

Electrochemistry

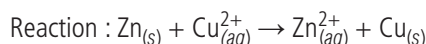


TRY YOURSELF

ANSWERS

1. K_2SO_4 , potassium sulphate is used as an electrolyte in the salt bridge of a Daniell cell.

2. The device in which electrical energy is produced from chemical reaction are called voltaic cells.



3. The cell reaction is split into two half reaction



When oxidation potential of both electrode taken into account.

$$E_{\text{cell}}^{\circ} = \text{Oxidation potential of anode} - \text{Oxidation potential of cathode}$$

$$= E_{\text{ox(anode)}}^{\circ} - E_{\text{ox(cathode)}}^{\circ}$$

$$= 0.76 \text{ V} - (-0.80) \text{ V} = 1.56 \text{ V}$$

4. Oxidation or reduction reactions cannot take place alone. Electrode potential of an electrode is a relative tendency and can be measured with respect to a reference electrode only.

$$5. \quad E_{\text{cell}}^{\circ} = E_{\text{Cathode}}^{\circ} - E_{\text{Anode}}^{\circ}$$

The standard reduction potential value of Cu is higher, it act as cathode whereas, the standard reduction potential value of Zn is low it act as anode.

$$E_{\text{cell}}^{\circ} = 0.34 - (-0.76) = 1.10 \text{ V}$$

$$6. \quad E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{6F} \ln \frac{[Cr^{3+}]^2}{[Cd^{2+}]^3} \quad (\because n=6)$$

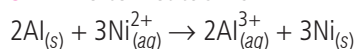
$$7. \quad \Delta G^{\circ} = -nFE_{\text{cell}}^{\circ} = -6 \times 96500 \times 0.89 = -515.3 \text{ kJ}$$

8. We know that,

$$E_{\text{cell}}^{\circ} = \frac{0.0591}{n} \log K_c, \text{ here value of } n = 2$$

$$\log K_c = \frac{0.235 \times 2}{0.0591} \Rightarrow K_c = 8.966 \times 10^7$$

9. The cell reaction is



$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log \frac{[Al^{3+}]^2}{[Ni^{2+}]^3}$$

$$E_{\text{cell}}^{\circ} = E_{\text{(Ni}^{2+}/\text{Ni)}}^{\circ} - E_{\text{(Al}^{3+}/\text{Al)}}^{\circ}$$

$$= -0.25 \text{ V} - (-1.66 \text{ V}) = 1.41 \text{ V}$$

$$n = 6$$

Substituting the values

$$E_{\text{cell}} = 1.41 - \frac{0.059}{6} \log \frac{(0.001)^2}{(0.50)^3}$$

$$= 1.41 + 0.05$$

$$= 1.46 \text{ V}$$

10. On increasing temperature,

(i) the electrical conduction of metals decreases

(ii) the electrical conduction of electrolytes increases.

11. Copper is conducting in nature because it contains free electrons. $CuSO_4$ is conducting in molten state or in aqueous solution because it gives ions only in the molten state or aqueous solution.

12. In metallic substances, flow of electricity is due to flow of electrons while in ionic substances in the aqueous solution or molten state, the flow of electricity is due to flow of ions.

$$13. \quad \kappa = \frac{1}{R} \times \frac{l}{A}$$

where, κ is the conductivity, R is resistance and l/A is the cell constant.

14. In finding the resistance of the solution of an electrolyte, the difficulty faced is that the solution cannot be connected in the bridge like a metallic wire or any solid conductor.

Hence, conductivity cells are used such that the solution is present between the two electrodes.

$$15. \quad R = \rho \frac{l}{A}$$

$$\kappa = \frac{1}{R} \frac{l}{A}$$

$$\kappa \propto \frac{1}{R}$$

Thus, solution B which has low conductivity with respect to solution A and will offer greater resistance to the flow of electrons.

16. Conductivity, $\kappa = 0.146 \times 10^{-3} \text{ S cm}^{-1}$

Resistance, $R = 1500 \text{ ohm}$

$$\text{Cell constant} = \frac{\text{Conductivity, } (\kappa)}{\text{Conductance } (G)}$$

$$= \text{Conductivity } (\kappa) \times \text{Resistance } (R)$$

$$\therefore \text{Cell constant} = 0.146 \times 10^{-3} \text{ ohm}^{-1} \text{ cm}^{-1} \times 1500 \text{ ohm} \\ = 0.219 \text{ cm}^{-1}$$

17. Molar conductivity of weak electrolytes keep on increasing with dilution and does not become constant even at very large dilutions.

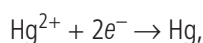
18. HCl because H^+ ions are smaller than Na^+ ions and hence H^+ ions have greater ionic mobility than Na^+ ions.

