



## Ordinary Thinking

### Objective Questions

#### Electrolytes and Electrolysis

- Which of the following will not conduct electricity in aqueous solution [AMU 1982, 83]
  - Copper sulphate
  - Sugar
  - Common salt
  - None of these
- Strong electrolytes are those which [MNR 1983]
  - Dissolve readily in water
  - Conduct electricity
  - Dissociate into ions at high dilution
  - Completely dissociate into ions at all dilutions
- In aqueous solution, strong electrolytes [AMU 1983, 84]
  - Are partially ionized
  - Do not ionise
  - Ionise almost completely
  - Form polymers
- An electrolyte [KCET 1984; MP PET/PMT 1988]
  - Forms complex ions in solution
  - Gives ions only when electricity is passed
  - Possesses ions even in solid state
  - Gives ions only when dissolved in water
- Electrolytes when dissolved in water dissociates into ions because [CPMT 1974, 78; MNR 1983]
  - They are unstable
  - The water dissolves it
  - The force of repulsion increases
  - The forces of electrostatic attraction are broken down by water
- Electrolyte can conduct electricity because
  - Their molecules contain unpaired electrons, which are mobile
  - Their molecules contain loosely held electrons which get free under the influence of voltage
  - The molecules break up into ions when a voltage is applied
  - The molecules are broken up into ions when the electrolyte is fused or is dissolved in the solvent
- Which one of the following metals could not be obtained on electrolysis of aqueous solution of its salts [IIT 1990]
  - Ag
  - Mg
  - Cu
  - Cr
- Which of the following aqueous solution will conduct an electric current quite well [MP PMT 1987]
  - Glycerol
  - HCl
  - Sugar
  - Pure water
- On the electrolysis of aqueous solution of sodium sulphate, on cathode we get
  - Na
  - H<sub>2</sub>
  - SO<sub>2</sub>
  - SO<sub>3</sub>
- Electrolysis involves oxidation and reduction respectively at [CPMT 1973; AMU 1983; NCERT 1983, 84; MH CET 2001]
  - Anode and cathode
  - Cathode and anode
  - At both the electrodes
  - None of the above
- Which of the following compounds will not undergo decomposition on passing electricity through aqueous solution
  - Sugar
  - Sodium Chloride
  - Sodium Bromide
  - Sodium Acetate
- During the electrolysis of an electrolyte, the number of ions produced, is directly proportional to the [AFMC 2002]
  - Time consumed
  - Electro chemical equivalent of electrolysis
  - Quantity of electricity passed
  - Mass of electrons
- 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
- In the electrolytic cell, flow of electrons is from [IIT Screening 2003]
  - Cathode to anode in solution
  - Cathode to anode through external supply
  - Cathode to anode through internal supply
  - Anode to cathode through internal supply
- An electric current is passed through an aqueous solution of the following. Which one shall decompose [NCERT 1972]
  - Urea
  - Glucose
  - AgNO<sub>3</sub>
  - Ethyl alcohol
- The electric conduction of a salt solution in water depends on the
  - Shape of its molecules
  - Size of its molecules
  - Size of solvent molecules
  - Extent of its ionization
- A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively [IIT 1987, 96]
  - H<sub>2</sub>, O<sub>2</sub>
  - O<sub>2</sub>, H<sub>2</sub>
  - O<sub>2</sub>, Na
  - O<sub>2</sub>, SO<sub>2</sub>

18. On electrolysing a solution of dilute  $H_2SO_4$  between platinum electrodes, the gas evolved at the anode is  
[NCERT 1977, 79; MNR 1980; CBSE PMT 1992]  
(a)  $SO_2$  (b)  $SO_3$   
(c)  $O_2$  (d)  $H_2$
19. The addition of a polar solvent to a solid electrolyte results in  
(a) Polarization (b) Association  
(c) Ionization (d) Non-liberation of heat
20. During the electrolysis of fused  $NaCl$ , which reaction occurs at anode [NCERT 1973; AFMC 1992; MP PMT 2002]  
(a) Chloride ions are oxidized  
(b) Chloride ions are reduced  
(c) Sodium ions are oxidised  
(d) Sodium ions are reduced
21. The amount of ion discharged during electrolysis is not directly proportional to [NCERT 1973]  
(a) Resistance  
(b) Time  
(c) Current  
(d) Chemical equivalent of the ion
22. Electrolysis of aqueous  $HCl$  solution produces [CPMT 1987]  
(a)  $H_2$  gas at the anode  
(b)  $H_2$  gas at the cathode  
(c)  $Cl_2$  gas at the cathode  
(d)  $Cl_2$  and  $O_2$  gases both at the anode
23. During electrolysis of  $NaCl$  solution, part of the reaction is  $Na^+ + e^- \rightarrow Na$ . This is termed as [NCERT 1984]  
(a) Oxidation (b) Reduction  
(c) Deposition (d) Cathode reaction
24. When a solution of an electrolyte is heated the conductance of the solution  
(a) Increases because of the electrolyte conducts better  
(b) Decreases because of the increased heat  
(c) Decreases because of the dissociation of the electrolyte is suppressed  
(d) Increases because the electrolyte is dissociated more
25. The passage of current liberates  $H_2$  at cathode and  $Cl_2$  at anode. The solution is [EAMCET 1979, 87]  
(a) Copper chloride in water (b)  $NaCl$  in water  
(c)  $H_2SO_4$  (d) Water
26. Pure water does not conduct electricity because it [Manipal MEE 1995]  
(a) Has a low boiling point  
(b) Is almost totally unionized  
(c) Is neutral  
(d) Is readily decomposed
27. Which is responsible for electrical conduction of molten sodium chloride [MADT Bihar 1995]  
(a) Free electrons  
(b) Free ions  
(c) Free molecules  
(d) Atoms of sodium and chlorine
28. In electrolysis of aqueous copper sulphate, the gas at anode and cathode is [AFMC 1995]  
(a)  $O_2$  and  $H_2$  (b)  $SO_2$  and  $H_2$   
(c)  $H_2$  and  $O_2$  (d)  $SO_3$  and  $O_2$
29. Use of electrolysis is [AFMC 1995]  
(a) Electroplating (b) Electrorefining  
(c) (a) and (b) both (d) None of these
30. Sodium is made by the electrolysis of a molten mixture of about 40%  $NaCl$  and 60%  $CaCl_2$  because [CBSE PMT 1995]  
(a)  $CaCl_2$  helps in conduction of electricity  
(b) This mixture has a lower melting point than  $NaCl$   
(c)  $Ca^{++}$  can displace  $Na$  from  $NaCl$   
(d)  $Ca^{++}$  can reduce  $NaCl$  to  $Na$
31. Electrolysis is a process in which the cations and anions of the electrolyte are  
(a) Hydrated (b) Hydrolysed  
(c) Charged (d) Discharged
32. Degree of ionisation of a solution depends upon [BHU 1998]  
(a) Temperature (b) Nature of the electrolyte  
(c) Nature of the solvent (d) None of these
33. Which of the following is non-electrolytes [KCET (Med.) 1999]  
(a)  $NaCl$  (b)  $CaCl_2$   
(c)  $C_{12}H_{22}O_{11}$  (d)  $CH_3COOH$
34. When a molten ionic hydride is electrolysed [JIPMER 1999]  
(a) Hydrogen is liberated at the cathode  
(b) Hydrogen is liberated at the anode  
(c) There is no reaction  
(d)  $H^-$  ions produced migrate to the cathode
35. During electrolysis, the species discharged at cathode are [AFMC 2000]  
(a) Ions (b) Cation  
(c) Anion (d) All of these
36. Electrolysis of molten anhydrous calcium chloride produces [AIIMS 2000]

## 508 Electrochemistry

- (a) Calcium (b) Phosphorus [EAMCET 1990; MP PET 1994, 97]  
(c) Sulphur (d) Sodium
37. Which of the following properties of pure metal makes it more useful than the corresponding alloy [RPET 2000]  
(a) It is harder than corresponding alloy  
(b) It has high density  
(c) It can be extracted easily  
(d) It conducts heat and electricity easily
38. Which of the following liberate hydrogen on reaction with dilute  $H_2SO_4$   
(a) Fe (b) Cu  
(c) Al (d) Hg
39. Which one of the following material conducts electricity [Kerala (Med.) 2003]  
(a) Diamond  
(b) Crystalline sodium chloride  
(c) Barium sulphate  
(d) Fused potassium chloride  
(e) Molten sulphur
40. Which of the following metals will give  $H_2$  on reaction with NaOH [RPET 2003]  
(a) Mg (b) Ba  
(c) Ca (d) Sr
41. Which of the following is not a non electrolyte [J & K 2005]  
(a) Acetic acid (b) Glucose  
(c) Ethanol (d) Urea
- (a) 9.89 g (b) 107.87 g  
(c) 1.0787 g (d) 1.002 g
5. The atomic weight of Al is 27. When a current of 5 Faradays is passed through a solution of  $Al^{+++}$  ions, the weight of Al deposited is  
(a) 27 gm (b) 36 gm  
(c) 45 gm (d) 39 gm
6. An apparatus used for the measurement of quantity of electricity is known as a [BHU 1979]  
(a) Calorimeter (b) Cathetometer  
(c) Coulometer (d) Colorimeter
7. The unit of electrochemical equivalent is [EAMCET 1980]  
(a) Gram (b) Gram/ampere  
(c) Gram/coulomb (d) Coulomb/gram
8. A certain current liberated 0.504 gm of hydrogen in 2 hours. How many grams of copper can be liberated by the same current flowing for the same time in a copper sulphate solution [NCERT 1973, 77; C  
(a) 12.7 gm (b) 15.9 gm  
(c) 31.8 gm (d) 63.5 gm
9. What weight of copper will be deposited by passing 2 Faradays of electricity through a cupric salt (Atomic weight of Cu = 63.5) [NCERT 1975; CPMT 1977, C  
(a) 2.0 gm (b) 3.175 gm  
(c) 63.5 gm (d) 127.0 gm
- If the current is passed into the solution of an electrolyte [AIIMS 1979]  
(a) Anions move towards anode, cations towards cathode  
(b) Anions and cations both move towards anode  
(c) Anions move towards cathode, cations towards anode  
(d) No movement of ions takes place

### Faraday's law of electrolysis

1. Amount of electricity that can deposit 108 gm of silver from  $AgNO_3$  solution is [AFMC 1993; MP PMT 2004]  
(a) 1 ampere (b) 1 coulomb  
(c) 1 faraday (d) None of the above
2. When 9.65 coulombs of electricity is passed through a solution of silver nitrate (atomic weight of Ag = 107.87 taking as 108) the amount of silver deposited is [EAMCET 1992; KCET 2000]  
(a) 10.8 mg (b) 5.4 mg  
(c) 16.2 mg (d) 21.2 mg
3. Three faradays electricity was passed through an aqueous solution of iron (II) bromide. The weight of iron metal (at. wt. = 56) deposited at the cathode (in gm) is [EAMCET 1991]  
(a) 56 (b) 84  
(c) 112 (d) 168
4. A silver cup is plated with silver by passing 965 coulombs of electricity, the amount of silver deposited is  
(a) 9.89 g (b) 107.87 g  
(c) 1.0787 g (d) 1.002 g
5. The atomic weight of Al is 27. When a current of 5 Faradays is passed through a solution of  $Al^{+++}$  ions, the weight of Al deposited is  
(a) 27 gm (b) 36 gm  
(c) 45 gm (d) 39 gm
6. An apparatus used for the measurement of quantity of electricity is known as a [BHU 1979]  
(a) Calorimeter (b) Cathetometer  
(c) Coulometer (d) Colorimeter
7. The unit of electrochemical equivalent is [EAMCET 1980]  
(a) Gram (b) Gram/ampere  
(c) Gram/coulomb (d) Coulomb/gram
8. A certain current liberated 0.504 gm of hydrogen in 2 hours. How many grams of copper can be liberated by the same current flowing for the same time in a copper sulphate solution [NCERT 1973, 77; C  
(a) 12.7 gm (b) 15.9 gm  
(c) 31.8 gm (d) 63.5 gm
9. What weight of copper will be deposited by passing 2 Faradays of electricity through a cupric salt (Atomic weight of Cu = 63.5) [NCERT 1975; CPMT 1977, C  
(a) 2.0 gm (b) 3.175 gm  
(c) 63.5 gm (d) 127.0 gm
- If the current is passed into the solution of an electrolyte [AIIMS 1979]  
(a) Anions move towards anode, cations towards cathode  
(b) Anions and cations both move towards anode  
(c) Anions move towards cathode, cations towards anode  
(d) No movement of ions takes place
11. Unit of Faraday is  
(a) Ampere (b) Coulomb  
(c) Coulomb  $mole^{-1}$  (d) Coulomb  $Sec^{-1}$
12. On passing 0.1 Faraday of electricity through aluminium chloride, the amount of aluminium metal deposited on cathode is (Al = 27) [MP PMT 1991]  
(a) 0.9 gm (b) 0.3 gm  
(c) 0.27 gm (d) 2.7 gm
13. Which of the following represents the first law of Faraday [MP PMT 1991]  
(a)  $E = mc^2$  (b)  $E = hv$   
(c)  $m = ect$  (d)  $PV = nRT$
14. 5 amperes is passed through a solution of zinc sulphate for 40 minutes. Find the amount of zinc deposited at the cathode [CBSE PMT 1996]  
(a) 40.65 gm (b) 4.065 gm  
(c) 0.4065 gm (d) 65.04 gm

15. In an electroplating experiment  $m$  g of silver is deposited, when 4 amperes of current flows for 2 minutes. The amount (in  $gms$ ) of silver deposited by 6 amperes of current flowing for 40 seconds will be [MNR 1991]  
 (a)  $4m$  (b)  $m/2$   
 (c)  $m/4$  (d)  $2m$
16. On passing 3 ampere of electricity for 50 minutes, 1.8 gram metal deposits. The equivalent mass of metal is [MP PMT 1992]  
 (a) 20.5 (b) 25.8  
 (c) 19.3 (d) 30.7
17. The desired amount of charge for obtaining one mole of  $Al$  from  $Al^{3+}$   
 (a)  $3 \times 96500$  C (b) 96500 C  
 (c)  $\frac{96500}{3}$  C (d)  $\frac{96500}{2}$  C
18. On passing one faraday of electricity through the electrolytic cells containing  $Ag^+$ ,  $Ni^{+2}$  and  $Cr^{+3}$  ions solution, the deposited  $Ag$  (At. wt. = 108),  $Ni$  (At. wt. = 59) and  $Cr$  (At. wt. = 52) is [AIIMS 1982]  

$Ag$	$Ni$	$Cr$
(a) 108 gm	29.5 gm	17.3 gm
(b) 108 gm	59.0 gm	52.0 gm
(c) 108.0 gm	108.0 gm	108.0 gm
(d) 108 gm	117.5 gm	166.0 gm
19. One Faraday of electricity when passed through a solution of copper sulphate deposits [CPMT 1978]  
 (a) 1 mole of  $Cu$  (b) 1 gm atom of  $Cu$   
 (c) 1 molecule of  $Cu$  (d) 1 gm equivalent of  $Cu$
20. When 1 coulomb of charge is passed through electrolyte solution, then the mass deposited is equal to  
 (a) Equivalent weight  
 (b) Atomic weight  
 (c) Electrochemical equivalent  
 (d) Chemical equivalent
21. The platinum electrodes were immersed in a solution of cupric sulphate and electric current passed through the solution. After some time it was found that colour of copper sulphate disappeared with evolution of gas at the electrode. The colourless solution contains [NCERT 1984]  
 (a) Platinum sulphate (b) Copper hydroxide  
 (c) Copper sulphate (d) Sulphuric acid
22. On passing  $C$  ampere of electricity through a 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)  $E = \frac{C \times m}{t \times 96500}$   
 (c)  $E = \frac{96500 \times m}{C \times t}$  (d)  $E = \frac{C \times t \times 96500}{m}$
23. How many Faradays are required to generate one gram atom of magnesium from  $MgCl_2$  [MADT Bihar 1982]  
 (a) 1 (b) 2  
 (c) 3 (d) 4
24. To deposit 0.6354 gm of copper by electrolysis of aqueous cupric sulphate solution, the amount of electricity required (in coulombs) is  
 (a) 9650 (b) 4825  
 (c) 3860 (d) 1930
25. In electrolysis of a fused salt, the weight of the deposit on an electrode will not depend on [CPMT 1973]  
 (a) Temperature of the bath  
 (b) Current intensity  
 (c) Electrochemical equivalent of ions  
 (d) Time for electrolysis
26. Faraday's laws of electrolysis will fail when [NCERT 1971]  
 (a) Temperature is increased  
 (b) Inert electrodes are used  
 (c) A mixture of electrolytes is used  
 (d) In none of these cases
27. According to the first law of Faraday, the weight of a substance discharge at the electrode is  
 (a)  $W = ZQ$  (b)  $W = eF$   
 (c)  $W = \frac{Z}{F} It$  (d)  $W = ZI$
28. When 0.04 faraday of electricity is passed through a solution of  $CaSO_4$ , then the weight of  $Ca^{2+}$  metal deposited at the cathode is [BHU 1996]  
 (a) 0.2 gm (b) 0.4 gm  
 (c) 0.6 gm (d) 0.8 gm
29. A current 2.0 A is passed for 5 hours through a molten metal salt deposits 22 g of metal (At. wt. = 177). The oxidation state of the metal in the metal salt is [KCET 1996]  
 (a) + 1 (b) + 2  
 (c) + 3 (d) + 4
30. How many atoms of calcium will be deposited from a solution of  $CaCl_2$  by a current of 25 milliamperes flowing for 60 seconds  
 (a)  $4.68 \times 10^{18}$  (b)  $4.68 \times 10^{15}$   
 (c)  $4.68 \times 10^{12}$  (d)  $4.68 \times 10^9$
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 gm (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 g (b) 1.32 g  
 (c) 33 g (d) 99 g

