Solution

CET25C2 ELECTROCHEMISTRY

Class 12 - Chemistry

1.

(c) the standard reduction potential of Li^+/Li is lower than that of Cu^{2+}/Cu

Explanation: Li is alkali metal E^o Li^+/Li = -3.045V.

2.

(d) 2.0

Explanation: Two faraday of electricity is passed through a solution of $CuSO_4$.

3.

(d) Cell emf

Explanation: EMF is the difference between the electrode potentials of two electrodes cathode and anode when no current is drawn through the cell.

4. **(a)** 6.02×10^{23}

Explanation: 6.02×10^{23}

1 mol of electrons contains avagadro number (6.02×10^{23}) of electrons

5. **(a)** All of these

Explanation: In a lead storage battery, all of these statements are true.

Explanation:

i. The lead storage battery is a **rechargeable battery**.
ii. This battery has a **pure lead** as anode; **lead(IV) oxide** as a cathode and **dilute sulfuric acid** as an electrolyte.
iii. When the battery is being used, lead(IV) oxide is **reduced** to $PdSO_4$ and lead is **oxidized** to $PdSO_4$.
iv. When the battery is being charged, $PdSO_4$ is again converted into **PbO_2 and Pb** as cathode and anode respectively.

6. **(a)** R = $H_{2(g)}$, $O_{2(g)}$: P = $H_2O(l)$

Explanation: R = $H_{2(g)}$, $O_{2(g)}$: P = $H_2O(l)$

7. **(a)** 124.0 S $cm^2 mol^{-1}$

Explanation: molar conductivity = k x Volume in cm^3 containing 1 mol of electrolyte.

so $\Lambda m = \frac{k}{c} \times 1000$

= $0.0248 \times 1000/0.2$ = 124Scm^2/mol

8. **(a)** 2.0

Explanation: 2.0

9.

(c) $E_{ext} > E_{cell}$

Explanation: $E_{ext} > E_{cell}$

10.

(c) 1×10^{10}

Explanation: $E^0_{cell} = \left(\frac{0.0591}{n}\right) \log K_c$

$0.295 = \frac{0.0591}{2} \log K_c$

$\log K_c = \frac{2 \times 0.295}{0.0591} = 10$

K_c = antilog 10 = 1×10^{10}

11.

(d) Y

Explanation: Y

12.

(b) $Zn | Zn^{2+} || Ag^{+} | Ag$

Explanation: $Zn | Zn^{2+} || Ag^{+} | Ag$ is correct cell representation

13.

(c) 0.80 V

Explanation: $E^{0}cell = E^{0}cathode - E^{0}$ anode

$0.80 = E^{0}Ag^{+}/Ag - E^{0}H^{+}/H_{2}$

$0.80 = E^{0}Ag^{+}/Ag - 0$

$E^{0}Ag^{+}/Ag = 0.80V$

The standard reduction potential of Ag^{+}/Ag electrode is 0.80V

14.

(b) 3F

Explanation: $\mathrm{Al_2O_3 \rightarrow 2Al^{3+} + 3O^{2-}}$

$\mathrm{Al^{3-} + 3e^{-} \rightarrow Al}$ (For 1 mole)

3F charge is required to obtain 1 mole Al from Al_2O_3. Hence to neutralize Al^{3+} ion 3F is required.

15. **(a)** -0.80 V

Explanation: $2Ag^{+}{}_{(aq)} + H_{2(g)} \rightarrow 2Ag_{(s)} + 2H^{+}{}_{(aq)}$

Formula: $E_o = E_{cathode} - E_{anode}$

$E_o = E_{Ag^{+}/Ag} - E_{H^{+}/H2}$

$0.08 = E_{Ag^{+}/Ag} - 0$

= -0.08V

16. **(a)** Cr^{3+}

Explanation: $\mathrm{E}^{\circ}_{\mathrm{Cr}^{3+}/\mathrm{Cr}}$ has the most negative value equal to -0.74 among given four choices. So, Cr^{3+} is the most stable oxidized species.

