CHAPTER

# 8

# Redox Reactions and Electrochemistry

**1.** How long (approximate) should water be electrolysed by passing through 100 amperes current so that the oxygen released can completely burn 27.66 g of diborane? (Atomic weight of B = 10.8 u)
(a) 6.4 hours (b) 0.8 hours
(c) 3.2 hours (d) 1.6 hours *(2018)*

**2.** When an electric current is passed through acidified water, 112 mL of hydrogen gas at N.T.P. was collected at the cathode in 965 seconds. The current passed, in ampere, is
(a) 2.0 (b) 1.0 (c) 0.1 (d) 0.5
*(Online 2018)*

**3.** When 9.65 ampere current was passed for 1.0 hour into nitrobenzene in acidic medium, the amount of *p*-aminophenol produced is
(a) 10.9 g (b) 98.1 g (c) 109.0 g (d) 9.81 g
*(Online 2018)*

**4.** Given :

$$E^\circ_{Cl_2/Cl^-} = 1.36\text{ V},\ E^\circ_{Cr^{3+}/Cr} = -0.74\text{ V}$$

$$E^\circ_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33\text{ V},\ E^\circ_{MnO_4^-/Mn^{2+}} = 1.51\text{ V}$$

Among the following, the strongest reducing agent is
(a) $Cr^{3+}$ (b) $Cl^-$ (c) Cr (d) $Mn^{2+}$
*(2017)*

**5.** Which of the following reactions is an example of a redox reaction?
(a) $XeF_6 + H_2O \rightarrow XeOF_4 + 2HF$
(b) $XeF_6 + 2H_2O \rightarrow XeO_2F_2 + 4HF$
(c) $XeF_4 + O_2F_2 \rightarrow XeF_6 + O_2$
(d) $XeF_2 + PF_5 \rightarrow [XeF]^+ PF_6^-$ *(2017)*

**6.** In which of the following reactions, hydrogen peroxide acts as an oxidizing agent?
(a) $I_2 + H_2O_2 + 2OH^- \rightarrow 2I^- + 2H_2O + O_2$
(b) $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$
(c) $2MnO_4^- + 3H_2O_2 \rightarrow 2MnO_2 + 3O_2 + 2H_2O + 2OH^-$
(d) $HOCl + H_2O_2 \rightarrow H_3O^+ + Cl^- + O_2$ *(Online 2017)*

**7.** What is the standard reduction potential ($E^\circ$) for $Fe^{3+} \rightarrow Fe$? Given that :

$$Fe^{2+} + 2e^- \rightarrow Fe;\ E^\circ_{Fe^{2+}/Fe} = -0.47\text{ V}$$

$$Fe^{3+} + e^- \rightarrow Fe^{2+};\ E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77\text{ V}$$

(a) + 0.057 V (b) + 0.30 V
(c) – 0.30 V (d) – 0.057 V *(Online 2017)*

**8.** To find the standard potential of $M^{3+}/M$ electrode, the following cell is constituted :
$Pt|\ M|\ M^{3+}\ (0.001\text{ mol L}^{-1})\ \|\ Ag^+(0.01\text{ mol L}^{-1})|Ag$
The emf of the cell is found to be 0.421 volt at 298 K. The standard potential of half reaction $M^{3+} + 3e^- \rightarrow M$ at 298 K will be

(Given : $E^\circ_{Ag^+/Ag}$ at 298 K = 0.80 Volt)
(a) 0.32 Volt (b) 0.66 Volt
(c) 0.38 Volt (d) 1.28 Volt *(Online 2017)*

**9.** Galvanization is applying a coating of
(a) Pb (b) Cr (c) Cu (d) Zn
*(2016)*

**10.** What will occur if a block of copper metal is dropped into a beaker containing a solution of 1 M $ZnSO_4$?
(a) The copper metal will dissolve with evolution of oxygen gas.
(b) The copper metal will dissolve with evolution of hydrogen gas.
(c) No reaction will occur.
(d) The copper metal will dissolve and zinc metal will be deposited.
*(Online 2016)*

**11.** Oxidation of succinate ion produces ethylene and carbon dioxide gases. On passing 0.2 Faraday electricity through an aqueous solution of potassium succinate, the total volume of gases (at both cathode and anode) at STP (1 atm and 273 K) is
(a) 8.96 L (b) 4.48 L
(c) 6.72 L (d) 2.24 L *(Online 2016)*

**12.** Identify the correct statement.
(a) Corrosion of iron can be minimized by forming a contact with another metal with a higher reduction potential.
(b) Iron corrodes in oxygen-free water.
(c) Corrosion of iron can be minimized by forming an impermeable barrier at its surface.
(d) Iron corrodes more rapidly in salt water because its electrochemical potential is higher.
*(Online 2016)*

**13.** Two Faradays of electricity are passed through a solution of $CuSO_4$. The mass of copper deposited at the cathode is (at. mass of Cu = 63.5 amu)
(a) 2 g (b) 127 g (c) 0 g (d) 63.5 g
*(2015)*

**14.** A variable, opposite external potential ($E_{ext}$) is applied to the cell : $Zn|Zn^{2+}(1\ M)||Cu^{2+}(1\ M)|Cu$, of potential 1.1 V. When $E_{ext} < 1.1$ V and $E_{ext} > 1.1$ V, respectively electrons flow from

(a) anode to cathode and cathode to anode
(b) cathode to anode and anode to cathode
(c) cathode to anode in both cases
(d) anode to cathode in both cases. *(Online 2015)*

**15.** At 298 K, the standard reduction potentials are 1.51 V for $MnO_4^-|Mn^{2+}$, 1.36 V for $Cl_2|Cl^-$, 1.07 V for $Br_2|Br^-$ and 0.54 V for $I_2|I^-$. At pH = 3, permanganate is expected to oxidize $\left(\frac{oq}{c} = 535: > b\right)$