## 510 Electrochemistry

33. On passing a current through  $KCl$  solution, 19.5 g of potassium is deposited. If the same quantity of electricity is passed through a solution of aluminium chloride, the amount of aluminium deposited is [EAMCET 1997]  
 (a) 4.5 g (b) 9.0 g  
 (c) 13.5 g (d) 27 g  
 (e) None is correct
34. Electrolysis rules of Faraday's states that mass deposited on electrode is proportional to [CBSE PMT 2000]  
 (a)  $m \propto I^2$  (b)  $m \propto Q$   
 (c)  $m \propto Q^2$  (d) None of these
35. A current being passed for two hour through a solution of an acid liberating 11.2 litre of oxygen at NTP at anode. What will be the amount of copper deposited at the cathode by the same current when passed through a solution of copper sulphate for the same time [BVP 2003]  
 (a) 16 g (b) 63 g  
 (c) 31.5 g (d) 8 g
36. In a metal oxide, there is 20% oxygen by weight. Its equivalent weight is [Pb. PMT 2000]  
 (a) 40 (b) 64  
 (c) 72 (d) 32
37. On the basis of the information available from the reaction  $\frac{4}{3}Al + O_2 \rightarrow \frac{2}{3}Al_2O_3, \Delta G = -827 kJmol^{-1}$  of  $O_2$ , the minimum *emf* required to carry out an electrolysis of  $Al_2O_3$  is ( $F = 96500 C mol^{-1}$ ) [CBSE PMT 2003]  
 (a) 8.56 V (b) 2.14 V  
 (c) 4.28 V (d) 6.42 V
38. Then during electrolysis of a solution of  $AgNO_3$ , 9650 coulombs of charge pass through the electroplating bath, the mass of silver deposited in the cathode will be [AIEEE 2003]  
 (a) 1.08 g (b) 10.8 g  
 (c) 21.6 g (d) 108 g
39. Total charge on 1 mole of a monovalent metal ion is equal to [DPMT 2001]  
 (a)  $9.65 \times 10^4$  Coulomb (b)  $6.28 \times 10^{18}$  Coulomb  
 (c)  $1.6 \times 10^{-19}$  Coulomb (d) None of these
40. When an electric current is passed through acidulated water 112 ml of hydrogen gas at N.T.P. collect at the cathode in 965 seconds. The current passed, in amperes is [MNR 1991; UPSEAT 2001]  
 (a) 1.0 (b) 0.5  
 (c) 0.1 (d) 2.0
41. How much chlorine will be liberated on passing one ampere current for 30 minutes through  $NaCl$  solution [BVP 2003]  
 (a) 0.66 mole (b) 0.33 mole  
 (c) 0.66 gm (d) 0.33 gm
42. The number of electrons involved in redox reactions when a Faraday of electricity is passed through an electrolyte in solution is [NCERT 1982; Pb. PMT 1992]  
 (a)  $6 \times 10^{23}$  (b)  $6 \times 10^{-23}$   
 (c) 96500 (d)  $8 \times 10^{19}$
43. Coulomb is equal to [Orissa JEE 2002]  
 (a) ampere  $\times$  second (b) ampere  $\times$  minute  
 (c) watt  $\times$  second (d) volt  $\times$  second
44. The energy required to release 1 electron from  $He^+$  is..... [Orissa JEE 2002]  
 (a) + 54.4 eV (b) - 13.6 eV  
 (c) + 27.2 eV (d) Cannot be predicted
45. Faraday's laws of electrolysis are related to the [IIT 1983]  
 (a) Atomic number of cation  
 (b) Atomic number of anion  
 (c) Equivalent weight of the electrolyte  
 (d) Speed of the cation
46. The electric charge for electrode decomposition of one gram equivalent of a substance is [IIT 1984; KCET 1992]  
 (a) One ampere per second  
 (b) 96500 coulombs per second  
 (c) One ampere for one hour  
 (d) Charge on one mole of electrons
47. The number of electrons passing per second through a cross-section of copper wire carrying  $10^{-6}$  amperes of current per second is found to be [EAMCET 1992]  
 (a)  $1.6 \times 10^{-19}$  (b)  $6 \times 10^{-35}$   
 (c)  $6 \times 10^{-16}$  (d)  $6 \times 10^{12}$
48. The electrolytic cells, one containing acidified ferrous chloride and another acidified ferric chloride are connected in series. The ratio of iron deposited at cathodes in the two cells when electricity is passed through the cells will be [CPMT 1989]  
 (a) 3 : 1 (b) 2 : 1  
 (c) 1 : 1 (d) 3 : 2
49. When 96500 coulomb of electricity is passed through a copper sulphate solution, the amount of copper deposited will be [MP PMT 1996]  
 (a) 0.25 mol (b) 0.50 mol  
 (c) 1.00 mol (d) 2.00 mol
50. During electrolysis of fused aluminium chloride 0.9 gm of aluminium was deposited on the cathode. The volume of chlorine liberated at the anode will be  
 (a) 2.24 litres (b) 11.2 litres  
 (c) 1.12 litres (d) 5.6 litres
51. Faraday has the dimensions of [MP PET 1995]  
 (a) Coulombs  
 (b) Coulomb equivalent

- (c) Coulomb per equivalent  
(d) Coulomb per degree Kelvin
52. The required charge for one equivalent weight of silver deposited on cathode 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 coulombs of electric current liberates from  $CuSO_4$  solution [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 amp was passed through  $CuSO_4$  solution for 6 minutes 26 seconds. The amount of copper deposited is  
(Atomic weight of Cu = 63.5)  
(1 faraday = 96500 coulombs)  
[EAMCET 1989; MP PET 1994]  
(a) 0.3175 g (b) 3.175 g  
(c) 0.635 g (d) 6.35 g
55. A certain quantity of electricity is passed through an aqueous solution of  $AgNO_3$  and cupric salt solution connected in series. The amount of Ag deposited is 1.08 gm, the amount of copper deposited is (atomic weight of Cu = 63.5; Ag = 108) [EAMCET 1986]  
(a) 0.6454 g (b) 6.354 g  
(c) 0.3177 g (d) 3.177 g
56. The number of electrons required to deposit 1 gm atom of aluminium (at. wt. = 27) from a solution of aluminium chloride will be (where  $N$  is Avogadro's number) [AIIMS 1992]  
(a) 1 N (b) 2 N  
(c) 3 N (d) 4 N
57. Three faradays of electricity are passed through molten  $Al_2O_3$ , aqueous solution of  $CuSO_4$  and molten  $NaCl$  taken in different electrolytic cells. The amount of Al, Cu and Na deposited at the cathodes will be in the ratio of [BHU 1990]  
(a) 1 mole : 2 mole : 3 mole  
(b) 3 mole : 2 mole : 1 mole  
(c) 1 mole : 1.5 mole : 3 mole  
(d) 1.5 mole : 2 mole : 3 mole
58. An electrolytic cell contains a solution of  $Ag_2SO_4$  and have platinum electrodes. A current is passed until 1.6 gm of  $O_2$  has been liberated at anode. The amount of silver deposited at cathode would be  
(a) 107.88 gm (b) 1.6 gm  
(c) 0.8 gm (d) 21.60 gm
59. The aqueous solution of which of the following decomposes on passing electric current [EAMCET 1973]  
(a) Canesugar (b) Urea  
(c) Methanol (d) Potassium iodide
60. The number of Faradays needed to reduce 4 gram equivalents of  $Cu^{++}$  to Cu metal will be [BHU 1981]  
[Roorkee 1995] (a) 1 (b) 2  
(c) 1/2 (d) 4
61. When electricity is passed through the solution of  $AlCl_3$ , 13.5 gm of Al are 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 Faraday is [AMU 1983; AFMC 1989; MP PET 2001]  
(a) 95500  $C mol^{-1}$  (b) 96550  $C mol^{-1}$   
(c) 96500  $C mol^{-1}$  (d) 98500  $C mol^{-1}$
63. The quantity of electricity needed to liberate 0.5 gram equivalent of an element is [CPMT 1988; MP PMT 1997]  
(a) 48250 Faradays (b) 48250 Coulombs  
(c) 193000 Faradays (d) 193000 Coulombs
64. The number of coulombs required for the deposition of 107.870 g of silver is [MP PET/PMT 1998]  
(a) 96,500 (b) 48,250  
(c) 1,93,000 (d) 10,000
65. When one of ampere current flows for 1 sec through a conductor, this quantity of electricity is known as [MP PMT 1993]  
(a) Faraday (b) Coulomb  
(c) E.M.F. (d) Ohm
66. The mass deposited at an electrode is directly proportional to [CPMT 1974; AFMC 1987; MP PET 2000]  
(a) Atomic weight (b) Equivalent weight  
(c) Molecular weight (d) Atomic number
67. From the solution of which of the following one faraday of electricity will liberate one gram atom of metal [MHCET 1999; MP PET 1993, 2000; AFMC 2000]  
(a) NaCl (b)  $BaCl_2$   
(c)  $CuSO_4$  (d)  $AlCl_3$
68. On electrolysis, 1 mole of aluminium will be deposited from its molten salt by [MH CET 2000]  
(a) 3 moles of electrons (b) 4 moles of electrons  
(c) 2 moles of electrons (d) 1 mole of electrons [CPMT 1971]
69. The atomic weight of Fe is 56. The weight of Fe deposited from  $FeCl_3$  solution by passing 0.6 Faraday of electricity is [MH CET 2000]  
(a) 5.6 g (b) 11.2 g  
(c) 22.4 g (d) 33.6 g

## 512 Electrochemistry

70. 2.5  $F$  of electricity are passed through a  $\text{CuSO}_4$  solution. The number of gm equivalent of  $\text{Cu}$  deposited on anode is  
[CPMT 1973; DPMT 1982; MP PMT 2001]  
(a) Zero (b) 1.25  
(c) 2.5 (d) 5.0
71. The equivalent weight of a certain trivalent element is 20. Molecular weight of its oxide is [KCET 2003]  
(a) 152 (b) 56  
(c) 168 (d) 68
72. Silver is removed electrically from 200 ml of a 0.1  $N$  solution of  $\text{AgNO}_3$  by a current of 0.1 ampere. How long will it take to remove half of the silver from the solution [AMU 1999]  
(a) 16 sec (b) 96.5 sec  
(c) 100 sec (d) 10 sec
73. In order to separate oxygen from one mole of  $\text{H}_2\text{O}$  the required quantity of coulomb would be [RPET 1999]  
(a)  $1.93 \times 10^5$  (b)  $9.6 \times 10^4$   
(c) 1.8 (d) 3.2
74. A current of 0.25A is passed through  $\text{CuSO}_4$  solution placed in voltameter for 45 minutes. The amount of  $\text{Cu}$  deposited on cathode is (At weight of  $\text{Cu} = 63.6$ ) [BHU 2001]  
(a) 0.20 g (b) 0.22 g  
(c) 0.25 g (d) 0.30 g
75. Faraday constant [KCET (Med.) 2001]  
(a) Is a numerical constant  
(b) Depends on equivalent  
(c) Depends upon the current passed  
(d) Depends on the number of electrons
76. If 0.5 amp current is passed through acidified silver nitrate solution for 10 minutes. The mass of silver deposited on cathode, is (eq. wt. of silver nitrate = 108) [AFMC 2001]  
(a) 0.235 g (b) 0.336 g  
(c) 0.536 g (d) 0.636 g
77. The unit for the electric current is [KCET (Med.) 2001]  
(a) Ohm (b) Volt  
(c) Ampere (d) Coulomb
78. The quantity of electricity required to liberate  $112 \text{ cm}^3$  of hydrogen at STP from acidified water is [KCET (Med.) 2001]  
(a) 0.1 Faraday (b) 1 Faraday  
(c) 965 Coulomb (d) 96500 Coulomb
79. Which solution will show highest resistance during the passage of current  
(a) 0.05  $N$   $\text{NaCl}$  (b) 2  $N$   $\text{NaCl}$   
(c) 0.1  $N$   $\text{NaCl}$  (d) 1  $N$   $\text{NaCl}$
80. 4 g of copper was dissolved in concentrated nitric acid. The copper nitrate solution on strong heating gave 5 g of its oxide. The equivalent weight of copper is [KCET 2004]  
(a) 23 (b) 32  
(c) 12 (d) 20
81. The amount of silver deposited by passing 241.25 coulomb of current through silver nitrate solution is [MHCET 2003]  
(a) 2.7 g (b) 2.7 mg  
(c) 0.27 g (d) 0.54 g
82. When 1F of electricity is passed through acidulated water,  $\text{O}_2$  evolved is  
(a)  $11.2 \text{ dm}^3$  (b)  $5.6 \text{ dm}^3$   
(c)  $22.4 \text{ dm}^3$  (d)  $1.0 \text{ dm}^3$
83. Charge required to liberate 11.5 g sodium is [AIIMS 1992; DCE 2002]  
(a) 0.5  $F$  (b) 0.1  $F$   
(c) 1.5  $F$  (d) 96500 coulombs
84. In the electrolysis of water, one Faraday of electrical energy would evolve  
(a) One mole of oxygen (b) One g atom of oxygen  
(c) 8 g of oxygen (d) 22.4 litres of oxygen
85. In a galvanic cell, the electrons flow from [KCET 2004]  
(a) Anode to cathode through the solution  
(b) Cathode to anode through the solution  
(c) Anode to cathode through the external circuit  
(d) Cathode to anode through the external circuit
86. An electric current is passed through silver nitrate solution using silver electrodes. 10.79 g of silver was found to be deposited on the cathode if the same amount of electricity is passed through copper sulphate solution using copper electrodes, the weight of copper deposited on the cathode is [Kerala PMT 2004]  
(a) 6.4 g (b) 2.3 g  
(c) 12.8 g (d) 1.6 g  
(e) 3.2 g
87. The law of electrolysis were proposed by [CPMT 1982; Pb.CET 2003]  
(a) Kohlraush (b) Faraday  
(c) Nernst (d) Berthelot
88. How many atoms of calcium will be deposited from a solution of  $\text{CaCl}_2$  by a current 0.25 mA following for 60 seconds [BHU 2004]  
(a)  $4.68 \times 10^{18}$  (b)  $4.68 \times 10^{15}$   
(c)  $4.68 \times 10^{12}$  (d)  $4.68 \times 10^9$   
[BHU 2001]
89. The number of coulombs required to reduce 12.3 g of nitrobenzene to aniline  
(a) 115800 C (b) 5790 C  
(c) 28950 C (d) 57900 C

90. During the process of electrolytic refining of copper, some metals present as impurity settle as 'anode mud'. These are  
[AIEEE 2005]  
(a) *Sn* and *Ag* (b) *Pb* and *Zn*  
(c) *Ag* and *Au* (d) *Fe* and *Ni*
91. A galvanic cell is set up from a zinc bar weighing 50g and 1.0 litre, 1.0M,  $\text{CuSO}_4$  solution. How long would the cell run, assuming it delivers a steady current of 1.0 ampere  
[Roorkee 2000]  
(a) 48 hrs (b) 41 hrs  
(c) 21 hrs (d) 1 hr
92. On passing electric current through molten aluminium chloride, 11.2 litre of  $\text{Cl}_2$  is liberated at NTP at anode. The quantity of aluminium deposited at cathode is (at. wt. of *Al* = 27)  
(a) 9 g (b) 18 g  
(c) 27 g (d) 36 g
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)  $56 \text{ cm}^3$  (b)  $550 \text{ cm}^3$   
(c)  $5.6 \text{ cm}^3$  (d)  $11.2 \text{ cm}^3$   
(e)  $22.4 \text{ cm}^3$
94. During electrolysis of aqueous  $\text{NaOH}$ , 4 g of  $\text{O}_2$  gas is liberated at NTP at anode,  $\text{H}_2$  gas liberated at cathode is  
[CBSE PMT 1998]  
(a) 2.8 litres (b) 5.6 litres  
(c) 11.2 litres (d) 22.4 litres
3. The factor which is not affecting the conductivity of any solution is  
(a) Dilution (b) Nature of electrolyte  
(c) Temperature (d) None of these
4. Specific conductance of 0.1 m nitric acid is  $6.3 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1}$ . The molar conductance of solution is  
[Kerala (Med.) 2003]  
(a)  $630 \text{ ohm}^{-1} \text{ cm}^2 \text{ mole}^{-1}$  (b)  $315 \text{ ohm}^{-1} \text{ cm}^2 \text{ mole}^{-1}$   
(c)  $100 \text{ ohm}^{-1} \text{ cm}^2 \text{ mole}^{-1}$  (d)  $6300 \text{ ohm}^{-1} \text{ cm}^2 \text{ mole}^{-1}$   
(e)  $63.0 \text{ ohm}^{-1} \text{ cm}^2 \text{ mole}^{-1}$
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) Decreases on dilution
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 2002]  
(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)  $\text{Sm mol}^{-1}$  (b)  $\text{Sm}^2 \text{ mol}^{-1}$   
(c)  $\text{S}^{-2} \text{ m}^2 \text{ mol}$  (d)  $\text{S}^2 \text{ m}^2 \text{ mol}^{-2}$
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 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

### Conductor and conductance

1. Which one of the following statements is correct  
[MP PET 1997]  
(a) The oxidation number of oxygen in  $\text{KO}_2$  is zero  
(b) The specific conductance of an electrolyte solution decreases with increase in dilution  
(c)  $\text{Sn}^{2+}$  oxidises  $\text{Fe}^{3+}$   
(d)  $\text{Zn}/\text{ZnSO}_4$  is a reference electrode
2. In infinite dilutions, the equivalent conductances of  $\text{Ba}^{2+}$  and  $\text{Cl}^-$  are 127 and  $76 \text{ ohm}^{-1} \text{ cm}^{-1} \text{ eqvt}^{-1}$ . The equivalent conductivity of  $\text{BaCl}_2$  at indefinite dilution is  
[CBSE 2000]  
(a) 101.5 (b) 139.5  
(c) 203.5 (d) 279.5



