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CHEMISTRY

Class 9th (KPK)

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ELECTROCHEMISTRY

(Topic Wise Questions)

Q1. What is electro chemistry? Discuss briefly

Ans: **Electro Chemistry:**

The branch of chemistry which deals with the study of conversion of electrical into chemical energy and chemical energy into electrical energy is called electro chemistry.

Explanation:

In electro chemistry we study the chemical change take place. When electric current is pass through a particular type of material.

Electrochemistry is used in this modern world to generate electricity that is used in different electronic devices like cell in remote control, mobile phones, camera and telecommunication. In industry, many important chemical and useful products are manufactured by electrochemical methods.

Q2. What is oxidation number? What are the rules of assigning oxidation number?

Ans: **2. Oxidation State:**

The apparent charge (+ive or –ive) on an atom of an element in a molecule or compound is called oxidation state.

Explanation:

Oxidation state of an element is described by its oxidation number. This number enables us to describe oxidation-reduction reaction, and a balancing Redox chemical reaction. An increase of oxidation number is oxidation where a decrease of oxidation state is reduction.

Rules for assigning oxidation number:

i. The oxidation number of all elements in the Free state is zero

Example: N_2, Cl_2, Na, Fe, P_4 and S_8 etc. is zero

ii. The oxidation number of a simple ions is the same as the charge on it.

Examples: $Na^+ = 1$, $Al^{+3} = +3$, Ca^{2+}, $Br^{-1} = -1$

iii. The oxidation number of hydrogens in its compounds is +1 except in the case of metal hydrides, where it is -1. E.g. Na^+ , H^{-1}

Example:

In HCl^+, $H2O^+$ is +1 where as in metal hydride NaH^{-1}, CaH_2^{-1}, LiH^{-1} etc. is -1

iv. The oxidation number of oxygen in its compounds in -2 except in this case of peroxide, where it is -1 and in case of OF_2, It is +2

Examples:

In Zn^{-2} O^{-2} is -2 and in peroxide $H_2O_2^{-1}$ is -1 while in $O+2F_2^{-2}$ is +2

v. The oxidation number of each element of group I, II and III are +1, +2 and +3 respectively

Examples:

Group IA Li^{+1} is +1 Group IIA Ca^{+2} is +2 Group IIIA Al^{+3} is +3

vi. The oxidation number of each element of group VIIA (Halogen) in their binary compound is -1.

Examples: Br^{-1} is -1

vii. In neutral molecules, the algebraic sum of the oxidation number of all the element is zero.

Examples: In H^{+1} Cl^{-1} ➔ +1 -1 = 0, H^{+1} N^{+5} O^{-6} ➔ (1+5-6) = 0

viii. In ions, the algebraic sum of oxidation number is equal to the charge on the ion.

Example:

CO_3^{-2} = (+4 x 1) + (-2 x 3) = -4 -6 = -2

ix. In any substance the more electronegative atom has the negative oxidation number.

x. The same element may show different oxidation number in different compounds.

Example: CO($C^{+2}O^{-2}$), CO_2 (C^{+4}, O_2^{-4})

Q3. What is the oxidation state of C in CO_2, O in CO_2, Sn in $SnCl_4$ and S in K_2SO_4?

Ans: i. C in CO_2

x + 2 (-2) = 0

x – 4 = 0

x = +4

ii. O in CO_2

4 + 2 (x) = 0

2x = -4 ÷ both sides on 2

$\frac{Zx}{2} = \frac{A}{2}2$

x = -2

iii. Sn in $SnCl_4$

Solution:

Oxidation No of Sn = x

Oxidation No. of Cl = -1

SnCl4

x + 4 (-1) = m0

x – 4 = 0

x - +4

Oxidation No. of Sn in $SnCl_4$ is +4

iv. S in K_2SO_4

Solution:

Oxidation No. of K = +1

Oxidation No. of S = x

Oxidation Number of O = -2

K_2SO_4

2(+1) + x + 4 (-2) = 0

2 + x -8 = 0

x -6 = 0

x = +6

Oxidation No. of S in K_2SO_4 is +6

Q4. What is oxidizing and reducing agents? Explain with examples.

Ans: **i. Oxidizing Agent:**

An oxidizing agent is the specie that oxidizes other substances and itself get reduced. For example, KMnO4, $K_2Cr_2O_7$, HNO3 and Cl2 etc.

According to classical concept, an oxidizing agent may be:

a. The donor of oxygen to a substance.

b. The acceptor of hydrogen from a substance.
c. The acceptor of an electron from a substance.
d. The oxidation number of an oxidizing agent is decreased during a redox reaction.
Consider a following reaction:
Zn (S) + CuO (S) ➔ ZnO (S) + Cu (S)
Oxidation number of Cu in CuO is +2
Oxidation number of Cu (s) is = 0
There is decrease in oxidation number Cu from +2 to 0. So, Cuo is an oxidizing agent in the given example.

ii. Reducing agent:

a reducing agent is the specie that reduces other substances and itself get oxidized.
For example, H_2S, SO_2, Na, Al and Mg etc.

According to classical concept, a reducing agent may be,

a. The acceptor of oxygen to a substance.
b. The donor of hydrogen forms a substance.
c. The donor of an electron from a substance.
d. The oxidation number of an oxidizing agent is increased during a redox reaction.
Consider a following reaction:
Br_2 (s) + H_2S (s) ➔ 2HBr (s) + S (s)
Oxidation number of S in H_2S is -2
Oxidation number of free S is 0
There is increase in oxidation number of S from -2 to 0. H_2S is a reducing agent in the given example:

OXIDIZING AGENTS	REDUCING AGENTS
Bromine (Br_2)	Carbon (C)
Chlorine (Cl_2)	Carbon monoxide (CO)
Concentrated sulfuric acid (H2SO4)	Hydrogen (H_2)
Nitric Acid (HNO_3)	Hydrogen Sulphide (H_2S)
Oxygen (O_2)	Metals
Potassium permanganate ($KMnO_4$)	Potassium iodide (Kl)
Potassium dichromate ($K_2Cr_2O_7$)	Sulphur dioxide (SO_2)

Q5. What are the oxidation and reduction reactions?

