

CHAPTER

7

ELECTROCHEMISTRY

SUBJECTIVE PART

LONG QUESTION ANSWERS

Q.1 a. Define electrochemistry

b. Differentiate between spontaneous and non-spontaneous reactions

Ans: a. Electrochemistry

The branch of Chemistry that deals with the relationship between electricity and chemical reactions is called electrochemistry.

OR

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

Electrochemistry involves oxidation and reduction reactions, which are also known as redox reactions. There are two types of redox reactions.

i. Spontaneous reaction

ii. Non-spontaneous reactions

b. Difference between spontaneous and non-spontaneous reactions:

i. Spontaneous reactions

Spontaneous reactions are those which take place on their own without an external agent.

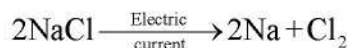
For example, reactions taking place in voltaic or galvanic cells and corrosion process



ii. Non-spontaneous reactions

Non-spontaneous reactions are those which take place in the presence of an external agent.

For example, reactions taking place in electrolytic cells like electrolysis of fused NaCl to produce sodium.



Q.2 Explain Oxidation and Reduction reactions on the basis of addition and removed of hydrogen and oxygen.

Ans: Oxidation and reduction reactions in term of loss or gain of hydrogen/oxygen:

Oxidation:

Oxidation is defined as addition of oxygen or removal of hydrogen during a chemical reaction.

Reduction:

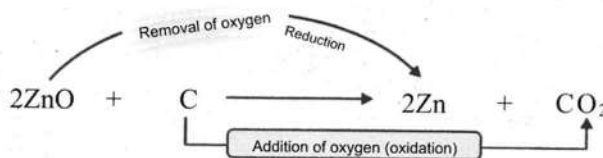
Reduction is defined as addition of hydrogen or removal of oxygen during a chemical reaction.

Both of these processes take place simultaneously in a redox reaction.

Examples:

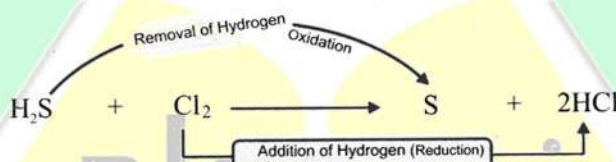
i. Addition or removal of oxygen:

A reaction between zinc oxide and carbon takes place by the removal of oxygen (reduction) from zinc oxide and addition of oxygen (oxidation) to carbon.



ii. Addition or removal of hydrogen:

A reaction between hydrogen sulphide and chlorine takes place by oxidation of hydrogen sulphide and reduction of chlorine. Hydrogen is being removed from H_2S and added to chlorine.



Q.3 Explain oxidation and reduction in terms of loss or gain of electrons.

Ans: Oxidation and reduction in terms of loss or gain of electrons.

There are many chemical reactions which do not involve oxygen or hydrogen but they are considered redox reactions. To deal with these reactions, a new concept "loss or gain of electrons" is used called electronic concept.

According to this concept:

Oxidation:

Oxidation is loss of electrons by an atom or an ion.

e.g.



Reduction:

Reduction is gain of electrons by an atom or ion.

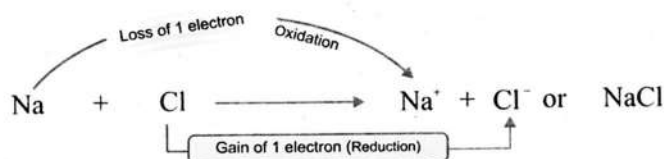
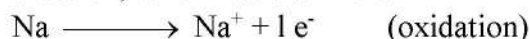
e.g.



Explanation:

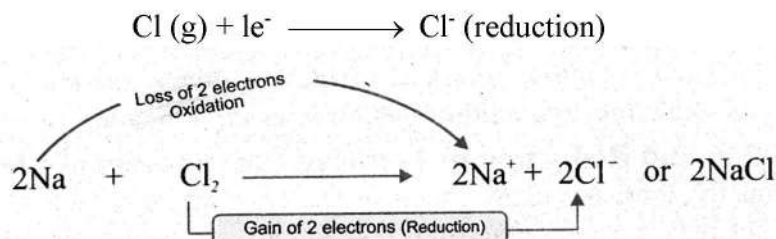
A reaction between sodium metal and chlorine takes place in three steps.