## 514 Electrochemistry

- (d) Degree of ionisation of the electrolyte
11. Which of the following conducts electricity [AFMC 1995]  
 (a) Fused  $\text{NaCl}$  (b)  $\text{CO}_2$   
 (c)  $\text{Br}_2$  (d)  $\text{Si}$
12. Which of the following shows electrical conduction [CBSE PMT 1999; AIIMS 1999]  
 (a) Potassium (b) Graphite  
 (c) Diamond (d) Sodium
13. The unit of equivalent conductivity is [CPMT 1999; BCECE 2005]  
 (a)  $\text{ohm cm}$   
 (b)  $\text{ohm}^{-1}\text{cm}^2 (\text{gm equivalent})^{-1}$   
 (c)  $\text{ohm cm}^2 (\text{gm equivalent})$   
 (d)  $\text{S cm}^{-2}$
14. It has been observed that gaseous hydrogen chloride is a very poor conductor of electricity but a solution of hydrogen chloride gas in water is a good conductor of electricity. This is due to the fact that [NCERT 1976]  
 (a) Water is good conductor of electricity  
 (b) Hydrogen chloride gas in water solution ionizes  
 (c) A gas is non-conductor but a liquid conducts electricity  
 (d) Gas does not obey Ohm's law whereas solution does
15. Electrolytic conduction differs from metallic conduction in that in the case of electrolytic conduction [KCET 1987; Bihar CEE 1992]  
 (a) The resistance increases with increasing temperature  
 (b) The resistance decreases with increasing temperature  
 (c) The flow of current does not generate heat  
 (d) The resistance is independent of the length of the conductor
16. The electrolytic conductance is a direct measure of [KCET 1990; CPMT 2003]  
 (a) Resistance (b) Potential  
 (c) Concentration (d) Dissociation
17. Conductivity of a strong electrolyte [KCET 1993]  
 (a) Increases on dilution  
 (b) Does not change considerably on dilution  
 (c) Decreases on dilution  
 (d) Depends on density
18. Which of the following statements is not applicable to electrolytic conductors [AIIMS 1991]  
 (a) New products show up at the electrodes  
 (b) Ions are responsible for carrying the current  
 (c) Show a positive temperature coefficient for conductance  
 (d) A single stream of electrons flows from cathode to anode
19. Which one is not a conductor of electricity [RPET 1999]  
 (a)  $\text{NaCl}$  (aqueous) (b)  $\text{NaCl}$  (solid)  
 (c)  $\text{NaCl}$  (molten) (d) Ag metal
20. 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
21. Which of the following is a poor conductor of electricity [EAMCET 1992]  
 (a)  $\text{CH}_3\text{COONa}$  (b)  $\text{C}_2\text{H}_5\text{OH}$   
 (c)  $\text{NaCl}$  (d)  $\text{KOH}$
22. The molar conductivity is maximum for the solution of concentration [DCE 2002]  
 (a) 0.001 M (b) 0.005 M  
 (c) 0.002 M (d) 0.004 M
23. The unit of molar conductivity is [DCE 2002]  
 (a)  $\Omega^{-1}\text{cm}^{-2}\text{mol}^{-1}$  (b)  $\Omega \text{cm}^{-2}\text{mol}^{-1}$   
 (c)  $\Omega^{-1}\text{cm}^2\text{mol}^{-1}$  (d)  $\Omega \text{cm}^2\text{mol}$
24. The highest electrical conductivity of the following aqueous solutions is of  
 (a) 0.1 M acetic acid (b) 0.1 M chloroacetic acid  
 (c) 0.1 M fluoroacetic acid (d) 0.1 M difluoroacetic acid
25. Given  $l/a = 0.5\text{cm}^{-1}$ ,  $R = 50\text{ohm}$ ,  $N = 1.0$ . The equivalent conductance of the electrolytic cell is [Orissa JEE 2002]  
 (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)  $300 \text{ohm}^{-1}\text{cm}^2 \text{gmeq}^{-1}$  (d)  $100 \text{ohm}^{-1}\text{cm}^2 \text{gmeq}^{-1}$
26. If equivalent conductance of 1M benzoic acid is  $12.8\text{ohm}^{-1}\text{cm}^2$  and if the conductance of benzoate ion and  $\text{H}^+$  ion are 42 and  $288.42 \text{ohm}^{-1}\text{cm}^2$  respectively. its degree of dissociation is  
 (a) 39% (b) 3.9%  
 (c) 0.35% (d) 0.039%
27. The unit  $\text{ohm}^{-1}$  is used for  
 (a) Molar conductivity (b) Equivalent conductivity  
 (c) Specific conductivity (d) Conductivity

### Cell constant and Electrochemical Cells

1. When electric current is passed through a cell having an electrolyte, the positive ions move towards 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]  
 (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 is  $E = E^\circ - \frac{RT}{nF} \ln a_{M^{n+}}$   
 (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)  $0.918 \text{ cm}^{-1}$  (d)  $1.12 \text{ cm}^{-1}$
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 [BHU 1981]  
 (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]  
 (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)  $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$   
 (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)  $PbO_2$  is dissolved  
 (b)  $H_2SO_4$  is regenerated  
 (c)  $PbSO_4$  is deposited on lead electrode  
 (d) Lead is deposited on lead electrode
9. 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  
 (b) Reduction  
 (c) Oxidation and reduction both  
 (d) Neutralisation
11. Which colourless gas evolves, when  $NH_4Cl$  reacts with zinc in a dry cell battery [Orissa JEE 2003]  
 (a)  $NH_4$  (b)  $N_2$   
 (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)  $H_2$  is evolved 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  $Zn | Zn^{2+} || Cu^{2+} | Cu$ , the negative electrode is

[MP PMT 1995]

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

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

19. Hydrogen-oxygen fuel cells are used in space-craft 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

$Zn^0|ZnSO_4||CuSO_4|Cu^0$  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  $Cu + 2Ag^+ \rightarrow Cu^{+2} + 2Ag$  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]

- (a) Weston cell (b) Daniel cell  
(c) Calomel cell (d) Faraday cell  
(e) Standard cell

25. The specific conductance of a solution is  $0.2 \text{ ohm}^{-1}\text{cm}^{-1}$  and conductivity is  $0.04 \text{ ohm}^{-1}$ . The cell constant would be [RPET 1999]

- (a)  $1 \text{ cm}^{-1}$  (b)  $0 \text{ cm}^{-1}$   
(c)  $5 \text{ cm}^{-1}$  (d)  $0.2 \text{ cm}^{-1}$

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  $KNO_3$  is used to make 'salt-bridge' 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

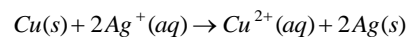
28. In balancing the half reaction  $S_2O_3^{2-} \rightarrow S_{(s)}$  the number of electrons that must be added is [DPMT 2000]

- (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



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  $CH_3COONa$  at infinite dilution are 126.45, 426.16 and  $91 \text{ ohm}^{-1}\text{cm}^2\text{mol}^{-1}$  respectively. The molar conductance of  $CH_3COOH$  at infinite dilution is [CBSE PMT 1999]

- (a)  $201.28 \text{ ohm}^{-1}\text{cm}^2\text{mol}^{-1}$   
(b)  $390.71 \text{ ohm}^{-1}\text{cm}^2\text{mol}^{-1}$   
(c)  $698.28 \text{ ohm}^{-1}\text{cm}^2\text{mol}^{-1}$   
(d)  $540.48 \text{ ohm}^{-1}\text{cm}^2\text{mol}^{-1}$

33. The electrodes of a conductivity cell are  $3\text{ cm}$  apart and have a cross-sectional area of  $4\text{ cm}^2$ . The cell constant of the cell (in  $\text{cm}^{-1}$ ) is  
 (a)  $4 \times 3$  (b)  $4/3$   
 (c)  $3/4$  (d)  $9/4$
34. The anode half-reaction occurring during the discharge of a lead storage battery is  
 (a)  $\text{Pb}(s) + \text{SO}_2 + \text{O}_2 \rightarrow \text{PbSO}_4(s)$   
 (b)  $\text{Pb}(s) + \text{SO}_4^{2-}(aq) \rightleftharpoons \text{PbSO}_4(s) + 2e^-$   
 (c)  $\text{PbO}_2(s) + 4\text{H}^+(aq) + 2e^- + \text{SO}_4^{2-}(aq) \rightleftharpoons \text{PbSO}_4(s) + 2\text{H}_2\text{O}$   
 (d)  $\text{Pb}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s)$
35. The unit of cell constant is [MP PET 1996]  
 (a)  $\text{ohm}^{-1}\text{ cm}^{-1}$  (b)  $\text{ohm cm}$   
 (c)  $\text{cm}$  (d)  $\text{cm}^{-1}$
36. In dry cell the reaction which takes place at the zinc anode is [MP PET 1996]  
 (a)  $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}(s)$   
 (b)  $\text{Zn}(s) \rightarrow \text{Zn}^{2+} + 2e^-$   
 (c)  $\text{Mn}^{2+} + 2e^- \rightarrow \text{Mn}(s)$   
 (d)  $\text{Mn}(s) \rightarrow \text{Mn}^{2+} + e^- + 1.5\text{V}$
37. The chemical reaction taking place at the anode of 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)  $\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}$   
 (b)  $\text{Pb}^{2+} + \text{SO}_4^{2-} \rightarrow \text{PbSO}_4$   
 (c)  $\text{Pb} \rightarrow \text{Pb}^{2+} + 2e^-$   
 (d)  $\text{PbSO}_4 + 2\text{H}_2\text{O} \rightarrow 2\text{PbO}_2 + 4\text{SO}_4^{2-} + 2e^-$
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)  $\text{SO}_2$  is evolved  
 (b) Lead sulphate is consumed  
 (c) Lead is formed  
 (d) Sulphuric acid is consumed
41. In electroplating, the article to be electroplated serves as [AMU 1982, 83]  
 (a) Cathode (b) Electrolyte  
 (c) Anode (d) Conductor
42. The position of some metals in the electrochemical series in decreasing electropositive character is given as  $\text{Mg} > \text{Al} > \text{Zn} > \text{Cu} > \text{Ag}$ . 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  $\text{Al}$   
 (b) An alloy of  $\text{Cu}$  and  $\text{Al}$  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
45. If a strip of  $\text{Cu}$  metal is placed in a solution of ferrous sulphate [NCERT 1974; CPMT 1977; MP PET 2000]  
 (a) Copper will precipitate out  
 (b) Iron will precipitate out  
 (c) Copper will dissolve  
 (d) No reaction will take place
46. Which of the following is not used to construct salt bridge [MP PET 2003]  
 (a)  $\text{CH}_3\text{COOK}$  (b)  $\text{KCl}$   
 (c)  $\text{NH}_4\text{NO}_3$  (d)  $\text{KNO}_3$
47. The reference electrode is made by using [MP PMT 2002]  
 (a)  $\text{ZnCl}_2$  (b)  $\text{CuSO}_4$   
 (c)  $\text{HgCl}_2$  (d)  $\text{Hg}_2\text{Cl}_2$
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.  $\lambda_{\text{ClCH}_2\text{COONa}} = 224\text{ ohm}^{-1}\text{ cm}^2\text{ gmeq}^{-1}$ ,

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

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

What is the value of  $\lambda_{ClCH_2COOH}$

(a)  $288.5 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$

(b)  $289.5 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$

(c)  $388.5 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$

(d)  $59.5 \text{ ohm}^{-1} \text{ cm}^2 \text{ gmeq}^{-1}$

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 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  $25^\circ C$  specific conductivity of a normal solution of  $KCl$  is  $0.002765 \text{ mho}$ . The resistance of cell is  $400 \text{ ohms}$ . The cell constant is

(a) 0.815

(b) 1.016

(c) 1.106

(d) 2.016

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

(a) Arsenic

(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)  $Pt | H_2(g), 1 \text{ bar} | 1M KCl(aq) | AgCl(s) | Ag(s)$

(b)  $Pt(s) | H_2(g), 1 \text{ bar} | 1M HCl(aq) || 1M Ag^+(aq) | Ag(s)$

[JEE Orissa 2004]

(c)  $Pt(s) | H_2(g), 1 \text{ bar} | 1M HCl(aq) | AgCl(s) | Ag(s)$

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

58. If the  $Zn^{2+} / Zn$  electrode is diluted to 100 times then the change in e.m.f. [DPMT 2005]

(a) Increase of  $59 \text{ mV}$

(b) Decrease of  $59 \text{ mV}$

[AIIMS 2004]

(c) Increase of  $29.5 \text{ mV}$

(d) Decrease of  $29.5 \text{ mV}$

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)  $0.177 \text{ V}$

(b)  $0.3 \text{ V}$

(c)  $0.052 \text{ V}$

(d)  $0.104 \text{ V}$

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, $E_{\text{cell}}$ , 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  $0.059 \text{ V}$ ) [KCET 1993, 2005]

(a)  $0.177 \text{ V}$

(b)  $-0.177 \text{ V}$

(c)  $0.087 \text{ V}$

(d)  $0.059 \text{ V}$

2. The standard electrode potentials of  $Zn^{2+} / Zn$  and  $Ag^+ / Ag$  are  $-0.76 \text{ V}$  and  $+0.799 \text{ V}$  respectively. [Pb.BMT 2004]

The standard potential of the cell is [KCET 1993]

(a)  $1.56 \text{ V}$

(b)  $0.036 \text{ V}$

(c)  $-1.562 \text{ V}$

(d)  $0.799 \text{ V}$

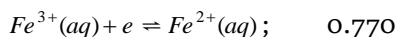
3. The standard reduction potentials at  $298 \text{ K}$  for the following half reactions are given against each

$Zn^{2+}(aq.) + 2e \rightleftharpoons Zn(s); \quad -0.762$

$Cr^{3+}(aq) + 3e \rightleftharpoons Cr(s); \quad -0.740$

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

[AIIMS 2005]



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)  $Ag$

6. 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 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

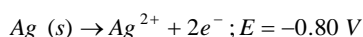
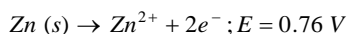
$$A = -0.250 \text{ V} \quad B = -0.136 \text{ V}$$

$$C = -0.126 \text{ V} \quad D = -0.402 \text{ V}$$

The element that displaces A from its compounds is

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

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



Which of the following reactions actually take place

[NCERT 1983, 84; KCET 2003]

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

9. Beryllium is placed above magnesium in the second group. Beryllium dust, therefore when added to  $MgCl_2$  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 ( $E^\circ$ ) 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

$$(a) E_{M^{n+}/M} = E^\circ_{M^{n+}/M} + \frac{0.0591}{n} \log(M^{n+})$$

$$(b) E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{0.0591}{n} \log(M^{n+})$$

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

- (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  $AgNO_3$ , 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  $-0.34 \text{ V}$ ,  $-3.05 \text{ V}$  and  $-1.66 \text{ V}$  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.  $E^\circ = \frac{RT}{nF} \ln K_{eq}$ . 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 -

## 520 Electrochemistry

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

[MP PMT/PET 1988 ; CPMT 1983; MNR 1993; UPSEAT 2002]

- (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]

- (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 1M) 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 an oxidising agent  
(d)  $Zn^{2+}$  is a reducing agent

21.  $2H^+(aq) + 2e^- \rightarrow H_2(g)$ . The standard electrode potential for the above reaction is (in volts) [CPMT 1988]

- (a) 0 (b) +1  
(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  $E^\circ$  for  $Ni^{2+} | Ni$  is  $-0.25 V$ ;  $E^\circ$  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 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  $E^\circ_{cell}$  will be respectively

- (a)  $-ve, > 1, +ve$  (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  $A | A^+(a=1) || B^+(a=1) | B$  will be

- (a) 1.25 V (b) -1.25 V  
(c) -0.25 V (d) 0.25 V

30. The standard reduction potential for  $Li^+/Li$ ;  $Zn^{2+}/Zn$ ;  $H^+/H_2$  and  $Ag^+/Ag$  is  $-3.05, -0.762, 0.00$  and  $+0.80 V$ . Which of the following has highest reducing capacity

