Chapter 7

ELECTROCHEMISTRY

CONCEPTUAL LINKAGE

Before reading this chapter, the student must know the:

- Charge concept.
- Basics of Chemical reactions
- Basic structure of periodic table

TIME ALLOCATION

Teaching periods	= 16
Assessment periods	= 03
Weightage	= 14%

LEARNING OUTCOMES

After reading this chapter, the Student will be able to:

- Define Oxidation and Reduction in terms of loss or gain of oxygen and hydrogen. (Understanding)
- Define oxidation and reduction in terms of loss or gain of electrons. (Understanding)
- Identify the oxidizing and reducing agents in redox reactions. (Analyzing)
- Identify the oxidizing and reducing agents in redox reactions. (Understanding)
- Define oxidation state. (Remembering)
- State the rules for assigning the oxidation numbers. (Remembering)
- Determine the oxidation numbers of element of any atom in a compound. (Applying)
- Describe the nature of electrochemical processes. (Understanding)
- Sketch an electrolytic cell, with labels. (Understanding)
- Identify the direction of movement of cations and anions towards respective electrodes. (Understanding)
- List the possible uses of an electrolytic cell. (Understanding)
- Sketch a daniell cell, with labeling. (Understanding)
- Describe how battery produces electrical energy. (Understanding)

7.0 Introduction to Electrochemistry

Electrochemistry is a branch of chemistry in which the conversion of chemical energy into electrical energy and electric change produced by an electric current are studied. The electrochemistry also deals with situations where oxidation and reduction reactions takes place, but the direct charge transfer in the form of electrons from one species to another is not the topic of electrochemistry.

Electrochemistry is an important branch of chemistry and many important topics like metallurgy of metals, analysis of chemical compounds, battery operations, electroplating etc are discussed here.

7.1 Oxidation and Reduction

Oxidation and Reduction may be defined in many ways, these definitions can be summarized in the following manner,

Oxidation:

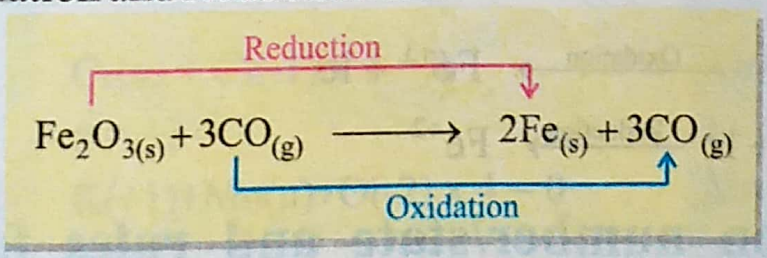
i) Oxidation is gain of oxygen
ii) Oxidation is loss of hydrogen
iii) Oxidation is loss of electrons
iv) Oxidation involve an increase in oxidation state

Reduction may be defined as:

i) Reduction is loss of oxygen
ii) Reduction is gain of hydrogen
iii) Reduction is gain of electron
iv) Reduction involve a decrease in oxidation state

These point are further explained by the help of following examples.

In the extraction of iron (Fe) from its ore Iron oxide (Fe_2O_3), the processes of Oxidation and Reduction are:



As shown here, 'Fe_2O_3' is reduced to Fe, and 'CO' is oxidized to 'CO_2', because here it can be seen that from reducing species 'Fe_2O_3' the oxygen is being eliminated, while from the oxidizing species i.e. the 'CO' the oxygen is being added.

In the following case of ethanol,

The Ethanol (or Ethyl alcohol) is oxidized by loss of hydrogen.

$$CH_3CH_2OH_{(l)} \longrightarrow CH_3CHO_{(l)} + H_{2(g)}$$

Oxidation by loss of hydrogen

And in this case of Ethanal (also called Acetaldehyde), the reduction is brought about by the gain of hydrogens.

Reduction by gain of hydrogen

$$H_{2(g)} + CH_3CHO_{(l)} \longrightarrow CH_3CH_2OH_{(l)}$$

The modern view of explaining the Oxidation and Reduction processes are in terms of electron transfer,

As described above the oxidation is loss of electrons, while Reduction is the gain of electrons.

Reduction by gain of electrons

$$Cu^{2+} + Mg_{(s)} \longrightarrow Cu_{(s)} + Mg^{2+}$$

Oxidation by loss of electrons

In this above example, the Copper (Cu) is being reduced to Cu from Cu^{+2}, by gaining electrons, while Magnesium is oxidized to 'Mg^{+2}'by the loss of electrons.

Similarly in the following case (a), the Fe^{+2} is oxidized to Fe^{+3} by the loss of one electron, while in case (b) the reaction is reversed and the process is the reduction.

(a) $Fe^{+2} \xrightarrow{\text{Oxidation}} Fe^{+3} + 1\bar{e}$

(b) $Fe^{+3} + 1\bar{e} \xrightarrow{\text{Reduction}} Fe^{+2}$

7.2 Oxidation number/state and rules for assigning oxidation numbers.

The charge which an atom bears in a compound, molecules or ions is called its oxidation number; the oxidation number may be positive or negative

depending upon the electronegativity of an atom.

Following rules are applied to calculate the oxidation number.

i) The oxidation state of a free element is always zero. For example, 'O_2', 'Cl_2', 'S_8', and 'Si' have zero oxidation state.

ii) The sum of the oxidation states of all the atoms or ions in a neutral compound is zero.

iii) The sum of the oxidation states of all the atoms in an ion is equal to the charge on the ion.

iv) The more electronegative element in a substance is given a negative oxidation state. The less electronegative one is given a positive oxidation state.

In the following table 7.1 common oxidation numbers of some elements are written.

Group	Elements	Oxidation State
1	Na, K etc.	'+1'
2	Mg, Ca, etc.	'+2'
17	F, Cl, etc.	Generally '-1'
16	O	Generally '-2'
1	H	+1 except in metallic hydride where its oxidation number is '-1'.

Table 7.1 Common oxidation states of some Groups in periodic table.

