Speed Test-45

1. **(a)** The given order of reduction potentials (or tendencies) is Z > Y > X. A spontaneous reaction will have the following characteristics
Z reduced and Y oxidised
Z reduced and X oxidised
Y reduced and X oxidised
Hence, Y will oxidise X and not Z.

2. **(c)** From the given data we find Fe^{3+} is strongest oxidising agent. More the positive value of E°, more is the tendency to get oxidized. Thus correct option is (c).

3. **(b)** $R = 100\,\Omega$, $\kappa = \frac{1}{R}\left(\frac{l}{A}\right)$,

$\frac{l}{A}$ (cell constant) $= 1.29 \times 100\ m^{-1}$
Given, $R = 520\,\Omega$, $C = 0.2$ M,
μ (molar conductivity) = ?

$\mu = \kappa \times V$ (κ can be calculated as $\kappa = \frac{1}{R}\left(\frac{l}{A}\right)$

now cell constant is known.)
Hence,

$$\mu = \frac{1}{520} \times 129 \times \frac{1000}{0.2} \times 10^{-6}\ m^3$$
$$= 12.4 \times 10^{-4}\ S\ m^2\ mol^{-1}$$

4. **(b)** For, $M^+ + X^- \longrightarrow M + X$, $E^\circ_{cell} = 0.44 - 0.33 = 0.11V$ is positive, hence reaction is spontaneous.

5. **(b)** RHS : $2H^+ + 2e^- \rightleftharpoons H_2(P_2)$
LHS : $H_2(P_1) \rightleftharpoons 2H^+ + 2e^-$
overall reaction : $H_2(P_1) \rightleftharpoons H_2(P_2)$

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

6. **(c)** $\Delta G^\circ = -nFE^\circ;\ E^\circ = \frac{-\Delta G^\circ}{nF}$;

$$E^\circ = \frac{-(-50.61J)}{2 \times 96500 \times 10^{-3}} = 0.26\ V$$

7. **(b)** $Ohm^{-1}\ cm^2\ (g\ eq)^{-1}$

8. **(a)** In case of equivalent conductance of strong electrolyte there is little increase with dilution.

9. **(d)** Here $n = 4$, and $[H^+] = 10^{-3}$ (as pH = 3)
Applying Nernst equation

$$E = E^o - \frac{0.059}{n}\log\frac{[Fe^{2+}]^2}{[H^+]^4(p_{O_2})}$$
$$= 1.67 - \frac{0.059}{4}\log\frac{(10^{-3})^2}{(10^{-3})^4 \times 0.1}$$
$$= 1.67 - \frac{0.059}{4}\log 10^7 = 1.67 - 0.103 = 1.567\ V$$

10. **(d)** Cell constant = l/a
Unit = $m/m^2 = m^{-1}$.

11. **(d)** $\Lambda^\circ_{m(NH_4Cl)} = \Lambda^\circ_{m_{NH_4^+}} + \Lambda^\circ_{m_{Cl^-}}$

$$\Lambda^\circ_{m(NaOH)} = \Lambda^\circ_{m_{Na^+}} + \Lambda^\circ_{m_{OH^-}}$$

$$\Lambda^\circ_{m(NaCl)} = \Lambda^\circ_{m_{Na^+}} + \Lambda^\circ_{m_{Cl^-}}$$

$$\therefore\ \Lambda^\circ_{m\left(NH_4^+\right)} + \Lambda^\circ_{m\left(OH^-\right)}$$

$$= \Lambda^\circ_{m\left(NH_4^+\right)} + \Lambda^\circ_{m\left(Cl^-\right)} + \Lambda^\circ_{m\left(Na^+\right)}$$

$$+\Lambda^\circ_{m\left(OH^-\right)} - \left[\Lambda^\circ_{m\left(Na^+\right)} + \Lambda^\circ_{m\left(Cl^-\right)}\right]$$

$$\Lambda^\circ_{m(NH_4OH)} = \Lambda^\circ_{m(NH_4Cl)} + \Lambda^\circ_{m(NaOH)} - \Lambda^\circ_{m(NaCl)}$$

12. **(d)**

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

$$Pb + PbO_2 + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O$$

The reaction indicates that 2 moles of H_2SO_4 corresponds to 2×96500 C and 2 moles $H_2SO_4 \equiv 4$ equiv. of H_2SO_4.
2×96500 C consumed 4 equiv. of H_2SO_4
and $100 \times 10^{-3} \times 9.65 \times 10^5$ C consumed

$$= \frac{4 \times 100 \times 10^{-3} \times 9.65 \times 10^5}{2 \times 96500} = 2\ \text{equiv. } H_2SO_4$$

$$\therefore\ \text{Decrease in normality} = \frac{2}{5} = 0.40$$

13. **(b)** For $Zn^{2+} \rightarrow Zn$

$$E_{Zn^{2+}/Zn} = E^\circ_{Zn^{2+}/Zn} - \frac{2.303RT}{nF}\log\frac{[Zn]}{[Zn^{2+}]}$$

$$= -0.76 - \frac{0.06}{2}\log\frac{1}{[0.1]} = -0.76 - 0.03$$

$$E_{Zn^{2+}/Zn} = -0.79V$$

14. **(c)** Reduction half reaction :

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$$

Oxidation half reaction :

$$Cr \longrightarrow Cr^{3+} + 3e^-$$

Overall reaction :

$$Cr_2O_7^{2-} + Cr + 14H^+ + 3e^- \longrightarrow 3Cr^{3+} + 7H_2O$$

3F of electricity = 3 moles of Cr^{3+}

1F of electricity = $\frac{3}{3}$ moles of Cr^{3+}

15. **(a)** $2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$

O.S. of Cr changes from +3 to +6 by loss of electrons. At anode oxidation takes place.

