CBSE Test Paper-01

Class 12 Chemistry (Electrochemistry)

1. Given
   1. $Cu^{2+} + 2e^- \rightarrow Cu$, $E^0 = 0.337$ V
   2. $Cu^{2+} + e^- \rightarrow Cu^+$, $E^0 = 0.153$ V

   Electrode potential E° for the reaction $Cu^{2+} + e^- \rightarrow Cu$ will be
   1. 0.52 V
   2. 0.38 V
   3. 0.30 V
   4. 0.90 V
2. How much electricity is required in coulomb for the oxidation of 1 mol of FeO to Fe_2O_3?
   1. 95000C
   2. 96000C
   3. 96487C
   4. 95550C
3. Correct arrangement of Al, Cu, Fe, Mg and Zn in the order in which they displace each other from the solution of their salts is
   1. Mg > Al > Cu > Fe > Zn
   2. Mg > Al > Zn > Cu > F
   3. Mg > Al > Zn > Fe > Cu
   4. Cu > Al >Zn > Fe > Mg
4. In the button cells widely used in watches and other devices the following reaction takes place:

   $Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \rightarrow Zn^{2+}_{(aq)} + 2Ag_{(s)} + 2OH^-_{(aq)}$ [Given E^0 Zn^{2+}/Zn = -0.76V, $E^0 Ag_2O/Ag$ = +0.344V]

   Determine E^o cell for the reaction.
   1. 1.104 V
   2. 1.005 V
   3. 0.913 V

   d. 1.159 V

5. Conductivity of 0.00241 M acetic acid is 7.896 × 10^{-5} S cm^{-1}. If $\Lambda\ ^{0}m$ for acetic acid is 390.5 S cm^2mol^{-1}, what is its dissociation constant?
   a. 1.75 × 10^{-5}
   b. 2.05 × 10^{-5}
   c. 1.95 × 10^{-5}
   d. 1.85 × 10^{-5}

6. Write the cell formulation and calculate the standard cell potential of the galvanic cell in which the following reaction takes place:
   Fe^{2+}(aq) + Ag^{+}(aq) $\rightarrow$ Fe^{3+}(aq) + Ag(s)
   Calculate $\Delta_r G^0$ for the above reaction
   [Given: $E^0_{Ag^+/Ag} = +0.80V$
   $E^0_{Fe^{3+}/Fe} = +0.77V$
   [1 F = 96500 C mol^{-1}]

7. What is the representation of a Daniell cell?

8. Why is the voltage of a mercury cell constant during its working?

9. Suggest a way to determine the Λ^0_m value of water?

10. Molar conductance of 1.5 M solution of an electrolyte is found to be 138.9 S cm^2 mol^{-1} What would be the specific conductance of this solution?

11. At 298 K, the molar conductivities at infinite dilution of NH_4Cl, NaOH and NaCl are 129.8, 217.4 and 108.9s cm^2mol^{-1} respectively. The molar conductivity of 0.01 M NH_4OH solution is 9.33s cm^2mol^{-1}, calculate the degree of dissociation of NH_4OH at this dilution?

12. Consider the reaction:
    $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
    What is the quantity of electricity in coulombs needed to reduce 1 mol of $Cr_2O_7^{2-}$?

13. a. The resistance of a conductivity cell containing 0.0001 M KCl solution at 298 K is 1500Ω. What is the cell constant if the conductivity of 0.001 M KCl solution at 298 K is $0.146 \times 10^{-3} S\,cm^{-1}$

    b. Predict the products of electrolysis in the following:

    A solution of H_2SO_4 with platinum electrode.

14. Zinc rod is dipped in 0.1 M solution of $ZnSO_4$

    The salt is 95% dissociated at is dilution at 298 K . Calculate the electrode potential.

    Given:

    $E^0(Zn^{2+}/Zn)$ = -0.76

15. Explain construction and working of standard Hydrogen electrode?

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Class 12 Chemistry (Electrochemistry)

Solutions

1. a. 0.52 V

   Explanation: E^0 cell = 2(0.337)-0.153

2. c. 96487C

   Explanation: For converting FeO to Fe_2O_3 1mol of electrons are required.

