Ordinary Thinking

Objective Questions

Electrolytes and Electrolysis

1. Which of the following will not conduct electricity in aqueous solution **[AMU 1982, 83]**

(a) Copper sulphate (b) Sugar

(c) Common salt (d) None of these

2. Strong electrolytes are those which **[MNR 1983]**

(a) Dissolve readily in water

(b) Conduct electricity

(c) Dissociate into ions at high dilution

(d) Completely dissociate into ions at all dilutions

3. In aqueous solution, strong electrolytes **[AMU 1983, 84]**

(a) Are partially ionized

(b) Do not ionise

(c) Ionise almost completely

(d) Form polymers

4. An electrolyte **[KCET 1984; MP PET/PMT 1988]**

(a) Forms complex ions in solution

(b) Gives ions only when electricity is passed

(c) Possesses ions even in solid state

(d) Gives ions only when dissolved in water

5. Electrolytes when dissolved in water dissociates into ions because **[CPMT 1974, 78; MNR 1983]**

(a) They are unstable

(b) The water dissolves it

(c) The force of repulsion increases

(d) The forces of electrostatic attraction are broken down by water

6. Electrolyte can conduct electricity because

(a) Their molecules contain unpaired electrons, which are mobile

(b) Their molecules contain loosely held electrons which get free under the influence of voltage

(c) The molecules break up into ions when a voltage is applied

(d) The molecules are broken up into ions when the electrolyte is fused or is dissolved in the solvent

7. Which one of the following metals could not be obtained on electrolysis of aqueous solution of its salts **[IIT 1990]**

(a) Ag (b) Mg

(c) Cu (d) Cr

8. Which of the following aqueous solution will conduct an electric current quite well **[MP PMT 1987]**

(a) Glycerol (b) HCl

(c) Sugar (d) Pure water

9. On the electrolysis of aqueous solution of sodium sulphate, on cathode we get

(a) Na (b) H_2

(c) SO_2 (d) SO_3

10. Electrolysis involves oxidation and reduction respectively at

[CPMT 1973; AMU 1983; NCERT 1983, 84; MH CET 2001]

(a) Anode and cathode (b) Cathode and anode

(c) At both the electrodes (d) None of the above

11. Which of the following compounds will not undergo decomposition on passing electricity through aqueous solution

(a) Sugar (b) Sodium Chloride

(c) Sodium Bromide (d) Sodium Acetate

12. During the electrolysis of an electrolyte, the number of ions produced, is directly proportional to the **[AFMC 2002]**

(a) Time consumed

(b) Electro chemical equivalent of electrolysis

(c) Quantity of electricity passed

(d) Mass of electrons

13. When the sample of copper with zinc impurity is to be purified by electrolysis, the appropriate electrodes are

[AIEEE 2002]

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

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

[IIT Screening 2003]

(a) Cathode to anode in solution

(b) Cathode to anode through external supply

(c) Cathode to anode through internal supply

(d) Anode to cathode through internal supply

15. An electric current is passed through an aqueous solution of the following. Which one shall decompose **[NCERT 1972]**

(a) Urea (b) Glucose

(c) $AgNO_3$ (d) Ethyl alcohol

16. The electric conduction of a salt solution in water depends on the

(a) Shape of its molecules

(b) Size of its molecules

(c) Size of solvent molecules

(d) Extent of its ionization

17. A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively **[IIT 1987,96]**

(a) H_2, O_2 (b) O_2, H_2

(c) O_2, Na (d) O_2, SO_2

18. On electrolysing a solution of dilute H_2SO_4 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

[NCERT 1973]

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]

[KCET 1991]

(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]

(a) Calcium (b) Phosphorus

(c) Sulphur (d) Sodium

37. Which of the following properties of pure metal makes it more useful then 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

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

[EAMCET 1990; MP PET 1994, 97]

(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]**

[Roorkee 2000]

(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,**

(a) 2.0 *gm* (b) 3.175 *gm*

(c) 63.5 *gm* (d) 127.0 *gm*

10. 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) 4.68×10^{18} (b) 4.68×10^{15}

(c) 4.68×10^{12} (d) 4.68×10^{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*

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 ($F = 96500C\ mol^{-1}$)**[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) 9.65×10^4 *Coulomb* (b) 6.28×10^{18} *Coulomb*
(c) 1.6×10^{-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**

(a) 6×10^{23} (b) 6×10^{-23}
(c) 96500 (d) 8×10^{19}

43. Coulomb is equal to **[Orissa JEE 2002]**

(a) *ampere* × *second* (b) *ampere* × *minute*
(c) *watt* × *second* (d) *volt* × *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**[EAMCE**

(a) 1.6×10^{-19} (b) 6×10^{-35}
(c) 6×10^{-16} (d) 6×10^{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 deposite on cathode is

(a) $9.65\times10^7 C$ (b) $9.65\times10^4 C$

(c) $9.65\times10^3 C$ (d) $9.65\times10^5 C$

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 : 1mole

(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*

70. 2.5 F of electricity are passed through a $CuSO_4$ solution. The number of *gm* equivalent of 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 $AgNO_3$ 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 H_2O the required quantity of coulomb would be **[RPET 1999]**

(a) 1.93×10^5 (b) 9.6×10^4
(c) 1.8 (d) 3.2

74. A current of 0.25*A* is passed through $CuSO_4$ solution placed in voltameter for 45 *minutes*. The amount of Cu deposited on cathode is (At weight of $Cu=63.6$)

[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 $112\ cm^3$ 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 $NaCl$ (b) 2 N $NaCl$
(c) 0.1 N $NaCl$ (d) 1 N $NaCl$

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 1*F* of electricity is passed through acidulated water, O_2 evolved is

(a) $11.2\ dm^3$ (b) $5.6\ dm^3$
(c) $22.4\ dm^3$ (d) $1.0\ dm^3$

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 $CaCl_2$ by a current 0.25 *mA* following for 60 seconds **[BHU 2004]**

(a) 4.68×10^{18} (b) 4.68×10^{15}
(c) 4.68×10^{12} (d) 4.68×10^{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 50*g* and 1.0*litre*, 1.0*M*, $CuSO_4$ 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 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 cm^3 (b) 550 cm^3

(c) 5.6 cm^3 (d) 11.2 cm^3

(e) 22.4 cm^3

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

Conductor and conductance

1. Which one of the following statements is *correct*

[MP PET 1997]

(a) The oxidation number of oxygen in KO_2 is zero

(b) The specific conductance of an electrolyte solution decreases with increase in dilution

(c) Sn^{2+} oxidises Fe^{3+}

(d) $Zn/ZnSO_4$ is a reference electrode

2. In infinite dilutions, the equivalent conductances of Ba^{2+} and Cl^- are 127 and 76 $ohm^{-1}cm^{-1}\ eqvt^{-1}$. The equivalent conductivity of $BaCl_2$ at indefinite dilution is **[CBSE 2000]**

(a) 101.5 (b) 139.5

(c) 203.5 (d) 279.5

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

[Kerala (Med.) 2003]

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

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

(e) 63.0 $ohm^{-1}cm^2mole^{-1}$

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) Depend upon density of electrolytes itself **[JIPMER 2002]**

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 2**

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

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

7. Conductivity (unit Siemen's) is directly proportional to area of the vessel and the concentration of the solution in it and is inversely proportional to the length of the vessel then the unit of the constant of proportionality is **[AIEEE 2002]**

(a) $Sm\ mol^{-1}$ (b) $Sm^2\ mol^{-1}$

(c) $S^{-2}m^2\ mol$ (d) $S^2m^2\ mol^{-2}$

8. If one end of a piece of a metal is heated, the other end becomes hot after some time. This is due to

[CBSE PMT 1995]

(a) Energised electrons moving to the other part of the metal

(b) Resistance of the metal

(c) Mobility of atoms in the metal

(d) Minor perturbation in the energy of atoms

9. Conductivity of a solution is directly proportional to

[KCET 1984]

(a) Dilution (b) Number of ions

(c) Current density (d) Volume of the solution

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

[MP PMT 1996]

(a) Ionic attraction

(b) Molecular attraction

(c) Degree of association of the electrolyte

(d) Degree of ionisation of the electrolyte

11. Which of the following conducts electricity **[AFMC 1995]**

(a) Fused $NaCl$ (b) CO_2

(c) Br_2 (d) 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) $ohm\ cm$

(b) $ohm^{-1} cm^2 (gm\ equivalent)^{-1}$

(c) $ohm\ cm^2 (gm\ equivalent)$

(d) $S\ cm^{-2}$

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) $NaCl$ (aqueous) (b) $NaCl$ (solid)

(c) $NaCl$ (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) CH_3COONa (b) C_2H_5OH

(c) $NaCl$ (d) 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) $\Omega^{-1} cm^{-2} mol^{-1}$ (b) $\Omega\ cm^{-2} mol^{-1}$

(c) $\Omega^{-1} cm^2 mol^{-1}$ (d) $\Omega\ cm^2 mol$

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 $l/a = 0.5 cm^{-1}, R = 50 ohm, N = 1.0$. The equivalent conductance of the electrolytic cell is **[Orissa J**

(a) $10\ ohm^{-1} cm^2 gm\ eq^{-1}$ (b) $20\ ohm^{-1} cm^2 gm\ eq^{-1}$

(c) $300\ ohm^{-1} cm^2 gmeq^{-1}$ (d) $100\ ohm^{-1} cm^2\ gmeq^{-1}$

26. If equivalent conductance of $1M$ benzoic acid is $12.8 ohm^{-1} cm^2$ and if the conductance of benzoate ion and H^+ ion are 42 and $288.42\ ohm^{-1} cm^2$ respectively. its degree of dissociation is

(a) 39% (b) 3.9%

(c) 0.35% (d) 0.039%

27. The unit 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, ΔG is related with electromotive force (E), as $\Delta G = -nFE$
(c) Nernst equation for single electrode potential is $E = E^o - \frac{RT}{nF} \ln a_{M^{n+}}$
(d) The efficiency of a hydrogen oxygen fuel cell is 23%

4. The specific conductance of a 0.1 *N KCl* solution at $23^o C$ is $0.012\ ohm^{-1} cm^{-1}$. 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) $0.142\ cm^{-1}$ (b) $0.66\ cm^{-1}$
(c) $0.918\ cm^{-1}$ (d) $1.12\ cm^{-1}$