(a) $Cl^-$, $Br^-$ and $I^-$ (b) $Cl^-$ and $Br^-$
(c) $Br^-$ and $I^-$ (d) $I^-$ only *(Online 2015)*

**16.** In which of the following reactions, $H_2O_2$ acts as a reducing agent?

(1) $H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$
(2) $H_2O_2 - 2e^- \longrightarrow O_2 + 2H^+$
(3) $H_2O_2 + 2e^- \longrightarrow 2OH^-$
(4) $H_2O_2 + 2OH^- - 2e^- \longrightarrow O_2 + 2H_2O$

(a) (2), (4) (b) (1), (2)
(c) (3), (4) (d) (1), (3) *(2014)*

**17.** Given below are the half-cell reactions:

$Mn^{2+} + 2e^- \longrightarrow Mn$; $E° = -1.18$ V
$2(Mn^{3+} + e^- \longrightarrow Mn^{2+})$; $E° = +1.51$ V

The $E°$ for $3Mn^{2+} \longrightarrow Mn + 2Mn^{3+}$ will be

(a) – 0.33 V; the reaction will occur
(b) – 2.69 V; the reaction will not occur
(c) – 2.69 V; the reaction will occur
(d) – 0.33 V; the reaction will not occur. *(2014)*

**18.** The equivalent conductance of NaCl at concentration $C$ and at infinite dilution are $\lambda_C$ and $\lambda_\infty$, respectively. The correct relationship between $\lambda_C$ and $\lambda_\infty$ is given as (where, the constant $B$ is positive)

(a) $\lambda_C = \lambda_\infty + (B)\sqrt{C}$ (b) $\lambda_C = \lambda_\infty + (B)C$
(c) $\lambda_C = \lambda_\infty - (B)C$ (d) $\lambda_C = \lambda_\infty - (B)\sqrt{C}$ *(2014)*

**19.** Resistance of 0.2 M solution of an electrolyte is 50 Ω. The specific conductance of the solution is 1.4 S $m^{-1}$. The resistance of 0.5 M solution of the same electrolyte is 280 Ω. The molar conductivity of 0.5 M solution of the electrolyte in S $m^2$ $mol^{-1}$ is

(a) $5 \times 10^2$ (b) $5 \times 10^{-4}$
(c) $5 \times 10^{-3}$ (d) $5 \times 10^3$ *(2014)*

**20.** Consider the following reaction,

$$xMnO_4^- + yC_2O_4^{2-} + zH^+ \rightarrow xMn^{2+} + 2yCO_2 + \frac{z}{2}H_2O$$

The values of $x$, $y$ and $z$ in the reaction are, respectively

(a) 5, 2 and 8 (b) 5, 2 and 16
(c) 2, 5 and 8 (d) 2, 5 and 16 *(2013)*

**21.** Given

$E°_{Cr^{3+}/Cr} = -0.74$ V; $E°_{MnO_4^-/Mn^{2+}} = 1.51$ V
$E°_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33$ V; $E°_{Cl/Cl^-} = 1.36$ V

Based on the data given above, strongest oxidising agent will be

(a) $MnO_4^-$ (b) $Cl^-$ (c) $Cr^{3+}$ (d) $Mn^{2+}$ *(2013)*

**22.** The standard reduction potentials for $Zn^{2+}/Zn$, $Ni^{2+}/Ni$, and $Fe^{2+}/Fe$ are – 0.76, – 0.23 and – 0.44 V respectively. The reaction $X + Y^{2+} \rightarrow X^{2+} + Y$ will be spontaneous when

(a) $X$ = Ni, $Y$ = Zn (b) $X$ = Fe, $Y$ = Zn
(c) $X$ = Zn, $Y$ = Ni (d) $X$ = Ni, $Y$ = Fe *(2012)*

**23.** The reduction potential of hydrogen half-cell will be negative if

(a) $p(H_2)$ = 1 atm and $[H^+]$ = 2.0 M
(b) $p(H_2)$ = 1 atm and $[H^+]$ = 1.0 M
(c) $p(H_2)$ = 2 atm and $[H^+]$ = 1.0 M
(d) $p(H_2)$ = 2 atm and $[H^+]$ = 2.0 M *(2011)*

**24.** The Gibbs energy for the decomposition of $Al_2O_3$ at 500°C is as follows :

$2/3Al_2O_3 \longrightarrow 4/3Al + O_2$, $\Delta_rG = +966$ kJ $mol^{-1}$.

The potential difference needed for electrolytic reduction of $Al_2O_3$ at 500°C is at least

(a) 5.0 V (b) 4.5 V (c) 3.0 V (d) 2.5 V *(2010)*

**25.** Given : $E°_{Fe^{3+}/Fe} = -0.036$ V, $E°_{Fe^{2+}/Fe} = -0.439$ V. The value of standard electrode potential for the change, $Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$ will be

(a) – 0.072 V (b) 0.385 V
(c) 0.770 V (d) – 0.270 V *(2009)*

**26.** Amount of oxalic acid present in a solution can be determined by its titration with $KMnO_4$ solution in the presence of $H_2SO_4$. The titration gives unsatisfactory result when carried out in the presence of HCl, because HCl

(a) oxidises oxalic acid to carbon dioxide and water
(b) gets oxidised by oxalic acid to chlorine
(c) furnishes $H^+$ ions in addition to those from oxalic acid
(d) reduces permanganate to $Mn^{2+}$. *(2008)*

**27.** Given $E°_{Cr^{3+}/Cr} = -0.72$ V, $E°_{Fe^{2+}/Fe} = -0.42$ V

The potential for the cell

$Cr\ |\ Cr^{3+}(0.1\ M)\ ||\ Fe^{2+}(0.01\ M)\ |\ Fe$ is

(a) –0.26 V (b) 0.26 V (c) 0.339 V (d) –0.339 V *(2008)*

**28.** The cell,

$Zn\ |\ Zn^{2+}\ (1\ M)\ ||\ Cu^{2+}\ (1\ M)\ |\ Cu$ ($E°_{cell}$ = 1.10 V)

was allowed to be completely discharged at 298 K. The relative concentration of $Zn^{2+}$ to $Cu^{2+}$ $\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is

(a) $9.65 \times 10^4$ (b) antilog(24.08)
(c) 37.3 (d) $10^{37.3}$ *(2007)*

**29.** The equivalent conductances of two strong electrolytes at infinite dilution in $H_2O$ (where ions move freely through a solution) at 25°C are given below:

$\Lambda°_{CH_3COONa}$ = 91.0 S $cm^2$/equiv.
$\Lambda°_{HCl}$ = 426.2 S$cm^2$/equiv.