3. c. Mg > Al > Zn > Fe > Cu

   Explanation: In application of electrochemical series, the metal which has lower reduction potential has higher tendency to get oxidised and would displace metals with lesser reduction potential from their salt solution.

4. a. 1.104 V

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

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

   = +0.344-(-0.76) = 1.104V

5. d. 1.85×10^{-5}

   Explanation: $\Lambda m = \frac{k}{c} \times 1000$ = $7.896 \times 10^{-5} \times 1000/0.00241$ = 32.76Scm^2/mol

   and $\alpha = \Lambda m/\Lambda^0 m$ = 32.76/390.5 = 0.084

   and K = $C\alpha^2/1-\alpha$ = $0.00241 \times (0.084)^2/1-0.084$ = 1.85×10^{-5}

6. $E_{cell}^0 = E_{Ag^+/Ag}^0 - E_{Fe^{3+}/Fe}^0$

   = +0.80 V - (+0.77) V

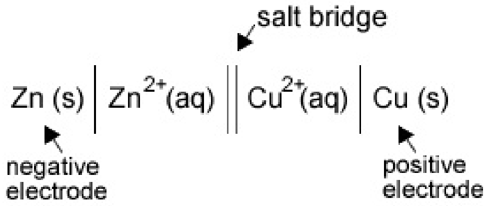
   = 0.03 V

   $\Delta_r G^0 = -nFE^o$

   $= -2 \times 0.03V \times 96500$

   = -5790 J mol^{-1}

   = 57.90 KJ mol^{-1}

7. 

8. As all the products and reactants are either in solid or liquid state, their concentration does not change with the use of the cell.

9. Applying Kohlrausch's law of independent migration of ions, the Λ_m^0 value of water can be determined as follows:
$$\Lambda^0_{m(H_2O)} = \lambda^0_{H^+} + \lambda^0_{OH^-}$$
$$= \left(\lambda^0_{H^+} + \lambda^0_{Cl^-}\right) + \left(\lambda^0_{Na^+} + \lambda^0_{OH^-}\right) - \left(\lambda^0_{Na^+} + \lambda^0_{Cl^-}\right)$$
$$= \lambda^0_{m(HCl)} + \lambda^0_{m(NaOH)} - \lambda^0_{m(NaCl)}$$
Hence, by knowing the λ_m^0 values of HCl, NaOH, and NaCl, the λ_m^0 value of water can be determined.

10. Given that

molarity = 1.5 m

molar conductivity = 138.9 S cm^2 mol^{-1}
$$\lambda_m = \frac{1000K}{M} = K = \frac{\lambda_m \times M}{1000}$$
$$= \frac{138.9\, S\, cm^2\, mol^{-1} \times 1.5 mol}{1000\, cm^3}$$
K = 0.20835 S cm^{-1}
$$= 2.0835 \times 10^{-1}\, Scm^{-1}$$

11. $\lambda_m{}^0(\mathrm{NH_4OH}) = \lambda_\mathrm{m}{}^\mathrm{o}(\mathrm{Na_4Cl}) + \lambda_\mathrm{m}{}^\mathrm{o}(\mathrm{NaOH}) - \lambda_\mathrm{m}{}^\mathrm{o}(\mathrm{NaCl})$

= 129.8 +217.4 - 108.9 = 237.3 5cm^2/mol

Degree of dissociation, $\alpha = \frac{\lambda m}{\lambda m^0} = \frac{9.335\mathrm{cm}^2/\mathrm{mol}}{237.35\mathrm{cm}^2/\mathrm{mol}}$

= 0.039 or 3.9%.

12. From the given reaction

1 mol of $Cr_2O_7^{2-}$ ions requires
$$6F = 6 \times 96500\ C$$
= 579000 C of electricity for reduction of Cr^{3+}.

13. a. At Anode:

$$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$$

At Cathode:

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

Overall reaction is:

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

Using nernst equation for n=2,

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

$$= +0.34V - (-2.37V) - \frac{0.0591}{2} \log \frac{0.001}{0.0001}$$

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

= 2.71V - 0.0295V

= 2.6805 V

For free gibbs energy

$$\Delta G = -nEF$$

$$= -2 \times 2.6805 \times 96500$$

= -517336.5 J mol^{-1}

= -517.34 KJ mol^{-1}

b. At cathode: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

At anode: $2OH^-(aq) \rightarrow O_2(g) + 2H^+(aq) + 4e^-$

H_2(g) is evolved at cathode and O_2(g) is evolved at anode.