- (a) Ag (b)  $H_2$   
(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.  $E^\circ$  values of  $Mg^{2+}/Mg$  is  $-2.37 V$ , of  $Zn^{2+}/Zn$  is  $-0.76 V$  and  $Fe^{2+}/Fe$  is  $-0.44 V$ . 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  $Fe^{2+} + Sn \rightarrow Fe + Sn^{2+}$ , the standard EMF is [IIT Screening 1990] (a)  $P > Q > R > S$  (b)  $Q > P > R > S$  (c)  $+0.30\text{ V}$  (d)  $+0.58\text{ V}$  (e)  $-0.58\text{ V}$  (f)  $-0.30\text{ V}$
34. Electrode potential of  $Zn^{2+}/Zn$  is  $-0.76\text{ V}$  and that of  $Cu^{2+}/Cu$  is  $+0.34\text{ V}$ . The EMF of the cell constructed between these two electrodes is [EAMCET 1992; BHU 2001; CBSE PMT 2001; KCET 1990; MHCET 1999, 2003; Pb. CET 2002; AFMC 2001; Pb. PMT 2004] (a)  $1.10\text{ V}$  (b)  $0.42\text{ V}$  (c)  $-1.1\text{ V}$  (d)  $-0.42\text{ V}$
35. EMF of a cell whose half cells are given below is  $Mg^{2+} + 2e^- \rightarrow Mg(s); E = -2.37\text{ V}$   $Cu^{2+} + 2e^- \rightarrow Cu(s); E = +0.33\text{ V}$  [EAMCET 1987; MP PET 1994; Pb. PMT 2000] (a)  $-2.03\text{ V}$  (b)  $1.36\text{ V}$  (c)  $2.7\text{ V}$  (d)  $2.03\text{ V}$
36. A cell constructed by coupling a standard copper electrode and a standard magnesium electrode has emf of  $2.7\text{ volts}$ . If the standard reduction potential of copper electrode is  $+0.34\text{ volt}$  that of magnesium electrode is [KCET 1989] (a)  $+3.04\text{ volts}$  (b)  $-3.04\text{ volts}$  (c)  $+2.36\text{ volts}$  (d)  $-2.36\text{ volts}$
37. When  $E_{Ag^+/Ag}^o = 0.8\text{ volt}$  and  $E_{Zn^{2+}/Zn}^o = -0.76\text{ volt}$ , which of the following is correct [MP PMT 1994] (a)  $Ag^+$  can be reduced by  $H_2$  (b)  $Ag$  can oxidise  $H_2$  into  $H^+$  (c)  $Zn^{2+}$  can be reduced by  $H_2$  (d)  $Ag$  can reduce  $Zn^{2+}$  ion
38. Adding powdered lead and iron to a solution that is  $1.0\text{ M}$  in both  $Pb^{2+}$  and  $Fe^{2+}$  ions, would result a reaction, in which [CPMT 1987] (a) More iron and  $Pb^{2+}$  ions are formed (b) More lead and  $Fe^{2+}$  ions are formed (c) Concentration of both  $Pb^{2+}$  and  $Fe^{2+}$  ions increases (d) There is no net change
39. Given standard electrode potentials  $Fe^{++} + 2e^- \rightarrow Fe; E^o = -0.440\text{ V}$   $Fe^{+++} + 3e^- \rightarrow Fe; E^o = -0.036\text{ V}$  The standard electrode potential ( $E^o$ ) for  $Fe^{+++} + e^- \rightarrow Fe^{++}$  is [AIIMS 1982] (a)  $-0.476\text{ V}$  (b)  $-0.404\text{ V}$  (c)  $+0.404\text{ V}$  (d)  $+0.771\text{ V}$
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)  $P > Q > R > S$  (b)  $Q > P > R > S$  (c)  $P > S > Q > R$  (d)  $Q > S > R > P$
41. Which of the following metal can deposit copper from copper sulphate solution [CPMT 1983; MP PMT 1989] (a) Mercury (b) Iron (c) Gold (d) Platinum
42. Standard electrode potential of  $Ag^+/Ag$  and  $Cu^+/Cu$  is  $+0.80\text{ V}$  and  $+0.34\text{ V}$  respectively. these electrodes are joint together by salt bridge if [AMU 2002] (a) Copper electrode is work like cathode, then  $E_{cell}^o$  is  $+0.45\text{ V}$  (b) Silver electrode is work like anode then  $E_{cell}^o$  is  $-0.34\text{ V}$  (c) Copper electrode is work like anode then  $E_{cell}^o$  is  $+0.46\text{ V}$  (d) Silver electrode is work like cathode then  $E_{cell}^o$  is  $-0.34\text{ V}$  (e) Silver electrode is work like anode then  $E_{cell}^o$  will be  $+1.14\text{ V}$
43. The reaction is spontaneous if the cell potential is [MP PET 1999] (a) Positive (b) Negative (c) Zero (d) Infinite
44. Which substance eliminates bromine from  $KBr$  solution [IIT 1981] (a)  $I_2$  (b)  $Cl_2$  (c)  $HI$  (d)  $SO_2$
45. A standard hydrogen electrode has zero electrode potential because [IIT 1997] (a) Hydrogen is easiest to oxidise (b) The electrode potential is assumed to be zero (c) Hydrogen atom has only one electron (d) Hydrogen is the lightest element
46. In the electrochemical cell  $H_2(g) | 1\text{ atm} | H^+(1\text{ M}) || Cu^{2+}(1\text{ M}) | Cu(s)$  Which one of the following statements is true [EAMCET 1997] (a)  $H_2$  is cathode;  $Cu$  is anode (b) Oxidation occurs at  $Cu$  electrode (c) Reduction occurs at  $H_2$  electrode (d)  $H_2$  is anode;  $Cu$  is cathode
47. Expression representing the cell potential ( $E_{cell}$ ) (a)  $E_{cathode} + E_{anode}$  (b)  $E_{anode} - E_{cathode}$  (c)  $E_{cathode} - E_{anode}$  (d)  $E_{left} - E_{right}$
48. Iron displaces copper from its salt solution, because [MP PMT 1996] (a) Atomic number of iron is less than that of copper



- (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 1M 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)  $E = E^{\circ} - \frac{RT}{nF} \ln \frac{[\text{product}]}{[\text{reactant}]}$   
 (b)  $E = E^{\circ} + \frac{RT}{F} \ln \frac{[\text{product}]}{[\text{reactant}]}$   
 (c)  $E = E^{\circ} - \frac{RT}{nF} \ln \frac{[\text{reactant}]}{[\text{product}]}$   
 (d)  $E = -\frac{RT}{F} \ln \frac{[\text{product}]}{[\text{reactant}]}$
53. Calculate standard free energy change for the reaction  $\frac{1}{2} Cu(s) + \frac{1}{2} Cl_2(g) \rightleftharpoons \frac{1}{2} Cu^{2+} + Cl^{-}$  taking place at  $25^{\circ}C$  in a cell whose standard e.m.f. is 1.02 volts [MP PMT 1997]  
 (a)  $-98430 J$  (b)  $98430 J$  (c)  $96500 J$  (d)  $-49215 J$
54. 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
55. Nernst equation is related with  
 (a) The electrode potential and concentration of ions in the solution  
 (b) Equilibrium constant and concentration of ions  
 (c) Free energy change and E.M.F. of the cell  
 (d) None of these
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) III (d) IV
57. Electrode potential data are given below :  
 $Fe^{3+}(aq) + e^{-} \rightarrow Fe^{2+}(aq); E^{\circ} = +0.77 V$   
 $Al^{3+}(aq) + 3e^{-} \rightarrow Al(s); E^{\circ} = -1.66 V$   
 $Br_2(aq) + 2e^{-} \rightarrow 2Br^{-}(aq); E^{\circ} = +1.08 V$   
 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 ( $E^{\circ}$ ) for  $Cl^{-}/\frac{1}{2}Cl_2$  respectively are 0.94 V and  $-1.36 V$ . The  $E^{\circ}$  value for  $OCl^{-}/\frac{1}{2}Cl_2$  will be [KCET 1998]  
 (a)  $-0.42 V$  (b)  $-2.20 V$   
 (c)  $0.52 V$  (d)  $1.04 V$
59. If the 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.  $E^{\circ}$  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  $Zn + Cu^{2+}(aq) \rightleftharpoons Cu + Zn^{2+}(aq)$  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  $25^\circ\text{C}$  of  $\text{Li}^+ | \text{Li}$ ,  $\text{Ba}^{2+} | \text{Ba}$ ,  $\text{Na}^+ | \text{Na}$  and  $\text{Mg}^{2+} | \text{Mg}$  are  $-3.05$ ,  $-2.90$ ,  $-2.71$  and  $-2.37$  volt respectively. Which one of the following is the strongest oxidising agent  
[CBSE PMT 1994; JIPMER 2002]
- (a)  $\text{Na}^+$  (b)  $\text{Li}^+$   
(c)  $\text{Ba}^{2+}$  (d)  $\text{Mg}^{2+}$
63. Which of the following displaces  $\text{Br}_2$  from an aqueous solution containing bromide ions  
[CBSE PMT 1994; JIPMER (Med.) 2002]
- (a)  $\text{Cl}_2$  (b)  $\text{Cl}^-$   
(c)  $\text{I}_2$  (d)  $\text{I}_3^-$
64. For the cell reaction  
 $\text{Cu}^{2+}(\text{C}_1\text{aq}) + \text{Zn}(\text{s}) = \text{Zn}^{2+}(\text{C}_2\text{aq}) + \text{Cu}(\text{s})$   
of an electrochemical cell, the change in free energy at a given temperature is a function of  
(a)  $\ln(\text{C}_1)$  (b)  $\ln(\text{C}_2)$   
(c)  $\ln(\text{C}_1 + \text{C}_2)$  (d)  $\ln(\text{C}_2 / \text{C}_1)$
65. The e.m.f. of the cell in which the following reaction  $\text{Zn}(\text{s}) + \text{Ni}^{2+}(a=1.0) \rightleftharpoons \text{Zn}^{2+}(a=10) + \text{Ni}(\text{s})$  occurs, is found to be  $0.5105\text{ V}$  at  $298\text{ K}$ . The standard e.m.f. of the cell is  
[Roorkee Qualifying 1998]
- (a)  $0.5400$  (b)  $0.4810\text{ V}$   
(c)  $0.5696\text{ V}$  (d)  $-0.5105\text{ V}$
66. For the redox reaction  
 $\text{Zn}(\text{s}) + \text{Cu}^{2+}(0.1\text{ M}) \rightarrow \text{Zn}^{2+}(1\text{ M}) + \text{Cu}(\text{s})$  taking place in a cell,  $E^\circ_{\text{cell}}$  is  $1.10\text{ volt}$ .  $E_{\text{cell}}$  for the cell will be  
 $\left(2.303 \frac{RT}{F} = 0.0591\right)$  [AIEEE 2003]
- (a)  $2.14\text{ volt}$  (b)  $1.80\text{ volt}$   
(c)  $1.07\text{ volt}$  (d)  $0.82\text{ volt}$
67. The emf of a Daniel cell at  $298\text{ K}$  is  $E_1$   
 $\text{Zn} | \text{ZnSO}_4(0.01\text{ M}) || \text{CuSO}_4(1.0\text{ M}) | \text{Cu}$  when the concentration of  $\text{ZnSO}_4$  is  $1.0\text{ M}$  and that of  $\text{CuSO}_4$  is  $0.01\text{ M}$ , the emf changed to  $E_2$ . What is the relationship between  $E_1$  and  $E_2$   
[CBSE PMT 2003]
- (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  
 $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-; E^\circ = 0.76\text{ V}$ ,  
 $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-; E^\circ = 0.44\text{ V}$  what will be the emf of cell, whose cell-reaction is  
 $\text{Fe}^{2+}(\text{aq}) + \text{Zn} \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Fe}$   
[MP PMT 2003]
- (a)  $-1.20\text{ V}$  (b)  $+0.32\text{ V}$   
(c)  $-0.32\text{ V}$  (d)  $+1.20\text{ V}$
69. The  $E^\circ$  for half cells  $\text{Fe} / \text{Fe}^{2+}$  and  $\text{Cu} / \text{Cu}^{2+}$  are  $-0.44\text{ V}$  and  $+0.32\text{ V}$  respectively. Then [MP PMT 2003]
- (a)  $\text{Cu}^{2+}$  oxidises  $\text{Fe}$  (b)  $\text{Cu}^{2+}$  oxidises  $\text{Fe}^{2+}$   
(c)  $\text{Cu}$  oxidises  $\text{Fe}^{2+}$  (d)  $\text{Cu}$  reduces  $\text{Fe}^{2+}$
70. What is  $E^\circ$  for electrode represented by  $\text{Pt}, \text{O}_2(1\text{ atm}) / 2\text{H}^+(\text{Im})$  [JIPMER 1997]
- (a) Unpredictable (b) Zero  
(c)  $0.018\text{ V}$  (d)  $0.118\text{ V}$
71. The cell potential of a cell in operation is  
(a) Zero (b) Positive  
(c) Negative (d) None of the above
72. Which of the following is displaced by  $\text{Fe}$  [Roorkee 1995]
- (a)  $\text{Ag}$  (b)  $\text{Hg}$   
(c)  $\text{Zn}$  (d)  $\text{Na}$
73. [CBSE PMT 1998] Standard electrode potential of the half cells are given below  
 $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}; E^\circ = -7.62\text{ V}$ ,  
 $\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}; E^\circ = -7.81\text{ V}$   
The emf of the cell  $\text{Fe}^{2+} + \text{Zn} \rightarrow \text{Zn}^{2+} + \text{Fe}$  is  
[CPMT 2003]
- (a)  $1.54\text{ V}$  (b)  $-1.54\text{ V}$   
(c)  $-0.19\text{ V}$  (d)  $+0.19\text{ V}$
74.  $\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}(\text{s}); E^\circ = -0.76$ ,  
 $\text{Fe}^{3+} + e^- \rightarrow \text{Fe}^{2+}; E^\circ = -0.77$ ,  
 $\text{Cr}^{3+} + 3e^- \rightarrow \text{Cr}; E^\circ = -0.79$ ,  
 $\text{H}^+ + 2e^- \rightarrow 1/2\text{H}_2; E^\circ = 0.00$   
Strongest reducing agent is  
(a)  $\text{Fe}^{2+}$  (b)  $\text{Zn}$   
(c)  $\text{Cr}$  (d)  $\text{H}_2$
75. Standard reduction electrode potentials of three metals A, B and C are respectively  $+0.5\text{ V}$ ,  $-3.0\text{ V}$  and  $-1.2\text{ V}$ . The reducing powers of these metals are [IIT 1998; AIEEE 2003]
- (a)  $B > C > A$  (b)  $A > B > C$   
(c)  $C > B > A$  (d)  $A > C > B$
76. For a cell reaction involving a two-electron change, the standard emf of the cell is found to be  $0.295\text{ V}$  at  $25^\circ\text{C}$ . The equilibrium constant of the reaction at  $25^\circ\text{C}$  will be  
[Roorkee 1999; AIEEE 2003; CBSE PMT 2004]
- (a)  $1 \times 10^{-10}$  (b)  $29.5 \times 10^{-2}$   
(c)  $10$  (d)  $1 \times 10^{10}$
77. For the electrochemical cell,  
 $\text{M} | \text{M}^+ || \text{X}^- | \text{X}, E^\circ(\text{M}^+ / \text{M}) = 0.44\text{ V}$  and  
 $E^\circ(\text{X} / \text{X}^-) = 0.33\text{ V}$ . 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)  $E_{\text{cell}} = 0.77 \text{ V}$   
 (d)  $E_{\text{cell}} = -0.77 \text{ V}$
- 78.** The standard potential at  $25^\circ \text{C}$  for the following half reactions are given against them  
 $\text{Zn}^{2+} + 2e \rightarrow \text{Zn}, E^\circ = -0.762 \text{ V}$   
 $\text{Mg}^{2+} + 2e \rightarrow \text{Mg}, E^\circ = -2.37 \text{ V}$   
 When zinc dust is added to the solution of  $\text{MgCl}_2$  [UPSEAT 2001]  
 (a)  $\text{ZnCl}_2$  is formed  
 (b) Zinc dissolves in the solution  
 (c) No reaction takes place  
 (d)  $\text{Mg}$  is precipitated
- 79.**  $\text{KMnO}_4$  acts as an oxidising agent in the neutral medium and gets reduced to  $\text{MnO}_2$ . The equivalent weight of  $\text{KMnO}_4$  in neutral medium [AMU 2001]  
 (a)  $\text{mol. wt}/2$  (b)  $\text{mol. wt}/3$   
 (c)  $\text{mol. wt}/4$  (d)  $\text{mol. wt}/7$
- 80.** Which of the following condition will increase the voltage of the cell, represented by the equation  
 $\text{Cu}_{(s)} + 2\text{Ag}^+_{(aq)} \rightarrow \text{Cu}^{2+}_{(aq)} + 2\text{Ag}_{(s)}$  [CBSE PMT 2001]  
 (a) Increase in the concentration of  $\text{Ag}^+$  ion  
 (b) Increase in the concentration of  $\text{Cu}^+$  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  
 $\text{Sn}_{(s)} + 2\text{Ag}^+_{(aq)} \rightarrow \text{Sn}^{2+}_{(aq)} + 2\text{Ag}_{(s)}$  [DPMT 2001]  
 (a) Increase in the concentration of  $\text{Ag}^+$  ions  
 (b) Increase in the concentration of  $\text{Sn}^{2+}$  ions  
 (c) Increase in size of the silver rod  
 (d) None of these
- 82.** The mass of the proton is 1840 times that of electron, its potential difference is  $V$ . The kinetic energy of proton is [DCE 2001]  
 (a)  $1840 \text{ KeV}$  (b)  $1 \text{ KeV}$   
 (c)  $\frac{1}{1840} \text{ KeV}$  (d)  $920 \text{ KeV}$
- 83.** What will be the  $\text{emf}$  for the given cell  
 $\text{Pt} | \text{H}_2(P_1) | \text{H}^+_{(aq)} || \text{H}_2(P_2) | \text{Pt}$  [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  $10^{-8} \text{ M H}^+$  and positive one in contact with  $0.025 \text{ M H}^+$  [MP PMT 2000]  
 (a)  $0.18 \text{ V}$  (b)  $0.28 \text{ V}$   
 (c)  $0.38 \text{ V}$  (d)  $0.48 \text{ V}$
- 85.** Will  $\text{Fe}_{(s)}$  be oxidised to  $\text{Fe}^{2+}$  by the reaction with  $1 \text{ M HCl}$  ( $E^\circ$  for  $\text{Fe} / \text{Fe}^{2+} = +0.44 \text{ V}$ ) [Pb. PMT 2000]  
 (a) Yes (b) No  
 (c) May be (d) Can't say
- 86.**  $\text{EMF}$  of a cell in terms of reduction potential of its left and right electrodes is  
 (a)  $E = E_{\text{left}} - E_{\text{right}}$  (b)  $E = E_{\text{left}} + E_{\text{right}}$   
 (c)  $E = E_{\text{right}} - E_{\text{left}}$  (d)  $E = -(E_{\text{right}} + E_{\text{left}})$
- 87.** Arrange the following in the order of their decreasing electrode potential  $\text{Mg}, \text{K}, \text{Ba}, \text{Ca}$  [JIPMER 2002]  
 (a)  $\text{K}, \text{Ba}, \text{Ca}, \text{Mg}$  (b)  $\text{Ca}, \text{Mg}, \text{K}, \text{Ba}$   
 (c)  $\text{Ba}, \text{Ca}, \text{K}, \text{Mg}$  (d)  $\text{Mg}, \text{Ca}, \text{Ba}, \text{K}$
- 88.** Which of the following has highest electrode potential [Pb. PMT 2000]  
 (a)  $\text{Li}$  (b)  $\text{Cu}$   
 (c)  $\text{Au}$  (d)  $\text{Al}$
- 89.** The cell reaction of a cell is  
 $\text{Mg}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Cu}_{(s)} + \text{Mg}^{2+}_{(aq)}$   
 If the standard reduction potentials of  $\text{Mg}$  and  $\text{Cu}$  are  $-2.37$  and  $+0.34 \text{ V}$  respectively. The  $\text{emf}$  of the cell is [EAMCET 1995; JIPMER (Med.) 2001; AFMC 2002; CBSE PMT 2002]  
 (a)  $2.03 \text{ V}$  (b)  $-2.03 \text{ V}$   
 (c)  $+2.71 \text{ V}$  (d)  $-2.71 \text{ V}$
- 90.** The element which can displace three other halogens from their compound is [EAMCET 1998]  
 (a)  $\text{Cl}$  (b)  $\text{F}$   
 (c)  $\text{Br}$  (d)  $\text{I}$
- 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  $\text{Fe}$ . But  $\text{Al}$  is less easily corroded than iron because [DCE 1999]  
 (a)  $\text{Al}$  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  
 $Pt(H_2, 1 \text{ atm}) / 0.1 \text{ M HCl}$   
 $|| 0.1 \text{ M acetic acid} / (H_2, 1 \text{ atm}) Pt$   
E.M.F. of this cell will not be zero because  
[CBSE PMT 1995]  
(a) The  $pH$  of  $0.1 \text{ M HCl}$  and  $0.1 \text{ M}$  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.  $E^\circ$  value for disproportionation of  $Cu^+$  is  
(Given  $E^\circ_{Cu^{2+}/Cu^+} = 0.15$ ,  $E^\circ_{Cu^{2+}/Cu} = 0.34 \text{ V}$ ) [IIT 1995]  
(a)  $-0.49 \text{ V}$  (b)  $0.49 \text{ V}$   
(c)  $-0.38 \text{ V}$  (d)  $0.38 \text{ V}$
96.  $E^\circ$  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
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  $0.01 \text{ M ZnSO}_4$  solution at  $258^\circ \text{C}$  ( $E^\circ = 0.763 \text{ V}$ )  
[AIIMS 2000; BHU 2000]  
(a)  $0.8221 \text{ V}$  (b)  $8.221 \text{ V}$   
(c)  $0.5282 \text{ V}$  (d)  $9.232 \text{ V}$
102. The emf of a galvanic cell, with electrode potentials of silver =  $+0.80 \text{ V}$  and that of copper =  $+0.34 \text{ V}$ , is [AIIMS 1999]  
(a)  $-1.1 \text{ V}$  (b)  $+1.1 \text{ V}$   
(c)  $+0.46 \text{ V}$  (d)  $+0.76 \text{ V}$
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