What additional information/quantity one needs to calculate $\Lambda^\circ$ of an aqueous solution of acetic acid?
(a) $\Lambda^\circ$ of chloroacetic acid ($ClCH_2COOH$)
(b) $\Lambda^\circ$ of NaCl (c) $\Lambda^\circ$ of $CH_3COOK$
(d) The limiting equivalent conductance of $H^+$ ($\lambda^\circ_{H^+}$).
*(2007)*

**30.** Resistance of a conductivity cell filled with a solution of an electrolyte of concentration 0.1 M is 100 Ω. The conductivity of this solution is 1.29 S $m^{-1}$. Resistance of the same cell when filled with 0.2 M of the same solution is 520 Ω. The molar conductivity of 0.02 M solution of the electrolyte will be
(a) $124 \times 10^{-4}$ S $m^2$ $mol^{-1}$ (b) $1240 \times 10^{-4}$ S $m^2$ $mol^{-1}$
(c) $1.24 \times 10^{-4}$ S $m^2$ $mol^{-1}$ (d) $12.4 \times 10^{-4}$ S $m^2$ $mol^{-1}$
*(2006)*

**31.** The molar conductivities $\Lambda^\circ_{NaOAc}$ and $\Lambda^\circ_{HCl}$ at infinite dilution in water at 25°C are 91.0 and 426.2 S $cm^2$/mol respectively. To calculate $\Lambda^\circ_{HOAc}$, the additional value required is
(a) $\Lambda^\circ_{H_2O}$ (b) $\Lambda^\circ_{KCl}$ (c) $\Lambda^\circ_{NaOH}$ (d) $\Lambda^\circ_{NaCl}$.
*(2006)*

**32.** Which of the following chemical reactions depicts the oxidising behaviour of $H_2SO_4$?
(a) $2HI + H_2SO_4 \rightarrow I_2 + SO_2 + 2H_2O$
(b) $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2H_2O$
(c) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$
(d) $2PCl_5 + H_2SO_4 \rightarrow 2POCl_3 + 2HCl + SO_2Cl_2$
*(2006)*

**33.**

| Electrolyte | KCl | $KNO_3$ | HCl | NaOAc | NaCl |
|---|---|---|---|---|---|
| (S $cm^2$ $mol^{-1}$) | 149.9 | 145.0 | 426.2 | 91.0 | 126.5 |

Calculate molar conductance of acetic acid using appropriate molar conductances of the electrolytes listed above at infinite dilution in $H_2O$ at 25°C.
(a) 517.2 (b) 552.7 (c) 390.7 (d) 217.5
*(2005)*

**34.** 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

$$Al^{3+} + 3e^- \rightarrow Al^0$$

To prepare 5.12 kg of aluminium metal by this method would require
(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) $5.49 \times 10^{10}$ C of electricity *(2005)*

**35.** 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.
*(2005)*

**36.** The $E^\circ_{M^{3+}/M^{2+}}$ 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) Cr (b) Mn (c) Fe (d) Co
*(2004)*

**37.** In a cell that utilizes the reaction,
$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$
addition of $H_2SO_4$ to cathode compartment, will
(a) lower the $E$ and shift equilibrium to the left
(b) lower the $E$ and shift the equilibrium to the right
(c) increase the $E$ and shift the equilibrium to the right
(d) increase the $E$ and shift the equilibrium to the left.
*(2004)*

**38.** The limiting molar conductivities $\Lambda^\circ$ for NaCl, KBr and KCl are 126, 152 and 150 S $cm^2$ $mol^{-1}$ respectively. The $\Lambda^\circ$ for NaBr is
(a) 128 S $cm^2$ $mol^{-1}$ (b) 176 S $cm^2$ $mol^{-1}$
(c) 278 S $cm^2$ $mol^{-1}$ (d) 302 S $cm^2$ $mol^{-1}$
*(2004)*

**39.** The standard e.m.f. of a cell, 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 $JK^{-1}mol^{-1}$)
(a) $1.0 \times 10^1$ (b) $1.0 \times 10^5$
(c) $1.0 \times 10^{10}$ (d) $1.0 \times 10^{30}$ *(2004)*

**40.** Consider the following $E^\circ$ values.
$E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77$ V ; $E^\circ_{Sn^{2+}/Sn} = -0.14$ V
Under standard conditions the potential for the reaction $Sn_{(s)} + 2Fe^{3+}_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + Sn^{2+}_{(aq)}$ is
(a) 1.68 V (b) 1.40 V (c) 0.91 V (d) 0.63 V
*(2004)*

**41.** In a hydrogen-oxygen fuel cell, combustion of hydrogen occurs to
(a) generate heat
(b) create potential difference between the two electrodes
(c) produce high purity water
(d) remove adsorbed oxygen from electrode surface.
*(2004)*

**42.** Among the properties ($A$) reducing ($B$) oxidising ($C$) complexing, the set of properties shown by $CN^-$ ion towards metal species is
(a) *A, B* (b) *B, C* (c) *C, A* (d) *A, B, C.*
*(2004)*

