CHEMISTRY (XI)

Chapter 10

Electrochemistry

Short Questions

1. What is Electrochemistry?

Ans: Electrochemistry is concerned with the conversion of electrical energy into chemical energy in electrolytic cells as well as the conversion of chemical energy into electrical energy as in galvanic or voltaic cells.

2. Define oxidation number or state.

Ans: It is the apparent charge on an atom of an element in a molecule or an ion. It may be positive or negative or zero. For example, Na^{+1}, S^{-2}, K^{+1}

3. State rules for assigning oxidation number (Mention any four as an answer to short question).

Ans: (i) The oxidation number of all elements in the free state is zero. This is often shown as a zero written on the symbol. For example, H_2^0, Na^0 , Mg^0

(ii) The oxidation number of an ion, consisting of a single element, is the same as the charge on the ion. For example, the oxidation number of K^+, Ca^{2+}, Al^{3+} , Br^- , S^{2-} are + 1, +2, +3, -1, -2, respectively.

(iii) The oxidation number of hydrogen in all its compounds except metal hydrides is +1. In metal hydrides it is -1. (Na^+H^- , Mg^{2+} H_2 $^{(-1)2}$)

(iv) The oxidation number of oxygen in all its compounds except in peroxides, OF_2 and in super oxides is -2. It is -1 in peroxides +2 in OF_2 and -1/2 in super oxides.

(v) In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.

(vi) In ions, the algebraic sum of oxidation number equals the charge on the ion.

(vii) In any substance the more electronegative atom has the negative oxidation number.

4. Calculate the oxidation no. of Mn in $KMnO_4$.

Ans: $KMnO_4$

$+1 + Mn + (-2)4 = 0$

$Mn + 1 - 8 = 0$

$Mn = +7$

5. Calculate the oxidation no. of Mn in Na_2MnO_4.

Ans: Na_2MnO_4

$(+1)2 + Mn + (-2)4 = 0$

$+2 + Mn - 8 = 0$

$Mn - 6 = 0$

$Mn = +6$

6. Calculate the oxidation no. of S in $Cr_2(SO_4)_3$ and SO_4^{-2}.

Ans: $\mathbf{Cr_2(SO_4)_3}$

$(+3)2 + (S)3 + (-2)12 = 0$

$+6 + (S)3 - 24 = 0$

$(S)3 - 18 = 0$

$S = +6$

$\mathbf{SO_4^{-2}}$

$S + (-2)4 = -2$

$S - 8 = -2$

$S = -2 + 8$

$S = +6$

7. Calculate the oxidation no. of Cr in $CrCl_3$.

Ans: $CrCl_3$

$Cr + (-1)3 = 0$

$Cr - 3 = 0$

$Cr = +3$

8. Calculate the oxidation no. of P in HPO_3.

Ans: HPO_3

$+1 + P + (-2)3 = 0$

$+1 + P - 6 = 0$

$P - 5 = 0$

$P = +5$

9. Calculate the oxidation no. of the elements underlined in the following compounds.

***Ans*:**

$Ca(\underline{Cl}O_3)_2$

Oxidation number of O =–2
Oxidation number of Ca = +2
Oxidation number of Cl = x

$$+2 + 2x + 6(-2) = 0$$
$$2x = +12 - 2 = +10$$
$$\mathbf{x = +5}$$

$Na_2\underline{C}O_3$

Oxidation number of O = –2
Oxidation number of Na = +1
Oxidation number of C = x

$$2(+1) + (x) + 3(-2) = 0$$
$$2 + x - 6 = 0$$
$$x = +6 - 2 = \mathbf{+4}$$

$Na_3\underline{P}O_4$

Oxidation number of O =–2
Oxidation number of Na = +1
Oxidation number of P = x = ?

$$3(+1) + x + 4(-2) = 0$$
$$+3 + x - 8 = 0$$
$$x = +8 - 3 = \mathbf{+5}$$

$H\underline{N}O_3$

Oxidation number of O = –2
Oxidation number of H = +1
Oxidation number of N = x = ?

$$(+1) + x + 3(-2) = 0$$
$$+1 + x - 6 = 0$$
$$x = +6 - 1 = \mathbf{+5}$$

$Cr_2(\underline{S}O_4)_3$

Oxidation number of O =–2
Oxidation number of Cr = +3

Oxidation number of S = x = ?

$$2(+3) + 3x + 12(-2) = 0$$
$$+6 + 3x - 24 = 0$$
$$3x = +2 - 6 = +18$$
$$x = \mathbf{+6}$$

(vi) **$H\underline{P}O_3$**

Oxidation number of O = –2
Oxidation number of H = +1
Oxidation number of P = x = ?

$$+1 + x + 3(-2) = 0$$
$$+1 + x - 6 = 0$$
$$x = +6 - 1 = \mathbf{+5}$$

(vii) **$K_2\underline{Mn}O_4$**

Oxidation number of O = –2
Oxidation number of K = +1
Oxidation number of Mn = x = ?

$$2(+1) + x + 4(-2) = 0$$
$$2 + x - 8 = 0$$
$$x = +8 - 2 = \mathbf{+6}$$

10. Define metallic conduction.

Ans: Most metals are conductors of electricity because of the relatively free movement of their electrons throughout the metallic lattice. This electronic conduction is simply called metallic conduction.