16. **(b)** In $H_2 - O_2$ fuel cell, the combustion of H_2 occurs to create potential difference between the two electrodes.

17. **(b)** $E^o_{cell} = \frac{0.059}{2} \log K_C$ or $\frac{1.10 \times 2}{0.059} = \log K_C$

$\therefore K_C = 1.9 \times 10^{37}$

18. **(a)** The value of $E^o_{M^{2+}/M}$ for given metal ions are

$E^o_{Mn^{2+}/Mn} = -1.18$ V,

$E^o_{Cr^{2+}/Cr} = -0.9$ V,

$E^o_{Fe^{2+}/Fe} = -0.44$ V and $E^o_{Co^{2+}/Co} = -0.28$ V.

The correct order of $E^o_{M^{2+}/M}$ values without considering negative sign would be

$$Mn^{2+} > Cr^{2+} > Fe^{2+} > Co^{2+}.$$

19. **(d)** For spontaneous reaction ΔG should be negative. Equilibrium constant should be more than one
($\Delta G = -2.303\, RT \log K_c$, If $K_c = 1$ then
$\Delta G = 0$; If $K_c < 1$
then $\Delta G = +ve$). Again $\Delta G = -nFE^o_{cell}$.

E^o_{cell} must be +ve to have ΔG –ve.

20. **(d)** Standard Gibbs free energy is given as $\Delta G° = -nE°F$
If $E°_{cell} < 0$ i.e., – ve
$\Delta G° > 0$
Further $\Delta G° = -RT \ln K_{eq}$
$\therefore \Delta G° > 0$ and $K_{eq} < 0$

21. **(c)** Conductivity of an electrolyte depends on the mobility of ions and concentration of ions. The motion of an ionic species in an electric field is retarded by the oppositely charged ions due to their interionic attraction. On dilution, concentration of electrolyte decreases and the retarding influence of oppositely charged ions decreases. Therefore mobility of ions increases.

22. **(d)** Higher the value of standard reduction potential, stronger is the oxidising agent, hence MnO_4^- is the strongest oxidising agent.

23. **(c)** As the value of reduction potential decreases the reducing power increases i.e.

C < B < D < A
(0.85). (0.6). (–0.76) (– 1.2)

24. **(b)** Specific conductance decreases with dilution.

25. **(a)** $\Lambda^{\infty}_m = 57 + 73 = 130\ S\,cm^2 mol^{-1}$

26. **(d)** $E_{cell} = 0$; when cell is completely discharged.

$$E_{cell} = E°_{cell} - \frac{0.059}{2} \log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$$

$$\text{or } 0 = 1.1 - \frac{0.059}{2} \log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right)$$

$$\log\left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right) = \frac{2 \times 1.1}{0.059} = 37.3$$

$$\therefore \left(\frac{[Zn^{2+}]}{[Cu^{2+}]}\right) = 10^{37.3}$$

27. **(c)** Reduction potential of Cu is more than H_2.

28. **(d)** Reducing character decreases down the series. Hence the correct order is

$Al < Fe^{2+} < Br^-$

29. **(c)** Writing the equation for pentane-oxygen fuel cell at respective electrodes and overall reaction, we get

At Anode:

$$\underset{\text{(pentane)}}{C_5H_{12}} + 10H_2O \rightarrow 5CO_2 + 32H^+ + 32e^-$$

At Cathode:

$$8O_2 + 32H^+ + 32e^- \rightarrow 16H_2O$$

$$\text{Overall}: C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$$

Calculation of $\Delta G°$ for the above reaction
$\Delta G° = [5 \times (-394.4) + 6 \times (-237.2)] - [-8.2]$
$= -1972.0 - 1423.2 + 8.2 = -3387.0$ kJ
$= -3387000$ Joules.
From the equation we find n = 32

Using the relation, $\Delta G° = -nFE^o_{cell}$ and substituting various values, we get

$$-3387000 = -32 \times 96500 \times E^o_{cell} \quad (F = 96500C)$$

$$\text{or } E^o_{cell} = \frac{3387000}{32 \times 96500}$$

$= \frac{3387000}{3088000}$ or $\frac{3387}{3088}\text{V} = 1.0968\text{ V}$

Thus option (c) is correct answer.