**43.** Standard reduction electrode potentials of three metals $A$, $B$ and $C$ are +0.5 V, –3.0 V and –1.2 V respectively. The reducing power of these metals are
(a) $B > C > A$ (b) $A > B > C$
(c) $C > B > A$ (d) $A > C > B$ *(2003)*

**44.** For a cell reaction involving a two-electron change, the standard e.m.f. of the cell is found to be 0.295 V at 25°C. The equilibrium constant of the reaction at 25°C will be
(a) $1 \times 10^{-10}$ (b) $29.5 \times 10^{-2}$
(c) 10 (d) $1 \times 10^{10}$ *(2003)*

**45.** For the redox reaction:

$$Zn_{(s)} + Cu^{2+}\,(0.1\ M) \rightarrow Zn^{+}\,(1M) + Cu_{(s)}$$

taking place in a cell, $E^\circ_{cell}$ is 1.10 volt. $E_{cell}$ for the cell will be
$\left(2.303\frac{RT}{F} = 0.0591\right)$
(a) 2.14 V (b) 1.80 V
(c) 1.07 V (d) 0.82 V *(2003)*

**46.** When during electrolysis of a solution of $AgNO_3$, 9650 coulombs of charge pass through the electroplating bath, the mass of silver deposited on the cathode will be

(a) 1.08 g (b) 10.8 g
(c) 21.6 g (d) 108 g *(2003)*

**47.** The heat required to raise the temperature of body by 1ºC is called

(a) specific heat (b) thermal capacity
(c) water equivalent (d) none of these. *(2002)*

**48.** Which of the following reaction is possible at anode?

(a) $2\,Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$
(b) $F_2 \rightarrow 2F^-$
(c) $(1/2)\,O_2 + 2H^+ \rightarrow H_2O$
(d) None of these. *(2002)*

**49.** What will be the emf for the given cell,

$Pt\,|\,H_2\,(P_1)\,|\,H^+_{(aq)}\,||\,H_2\,(P_2)\,|\,Pt$ ?

(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.
*(2002)*

**50.** If $\phi$ denotes reduction potential, then which is true?

(a) $E^\circ_{cell} = \phi_{right} - \phi_{left}$ (b) $E^\circ_{cell} = \phi_{left} + \phi_{right}$
(c) $E^\circ_{cell} = \phi_{left} - \phi_{right}$ (d) $E^\circ_{cell} = -(\phi_{left} + \phi_{right})$
*(2002)*

**51.** Conductivity (unit Siemen's 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

(a) $S\,m\,mol^{-1}$ (b) $S\,m^2\,mol^{-1}$
(c) $S^{-2}m^2\,mol$ (d) $S^2\,m^2\,mol^{-2}$ *(2002)*

**52.** Which of the following is a redox reaction?

(a) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$
(b) $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$
(c) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2NH_4OH$
(d) $Zn + 2AgCN \rightarrow 2Ag + Zn(CN)_2$ *(2002)*

**53.** When $KMnO_4$ acts as an oxidising agent and ultimately forms $[MnO_4]^-$, $MnO_2$, $Mn_2O_3$, $Mn^{2+}$ then the number of electrons transferred in each case respectively is

(a) 4, 3, 1, 5 (b) 1, 5, 3, 7
(c) 1, 3, 4, 5 (d) 3, 5, 7, 1 *(2002)*

**54.** 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) $E = -(E_{right} + E_{left})$
*(2002)*

## ANSWER KEY

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

# Explanations

**1.** **(c) :** $\underset{27.66\text{ g}}{B_2H_6} + \underset{96\text{ g}}{3O_2} \longrightarrow B_2O_3 + 3H_2O$

Thus, for combustion of 27.66 g of $B_2H_6$ oxygen required is 96 g. According to Faraday's law of electrolysis,

$$w = ZIt = \frac{EIt}{96500} \quad \Rightarrow \quad 96\text{g} = \frac{8\times100\times t}{96500}$$

$$t = \frac{96\times96500}{8\times100} = 11{,}580\text{ s} = \frac{11580}{3600} = 3.2\text{h}$$

**2.** **(b) :** $2H^+ + 2e^- \to H_2$ (at cathode)

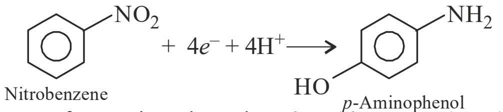
$$w = ZIt = \frac{EIt}{96500}$$

Moles of $H_2$ deposited $= \dfrac{112}{22400}$

Mass of $H_2$ deposited ($w$) = Moles × Molar mass $= \dfrac{112}{22400}\times 2$

Thus, $\dfrac{112}{22400}\times2 = \dfrac{1\times I\times965}{96500} \Rightarrow I = 1$ A

**3.** **(d) :**



Molar mass of $p$-aminophenol $= 6 \times 12 + 7 \times 1 + 14 + 16 = 109$ g mol$^{-1}$

$$\text{Eq. wt.} = \frac{W}{Q}\times96500 = \frac{W}{I\times t}\times96500$$

$$W = \frac{\text{Eq. wt.}\times I\times t}{96500} = \frac{109}{4}\times\frac{9.65\times1\times60\times60}{96500} = 9.81\text{ g}$$

**4.** **(c) :** More negative the $E°$ value of the species, more stronger is the reducing agent.

Since, $E^\circ_{Cr^{3+}/Cr} = -0.74$ V (most negative among the given examples), Cr is the strongest reducing agent.

**5.** **(c) :** (a) $\overset{+6\ -1}{XeF_6} + \overset{+1\ -2}{H_2O} \longrightarrow \overset{+6\ -2\ -1}{XeOF_4} + \overset{+1-1}{2HF}$

No change in oxidation numbers, hence, no redox reaction occurs.