14. $\left[Zn^{2+}\right] = 0.1 \times \frac{95}{100} = 0.095M$

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

$$E_{(Zn^+/Zn)} = E^0_{(Zn^+/Zn)} - \frac{0.0591}{2} \log \frac{1}{[Zn^{2+}]}$$

$$= -0.76V - \frac{0.0591}{2} \log \frac{1}{0.095}$$

$$= -0.76V - \frac{0.0591}{2} [\log 1000 - \log 95]$$

$$= -0.76 - \frac{0.0591}{2} [3.000 - 1.9777]$$

$$= -0.76V - \frac{0.0591}{2} \times 1.0223$$

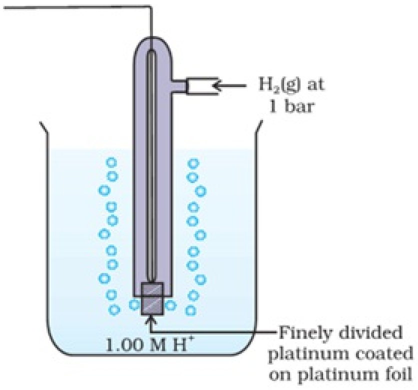
$$= -0.76V - \frac{0.0604}{2} = 0.76 - 0.0302$$

= - 0.7902 V

15. **Construction:** SHE consists of a platinum electrode coated with platinum black. The electrode is dipped in an acidic solution and pure Hydrogen gas is bubbled through it.

The concentration of both the reduced and oxidized. Forms of Hydrogen is maintained at unity i.e) pressure of H_2 gas is 1 bar and concentration of Hydrogen ions in the solution is 1 molar.

Working – The reaction taking place in SHE is $H^+(aq) + e^- \rightarrow 1/2H_2(g)$ At 298 K, the emf of the cell constructed by taking SHE as anode and other half-cell as cathode, gives the reduction potential of the other f cell whereas for a cell constructed by taking SHE as anode gives the oxidation potential of other half cell as conventionally the electrode potential of SHE is zero.



CBSE Test Paper-04

Class 12 Chemistry (Electrochemistry)

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1. Calculate the standard cell potentials of galvanic cell, $\Delta_r G^o$ and equilibrium constant of the reactions if the reaction is
   $$Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s) \left[\frac{E^0 Fe^{3+}}{Fe^{2+}} = 0.78V, \frac{E^0 Ag^{+}}{Ag} = 0.8V\right]$$
   a. 0.04V, -2.955 kJ, 3.2
   b. 0.02V, - 1.930 kJ, 2.180
   c. 0.01V, - 2.800 kJ, 3.2
   d. 0.02V, -2.850 kJ, 3.2
2. The SI units of molar conductance are:
   a. $Sm^{3}mol^{-1}$
   b. $Sm-^{1}mol^{-1}$
   c. $Sm^{2}mol^{-1}$
   d. $Sm^{-2}mol$
3. Li occupies higher position in the electrochemical series of metals as compared to Cu since
   a. Li is smaller in size as compared to Cu
   b. the standard oxidation potential of Li^{+}/Li is lower than that of Cu^{2+}/Cu
   c. the standard reduction potential of Li^{+}/Li is lower than that of Cu^{2+}/Cu
   d. the standard reduction potential of Cu^{2+}/Cu is lower than that of Li^{+}/Li
4. The e.m.f of the cell in which the reaction:
   $$2Ag^{+}(aq) + H_2(g) \rightarrow 2Ag(s) + 2H^{+}(aq)$$
   Occurs is 0.80 V. The standard reduction potential of Ag^{+}/Ag electrode is:
   a. 0.40 V
   b. -0.80 V
   c. 0.80 V
   d. -0.40 V
5. Electrochemistry deals with the production of electricity from
   a. energy released during spontaneous chemical reactions and the use of electrical

energy to bring about non-spontaneous chemical transformations

   b. energy released during spontaneous chemical reactions and the use of heat energy to bring about non-spontaneous chemical transformations
   c. nuclear energy and the use of electrical energy to bring about non-spontaneous chemical transformations
   d. nuclear energy and the use of heat energy to bring about non-spontaneous chemical transformations

