BHU 1981]
[]	Rognkee 1995]	(b) 2
	(c) 1/2	(d) 4
61.	When electricity is pas $AlCl_3$ , 13.5 gm of Al are	sed through the solution of e deposited. The number of
	Faraday must be	
	[NCERT 1974;	MP PET 1992; MP PMT 1994]
	(a) 0.50	(b) 1.00
-	(c) 1.50	(d) 2.00
62.	The value of one Farada	ay is
	[AMU 198	3; AFMC 1989; MP PET 2001]
	(a) 95500 C mol	(b) 96550 C mol
	(c) 96500 $C mol^{-1}$	(d) 98500 $C mol^{-1}$
63.	The quantity of electri gram equivalent of an e	city needed to liberate 0.5 element is <b>[CPMT 1988; MP PMT 199</b> 7
	(a) 48250 <i>Faradays</i>	(b) 48250 <i>Coulombs</i>
	(c) 193000 Faradays	(d) 193000 <i>Coulombs</i>
64.	The number of cou deposition of 107.870 g	lombs required for the of silver is[ <b>MP PET/PMT 1998]</b>
	(a) 96,500	(b) 48,250
	(c) 1,93,000	(d) 10,000
65.	When one of amp 1 sec through a cond	ere current flows for uctor, this quantity of
	electricity is known as	
	(a) Tamadan	[MP PMT 1993]
	(a) Faraday	
~~	(C) E.M.F.	(a) Onm
00.	proportional to	t all electrode is directly
	[CPMT 197	4; AFMC 1987; MP PET 2000]
	(a) Atomic weight	(b) Equivalent weight
	(c) Molecular weight	(d) Atomic number
67.	From the solution of w faraday of electricity w of metal	vhich of the following one vill liberate one gram atom
	[MHCET 1999; MP	PET 1993,2000; AFMC 2000]
	(a) NaCl	(b) $BaCl_2$
	(c) $CuSO_4$	(d) $AlCl_3$
68.	On electrolysis, 1 mo deposited from its molt	le of aluminium will be en salt by <b>[MH CET 2000]</b>
	(a) 3 moles of electrons	s (b) 4 moles of electrons
69.	(c) 2 moles of electrons [CPMT 1971] The atomic weight of <i>L</i> denosited from <i>E</i> - <i>C</i>	s (d) 1 mole of electrons Fe is 56. The weight of Fe
	Faraday of electricity is	Solution by pussing 0.0

(a) 5.6 g	(b) 11.2 <i>g</i>
(c) 22.4 <i>g</i>	(d) 33.6 <i>g</i>

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[MH	СЕТ	2000]	

ectro	oche	mistr	v 51

7 <b>0.</b>	2.5 $F$ of electricity are	passed through a $CuSO_4$		heating gave 5 $g$ of it	s oxide. The equivalent
	solution. The number	of gm equivalent of Cu		weight of copper is	[KCET 2004]
	deposited on anode is			(a) 23	(D) 32
	[CPMT 1973]	; DPMT 1982; MP PMT 2001]	04	(c) 12 The end of eileren der	(a) 20
	(a) Zero	(b) 1.25	81.	The amount of silver dep	posited by passing 241.25
	(c) 2.5	(d) 5.0		is [MHCET 2002]	ign silver intrate solution
71.	The equivalent weight	of a certain trivalent		(a) 2.7 a	(b) $2.7 ma$
	element is 20. Molecular	weight of its oxide is[KCET 2	003]	(a) $2.7 g$	(d) $0.54 a$
	(a) 152	(b) 56	82	When $1E$ of electrici	ity is passed through
	(c) 168	(d) 68	02.	acidulated water <i>Q</i> , evol	lved is
72.	Silver is removed electric N solution of $AgNO_3$ by	a current of 0.1 <i>ampere</i> .		(a) $11.2 dm^3$	(b) $5.6 dm^3$
	How long will it take to	remove half of the silver			
	from the solution	[AMU 1999]		(c) $22.4 \ dm^3$	(d) $1.0 \ dm^3$
	(a) 16 sec	(b) 96.5 <i>sec</i>	83.	Charge required to libera	ate 11.5 <i>g</i> sodium is
	(c) 100 <i>sec</i>	(d) 10 <i>sec</i>			[AIIMS 1992; DCE 2002]
73.	In order to separate o	xygen from one mole of		(a) 0.5 <i>F</i>	(b) 0.1 <i>F</i>
	$H_2O$ the required quant	ity of coulomb would be [RPE	T 1999	$P_{(c) 1.5 F}$	(d) 96500 coulombs
	(a) $1.93 \times 10^5$	(b) $9.6 \times 10^4$	84.	In the electrolysis of	water, one Faraday of
	(c) 1.8	(d) 3.2		electrical energy would e	evolve
74.	A current of 0.25A is	s passed through CuSO $_4$		(a) One mole of oxygen	(b) One $g$ atom of
	solution placed in volta	meter for 45 <i>minutes</i> . The			
	amount of Cu deposited of $C_{u}$ (2.6)	l on cathode is (At weight	0-	(c) 8 g of oxygen	(d) 22.4 <i>ittres</i> of oxygen
	$01 \ Cu = 05.0$	[BHII 2001]	85.	In a galvanic cell, the ele	ctrons flow from [KCET 2004]
	(a) 0.20 a	(b) 0.22 <i>a</i>		(a) Anode to cathode thro	ough the solution
	(c) $0.25 q$	(d) $0.30 a$		(b) Cathode to anode thr	ough the solution
75.	Faraday constant	[KCET (Med.) 2001]		(c) Anode to cathode three	ough the external circuit
/3	(a) Is a numerical constant	ant		(d) Cathode to anode three	ough the external circuit
	(b) Depends on equivale	nt	86.	An electric current is	passed through silver
	(c) Depends upon the cu	rrent passed		nitrate solution using sil	ver electrodes. 10.79 $g$ of
	(d) Depends on the num	ber of electrons		the same amount of elec	ctricity is passed through
76.	If 0.5 <i>amp</i> current is	passed through acidified		copper sulphate solution	using copper electrodes.
	silver nitrate solution fo	or 10 minutes. The mass of		the weight of copper dep	osited on the cathode is
	silver deposited on cath	node, is ( <i>eq. wt</i> . of silver			[Kerala PMT 2004]
	nitrate = 108)	[AFMC 2001]		(a) 6.4 <i>q</i>	(b) 2.3 <i>q</i>
	(a) 0.235 g	(b) 0.336 g		(c) $12.8 q$	(d) 1.6 <i>a</i>
	(c) $0.536 g$	(d) 0.636 g		(e) 3.2 <i>a</i>	
77.	The unit for the electric	current is[KCET (Med.) 2001]	87.	The law of electrolysis w	ere proposed by
	(a) Onm	(b) Volt	07.		
-0	(c) Ampere	(d) Coulomb		(a) Kohlrauch	(b) Faraday
78.	The qualitity of electricity $112 \text{ cm}^3$ of bydrogon at	STD from acidified water		(a) Nornet	(d) Portholot
	in or injuriogen at	STP from actumed water	00	(c) Nerrist	(u) berthelot
	15	[KCET (Med.) 2001]	ōð.	from a solution of CaC	$Cl_2$ by a current 0.25 mA
	(a) 0.1 <i>Faraday</i>	(b) 1 Faraday		following for 60 seconds	[BHU 2004]
	(c) 965 Coulomb	(d) 96500 <i>Coulomb</i>		(a) $4.68 \times 10^{18}$	<b>(b)</b> $4.68 \times 10^{15}$
7 <b>9</b> .	Which solution will a	show highest resistance		(c) $4.68 \times 10^{12}$	(d) $4.68 \times 10^9$
	$(a) \cap C = N N C$	$(b) \geq N N_{a}C'$	89.	The number of coulombs	required to reduce 12.3 a
	(a) 0.05 N NaCl	$(0) \ge IN INACI$ $(d) = 1 N NaCI$	- 01	of nitrobenzene to aniline	e
80	(c) U.I IV IVUCI	ved in concentrated nitric		(a) 115800 C	(b) 5790 C
00.	acid. The copper nit	rate solution on strong		(c) 28950 C	(d) 57900 C
	inter experiment	services on burning			(-, ), ), ) < <b>&gt;</b>

**90.** During the process of electrolytic refining of copper, some metals present as impurity settle as 'anode mud'. These are

# [AIEEE 2005]

(a) <i>Sn</i> and <i>Ag</i>	(b) Pb and Zn
(c) Ag and Au	(d) Fe and Ni

**91.** A galvanic cell is set up from a zinc bar weighing 50*g* and 1.0*litre*, 1.0*M*, *CuSO*<sub>4</sub> solution. How long would the cell run, assuming it delivers a steady current of 1.0 *ampere* 

[Roorkee	2000]
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(a) 48 hrs	(b) 41 <i>hrs</i>
(c) 21 <i>hrs</i>	(d) 1 <i>hr</i>

**92.** On passing electric current through molten aluminium chloride, 11.2 *litre* of  $Cl_2$  is liberated at NTP at anode. The quantity of aluminium deposited at cathode is (at. *wt*. of Al = 27)

(a) 9 <i>g</i>	(b) 18 <i>g</i>
(c) 27 g	(d) 36 <i>g</i>

**93.** An electric current is passed through silver voltameter connected to a water voltameter. The cathode of the silver voltameter weighed 0.108 g more at the end of the electrolysis. The volume of oxygen evolved at STP is

	[Kerala (Med.) 2003]
(a) 56 cm <sup>3</sup>	(b) 550 cm <sup>3</sup>
(c) 5.6 <i>cm</i> <sup>3</sup>	(d) 11.2 <i>cm</i> <sup>3</sup>
2	

(e) 22.4 *cm*<sup>3</sup>

**94.** During electrolysis of aqueous NaOH, 4g of  $O_2$  gas is liberated at *NTP* at anode,  $H_2$  gas liberated at cathode is

	[CBSE PMT 1998]
(a) 2.8 <i>litres</i>	(b) 5.6 <i>litres</i>
(c) 11.2 <i>litres</i>	(d) 22.4 <i>litres</i>

# Conductor and conductance

- 1. Which one of the following statements is *correct*[MP PET 1997]
  - (a) The oxidation number of oxygen in  $KO_2$  is zero
  - (b) The specific conductance of an electrolyte solution decreases with increase in dilution
  - (c)  $Sn^{2+}$  oxidises  $Fe^{3+}$
  - (d)  $Zn/ZnSO_4$  is a reference electrode
- 2. In infinite dilutions, the equivalent conductances of  $Ba^{2+}$  and  $Cl^{-}$  are 127 and 76  $ohm^{-1}cm^{-1}$   $eqvt^{-1}$ . The equivalent conductivity of  $BaCl_2$  at indefinite

[CBSE 2000]
(b) 139.5
(d) 279.5

3.	The factors of any so	or which is no olution is	ot af	fecti	ng t	he cond	ductiv	ity
	(a) Dilut	ion	(	b) Na	atur	e of ele	ctroly	te
	(c) Temj	perature	(	d) No	one	of these	e	
4.	Specific	conductance	of	0.1	т	nitric	acid	is
	2	· -1 -1	,					

 $6.3 \times 10^{-2} ohm^{-1} cm^{-1}$ . The molar conductance of solution is

# [Kerala (Med.) 2003]

- (a) 630  $ohm^{-1}cm^2mole^{-1}$  (b) 315  $ohm^{-1}cm^2mole^{-1}$
- (c) 100  $ohm^{-1}cm^{2}mole^{-1}$  (d) 6300  $ohm^{-1}cm^{2}mole^{-1}$

(e) 63.0 *ohm*<sup>-1</sup>*cm*<sup>2</sup>*mole*<sup>-1</sup>

- 5. The conductivity of strong electrolyte is[CPMT 2003]
  - (a) Increase on dilution slightly
    - (b) Decrease on dilution
    - (c) Does not change with dilution
  - (d) DEPEMER page abusity of electrolytes itself
- 6. If X is the specific resistance of the solution and M is the molarity of the solution, the molar conductivity of the solution is given by[Kurukshetra CEE 24]

(a) 
$$\frac{1000 X}{M}$$
 (b)  $\frac{1000}{MX}$   
(c)  $\frac{1000 M}{X}$  (d)  $\frac{MX}{1000}$ 

- 7. Conductivity (unit Siemen's) is directly proportional to area of the vessel and the concentration of the solution in it and is inversely proportional to the length of the vessel then the unit of the constant of proportionality is [AIEEE 2002]
  - (a)  $Sm \ mol^{-1}$  (b)  $Sm^2 \ mol^{-1}$
  - (c)  $S^{-2}m^2 mol$  (d)  $S^2m^2 mol^{-2}$
- **8.** If one end of a piece of a metal is heated, the other end becomes hot after some time. This is due to

# [CBSE PMT 1995]

- (a) Energised electrons moving to the other part of the metal
- (b) Resistance of the metal
- (c) Mobility of atoms in the metal
- (d) Minor perturbation in the energy of atoms
- **9.** Conductivity of a solution is directly proportional to

# [KCET 1984]

(a) Dilution	(b) Number o	of ions	
(c) Current density	(d) Volume	of	the
solution			

**10.** The increase in equivalent conductance of an electrolyte solution with dilution is due to the increase in

#### [MP PMT 1996]

- (a) Ionic attraction
- (b) Molecular attraction
- (c) Degree of association of the electrolyte

	(d) Degree of ionisation of the electrolyte		(c) Show a positive temperature coefficient for
11.	Which of the following conducts electricity[AFMC 199	95]	conductance
	(a) Fused $NaCl$ (b) $CO_2$		(d) A single stream of electrons flows from
	(c) $Br_2$ (d) Si	4.0	cathode to anode
12.	Which of the following shows electrical	19.	which one is not a conductor of electricity[ <b>RPET 19</b> (a) $M \subset C$ (colid)
	conduction		(a) $NaCl$ (aqueous) (b) $NaCl$ (solid)
	[CBSE PMT 1999; AIIMS 1999]		(c) NaCl (molten) (d) Ag metal
	(a) Potassium (b) Graphite	20.	Solid sodium chloride is bad conductor of
	(c) Diamond (d) Sodium		(a) It contains only molecules
13.	The unit of equivalent conductivity is		(a) It contains only molecules
	[CPMT 1999; BCECE 2005]		(b) It does not possess folls
	(a) ohm cm		(d) It does not contain free molecules
	(b) $ohm^{-1}cm^2$ (gm equivalent) <sup>-1</sup>	21	Which of the following is a poor conductor of
	(c) $ohm \ am^2(am \ againstant)$	21.	electricity
			[EAMCET 1992]
	(d) $S \ cm^{-2}$		(a) $CH_2COONa$ (b) $C_2H_4OH$
14.	It has been observed that gaseous hydrogen		(c) $NaCl$ (d) $KOH$
	chloride is a very poor conductor of electricity	22	The molar conductivity is maximum for the
	but a solution of hydrogen chloride gas in water is		solution of concentration [DCE 2002]
	fact that		(a) 0.001 <i>M</i> (b) 0.005 <i>M</i>
	(a) Water is good conductor of electricity		(c) 0.002 <i>M</i> (d) 0.004 <i>M</i>
	(h) Hydrogen chloride gas in water solution	23.	The unit of molar conductivity is [DCE 2002]
ioniz	es		(a) $\Omega^{-1} cm^{-2} mol^{-1}$ (b) $\Omega cm^{-2} mol^{-1}$
	(c) A gas is non-conductor but a liquid conducts		(c) $Q^{-1}cm^2mol^{-1}$ (d) $Q^{-mol}cm^2mol$
elect	ricity		The highest cleatrical conductivity of the
	(d) Gas does not obey Ohm's law whereas solution	24.	following aqueous solutions is of
does	•		(a) $0.1 M$ acetic acid (b) $0.1 M$ chloroacetic
15.	Electrolytic conduction differs from metallic	acid	
	conduction in that in the case of electrolytic		(c) 0.1 <i>M</i> fluoroacetic acid (d) 0.1 <i>M</i>
	conduction	diflu	loroacetic acid
	[KCET 1987; Bihar CEE 1992]	25.	Given $l/a = 0.5 cm^{-1}$ , $R = 50 ohm$ , $N = 1.0$ . The
+	(a) The resistance increases with increasing	5	equivalent conductance of the electrolytic cell is [0
temp	(b) The register of degrapping with increasing		(a) 10 $chm^{-1}cm^{2}am ca^{-1}$ (b) 20 $chm^{-1}cm^{2}am ca^{-1}$
temr	(b) The resistance decreases with increasing		
temp	(c) The flow of current does not generate heat		(c) $300 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$ (d) $100 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$
	(d) The resistance is independent of the length of	26.	If equivalent conductance of $1M$ benzoic acid is
	the conductor		$12.8 ohm^{-1} cm^2$ and if the conductance of benzoate
16.	The electrolytic conductance is a direct measure		ion and $H^+$ ion are 42 and 288.42 $ohm^{-1}cm^2$
	of		respectively. its degree of dissociation is
	[KCET 1990; CPMT 2003]		(a) 39% (b) 3.9%
	(a) Resistance (b) Potential		(c) 0.35% (d) 0.039%
	(c) Concentration (d) Dissociation	27.	The unit $ohm^{-1}$ is used for
17.	Conductivity of a strong electrolyte [KCET 1993]	-	(a) Molar conductivity (b) Equivalent
	(a) Increases on dilution	cond	luctivity
	(b) Does not change considerably on dilution		(c) Specific conductivity (d) Conductivity
	(c) Decreases on dilution		Coll constant and Electrophomical Colla
	(d) Depends on density		Cen constant and Electrochemical Cens
18.	Which of the following statements is not	1	When electric current is passed through a call
	applicable to electrolytic conductors [AIIMS 1991]	1.	having an electrolyte, the positive ions move
	(a) New products show up at the electrodes		towards the cathode and the negative ions
	(p) jons are responsible for carrying the current		6

conductance	
(d) A single stream of electrons flows from	om
cathode to anode	
Which one is not a conductor of electricity[ <b>RPET</b>	1999]
(a) <i>NaCl</i> (aqueous) (b) <i>NaCl</i> (solid)	
(c) <i>NaCl</i> (molten) (d) <i>Ag</i> metal	
Solid sodium chloride is bad conductor	of

- electricity because
  - (a) It contains only molecules
  - (b) It does not possess ions
  - (c) The ions present in it are not free to move
  - d) It does not contain free molecules
- Which of the following is a poor conductor of electricity

		LEAMCET 19	92]
	(a) <i>CH</i> <sub>3</sub> <i>COONa</i>	(b) $C_2 H_5 OH$	
	(c) NaCl	(d) <i>KOH</i>	
2.	The molar conductivit	y is maximum for t	the
	solution of concentration	n [DCE 200	02]
	(a) 0.001 <i>M</i>	(b) 0.005 <i>M</i>	

- c) 0.002 M (d) 0.004 M The unit of molar conductivity is [DCE 2002] (a)  $\Omega^{-1} cm^{-2} mol^{-1}$ (b)  $\Omega \ cm^{-2} mol^{-1}$ 
  - c)  $\Omega^{-1}cm^2mol^{-1}$ (d)  $\Omega cm^2 mol$

 $l/a = 0.5 cm^{-1}, R = 50 ohm, N = 1.0$ . Given The equivalent conductance of the electrolytic cell is [Orissa JI (a)  $10 \text{ ohm}^{-1} \text{ cm}^2 \text{ gm eq}^{-1}$  (b)  $20 \text{ ohm}^{-1} \text{ cm}^2 \text{ gm eq}^{-1}$ 

- (c)  $300 \ ohm^{-1}cm^2 \ gmeq^{-1}$  (d)  $100 \ ohm^{-1}cm^2 \ gmeq^{-1}$
- f equivalent conductance of 1M benzoic acid is  $12.8 ohm^{-1} cm^2$  and if the conductance of benzoate on and  $H^+$  ion are 42 and 288.42  $ohm^{-1}cm^2$ respectively. its degree of dissociation is (a) 39% (b) 3.9%
  - (c) 0.35% (d) 0.039%
- The unit  $ohm^{-1}$  is used for (a) Molar conductivity (b) Equivalent ctivity

# **Cell constant and Electrochemical Cells**

When electric current is passed through a cell naving an electrolyte, the positive ions move cowards the cathode and the negative ions towards the anode. If the cathode is pulled out of the solution [AIIMS 1980]

- (a) The positive and negative ions will move towards the anode
- (b) The positive ions will start moving towards the anode, the negative ions will stop moving
- (c) The negative ions will continue to move towards the anode and the positive ions will stop moving
- (d) The positive and negative ions will start moving randomly
- **2.** If the half cell reaction  $A + e^- \rightarrow A^-$  has a large negative reduction potential, it follows that
  - [MNR 1992; UPSEAT 2000, 02]
  - (a) *A* is readily reduced (b) *A* is readily oxidised
- (c)  $A^-$  is readily reduced (d)  $A^-$  is readily oxidised 3. Mark the *false* statement [MP PET 1997]
  - Mark the *false* statement [MP PET 1997] (a) A salt bridge is used to eliminate liquid junction potential
    - (b) The Gibbs free energy change,  $\Delta G$  is related with electromotive force (*E*), as  $\Delta G = -nFE$
    - (c) Nernst equation for single electrode potential RT.

is 
$$E = E^o - \frac{RT}{nF} \ln a_{M^{n+}}$$

(d) The efficiency of a hydrogen oxygen fuel cell is 23%

4. The specific conductance of a 0.1 *N KCl* solution at  $23^{\circ}C$  is  $0.012 \text{ ohm}^{-1} \text{ cm}^{-1}$ . The resistance of cell containing the solution at the same temperature was found to be 55 ohm. The cell constant will be [CBSE PMT 1999, 2000; KCET 2001]

(a)  $0.142 \text{ cm}^{-1}$  (b)  $0.66 \text{ cm}^{-1}$ 

(c) 
$$0.918 \ cm^{-1}$$
 (d)  $1.12 \ cm^{-1}$ 

- **5.** Which of the following reactions occurs at the cathode of a common dry cell
  - (a)  $Mn \rightarrow Mn^{2+} + 2e^{-}$
  - (b)  $2MnO_2 + Zn^{2+} + 2e^- \rightarrow ZnMn_2O_4$
  - (c)  $2ZnO_2 + Mn^{2+} + 2e^- \rightarrow MnZn_2O_4$
  - (d)  $Zn \rightarrow Zn^{2+} + 2e^{-}$
- 6. In Cu Zn cell
  - (a) Reduction occurs at the copper cathode
  - (b) Oxidation occurs at the copper cathode
  - (c) Reduction occurs at the anode
  - (d) Chemical energy is converted to light energy
- **7.** Which of the following reaction is used to make a fuel cell

# [AIIMS 2003]

[BHU 1981]

(a)  $Cd(s) + 2Ni(OH)_3(s) \rightarrow CdO(s) + 2Ni(OH) + H_2O(l)$ 

(b) 
$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

- (c)  $2H_2(g) + O_2(g) \to 2H_2O(l)$
- (d)  $2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l)$
- 8. When lead storage battery is charged[MP PET 2003](a) *PbO*<sub>2</sub> is dissolved

- (b)  $H_2SO_4$  is regenerated
- (c)  $PbSO_4$  is deposited on lead electrode
- (d) Lead is deposited on lead electrode
- When lead storage battery is charged
  - [MP PET 1993; MP PMT 2000]
  - (a) Lead dioxide dissolves
  - (b) Sulphuric acid is regenerated

(c) The lead electrode becomes coated with lead sulphate

- (d) The amount of sulphuric acid decreases
- **10.** The electrolytic decomposition of dilute sulphonic acid with platinum electrode in cathodic reaction is

# [MNR 1988; UPSEAT 1999, 2002]

(a) Oxidation

9.