(b) $\overset{+6\ -1}{XeF_6} + \overset{+1\ -2}{2H_2O} \longrightarrow \overset{+6\ -2\ -1}{XeO_2F_2} + \overset{+1-1}{4HF}$

No change in oxidation numbers hence, no redox reaction occurs.

decrease in O.No.
(reduction)

(c) $\overset{+4\ -1}{XeF_4} + \overset{+1\ -1}{O_2F_2} \longrightarrow \overset{+6\ -1}{XeF_6} + \overset{0}{O_2}$

increase in O.No.
(oxidation)

Hence, it is a redox reaction.

(d) $\overset{+2\ -1}{XeF_2} + \overset{+5\ -1}{PF_5^-} \longrightarrow \overset{+2\ -1}{[XeF]^+}\ \overset{+5\ -1}{PF_6^-}$

No change in oxidation numbers, hence, no redox reaction occurs.

**6.** **(b) :** $\overset{+2\ -2}{PbS} + \overset{-1}{4H_2O_2} \longrightarrow \overset{+2\ +6}{PbSO_4} + \overset{-2}{4H_2O}$

gain of electron

Since, $H_2O_2$ gains electron to convert into $H_2O$. Hence, $H_2O_2$ acts as an oxidizing agent.

**7.** **(d) :** For the given reaction :

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$; $E°_{Fe^{3+}/Fe^{2+}} = +0.77$ V; $\Delta G° = -nFE°$

$\Delta G_1^\circ = -0.77\,F$ ...(i)

For the following reaction :

$Fe^{2+} + 2e^- \longrightarrow Fe$; $E°_{Fe^{2+}/Fe} = -0.47$ V;

$\Delta G_2^\circ = -2(-0.47)F = 0.47 \times 2\,F$ ...(ii)

Overall reaction : $Fe^{3+} + 3e^- \longrightarrow Fe$;

$\Delta G_3^\circ = -3FE_3^\circ$ ...(iii)

$\Delta G_3^\circ = \Delta G_1^\circ + \Delta G_2^\circ - 3FE_3^\circ = -0.77\,F + 0.47 \times 2F$

$$E_3^\circ = -\frac{0.17}{3} = -0.057\text{ V}$$

**8.** **(a) :** $E_{\text{cell}} = E^\circ_{\text{cell}} - \dfrac{0.059}{n}\log\dfrac{[M^{3+}]}{[Ag^+]^3}$

$$0.421 = E^\circ_{\text{cell}} - \frac{0.059}{3}\log\frac{0.001}{(0.01)^3}$$

$$E^\circ_{\text{cell}} = 0.421 + \frac{0.059}{3}\times 3 = 0.480\text{ V}$$

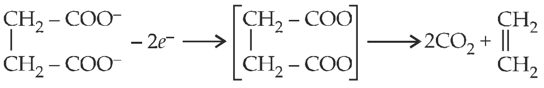
$E^\circ_{\text{cell}} = E^\circ_{Ag+/Ag} - E^\circ_{M^{3+}/M}$

$0.480\text{ V} = 0.8\text{ V} - E^\circ_{M^{3+}/M} \Rightarrow E^\circ_{M^{3+}/M} = 0.32\text{ V}$

**9.** **(d) :** Galvanization is a method of rust prevention by applying zinc coating which acts as a sacrificial metal.

**10.** **(c) :** No reaction will occur. As reduction potential of $Zn^{2+}$ ions to Zn atom is lower than that for $Cu^{2+}$ ions. Hence, Cu metal cannot displace $Zn^{2+}$ ions in $ZnSO_4$ solution.

**11.** **(a) :** At anode :



$$\begin{array}{l}CH_2-COO^-\\ |\\ CH_2-COO^-\end{array} - 2e^- \longrightarrow \underset{\text{(unstable)}}{\left[\begin{array}{l}CH_2-COO\\ |\\ CH_2-COO\end{array}\right]} \longrightarrow 2CO_2 + \begin{array}{c}CH_2\\ \|\\ CH_2\end{array}$$

2 F produce 1 mole of ethene gas and 2 moles of $CO_2$ gas. Therefore, 0.2 F will produce 0.1 mole of ethene gas and 0.2 mole of $CO_2$ gas.

At cathode : $2H^+ + 2e^- \longrightarrow 2[H] \longrightarrow H_2$

2 F produce 1 mole of $H_2$. Therefore, 0.2 F will produce 0.1 mole of $H_2$ gas.

Hence, total no. of moles of gases produced at anode and cathode is 0.2 + 0.1 + 0.1 = 0.4

$$V = \frac{nRT}{P} = \frac{0.4\times0.0821\times273}{1} = 8.96\text{ L}$$

**12.** **(c) :** When an impermeable barrier is formed at the surface of iron then oxygen and moisture cannot attack the metal hence, its corrosion is prevented.

**13.** **(d) :** $Cu^{2+} + 2e^- \to Cu$

2 F charge deposit 1 mol of Cu *i.e.*, 2 F of electricity deposit 63.5 g mass of Cu at the cathode.

**14.** **(d) :** EMF of galvanic cell = 1.1 V

If $E_{\text{ext}}$ < EMF, then electrons flow steadily from anode to cathode while if $E_{\text{ext}}$ > EMF, then electrons flow in reverse direction and cell behaves as an electrolytic cell in which copper electrode behaves as anode and zinc electrode behaves as cathode.

**15. (c)** : $MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$

$$b = b^\circ - \frac{53\!5\!:>}{}x\{ s \frac{gT\ z^{7+}i}{gT\ zl\ _9^-igO^+i^=}$$

Taking $[Mn^{2+}]$ and $[MnO_4^-]$ in standard state *i.e.*, 1 M.

$$b = b^\circ - \frac{53\!5\!:>}{}x\{ s \frac{6}{gO^+i^=} = 63\!:6 - \frac{53\!5\!:>}{} \times = \times | O$$

$$= 63\!:6 - \frac{53\!5\!:>}{:} \times = \times 8 = 63\!77 < b$$

So, $MnO_4^-$ will oxidise only $Br^-$ and $I^-$ as standard reduction potential of $Cl_2/Cl^-$ is 1.36 V which is greater than that of $MnO_4^-/Mn^{2+}$.

