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ELECTRO CHEMISTRY XII (CHEMISTRY) CHAPTER - 3

1. If the cathode in a Cu–Zn cell is replaced by an Ni rod immersed in a 1 M solution of $NiSO_4$, how will the potential of the cell be affected? Will the potential of the Ni–Zn cell be greater, smaller or the same as that of the Cu–Zn cell?
2. Consider the cell

   $Zn \mid Zn^{2+}$ (0.01 M) $\|$ Fe^{2+} (0.01 M), Fe^{3+} (0.01 M) $\mid$ Pt

   (a) Write the equation for the spontaneous cell reaction.

   (b) Which way do electrons flow in the external circiut during the spontaneous reaction?

   (c) Write the reaction at each electrode in the spontaneous reaction.

   (d) Evaluate $E^\ominus_{cell}$ at 25°C and also E_{cell} at 25°C using the Nernst equation.
3. For a cell in which the reaction

   $A + B \rightarrow C + D$

   takes place, the emf is negative. What will be the direction of spontaneity of the reaction? Explain.
4. What do you understand by corrosion?
5. What is cathodic protection? How is it different from galvanising?
6. What is electrolysis?
7. What are the products of electrolysis of the following?

   (a) Molten NaCl (b) Aqueous solution of NaCl
8. A voltaic cell with a basic aqueous electrolyte is based on the oxidation of Cd(s) to $Cd(OH)_2$(s) and the reduction of MnO_4^-(aq) to MnO_2(s).

   (a) Write the half-reactions for the cells at the anode and cathode.

   (b) Write the overall balanced cell reaction.

   (c) Draw a diagram of the cell indicating all the details.

Long-Answer Questions

1. Describe the circuit for measuring conductance.
2. Explain the term emf of a cell.
3. Describe the standard hydrogen electrode.
4. How is a cell represented?
5. Describe the two types of electrochemical cells, giving examples.
6. Explain corrosion as an electrochemical process.
7. How can you prevent corrosion?
8. Write short notes on (i) galvanisation, and (ii) cathodic protection.
9. Write a short note on fuel cells.
10. Give the cell reactions for the following cells.

    (a) Pt(s) $\mid Fe^{2+}, Fe^{3+} :: Cl^- \mid$ AgCl(s) $\mid$ Ag(s)

    (b) Pt(s) $\mid H_2$(g, 1 atm) $\mid H_2SO_4 \mid PbSO_4$(s) $\mid$ Pb(s)

    (c) Pt(s) $\mid H_2$(g, 1 atm) $\mid$ HCl $\mid Hg_2Cl_2$(s) $\mid$ Hg(l)

    (d) $Zn \mid Zn^{2+} \| Fe^{3+}, Fe^{2+} \mid Pt$

    (e) $Ag \mid Ag^+ \mid Br^- \mid AgBr(s) \mid Ag$

Numericals

1. The conductivity of a 0.1 M KCl solution is 1.289 $S\ m^{-1}$ at 25°C. What is the resistance of a conductance cell containing this solution, given that the electrodes are separated by 0.5 cm and have an area of cross section of 1 cm^2.
2. The conductance of a 0.1 M KCl solution was found to be 0.012 S. If the molar conductivity is 0.06 $S\ m^2\ mol^{-1}$, calculate the cell constant of the cell.

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3. The conductance of a 0.01 M KCl solution is 0.00141 S. Calculate its molar conductivity if the cell constant is $100\ m^{-1}$.
4. Calculate the limiting molar conductivity of ammonium hydroxide (NH_4OH) if the limiting molar conductivities of NaCl, NaOH and NH_4Cl are 126.45, 247.8 and 149.7 $S\ cm^2\ mol^{-1}$ respectively.

   [Hint: Write all the formulae in ionised form and obtain $NH_4^+ + OH^-$ from a combination of NaCl, NaOH and NH_4Cl.

   $$\Lambda^0_{NH_4OH} = \lambda^0_{NH_4^+} + \lambda^0_{OH^-}\ ;\ \Lambda^0_{NH_4Cl} = \lambda^0_{NH_4^+} + \lambda^0_{Cl^-}\ ;\ \Lambda^0_{NaCl} = \lambda_{Na^+} + \lambda_{Cl^-}\ ;\ \Lambda^0_{NaOH} = \lambda^0_{Na^+} + \lambda^0_{OH^-}$$