- (b) Reduction
- (c) Oxidation and reduction both
- (d) Neutralisation
- 11. Which colourless gas evolves, when NH<sub>4</sub>Cl reacts with zinc in a dry cell battery [Orissa JEE 2003]
  (a) NH<sub>4</sub>
  (b) N<sub>2</sub>
  - (c)  $H_2$  (d)  $Cl_2$
- **12.** Which of the substances *Na*, *Hg*, *S*, *Pt* and graphite can be used as electrodes in electrolytic cells having aqueous solutions
  - (a) *Na*, *Pt* and graphite (b) *Na* and *Hg*
  - (c) *Pt* and graphite only (d) *Na* and *S* only
- **13.** In electrolysis of dilute  $H_2SO_4$  using platinum electrodes

[DPMT 1983; IIT 1983; Kurukshetra CET 2002; AFMC 2005] (a)  $H_{2}^{NCE}$  (a)  $H_{2}^{NCE}$  (a) at cathode

- (b)  $NH_3$  is produced at anode
- (c)  $Cl_2$  is obtained at cathode
- (d)  $O_2$  is produced
- **14.** For cell reaction,  $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ , cell representation is [BCECE 2005] (a)  $Zn | Zn^{2+} | |Cu^{2+}|Cu$  (b)  $Cu | Cu^{2+} | |Zn^{2+}| Zn$ (c)  $Cu | Zn^{2+} | | Zn | Cu^{2+} (d) Cu^{2+} | Zn | | Zn^{2+} | Cu$
- **15.** Which one is not called a anode reaction from the following

(a) 
$$Cl^{-} \rightarrow \frac{1}{2}Cl_{2} + e^{-}$$
 (b)  $Cu \rightarrow Cu^{++} + 2e^{-}$   
(c)  $Hg^{+} \rightarrow Hg^{++} + e^{-}$  (d)  $Zn^{2+} + 2e^{-} \rightarrow Zn$ 

- 16. A cell from the following which converts electrical energy into chemical energy
  (a) Dry cell
  (b) Electrochemical cell
  (c) Electrolytic cell
  (d) None of these
- **17.** In the cell  $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$ , the negative electrode is

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(a) Cu (b)  $Cu^{2+}$ (c) Zn (d)  $Zn^{2+}$ 

(c) Zn (d)  $Zn^{2+}$ Which of the following statements is correct ?

- 18. Which of the following statements is correct ? Galvanic cell converts [KCET 1991; MP PMT 1993]
  (a) Chemical energy into electrical energy
  - (b) Electrical energy into chemical energy

(c) Metal from its elemental state to the combined state

(d) Electrolyte into individual ions

- Hydrogen-oxygen fuel cells are used in spacecraft to supply [MP PMT 1993; MP PET 1999]
  - (a) Power for heat and light
  - (b) Power for pressure
  - (c) Oxygen
  - (d) Water
- **20.** The standard cell potential of

 $Zn |Zn^{2+}_{(aq)}| Cu^{2+}_{(aq)} |Cu$  cell is 1.10 V. The

maximum work obtained by this cell will be

- (a) 106.15 kJ (b) 212.30 kJ
- (c) -318.45 kJ (d) -424.60 kJ
- 21. The relationship between standard reduction potential of cell and equilibrium constant is shown by [MP PET 2002]

(a) 
$$E_{cell}^0 = \frac{n}{0.059} \log K_c$$
 (b)  $E_{cell}^0 = \frac{0.059}{n} \log K_c$ 

(c) 
$$E_{cell}^0 = 0.059 \ n \log K_c$$
 (d)  $E_{cell}^0 = \frac{\log K_c}{n}$ 

- **22.** Consider the Galvanic cell  $Zn^{\Theta} | ZnSO_4 || CuSO_4 | Cu^{\oplus}$  the reaction at cathode is **[AMU 2000]** 
  - (a)  $Zn^{2^+} + 2e^- \rightarrow Zn$
  - (b)  $Cu^{2+} + 2e^- \rightarrow Cu$
  - (c)  $Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$
  - (d)  $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$
- **23.** The cell reaction  $Cu + 2Ag^+ \rightarrow Cu^{+2} + Ag$  is best represented by [AMU 2000]
  - (a)  $Cu_{(s)} | Cu^{+2}_{(aq)} || Ag^{+}_{(aq)} | Ag_{(s)}$
  - (b)  $Pt | Cu^{+2} || Ag^{+}_{(aq)} | Ag_{(s)}$
  - (c)  $Cu^{+2} | Cu || Pt | Ag$
  - (d) None of the above representations
- 24.  $Zn_{(s)} | Zn^{2+}_{(aq)} | | Cu^{2+}_{(aq)} | Cu_{(s)}$  is [Kerala (Engg.) 2002] (anode) (cathode)
  - (a) Weston cell (b) Daniel cell
  - (c) Calomel cell (d) Faraday cell
  - (e) Standard cell

- The specific conductance of a solution is 0.2 25.  $ohm^{-1}cm^{-1}$  and conductivity is 0.04  $ohm^{-1}$ . The cell constant would be [RPET 1999] (a) 1 cm<sup>-1</sup> (b) 0 cm<sup>-1</sup> (c) 5 cm<sup>-1</sup> (d) 0.2 cm<sup>-1</sup> 26. If the conductance and specific conductance of a solution is one then its cell constant would be [RPET 1999] (a) 1 (b) Zero (c) 0.5 (d) 4 27. Saturated solution of KNO<sub>3</sub> is used to make 'saltbridge' because [IIT Screening 2002] (a) Velocity of  $K^+$  is greater than that of  $NO_3^-$ (b) Velocity of  $NO_3^-$  is greater than that of  $K^+$ (c) Velocities of both  $K^+$  and  $NO_3^-$  are nearly the same (d)  $KNO_3$  is highly soluble in water **28.** In balancing the half reaction  $S_2 O_3^{2-} \rightarrow S_{(s)}$  the number of electrons that must be added is[DPMT 2000] [MP PET 2002] (a) 4 on the left (b) 3 on the right (c) 2 on the left (d) 2 on the right 29. Which one of the following statement is true for a electrochemical cell [Pb. PMT 1999; KCET 1999] (a)  $H_2$  is cathode and Cu is anode (b)  $H_2$  is anode and Cu is cathode (c) Reduction occurs at  $H_2$  electrode
  - (d) Oxidation occurs at *Cu* electrode
- **30.** In the reaction  $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

The reduction half-cell reaction is

- (a)  $Cu + 2e^- \rightarrow Cu^{2-}$  (b)  $Cu 2e^- \rightarrow Cu^{2+}$
- (c)  $Ag^+ + e^- \rightarrow Ag$  (d)  $Ag e^- \rightarrow Ag^+$
- 31. Which of the following statements about galvanic cell is incorrect [JIPMER 1997]
  (a) Anode is positive
  - (b) Oxidation occurs at the electrode with lower reduction potential
  - (c) Cathode is positive
  - (d) Reduction occurs at cathode
- **32.** The molar conductances of *NaCl*, *HCl* and *CH*<sub>3</sub>*COONa* at infinite dilution are 126.45, 426.16 and  $91 ohm^{-1} cm^2 mol^{-1}$  respectively. The molar conductance of *CH*<sub>3</sub>*COOH* at infinite dilution is[CBSE PMT]
  - (a)  $201.28 \text{ ohm}^{-1} \text{ cm}^{2} \text{ mol}^{-1}$
  - (b)  $390.71 ohm^{-1} cm^2 mol^{-1}$
  - (c)  $698.28 \text{ ohm}^{-1} \text{ cm}^{2} \text{ mol}^{-1}$
  - (d)  $540.48 \text{ ohm}^{-1} \text{cm}^{2} \text{ mol}^{-1}$

- The electrodes of a conductivity cell are 3 cm 33. apart and have a cross-sectional area of  $4 cm^2$ . The cell constant of the cell (in  $cm^{-1}$ ) is (a)  $4 \times 3$ (b) 4/3 (d) 9/4(c) 3/4
- The anode half-reaction occurring during the 34. discharge of a lead storage battery is  $\rightarrow PhSO$ (a)  $Pb(s) \pm SO \pm O$

(a) 
$$PD(s) + SO_2 + O_2 \rightarrow PDSO_4(s)$$

- (b)  $Pb(s) + SO_4^{2-}(aq) \Rightarrow PbSO_4(s) + 2e^{-1}$
- (c)  $PbO_2(s) + 4H^+(aq) + 2e^- + SO_4^{2-}(aq) \Rightarrow$

$$PbSO_4(s) + 2H_2O$$

- (d)  $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$
- The unit of cell constant is 35. [MP PET 1996] (a)  $a^{hm} - 1 a^{-1}$ (b) .....

(c) 
$$cm$$
 (d)  $cm^{-1}$ 

**36.** In dry cell the reaction which takes place at the zinc anode is

[MP PET 1996]

(a)  $Zn^{2+} + 2e^- \rightarrow Zn(s)$ 

(c) *cm* 

(b) 
$$Zn(s) \rightarrow Zn^{2+} + 2e^{-}$$

- (c)  $Mn^{2+} + 2e^- \rightarrow Mn(s)$
- (d)  $Mn(s) \rightarrow Mn^+ + e^- + 1.5V$
- The chemical reaction taking place at the anode of 37. a cell is

[MP PET 1996]

(a) Ionisation	(b) Reduction
(c) Oxidation	(d) Hydrolysis

38. Which of the following reactions occurs at the cathode during the charging of a lead storage battery

[Manipal MEE 1995; MP PET 2002]

(a) 
$$Pb^{2+} + 2e^- \rightarrow Pb$$

(b) 
$$Pb^{2+} + SO_4^{2-} \rightarrow PbSO_4$$

- (c)  $Pb \rightarrow Pb^{2+} + 2e^{-}$
- (d)  $PbSO_4 + 2H_2O \rightarrow 2PbO_2 + 4SO_4^{2-} + 2e^{-}$
- 39. A depolarizer used in dry cell batteries is[NCERT 1981] (a) Ammonium chloride (b) Manganese dioxide (c) Potassium hydroxide (d) Sodium phosphate
- **40.** When a lead storage battery is discharged

[IIT 1987; MP PMT 2004]

- (a)  $SO_2$  is evolved
- (b) Lead sulphate is consumed
- (c) Lead is formed
- (d) Sulphuric acid is consumed

In electroplating, the article to be electroplated 41. serves as

[AMU 1982, 83]

- (a) Cathode (b) Electrolyte
- (c) Anode (d) Conductor
- The position of some metals in the 42. series decreasing electrochemical in electropositive character is given as Mg > Al > Zn > Cu > Ag. What will happen, if a copper spoon is used to stir a solution of aluminium nitrate

# [NCERT 1977]

- (a) The spoon will get coated with Al
- (b) An alloy of Cu and Al is formed
- (c) The solution becomes blue
- (d) There is no reaction
- **43.** In a electrochemical cell [AFMC 1989]
  - (a) Potential energy changes into kinetic energy (b) Kinetic energy changes into potential energy
- (c) Chemical energy changes into electrical energy

(d) Electrical energy changes into chemical energy

- 44. In galvanic cell, the salt bridge is used to[MP PMT 2002] (a) Complete the circuit
  - (b) Reduce the electric resistance in the cell
  - (c) Separate cathode from anode
  - (d) Carry salts for the chemical reaction
- If a strip of *Cu* metal is placed in a solution of 45. ferrous sulphate[NCERT 1974; CPMT 1977; MP PET 2000]
  - (a) Copper will precipitate out
- 46. Which of the following is not used to construct salt bridge

# [MP PET 2003]

- (a)  $CH_3COOK$ (b) *KCl*
- (c)  $NH_4NO_3$ (d)  $KNO_3$
- The reference electrode is made by using[MP PMT 2002] (a)  $ZnCl_2$ (b)  $CuSO_4$
- 48. In a hydrogen - oxygen fuel cell, combustion of hydrogen occurs to [AIEEE 2004] (a) Produce high purity water

(b) Create potential difference between the two

- electrodes
- (c) Generate heat

(d) Remove adsorbed oxygen from electrode surfaces

**49.**  $\lambda_{CICH_2COONa} = 224 \ ohm^{-1} cm^2 gmeq^{-1}$ ,

- - - 47.

      - (d)  $Hg_2Cl_2$ (c)  $HgCl_2$

- (b) Iron will precipitate out

- - - (c) Copper will dissolve
    - (d) No reaction will take place

 $\lambda_{NaCl} = 38.2 \, ohm^{-1} cm^2 gmeq^{-1}$ ,

 $\lambda_{HCl} = 203 \ ohm^{-1} cm^2 gmeq^{-1}$  ,

What is the value of  $\lambda_{CICH_2COOH}$ 

- (a)  $288.5 \ ohm^{-1}cm^2gmeq^{-1}$
- (b)  $289.5 \ ohm^{-1} cm^2 gmeq^{-1}$
- (c)  $388.5 \text{ ohm}^{-1} \text{ cm}^{2} \text{ gmeg}^{-1}$
- (d) 59.5  $ohm^{-1}cm^{2}gmeq^{-1}$
- **50.** Which of the following statement is true for the electrochemical Daniel cell

(a) Electrons flow from copper electrode to zinc electrode

- (b) Current flows from zinc electrode to copper electrode
- (c) Cations move toward copper electrode which is cathode
- (d) Cations move toward zinc electrode
- **51.** Which of the following statement is true for an<br/>electrochemical cell[Pb.CET 2002]
  - (a)  $H_2$  is cathode and *Cu* is anode
  - (b)  $H_2$  is anode and Cu is cathode
  - (c) Reduction occurs at  $H_2$  electrode
  - (d) Oxidation occurs at Cu electrode
- **52.** Which of the following statements are true for an fuel cells

# [DPMT 2004]

- (a) They run till the reactants are active
- (b) They are free from pollution
- (c) They are more efficient
- (d) All of the above
- 53. For gold plating, the electrolyte used is[Pb.CET 2004]
  - (a)  $AuCl_3$  (b)  $HAuCl_4$
  - (c)  $k[Au(CN)_2]$  (d) None of these
- 54. The acid used in lead storage battery is[Pb.CET 2003]

(a)	$H_2SO_4$	(b)	$H_3PO_4$
(c)	HCl	(d)	$HNO_3$

**55.** At 25°C specific conductivity of a normal solution of *KCl* is 0.002765 *mho*. The resistance of cell is 400 *ohms*. The cell constant is

(a) 0.815	(b) 1.016
(c) 1.106	(d) 2.016

 Which of the following is used widely in the manufacture of lead storage battery [BHU 2004]

(a) Ar	senic	(b)	Lithium

- (c) Bismuth (d) Antimony
- **57.** The chemical reaction,  $2AgCl(s)+H_2(g) \rightarrow 2HCl(aq)+2Ag(s)$  taking place in a galvanic cell is represented by the notation

- (a)  $Pt \mid H_2(g)$ , 1bar  $\mid 1M \text{ KCl}(aq) \mid Ag \mid Cl(s) \mid Ag(s)$
- **(b)**  $Pt(s)|H_2(g), 1bar| 1M HCl(aq)|| 1M Ag^+(aq)|Ag(s)$
- [JEE Orissa 2004], 1bar | 1M HCl(aq)| AgCl(s)| Ag(s)

(d)  $Pt(s)|H_2(g)$ , 1bar | 1M HCl(aq)| Ag(s)|AgCl(s)

- **58.** If the  $Zn^{2+}/Zn$  electrode is diluted to 100 times then the change in e.m.f. [DPMT 2005]
  - (a) Increase of 59 mV
  - (b) Decrease of 59 mV [AIIMS 2004]
    - (c) Increase of 29.5 mV
    - (d) Decrease of  $29.5 \, mV$
- **59.** If hydrogen electrode dipped in 2 solution of pH = 3 and pH = 6 and salt bridge is connected the e.m.f. of resulting cell is **[DPMT 2005]** 
  - (a) 0.177 V (b) 0.3 V
  - (c) 0.052 V (d) 0.104 V
- **60.** The tendency of an electrode to lose electrons is known as

# [J & K 2005]

- (a) Electrode potential (b) Reduction potential
- (c) Oxidation potential (d) e.m.f.
- 61. When electric current is supplied through an ionic hydride of fused state, then [Kerala CET 2005]
  - (a) Hydrogen is obtained at anode
  - (b) Hydrogen is obtained at cathode
  - (c) No change
  - (d) Hydride ion moves towards cathode
  - (e) hydride ion present in solution

# Electrode potential, Ecell, Nernt equation and ECS

1. The hydrogen electrode is dipped in a solution of pH = 3 at  $25^{\circ}C$ . The potential of the cell would be (the value of 2.303 RT / F is 0.059 V) [KCET 1993,2005] (a) 0.177 V (b) - 0.177 V(c) 0.087 V (d) 0.059 V

2. The standard electrode potentials of  $Zn^{2+}/Zn$  and  $Ag^+/R^{b-R} = 20\% B V$  and +0.799 V respectively. The standard potential of the cell is [KCET 1993] (a) 1.56 V (b) 0.036 V(c) -1.562 V (d) 0.799 V

**3.** The standard reduction potentials at 298 *K* for the following half reactions are given against each

 $Zn^{2+}(aq.) + 2e \Rightarrow Zn(s);$  - 0.762  $Cr^{3+}(aq) + 3e \Rightarrow Cr(s);$  - 0.740

 $2H^+(aq) + 2e \Rightarrow H_2(g);$  0.00

[AIIMS 2005]

 $Fe^{3+}(aq) + e \Rightarrow Fe^{2+}(aq);$  0.770 Which is the strongest reducing agent [IIT 1981; MP PET/PMT 1988; MP PMT 1989;

MH CET 2001]

(a) Zn(s) (b) Cr(s)

(c)  $H_2(g)$  (d)  $Fe^{2+}(aq)$ 

**4.** When Zn piece is kept in  $CuSO_4$  solution, the copper get precipitated due to standard potential of zinc is

# [CPMT 1999]

- (a) > copper (b) < copper
- (c) > sulphate (d) < sulphate
- 5. Which of the following metal does not react with the solution of copper sulphate [CPMT 1999]
  (a) Mg
  (b) Fe
  (c) Zn
  (d) Ag
  - A solution containing one mole per litre of each  $Cu(NO_3)_2$ ,  $AgNO_3$ ,  $Hg_2(NO_3)_2$  and  $Mg(NO_3)_2$ , is

6.

being electrolysed by using inert electrodes. The values of standard electrode potentials in volts (reduction potentials) are  $Ag / Ag^+ = +0.80, 2Hg / Hg_2^{2+} = +0.79, Cu / Cu^{2+} = +0.34,$ 

 $Mg/Mg^{2+} = -2.37$  with increasing voltage, the sequence of deposition of metals on the cathode will be

[IIT 1984; AMU 1999; Kerala PMT 2004]

(a) Ag, Hg, Cu, Mg
(b) Mg, Cu, Hg, Ag
(c) Ag, Hg, Cu
(d) Cu, Hg, Ag

**7.** The standard reduction electrode potentials of four elements are

A = -0.250 V	B = -0.136 V
C = -0.126 V	D = -0.402 V

The element that displaces A from its compounds is

(a) B	(b) C
(c) D	(d) None of these

The standard oxidation potential of zinc and silver in water at 298 K are

 $Zn(s) \rightarrow Zn^{2+} + 2e^{-}; E = 0.76 V$ 

 $Ag(s) \rightarrow Ag^{2+} + 2e^{-}; E = -0.80 V$ 

Which of the following reactions actually take place

#### [NCERT 1983, 84; KCET 2003]

(a)  $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{++}(aq) + 2Ag(s)$ 

- (b)  $Zn^{++}(aq) + 2Ag(s) \rightarrow 2Ag^{+}(aq) + Zn(s)$
- (c)  $Zn(s) + Ag(s) \rightarrow Zn^{++}(aq) + Ag^{+}(aq)$
- (d)  $Zn^{++}(aq) + Ag^{+}(aq) \rightarrow Zn(s) + Ag(s)$

**9.** Beryllium is placed above magnesium in the second group. Beryllium dust, therefore when added to  $M_gCl_2$  solution will

[CPMT 1977]

- (a) Have no effect
- (b) Precipitate Mg metal
- (c) Precipitate MgO
- (d) Lead to dissolution of *Be* metal
- **10.** The name of equation showing relation between electrode potential (E) standard electrode potential  $(E^{o})$  and concentration of ions in solution is

(a) Kohlrausch's equation (b) Nernst's equation

- (c) Ohm's equation (d) Faraday's equation
- 11. The correct representation of Nernst's equation is

(a) 
$$E_{M^{n+}/M} = E^{o}_{M^{n+}/M} + \frac{0.0591}{n} \log(M^{n+})$$
  
(b)  $E_{M^{n+}/M} = E^{o}_{M^{n+}/M} - \frac{0.0591}{n} \log(M^{n+})$   
(c)  $E_{m^{n+}} = E^{o}_{M^{n+}/M} + \frac{n}{m} \log(M^{n+})$ 

(c) 
$$E_{M^{n+}/M} = E^{o}_{M^{n+}/M} + \frac{n}{0.0591} \log(M^{n+})$$

(d) None of the above

- 12. Standard electrode potential of *NHE* at 298 K is
  - (a) 0.05 V (b) 0.1 V
  - (c) 0.00 V (d) 0.11 V
- 13. When a copper wire is placed in a solution of  $A_{gNO_3}$ , the solution acquires blue colour. This is due to the formation of

# [Roorkee 1989]

- (a)  $Cu^{2+}$  ions
- (b)  $Cu^+$  ions
- (c) Soluble complex of copper with  $AgNO_3$

(d)  $Cu^-$  ion by the reduction of Cu

14. Consider the reaction  $M_{(aq)}^{n+} + ne^- \rightarrow M_{(s)}$ . The standard reduction potential values of the elements  $M_1, M_2$  and  $M_3$  are -0.34V, -3.05V and -1.66V respectively. The order of their reducing power will be [NCERT 1990]

(a) 
$$M_1 > M_2 > M_3$$
 (b)  $M_3 > M_2 > M_1$ 

(c) 
$$M_1 > M_3 > M_2$$
 (d)  $M_2 > M_3 > M_1$ 

**15.** 
$$E^0 = \frac{RT}{nF}$$
 In  $K_{eq}$ . This is called

# [CPMT 1988; MP PET 2000]

- (a) Gibb's equation
- (b) Gibb's-Helmholtz equation
- (c) Nernst's equation
- (d) Vander Waal's equation
- **16.** Four alkali metals A, B, C and D are having respectively standard electrode potential as –