11. Define ionization. Give an example.

Ans: Ionization is the process in which ionic compounds when fused or dissolved in water split up into charged particles called ions.

$$NaCl\,(s) \underset{}{\overset{H_2O}{\rightleftharpoons}} Na^+\,(aq) + Cl^-\,(aq)$$

12. What is an electrolytic cell?

Ans: The movement of ionic charges through the liquid brought by the application of electricity is called electrolytic conduction and the apparatus used is known as electrolytic cell.

13. Differentiate between electrolytic and galvanic cell.

Ans:

Electrolytic Cell	Voltaic cell
1. The electrochemical cell in which electrical energy is converted into chemical energy is called Electrolytic cell. **2.** In this cell, non- spontaneous reaction occurs. **3.** Electric current is used to drive the non-spontaneous oxidation reduction reaction. **4.** Electrolysis takes place in this cell. **Example:** Down's cell, Nelson's cell	**1.** The electrochemical cell in which chemical energy is converted into electrical energy is called Voltaic cell. **2.** In this cell, spontaneous reaction occurs. **3.** Electric current is produced due to spontaneous reaction. **4.** Electric conduction takes place in this cell. **Example:** Daniel's cell, Fuel cells

14. Define electrolysis OR Define electrolysis. Give one example.

Ans: The electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis.

OR

Electro comes from electricity and lysis means breakdown, so electrolysis is the breakdown of salts by passing electric current. Moreover, the electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis. The process is carried out in an electrolytic cell.

Example: Electrolysis of Fused salts

When a fused salt is electrolyzed the metal ions called cations move to cathode and get discharged by picking up electrons hence undergoing reduction. The anions move towards anode and also get discharged by losing electrons hence undergoing oxidation.

For example in case of fused lead chloride, the equations for electrode half reactions are as follows:

At Cathode: $Pb^{+2} + 2e^- \rightarrow Pb$ (reduction)

At Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$ (oxidation)

15. Differentiate between Ionization and electrolysis.

Ans:

Ionization	Electrolysis
1. The process in which ionic compounds when fused or dissolved in water split up into charged particles called ions. **2.** Electrodes are not needed **3.** Electricity is not needed **4.** Since there are no electrodes, therefore, ions do not move towards electrodes **5.** After ionization, ions are not discharged **Example:** $NaCl_{(s)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$	**1.** The process in which electricity is used to carry out a non-spontaneous reaction is called electrolysis. **2.** Electrodes are required **3.** Electricity is required **4.** Ions move towards their respective electrodes **5.** Ions are discharged at electrodes to give neutral products. **Example:** $NaCl_{(l)} \longrightarrow 2Na_{(s)} + Cl_{2(g)}$ At cathode At anode

16. Differentiate between conduction through metals and molten electrolytes.

Ans:

Conduction through metals	Conduction through molten electrolyte
1. Electrical conduction takes place due to free electrons.	**1.** Electrolytic conduction takes place due to ions.

2. In this case, conductance decreases with increase in temperature. **3.** No chemical reaction occurs during conduction. **4.** Chemical composition of metal is not changed during conduction and no new substances are produced. **Example:** All metals are conductors. e.g. Fe, Pb etc.	**2.** In this case, conductance increase with increase in temperature **3.** Redox reactions occurs take place during conduction **4.** Since chemical reactions occur, therefore new substances are produced **Example:** Molten salt e.g. NaCI (l) or their aqueous solutions, acids, bases etc.

17. How electrolysis takes place in fused salts?

Ans: When a fused salt is electrolyzed, the metal ions called cations arrive at the cathode which being negatively charged supply electrons to them and thus discharge the cations. The anions move towards the anode, give up their electrons and are thus discharged. In the case of fused lead chloride, the equations for electrode processes are given as under:

cathode: $Pb^{2+}_{(\ell)} + 2e^- \rightarrow Pb_{(s)}$ reduction)

anode: $2Cl^-_{(\ell)} \rightarrow Cl_{2(g)} + 2e^-$ oxidation

So, oxidation happens at anode and reduction at the cathode. Similarly, for fused NaCl and fused $PbBr_2$ the electrolytes are decomposed during electrolysis. Fused Pb and Na are deposited at cathode and Cl_2 (g) and Br_2 (l) at anode. Electrons flow through the external circuit from anode to cathode. The electric current is conducted through the cell by the ions and through the external circuit by the electrons.

18. How electrolysis takes place in aqueous solution of salts? OR Explain electrolysis of aqueous solution of sodium nitrate.

