7 ELECTROCHEMISTRY

MULTIPLE CHOICE QUESTIONS

OXIDATION AND REDUCTION

1. **The oxidation state of Group IA:**
   (a) +1 (b) +2
   (c) +3 (d) -1

2. **Oxidation is _____**
   (a) Gain of Oxygen (b) Loss of hydrogen
   (c) Both a and b (d) None of these

3. **Oxidation state of Halogen _____**
   (a) +1 (b) -1
   (c) +2 (d) -2

4. **Complete this equation** $2Cl^{-1} \longrightarrow Cl_2 + ?$
   (a) $2Cl \longrightarrow Cl_2 + 2e^-$ (b) $2Cl \longrightarrow Cl_2 + 2e$
   (c) Both (a) & (b) (d) None of these

ELECTROCHEMICAL CELLS

5. **NaOH is prepared in _____**
   (a) Nelson's Cell (b) Down's cell
   (c) Both a & b (d) None of these

6. **Galvanic cell is also called the**
   (a) Nelson cell (b) Daniel cell
   (c) Electronic cell (d) None of these

7. **On dry cell graphite act an cathode.**
   (a) Zn cup (b) graphite rod
   (c) paste (d) steel rod

8. **In Zn-Cu galvanic cell, Zn is used as**
   (a) cathode (b) anode
   (c) electrode (d) all of above

9. **The reactant that containing the element that is reduced in a chemical reaction is called**
   (a) Oxidizing Agent (b) Galvanic cell
   (c) Both a and b (d) None of these

10. **Nacl is the example of**
    (a) Electrolyte (b) Galvanic cell
    (c) Electrolytic cell (d) None of these

11. **Sucrose is the example of**
    (a) Electrolytic call (b) Non-electrolyte
    (c) Galvanic cell (d) None of these

12. **Cations are _____ ions:**
    (a) negative (b) positive
    (c) neutral (d) none of these

13. Electrolysis of NaCl is done in the cell:
(a) Electrolyte (b) Voltaic
(c) down's (d) for a day's

14. Day cell battery are used for: -
(a) Flesh light (b) Small appliances
(c) Toys (d) All of these

15. Complete this chemical equation Zn → ?
(a) $Zn \rightarrow Zn^{+2} + 2e^-$ (b) $Zn \rightarrow Zn^{+2} + 1e^-$
(c) $Zn \rightarrow Zn^{+2} + 2O^2$ (d) $Zn \rightarrow Zn^{+2} + 3e^-$

16. Large scale sodium metal is produced by the electrolysis of fused:-
(a) Cl (b) Fe
(c) Br (d) NaCl

17. Electrode through which electrons enter the external circuit:
(a) anode (b) cathode
(c) electrode (d) none of these

ELECTROPLATING

18. Electrode through which electrons leave the external circuit.
(a) anode (b) cathode
(c) graphite (d) electroyte

19. Zinc plates are made by
(a) Iron (b) Steel
(c) Both (a) and (b) (d) None of these

20. Food cans are generally:-
(a) Tin plated (b) Zinc Plated
(c) Iron Plated (d) None of these

21. Metal reacts with Oxygen and moisture in the atmosphere is called:-
(a) Corrosion (b) Prevention
(c) Chromium (d) None of these

22. In reduction reaction electrons are:
(a) lost (b) absorbed
(c) kept constant (d) all of these

23. In an oxidation reaction electrons are:
(a) absorbed (b) lost
(c) moved (d) increased

24. In an oxidation reaction electrons are:
(a) absorbed (b) lost (c) moved (d) increased

ANSWERS KEYS

1	a	6	b	11	b	16	d	21	a
2	c	7	b	12	b	17	c	22	b
3	a	8	b	13	c	18	a	23	b
4	b	9	a	14	d	19	b	24	b
5	a	10	a	15	a	20	a		

SHORT QUESTIONS

OXIDATION AND REDUCTION

Q.1 Define Oxidation & Reduction Terms of Loss OR Gain of Electrons.

Ans: **Oxidation & Reduction in Terms of Loss OR Gain of Electrons:-**

A process that involves the loss of electrons by an element is called oxidation.

$$Na \longrightarrow Na^{+} + e^{-}$$

$$Ca \longrightarrow Ca^{+2} + 2e^{-}$$

a. The process that involves the gain of electrons by a substance is called Reduction.

$$Cl + e^{-} \longrightarrow Cl^{-}$$

$$O + 2e^{-} \longrightarrow O^{-2}$$

Q.2 Define Reduction?

Ans: The process that involves the gain of electrons gain of hydrogen and removal of oxygen by a substance is called Reduction.

Q.3 Differentiate between oxidation and reduction

Ans:

Oxidation	Reduction
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electron
Increase in oxidation number	Decrease in oxidation number.

Q.4 Explain the term oxidation on the basic of electronic concept with an example.

Ans: According to electronic concept: The loss of electron by an atom or an ion is called oxidation. For example:

$$Mg \xrightarrow{\text{oxidation}} Mg^{+2} + 2e^{-}$$

Q.5 In a compound MX_3, find out the oxidation number of M and X.

Ans. **Compound = MX_3**

If we consider that the oxidation no. of M = +3

[Oxidation no. of M] + [oxidation no. of X] 3 = 0

(+3) + [Oxidation no. of "X"]3 = 0

[Oxidation no. of "X"] 3 = - 3

$$\text{Oxidation no. of "X"} = \frac{-3}{3}$$

Oxidation no. of X = -1

Q.6 Define oxidizing Agent?

Ans: Is the reactant containing the elements that is reduced (gain electrons) in a chemical reaction.

ELECTROCHEMICAL CELLS

Q.7 Define battery?

Ans: A battery is a galvanic cell or group of galvanic cells joined an series.

Q.8 Define voltaic cell?

Ans: The cell which involves spontaneous redox reaction to generate electricity is called a voltaic cell

Q.9 What are the 3 important products are produced by brine electrolysis?

Ans: (i) Cl gas (ii) NaOH, (iii) Hydrogen gas.

ELECTROLYTIC CELLS

Q.10 Which type of chemical reaction takes place in electrolyte cell?

Ans: In electrolytic cell, non spontaneous reaction takes place.

Q.11 What type of reaction takes place at anode in electrolyte cell?

Ans: Oxidation takes place at anode. Anode in positive charge electrode. The atoms on this electrode release electrons as:

$$2Cl^- \longrightarrow 2Cl + 2e^-$$

Q.12 Why the positively charged electrode is called anode in electrolytic cell?

Ans: The positive charged electrode is called anode increase it connected to the C+) terminal of the battery and all electron move towards it as

$$2Cl^- \longrightarrow 2Cl + 2e^-$$

VOLTAIC CELLS

Q.13 Define Galvanic cell?

Ans: An electrochemical cell that converts chemical energy into electricity is called Galvanic cell.

Q.14 Differentiate between Electrolyte and non-electrolyte light

Ans:

Electrolyte	Non-electrolyte
A substance that conduct electrically when it is dissolved in water or in the molten state is called electrolyte. **e.g:** NaCl, KCl, HCl, NaOH etc.	A substance that cannot conduct electricity when dissolved in water or in a molten state is called Non-electrolyte. **e.g:** Urea, sucrose, benzene etc.

Q.15 Towards which electrode of the electrolytic cell moves the catious and what does they do there?

Ans: Catious ions carry (+) charge, they move towards the cathode in an electrolytic cell. They gain electrons at cathode and oxidized.

Q.16 Where does the sodium metal is collected in down cell?

Ans: 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.

Q.17 Define Galvanic cells,

Ans: Galvanic cells:

The cells which involves spontaneous redox reaction to generates electricity is called a galvanic cell or voltaic cell.

Q.18 Define Battery?

Ans: Battery:-

A battery is a galvanic cell or a group of galvanic cells joined in series. It generates electric current by redox reaction.

Q.19 Discuss construction of galvanic cell?

Ans: Galvanic cell consist of following parts:

- A zinc bar dipped into a 1M $ZnSO_4$ solution
- A copper bar dipped into 1M $CuSO_4$ solution
- A salt bridge which is used containing an inert electrolyte ear as KCl, it ions do not react with electrodes.
- A voltmeter to measure current.