3.05,-1.66,-0.40	and	0.80.	Which	one	will	be th	ıe
most reactive							

[MP PMT/PET 1988 ; CPMT 1983;
MNR 1993; UPSEAT 2002]
(b) B

(a) A (b) B (c) C (d) D

- **17.** Which one of the following metals cannot evolve  $H_2$  from acids or  $H_2O$  or from its compounds
  - [MP PET/PMT 1988; CPMT 1996;
  - AFMC 1998, 99; Pb. PET 1999; BVP 2003]
  - (a) Hg (b) Al
- (c) *Pb*(d) *Fe*18. Which one of the following reaction is not possible

[MP PMT 1991]

electrode

- (a)  $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$
- (b)  $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$
- (c)  $2KBr + I_2 \rightarrow 2KI + Br_2$
- (d)  $CuO + H_2 \rightarrow Cu + H_2O$
- **19.** When a rod of metal *A* is dipped in an aqueous solution of metal *B* (concentration of  $B^{2+}$  ion being 1*M*) at  $25^{\circ}C$ , the standard electrode potentials are  $A^{2+}/A = -0.76$  volts,  $B^{2+}/B = +0.34$  volts [KCET 1992]
  - (a) A will gradually dissolve
  - (b) B will deposit on A
  - (c) No reaction will occur
  - (d) Water will decompose into  $H_2$  and  $O_2$
- **20.** The reaction  $Zn^{2+} + 2e^- \rightarrow Zn$  has a standard potential of -0.76 V. This means [KCET 1992]
  - (a) Zn can't replace hydrogen from acids
  - (b) *Zn* is a reducing agent
  - (c) Zn is a oxidising agent
    (d) Zn<sup>2+</sup> is a reducing agent
- **21.**  $2H^+(aq) + 2e^- \rightarrow H_2(g)$ . The standard
  - - (c) 1 (d) None of these
- 22. *K*, *Ca* and *Li* metals may be arranged in the decreasing order of their standard electrode potentials as [CPMT 1990]
  (a) *K*, *Ca*, *Li*(b) *Ca*, *K*, *Li*(c) *Li*, *Ca*, *K*(d) *Ca*, *Li*, *K*
- **23.** The correct order of chemical reactivity with water according to electrochemical series[**MP PMT 1991**] (a) K > Mg > Zn > Cu (b) Mg > Zn > Cu > K(c) K > Zn > Mg > Cu (d) Cu > Zn > Mg > K
- **24.** *EMF* of cell  $Ni | Ni^{2+} (1.0 M) || Au^{3+} (1.0 M) || Au$ (Where  $E^o$  for  $Ni^{2+} | Ni$  is -0.25 V;  $E^o$  for  $Au^{+3} | Au$  is 1.50 V) is [MP PET 1993; MP PMT 2000]

- (a) + 1.25 V (b) -1.75 V
- (c) + 1.75 V (d) + 4.0 V
- 25. Oxidation and reduction take place in a cell, then<br/>its electromotive force will be[RPET 1999](a) Positive(b) Negative
  - (c) Zero (d) Stable
- **26.** For a spontaneous reaction the  $\Delta G$ , equilibrium constant (*K*) and  $E_{Cell}^o$  will be respectively

(a) - <i>ve</i> ,>1,+ <i>ve</i>	(b) $+ve, >1, -ve$

- (c) -ve, <1, -ve (d) -ve, >1, -ve
- **27.** The reference electrode is made from which of the following

	[MP PET/PMT 1988]
(a) $ZnCl_2$	(b) $CuSO_4$
(c) $Hg_2Cl_2$	(d) $HgCl_2$

- **28.** The charge over anode in a galvanic cell is
  - (a) Negative
  - (b) Positive
  - (c) No charge
  - (d) Sometimes negative and sometimes positive
- **29.** The standard electrode potential for the two electrode  $A^+/A$  and  $B^+/B$  are respectively 0.5 *V* 
  - and 0.75 V. The *emf* of the given cell  $A \mid A^+(a=1) \parallel B^+(a=1) \mid B$  will be
    - (a) 1.25 V (b) 1.25 V
  - (c) -0.25 V (d) 0.25 V
- 30. The standard reduction potential for Li<sup>+</sup>/Li; Zn<sup>2+</sup>/Zn; H<sup>+</sup>/H<sub>2</sub> and Ag<sup>+</sup>/Ag is -3.05, -0.762, 0.00 and +0.80 V. Which of the following has highest reducing capacity
  (a) Ag
  (b) H<sub>2</sub>
  - (c) Zn (d) Li
- **31.** If an iron rod is dipped in  $CuSO_4$  solution

# [MADT Bihar 1984]

- (a) Blue colour of the solution turns green
- (b) Brown layer is deposited on iron rod
- (c) No change occurs in the colour of the solution
- (d) Blue colour of the solution vanishes
- (e) None of the above

**32.**  $E^{o}$  values of  $Mg^{2+}/Mg$  is -2.37V, of  $Zn^{2+}/Zn$  is -0.76V and  $Fe^{2+}/Fe$  is -0.44V. Which of the following statements is correct [EAMCET 1989]

- (a) Zn will reduce  $Fe^{2+}$
- (b) Zn will reduce  $Mg^{2+}$
- (c) Mg oxidises Fe
- (d) Zn oxidises Fe
- **33.** The standard reduction potential for  $Fe^{2+}/Fe$  and  $Sn^{2+}/Sn$  electrodes are -0.44 and -0.14 volt

respectively. For the given cell reaction (b) Q > P > R > S(a) P > O > R > S $Fe^{2+} + Sn \rightarrow Fe + Sn^{2+}$ , the standard EMF is[IIT Screening 1990(MP/PMD 2003]P (d) P > S > Q > R(b) - 0.58 V (a) + 0.30 VWhich of the following metal can deposit copper 41. from copper sulphate solution[CPMT 1983; MP PMT 1989] (c) + 0.58 V(d) - 0.30 V (a) Mercury (b) Iron **34.** Electrode potential of  $Zn^{2+}/Zn$  is -0.76V and that (c) Gold (d) Platinum of  $Cu^{2+}/Cu$  is +0.34V. The *EMF* of the cell constructed between these two electrodes is 42. Standard electrode potential of  $Ag^+/Ag$ and [EAMCET 1992; BHU 2001; CBSE PMT 2001; KCET 1990;  $Cu^+/Cu$  is +0.80V and +0.34V respectively, these MHCET 1999, 2003; Pb. CET 2002; electrodes are joint together by salt bridge if[AMU 2002] AFMC 2001; Pb. PMT 2004] (a) Copper electrode is work like cathode, then (a) 1.10 V (b) 0.42 V  $E_{cell}^{o}$  is +0.45V (c) - 1.1V(d) - 0.42 V (b) Silver electrode is work like anode then  $E_{cell}^{o}$ **35.** *EMF* of a cell whose half cells are given below is  $Mg^{2+} + 2e^{-} \rightarrow Mg(s); E = -2.37 V$ is -0.34V(c) Copper electrode is work like anode then  $E_{cell}^{o}$  $Cu^{2+} + 2e^- \rightarrow Cu(s); E = +0.33 V$ is +0.46V [EAMCET 1987; MP PET 1994; Pb. PMT 2000] (d) Silver electrode is work like cathode then  $E_{cell}^{o}$ (a) - 2.03 V (b) 1.36 V (c) 2.7 V (d) 2.03 V is -0.34V A cell constructed by coupling a standard copper 36. (e) Silver electrode is work like anode then  $E_{cell}^{o}$ electrode and a standard magnesium electrode will be +1.14Vhas emf of 2.7 volts. If the standard reduction The reaction is spontaneous if the cell potential is 43. potential of copper electrode is + 0.34 volt that of [MP PET 1999] magnesium electrode is [KCET 1989] (a) Positive (b) Negative (b) - 3.04 volts (a) + 3.04 volts (c) Zero (d) Infinite (c) + 2.36 volts (d) - 2.36 volts Which substance eliminates bromine from KBr 44. **37.** When  $E^o_{Ag^+/Ag} = 0.8$  volt and  $E^o_{Zn^{2+}/Zn} = -0.76$  volt, solution which of the following is correct [MP PMT 1994] [IIT 1981] (a)  $Ag^+$  can be reduced by  $H_2$ (a) *I*<sub>2</sub> (b) *Cl*<sub>2</sub> (b) Ag can oxidise  $H_2$  into  $H^+$ (c) *HI* (d)  $SO_2$ (c)  $Zn^{2+}$  can be reduced by  $H_2$ 45. A standard hydrogen electrode has zero electrode potential because [IIT 1997] (d) Ag can reduce  $Zn^{2+}$  ion (a) Hydrogen is easiest to oxidise **38.** Adding powdered lead and iron to a solution that (b) The electrode potential is assumed to be zero is 1.0 *M* in both  $Pb^{2+}$  and  $Fe^{2+}$  ions, would result a (c) Hydrogen atom has only one electron reaction, in which (d) Hydrogen is the lightest element [CPMT 1987] **46.** In the electrochemical cell (a) More iron and  $Pb^{2+}$  ions are formed  $H_2(g)1 atm | H^+(1M) || Cu^{2+}(1M) Cu(s)$ (b) More lead and  $Fe^{2+}$  ions are formed (c) Concentration of both  $Pb^{2+}$  and  $Fe^{2+}$  ions Which one of the following statements is true increases [EAMCET 1997] (d) There is no net change (a)  $H_2$  is cathode; Cu is anode Given standard electrode potentials 39. (b) Oxidation occurs at Cu electrode  $Fe^{++} + 2e^{-} \rightarrow Fe; E^{o} = -0.440 V$ (c) Reduction occurs at  $H_2$  electrode  $Fe^{+++} + 3e^{-} \rightarrow Fe; E^{o} = -0.036 V$ (d)  $H_2$  is anode; Cu is cathode The standard electrode potential  $(E^{o})$  for **47.** Expression representing the cell potential (*E cell*)  $Fe^{+++} + e^- \rightarrow Fe^{++}$  is [AIIMS 1982]

- (a) 0.476 V (b) - 0.404 V (c) + 0.404 V(d) + 0.771 V
- **40.** Reduction potential of four elements P, Q, R, S is -2.90, +0.34, +1.20 and -0.76. Reactivity decreases in the order [MP PET 1989; UPSEAT 2001]
- (a) Atomic number of iron is less than that of copper

(a)  $E_{\text{cathode}} + E_{\text{anode}}$ (b) Eanode - Ecathode

- (d)  $E_{\text{left}} E_{\text{right}}$ (c)  $E_{\text{cathode}} - E_{\text{anode}}$
- 48. Iron displaces copper from its salt solution, because

#### [MP PMT 1996]

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- (b) The standard reduction potential of iron is less than that of copper
- (c) The standard reduction potential of iron is more than that of copper
- (d) The iron salt is more soluble in water than the copper salt
- **49.** (i) Copper metal dissolves in 1*M* silver nitrate solution and crystals of silver metal get deposited.
  - (ii) Silver metal does not react with 1 *M* zinc nitrate solution
  - (iii) Zinc metal dissolves in 1M copper sulphate solution and copper metal gets deposited

Hence the order of decreasing strength of the three metals as reducing agents will be

(a) Cu > Ag > Zn (b) Ag > Cu > Zn

(c) Zn > Cu > Ag (d) Cu > Zn > Ag

- **50.** Standard electrode potentials of Zn and Fe are known to be (i) -0.76 V and (ii) -0.44 V respectively. How does it explain that galvanization prevents rusting of iron while zinc slowly dissolves away
  - (a) Since (i) is less than (ii), zinc becomes the cathode and iron the anode
  - (b) Since (i) is less than (ii), zinc becomes the anode and iron the cathode
  - (c) Since (i) is more than (ii), zinc becomes the anode and iron the cathode
  - (d) Since (i) is more than (ii), zinc becomes the cathode and iron the anode
- **51.** Amongst the following electrodes the one with zero electrode potential is
  - (a) Calomel electrode
  - (b) Standard hydrogen electrode
  - (c) Glass electrode
  - (d) Gas electrode
- **52.** Which of the following is correct expression for electrode potential of a cell

(a) 
$$E = E^{o} - \frac{RT}{nF} \ln \frac{[\text{product}]}{[\text{reactant}]}$$
  
(b)  $E = E^{o} + \frac{RT}{F} \ln \frac{[\text{product}]}{[\text{reactant}]}$ 

(c) 
$$E = E^o - \frac{RT}{nF} \ln \frac{[\text{reactant}]}{[\text{product}]}$$

(d) 
$$E = -\frac{RT}{F} \ln \frac{[\text{product}]}{[\text{reactant}]}$$

53. Calculate standard free energy change for the reaction  $\frac{1}{2}Cu(s) + \frac{1}{2}Cl_2(g) \Rightarrow \frac{1}{2}Cu^{2+} + Cl^-$  taking place at 25 ° C in a cell whose standard e.m.f. is 1.02 volts [MP PMT 1997] (a) - 98430 J (b) 98430 J In which cell the free energy of a chemical reaction is directly converted into electricity ?[MP PET/PM (a) Leclanche cell (b) Concentration cell
(c) Fuel cell (d) Lead storage battery
Nernst equation is related with
(a) The electrode potential and concentration of ions in the solution
(b) Equilibrium constant and concentration of
(c) Free energy change and E.M.F. of the cell
(d) None of these

(d) - 49215 J

**56.** The standard reduction potentials of 4 elements are given below. Which of the following will be the most suitable reducing agent

I = - 3.04 V, II = - 1.90 V, III = 0 V, IV = 1.90 V

- [CPMT 1999]
- (a) I (b) II

(c) 96500 J

54.

55.

ions

(c) III (d) IV

57. Electrode potential data are given below :  

$$Fe^{3+}(aq) + e^- \rightarrow Fe^{-1}(aq); E^o = +0.77 V$$
  
 $Al^{3+}(aq) + 3e^- \rightarrow Al(s); E^o = -1.66 V$   
 $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq); E^o = +1.08 V$   
Based on the data given above, reducing power of  
 $Fe^{2+}, Al$  and  $Br^-$  will increase in the order[Pb. PMT 1998]  
(a)  $Br^- < Fe^{2+} < Al$  (b)  $Fe^{2+} < Al < Br^-$   
(c)  $Al < Br^- < Fe^{2+}$  (d)  $Al < Fe^{2+} < Br^-$ 

- **58.** The standard electrode potential  $(E^{o})$  for
  - [MPEMT  $G_{997}$  and  $Cl^{-}/\frac{1}{2}Cl_{2}$  respectively are 0.94 V

and -1.36 V. The  $E^o$  value for  $OCl^- / \frac{1}{2}Cl_2$  will be[KCET 19

(a) -0.42 V	(b) −2.20 V
(c) 0.52 V	(d) 1.04 V

59 [MF PMT reduction potential is more, then [CPMT 1996]

- (a) It is easily oxidised
- (b) It is easily reduced
- (c) It acts as oxidising agent
- (d) It has redox nature
- **60.** One of the following is false for Hg [BHU 1998]
  - (a) It can evolve hydrogen from  $H_2S$
  - (b) It is a metal
  - (c) It has high specific heat
  - (d) It is less reactive than hydrogen

**61.**  $E^{o}$  for the cell  $Zn | Zn^{2+}(aq) || Cu^{2+}(aq) || Cu$  is 1.10 V

at  $25^{\circ}C$ , the equilibrium constant for the reaction  $Zn + Cu^{2+}(aq) = Cu + Zn^{2+}(aq)$  is of the order of

[CBSE PMT 1997]

- (a)  $10^{-28}$  (b)  $10^{-37}$
- (c)  $10^{+18}$  (d)  $10^{+17}$

62.	Standard reduction	potentials at 25°C of
	$Li^+ \mid Li, Ba^{2+} \mid Ba, Na^+ \mid Na$	$a$ and $Mg^{2+} \mid Mg$ are
	-3.05, -2.90, -2.71 and	d –2.37 <i>volt</i> respectively.
	Which one of the f	following is the strongest
	oxidising agent	
	[0	CBSE PMT 1994; JIPMER 2002]
	(a) Na <sup>+</sup>	(b) <i>Li</i> <sup>+</sup>
	(c) $Ba^{2+}$	(d) $Mg^{2+}$
63.	Which of the follow:	ing displaces Br <sub>2</sub> from an
	aqueous solution conta	aining bromide ions
	[CBSE PN	IT 1994; JIPMER (Med.) 2002]
	(a) $Cl_2$	(b) <i>Cl</i> <sup>-</sup>
	(c) I <sub>2</sub>	(d) $I_3^-$
64.	For the cell reaction	
	$Cu^{2+}(C_1aq) + Zn(s) = Zn^{2+}$	$(C_2aq) + Cu(s)$
	of an electrochemica	l cell, the change in free
	energy at a given temp	perature is a function of
	(a) $\ln (C_1)$	(b) $\ln (C_2)$
	(c) $\ln (C_1 + C_2)$	(d) $\ln (C_2 / C_1)$
65.	The e.m.f. of the ce	ell in which the following
	reaction $Zn(s) + Ni^{2+}(s)$	$a = 1.0) \rightleftharpoons Zn^{2+}(a = 10) + Ni(s)$
	occurs, is found to	be $0.5105 V$ at $298 K$ . The
	standard e.m.f. of the	cell is
		[Roorkee Qualifying 1998]
	(a) 0.5400	(b) 0.4810 V
	(c) 0.5696 V	(d) – 0.5105 V
66.	For the redox reaction	
	$Zn(s) + Cu^{2+}(0.1M) \to Zn^{4}$	$C^{+}(1M) + Cu(s)$ taking place in
	a cell, $E_{cell}^o$ is 1.10 volt	. $E_{cell}$ for the cell will be
	$\left(2.303  \frac{RT}{F} = 0.0591\right)$	[AIEEE 2003]
	(a) 2.14 <i>volt</i>	(b) 1.80 <i>volt</i>
	(c) 1.07 <i>volt</i>	(d) 0.82 <i>volt</i>
67.	The <i>emf</i> of a Dan	iel cell at 298K is E

**57.** The *emf* of a Daniel cell at 298K is  $E_1$  $Zn|ZnSO_4||CuSO_4|Cu$  when the concentration of (0.01 M) (1.0 M)

 $ZnSO_4$  is 1.0 *M* and that of  $CuSO_4$  is 0.01 *M*, the *emf* changed to  $E_2$ . What is the relationship between  $E_1$  and  $E_2$ 

(a)  $E_2 = 0 \neq E_1$  (b)  $E_1 > E_2$ (c)  $E_1 < E_2$  (d)  $E_1 = E_2$ 

**68.** The oxidation potentials of following half-cell reactions are given

 $Zn \rightarrow Zn^{2+} + 2e^-; E^o$  = 0.76 V ,

 $Fe \rightarrow Fe^{2+} + 2e^-; E^o = 0.44 V$  what will be the *emf* of cell, whose cell-reaction is

 $Fe^{2+}(aq) + Zn \rightarrow Zn^{2+}(aq) + Fe$ 

[MP PMT 2003]

	(a) – 1.20 V	(b) + 0.32 <i>V</i>
	(c) - 0.32 V	(d) + 1.20 V
69.	The $E^{\circ}$ for half cells Fe	$e/Fe^{2+}$ and $Cu/Cu^{2+}$ are -
	0.44 V and + 0.32 V resp	pectively. Then[MP PMT 2003]
	(a) $Cu^{2+}$ oxidises Fe	(b) $Cu^{2+}$ oxidises $Fe^{2+}$
	(c) Cu oxidises $Fe^{2+}$	(d) $Cu$ reduces $Fe^{2+}$
7 <b>0.</b>	What is $E^{o}$ for elements	ectrode represented by
	$Pt, O_2(1 atm)/2H^+(\text{Im})$	[JIPMER 1997]
	(a) Unpredictable	(b) Zero
	(c) 0.018 V	(d) 0.118 V
71.	The cell potential of a ce	ll in operation is
	(a) Zero	(b) Positive
	(c) Negative	(d) None of the above
72.	Which of the following is	s displaced by <i>Fe</i> [Roorkee 1995]
	(a) <i>Ag</i>	(b) <i>Hg</i>
	(c) <i>Zn</i>	(d) <i>Na</i>
73.[0	en strandegistjelectrode p	ootential of the half cells
	are given below	
	$Zn^{2+} + 2e^- \rightarrow Zn; E = -7.62$	<i>V</i> ,
	$Fe^{2+} + 2e^- \rightarrow Fe; E = -7.81$	V
	The <i>emf</i> of the cell $Fe^{2+}$	$+Zn \rightarrow Zn^{2+} + Fe$ is
		[CPMT 2003]
	(a) 1.54 V	(b) – 1.54 V
	(c) - 0.19 V	(d) + 0.19 $V$
74.	$Zn^{2+} + 2e^- \rightarrow Zn(s); E^o = -0$	.76 ,
	$Fe^{3+} + e^{-} \rightarrow Fe^{2+}; E^{o} = -0.7$	77,
	$Cr^{3+} + 3e^- \rightarrow Cr; E^o = -0.79$	9,
	$H^+ + 2e^- \rightarrow 1/2H_2; E^o = 0.$	00
	Strongest reducing agen	t is
	(a) $Fe^{2+}$	(b) <i>Zn</i>
	(c) <i>Cr</i>	(d) $H_2$
75.	Standard reduction electron metals <i>A</i> , <i>B</i> and <i>C</i> are read and – 1.2 <i>V</i> . The reducinare <b>[IIT 1998; AIEEE 2003</b> ]	trode potentials of three espectively + 0.5V, - 3.0V ng powers of these metals ]
	(a) $B > C > A$	(b) $A > B > C$
	(c) $C > B > A$	(d) $A > C > B$
76.	For a cell reaction i change, the standard <i>em</i>	nvolving a two-electron f of the cell is found to be
	0.295 V at $25^{\circ}C$ . The ed	quilibrium constant of the
	reaction at $25^{\circ}C$ will b	e
	[Roorkee 1999; Al	EEE 2003; CBSE PMT 2004]
	(a) $1 \times 10^{-10}$	<b>(b)</b> $29.5 \times 10^{-2}$
	(c) 10	(d) $1 \times 10^{10}$
77.	For the electrochemical	cell,
	$M \mid M^+ \parallel X^- \mid X, E^o(M^+ / M)$	$(1) = 0.44 \ V \text{ and}$

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 $M \mid M \mid X \mid X, E (M \mid M) = 0.44$  V and  $E^{o}(X \mid X^{-}) = 0.33$  V. From this data one can deduce that

[IIT-JEE (Screening) 2000]