Ans: While electrolyzing aqueous sodium nitrate ($NaNO_3$) solution, sodium ions present are not discharged at the cathode. A small concentration of hydronium and hydroxyl ions arises from the dissociation of water:

$$NaNO_3 \rightarrow Na^+ + NO_3^-$$

$$2H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

Hydronium ions accept electrons from the cathode to form hydrogen atoms:

$$\text{At cathode: } H_3O^+_{(aq)} + e^- \rightarrow H_{(g)} + H_2O_{(\ell)} \text{ (reduction)}$$

Subsequently, hydrogen atoms combine rapidly to form hydrogen molecules at the cathode.

$$H_{(g)} + H_{(g)} \rightarrow H_{2(g)}$$

So, H_2 gas evolves at the cathode.

At the anode, both nitrate and hydroxide ions are present. Hydroxide ions are easier to discharge than nitrate ions. Nitrate ions remain in solution while the electrode reaction is:

$$\text{At anode: } OH^-_{(aq)} \rightarrow OH_{(aq)} + e^- \text{ (oxidation)}$$

The OH groups combine to give O_2 gas as follows.

$$4OH \rightarrow O_{2(g)} + 2H_2O_{(\ell)} \text{ (anode)}$$

So, O_2 gas evolves at the anode.

19. Give the chemistry of electrolysis of aqueous solution of sodium chloride.

Ans: Caustic Soda is obtained on commercial scale by the electrolysis of concentrated aqueous solution of sodium chloride using Titanium anode and Mercury or Steel cathode. This electrolysis is carried out in Nelson's cell or Castner- Kellner cell or Hg- cell.

$$NaCl \longrightarrow Na^+ + Cl^-$$

At Anode:

$2Cl^- \rightarrow Cl_2 + 2e^-$ (oxidation)

At Cathode:

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (reduction)

Overall reaction: (combining Na+ ions)

$2Na^+ + 2Cl^- + 2H\ O \rightarrow Cl + H\ + 2Na^+ + 2OH^-$

20. Explain extraction of sodium by Down's cell.

Ans: Extraction of sodium by the electrolysis of fused sodium chloride is carried out in Down's cell. In this case, molten sodium chloride is electrolyzed between iron cathode and graphite anode. The cell is planted to get sodium metal commercially. Chlorine is obtained as a by-product.

$$2NaCl(s) \rightarrow 2Na^+(\ell) + 2Cl^-(\ell)$$

At cathode

$2Na^+(\ell) + 2e^- \rightarrow 2Na(s)$ reduction

At anode

$2Cl^-(\ell) \rightarrow 2Cl(g) + 2e^-$ oxidation

$Cl(g) + Cl(g) \rightarrow Cl_2(g)$

By adding the two reactions at anode and cathode, the overall reaction is

$$2Na^+(\ell) + 2Cl^-(\ell) \rightarrow 2Na^\circ(s) + Cl_2{}^\circ(g)$$

21. How the process of electrolysis is applied in the production of caustic soda?

Ans: Caustic soda is obtained on industrial scale by the electrolysis of concentrated aqueous solution of sodium chloride using titanium anode and mercury or steel cathode. This electrolysis is carried out in Nelson cell and Castner- Kellner cell or Hg- cell.

$$NaCl\,(s) \rightleftharpoons Na^{+}\,(aq) + Cl^{-}\,(aq)$$

At anode

$$2Cl^{-}(g) \rightarrow Cl_2(g) + 2e^{-} \qquad \text{(oxidation)}$$

At cathode

$$2H_2O(\ell) + 2e^{-} \rightarrow H_2(g) + 2OH^{-}(aq) \qquad \text{(reduction)}$$

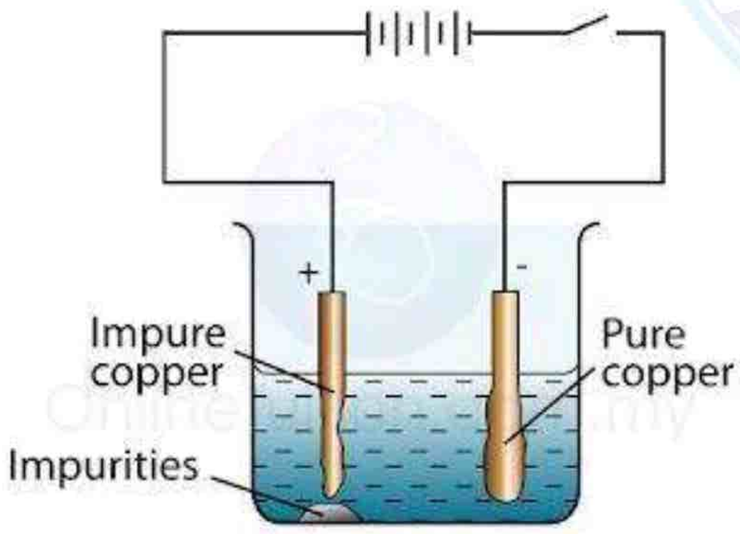
By combining, the electrode reactions and including Na^{+} ions, the overall reaction is:

$$2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + 2H_2O_{(l)} \rightarrow Cl_{2(g)} + H_{2(g)} + 2Na^{+}_{(aq)} + 2OH^{-}_{(aq)}$$

Here, chlorine and hydrogen are obtained as by-products and Na^{+} is not discharged at cathode.