Q.20 What is the advantage of galvanizing?

Ans: A big advantage of galvanizing is that zinc protects the iron against corrosion even after a coating surface is broken.

ELECTROPLATING

Q.21 Define electroplating?

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

Q.22 Why is the anode made up of a metal to be deposited during electrolysis?

Ans: Because when current passed, the metal from anode dissolved in solution and metallic ions migrate to the cathode and discharge or deposit on the object. As the result of this charge, a thin layer of metal deposits on the object.

Q.23 State the best method for protection of metal from corrosion.

Ans. The best method for protection of metal from corrosion is the coating of highly resistant metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion.

Again

Q.24 What is the advantage of galvanizing?

Ans. A big advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

Q.25 Why is the anode made up of a metal to be deposited during electrolysis

Ans. Because, when the current passed, the metal from anode dissolved in the solution and metallic ions migrate to the cathode and discharge or deposit on the object. As a result of this charge, a thin layer of metal deposits on the object.

LEARNING OUTCOMES

Understanding:

Students will be able to:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen. (Remembering)
- Define oxidation and reduction in terms of loss or gain of electrons. (Understanding)
- Identify the oxidizing and reducing agents in a redox reaction. (Analyzing)
- Define oxidizing and reducing agent in a redox reaction. (Understanding)
- Define oxidation state. (Remembering)
- State the common rules for assigning oxidation numbers to free elements, ions, simple and complex molecules, atoms. (Remembering)
- Determine the oxidation number of an atom of any element in a compound. (Applying)
- Describe the nature of electrochemical processes. (Understanding)

LONG QUESTIONS

Q.No.1 : What is the importance of Electro Chemistry or Redox Reactions On Industrial Level?

Ans: The branch of chemistry which deals with the conversion of electrical energy into chemical energy and chemical energy into electrical energy is called ELECTROCHEMISTRY.

Electrochemistry plays a vital role on Industrial level for the production of Industrial products.

For Example

1) Electrochemical Industries produce millions of tons of Important metals such as Copper, Aluminium, Magnesium, Sodium and Zinc etc. through electrolysis processes.
2) They also produces Caustic Soda, Chlorine, Silicon carbide etc.
3) Important processes such as Rusting of Iron objects, combustion of fuel in automobile engine, forest fire and metabolism of food in human and animal bodies are involving Redox reactions.
4) The turning on of a flash light, mobile, calculator, electronic toys etc also generates out because of a current of electricity.
5) House hold bleaching agents decolourize the color bearing substances in strains involve Redox reactions.
6) Dying of Aluminium can produce metallic red, blue or other colours on the metal surface because of process Electroplating.
7) Anodized Aluminium sheets are widely used in buildings now-a-days for absorbing dyes because of Electrolysis process.

Hence due to the detail discussion, we conclude that Electrochemistry plays a vital role in Industries.

Q.No.2 : What is meant by Anodized Aluminium?

Ans: Process of anodic oxidation of metals (Aluminium) with the object of giving protection against corrosion called anodized Aluminium.

Q.No.3 : a) How many ways of expressing Oxidation and Reduction?

b) Justify the practical use of oxidation/ Reduction reactions on Industrial levels with reference of Loss or gain of oxygen and Hydrogen.

Ans:

a) Ways of Expressing Redox Reactions

There are three different ways of expressions of oxidation and Reduction reactions.

(i) Loss or gain of oxygen and hydrogen.
(ii) Loss or gain of electron (s).
(iii) Oxidation number decreases or Increases.

b) Practical use of Redox Reaction

Redox is stand for Reduction-Oxidation reactions simultaneously observed in same chemical reaction. These reactions plays an important role as follows:

1) In Steel mills, iron ores usually oxides of Iron are converted to the pure metal on Industrial level by the following reaction in the blast furnace.

$$C + O_2 \xrightarrow{1500C^\circ} CO_2 + \text{heat (Exothermic reation)}$$

$$CO_2 + C \xrightarrow{1100C^\circ} 2CO \text{ (Endothermic reaction)}$$

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

In this reaction, Fe_2O_3 loses oxygen which gains by CO and oxidizes into CO_2. Thus Fe_2O_3 is reduced into Fe and CO is oxidized into CO_2.

2) Acetylene (C_2H_2) is commercially used for cutting and welding metals. When acetylene burns, it produces a very hot flame known as oxy-acetylene flame and the following reaction takes place.

$$C_2H_2 + \frac{5}{2} O_2 \longrightarrow 2CO_2 + H_2O + \text{heat (Exothermic reaction)}$$

Or

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O + \text{heat}$$

In this reaction, Acetylene loses hydrogen and oxygen gains hydrogen to form water. Thus Acetylene gets oxidized (loses hydrogen) into CO_2 and oxygen is reduced into H_2O (gain Hydrogen)

3) Coal is burned in Thermal Power Stations to produce electricity. The following reaction occurs, when it burns $C + O_2 \longrightarrow CO_2 + \text{heat}$.

In this reaction, carbon atom gain oxygen and gets oxidized into CO_2.

4) The formation of a loose flaky layer of hydrated Iron (III) oxide, Fe_2O_3 on the surface of Iron is called Rusting. In the rusting of Iron, following reaction occurs

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

In this reaction, iron gains oxygen and gets oxidized into Fe_2O_3.

Example 7.1

Identifying the elements undergoing oxidation

Following reaction occurs when we burn Sui gas (Methane, CH_4)

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + \text{heat}$$

Identify the element undergoing oxidation.

Problem solving strategy

Identify the substance that gains O-atoms or loses H-atoms.

Solution

Since C in CH_4 which loses H-atoms and combines with oxygen atoms and form CO_2 thus C atoms undergo oxidation. At the same time O-atoms combine with H-atoms to form H_2O, thus O-atoms undergo reduction.

Q.No.4:a) Define Oxidation and Reduction in terms of loss or gain of oxygen and hydrogen.
b) Define Oxidation and Reduction in terms of loss or gain of electron (s).
c) Define oxidation Number or oxidation State and explain oxidation and Reduction in terms of decrease or Increase of Oxidation State.

Ans: a) Oxidation

Oxidation is defined as the gain of oxygen atoms

Example:

i) oxidation of C into CO_2 by combustion

$$C + O_2 \longrightarrow CO_2$$

(ii) Oxidation of Al into Al_2O_3

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

Similarly oxidation is also defined as the loss of hydrogen atoms by a molecule.

Example:

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O + \text{heat}$$

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + \text{heat}$$

Since C_2H_2 and CH_4 loss H-atoms and combined with oxygen atoms, thus c-atoms undergo oxidation and form H_2O.

Reduction

Reduction is defined as the gain of Hydrogen atoms.

Example:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

In these reactions, O_2 and N_2 molecules gain hydrogen and forms H_2O and NH_3 respectively similarly, reduction is also defined as the loss of O-atoms.

Example:

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

$$Zn + CuO \longrightarrow ZnO + Cu$$

In these examples, Fe_2O_3 and CuO loses oxygen and forms Fe and Cu respectively.

b) Oxidation

A process that involves the loss of electrons by an element is called oxidation.

For Example

Group l-A Alkali Metals (Li, Na, K) and Groups ll-A Alkaline earth metals (Be, Mg, Ca, Sr, Ba) loses one or two electrons respectively to form positive charges called cations i,e M^{+1} and M^{+2}.

In doing so, these metals undergo oxidation

Examples:

$$Li \longrightarrow Li^{+1} + e^- \qquad Mg \longrightarrow Mg^{+2} + 2e^-$$

$$Na \longrightarrow Na^{+1} + e^- \qquad Ca \longrightarrow Ca^{+2} + 2e^-$$

$$K \longrightarrow K^{+1} + e^- \qquad Sr \longrightarrow Sr^{+2} + 2e^-$$

Other examples of oxidation are:

$$Fe \longrightarrow Fe^{+2} + 2e^-$$

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

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

Reduction

A process that involves the gain of electron (s) by a substance is called Reduction.

For Example

Group VA (N,P),VIA (O,S) and VII A (F, C*l*) gain 3, 2, 1 electrons respectively to form Negative lons called Anions i,e N^{-3}, O^{-2}, F^{-1} respectively. They undergo Reduction.