## 526 Electrochemistry

- (c) Lower the  $E$  and shift equilibrium to the left  
(d) Increase the  $E$  and shift equilibrium to the left
- 108.** For the electrochemical cell,  $M|M^+||X^-|X$ ,  $E^\circ(M^+|M) = 0.44\text{ V}$ ,  $E^\circ(X|X^-) = 0.33\text{ V}$ . From this data, one can deduce that  
(a)  $E^\circ_{\text{cell}} = -0.77\text{ V}$   
(b)  $M^+ + X^- \rightarrow M + X$  is the spontaneous reaction  
(c)  $M + X \rightarrow M^+ + X^-$  is the spontaneous reaction  
(d)  $E^\circ_{\text{cell}} = .77\text{ V}$
- 109.** The standard  $e.m.f.$  of a cell, involving one electron change is found to be  $0.591\text{ V}$  at  $25^\circ\text{C}$ . The equilibrium constant of the reaction is ( $F = 96,500\text{ C mol}^{-1}$ ;  $R = 8.314\text{ JK}^{-1}\text{mol}^{-1}$ ) [AIEEE 2004]  
(a)  $1.0 \times 10^{10}$  (b)  $1.0 \times 10^5$   
(c)  $1.0 \times 10^1$  (d)  $1.0 \times 10^{30}$
- 110.** Standard electrode potential of cell  $H_2|H^+||Ag^+|Ag$  is [AIEEE 2004]  
(a)  $0.8\text{ V}$  (b)  $-0.8\text{ V}$   
(c)  $-1.2\text{ V}$  (d)  $1.2\text{ V}$
- 111.** A galvanic cell with electrode potential of 'A' =  $+2.23\text{ V}$  and 'B' =  $-1.43\text{ V}$ . The value of  $E^\circ_{\text{cell}}$  is [Pb.CET 2003]  
(a)  $3.66\text{ V}$  (b)  $0.80\text{ V}$   
(c)  $-0.80\text{ V}$  (d)  $-3.66\text{ V}$
- 112.** The  $e.m.f.$  of a cell whose half cells are given below is  
 $Mg^{2+} + 2e^- \rightarrow Mg(s)$   $E^\circ = -2.37\text{ V}$   
 $Cu^{2+} + 2e^- \rightarrow Cu(s)$   $E^\circ = +0.34\text{ V}$  [Pb.CET 2001]  
(a)  $+1.36\text{ V}$  (b)  $+2.71\text{ V}$   
(c)  $+2.17\text{ V}$  (d)  $-3.01\text{ V}$
- 113.** For the cell reaction,  $2Ce^{4+} + Co \rightarrow 2Ce^{3+} + Co^{2+}$   $E^\circ_{\text{cell}}$  is  $1.89\text{ V}$ . If  $E^\circ_{Ce^{4+}/Ce^{3+}}$  [Pb.CET 2000]  
(a)  $-1.64\text{ V}$  (b)  $+1.64\text{ V}$   
(c)  $-2.08\text{ V}$  (d)  $+2.17\text{ V}$
- 114.** If the  $\Delta G$  of a cell reaction  $AgCl + e^- \rightarrow Ag + Cl^-$  is  $-21.20\text{ KJ}$ ; the standard  $e.m.f.$ , of cell is [MP PMT 2004]  
(a)  $0.229\text{ V}$  (b)  $0.220\text{ V}$   
(c)  $-0.220\text{ V}$  (d)  $-0.110\text{ V}$
- 115.** The  $e.m.f.$  of the cell  $Ag|Ag^+(0.1M)||Ag^+(1M)|Ag$  at  $298\text{ K}$  is [DCE 2003]  
(a)  $0.0059\text{ V}$  (b)  $0.059\text{ V}$   
(c)  $5.9\text{ V}$  (d)  $0.59\text{ V}$
- 116.** The  $e.m.f.$  of the cell  $Zn|Zn^{2+}(0.01M)||Fe^{2+}(0.001M)|Fe$  at  $298\text{ K}$  is  $0.2905$  then the value of equilibrium for the cell reaction is [IIT-JEE Screening 2004]  
(a)  $\frac{0.32}{e^{0.0295}}$  (b)  $\frac{0.32}{10^{0.0295}}$   
(c)  $\frac{0.26}{10^{0.0295}}$  (d)  $\frac{0.32}{10^{0.0591}}$
- 117.** Aluminium displaces hydrogen from dilute  $HCl$  whereas silver does not. The  $e.m.f.$  of a cell prepared by combining  $Al/Al^{3+}$  and  $Ag/Ag^+$  is  $2.46\text{ V}$ . The reduction potential of silver electrode is  $+0.80\text{ V}$ . The reduction potential of aluminium electrode is [KCET 2004]  
(a)  $+1.66\text{ V}$  (b)  $-3.26\text{ V}$   
(c)  $3.26\text{ V}$  (d)  $-1.66\text{ V}$
- 118.** Consider the following  $E^\circ$  values :  
 $E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77\text{ V}$   
 $E^\circ_{Sn^{2+}/Sn} = -0.14\text{ V}$   
Under standard conditions the potential for the reaction  $Sn_{(s)} + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{2+}(aq)$  is [AIEEE 2000]  
(a)  $0.91\text{ V}$  (b)  $1.40\text{ V}$   
(c)  $1.68\text{ V}$  (d)  $0.63\text{ V}$
- 119.**  $Cr_2O_7^{2-} + I^- \rightarrow I_2 + Cr^{3+}$   
 $E^\circ_{\text{cell}} = 0.79\text{ V}$   
 $E^\circ_{Cr_2O_7^{2-}} = 1.33\text{ V}$ ,  $E^\circ_{I_2}$  is [BVP 2004]  
(a)  $-0.10\text{ V}$  (b)  $+0.18\text{ V}$   
(c)  $-0.54\text{ V}$  (d)  $0.54\text{ V}$
- 120.**  $Zn(s) + Cl_2(1\text{ atm}) \rightarrow Zn^{2+} + 2Cl^-$ .  $E^\circ_{\text{cell}}$  of the cell is  $2.12\text{ V}$ . To increase  $E$  [BVP 2004]  
(a)  $[Zn^{2+}]$  should be increased  
(b)  $[Zn^{2+}]$  should be decreased  
(c)  $[Cl^-]$  should be decreased  
(d)  $P_{Cl_2}$  should be decreased
- 121.** The  $E^\circ_{M^{3+}/M^{2+}}$  values for  $Cr, Mn, Fe$  and  $Co$  are  $-0.41, +1.57, +0.77$  and  $+1.97\text{ V}$  respectively. For which one of these metals the change in oxidation state from  $+2$  to  $+3$  is easiest  
(a)  $Fe$  (b)  $Mn$   
(c)  $Cr$  (d)  $Co$
- 122.** The rusting of iron takes place as follows  
 $2H^+ + 2e^- + \frac{1}{2}O_2 \longrightarrow H_2O(l)$ ;  
 $E^\circ = +1.23\text{ V}$   
 $Fe^{2+} + 2e^- \longrightarrow Fe(s)$ ;  $E^\circ = -0.44\text{ V}$

- Calculate  $\Delta G^\circ$  for the net process [IIT 2005]  
 (a)  $-322 \text{ kJ mol}^{-1}$  (b)  $-161 \text{ kJ mol}^{-1}$   
 (c)  $-152 \text{ kJ mol}^{-1}$  (d)  $-76 \text{ kJ mol}^{-1}$
123. When an acid cell is charged, then [AFMC 2005]  
 (a) Voltage of cell increases  
 (b) Electrolyte of cell dilutes  
 (c) Resistance of cell increases  
 (d) None of these
124. The standard electrode potential is measured by [KCET 2005]  
 (a) Electrometer (b) Voltmeter  
 (c) Pyrometer (d) Galvanometer
125. Aluminium displaces hydrogen from acids but copper does not. A galvanic cell prepared by combining  $\text{Cu}/\text{Cu}^{2+}$  and  $\text{Al}/\text{Al}^{3+}$  has an e.m.f. of 2.0 V at 298 K. If the potential of copper electrode is + 0.34 V, that of aluminium is [CPMT 2001; KCET 2001]  
 (a) + 1.66 V (b) - 1.66 V  
 (c) + 2.34 V (d) - 2.3 V
126. If the standard electrode potential of  $\text{Cu}^{2+}/\text{Cu}$  electrode is 0.34 V, what is the electrode potential of 0.01 M concentration of  $\text{Cu}^{2+}$  ( $T = 298 \text{ K}$ ) [EAMCET 2003]  
 (a) 0.399 V (b) 0.281 V  
 (c) 0.222 V (d) 0.176 V
127. Calculate the electrode potential at 298°K for  $\text{Zn}|\text{Zn}^{++}$  electrode in which the activity of zinc ions is 0.001 M and  $E^\circ_{\text{Zn}/\text{Zn}^{++}}$  is -0.74 volts [AMU 2002]  
 (a) 0.38 volts (b) 0.83 volts  
 (c) 0.40 volts (d) 0.45 volts
128. Which of the following expression is correct [Orissa JEE 2005]  
 (a)  $\Delta G^\circ = -nFE^\circ_{\text{cell}}$  (b)  $\Delta G^\circ = +nFE^\circ_{\text{cell}}$   
 (c)  $\Delta G^\circ = -2.303 RT nFE^\circ_{\text{cell}}$  (d)  $\Delta G^\circ = -nF \log K_C$
129. For the feasibility of a redox reaction in a cell, the e.m.f. should be [J & K 2002]  
 (a) Positive (b) Fixed  
 (c) Zero (d) Negative
- (a)  $\text{FeCl}_2$  (b)  $\text{PbCl}_2$   
 (c)  $\text{Hg}_2\text{Cl}_2$  (d)  $\text{HgCl}_2$
4. Corrosion of iron is essentially an electrochemical phenomenon where the cell reactions are [KCET 2005]  
 (a) Fe is oxidised to  $\text{Fe}^{2+}$  and dissolved oxygen in water is reduced to  $\text{OH}^-$   
 (b) Fe is oxidised to  $\text{Fe}^{3+}$  and  $\text{H}_2\text{O}$  is reduced to  $\text{O}_2^{2-}$   
 (c) Fe is oxidised to  $\text{Fe}^{2+}$  and  $\text{H}_2\text{O}$  is reduced to  $\text{O}_2^-$   
 (d) Fe is oxidised to  $\text{Fe}^{2+}$  and  $\text{H}_2\text{O}$  is reduced to  $\text{O}_2$

## Critical Thinking

### Objective Questions

- The limiting molar conductivities  $\Lambda^\circ$  for NaCl, KBr and KCl are 126, 152 and 150  $\text{S cm}^2\text{mol}^{-1}$  respectively. The  $\Lambda^\circ$  for NaBr is  
 (a) 278  $\text{S cm}^2\text{mol}^{-1}$  (b) 176  $\text{S cm}^2\text{mol}^{-1}$   
 (c) 128  $\text{S cm}^2\text{mol}^{-1}$  (d) 302  $\text{S cm}^2\text{mol}^{-1}$
2. On the basis of the electrochemical theory of aqueous corrosion, the reaction occurring at the cathode is [MP PET 1994; UPSEAT 2001]  
 (a)  $\text{O}_{2(g)} + 4\text{H}^+_{(aq)} + 4e^- \rightarrow 2\text{H}_2\text{O}_{(l)}$   
 (b)  $\text{Fe}_{(s)} \rightarrow \text{Fe}^{2+}_{(aq)} + 2e^-$   
 (c)  $\text{Fe}^{2+}_{(aq)} \rightarrow \text{Fe}^{3+}_{(aq)} + e^-$   
 (d)  $\text{H}_{2(g)} + 2\text{OH}^-_{(aq)} \rightarrow 2\text{H}_2\text{O}_{(l)} + 2e^-$
3. The reaction  $\frac{1}{2} \text{H}_2(g) + \text{AgCl}(s) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) + \text{Ag}(s)$  occurs in the galvanic cell [IIT 1985; AMU 2002; KCET 2003]  
 (a)  $\text{Ag} / \text{AgCl}(s) \text{KCl}(\text{soln}) || \text{AgNO}_3(\text{soln}) / \text{Ag}$   
 (b)  $\text{Pt} / \text{H}_2(g) \text{HCl}(\text{soln}) || \text{AgNO}_3(\text{soln}) / \text{Ag}$   
 (c)  $\text{Pt} / \text{H}_2(g) \text{HCl}(\text{soln}) || \text{AgCl}(s) / \text{Ag}$   
 (d)  $\text{Pt} / \text{H}_2(g) \text{KCl}(\text{soln}) || \text{AgCl}(s) / \text{Ag}$
4. The standard reduction potential  $E^\circ$  for the half reactions are as  
 $\text{Zn} = \text{Zn}^{2+} + 2e^-$ ;  $E^\circ = +0.76 \text{ V}$   
 $\text{Fe} = \text{Fe}^{2+} + 2e^-$ ;  $E^\circ = +0.41 \text{ V}$   
 The EMF for cell reaction  $\text{Fe}^{2+} + \text{Zn} \rightarrow \text{Zn}^{2+} + \text{Fe}$  is

## Corrosion

1. Corrosion is basically a [Kerala (Med.) 2002]  
 (a) Altered reaction in presence of  $\text{H}_2\text{O}$   
 (b) Electrochemical phenomenon  
 (c) Interaction  
 (d) Union between light metal and heavy metal
2. Rusting of iron is catalysed by which of the following [MNR 1990; UPSEAT 2001]  
 (a) Fe (b)  $\text{O}_2$   
 (c) Zn (d)  $\text{H}^+$
3. Which of the following is a highly corrosive salt [AFMC 2005]

[IIT 1988; CBSE PMT 1993, 96; BHU 1995, 2000; CPMT 2000; KCET 2000; AIIMS 2001; Orissa JEE 2002]

- (a)  $-0.35\text{ V}$  (b)  $+0.35\text{ V}$   
(c)  $+1.17\text{ V}$  (d)  $-1.17\text{ V}$

5. The number of electrons to balance the following equation  $\text{NO}_3^- + 4\text{H}^+ + e^- \rightarrow 2\text{H}_2\text{O} + \text{NO}$  is [IIT Screening 1991]

- (a) 5 (b) 4  
(c) 3 (d) 2

6. The standard EMF for the given cell reaction  $\text{Zn} + \text{Cu}^{2+} = \text{Cu} + \text{Zn}^{2+}$  is  $1.10\text{ V}$  at  $25^\circ\text{C}$ . The EMF for the cell reaction, when  $0.1\text{ M Cu}^{2+}$  and  $0.1\text{ M Zn}^{2+}$  solutions are used, at  $25^\circ\text{C}$  is

[MNR 1994; AMU 1999; UPSEAT 2002]

- (a)  $1.10\text{ V}$  (b)  $0.110\text{ V}$   
(c)  $-1.10\text{ V}$  (d)  $-0.110\text{ V}$

7. A gas  $X$  at  $1\text{ atm}$  is bubbled through a solution containing a mixture of  $1\text{ M Y}^-$  and  $1\text{ M Z}^-$  at  $25^\circ\text{C}$ . If the reduction potential of  $\text{Z} > \text{Y} > \text{X}$ , then [IIT 1999]

- (a)  $\text{Y}$  will oxidize  $X$  and not  $Z$   
(b)  $\text{Y}$  will oxidize  $Z$  and not  $X$   
(c)  $\text{Y}$  will oxidize both  $X$  and  $Z$   
(d)  $\text{Y}$  will reduce both  $X$  and  $Z$

8. The oxidation potential of a hydrogen electrode at  $\text{pH} = 10$  and  $\text{pH}_1 = 1$  [JIPMER 2000]

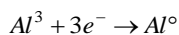
- (a)  $0.059\text{ V}$  (b)  $0.59\text{ V}$   
(c)  $0.00\text{ V}$  (d)  $0.51\text{ V}$

9. The decomposition of hydrogen peroxide is an example of

[Roorkee 2000]

- (a) Exothermic reaction (b) Endothermic reaction  
(c) Negative catalysis (d) Auto-oxidation

10. Aluminium oxide may be electrolysed at  $1000^\circ\text{C}$  to furnish aluminium metal (At. Mass =  $27\text{ amu}$ ;  $1\text{ Faraday} = 96,500\text{ Coulombs}$ ). The cathode reaction is



To prepare  $5.12\text{ kg}$  of aluminium metal by this method would require [AIEEE 2005]

- (a)  $5.49 \times 10^7\text{ C}$  of electricity  
(b)  $1.83 \times 10^7\text{ C}$  of electricity  
(c)  $5.49 \times 10^4\text{ C}$  of electricity  
(d)  $5.49 \times 10^1\text{ C}$  of electricity

Electrolyte :	KCl	KNO <sub>3</sub>	HCl	NaOAc	NaCl
$\Lambda^\infty (\text{Scm}^2\text{mol}^{-1})$	149.	145.	426.	91.0	126.