22. How impure copper is purified by the process of electrolysis?

Ans: Electrolytic cell can also be used for the purification of copper. Impure copper is made the anode and a thin sheet of pure copper is made the cathode. Copper sulphate solution is used as an electrolyte. The atoms of Cu from impure Cu- anode are converted to Cu^{2+} ions and migrate to cathode which is made up of pure Cu. In this way, Cu anode is purified. Impurities are left at anode.



23. What is anodized aluminium?

Ans: Anodized aluminium is prepared by making it an anode in an electrolytic cell containing sulphuric acid or chromic acid, which coats a thin layer of oxide on it. The aluminium oxide layer resists attack for corrosive agents. The freshly anodized aluminium is hydrated and can absorb dyes.

24. What is the function of salt bridge?

Ans: The purpose of the salt bridge is to prevent any net charge accumulation in either beaker by allowing negative ions to leave the right beaker, diffuse through the bridge and enter the left beaker. If this diffusional exchange of ions does not occur, the net charge accumulating in the beakers would immediately stop the flow of electrons through the external circuit and the oxidation-reduction reaction would stop.

25. A salt bridge maintains the electrical neutrality in the cell. Explain.

Ans: In galvanic cell the salt bridge connects the two half cells together and maintains the electrical neutrality in both the half cells by transferring certain ions from one half cell to the other and thus preventing the accumulation of ions in either of the half cell. If salt bridge is not connected the cell will stop working very soon.

26. What is a voltaic cell?

Ans: A voltaic cell is one in which chemical energy is converted to electrical energy. It is a spontaneous cell.

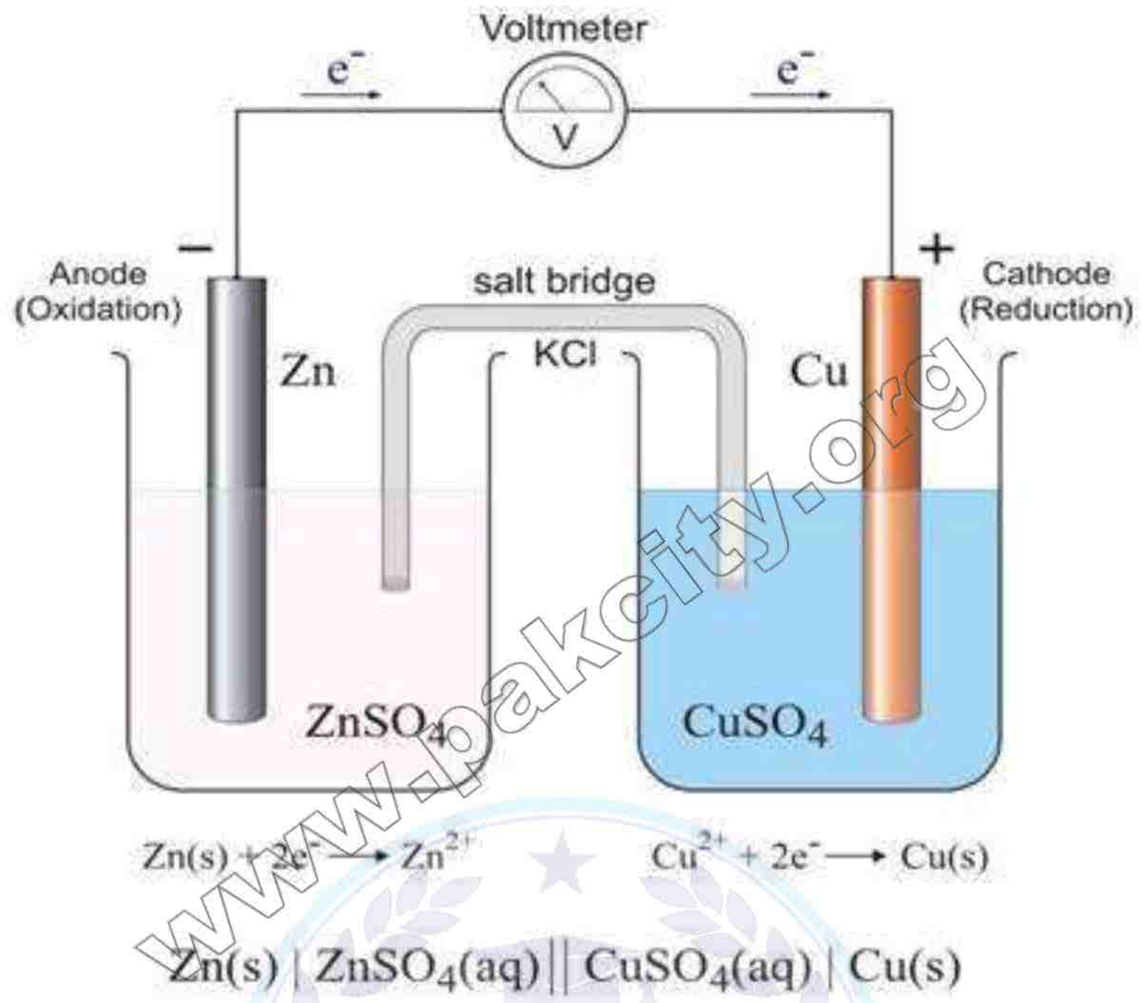
A voltaic or a galvanic cell consists of two half-cells that are electrically connected. Each half cell is a portion of the total cell in which a half reaction takes place. The left half cell consists of a strip of zinc metal dipped in 1.0 M solution of zinc sulphate giving the following equilibrium:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

The right half-cell is a copper metal strip that dips into 1.0 M copper sulphate solution and the equilibrium here is represented as follows:

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

These half-cells are connected electrically by a salt bridge. Oxidation takes place at Zinc and



reduction takes place at Copper.