i. First sodium atom loses an electron, to form sodium ion.



ii. Simultaneously, this electron is accepted by chlorine atom (reduction process), as chlorine atom needs one electron to complete its octet. As a result chlorine atom

changes to chloride ion.



iii. Ultimately, both these ions attract each other to form sodium chloride. Complete redox reaction is sum of the oxidation and reduction reactions between sodium and chlorine atoms and it is represented as;



Summary of concept of oxidation and reduction

Oxidation	Reduction
Addition of oxygen	Removal of oxygen
Removal of hydrogen	Addition of hydrogen
Loss of electrons	Gain of electrons

Q.4 Describe the rules for assigning oxidation states.

(Ex.Q.1)

Oxidation states:

Oxidation state or oxidation number (O.N.) is the apparent charge assigned to an atom of an element in a molecule or in an ion.

It may be positive or negative or whole number or in fraction or zero.

Examples:

In HCl, oxidation number of H is +1 and that of Cl is -1.

Rules for assigning oxidation numbers (O.N):

i. **O.N of elements in free state:**

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

ii. **O. N. of simple ions:**

The oxidation number of an ion consisting of a single element is the same as the charge on the ion.

iii. **O. N of elements in complex ions:**

The oxidation number of different elements in Group-I is +1, in Group-2 is +2 and in Group-3 is +3.

iv. **O.N of H:**

The oxidation number of hydrogen in all its compounds is + 1 but in metal hydrides it is -1.

v. **O. N. of oxygen:**

The oxidation number of oxygen in all its compounds is -2 but it is -1 in peroxides and +2 in OF_2 .

vi. **Negative O.N:**

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

vii. **O.N. of neutral molecules:**

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

viii. **O.N. in ions:**

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

Q.5 Explain oxidizing and reducing agents with the help of suitable examples.

Ans: Oxidizing agents:

An oxidizing agent is the species that oxidizes a substance by taking electrons from it.

OR

The substance (atom or ion) which is reduced itself by gaining electrons from other substance is also called oxidizing agent.

Examples: Non-metals (O_2, Cl_2, F_2) are oxidizing agents because they accept electrons being more electronegative elements. Strong acids like HNO_3, H_2SO_4 etc are also metals.

Reducing agents:

Reducing agent is the species that reduces a substance by donating electron to it.

OR

The substance (atom or ion) which is oxidized by losing electrons to other substance is also called reducing agent.

Examples: Metals like zinc, iron, aluminum etc. are good reducing agent.

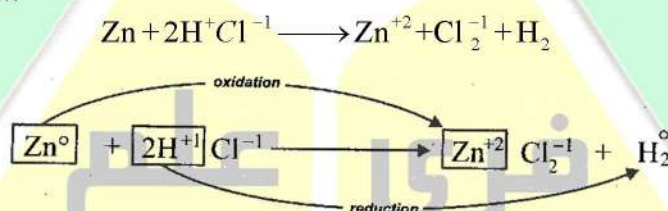
Explanation:

i. Reaction between Zn and HCl :

Let us discuss a reaction of Zn metal with hydrochloric acid.



The oxidation states or oxidation numbers of all the atoms or ions in this reaction are indicated below



Let us find the atoms, which are oxidized or reduced or whether there is a change in their oxidation state. It is indicated as follows.

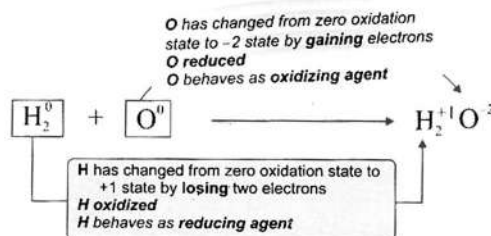
In this reaction, zinc is oxidized and acts as a reducing agent. Hydrogen is reduced and acts as an oxidizing agent.

ii. Reaction between H & O:

In case of formation of water from hydrogen and oxygen gases, redox reaction takes place as follows.



In this reaction, H is oxidized and acts as a reducing agent. O is reduced and acts as an oxidizing agent.



Q.6 What are electrochemical cells? Explain the concept of electrolytes.

Ans: Electrochemical cell:

Electrochemical cell is a system in which two electrodes are dipped in the solution of an electrolyte or molten mass.

Electrochemical cell is an energy storage device in which either a chemical reaction takes place by using electric current (electrolysis) or chemical reaction produces electric current.