5. Which of the following reactions occurs at the cathode of a common dry cell
(a) $Mn \rightarrow Mn^{2+} + 2e^-$
(b) $2MnO_2 + Zn^{2+} + 2e^- \rightarrow ZnMn_2O_4$
(c) $2ZnO_2 + Mn^{2+} + 2e^- \rightarrow MnZn_2O_4$
(d) $Zn \rightarrow Zn^{2+} + 2e^-$

6. In $Cu - Zn$ cell **[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]**
[NCERT 1978]
(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) $Zn \mid Zn^{2+} \mid\mid Cu^{2+} \mid Cu$ (b) $Cu \mid Cu^{2+} \mid\mid Zn^{2+} \mid Zn$
(c) $Cu \mid Zn^{2+} \mid\mid Zn \mid Cu^{2+}$ (d) $Cu^{2+} \mid Zn \mid\mid Zn^{2+} \mid Cu$

15. Which one is not called a anode reaction from the following
(a) $Cl^- \rightarrow \frac{1}{2} Cl_2 + e^-$ (b) $Cu \rightarrow Cu^{++} + 2e^-$
(c) $Hg^+ \rightarrow Hg^{++} + e^-$ (d) $Zn^{2+} + 2e^- \rightarrow Zn$

16. A cell from the following which converts electrical energy into chemical energy
(a) Dry cell (b) Electrochemical cell
(c) Electrolytic cell (d) None of these

17. In the cell $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$, the negative electrode is

[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) $E^0_{cell} = \dfrac{n}{0.059}\log K_c$ (b) $E^0_{cell} = \dfrac{0.059}{n}\log K_c$

(c) $E^0_{cell} = 0.059\, n \log K_c$ (d) $E^0_{cell} = \dfrac{\log K_c}{n}$

22. Consider the Galvanic cell $Zn^{\ominus}|\ ZnSO_4\|\ \ CuSO_4|\ Cu^{\oplus}$ the reaction at cathode is **[AMU 2000]**

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

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

(c) $Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$

(d) $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$

23. The cell reaction $Cu + 2Ag^+ \rightarrow Cu^{+2} + Ag$ is best represented by **[AMU 2000]**

(a) $Cu_{(s)}|\ Cu^{+2}{}_{(aq)}\|\ \ Ag^+{}_{(aq)}|\ Ag_{(s)}$

(b) $Pt|\ Cu^{+2}\|\ \ Ag^+{}_{(aq)}|\ Ag_{(s)}$

(c) $Cu^{+2}|\ Cu\|\ \ Pt|\ Ag$

(d) None of the above representations

24. $\underset{\text{(anode)}}{Zn_{(s)}}|\ Zn^{2+}{}_{(aq)}||\ \underset{\text{(cathode)}}{Cu^{2+}{}_{(aq)}}|\ Cu_{(s)}$ 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 $ohm^{-1}cm^{-1}$ and conductivity is 0.04 ohm^{-1}. The cell constant would be **[RPET 1999]**

(a) 1 cm^{-1} (b) 0 cm^{-1}

(c) 5 cm^{-1} (d) 0.2 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]** **[MP PET 2002]**

(a) 4 on the left (b) 3 on the right

(c) 2 on the left (d) 2 on the right

29. Which one of the following statement is true for a electrochemical cell **[Pb. PMT 1999; KCET 1999]**

(a) H_2 is cathode and Cu is anode

(b) H_2 is anode and Cu is cathode

(c) Reduction occurs at H_2 electrode

(d) Oxidation occurs at Cu electrode

30. In the reaction

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

The reduction half-cell reaction is

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

(c) $Ag^+ + e^- \rightarrow Ag$ (d) $Ag - e^- \rightarrow Ag^+$

31. Which of the following statements about galvanic cell is incorrect **[JIPMER 1997]**

(a) Anode is positive

(b) Oxidation occurs at the electrode with lower reduction potential

(c) Cathode is positive

(d) Reduction occurs at cathode

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

(a) $201.28\, ohm^{-1}\, cm^2\, mol^{-1}$

(b) $390.71\, ohm^{-1} cm^2\, mol^{-1}$

(c) $698.28\, ohm^{-1}\, cm^2\, mol^{-1}$

(d) $540.48\, ohm^{-1} cm^2\, mol^{-1}$

33. The electrodes of a conductivity cell are $3\,cm$ apart and have a cross-sectional area of $4\,cm^2$. The cell constant of the cell (in cm^{-1}) is

(a) 4×3 (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) $Pb(s) + SO_2 + O_2 \rightarrow PbSO_4(s)$

(b) $Pb(s) + SO_4^{2-}(aq) \rightleftharpoons PbSO_4(s) + 2e^-$

(c) $PbO_2(s) + 4H^+(aq) + 2e^- + SO_4^{2-}(aq) \rightleftharpoons PbSO_4(s) + 2H_2O$

(d) $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

35. The unit of cell constant is **[MP PET 1996]**

(a) $ohm^{-1}\,cm^{-1}$ (b) $ohm\,cm$

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

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

[MP PET 1996]

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

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

(c) $Mn^{2+} + 2e^- \rightarrow Mn(s)$

(d) $Mn(s) \rightarrow Mn^+ + e^- + 1.5V$

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

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

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

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

39. A depolarizer used in dry cell batteries is **[NCERT 1981]**

(a) Ammonium chloride (b) Manganese dioxide

(c) Potassium hydroxide (d) Sodium phosphate

40. When a lead storage battery is discharged

[IIT 1987; MP PMT 2004]

(a) SO_2 is evolved

(b) Lead sulphate is consumed

(c) Lead is formed

(d) Sulphuric acid is consumed

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 $Mg > Al > Zn > Cu > Ag$. What will happen, if a copper spoon is used to stir a solution of aluminium nitrate

[NCERT 1977]

(a) The spoon will get coated with Al

(b) An alloy of Cu and Al is formed

(c) The solution becomes blue

(d) There is no reaction

43. In a electrochemical cell **[AFMC 1989]**

(a) Potential energy changes into kinetic energy

(b) Kinetic energy changes into potential energy

(c) Chemical energy changes into electrical energy

(d) Electrical energy changes into chemical energy

44. In galvanic cell, the salt bridge is used to **[MP PMT 2002]**

(a) Complete the circuit

(b) Reduce the electric resistance in the cell

(c) Separate cathode from anode

(d) Carry salts for the chemical reaction

45. If a strip of 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) CH_3COOK (b) KCl

(c) NH_4NO_3 (d) KNO_3

47. The reference electrode is made by using **[MP PMT 2002]**

(a) $ZnCl_2$ (b) $CuSO_4$

(c) $HgCl_2$ (d) Hg_2Cl_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. $\lambda_{ClCH_2COONa} = 224\ ohm^{-1} cm^2 gmeq^{-1}$,

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

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

What is the value of λ_{ClCH_2COOH}

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

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

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

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

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

(a) 0.815 (b) 1.016

(c) 1.106 (d) 2.016

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) $Pt|\ H_2(g), 1bar\ |\ 1M\ KCl(aq)|Ag\ Cl(s)|\ Ag(s)$

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

[JEE Orissa 2004]

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

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

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

(a) Increase of $59\,mV$

(b) Decrease of $59\,mV$

[AIIMS 2004]

(c) Increase of $29.5\,mV$

(d) Decrease of $29.5\,mV$

59. If hydrogen electrode dipped in 2 solution of $pH = 3$ and $pH = 6$ and salt bridge is connected the e.m.f. of resulting cell is **[DPMT 2005]**

(a) 0.177 V (b) 0.3 V

(c) 0.052 V (d) 0.104 V

60. The tendency of an electrode to lose electrons is known as

[J & K 2005]

(a) Electrode potential (b) Reduction potential

(c) Oxidation potential (d) e.m.f.

61. When electric current is supplied through an ionic hydride of fused state, then **[Kerala CET 2005]**

(a) Hydrogen is obtained at anode

(b) Hydrogen is obtained at cathode

(c) No change

(d) Hydride ion moves towards cathode

(e) hydride ion present in solution

Electrode potential, E_{Cell}, Nernt equation and ECS

1. The hydrogen electrode is dipped in a solution of $pH = 3$ at 25^oC. The potential of the cell would be (the value of $2.303\,RT/F$ is 0.059 V) **[KCET 1993,2005]**

(a) 0.177 V (b) – 0.177 V

(c) 0.087 V (d) 0.059 V

2. The standard electrode potentials of Zn^{2+}/Zn and Ag^+/Ag are $-0.763\,V$ and $+0.799\,V$ respectively. **[Pb.PMT 2004]** The standard potential of the cell is **[KCET 1993]**

(a) 1.56 V (b) 0.036 V

(c) – 1.562 V (d) 0.799 V

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

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

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

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

[AIIMS 2005]

$Fe^{3+}(aq) + e \rightleftharpoons Fe^{2+}(aq)$;   0.770

Which is the strongest reducing agent

[IIT 1981; MP PET/PMT 1988; MP PMT 1989; MH CET 2001]

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

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

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

[CPMT 1999]

(a) > copper   (b) < copper

(c) > sulphate   (d) < sulphate

5. Which of the following metal does not react with the solution of copper sulphate **[CPMT 1999]**

(a) Mg (b) Fe

(c) Zn (d) Ag

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

$A = -0.250\ V$ $B = -0.136\ V$

$C = -0.126\ V$ $D = -0.402\ V$

The element that displaces A from its compounds is

(a) B   (b) C

(c) D   (d) None of these

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

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

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

Which of the following reactions actually take place

[NCERT 1983, 84; KCET 2003]

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

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

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

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

9. Beryllium is placed above magnesium in the second group. Beryllium dust, therefore when added to $MgCl_2$ 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^o) and concentration of ions in solution is

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

(c) Ohm's equation   (d) Faraday's equation

11. The correct representation of Nernst's equation is

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

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

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

(d) None of the above

12. Standard electrode potential of NHE at 298 K is

(a) 0.05 V (b) 0.1 V

(c) 0.00 V (d) 0.11 V

13. When a copper wire is placed in a solution of $AgNO_3$, 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 $M^{n+}_{(aq)} + ne^- \rightarrow M_{(s)}$. The standard reduction potential values of the elements M_1, M_2 and M_3 are $-0.34\,V, -3.05\ V$ and $-1.66\ V$ 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. $E^0 = \frac{RT}{nF}$ In K_{eq}. This is called

[CPMT 1988; MP PET 2000]

(a) Gibb's equation

(b) Gibb's–Helmholtz equation

(c) Nernst's equation

(d) Vander Waal's equation

16. Four alkali metals A, B, C and D are having respectively standard electrode potential as –

3.05,–1.66,–0.40 and 0.80. Which one will be 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^oC, the standard electrode potentials are A^{2+}/A =–0.76 volts, B^{2+}/B = + 0.34 volts **[KCET 1992]**