Ans: **Oxidation Reduction reactions:**

Redox reaction (Red means reduction, Oxi means oxidation)
Those chemical reactions in which Oxidation-reduction (loss and gain of electrons) takes place simultaneously are called oxidation-reactions or Redox reactions.
$CuSO_4$ (aq) + Zn (s) ➔ $ZnSO_4$ (aq) + Cu (s)
The overall ionic equation is
The Redox reactions are actually made up two half reaction, for example consider the reaction.
$2Na^0 + l_2^0 \rightarrow 2Na^{+1} + 21^{-1}$
The reaction can be written in form of two half reaction as,
$2Na^0 \rightarrow 2Na^{+1} + 2e^-$ (oxidation)
$l_2^0 + 2e \rightarrow 21$ (reduction)

In the 1st half reaction, loss of electrons take place. This reaction is called an oxidation reaction.
In the 2nd half reaction, gain of electrons takes place. This reaction is called a reduction reaction.

Q6. What are electrochemical cells? Explain electrolytic cell in detail.

Ans: **Electrochemical Cells:**

A device in which inter conversion of electrical and chemical energies take place is called an electrochemical cell.

It is an energy device, in which either a chemical reaction take place by using electric current (such as electrolysis) or chemical reaction produces electric current (such as electric conductance).

Electro chemical cells are of two types.

i. **Electrolytic Cell**
ii. **Galvanic or Voltaic cell**

Electrolytic Cell:

A device in which electric current is used to produce Redox reaction.

(OR)

It is a device in which electrical energy is converted into chemical energy by non-spontaneous redox reaction.

Explanation:

In this cell electric current is produced due to the presence of potential difference. There reaction involve the gaining of electrons (reduction) and the losing of electrons (oxidation).

Construction:

An electrolytic cell consists of solution of an electrolytes. Two metallic called electrodes i.e. anode and cathode are dipped in the electrolytic solution. These electrodes are connected to the terminals of the battery. The electrode which is connected to the positive terminal of the battery is called anode and electrode which is connected with negative terminal of battery is called cathode.

Working of an electrolytic cell:

When the electrodes are connected to the battery and electric current is passed in the electrolytic cell, the ions in the electrolyte moves towards their respective electrodes. The anions liberate electrons at anode. These electrons pass through outer circuit to the cathode. The cations which surround the cathode, consume those electrons. Hence, the number of electrons lost is always equal to the number of electrons gained.

The battery can be thought of as an electron pump, simultaneously supplying electrons to the cathode and receiving electrons form the anode. Te anions move toward anode and discharge their electron (s) there and thus oxidation take place at anode. The cations move towards cathode and gain the electron (s) there are thus oxidation takes place at cathode.

For example, when electric current passed from the fused sodium chloride (NaCl), the following reactions take place during the process.

$2NaCl\ (l) \rightarrow 2Na^{+}\ (l) + 2Cl^{-}\ (l)$

At anode:

$2Cl^{-}\ (l) \rightarrow Cl_2\ (g) + 2e$

At cathode:

$2Na^{+}$ (l) + 2e ➔ 2Na (l)

Overall reaction:

$2Na^{+}$ (l) + 2Cl (l) ➔ 2Na (l) + Cl_2 (g)

Diagram:

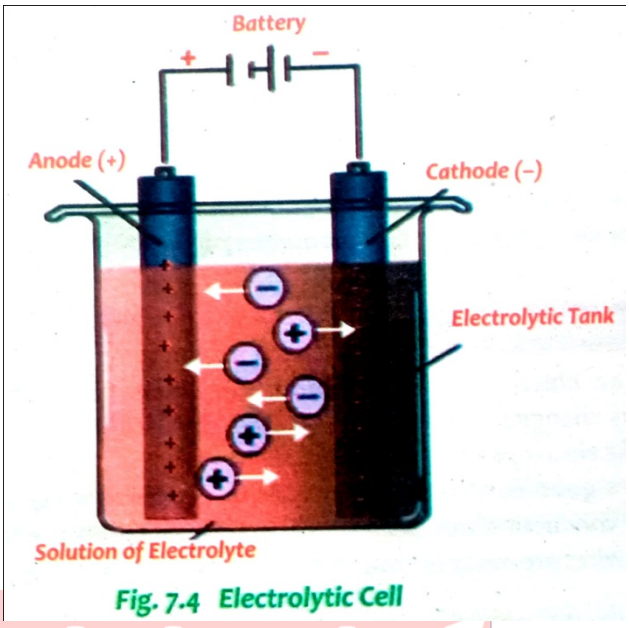


Fig. 7.4 Electrolytic Cell

Q7. What is galvanic or Voltaic cell? Explain in detail.

Ans: **Galvanic Cell:**

The cell in which the chemical energy is converted into electrical energy by spontaneous redox reaction is called voltaic cell (Galvanic cell)

Explanation:

Voltaic (Galvanic) cell consist of two separate compartments called “half cells” containing electrolytic solution and electrodes one is for oxidation and other is for reduction. The two half-cells are connected internally by a salt bridge and externally by a wire to which galvanometer or voltmeter is connected which detect or record the current. The best example of voltaic or galvanic cell in the Daniel cell.

Daniel Cell:

Construction of Daniel Cell:

A Galvanic cell consist of two separate containers, each container is called as half-cell. In each half-cell, an electrode is dipped in 1M solution of its own salt. The left half-cell consists of zinc electrode dipped in 1M solution of zinc sulphate ($ZnSo_4$) and right half copper electrode dipped in 1m of $CuSo_4$ connected to a wire to an external circuit to which a galvanometer or voltmeter is attached. The solutions in different containers are connected with the bridge. This bridge is known as “salt bridge”. A salt bridge is a U-shaped tube. This tube is filled with the electrolyte gel, such as K_2SO_4 or Na_2SO_4 is called as “agar”. The salt bridge inter-connects the two solutions in the anode container and the cathode container. A salt bridge performs three functions.

i. It allows electrical contact between the two solutions.

ii. It prevents the mixing of the two solutions.

iii. It keeps electrical neutrality in each half-cell.