The oxidation state of some elements are calculated in neutral compounds are given below in the table 7.2:

Substance	Oxidation Number Calculation	Oxidation Number of the Elements
$\underline{Cr}Cl_3$	$Cr(n) + Cl(-1) \times 3 = 0$ $n = +3$	$Cr = +3$
$K\underline{Mn}O_4$	$K(+1) + Mn(n) + O(-2) \times 4 = 0$ $n = +7$	$Mn = +7$
$K_2\underline{Cr_2}O_7$	$K(+1) \times 2 + Cr(n) \times 2 + O(-2) \times 7 = 0$ $n = +6$	$Cr = +6$

Table 7.2: Oxidation numbers of selected elements in some compounds.

Activity

Calculate the oxidation number of 'N' in following:

(i) HNO_3 (ii) NO_2 (iii) NOCl

7.3 Oxidizing and Reducing Agents

Substances that bring oxidation process are known as oxidizing agents. In a broad sense, the oxidation, which is defined as the loss of electrons, so any substance which brings loss of electrons or accepts electrons during the chemical reactions is termed as the oxidizing agents, oxidants or oxidizers. The oxidant removes electrons from another substance, and is thus reduced itself. And because it "accepts" electrons it is also called an electron acceptor. Oxidants are usually chemical substances with elements in high oxidation numbers (e.g., H_2O_2, H_2SO_4, $K_2Cr_2O_7$, MnO_4^-, CrO_3, $Cr_2O_7^{2-}$, OsO_4) or highly electronegative substances that can gain one or two extra electrons by oxidizing a substance (O, F, Cl, Br).

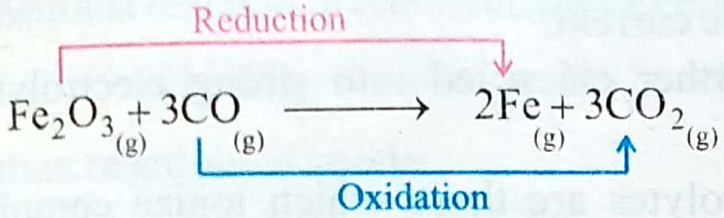
The reducing agents are just opposite to the oxidizing agents. The substances that have the ability to reduce other substances are said to be reductive and are known as reducing agents, reductants, or reducers. The reductant transfers electrons to another substance, and is thus oxidized itself. It "donates" electrons it is also called an electron donor. Metals (Li, Na, Mg, Fe, Zn, Al) can be used as reductants. These metals donate or give away electrons readily. The hydrogen gas (H_2) with a palladium, platinum, or nickel catalyst can also be used as the reducing agent. These catalytic reductions are primarily used in the reduction of carbon-carbon double or triple bonds.

Redox process is that the reductant transfers electrons to the oxidant. Thus, in the reaction, the reductant or reducing agent loses electrons and is oxidized and the oxidant or oxidizing agent gains electrons and is itself reduced.

7.4 The Oxidation-reduction Reactions (Redox Reaction)

As described earlier, the oxidation and reduction go side by side in a reaction. A substance which oxidizes the other is itself reduced and vice versa.

For example:

i) 

ii) Another example is the reaction between the H_2 and F_2

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

Here, The H_2 loses two electrons, which are accepted by the F_2,

$$H_{2(g)} \longrightarrow 2H^+ + 2\bar{e} \quad \text{(Oxidation)}$$

$$F_{2(g)} + 2\bar{e} \longrightarrow 2F^- \quad \text{(Reduction)}$$

$$2H^+ + 2F^- \longrightarrow 2HF_{(g)}$$

$$H_{2(g)} + F_{2(g)} \longrightarrow 2HF_{(g)} \quad \text{(Overall reaction)}$$

As seen, the overall reaction results in the formation of hydrogen fluoride (HF).

The redox reactions are very important to understand, both in research and industrial applications, as most of reactions involve the transfer of electrons from one species to the other, and are thus categorized as the redox reactions. There understanding hence is used to increase the desired products by suitable procedures.

7.5 The Electrochemical Cell

An electrochemical cell is a device containing at least two electrodes in the solution of an electrolyte, and here in this cell the oxidation-reduction (Redox) reactions takes place.

The electrochemical cells are of two types, i.e.

i) Electrolytic cell
ii) Galvanic cell

These types are described in detail in the following pages, but let's first understand the term electrolytes.

7.5.1 The Concept of Electrolytes

The electrolyte is a substance which is capable of conducting electricity in its solution or molten state. The electrolytes are generally

consists of ions which serve as the carrier of electrons and hence conduct easily the electric current.

They may be further classified into strong electrolytes and weak electrolytes.

The strong electrolytes are those which ionize completely in their aqueous solution, e.g. NaCl, HCl , and NaOH etc are some strong electrolytes.

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

The weak electrolytes are those which do not ionize in greater extent, e.g. acetic acid.

$$CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

The electrolytes although being composed of ions do not conduct electricity in the solid state because in solid state the mobility of ions is very low.

7.5.2 Electrolytic Cells

An electrolytic cell is the cell that contains a solution of electrolytes and has been fitted with two electrodes, anode and the cathode which are attached with a source of direct current (DC). In this cell a non spontaneous reaction takes place using the electrical energy and some products are formed.

In the electrolytic cell when electricity is passed the ions present in the solution are attracted towards the respective electrode, i.e. anions towards anode and cations toward cathode. The electrode that attracts anions is called anode and the electrode that attract cations is called cathode.

In case of the NaCl,the Cl^- ions move towards the anode and loss their electrons there, hence oxidation reaction occurs here, while the Na^+ ions moves toward cathode where reduction reaction occurs, and Na^+ gains electron yielding neutral products.

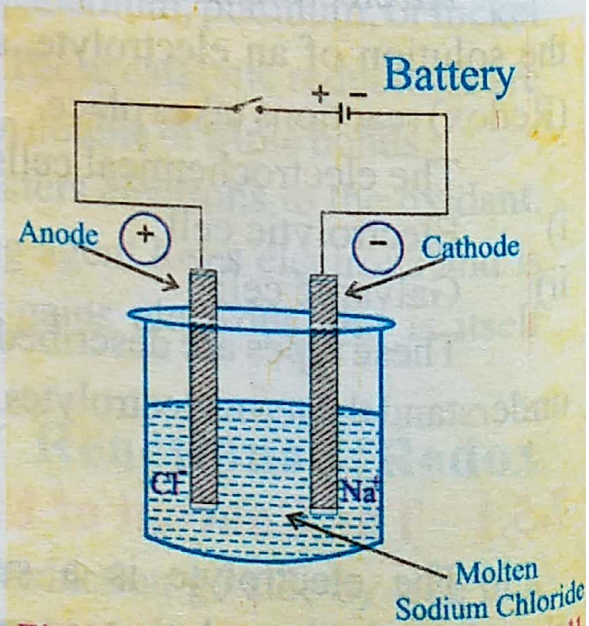


Figure 7.1: A typical electrolytic cell

The chemical reaction in this electrolytic cell are following:

$$NaCl_{(s)} \longrightarrow Na^{+}_{(l)} + Cl^{-}_{(l)}$$

Oxidation reactions at anode:

$$2Cl^{-}_{(l)} \longrightarrow Cl_{2(g)} + 2e$$

Reduction reaction at cathode:

$$2Na^{+}_{(l)} + 2e \longrightarrow 2Na_{(l)}$$

Overall the whole process can be written as:

$$2NaCl_{(s)} \xrightarrow[\text{Heat}]{\text{Electrotytic cell}} 2Na_{(l)} + Cl_{2(g)}$$

The electrolytic cells are very important. They are used in electrorefining process, through which high grade metals and other substances are produced. These cells are also used in electroplating processes.

7.5.3 Galvanic Cells or the Daniell Cell

The Galvanic cell is also a type of electrochemical cell like the electrolytic cells but differs from that in the sense that it is used for the production of electric current or flow of electron current from the spontaneous redox reaction. This cell is also known as Daniell cell or Voltaic cell, named after Luigi Galvani.

This consists of two different metals connected by a salt bridge or a porous disk between the individual half-cells. A Galvanic cell is show in figure 7.2.

Luigi Galvani
(Italy)
Born 1737—died 1798

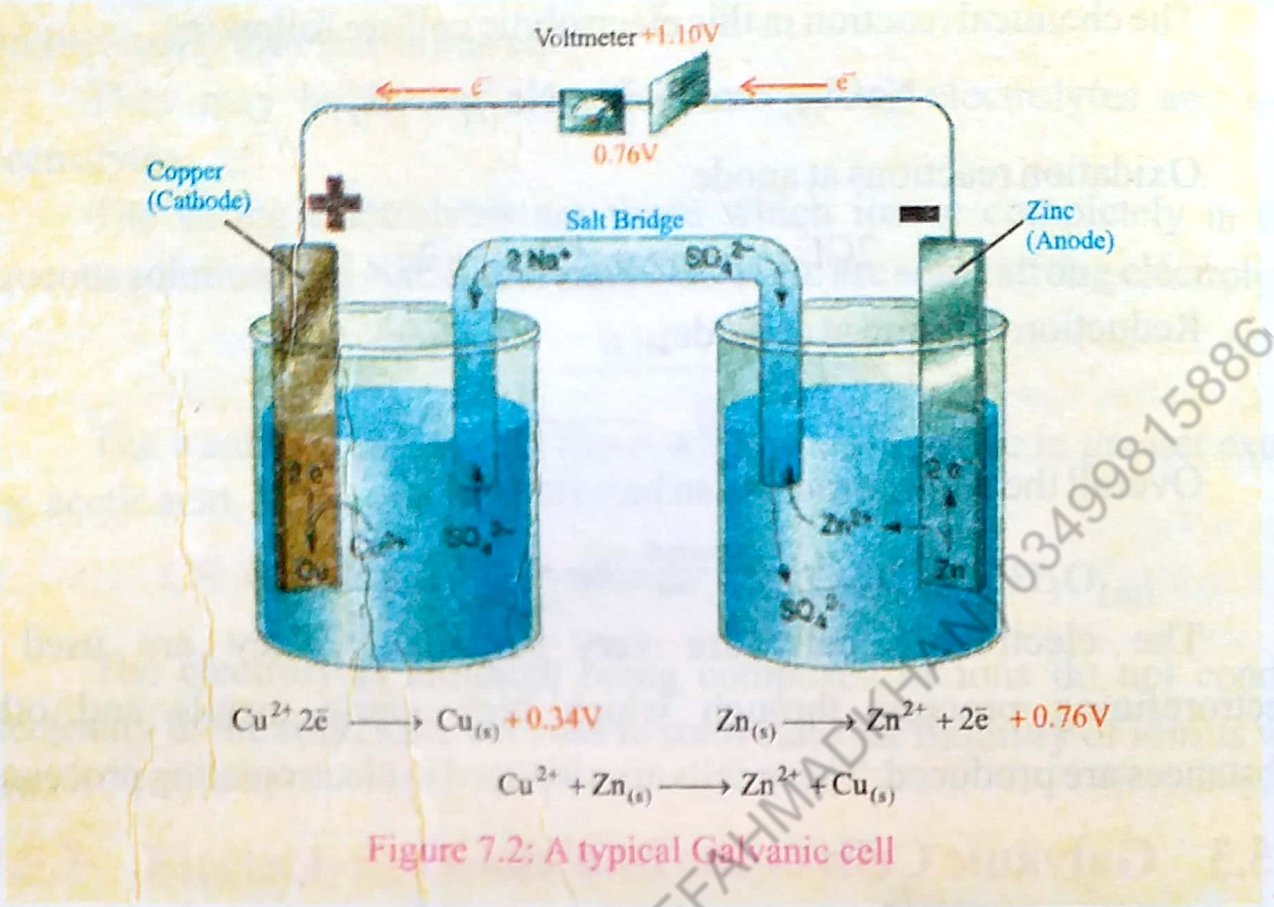


Figure 7.2: A typical Galvanic cell

In the Galvanic cell, the Zn metal has a high tendency to lose electrons than Cu, as a result oxidation reaction takes place at Zn electrode. The electrons flow from Zn electrode through the external wire in the circuit to Cu electrode. These electrons are gained by the Cu ions of the solution and copper atoms deposit at the electrode. As a result of this electric current is produced.