(a) A will gradually dissolve
(b) B will deposit on A
(c) No reaction will occur
(d) Water will decompose into H_2 and O_2

20. The reaction $Zn^{2+} + 2e^- \rightarrow Zn$ has a standard potential of $-0.76\ V$. This means **[KCET 1992]**

(a) Zn can't replace hydrogen from acids
(b) Zn is a reducing agent
(c) Zn is a oxidising agent
(d) Zn^{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^o for $Ni^{2+}|\ Ni$ is -0.25 V; E^o for $Au^{+3}|\ Au$ is 1.50 V) is **[MP PET 1993; MP PMT 2000]**

(a) + 1.25 V
(b) -1.75 V
(c) + 1.75 V
(d) + 4.0 V

25. Oxidation and reduction take place in a cell, then its electromotive force will be **[RPET 1999]**

(a) Positive
(b) Negative
(c) Zero
(d) Stable

26. For a spontaneous reaction the ΔG, equilibrium constant (K) and E^o_{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^o 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 $Fe^{2+} + Sn \to Fe + Sn^{2+}$, the standard EMF is **[IIT Screening 1990; MP PMT 2003]**

(a) + 0.30 V (b) – 0.58 V

(c) + 0.58 V (d) – 0.30 V

34. Electrode potential of Zn^{2+}/Zn is $-0.76\ V$ and that of Cu^{2+}/Cu is $+0.34\,V$. 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 V (b) 0.42 V

(c) – 1.1V (d) – 0.42 V

35. *EMF* of a cell whose half cells are given below is

$Mg^{2+} + 2e^- \to Mg(s);\ E = -2.37\ V$

$Cu^{2+} + 2e^- \to Cu(s);\ \ E = +0.33\ V$

[EAMCET 1987; MP PET 1994; Pb. PMT 2000]

(a) – 2.03 V (b) 1.36 V

(c) 2.7 V (d) 2.03 V

36. A cell constructed by coupling a standard copper electrode and a standard magnesium electrode has *emf* of 2.7 *volts*. If the standard reduction potential of copper electrode is + 0.34 *volt* that of magnesium electrode is **[KCET 1989]**

(a) + 3.04 *volts* (b) – 3.04 *volts*

(c) + 2.36 *volts* (d) – 2.36 *volts*

37. When $E^o_{Ag^+/Ag} = 0.8$ *volt* and $E^o_{Zn^{2+}/Zn} = -0.76$ *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 *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

$Fe^{++} + 2e^- \to Fe;\ \ E^o = -0.440\ \ V$

$Fe^{+++} + 3e^- \to Fe;\ \ E^o = -0.036\ \ V$

The standard electrode potential (E^o) for $Fe^{+++} + e^- \to Fe^{++}$ is **[AIIMS 1982]**

(a) – 0.476 V (b) – 0.404 V

(c) + 0.404 V (d) + 0.771 V

40. Reduction potential of four elements P, Q, R, S is $-2.90, +0.34, +1.20$ and -0.76. Reactivity decreases in the order **[MP PET 1989; UPSEAT 2001]**

(a) $P > Q > R > S$ (b) $Q > P > R > S$

(c) $R > Q > S > P$ (d) $P > S > Q > R$

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 $+0.80\,V$ and $+0.34\,V$ respectively. these electrodes are joint together by salt bridge if **[AMU 2002]**

(a) Copper electrode is work like cathode, then E^o_{cell} is $+0.45\,V$

(b) Silver electrode is work like anode then E^o_{cell} is $-0.34\,V$

(c) Copper electrode is work like anode then E^o_{cell} is $+0.46\,V$

(d) Silver electrode is work like cathode then E^o_{cell} is $-0.34\,V$

(e) Silver electrode is work like anode then E^o_{cell} will be $+1.14\,V$

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

$H_2(g) 1\,atm |\ H^+(1M) ||\ Cu^{2+}(1M) |\ Cu(s)$

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

(c) $E_{\text{cathode}} - E_{\text{anode}}$ (d) $E_{\text{left}} - E_{\text{right}}$

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 1*M* silver nitrate solution and crystals of silver metal get deposited.

(ii) Silver metal does not react with 1 *M* zinc nitrate solution

(iii) Zinc metal dissolves in 1*M* copper sulphate solution and copper metal gets deposited

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

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

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

50. Standard electrode potentials of Zn and Fe are known to be (i) $-0.76\,V$ and (ii) $-0.44\,V$ respectively. How does it explain that galvanization prevents rusting of iron while zinc slowly dissolves away

(a) Since (i) is less than (ii), zinc becomes the cathode and iron the anode

(b) Since (i) is less than (ii), zinc becomes the anode and iron the cathode

(c) Since (i) is more than (ii), zinc becomes the anode and iron the cathode

(d) Since (i) is more than (ii), zinc becomes the cathode and iron the anode

51. Amongst the following electrodes the one with zero electrode potential is

(a) Calomel electrode

(b) Standard hydrogen electrode

(c) Glass electrode

(d) Gas electrode

52. Which of the following is correct expression for electrode potential of a cell

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

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

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

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

53. Calculate standard free energy change for the reaction $\frac{1}{2}Cu(s) + \frac{1}{2}Cl_2(g) \rightleftharpoons \frac{1}{2}Cu^{2+} + Cl^-$ taking place at 25^oC 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 :

$Fe^{3+}(aq) + e^- \rightarrow Fe^{-1}(aq); E^o = +0.77\,V$

$Al^{3+}(aq) + 3e^- \rightarrow Al(s); E^o = -1.66\,V$

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq); E^o = +1.08\,V$

Based on the data given above, reducing power of Fe^{2+}, Al and Br^- will increase in the order**[Pb. PMT 1998]**

(a) $Br^- < Fe^{2+} < Al$ (b) $Fe^{2+} < Al < Br^-$

(c) $Al < Br^- < Fe^{2+}$ (d) $Al < Fe^{2+} < Br^-$

58. The standard electrode potential (E^o) for OCl^-/Cl^- and $Cl^-/\frac{1}{2}Cl_2$ respectively are $0.94\,V$ and $-1.36\,V$. The E^o value for $OCl^-/\frac{1}{2}Cl_2$ will be**[KCET 19** **[MP PMT 1997]**

(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]** **[MP PMT 1997]**

(a) It is easily oxidised

(b) It is easily reduced

(c) It acts as oxidising agent

(d) It has redox nature

60. One of the following is false for Hg **[BHU 1998]**

(a) It can evolve hydrogen from H_2S

(b) It is a metal

(c) It has high specific heat

(d) It is less reactive than hydrogen

61. E^o for the cell $Zn|\ Zn^{2+}(aq)\|\ Cu^{2+}(aq)|\ Cu$ is $1.10\,V$ at 25^oC, 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 $25^o C$ of $Li^+ | Li, Ba^{2+} | Ba, Na^+ | Na$ and $Mg^{2+} | Mg$ 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) Na^+ (b) Li^+

(c) Ba^{2+} (d) Mg^{2+}

63. Which of the following displaces Br_2 from an aqueous solution containing bromide ions

[CBSE PMT 1994; JIPMER (Med.) 2002]

(a) Cl_2 (b) Cl^-

(c) I_2 (d) I_3^-

64. For the cell reaction

$Cu^{2+}(C_1 aq) + Zn(s) = Zn^{2+}(C_2 aq) + Cu(s)$

of an electrochemical cell, the change in free energy at a given temperature is a function of

(a) ln (C_1) (b) ln (C_2)

(c) ln $(C_1 + C_2)$ (d) ln (C_2 / C_1)

65. The e.m.f. of the cell in which the following reaction $Zn(s) + Ni^{2+}(a = 1.0) \rightleftharpoons Zn^{2+}(a = 10) + Ni(s)$ occurs, is found to be $0.5105\ V$ at $298 K$. The standard e.m.f. of the cell is

[Roorkee Qualifying 1998]

(a) 0.5400 (b) 0.4810 *V*

(c) 0.5696 *V* (d) – 0.5105 *V*

66. For the redox reaction

$Zn(s) + Cu^{2+}(0.1M) \rightarrow Zn^{2+}(1M) + Cu(s)$ taking place in a cell, E^o_{cell} is 1.10 *volt*. E_{cell} for the cell will be

$\left(2.303 \dfrac{RT}{F} = 0.0591\right)$ **[AIEEE 2003]**

(a) 2.14 *volt* (b) 1.80 *volt*

(c) 1.07 *volt* (d) 0.82 *volt*

67. The *emf* of a Daniel cell at 298*K* is E_1 $Zn | \underset{(0.01\ M)}{ZnSO_4} \| \underset{(1.0\ M)}{CuSO_4} | Cu$ when the concentration of $ZnSO_4$ is 1.0 *M* and that of $CuSO_4$ is 0.01 *M*, the *emf* changed to E_2. What is the relationship between E_1 and E_2

[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

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

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

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

[MP PMT 2003]

(a) – 1.20 *V* (b) + 0.32 *V*

(c) – 0.32 *V* (d) + 1.20 *V*

69. The E^o for half cells Fe / Fe^{2+} and Cu / Cu^{2+} are – 0.44 *V* and + 0.32 *V* respectively. Then**[MP PMT 2003]**

(a) Cu^{2+} oxidises Fe (b) Cu^{2+} oxidises Fe^{2+}

(c) Cu oxidises Fe^{2+} (d) Cu reduces Fe^{2+}

70. What is E^o for electrode represented by $Pt, O_2(1\,atm) / 2H^+(\mathrm{Im})$ **[JIPMER 1997]**

(a) Unpredictable (b) Zero

(c) 0.018 *V* (d) 0.118 *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 Fe **[Roorkee 1995]**

(a) Ag (b) Hg

(c) Zn (d) Na

73.[CBSE PMT 1998]The standard electrode potential of the half cells are given below

$Zn^{2+} + 2e^- \rightarrow Zn; E = -7.62\ V$,

$Fe^{2+} + 2e^- \rightarrow Fe; E = -7.81\ V$

The *emf* of the cell $Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$ is

[CPMT 2003]

