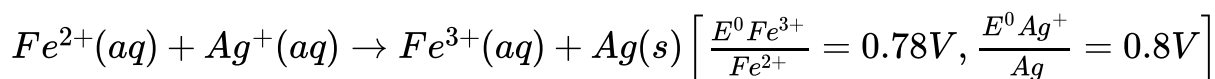


CBSE Test Paper-04
Class 12 Chemistry (Electrochemistry)

1. Calculate the standard cell potentials of galvanic cell, $\Delta_r G^\circ$ and equilibrium constant of the reactions if the reaction is



- 0.04V, -2.955 kJ, 3.2
 - 0.02V, -1.930 kJ, 2.180
 - 0.01V, -2.800 kJ, 3.2
 - 0.02V, -2.850 kJ, 3.2
2. The SI units of molar conductance are:
- Sm^3mol^{-1}
 - $Sm^{-1}mol^{-1}$
 - Sm^2mol^{-1}
 - $Sm^{-2}mol$
3. Li occupies higher position in the electrochemical series of metals as compared to Cu since
- Li is smaller in size as compared to Cu
 - the standard oxidation potential of Li^+/Li is lower than that of Cu^{2+}/Cu
 - the standard reduction potential of Li^+/Li is lower than that of Cu^{2+}/Cu
 - 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:
- 0.40 V
 - 0.80 V
 - 0.80 V
 - 0.40 V
5. Electrochemistry deals with the production of electricity from
- 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 $T=298 \text{ K}$.
$$Fe(s) + Cd^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cd(s)$$

$$\left[\text{Given, } E_{Cd^{2+}/cd}^{\ominus} = -0.40V, E_{Fe^{2+}/Fe}^{\ominus} = -0.44V \right]$$
13. The cell in which the following reaction occurs:
$$Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$$
 has $E^{\ominus}_{(Fe^{3+}/Fe^{2+})} = +0.77 \text{ V}$ and $E^{\ominus}_{(Ag^{+}/Ag)} = +0.80 \text{ V}$ at 298 K. Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction. ($F = 96500 \text{ C mol}^{-1}$)
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: $E_{cell}^0 = E_{cathode}^0 - E_{anode}^0$

$$= \frac{E^0_{Ag}}{Ag} - \frac{E^0_{Fe^{3+}}}{Fe^{2+}} = +0.8 - 0.78 = +0.02V$$

$$\Delta_r G^0 = -nFE^0_{cell} = -1 \times 96500 \times 0.02 = -1930J = 1.930KJ$$

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

$$\log K = \frac{n \times E^0_{cell}}{0.0591} = \frac{1 \times 0.02}{0.0591} = 0.3384$$

$$k = 10^{0.3384} = 2.180$$

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

Explanation: Sm^2mol^{-1}

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

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

$$0.8 = E^0_{Ag^+/Ag} - E^0_{H^+/H_2}$$

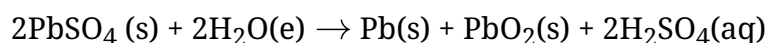
$$0.8 = E^0_{Ag^+/Ag} - 0$$

$$E^0_{Ag^+/Ag} = 0.8V$$

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 -



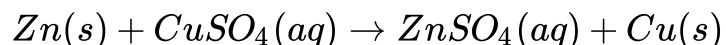
7. When the concentration of all the species involved in a half-cell is unity, then the electrode potential is called standard electrode potential.

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

9. Zinc is more reactive than copper.

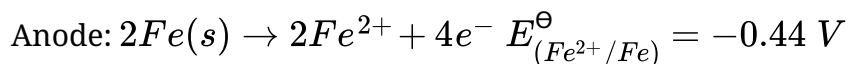
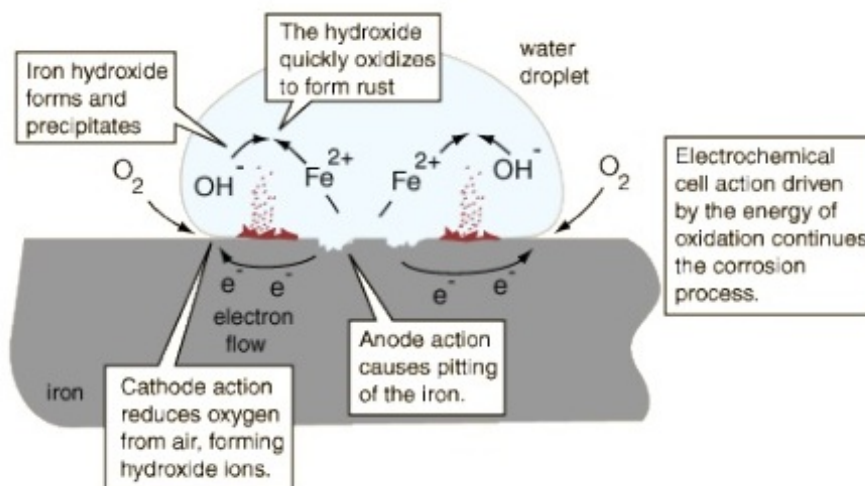


Hence, it displaces copper from copper sulphate solution as follows.