Working of the cell:

The Zn metal has the tendency to lose electron more readily than copper when the circuit is completed. As a result, oxidation take place at Zn electrode. The electron flows from Zn electrode through the external circuit to copper electrode. These electrons are gained by the copper ions of the solution at the cathode and deposited as a copper atom at the cathode.

Cell Reaction:

The flow of electrons forms one electrode to other in the cell is due to the half-cell reaction taking place in the anode and cathode compartments. The net chemical changes obtained by adding the two half-cell reactions are called cell reaction. Thus, we have half-cell reaction i.e. oxidation and reduction processes, going on two electrodes simultaneously. Electrons travel in external circuit, while ions move through the salt bridge and this way electric current is produced. These reactions are as follow:

At anode:

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

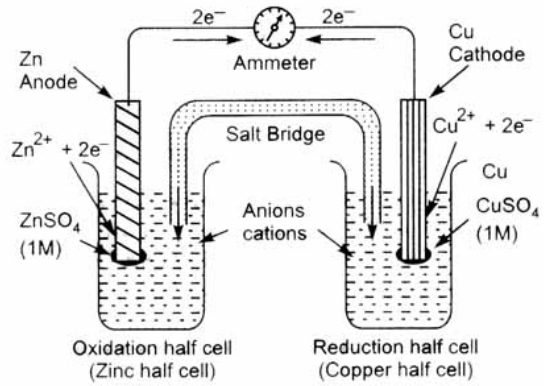
At cathode:

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

Overall reaction:

$Zn\ (s) + Cu^{+}\ (Aq) \rightarrow Zn^{+}\ (aq) + Cu\ (s)$

Diagram:



Daniell cell

Q8. Define electrolytes, Non electrolytes and electrodes.

Ans: **Electrolytes:**

The substance which dissociate into (+ive) and (-ive) ions in an aqueous solution and conduct electric current easily are called electrolytes.

Those substances which in aqueous solution or in molten state allow the electric current to pass through them are called electrolytes.

Examples:

Acids, Bases and salts,

Acids = Hcl, H_2SO_4, HNO_3, CH_3COOH

Bases = KOH, NaOH, NH_4OH

Salts = NaCl, N_4HCl etc. are electrolytes

Types of electrolytes:

Weak Electrolytes:

Those substance which ionize partially (upto limit extent) and conduct electric current poorly due to small ionization in aqueous solutions are called weak electrolytes

Examples:

NH_4OH, CH_3COOH, $Ca(OH)_2$, $Fe(OH)_3$ etc.

NH_4OH (aq) <===> NH_4 (aq) + OH (aq) (1.52%)

(Weak base)

CH_3COOH (aq) <===> CH_3COO (aq) + H^+ (aq) (1.5%)

(weak acid)

Examples:

NaOH, HCL, KOH, H2SO4 etc.

= NaOH (aq) ➔ Na+ (Aq) + OH (aq) (93%) (Strong base)

= HCl (aq) ➔ H+ (aq) + Cl (aq) (84%) (strong acid)

Strong Electrolytes:

Those substance which ionize completely (a large extent) in aqueous solution and conduct electricity to a large extent are called strong electrolytes.

Examples:

NaOH, HCL, KOH, H_2SO_4, etc.

= NaOH (aq) ➔ Na^+ (aq) + OH^- (aq) (93%) (strong base)

= HCL (aq) ➔ H^+ (Aq) + Cl^- (aq) (84%) (strong acid)

Non-electrolytes:

The substance which do not ionized in aqueous solution and do not allow the electric current to pass through them are called non-electrolytes. Most of the organic compounds are non-electrolytes.

Examples: Petrol, Sugar, Glucose, Urea Pure water etc.

Electrodes:

Electrodes are the conductor i.e. metallic plates, wires or rods through which electrons enter or leave the electrolyte in a cell are called electrodes.

There are two types of electrodes, Anode and Cathode.

a. Anode:

The anode is the positive electrode at which anion gathers and leaves the electron in the electrolytic cell.

Anions are the negatively charged particles e.g. Cl, OH etc.

b. Cathode:

The cathode is the negative electrode at which cations gathers and gains the electron in the electrolytic cell.

Cations are the positively charged particles e.g. Na^+, NH_4^+, etc.

Q9. Write the electrolytic refining of copper Cu.

Ans: **Electrolytic refining of copper Cu.**

Pure copper is very good conductor of electricity and used in electrical instruments.

Copper is purified by electro-refining.

Construction:

Large block of the blistered (impure) copper (99% pure) are suspended as anode in the copper Sulphate ($CuSO_4$) and sulphuric acid (H_2SO_4) between the thin sheets of pure copper which acts as cathode. The operation is performed at 50^0C and applied voltage of about 0.3 volts and optimum current density used is 160-400 A/m^2.

Procedure:

When the electric current is passed copper dissolves from the impure copper anode to give Cu^{+2} ions.

Reaction at anode (oxidation):

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

At the cathode, all the cu^{+2} ions from the solution are reduced to the metallic copper and get deposited at the cathode.

Reaction at anode (oxidation):

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

As the electrolysis continued, copper from the anode goes into solution. Traces of more active metals like Zn, Fe, etc. are also dissolved. The less active metals, for example Au, Ag remain un-dissolved and settled at the bottom of the cell as "Anode Sludge", which is processed to recover these precious metals. The voltage and temperature conditions are such that only copper is deposited at the cathode. By electrolytic refining up to 99.99% pure copper is obtained.