Examples

$$N + 3e^- \longrightarrow N^{-3}\,(\text{Nitride}) \qquad P + 3e^- \longrightarrow P^{-3}\,(\text{Phosphide})$$

$$O + 2e^- \longrightarrow O^{-2}\,(\text{Oxides}) \qquad S + 2e^- \longrightarrow S^{-2}\,(\text{Sulphides})$$

$$F + e^- \longrightarrow F^{-1}\,(\text{Flourides}) \qquad Cl + e^- \longrightarrow Cl^{-1}\,(\text{Chlorides})$$

c) **Oxidation Number or Oxidation State**

It is defined as the number of charges on an atoms, molecule or compound is called Oxidation state.

For Example

i) In Na^{+1}, the oxidations Number of Na is +1 while "1" is a valency.

ii) In O^{-2}, the oxidation number of O is -2 while "2" is a valency.

iii) In Cl atom, the oxidation Number is Zero because it is a neutral atom.

Oxidation

The element that shows an increase in oxidation number are oxidized and the process will be called oxidation.

Example

$Na^{o} \longrightarrow Na^{+1} + e^{-}$

$Fe \longrightarrow Fe^{+2} + 2e^{-}$

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

Reduction

The element that show a decrease in Oxidation number are reduced and the process will be called Reduction.

For Example

$Cl^{o} + e^{-} \longrightarrow Cl^{-1}$

$O^{o} + 2e^{-} \longrightarrow O^{-2}$

$H + \bar{e} \longrightarrow H^{-1}$

Similarly, $H_2 + Cl_2 \longrightarrow 2HCl$

In this reaction H in HCl is oxidized from H_2^{o} to $H^{+1}Cl^{-1}$ which Cl in H Cl is reduced from Cl_2^{o} to $H^{+1}Cl^{-1}$

Process leading to oxidation and reduction

Oxidation	Reduction
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
Increase in oxidation number/state	Decrease in oxidation number/state

Q.No.5: Justify that In Photography, oxidation-Reduction reaction is involved.

Ans: **Redox in photography**

A photographic film is basically an emulsion of silver bromide, (AgBr) in gelatin. When the film is exposed to light, Silver bromide granules become activated. This activation depends on the intensity of the light falling upon them. When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is a mild reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver. Reduced silver atoms form image.

$$Ag^{+} + 1e^{-} \longrightarrow Ag^{o}_{(s)} \quad -(\text{Reduction})$$

Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to the light appear darkest because they have the highest concentration of metallic Silver. Thus photography involves oxidation- reduction reaction.

Example 7.2

Identifying the element oxidized or reduced

In the following reaction identify which element is oxidized and which element is being reduced.

$$2Ca + O_2 \longrightarrow 2CaO$$

Problem solving strategy

Ca being metal will form cation by losing electrons (oxidation) and oxygen being non-metal will gain electrons (reduction) to form anions.

Solution

We have learned that Group IIA metals form M^{+2} cations, and that Group VIA non-metals form X^{-2} anions. This means in this reaction each Ca atom loses two electrons to form Ca^{+2}, so it is oxidized. Each oxygen atom gains two electrons to form O^{-2}, so it is reduced.

$$Ca \longrightarrow Ca^{+2} + 2e^{-} \text{ (oxidation)}$$

$$O_2 + 4\bar{e} \longrightarrow 20^{-2} \text{ (reduction) or } O + 2e^{-} \longrightarrow O^{-2}$$

Q.No.6: Describe the General Rules For Assigning Oxidation States or Number?

Ans: Rules for Assigning Oxidation states or Numbers

1) The oxidation state of any uncombined or free elements is always zero e.g., oxidation state of Zn, Na, H in H_2, S in S_8 etc is zero.
2) In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^{+1} and Ca in Ca^{-2} are +1 and +2 respectively.
3) In a complex ion the total sum of oxidation states of atoms is equal to the charge on their ion. e.g., in CO_3^{-2} the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^{+1}, the sum of oxidation states of N and 4H atoms is +1.
4) The oxidation number of each of the atoms in a molecule or compound counts separately and their algebraic sum is zero e.g., In HCl, the sum of oxidation states of H and Cl atoms is zero. Similarly in CO_2, the sum of oxidation states of one C and 2 oxygen atoms is zero.

Table shows the oxidation states of some of the elements in binary compounds which rarely change.

Table: Oxidation states of some elements in binary compounds that rarely change.

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2 (except peroxides^{-1}) and in OF_2 it is +2

Example 7.3

Determining oxidation number In a Compound:

A device called Breath Alyzer is used by police to test a person's breath for alcohol. It contains an acidic solution of potassium dichromate $K_2Cr_2O_7$. It is a strong oxidizing agent. Determine oxidation state of Cr in it.

Problem Solving Strategy

Use rules 1 to 4 and table above to get as many oxidation numbers as you can. Use rule 4 to get oxidation number that has not been assigned.

Solution

1) The oxidation number of K is +1, since it belongs to Group-IA. There are 2 K atoms therefore, overall oxidation number for k is 2(+1) = +2
2) There are 7 oxygen atoms, therefore overall oxidation state for O is 7(-2) = -14
3) Suppose oxidation for Cr is x, since there two Cr atoms, therefore, overall oxidation state for Cr is 2x.
4) The sum of oxidation numbers must be zero.

$$+2 + 2x + (-14) = 0$$
$$2x - 12 = 0$$
$$2x = 12$$
$$x = +6$$
$$x = +3$$

Thus oxidation state for Cr in $K_2Cr_2O_7$ is +6

Example 7.4

Determining oxidation state

Boric acid H_3BO_3 is used in eye wash. What is the oxidation state of B in this acid?

Problem solving strategy

Use rules and table 7.2 to get the oxidation state of H and O-atoms. Use rule 4 to get the oxidation state of B.

Solution

1) There are 3 H-atoms, therefore, overall oxidation state for H is 3(+1) = +3
2) There are 3 O-atoms, therefore, overall oxidation state for O is 3(-2) = -6
3) Suppose the oxidation state for B is x.
4) The total oxidation states for all the atoms must be zero.

$$+3 + x + (-6) = 0$$
$$+3 + x - 6 = 0$$
$$x - 3 = 0$$
$$x = +3$$

Thus the oxidation state for B in H_3BO_3 is +3.

Q.No.7 Define Oxidizing and Reducing Agents with Examples.

Ans: Oxidizing Agent

An oxidizing agent is the reactant containing the element that is reduced (gains electrons) in a chemical reaction.

Reducing Agent

A Reducing agent is the reactant containing the element that is oxidized (loses electrons) in a chemical reaction.

For Example

Sodium and chlorine reacts to form Sodium chloride.

$$2\overset{0}{Na} + Cl^{0}_{2} \longrightarrow 2Na^{+1}\overset{-1}{Cl}$$

Na is reducing agent because it is being oxidized as its oxidation no increases from

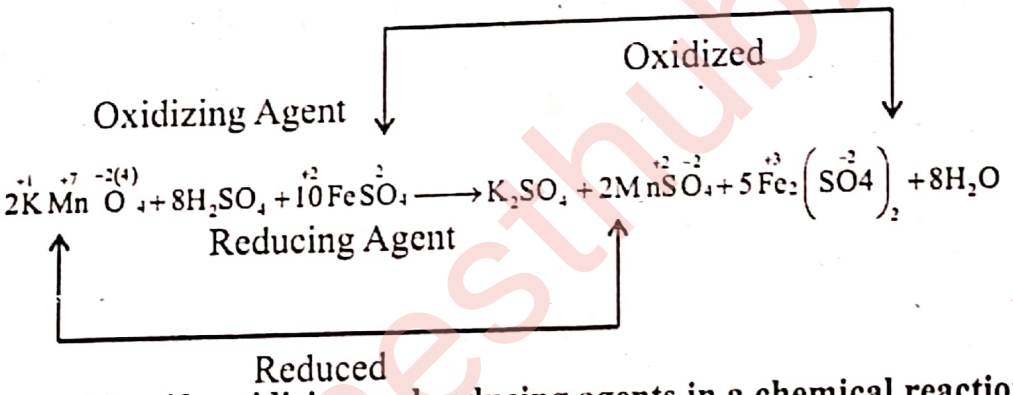
$$2\overset{0}{Na} \longrightarrow \overset{+1}{Na}Cl$$ while cl_2 is oxidizing agent because it is being reduced as its oxidation No decreases from $Cl_2^{0} \longrightarrow NaCl^{-1}$

Identifying Oxidizing and Reducing Agents

Activity

Prepare solutions of ferrous sluphate $(FeSO_4)$ and potassium permanganate $(KMnO_4)$ in separate beakers. Transfer about $10cm^3$ of ferrous sulphate solution in a test tube. Add about $10cm^3$ of dill. H_2SO_4 in it. Now add few drops of $KMnO_4$ solution in the test tube. What happens?