6. Give the reaction taking place in lead storage battery when it is on charging?

7. Define the term - standard electrode potential?

8. Suggest two materials other than hydrogen that can be used as fuels in fuel cells.

9. Can you store copper sulphate solution in a zinc pot?

10. Explain rusting of iron.

11. The conductivity of 0.20 M solution of KCl at 298 K is 0.0248 S cm^{-1}. Calculate its molar conductivity.

12. Calculate the equilibrium constant for the reaction at T=298 K.
    $$Fe(s) + Cd^{2+}(aq) \rightarrow\ Fe^{2+}(aq) + Cd(s)$$
    $$\left[\text{Given, } E^{\Theta}_{Cd^{2+}/cd} = -0.40V, E^{\Theta}_{Fe^{2+}/Fe} = -0.44V\right]$$

13. The cell in which the following reaction occurs:
    $Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$ has $E^{0}{}_{(Fe^{3+}/Fe^{2+})} = +0.77\ V$ and $E^{0}{}_{(Ag^{+}/Ag)} = +0.80\ V$ at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction. (F = 96500 C mol^{-1})

14. What is cell constant? How is it determined?

15. Explain with an example how weak and strong electrolytes can be distinguished.

CBSE Test Paper-04

Class 12 Chemistry (Electrochemistry)

Solutions

1. b. 0.02V, - 1.930 kJ , 2.180

   Explanation: $E^0_{cell} = E^0_{cathode} - E^o_{anode}$
   $= \frac{E^0\,Ag}{Ag} - \frac{E^0\,Fe^{3+}}{Fe^{2+}} = +0.8 - 0.78 = +0.02V$
   $\Delta_r G^0 = -nFE^0 cell = -1 \times 96500 \times 0.02$ = -1930J = 1.930KJ
   $E^0_{cell} = \frac{0.0591}{n}\log K$
   $\log K = \frac{n \times E^0_{cell}}{0.0591} = \frac{1 \times 0.02}{0.0591} = 0.3384$
   k = Al 0.3384 = 2.180

2. c. Sm^2mol^{-1}

   Explanation: Sm^2mol^{-1}

   Λm = k × 1000/C = $Sm^{-1} \times m^3$ /mol/L = Sm^2/mol

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

   Explanation: Li is alkali metal E^0 Li^+/Li = -3.045V.

4. c. 0.80 V

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

   0.8 = E^0Ag^+/Ag - E^0H^+/H_2

   0.8 = E^0Ag^+/Ag - 0

   E^0Ag^+/Ag = 0.8V

5. a. energy released during spontaneous chemical reactions and the use of electrical energy to bring about non-spontaneous chemical transformations

   Explanation: Electrochemistry is branch of chemistry that deals with production of electricity from a redox chemical reaction or carrying out a non spontaneous reaction using electrical energy.

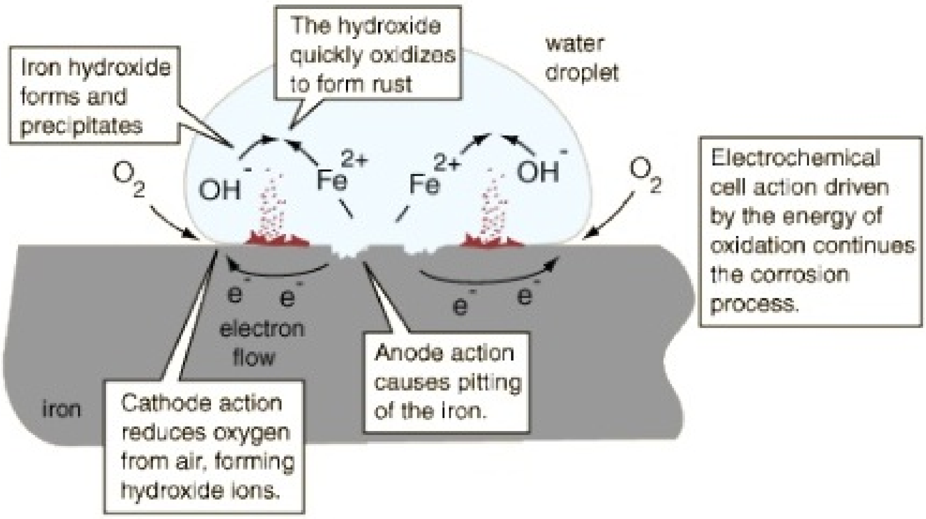
6. When the lead storage battery is on charging -
   $2PbSO_4$ (s) + $2H_2O$(e) $\rightarrow$ Pb(s) + PbO_2(s) + $2H_2SO_4$(aq)