The reaction of this cell are shown here in following.

$$Zn_{(s)} \xrightarrow[E^\circ = 0.76V]{\text{Oxidation}} Zn^{+2} + 2\bar{e}s- \qquad \text{(At the anode)}$$

$$Cu^{++} + 2\bar{e}s \xrightarrow[E^\circ = 0.34V]{\text{Reduction}} Cu_{(s)} \qquad \text{(At the cathode)}$$

That the overall reaction could be:

$$Zn_{(g)} + Cu^{++} \longrightarrow Cu_{(s)} + Zn^{++}$$

The cell voltage can be calculated as:

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

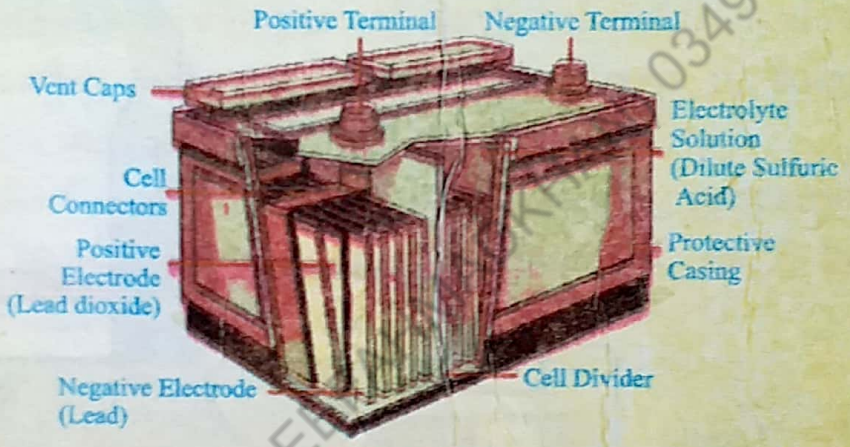
$$E^\circ_{cell} = 0.76 + 0.34 = 1.10V$$

The Galvanic cells are extensively used in the production of electricity through the chemical reaction and the lead––acid batteries are based on this cell.

7.6 Electrochemical Industries

The Electrochemistry occupies a great place in the life standards of modern world. Every institution, whether it is a home or a factory, utilizes the applications based upon the electrochemical instrumentations. For example, the batteries, whether they are the conventional Lead-acid batteries (without

(b) Different Dry Batteries



(a) Lead-Acid Battery

Figure: 7.3 (Different Types of Batteries)

which no vehicle can run) or dry cells batteries or the lithium batteries, (which provides the source of electricity for the most of electronic equipments such as the wrist watches, radios, and even the computers) or the most modern used the sealed batteries (to which you are familiar in your rechargeable emergency lights), all these are the direct result of the electrochemical applications.

The electroplating methods, which are used widely for the prevention of corrosion, and also for the fine looking of the objects (e.g. the gold and other electroplating) also use the electrochemical principles.

The process of electrolysis is also the important example of the electrochemical application. This

Figure 7.4
A Gold plated ornament

process is frequently used in the preparation of many important compounds and for the decomposition of many substances into simpler fragments.

7.6.1 Manufacture Of Sodium Metal from fused NaCl

The Sodium metal which holds an important place is obtained from its common salt, the NaCl. This process is commercially performed in a specialized electrolytic cell called Down's cell.

The Down's cell is shown in figure 7.5.

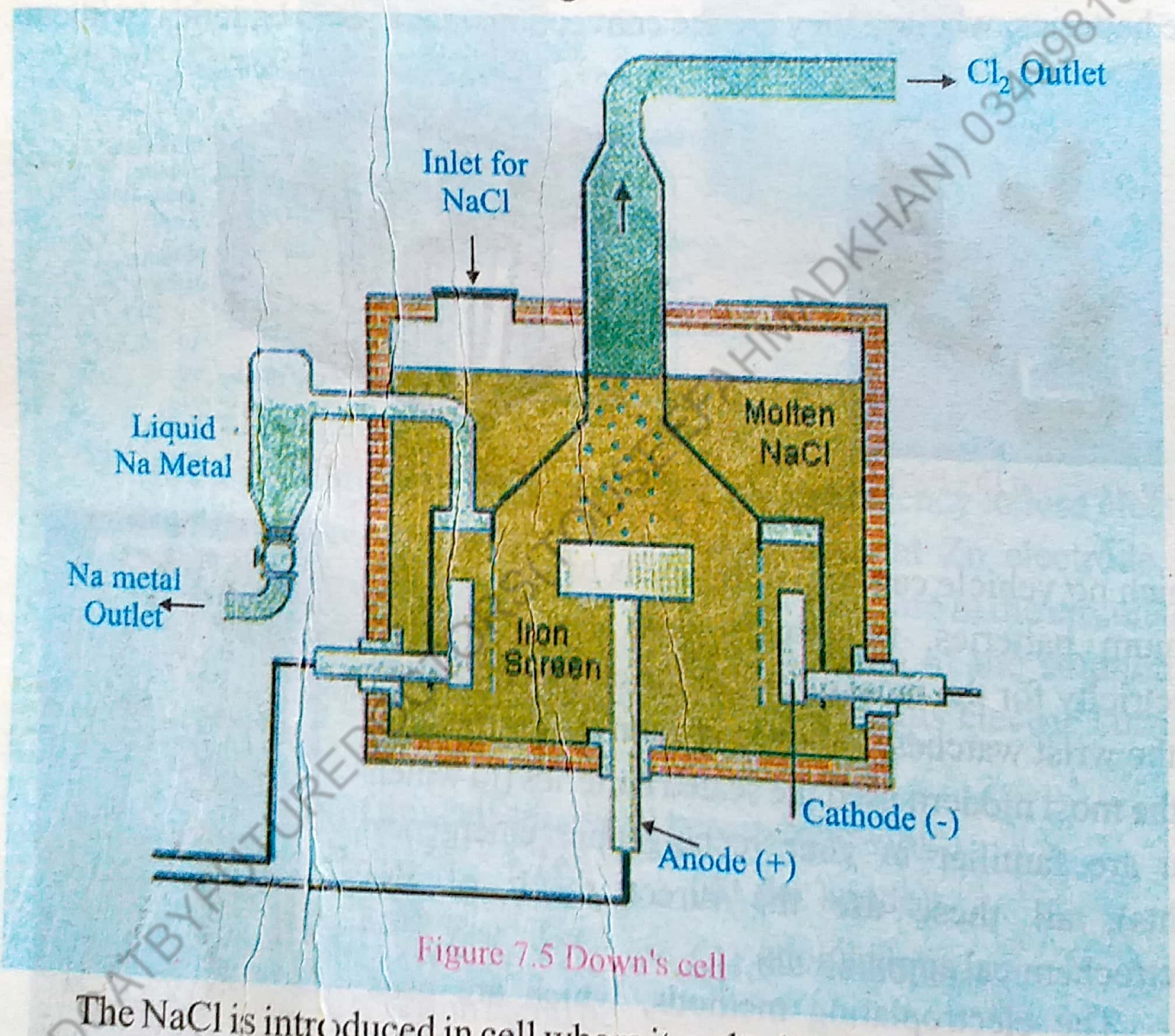


Figure 7.5 Down's cell

The NaCl is introduced in cell where it melted. The electricity is passed through the electrodes. Different reactions of the process are:

$$NaCl_{(s)} \xrightarrow{\text{Heat}} Na^{+}_{(l)} + Cl^{-}_{(l)} \times [2]$$

Oxidation reactions at anode.