$FeSO_4$ reduces $KMnO_4$. So its purple colour is discharged. $KMnO_4$ Oxidizes $FeSO_4$ in this reaction. $FeSO_4$ is reducing agent whereas $KMnO_4$ is oxidizing agent.



How can you identify oxidizing and reducing agents in a chemical reaction?

Consider the following reaction that takes place in the manufacture of steel.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

To identify the oxidizing and reducing agents, work out the oxidation states of all the elements involved in the reaction.

$$\overset{2(+3)}{Fe}_2\overset{3(-2)}{O}_3 + 3\overset{+2}{C}\overset{-2}{O} \longrightarrow 2\overset{0}{Fe} + 3\overset{+4}{C}\overset{2(-2)}{O}_2$$

i) Carbon is being oxidized because there is an increase in its oxidation state from +2 to +4.

ii) Fe is being reduced because there is a decrease in its oxidation state from +3 to zero.

iii) The reactant Fe_2O_3 contains the Fe that is being reduced. So Fe_2O_3 is oxidizing agent. Oxidizing or reducing agent is the whole molecule or formula unit and not the atom that has undergone change in oxidation number.

Example 7.5

Identifying the oxidizing and reducing agents

Tungsten is used to make filaments for electric bulbs because it has the highest melting point and high electrical resistance. This metal is obtained from tungsten (VI) oxide, WO_3 by reducing it with hydrogen gas.

$$WO_3 + 3H_2 \longrightarrow W + 3H_2O$$

Identifying the oxidizing and reducing agents in this reaction.

Problem solving strategy

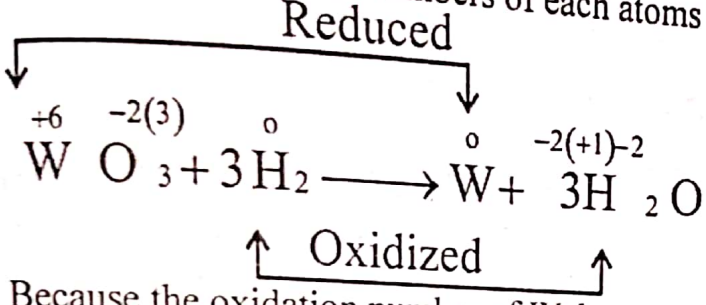
Step 1: Workout the oxidation states of all the elements involved in the reaction.

Step 2: Note the element that is undergoing an increase in its oxidation state. Since it is being oxidized. The reactant that contains this element is reducing agent.

Step 3: Note the element that is undergoing a decreases in its oxidation state. Since it is being reduced. The reactant that contains this element is oxidizing agent.

Solution

First assign oxidation Numbers of each atoms



Because the oxidation number of W decreases, So WO_3 is an oxidizing agent. Similarly the oxidation number of H increases, therefore, H_2 is reducing agent.

Q.No.8: a) What are Redox Reactions? How they are produced?

b) Describe the method of recovering metal from its Ores.

Ans:

a) Redox Reactions

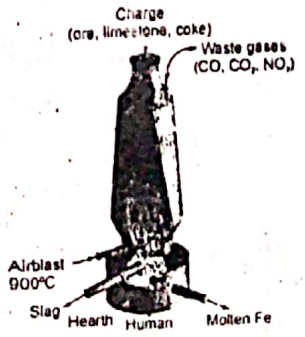
Cations are formed from metallic elements after removal of electron (s) and Anions are formed from non- metallic elements after gaining of electron (s) as a result of combining of cations and Anions, Ionic compound are formed e.g: Calcium (metal) and Chlorine (non-metal) contain neutral atoms. After transference of elements from calcium atoms to chlorine atoms and form calcium chloride involved oxidation reduction reactions simultaneous called REDOX Reaction (Red for Reduction and Ox for Oxidation)

$$\overset{o}{Ca} + \overset{o}{Cl_2} \longrightarrow \overset{+2}{Ca}\overset{(-1)2}{Cl_2}$$

$$\overset{o}{Ca} \longrightarrow \overset{+2}{Ca} \text{ oxidation}$$

$$\overset{o}{Cl_2} \longrightarrow 2\overset{-1}{Cl} \text{ Reduction}$$

b) Method of Recovering Metals from their Ores



A blast furnace

Redox Reactions are commercially very important. Most of the metals are recovered from their ores by Redox Reactions because most of the metals are found in nature as oxides or Sulphide ores. After mining the ore, desired mineral is separated from the other materials. Purified Metal Oxides are reduced to free metals by using Reducing Agent. Aluminium, Coke, Carbon, Monoxide gas and Hydrogen gas are generally used as Reducing Agents.

Iron can be extracted from their ore Haematite (Fe_2O_3) by reducing agent in a Blast Furnnace as shown in figure.

Iron ore (Fe_2O_3), Lime stone ($CaCO_3$) and Coke are introduced into the Blast Furnace from the top. A blast of hot air is forced up the furnace from the bottom so called Blast Furnace. The oxygen gas reacts with Coke to form mostly carbon monoxide and some

a

carbon dioxide. These reactions are highly exothermic. As hot Co rises, it reacts with Iron oxide and reduces it to iron. Molten Iron collects at the bottom of the furnace. Limestone removes impurities i,e. SiO_2, AlO_3 Iron as slag. The reactions are given below:

$$C + O_2 \longrightarrow CO_2 + \text{heat}$$

$$CO_2 + C \longrightarrow 2CO + \text{heat}$$

$$3CO + Fe_2O_3 \longrightarrow 2Fe + 3CO_2$$

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

$$\left.\begin{array}{l} CaO + SiO_2 \longrightarrow CaSiO_3 \\ CaO + Al_2O_3 \longrightarrow Ca(AlO_2)_2 \end{array}\right] \text{Slag}$$

Similarly, Lead and Zinc metals are extracted by their ores. These ores are first converted into corresponding oxide by heating in oxygen. In this process, Sulphur is oxidized and oxygen is reduced. These oxides are then reduced by Coke, Co or H_2 as reducing agents as shows in the following reactions.

$$\overset{+2\ -2}{PbO} + \underset{\text{R Agent}}{\overset{+2\ -2}{CO}} \longrightarrow \overset{0}{Pb} + \overset{+4\ -2(2)}{CO_2}$$

$$\overset{+2\ 2}{ZnO} + \overset{0}{C} \longrightarrow \overset{0}{Zn} + \overset{+2\ -2}{CO}$$

$$\underset{\text{O Agent}}{\overset{+2\ \ -2}{CuO}} + \underset{\text{R Agent}}{\overset{0}{H_2}} \longrightarrow \overset{0}{Cu} + \overset{+1(2)-2}{H_2O}$$

Q.No.9: a) Define Electro Chemical Cells.

b) Differentiate between Electrolytic and Galvanic Cells.

c) Why Electrochemical Process Involves Redox Reaction?

Ans:

a) Electrochemical Cells

An arrangement which consists of electrodes dipped into an electrolyte in which chemical reaction uses or generates electric current is called Electrochemical Cell. e.g. Electrolytic Cells (Down's and Nelson's cell) Galvanic cells (Daniel's and Ni-Cd cell)

b) Electrolytic Cell

1) The cell in which a reaction occurs with the help of Electric current is called Electrolytic cell.

e.g.: Down's Cell

$$2NaCl \xrightarrow{EC} 2Na_{(s)} + Cl_{2(g)}$$

Nelson's Cell

$$2NaC_{(l)(eq)} + 2H_2O \longrightarrow 2N^+a + 2OH^- + H_2 + Cl_2$$

2) Non-Spontaneous oxidation-reduction reactions takes place

Galvanic or Voltaic Cell

1) The cell in which a reaction generates electric current is called voltaic or Galvanic cells. e.g., Daniel's cell, Lead storage battery, fuel cell etc.