(a) 1.54 *V* (b) – 1.54 *V*

(c) – 0.19 *V* (d) + 0.19 *V*

74. $Zn^{2+} + 2e^- \rightarrow Zn(s); E^o = -0.76$,

$Fe^{3+} + e^- \rightarrow Fe^{2+}; E^o = -0.77$,

$Cr^{3+} + 3e^- \rightarrow Cr; E^o = -0.79$,

$H^+ + 2e^- \rightarrow 1/2H_2; E^o = 0.00$

Strongest reducing agent is

(a) Fe^{2+} (b) Zn

(c) Cr (d) H_2

75. Standard reduction electrode potentials of three metals *A*, *B* and *C* are respectively + 0.5*V*, – 3.0*V* and – 1.2 *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 *V* at $25^o C$. The equilibrium constant of the reaction at $25^o C$ will be

[Roorkee 1999; AIEEE 2003; CBSE PMT 2004]

(a) 1×10^{-10} (b) 29.5×10^{-2}

(c) 10 (d) 1×10^{10}

77. For the electrochemical cell, $M | M^+ \| X^- | X$, $E^o(M^+ / M)$ = 0.44 *V* and $E^o(X / X^-)$ = 0.33 *V*. From this data one can deduce that

[IIT-JEE (Screening) 2000]

(a) $M + X \rightarrow M^+ + X^-$ is the spontaneous reaction

(b) $M^+ + X^- \rightarrow M + X$ is the spontaneous reaction

(c) E_{cell} = 0.77 V

(d) E_{cell} = – 0.77 V

78. The standard potential at $25°C$ for the following half reactions are given against them

$Zn^{2+} + 2e \rightarrow Zn, E^o = -0.762\,V$

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

When zinc dust is added to the solution of $MgCl_2$ **[UPSEAT 2001]**

(a) $ZnCl_2$ is formed

(b) *Zinc* dissolves in the solution

(c) No reaction takes place

(d) Mg is precipitated

79. $KMnO_4$ acts as an oxidising agent in the neutral medium and gets reduced to MnO_2. The equivalent weight of $KMnO_4$ in neutral medium **[AMU 2001]**

(a) *mol. wt*/2 (b) *mol.wt*/3

(c) *mol. wt*/4 (d) *mol .wt*/7

80. Which of the following condition will increase the voltage of the cell, represented by the equation $Cu_{(s)} + 2Ag^+{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2Ag_{(s)}$ **[CBSE PMT 2001]**

(a) Increase in the concentration of Ag^+ ion

(b) Increase in the concentration of Cu^+ ion

(c) Increase in the dimension of silver electrode

(d) Increase in the dimension of copper electrode

81. Which will increase the voltage of the cell $Sn_{(s)} + 2Ag^+{}_{(aq)} \rightarrow Sn^{2+}{}_{(aq)} + 2Ag_{(s)}$ **[DPMT 2001]**

(a) Increase in the concentration of Ag^+ ions

(b) Increase in the concentration of Sn^{2+} ions

(c) Increase in size of the silver rod

(d) None of these

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 *KeV* (b) 1 *KeV*

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

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

(a) $\frac{RT}{f}\log\frac{P_1}{P_2}$ (b) $\frac{RT}{2f}\log\frac{P_1}{P_2}$

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

84. What is the potential of a cell containing two hydrogen electrodes the negative one in contact with $10^{-8}\,M\ H^+$ and positive one in contact with 0.025 $M\ H^+$ **[MP PMT 2000]**

(a) 0.18 V (b) 0.28 V

(c) 0.38 V (d) 0.48 V

85. Will $Fe_{(s)}$ be oxidised to Fe^{2+} by the reaction with 1 M HCl (E^o for Fe/Fe^{2+} = + 0.44 V)**[Pb. PMT 2000]**

(a) Yes (b) No

(c) May be (d) Can't say

86. *EMF* of a cell in terms of reduction potential of its left and right electrodes is

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

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

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

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

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

88. Which of the following has highest electrode potential **[Pb. PMT 2000]**

(a) Li (b) Cu

(c) Au (d) Al

89. The cell reaction of a cell is

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

If the standard reduction potentials of Mg and Cu are -2.37 and $+0.34\,V$ respectively. The *emf* of the cell is **[EAMCET 1995; JIPMER (Med.) 2001; AFMC 2002; CBSE PMT 2002]**

(a) 2.03 V (b) – 2.03 V

(c) + 2.71 V (d) – 2.71 V

90. The element which can displace three other halogens from their compound is **[EAMCET 1998]**

(a) Cl (b) F

(c) Br (d) I

91. Which of the following has been universally accepted as a reference electrode at all temperatures and has been assigned a value of zero volt **[AIIMS 1998]**

(a) Graphite electrode

(b) Copper electrode

(c) Platinum electrode

(d) Standard hydrogen electrode

92. Aluminium is more reactive than Fe. But Al is less easily corroded than iron because **[DCE 1999]**

(a) *Al* is noble metal

(b) Iron forms both mono and divalent ions

(c) Oxygen forms a protective oxide layer

(d) Fe undergoes reaction easily with H_2O

93. Zinc displaces copper from the solution of its salt because

[MP PET 1995]

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

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

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

(d) Zinc is placed higher than copper in electro-chemical series

94. An electrochemical cell is set up as follows

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

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

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

[CBSE PMT 1995]

(a) The pH of 0.1 M HCl and 0.1 M acetic acid is not the same

(b) Acids used in two compartments are different

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

(d) The temperature is constant

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

(Given $E^o_{Cu^{2+}/Cu^+} = 0.15$, $E^o_{Cu^{2+}/Cu} = 0.34\,V$) **[IIT 1995]**

(a) – 0.49 V (b) 0.49 V

(c) – 0.38 V (d) 0.38 V

96. E^o of a cell $aA + bB \rightarrow cC + dD$ is **[CPMT 1997]**

(a) $-\frac{RT}{nF}\log\frac{[C]^c[D]^d}{[A]^a[B]^b}$ (b) $-RT\log\frac{[a]^A[b]^B}{[a]^C[d]^D}$

(c) $-\frac{RT}{nF}\log\frac{[C]^c[d]^D}{[A]^a[B]^b}$ (d) $-\frac{RT}{nF}\log\frac{[C]^c[d]^D}{[a]^A[B]^b}$

97. In the experiment set up for the measurement of EMF of a half cell using a reference electrode and a salt bridge, when the salt bridge is removed, the voltage

[NCERT 1984; CPMT 1985]

(a) Does not change

(b) Decreases to half the value

(c) Increase to maximum

(d) Drops to zero

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 0.01m $ZnSO_4$ solution at 258^oC

$(E^o = 0.763\,V)$

[AIIMS 2000; BHU 2000]

(a) 0.8221 V (b) 8.221 V

(c) 0.5282 V (d) 9.232 V

102. The *emf* of a galvanic cell, with electrode potentials of silver = +0.80V and that of copper = + 0.34 V, is**[AIIMS 19**

(a) – 1.1 V (b) + 1.1 V

(c) + 0.46 V (d) + 0.76 V

103. Copper cannot replace..... from solution**[DPMT 2002]**

(a) Fe (b) Au

(c) Hg (d) Ag

104. The strongest reducing agent of the alkali metal is

[CBSE PMT 2000]

(a) Li (b) Na

(c) K (d) Cs

105. Which of the following is the most electropositive element

[Pb. PMT 2000]

(a) Carbon (b) Calcium

(c) Chlorine (d) Potassium

106. The metal that forms a self protecting film of oxide to prevent corrosion, is **[BHU 1999]**

(a) Cu (b) Al

(c) Na (d) Au

107. In a cell that utilises the reaction $Zn_{(s)} + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_{2(g)}$ addition of H_2SO_4 to cathode compartment, will **[AIEEE 2004]**

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

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

(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, $M|\,M^+\|\ X^-|\,X$, $E°(M^+|\,M)=0.44\ V$ $E°(X|\,X^-)=0.33\ V$ From this data, one can deduce that

(a) $E°_{cell}=-0.77\ V$
(b) $M^++X^-\to M+X$ is the spontaneous reaction
(c) $M+X\to M^++X^-$ is the spontaneous reaction
(d) $E°_{cell}=.77\ V$

109. The standard *e.m.f.* of a call, involving one electron change is found to be $0.591\ V$ at 25°*C*. The equilibrium constant of the reaction is ($F=96{,}500\ C\ mol^{-1}$; $R=8.314\ JK^{-1}mol^{-1}$) **[AIEEE 2004]**

(a) 1.0×10^{10}
(b) 1.0×10^{5}
(c) 1.0×10^{1}
(d) 1.0×10^{30}

110. Standard electrode potential of cell $H_2|\,H^+\|\ Ag^+|\,Ag$ is **[AIEEE 2004]**

(a) 0.8 V
(b) – 0.8 V
(c) – 1.2 V
(d) 1.2 V

111. A galvanic cell with electrode potential of $'A'=+2.23\ V$ and $'B'=-1.43\ V$. The value of $E°_{cell}$ is**[Pb.CET 2003]**

(a) 3.66 V
(b) 0.80 V
(c) – 0.80 V
(d) – 3.66 V

112. The *e.m.f.* of a cell whose half cells are given below is

$Mg^{2+}+2e^-\to Mg(s)\ E°=-2.37\ V$

$Cu^{2+}+2e^-\to Cu(s)\ E°=+0.34\ V$ **[Pb.CET 2001]**

(a) + 1.36 V
(b) + 2.71 V
(c) + 2.17 V
(d) – 3.01 V

113. For the cell reaction, $2Ce^{4+}+Co\to 2Ce^{3+}+Co^{2+}$ $E°_{cell}$ is 1.89 V. If $E°_{Ce^{4+}/Ce^{3+}}$ **[Pb.CET 2000]**

(a) – 1.64 V
(b) + 1.64 V
(c) – 2.08 V
(d) + 2.17 V

114. If the ΔG of a cell reaction $AgCl+e^-\to Ag+Cl^-$ is $-21.20\ KJ$; the standard *e.m.f.*, of cell is **[MP PMT 2004]**

(a) 0.229 V
(b) 0.220 V
(c) – 0.220 V
(d) – 0.110 V

115. The *e.m.f.* of the cell $Ag|\,Ag^+(0.1M)\|\ Ag^+(1M)|\,Ag$ at 298 K is **[DCE 2003]**

(a) 0.0059 V
(b) 0.059 V
(c) 5.9 V
(d) 0.59 V

116. The *e.m.f.* of the cell $Zn|\,Zn^{2+}(0.01M)\|\ Fe^{2+}(0.001M)|\,Fe$ at 298 K is 0.2905 then the value of equilibrium for the cell reaction is **[IIT-JEE Screening 2004]**

(a) $\dfrac{0.32}{e^{0.0295}}$
(b) $\dfrac{0.32}{10^{0.0295}}$
[Pb.CET 2004]
(c) $\dfrac{0.26}{10^{0.0295}}$
(d) $\dfrac{0.32}{10^{0.0591}}$