27. Is voltaic cell a reversible cell?

Ans: If the external circuit is replaced by a source of electricity that opposes the voltaic cell, the electrode reactions can be reversed. Now, the external source pushes the electrons in the opposite direction and supplies energy or work to the cell so that the reverse non-spontaneous reaction occurs. Such, a cell is called a reversible cell. For the zinc-copper cell, the half-cell reactions are reversed to give.

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)} \quad \text{(reduction)}$$

$$Cu_{(s)} \rightarrow Cu^{2+} + 2e^- \quad \text{(oxidation)}$$

and the overall reaction being reversed, becomes

$$Zn^{2+}_{(aq)} + Cu_{(s)} \rightarrow Zn_{(s)} + Cu^{2+}_{(aq)}$$

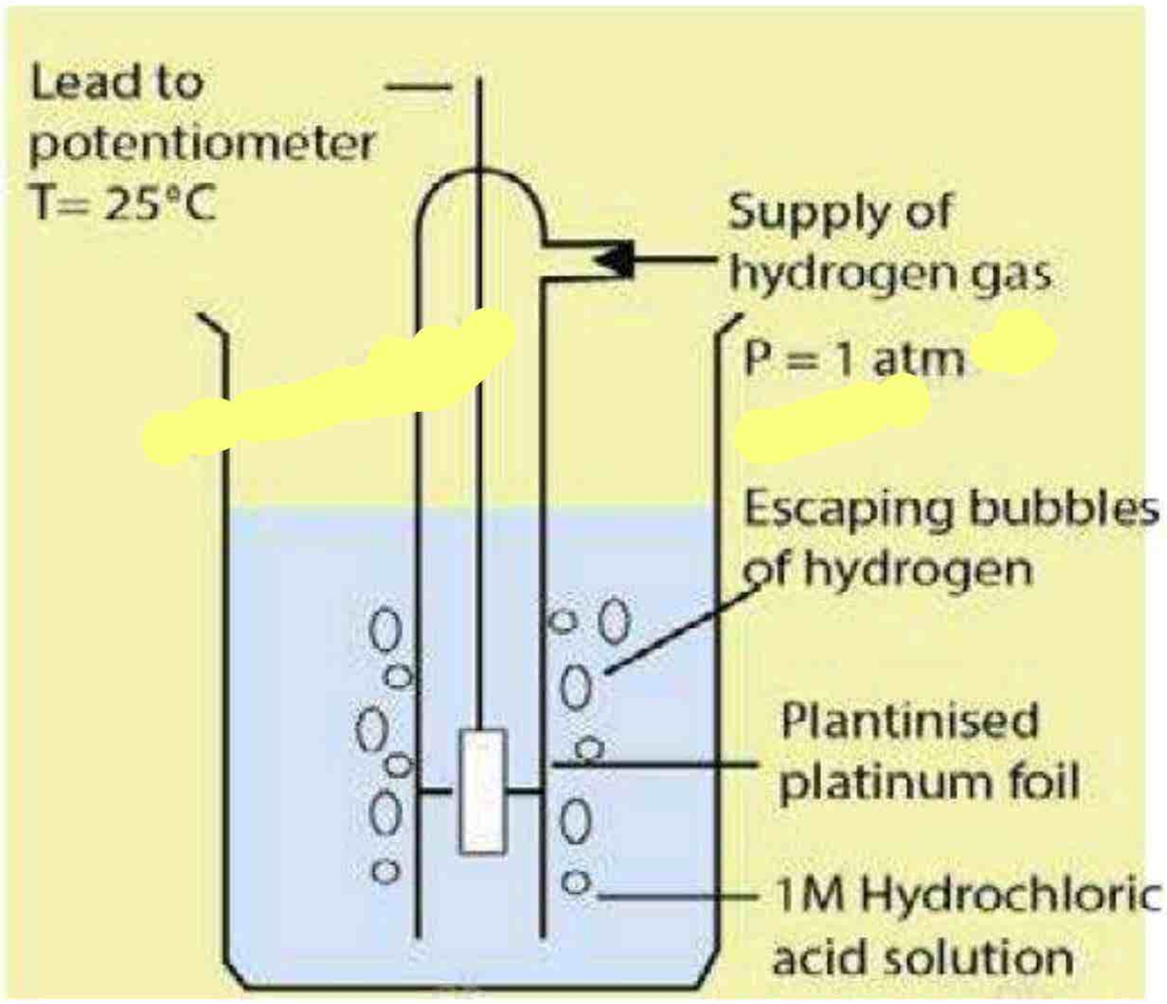
28. Define electrode potential OR What is standard electrode potential?