2) Spontaneous oxidation-reduction reaction takes place.

c) Nature of Electro Chemical process

Electrochemical processes are oxidation-reeducation reactions in which chemical energy released by a spontaneous reaction is converted to electricity or in which electrical energy is used to drive a non-spontaneous reaction. Whether an electrochemical process releases or requires energy, it always involves the transfer of electrons from one substance to another. This means that this process always involves an oxidation-reduction or a redox reaction.

Q.No.10 : a) What are Electrolytes and non-Electrolytes?
b) Differentiate between weak and strong Electrolytes.
c) Differentiate between spontaneous and Non-Spontaneous Process.

Ans:

a) Electrolytes

The substances which undergoes partial or complete dissociation into ions in solution or molten form passing electricity are called Electrolytes.
e.g.: NaCl solution, KCl solution, HCl, NaOH etc.

Non-Electrolytes

A substances that cannot conduct electricity when dissolved in water or in the molten state is called Non-electrolytes.
e.g.: urea, glucose, Sucrose, benzene etc.

b) Weak Electrolyte

The substances which gives poorly conducting solutions when dissolved in water are called weak electrolytes. e.g: Ammonia, Acetic Acid, carbonic Acid and organic bases.

Strong Electrolyte

The electrolytes which give highly conducting solutions when dissolved in water are called Strong Electrolytes e.g: HCl, H_2SO_4,NaCl etc.

c) Spontaneous Process

A physical or chemical change that occurs by itself is called a spontaneous process. Spontaneous process do not require a source of energy to make them happen. For instance water flows from higher level to lower level. Iron placed in moist air, rusts. The flow of electrons through a conductor can be obtained from a spontaneous oxidation-reduction reaction. This is the basis for how batteries work.

Non-spontaneous Process

On the other hand a physical or chemical change that requires a source of energy to make them happen is called non-spontaneous process. For example water can be made to flow from lower level to higher level by using a pump.

Activity 7.1

Passing electricity through aqueous solution

You will need

- Beaker
- Graphite rod
- A battery and Conducting wires
- Aqueous solutions of NaCl, NaOH, urea and glucose

Carry out the following

1) Set up the circuit as sown in the figure 7.4.
2) Fill beaker with one of the given solutions.
3) Dip graphite rods in the solution.
4) Note if the bulb is lighted or not.

Do You Know

An electrode is a conductor in a circuit that carries electrons to or form a substance other that a metal.

e.g. Graphite rods caressed as electrodes.

5) Repeat the experiment with other solution one by one.
6) Identify electrolytes.

Figure shows an easy method of distinguishing between electrolytes and non-electrolytes. When graphite electrolytes are dipped in pure water, no current flows. But if you add a small quantity of solution of KCl, the bulb glows. KCl is an ionic compound. It breaks up into K^+ and Cl^{-1} ions, as its dissolves in water. These ions move towards the opposite electrodes. Thus current flows through the solution and the following reactions will be observed.

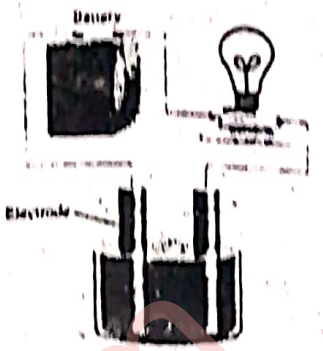


Figure: Assembly for distinguishing electrolytes and non-electrolytes

$$KCl_{(aq)} \xrightarrow{H_2O} K^+_{(aq)} + Cl^-_{(aq)}$$

At Anode (Oxidation Occurs

$$2Cl^- \longrightarrow Cl_2 + 2e^-$$

At Cathode (Reduction Occur)

$$2K^+ + 2e^- \longrightarrow 2K$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

Q.No.11 : Sketch a electrolytic Cell, Labelling the Cathode, Anode and the direction of flow of the electrons.

Ans: An electrochemical Cell in which electrical energy is used to drive a chemical reaction is called an Electrolytic Cell. Figure shows the sketch of an Electrolytic Cell.

An electrolytic cell consists of

a) A vessel containing an electrolyte (MX).
b) Two inert electrodes
c) A battery

The figure shows that electrons move from anode to cathode in the outer circuit, in the solution the cations move towards cathode and anions towards anode. At anode anions oxidize by loosing electrons. At cathode cations reduce by gaining electrons. This means oxidation occurs at anode and reduction at cathode.

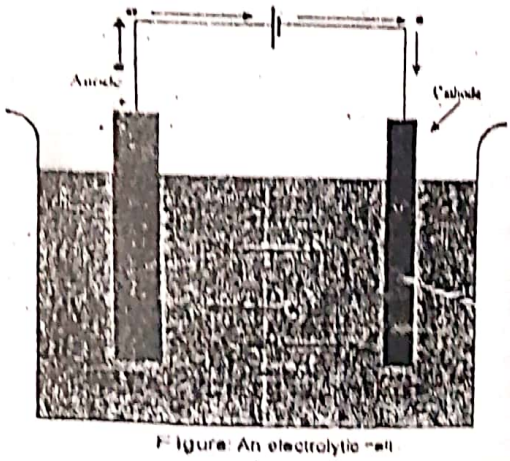


Figure: An electrolytic cell

$$MX_{(aq)} \xrightarrow{H_2O} M^+_{(aq)} + \overline{X}_{(aq)}$$

At Anode $X^- \longrightarrow X + e^-$ (oxidation occurs)

At Cathode $M^+ + e^- \longrightarrow M$ (reduction occurs)

Q.No.12 : List the possible uses of Electrolytic Cells

Ans:

Uses of Electrolytic Cells

Possible uses of electrolytic cells are as follows:

a) Down's Cell is used for the commercial preparation of sodium metal. It produces chlorine gas as by product.
b) Nelson's Cell is used for the commercial preparation of sodium hydroxide. It also produces chlorine and hydrogen gas as by product.

c) Electrolysis cells are used for the commercial preparation of calcium and magnesium metals.
d) It is used to produce aluminium metal commercially.
e) It is used for the purification of copper.
f) Electrolytic cells are used to electroplate metals such as tin, silver, nickel etc on steel.
g) Electrolytic cells are used to prepare anodized aluminium. Anodized aluminium can absorb dyes. Dyeing of anodized aluminium can produce metallic red, metallic blue or other metallic colours on the metal surface.

Q.No.13 : a) Define Salt bridge.
b) What is the function of Salt bridge?
c) Describe the Sketch of Galvanic (Daniel) Cells.

Ans:

a) **Salt Bridge**
It is a U-shaped glass tube having a Saturated Solution of some strong electrolyte like KCl, K_2SO_4 or KNO_3. The glass tube is sealed at both of its ends by a porous material like glass wool or cotton play. It prevents the physical contact between the two electrolytic solutions.

b) **Function of Salt Bridge**
It has two major functions:
1) It connects the solutions in two half cells and completes the cell circuit.
2) It maintains the electrical neutrality by the diffusion of ions through it.

c) **Galvanic Cells (Daniel Cells)**
The cell which involves spontaneous redox reaction to generates electricity is called a galvanic or voltaic cell. The name Voltaic is given to this cell because Alessandro Volta discovered first such cell. The English chemist Fredrick Daniel Constructed first Voltaic cell using zinc (Zn) and copper (Cu) electrodes. Therefore this cells is named as Daniel Cell. A galvanic or Daniel cell is shown in figure.

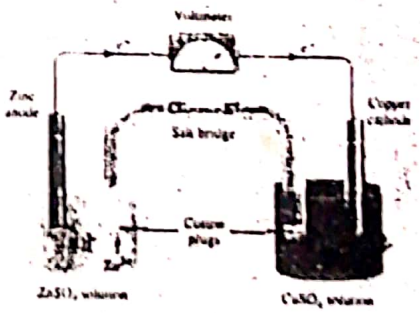


Figure: Galvanic cell

A galvanic cell consists of the following parts:
1) A zinc bar dipped into a 1 M $ZnSo_4$ solution.
2) A copper bar dipped into a 1 M $CuSO_4$ solution.
3) A salt bridge which is inverted U tube containing an inert electrolyte such as KCl. Its ions do not react with the electrodes or with the ions in solution. It makes the electrical contacts between the solutions through which ions can move.
4) A voltmeter to measure current.