7. When the concentration of all the species involved in a half-cell is unity, then the electrode potential is called standard electrode potential.

8. Methane and methanol can be used as fuels in fuel cells.

9. Zinc is more reactive than copper.
   Zn (s) E° = -0.76 V Cu(s) E° = - 0.34 V
   Hence, it displace copper from copper sulphate solution as follows.
   $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
   We cannot store copper sulphate sol. in a zinc pot.

10. **Rusting of Iron:** Corrosion slowly coats the surfaces of metallic objects with oxides or other salt of the metal. In corrosion, a metal is oxidized by loss of electrons to oxygen and formation of oxides. Corrosion of iron (commonly known as rusting) occurs in presence of water and air. The chemistry of corrosion is quite complex but it may be considered essentially as an electrochemical phenomenon. At a particular spot (in the given figure) of an object made of iron, oxidation takes place and that spot behaves as anode and we can write the reaction



Anode: $2Fe(s) \rightarrow 2Fe^{2+} + 4e^- \ E^{\ominus}_{(Fe^{2+}/Fe)} = -0.44\ V$
Electrons released at anode spot move through the metal and go to another spot on the metal and reduce oxygen in presence of H^+ (which is believed to be available from H_2CO_3 formed due to dissolution of carbon dioxide from air into water. Hydrogen ion in water may also be available due to dissolution of other acidic oxides from their

atmosphere). This spot behaves as cathode with the reaction.

Cathode: $O_2(g)+4H^+(aq)+4e^-\rightarrow 2H_2O(l)\ E^O_{(H^+/O_2/H_2O)}=1.23\ V$

The overall reaction being:

$2Fe(s)O_2(g)+4H^+(aq)\rightarrow 2Fe^{2+}(aq)+2H_2O(l)\ E^O_{(cell)}=1.67\ V$

The ferrous ion are further oxidized by atmospheric oxygen to ferric ion which comes out as rust in the form of hydrated ferric oxide (Fe_2O_3. x H_2O) and with further production of hydrogen ions.

Prevention of corrosion is of prime importance. It not only saves money but also helps in preventing accidents such as a bridge collapse or failure of a key component due to corrosion. One of the simplest methods of preventing corrosion is to prevent the surface of the metallic objects to come in contact with atmosphere. This can be done by covering the surface methods is to cover the surface by other metals (Sn, Zn, etc.) that are inert or react to save the object. An electrochemical method is to provide a sacrificial electrode of another metal (like Mg, Zn etc.) which corrodes itself but saves the object.

11. Given,

K= 0.0248 S cm^{-1}

c = 0.20 M

Therefore, Molar conductivity, $A_m=\frac{K\times 1000}{c}$

$=\frac{0.0248\times 1000}{0.2}$

= 124S cm^2 mol^{-1}

12. We have,

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

In this reaction, iron is oxidized and cadmium is reduced.

Half cell reaction of this cell is:

At Cathode (reduction):

$Cd^{2+}(aq)+2e^-\rightarrow Cd(s)$

At Anode (oxidation):

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

For this reaction, **n = 2 moles of electrons** takes place.