Types of electrochemical cells:

- i. Electrolytic cells
- ii. Galvanic cells or Voltaic cells

i. Electrolytic cells:

The type of electrochemical reaction in which a non-spontaneous chemical reaction takes place when electric current is passed through an electrolyte is called an electrolytic cell.

Examples: Down's cell, Nelson cell etc.

ii. Galvanic cells or Voltaic cells:

The type of electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called Galvanic or Voltaic cell.

Example: Daniel cell, Dry cell.

Concept of Electrolytes:**a. Electrolytes**

The substances, which can conduct electricity in their solutions or molten states, are called electrolytes.

Solutions of salts, acids or bases are good electrolytes. The electricity cannot pass through solid NaCl but its aqueous solution or molten NaCl are good electrolytes.

Classification of electrolytes:

Electrolytes are classified into two groups depending upon their extent of ionization in solution.

i. Strong Electrolytes

The electrolytes which ionize completely in solution and produce more ions, are called strong electrolytes.

Examples: Strong electrolytes are aqueous solutions of NaCl, NaOH and H₂SO₄ etc.

**ii. Weak electrolytes:**

The substances which ionize to a small extent when dissolved in water and produce less ions are called weak electrolytes.

Examples: Acetic acid (CH₃COOH), and Ca(OH)₂ when dissolved in water, ionizes to a small extent. These are good examples of weak electrolytes. Weak electrolytes do not ionize completely. For example, ionization of acetic acid in water produces less ions:



As a result the weak electrolyte is a poor conductor of electricity.

b. Non-Electrolytes

The substances, which do not ionize in solution and do not allow the current to pass through their solutions, are called non-electrolytes.

Example: Sugar solution and benzene are non-electrolytes.

Q.7 (Ex. Q.s) What is electrolysis? Write a note on electrolytic cells.

OR

How can a non-spontaneous reaction be carried out in an electrolytic cell? Discuss in detail.

Ans: Electrolysis:

The chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound is called electrolysis.

Example: Electrolysis of NaCl into Na and Cl.

Electrolytic Cell:

The type of electrochemical cell in which a non-spontaneous chemical reaction takes place when electric current is passed through the electrolyte, is called an electrolytic cell.

Example: Down's cell, Nelson's cell, Hoffmann's Voltameter.

Construction of an Electrolytic Cell:

An electrolytic cell consists of a solution of an electrolyte, two electrodes (anode and cathode) that are dipped in the electrolytic solution and connected to the battery.

Anode:

The electrode connected to positive terminal is called anode

Cathode:

Electrode connected to the negative terminal is called cathode.

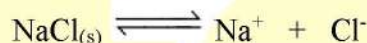
Working of an Electrolytic Cell:

When electric current is applied from battery, the ions in the electrolyte migrate to their respective electrodes. The anions, which are negatively charged, move towards the anode and discharge there by losing their electrons. Thus oxidation takes place at anode. While cations, which are positively charged ions, move towards cathode. Cations gain electrons from the cathode and as a result reduction takes place at cathode.

Example (Electrolysis of fused sodium chloride):

When fused salt of sodium chloride is electrolyzed the following reactions take place during this process.

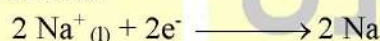
Ionization reaction:



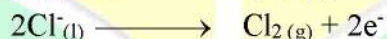
Oxidation reaction at anode:



Reduction reaction at cathode:



Overall reaction:



Q.8 Discuss the electrolysis of water.

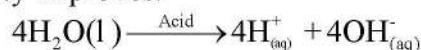
Ans: Electrolysis:

The chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound is called electrolysis.

Electrolysis of water:

Extent of ionization of water:

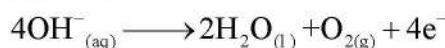
Pure water is a very weak electrolyte. It ionizes to a very small extent. The concentrations of hydrogen ions (H^+) and hydroxide ions (OH^-) are both 10^{-7} respectively. Increase in conductivity of water. When a few drops of an acid are added in water, its conductivity improves.