17.

(d) $\mathrm{2Fe^{3+}_{(aq)} + Sn_{(s)} \rightarrow 2Fe^{2+}_{(aq)} + Sn^{2+}_{(aq)} E^{\circ} = +0.91\ V}$

Explanation: $\mathrm{2Fe^{3+}_{(aq)} + Sn_{(s)} \rightarrow 2Fe^{2+}_{(aq)} + Sn^{2+}_{(aq)} E^{\circ} = +0.91\ V}$

18.

(d) Mn^{2+}

Explanation: Mn^{2+} is most stable in its reduced form due to highest E^{o} cell value.

19.

(c) an electrochemical process

Explanation: Corrosion of an iron is an electrochemical process.

20.

(b) energy of combustion of fuel is converted to electrical energy.

Explanation: energy of combustion of fuel is converted to electrical energy.

21. **(a)** E_a

Explanation: E_a Activation energy is affected by the catalyst.

22.

(d) Fe^{3+}

Explanation: Strongest oxidizing agent is one having more positive or less negative reduction potential. The reduction potential of Fe^{3+}/Fe^{2+} = 0.77V. so it has more tendency to get reduced or gain electrons. hence it is good oxidising agent.

23.

(c) 1, 3, 4, 5

Explanation: No. of electrons = change in oxidation number of active element.

24.

(c) 1.0 M

Explanation: 1.0 M

25.

(c) 0.53 V

Explanation: $E_{cell} = E^0_{cell} - (0.0591/n)\ \log\,[Fe^{2+}]/\,[H^+]^2$

$E^0_{cell} = E^0_{H^+/H2} - E^0_{Fe^{2+}/Fe} =$ 0 - (-0.44)V = +0.44V

$E_{cell} = 0.44 - \frac{0.0591}{2}\log\frac{10^{-3}}{(1)^2} = \frac{0.44 + 0.0591 \times (-3)}{2}$

E_{cell} = 0.44 + 0.089V

E_{cell} = +0.53V

26.

(b) 10^{30}

Explanation: $E^0 = \left(\frac{0.0591}{n}\right)\log k$ o.59 = 0.0591/3 log K

log K = 3 × 0.59/0.0591 = 30

K = AL 30 = 10^{30}

27.

(b) $\Lambda^0_m(\mathrm{NH_4Cl})^+\Lambda^0_{m(\mathrm{NaOH})} - \Lambda^0_{(\mathrm{NaCl})}$

Explanation: $\mathrm{NH_4Cl} \rightleftharpoons \mathrm{NH_4^+} + \mathrm{Cl}$ ...(i)

$\mathrm{NaCl} \rightleftharpoons \mathrm{Na^+} + \mathrm{Cl^-}$ ....(ii)

$\mathrm{NaOH} \rightleftharpoons \mathrm{Na^+} + \mathrm{OH}$ ...(iii)

$\mathrm{NH_4OH} \rightleftharpoons \mathrm{NH_4^+} + \mathrm{OH}$ ......(iv)

Therefore, $\Lambda^\circ_{\mathrm{m(NH_4OH)}}$ is equal to

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

28.

(c) 3F

Explanation: For reduction of 1 mol of Al^{3+} to Al , Al^{3+} + 3e gives Al ,

, 3 mol of electrons are required to deposit 1 mol of Al .so total charge required will be 3F.

29. **(a)** 1.104 V

Explanation: $E^0_{cell} = E^0_{cathode} - E^0_{anode}$

E^0cell = $E^0 Ag_2O/Ag$ - E^0 Zn^{2+}/Zn

= +0.344-(-0.76) = 1.104V

30.

(d) I and III

Explanation: On electrolysis of brine solution

at anode $2Cl^-$ → Cl_2 + $2e^-$ (oxidation)

at cathode Na^+ + e^- → Na

31.