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 V. The reduction potential of silver electrode is $+0.80\ V$. The reduction potential of aluminium electrode is **[KCET 2004]**

(a) $+1.66\ V$
(b) $-3.26\ V$
(c) $3.26\ V$
(d) $-1.66\ V$

118. Consider the following E^0 values :

$E^0_{Fe^{3+}/Fe^{2+}}=+0.77\ V$

$E^0_{Sn^{2+}/Sn}=-0.14\ V$

Under standard conditions the potential for the reaction $Sn_{(s)}+2Fe^{3+}(aq)\to 2Fe^{2+}(sq)+Sn^{2+}(aq)$ is**[AIEEE 200**

(a) 0.91 V
(b) 1.40 V
(c) 1.68 V
(d) 0.63 V

119. $Cr_2O_7^{2-}+I^-\to I_2+Cr^{3+}$

$E^0_{cell}=0.79\ V$

$E^0_{Cr_2O_7^{2-}}=1.33\ V,\ E^0_{I_2}$ is **[BVP 2004]**

(a) $-0.10\ V$
(b) $+0.18\ V$
(c) $-0.54\ V$
(d) $0.54\ V$

120. $Zn(s)+Cl_2(1\text{ atm})\to \text{Zn}^{2+}+2Cl^-$. E^0_{cell} of the cell is 2.12 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 $E^0_{M^{3+}/M^{2+}}$ values for Cr, Mn, Fe and Co are $-0.41, +1.57, +0.77$ and $+1.97\ V$ 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

$2H^+ + 2e^- + \tfrac{1}{2}O_2 \longrightarrow H_2O(l)$;

$E° = +1.23\ V$

$Fe^{2+} + 2e^- \longrightarrow Fe(s)$; $E° = -0.44\ V$

Calculate ΔG^{o} for the net process **[IIT 2005]**

(a) $-322\ kJ\ mol^{-1}$ (b) $-161\ kJ\ mol^{-1}$

(c) $-152\ kJ\ mol^{-1}$ (d) $-76\ kJ\ mol^{-1}$

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 Cu/Cu^{2+} and Al/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 Cu^{2+}/Cu electrode is 0.34V, what is the electrode potential of 0.01M concentration of Cu^{2+} $(T = 298\ K)$ **[EAMCET 2003]**

(a) 0.399 V (b) 0.281 V

(c) 0.222 V (d) 0.176 V

127. Calculate the electrode potential at $298^{o}K$ for $Zn|\ Zn^{++}$ electrode in which the activity of zinc ions is 0.001 M and $E^{o}{}_{Zn/Zn^{++}}$ 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) $\Delta G^{o} = -nFE^{o}_{cell}$ (b) $\Delta G^{o} = +nFE^{o}_{cell}$

(c) $\Delta G^{o} = -2.303\,RT\,nFE^{o}_{cell}$ (d) $\Delta G^{o} = -nF\log K_{C}$

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

Corrosion

1. Corrosion is basically a **[Kerala (Med.) 2002]**

(a) Altered reaction in presence of H_2O

(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) O_2

(c) Zn (d) H^{+}

3. Which of the following is a highly corrosive salt **[AFMC 2005]**

(a) $FeCl_2$ (b) $PbCl_2$

(c) Hg_2Cl_2 (d) $HgCl_2$

4. Corrosion of iron is essentially an electrochemical phenomenon where the cell reactions are **[KCET 2005]**

(a) Fe is oxidised to Fe^{2+} and dissolved oxygen in water is reduced to $\overset{\text{📫}}{OH}$

(b) Fe is oxidised to Fe^{3+} and H_2O is reduced to O_2^{2-}

(c) Fe is oxidised to Fe^{2+} and H_2O is reduced to O_2^{-}

(d) Fe is oxidised to Fe^{2+} and H_2O is reduced to O_2

Critical Thinking

Objective Questions

1. The limiting molar conductivities $\wedge^{0}$ for *NaCl, KBr* and *KCl* are 126, 152 and 150 $S\ cm^{2}mol^{-1}$ respectively. The $\wedge^{0}$ for *NaBr* is

(a) $278\ S\ cm^{2}mol^{-1}$ (b) $176\ S\ cm^{2}mol^{-1}$

(c) $128\ S\ cm^{2}mol^{-1}$ (d) $302\ S\ cm^{2}mol^{-1}$

2. On the basis of the electrochemical theory of aqueous corrosion, the reaction occurring at the cathode is **[MP PET 1994; UPSEAT 2001]**

(a) $O_{2(g)} + 4H^{+}_{(aq)} + 4e^{-} \rightarrow 2H_2O_{(l)}$

(b) $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^{-}$

(c) $Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + e^{-}$

(d) $H_{2(g)} + 2OH^{-}_{(aq)} \rightarrow 2H_2O_{(l)} + 2e^{-}$

3. The reaction

$\frac{1}{2}H_2(g) + AgCl(s) \rightarrow H^{+}(aq) + Cl^{-}(aq) + Ag(s)$

occurs in the galvanic cell **[IIT 1985; AMU 2002; KCET 2003]**

(a) $Ag/AgCl(s)KCl(\text{soln})\|\ AgNO_3(\text{soln})/Ag$

(b) $Pt/H_2(g)HCl(\text{soln})\|\ AgNO_3(\text{soln})/Ag$

(c) $Pt/H_2(g)HCl(\text{soln})\|\ AgCl(s)/Ag$

(d) $Pt/H_2(g)KCl(\text{soln})\|\ AgCl(s)/Ag$

4. The standard reduction potential E^{o} for the half reactions are as

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

$Fe = Fe^{2+} + 2e^{-};\ E^{o} = +0.41\ V$

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

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

(a) $-0.35\ V$ (b) $+0.35\ V$

(c) $+1.17\ V$ (d) $-1.17\ V$

5. The number of electrons to balance the following equation $NO_3^- + 4H^+ + e^- \rightarrow 2H_2O + NO$ is **[IIT Screening 1991]**

(a) 5 (b) 4

(c) 3 (d) 2

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

[MNR 1994; AMU 1999; UPSEAT 2002]

(a) $1.10\ V$ (b) $0.110\ V$

(c) $-1.10\ V$ (d) $-0.110\ V$

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

(a) Y will oxidize X and not Z

(b) Y will oxidize Z and not X

(c) Y will oxidize both X and Z

(d) Y will reduce both X and Z

8. The oxidation potential of a hydrogen electrode at pH = 10 and $pH_1 = 1$ **[JIPMER 2000]**

(a) 0.059 V (b) 0.59 V

(c) 0.00 V (d) 0.51 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°C to furnish aluminium metal (At. Mass = 27 *amu*; 1 *Faraday* = 96,500 *Coulombs*). The cathode reaction is

$Al^3 + 3e^- \rightarrow Al^\circ$

To prepare 5.12*kg* of aluminium metal by this method would require **[AIEEE 2005]**

(a) $5.49 \times 10^7 C$ of electricity

(b) $1.83 \times 10^7 C$ of electricity

(c) $5.49 \times 10^4 C$ of electricity

(d) $5.49 \times 10^1 C$ of electricity

11.

Electrolyte :	KCl	KNO_3	HCl	$NaOAc$	$NaCl$
$\Lambda^{\infty}(Scm^2mol^-$	149.	145.	426.	91.0	126.

Calculate Λ^{∞}_{HOAc} using appropriate molar conductances of the electrolytes listed above at infinite dilution in H_2O at 25°C **[AIEEE 2005]**

(a) 517.2 (b) 552.7

(c) 390.7 (d) 217.5

12. The mass of carbon anode consumed (giving only carbondioxide) in the production of 270*kg* of aluminium metal from bauxite by the Hall process is **[CBSE PMT 2005]**

(a) 180*kg* (b) 270*kg*

(c) 540*kg* (d) 90*kg*

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

[CBSE PMT 2005]

(a) 22.4 L (b) 44.8 L

(c) 5.6 L (d) 11.2 L

14. What amount of Cl_2 gas liberated at anode, if 1 *amp*. current is passed for 30 *min*. from *NaCl* solution.

[BHU 2005]

(a) 0.66 moles (b) 0.33 moles

(c) 0.66 g (d) 0.33 g

Assertion & Reason

For AIIMS 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 mole of silver is different from that required for 1 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 in 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 uints of resistivity are ohm metre (Ωm) 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 $\lambda^0_{Na^+} + \lambda^0_{Cl^-}$ are molar limiting conductivity of the sodium and chloride ions respectively, then the limiting molar conducting for sodium chloride is given by the equation: $\wedge^0_{NaCl} = \lambda^0_{Na^+} + \lambda^0_{Cl^-}$.

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 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]**

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 $E^o_{Au^{3+}/Au} = 1.50$ and $E^o_{Ni^{2+}/Ni} = 0.25V$.

Reason : Emf of the cell $= E^o_{\text{cathode}} - E^o_{\text{anode}}$.

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 & 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				

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 $(E^o = -0.83V)$. Hence their ions in the aqueous solution cannot be reduced instead water will be reduced

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

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: $2H_2O + 2e^- \to H_2 + 2OH^-$

Anode : $H_2O \to 2H^+ + \frac{1}{2}O_2 + 2e^-$.

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: $2H^+ + 2e \to H_2$,

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

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) $2H^+ + 2e^- \to H_{2(g)}$ at cathode.

23. (b) $\underset{+1}{Na^+} + e^- \to \underset{0}{Na}$, 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: $H_2O + e^- \to \frac{1}{2}H_2 + OH^-$

Anode: $Cl^- \to \frac{1}{2}Cl_2 + e^-$.

26. (b) Because it does not have ions.

27. (b) $NaCl \rightleftharpoons Na^+ + Cl^-$.

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

Anode : $H_2O \to 2H^+ + \frac{1}{2}O_2 + 2e^-$.

29. (c) Electrolysis use for electroplating and electrorefining.

32. (d) The degree of ionization depend upon the nature of the solute the size of the solute molecules and the concentration of the solution.

33. (c) $C_{12}H_{22}O_{11}$ is an non-electrolyte.

34. (b) On electrolysis molten ionic hydride liberate H_2 at the anode.

35. (b) During electrolysis cation discharged at cathode and anion discharged at anode.

36. (a) Calcium is produces when molten anhydrous calcium chloride is electrolysed.

37. (d) All metals conducts heat and electricity.

38. (c) $2Al + dil. H_2SO_4 \to Al_2SO_4 + H_2 \uparrow$.

39. (d) Generally fussed potassium chloride flow the electric conductivity.