Now, we have

$$E^{\Theta}_{cell} = E^{\Theta}_{Cd^{2+}/Cd} - E^{\Theta}_{Fe^{2+}/Fe}$$

= -0.40 - (-0.44) = 0.04 V

We know that,

$$E^{\ominus}_{cell} \ = \ \frac{2.303RT}{nF} \log K_c$$

for n=2 and at T=298 K

$$E^{\theta}{}_{cell} = \frac{0.059}{2} \log K_c = 0.04$$

$$or \log K_c = \frac{2 \times 0.04}{0.059} = 1.356$$

K_c = Antilog 1.356

$$K_c = 2.270 \times 10^1 = 22.7$$

13. Given that the cell reaction is

$$Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$$

In this cell, Fe^{2+} is oxidized to Fe^{3+} and Ag^{+} is reduced to Ag(s). Therefore

At Cathode (Reduction):

$$Ag^{+}(aq) + 1e^{-} \rightarrow \ Ag(s)$$

At Anode (Oxidation):

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + 1e^{-}$$

Therefore number of moles of electrons is **n = 1 mole** for this cell.

Standard emf of the cell is

$$E^{\Theta}_{cell} = E^{\Theta}_{redution} - E^{\Theta}_{oxidation}$$

= +0.80 - (+0.77)

= +0.03 V

Now, for **standard free Gibb's energy,** we have

$$\Delta G^{\Theta} = -nFE^{\Theta}$$

$$= -1 \ mol \times 96500 \ Cmol^{-1} \times 0.03 \ V$$

= -2895 J

= -2.895 kJ

For **Equilibrium constant,** we have,

$$\Delta G^{\Theta} = -2.303RT \log \ K$$

$$\log K = -\frac{\Delta G^{\Theta}}{2.303RT}$$

$$= -\frac{2895J}{2.303 \times 8.314JK^{-1}mol^{-1} \times 298K}$$

= - 0.507374

$\log K = -0.507374$

K = $10^{-0.507374}$

K = 0.311

14. It is ratio of distance between electrodes (l) to the cross sectional area between electrodes.
    Cell constant $= \frac{l}{A}$
    Or conductivity (K) = Conductance (G) $\times$Cell constant (G*)
    (K) = (G) $\times$ (G*)
    Cell constant of any particular cell can be found by measuring the conductance of a solution whose conductivity is known.

15. Strong electrolysis: Those electrolytes which dissociate into ions completely into aqueous solution are called strong electrolytes. for example:
    KCl, NaOH, H_2SO_4

    Weak electrolytes: Those electrolytes which do not dissociate into ions completely into aqueous solution are weak electrolytes. for example: CH_3COOH, NH_4Cl.

CBSE Test Paper-05

Class 12 Chemistry (Electrochemistry)

1. The standard emf of galvanic cell involving 3 moles of electrons in its redox reaction is 0.59 V. The equilibrium constant for the reaction of the cell is
   a. 10^{25}
   b. 10^{30}
   c. 10^{15}
   d. 10^{20}
2. The standard reduction potential E_o for half reactions are
   $E^0_{cell} = E^0_{cathode} - E^0_{anods}$
   The EMF of the cell reaction Fe^{2+} + Zn = Zn^{2+} + Fe is--- [Given E^0Zn^{2+}/Zn = -0.76V; E^0Fe^{2+}/Fe = -0.44V]
   a. -1.17 V
   b. -0.32 V
   c. + 0.32 V
   d. +1.17 V
3. An increase in equivalent conductance of a strong electrolyte with dilution is mainly due to
   a. increase in concentration of electrolyte
   b. decrease in ionic mobility
   c. increase in ionic mobility and number of ions
   d. decrease in both i.e. number of ions and mobility of ions
4. Electrolytic conduction is due to the movement of:
   a. molecules
   b. ions
   c. atoms
   d. electrons
5. Relationship between equilibrium constant of the reaction and standard electrode potential of electrochemical cell in which that reaction takes place is
   a. $E^0_{cell} = \frac{2.303RT}{nF} \log K_c$

   b. $E^0_{cell} = \frac{2.03RT}{nF} \log K_c$
   c. $E^0_{cell} = \frac{2.230RT}{nF} \log K_c$
   d. $E^0_{cell} = \frac{2.303RT}{F} \log K_c$

6. What do you understand by corrosion?

7. Two metals A and B have reduction potential values -0.76 V and +0.34 V respectively. Which of these will liberate H_2 from dil H_2SO_4?