Calculate  $\Lambda_{\text{HOAc}}^\infty$  using appropriate molar conductances of the electrolytes listed above at infinite dilution in  $\text{H}_2\text{O}$  at  $25^\circ\text{C}$  [AIEEE 2005]

- (a)  $517.2$  (b)  $552.7$   
(c)  $390.7$  (d)  $217.5$

The mass of carbon anode consumed (giving only carbondioxide) in the production of  $270\text{ kg}$  of aluminium metal from bauxite by the Hall process is [CBSE PMT 2005]

- (a)  $180\text{ kg}$  (b)  $270\text{ kg}$   
(c)  $540\text{ kg}$  (d)  $90\text{ kg}$

13.  $4.5\text{ g}$  of aluminium (at mass  $27\text{ amu}$ ) is deposited at cathode from  $\text{Al}^{3+}$  solution by a certain quantity of electric charge. The volume of hydrogen produced at STP from  $\text{H}^+$  ions in solution by the same quantity of electric charge will be

[CBSE PMT 2005]

- (a)  $22.4\text{ L}$  (b)  $44.8\text{ L}$   
(c)  $5.6\text{ L}$  (d)  $11.2\text{ L}$

What amount of  $\text{Cl}_2$  gas liberated at anode, if  $1\text{ amp.}$  current is passed for  $30\text{ min.}$  from  $\text{NaCl}$  solution.

[BHU 2005]

- (a)  $0.66\text{ moles}$  (b)  $0.33\text{ moles}$   
(c)  $0.66\text{ g}$  (d)  $0.33\text{ g}$

## Assertion & Reason

For AIEEE Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

- (a) If both assertion and reason are true and the reason is the correct explanation of the assertion.  
(b) If both assertion and reason are true but reason is not the correct explanation of the assertion.  
(c) If assertion is true but reason is false.  
(d) If the assertion and reason both are false.  
(e) If assertion is false but reason is true.

1. Assertion : Sodium ions are discharged in preference to hydrogen ions at a mercury cathode.

Reason : The nature of the cathode can effect the order of discharge of ions.

2. Assertion : In electrolysis, the quantity of electricity needed for depositing  $1\text{ mole}$  of silver is different from that required for  $1\text{ mole}$  of copper.

Reason : The molecular weights of silver and copper are different

3. Assertion : Equivalent conductance of all electrolytes decreases with the increases in concentration.

- Reason : Lesser number of ions are available per gram equivalent at higher concentration.
4. Assertion : Copper reacts with hydrochloric acid and liberates hydrogen from the solution of dilute hydrochloric acid.
- Reason : Hydrogen is below copper in the electrochemical series.
5. Assertion : Copper liberates hydrogen from a solution of dil.  $HCl$ .
- Reason : Hydrogen is below copper in the reactivity series.
6. Assertion :  $Zn$  metal is formed when a  $Cu$  plate is dipped in  $ZnSO_4$  solution.
- Reason :  $Cu$  being placed above  $Zn$  in electrochemical series.
7. Assertion : Electrical conductivity of copper increases with increase in temperature.
- 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  $HCl$  and liberates hydrogen.
- Reason : Hydrogen is present above  $Cu$  in the reactivity series.
10. Assertion :  $K$  and  $Cs$  are used in photoelectric cells.
- Reason :  $K$  and  $Cs$  emit electrons on exposure to light.
11. Assertion : A large dry cell has high  $e.m.f.$
- Reason : The  $e.m.f.$  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  $SI$  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 place.
- 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 sodium chloride is given by the equation:  $\Lambda_{NaCl}^0 = \lambda_{Na^+}^0 + \lambda_{Cl^-}^0$ .
- Reason : This is according to Kohlrausch law of independent migration of ions.
15. Assertion : One coulomb of electric charge deposits weight equal to the electrochemical equivalent of the substance.
- Reason : One Faraday deposits one mole of the substance.
16. 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.
17. Assertion : For a cell reaction  $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ ; at the equilibrium, voltmeter gives zero reading.
- Reason : At the equilibrium, there is no change in the concentration of  $Cu^{2+}$  and  $Zn^{2+}$  ions.
18. 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.
19. Assertion : Weston is a standard cell.
- Reason : Its  $e.m.f.$  does not change with temperature.
20. 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.
21. Assertion : Zinc displaces copper from copper sulphate solution.
- Reason : The  $E^0$  of zinc is  $-0.76 V$  and that of copper is  $+0.34 V$ .
22. Assertion : Identification of cathode and anode is done by the use of a thermometer.
- Reason : Higher is the value of reduction potential, greater would be its reducing power.
23. 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.
24. Assertion : Galvanised iron does not rust.
- Reason : Zinc has a more negative electrode potential than iron. [AIIMS 2005]



## 530 Electrochemistry

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 :  $Ni / Ni^{2+} (1.0M) || Au^{3+} (1.0M) | Au$ , for this cell emf is 1.75 V if  $E_{Au^{3+}/Au}^{\circ} = 1.50$  and  $E_{Ni^{2+}/Ni}^{\circ} = 0.25 V$ .  
Reason : Emf of the cell =  $E_{cathode}^{\circ} - E_{anode}^{\circ}$ .
27. Assertion : Salts like  $KCl, KNO_3$  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	c	5	d
6	d	7	b	8	b	9	b	10	a
11	a	12	c	13	d	14	d	15	c
16	d	17	a	18	c	19	c	20	a
21	a	22	b	23	b	24	d	25	b
26	b	27	b	28	a	29	c	30	b
31	d	32	d	33	c	34	b	35	b
36	a	37	d	38	c	39	d	40	a
41	a								

## Faraday's law of electrolysis

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

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

## Conductor and Conductance

1	b	2	b	3	d	4	d	5	a
6	b	7	b	8	a	9	b	10	d
11	a	12	b	13	b	14	b	15	b
16	d	17	b	18	d	19	b	20	c
21	b	22	a	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	a	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	c	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	a	55	c
56	d	57	b	58	a	59	a	60	c
61	a								

## Electrode potential, $E_{Cell}$ , Nernst equation and ECS

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

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

## Corrosion

1	b	2	d	3	d	4	a		
---	---	---	---	---	---	---	---	--	--

## Critical Thinking Questions

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

## Assertion &amp; Reason

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

# AS Answers and Solutions

## Electrolytes and Electrolysis

1. (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 ( $E^\circ = -0.83V$ ). 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:  $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$
10. (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)  $NaCl \rightleftharpoons Na^+ + Cl^-$ .
28. (a) Cathode:  $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$   
 Anode:  $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$
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 a 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 produced when molten anhydrous calcium chloride is electrolysed.
37. (d) All metals conduct heat and electricity.
38. (c)  $2Al + dil. H_2SO_4 \rightarrow Al_2SO_4 + H_2 \uparrow$ .
39. (d) Generally fused potassium chloride flows 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)  $Ag^+ + e^- \rightarrow Ag$ ;  $E_{Ag} = \frac{\text{Atomic Mass}}{1} = 108$   
 Number of faraday =  $\frac{W_{Ag}}{E_{Ag}} = \frac{108}{108} = 1$ .
2. (a)  $W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{108 \times 9.65}{96500}$   
 $= 1.08 \times 10^{-2} \text{ gm} = 10.8 \text{ mg}$
3. (b)  $Fe^{2+} + 2e^- \rightarrow Fe$ ;  $E_{Fe} = \frac{56}{2} = 28$   
 $W_{Fe} = E_{Fe} \times \text{Number of faraday}$   
 $= 28 \times 3 = 84 \text{ gm}$
4. (c)  $W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{107.87 \times 965}{96500} = 1.0787 \text{ gm}$
5. (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 known charge through solution.

## 530 Electrochemistry

8. (b)  $\frac{\text{Weight of Cu}}{\text{Weight of H}_2} = \frac{\text{Eq. weight of Cu}}{\text{Eq. weight. of H}}$   
 $\frac{\text{Weight of Cu}}{0.50} = \frac{63.6/2}{1}$   
 Weight of Cu = 15.9 gm.
9. (c)  $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$   
 2 Faradays will deposit  
 = 1 g atom of Cu = 63.5 g.
12. (a) At cathode;  $\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$   
 $E_{\text{Al}} = \frac{27}{3} = 9$   
 $W_{\text{Al}} = E_{\text{Al}} \times \text{No. of faradays} = 9 \times 0.1 = 0.9 \text{ gm}.$
14. (b)  $W = Zit$ ;  $W = \frac{32.69 \times 5 \times 60 \times 40}{96500} = 4.065 \text{ gm}.$
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)  $W_{\text{metal}} = \frac{E \times I \times t}{96500} = \frac{E \times 3 \times 50 \times 60}{96500}$   
 $E = \frac{96500 \times w}{3 \times 50 \times 60} = \frac{96500 \times 1.8}{3 \times 50 \times 60} = 19.3.$
17. (a)  $\text{Al} \rightarrow \text{Al}^{3+} + 3e^-$ .  
 The charged obtained is  $3 \times 96500 \text{ C}.$
18. (a) Wt. of Ag deposited = Eq. wt. of Ag = 108 gm  
 Wt. of Ni deposited = Eq. wt. of Ni = 29.5 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  $\text{CuSO}_4$ ,  $\text{Cu}^{2+}$  gets discharged at cathode and  $\text{OH}^-$  at anode. Thus solution becomes acidic due to excess of  $\text{H}^+$  and  $\text{SO}_4^{2-}$  or  $\text{H}_2\text{SO}_4$ .
23. (b) 1 mole of electrons = 1 faraday  
 $\text{Mg}^{++} + 2e^- \rightarrow \text{Mg}$ ; 2 moles of electrons = 2 faraday.
24. (d)  $\text{Cu}^{++} + 2e^- \rightarrow \text{Cu}$   
 $E_{\text{Cu}} = \frac{63.54}{2} = 31.77$   
 Amount of electricity required to deposit .6354 gm of Cu  
 $= \frac{96500 \times 0.6354}{31.77} = 1930 \text{ Coulombs}.$
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)  $\text{Ca}^{++} + 2e^- \rightarrow \text{Ca}$   
 $E_{\text{Ca}} = \frac{40}{2} = 20$   
 $W_{\text{Ca}} = E_{\text{Ca}} \times \text{No. of faradays} = 20 \times 0.04 = 0.8 \text{ gm}.$
29. (c)  $E_{\text{metal}} = \frac{\text{Weight of metal} \times 96500}{\text{Number of coulombs}}$   
 $= \frac{22.2 \times 96500}{2 \times 5 \times 60 \times 60} = 59.5$   
 Oxidation number of the metal =  $\frac{177}{59.5} = +3$
30. (a) Quantity of electricity passed =  $\frac{25}{1000} \times 60 = 1.5$   
 $2F = 2 \times 96500 \text{ C deposit Ca} = 1 \text{ mole}$   
 $\therefore 1.5 \text{ C will deposit Ca} = \frac{1}{2 \times 96500} \times 1.5 \text{ mole}$   
 $= \frac{1}{2 \times 96500} \times 1.5 \times 6.023 \times 10^{23} \text{ atom} = 4.68 \times 10^{18}.$
31. (b) Equivalent of Cl deposited  
 = No. of Faraday passed = 0.5  
 Wt. of Cl =  $0.5 \times \text{Eq. wt.} = 0.5 \times 35.5 = 17.75 \text{ gm}.$
32. (b) At Andoe At Anode  $\text{Cl}^- \rightarrow \frac{1}{2} \text{Cl}_2 + e^-$   
 $E_{\text{Cl}_2} = \frac{35.5 \times 2}{2} = 35.5$   
 $W_{\text{Cl}_2} = \frac{E_{\text{Cl}_2} \times I \times t}{96500} = \frac{35.5 \times 2 \times 30 \times 60}{96500} = 1.32 \text{ gm}.$
34. (b) It is Faraday's law.
35. (b) Equivalent wt. of  $\text{O}_2$  = Equivalent wt. of Cu
36. (d)  $\text{O}_2\% = 20\%$   
 Metal% =  $80\% = \frac{80}{20} \times 8 = 32 \text{ g of metal}.$
37. (b)  $V = \frac{827 \times 10^3}{4 \times 96500} = 2.14 \text{ V}.$
38. (b)  $\text{Ag}^+ + e^- \rightarrow \text{Ag}$ , 96500 C will liberate silver = 108 gm. 9650C 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)  $\times$  second (S).
44. (b)  $E = -\frac{13.6}{n^2}$  for  $\text{He}^+$   $n = 1$   
 $E = -\frac{13.6}{1^2} = -13.6 \text{ eV}.$
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;  
 $\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}$ ;  $\text{Fe}^{3+} + 3e^- \rightarrow \text{Fe}$   
 $(E_{\text{Fe}})_1 = \frac{\text{Atomic.weight}}{2}$ ;  $(E_{\text{Fe}})_2 = \frac{\text{Atomic.weight}}{3}$   
 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.

52. (b) For deposition of one equivalent silver required charged is 96500 C.

53. (b)  $Cu^{++} + 2e^{-} \rightarrow Cu$ ;  $E_{Cu} = \frac{63.55}{2} = 31.75 \text{ gm Cu}.$

54. (a)  $Q = 2.5 \times 386 = 96500 \text{ C}$

$2F(2 \times 96500 \text{ C})$  deposited  $Cu = 63.5 \text{ g}$

$\therefore$  Hence 965 C will deposited;  $Cu = 0.3175 \text{ gm}.$

55. (c)  $\frac{\text{Wt. of Cu}}{\text{Wt. of Ag}} = \frac{\text{Eq. wt. of Cu}}{\text{Eq. wt. of Ag}}; \frac{\text{Wt. of Cu}}{1.08} = \frac{63.5/2}{108}$

$\text{Wt. of Cu} = 0.3177 \text{ gm}.$

56. (c) 1 g atom of Al = 3 equivalent of Al = 3 faraday charge

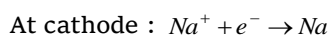
3 mole electrons = 3 N electron.

57. (c) At cathode :  $Al^{3+} + 3e^{-} \rightarrow Al$

$E_{Al} = \frac{\text{Atomic mass}}{3}$



$E_{Cu} = \frac{\text{Atomic mass}}{2}$



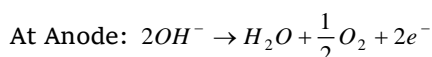
$E_{Na} = \frac{\text{Atomic mass}}{1}$

For the passage of 3 faraday;  
mole atoms of Al deposited = 1

mole atoms of Cu deposited =  $\frac{1 \times 3}{2} = 1.5$

mole atoms of Na deposited =  $1 \times 3 = 3.$

58. (d) At cathode:  $Ag^{+} + e^{-} \rightarrow Ag$



$E_{Ag} = \frac{108}{1} = 108; E_{O_2} = \frac{\frac{1}{2} \times 32}{2} = 8$

$\frac{W_{Ag}}{E_{Ag}} = \frac{W_{O_2}}{E_{O_2}}; W_{Ag} = \frac{1.6 \times 108}{8} = 21.6 \text{ gm}.$

59. (d) KI is an electrolyte.

60. (d) Number of gm equivalent = Number of faraday pass

$4 \text{ gm} = 4 \text{ faraday}.$

61. (c) Eq. of Al =  $\frac{13.5}{27/3} = 1.5.$

Thus 1.5 Faraday is needed.

63. (b) Electricity required

= No. of gm equivalent  $\times$  96500 coulombs

=  $0.5 \times 96500 = 48250 \text{ C}.$

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 1g atom of Na is deposited by one Faraday of current.

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

70. (c)  $\therefore$  1F obtained from 1 g equivalent

$\therefore$  2.5 F obtained from 2.5 g equivalent.

75. (c) Faraday constant depends upon the current passed.

80. (b) In 5 gm CuO, 4 gm Cu and 1 gm O be present.

Element	Wt.	At. Wt.	Wt. / At. Wt. $\neq x$	Ratio
Cu	4 gm	63.5	$4/63.5 = .0625$	$\frac{.0625}{.0625} = 1$
O	1 gm	16	$1/16 = .0625$	$\frac{.0625}{.0625} = 1$

Empirical formula = CuO of oxide

In this oxide, oxidation no. of Cu = +2

Equivalent weight

=  $\frac{\text{Molecular weight}}{\text{Oxidation no.}} = \frac{63.5}{2} \approx 31.75$  but Equivalent

weight should be an integral no. = 32

81. (c) Given, Current = 241.25 coulomb

1 coulomb current will deposit

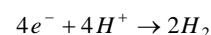
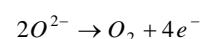
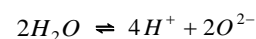
=  $1.118 \times 10^{-3} \text{ gm Ag}.$

$\therefore$  241.25 current will deposit

=  $1.118 \times 10^{-3} \times 241.25$

= 0.27 gm silver.

82. (b) Reaction for electrolysis of water is



$\therefore$   $n = 4$  so 4 Faraday charge will liberate

1 mole =  $22.4 \text{ dm}^3$  oxygen

$\therefore$  1 Faraday charge will liberate

$\frac{22.4}{4} = 5.6 \text{ dm}^3 O_2.$

83. (a)  $Na^{+} + e^{-} \rightarrow Na$

Charge (in F) = moles of  $e^{-}$  used = moles of Na deposited

=  $\frac{11.5}{23} \text{ gm} = 0.5 \text{ Faraday}.$

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

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

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

85. (c) In a galvanic cell, the electrons flow from 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.