$$2Cl^{-}_{(l)} \xrightarrow{\text{oxidation}} Cl_{2(g)} + 2e$$

Reduction reaction at cathode.

$$2Na^{+}_{(l)} + 2e^{-} \xrightarrow{\text{reduction}} 2Na_{(l)}$$

Overall the whole process can be written as:

$$2NaCl_{(s)} \xrightarrow[\text{Heat to melt}]{\text{electrolysis}} 2Na_{(l)} + Cl_{2(g)}$$

As a result Cl_2 is given out from the anode and molten sodium is collected at the cathode.

7.6.2 Manufacture of NaOH from Brine

The NaOH or Caustic soda is a very important chemical and is used in countless industries like in leather industry, in fabric industry, in soap industry, paper industry and many more.

It is commonly known as Caustic soda and is industrially prepared from the electrolysis of brine (concentrated aqueous solution of NaCl) in Nelson's cell which is a specialized electrolytic cell.

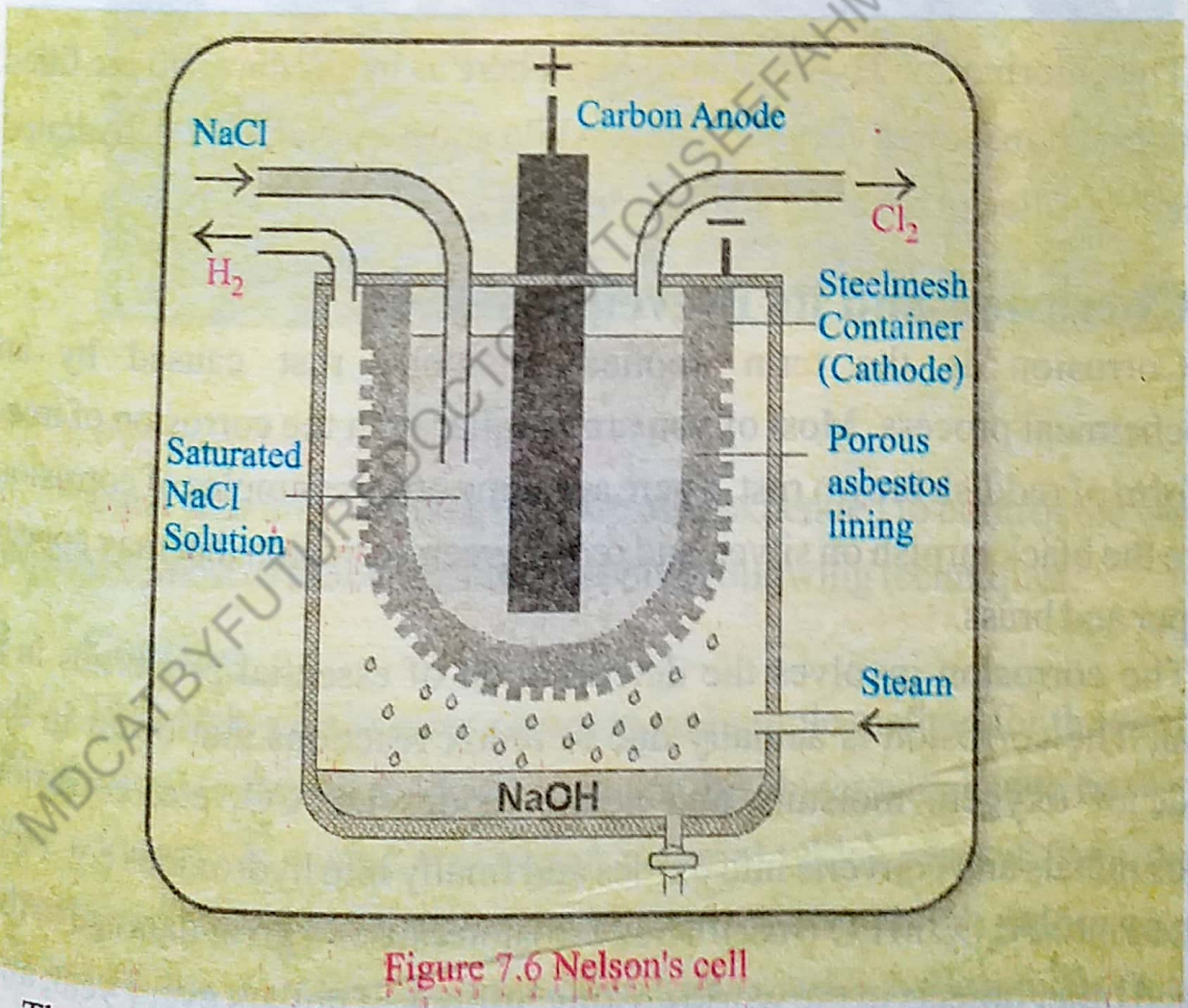


Figure 7.6 Nelson's cell

The process differs from the Down's method in the sense that in brine

some H_2O is also present which is reduced during the process and form OH^- ions. Na^+ and OH^- combine and give NaOH.

The Nelson cell is shown in the figure 7.7.

During the electrolysis Cl_2 is given out at the anode, H_2 at the cathode and NaOH is collected from the base of the cell.