Reaction at Cathode (reduction)

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

Diagram:

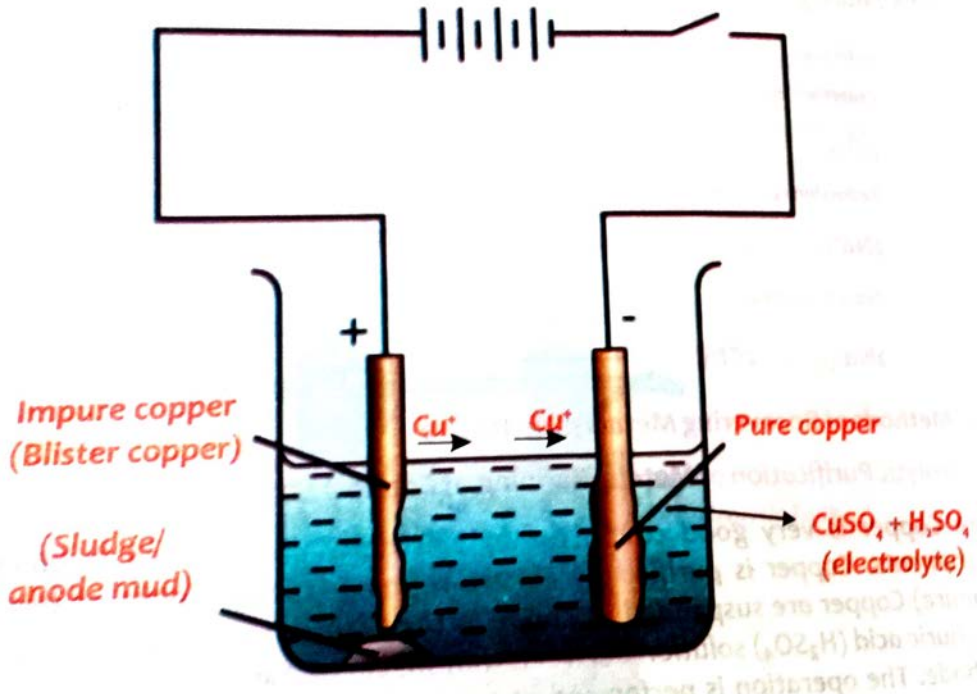


Fig. 7.5: Refining of Copper

Q10. What is battery? Explain dry cell.

Ans: **Battery:**

A group of combination of galvanic cell joined in a series is called battery.

Explanation:

A battery is a self-contained, chemical power pack than can produce a limited amount of electrical energy, whenever it is needed. It covers chemical energy into electrical energy, for specific period of time. Car batteries consist of six or more identical voltaic cells connected in series.

Dry Cell:

The dry cell was prepared by Leclanche in 1887.

Construction of dry cell:

The dry cell consists of metallic container. Its container is made up of Zinc (zn) which acts as anode. The Zinc casing is consumed during the chemical reaction. A graphite rod is placed in the center of container which acts as cathode. This container is also filled with the mixture of Ammonium chloride (NH_4Cl), Magnesium dioxide (MnO_2) and Carbon (C) which is in the form of paste. The cell is water proofed with the wax. The voltage produces by the dry cell is 1.25v.

Reaction of the Cell:

Oxidation and reduction reactions occurs in the cell to produce the electric current.

At Anode:

The zinc acts as anode in the cell. The zinc is oxidized by losing two electrons.

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

At Cathode:

The graphite acts as cathode in the cell. In the cell NH_4Cl and MnO_2 are reduced to Mn_2O_3 and NH_4.

$2MnO_2 + 2NH4 + 2e \rightarrow Mn_2O_3 + 2NH_3 + H_2O$

$2NH_4Cl \rightarrow 2NH_4^+ + 2Cl$

Overall reaction:

$Zn^{+2} + 2Cl$ ==. $ZnCl_2$

Q11. Write a note on electrochemical industries.

Ans: **Electrochemical industries:**

Electrochemical industries are based on many electrochemical operations. Some of these are as follow:

i. Electrochemical cells or batteries constructed with different electrodes are available in the market, which are widely used to power toys, flashlights, electronic calculators, pacemakers, radios, tape-recorders, automobiles etc.

ii. Electroplating of metals is the deposition of one metal on other metal electrolytically. This is done for the purpose of its durability, beauty or repair.

iii. Electrolytic production of metals (e.g. Na) and electrolytic refining of metals (e.g. Cu) are the popular methods for obtaining metals in their pure form.

iv. Many important chemicals are manufactured by electrochemical process, e.g. NaOH.

Q12. Write a note on prevention of corrosion and its techniques.

Ans: **Prevention of Corrosion:**

Corrosion can be prevented by a number of methods depending on the circumstances of corroded metal. Corrosion prevention techniques are generally classified into six groups.
These techniques are following:

a. Removal of Stains:

The regions of stains in an iron act as site for corrosion. When the surface of iron is properly cleaned and stains are removed, it prevents the process of rusting.

b. Paints and Coatings:

Paints and other organic coatings are used to protect metals from the corrosion effects. Beside these modern paints contain a combination of chemical called stabilizers. These stabilizers provide prevention against not only corrosion but also against not only corrosion but also against weathering and other atmospheric effects.

c. Alloying:

Alloying also helps to protect the corrosion of metals. The best example of alloying is the stainless steel, which is a solid mixture of iron, chromium a nickel. Stainless steel strongly resists the corrosion. The development of new alloys are constantly under production.

d. Metallic Coating or Plating:

Metallic coatings or plating, can be applied to inhibit corrosion as well as provide aesthetic and decorative finish. A thin coating of one metal on another can be applied by spraying, galvanizing (deposition of Zinc on other metal by dipping) or electroplating, for example iron articles are protected from rusting by Nickel (Nil), chromium (Cr) or tin (Sn) plating.

e. Corrosion Inhibitors:

Corrosion inhibitors are chemicals that react with the metal's surface or with the environmental gases which cause corrosion. They interrupt the chemical reaction that causes corrosion. These chemicals can be applied as a solution or as a protective coating via dispersion techniques e.g. glycine, polyethylene etc.

f. Cathodic Protection:

Cathodic protection is a method usually to protect iron buried fuel tanks and pipelines. An active metal, such as magnesium or zinc, is connected by a wire to the pipeline or tank to be protected. It is because the magnesium or zinc is a better reducing agent than iron, electrons are supplied by the magnesium or zinc than by iron, keeping the iron form being oxidized. As oxidation occurs, the magnesium or zinc anode dissolves and so it must be replaced periodically.
Ships hulls are protected in a similar way by attaching bars of titanium metal to steel hull, in salt water the titanium acts as anode and is oxidized instead of the steel hull (the cathode).