## 532 Electrochemistry

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

$$\text{Equivalent weight of Cu} = \frac{63.6}{2} = 31.8$$

$$\frac{M_1}{M_2} = \frac{E_1}{E_2}; \quad \frac{10.79}{M_{Cu}} = \frac{108}{31.8}$$

$$M_{Cu} = \frac{10.79 \times 31.8}{108} = 3.2 \text{ gm.}$$

87. (b) Laws of electrolysis were proposed by Faraday.

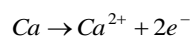
88. (a) Given, Current ( $i$ ) = 25 mA = 0.025 A

Time ( $t$ ) = 60 sec

$$Q = i t = 60 \times 0.025 = 1.5 \text{ coulombs}$$

$$\text{No. of electrons} = \frac{1.5 \times 6.023 \times 10^{23}}{96500}$$

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



$2e^-$  are required to deposit one Ca atom

$9.36 \times 10^{18} e^-$  will be used to deposit

$$= \frac{9.36 \times 10^{18}}{2} = 4.68 \times 10^{18}.$$

89. (d)  $C_6H_5NO_2 + 6H^+ + 6e^- \rightarrow C_6H_5NH_2 + 2H_2O$

1 mole = 123 gm nitrogen requires 6 mole electron  $e$

$$= 6 \times 96500 \text{ coulomb charge}$$

$\therefore$  12.3 gm nitrobenzene will require

$$= \frac{6 \times 96500 \times 12.3}{123} = 6 \times 9650 = 57900 C.$$

90. (c) Au and Ag settle down below the anode as anode mud during the process of electrolytic refining of copper.

16. (d) Because conductance is increase when the dissociation is more.

17. (b) Strong electrolyte ionize completely at all 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.

18. (d) In electrolytic conductors, a single stream of electrons flow from cathode to anode.

19. (b) In solid state NaCl does not dissociate into ions so it does not conduct electricity.

20. (c) The ions are not free to move in solid state and held up in lattice due to strong coulombic forces of attraction.

21. (b)  $C_2H_5OH$  being non electrolyte so does not ionize.

22. (a) Since molar conductance  $\propto \frac{1}{\text{Molarity}}$ .

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

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

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

$$\rho = \frac{Ra}{l} = \frac{50}{0.5} = 100$$

$$\Lambda = k \times \frac{1000}{N} = \frac{1}{\rho} \times \frac{1000}{N} = \frac{1}{100} \times \frac{1000}{1}$$

$$10 \text{ ohm}^{-1} \text{ cm}^2 \text{ gm eq}^{-1}$$

26. (b)  $\Lambda_{m(C_6H_5COOH)}^o = \Lambda_{(C_6H_5COO^-)}^o + \Lambda_{(H^+)}^o$

$$= 42 + 288.42 = 330.42$$

$$\alpha = \frac{\Lambda_m^c}{\Lambda_m^o} = \frac{12.8}{330.42} = 3.9\%$$

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

## Conductor and Conductance

2. (b)  $\lambda^\infty BaCl_2 = \frac{1}{2} \lambda^\infty Ba^{2+} + \lambda^\infty Cl^-$

$$= \frac{127}{2} + 76 = 139.5 \text{ ohm}^{-1} \text{ cm}^{-1} \text{ eq}^{-1}.$$

3. (d) Dilution, temperature and nature of electrolyte affect the conductivity of solution.

5. (a) Generally strong electrolyte on dilution shows conductivity characters.

6. (b) Molar conductivity =  $\frac{1000}{MX}$ .

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

9. (b) Conductivity of a solution is directly proportional to the number of ions.

11. (a)  $NaCl \rightleftharpoons Na^+ + Cl^-$ . So it conduct electricity.

12. (b) Graphite is a good conductor of electricity.

15. (b) Electrolytic conduction resistance decreases with increasing temperature.

## Cell constant and Electrochemical cells

1. (d) In the absence of electric field the ions in the solution move randomly due to thermal energy.

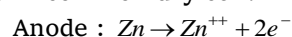
2. (d) Since  $E_{A/A^-}^o$  has large negative value, the tendency of A to be reduced to  $A^-$  is very small. In other words tendency of  $A^-$  to be oxidized to A is very large.

3. (d) Practically only 60-70% efficiency has been attained.

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

$$\text{Cell constant} = K \times R; 0.012 \times 55 = 0.66 \text{ cm}^{-1}.$$

5. (b) In common dry cell.



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

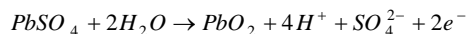
6. (a) Because the reduction potential of Cu is highest.

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:  $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$

Cathode:



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



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)  $E_{\text{cell}}^{\circ} = \frac{2.303 RT}{nF} \log K = \frac{0.0591}{n} \log K_c \text{ at } 298 K$ .

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

Reductio

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

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

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

27. (c) Velocities of both  $K^+$  and  $NO_3^-$  are nearly the same in  $KNO_3$  so it is used to make salt-bridge.

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)  $\wedge_m^{\circ}(CH_3COOH) =$

$$\wedge^{\circ}(CH_3COONa) + \wedge^{\circ}(HCl) - \wedge^{\circ}(NaCl)$$

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

36. (b) At anode:  $Zn_{(s)} \rightarrow Zn^{2+} + 2e^-$ .

38. (d)  $PbSO_4 + 2H_2O \rightarrow 2PbO_2 + 4SO_4^{--} + 2e^-$ .

39. (b)  $MnO_2$  is used in dry batteries cell.

40. (d)  $Pb + PbO_2 + 2H_2SO_4 \xrightleftharpoons[\text{Recharge}]{\text{Discharge}} 2PbSO_4 + 2H_2O$ .  
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.

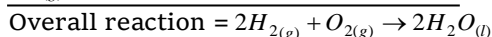
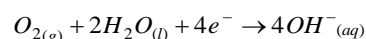
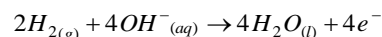
43. (c) In the electrochemical cell chemical energy changes into electrical energy.

44. (a) In galvanic cell, the salt bridge used to complete the circuit.

45. (d)  $Cu + FeSO_4 \rightarrow$  No reaction Because Cu has  $E^{\circ} = 0.34 \text{ volt}$  and Fe has  $E^{\circ} = -0.44 \text{ volt}$ .

47. (d) Calomel electrode as reference electrode is made by using  $Hg_2Cl_2$ .

48. (b) In hydrogen-oxygen fuel cell following reactions take place to create potential difference between two electrodes.



the net reaction is the same as burning (Combustion) of hydrogen to form water.

49. (c)  $ClCH_2COONa + HCl \rightarrow ClCH_2COOH + NaCl$

$$\lambda_{ClCH_2COONa} + \lambda_{HCl} = \lambda_{ClCH_2COOH} + \lambda_{NaCl}$$

$$224 + 203 = \lambda_{ClCH_2COOH} + 38.2$$

$$\lambda_{ClCH_2COOH} = 427 - 38.2 = 388.8 \text{ ohm}^{-1} \text{cm}^2 \text{gmeq}^{-1}$$

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. It is 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  $K[Au(CN)_2]$ .

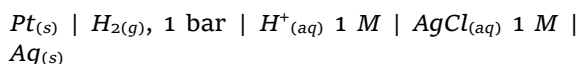
54. (a) Dil.  $H_2SO_4$  is used in lead in lead storage battery as electrolyte.

55. (c) Cell constant =  $\frac{\text{Specific conductivity}}{\text{Observed conductance}}$   
 $= \frac{0.002765}{1/R} = 0.002765 \times 400 = 1.106$ .

57. (b)  $2AgCl_{(s)} + H_{2(g)} \rightarrow 2HCl_{(aq)} + 2Ag_{(s)}$

The activities of solids and liquids are taken as unity and at low concentrations, the activity of a solute is approximated to its molarity. The cell reaction will be

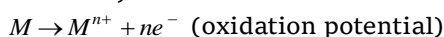
## 534 Electrochemistry



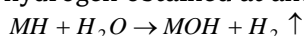
58. (a)  $E_{cell} = \frac{0.059}{n} \log \frac{1}{C} = -\frac{0.059}{2} \log \frac{1}{100}$   
 $= -\frac{0.059}{2} (-2) = 0.059 V = 59 mV$  . (increase)

59. (a)  $E_{cell} = -0.059 \log \frac{10^{-6}}{10^{-3}} = -0.059 \log 10^{-3}$   
 $= -0.059 \times (-3) = 0.177 V$  .

60. (c) The magnitude of the electrode potential of a 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).



61. (a) On electrolysis of fused ionic hydride ( $LiH$ ), hydrogen obtained at anode.



### Electrode potential, $E_{cell}$ , Nernst equation and ECS

1. (b) Reduction potential of hydrogen electrode,

$$E_H = \frac{-2.303 RT}{F} \log \frac{1}{[H^+]}$$

$$= -0.059 pH = -0.059 \times 3 = -0.177 V$$
 .

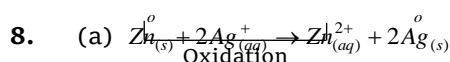
2. (a)  $E_{cell}^o = E_{cathode}^o - E_{anode}^o = 0.799 - (-0.763) = 1.562 V$   
 3. (a) More negative is the reduction potential, higher will be the reducing property, i.e. the power to give up electrons.  
 4. (b) Standard potential of Zinc < Copper.  
 6. (c) A cation having highest reduction potential will be reduced first and so on. However,  $Mg^{2+}$  in aqueous solution will not be reduced

$$\left( E_{Mg^{2+}/Mg}^o < E_{H_2O/\frac{1}{2}H_2+OH^-} \right) . \text{ Instead water would}$$

be reduced in preference.

7. (c) A is displace from D because D have a  
 $E^o = -0.402 V$

Reduction



In this reaction zinc act as a anode and Ag act as a cathode.

9. (a) No doubt Be is placed above Mg in the second group of periodic table but it is below Mg in electrochemical series.  
 10. (b) Nernst's equation shows relation between E and  $E^o$  .

11. (a)  $E = E^o - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$  ;  $E = E^o + \frac{RT}{nF} \ln [M^{n+}]$

$$E = E^o + \frac{2.303 RT}{nF} \log [M^{n+}]$$

Substituting the value of R, T (298K) and F we get

$$E = E^o + \frac{0.0591}{n} \log (M^{n+}) .$$

12. (c) At 298 K standard electrode potential of NHE electrode is 0.00 V.  
 13. (a) Since,  $Ag^+$  ions are reduced to Ag and  
 $E_{Ag^+/Ag}^o > E_{Cu^{++}/Cu}^o$  Cu is oxidized to  $Cu^{++}$  .  
 14. (d) The reducing power decreases as the reduction potential increase (becomes less negative).  
 15. (c) Actually the equation is derived from Nernst equation assuming equilibrium condition in a cell reaction, when  $E = 0$  .  
 16. (a) More negative is the standard reduction potential, greater is the tendency to lose electrons and hence greater reactivity.  
 17. (a) Hg 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  $Fe^{2+}$  .  
 19. (b) Since  $E_{A^{2+}/A}^o < E_{B^{2+}/B}^o$  . A has greater tendency to be oxidized.  
 $A + B^{2+} \rightarrow A^{2+} + B$  .  
 20. (b) Since  $E_{Zn^{++}/Zn}^o$  is negative, so Zn has greater tendency to be oxidized than hydrogen. Hence it can act as reducing agent.  
 21. (a) Standard electrode potential of Hydrogen is zero.  
 22. (b) According to electrochemical series.  
 23. (a) The standard reduction potential of  $K^+$ ,  $Mg^{2+}$ ,  $Zn^{2+}$ ,  $Cu^{2+}$  increase in this order.  
 24. (c)  $E_{cell} = E_{Au^{3+}/Au}^o - E_{Ni^{2+}/Ni}^o = 1.50 - (-0.25) = 1.75 V$  .  
 25. (a) Electromotive force is +ve if oxidation and reduction both takes place in a cell.  
 28. (a) In galvanic cell anode always made up of negative electrode.  
 29. (d)  $A | A^+(a=1) || B^+(a=1) | B$   
 $EMF = E_{cathode} - E_{anode} = 0.75 - (0.5) ; EMF = 0.25 V$  .  
 30. (d)  $E^o = -3.05 Li^+/Li$  is most negative (minimum) and hence Li has maximum tendency to lose electrons or it is the strongest reducing agent.  
 31. (b) Brown layer is deposited on iron rod because Cu has greater than reduction potential than that of  $Fe^{2+}$  .  
 32. (a)  $E_{Zn^{++}/Zn}^o < E_{Fe^{++}/Fe}^o$  , so Zn will reduce  $Fe^{++}$  . Zn cannot reduce  $Mg^{2+}$  because  $E_{Zn^{++}/Zn}^o > E_{Mg^{++}/Mg}^o$



On similar reason *Mg* and *Zn* cannot oxidize *Fe*.

33. (d) For the cell reaction, *Fe* acts as cathode and *Sn* as anode. Hence,

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = -0.44 - (-0.14) = -0.30 \text{ V}$$

The negative *EMF* suggests that the reaction goes spontaneously in reversed direction.

34. (a)  $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = 0.34 - (-0.76) = 1.10 \text{ V}$ .

35. (c)  $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$ ;  $E_{\text{cell}}^{\circ} = 0.34 - (-2.37)$   
 $E_{\text{cell}}^{\circ} = 2.71 \text{ V}$ .

36. (d) *Mg* lies above *Cu* in electrochemical series and hence *Cu* electrode acts as cathode

$$E_{\text{cell}}^{\circ} = E_{\text{Cu}^{++}/\text{Cu}}^{\circ} - E_{\text{Mg}^{++}/\text{Mg}}^{\circ}$$

$$2.70 \text{ V} = 0.34 - E_{\text{Mg}^{++}/\text{Mg}}^{\circ}; E_{\text{Mg}^{++}/\text{Mg}}^{\circ} = -2.36 \text{ V}$$

37. (a) Because  $H_2$  has greater reduction potential so it reduced the  $Ag^+$ .

39. (d)  $\Delta G^{\circ} = -nE^{\circ}F$

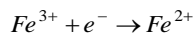


$$\Delta G^{\circ} = -2 \times F \times (-0.440 \text{ V}) = 0.880 F$$



$$\Delta G^{\circ} = -3 \times F \times (-0.036) = 0.108 F$$

On subtracting equation (i) from (ii)



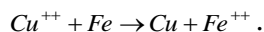
$$\Delta G^{\circ} = 0.108 F - 0.880 F = -0.772 F$$

$$E^{\circ} \quad \text{for the reaction} = -\frac{\Delta G^{\circ}}{nF}$$

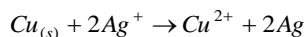
$$= -\frac{(-0.772 F)}{1 \times F} = +0.772 \text{ V}$$

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.



42. (c) Cell reaction is



Two half cell reaction is



$$E_{\text{cell}} = E_{\text{ox}} - E_{\text{red}} = 0.80 - 0.34 = +0.46 \text{ V}$$

43. (a)  $EMF = [\text{s.r.p. of cathode} - \text{s.r.p. of anode}]$

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

(b) Cathode is made of copper

(c) Anode is made of Zinc

$$EMF = 0.34 - (-0.76) = 1.1 \text{ volt}$$

46. (d)  $H_2$  is anode because oxidation takes place.

*Cu* is cathode because reduction takes place.

47. (c)  $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$ .

51. (b) Standard hydrogen electrode have zero electrode potential.

53. (a)  $\Delta G = -nFE^{\circ}$

$$\Delta G = -1 \times 96500 \times 1.02; \Delta G = -98430$$

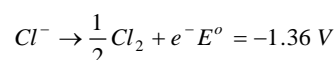
54. (c) Fuel cell converts the chemical energy into electrical energy.

55. (a)  $E = E^{\circ} - \frac{2.303 RT}{nF} \log \frac{[M]}{[M^{n+}]}$ .

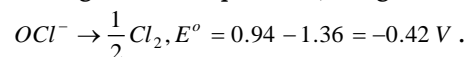
56. (a) Less is the reduction potential stronger is the reducing agent.

57. (a) Reducing power, *i.e.* the tendency to lose electrons increases as the reduction potential decreases.

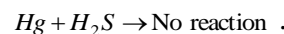
58. (a)  $OCl^- \rightarrow Cl^-$ ,  $E^{\circ} = 0.94 \text{ V}$



adding the two equations, we get



60. (a) It cannot evolved  $H_2$  from  $H_2S$



61. (b)  $E_{\text{cell}}^{\circ} = \frac{0.059}{n} \log K$

$$\log K = \frac{1.10 \times 2}{0.059} = 37.2881 \quad \text{or } K = 10^{-37}$$

62. (d) The oxidizing character *i.e.* acceptance of electrons increases with the reduction potential.

63. (a) According to electrochemical series.

64. (d)  $E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \frac{C_2}{C_1}$  and  $\Delta G = -nF E_{\text{cell}}$

hence  $\Delta G$  is the function of  $\ln \left( \frac{C_2}{C_1} \right)$ .

66. (c)  $E = E^{\circ} - \frac{0.059}{n} \log \frac{[Zn^{++}]}{[Cu^{++}]} = 1.10 - \frac{0.059}{2} \log \frac{1}{0.1}$

$$= 1.10 - 0.0295 \log 10 = 1.07 \text{ volt}$$

67. (b)  $E_1 = E_o - \frac{0.0591}{2} \log \frac{0.01}{1} = E_o + \frac{0.0591}{2} \times 2$

$$E_2 = E_o - \frac{0.0591}{2} \log \frac{100}{0.01} = E_o - \frac{0.0591}{2} \times 4$$

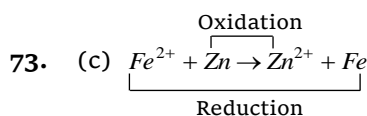
$$\therefore E_1 > E_2$$

68. (b) 
$$\begin{array}{c} \text{Oxidation} \\ \uparrow \\ Fe^{+2} + Zn \rightarrow Zn^{2+} + Fe \\ \downarrow \\ \text{Reduction} \end{array}$$

$$EMF = E_{\text{cathode}} - E_{\text{anode}} = 0.44 - (0.76) = +0.32 \text{ V}$$

69. (a) *Fe* is more electropositive than copper. Hence  $Cu^{2+}$  can oxidise *Fe*.