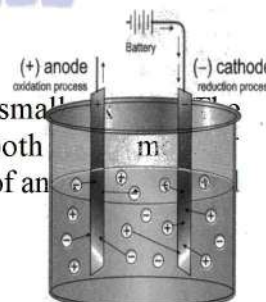
Working:

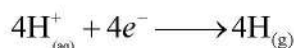
When an electric current is passed through this acidified water, OH^- (**anions**) move towards positive electrode (anode) and H^+ (**cations**) move towards negative electrode (cathode) and discharge takes place at these electrodes. They produce oxygen and hydrogen gases respectively at anode and cathode.

Oxidation reaction at anode:



Reduction reaction at cathode:





Overall reaction:

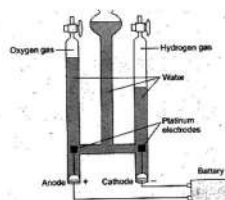
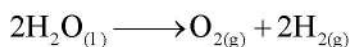


Fig. 7.2 Electrolytic cell showing electrolysis of water

Q.9 Discuss the construction and working of a cell in which electricity is produced.
OR

What is meant by Galvanic cell? Write construction and working of Daniel's cell.

Ans: Galvanic cell:

The electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called Galvanic or Voltaic cell.

Examples: Daniel cell, Dry cell etc.

Introduction:

A. Volta (1745-1827) was an Italian physicist known especially for the development of the first electric cell in 1800.

Construction and working of Daniel cell:

a. Construction: galvanic cell consists of two cells, each called as half cell, connected electrically by a salt-bridge. In each of the half-cell, an electrode is dipped in 1M solution of its own salt and connected through a wire to an external circuit.

i. Left half cell (Oxidation half-cell):

The left half cell consists of an electrode of zinc metal dipped in 1 M solution of zinc sulphate.

ii. Right half Cell (Reduction half cell):

The right half cell is a copper electrode dipped in 1M solution of copper sulphate.

iii. Salt bridge:

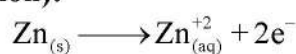
Salt bridge is a U shaped glass tube. It consists of saturated solution of strong electrolyte supported in a jelly type material. The ends of the U tube are sealed with a porous material like glass wool.

Function of the salt bridge:

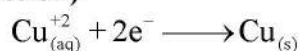
To keep the solutions of two half cells neutral by providing a pathway for migration of Ions.

b. Working of the Cell: The Zn metal has tendency to lose electrons more readily than copper. As a result oxidation takes place at Zn-electrode. The electrons flow from Zn-electrode through the external wire in a circuit to copper electrode. These electrons are gained by the copper ions of the solution and copper atoms deposit at the electrode. The respective oxidation and reduction processes going on at two electrodes are as follows:

Reaction at anode (oxidation):



Reaction at cathode (reduction)



Overall galvanic reaction is the sum of these two half cell reactions:-



As a result of redox reaction electric current is produced.

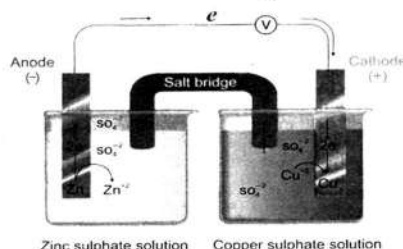


Fig. 7.3 A Daniel Cell

Uses:

The batteries which are used for starting automobiles, running calculators and toys etc. work on the principle of Galvanic cell.

Q.10 Differentiate between electrolytic and galvanic cell

Ans: The Comparison of Electrolytic and Galvanic Cells

Electrolytic cell	Galvanic cell
i. It consists of one complete cell, connected to a battery.	i. It consist of two half cells connected, through a salt bridge.
ii. Anode has positive charge while cathode has negative charge.	ii. Anode has negative charge while cathode has positive charge.
iii. Electrical energy is converted into chemical energy.	iii. Chemical energy converted into electrical energy.
iv. Current is use for a non-spontaneous chemical reaction to take a place.	iv. Redox reaction takes place spontaneously and produce electric current.
Examples: Down's cell, Nelson cell	Examples: Daniel cell, dry Cell etc.

Q.11 How sodium metal is manufactured from fused NaCl?

Ans: Manufacture of Sodium Metal from Fused NaCl:

Principle:

On the industrial scale molten sodium metal is obtained by the electrolysis of fused NaCl in the Downs cell. Construction of Down's cell.