8. Define the term specific resistance and give its SI unit.

9. The resistance of a conductivity cell containing 0.001 M KCl solution at 298 K is $1500\,\Omega$. What is the cell constant if conductivity of 0.001 M KCl solution at 298 K is $0.146 \times 10^{-3} Scm^{-1}$.

10. How much electricity in terms of Faraday is required to produce
    i. 20.0 g of Ca from molten $CaCl_2$.
    ii. 40.0 g of Al from molten Al_2O_3.

11. Calculate E_{cell} for following:
    $2Cr(s) + 3Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Fe(s)$
    $Cr(s)|Cr^{3+}(aq)(0.1M)||Fe^{2+}(aq)(0.01M)|Fe(s)$
    $E^{\ominus}_{(Cr^{3+}/Cr)} = -0.74\ V$
    $E^{\ominus}_{(Fe^{2+}/Fe)} = -0.44\ V$

12. What type of a cell is the lead storage battery? Write the anode and cathode reactions and the overall reaction occurring in a lead storage battery while operating.

13. Write mathematical expression for Kohlrausch's law.

14. What is a salt bridge? What is it used for?

15. Calculate the standard cell potentials of galvanic cells in which the following reactions take place:
    i. $2Cr(s) + 3Cd^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cd$
    ii. $Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$
    Calculate the$\Delta_r G^{\Theta}$, and equilibrium constant of the reactions.

CBSE Test Paper-05

Class 12 Chemistry (Electrochemistry)

Solutions

1. b. 10^{30}

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

   0.59 = 0.0591/3 log K

   log K = 3 × 0.59/0.0591 = 30

   K = AL 30 = 10^{30}

2. 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.

3. a. increase in ionic mobility and number of ions

   Explanation: Equivalent conductance increases on dilution for a strong electrolyte as interionic attraction also decreases along with dilution. So ionic mobility increases which in turn increases the equivalent conductance.

4. c. atoms

   Explanation: atoms

5. a. $E^0_{cell} = \frac{2.303RT}{nF} \log K_c$

   Explanation: ΔG^0 = - 2.303 RT log Kc ; ΔG^0 = -nFE^0cell

   Equating, $E^0_{cell} = \frac{2.303RT}{nF} \log Kc$

6. Corrosion is an electrochemical phenomenon in which metal gets decomposed in the presence of air and water and forms compounds like oxides, sulphates, carbonates, sulphides etc.
7. Metals having higher oxidation potential (or Lower reduction poetial) will liberate H_2 from H_2SO_4. Thus, A will liberate H_2 from H_2SO_4.
8. The specific resistance of a substance is its resistance when it is one-meter-long and its area of cross Section is one m^2. Its SI unit is Ωm(ohm meter).
9. Given,

Conductivity, $K = 0.146 \times 10^{-3} Scm^{-1}$

Resistance, R = $1500\,\Omega$

Therefore, Cell constant = $K \times R$

$= 0.146 \times 10^{-3} \times 1500 = 0.219 cm^{-1}$

10. i. According to the question,

$Ca^{2+} + 2e^{-1} \rightarrow \underset{40\,g}{Ca}$

Electricity required to produce 40 g of calcium = 2 F

Therefore, electricity required to produce 20 g of calcium = $\frac{2\times20}{40}F$

= 1 F

ii. According to the question,

$Al^{3+} + 3e^{-1} \rightarrow \underset{27g}{Al}$

Electricity required to produce 27 g of Al = 3 F

Therefore, electricity required to produce 40 g of Al = $\frac{3\times40}{27}F$

= 4.44 F

11. We have

2Cr (s) + $3Fe^{2+}$(aq)$\rightarrow$ $2Cr^{3+}$(aq) + 3Fe(s)

$Cr(s)|Cr^{3+}(aq)(0.1M)||Fe^{2+}(aq)(0.01M)|Fe(s)$

Anode half reaction:

$[Cr \rightarrow Cr^{3+} + 3e^{-}] \times 2$

Cathode half reaction:

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

Net cell reaction:

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

Also standard emf of the cell is

$E^{\Theta}_{cell} \;=\; E^{\Theta}_{(Fe^{2+}/Fe)} \;-\; E^{\Theta}_{(Cr^{3+}/Cr)}$