- (a)  $M + X \rightarrow M^+ + X^-$  is the spontaneous reaction
- (b)  $M^+ + X^- \rightarrow M + X$  is the spontaneous reaction
- (c)  $E_{cell} = 0.77 V$
- (d)  $E_{cell} = -0.77 V$
- **78.** The standard potential at  $25^{\circ}C$  for the following half reactions are given against them  $Zn^{2+} + 2e \rightarrow Zn, E^{\circ} = -0.762 V$

$$Mg^{2+} + 2e \rightarrow Mg, E^o = -2.37 V$$

When zinc dust is added to the solution of  $MgCl_2$ 

(a)  $ZnCl_2$  is formed

- (b) Zinc dissolves in the solution
- (c) No reaction takes place
- (d) *Mg* is precipitated
- **79.**  $KMnO_4$  acts as an oxidising agent in the neutral medium and gets reduced to  $MnO_2$ . The equivalent weight of  $KMnO_4$  in neutral medium [AMU 2001]

(a) <i>mol. wt</i> /2	(b) <i>mol.wt</i> /3
(c) <i>mol. wt</i> /4	(d) <i>mol</i> .wt/7

- **80.** Which of the following condition will increase the voltage of the cell, represented by the equation  $Cu_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$ [CBSE PMT 2001]
  - (a) Increase in the concentration of  $Ag^+$  ion
  - (b) Increase in the concentration of  $Cu^+$  ion
  - (c) Increase in the dimension of silver electrode
  - (d) Increase in the dimension of copper electrode
- **81.** Which will increase the voltage of the cell  $Sn_{(s)} + 2Ag^+_{(aq)} \rightarrow Sn^{2+}_{(aq)} + 2Ag_{(s)}$  [DPMT 2001]
  - (a) Increase in the concentration of  $Ag^+$  ions
  - (b) Increase in the concentration of  $Sn^{2+}$  ions
  - (c) Increase in size of the silver rod
  - (d) None of these

(a) 1840 KeV

**82.** The mass of the proton is 1840 times that of electron, its potential difference is *V*. The kinetic energy of proton is

(b) 1 KeV

[DCE 2001]

[UPSEAT 2001]

(c) 
$$\frac{1}{1840}$$
 KeV (d) 920 KeV

83. What will be the *emf* for the given cell  $Pt \mid H_2(P_1) \mid H^+(aq) \mid \mid H_2(P_2) \mid Pt$  [AIEEE 2002]

(a) 
$$\frac{RT}{f} \log \frac{P_1}{P_2}$$
 (b)  $\frac{RT}{2f} \log \frac{P_1}{P_2}$   
(c)  $\frac{RT}{f} \log \frac{P_2}{P_1}$  (d) None of these

**84.** What is the potential of a cell containing two hydrogen electrodes the negative one in contact

with  $10^{-8} M H^+$  and positive one in contact with 0.025  $M H^+$ 

# [MP PMT 2000]

[Pb. PMT 2000]

- (a) 0.18 V
  (b) 0.28 V
  (c) 0.38 V
  (d) 0.48 V
- **85.** Will  $Fe_{(s)}$  be oxidised to  $Fe^{2+}$  by the reaction with

1 *M HCl* (
$$E^o$$
 for  $Fe/Fe^{2+} = + 0.44$  V)[Pb. PMT 2000]

- (c) May be (d) Can't say
- **86.** *EMF* of a cell in terms of reduction potential of its left and right electrodes is

(a) 
$$E = E_{left} - E_{right}$$
 (b)  $E = E_{left} + E_{right}$ 

(c) 
$$E = E_{right} - E_{left}$$
 (d)  $E = -(E_{right} + E_{left})$ 

**87.** Arrange the following in the order of their decreasing electrode potential *Mg*, *K*, *Ba*, *Ca*[JIPMER 2002]

(a) 
$$K, Ba, Ca, Mg$$
 (b)  $Ca, Mg, K, Ba$ 

(c) 
$$Ba, Ca, K, Mg$$
 (d)  $Mg, Ca, Ba, K$ 

**88.** Which of the following has highest electrode potential

**89.** The cell reaction of a cell is

 $Mg_{(s)} + Cu^{2+}(aq) \rightarrow Cu_{(s)} + Mg^{2+}(aq)$ 

If the standard reduction potentials of  $M_g$  and Cu are -2.37 and +0.34 V respectively. The *emf* of the cell is

# [EAMCET 1995; JIPMER (Med.) 2001;

#### AFMC 2002; CBSE PMT 2002]

(a) 2.03 V	(b) – 2.03 V

- (c) + 2.71 V (d) 2.71 V
- **90.** The element which can displace three other halogens from their compound is **[EAMCET 1998]** 
  - (a) *Cl* (b) *F*

(c) Br (d) I

- 91. Which of the following has been universally accepted as a reference electrode at all temperatures and has been assigned a value of zero volt [AIIMS 1998]
  - (a) Graphite electrode
  - (b) Copper electrode
  - (c) Platinum electrode
  - (d) Standard hydrogen electrode
- **92.** Aluminium is more reactive than *Fe*. But *Al* is less easily corroded than iron because **[DCE 1999]** 
  - (a) *Al* is noble metal
  - (b) Iron forms both mono and divalent ions
  - (c) Oxygen forms a protective oxide layer

(d) *Fe* undergoes reaction easily with  $H_2O$ 

- **93.** Zinc displaces copper from the solution of its salt because
  - [MP PET 1995]

(a) Atomic number of zinc is more than that of copper

(b) Zinc salt is more soluble in water than the copper salt

(c) Gibbs free energy of zinc is less than that of copper

- (d) Zinc is placed higher than copper in electrochemical series
- 94. An electrochemical cell is set up as follows

 $Pt(H_2, 1 atm)/0.1 M HCl$ 

|| 0.1 M acetic acid /( $H_2$ , 1 atm) Pt

E.M.F. of this cell will not be zero because

#### [CBSE PMT 1995]

- (a) The *pH* of 0.1 *M HCl* and 0.1 *M* acetic acid is not the same
- (b) Acids used in two compartments are different

(c) E.M.F. of a cell depends on the molarities of acids used

(d) The temperature is constant

**95.**  $Cu^+$  ion is not stable in aqueous solution because of disproportionation reaction.  $E^o$  value for disproportionation of  $Cu^+$  is

(Given  $E_{Cu^{2+}/Cu^{+}}^{o} = 0.15$ ,  $E_{Cu^{2+}/Cu}^{o} = 0.34V$ ) [IIT 1995] (a) - 0.49 V (b) 0.49 V (c) - 0.38 V (d) 0.38 V

- $E^{o}$  of a cell  $aA + bB \rightarrow cC + dD$  is [CPMT 1997]
- (a)  $-\frac{RT}{nF}\log\frac{[C]^{c}[D]^{d}}{[A]^{a}[B]^{b}}$  (b)  $-RT\log\frac{[a]^{A}[b]^{B}}{[a]^{C}[d]^{D}}$ (c)  $-\frac{RT}{nF}\log\frac{[C]^{c}[d]^{D}}{[A]^{a}[B]^{b}}$  (d)  $-\frac{RT}{nF}\log\frac{[C]^{c}[d]^{D}}{[a]^{A}[B]^{b}}$
- **97.** In the experiment set up for the measurement of *EMF* of a half cell using a reference electrode and a salt bridge, when the salt bridge is removed, the voltage

# [NCERT 1984; CPMT 1985]

- (a) Does not change
- (b) Decreases to half the value
- (c) Increase to maximum

(d) Drops to zero

96.

**98.** Electrode potentials of five elements *A*,*B*,*C*,*D* and *E* are respectively – 1.36 , – 0.32, 0, – 1.26 and – 0.42. The reactivity order of these elements are in the order of

#### [MP PMT 1995]

(a) *A*, *D*, *E*, *B* and *C*(b) *C*, *B*, *E*, *D* and *A*(c) *B*, *D*, *E*, *A* and *C*(d) *C*, *A*, *E*, *D* and *B*

**99.** What is wrongly stated about electrochemical series

#### [DCE 1999]

- (a) It is the representation of element in order of increasing or decreasing standard electrode reductional potential
- (b) It does not compare the relative reactivity of metals
- (c) It compares relative strengths of oxidising agents
- (d)  $H_2$  is centrally placed element
- **100.** Which of the following statements is true for fuel cells

# [KCET (Med.) 1999; AFMC 2000]

- (a) They are more efficient
- (b) They are free from pollution
- (c) They run till reactants are active
- (d) All of these
- **101.** What is the potential of a half-cell consisting of zinc electrode in 0.01*m*  $ZnSO_4$  solution at 258 °*C* ( $E^o = 0.763 V$ )

	[AIIMS 2000; BHU 2000]
(a) 0.8221 V	(b) 8.221 V
(c) 0.5282 V	(d) 9.232 V

- **102.** The *emf* of a galvanic cell, with electrode potentials of silver = +0.80V and that of copper = + 0.34 V, is[AIIMS 19]
  - (a) 1.1 V (b) + 1.1 V
  - (c) + 0.46 V (d) + 0.76 V
- **103.** Copper cannot replace..... from solution[**DPMT 2002**]

(a) 
$$Fe$$
 (b)  $Au$   
(c)  $Hg$  (d)  $Ag$ 

**104.** The strongest reducing agent of the alkali metal is

- [CBSE PMT 2000]
- (a) *Li* (b) *Na* (c) *K* (d) *Cs*

**105.** Which of the following is the most electropositive element

- [Pb. PMT 2000]
- (a) Carbon (b) Calcium
- (c) Chlorine (d) Potassium
- 106. The metal that forms a self protecting film of oxide to prevent corrosion, is [BHU 1999]
  (a) Cu
  (b) Al
  - (c) *Na* (d) *Au*

**107.** In a cell that utilises the reaction  $Zn_{(s)} + 2H^+(aq) \rightarrow$ 

 $Zn^{2+}(aq) + H_{2(g)}$  addition of  $H_2SO_4$  to cathode

compartment, will [AIEEE 2004]

(a) Increase the E and shift equilibrium to the right

(b) Lower the E and shift equilibrium to the right

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	<ul><li>(c) Lower the <i>E</i> and shi</li><li>(d) Increase the <i>E</i> and s</li></ul>	ft equilibrium to the le shift equilibrium to the	eft e left	0.2905 then the value reaction is	of equilibrium for the cell [IIT-JEE Screening 2004]
108.	For the electrochemic $E^{\circ}(M^+ M) = 0.44 \text{ V} E^{\circ}(X)$	ical cell, $M \mid M^+ \parallel X$	X,	(a) $\frac{0.32}{e^{0.0295}}$	(b) $\frac{0.32}{10^{0.0295}}$
	$E(M \mid M) = 0.44$ V $E(X \mid data, one can deduce th$	at		0.26 Pl <b>í cìe t<del>-2004 ]</del></b>	(d) $\frac{0.32}{0.32}$
	(a) $E^{\circ}_{call} = -0.77 V$		Ľ	10 <sup>-0002951</sup>	$10^{0.0591}$
	(b) $M^+ + X^- \rightarrow M + X$ is	the spontaneous reacti	117. Ion	Aluminium displaces whereas silver does	hydrogen from dilute <i>HCl</i> not. The <i>e.m.f.</i> of a cell
	(c) $M + X \rightarrow M^+ + X^-$ is t (d) $E^{\circ}_{cell} = .77 V$	he spontaneous reactio	n	prepared by combini 2.46 V. The reduction is +0.80 V. The reduct	ng $Al / Al^{3+}$ and $Ag / Ag^+$ is potential of silver electrode tion potential of aluminium
	<b>109.</b> The standard <i>e</i> .	<i>m.f.</i> of a call, involving	g one	electrode is	[KCET 2004]
	electron change is f	ound to be $0.591 V$ at 2	25°C.	(a) +1.66 V	(b) -3.26 V
	( $F = 96,500 \ C \ mol^{-1}$ ;	$R = 8.314 \ JK^{-1}mol^{-1}$ ) [A	IEEE	(c) 3.26 V	(d) -1.66 V
		2	2004] 118.	Consider the following	$E^0$ values :
	(a) $1.0 \times 10^{10}$	(b) $1.0 \times 10^5$		$E^0 F e^{3+} / F e^{2+} = +0.77 V$	
110.	(c) $1.0 \times 10^{1}$ Standard electrode	(d) $1.0 \times 10^{30}$ potential of	cell	$E^{0}_{Sn^{2+}/Sn} = -0.14 V$	
	$H_2 \mid H^+ \parallel Ag^+ \mid Ag$ is	[AIEEE 2	2004]	Under standard condi reaction $Sn_{(s)} + 2Fe^{3+}(aq)$	tions the potential for the $q \rightarrow 2Fe^{2+}(sq) + Sn^{2+}(aq)$ is[AIEEE 200
	(a) 0.8 V	(b) – 0.8 V		(a) 0.91 V	(b) 1.40 V
	(c) – 1.2 V	(d) 1.2 <i>V</i>		(c) 1.68 V	(d) 0.63 V
111.	A galvanic cell with $A' = +2.23 V$ and $B' = -1.4$	electrode potentia 13 V. The value of $E^{\circ}_{ce}$	l of <b>119.</b> " is <b>[Pb.CET 2</b> 0	$Cr_2O_7^{2-} + I^- \rightarrow I_2 + Cr^{3+}$ <b>DO3</b>	
	(a) 3.66 V	(b) 0.80 V		$E^0_{\ cell} = 0.79 \ V$	
	(c) - 0.80 V	(d) – 3.66 V		$E_{Cr_2O^{2-}}^0 = 1.33 V, E_{I_2}^0$ is	[BVP 2004]
112.	The <i>e.m.f.</i> of a cell w below	hose half cells are g	given is	(a) -0.10 V	(b) +0.18 V
	$Mg^{2+} + 2e^- \to Mg(s) E^\circ = -$	2.37 V		(c) -0.54 V	(d) 0.54 V
	$Cu^{2+} + 2e^- \to Cu(s) E^\circ = +0$	.34 V [Pb.CET 2	2001] 120.	$Zn(s) + Cl_2(1 \text{ atm}) \rightarrow Zn^{2+}$	$+2Cl^{-}$ . $E^{0}_{cell}$ of the cell is
	(a) + 1.36 V	(b) + 2.71 V		2.12 V. To increase E	[BVP 2004]
	(c) + 2.17 V	(d) – 3.01 V		(a) $[Zn^{2+}]$ should be in	creased
113.	For the cell reaction, $E^{\circ}_{cell}$ is 1.89 V. If $E^{\circ}_{Ce^{4+}/C$	$2Ce^{4+} + Co \rightarrow 2Ce^{3+} +$ $Ce^{3+} \qquad [Pb.CET 2]$	<i>Co</i> <sup>2+</sup>	(b) $[Zn^{2+}]$ should be defined as	ecreased
	(a) - 1.64 V	(b) + 1.64 V		(c) $[Cl^{-}]$ should be decr	reased
	(c) - 2.08 V	(d) + 2.17 V		(d) $P_{Cl_2}$ should be decr	eased
114.	If the $\Delta G$ of a cell read $-21.20 \text{ KJ}$ ; the standard	tion $AgCl + e^- \rightarrow Ag + C$ e.m.f., of cell is [MP	cl <sup>–</sup> is <b>121.</b> PMT 2004]	The $E^0_{M^{3+}/M^{2+}}$ values -0.41 +1.57 +0.77 and	for <i>Cr</i> , <i>Mn</i> , <i>Fe</i> and <i>Co</i> are
	(a) 0.229 V	(b) 0.220 V		which one of these me	tals the change in oxidation
	(c) - 0.220 V	(d) - 0.110 V		state from +2 to +3 is	easiest
115.	The <i>e.m.f.</i> of the cell	$Ag \mid Ag^{+}(0.1M) \parallel Ag^{+}(1M)$	) Ag	(a) <i>Fe</i>	(b) <i>Mn</i>
	at 298 <i>K</i> is	[DCE 2	2003]	(c) <i>Cr</i>	(d) <i>Co</i>
	(a) 0.0059 V	(b) 0.059 V	122.	The rusting of iron tak	es place as follows
	(c) 5.9 V	(d) 0.59 V		$2H^+ + 2e^- + \frac{1}{2}O_2$	$\rightarrow H_2O(l);$
116.	The <i>e.m.f.</i> of the cell			$E^{\circ} = +1.23 V$	<b>T</b> 0 <b>I</b> 1
	$Zn  Zn^{2+}(0.01M)  Fe^{2+}(0.01M) $	001M  Fe at 298 K is		$Fe^{2^+} + 2e^- \longrightarrow Fe(s)$	; $E^{\circ} = -0.44 V$

				10 /
	Calculate $\Delta G^{\circ}$ for the net process	[IIT 2005]	(a) <i>FeCl</i> <sub>2</sub>	(b) <i>PbCl</i> <sub>2</sub>
	(a) $-322 \ kJ \ mol^{-1}$ (b) $-161 \ kJ \ mol^{-1}$	-1	(c) $Hg_2Cl_2$	(d) $HgCl_2$
123.	$(c) -152 \text{ kJ mol}^2$ $(d) -76 \text{ kJ mol}^2$ When an acid cell is charged, then [A	FMC 2005] 4.	Corrosion of iron is phenomenon where	essentially an electrochemical the cell reactions are <b>[KCET 20</b>
	(a) Voltage of cell increases		(a) Fe is oxidised	to $Ee^{2+}$ and dissolved oxygen
	(b) Electrolyte of cell dilutes		(d) It is oxidised	
	(c) Resistance of cell increases		in water is redu	aced to OH
124	(d) None of these The standard electrode notential is more	urad by	(b) <i>Fe</i> is oxidised	to $Fe^{3+}$ and $H_2O$ is reduced to
124.	The standard electrode potential is meas	$Q_2^{2-}$		
	(a) Electrometer (b) Voltmeter			- 21
	(c) Pyrometer (d) Galvanomet	er	(c) <i>Fe</i> is oxidised	to $Fe^{2+}$ and $H_2O$ is reduced to
125.	Aluminium displaces hydrogen from	acids but $O_2^-$		
	copper does not. A galvanic cell pre	epared by	(d) $E_{\alpha}$ is ovidised t	$E_{a}^{2+}$ and $H_{a}$ is reduced to
	combining $C_{\mu}/C_{\mu}^{2+}$ and $A_{\mu}/A_{\mu}^{3+}$ has an	e.m.f. of		$H_2O$ is reduced to
	2.0 V at 298 K. If the potential of copper	electrode	$O_2$	
	is $+ 0.34 V$ , that of aluminium is			
	[CPMT 2001; H	(CET 2001]		aal Thinking
	(a) + 1.66 V (b) - 1.66 V			cal Ininking
	(c) + 2.34 V (d) - 2.3 V			
126.	If the standard electrode potential of	$Cu^{2+}/Cu$		Objective Questions
	electrode is 0.34 <i>V</i> , what is the electrode	potential		
	of 0.01 <i>M</i> concentration of $Cu^{2+}$ ( <i>T</i> = 298)	K) [EAMCET 2003]	The limiting molar	conductivities $\wedge^0$ for NaCl KBr
	(a) $0.200 V$ (b) $0.281 V$	1.0	and KCl are 126	$5$ 152 and 150 S $cm^2mol^{-1}$
	(a) $0.399V$ (b) $0.201V$ (c) $0.222V$ (d) $0.176V$			
	$ \begin{array}{c} (c) & 0.222 \\ \end{array} \\ \begin{array}{c} (c) & 0.222 \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ \end{array} \\ \end{array} \\ \end{array} \\ \begin{array}{c} (c) & 0.170 \\ \end{array} \\ $	$000^{\circ}V$ for	respectively. The $\wedge$	for NaBr is
127.	Calculate the electrode potential at $2$	.98 K 101	(a) 278 $S \ cm^2 mol^{-1}$	(b) 176 $S \ cm^2 mol^{-1}$
	$Zn \mid Zn$ electrode in which the activit	y of zinc	(c) 128 $S cm^2 mol^{-1}$	(d) $302 \ S \ cm^2 mol^{-1}$
	ions is 0.001 <i>M</i> and $E^{o}_{Z_{l}/Z_{l}^{++}}$ is -0.74 vol	ts [AMU 2002]		
	(a) 0. 38 <i>volts</i> (b) 0.83 <i>volts</i>	2.	On the basis of t	the electrochemical theory of
	(c) 0.40 <i>volts</i> (d) 0.45 <i>volts</i>		aqueous corrosion,	the reaction occurring at the
128.	Which of the following expression is cor	rect	catiloue is	[MD DET 1004, UDSEAT 2001]
	[Orissa	JEE 2005]		[MP PET 1994, UPSEAT 2001]
	(a) $\Delta G^{\circ} = -nFE^{\circ}_{cell}$ (b) $\Delta G^{\circ} = +nFE^{\circ}_{cell}$	ell	(a) $O_{2(g)} + 4H_{(aq)} + 4e$	$e^- \rightarrow 2H_2O_{(l)}$
	(c) $\Delta G^{\circ} = -2.303 RT nFE^{\circ}_{coll}$ (d) $\Delta G^{\circ} = -nF \log nFE^{\circ}_{coll}$	$g K_c$	(b) $Fe_{(s)} \rightarrow Fe_{(aq)}^{2+} + 2e$	-
120	For the feasibility of a redox reaction in	a cell the	(c) $E_{2}^{2+}$ $E_{3}^{3+}$	_
29.	e.m.f. should be	& K 2002]	(c) $Fe_{(aq)} \rightarrow Fe_{(aq)} + e$	
	(a) Positive (b) Fixed		(d) $H_{2(g)} + 2OH_{(aa)}^{-} \rightarrow$	$2H_2O_{(l)} + 2e^{-l}$
	(c) Zero (d) Negative	2	The reaction	- ()
	(a) reguire	3.	$\frac{1}{1} H(x) + A - C (x) + H$	$+(a_{2}) + C + (a_{2}) + A_{2} + (a_{3})$
	Corrosion		$\frac{1}{2}H_2(g) + AgCl(s) \rightarrow H$	(aq) + Cl (aq) + Ag(s)
			occurs in the galvar	nic cell
1.	Corrosion is basically a [Kerala (M	[ed.) 2002]	[	IIT 1985; AMU 2002; KCET 2003]
	(a) Altered reaction in presence of $H_2O$		(a) $Ag / AgCl(s) KCl(s)$	oln) $\  AgNO_3 (\text{soln}) / Ag$
	(b) Electrochemical phenomenon		(b) $Pt/H_2(g)HCl$ (solr	$\ AgNO_3(soln)/Ag$

4.