41. (a) The substances whose aqueous solutions allow the passage of electric current and are chemically decomposed, are termed electrolytes. Electrolytic substances are classified as strong or weak according to how readily they dissociate into conducting ions. Acetic acid is a weak electrolyte. Glucose, ethanol and urea are non-electrolytes.

Faraday's law of electrolysis

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

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

2. (a) $W_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{108 \times 9.65}{96500}$

$= 1.08 \times 10^{-2} gm = 10.8\, mg$

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

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

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

5. (c) $Al^{3+} + 3e^- \to Al$

$E_{Al} = \frac{27}{3} = 9$

$W_{Al} = E_{Al} \times \text{No. of faradays} = 9 \times 5 = 45\ gm$.

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

8. (b) $\frac{\text{Weight of } Cu}{\text{Weight of } H_2} = \frac{\text{Eq. weight of } Cu}{\text{Eq. weight. of } H}$

$\frac{\text{Weight of } Cu}{0.50} = \frac{63.6/2}{1}$

Weight of $Cu = 15.9$ *gm*.

9. (c) $Cu^{2+} + 2e^- \rightarrow Cu$

2 Faradays will deposit $= 1\, g$ atom of $Cu = 63.5\, g$.

12. (a) At cathode; $Al^{3+} + 3e^- \rightarrow Al$

$E_{Al} = \frac{27}{3} = 9$

$W_{Al} = E_{Al} \times \text{No. of faradays} = 9 \times 0.1 = 0.9\, gm$.

14. (b) $W = zit$; $W = \frac{32.69 \times 5 \times 60 \times 40}{96500} = 4.065\, gm$.

15. (b) $m = Z \times 4 \times 120$; $M = Z \times 6 \times 40$

$\frac{M}{m} = \frac{6 \times 40}{4 \times 120} = \frac{1}{2}$; $M = m/2$.

16. (c) $W_{\text{metal}} = \frac{E \times I \times t}{96500} = \frac{E \times 3 \times 50 \times 60}{96500}$

$E = \frac{96500 \times w}{3 \times 50 \times 60} = \frac{96500 \times 1.8}{3 \times 50 \times 60} = 19.3$.

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

The charged obtained is $3 \times 96500\ C$.

18. (a) Wt.of Ag deposited $= Eq.wt\ of\ Ag = 108\ gm$

Wt.of 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 $CuSO_4$. Cu^{2+} gets discharged at cathode and OH^- at anode. Thus solution becomes acidic due to excess of H^+ and SO_4^{2-} or H_2SO_4.

23. (b) 1 mole of electrons = 1 faraday

$Mg^{++} + 2e^- \rightarrow Mg$; 2 moles of electrons = 2 faraday.

24. (d) $Cu^{++} + 2e^- \rightarrow Cu$

$E_{Cu} = \frac{63.54}{2} = 31.77$

Amount of electricity required to deposit .6354 *gm* of *Cu*

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

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

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

28. (d) $Ca^{++} + 2e^- \rightarrow Ca$

$E_{Ca} = \frac{40}{2} = 20$

$W_{Ca} = E_{Ca} \times$ No. of faradays $= 20 \times 0.04 = 0.8\, gm$.

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 $= \frac{25}{1000} \times 60 = 1.5$

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

∴ 1.5 *C* will deposit $Ca = \frac{1}{2 \times 96500} \times 1.5\, mole$

$= \frac{1}{2 \times 96500} \times 1.5 \times 6.023 \times 10^{23}$ *atom* $= 4.68 \times 10^{18}$.

31. (b) Equivalent of *Cl* deposited

= No. of *Faraday* passed = 0.5

Wt. of $Cl = 0.5 \times Eq.wt. = 0.5 \times 35.5 = 17.75\ gm$.

32. (b) At Andoe *At Anode* $Cl^- \rightarrow \frac{1}{2} Cl_2 + e^-$

$E_{Cl_2} = \frac{35.5 \times 2}{2} = 35.5$

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

34. (b) It is Faraday's law.

35. (b) Equivalent wt. of O_2 = Equivalent wt. of *Cu*

36. (d) $O_2\% = 20\%$

Metal% = 80% $= \frac{80}{20} \times 8 = 32\ g$ of metal.

37. (b) $V = \frac{827 \times 10^3}{4 \times 96500} = 2.14\ V$.

38. (b) $Ag^+ \xrightarrow{+e^-} Ag$, 96500 *C* will liberate silver =108 *gm*. 9650*C* will liberate silver = 10.8 *gm*.

39. (a) One mole of monovalent metal ion means charge of *N* electron i.e. 96500 *C* or 1 *Faraday*.

42. (a) 1 Faraday involves charge of 1 *mole* electrons.

43. (a) *Coulomb* = *ampere* (A) × *second* (S).

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

$E = -\frac{13.6}{1^2} = -13.6\, eV$.

45. (c) $w \propto E$ if *i* and *t* are constant.

47. (d) Charge (Coulombs) pass per second $= 10^{-6}$ number of electrons passed per second

$= \frac{10^{-6}}{1.602 \times 10^{-19}} = 6.24 \times 10^{12}$.

48. (d) At cathode;

$Fe^{2+} + 2e^- \rightarrow Fe$; $Fe^{3+} + 3e^- \rightarrow Fe$

$(E_{Fe})_1 = \frac{\text{Atomic.weight}}{2}$; $(E_{Fe})_2 = \frac{\text{Atomic.weight}}{3}$

Ratio of weight of *Fe* liberated

$= \frac{\text{Atomic weight}}{3} : \frac{\text{Atomic weight}}{2} = 3 : 2$.

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

31.75 *gm* of Cu is equal to 0.5 mole of Cu deposited at cathode on passing 1F of current.

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

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

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

$2F(2\times 96500\ C)$ deposited $Cu=63.5\,g$

∴Hence 965 C will deposited; $Cu=0.3175\ gm$.

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

Wt. of $Cu=0.3177\ gm$.

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

3 *mole* electrons = 3 N electron.

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

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

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

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

At cathode : $Na^+ + e^- \to Na$

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

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

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

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

58. (d) At cathode: $Ag^+ + e^- \to Ag$

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

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

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

59. (d) KI is an electrolyte.

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

4 *gm* = 4 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 × 96500 *coulombs*

= 0.5 × 96500 = 48250 *C*.

64. (a) Equivalent weight of silver = 107.870 *g*.

1 Faraday = 96500 *coulomb*.

67. (a) Equivalent weight and atomic weight of Na metal are the same, so 1*g* atom of Na is deposited by one Faraday of current.

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

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

∴ 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	$\frac{.0625}{.0625}=1$

Emperical 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 integeral no. = 32

81. (c) Given, Current = 241.25 *columb*

1 *coulomb* current will deposite

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

∴ 241.25 current will deposite

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

$=0.27\ gm$ silver.

82. (b) Reaction for electrolysis of water is

$2H_2O \rightleftharpoons 4H^+ + 2O^{2-}$

$2O^{2-} \to O_2+4e^-$

$4e^- + 4H^+ \to 2H_2$

∵ $n=4$ so 4 Faraday charge will liberate

1 mole $=22.4\ dm^3$ oxygen

∴ 1 Faraday charge will liberate $\frac{22.4}{4}=5.6\ dm^3\ O_2$.

83. (a) $Na^+ + e^- \to Na$

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

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

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

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

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

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.

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

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

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

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

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

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

Time (t) = 60 *sec*

Q = $i\ t = 60 \times 0.025 = 1.5$ *coulombs*

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

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

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

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

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

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

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*

$= 6 \times 96500\ coulomb$ *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.

Conductor and Conductance

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

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

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

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.

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 condcutivity $= \frac{1}{\rho M}$

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

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

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

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

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

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

$= 42 + 288.42 = 330.42$

$$\alpha = \frac{\Lambda^c_m}{\Lambda^o_m} = \frac{12.8}{330.42} = 3.9\%$$

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

or *mho*

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^o_{A/A^-} 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) $K = \frac{1}{R} \times$ Cell constant

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

5. (b) In common dry cell.

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

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

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

7. (c) Overall reaction - $2H_{2(g)} + O_{2(g)} \to 2H_2O(l)$.

9. (b) During charging of a lead storage battery, the reaction at the anode and cathode are

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

Cathode:

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

In both the reactions H_2SO_4 is regenerated.

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

13. (a) When platinum electrodes are dipped in dilute solution H_2SO_4 than H_2 is evolved at cathode.

14. (a) Electrode on which oxidation occurs is written on $L.H.S.$ and the other on the $R.H.S.$ as represented by

$Zn|\ Zn^{2+}\|\ Cu^{2+}|\ Cu$.

Reduction

15. (d) $Zn^{2+} + 2e^- \to Zn$. 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) $E^o_{\text{cell}} = \frac{2.303\ RT}{nF}\log K = \frac{0.0591}{n}\log K_c\ at\ 298\,K$.

22. (b) $\underbrace{Cu^{2+} + 2e^- \to Cu}_{\text{Reductio}}$

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

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

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

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) $\wedge^o_m\ (CH_3COOH) =$

$\wedge^o\ (CH_3COONa)\ + \wedge^o (HCl) - \wedge^o\ (NaCl)$

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

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

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

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

40. (d) $Pb + PbO_2 + 2H_2SO_4 \underset{\text{Recharge}}{\overset{\text{Discharge}}{\rightleftharpoons}} 2PbSO_4 + 2H_2O$.

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) $Cu + FeSO_4 \to \text{No reaction}$ Because Cu has $E^o = 0.34$ *volt* and Fe has $E^o = -0.44$ *volt*.

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

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

$2H_{2(g)} + 4OH^-_{(aq)} \to 4H_2O_{(l)} + 4e^-$

$O_{2(g)} + 2H_2O_{(l)} + 4e^- \to 4OH^-_{(aq)}$

Overall reaction $= 2H_{2(g)} + O_{2(g)} \to 2H_2O_{(l)}$

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

49. (c) $Cl\,CH_2COONa + HCl \to ClCH_2COOH + NaCl$

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

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

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

50. (c) In daniel cell copper rod acts as cathode so there cations move towards copper electrode and reduction take place on copper rod.

51. (b) Elements with lower reduction potential act as anode. His placed above Cu in electrochemical series so it has lesser reduction potential and thus act as anode and Cu act as cathode.

52. (d) Fuel cells are more efficient as they are free from pollution and hence they run till the reactants are active. They have longer life than lead storage cells.

53. (c) For gold plating, the used electrolyte is $K[Au(CN)_2]$.