Q13. Write a note on zinc plating, chrome plating and tin plating.

Ans:

a. Electroplating of Tin:

The target metal is cleaned with caustic soda, treated with acids, in order to remove the rusts and oils/greases if any present on it. Then it is washed thoroughly with water. The electroplating of tin is carried out in electrolytic cell. In this process, pure piece of tin acts as anode and is dipped in sodium stannate ($Na_2SnO_3.3H_2O$) used as electrolytic solution. The cathode is the object to be coated with tin. When the electric current is passed through the cell, the anode starts dissolving and converted into Sn^{+2} ions. These Sn^{+2} ions move towards the cathode. At the cathode they discharged a deposited on the object.
The following reaction occurs:

At Anode:

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

b. Electroplating of Zinc

The target is cleaned and washed thoroughly with water. This object is dipped in zinc sulphate ($ZnSO_4$) container a small amount of sulphuric acid (H_2SO_4) solutions which acts as an electrolyte. The object to be electroplated acts as cathode, while anode is made of zinc plate or rod. When the electric current is passed, zinc anode dissolves and converted into Zn^{-2} ions. The electron moves to the cathode, where Zn^{+2} ions discharged and deposited as Zn metal.

Reactions taking place at electrodes are:

Anode: $Zn\ (s) \rightarrow Zn^{+2}\ (aq) + 2e$

Cathode: $Zn^{+2}\ (aq) + 2e \rightarrow Zn\ (S)$

c. Electroplating of chromium:

The electroplating of chromium is carried out in electrolytic cell. In this process pure sheet of chromium act as anode a dipped in chromium Acid (H_2CrO_4) solution containing small amount of sulphuric acid (H_2SO4). The object to be electroplated acts as cathode. The electrolyte ionizes and produces Cr+3 ions, at cathode they discharged and deposited on the object.

Reaction taking place at electrodes are:

At Anode: $Cr(s) \rightarrow Cr^{+3}\ (aq) + 3e$

At Cathode: $Cr^{+3}\ (Aq) + 3e \rightarrow Cr\ s(s)$

ELECTROCHEMISTRY

(Long Questions Answers)

Q1. a. What is electroplating?

a. Electroplating:

It is the process in which a thin layer of one metal is deposited on another metal by using electricity is called electroplating.

(OR)

The process in which a layer of superior metal is to be deposited on inferior metal by using electricity is called electroplating.

Procedure of electroplating:

i. The metallic substance must be deposited must be cleaned well and pure, washed with sand paper or with water or caustic soda.

ii. The object on which the layer of another metal is to be deposited is made cathode (electrode) by connecting it with negative terminal of battery.

iii. A pure sheet of metal which is to be electroplated is made the anode by connecting it with positive terminal of battery.

iv. A soluble salt of the metal to be electroplated is used as an electrolyte in the cell.

v. Electroplating is carried out in a tank made of cement, wood or glass.

vi. When current is passed, the metal atoms from anode are deposited over the cathode.

Q1. b. Distinguish between the nature of the anode and cathode in such a process.

Ans: **b. Nature of Anode:**

A sheet or rod of pure metal which is to be deposited is made anode. It is connected to the positive terminal of battery. When the electric current is passed, from anode converted into ions in solution. Thus, the mass of anode starts decreasing.

Examples:

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

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

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

Nature of cathode:

The cathode is made of metal which is to be deposit with a superior metal. The metallic ions of superior metal move towards the cathode and are deposited on the object to be coated.

Examples:

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

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

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

Q2. Differentiate between the process of oxidation and reduction. Write an equation to illustrate each.

Ans: There are different concepts to explain the oxidation and reduction.

Oxidation:

1. Classical concept:

According to the classical concept oxidation is a chemical reaction which involves the addition of oxygen.

Examples:

In a given reactions addition of oxygen take place to carbon and hydrogen to form CO_2 and H_2O.

$C + O_2 \rightarrow CO_2$

$2H_2 + O_2 \rightarrow 2H_2O$

(OR)

Removal of hydrogen is also called oxidation according to the classical concept.

Examples:

In a given reactions NH_3 is oxidized to N_2 and H_2S is to S.

$2NH_3 + 3Cl_2 \rightarrow N_2 + 6HCL$

$2H_2S + 2HNO_3 \rightarrow 3S + 2No + 4H_2O$

2. Modern electronic concept:

The modern electronic concept states that a substance which loses electrons in a chemical reaction is said to be oxidized and the reaction is called oxidation.

Example:

In a given reaction ferrous ion is oxidized to ferric ion by losing one electron.

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

Reduction:

Classical concept:

According to the classical concept reduction is a chemical reaction which involves the removal of oxygen.

It is an opposite phenomenon of oxidation

Examples: In a given reactions removal of oxygen take place to form metals.

$Cuo + H_2 \rightarrow Cu + H_2O$

$2HgO \rightarrow 2Hg + O_2$ (OR)

ii. When addition of hydrogen occurs in a chemical reaction is also called reduction according to the classical concept.

Examples:

$N_2 + 3H_2 \rightarrow 2NH_3$

$C + 2H_2 \rightarrow CH_4$

Modern electronic concept:

The modern electronic concept states that a substance which gains electrons in a chemical reaction is said to be reduced and the reaction is called reduction.