- (b) Electrochemical phenomenon
- (c) Interaction
- (d) Union between light metal and heavy metal
- Rusting of iron is catalysed by which of the 2. following

# [MNR 1990; UPSEAT 2001]

(a) *Fe* (b) O<sub>2</sub> (d)  $H^+$ (c) *Zn* 

Which of the following is a highly corrosive salt 3. [AFMC 2005]  $Fe = Fe^{2+} + 2e^{-}; E^{o} = +0.41 V$ 

 $Zn = Zn^{2+} + 2e^{-}; E^{o} = +0.76 V$ 

reactions are as

(c)  $Pt/H_2(g)HCl(soln) \parallel AgCl(s)/Ag$ 

(d)  $Pt/H_2(g)KCl(soln) \parallel AgCl(s)/Ag$ 

The *EMF* for cell reaction  $Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$  is

The standard reduction potential  $E^{o}$  for the half

	[IIT 1988; CBSE PM	IT 1993, 96; BHU 1995, 2000;		Calculate	$\Lambda^{\infty}_{HOAc}$ usin	ng appropriate	molar
	CPMT 2000; KCET 2000; A	AIIMS 2001; Orissa JEE 2002]		conductance	es of the ele	ctrolytes listed ab	ove at
	(a) $-0.35 V$	(b) $+0.35 V$		infinite dilu	tion in $H_2O$	at 25°C [AIEEE	2005]
	(c) +1.17 V	(d) -1.17 V		(a) 517.2		(b) 552.7	
5٠	The number of electror	ns to balance the following		(c) 390.7		(d) 217.5	
	equation $NO_3^- + 4H^+ + e^-$	$\rightarrow 2H_2O + NO$ is[IIT Screeni	ng 1 <b>991</b> ]	The mass of	carbon ano	le consumed (givin	g only
	(a) 5	(b) 4		carbondioxi	de) in the	production of 270	kg of
	(c) 3	(d) 2		is <b>[CBSE PM</b>	IT 2005]	auxite by the half p	100055
6.	The standard EMF fo	or the given cell reaction		(a) 180 <i>kg</i>		(b) 270 <i>kg</i>	
	$Zn + Cu^{2+} = Cu + Zn^{2+}$ is	$1.10V$ at $25^{\circ}C$ . The EMF		(c) 540 <i>kg</i>		(d) 90 <i>kg</i>	
	for the cell reaction	n, when $0.1M Cu^{2+}$ and	13.	4.5 <i>g</i> of alum	ninium (at n	nass 27 <i>amu</i> ) is dep	osited
	$0.1M Zn^{2+}$ solutions are	used, at $25^{\circ}C$ is		at cathode	from Al <sup>3+</sup>	solution by a c	ertain
	[MNR 199	94; AMU 1999; UPSEAT 2002]		quantity of	electric c	harge. The volur	ne of
	(a) 1.10 V	(b) 0.110 V		solution by	the same ou	antity of electric (	charge
	(c) -1.10 V	(d) -0.110 V		will be	1.		0
7.	A gas X at $1 atm$ is b	ubbled through a solution		(2) 22 4 I		[CBSE PMT	2005]
	containing a mixture	of $1MY^-$ and $1MZ^-$ at		(c) 5.6 L		(d) 11.2 L	
	$25^{\circ}C$ If the reduction	notential of $7 > Y > X$ then	14.	What amound	nt of $Cl_2$ gas	liberated at anod	e, if 1
	(a) Y will oxidize X a	nd not $Z$		amp. currer	it is passed	for 30 min. from	NaCl
	(b) Y will oxidize Z a	nd not X		solution.		[BHU	2005]
	(c) <i>Y</i> will oxidize both	X and $Z$		(a) 0.66 mo	les	(b) 0.33 moles	
	(d) <i>Y</i> will reduce both	X and $Z$		(c) 0.66 g		(d) 0.33 g	
8.	The oxidation potential	of a hydrogen electrode at					
	$pH = 10$ and $pH_1 = 1$	[JIPMER 2000]			cortion	8 Doocon	
	(a) 0.059 V	(b) 0.59 V		R AB	5011011	α πεαδυπ	
	(c) 0.00 V	(d) 0.51 V				For ANMS Asp	irants
9.	The decomposition of	hydrogen peroxide is an			_		
	example of		Read	d the assertion	on and reaso	on carefully to main	rk the
		[Roorkee 2000]	COLL				
	(a) Exothermic reaction	(b) Endothermic	(a)	If both ass	ertion and i	reason are true a	nd the
	(a) Exothermic reaction reaction	n (b) Endothermic	(a)	If both ass reason is the	ertion and a correct expl	reason are true ar anation of the asser	id the tion.
	<ul><li>(a) Exothermic reaction</li><li>reaction</li><li>(c) Negative catalysis</li></ul>	n (b) Endothermic (d) Auto-oxidation	(a) (b)	If both ass reason is the If both asse	ertion and retrion and retrion and retrievely between the second se	reason are true an anation of the asser son are true but rea	nd the tion. Ison is
10.	<ul><li>(a) Exothermic reaction</li><li>(c) Negative catalysis</li><li>Aluminium oxide may</li></ul>	n (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C	(a) (b)	If both ass reason is the If both asse not the corre	ertion and n e correct expl rtion and rea ect explanatio	reason are true ar anation of the asser son are true but rea on of the assertion.	nd the tion. Ison is
10.	<ul> <li>(a) Exothermic reaction</li> <li>reaction</li> <li>(c) Negative catalysis</li> <li>Aluminium oxide may</li> <li>to furnish aluminium n</li> </ul>	<ul> <li>(b) Endothermic</li> <li>(d) Auto-oxidation</li> <li>be electrolysed at 1000°C</li> <li>netal (At. Mass = 27 amu; 1</li> <li>Caulemba) The arthodo</li> </ul>	(a) (b) (c)	If both ass reason is the If both asse not the corre If assertion	ertion and re e correct expl rtion and rea ect explanation is true but re	reason are true and anation of the asser son are true but rea on of the assertion. ason is false.	nd the tion. Ison is
10.	<ul> <li>(a) Exothermic reaction</li> <li>reaction</li> <li>(c) Negative catalysis</li> <li>Aluminium oxide may</li> <li>to furnish aluminium n</li> <li>Faraday = 96,500</li> <li>reaction is</li> </ul>	n (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C netal (At. Mass = 27 <i>amu</i> ; 1 <i>Coulombs</i> ). The cathode	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion	ertion and re e correct expl rtion and rea ect explanation is true but re ion and reason is false but re	reason are true an anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true.	nd the tion. Ison is
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3a^- \rightarrow Al^2$	n (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C netal (At. Mass = 27 <i>amu</i> ; 1 <i>Coulombs</i> ). The cathode	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> </ul>	If both ass reason is the If both asse not the corre If assertion If the assert If assertion	ertion and re e correct expl rtion and rea ect explanation is true but re ion and reason is false but re	reason are true an anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true.	nd the tion. ason is
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5 12kg of	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 <i>amu</i> ; 1 <i>Coulombs</i> ). The cathode	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion Assertion :	ertion and rea e correct expl rtion and rea ect explanation is true but re ion and reason is false but re Sodium io	reason are true ar anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true.	nd the tion. ason is
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 <i>amu</i> ; 1 <i>Coulombs</i> ). The cathode aluminium metal by this [AIEEE 2005]	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion Assertion :	ertion and re e correct expl rtion and rea ect explanation is true but re ion and reaso is false but re Sodium ion preference mercury ca	reason are true an anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true. ons are discharg to hydrogen ions	ed in
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) 5.49 × 10 <sup>7</sup> C of electr	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 <i>amu</i> ; 1 <i>Coulombs</i> ). The cathode aluminium metal by this [AIEEE 2005] icity	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion Assertion : Reason :	ertion and rea e correct expl rtion and rea ect explanation is true but re ion and reaso is false but re Sodium ion preference mercury ca The nature	reason are true ar anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true. ons are discharg to hydrogen ions thode. of the cathode can	ed in effect
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10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) $5.49 \times 10^7 C$ of electr (b) $1.83 \times 10^7 C$ of electr (c) $5.49 \times 10^4 C$ of electr (d) $5.49 \times 10^1 C$ of electr	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 amu; 1 Coulombs). The cathode aluminium metal by this [AIEEE 2005] icity ricity icity icity	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> <li>2.</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion Assertion : Reason : Assertion :	ertion and re ecorrect expl rtion and rea ect explanation is true but re ion and reaso is false but re Sodium ion preference mercury ca The nature the order on In electro electricity mole of sil	reason are true an anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true. ons are discharg to hydrogen ions thode. of the cathode can of discharge of ions. olysis, the quanti needed for deposi ver is different from	ed in ed in effect ty of ting 1 n that
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) $5.49 \times 10^7 C$ of electr (b) $1.83 \times 10^7 C$ of electr (c) $5.49 \times 10^4 C$ of electr (d) $5.49 \times 10^1 C$ of electr	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 amu; 1 <i>Coulombs</i> ). The cathode aluminium metal by this [AIEEE 2005] icity icity icity icity	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> <li>2.</li> </ul>	If both ass reason is the If both asser not the corre If assertion If the assert If assertion Assertion : Reason : Assertion :	ertion and re e correct expl rtion and rea ect explanatio is true but re ion and reaso is false but re Sodium io preference mercury ca The nature the order o In electro electricity mole of sil required fo	reason are true at anation of the asser son are true but rea- on of the assertion. ason is false. on both are false. cason is true. ons are discharg to hydrogen ions thode. of the cathode can of discharge of ions. olysis, the quanti needed for deposi ver is different from or 1 mole of copper.	ed in ed in effect ty of ting 1 n that
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) $5.49 \times 10^7 C$ of electr (b) $1.83 \times 10^7 C$ of electr (c) $5.49 \times 10^4 C$ of electr (d) $5.49 \times 10^1 C$ of electr Electrolyte: <i>KCl</i>	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 amu; 1 Coulombs). The cathode aluminium metal by this [AIEEE 2005] icity ficity ficity icity icity icity	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> <li>2.</li> </ul>	If both ass reason is the If both asset not the corre If assertion If the assert If assertion Assertion : Reason : Assertion : Reason :	ertion and rea ecorrect expl rtion and rea ect explanation is true but re ion and reaso is false but re Sodium ion preference mercury ca The nature the order on In electro electricity mole of sill required for The molecu	reason are true an anation of the asser son are true but rea on of the assertion. ason is false. on both are false. cason is true. ons are discharg to hydrogen ions thode. of the cathode can of discharge of ions. olysis, the quanti needed for deposi ver is different from or 1 mole of copper. alar weights of silv different	ed the tion. ason is ed in a at a effect ty of ting 1 n that er and
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) $5.49 \times 10^7 C$ of electr (b) $1.83 \times 10^7 C$ of electr (c) $5.49 \times 10^4 C$ of electr (d) $5.49 \times 10^1 C$ of electr (d) $5.49 \times 10^1 C$ of electr (electrolyte: <u>KCl</u> <u>A<sup>x</sup>(Scm<sup>2</sup>mol<sup>2</sup></u> 149.	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 amu; 1 Coulombs). The cathode aluminium metal by this [AIEEE 2005] icity	(a) (b) (c) (d) (e) <b>1.</b> <b>2.</b> <b>3.</b>	If both ass reason is the If both asser not the corre If assertion If the assert If assertion Assertion : Reason : Assertion : Reason : Assertion :	ertion and re e correct expl rtion and rea ect explanatio is true but re ion and reaso is false but re Sodium io preference mercury ca The nature the order o In electro electricity mole of sil required fo The molecu copper are Equivalent	reason are true an anation of the asser son are true but rea- on of the assertion. ason is false. on both are false. eason is true. ons are discharge to hydrogen ions thode. of the cathode can of discharge of ions. olysis, the quanti needed for deposi ver is different from or 1 mole of copper. alar weights of silv different conductance o	ed the tion. son is ed in at a effect ty of ting 1 n that er and f all
10.	(a) Exothermic reaction reaction (c) Negative catalysis Aluminium oxide may to furnish aluminium m <i>Faraday</i> = 96,500 reaction is $Al^3 + 3e^- \rightarrow Al^\circ$ To prepare 5.12kg of method would require (a) $5.49 \times 10^7 C$ of electr (b) $1.83 \times 10^7 C$ of electr (c) $5.49 \times 10^4 C$ of electr (d) $5.49 \times 10^1 C$ of electr (d) $5.49 \times 10^1 C$ of electr <u>Electrolyte</u> : <u><i>KCl</i></u> $\Lambda^{\alpha}(Scm^2mol^-$ 149.	h (b) Endothermic (d) Auto-oxidation be electrolysed at 1000°C hetal (At. Mass = 27 amu; 1 Coulombs). The cathode aluminium metal by this [AIEEE 2005] icity ricity ricity icity $\frac{KNO_3 HCl NaOAc NaCl}{145. 426. 91.0 126.}$	<ul> <li>(a)</li> <li>(b)</li> <li>(c)</li> <li>(d)</li> <li>(e)</li> <li>1.</li> <li>2.</li> <li>3.</li> </ul>	If both ass reason is the If both asser not the corre If assertion If the assert If assertion Assertion : Reason : Assertion : Reason : Assertion :	ertion and rea ecorrect expl rtion and rea ect explanation is true but re- ion and reaso is false but re- Sodium ion preference mercury ca The nature the order on In electro electricity mole of sill required for The molecu copper are Equivalent electrolyte	reason are true ar anation of the asser son are true but rea on of the assertion. ason is false. on both are false. eason is true. ons are discharg to hydrogen ions thode. of the cathode can of discharge of ions. olysis, the quanti needed for deposi ver is different from or 1 mole of copper. alar weights of silv different conductance of s decreases with	ed the tion. ason is ed in a at a effect ty of ting 1 n that er and f all n the

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sodium chloride is given by the

	Reason :	Lesser number of ions are available per gram equivalent at higher
4.	Assertion :	Concentration. Copper reacts with hydrochloric acid and liberates hydrogen from
		the solution of dilute hydrochloric acid.
	Reason :	Hydrogen is below copper in the
5.	Assertion :	Copper liberates hydrogen from a solution of dil. <i>HCl</i> .
	Reason :	Hydrogen is below copper in the
6.	Assertion :	Zn metal is formed when a $Cu$ plate in dipped in $ZnSO_4$ solution.
	Reason :	<i>Cu</i> being placed above <i>Zn</i> in electrochemical series.
7.	Assertion :	Electrical conductivity of copper increases with increase in
	Reason :	The electrical conductivity of metals is due to the motion of
		electrons.
8.	Assertion :	A small amount of acid or alkali is added before electrolysis of water.
	Reason :	Pure water is weak electrolyte.
9.	Assertion :	Copper reacts with <i>HCl</i> and liberates bydrogen
	Reason :	Hydrogen is present above <i>Cu</i> in the reactivity series.
10.	Assertion : cells.	K and Cs are used in photoelectric
	Reason :	<i>K</i> and <i>Cs</i> emit electrons on exposure to light.
11.	Assertion :	A large dry cell has high <i>e.m.f</i> .
	Reason :	The <i>e.m.f.</i> of a dry cell is proportional to its size.
12.	Assertion :	The resistivity for a substance is its resistance when its is one meter long and its area of cross section is one square meter
	Reason :	The <i>SI</i> units of resistivity are ohm metre $(\Omega m)$ and ohm centimeter $(\Omega cm)$ .
13.	Assertion :	When acidified zinc sulphate solution is electrolysed between zinc electrodes, it is zinc that is deposited at the cathode and hydrogen evolution does not take
	Reason :	The electrode potential of zinc is more negative than hydrogen as the overvoltage for the hydrogen as the evolution on zinc is quite large.
14.	Assertion :	If $\lambda_{Na^+}^0 + \lambda_{Cl^-}^0$ are molar limiting
		conductivity of the sodium and chloride ions respectively, then the limiting molar conducting for

	equation: $\wedge^0_{NaCI} = \lambda^0_{Na^+} + \lambda^0_{CI^-}$ .
Reason :	This is according to Kohlrausch law
Accortion :	One coulomb of electric charge
Assertion .	deposite weight equal to the
	deposits weight equal to the
	electrochemical equivalent of the
	substance.
Reason :	One Faraday deposits one mole of
	the substance.
Assertion :	Auric chloride $(AuCl_3)$ solution
	cannot be stored in a vessel made of
	copper, iron, nickel, chromium, zinc
	or tin.
Reason :	Gold is a very precious metal.
Assertion :	For a cell reaction $Zn(s) + Cu^{2+}(aa) \rightarrow$
	$Zn^{2+}(aa) + Cu(s)$ : at the equilibrium
	voltmeter gives zero reading
Reason ·	At the equilibrium there is no
Reason .	abanga in the concentration of
	$Cu^{2+}$ and $Zn^{2+}$ ions.
Assertion :	A negative value of standard
	reduction potential means that
	reduction take place on this
	electrode with reference to
	standard hydrogen electrode.
Reason :	The standard electrode potential of
	a half cell has a fixed value.
Assertion :	Weston is a standard cell.
Reason :	Its <i>e.m.f.</i> does not change with
	temperature.
Assertion :	Galvanic cells containing hydrogen,
	methane, methanol etc. as fuels are
	called fuel cells.
Reason :	They are designed to convert the
	energy of combustion of fuels
	directly into electrical energy.
Assertion ·	Zinc displaces copper from copper
1350111011 .	sulphate solution.
Reason :	The $E^0$ of zinc is $-0.76$ V and that of
	copper is +0.34 V.
Accortion .	Identification of cathodo and anodo
Assertion .	ic done by the use of a
	is done by the use of a
_	thermometer.
Reason :	Higher is the value of reduction
	potential, greater would be its
	reducing power.
Assertion :	An electrochemical cell can be set
	up only if the redox reaction is
	spontaneous.
Reason :	A reaction is spontaneous if free
	energy change is negative.
Assertion :	Galvanised iron does not rust.

24. Assertion : Galvanised iron does not rust. Reason : Zinc has a more negative electrode potential than iron. [AIIMS 2005]

25.	Assertion :	In an electrochemical cell anode and cathode are respectively negative and positive electrodes.							
	Reason :	At anode oxidation takes place and at cathode reduction takes place.							
26.	Assertion :	$Ni/Ni^{2+}(1.0M) \parallel Au^{3+}(1.0M) \parallel Au$ , for							
		this cell emf is 1.75 V if							
		$E^o_{Au^{3+}/Au} = 1.50$ and $E^o_{Ni^{2+}/Ni} = 0.25 V$ .							
	Reason :	Emf of the cell $= E_{\text{cathode}}^o - E_{\text{anode}}^o$ .							
27.	Assertion :	Salts like <i>KCl</i> , <i>KNO</i> <sub>3</sub> i.e., inert							
		electrolytes are used in salt bridge.							
	Reason :	An inert electrolyte can easily be filled in the U-tube.							
28.	Assertion :	Emf and potential difference are same for cell.							
	Reason :	Both gives the difference in electrode potential under any condition.							

Answers

# **Electrolytes and Electrolysis**

1	b	2	d	3	c	4	с	5	d
6	d	7	b	8	b	9	b	10	а
11	а	12	C	13	d	14	d	15	C
16	d	17	а	18	С	19	С	20	а
21	а	22	b	23	b	24	d	25	b
26	b	27	b	28	a	29	c	30	b
31	d	32	d	33	С	34	b	35	b
36	а	37	d	38	с	39	d	40	a
41	a								

# Faraday's law of electrolysis

1	с	2	а	3	b	4	с	5	с
6	С	7	C	8	b	9	C	10	a
11	C	12	а	13	C	14	b	15	b
16	C	17	a	18	а	19	d	20	C
21	d	22	C	23	b	24	d	25	а
26	d	27	а	28	d	29	C	30	а
31	b	32	b	33	а	34	b	35	b
36	b	37	b	38	b	39	a	40	а

41	с	42	а	43	а	44	b	45	с
46	b	47	d	48	d	49	b	50	а
51	C	52	b	53	b	54	а	55	C
56	c	57	c	58	d	59	d	60	d
61	C	62	C	63	b	64	а	65	b
66	b	67	а	68	а	69	b	70	C
71	C	72	b	73	a	74	b	75	C
76	b	77	C	78	С	79	b	80	b
81	C	82	b	83	а	84	C	85	C
86	е	87	b	88	а	89	d	90	C
91	C	92	а	93	C	94	b		

# **Conductor and Conductance**

1	b	2	b	3	d	4	d	5	а
6	b	7	b	8	а	9	b	10	d
11	а	12	b	13	b	14	b	15	b
16	d	17	b	18	d	19	b	20	C
21	b	22	а	23	C	24	d	25	a
26	b	27	d						

# **Cell constant and Electrochemical cells**

1	d	2	d	3	d	4	b	5	b
6	а	7	C	8	C	9	b	10	b
11	C	12	C	13	a	14	a	15	d
16	C	17	C	18	a	19	b	20	b
21	b	22	b	23	a	24	b	25	C
26	a	27	С	28	a	29	b	30	c
31	a	32	b	33	c	34	b	35	d
36	b	37	C	38	d	39	b	40	d
41	a	42	d	43	C	44	a	45	d
46	a	47	d	48	b	49	C	50	C
51	b	52	d	53	C	54	а	55	C
56	d	57	b	58	а	59	а	60	C
61	а								

# Electrode potential, E<sub>Cell</sub>, Nernst equation and ECS

1	b	2	а	3	а	4	b	5	d
6	С	7	C	8	a	9	а	10	b
11	а	12	C	13	а	14	d	15	C
16	а	17	а	18	С	19	b	20	b
21	а	22	b	23	а	24	C	25	а
26	a	27	C	28	a	29	d	30	d
31	b	32	а	33	d	34	а	35	C
36	d	37	a	38	b	39	d	40	d

41	b	42	С	43	a	44	b	45	b
46	d	47	С	48	b	49	C	50	b
51	b	52	а	53	a	54	c	55	a
56	a	57	а	58	a	59	c	60	a
61	b	62	d	63	a	64	d	65	b
66	C	67	b	68	b	69	а	70	b
71	b	72	ab	73	c	74	с	75	a
76	d	77	b	78	C	79	b	80	a
81	a	82	b	83	b	84	С	85	a
86	C	87	а	88	C	89	С	90	b
91	d	92	C	93	d	94	а	95	d
96	a	97	d	98	a	99	b	100	d
101	a	102	C	103	a	104	а	105	d
106	b	107	а	108	b	109	а	110	a
111	а	112	b	113	b	114	b	115	b
116	b	117	d	118	a	119	d	120	b
121	c	122	а	123	a	124	b	125	b
126	b	127	b	128	а	129	а		

# Corrosion

3

1 b

2

d

d 4

а

# **Critical Thinking Questions**

1	C	2	a	3	C	4	b	5	C
6	а	7	а	8	b	9	acd	10	а
11	C	12	d	13	C	14	С		

# Assertion & Reason

1	а	2	b	3	а	4	d	5	d
6	d	7	е	8	а	9	е	10	а
11	d	12	b	13	а	14	а	15	С
16	b	17	а	18	е	19	а	20	а
21	а	22	d	23	b	24	а	25	а
26	а	27	C	28	d				



# **Electrolytes and Electrolysis**

- (b) Sugar solution does not form ion; hence does not conduct electricity in solution.
- **3.** (c) Strong electrolytes are almost completely ionised in polar solvent.
- 7. (b) The reduction potential of Mg is less than that of water ( $E^{o} = -0.83V$ ). Hence their ions in the aqueous solution cannot be reduced instead water will be reduced

 $2H_2O+2e^- \rightarrow H_2+2OH^-$  .

- **8.** (b) *HCl* is an electrolyte.
- **9.** (b) Water is reduced at the cathode and oxidized at the anode instead of  $Na^+$  and  $SO_4^{2-}$ .

Cathode:  $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ 

Anode: 
$$H_2 O \to 2H^+ + \frac{1}{2}O_2 + 2e^-$$
.