Construction and working of Down's cell:

a. Construction

This electrolytic cell is a circular furnace. In the center there is a large block of graphite, which acts as an anode while cathode around it is made of iron.

b. Working:

i. Ionization:

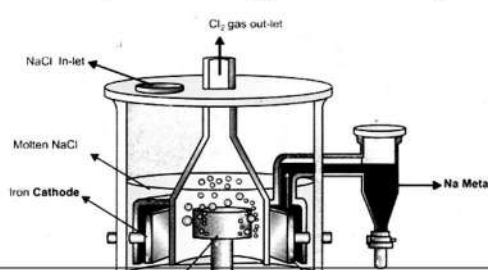
The fused NaCl produces Na^+ and Cl^- ions, which migrate to their respective electrodes on passage electric current.

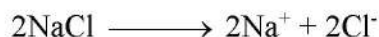
ii. Separation of electrodes:

The electrodes are separated by steel gauze to prevent the contact between the products.

iii. Reactions:

The Cl^- ions are oxidized to give Cl_2 gas at the anode. It is collected over the anode within an inverted cone-shaped structure while Na^+ are reduced at cathode and molten Na metal floats on the denser molten salt mixture from where it is collected in a side tube. Following reactions take place during the electrolysis of the molten sodium chloride:



Ionization of molten NaCl**Reaction at anode (oxidation)****Reaction at cathode (reduction)****Overall reaction**

Q.12 How can we prepare NaOH on commercial scale? Discuss its chemistry along with diagram?

Ans: Manufacture of NaOH (Caustic soda):

Principle:

On industrial scale caustic soda, sodium hydroxide NaOH, is produced in Nelson's cell by the electrolysis of aqueous solution of NaCl, called brine.

Construction and working of Nelson's cell:

a. Construction: It consists of a steel tank in which graphite anode is suspended in the center of a U shaped perforated iron cathode. This iron cathode is internally lined with asbestos diaphragm. Electrolyte (brine) is present inside the iron cathode.

b. Working:**i. Ions Present:**

Aqueous solution of sodium chloride consists of Na^+ , Cl^- , H^+ and OH^- ions. These ions move towards their respective electrodes and redox reactions take place at these electrodes. When electrolysis takes place Cl^- ions are discharged at anode and Cl_2 gas rises into the dome at the top of the cell. The H^+ ions are discharged at cathode and H_2 gas escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.

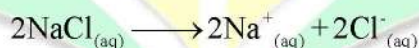
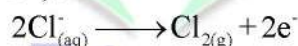
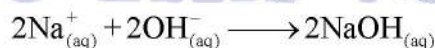
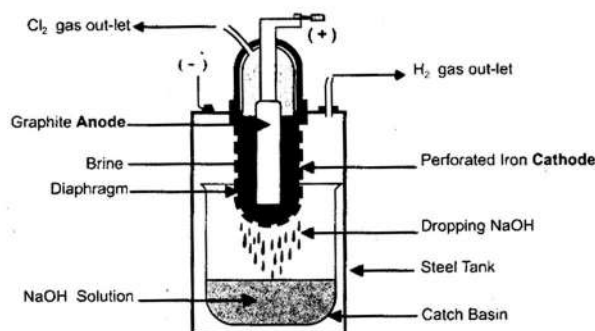
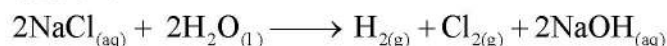
Ionization of Brine:**Reaction at anode (oxidation):****Reaction at cathode (reduction):****Overall cell reaction:**

Fig. 7.5 Nelson's Cell for production of NaOH

Q.13 What is Corrosion? How iron gets rusted?

Ans: Corrosion:

Corrosion is slow and continuous eating away of a metal by the surrounding medium. Corrosion is a redox chemical reaction that takes place by the action of air and moisture with the metals. The **most common example** of corrosion is **rusting of iron**.

Rusting of Iron:

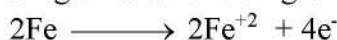
Corrosion is a general term but corrosion of iron is called rusting. Formation of hydrated iron oxide ($\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$) at surface of iron is called rusting.

Conditions for rusting:

The important condition for rusting is moist air (air having water vapours in it). There will be no rusting in water vapours free of air or air free of water.