19. Conductivity (κ) decreases whereas, conductance (G), molar conductivity (Λ_m) and equivalent conductivity (Λ_{eq}) increase with dilution.

$$20. \Lambda_m (\text{Molar conductivity}) = \frac{\kappa (\text{Conductivity})}{M (\text{Molarity})} = \frac{\kappa \times V (\text{Volume})}{1 \text{ mole}}$$

Solution B will have greater molar conductivity with dilution, κ decreases but V increases much more so that product increases.

21. The reaction is represented as



$$Q = I \times t$$

$$Q = 2 \times (3 \times 3600) = 21600 \text{ C}$$

$2F = 2 \times 96500 \text{ C}$ will produce 1 mole Hg

$$\therefore 21600 \text{ C will produce Hg} = \frac{21600}{2 \times 96500} = 0.112 \text{ mol}$$

22. Quantity of electricity passed $Q = I \times t$

$$Q = 0.2 \times 3 \times 60 \times 60 = 2160 \text{ C}$$

For the reaction, $Ag^+ + e^- \rightarrow Ag$

96500 C deposit Ag = 107.92 g

$$2160 \text{ C will deposit Ag} = \frac{107.92}{96500} \times 2160 \text{ g} = 2.415 \text{ g}$$

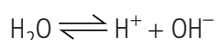
23. $Q = I \times t = 1.2 \times 15 \times 60 = 1080 \text{ C}$

1080 C of charge deposit = 0.3605 g

$$96500 \text{ C of charge will deposit} = \frac{0.3605}{1080} \times 96500 \text{ C} = 32.21 \text{ g}$$

$$\text{Valency} = \frac{\text{Atomic weight}}{\text{Equivalent weight}} = \frac{96}{32.21} = 2.98 \approx 3$$

24. For the reaction, $H_2SO_{4(aq)} \rightleftharpoons 2H^+_{(aq)} + SO_4^{2-}_{(aq)}$



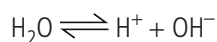
At cathode : $H^+ + e^- \rightarrow H$, $H + H \rightarrow H_{2(g)}$

At anode : $OH^- \rightarrow OH + e^-$, $4OH \rightarrow 2H_2O_{(l)} + O_{2(g)}$

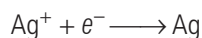
Therefore, H_2 is liberated at cathode and O_2 at the anode.

25. The H^+ ions produced from ionisation of water have lower discharge potential than Na^+ ions produced from ionization of NaCl or reduction potential of water is greater than that of sodium.

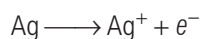
26. For the reaction, $AgNO_{3(aq)} \rightarrow Ag^+_{(aq)} + NO_3^-_{(aq)}$



At cathode : Ag^+ ions have lower discharge potential than H^+ ions. Hence, Ag^+ ions will be deposited on the cathode.



At anode : Ag anode will be attacked by NO_3^- ions. Hence Ag anode will dissolve to form Ag^+ ions in the solution



27. Those cells are rechargeable in which the products formed during discharge are deposited on the electrode and these can be decomposed to give the original substances when electrical energy is supplied.

28. With time, concentrations of the electrolytic solutions change. Hence, their electrode potentials change. When the electrode potentials of the two half-cells become equal, the cell stops working.

29. The cells which can be recharged by passing an electric current through them and hence can be used over and over again. The two known examples of this type of cells are lead storage battery and nickel-cadmium storage cell.

30. At anode : $2H_{2(g)} + 4OH^-_{(aq)} \rightarrow 4H_2O_{(l)} + 4e^-$

At cathode : $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$

Overall reaction : $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

31. (i) They do not cause any pollution unlike thermal plants which burn fossil fuels like coal, gas oil etc.

(ii) The fuel cells provide better efficiency as compared to thermal power plants in which fuels are burnt to produce heat which then changes water into steam to run the turbine.

(iii) Because of continuous supply, such cells never become dead.

32. (i) The corrosiveness of the electrolytes used.

(ii) High cost of the catalyst needed for the electrode reactions.

(iii) Problem of handling gaseous fuel at low temperatures or high pressures.

(iv) Providing of contact between the three phase needed in a fuel cell i.e., the gaseous fuel, the liquid electrolyte and the solid electrolyte.

33. $2Fe_{(s)} + O_{2(g)} + 2H^+_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)}$

34. Zinc protects better because oxidation potential of zinc is greater but that of tin is less than that of iron.

35. Rusting of iron takes place in presence of H^+ ions. Alkaline medium neutralizes the H^+ ions thereby inhibits rusting.