   Hence $\Lambda^0_{NH_4OH} = \Lambda^0_{NH_4Cl} - \Lambda^0_{NaCl} + \Lambda^0_{NaOH}$.]
5. Calculate the degree of dissociation of a 0.01 M solution of an acid whose molar conductivity is $0.00163\ S\ m^2\ mol^{-1}$ and molar conductivity at infinite dilution (Λ^0_m) is $0.03907\ S\ m^2\ mol^{-1}$. Also find the equilibrium constant for the dissociation of this acid.
6. The dissociation constant of a 0.05 M solution of a weak acid was found to be 1×10^{-5}. What per cent of the acid remained undissociated at this concentration? If the molar conductivity at infinite dilution is $0.039\ S\ m^2\ mol^{-1}$, what is the molar conductivity of this acid solution?
7. Calculate the molar conductivity at infinite dilution for a weak acid HX given that the conductance of a 0.05 M solution of it is 1.5×10^{-5} S using a cell of cell constant $1.5\ cm^{-1}$ and that the acid dissociation constant is 1×10^{-4}.
8. Calculate the amount of Cu deposited at an electrode due to the reaction $Cu^{2+} + 2e \rightarrow Cu$ when 1 A of current is passed through the solution for 30 min. The atomic mass of Cu is 63.5.
9. How much time is required to deposit 0.5 mol of Al according to the electrode reaction $Al^{3+} + 3e^- \rightarrow Al$ when 1.5 A of current is passed through the cell?
10. How much charge has to be passed through a cell containing a Zn^{2+}/Zn electrode so that 130.8 g of Zn is deposited at the electrode? Atomic mass of Zn = 65.4.
11. How much charge has to be passed through a cell containing molten NaCl so that 2.5 mol of chlorine gas is evolved at the corresponding electrode?
12. What is the mass of chlorine produced by the electrolysis of molten NaCl when a current of 5 A is passed through the melt for 20 min? Atomic mass of chlorine = 35.5.

    [Note: Chlorine is liberated as Cl_2]
13. How much time is required for the deposition of 98.1 g of Zn according to the electrode reaction $Zn^{2+} + 2e \rightarrow Zn$ when 2 A current is passed through the cell. Atomic mass of zinc = 65.4.
14. How much current should be passed for one hour through a cell containing a Cu^{2+}/Cu electrode so that 4 g of Cu is deposited at the electrode. Atomic mass of Cu = 63.5.
15. How many hours does it take to reduce 3 mol of Fe^{3+} to Fe^{2+} with 2 A of current? [CBSE]
16. The standard potentials for the two electrodes in a cell are given as $E^\ominus_{Cu^{2+}/Cu} = 0.34$ V; $E^\ominus_{Ag^+/Ag} = 0.8$ V. Calculate the cell potential (E) for the cell containing 0.1 M Ag^+ and 4 M Cu^{2+} at 298 K. [CBSE]
17. Calculate the potential of a cell in which the following cell reaction occurs at 298 K.

    $$Sn^{4+}\ (1.5\ M) + Zn(s) \rightarrow Sn^{2+}\ (0.5\ M) + Zn^{2+}\ (2\ M)$$

    The standard cell potential $E^\ominus$ of the cell is 0.89 V. State whether the cell potential will increase or decrease if the concentration of Sn^{4+} is increased in the cell. ($R = 8.314\ J\ K^{-1}\ mol^{-1}$, $F = 96{,}500\ C\ mol^{-1}$). [CBSE]
18. Consider a cell composed of two half-cells $Cu(s) | Cu^{2+}(aq)$ and $Ag(s) | Ag^+(aq)$. Calculate (a) the standard cell potential and (b) the cell potential when $[Cu^{2+}]$ is 2 M and $[Ag^+]$ is 0.05 M.

    Given $E^\ominus(Cu^{2+}/Cu) = 0.34$ V

    $E^\ominus(Ag^+/Ag) = 0.80$ V, $R = 8.314\ J\ K^{-1}\ mol^{-1}$

    $F = 96{,}500\ C\ mol^{-1}$. {CBSE]
19. Calculate the emf of the following cell at 298 K.

    $$Cd | Cd^{2+}\ (0.1\ m) \| Ag^+\ (0.1\ m) | Ag$$

    Given $E^\ominus_{Cd^{2+}/Cd} = -0.4$ V; $E^\ominus_{Ag^+/Ag} = 0.8$ V. [CBSE]

20. Calculate the emf of a cell with the following cell reaction.

$$Co^{2+}(aq) + Ba(s) \rightarrow Co(s) + Ba^{2+}(aq)$$

Given that $E^{\ominus}_{Co^{2+}/Co} = -0.28$ V and $E^{\ominus}_{Ba^{2+}/Ba} = -2.9$ V.