Each compartment of cell is called a half cell. Thus a Daniel cell consists of two half cell joined in series. When circuit is complete electrons begin to flow from Zn rod through the external wires to Cu rod. Thus Zn half cell acts as anode and Cu half cell as cathode. Note that a half cell consists of a metal rod dipped in the solution of its salt.

Q.No.14 : Describe the Reactions In a Daniel Cell.

Ans: Reactions in a Daniel Cell
In Daniel cell, electrons flow from Zn rod, through the external wire to Cu rod. This is because Zn has more tendency to undergo oxidation than Cu. Zn atoms from the rod go into the solution as Zn^{+2} ions leaving electrons on the rod. These electrons flow in the

external circuit. Thus oxidation half reaction occurs at anode compartment. Cu^{+2} ions in copper sulphate solution capture electrons from Cu electrode and are reduced. Reduction half cell reduction occurs at the cathode compartment. Such oxidation and reduction reactions are called half cell reactions.

At Anode (Oxidation half reaction):

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

At Cathode (Reduction half reaction):

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

Q.No.15 : a) Define Cell and Battery.

b) Describe how a battery produces Electrical Energy in a Dry cell.

Ans:

a) Cell and Battery

A cell is combination of two metal plates, one acts as cathode (negative terminal) and other act as Anode (positive terminal). It is represented as $+ \;||^-$ and its potential is 2 volt. Conventionally, Anode line is bigger than cathode.

A Battery is a Galvanic cell groups joined in series of a circuit. It generates electric current by a Redox Reaction. It is represented as $^+||\,|\,|\,|\,||^-$

Examples of batteries are dry cell, storage cell, mercury battery, Ni Cd battery etc...

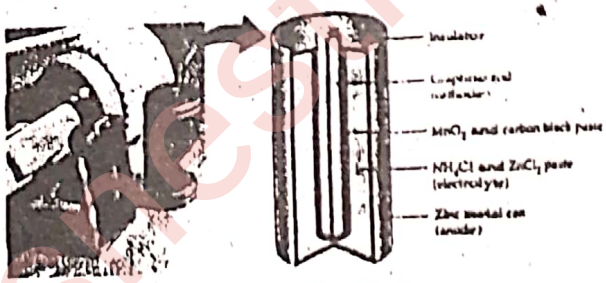


Figure: A dry Cell

b) Dry Cell

The dry cell batteries are used to power many flashlights, toys and small appliances. The anode is the zinc metal of the container and the cathode is an inert graphite rod at the center of the container in contact with a mixture of MnO_2 and carbon (charcoal) see fig below. The electrolyte is a mixture of moist NH_4Cl and $ZnCl_2$. Following reactions take place in it.

At Anode (oxidation takes place which means loss of electrons)

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

At Cathode (Reduction takes place which means gain of electrons)

$$2NH^+_4 + 2MnO_2 + 2e^- \longrightarrow Mn_2O_3 + 2NH_3 + H_2O$$

This cell produces a potential of 1.5V

Q.No.16 : a) What are the functions of Electrochemical Industries
b) Describe the Manufacture of Sodium Metal from fused sodium chloride by Down's Cell.
c) Describe the Manufacture of Sodium Hydroxide from Brine Solution by Nelson's Cell.

Ans:

a) Functions of Electrochemical Industries

Electrical energy is extensively consumed by the chemical process industries. Electrochemical industries use electricity to bring chemical change and produce wide variety of substances. Such industries are called electrolytic industries. These industries produce substances that cannot be made economically in another way. For example, aluminium and calcium carbide, sodium hydroxide, magnesium, and hydrogen. These industries also include electrolytic refining and electroplating industries. Electricity is also used to produce heat that produces high temperature required in the electrothermal chemical industries.

We will discuss the electrochemical preparation of the following.

1) Manufacture of sodium metal from fused sodium chloride.
2) Manufacture of NaOH from brine.
3) Electrolytic refining of copper.
4) Electroplating of Zinc, tin and Chromium on steel.

b) Manufacture of Sodium Metal from Fused Sodium Chloride

On the large scale sodium metal is produced by the electrolysis of fused NaCl. An electrolytic cell called Down's Cell is used for this purpose as shown below. The electrodes are iron cathode and graphite anode. Chlorine is obtained as by-product. In molten sodium chloride (NaCl), Na^+ ions are free to move about. Under the influence of electric current, Na^+ ions move towards the cathode and Cl^- ions towards the anode. At the electrodes following reactions occur.

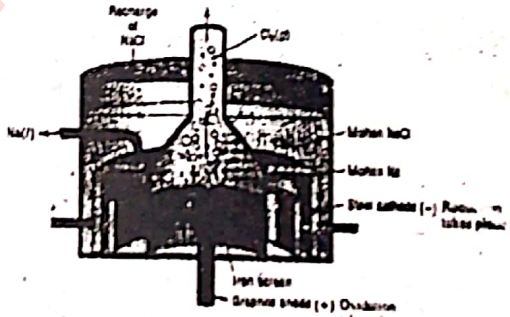


Figure Down's Cell

At anode $2Cl^- \longrightarrow Cl_{2\,(g)} + 2e^-$ (oxidation)

At cathode $2Na^+ + 2e^- \longrightarrow 2Na$ (reduction)

Molten sodium is collected into a sodium collecting ring, from where it is periodically drained. Whereas, chlorine gas in collected into the funnel at the top of the cell. Other alkali metals can also be obtained by the electrolysis of their fussed salts. For this purpose ores are first purified than electrolyzed.

c) Manufacture of Sodium Hydroxide from Brine

Electrolysis of brine, a concentrated aqueous solution of sodium chloride is used for the industrial production of sodium hydroxide. Electrolysis of brine produces simultaneously three important industrial chemicals, chlorine gas, hydrogen gas and sodium hydroxide. The electrolytic cell called Nelson's Cell as shown in figure. The solution contains four types of Ions i.e, Na^+, Cl^-, H^+, OH^- ions.

During electrolysis chloride ions move towards anode. At anode chloride ions are oxidized to produce chlorine gas. Sodium ions move towards cathode. However sodium ions do not reduce to sodium metal in this process. This is because water molecules are more easily reduced than sodium ions. The reduction of water molecules produces hydroxide ions and hydrogen gas. Thus the electrolyte in solution becomes sodium hydroxide (NaOH). Following reactions occur in the electrolytic cell.

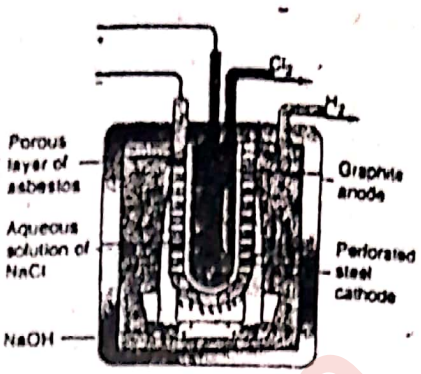


Figure Nelson's Cell

At anode (oxidation) $2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$

At cathode (Reduction) $2H_2O_{(l)} + 2e^- \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$

$2Na^+ + 2e^- \longrightarrow 2Na$

Overall Reaction

$$2Na_{(aq)} + 2Cl_{(aq)} + 2H_2O_{(l)} \longrightarrow Cl_2(g) + H_{2(g)} + 2OH^-_{(aq)} + 2Na^+_{(aq)}$$

$$2Na^+ + 2OH^- \longrightarrow 2NaOH$$

The solution contains Na^+ and OH^- ions. Evaporation of water from this solution produces relatively pure solid sodium Hydroxide. H_2 and Cl_2 are obtained as useful by products.

Q.No.17 : Explain Electrolytic Refining of Copper.

Ans:

Electrolytic Refining of Copper.