Following reaction occur in the cell.

$$NaCl_{(s)} \xrightarrow[H_2O]{\text{Ionization}} Na^+_{(aq)} + Cl^-_{(aq)} \times [2]$$

$$2Cl^-_{(aq)} \xrightarrow[\text{Oxidation}]{\text{At Anode}} Cl_2 + 2e^-$$

$$2H_2O + 2e^- \xrightarrow[\text{Reduction}]{\text{At Cathode}} H_2 + 2OH^-$$

$$2Na^+ + 2OH^- \xrightarrow{\text{In the cell}} 2NaOH$$

$$2NaCl_{(s)} + 2H_2O_{(l)} \xrightarrow{\text{electricity}} 2NaOH_{(aq)} + Cl_{2(g)} + H_{2(g)}$$

The chlorine and Hydrogen produced here as by product and are filled in cylinders, transported for marketing, while aqueous NaOH is dehydrated and dried up, filled in bags and is sold in market.

7.7 Corrosion and its Prevention

Corrosion is the term applied to metal rust caused by an electrochemical process. Most of you are familiar with the corrosion of iron, in the form of reddish brown rust. There are some other examples of corrosion too, like the black tarnish on silver, and red or green corrosion that may appear on copper and brass.

The corrosion involves the deterioration of essential properties in a material. The corrosion is actually due to redox reactions that occur in the presence of oxygen, moisture and acidic oxides i.e. CO_2 etc. corrosion damages metals and converts into oxides and finally into hydroxides e.g. Fe is converted into $Fe_2O_3 . nH_2O$ (rust). The corrosion causes a great damage to the metal and other objects in the world costing billions of dollars every year, and also causes accidents resulting in the loss of precious loss of lives.

7.7.1 Rusting Of Iron

The rusting of iron chemically means the electrochemical reaction of iron with oxygen and water to form $Fe(OH)_3$, $Fe(OH)_2$, or even $Fe_2O_3.nH_2O$.

Chemically the corrosion of Iron is written as:

$$4Fe_{(s)} + 3O_{2\,(g)} + H_2O_{(\ell)} \longrightarrow \underset{\text{(Hydrated Ferric oxide)}}{2Fe_2O_3.nH_2O_{(s)}}$$

You may have observed this rusting around yourself at many places, the rusting of Iron which causes great loss of money and efforts can be reduced by applying following operations.

1. By painting the surface.
2. By coating the surface with a thin film of oil or grease.
3. By electroplating iron with some non-corrosive metal such as nickel, chromium, copper, etc.

Figure 7.7 Rust in Iron

7.7.2 Prevention of Corrosion

Each year billions of dollars are also expended to control the corrosion processes, most of these processes involve following techniques.

i) Coating

The coating is the most common and simple method for the prevention of corrosion. In this method, the metal or the object which is to be prevented from corrosion is coated with some material which seals it from the outer atmosphere; the most common coatings are done with the **paints or grease**, this cut off the metal from atmosphere and reaction between the O_2 and the metal is prevented.

ii) Electroplating

The electroplating, in which the metal which is being protected by corrosion is electroplated by some substance which is resistant to corrosion, e.g the ornaments are electroplated with the gold, or some objects with the copper. This reduces the corrosion significantly.

iii) Alloys Forming

In this method, two or some time more metals are mixed to obtain a mixture of these metals, called **alloy**, which is very resistant to corrosion, e.g. the steel is an alloy of iron, chromium and nickel, Brass is an alloy of copper and zinc. These alloys are very resistant to corrosion.

7.8 The Electroplating Process

The Electroplating is the process of coating a metal object with a thin layer of another metal using the process of electrolysis. Usually the electroplated coating is about 0.002 inch (0.05 mm) thick.

Electroplating is carried out by constructing an electroplating cell in which the metal to be plated acts as a cathode and the metal to be deposited acts as anode, which is immersed in an aqueous solution of its salt.

During electrolysis following reaction occur.

$$M \xrightarrow[\text{Oxidation}]{\text{At Anode}} M^{++} + 2\bar{e} \qquad \text{(Oxidation)}$$

$$M^{++} + 2\bar{e} \xrightarrow[\text{Reduction}]{\text{At Cathode}} M \qquad \text{(Reduction)}$$

Electroplating is used to give metal objects a better appearance or to protect them from corrosion, wear, or rust. Tableware, trays, decorative pieces, and jewelry are plated with gold or silver to make them more attractive.

There are several types of electroplating processes, classified on the basis of type of compound which is used for electroplating coatings; some of these electroplating processes are described below:

7.8.1 Electroplating of Tin

Tin is a malleable, ductile, highly crystalline, silvery-white metal. This metal resists corrosion from distilled, sea, or soft tap water, but can be attacked by strong acids, alkalis, and by acid salts.

Tin Plating is used extensively to protect both ferrous and nonferrous surfaces. Metals plated using tin are useful for the food processing industry since they are non-toxic, ductile and corrosion resistant. The excellent ductility of tin allows a coated base metal sheet to be formed into a variety of shapes without damage to the surface tin plating layer. Plating with Tin provides sacrificial protection for copper, nickel and other non-ferrous metals, but not for steel.

During electroplating, a mixture of tin salts is dissolved in a water. When an electric current is passed through water, the tin ions will be attracted to a metal object placed in the water, creating a permanent coating of tin.