Example:

In a given reaction stannic ions is reduced to stannous ion by gaining two electrons.

$Sn^{+4} + 2e \rightarrow Sn^{+2}$

Q3.a. What is corrosion? Explain the rusting of iron as an example of corrosion.

Ans: **a. Corrosion:**

The conversion of any metal into its oxide by the action of environment is called corrosion.

(OR)

It is an oxidation-reduction process which takes place by the action of air in the presence of moisture with the metals.

Explanation:

The meaning of corrosion to a number of people is rust. The term rust and corrosion are almost synonymous. Corrosion is usually starting at the exposed surface of the metal. Corrosion process involve oxidation reduction reaction. Corrosion is a naturally occurring spontaneous phenomenon and it drives the materials to its lowest possible energy state.

Rusting of iron:

Pure iron is silvery metal but when exposed to moist air its surface is corroded and converted to a reddish-brown mass known as rust. Chemically rust is hydrate iron (III) oxide. For rust information there must be a thin film of water on the surface of the metal and air in surrounding. The impurities or the stained portions are responsible for the formation of small electrolytic cells, with anode of pure iron and cathode of impure or strained portions. Iron is oxidized at the anode, producing Fe (II) ions and electrons. It moves along the surface of the metal to cathode where it reacts with water and oxygen to form hydroxide ions.

At anode $2Fe \rightarrow 2Fe^{+2}(aq) + 4e$

At Cathode $2H_2O + O2\ 4e \rightarrow 4OH^-$

Fe (II) hydroxide which is further oxidized by atmospheric oxygen to form hydrated Fe (III) oxide, rust

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

$Fe^{+3}(aq) + 3OH(aq) \rightarrow Fe(OH)_3(s)$ (rust)

The rust mass is soft and porous in nature and therefore cannot prevent further deeper atmospheric action.

Q3. **b. Differentiate between electrolytic cell and Galvanic cell.**

Ans: **b.**

Electrolytic cell	Galvanic cell (voltaic cell)
An electrolytic cell converts electrical energy into chemical energy.	A voltaic cell converts chemical energy into electrical energy.
The redox reaction is non spontaneous and electrical energy has to be supplied to initiate the reaction.	The redox reaction is spontaneous and is responsible for the production of electrical energy.
Both the electrodes are placed in a same container in the solution of molten electrolyte. Salt bridge is not required.	The two half cells are set up in different cells containers, being connected through the salt bridge or porous partition. Salt bridge is required.
The anode is positive and cathode is the negative electrode.	The anode is negative and cathode is the positive electrode.
The reaction at the anode is oxidation and the reaction at cathode is reduction.	The reaction at the anode is oxidation and the reaction at cathode is reduction
The external battery supplies the electrons. They enter through the cathode and come out through the anode.	The electrons are supplied by the species getting oxidized. They move from the anode to the cathode in the external circuit.

Q3. **c. Discuss the method of recovering/extracting of metal from its ore.**

Ans: c. The process of extracting metals from their ores is called metallurgy. Electrolytic cell is used as a device for this purpose. Down cell is one the example of recovering of metal from is its ore is which sodium is extracted from NaCl.

Manufacture of sodium metal from fused NaCl:

Sodium metal was first discovered by an English chemist, Sir Humphrey Davey in 1807 by the electrolysis of fused Sodium Hydroxide (NaoH). Commercially sodium metal is obtained from the electrolysis of molten Sodium Chloride (NaCl) in the down cell.

Construction of Down's Cell:

Sodium chloride being a strong electrolyte in the molten state give us Na and Cl ions which are free to move towards their respective electrodes. The cell used for the electrolysis of fused NaCl is called Down's cell. Graphite is used as a "Anode" and a steel electrode at both sides are used as a cathode. Anode is present in the center of the cell. The electrolysis yield sodium (Na) metal and chlorine gas as product. So, it is necessary to keep these products separated, otherwise they will react to give sodium chloride (NaCl) again.

At the cathode, one Na^+ ion pick up one electron and is changed into Na atom. Cl ion will move towards the Anode and give an electron and will change into chlorine atom.

Reaction:

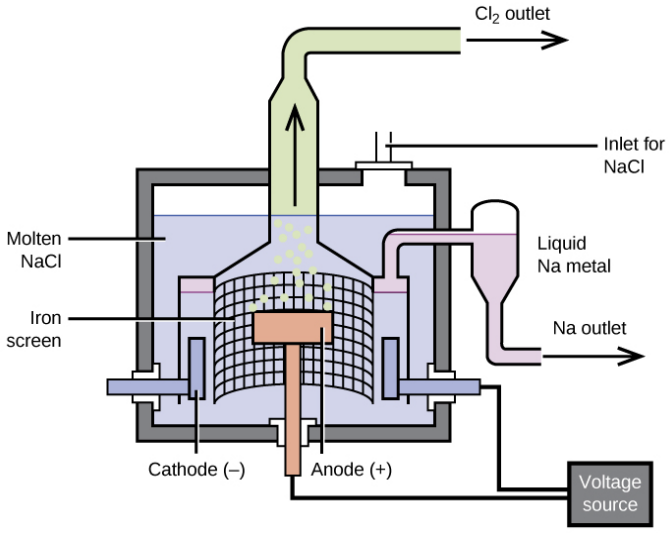
$2Na^+ + 2e^- \rightarrow 2Na$ (reduction)

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

Net Reaction:

$2Nacl\ (fused) + 2Cl \rightarrow 2Na\ (l) + Cl_2\ (g)$

Diagram:



Q4. **Discuss the preparation of Sodium Hydroxide (NaOH) from brine along with diagram and reactions at cathode and anode.**

Ans: **Manufacture of NaOH from brine:**

A concentrated aqueous solution of sodium chloride, NaCl (Brine) is placed in a special apparatus, known as Nelson cell for the manufacture of NaOH.