Process of rusting:**i. Anodic region:**

Stains and dents on the surface of the iron provide the sites for this process to occur. This region is called anodic region and following oxidation reaction takes place here:



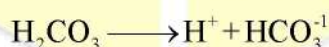
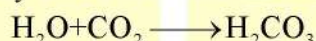
This loss of electrons damages the object.

ii. Cathodic region

The free electrons move through iron sheet, until they reach to a region of relatively high O_2 concentration near the surface surrounded by water layer. This region acts as cathode and electrons reduce the oxygen molecule in the presence of H^+ ions:

Provision of H^+ ions:

The H^+ ions are provided by the carbonic acid, which are formed because of presence of CO_2 in water. That's why acidic medium accelerates the process of rusting



The overall redox process is completed without the formation of rust

**Formation of rust:**

The Fe^{+2} formed spreads through out the surrounding water and react with O_2 to form the salt $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ which is called **rust**. It is also a redox reaction.

The rust layer of iron is porous and does not prevent further corrosion. Thus rusting continues until all the piece of iron is eaten up

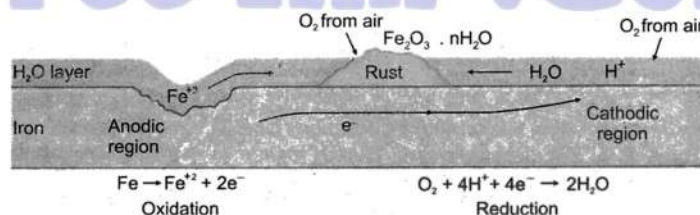


Fig. 7.6 Rusting of iron.

Q.14 Describe the methods for the prevention of corrosion.**Ans: Corrosion:**

Correction is a slow and continuous eating away of metal by the surrounding medium. The most common example of corrosion is rusting of iron.

Methods for the prevention of corrosion:**i. Removal of stains**

The regions of stains in an iron rod act as the site for corrosion. If the surface of iron is properly cleaned and stains are removed. It would prevent corrosion.

ii. Paints and greasing

Greasing, Polishing or painting of the surface can prevent the corrosion of iron. With development of technologies, modern paints contain a combination of chemicals called stabilizers that provide protection against the corrosion in addition to prevention against the weathering and other atmospheric effects.

iii. Alloying

Alloy is a homogeneous mixture of one metal with one or more other metals or non-metals. Alloying of iron with other metals has proved to be very successful technique against rusting. The best example of alloying is the 'stainless steel', which is a good combination of iron, chromium and nickel.

iv. Metallic coating

The best method for protection against the corrosion of metals exposed to acidic conditions is coating the metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion. It is the most widely applied technique in the food industry where food is 'tin-packed'. The containers of iron are coated with tin or chromium to give it a longer life.

Method of metallic coating:

a. Physical Methods

b. Chemical method (Electroplating)

1. Zinc coating or Galvanizing

"The process of coating a thin layer of zinc on iron is called galvanizing."

This process is carried out by dipping a clean iron sheet in a zinc chloride bath and then heating it. After this iron sheet is removed, rolled into molten zinc metal bath and finally air-cooled. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

2. Tin Coating or Tinning:

The process of coating a thin layer of tin on iron is called tinning.

It involves the dipping of the clean sheet of iron in a bath of molten tin and then passing it through hot pairs of rollers. Such sheets are used in the beverage and food cans. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken. The tin protects the iron only as long as its protective layer remains intact. Once it is broken and the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Q.15 What is electroplating? Write down procedure of electroplating?

Ans: Electroplating:

Electroplating is depositing of one metal over the other by means of electrolysis.

Objective

This process is used to protect metals against corrosion and to improve their appearance, shine and beauty.

Principle:

Principle of electroplating is to establish an electrolytic cell in which anode is made of the metal to be deposited and cathode of the object on which metal is to deposit. The electrolyte is in aqueous solution of a salt of the respective metal.

Procedure for Electroplating:

i. Cleaning of metal:

In this process the object to be electroplated is cleaned with sand and washed with caustic soda solution.

ii. Anode: The anode is made of the metal, which is to be deposited like Cr, Ni. The

iii. Cathode: The cathode is made up of the object that is to be electroplated like some sheet made up of iron.

iv. Electrolyte: The electrolyte in this system is a salt of the metal being deposited.

The

v. Electrolytic Tank: The electrolytic tank is made of cement, glass or wood in which anode and cathode are suspended.

The electrodes are connected with a battery. When the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and discharge or deposit on the cathode (object). As a result of this discharge, a thin layer of metal deposits on the object, which then pulled out and cleaned.