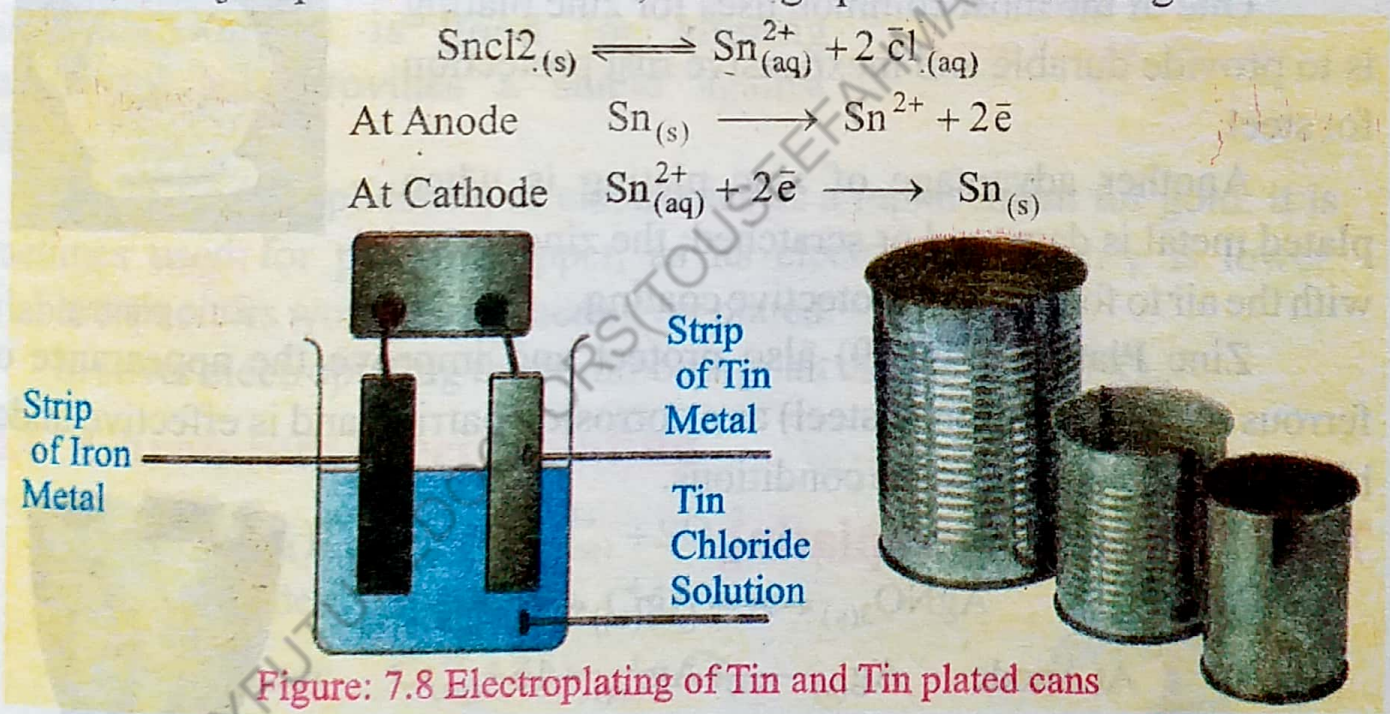


Figure: 7.8 Electroplating of Tin and Tin plated cans

Tin plating is an effective method of protecting base metals like iron or steel from rust and corrosion. While zinc plating can also be used to improve corrosion resistance, zinc is too toxic for food or medical applications, so tin plating must be used for these types of projects. Tin plating also increases the ductility of a metal object, making it easier to weld to another surface.

One potential drawback to tin plating is that it is not as effective as zinc plating, or galvanizing, when it comes to protecting materials from rust and

corrosion. When galvanized metal is damaged or scratched, the zinc reacts with the air to form a new protective coating. When tin plated metal is damaged or scratched, the underlying metal is left exposed to rust and corrosion. Tin is also a relatively soft metal, making it highly susceptible to damage. This material cannot hold up under extremely cold temperatures, and should not be used in freezing conditions.

7.8.2 Zinc Plating

$$ZnCl_{2(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2\bar{e}$$

At Anode $\quad Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2\bar{e}$

At Cathode $\quad Zn^{2+}_{(aq)} + 2\bar{e} \longrightarrow Zn_{(s)}$

Zinc plating is also an important plating process which involves less hazardous chemicals than most other forms of plating, and is successfully done in small shops and even at homes.

One of the most common uses for zinc plating is to provide durable and inexpensive rust protection for steel.

Another advantage of zinc plating is when plated metal is damaged or scratched; the zinc reacts with the air to form a new protective coating.

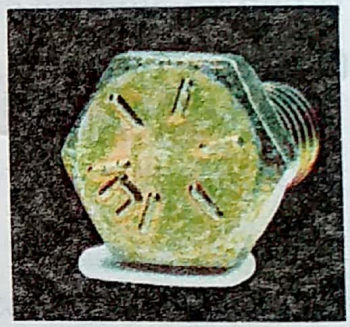


Figure 7.9: Zinc plating

Zinc Plating (Fig.7.9) also protect and improve the appearance of ferrous metals, (i.e. iron & steel) as a corrosion barrier, and is effective under high humidity and moisture conditions.

7.8.3 Silver Electroplating

$$AgNO_{3(s)} \rightleftharpoons Ag^{+}_{(aq)} + NO^{-}_{3(aq)}$$

At Anode $\quad Ag_{(s)} \longrightarrow Ag^{+}_{(aq)} + 1\bar{e}$

At Cathode $\quad Ag^{+}_{(aq)} + \bar{e} \longrightarrow Ag_{(s)}$

Silver is a white colored semi precious metal with an oxidizing property. Due to its property of high electrical conductivity it is used in electronics and semiconductor industries. Silver is also an excellent conductor of heat and offers great solderability. It is malleable and has a high degree of ductibility so it can be

Figure 7.10: Silver platted utensil

bent into any shape and is hence preferred by jewelry and flatware industries for its decorative appeal.

The silver also acts as an economically beneficial metal that may be put to good use for a relatively cheaper price as compared to other precious metals like gold or platinum. (Figure:7.10)

The most common usage of silver plating is for ornamentation of jewelry, plates, cups, trophies and medals. The process adds a perfect polish to the surface on which the plating is done.

Silver plating improves solderability by hardening the surface and increasing wear-ability. It is also used to inhibit corrosion. The plating helps in reducing friction and to improve paint adhesion. It is good for altering conductivity and provides a shield against radiation.

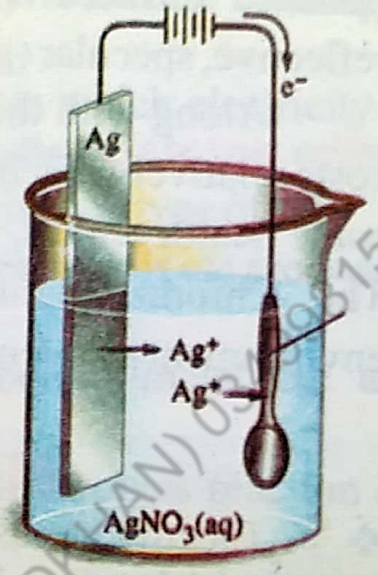


Figure 7.11:
Silver Plating Process

Silver finds application in electronics as a replacement for gold, it is sometimes used for plating copper, as its electrical resistance is lower. Variable capacitors works best when silver plated.

In silver electroplating the common Ag salt used is the $AgNO_3$.