(c) ΔG = -ve, E^o = +ve

Explanation: For spontaneity of cell,

ΔG = -ve

E^o = +ve

32.

(d) 1.57 V

Explanation: $E_{cell} = E^0_{cell} - \left(\frac{0.0591}{n}\right)\log\left(\frac{[Fe^{2+}]^2}{[H^+]^4(P_{O2})}\right)$

$E_{cell} = 1.67 - \frac{0.0591}{4}\log\frac{(10^{-3})^2}{(10^{-3})^4(10^{-1})}$ = 1.57V

33. **(a)** $\wedge = \wedge_0$ as $\mathrm{C} \longrightarrow 0$

Explanation: $\wedge = \wedge_0$ as $\mathrm{C} \longrightarrow 0$

34. **(a)** $Mn^{2+} < Cl^- < Cr^{3+} < Cr$

Explanation: The lower the reduction potential, the higher will its reducing power. Therefore, The order of reducing power is $Mn^{2+} < Cl^- < Cr^{3+} < Cr$

35.

(b) Pt(s) I H_2 (g, 1 bar) I H^+ (aq.,1 M) II Cu^{2+} (aq.,1 M) I Cu

Explanation: When the copper electrode is connected to the standard hydrogen electrode, it acts as the cathode and its standard electrode potential can be measured.

$E^o = E^o_R - E^o_L = E^o_R - O = E^o_R$

Pt(s) | H_2 (g, 1 bar)| H^+ (aq., 1M) || Cu^{2+} (aq., 1M)| Cu will measure standard electrode potential of copper electrode.

To calculate the standard electrode potential of the given cell it is coupled with the standard hydrogen electrode in which pressure of hydrogen gas is one bar and the concentration of H^+ ion in the solution is one molar and also the concentrations of the oxidized and the reduced forms of the species in the right-hand half-cell are unity. Hydrogen gas can reduce copper under standard conditions.

36.

(b) 5F

Explanation: For reduction of MnO_4^- to Mn^{2+} 5 mol of electrons are required to deposit 1 mol of Mn^{2+} ion. so total charge required will be 5F.

$5e^- + 8H^+ + MnO_4^- = Mn^{2+} + 4H_2O$

37.

(b) 0.03V, -2.895 kJ/mol, 3.22

Explanation: $E^0_{cell} = E^0_{cathode} - E^o_{anode} = \frac{E^0 Ag}{Ag} - \frac{E^0 Fe^{3+}}{Fe^{2+}} = +0.8 - (-0.78) = +0.03V$

$\Delta_r G^0 = -nFE^0 cell = -1 \times 96500 \times 0.03$ = -2895J/mol = -2.895kJ/mol

$E^0_{cell} = \frac{0.0591}{n}\log K_c$

$\log K_c = \frac{n \times E^0_{cell}}{0.0591} = \frac{1 \times 0.03}{0.0591} = 0.508$

$k_c = 3.22$

38.

(b) $10^{\frac{0.32}{0.0295}}$

Explanation: For cell Zn|Zn^{2+} (a=0.1M) ‖ Fe^{2+}(a=0.01M)|Fe

$E^0_{cell} = \left(\frac{0.0591}{n}\right)\log k$

For this cell, reaction is $Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe$

$E = E^0 - \frac{0.0591}{n}\log\frac{C_1}{C_2}$

$E^0 = E + \frac{0.0591}{n}\log\frac{C_1}{C_2}$

$= 0.2905 + \frac{0.0591}{2}\log\frac{10^{-1}}{10^{-2}} = 0.32$

$E^0 = \frac{0.0591}{2}\log K$

$\log K = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$

$K = 10^{\frac{0.32}{0.0295}}$

39.

(d) Zn + 2AgCN → 2Ag + Zn $(CN)_2$

Explanation: In redox both oxidation and reduction take place simultaneously.

zinc is oxidised and silver is reduced.

40.