**16. (a)** : The reducing agent itself gets oxidised *i.e.* the oxidation number increases.

(1) $\overset{-1}{H_2O_2} + 2H^+ + 2e^- \longrightarrow \overset{-2}{2H_2O}$

(2) $\overset{-1}{H_2O_2} - 2e^- \longrightarrow \overset{0}{O_2} + 2H^+$

(3) $\overset{-1}{H_2O_2} + 2e^- \longrightarrow \overset{-2}{2OH^-}$

(4) $\overset{-1}{H_2O_2} + 2OH^- - 2e^- \longrightarrow \overset{0}{O_2} + 2H_2O$

**17. (b)** : Overall reaction:

$Mn^{2+} + 2e^- \longrightarrow Mn$; $E^\circ = -1.18$ V

$2Mn^{2+} \longrightarrow 2Mn^{3+} + 2e^-$; $E^\circ = -1.51$ V

$3Mn^{2+} \longrightarrow Mn + 2Mn^{3+}$; $E^\circ = -1.18 + (-1.51) = -2.69$ V

As $E^\circ$ is negative, the reaction will not occur.

**18. (d)** : According to Debye—Hückel's theory, for a strong electrolyte (like NaCl), $\lambda_C = \lambda_\infty - (B)\sqrt{C}$

**19. (b)** : Case I : $C = 0.2$ M, $R = 50\ \Omega$, $\kappa = 1.4$ S m$^{-1}$

$$\kappa = \frac{l}{A}\cdot\frac{1}{R} \Rightarrow 1.4 = \frac{l}{A}\cdot\frac{1}{50} \Rightarrow \frac{l}{A} = 1.4 \times 50 = 70\ \text{m}^{-1}$$

Case II : $\frac{l}{A} = 70\ \text{m}^{-1}$, $C = 0.5$ M, $R = 280\ \Omega$,

$$R = \rho\frac{l}{A} \Rightarrow \frac{1}{\rho} = \frac{1}{R}\times\frac{l}{A} \Rightarrow \frac{1}{\rho} = \frac{1}{280}\times 70 \Rightarrow \kappa = \frac{1}{\rho} = 0.25\ \text{S m}^{-1}$$

Now, $\Lambda_m = \kappa \times \frac{1000}{C}$

If molarity is in mol L$^{-1}$, then

$$\Lambda_m\ (\text{S m}^2\ \text{mol}^{-1}) = \frac{\kappa(\text{S m}^{-1})}{1000\ \text{L m}^{-3} \times \text{Molarity (mol L}^{-1})}$$

$$= \frac{0.25\ \text{S m}^{-1}}{1000\ \text{L m}^{-3} \times 0.5\ \text{mol L}^{-1}} = 5 \times 10^{-4}\ \text{S m}^2\ \text{mol}^{-1}$$

**20. (d)** : $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

$\therefore\ x = 2, y = 5, z = 16$

**21. (a)** : Greater the reduction potential of a substance, stronger is the oxidising agent.

$\therefore$ $MnO_4^-$ is the strongest oxidising agent.

**22. (c)** : The elements with high negative value of standard reduction potential are good reducing agents and can be easily oxidised. Thus $X$ should have high negative value of standard potential than $Y$ so that it will be oxidised to $X^{2+}$ by reducing $Y^{2+}$ to $Y$.

$X$ = Zn, $Y$ = Ni ; $Zn + Ni^{2+} \rightarrow Zn^{2+} + Ni$

Alternatively, for a spontaneous reaction $E^\circ$ must be positive.

$E^\circ = E^\circ_{\text{reduced species}} - E^\circ_{\text{oxidised species}} = -0.23 - (-0.76) = +0.53$ V

**23. (c)** : $2H^+_{(aq)} + 2e^- \longrightarrow H_{2\,(g)}$

$$E_{\text{red}} = E^\circ_{\text{red}} - \frac{0.0591}{n}\log\frac{p_{H_2}}{[H^+]^2} = 0 - \frac{0.0591}{2}\log\frac{2}{(1)^2}$$

$E_{\text{red}}$ will only be negative when $p_{H_2} > [H^+]$.

**24. (d)** : The ionic reactions are :

$2/3Al_2^{3+} + 4e^- \longrightarrow 4/3Al$

$2/3O_3^{2-} \longrightarrow O_2 + 4e^-$

Thus, no. of electrons transferred = 4 = $n$

$\Delta G = -nFE = -4 \times 96500 \times E \Rightarrow 966 \times 10^3 = -4 \times 96500 \times E$

$$\Rightarrow\ E = -\frac{966 \times 10^3}{4 \times 96500} = -2.5\ \text{V}$$

**25. (c)** : Given,

$Fe^{3+} + 3e^- \longrightarrow Fe$ ; $E_1^\circ = -0.036$ V

$Fe^{2+} + 2e^- \longrightarrow Fe$ ; $E_2^\circ = -0.439$ V

Required equation is

$Fe^{3+} + e^- \longrightarrow Fe^{2+}$ ; $E_3^\circ = ?$

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

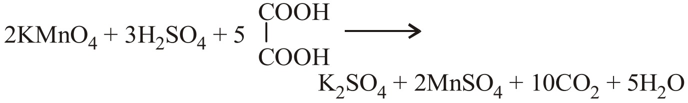
$\therefore\ \Delta G_3^\circ = \Delta G_1^\circ - \Delta G_2^\circ$

$(-n_3FE_3^\circ) = (-n_1FE_1^\circ) - (-n_2FE_2^\circ)$

$E_3^\circ = 3E_1^\circ - 2E_2^\circ = 3 \times (-0.036) - 2 \times (-0.439)$

$E^\circ_3 = -0.108 + 0.878 = 0.77$ V

**26. (d)** : Oxalic acid present in a solution can be determined by its titration with $KMnO_4$ solution in the presence of $H_2SO_4$.



$$2KMnO_4 + 3H_2SO_4 + 5\begin{array}{c}COOH\\|\\COOH\end{array} \longrightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + 5H_2O$$

Titration cannot be done in the presence of HCl because $KMnO_4$ being a strong oxidizing agent oxidises HCl to $Cl_2$ and get itself reduced to $Mn^{2+}$. So actual amount of oxalic acid in solution cannot be determined.