30. (d) From the given representation of the cell, E_{cell} can be found as follows.

$$E_{cell} = \left(E^{o}_{\text{Fe}^{2+}/\text{Fe}} - E^{o}_{\text{Cr}^{3+}/\text{Cr}}\right) - \frac{0.059}{6}\log\frac{[\text{Cr}^{3+}]^2}{[\text{Fe}^{2+}]^3}$$

[Nernst -Equ.]

$$= -0.42 - (-0.72) - \frac{0.059}{6}\log\frac{(0.1)^2}{(0.01)^3}$$

$$= -0.42 + 0.72 - \frac{0.059}{6}\log\frac{0.1\times 0.1}{0.01\times 0.01\times 0.01}$$

$$= 0.3 - \frac{0.059}{6}\log\frac{10^{-2}}{10^{-6}} = 0.3 - \frac{0.059}{6}\times 4$$

$= 0.30 - 0.0393 = 0.26\text{ V}$

Hence option (d) is correct answer.

31. (b) $H_2O \rightleftharpoons H^+ + OH^-$

$H^+ + e^- \longrightarrow \frac{1}{2}H_2$

$\therefore$ 0.5 mole of H_2 is liberated by 1 F = 96500 C

0.01 mole of H_2 will be liberated by

$= \frac{96500}{0.5}\times 0.01 \quad = 1930\text{ C}$

$Q = I \times t$

$$t = \frac{Q}{I} = \frac{1930\text{ C}}{10\times 10^{-3}\text{ A}} = 19.3\times 10^4\text{ sec}$$

32. (d)

33. (b) Conductivity (X) = conductance (c) × cell constant

$\therefore$ Cell constant $= \frac{X}{Y}$

Conductivity of NaOH $= \frac{X}{Y}.Z$

$$\Delta m\,(\text{NaOH}) = \frac{X}{Y}.Z\times\frac{1000}{0.1} = \frac{XZ}{Y}10^4$$

34. (b) Total of 6 electrons are required to form 2 moles of Cr^{3+} therefore to form 1 mole of Cr^{3+} 3F of charge is required.

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

36. (a) Specific conductance = Conductance × Cell constant

$$k = \frac{1}{2.5\times 10^3}\times 1.15\,;$$

$$\Lambda_{eq} = \frac{1.15}{2.5\times 10^3}\times\frac{1000}{0.1} = 4.6$$

37. (a) Thus difluoro acetic acid being strongest acid will furnish maximum number of ions showing highest electrical conductivity. The decreasing acidic strength of the carboxylic acides given is difluoro acetic acid > fluoro acetic acid > chloro acetic acid > acetic acid.

38. (a) No. of moles of silver $= \frac{9650}{96500} = \frac{1}{10}$ moles

$\therefore$ Mass of silver deposited $= \frac{1}{10}\times 108 = 10.8\text{ g}$

39. (c) $H^+ + e^- \longrightarrow \frac{1}{2}H_2$

$$E = E^{o} - \frac{0.059}{1}\log\frac{[P(H_2)]^{1/2}}{[H^+]}$$

Now if $p_{H_2} = 2$ atm and $[H^+] = 1M$

$$\text{then } E = 0 - \frac{0.059}{1}\log\frac{2^{1/2}}{1} = \frac{-0.059}{2}\log 2$$

40. (a) 112 mL of H_2 at STP $= \frac{2\times 112\text{g}}{22400}$

(Since 22400 mL at STP = M.wt)

Amount deposiled $= \frac{\text{Eq.wt}\times i\times t}{96500}$

$$\therefore \frac{2\times 112}{22400} = \frac{1\times 965\times i}{96500};$$

$i = 1$ amp

41. (d) $\frac{W_A}{E_A} = \frac{W_B}{E_B};\ \frac{1.6}{8} = \frac{\text{Wt. of Ag}}{108}$

$\therefore$ Wt. of Ag = 21.6 g

42. (d) Correct matching for pair (iii) will be [G (conductance) – siemens or ohm^{-1}(S).]

43. (a) The charge carried by 1 mole of electrons is one faraday. Thus for a reaction

$M^{n+} + ne^- \longrightarrow M$

nF = 1 mole of M

$$\underset{\substack{3F\\1F}}{Al^{3+} + 3e^-} \longrightarrow \underset{\substack{1\text{ mole}\\1/3\text{ mole}}}{Al}$$

$$\underset{\substack{2F\\1F}}{Cu^{2+} + 2e^-} \longrightarrow \underset{\substack{1\text{ mole}\\1/2\text{ mole}}}{Cu}$$

$$\underset{1F}{Na^{+} + e^-} \longrightarrow \underset{1\text{ mole}}{Na}$$

The mole ratio of Al, Cu and Na deposited at the respective cathode is $\frac{1}{3}:\frac{1}{2}:1$ or 2 : 3 : 6.

44. (a) Reciprocal of resistance is conductance. We can add the conductance and not resistance.

Hence $\frac{1}{r} = \frac{1}{\rho} + \frac{1}{\rho} + ...$

45. (c) The correct relation is $\Delta G^{\circ} = -nFE^{\circ}$