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

55. (c) Cell constant $= \frac{\text{Specific conductivity}}{\text{Observed conductance}}$

$= \frac{0.002765}{1/R} = 0.002765 \times 400 = 1.106$.

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

The activities of solids and liquids are takes as unity and at low concentrations, the activity of a solute is approximated to its molarity.

Th cell reaction will be

$Pt_{(s)}$ | $H_{2(g)}$, 1 bar | $H^{+}_{(aq)}$ 1 M | $AgCl_{(aq)}$ 1 M | $Ag_{(s)}$

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).

$M \rightarrow M^{n+} + ne^{-}$ (oxidation potential)

$M^{n+} + ne^{-} \rightarrow M$ (reduction potential)

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

$MH + H_2O \rightarrow MOH + H_2\uparrow$

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) $E^{o}_{cell} = E^{o}_{cathode} - E^{o}_{anode} = 0.799 - (-0.763) = 1.562\ V$

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 $\left(E^{0}_{Mg^{2+}/Mg} < E_{H_2O/\frac{1}{2}H_2+OH^{-}}\right)$. Instead water would be reduced in preference.

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

8. (a) Reductio

$\overset{o}{Zn}_{(s)} + 2Ag^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2\overset{o}{Ag}_{(s)}$

Oxidation

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 $E^{o}_{Ag^{+}/Ag} > E^{o}_{Cu^{++}/Cu}$ 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 Nerst 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 $E^{o}_{A^{2+}/A} < E^{o}_{B^{2+}/B}$. A has greater tendency to be oxidized.

$A + B^{2+} \rightarrow A^{2+} + B$.

20. (b) Since $E^{o}_{Zn^{++}/Zn}$ 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) $E_{cell} = E^{o}_{Au^{3+}/Au} - E^{o}_{Ni^{2+}/Ni} = 1.50 - (-0.25) = 1.75\ V$.

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) $\overset{\text{Anode}}{A|\ A^{+}(a=1)}\|\ \overset{\text{Cathode}}{B^{+}(a=1)|\ B}$

$EMF = E_{cathode} - E_{anode} = 0.75 - (0.5)$; $EMF = 0.25\,V$.

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) $E^{o}_{Zn^{++}/Zn} < E^{o}_{Fe^{++}/Fe}$, so Zn will reduce Fe^{++} . Zn cannot reduce Mg^{2+} because $E^{o}_{Zn^{++}/Zn} > E^{o}_{Mg^{++}/Mg}$

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

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

$E^o_{cell} = E^o_{\text{cathode}} - E^o_{\text{anode}}$ = – 0.44 – (– 0.14)= – 0.30*V*

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

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

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

$E^o_{\text{cell}} = 2.71 \ V$.

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

$E^o_{cell} = E^o_{Cu^{++}/Cu} - E^o_{Mg^{++}/Mg}$

$2.70 \ V = 0.34 - E^o_{Mg^{++}/Mg}$; $E^o_{Mg^{++}/Mg} = -2.36 \ V$.

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

39. (d) $\Delta G^o = -nE^oF$

$Fe^{2+} + 2e^- \rightarrow Fe$ .....(i)

$\Delta G^o = -2 \times F \times (-0.440 \ V) = 0.880 \ F$

$Fe^{3+} + 3e^- \rightarrow Fe$ .....(ii)

$\Delta G^o = -3 \times F \times (-0.036) = 0.108 \ F$

On subtracting equation (i) from (ii)

$Fe^{3+} + e^- \rightarrow Fe^{2+}$

$\Delta G^o = 0.108 \ F - 0.880 \ F = -0.772 \ F$

E^o for the reaction $= -\dfrac{\Delta G^o}{nF}$

$= -\dfrac{(-0.772 \ F)}{1 \times F} = +0.772 \ V$.

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.

$Cu^{++} + Fe \rightarrow Cu + Fe^{++}$.

42. (c) Cell reaction is

$Cu_{(s)} + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

Two half cell reaction is

$Cu \rightarrow Cu^{2+} + 2e^-$ Oxidation (anode)

$Ag^+ + e^- \rightarrow Ag$ Reduction (cathode)

$E_{Cell} = E_{ox} - E_{\text{Re}\,d} = 0.80 - 0.34 = +0.46 \, V$

43. (a) EMF = [s.r.p. of cathode–s.r.p of anode]

Where s.r.p. = Standard reduction potential

If EMF is positive then the reaction is spontaneous

For *e.g.* in Galvanic cell

(a) EMF = 1.1 *volt*

(b) Cathode is made of copper

(c) Anode is made of Zinc

EMF = 0.34 –(– 0.76) = 1.1 *volt*.

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

Cu is cathode because reduction is takes place.

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

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

53. (a) $\Delta G = -nFE^o$

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

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

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

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) $OCl^- \rightarrow C^-, E^o = 0.94 \, V$

$Cl^- \rightarrow \frac{1}{2} Cl_2 + e^- E^o = -1.36 \ V$

adding the two equations, we get

$OCl^- \rightarrow \frac{1}{2} Cl_2, E^o = 0.94 - 1.36 = -0.42 \ V$.

60. (a) It cannot evolved H_2 from H_2S

$Hg + H_2S \rightarrow \text{No reaction}$.

61. (b) $E^o_{cell} = \dfrac{0.059}{n} \log \ K$

$\log \ K = \dfrac{1.10 \times 2}{0.059} = 37.2881$ or $K = 10^{-37}$.

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

63. (a) According to electrochemical series.

64. (d) $E_{cell} = E^o_{cell} - \dfrac{RT}{nF} \ln \dfrac{C_2}{C_1}$ and $\Delta G = -nF \, E_{\text{cell}}$

hence ΔG is the function of $\ln \left(\dfrac{C_2}{C_1} \right)$.

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

$= 1.10 - 0.0295 \log 10 = 1.07 \, volt$.

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

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

$\therefore E_1 > E_2$.

68. (b)

Oxidation

$Fe^{+2} + Zn \rightarrow Zn^{2+} + Fe$

Reduction

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

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

70. (b) $E^o = 0$ because hydrogen have zero potential.

71. (b) Cell potential of the cell is positive.

72. (a,b) Because these comes after the *Fe* in electrochemical series.

73. (c)

Oxidation

$$Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$$

Reduction

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

$EMF = -0.19\,V$.

74. (c) $Cr^{3+} > Zn^{2+} > H > Fe^{3+}$.

Reducing nature decreasing order.

75. (a) More is reduction potential, more is the power to get itself reduced or lesser is reducing power or greater is oxidising power.

76. (d) $\Delta G = -nFE^o$

$\Delta G = -2.303\,RT \log K$; $nFE^o = 2.303\,RT \log K$

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

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

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

RHS: reduction $X + e^- \rightarrow X^-$ .....(i)

LHS: Oxidation $M \rightarrow M^+ + e^-$ .....(ii)

Add (i) and (ii) $M + X \rightarrow M^+ + X^-$

The cell potential = $-0.11\,V$

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

78. (c) $\overset{0}{Zn} + \overset{+2}{MgCl_2} \rightarrow \overset{+2}{ZnCl_2} + Mg$ No reaction

Oxidation

Reduction

This type of reaction does not occur because $Mg^{2+}E^o = -2.37\,V$ while $Zn^{2+}E^o = -0.76\,V$.

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) $E_{\text{cell}} = E^o_{\text{cell}} + \frac{0.059}{2} \log \frac{(Ag^+)}{(Sn^{2+})}$.

82. (b) The *K.E.* of proton is 1 *KeV*.

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

Cathodic reaction : $2H^+ \rightarrow H_2(P_2)$

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

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

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

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

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

$H^+ (0.025\,M) + e^- \rightarrow \frac{1}{2} H_2 (\text{reduction})$

Cell reaction is :

$H^+ (0.025\,M) \rightarrow H^+ (10^{-8}\,M)$; $E_{\text{cell}} = 0.38\,V$.

85. (a) E^o for $Fe/Fe^{2+} = 0.44$ V .

86. (c) (Reduction potential of cathode) – (reduction potential of anode).

87. (a) The correct decreasing electrode potential order is : K, Ba, Ca, Mg.

89. (c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.34 - (-2.37) = +2.71\,V$.

90. (b) Because flourine is most powerful reducing agent than other halogens.

92. (c) Aluminium forms a protective oxide layer but iron does not.

93. (d) The reduction potential of *Zn* is very higher than *Cu*.

94. (a) The *pH* of 0.1 *M* *HCl* and 0.1 *M* acetic acid is not the same, because *HCl* is a strong acid so its *pH* is less and CH_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.

$Cu^{++} + e^- \rightarrow Cu^+$; $E^o_{Cu^{++}/Cu^+} = 0.15\,V$ .....(i)

$Cu^{++} + 2e^- \rightarrow Cu$; $E^o_{Cu^{++}/Cu} = 0.34\,V$ .....(ii)

Multiplying eq. (i) by 2 we get

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

$\Delta G_1 = -nFE = -2 \times F \times 0.15$

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

$\Delta G_2 = -nFE = -2 \times F \times 0.34$

Subtract the *eq*. (iv) from (iii)

$Cu^{++} + Cu \rightarrow 2Cu^+$

$\Delta G_3 = -nFE = -1 \times F \times E^o$

Also $\Delta G_3 = \Delta G_1 - \Delta G_2$

$-1FE^o = (-2F \times 0.15) - (-2F \times 0.34)$

$E^o = -0.38$

This is the value for the reaction

$Cu^{++} + Cu \rightarrow 2Cu^+$

But the given reaction is just reverse of it

$\therefore E_{\text{cell}}$ for given reaction = + 0.38*V*.

97. (d) It connect two solutions and complete the circuit.

98. (a) Greater the oxidation potential, greater is the reactivity.

99. (b) Electrochemical series compare the relative reactivity of metals.

100. (d) Fuel cells are more efficient, free from pollution and they run till reactants are active.

102. (c) $E^o = E^o_{Ag^{2+}/Ag} + E^o_{Cu/Cu^{2+}} = -0.34 + 0.80 = +0.46\ V$.

103. (a) *Fe* is placed above *Cu* in electrochemical series.

104. (a) Lithium is the strongest reducing agent of the alkali metals.

105. (d) Potassium is more electropositive element, because it is the only alkali metal among the given elements.

106. (b) Aluminium forms a self protecting film of oxide to prevent corrosion.