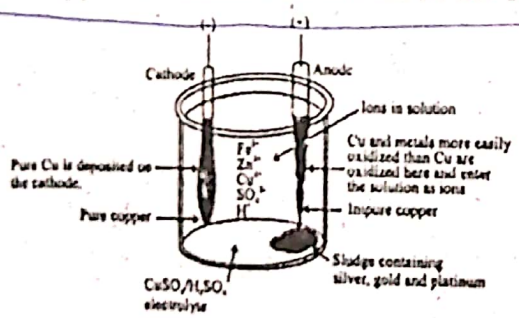
Purification of copper

Electrolytic cell can be used to purify copper. For this purpose, an impure copper is made the anode and a thin sheet of pure copper is made a cathode. The electrolyte used is copper sulphate ($CuSO_4$) solution.

The impurities are mostly silver (Ag), gold (Au), platinum (Pt), iron (Fe) and zinc (Zn) and they settle down as anode mud. Presence of these metals decreases the electrical conductivity of copper (Cu).

Cathode is a very thin sheet of very pure copper. When the cell is operated at the given value, the less active metals simply fall of the electrode and settle to the bottom of container. At cathode, Cu^{+2} is reduced but Zn^{+2} and Fe^{+2} remain in the solution.

In this way, 99.5% pure copper is obtained as shown in the following figure.



Following reaction occurs in this process.

At Anode, oxidation occurs which means loss of electrons

$$Cu_{(s)} \longrightarrow Cu^{+2}_{(aq)} + 2e^-$$

At Cathode, Reduction occur which means gain of electrons.

$$Cu^{+2}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

Q.No.18 : a) Define Electroplating. What is its purpose?
b) What are the conditions for Good Electroplating?
c) Discuss the Electroplating of silver on spoon and Electroplating on steel.

Ans:

a) Electroplating

It is the art of depositing one metal over the surface of another metal with the help of Electric Current.

Purpose of Electroplating

The electroplating is done for two purposes.

1) Preservation from rusting and enhancing the life of metals.
2) Decoration

b) Conditions for Good Electroplating

(i) High current density
(ii) Low temperature
(iii) High concentrate ion of metal in its electrolyte.

c) Electroplating of Silver on Teaspoon

In electroplating of silver, the electrolyte used in the solution of potassium argento cyanide. $K[Ag(CN_2]$. Anode is made up of pure silver. The article to be plated is made cathode. The ions which are formed in the electrolytic solution are:

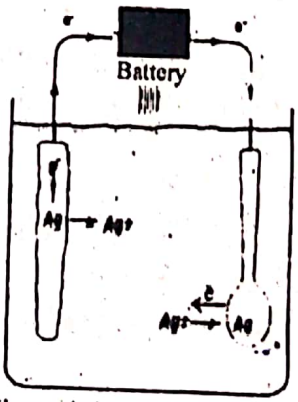
$$K\left[Ag(CN)_2\right]_{(aq)} \rightleftharpoons K^{\oplus}_{(aq)} + \left[Ag(CN)_2\right]^{\ominus}_{(aq)}$$

$$\left[Ag(CN)_2\right]^{\ominus}_{(aq)} \rightleftharpoons Ag^{\oplus}_{(aq)} + 2CN^{\ominus}_{(aq)}$$

When the electric current is passed, silver ions $(Ag)^{\oplus}$ migrate towards the cathode and deposit after picking up electrons. Cyanide ions $(CN)^{\ominus}$ move towards the anode, where they react with silver to form AgCN. It then combines with KCN to reproduce the electrolyte, $K[Ag(CN)_2]$

$$Ag^{\oplus}_{(aq)} + 1e^{\ominus} \longrightarrow Ag^{o}_{(s)} \text{ (deposited at cathode)}$$

Thus silver is deposited on the object (say spoon or an ornament) during the process of electroplating of silver.



Electroplating of silver on spoon

Electroplating on Steel

Steel objects are often protected from corrosion by electroplating with zinc, tin and chromium.

These electroplating are discussed below on steel.

(i) Zinc Plating

Zinc plating on steel is done by using zinc metal as anode. A solution of potassium zinc cyanide $K_2[Zn(Cn)_4]$ containing little sodium cyanide. The steel object is made cathode. During the electrolysis zinc at the anode enters the solution as Zn^{+2} ions, which are deposited at the cathode.

The electrolyte ionizes as follows.

$$K_2[Zn(CN)_4]_{(aq)} \rightleftharpoons 2K^{+1}_{(aq)} + [Zn(CN)_4]^{-2}_{(aq)}$$

$$K_2[Zn(CN)_4]^{-2}_{(aq)} \rightleftharpoons Zn^{+2}_{(aq)} + 4CN^{-1}_{(aq)}$$

Following reactions occur at the electrodes:

At anode: (Oxidation) $Zn_{(S)} \longrightarrow Zn^{+2}_{(aq)} + 2e^-$

At cathode: (Reduction) $Zn^{+2}_{(aq)} + 2e^- \longrightarrow Zn_{(s)}$

(ii) Tin Plating

Food cans are generally tin plated. Tin plating on steel is done by using tin as anode and a solution of stannous sulphate, ($SnSO_4$) as electrolyte. Few drops of dil H_2SO_4 are added in the electrolyte to prevent its hydrolysis. The electrolyte ionizes as follows.

$$SnSO_{4(aq)} \rightleftharpoons Sn^{+2}_{(aq)} + SO4^{-2}_{(aq)}$$

During the electrolysis following reactions occur.

At anode: (oxidation) $Sn_{(S)} \longrightarrow Sn^{+2}_{(aq)} + 2e^-$

At cathode: (Reduction) $Sn^{+2}_{(aq)+} 2e^- \longrightarrow Sn_{(S)}$

(iii) Chromium Plating

Chromium metal does not adhere strongly to the steel therefore steel is first plated with copper or nickel and then chromium. For electroplating chromium, chromium metal is used as anode and chromium sulphate, $Cr_2(SO_4)_3$ as an electrolyte. A few drops of dil H_2SO_4 are added in the electrolyte to prevent its hydrolysis. The electrolyte ionizes as follows:

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

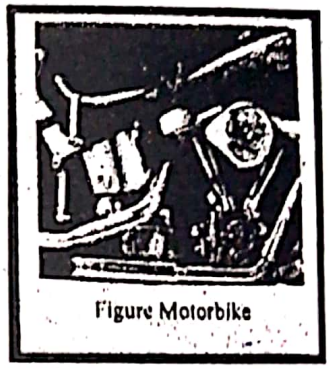


Figure Motorbike

The following electrolytic reactions will occur at the electrodes.

At anode: (oxidation) $Cr_{(S)} \rightleftharpoons Cr^{+3}_{(aq)} + 3e^-$

At cathode (Reduction) $Cr^{+3}_{(aq)} + 3e- \rightleftharpoons Cr_{(S)}$

Chromium plated steel is used to make automobile parts as shown in the figure above.

Q.No.19 a) Why cyanide Ions are toxic in nature?

b) Why we prefer tin plating over iron plating on industrial level?

Ans:

a) Cyanide ions are extremely toxic because sodium cyanide prevents the hydrolysis of the electrolyte, therefore solution containing cyanide ions must not be dropped into rivers and streams. This is responsible for killing fish and other animals.

b) Tin plated steel is used to make food cans. Food and beverages industries use tinplated steel cans. This is because the components of food beverages and the preservatives contain organic acids or their salts. They may form toxic substances by reacting with iron. These acids and salts are corrosive. Tin plating is non-poisonous and prevents corrosion.

Q.No.20 a) What is meant by corrosion?
b) Describe rusting of iron and corrosion of aluminium metals.

Ans:

a) Corrosion

The eating away of metals by environmental gases and water is called corrosion. It is a natural Redox Process that oxidizes metals to their oxides and sulphides under atmospheric conditions.

b) Rusting of Iron

Rust chemically is hydrated Ferric Oxide which appears on the surface of metal. Oxygen, water and CO_2 are necessary for iron to rust. A region of metal surface that has relatively less moisture, act as anode and Fe oxidizes into Fe^{+2} in this region as shown below.

$$Fe \longrightarrow Fe_{(aq)}^{+2} + 2e^-$$

Another region on the surface of metal that has relatively more moisture act as cathode. The electrons released in the oxidation process reduces atmospheric oxygen to hydrogen ions.