(c) 1.8245g

Explanation: $W = \frac{(E \times I \times t)}{96500}$

equivalent mass of Ni $= \frac{atomicmass}{valency} = \frac{58.69}{2}$ = 29.345

$w\ = \frac{29.345 \times\ 5 \times\ 20 \times\ 60}{96500}$

= 1.8245g

41. **(a)** + 0.32 V

Explanation: E^0 cell = E^0cathode - E^0anode

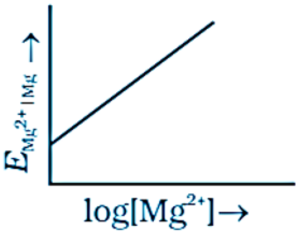
= -0.44 - (-0.76) = +0.32V

The emf of the cell is 0.32V

42. **(a)** from Cu to Zn outside the cell

Explanation: In Daniell cell, Zn acts an anode and Cu acts as cathode, electron flow takes place from anode to cathode and hence, electricity flows from Cu to Zn outside the cell.

43. **(a)**



Explanation: $E_{Mg^{2+}/Mg} = E^{\circ}_{Mg^{2+}/Mg} - \frac{0.059}{2}\log\left[Mg^{2+}\right]$

Compare this equation with the equation of straight line y = mx + c. The graph of $E_{Mg^{2+}/Mg}$ versus log $[Mg^{2+}]$ is a straight line with a positive slope and the intercept $E_{Mg^{2+}/Mg}$.

44.

(c) remains constant for a cell

Explanation: The cell constant of a conductivity (k) cell remains constant for a battery cell.

45.

(c) 2.651 V

Explanation: $E_{cell} = E^{\circ}_{cell} - \frac{0.059}{n}\log\frac{\left[Mg^{2+}\right]}{\left[Cu^{2+}\right]}$

$= 2.71 - \frac{0.059}{2}\log\frac{0.1}{0.001}$

$= 2.71 - \frac{0.059}{2}\log 10^2$

$E_{cell} = 2.651$ V

46.

(c) + 0.32 V

Explanation: $E^0_{cell} = E^0_{cathods} - E^0_{anods}$

E^0cell = E^0Fe^{2+}/Fe - E^0Zn^{2+}/Zn

= -0.44-(-0.76)= +0.32V.

47. **(a)** $6.25 \times 10^{-4} Sm^2 mol^{-1}$

Explanation: κ = G × cell constant and G= $\frac{1}{R}$

For 0.2M solution, R = 50 ohms, k= $1.3Sm^{-1}$

cell constant $\frac{l}{A}$ = 50 × 1.3 = $65m^{-1}$

$\Lambda m = \frac{k}{c}$ for 0.4M solution

molar conductivity = ? , R = 260 ohms, cell constant = $65m^{-1}$

so K = 65/260 = 0.25 Sm^{-1}

molar conductivity of the solution = K × 1000/M = 0.25/0.4 × 1000 = 6.25×10^{-4} Sm^2/mol^{-1}

48. **(a)** reduced form is more stable compared to hydrogen gas

Explanation: The standard hydrogen electrode is a reduction-oxidation electrode which forms the standard of the scale of oxidation-reduction potentials. It is given a scale of zero.

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

Now if an electrode is having greater electrode potential than standard or zero scales then we can infer that its reduced form is more stable compared to hydrogen (H_2) gas. It is more likely that metals ions will be formed than the hydrogen gas.

49.

(d) 2.68 V

Explanation: $E_{cell} = E^0_{cell} - \frac{0.0591}{n} \log \frac{[Mg^{2+}]}{[Cu^{2+}]}$

$E^0_{cell} = \frac{E^0 Cu^{2+}}{Cu} - \frac{E^0 Mg^{2+}}{Mg} = +0.34 - (-2.37) = 2.71V$

Here, E_{cell} = 2.71V, $[Mg^{2+}]$ = 0.1M, $[Cu^{2+}]$ =0.01M, n =2

$E_{cell} = 2.71 - \frac{0.0591}{2} \log \frac{0.1}{0.01}$

$= 2.71 - \frac{0.0591}{2} \log 10$

= 2.68 V

50. **(a)** Zn is readily oxidized.