Ans: The potential set up when an electrode is in contact with one molar solution of its own ions at 298 K is known as standard electrode potential or standard reduction potential of the element. It is represented as E°.

29. What is SHE?

Ans: It stands for standard hydrogen electrode which is used as a standard. It consists of a piece of platinum foil, which is coated electrolytically with finely divided platinum black, to give a large surface area and suspended in one molar solution of HCl.

Pure hydrogen gas at one atmosphere pressure is continuously bubbled into 1M HCl solution. The platinum acts as an electrical conductor and also facilitates the attainment of equilibrium between the gas and its ions in solution. The potential of this electrode is arbitrarily taken as zero.



30. The equilibrium is set up between metal atoms of electrode and ions of metal in a cell. How?

Ans: The electrode potential can be explained on the basis of equilibrium which is established between atoms of metal deposited on the surface of metal through reduction process and the ions released by metal atoms into the solution due to oxidation of metal atoms. Whichever process is greater or faster determines whether the metal becomes positively or negatively charged. At equilibrium both the processes occur at the same rate and no further change occurs in the potential difference.

31. SHE acts as anode when connected with Cu but acts as cathode when connected with Zn.

Justify your answer with equations.

Ans: The reduction potential of Copper is +0.34V i.e a positive value, so it brings about reduction and act as cathode and SHE as anode.

Anode: $H_2 \rightarrow 2H^+ + 2e^-$ (oxidation)

Cathode: $Cu^{+2} + 2e^- \rightarrow Cu$ (reduction)

The oxidation potential of Zn is +0.76V i.e a positive value so Zn brings about oxidation and act as anode while SHE act as cathode.

Anode: $Zn \rightarrow Zn^{+2} + 2e^-$ (oxidation)

Cathode: $2H^+ + 2e^- \rightarrow H_2$ (reduction)

32. A porous plate or a salt bridge is not required in Lead acid storage battery. Justify.

Ans: A porous plate or a salt bridge is not required in a Lead storage battery because all the cells are dipped in the same electrolyte (**30% H_2SO_4 solution/ d= 1.25gcm^{-3}**). Salt bridge usually separates the two electrodes in which different electrolytes are used.

33. Write down reactions taking place at the electrodes during discharging of Nickel Cadmium cell.

Ans: At Anode:

$Cd + 2OH^- \longrightarrow Cd(OH)_2 + 2e^-$

(oxidation) At Cathode:

$NiO2 + 2H_2O + 2e^- \longrightarrow Ni(OH)_2 + 2OH^-$ (reduction)

Net reaction:

$Cd + NiO_2 + 2H_2O \longrightarrow Cd(OH)2 + Ni(OH)_2$

34. Give chemical reactions taking place at anode and cathode of Fuel cell.

Ans:

At Anode:

$[H_2 + 2OH^- \rightarrow 2H_2O + 2e^-]2$

At Cathode:

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Overall reaction:

$2H_2 + O_2 \rightarrow 2H_2O$

35. How Fuel cells produce electricity?

Ans: In fuel cells H_2 is oxidized and O_2 is reduced. The electrolyte of the cell is aqueous KOH solution. Electrodes of the cell are made of porous carbon impregnated with platinum. The following reactions take place at the respective electrodes to produce electricity.

At Anode:

$[H_2 + 2OH^- \rightarrow 2H_2O + 2e^-]$ x 2

At Cathode:

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Overall reaction:

$2H_2 + O_2 \rightarrow 2H_2O$

36. Write two advantages of Fuel cells.

Ans: Advantages of fuel cells are as follows.

a. Fuel cells are light, portable and produce electricity and pure water during space flights.

b. These are environment friendly, efficient and convert about 75% fuel bond energy into electrical energy.

37. Write recharging of lead accumulator battery.

Ans: During recharging the lead accumulator battery is connected to an external battery through the electrodes. As a result, the electrode half reactions are reversed as follows:

At Anode: $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{-2}$ (reduction)

At Cathode: $PbSO_4 + 2H_2O \rightarrow PbO_2 + 4H^+ + SO_4^{-2} + 2e^-$ (oxidation)

Overall reaction: $2PbSO_4 + 2H_2O \rightarrow Pb + PbO_2 + 4H^+ + 2SO_4^{-2}$

Both the density of the acid and voltage of the battery are restored.

38. Lead accumulator is a chargeable battery. Justify.

Ans: Lead accumulator is a chargeable battery and is based on secondary cells. The battery is chargeable the reactions of which can be reversed by applying external electrical source. During the process of discharging, the density of H_2SO_4 decreases from 1.25 g cm^{-3} to 1.15 g cm^{-3} and volts drop down from 12V. After recharging the density of sulphuric acid again increases up to 1.25 g cm^{-3}, electrodes regain original physical states and cell restarts providing E.M.F of 12V.