= -0.44-(-0.74) = 0.30 V

According to Nernst equation

[Here **n=6** moles of electrons]

$$E_{cell} = E^{\Theta}_{cell} - \frac{0.059}{n}\log\frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$$

$$E_{cell} = E^{\Theta}_{cell} \;-\; \frac{0.059}{6}\log\left[\frac{(0.1)^2}{(0.01)^3}\right]$$

$= 0.30 \ - \ \frac{0.059}{6} \log \left[\frac{10^{-2}}{10^{-6}}\right]$

$= 0.30 - \frac{0.059}{6} \log 10^4$

$= 0.3 - \frac{0.059}{6} \times 4$

= 0.30 - 0.039 = 0.261 V

$\therefore E_{cell} \ = \ 0.261 \ V$

12. The lead storage battery is the most important secondary cell. The cell reaction when the battery is in use are given below:

    At anode: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$

    At cathode: $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow 2PbSO_4(s) + 2H_2O(l)$

    The overall cell reaction is:

    $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$
13. Mathematical expression for Kohlrausch's law is

    Λ_m^∞ or μ^∞ = Molar conductance at infinite dilution

    $= m\lambda_+^\infty + n\lambda_-^\infty$

    Where m and n are number of ions formed.
14. A salt bridge is a U-shaped tube containing concentrated solution of an inert electrolyte like KCl, KNO_3 etc. or solidified solution of such an electrolyte in agar-agar and gelatine.

    It is used for:

    i. To complete the electrical circuit by allowing ions to flow from one solution to the other without mixing the two solutions.

    ii. To maintain the electrical neutrality of the solution in the two half cells.
15. i. $E^\Theta{}_{Cr^{3+}/Cr} = 0.74V$

    $E^\Theta{}_{Cd^{2+}/Cd} = 0.40V$

    The galvanic cell of the given reaction is depicted as:

    $Cr_{(s)}|Cr^{3+}{}_{(aq)}||Cd^{2+}{}_{(aq)}|Cd_{(s)}$

    Now, the standard cell potential is $E_{cell}^\Theta = E_R^\Theta - E_L^\Theta$

    = 0.40 - (-0.74)

    = +0.34 V

    $\Delta_r G^\Theta = -nFE_{cell}^\Theta$

    In the given equation,

    n = 6

F = 96487 C mol^{-1}

E_{cell}^{Θ} = +0.34 V

Then, $\Delta_r G^{\Theta} = -6 \times 96487 C mol^{-1} \times 0.34 V$

= -196833.48 CV mol^{-1}

= -196833.48 J mol^{-1}

= -196.86 kJ mol^{-1}

Again, $\Delta_r G^{\Theta} = -RT\, ln\, K$

$\Delta_r G^{\Theta} = -2.303\, RT\, ln\, K$

log K = $-\frac{\Delta_r G^{\Theta}}{2.303 RT}$

$= \frac{196.83 \times 10^3}{2.303 \times 8.314 \times 298}$

= 34.496

Therefore, K = antilog (34.496)

$= 3.13 \times 10^{34}$

ii. $E^{\Theta}{}_{Fe^{3+}/Fe^{2+}} = 0.77\, V$

$E^{\Theta}{}_{Ag^{+}/Ag} = 0.80\, V$

The galvanic cell of the given reaction is depicted as:

$Fe^{2+}{}_{(aq)} | Fe^{3+}{}_{(aq)} || Ag^{+}{}_{(aq)} | Ag_{(s)}$

Now, the standard cell potential is $E_{cell}^{\Theta} = E_{R}^{\Theta} - E_{L}^{\Theta}$

= 0.80 - 0.77 = 0.03 V

Here, n = 1.

Then,$\Delta_r G^{\Theta} = -nFE_{cell}^{\Theta}$

=$-1 \times 96487 C mol^{-1} \times 0.03 V$

= -2894.61 J mol^{-1}

= -2.89 kJ mol^{-1}

Again,$\Delta_r G^{\Theta} = 2.303\, RT\, In\, K$

log $K = -\frac{\Delta_r G}{2.303\, RT}$

$= \frac{-2894.61}{2.303 \times 8.314 \times 298}$

= 0.5073

Therefore, K = antilog (0.5073)

= 3.2 (approximately)