107. (a) $Zn_{(s)} + 2H^+{}_{(aq)} \rightleftharpoons Zn^{2+}_{(aq)} + H_{2(g)}$

$$E_{Cell} = E^0_{Cell} - \frac{.059}{2}\log\frac{[Zn^{2+}]}{[H^+]^2}$$

When H_2SO_4 is added then $[H^+]$ will increase therefore E_{Cell} will also increases and equilibrium will shift towards right.

108. (b) For $M^+ + X^- \rightarrow M + X$

$$E^0_{Cell} = E^0_{Cathode} + E^0_{Anode} = 0.44 - 0.33 = +0.11\ V$$

Since $E^0_{Cell} = (+)\,0.11\ V$ is positive hence this reaction should be spontaneous.

109. (a) $E_{Cell} = E^0_{Cell} - \dfrac{0.0591}{n}\log K_c$

At 298 K $E_{Cell} = 0$ $O = 0.591 - \dfrac{0.0591}{n}\log K_c$

$$\log K_c = \frac{0.591 \times 1}{0.0591} = 10\ ;\ K_c = \text{Antilog}\,10 = 1 \times 10^{10}.$$

110. (a) $\frac{1}{2}H_2 |\ H^+ |\ \ |\ Ag^+ |\ Ag |$

$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = E^0_{Ag^+/Ag} - E^0_{H^+/\frac{1}{2}H_2}$$

$$(0.80) - (0.0) = 0.80\ V.$$

111. (a) $E_A = 2.23\,V > E_B = 1.43\ V$

So A will act as cathode in galvanic cell. Hence

$$E^0_{Cell} = E_{Cathode} - E_{Anode} = E_A - E_B$$

$$= (2.23) - (-1.43) = 3.66\ V.$$

112. (b) $E^0_{Cu} > E^0_{Mg}$ hence *Cu* acts as cathode and *Mg* acts as anode.

$$E^0_{Cell} = E^0_{Cu} - E^0_{Mg} = (0.34) - (-2.37) = +2.71\ V.$$

113. (b) In this cell *Co* is oxidised and it acts as anode and *Ce* acts as cathode.

$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = 1.89 = E^0_{Cell} - (-0.28)$$

$$E^0_{Cell} = 1.89 - 0.28 = 1.61\ Volts.$$

114. (b) Given: $\Delta G = -21.20\ kJ = 21200\ J$

$\because\ \Delta G = -nFE$

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

115. (b) $Ag |\ Ag^+(.1m) \|\ Ag^+ 1M |\ Ag |$

$$E_{Cell} = \frac{2.303\,RT}{nF}\log\frac{c_1}{c_2} = \frac{0.059}{1}\log\frac{1}{0.1}$$

$$= 0.059\log 10 = 0.059\ Volt.$$

116. (b) For this cell, reaction is: $Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe$

$$E = E^0 - \frac{0.0591}{n}\log\frac{c_1}{c_2}\ ;\ E^0 = E + \frac{0.0591}{n}\log\frac{c_1}{c_2}$$

$$= 0.2905 + \frac{0.0591}{2}\log\frac{10^{-2}}{10^{-3}} = 0.32\ V.$$

$$E^0 = \frac{0.0591}{2}\log K_c\ ;\ \log K_c = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$$

$$\therefore\ K_c = \frac{0.32}{10^{0295}}.$$

117. (d) *Al* displaces *H* from *HCl* but silver cannot it means *Al* is situated above the *Ag* in *ECS*, hence *Al* will acts as anode and *Ag* will act as cathode.

$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = E^0_{Ag^+/Ag} - E^0_{Al^{3+}/Al}$$

$$2.46 = 0.8 - E^0_{Al^{3+}/Al}\ ;\ E^0_{Al} = 0.8 - 2.46 = -1.66\ V.$$

118. (a) For $Sn_{(s)} + 2Fe^{3+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + Sn^{2+}_{(aq)}$

$$E^0_{Cell} = E^0_{Sn/Sn^{2+}} + E^0_{Fe^{3+}/Fe^{2+}} = (0.14) + (0.77)$$

$$E_{Cell} = 0.91\ Volts.$$

119. (d) I^- get oxidised to I_2 hence will form anode and $Cr_2O_7^{2-}$ get reduced to Cr^{3+} hence will form cathode.

$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode}\ ;\ E^0_{Cell} = E_{Cr_2O_7^{-2}} - E^0_{I_2}$$

$$0.79 = 1.33 - E^0_{I_2}\ ;\ E^0_{I_2} = 1.33 - 0.79\ ;\ E^0_{I_2} = 0.54\ V.$$

120. (b) According to nernst's equation

$$E_{Cell} = E^0_{Cell} - \frac{nRT}{F}\log\frac{c_1}{c_2}$$

For $Zn_{(s)} + Cl_{2(1\ atm)} \rightarrow Zn^{2+} + 2Cl^-$

$c_1 = [Zn^{2+}]$ and $c_2 = [Cl^-]$

Hence to increase E, c_1 should be decreased and c_2 should be increased is $[Zn^{2+}]$ should be decreased and Cl should be increased.

121. (c)

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

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

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

$$\underline{2H^+ + 2e^- + ½O_2 \longrightarrow H_2O(l)\ ;\ \Delta G_2^o}$$

$$Fe(s) + 2H^+ + ½O_2 \longrightarrow Fe^{2+} + H_2O\ ; \Delta G_3^o$$

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

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

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

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

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

128. (a) $\Delta G^o = -2.303\,RT \log K_{eq}$ or $\Delta G^o = -nFE^o_{cell}$

129. (a) Any redox reaction would occur spontaneously if the free energy change (ΔG) is negative.

$\Delta G^o = nFE^o$

Where n is the number of electrons involved, F is the value of Faraday and E^o is the cell emf. ΔG^o can be negative if E^o is positive.

Corrosion

2. (d) Rusting of iron is catalysed by $[H^+]$.

3. (d) $HgCl_2$ has corrosive action. It is highly poisonous. It sublimes on heating. It is, therefore, known as corrosive sublimate.

4. (a) $Fe \to Fe^{2+} + 2e$ (anode reaction)

$O_2 + 2H_2O + 4e \to 4OH^-$ (cathode reaction)

The overall reaction is

$2Fe + O_2 + 2H_2O \to 2Fe(OH)_2$

$Fe(OH)_2$ may be dehydrated to iron oxide FeO, or further oxidised to $Fe(OH)_3$ and then dehydrated to iron rust, Fe_2O_3.

Critical Thinking Questions

1. (c) $(126\ scm^2)\wedge^0_{NaCl} = \wedge^0_{Na^+} + \wedge^0_{Cl^-}$ .....(1)

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

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

By equation (1)+(2) – (3)

$\because\ \wedge^0_{NaBr} = \wedge^0_{Na^+} + \wedge^0_{Br^-}$

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

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

$$\underline{2H + \frac{1}{2}O_2 \to H_2O}$$

$$2H^+ + \frac{1}{2}O_2 + 2e^- \to H_2O$$

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

4. (b) In this reaction $\overset{}{Fe^{2+}} + \overset{0}{Zn} \to \overset{2+}{Zn} + \overset{0}{Fe}$

Oxidation

Reduction

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

$EMF = +0.35\ V.$

5. (c) $NO_3^- + 4H^+ + 4e^- \to 2H_2O + NO$. In this equation all the atoms are balanced. For balancing added $3e^-$ to $L.H.S.$ we have, $NO_3^- + 4H^+ + 3e^- \to 2H_2O + NO$.

6. (a) $E_{cell} = E^o_{cell} - \frac{0.059}{2} \log \frac{(Zn^{++})}{(Cu^{++})}$

$= 1.10 - \frac{0.059}{2} \log \frac{0.1}{0.1} = 1.10\ V$.

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

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

8. (b) $E_{OP} = E^o_{OP} - \frac{0.059}{1} \log \frac{[H^+]}{P_{H_2}}$

$\therefore [H^+] = 10^{-10}$; $P_{H_2} = 1\ atm$; $E_{OP} = 0.59\ V$.

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 × 96500 *C*.

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

$\therefore$ 5.12 × 10^3 *gm* of *Al* is obtained by passing a current of $3 \times \frac{96500}{27} \times 5.12 \times 1000$

= 1.83 × 10^7 *C* × 3 = 5.49 × 10^7*C*.

11. (c) $\Lambda^\infty_{HOAC} = \Lambda^\infty_{NaOAC} + \Lambda^\infty_{HCl} - \Lambda^\infty_{NaCl}$

$= 91.0 + 426.2 - 126.5 = 390.7$

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

13. (c) $Eq\ of\ Al = eq\ of\ H_2$

$\frac{4.5}{\frac{27}{3}} = eq\ of\ H_2$; $\frac{4.5}{9} = eq\ of\ H_2$

$2H^+ + 2e^- \to H_2$

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

$0.5 = n_{H_2} \times 2$

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

14. (c) The reaction taking place at anode is

$2Cl^- \to Cl_2 + 2e^-$

1 *mole* 2 × 96500 coulomb

Q = *i* . *t* = 1 × 30 × 60 = 1800 coulomb.

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

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

Assertion & Reason

1. (a) The nature of the cathode can affect the order of discharge of ions.

2. (b) 1 mole of silver = 1*g* equivalent of silver

1 mole of copper = 2*g* 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 ρ is called resistivity. If $l = 1m$ and $A = 1m^2$, then $R = \rho$ *i.e.* Resistance = Resistivity.

14. (a) According to Kohlrausch law, "Limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of the anion and cation of the electrolyte".

15. (c) One Faraday deposite one gram equivalent of the substance.

16. (b) Gold has higher reduction potential than the given metals. Hence $AuCl_3$ will react with these metals.

17. (a) $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

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, ΔG is *–ve* and hence E^0 is positive. $-\Delta G^0 = nFE^0$ cell

24. (a) Zinc metal which has a more negative electrode potential than iron will provide electrons in preference of the iron, and therefore corrode first. Only when all the zinc has been oxidised does the iron start to rust.

25. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

Anode (oxidation) for eq. $Zn \longrightarrow Zn^{2+} + 2e^-$, so

Excess of electrons and hence negatively charged while cathode is positively charged.

26. (a) Both assertion and reason are true and reason is the correct explanation of assertion.

$E^0_{Au^{3+}/Au} - E^0_{Ni/Ni^{2+}} = 1.50 - (-0.25) = 1.75\ V$

27. (c) Assertion is true but reason is false. Ions of inert electrolytes are not involved in any electrochemical change until they react chemically with the electrolytes in the two half-cells.

28. (d) Both assertion and reason are false. Potential difference is the difference between the electrode potential of the two electrodes of the cell when cell is under operation while emf is the potential difference generated by a cell when there is zero electron flow.