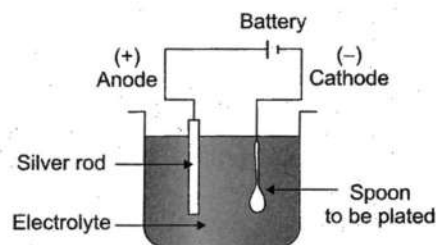


Fig. 7.7 Electroplating of an object.

Q.16 Describe electroplating of Silver, Chromium, Zinc and Tin in detail.

Ans: a. Electroplating of Silver

Principle:

The electroplating of silver is carried out by establishing an electrolytic cell.

Anode: The pure piece of silver strip acts as anode that is dipped in silver nitrate solution.

Cathode: The cathode is the metallic object to be coated such as silver spoon.

Chemical reaction: When the current is passed through the cell, the Ag^+ ions dissolve at the anode and migrate towards the cathode where, they discharge and deposit on the object e.g. spoon. The chemical reaction can be represented as:

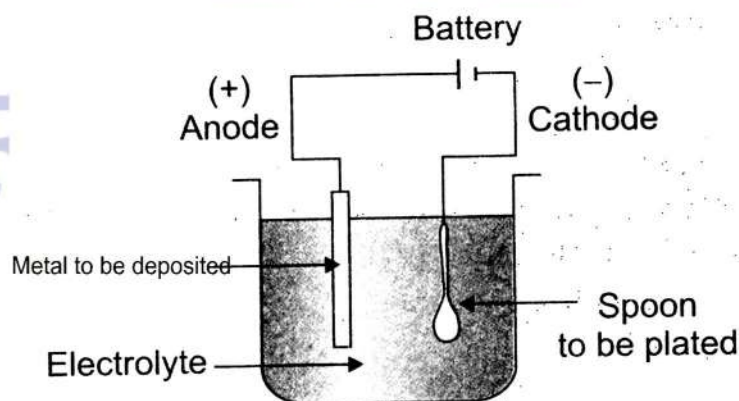
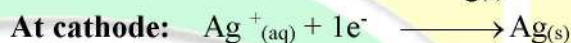
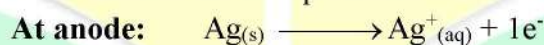


Fig. 7.7 Electroplating of an object.

Uses:

Common examples of silver plating are table wares, cutlery, jewelry and steel objects.

b. Electroplating of Chromium

Principle: The electroplating of chromium is carried out by establishing an electrolytic cell.

Electrolyte: Aqueous solution of chromium sulphate containing a little sulphuric acid, acts as an electrolyte.

Cathode: The object to be electroplated acts as cathode

Anode: Anode is made of antimonial lead.

Chemical reactions: The electrolyte ionizes and provides Cr^{3+} ions, which reduce and deposit at cathode. Electrolyte produces the following ions.



Reactions at the electrodes are as follows.

At anode:



At cathode:

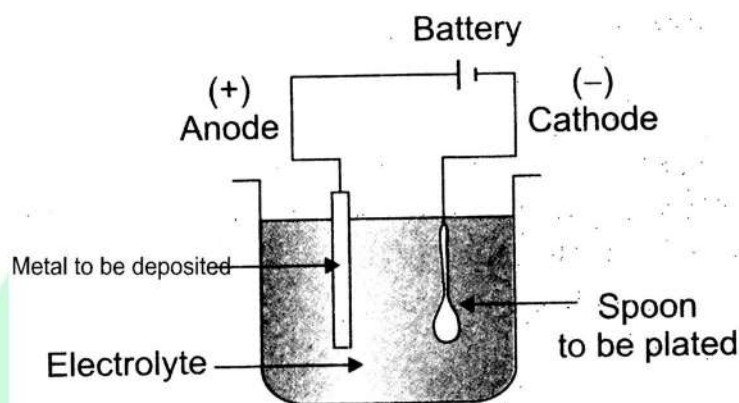


Fig. 7.7 Electroplating of an object.

Why steel is usually plated first with nickel or copper and then by chromium?

For practical convenience, the steel is usually plated first with nickel or copper and then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off. The nickel or copper provides adhesion and then chromium deposited over the adhesive layer of copper lasts longer. This type of electroplating resists corrosion and gives a bright silvery appearance to the object.

c. Electroplating of zinc

Principle:

The electroplating of zinc is carried out by establishing an electrolytic cell.