The concentrations of $Ba^{2+}(aq)$ and $Co^{2+}(aq)$ are 1×10^{-5} M and 0.1 M respectively. [CBSE

21. Calculate the emf of a cell with the following cell reaction.

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

Given that $E^{\ominus}_{Cu^{2+}/Cu} = 0.34$ V and $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V.

The concentrations of Cu^{2+} and Ag^{+} are 0.05 M and 0.01 M respectively. [CBSE

22. Calculate the emf of a cell whose half-cell reactions are as follows.

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

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

The standard reduction potentials of the two half-cells are $E^{\ominus}_{Zn^{2+}/Zn} = -0.76$ V and $E^{\ominus}_{Cu^{2+}/Cu} = 0.34$ V.

The concentrations of Zn^{2+} and Cu^{2+} are both 0.1 M.

23. Calculate the emf of the following cell at 298 K.

$$Co \,|\, Co^{2+}\ (0.1\ M) \,||\, Ag^{+}\ (0.1\ M) \,|\, Ag$$

Given that $E^{\ominus}_{Co^{2+}/Co} = -0.28$ V and $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V. [CBSE

24. Calculate the emf of the following cell at 298 K.

$$Fe \,|\, Fe^{2+}\ (0.1\ M) \,||\, Ag^{+}\ (0.1\ M) \,|\, Ag$$

Given that $E^{\ominus}_{Fe^{2+}/Fe} = -0.44$ V and $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V. [CBSE

25. Calculate the emf of the following cell at 298 K.

$$Sn \,|\, Sn^{2+}\ (0.1\ M) \,||\, Ag^{+}\ (0.1\ M) \,|\, Ag$$

Given that $E^{\ominus}_{Sn^{2+}/Sn} = -0.14$ V; $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V. [CBSE

26. Calculate the emf of the following cell.

$$Mg \,|\, Mg^{2+}\ (0.001\ M) \,||\, Cu^{2+}\ (0.0001\ M) \,|\, Cu$$

Given that $E^{\ominus}_{Cu^{2+}/Cu} = 0.337$ V and $E^{\ominus}_{Mg^{2+}/Mg} = -2.37$ V.

Also determine the value of the standard free energy change ($\Delta G^{\ominus}$) for the cell. [CBSE

27. The emf of a cell whose cell reaction at 298 K is

$$Mg + 2Ag^{+} \rightarrow Mg^{2+} + 2Ag$$

is 3.38 V. Given that the concentration of Mg^{2+} ions is 25 times that of Ag^{+} ions and that the standard reductio potentials are $E^{\ominus}_{Mg^{2+}/Mg} = -2.37$ V and $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V, find the concentration of Ag^{+} ions and Mg^{2+} ions in the cell.

28. The potential of the cell

$$Pt(s) \,|\, H_2(g, 1\ atm) \,|\, H^{+}\ (1\ M) \,|\, Fe^{3+}, Fe^{2+} \,|\, Pt(s)$$

is 0.8 V. If the concentration of Fe^{2+} is 0.0323 M, what is the concentration of Fe^{3+}? $E^{\ominus}_{Fe^{3+}/Fe^{2+}} = 0.771$ V.

29. Calculate the cell potential of the given cell at 25°C ($R = 8.34$ J K^{-1} mol^{-1}, $F = 96{,}500$ C mol^{-1}).

$$Ni(s) \,|\, Ni^{2+}\ (0.01\ M) \,||\, Cu^{2+}\ (0.1\ M) \,|\, Cu(s)$$

Given that $E^{\ominus}_{Cu^{2+}/Cu} = 0.34$ V and $E^{\ominus}_{Ni^{2+}/Ni} = -0.25$ V. [CBSE

30. For what concentration of $Ag^{+}(aq)$ will the emf of the given cell be zero at 25°C, if the concentration of $Cu^{2+}(aq)$ i 0.1 M?

$$Cu(s) \,|\, Cu^{2+}\ (0.1\ M) \,||\, Ag^{+}(aq) \,|\, Ag(s)$$

Given that $E^{\ominus}_{Ag^{+}/Ag} = 0.8$ V and $E^{\ominus}_{Cu^{2+}/Cu} = 0.34$ V. [CBSE

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