**27. (b)** : $Cr \rightarrow Cr^{3+} + 3e^-$ ; $E^\circ_{\text{red}} = -0.72$ V

$\underline{Fe^{2+} + 2e^- \rightarrow Fe}$ ; $E^\circ_{\text{red}} = -0.42$ V

$2Cr + 3Fe^{2+} \rightarrow 2Cr^{3+} + 3Fe$

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.42 - (-0.72) = 0.3$

According to Nernst equation, $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.059}{n_{\text{cell}}}\log_{10}\frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$

$$E_{\text{cell}} = 0.3 - \frac{0.059}{6}\log_{10}\frac{(0.1)^2}{(0.01)^3} = 0.3 - \frac{0.059}{6}\log_{10}10^4$$

$= 0.3 - 0.039 = 0.261$ V

**28. (d)** : $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.059}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

When the cell is completely discharged, $E_{\text{cell}} = 0$

$$0 = 1.1 - \frac{0.059}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]} \Rightarrow \log\frac{[Zn^{2+}]}{[Cu^{2+}]} = \frac{2\times 1.1}{0.059}$$

or, $\log\frac{Zn^{2+}}{Cu^{2+}} = 37.3 \Rightarrow \frac{Zn^{2+}}{Cu^{2+}} = 10^{37.3}$

**29. (b)** : According to Kohlrausch's law, the molar conductivity at infinite dilution ($\Lambda^\circ$) for weak electrolyte, $CH_3COOH$ is

$$\overset{\circ}{\Lambda}_{CH_3COOH} = \overset{\circ}{\Lambda}_{CH_3COONa} + \overset{\circ}{\Lambda}_{HCl} - \overset{\circ}{\Lambda}_{NaCl}$$

So, for calculating the value of $\overset{\circ}{\Lambda}_{CH_3COOH}$, value of $\overset{\circ}{\Lambda}_{NaCl}$ should also be known.

**30. (a)** : $\kappa = \frac{1}{R}\left(\frac{l}{a}\right) \Rightarrow 1.29 = \frac{1}{100}\left(\frac{l}{a}\right) \Rightarrow l/a = 129\ \text{m}^{-1}$

$R = 520\ \Omega$ for 0.2 M, $C = 0.02$ M

$$\lambda_m = \kappa \times \frac{1000}{\text{molarity}} = \frac{1\times 129}{520}\times\frac{1000}{0.02}\times 10^{-6}\ \text{m}^3 = 124 \times 10^{-4}\ \text{S m}^2\ \text{mol}^{-1}$$

**31. (d)** : $CH_3COONa + HCl \rightarrow CH_3COOH + NaCl$

From the reaction, $\Lambda^\circ_{CH_3COONa} + \Lambda^\circ_{HCl} = \Lambda^\circ_{CH_3COOH} + \Lambda^\circ_{NaCl}$

or, $\Lambda^\circ_{CH_3COOH} = \Lambda^\circ_{CH_3COONa} + \Lambda^\circ_{HCl} - \Lambda^\circ_{NaCl}$

Thus to calculate the value of $\Lambda^\circ_{CH_3COOH}$ one should know the value of $\Lambda^\circ_{NaCl}$ along with $\Lambda^\circ_{CH_3COONa}$ and $\Lambda^\circ_{HCl}$.

**32. (a) :** In the reaction,

$$\underset{+1-1}{2HI} + \underset{+1\ +6\ -2}{H_2SO_4} \rightarrow \underset{0}{I_2} + \underset{+4}{SO_2} + \underset{+1\ -2}{2H_2O}$$

I$^{st}$ half reaction : $\underset{-1}{2HI} \rightarrow \underset{0}{I_2}$

In this reaction oxidation number of I increases by one, thus this is an oxidation reaction and HI behaves as a reducing agent.

II$^{nd}$ half reaction : $\underset{+6}{H_2SO_4} \rightarrow \underset{+4}{SO_2}$

In this reaction oxidation number of S decreases by two, thus this is a reduction reaction and $H_2SO_4$ behaves as oxidising agent.

**33. (c) :** $\Lambda^\infty_{AcOH} = \Lambda^\infty_{AcONa} + \Lambda^\infty_{HCl} - \Lambda^\infty_{NaCl}$
$= 91.0 + 426.2 - 126.5 = 390.7\ \text{S cm}^2\ \text{mol}^{-1}$

**34. (a) :** From Faraday's 1st law,
$W = Z \times Q$ [$W$ = weight, $Z$ = electrochemical equivalent, $Q$ = quantity of electricity]
Now $E = Z \times F$ [$E$ = equivalent weight , $F$ = Faraday]

$$W = \frac{E}{F} \times Q \Rightarrow Q = \frac{W \times F}{E} = \frac{W \times F}{\frac{A}{n}}$$

[$A$ = Atomic weight , $n$ = valency of ion]

or $Q = \dfrac{n \times w \times F}{A} = \dfrac{3 \times 5.12 \times 10^3 \times 96500}{27} = 5.49 \times 10^7 \text{C}$

**35. (d) :** Higher the acidity, higher will be the tendency to release protons and hence lighter will be the electrical conductivity. Difluoroacetic acid will be strongest acid due to electron withdrawing effect of two fluorine atoms so as it will show maximum electrical conductivity.