- (a) In electrolysis process oxidation occurs at anode and reduction occurs at cathode.
- **11.** (a) Because in it covalent bonding is present.
- **12.** (c) According to Faraday's law.
- **13.** (d) Impure metal made anode while pure metal made cathode.
- 14. (d) In electrolytic cell, cathode acts as source of electrons.
- **15.** (c)  $AgNO_3$  is an electrolyte.
- 17. (a) At cathode:  $2H^+ + 2e \rightarrow H_2$ ,

At anode :  $2OH^- \rightarrow H_2O + \frac{1}{2}O_2 + 2e$ 

- **18.** (c) In between dilute  $H_2SO_4$  and platinum electrode  $O_2$  gas evolve at anode.
- **19.** (c) When polar solvent added in to solid electrolyte than it is ionised.
- **20.** (a) In fused *NaCl* chloride ions are oxidized at anode and it is called oxidation.
- **21.** (a) w = zit, Q = it.
- **22.** (b)  $2H^+ + 2e^- \rightarrow H_{2(g)}$  at cathode.
- **23.** (b)  $Na^+_{+1} + e^- \rightarrow Na_0$ , means oxidation number is decreased so the reaction is reduction.
- **24.** (d) Degree of dissociation of weak electrolyte increases on increasing temperature.
- **25.** (b) Since discharge potential of water is greater than that of sodium so water is reduced at cathode instead of  $Na^+$

Cathode: 
$$H_2O + e^- \rightarrow \frac{1}{2}H_2 + OH^-$$
  
Anode:  $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$ .

- **26.** (b) Because it does not have ions.
- **27.** (b)  $NaCl \Rightarrow Na^+ + Cl^-$ .

A

**28.** (a) Cathode:  $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ 

anode : 
$$H_2 O \to 2H^+ + \frac{1}{2}O_2 + 2e^-$$

- **29.** (c) Electrolysis use for electroplating and electrorefining.
- 32. (d) The degree of ionization depend upon the nature of the solute the size of the solute molecules and the concentration of the solution.
- **33.** (c)  $C_{12}H_{22}O_{11}$  is an non-electrolyte.
- **34.** (b) On electrolysis molten ionic hydride liberate  $H_2$  at the anode.
- **35.** (b) During electrolysis cation discharged at cathode and anion discharged at anode.
- **36.** (a) Calcium is produces when molten anhydrous calcium chloride is electrolysed.
- **37.** (d) All metals conducts heat and electricity.
- **38.** (c)  $2Al + dil H_2 SO_4 \rightarrow Al_2 SO_4 + H_2 \uparrow$ .
- **39.** (d) Generally fussed potassium chloride flow the electric conductivity.
- 41. (a) The substances whose aqueous solutions allow the passage of electric current and are chemically decomposed, are termed electrolytes. Electrolytic substances are classified as strong or weak according to how readily they dissociate into conducting ions. Acetic acid is a weak electrolyte. Glucose, ethanol and urea are non-electrolytes.

# Faraday's law of electrolysis

**1.** (c) 
$$Ag^+ + e^- \to Ag; E_{Ag} = \frac{Atomic \ Mass}{1} = 108$$

Number of faraday 
$$= \frac{W_{Ag}}{E_{Ag}} = \frac{108}{108} = 1$$
.

2. (a) 
$$W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{108 \times 9.65}{96500}$$
  
= 1.08 × 10<sup>-2</sup> gm = 10.8 mg

3

5٠

**3.** (b) 
$$Fe^{2+} + 2e^- \rightarrow Fe$$
;  $E_{Fe} = \frac{56}{2} = 28$ 

 $W_{Fe} = E_{Fe}$  × Number of faraday = 28 × 3 = 84 gm.

4. (c) 
$$W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{107.87 \times 965}{96500} = 1.0787 \ gm$$
.

(c) 
$$Al^{3+} + 3e^- \rightarrow Al$$
  
 $E_{Al} = \frac{27}{3} = 9$   
 $W_{Al} = E_{Al} \times \text{No. of faradays} = 9 \times 5 = 45 \text{ gm}.$ 

**6.** (c) *Cu* voltameter or *Cu* or *Ag* coulometer are used to detect the amount deposited on an electrode during passage of know charge through solution.

8. (b) Weight of 
$$L_2 = \frac{Eq. weight of Cu}{Eq. weight. of H}$$
  
Weight of  $L_2 = \frac{63.6/2}{1}$   
Weight of  $Cu = 15.9 \ gm.$   
9. (c)  $Cu^{2+} + 2e^- \rightarrow Cu$   
2 Faradays will deposit  
 $= 1 g \operatorname{atom}$  of  $Cu = 63.5 g$ .  
12. (a) At cathode;  $Al^{3+} + 3e^- \rightarrow Al$   
 $E_{Al} = \frac{27}{3} = 9$   
 $W_{Al} = E_{Al} \times \operatorname{No.of} \operatorname{faradays} = 9 \times 0.1 = 0.9 \ gm.$   
14. (b)  $W = zit$ ;  $W = \frac{32.69 \times 5 \times 60 \times 40}{96500} = 4.065 \ gm.$   
15. (b)  $m = Z \times 4 \times 120$ ;  $M = Z \times 6 \times 40$   
 $\frac{M}{m} = \frac{6 \times 40}{4 \times 120} = \frac{1}{2}$ ;  $M = m/2$ .  
16. (c)  $W_{meal} = \frac{E \times I \times i}{96500} = \frac{96500 \times 1.8}{3 \times 50 \times 60} = 19.3$ .  
17. (a)  $Al \rightarrow Al^{3+} + 3e^-$ .  
The charged obtained is  $3 \times 96500 \ C.$   
18. (a) Wt. of  $Ag$  deposited  $= Eq. wt. of \ Ag = 108 \ gm$   
Wt. of  $Cr$  deposited  $= Eq. wt. of \ Cr = 17.3 \ gm.$   
19. (d) One Faraday = 1  $gm$  of equivalent of  $Cu$ .  
20. (c)  $W = Zit$ ;  $Z = \frac{E}{96500}$ .  
21. (d) During electrolysis of  $CuSO_4 \cdot Cu^{2+}$  gets  
discharged at cathode and  $OH^-$  at anode. Thus  
solution becomes acidic due to excess of  $H^+$   
and  $SO_4^{2-}$  or  $H_2SO_4$ .  
23. (b) 1 mole of electrons = 1  $faraday$   
 $Mg^{++} + 2e^- \rightarrow Cu$   
 $E_{Cu} = \frac{63.54}{2} = 31.77$   
Amount of electricity required to deposit  
 $.6354 \ gm$ 

$$=\frac{96500 \times 0.6354}{31.77} = 1930 \ Coulombs \ .$$

**25.** (a) The amount deposited is directly proportional to current intensity, electrochemical equivalent of ions and the time for electrolysis and is independent of the temperature.

**27.** (a) 
$$W = ZQ$$
;  $W = Zit$ .

**28.** (d) 
$$Ca^{++} + 2e^{-} \rightarrow Ca$$
  
 $E_{Ca} = \frac{40}{2} = 20$   
 $W_{Ca} = E_{Ca} \times \text{No. of faradays} = 20 \times 0.04 = 0.8 \text{ gm}$ .

29. (c)  $E_{\text{metal}} = \frac{\text{Weight of metal} \times 96500}{\text{Number of coulombs}}$ =  $\frac{22.2 \times 96500}{100} = 59.5$ 

$$2 \times 5 \times 60 \times 60$$
  
Oxidation number of the metal  $= \frac{177}{59.5} = +3$ 

**30.** (a) Quantity of electricity passed = 
$$\frac{25}{1000} \times 60 = 1.5$$

 $2F = 2 \times 96500 C$  deposit Ca = 1 mole

$$\therefore 1.5 C \text{ will deposit } Ca = \frac{1}{2 \times 96500} \times 1.5 \text{ mole}$$
$$= \frac{1}{2 \times 96500} \times 1.5 \times 6.023 \times 10^{23} \text{ atom} = 4.68 \times 10^{18} \text{ .}$$

**31.** (b) Equivalent of *Cl* deposited = No. of *Faraday* passed = 0.5 Wt. of  $Cl = 0.5 \times Eq.wt$ . =  $0.5 \times 35.5 = 17.75 gm$ .

32. (b) At Andoe At Anode 
$$Cl^- \to \frac{1}{2}Cl_2 + e^-$$
  
 $E_{--} = \frac{35.5 \times 2}{2} = 35.5$ 

$$E_{Cl_2} = \frac{1}{2} = 35.5$$
$$W_{Cl_2} = \frac{E_{Cl_2} \times I \times t}{96500} = \frac{35.5 \times 2 \times 30 \times 60}{96500} = 1.32 \ gm$$

- **34.** (b) It is Faraday's law.
- **35.** (b) Equivalent wt. of  $O_2$  = Equivalent wt. of Cu

36. (d) 
$$O_2 \% = 20\%$$
  
Metal $\% = 80\% = \frac{80}{20} \times 8 = 32 \text{ g of metal}$ .  
27. (b)  $V = \frac{827 \times 10^3}{20} = 2.14 \text{ V}$ 

**37.** (b) 
$$V = \frac{827 \times 10}{4 \times 96500} = 2.14 V$$
.

- **38.** (b)  $Ag^+ \xrightarrow{+e^-} Ag$ , 96500 *C* will liberate silver = 108 *gm*. 9650*C* will liberate silver = 10.8 *gm*.
- 39. (a) One mole of monovalent metal ion means charge of N electron i.e. 96500 C or 1 Faraday.
- **42.** (a) 1 Faraday involves charge of 1 *mole* electrons.
- **43.** (a) Coulomb = ampere (A) × second (S).

44. (b) 
$$E = -\frac{13.6}{n^2}$$
 for  $He^+ n = 1$   
 $E = -\frac{13.6}{1^2} = -13.6 \, eV$ .

- **45.** (c)  $w \propto E$  if *i* and *t* are constant.
- 47. (d) Charge (Coulombs) pass per second =  $10^{-6}$ number of electrons passed per second =  $\frac{10^{-6}}{1.602 \times 10^{-19}} = 6.24 \times 10^{12}$ .

**48.** (d) At cathode;  

$$Fe^{2^+} + 2e^- \rightarrow Fe \; ; \; Fe^{3^+} + 3e^- \rightarrow Fe$$
  
 $(E_{Fe})_1 = \frac{\text{Atomic.weight}}{2}; (E_{Fe})_2 = \frac{\text{Atomic.weight}}{3}$   
Ratio of weight of *Fe* liberated  
 $= \frac{\text{Atomic weight}}{3} : \frac{\text{Atomic weight}}{2} = 3 : 2.$ 

**49.** (b) 31.75 *g* copper gets deposited at cathode on passing 96500 *coulomb* charge. We know that

31.75 gm of Cu is equal to 0.5 mole of Cu deposited at cathode on passing 1F of current. (b) For deposition of one equivalent silver 52. required charged is 96500 C. (b)  $Cu^{++} + 2e^- \rightarrow Cu$ ;  $E_{Cu} = \frac{63.55}{2} = 31.75 \ gm \ Cu$ . 53. (a)  $Q = 2.5 \times 386 = 96500 C$ 54.  $2F(2 \times 96500 C)$  deposited Cu = 63.5 g: Hence 965 C will deposited; Cu = 0.3175 gm. (c)  $\frac{\text{Wt.of } Cu}{\text{Wt.of } Ag} = \frac{\text{Eq. wt.of } Cu}{\text{Eq. wt.of } Ag} ; \frac{\text{Wt. of } Cu}{1.08} = \frac{63.5/2}{108}$ 55. Wt. of  $Cu = 0.3177 \ gm$ . 56. (c) 1 g atom of Al = 3 equivalent of Al = 3 faraday charge 3 *mole* electrons = 3 *N* electron. (c) At cathode :  $Al^{3+} + 3e^- \rightarrow Al$ 57.  $E_{Al} = \frac{\text{Atomic mass}}{3}$ At cathode :  $Cu^{2+} + 2e^- \rightarrow Cu$  $E_{Cu} = \frac{\text{Atomic mass}}{2}$ At cathode :  $Na^+ + e^- \rightarrow Na$  $E_{Na} = \frac{\text{Atomic mass}}{1}$ For the passage of 3 faraday; mole atoms of Al deposited = 1 mole atoms of *Cu* deposited  $=\frac{1\times3}{2}=1.5$ mole atoms of *Na* deposited  $= 1 \times 3 = 3$ . (d) At cathode:  $Ag^+ + e^- \rightarrow Ag$ 58. At Anode:  $2OH^- \rightarrow H_2O + \frac{1}{2}O_2 + 2e^ E_{Ag} = \frac{108}{1} = 108; E_{O_2} = \frac{\frac{1}{2} \times 32}{2} = 8$  $\frac{W_{Ag}}{E_{Ag}} = \frac{W_{O_2}}{E_{O_2}}; \ W_{Ag} = \frac{1.6 \times 108}{8} = 21.6 \ gm .$ (d) *KI* is an electrolyte. 59. (d) Number of gm equivalent = Number of 60. faraday pass 4 gm = 4 faraday. (c) Eq. of  $Al = \frac{13.5}{27/3} = 1.5$ . 61. Thus 1.5 Faraday is needed. (b) Electricity required 63. = No. of gm equivalent × 96500 coulombs  $= 0.5 \times 96500 = 48250 C.$ 64. (a) Equivalent weight of silver = 107.870 g. 1 Faraday = 96500 *coulomb*. 67. (a) Equivalent weight and atomic weight of Na

metal are the same, so 1q atom of Na is deposited by one Faraday of current.

(a)  $Al \to Al^{3+} + 3e^{-}$ . 68.

**70.** (c) :: 1*F* obtained from 1 *g* equivalent

- $\therefore$  2.5 *F* obtained from 2.5 *q* equivalent.
- (c) Faraday constant depends upon the current 75. passed.
- (b) In 5 gm CuO, 4 gm Cu and 1 gm O be present. 80.

Eleme nt	Wt.	At Wt.	$Wt./At.Wt. \neq x$	Ratio
Си	4 gm	63.5	4/63.5=.0625	$\frac{.0625}{.0625} = 1$
0	1 gm	16	1/16 =.0625	$\frac{.0625}{.0625} = 1$

Emperical formula = CuO of oxide In this oxide, oxidation no. of Cu = +2Equivalent weight  $= \frac{\text{Molecular weight}}{\text{Oxidation no.}} = \frac{63.5}{2} \approx 31.75 \text{ but Equivalent}$ weight should be an integeral no. = 32 (c) Given, Current = 241.25 columb 1 coulomb current will deposite  $= 1.118 \times 10^{-3} gm Ag$ . ∴ 241.25 current will deposite  $= 1.118 \times 10^{-3} \times 241.25$ = 0.27 gm silver. 82. (b) Reaction for electrolysis of water is  $2H_2O \rightleftharpoons 4H^+ + 2O^{2-}$  $2O^{2-} \rightarrow O_2 + 4e^ 4e^- + 4H^+ \rightarrow 2H_2$  $\therefore$  *n* = 4 so 4 Faraday charge will liberate 1 mole = 22.4  $dm^3$  oxygen ∴ 1 Faraday charge will liberate  $\frac{22.4}{4} = 5.6 \ dm^3 \ O_2$ . **83.** (a)  $Na^+ + e^- \to Na$ 

81.

Charge (in F) = moles of  $e^-$  used = moles of Na deposited  $=\frac{11.5}{23}$  gm = 0.5 Faraday.

**84.** (c) Hydrolysis of water :  $2H_2O \Rightarrow 4H^+ + 4e^- + O_2$ 

4 F charge will produce = 1 mole  $O_2 = 32 \ gm \ O_2$ 

1 *F* charge will produce  $=\frac{32}{4}=8 gm O_2$ .

- (c) In a galvanic cell, the electrons flow from 85. anode to cathode through the external circuit. At anode (-ve pole) oxidation and at cathode (+ pole) reduction takes place.
- 86. (e) Number of equivalents of silver formed =Number of equivalents of copper formed. In  $AgNO_3$ , Ag is in +1 oxidation state.

In  $CuSO_4$ , Cu is in +2 oxidation state.

Equivalent weight of  $Ag = \frac{108}{1} = 108$ 

Equivalent weight of 
$$Cu = \frac{65.0}{2} = 31.8$$

$$\frac{M_1}{M_2} = \frac{E_1}{E_2} ; \frac{10.79}{M_{Cu}} = \frac{108}{31.8}$$
$$M_{Cu} = \frac{10.79 \times 31.8}{108} = 3.2 \ gm .$$

- (b) Laws of electrolysis were proposed by 87. Faraday.
- (a) Given, Current (i) = 25 mA = 0.025 A 88. Time (*t*) = 60 *sec* 
  - $Q = i t = 60 \times 0.025 = 1.5$  coulombs  $1.5 \times 6.023 \times 10^{23}$

No. of electrons = 
$$\frac{1.5 \times 0.025 \times 10}{96500}$$

$$e^{-} = 9.36 \times 10^{18}$$

 $Ca \rightarrow Ca^{2+} + 2e^{-}$ 

 $2e^-$  are required to deposite one *Ca* atom

 $9.36 \times 10^{18} e^{-10}$  will be used to deposite  $=\frac{9.36\times10^{18}}{2}$  = 4.68×10<sup>18</sup>.

89. (d) 
$$C_6H_5NO_2 + 6H^+ + 6e^- \rightarrow C_6H_5NH_2 + 2H_2O_1$$
 mole = 123 *am* nitrogen requires 6

123 gm nitrogen requires 6 mole 1 mole electron e

 $=6 \times 96500$  coulomb charge

123

- ∴ 12.3 *qm* nitrobenzene will require  $=\frac{6\times96500\times12.3}{6\times9650}=6\times9650=57900\ C.$
- (c) Au and Ag settle down below the anode as 90. anode mud during the process of electrolytic refining of copper.

# **Conductor and Conductance**

2. (b) 
$$\lambda^{\infty} BaCl_2 = \frac{1}{2}\lambda^{\infty}Ba^{2+} + \lambda^{\infty}Cl^{-}$$
  
=  $\frac{127}{2} + 76 = 139.5 \ ohm^{-1} \ cm^{-1} \ eq^{-1}$ .

- nature 3. (d) Dilution, temperature and of electrolyte affect the conductivity of solution.
- (a) Generally strong electrolyte on dilution shows 5. conductivity characters.
- (b) Molar conductivity =  $\frac{1000}{MX}$ 6.

7. (b) 
$$C = \frac{K[A]A}{l}, K = \frac{C \times l}{[A]A} = \frac{Sm}{mol \ m^{-3} \ m^2} = Sm^2 mol^{-1}.$$

- (b) Conductivity of a solution is directly 9. proportional to the number of ions.
- (a)  $NaCl \Rightarrow Na^+ + Cl^-$ . So it conduct electricity. 11.
- (b) Graphite is a good conductor of electricity. 12.
- (b) Electrolytic conduction resistance decreases 15. with increasing temperature.

- 16. (d) Because conductance is increase when the dissociation is more.
- (b) Strong electrolyte ionize completely at all 17. dilutions and the number of ions does not increase on dilution. A small increase in  $\wedge_m$ volume with dilution is due to the weakening of electrostatic attraction between the ions on dilution.
- (d) In electrolytic conductors, a single stream of 18. electrons flow from cathode to anode.
- (b) In solid state NaCl does not dissociate into 19. ions so it does not conduct electricity.
- (c) The ions are not free to move in solid state 20. and held up in lattice due to strong coulombic forces of attraction.
- (b)  $C_2H_5OH$  being non electrolyte so does not 21. ionize.
- (a) Since molar conductance  $\propto \frac{1}{\text{Molarity}}$ . 22.

**23.** (c) Molar condcutivity 
$$=\frac{1}{\rho M}$$

So its unit will be  $\Omega^{-1} cm^2 mol^{-1}$ .

**25.** (a) 
$$l/a = 0.5 \ cm^{-1}$$
,  $R = 50 \ ohm$ 

$$p = \frac{Ra}{l} = \frac{50}{0.5} = 100$$
$$\Lambda = k \times \frac{1000}{N} = \frac{1}{p} \times \frac{1000}{N} = \frac{1}{100} \times \frac{1000}{1}$$
$$10 \ ohm^{-1} cm^{2} \ gm \ eq^{-1}$$

26. (b) 
$$\Lambda^{o}_{m(C_{6}H_{5}COOH)} = \Lambda^{o}_{(C_{6}H_{5}COO^{-})} + \Lambda^{o}_{(H^{+})}$$
  
= 42 + 288.42 = 330.42  
 $\alpha = \frac{\Lambda^{c}_{m}}{\Lambda^{o}_{m}} = \frac{12.8}{330.42} = 3.9\%$ 

(d) Conductance =  $\frac{1}{1}$ 27. resistance or mho

# Cell constant and Electrochemical cells

- (d) In the absence of electric field the ions in the 1. solution move randomly due to thermal energy.
- (d) Since  $E^o_{A/A^-}$  has large negative value, the 2. tendency of A to be reduced to  $A^-$  is very small. In other words tendency of  $A^-$  to be oxidized to A is very large.
- (d) Practically only 60-70% efficiency has been 3. attained.

**1.** (b) 
$$K = \frac{1}{R} \times \text{Cell constant}$$

Cell constant =  $K \times R$ ; 0.012 × 55 = 0.66 cm<sup>-1</sup>.

(b) In common dry cell. 5. Anode :  $Zn \rightarrow Zn^{++} + 2e^{-1}$ 

Cathode:  $2MnO_2 + Zn^{++} + 2e^- \rightarrow ZnMn_2O_4$ .

- **6.** (a) Because the reduction potential of Cu is highest.
- 7. (c) Overall reaction  $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O(l)$ .
- 9. (b) During charging of a lead storage battery, the reaction at the anode and cathode are Anode: PbSO 4 + 2e<sup>-</sup> → Pb + SO 4<sup>-</sup> Cathode:

$$PbSO_4 + 2H_2O \rightarrow PbO_2 + 4H^+ + SO_4^{2-} + 2e^-$$

In both the reactions  $H_2SO_4$  is regenerated.

**11.** (c)  $2NH_4Cl + Zn \rightarrow 2NH_3 + ZnCl_2 + H_2$   $\uparrow$ .

- **13.** (a) When platinum electrodes are dipped in dilute solution  $H_2SO_4$  than  $H_2$  is evolved at cathode.
- 14. (a) Electrode on which oxidation occurs is written on *L.H.S.* and the other on the *R.H.S.* as represented by

 $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$  .

Reduction

- **15.** (d)  $Zn^{2+} + 2e^- \rightarrow Zn$ . It shows reduction reaction.
- **16.** (c) In the electrolytic cell electrical energy change into chemical energy.
- 17. (c) In the cell  $Zn | Zn^{2+} || Cu^{2+} | Cu$  the negative electrode (anode) is Zn. In electrochemical cell representation anode is always written on left side while cathode on right side.
- **18.** (a) Galvanic cell converts the chemical energy into electrical energy.
- **19.** (b) Fuel-cells are used to provide power and drinking water to astronauts in space programme.