Explanation: Since $E^o_{Zn^{2+}/Zn}$ is negative, Zn has greater tendency to be oxidized than hydrogen.

51.

(d) 0.1 M difluoroacetic acid

Explanation: Acidity increases on attaching electron withdrawing group because of stability of conjugate base. Conduction of electric charge depends on the ions. so higher the number of ions higher is electrical conductivity. As fluoro group causes negative inductive effect increasing ionisation, it gives stable ions in solution. also in difluoroacetic acid, the presence of 2 fluorine atom increases the negative inductive effect.

52. **(a)** 0.002 N

Explanation: Specific conductance or conductivity decreases with decrease in concentration of the electrolytic solution as the number of ions per unit volume that carry the current in a solution decrease on dilution.

53.

(d) 124.66 S cm^2/mol

Explanation: 124.66×10^{-4} Sm^2 mol^{-1}

Molar conductance of NaCl = $\lambda^+{}_{Na^+} + \lambda^+{}_{Cl^-}$

$= 51.12 \times 10^{-4} + 73.54 \times 10^{-4}$

$= 124.66 \times 10^{-4}$ Sm^2 mol^{-1}

54.

(c) S cm^2 mol^{-1}

Explanation: Unit of molar conductivity is S cm^2 mol^{-1}

55. **(a)** Ag

Explanation: Ag is less reactive than Hydrogen.

56.

(b) 9.65 Amp

Explanation: $W = \frac{(E \times I \times t)}{96500}$ volume occupied by 1mol of any gas at STP = 22,400cc

2g of hydrogen, the volume occupied is 22,400cc

so for 1.12cc, the mass of hydrogen will be = 1×10^{-4}g

$I = \frac{96500 \times 10^{-4}}{1} = 9.65A$

current to be passed in 1 sec = 9.65A

57.

(d) velocity of both K^+ and $NO_3{}^-$ are nearly the same.

Explanation: Saturated solution of KNO_3 is used in salt bridge because of same mobility of K^+ and NO_3^- . if the velocities would not have been the same then the ions would not have been able to neutralise the charge equally in both half cells. consequently, an opposing force would be developed that restricts the flow of current.

58.

(d) 0.219 cm^{-1}

Explanation: Conductivity k = 1/R $\times$ cell constant.

cell constant l/A = conductivity $\times$ R = $0.146 \times 10^{-3} \times 1500 = 0.219\ cm^{-1}$

59.

(b) Potentiometer

Explanation: The standard electrode potential is measured by the potentiometer. It is the potential difference when no current is flowing in the circuit.

60.

(d) sulphuric acid is consumed

Explanation: Discharging the stored energy relies on both the positive and negative plates becoming lead (II) sulphate and the electrolyte losing much of its dissolved sulphuric acid. Hence, sulphuric acid is consumed.

61.

(c) ions

Explanation: The flow of electricity, through solutions of electrolytes, is due to the migration of ions when a potential difference is applied between the two electrodes.

62.

(d) Lead- storage battery cell

Explanation: Lead storage battery is a secondary cell while leclanche cell and mercury cell are examples of primary batteries.

63.

(c) specific conductance

Explanation: Specific conductance decreases with dilution as a number of ions decreases per unit volume. (Conductivity through a solution is generally because of the ions that are present in it).

64.

(b) 2F

Explanation: 2F

65.

(d) $1.93 \times 10^5 s$

Explanation: $W = \frac{(E \times I \times t)}{96500}$

W is mass of hydrogen gas liberated.

w= 0.01mol = 0.02g, E = 1.008 , I = 10 milliamp =10 x 10^{-3}A ; t= ?