**36. (a) :** $Cr^{2+} \mid Cr^{3+} = +0.41$ V, $Mn^{2+} \mid Mn^{3+} = -1.57$ V
$Fe^{2+} \mid Fe^{3+} = -0.77$ V, $Co^{2+} \mid Co^{3+} = -1.97$ V
More is the value of oxidation potential more is the tendency to get oxidised.
As Cr will have maximum oxidation potential value, therefore its oxidation will be easiest.

**37. (c) :** $Zn_{(s)} + 2H^+_{(aq)} \rightleftharpoons Zn^{2+}_{(aq)} + H_{2\,(g)}$

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{[Zn^{2+}] \times p_{H_2}}{[H^+]^2}$$

On adding $H_2SO_4$ the $[H^+]$ will increase therefore $E_{cell}$ will also increase and the equilibrium will shift towards the right.

**38. (a) :** $\Lambda^\circ_{NaCl} = \Lambda^\circ_{Na^+} + \Lambda^\circ_{Cl^-}$ ...(i)
$\Lambda^\circ_{KBr} = \Lambda^\circ_{K^+} + \Lambda^\circ_{Br^-}$ ...(ii)
$\Lambda^\circ_{KCl} = \Lambda^\circ_{K^+} + \Lambda^\circ_{Cl^-}$ ...(iii)
Equation (i) + (ii) – (iii)
$\Lambda^\circ_{NaBr} = \Lambda^\circ_{Na^+} + \Lambda^\circ_{Br^-} = \Lambda^\circ_{NaCl} + \Lambda^\circ_{KBr} - \Lambda^\circ_{KCl}$
$= 126 + 152 - 150 = 128\ \text{S cm}^2\ \text{mol}^{-1}$

**39. (c) :** $E_{cell} = E^\circ_{cell} - \dfrac{0.0591}{n}\log K_c$

$0 = 0.591 - \dfrac{0.0591}{1}\log K_c$

$\Rightarrow\ -0.591 = -0.0591 \log K_c \Rightarrow \log K_c = \dfrac{0.591}{0.0591} = 10$

$\therefore\ K_c = \text{antilog } 10 = 1 \times 10^{10}$

**40. (c) :** $E^\circ_{cell} = E^\circ_{Sn/Sn^{2+}} + E^\circ_{Fe^{3+}/Fe^{2+}} = 0.14 + 0.77 = 0.91$ V

**41. (b) :** Direct conversion of chemical energy to electric energy can be made considerably more efficient (*i.e.* upto 75%) than the 40% maximum now obtainable through burning of fuel and using the heat to form steam for driving turbines. Furthermore, the water obtained as a byproduct may be used for drinking by the astronauts.
At anode : $2H_{2(g)} + 4OH^-_{(aq)} \rightarrow 4H_2O_{(l)} + 4e^-$
At cathode : $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$
$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

**42. (c) :** $CN^-$ ions act both as reducing agent as well as good complexing agent.

**43. (a) :**

| | *A* | *B* | *C* |
|---|---|---|---|
| $E^\circ_{red}$ | +0.5 V | –3.0 V | –1.2 V |

More is the value of reduction potential, more is the tendency to get reduced, *i.e.* less is the reducing power.
The reducing power follows the following order: $B > C > A$.

**44. (d) :** $E^\circ_{cell} = \dfrac{0.0591}{n}\log K_c \Rightarrow 0.295 = \dfrac{0.0591}{2}\log K_c$

$0.295 = 0.0295 \log K_c \Rightarrow K_c = \text{antilog } 10 = 1 \times 10^{10}$

**45. (c) :** $E_{cell} = E^\circ_{cell} - \dfrac{0.0591}{n}\log\dfrac{1}{0.1}$

Here $n = 2$, $E^\circ_{cell} = 1.10$ V

$E_{cell} = 1.10 - \dfrac{0.0591}{2}\log 10 = 1.10 - 0.0295 = 1.0705$ V

**46. (b) :** The mass of silver deposited on the cathode $= \dfrac{108 \times 9650}{96500} = 10.8$ g

**47. (b) :** It is also known as heat capacity.

**48. (a) :** Here $Cr^{3+}$ is oxidised to $Cr_2O_7^{2-}$.

**49. (b) :** $2H^+ + 2e^- \longrightarrow H_2(P_2)$
$H_2(P_1) \longrightarrow 2H^+ + 2e^-$
Overall reaction : $H_2(P_1) \longrightarrow H_2(P_2)$

$$E = E^\circ - \frac{RT}{nF}\log\frac{P_2}{P_1} = 0 - \frac{RT}{nF} og \frac{P_2}{P_1} = \frac{RT}{nF}\log\frac{P_1}{P_2}$$

**50. (a) :** $E_{cell} = E_{right\ (cathode)} - E_{left\ (anode)}$.

**51. (b) :** $S \propto A$ ($A$ = area); $S \propto C$ ($C$ = concentration)
$S \propto \dfrac{1}{L}$ ($L$ = length)

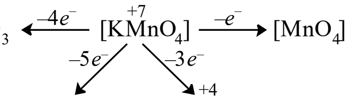
Combining we get, $S \propto \dfrac{AC}{L}$

or $S = K\dfrac{AC}{L}$ [$K$ = constant of proportionality]

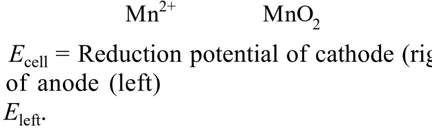
$K = \dfrac{SL}{AC}$

$\therefore$ Unit of $K = \dfrac{\text{S} \times \text{m}}{\text{m}^2 \times \dfrac{\text{mol}}{\text{m}^3}} = \dfrac{\text{S} \times \text{m} \times \text{m}^3}{\text{m}^2 \times \text{mol}} = \text{S m}^2\ \text{mol}^{-1}$

**52. (d) :** The oxidation states show a change only in reaction (d).



**53. (c) :**



**54. (c) :** $E_{cell}$ = Reduction potential of cathode (right) – reduction potential of anode (left)
$= E_{right} - E_{left}$.