39. What is alkaline battery?

Ans: A dry alkaline battery is that which uses an alkali i.e KOH as an electrolyte. Zinc rod serves as anode and manganese dioxide as cathode. The battery is enclosed in a steel container. The voltage of the cell is 1.5V. The electrode reactions are as follows:

Anode: $Zn + 2OH^- \rightarrow Zn(OH)_2 + 2e^-$

Cathode: $2MnO_2 + H_2O + 2e^- \rightarrow Mn_2O_3 + 2OH^-$

Overall reaction: $Zn + 2MnO_2 + H_2O \rightarrow Zn(OH)_2 + Mn_2O_3$

40. Write reactions of discharging of lead accumulator battery.

Ans: At the anode the lead atoms release two electrons each to be oxidized to Pb^{2+} ions, which combine with SO_4^{2-} ions present in the electrolyte and get deposited on the anode as $PbSO_4$.

$$PbO_2\ (s) + 4H^+\ (aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow PbSO_4\ (s) + 2H_2O\ (l)\ \text{(reduction)}$$

$$Pb\ (s) + SO_4^{2-}(aq) \longrightarrow PbSO_4\ (s) + 2e^-\text{(oxidation)}$$

At the cathode the electrons from the anode are accepted by PbO_2 and hydrogen ions from the electrolyte then undergo a redox reaction to produce lead ions and water as follows: The Pb^{2+} ions then combine with the SO_4^{2-} ions and they both deposit at the cathode as $PbSO_4$. When both electrodes are completely covered with $PbSO_4$ deposits, the cell will cease to discharge any more current until it is recharged. The overall reaction is:

$$Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq)} + 2SO_4^{2-}{}_{(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(\ell)}$$

41. Explain the working of silver oxide battery.

Ans: The cathode is of silver oxide, Ag_2O and the anode is of zinc metal. The following reactions occur in a basic electrolyte.

At the anode

$$Zn_{(s)} + 2OH^-_{(aq)} \rightarrow Zn(OH)_{2(s)} + 2e^- \qquad \text{(oxidation)}$$

At the cathode

$$Ag_2O_{(s)} + H_2O_{(\ell)} + 2e^- \rightarrow 2Ag_{(s)} + 2OH^-_{(aq)}\ \text{(reduction).}$$

The overall reaction is:

$$Zn_{(s)} + Ag_2O_{(s)} + H_2O_{(l)} \rightarrow Zn(OH)_2 + 2Ag_{(s)}$$

The voltage of silver oxide battery is about 1.5 V

42. Define electrochemical series.

Ans: When elements are arranged in the order of their standard, electrode potentials on the hydrogen scale the resulting list is known as electrochemical series.

43. How electrochemical series predict feasibility of a reaction?

Ans: Electrochemical series help to predict whether a particular reaction will take place or not. For example, Cu^{2+} (aq) can oxidize solid zinc but Zn^{2+} (aq) cannot oxidize solid copper. Standard

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad E^\circ = +0.34 \text{volt}$$

$$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)} \qquad E^\circ = -0.76 \text{ volts}$$

reduction potential values of copper and zinc can explain this.

Since zinc is being oxidized so the reverse reaction will be considered.

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- \qquad E^\circ = -0.76 \text{ volts (oxidation)}$$

The overall reaction will thus be

$$Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)} \qquad E^\circ_{cell} = 1.10 \text{ volts}$$

The overall positive value for the reaction potential suggests that the process is energetically feasible. If the sum of E° values of the two half-cell reactions is negative, then the reaction will not be feasible.

44. The standard oxidation potential of Zinc is 0.76V and its reduction potential is -0.76. Why?

Ans: When Zn electrode is connected to S.H.E, Zn gives electrons to Hydrogen electrode and gets oxidized:

$$Zn \text{ ---> } Zn^{+2} + 2e^- \qquad E^o = +0.76$$

The standard oxidation potential of Zn is +0.76V as it is a spontaneous process. Its reduction will be non-spontaneous process with reduction potential of -0.76V.

45. How electrochemical series helps to calculate voltage or emf?

Ans: In a galvanic cell, the electrode occupying a higher position in the electrochemical series, will act as anode and oxidation takes place on it. Similarly, the electrode occupying the lower position in the series will act as a cathode and reduction will take place on it. The half-cell reactions are:

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e \qquad \text{(oxidation half reaction)}$$

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad \text{(reduction half reaction)}$$

$$Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)} \qquad \text{(complete cell reaction)}$$

The oxidation potential of Zn is positive. The reduction potential of Cu^{2+} is also positive. The cell voltage or emf of the cell is given by

$$E^\circ_{cell} = E^\circ_{oxi} + E^\circ_{red}$$

$$E^\circ_{cell} = 0.76 + 0.34 = 1.10 \text{ volts}$$

The cell voltage or emf measures the force with which electrons move in the external circuit and therefore measures the tendency of the cell reaction to take place. Galvanic cells, thus, give quantitative measure of the relative tendency of the various reactions to occur.