7.8.4 Chromium Electroplating

$$Cr_2(SO_4)_{3(s)} \rightleftharpoons 2Cr^{3+}_{(aq)} + SO^{-2}_{4(aq)}$$

At Anode $\quad Cr_{(s)} \longrightarrow Cr^{+3}_{(aq)} + 3\bar{e}$

At Cathode $\quad Cr^{3+}_{(aq)} + 3\bar{e} \longrightarrow Cr_{(s)}$

The chrome plating involves the electroplating of a thin veneer of chromium onto an underlying metal. Sometimes used for decorating purposes, chrome plating can also serve as a layer of protection as well. Often, chrome plating is applied to metal objects when there is an increased opportunity for corrosion to occur.

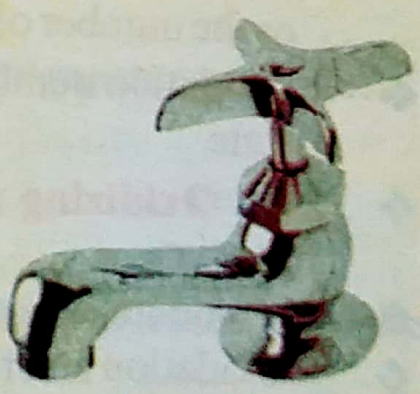


Figure 7.12:
Chromium plated tap

This plating provides corrosion resistance, eases cleaning procedures, increases the surface hardness and also offers a decorative shine.

Though differentiating between chrome plated surface and other polished surfaces is tough; yet it is possible as chrome plating is more reflective, specular (gives mirror like reflection) and bluer than others.

Along with this, chrome plating also have some disadvantages e.g. it costs relatively much higher then the zinc and tin plating, moreover it also has some health effects, as Chromium VI is poisonous and causes even cancer. The Chrome plating also produces some compounds which create environment problems.

SUMMARY OF THE CHAPTER

- Electrochemistry is a branch of chemistry that studies the reactions which take place at the interface of an electronic conductor called electrode.
- In the electrochemistry study of oxidation and reduction processes is carried out.
- Oxidation is the process of adding oxygen, removing hydrogen or removing electrons.
- Reduction is a process of adding hydrogen, removing oxygen or adding electrons.
- The apparent charge upon an atom in a compound is called its **Oxidation state** or **Oxidation number**. Remember it is **quite different from the valency** which is the combing capacity of an atom or the number of unpaired electrons in an atom.
- Oxidation number of an atom /element is always zero when it is in free state.
- An **Oxidizing agent** is the species that causes oxidation of another species.
- A reducing agent is the species that reduces another species.
- Oxidation and reduction reactions involve migration of electrons from one species to another species, these reactions always occur side by side.
- Collectively the oxidation and reduction reactions are called **Redox reactions**.

- The cells where the chemical reaction uses or produces electricity are called **Electrolytic cells**.
- The electrochemical cells are of two types, one that uses electricity is called **electrolytic cells**, in such cells electricity is used to breakup a compound into its constituents.
- Second type of electrochemical cells are those in which electricity is produced by chemical reactions, such cells are called **Galvanic cells** or **Daniell cells**, the Lead-acid battery (car battery) is such type of cell.
- Many important industries e.g electroplating, electrolysis etc are based on electrochemical cells.
- Many important compounds like NaOH, Sodium metal etc are obtained by using electrolytic cells.
- **Corrosion** of metals is a redox reaction, which involves reaction of metals with oxygen, moisture of air.
- Corrosion can be reduced by or prevented by making alloys, by painting or by electroplating processes.

EXERCISE

Q1. Fill in the blanks with appropriate words.

i) Reduction is the process which involves ..... of electrons.

ii) In the free state of an element, the oxidation state has always ..... value.

iii) The Galvanic cell produces ..... using redox reaction.

iv) Coating with paints prevents .....

v) Dry cells are also a type of .....

vi) Electroplating brings both the ..... and resistance to .....

vii) The food cans are electroplated with .....

viii) Caustic soda is industrially prepared by ..... cell.

ix) $K_2Cr_2O_7$ is a strong ..... agent.

x) Collectively the oxidation and reduction processes are called .....

Q2. Choose the correct answer.

i) The Semiconductors are the species that:

(a) conduct electricity easily

(b) do not conduct electricity easily
(c) conduct electricity but not heat
(d) conduct heat but not electricity

ii) Oxidation means:
(a) addition of electrons (b) removal of electrons
(c) addition of protons (d) removal of protons

iii) Oxidation state is the:
(a) apparent +ve charge (b) apparent –ve charge
(c) apparent charge (d) not related with charge

iv) The electroplating prevents:
(a) oxidation (b) reduction
(c) both oxidation and reduction
(d) neither oxidation nor reduction

v) Steel is the alloy of:
(a) iron and mercury (b) mercury and copper
(c) iron and copper (d) iron, chromium and nickel

Q3. Answer the following questions in short.

i) Define electrochemistry. Also state its importance.
ii) What is oxidation? Illustrate with an example.
iii) Explain the term reduction with an example.
iv) Draw a sketch of an electrolytic cell, showing its different parts.
v) What are oxidizing and reducing agents? Explain with example.
vi) What is the difference between weak and strong electrolytes explain with examples.
vii) What is electroplating? Also describe its importance.
viii) What is electrolysis?
ix) Describe importance of Alloy formation in industries.
x) Explain why the food articles are packed in tin plated cans rather than chrome plated cans.

Q4. Answer the following questions with reasoning.

i) Differentiate between oxidation number and valency.
ii) How electricity is produced from chemical reactions?
iii) Explain how oxidation and reduction goes side by side in chemical processes?

iv) Explain how NaCl yields two different products i.e. Na-metal and NaOH by the same process of electrolysis.

v) How you can prevent a piece of iron from corrosion?

Q5. Distinguish between the oxidation and reduction processes?

Q6. What is the oxidation number? Describe the rules to calculate the oxidation number.

Q7. Calculate oxidation number of Mn in $KMnO_4$.

Q8. What do you know about the electrochemical industries?

Q9. What is corrosion? How can you prevent the corrosion process?

Q10. Describe the working of an electrolytic cell, How Sodium metal can be purified by using electrolytic cell?

Q11. What is the electrochemical cell? Discuss the different types of electrochemical cells?

Q12. Discuss the importance of Caustic soda? How it is prepared on industrial scale?

Q13. How the electricity is prepared by chemical means?

Q14. Write a note on electroplating process.