Cleaning of target metal: The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales.

Anode: Zinc metal is used as anode.

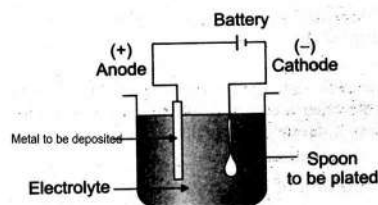


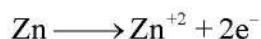
Fig. 7.7 Electroplating of an object.

Cathode: The metal to be plated is used as cathode.

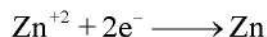
Electrolyte: Aqueous solution of zinc sulphate (ZnSO_4) is used as an electrolyte.

Next, the zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied, which results in zinc being deposited on the target metal i.e. cathode.

At anode:



At cathode:



d. Tin coating:

Principle:

The electroplating of tin is carried out by establishing an electrolytic cell.

Cleaning of target metal: The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales.

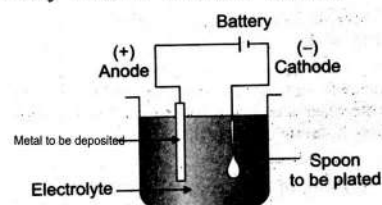


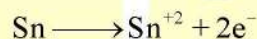
Fig. 7.7 Electroplating of an object.

Anode: Tin metal is used as anode.

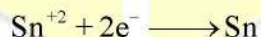
Cathode: The metal to be plated is used as cathode.

Electrolyte: Aqueous solution of tin sulphate (SnSO_4) is used as an electrolyte. Tin is usually electroplated on steel by placing the steel into a container containing a solution of tin salt. The steel is connected to an electrical circuit, acting as cathode. While the other electrode made of tin metal acts as anode. When an electrical current passes through the circuit, tin metal ions present in the solution deposit on steel.

At anode:



At cathode:



d. Electrolytic refining of copper

Impure copper is refined by the electrolytic method in the electrolytic cell.

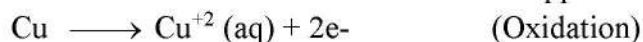
Anode: Impure copper acts as anode

Cathode: a pure copper plate acts as cathode

Electrolyte: Copper sulphate solution in water is used as an electrolyte.

Reaction at anode:

Oxidation reaction takes place at the anode. Copper atoms from the impure copper lose electrons to the anode and dissolve in solution as copper ions:



Reaction at cathode:

Reduction reaction takes place at the cathode. The copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.



In the process impure copper is eaten up and purified copper atoms deposit on the cathode. 99.9% pure copper is obtained in this process.

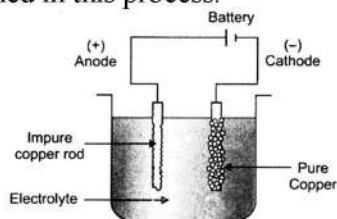


Fig. 7.8 Refining of copper in an electrolytic cell.

A comparison of effect of Al_2O_3 and Fe_2O_3 on parent metal

Aluminium has a great tendency to corrosion. However, aluminium corrosion is aluminium oxide (Al_2O_3), a very hard material that actually protects the aluminium from further corrosion. Aluminium oxide corrosion also looks a lot more like aluminium, so it isn't as easy to notice.

When iron corrodes the color changes and it actually expands. This expanding and color change can produce large red flakes that we all know as rust. Unlike aluminium oxide, the expanding and flaking of rust in iron exposes new metal to further rusting.

Chemistry of Photography

In early nineteenth century photographers produced crude images using papers covered with silver nitrate or silver chloride. The exposure of light on photographic plate initiated chemical reaction. The light exposed portion became dark depending the amount or time of exposure. That exposed plate was later on developed to show the image. Those early days "photographs" darkened with time because of ongoing chemical reaction on them. Later on procedures were developed to make the image permanent by use of mercury vapors, followed by washing with sodium hyposulfite ($\text{Na}_2\text{S}_2\text{O}_3$). It dissolved away the silver iodide from the unexposed portion of the plate, and stopped the reaction further. Although technologically more advanced, the basic procedures developed originally are still used in all silver-based photography today.

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