Construction of Nelson Cell:

It consists of U-shaped tube. This tube is made of steel. It is perforated. This perforated tube acts as cathode. A graphite anode is suspended in the U-shaped tube. The cathode is coated with asbestos. The asbestos separates the anode from the cathode.
During the electrolysis, the chlorine is produced at the anode. It is collected at the chlorine outlet. Hydrogen gas is produced at the cathode. It is collected at Hydrogen outlet. During this reaction sodium hydroxide is also produced. The sodium hydroxide is collected in the catch basin, placed under the U-shaped tube. In this process, the Hydrogen, chlorine and sodium hydroxide is produced at the same time.

$2Nacl \rightarrow 2Na\ (aq) + 2Cl\ (aq)$

$2H_2O\ (l) + 2e \leftrightarrows H^+\ (aq) + OH^-\ (aq)$

Working of Nelson cell:

When the electrodes are connected to the battery, the positive ions, Na^+ and H^+ move towards the cathode. Since, H^+ have a greater tendency to pick up electrons to form H_2 gas. Na^+ ions are not reduced instead with OH ions, present in the solution to form caustic soda (NaOH) which make the solution alkaline, while Cl ions move towards anode where thy give electrons to the electrode.

Reaction at Anode:

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

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

At Cathode:

$2H^+\ (aq) + 2e \rightarrow 2H\ (aq)$ (reduction)

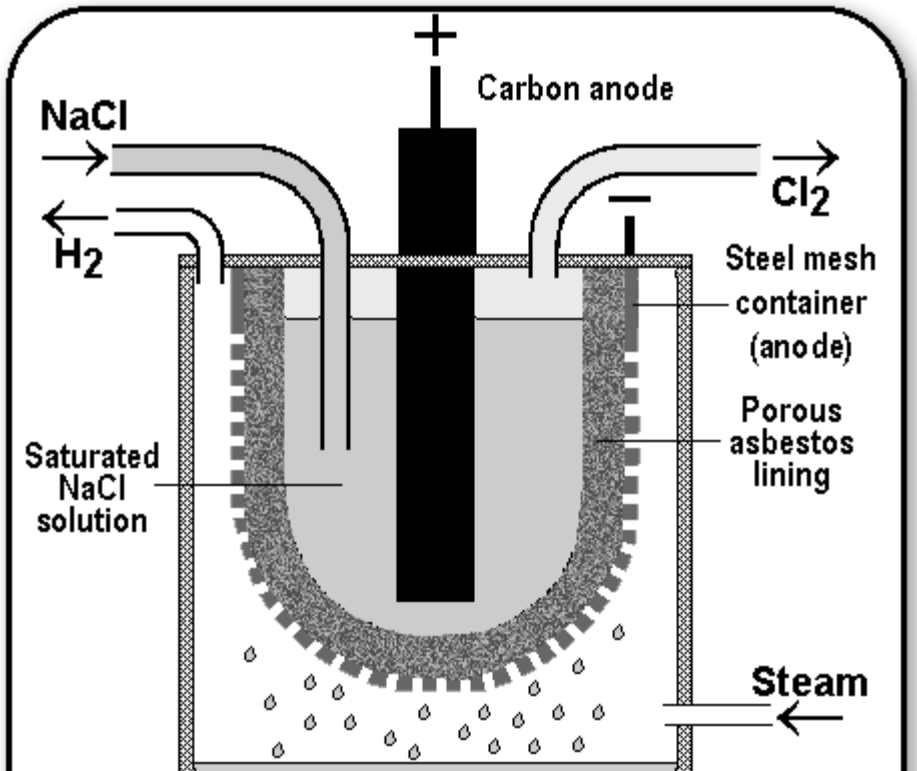
$H + H \rightarrow H_2$

$2Na^+\ (aq) + 2OH\ (aq) \rightarrow 2NaOH\ (aq)$

The overall reaction:

$2NaCl\ (aq) + 2H_2O\ (l) - 2NaOH\ (aq) + Cl_2\ (g) + H_2\ (g)$

Diagram:



ELECTROCHEMISTRY

(Short Questions Answers)

Answer briefly the following questions.

Q1. Indicate which element is reduced in the following reactions.

a. $Ca\ (s) + Br_2\ (g) \rightarrow CaBr_2(s)$

b. $8H + MnO_4 + 5e \rightarrow Mn^2 + 4H_2O$

Ans: **a. $Ca\ (s) + Br_2\ (g) \rightarrow CaBr_2(s)$**

Reduction is decreased in oxidation number of an element in a chemical reaction. Hence in above reaction Bromine is reduced because its oxidation number is reduced from 0 to -1.

b. $8H + MnO_4 + 5e \rightarrow Mn^2 + 4H_2O$

In above equation Manganese is reduced because its oxidation number is reduced from +7 to +2.

Q2. What is Oxidation number of silvers on each side of the following equation?

$4Ag\ (s) + O_2\ (g) \rightarrow 2Ag_2O\ (s)$

Ans: **$4Ag\ (s) + O_2\ (g) \rightarrow 2Ag_2O\ (s)$**

Oxidation number of silvers on reactant hand side is zero because it is present in Free state. While oxidation number of silvers on right hand side is +1 due to loss of one electron.

Q3. Why NaOH is a strong but NH_4OH is weak electrolytes?

Ans: NaOH (Sodium Hydroxide) is a strong electrolyte. Strong electrolytes are those electrolytes which

Are dissociate completely in aqueous solution and can conduct electricity to a large extent.

NaOH (aq) ➔ Na^+ (aq) + OH (aq) (84%)

While NH4OH (sodium hydroxide) is weak electrolyte which do not dissociate completely in water.

NH_4OH (aq) <===> $NH4^+$ (Aq) + OH (aq) (1.5%)

Q4. How to prevent corrosion? Enlist few of methods.