**21.** (b) 
$$E_{\text{cell}}^o = \frac{2.303 \ RT}{nF} \log K = \frac{0.0591}{n} \log K_c \ at \ 298 \ K$$
.

**22.** (b)  $Cu^{2+} + 2e^{-} \rightarrow Cu$   $\square$ Reductio

Reductio

**24.** (b) The cell in which *Cu* and *Zn* roads are dipped in its solutions called Daniel cell.

**25.** (c) 
$$K = C \times \text{Cell constant} = \frac{K}{C} = \frac{0.2}{0.04} = 5 \ cm^{-1}$$
.

**26.** (a)  $\frac{K}{C}$  = Cell Constant .

- **27.** (c) Velocities of both  $K^+$  and  $NO_3^-$  are nearly the same in  $KNO_3$  so it is used to make saltbridge.
- **28.** (a) In this reaction 4 electrons are needed for the reaction volume.
- **29.** (b) In electrochemical cell  $H_2$  release at anode and *Cu* is deposit at the cathode.
- **31.** (a) Anode has negative polarity.

**32.** (b)  $\wedge_m^o(CH_3COOH) =$ 

$$\wedge^{o} (CH_{3}COONa) + \wedge^{o} (HCl) - \wedge^{o} (NaCl)$$

 $= 91 + 426.16 - 126.45 = 390.71 \ ohm^{-1}cm^{2}mol^{-1}$ .

- **36.** (b) At anode:  $Zn_{(s)} \to Zn^{2+} + 2e^{-}$ .
- **38.** (d)  $PbSO_4 + 2H_2O \rightarrow 2PbO_2 + 4SO_4^{--} + 2e^{-}$ .
- **39.** (b)  $MnO_2$  is used in dry batteries cell.
- **40.** (d)  $Pb + PbO_2 + 2H_2SO_4 \xrightarrow{\text{Discharge}} 2PbSO_4 + 2H_2O$ . Sulphuric acid is consumed on discharging.
- **42.** (d) The metal placed below in electrochemical series does not react with that metal salt solution which metal is placed above in series.
- **43.** (c) In the electrochemical cell chemical energy changes into electrical energy.
- **44.** (a) In galvanic cell, the salt bridge used to complete the circuit.

**45.** (d)  $Cu + FeSO_4 \rightarrow No$  reaction Because Cu has  $E^o = 0.34$  volt and Fe has  $E^o = -0.44$  volt.

- **47.** (d) Calomel electrode as reference electrode is made by using  $H_{g_2}Cl_2$ .
- **48.** (b) In hydrogen-oxygen fuel cell following reactions take place to create potential difference between two electrodes.  $2H_{2(g)} + 4OH^{-}_{(aq)} \rightarrow 4H_2O_{(l)} + 4e^{-}$

$$O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^{-}_{(aq)}$$

Overall reaction =  $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$ 

the net reaction is the same as burning (Combustion) of hydrogen to form water.

**49.** (c) 
$$Cl CH_2COONa + HCl \rightarrow ClCH_2COOH + NaCl \lambda_{ClCH_2COONa} + \lambda_{HCl} = \lambda_{ClCH_2COOH} + \lambda_{NaCl} + \lambda_{2Cl} + 203 = \lambda_{ClCH_2COOH} + 38.2$$

 $\lambda_{CICH_2COOH} = 427 - 38.2 = 388.8 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$ .

- **50.** (c) In daniel cell copper rod acts as cathode so there cations move towards copper electrode and reduction take place on copper rod.
- 51. (b) Elements with lower reduction potential act as anode. His placed above *Cu* in electrochemical series so it has lesser reduction potential and thus act as anode and *Cu* act as cathode.
- **52.** (d) Fuel cells are more efficient as they are free from pollution and hence they run till the reactants are active. They have longer life than lead storage cells.
- **53.** (c) For gold plating, the used electrolyte is  $K[Au(CN)_2]$ .
- **54.** (a) Dil.  $H_2SO_4$  is used in lead in lead storage battery as electrolyte.
- **55.** (c) Cell constant  $=\frac{\text{Specific conductivity}}{\text{Observed conductanc e}}$

$$=\frac{0.002765}{1/R} = 0.002765 \times 400 = 1.106$$

57. (b)  $2AgCl_{(s)} + H_{2(g)} \rightarrow 2HCl_{(aq.)} + 2Ag_{(s)}$ The activities of solids and liquids are takes as unity and at low concentrations, the activity of a solute is approximated to its molarity. Th cell reaction will be

 $Pt_{(s)} \mid H_{2(g)}, 1 \text{ bar } \mid H^+_{(aq)} 1 M \mid AgCl_{(aq)} 1 M \mid$  $Aq_{(s)}$ 

(a)  $E_{cell} = \frac{0.059}{n} \log \frac{1}{C} = -\frac{0.059}{2} \log \frac{1}{100}$ 58.  $=-\frac{0.059}{2}(-2)=0.059 V=59 mV$ . (increase) (a)  $E_{cell} = -0.059 \log \frac{10^{-6}}{10^{-3}} = -0.059 \log 10^{-3}$ 59.

 $= -0.059 \times (-3) = 0.177 V$ .

(c) The magnitude of the electrode potential of a 60. metal is a measure of its relative tendency to lose or gain electrons. i.e., it is a measure of the relative tendency to undergo oxidation (loss of electrons) or reduction (gain of electrons).

> $M \rightarrow M^{n+} + ne^{-}$  (oxidation potential)  $M^{n+} + ne^- \rightarrow M$  (reduction potential)

- (a) On electrolysis of fused ionic hydride (*LiH*),
- 61. hydrogen obtained at anode.  $MH + H_2O \rightarrow MOH + H_2 \uparrow$

# Electrode potential, Ecell, Nernst equation and ECS

1. (b) Reduction potential of hydrogen electrode,  $E_H = \frac{-2.303 \text{ RT}}{E} \log \frac{1}{[H^+]}$ 

$$= -0.059 \ pH = -0.059 \times 3 = -0.177 \ V$$
.

(a)  $E_{\text{cell}}^o = E_{\text{cathode}}^o - E_{\text{anode}}^o = 0.799 - (-0.763) = 1.562 V$ 2.

- (a) More negative is the reduction potential, 3. higher will be the reducing property, i.e. the power to give up electrons.
- (b) Standard potential of Zinc < Copper. 4.
- 6. (c) A cation having highest reduction potential will be reduced first and so on. However,  $Mg^{2+}$  in aqueous solution will not be reduced

 $\left( E^{0}_{Mg^{2^{+}}/Mg} < E_{H_{2}O/\frac{1}{2}H_{2}+OH^{-}} \right)$ . Instead water would

be reduced in preference.

(c) A is displace from D because D have a 7.  $E^o = -0.402$  V Reductio

8. (a) 
$$Z_{\mu_{(s)}}^{o} + 2Ag_{(qq)}^{+} \rightarrow Z_{\mu_{(aq)}}^{2+} + 2Ag_{(s)}^{o}$$
  
Oxidation

In this reaction zinc act as a anode and  $A_g$  act as a cathode.

- (a) No doubt Be is placed above Mg in the 9. second group of periodic table but it is below  $M_g$  in electrochemical series.
- 10. (b) Nernst's equation shows relation between E and  $E^{o}$ .

**11.** (a) 
$$E = E^o - \frac{RT}{nF} \ln \frac{1}{[M^{n+1}]}$$
;  $E = E^o + \frac{RT}{nF} \ln[M^{n+1}]$ 

$$E = E^o + \frac{2.303 \, RT}{nF} \log[M^{n+}]$$

Substituting the value of *R*, *T* (298*K*) and *F* we get

$$E = E^{o} + \frac{0.0591}{n} \log(M^{n+})$$
.

- (c) At 298 K standard electrode potential of NHE 12. electrode is 0.00 V.
- (a) Since,  $Ag^+$  ions are reduced to Ag and 13.

 $E_{Ag^+/Ag}^o > E_{Cu^{++}/Cu}^o$  Cu is oxidized to Cu<sup>++</sup>.

- (d) The reducing power decreases as 14. the reduction potential increase (becomes less negative).
- (c) Actually the equation is derived from Nerst 15. equation assuming equilibrium condition in a cell reaction, when E = 0.
- (a) More negative is the standard reduction 16. potential, greater is the tendency to lose electrons and hence greater reactivity.
- 17. (a)  $H_g$  has greater reduction potential than that of  $H^+$  and hence cannot displace hydrogen from acid.
- 18. (c) Brown layer is deposited on iron rod because Cu has greater reduction potential than that of  $Fe^{2+}$ .
- (b) Since  $E^o_{A^{2+}/A} < E^o_{B^{2+}/B}$ . A has greater tendency 19. to be oxidized.  $A + B^{2+} \rightarrow A^{2+} + B$
- (b) Since  $E_{Z_{n}^{++}/Z_{n}}^{o}$  is negative, so Zn has greater 20. tendency to be oxidized than hydrogen. Hence it can act as reducing agent.
- (a) Standard electrode potential of Hydrogen is 21. zero.
- (b) According to electrochemical series. 22.

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(a) The standard reduction potential of  $K^+$ , 23.  $Mg^{2+}$ ,  $Zn^{+2}$   $Cu^{2+}$  increase in this order.

**24.** (c) 
$$E_{\text{cell}} = E_{Au^{3+}/Au}^o - E_{Ni^{2+}/Ni}^o = 1.50 - (-0.25) = 1.75 \text{ V.}$$

- (a) Electromotive force is +ve if oxidation and 25. reduction both takes place in a cell.
- 28. (a) In galvanic cell anode always made up of negative electrode.

9. (d) 
$$A \mid A^+(a=1) \parallel B^+(a=1) \mid B$$
  
 $EMF = E_{cathode} - E_{anode} = 0.75 - (0.5); EMF = 0.25 V.$ 

- (d)  $E^o = -3.05 Li^+ / Li$  is most negative (minimum) 30. and hence Li has maximum tendency to lose electrons or it is the strongest reducing agent.
- (b) Brown layer is deposited on iron rod because 31. *Cu* has greater than reduction potential than that of  $Fe^{2+}$ .
- 32. (a)  $E^o_{Zn^{++}/Zn} < E^o_{Fe^{++}/Fe}$ , so Zn will reduce  $Fe^{++}$ . Zn cannot reduce  $Mg^{2+}$  because  $E^o_{Z_n^{++}/Z_n} > E^o_{Ma^{++}/Ma}$

On similar reason Mg and Zn cannot oxidize Fe.

**33.** (d) For the cell reaction, *Fe* acts as cathode and *Sn* as anode. Hence,

 $E_{cell}^{o} = E_{cathode}^{o} - E_{anode}^{o} = - 0.44 - (- 0.14) = - 0.30V$ 

The negative *EMF* suggests that the reaction goes spontaneously in reversed direction.

**34.** (a) 
$$E_{\text{cell}}^o = E_{\text{cathode}}^o - E_{\text{anode}}^o = 0.34 - (-0.76) = 1.10 V$$
.

- 35. (c)  $E_{\text{cell}}^o = E_{\text{cathode}}^o E_{\text{anode}}$ ;  $E_{\text{cell}}^o = 0.34 (-2.37)$  $E_{\text{cell}}^o = 2.71 \text{ V}$ .
- **36.** (d) *Mg* lies above *Cu* in electrochemical series and hence *Cu* electrode acts as cathode

$$E_{cell}^{o} = E_{Cu^{++}/Cu}^{o} - E_{Mg^{++}/Mg}^{o}$$
  
2.70 V = 0.34 -  $E_{Mg^{++}/Mg}^{o}$ ;  $E_{Mg^{++}/Mg}^{o}$  = -2.36 V.

**37.** (a) Because  $H_2$  has greater reduction potential so it reduced the  $Ag^+$ .

**39.** (d) 
$$\Delta G^{\circ} = -nE^{\circ}F$$
  
 $Fe^{2+} + 2e^{-} \rightarrow Fe$   
 $\Delta G^{\circ} = -2 \times F \times (-0.440 V) = 0.880 F$ 

 $Fe^{3+} + 3e^{-} \rightarrow Fe \qquad \dots \dots (ii)$   $\Delta G^{o} = -3 \times F \times (-0.036) = 0.108 F$ On subtracting equation (i) from (ii)  $Fe^{3+} + e^{-} \rightarrow Fe^{2+}$   $\Delta G^{o} = 0.108 F - 0.880 F = -0.772 F$   $E^{o} \qquad \text{for} \qquad \text{the} \qquad \text{reaction} \qquad = -\frac{\Delta G^{o}}{nF}$  $= -\frac{(-0.772 F)}{1 \times F} = +0.772 V.$ 

.....(i)

- **40.** (d) Reducing power *i.e.* the tendency to lose electrons increases as the reduction potential decreases.
- **41.** (b)  $Cu^{++}$  will be reduced and Fe will be oxidized.  $Cu^{++} + Fe \rightarrow Cu + Fe^{++}$ .

(c) Cell reaction is  

$$Cu_{(s)} + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$$
  
Two half cell reaction is  
 $Cu \rightarrow Cu^{2+} + 2e^-$  Oxidation (anode)  
 $Ag^+ + e^- \rightarrow Ag$  Reduction (cathode)  
 $E_{Cell} = E_{ox} - E_{\text{Re}\,d} = 0.80 - 0.34 = +0.46V$ 

- **43.** (a) EMF = [s.r.p. of cathode-s.r.p of anode] Where s.r.p. = Standard reduction potential
  - If EMF is positive then the reaction is spontaneous For *e.g.* in Galvanic cell

(a) EMF = 1.1 *volt* 

42.

(b) Cathode is made of copper

(c) Anode is made of Zinc

EMF = 0.34 -(- 0.76) = 1.1 *volt*.

- **46.** (d)  $H_2$  is anode because oxidation takes place. *Cu* is cathode because reduction is takes place.
- 47. (c)  $E_{\text{cell}}^o = E_{\text{cathode}} E_{\text{anode}}$ .
- **51.** (b) Standard hydrogen electrode have zero electrode potential.
- **53.** (a)  $\Delta G = -nFE^{\circ}$  $\Delta G = -1 \times 96500 \times 1.02$ ;  $\Delta G = -98430$ .
- 54. (c) Fuel cell converts the chemical energy into electrical energy.

**55.** (a) 
$$E = E^o - \frac{2.303 \ RT}{nF} \log \frac{[M]}{[M^{n+1}]}$$

- **56.** (a) Less is the reduction potential stronger is the reducing agent.
- **57.** (a) Reducing power, *i.e.* the tendency to lose electrons increases as the reduction potential decreases.

**58.** (a) 
$$OCl^- \to C^-, E^o = 0.94 V$$

$$Cl^{-} \rightarrow \frac{1}{2}Cl_{2} + e^{-}E^{o} = -1.36 V$$

adding the two equations, we get

$$OCl^{-} \rightarrow \frac{1}{2}Cl_{2}, E^{o} = 0.94 - 1.36 = -0.42 V.$$

**60.** (a) It cannot evolved 
$$H_2$$
 from  $H_2S$ 

 $Hg + H_2 S \rightarrow \text{No reaction}$ .

**61.** (b) 
$$E_{cell}^{o} = \frac{0.059}{n} \log K$$
  
 $\log K = \frac{1.10 \times 2}{0.059} = 37.2881$  or  $K = 10^{-37}$ .

**62.** (d) The oxidizing character *i.e.* acceptance of electrons increases with the reduction potential.

64. (d) 
$$E_{cell} = E_{cell}^o - \frac{KI}{nF} \ln \frac{C_2}{C_1}$$
 and  $\Delta G = -nF E_{cell}$   
hence  $\Delta G$  is the function of  $\ln \left(\frac{C_2}{C_1}\right)$ .

**66.** (c) 
$$E = E^o - \frac{0.059}{n} \log \frac{[Zn^{++}]}{[Cu^{++}]} = 1.10 - \frac{0.059}{2} \log \frac{1}{0.1}$$
  
= 1.10 - 0.0295 log 10 = 1.07 volt.

67. (b) 
$$E_1 = E_o - \frac{0.0591}{2} \log \frac{0.01}{1} = E_o + \frac{0.0591}{2} \times 2$$
  
 $E_2 = E_o - \frac{0.0591}{2} \log \frac{100}{0.01} = E_o - \frac{0.0591}{2} \times 4$   
 $\therefore E_1 > E_2$ .  
Oxidation

**68.** (b) 
$$Fe^{+2} + Zn \rightarrow Zn^{2+} + Fe$$
  
Reduction

 $EMF = E_{\text{cathode}} - E_{\text{anode}} = 0.44 - (0.76) = +0.32 V.$ 

**69.** (a) Fe is more electropositive than copper. Hence  $Cu^{2+}$  can oxidise Fe.

- **70.** (b)  $E^o = 0$  because hydrogen have zero potential.
- **71.** (b) Cell potential of the cell is positive.
- **72.** (a,b) Because these comes after the *Fe* in electrochemical series.

73. (c) 
$$Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$$
  
Reduction  
$$EMF = E_{cathode} - E_{anode} = -7.81 - (-7.62)$$
$$FMF = -0.19 V.$$

74. (c)  $Cr^{3+} > Zn^{2+} > H > Fe^{3+}$ .

Reducing nature decreasing order.

**75.** (a) More is reduction potential, more is the power to get itself reduced or lesser is reducing power or greater is oxidising power.

**76.** (d) 
$$\Delta G = -nFE^{\circ}$$

$$\Delta G = -2.303 RT \log K$$
;  $nFE^{\circ} = 2.303 RT \log K$ 

$$\log K = \frac{nFE^{\circ}}{2.303 RT} = \frac{2 \times 96500 \times 0.295}{2.303 \times 8.314 \times 298}$$
$$\log K = 9.97 = K = 1 \times 10^{10} .$$

- 77. (b) For the given cell  $M | M^+ || X^- | X$ , the cell reaction is derived as follows: RHS: reduction  $X + e^- \rightarrow X^-$  .....(i)
  - LHS: Oxidation  $M \rightarrow M^+ + e^-$  .....(ii) Add (i) and (ii)  $M + X \rightarrow M^+ + X^-$ The cell potential = -0.11 VSince  $E_- - - ve$  the cell reaction derive

Since  $E_{cell} = -$  ve, the cell reaction derived above is not spontaneous. In fact, the reverse reaction will occur spontaneously.

**78.** (c)  $Z_{n+M_gCl_2}^{0} \xrightarrow{+2} Z_nCl_2 + M_g$  No reaction Reduction

This type of reaction does not occur because  $Mg^{2+}E^o = -2.37V$  while  $Zn^{2+}E^o = -0.76V$ .

- **79.** (b) In neutral medium  $Mn^{+7}$  oxidation state change into +4 oxidation state, hence equivalent weight of  $KMnO_4 = \frac{M}{3}$ .
- **80.** (a) Increase in the concentration of  $Ag^+$  ion increase the voltage of the cell.

**81.** (a) 
$$E_{\text{cell}} = E_{\text{cell}}^o + \frac{0.059}{2} \log \frac{(Ag^+)}{(Sn^{2+})}$$
.

- 82. (b) The K.E. of proton is 1 KeV.
- **83.** (b) Anodic reaction :  $H_2(P_1) \rightarrow 2H^+$

Cathodic reaction :  $2H^+ \rightarrow H_2(P_2)$ 

$$E_{cathode} = -\frac{RT}{2F} \ln \frac{P_2}{[H^+]^2}$$
;  $E_{anode} = -\frac{RT}{2F} \ln \frac{[H^+]^2}{P_1}$ 

$$E_{inf} = E_{anode} + E_{cathode}$$
  
=  $-\frac{RT}{2F} \ln \frac{(H^+)^2}{P_1} - \frac{RT}{2F} \ln \frac{P_2}{(H^+)^2}$   
=  $-\frac{RT}{2F} \ln \frac{P_2}{P_2} = \frac{RT}{2F} \ln \frac{P_1}{P_1}$ .

**84.** (c) 
$$\frac{1}{2}H_2 \rightarrow H^+(10^{-8}M) + e^-(\text{oxidation})$$

$$H^+(0.025 M) + e^- \rightarrow \frac{1}{2}H_2$$
 (reduction)

Cell reaction is :

 $H^+(0.025 \ M) \to H^+(10^{-8} \ M)$ ;  $E_{\text{cell}} = 0.38 \ V$ .

- **85.** (a)  $E^{o}$  for Fe / Fe<sup>2+</sup> = 0.44 V .
- 86. (c) (Reduction potential of cathode) (reduction potential of anode).
- **87.** (a) The correct decreasing electrode potential order is : *K*, *Ba*, *Ca*, *Mg*.

**89.** (c) 
$$E_{cell}^o = E_{cathode}^o - E_{anode}^o$$
  
= 0.34 -(-2.37) = +2.71 V.

- **90.** (b) Because flourine is most powerful reducing agent than other halogens.
- **92.** (c) Aluminium forms a protective oxide layer but iron does not.
- **93.** (d) The reduction potential of *Zn* is very higher than *Cu*.
- 94. (a) The *pH* of 0.1 *M HCl* and 0.1 *M* acetic acid is not the same, because *HCl* is a strong acid so its *pH* is less and *CH*<sub>3</sub>*COOH* is a weak acid, so its *pH* is more.
- (d) The required reaction  $(Cu^{++} + Cu \rightarrow 2Cu^{+})$  can 95. be obtained by using the following reactions.  $Cu^{++} + e^{-} \rightarrow Cu^{+}; E^{o}_{Cu^{++}/Cu^{+}} = 0.15 V$ ....(i)  $Cu^{++} + 2e^{-} \rightarrow Cu; \ E^{o}_{Cu^{++}/Cu} = 0.34 \ V$ .....(ii) Multiplying eq. (i) by 2 we get  $2Cu^{++}+2e^- \rightarrow 2Cu^+$ .....(iii)  $\Delta G_1 = -nFE = -2 \times F \times 0.15$  $Cu^{++} + 2e^- \rightarrow Cu$ .....(iv)  $\Delta G_2 = -nFE = -2 \times F \times 0.34$ Subtract the eq. (iv) from (iii)  $Cu^{++} + Cu \rightarrow 2Cu^{+}$  $\Delta G_3 = -nFE = -1 \times F \times E^o$ Also  $\Delta G_3 = \Delta G_1 - \Delta G_2$  $-1FE^{\circ} = (-2F \times 0.15) - (-2F \times 0.34)$  $E^{o} = -0.38$ This is the value for the reaction  $Cu^{++} + Cu \rightarrow 2Cu^{+}$ But the given reaction is just reverse of it  $\therefore E_{\text{cell}}$  for given reaction = + 0.38*V*.
- **97.** (d) It connect two solutions and complete the circuit.
- **98.** (a) Greater the oxidation potential, greater is the reactivity.