46. How relative tendency of elements for redox reactions can be compared by using electrochemical series?

Ans: The value of the reduction potential of a metal or a nonmetal tells us the tendency to lose electrons and act as a reducing agent. It also gives the information about the tendency of a species to gain electrons and act as a oxidizing agent. Greater the value of standard reduction potential of a given species, greater is its tendency to accept electrons to undergo reduction and hence to act as an oxidizing agent. For example, ions like Au^{3+}, Pt^{2+}, Hg^{2+}, Ag^{+}, Cu^{2+} and the nonmetals elements like F_2, Cl_2, Br_2 and I_2 which lie below the SHE, have a strong tendency to gain electrons and undergo reduction. The series tell us that strong oxidizing agents like F_2, Cl_2, Br_2, etc. have a large positive value of standard reduction potentials, while strong reducing agents have large negative values like Li, K, Ca, Na, etc. which lie above SHE.

47. How can relative chemical reactivity of metals be compared by using electrochemical series?

Ans: Greater the value of standard reduction potential of a metal, smaller is its tendency to lose electrons to change into a positive ion and hence lower will be its reactivity. For example, metals like Li, Na, K and Rb are highly reactive. Coinage metals, Cu, Ag, and Au are the least reactive because they have positive reduction potentials. Similarly, metals like Pb, Sn, Ni, Co and Cd which

are very close to SHE react very slowly with steam to liberate hydrogen gas, while the metals like Fe, Cr, Zn, Mn, Al and Mg which have more negative reduction potentials react with steam to produce the metallic oxides and hydrogen gas.

48. How reaction of metals with dilute acids can be compared by electrochemical series?

Ans: Greater the value of standard reduction potential of a metal, lesser is its tendency to lose electrons to form metal ions and so weaker is its tendency to displace H^+ ions from acids as H_2 gas. For example, metals like Au, Pt, Ag and Cu which have sufficiently high positive values of reduction potentials, do not liberate hydrogen gas from acids. While, metals like Zn, Mg and Ca which are close to the top of the series and have very low reduction potentials, liberate hydrogen gas, when they react with acids.

49. Na and K can displace hydrogen from acids but Pt, Pd and Cu cannot.

Ans: The elements whose reduction potential are negative show oxidation and are considered as oxidizable elements. On the other hand, the elements with positive reduction potential exhibit reduction and fall under the category of reducible elements. Therefore, Na and K have negative reduction potential so they can react with dilute acids like HCl and H_2SO_4 releasing H_2 gas. Pt, Pd and Cu cannot react with dilute acids.

50. How displacement of one metal by another can be studied by electrochemical series?

Ans: One metal will displace another metal from the aqueous solution of its salt if it lies above in the electrochemical series. For example, Fe can displace Cu from $CuSO_4$, Zn does not displace Mg from solution of $MgSO_4$.

51. Give two applications of electrochemical series.

Ans:

Following are the two applications of electrochemical series:

1. Prediction of feasibility of a chemical reaction

When we look at the electrochemical series, it is easy to predict whether a particular reaction will take place or not. For example Cu^{+2} ion can oxidize solid Zinc but Zn^{+2} ion cannot oxidize solid copper. Standard reduction potential values of Copper and Zinc are shown below.

$Cu^{+2} + 2e^{-} \text{----------}\rightarrow Cu \qquad E^{o} = +0.34V$

$Zn^{+2} + 2e^{-} \text{----------}\rightarrow Zn \qquad E^{o} = -0.76V$

2 .Displacement of one metal by another from its solution

Metal will displace another metal from the aqueous solution of its salt, if it lies above the electrochemical series. For example Fe can displace Cu from $CuSO_4$ but Zn does not displace Mg from solution of $MgSO_4$.

52. Is the reaction $Fe^{3+} + Ag \longrightarrow Fe^{2+} + Ag^{+}$ spontaneous? If not, write spontaneous reaction involving these species.

Ans:

$Fe^{3+} + Ag \longrightarrow Fe^{2+} + Ag^{+}$

Fe is reduced and Ag is oxidized. Hence, Fe^{+3} will act as cathode while Ag^{o} as anode.

Emf of the cell can be calculated as:

$E^{o}cell = E^{o}(ox) + E^{o}(red)$

$E^{o}cell = - 0.7994 + (- 0.44)$

$E^{o}cell = - 0.7994 - 0.44$

$E^{o}cell = - 1.2394$

emf of cell is negative so the cell-reaction is non-spontaneous.

On reversing the electrodes the cell-reaction becomes spontaneous

$Fe^{2+} + Ag^{+} \longrightarrow Fe^{3+} + Ag^{o}$

53. What is the difference between single electrode potential and standard electrode potential? Give its importance.

Ans:

Single Electrode potential:

The potential set up when an electrode is in contact with one molar solution of its own ions is known as standard electrode potential or standard reduction potential of the element. It is represented as E.

Standard electrode potential:

The potential set up when an electrode is in contact with one molar solution of its own ions at 298 K is known as standard electrode potential or standard redox potential of the element. It is represented as E^o.

Importance:

1. The standard electrode potential value of an element indicates the potential generated when an element is reduced or oxidized.

2. Total Emf of the cell can also be calculated by standard electrode potential of the elements working as electrode.