Ans: **Prevention of corrosion:**

Corrosion can be prevented by:

i. Metallic Coating:

Metallic coating is used to prevent the metal from the atmosphere effect

Example:

Electroplating (iron particles are protected from rusting by Ni, Cr or tin plating)

ii. Paint Coating:

We can prevent corrosion by paint which as a protective coating.

iii. Alloying:

By alloying the metal.

Example:

Stainless steel is the alloy of iron with Ni, Cr and Si

iv. Corrosion inhibitors:

Corrosion inhibitors are chemicals that react with the metal's surface or with the environmental gases which cause corrosion.

Example: Glycine, polyethylene etc.

v. Cathodic Protection:

Cathodic protection is a method usually used to protect iron in buried fuel tanks and pipelines. For this purpose mostly active metals like zinc or magnesium is used.

Q5. Write chemical reactions that occur in Nelson's cell.

Ans: **Reactions that occur in Nelson's are:**

The saturated brine ionized as follows:

2NaCl ➔ 2Na (aq) + 2Cl (aq)

$2H_2O$ (l) ➔ ➔ H + (aq) + OH^- (aq)

At Anode:

2Cl- (Aq) ➔ Cl_2 (g) + 2e (oxidation)

Cl (g) + Cl (g) ➔ Cl_2 (g)

At Cathode:

$2H^+$ (aq) + 2e ➔ 2H (aq)

H + H ➔ H_2

$2Na^+$ (aq) + 2_0H (aq) ➔ 2NaOH (aq)

The overall reaction:

2NaCl (aq) + $2H_2O$ (l) ➔ 2NaOH (aq) + Cl_2 (g) + H_2 (g)

Q6. Write an example from daily life which involves the oxidation-reduction reaction.

Ans: Photosynthesis in plants is the example of oxidation-reduction reaction. It converts the light energy into chemical energy. This process consists of series of chemical reactions that require carbon dioxide, water and store chemical energy in the form of sugar, light energy from light drives the reactions.

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

Photosynthesis transfers electrons from water to carbon dioxide molecules. This transfer of electrons is the example of an oxidation-reduction process. Water is oxidized by losses electrons and carbon dioxide is reduced by gaining electrons.

Q7. Assign oxidation numbers to each atom in the following compounds.

a. Hl b. PBr_3 c. $CaCo_3$ d. H_2PO_4 e. As_3O_5 f. H_2SO_4

Ans: a. Hl – H = +1, 1 = -1

b. PBr_3. P = +3, Br = -1 (3Br = -3)

c. $CaCo_3$ – Ca = +2, C = +4, O = -2 (30 = -6)

d. H_3PO_4. H = +1 (3H = +3), P = +5, 0 = -2 (40 = -8)

e. As_3O_5. As = +3.33 (3As = +1U), (50 = -10)

f. H_2SO_4. H = +1 (2H = +2), S = +6, 0 = -2 (40 = -8)

Q8. Why oxygen is necessary for rusting?

Ans: Oxygen is necessary for rusting. It converts metal into hydrated rust. Dents and strains present on the surface of iron acts as an anode which oxidizes the iron.

$2Fe\ (s) \rightarrow 2Fe^{+2}\ (aq) + 4e$

While that part on the surface of iron where oxygen and water are present acts as cathode. The free electrons move to this part and following reaction occurs.

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

$Fe^{+2}$is further oxidized by atmospheric oxygen to form hydrated Fe (III) oxide which is known as rust.

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

$Fe^{+3} + 3OH^-\ (aq) \rightarrow Fe\ (OH)_3\ (s)$.

Q9. Sketch the Daniel cell, labeling the cathode, anode and the direction of flow of electrons.

Ans:

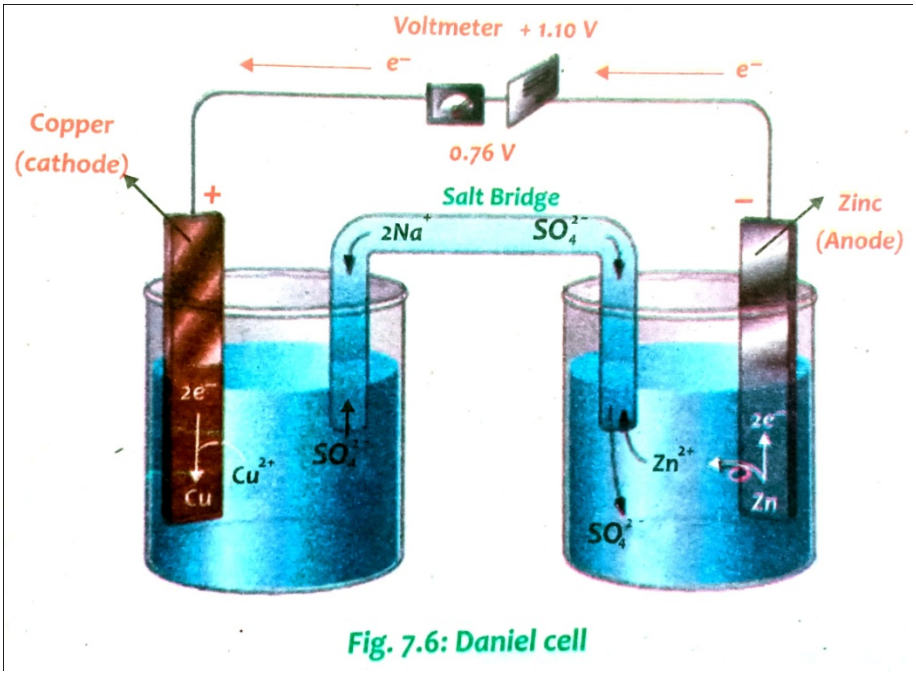


Fig. 7.6: Daniel cell

Q10. Write down some possible uses of an electrolytic cell.

Ans: **Uses of an electrolytic cell:**

i. It is used for the commercial preparation of sodium metal.
ii. It is used for the conversion of electrical energy into chemical energy.
iii. It is used for the purification of impure metals.
iv. It is used in the extraction of metals from their ores.
v. It is also used in the process of electroplating.