- **99.** (b) Electrochemical series compare the relative reactivity of metals.
- 100. (d) Fuel cells are more efficient, free from pollution and they run till reactants are active.
- **102.** (C)  $E^o = E^o_{Ag^{2+}/Ag} + E^o_{Cu/Cu^{2+}} = -0.34 + 0.80 = +0.46 V$ .
- **103.** (a) *Fe* is placed above *Cu* in electrochemical series.
- **104.** (a) Lithium is the strongest reducing agent of the alkali metals.
- 105. (d) Potassium is more electropositive element, because it is the only alkali metal among the given elements.
- **106.** (b) Aluminium forms a self protecting film of oxide to prevent corrosion.
- **107.** (a)  $Zn_{(s)} + 2H^{+}_{(aq)} \approx Zn_{(aq)}^{2+} + H_{2(g)}$  $E_{Cell} = E_{Cell}^{0} - \frac{.059}{2} \log \frac{[Zn^{2+}]}{[H^{+}]^{2}}$

When  $H_2SO_4$  is added then  $[H^+]$  will increase therefore  $E_{Cell}$  will also increases and equilibrium will shift towards right.

**108.** (b) For  $M^+ + X^- \rightarrow M + X$ 

 $E_{Cell}^{0} = E_{Cathode}^{0} + E_{Anode}^{0} = 0.44 - 0.33 = +0.11 V$ 

Since  $E_{Cell}^0 = (+) 0.11 V$  is positive hence this reaction should be spontaneous.

**109.** (a) 
$$E_{Cell} = E_{Cell}^0 - \frac{0.0591}{n} \log K_c$$
  
At 298  $K E_{Cell} = 0$   $O = 0.591 - \frac{0.0591}{n} \log K_c$   
 $\log K_c = \frac{0.591 \times 1}{0.0591} = 10$ ;  $K_c = \text{Antilog } 10 = 1 \times 10^{10}$ .

- **110.** (a)  $\frac{1}{2}H_2 \mid H^+ \mid \mid Ag^+ \mid Ag \mid$   $E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = E^0_{Ag^+/Ag} - E^0_{H^+/-H_2}$ 
  - (0.80) (0.0) = 0.80 V.
- 111. (a)  $E_A = 2.23 V > E_B = 1.43 V$ So A will act as cathode in galvanic cell. Hence  $E_{Cell}^0 = E_{Cathode} - E_{Anode} = E_A - E_B$ = (2.23) - (-1.43) = 3.66 V.
- **112.** (b)  $E_{Gu}^0 > E_{Mg}^0$  hence *Cu* acts as cathode and *Mg* acts as anode.

$$E_{Cell}^0 = E_{Cu}^0 - E_{Mg}^0 = (0.34) - (-2.37) = +2.71 V.$$

**113.** (b) In this cell *Co* is oxidised and it acts as anode and *Ce* acts as cathode.

$$E_{Cell}^{0} = E_{Cathode}^{0} - E_{Anode}^{0} = 1.89 = E_{Cell}^{0} - (-0.28)$$

$$E_{Cell}^0 = 1.89 - 0.28 = 1.61$$
 Volts.

**114.** (b) Given:  $\Delta G = -21.20 \ kJ = 21200 \ J$  $\therefore \ \Delta G = -nFE$   $E = \frac{21200}{1 \times 96500} = 0.2196 \ V = 0.22 \ V.$ 

**115. (b)** 
$$Ag \mid Ag^{+}(.1m) \parallel Ag^{+}1M \mid Ag \mid$$
  
 $E_{Cell} = \frac{2.303 RT}{nF} \log \frac{c_1}{c_2} = \frac{0.059}{1} \log \frac{1}{0.1}$   
 $= 0.059 \log 10 = 0.059 Volt$ .

**116.** (b) For this cell, reaction is:  $Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe$ 

$$E = E^{0} - \frac{0.0591}{n} \log \frac{c_{1}}{c_{2}}; E^{0} = E + \frac{0.0591}{n} \log \frac{c_{1}}{c_{2}}$$
$$= 0.2905 + \frac{0.0591}{2} \log \frac{10^{-2}}{10^{-3}} = 0.32 V .$$
$$E^{0} = \frac{0.0591}{2} \log K_{c}; \log K_{c} = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$$
$$\therefore K_{c} = \frac{0.32}{10^{0295}} .$$

117. (d) Al displaces H from HCl but silver cannot it means Al is situated above the Ag in ECS, hence Al will acts as anode and Ag will act as cathode.

$$E^{0}_{Cell} = E^{0}_{Cathode} - E^{0}_{Anode} = E^{0}_{Ag^{+}/Ag} - E^{0}_{Al^{3+}/Al}$$
  
2.46 = 0.8 -  $E^{0}_{Al^{3+}/Al}$ ;  $E^{0}_{Al} = 0.8 - 2.46 = -1.66 V$ 

**118.** (a) For 
$$Sn_{(s)} + 2Fe_{(aq)}^{3+} \rightarrow 2Fe_{(aq)}^{2+} + Sn_{(aq)}^{2+}$$
  
 $E_{Cell}^{0} = E_{Sn/Sn^{2+}}^{0} + E_{Fe^{3+}/Fe^{2+}}^{0} = (0.14) + (0.77)$   
 $E_{Cell} = 0.91$  Volts.

**119.** (d)  $I^-$  get oxidised to  $I_2$  hence will form anode and  $Cr_2O_7^{2-}$  get reduced to  $Cr^{3+}$  hence will form cathode.

$$E_{Cell}^{0} = E_{Cathode}^{0} - E_{Anode}^{0} ; E_{Cell}^{0} = E_{Cr_{2}O_{7}^{-2}}^{0} - E_{I_{2}}^{0}$$
  
0.79 = 1.33 -  $E_{I_{2}}^{0} ; E_{I_{2}}^{0} = 1.33 - 0.79 ; E_{I_{2}}^{0} = 0.54 V$ 

120. (b) According to nernst's equation

$$E_{Cell} = E_{Cell}^{0} - \frac{nRT}{F} \log \frac{c_1}{c_2}$$
  
For  $Zn_{(s)} + Cl_{2(1 atm)} \rightarrow Zn^{2+} + 2Cl^{-}$   
 $c_1 = [Zn^{2+}]$  and  $c_2 = [Cl^{-}]$ 

Hence to increase *E*,  $c_1$  should be decreased and  $c_2$  should be increased is  $[Zn^{2+}]$  should be decreased and *Cl* should be increased.



Reduction $E_0 M^{3+} / M^{2+}$	Cell re	action	$E_0 M^{2+} / M^{3+}$ (Oxidation)
41 V	Cr <sup>2+</sup>	Cr <sup>3+</sup>	+.41 V
+ 1.57 V	Mn <sup>2+</sup>	Mn <sup>3+</sup>	- 1.57 V
+ 0.77 V	Fe <sup>2+</sup>	Fe <sup>3+</sup>	- 0.77 V
+ 1.97 V	Co <sup>2+</sup>	Co <sup>3+</sup>	– 1.97 V

As *Cr* has maximum oxidation potential value, therefore its oxidation should be easiest.

- 122. (a)  $Fe(s) \longrightarrow Fe^{2^+} + 2e^-$ ;  $\Delta G_1^o$   $\frac{2H^+ + 2e^- + \frac{1}{2}O_2 \longrightarrow H_2O(l) ; \Delta G_2^o}{Fe(s) + 2H^+ + \frac{1}{2}O_2 \longrightarrow Fe^{2^+} + H_2O ; \Delta G_3^o}$ Applying,  $\Delta G_1^o + \Delta G_2^o = \Delta G_3^o$   $\Delta G_3^o = (-2F \times 0.44) + (-2F \times 1.23)$   $\Delta G_3^o = -(2 \times 96500 \times 0.44 + 2 \times 96500 \times 1.23)$   $\Delta G_3^o = -322310 J$   $\therefore \Delta G_3^o = -322 kJ$ 128. (a)  $\Delta G^o = -2.303 RT \log K_{eq}$  or  $\Delta G^o = -nFE_{cell}^o$
- **129.** (a) Any redox reaction would occur spontaneously if the free energy change  $(\Delta G)$  is negative.
  - $\Delta G^{o} = nFE^{o}$

Where *n* is the number of electrons involved, *F* is the value of Faraday and  $E^o$  is the cell emf.  $\Delta G^o$  can be negative if  $E^o$  is positive.

# Corrosion

- **2.** (d) Rusting of iron is catalysed by  $[H^+]$ .
- 3. (d)  $HgCl_2$  has corrosive action. It is highly poisonous. It sublimes on heating. It is, therefore, known as corrosive sublimate.
- 4. (a)  $Fe \rightarrow Fe^{2+} + 2e$  (anode reaction)

 $O_2 + 2H_2O + 4e \rightarrow 4OH^-$  (cathode reaction)

The overall reaction is

 $2Fe+O_2+2H_2O\rightarrow 2Fe(OH)_2$ 

 $Fe(OH)_2$  may be dehydrated to iron oxide FeO, or further oxidised to  $Fe(OH)_3$  and then dehydrated to iron rust,  $Fe_2O_3$ .

# **Critical Thinking Questions**

1. (c) 
$$(126 \ scm^2) \wedge_{NaCl}^0 = \wedge_{Na^+}^0 + \wedge_{Cl^-}^0$$
 .....(1)  
 $(152 \ scm^2) \wedge_{KBr}^0 = \wedge_{K^+}^0 + \wedge_{Br^-}^0$  .....(2)  
 $(150 \ scm^2) \wedge_{KCl}^0 = \wedge_{K^+}^0 + \wedge_{Cl^-}^0$  .....(3)  
By equation (1)+(2) - (3)  
 $\because \wedge_{NaBr}^0 = \wedge_{Na^+}^0 + \wedge_{Br^-}^0$   
 $= 126 + 152 - 150 = 128 \ Scm^2 mol^{-1}$   
2. (a) At cathode :  $2H^+(aq) + 2e^- \rightarrow 2H$   
 $2H + \frac{1}{2}O_2 \rightarrow H_2O$   
 $\overline{2H^+ + \frac{1}{2}O_2 + 2e^-} \rightarrow H_2O$ 

3. (c)  $H_2$  undergoes oxidation and  $AgCl(Ag^+)$ undergoes reduction  $AgCl(Ag^+)$ 

4. (b) In this reaction 
$$Fe^{2+} + Zn \rightarrow Zn + Fe$$
  
Reduction

 $EMF = E_{\text{cathode}} - E_{\text{anode}} = -0.41 - (-0.76)$ EMF = +0.35 V.

5. (c)  $NO_3^- + 4H^+ + 4e^- \rightarrow 2H_2O + NO$ . In this equation all the atoms are balanced. For balancing added  $3e^-$  to *L.H.S.* we have,  $NO_3^- + 4H^+ + 3e^- \rightarrow 2H_2O + NO$ .

6. (a) 
$$E_{cell} = E_{cell}^o - \frac{0.059}{2} \log \frac{(Zn^{++})}{(Cu^{++})}$$
  
=  $1.10 - \frac{0.059}{2} \log \frac{0.1}{0.1} = 1.10 V$ .

7. (a) The tendency to gain electron is in the order Z > Y > XThus  $Y + e \rightarrow Y^{-}$ ;  $X \rightarrow X^{+} + e$ .

8. (b) 
$$E_{OP} = E_{OP}^{o} - \frac{0.059}{1} \log \frac{[H^+]}{P_{H_2}}$$

 $\therefore [H^+] = 10^{-10}; P_{H_2} = 1 \text{ atm}; E_{OP} = 0.59 V.$ 

- **9.** (a.c,d)Decomposition of  $H_2O_2$  is an example of exothermic reaction, negative catalysis and auto-oxidation.
- **10.** (a) 27 gm of Al is obtained by passing a current of  $3 \times 96500$  C.

∴ 1 gm of Al is obtained by passing a current of  $3 \times \frac{96500}{27} C$ .

∴ 5.12 × 10<sup>3</sup> gm of Al is obtained by passing a current of  $3 \times \frac{96500}{27} \times 5.12 \times 1000$ 

$$= 1.83 \times 10^{7} C \times 3 = 5.49 \times 10^{7} C.$$

$$= 91.0 + 426.2 - 126.5 = 390.7$$

**12.** (d) 
$$\frac{\omega_1}{E_1} = \frac{\omega_2}{E_2}$$
;  $\frac{\omega_1}{3} = \frac{270}{93}$ ;  $\omega_1 = 90 \ kg$ .

11.

13. (c) Eq of 
$$Al = eq of H_2$$
  
 $\frac{4.5}{\frac{27}{3}} = eq of H_2; \quad \frac{4.5}{9} = eq of H_2$   
 $2H^+ + 2e^- \rightarrow H_2$   
 $eq. of H_2 = Number of moles × n factor$   
 $0.5 = n_{H_2} \times 2$   
 $V_{H_2} = \frac{0.5}{2} \times 22.4; \quad V_{H_2} = 5.6 L$ 

14. (c) The reaction taking place at anode is  

$$2Cl^{-} \rightarrow Cl_{2} + 2e^{-}$$

$$1 \text{ mole } 2 \times 96500 \text{ coulomb}$$

$$Q = i \cdot t = 1 \times 30 \times 60 = 1800 \text{ coulomb}.$$
The amount of chlorine liberated by passing  

$$1800 \quad \text{coulomb} \quad \text{of electric charge}$$

$$= \frac{1 \times 1800 \times 71}{2 \times 96500} = 0.66 \text{ g}.$$

# Assertion & Reason

- 1. (a) The nature of the cathode can affect the order of discharge of ions.
- 2. (b) 1 mole of silver = 1g equivalent of silver
  1 mole of copper = 2g equivalent of copper
  We know from Faraday's law of electrolysis that "The weight of ion deposited on an electrode is directly proportional to the quantity of electricity passed".
- (d) Copper is present below hydrogen therefore hydrogen from *HCl* cannot be liberated by treating with copper. Hence assertion is false while reason is true.
- 5. (d) Copper cannot liberate hydrogen from a dil. *HCl* solutions because it is situated below hydrogen in the reactivity series. Here both assertion and reason are false.
- 6. (d) The formation of Zn is not possible by placing Cu plate in  $ZnSO_4$  solution because Zn is placed above Cu in electrochemical series. Therefore, both assertion and reason are false.
- (e) Electrical conductivity of copper decreases with increase in temperature because the metallic conductivity is due to the motion of electrons. On increasing temperature the motion of electron increases which hinder in conductance of current. Hence, here assertion is false but the reason is true.
- **8.** (a) Dry air is heavier than wet air because the density of dry air is more than water.
- **9.** (e) Copper is present below hydrogen therefore hydrogen from *HCl* cannot be liberated by treating with copper. Hence, assertion is false while reason is true.
- 10. (a) K and Cs emit electrons on exposure to light hence, both are used in photoelectric cells. Here, assertion and reason are true and reason is a correct explanation.

**12.** (b) We know, 
$$R \propto \frac{l}{A}$$
 or  $R = \rho\left(\frac{l}{A}\right)$ , where

proportionality constant  $\rho$  is called resistivity. If l=1m and  $A=1m^2$ , then  $R=\rho$ *i.e.* Resistance = Resistivity.

- 14. (a) According to Kohlrausch law, "Limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte".
- **15.** (c) One Faraday deposite one gram equivalent of the substance.
- 16. (b) Gold has higher reduction potential than the given metals. Hence  $AuCl_3$  will react with these metals.
- **17.** (a)  $Zn(s) + Cu^{2+}(aq) \to Zn^{2+}(aq) + Cu(s)$

As the time passes, the concentration of  $Zn^{2+}$  keeps on increasing while the concentration of  $Cu^{2+}$  keeps on decreasing. At the same time voltage of the cell keeps on decreasing. When there is no change in concentration of  $Cu^{2+}$  and  $Zn^{2+}$  ions, voltmeter gives zero reading and this state is known as equilibrium.

- 18. (e) A negative value of standard reduction potential means that oxidation takes place on this electrode with reference to SHE.
- **19.** (a) A standard cell is one whose *e.m.f.* almost does not change with temperature.
- 22. (d) Identification of cathode and anode is done by the use of ammeter/voltmeter. Higher is the value of reduction potentials greater would be its oxidising power.
- **23.** (b) If redox reaction is spontaneous,  $\Delta G$  is -ve and hence  $E^0$  is positive.  $-\Delta G^0 = nFE^0$  cell
- 24. (a) Zinc metal which has a more negative electrode potential than iron will provide electrons in preference of the iron, and therefore corrode first. Only when all the zinc has been oxidised does the iron start to rust.
- 25. (a) Both assertion and reason are true and reason is the correct explanation of assertion. Anode (oxidation) for eq.  $Zn \longrightarrow Zn^{2+} + 2e^{-}$ , so Excess of electrons and hence negatively charged while cathode is positively charged.
- **26.** (a) Both assertion and reason are true and reason is the correct explanation of assertion.

$$E^{0}_{Au^{3+}/Au} - E^{0}_{Ni/Ni^{2+}} = 1.50 - (-0.25) = 1.75 V$$

- 27. (c) Assertion is true but reason is false. Ions of inert electrolytes are not involved in any electrochemical change until they react chemically with the electrolytes in the two half-cells.
- **28.** (d) Both assertion and reason are false. Potential difference is the difference between the electrode potential of the two electrodes of the cell when cell is under operation while emf is the potential difference generated by a cell when there is zero electron flow.

# Electrochemistry

- The mass of copper deposited from a solution of *CuSO*<sub>4</sub> by passage of 5 *A* current for 965 second is (*Mol. wt.* of Copper = 63.5)
  - (a) 15.875 g (b) 1.5875 g
  - (c) 4825 *g* (d) 96500 *g*
- 2. The current in a given wire is 1.8 *A*. The number of coulombs that flow in 1.36 minutes will be [AIIMS 2001]
  - (a) 100 *C* (b) 147 *C*
  - (c) 247 *C* (d) 347 *C*
- 3. A solution of a salt of a metal was electrolysed for 150 minutes with a current of 0.15 amperes. The weight of metal deposited was 0.783 gm. The equivalent weight of the metal is [AFMC 2001]

(a) 55.97 <i>gm</i>	(b) 65.97 gm
(c) 75.97 gm	(d) 85.97 gm

- 4. The resistance of 0.01*N* NaCl solution at 25 ° C is 200  $\Omega$ . Cell constant of conductivity cell is 1 cm<sup>-1</sup>. The equivalent conductance is
  - (a)  $5 \times 10^2 \Omega^{-1} cm^2 eq^{-1}$  (b)  $6 \times 10^3 \Omega^{-1} cm^2 eq^{-1}$ (c)  $7 \times 10^4 \Omega^{-1} cm^2 eq^{-1}$  (d)  $8 \times 10^5 \Omega^{-1} cm^2 eq^{-1}$
- 5. Which of the following reaction is possible at anode

[AIEEE 2002]

(a)  $2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$ (b)  $F_2 \rightarrow 2F^-$ (c)  $\frac{1}{2}O_2 + 2H^+ \rightarrow H_2O$ 

(d) None of these

**6.** What is the standard cell potential for the cell

$$Zn / Zn^{2+} (1M) \| Cu^{2+} (1M) / Cu$$

 $E^{o}$  for  $Zn/Zn^{2+}(1M) = -0.76 V \& Cu^{2+}/Cu = +0.34 V$ 

[AIIMS 1980]

(a) -0.76 + (-0.34) = -0.42 V

(b) -0.34 + 0.76 = +0.42 V

(c) 0.34 - (-0.76) = 1.10 V

- (d) -0.76 (+0.34) = -1.10 V
- [AIIMS 2001]
  7. Normal aluminium electrode coupled with normal hydrogen electrode gives an *emf* of 1.66 *volts*. So the standard electrode potential of aluminium is[KCET 198]

(a) 
$$- 1.66 V$$
 (b)  $+ 1.66 V$   
(c)  $- 0.83 V$  (d)  $+ 0.83 V$ 

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**8.** Which one among the following is the strongest reducing agent

 $Fe^{2+} + 2e^{-} \rightarrow Fe(-0.44 V)$   $Ni^{2+} + 2e^{-} \rightarrow Ni(-0.25 V)$   $Sn^{2+} + 2e^{-} \rightarrow Sn(-0.14 V)$   $Fe^{3+} + e^{-} \rightarrow Fe^{2+}(-0.77 V)$ [BHU 1998]
(a) Fe
(b)  $Fe^{2+}$ (c) Ni(b)  $Fe^{2+}$ 

- (c) *Ni* **(CBSE PMT 1999]** (d) *Sn*
- 9. The cell reaction of the galvanic cell  $Cu_{(s)} | Cu^{2+}_{(aq)} | Hg^{2+}_{(aq)} | Hg_{(l)}$  is [EAMCET 2003]
  - (a)  $Hg + Cu^{2+} \rightarrow Hg^{2+} + Cu$
  - (b)  $Hg + Cu^{2+} \rightarrow Cu^+ + Hg^+$
  - (c)  $Cu + Hg \rightarrow CuHg$
  - (d)  $Cu + Hg^{2+} \rightarrow Cu^{2+} + Hg$
- 10. The specific conductivity of  $N/10 \ KCl$  solution at  $20^{\circ}C$  is  $0.0212 \ ohm^{-1} \ cm^{-1}$  and the resistance of cell containing this solution at  $20^{\circ}C$  is 55 ohm. The cell constant is

# [AIIMS 1999]

- (a) 1.166  $cm^{-1}$ (b) 2.173  $cm^{-1}$ (c) 3.324  $cm^{-1}$ (d) 4.616  $cm^{-1}$
- **11.** The oxide which is not reduced by hydrogen is



= 
$$5 \times 965 = 4825C$$
. Since 96500 *coulombs* will  
deposit  $\frac{63.5}{2}g$  of copper therefore  $4825$  **8.** (a)

coulombs will deposit

$$= \frac{63.5 \times 4825}{96500 \times 2} = 1.5875 \ g$$

- (b)  $Q = I \times t$ ;  $1.8 \times 1.36 \times 60 = 147 C$ . 2.
- (a) Time (t) = 150 min = 9000 sec 3. Current (I) = 0.15 AWeight of metal (w) = 0.783 g. We know  $Q = I \times t = 0.15 \times 9000 = 1350 C$ . Since 1350 *C* of electricity will deposited 0.783 *g* of metal, so, 96500 C of electricity will deposited  $\frac{0.783 \times 96500}{1350} = 55.97 \ g$ . (a)  $\lambda = k \times V = \frac{1}{R} \times \frac{l}{a} \times V = \frac{1}{200} \times 1 \times 10,000$ 4.

$$=5 \times 10^{2} \Omega^{-1} cm^{2} eq.^{-1}$$

(a) Oxidation always occurs at anode. 5٠

(a) 
$$E_{\text{cell}}^o = 1.66 = E_{H^+/H_2}^o - E_{Al^{3+}/Al}^o$$
  
=  $O - E_{Al^{3+}/Al}^o$  or  $E_{Al^{3+}/Al} = -1.66 V$ 

- (a) The reduction potential of Fe is very high, so 8. it is a strongest reducing agent.
- (b)  $Cu_{(s)} | Cu_{(Ag)}^{2+} || Hg_{(Ag)}^{2+} | Hg_{(l)}$ anode oxidation cathode reduction 9.

$$Cu + Hg^{2+} \rightarrow Cu^{2+} + Hg$$
.  
Oxidation

10. (a)  $K = \frac{1}{R} \times \text{cell constant}$ 

 $= K \times R = 0.0212 \times 55 = 1.166 \ cm^{-1}$ .

(b) On the basis of electrochemical series  $K_2O$  is 11. not reduced by hydrogen.

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