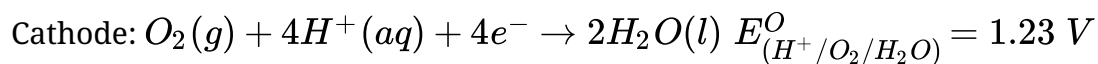
We cannot store copper sulphate sol. in a zinc pot.

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

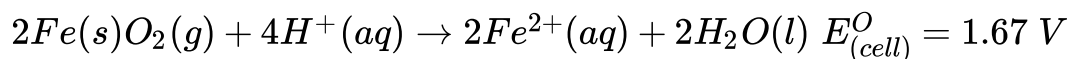


Electrons released at anode spot move through the metal and go to another spot on the metal and reduce oxygen in presence of H^+ (which is believed to be available from H_2CO_3 formed due to dissolution of carbon dioxide from air into water. Hydrogen ion in water may also be available due to dissolution of other acidic oxides from their

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



The overall reaction being:



The ferrous ion are further oxidized by atmospheric oxygen to ferric ion which comes out as rust in the form of hydrated ferric oxide ($Fe_2O_3 \cdot x H_2O$) 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.

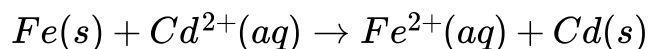
11. Given,

$$K = 0.0248 \text{ S cm}^{-1}$$

$$c = 0.20 \text{ M}$$

$$\begin{aligned} \text{Therefore, Molar conductivity, } A_m &= \frac{K \times 1000}{c} \\ &= \frac{0.0248 \times 1000}{0.2} \\ &= 124 \text{ S cm}^2 \text{ mol}^{-1} \end{aligned}$$

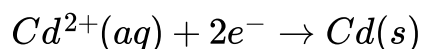
12. We have,



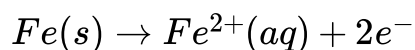
In this reaction, iron is oxidized and cadmium is reduced.

Half cell reaction of this cell is:

At Cathode (reduction):



At Anode (oxidation):



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

Now, we have

$$E_{cell}^{\ominus} = E_{Cd^{2+}/Cd}^{\ominus} - E_{Fe^{2+}/Fe}^{\ominus}$$

$$= -0.40 - (-0.44) = 0.04 \text{ V}$$

We know that,

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

for n=2 and at T=298 K

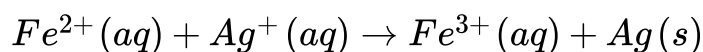
$$E_{cell}^{\ominus} = \frac{0.059}{2} \log K_c = 0.04$$

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

$$K_c = \text{Antilog } 1.356$$

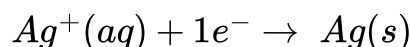
$$K_c = 2.270 \times 10^1 = 22.7$$

13. Given that the cell reaction is

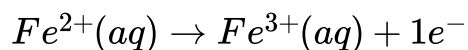


In this cell, Fe^{2+} is oxidized to Fe^{3+} and Ag^+ is reduced to $Ag(s)$. Therefore

At Cathode (Reduction):



At Anode (Oxidation):



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

Standard emf of the cell is

$$E_{cell}^{\ominus} = E_{reduction}^{\ominus} - E_{oxidation}^{\ominus}$$

$$= +0.80 - (+0.77)$$

$$= +0.03 \text{ V}$$

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

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

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

$$= -2895 \text{ J}$$

$$= -2.895 \text{ kJ}$$

For **Equilibrium constant**, we have,

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

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

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

$$= -0.507374$$

$$\log K = -0.507374$$

$$K = 10^{-0.507374}$$

$$K = 0.311$$

14. It is ratio of distance between electrodes (l) to the cross sectional area between electrodes.

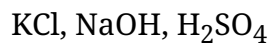
$$\text{Cell constant} = \frac{l}{A}$$

$$\text{Or conductivity (K)} = \text{Conductance (G)} \times \text{Cell constant (G}^*)$$

$$(K) = (G) \times (G^*)$$

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:



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