The time required to liberate 0.01 mol of H_2 gas at the cathode is

$t = \frac{w \times 96500}{E \times I} = \frac{0.02 \times 96500}{1.008 \times 0.01} = 1.93 \times 10^5 s$

66.

(d) 0.521 V

Explanation:

$\Delta G^0 = -nFE^0$

For $Cu^{2+} + 2e \rightarrow Cu$;ΔG_1^0 =-2F(+0.337)---(i)

$Cu^{2+} + e \rightarrow Cu^+$; ΔG^0_2=-F(+0.153) -----(ii)

Subtracting (ii) from (i) $Cu^+ + e \rightarrow Cu$.----(iii)

ΔG^0 = -0.674F + 0.153F

$-nFE^0$ = -0.521F

Since for (iii) , n=1, so E^0 = +0.521V.

67. **(a)** Graphite rod

Explanation: Graphite rod

68. **(a)** negative, positive

Explanation: ΔG and E°cell for spontaneous reaction will be negative, positive.

69.

(c) H_2 - O_2 Fuel cell

Explanation: H_2 - O_2 Fuel cell is used in Apollo space project.

70.

(b) Fe_2O_3 and $Fe(OH)_3$

Explanation: Fe_2O_3 and $Fe(OH)_3$ rust has general formula $Fe_2O_3.xH_2O$. it is formed when iron compounds corrode in the presence of air and oxygen. So it is a mixture of oxides and hydroxides.

71.

(c) –1.66 V

Explanation: $E^o_{cell} = E^o_{cathods} - E^o_{anods}$

72.

(c) -0.177 V

Explanation: pH = 3

$E_{cell} = E°_{cell} - 0.059/n \log(P_{H2}/[H^+]^2)$

$E_{cell} = -0.177V$

73.

(b) 863s, Copper 0.426g, Zinc 0.438 g

Explanation: $Ag^+ + e = Ag$ 96500C of current deposits 108g of silver . so 1.45g of silver would be deposited by = 96500 × 1.45/108 = 1295.6C

Q = I × t 1295.6C = 1.5 × t therefore t = 863s

$W = \frac{(E\times I\times t)}{96500}$ Equivalent mass of copper = 31.75, equivalent mass of zinc = 32.65

mass of zinc deposited = 32.65 × 1.5× 863/96500 = 0.438g

mass of copper deposited = 31.75 × 1.5× 863/96500 = 0.426g

74. **(a)** $1.93 \times 10^5 C$

Explanation: $H_2O = 2H^+ + 1/2\ O_2 + 2e^-$

2 mol of electrons = 2 × 96500 = 1.93×10^5 C

75. **(a)** 0.52 V

Explanation: Gibb's free energy is an additive property.

$\Delta G^\circ = -nFE^0$

For reaction, $Cu^{2+} + 2e^- \rightarrow Cu$;

$\Delta G^\circ = 2 \times F \times 0.337$ .... (i)

For reaction, $Cu^+ \rightarrow Cu^{2+} + e^-$;

$\Delta G^\circ = +1 \times F \times 0.153$ ..... (ii)

Adding Eqs. (i) and (ii), we get

$Cu^+ + e^- \rightarrow Cu$; ΔG° = -0.521 F

$\Delta G^\circ = -nFE^0$

-0.521 F = $-nFE^o$

E^o = 0.52V

Gibb's free energy is an additive property.

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

For reaction, $Cu^{2+} + 2e^- \rightarrow Cu$;

$\Delta G^\circ = 2 \times F \times 0.337$ .... (i)

For reaction, $Cu^+ \rightarrow Cu^{2+} + e^-$;

$\Delta G^o = +1 \times F \times 0.153$

Adding Eqs. (i) and (ii), we get

$Cu^+ + e^- \rightarrow Cu$; ΔG° = -0.521F

$\Delta G^\circ = -nFE^0$

-0.521 F = -nFE^o

E^o = 0.52 V