70. (b)  $E^\circ = 0$  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.



$$EMF = E_{\text{cathode}} - E_{\text{anode}} = -7.81 - (-7.62)$$

$$EMF = -0.19 \text{ V}.$$

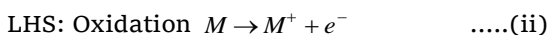
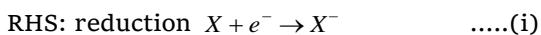
74. (c)  $Cr^{3+} > Zn^{2+} > H > Fe^{3+}$ .  
 Reducing nature decreasing order.  
 75. (a) More is reduction potential, more is the power to get itself reduced or lesser is reducing power or greater is oxidising power.

76. (d)  $\Delta G = -nFE^\circ$   
 $\Delta G = -2.303 RT \log K$ ;  $nFE^\circ = 2.303 RT \log K$

$$\log K = \frac{nFE^\circ}{2.303 RT} = \frac{2 \times 96500 \times 0.295}{2.303 \times 8.314 \times 298}$$

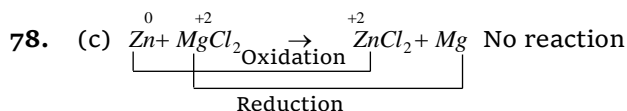
$$\log K = 9.97 = K = 1 \times 10^{10}.$$

77. (b) For the given cell  $M | M^+ || X^- | X$ , the cell reaction is derived as follows:



The cell potential =  $-0.11 \text{ V}$

Since  $E_{\text{cell}} = -$  ve, the cell reaction derived above is not spontaneous. In fact, the reverse reaction will occur spontaneously.



This type of reaction does not occur because

$$Mg^{2+} E^\circ = -2.37 \text{ V} \text{ while } Zn^{2+} E^\circ = -0.76 \text{ V}.$$

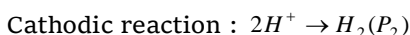
79. (b) In neutral medium  $Mn^{+7}$  oxidation state change into  $+4$  oxidation state, hence equivalent weight of  $KMnO_4 = \frac{M}{3}$ .

80. (a) Increase in the concentration of  $Ag^+$  ion increase the voltage of the cell.

81. (a)  $E_{\text{cell}} = E_{\text{cell}}^\circ + \frac{0.059}{2} \log \frac{(Ag^+)}{(Sn^{2+})}$ .

82. (b) The  $K.E.$  of proton is  $1 \text{ KeV}$ .

83. (b) Anodic reaction :  $H_2(P_1) \rightarrow 2H^+$



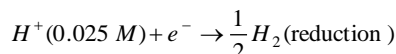
$$E_{\text{cathode}} = -\frac{RT}{2F} \ln \frac{P_2}{[H^+]^2}; E_{\text{anode}} = -\frac{RT}{2F} \ln \frac{[H^+]^2}{P_1}$$

$$E_{\text{inf}} = E_{\text{anode}} + E_{\text{cathode}}$$

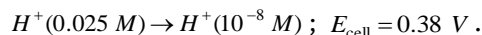
$$= -\frac{RT}{2F} \ln \frac{(H^+)^2}{P_1} - \frac{RT}{2F} \ln \frac{P_2}{(H^+)^2}$$

$$= -\frac{RT}{2F} \ln \frac{P_2}{P_1} = \frac{RT}{2F} \ln \frac{P_1}{P_2}.$$

84. (c)  $\frac{1}{2} H_2 \rightarrow H^+(10^{-8} M) + e^-$  (oxidation)



Cell reaction is :



85. (a)  $E^\circ$  for  $Fe / Fe^{2+} = 0.44 \text{ V}$ .  
 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)  $E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ$   
 $= 0.34 - (-2.37) = +2.71 \text{ V}$ .  
 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  $0.1 \text{ M HCl}$  and  $0.1 \text{ M}$  acetic acid is not the same, because  $HCl$  is a strong acid so its  $pH$  is less and  $CH_3COOH$  is a weak acid, so its  $pH$  is more.  
 95. (d) The required reaction ( $Cu^{++} + Cu \rightarrow 2Cu^+$ ) can be obtained by using the following reactions.  
 $Cu^{++} + e^- \rightarrow Cu^+; E_{Cu^{++}/Cu^+}^\circ = 0.15 \text{ V}$  .....(i)  
 $Cu^{++} + 2e^- \rightarrow Cu; E_{Cu^{++}/Cu}^\circ = 0.34 \text{ V}$  .....(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)  
 $Cu^{++} + Cu \rightarrow 2Cu^+$   
 $\Delta G_3 = -nFE = -1 \times F \times E^\circ$   
 Also  $\Delta G_3 = \Delta G_1 - \Delta G_2$   
 $-1FE^\circ = (-2F \times 0.15) - (-2F \times 0.34)$   
 $E^\circ = -0.38$   
 This is the value for the reaction  
 $Cu^{++} + Cu \rightarrow 2Cu^+$   
 But the given reaction is just reverse of it  
 $\therefore E_{\text{cell}}^\circ$  for given reaction =  $+0.38 \text{ V}$ .  
 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)  $E^{\circ} = E_{\text{Ag}^{2+}/\text{Ag}}^{\circ} + E_{\text{Cu}/\text{Cu}^{2+}}^{\circ} = -0.34 + 0.80 = +0.46 \text{ V}$ .
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)  $\text{Zn}_{(s)} + 2\text{H}^{+}_{(aq)} \rightleftharpoons \text{Zn}^{2+}_{(aq)} + \text{H}_{2(g)}$

$$E_{\text{Cell}} = E_{\text{Cell}}^{\circ} - \frac{0.059}{2} \log \frac{[\text{Zn}^{2+}]}{[\text{H}^{+}]^2}$$

When  $\text{H}_2\text{SO}_4$  is added then  $[\text{H}^{+}]$  will increase therefore  $E_{\text{Cell}}$  will also increase and equilibrium will shift towards right.

108. (b) For  $\text{M}^{+} + \text{X}^{-} \rightarrow \text{M} + \text{X}$
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} + E_{\text{Anode}}^{\circ} = 0.44 - 0.33 = +0.11 \text{ V}$$
- Since  $E_{\text{Cell}}^{\circ} = (+) 0.11 \text{ V}$  is positive hence this reaction should be spontaneous.
109. (a)  $E_{\text{Cell}} = E_{\text{Cell}}^{\circ} - \frac{0.0591}{n} \log K_c$
- $$\text{At } 298 \text{ K } E_{\text{Cell}} = 0 \quad 0 = 0.591 - \frac{0.0591}{n} \log K_c$$
- $$\log K_c = \frac{0.591 \times 1}{0.0591} = 10 ; K_c = \text{Antilog } 10 = 1 \times 10^{10}.$$
110. (a)  $\frac{1}{2} \text{H}_2 | \text{H}^{+} || \text{Ag}^{+} | \text{Ag}$
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ} = E_{\text{Ag}^{+}/\text{Ag}}^{\circ} - E_{\text{H}^{+}/\frac{1}{2}\text{H}_2}^{\circ}$$
- $$(0.80) - (0.0) = 0.80 \text{ V}.$$
111. (a)  $E_A = 2.23 \text{ V} > E_B = 1.43 \text{ V}$
- So A will act as cathode in galvanic cell. Hence
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ} = E_A - E_B$$
- $$= (2.23) - (-1.43) = 3.66 \text{ V}.$$
112. (b)  $E_{\text{Cu}}^{\circ} > E_{\text{Mg}}^{\circ}$  hence Cu acts as cathode and Mg acts as anode.
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cu}}^{\circ} - E_{\text{Mg}}^{\circ} = (0.34) - (-2.37) = +2.71 \text{ V}.$$
113. (b) In this cell Co is oxidised and it acts as anode and Ce acts as cathode.
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ} = 1.89 = E_{\text{Cell}}^{\circ} - (-0.28)$$
- $$E_{\text{Cell}}^{\circ} = 1.89 - 0.28 = 1.61 \text{ Volts}.$$
114. (b) Given:  $\Delta G = -21.20 \text{ kJ} = 21200 \text{ J}$
- $$\therefore \Delta G = -nFE$$

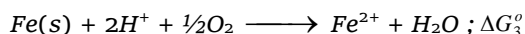
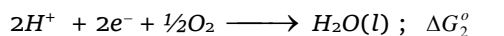
$$E = \frac{21200}{1 \times 96500} = 0.2196 \text{ V} = 0.22 \text{ V}.$$

115. (b)  $\text{Ag} | \text{Ag}^{+}(1\text{M}) || \text{Ag}^{+} 1\text{M} | \text{Ag}$
- $$E_{\text{Cell}} = \frac{2.303 RT}{nF} \log \frac{c_1}{c_2} = \frac{0.059}{1} \log \frac{1}{0.1}$$
- $$= 0.059 \log 10 = 0.059 \text{ Volt}.$$
116. (b) For this cell, reaction is:  $\text{Zn} + \text{Fe}^{2+} \rightarrow \text{Zn}^{2+} + \text{Fe}$
- $$E = E^{\circ} - \frac{0.0591}{n} \log \frac{c_1}{c_2}; E^{\circ} = E + \frac{0.0591}{n} \log \frac{c_1}{c_2}$$
- $$= 0.2905 + \frac{0.0591}{2} \log \frac{10^{-2}}{10^{-3}} = 0.32 \text{ V}.$$
- $$E^{\circ} = \frac{0.0591}{2} \log K_c; \log K_c = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$$
- $$\therefore K_c = \frac{0.32}{10^{0.295}}.$$
117. (d) Al displaces H from HCl but silver cannot it means Al is situated above the Ag in ECS, hence Al will act as anode and Ag will act as cathode.
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ} = E_{\text{Ag}^{+}/\text{Ag}}^{\circ} - E_{\text{Al}^{3+}/\text{Al}}^{\circ}$$
- $$2.46 = 0.8 - E_{\text{Al}^{3+}/\text{Al}}^{\circ}; E_{\text{Al}}^{\circ} = 0.8 - 2.46 = -1.66 \text{ V}.$$
118. (a) For  $\text{Sn}_{(s)} + 2\text{Fe}^{3+}_{(aq)} \rightarrow 2\text{Fe}^{2+}_{(aq)} + \text{Sn}^{2+}_{(aq)}$
- $$E_{\text{Cell}}^{\circ} = E_{\text{Sn}/\text{Sn}^{2+}}^{\circ} + E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^{\circ} = (0.14) + (0.77)$$
- $$E_{\text{Cell}} = 0.91 \text{ Volts}.$$
119. (d)  $\text{I}^{-}$  get oxidised to  $\text{I}_2$  hence will form anode and  $\text{Cr}_2\text{O}_7^{2-}$  get reduced to  $\text{Cr}^{3+}$  hence will form cathode.
- $$E_{\text{Cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ}; E_{\text{Cell}}^{\circ} = E_{\text{Cr}_2\text{O}_7^{2-}}^{\circ} - E_{\text{I}_2}^{\circ}$$
- $$0.79 = 1.33 - E_{\text{I}_2}^{\circ}; E_{\text{I}_2}^{\circ} = 1.33 - 0.79; E_{\text{I}_2}^{\circ} = 0.54 \text{ V}.$$
120. (b) According to nernst's equation
- $$E_{\text{Cell}} = E_{\text{Cell}}^{\circ} - \frac{nRT}{F} \log \frac{c_1}{c_2}$$
- For  $\text{Zn}_{(s)} + \text{Cl}_{2(1 \text{ atm})} \rightarrow \text{Zn}^{2+} + 2\text{Cl}^{-}$
- $$c_1 = [\text{Zn}^{2+}] \text{ and } c_2 = [\text{Cl}^{-}]$$
- Hence to increase  $E$ ,  $c_1$  should be decreased and  $c_2$  should be increased is  $[\text{Zn}^{2+}]$  should be decreased and Cl should be increased.
121. (c)

Reduction $E_0 \text{M}^{3+} / \text{M}^{2+}$	Cell reaction		$E_0 \text{M}^{2+} / \text{M}^{3+}$ (Oxidation)
- .41 V	$\text{Cr}^{2+}$	$\text{Cr}^{3+}$	+ .41 V
+ 1.57 V	$\text{Mn}^{2+}$	$\text{Mn}^{3+}$	- 1.57 V
+ 0.77 V	$\text{Fe}^{2+}$	$\text{Fe}^{3+}$	- 0.77 V
+ 1.97 V	$\text{Co}^{2+}$	$\text{Co}^{3+}$	- 1.97 V

As Cr has maximum oxidation potential value, therefore its oxidation should be easiest.

122. (a)
- $Fe(s) \longrightarrow Fe^{2+} + 2e^-$
- ;
- $\Delta G_1^\circ$

Applying,  $\Delta G_1^\circ + \Delta G_2^\circ = \Delta G_3^\circ$ 

$$\Delta G_3^\circ = (-2F \times 0.44) + (-2F \times 1.23)$$

$$\Delta G_3^\circ = -(2 \times 96500 \times 0.44 + 2 \times 96500 \times 1.23)$$

$$\Delta G_3^\circ = -322310 J$$

$$\therefore \Delta G_3^\circ = -322 kJ$$

128. (a)
- $\Delta G^\circ = -2.303 RT \log K_{eq}$
- or
- $\Delta G^\circ = -nFE_{cell}^\circ$

129. (a) Any redox reaction would occur spontaneously if the free energy change (
- $\Delta G$
- ) is negative.

$$\Delta G^\circ = nFE^\circ$$

Where  $n$  is the number of electrons involved,  $F$  is the value of Faraday and  $E^\circ$  is the cell emf.  $\Delta G^\circ$  can be negative if  $E^\circ$  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)
- $(126 \text{ scm}^2) \wedge_{NaCl}^0 = \wedge_{Na^+}^0 + \wedge_{Cl^-}^0$
- .....(1)

$$(152 \text{ scm}^2) \wedge_{KBr}^0 = \wedge_{K^+}^0 + \wedge_{Br^-}^0$$
 .....(2)

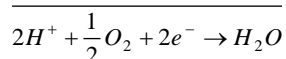
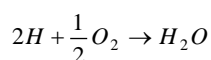
$$(150 \text{ scm}^2) \wedge_{KCl}^0 = \wedge_{K^+}^0 + \wedge_{Cl^-}^0$$
 .....(3)

By equation (1)+(2) - (3)

$$\therefore \wedge_{NaBr}^0 = \wedge_{Na^+}^0 + \wedge_{Br^-}^0$$

$$= 126 + 152 - 150 = 128 \text{ Scm}^2 \text{ mol}^{-1}$$

2. (a) At cathode:
- $2H^+_{(aq)} + 2e^- \rightarrow 2H$



3. (c)
- $H_2$
- undergoes oxidation and
- $AgCl (Ag^+)$
- undergoes reduction.

4. (b) In this reaction
- $Fe^{2+} + Zn \xrightarrow[\text{Reduction}]{\text{Oxidation}} Zn^{2+} + Fe$

$$EMF = E_{\text{cathode}} - E_{\text{anode}} = -0.41 - (-0.76)$$

$$EMF = +0.35 V.$$

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)
- $E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.059}{2} \log \frac{(Zn^{++})}{(Cu^{++})}$
- 
- $= 1.10 - \frac{0.059}{2} \log \frac{0.1}{0.1} = 1.10 V$
- .

7. (a) The tendency to gain electron is in the order
- $Z > Y > X$

Thus  $Y + e \rightarrow Y^-$ ;  $X \rightarrow X^+ + e$ .

8. (b)
- $E_{OP} = E_{OP}^\circ - \frac{0.059}{1} \log \frac{[H^+]}{P_{H_2}}$
- 
- $\therefore [H^+] = 10^{-10}$
- ;
- $P_{H_2} = 1 \text{ atm}$
- ;
- $E_{OP} = 0.59 V$
- .

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
- $3 \times 96500 C$
- .

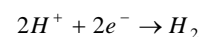
$\therefore 1 \text{ gm of Al}$  is obtained by passing a current of  $3 \times \frac{96500}{27} C$ .

$\therefore 5.12 \times 10^3 \text{ gm of Al}$  is obtained by passing a current of  $3 \times \frac{96500}{27} \times 5.12 \times 1000$   
 $= 1.83 \times 10^7 C \times 3 = 5.49 \times 10^7 C$ .

11. (c)
- $\Lambda_{HOAC}^\infty = \Lambda_{NaOAC}^\infty + \Lambda_{HCl}^\infty - \Lambda_{NaCl}^\infty$
- 
- $= 91.0 + 426.2 - 126.5 = 390.7$

12. (d)
- $\frac{\omega_1}{E_1} = \frac{\omega_2}{E_2}$
- ;
- $\frac{\omega_1}{3} = \frac{270}{93}$
- ;
- $\omega_1 = 90 \text{ kg}$
- .

13. (c)
- $Eq \text{ of Al} = eq \text{ of } H_2$
- 
- $\frac{4.5}{27} = eq \text{ of } H_2$
- ;
- $\frac{4.5}{9} = eq \text{ of } H_2$
- 
- $3$

eq. of  $H_2 = \text{Number of moles} \times n \text{ factor}$ 

$$0.5 = n_{H_2} \times 2$$

$$V_{H_2} = \frac{0.5}{2} \times 22.4; V_{H_2} = 5.6 L$$

14. (c) The reaction taking place at anode is
- $2Cl^- \rightarrow Cl_2 + 2e^-$

$$1 \text{ mole} \quad 2 \times 96500 \text{ coulomb}$$

$$Q = i \cdot t = 1 \times 30 \times 60 = 1800 \text{ coulomb.}$$

The amount of chlorine liberated by passing 1800 coulomb of electric charge

$$= \frac{1 \times 1800 \times 71}{2 \times 96500} = 0.66 \text{ g.}$$

## Assertion &amp; 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".
  4. (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.
  7. (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 deposits 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)  $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
- 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.  
 $E_{Au^{3+}/Au}^0 - E_{Ni/Ni^{2+}}^0 = 1.50 - (-0.25) = 1.75 V$
  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.