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

The irons ions, Fe^{+2} formed at the anodic regions flow to the cathodic regions through the moisture on the surface. Here Fe^{+2} ion further react with oxygen to form rust, $Fe_2O_3 . xH_2O$

Corrosion of Aluminium

Corrosion is not limited to iron. Aluminium is extensively used in the construction of aircraft, ships, cars, coking utensil, window frames, soda cans etc.. Aluminium has much higher tendency to oxidize than iron. Therefore, we might expect to see aircrafts, ships, cars, cooking utensils, soda cans transformed into piles of corroded aluminium. However, this does not occur. This is because a tough layer of insoluble aluminium oxide (Al_2O_3) forms on its surface when metal is exposed to air. This layer firmly adheres to the metal and serves to protect the underlying aluminium layers from further corrosion. On the other hand, the insoluble layer of rust, $Fe_2O_3.xH_2O$ that forms on the surface of iron is too porous to protect the underlying metal. This layer flakes away and exposes metal for further corrosion.

Q.No.21 : How rusting can be prevented?

Ans: Prevention of Corrosion or Rusting

Prevention of corrosion is an important way of conserving our natural resources.

Following methods have been devised to protect metals from corrosion:

1) Coating with paint

Corrosion can be prevented by painting the metal, so that is does not come in contact with oxygen and moisture and other harmful agents. Paint is cheap and easily applied. Paint is used to protect many everyday steel objects such as cars, trucks, trains, bikes, bridges etc. Paint also provides visual appeal.

2) Alloying

The tendency of iron to oxidize can be greatly reduced by alloying it with other metals. For example, stainless steel is an alloy of iron, chromium and nickel. It is protected from corrosion by an outer layer of Cr_2O_3.

3) Coating with a thin layer of another metal

Metals that readily corrode can be protected by coating with a thing layer of another metal that resists corrosion. This can be done by:

a) Tinning b) Galvanizing c) Electroplating

a) Tinning

In the process of tin plating, clean iron sheet is dipped in a bath of molten tin. It is then passed through hot pair of rollers. Tin protects iron effectively, since, it is very stable.

b) Galvanizing (Coating with Zinc)

The process of galvanizing consists of dipping a clean iron sheet in a hot zinc chloride bath and heating. After this sheet is rolled into zinc bath and cooled.

c) Electroplating

In electroplating an electrolytic process is used to deposit one metal on another metal.

4) Cathodic Protection

Cathodic protection is the process in which the metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminium. These metals are more active than iron, so they act as anode and iron as cathode. The more active metals themselves oxidize and save iron from corrosion. Cathodic protection is employed to prevent iron and steel structures such as pipes, tanks, oil rigs etc. in the moist underground and marine environment as shown in the figure.

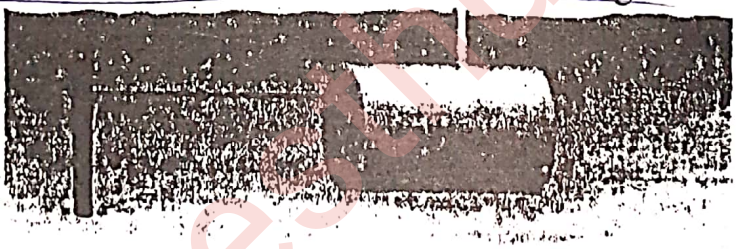


Figure Cathode protection

Society, Technology and Science

Silver is very soft metal. Silver atoms have weak interactions and are loosely packed together. Silver tarnishes in air when it comes in contact with trace quantities of H_2S or SO_2 in the atmosphere or food such as eggs, that are rich in sulphur compounds. Silver tarnish is silver sulphide that gives silver blackish appearance. Due to this reason decorative and practical objects made of solid silver gradually turn black and lose shining appearance. Decorative and practical object are plated with a thin layer of silver. Atoms in thin layers firmly adhere to the metal atoms of the object and form a durable layer. An article thickly plated with silver contains many layers of silver atoms. Such layers form soft covering. These layers gradually turn black.

KEY POINTS

- Oxidation is the gain of oxygen atom or loss of hydrogen atom or loss of electrons by a substance.
- Reduction is the loss of oxygen atom or gain of hydrogen atom or gain of electrons by a substance.
- Oxidation state or oxidation number is defined as the number of apparent charges that an atom will have in a molecule.
- The sum of oxidation state of all the atoms in a molecule of compound is zero.
- An oxidizing agent is the reactant containing the element that is reduced in a reaction.
- An electrochemical cell in which electrical energy is used to drive a chemical reaction is called an electrolytic cell.
- Oxidation always occurs at the anode.
- Reduction always occurs at cathode.
- An electrochemical cell in which a spontaneous oxidation reaction generates electricity is called a galvanic or voltaic cell.
- A galvanic cell consists of two half cells which are joined in series.
- In a galvanic cell oxidation half reaction occurs in anode compartment and reduction half reaction occurs in the cathode compartment.
- A battery is a galvanic cell or a group of galvanic cells joined in series.
- On the large scale, sodium metal is produced by the electrolysis of fused sodium chloride.
- Electrolysis of brine is used for the industrial production of sodium hydroxide.
- Corrosion is the process in which a metal reacts with oxygen and moisture in the atmosphere.
- Electrolytic process used to deposit one metal on another metal is called electroplating.
- Cathodic protection is the process in which metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminium.

TEXT BOOK EXERCISE

Q.1: Encircle the correct answer.

(i) In which of the following changes the nitrogen atom is reduced?

a) N_2 to NO b) N_2 to NO_2 **c) N_2 to NH_3** d) N_2 to HNO_3

(ii) Which of the following changes reaction is an example of oxidation?

a) Chlorine molecule to chloride ion **b) Silver atoms to silver (I) ion**

c) Oxygen molecule to oxide ion d) Iron (III) ion to iron (II) ion

(iii) Which of the following elements in the given reaction is reduced?

$$ZnO + H_2 \longrightarrow Zn + H_2O$$

a) H_2 b) ZnO **c) Zn** d) O

(iv) Consider the following reaction:

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

In this reaction what does H_2S behave as?

a) Reducing agent b) Oxidizing agent c) Catalyst d) Electrolyte

(v) The oxidation state of Cr in $K_2Cr_2O_7$ is:

a) +12 **b) +6** c) +3 d) Zero

(vi) Which of the following statement is not correct about the galvanic cell?

a) Cations are reduced at cathode

b) Anions are oxidized at anode

c) Electrons flow from cathode to anode

d) Oxidation occurs at anode

(vii) Which of the following is not true about the Daniel's cell?

a) Half cell of an active metal acts as cathode

b) Half cell contains an element in contact with its ions in aqueous solution

c) A salt bridge connects the two half cells

d) A spontaneous oxidation-reduction reaction generates electricity

(viii) Which of the following do not involve electrolytic process?

a) Refining of copper

b) Manufacture of sodium from NaCl

c) Electroplating of steel

d) Reduction of metal oxide by a reducing agent

(ix) Galvanizing is:

a) Coating with Sn **b) Coating with Zn**

c) Coating with Cr d) Coating with Cu

(x) Which of the following is true for the Nelson Cell?

a) Sodium metal is produced at anode

b) Chloride gas is produced at anode

c) Hydrogen gas is produced at anode

d) Sodium ions are not reduced at cathode

Q.2: Give the short answers.

(i) What is oxidation state?

Ans:

See Question 4C.

(ii) What is the oxidation number of Cr in chromic acid (H_2CrO_4)?

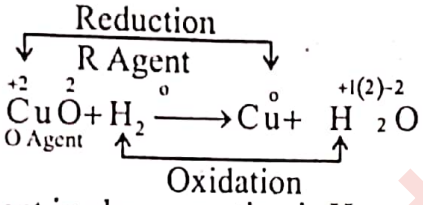
Ans:

$$2H + Cr + 4(O) = 0$$
$$2(+1) + Cr + 4(-2) = 0$$
$$+2 + Cr - 8 = 0$$
$$Cr - 6 = 0$$
$$Cr = +6$$

(iii) Identify reducing agent in the following reaction

$$CuO + H_2 \longrightarrow Cu + H_2O$$

Ans:



So reducing agent in above reaction is H_2.

(iv) Write chemical reactions that occur in Nelson's Cell.

Ans:

See Question 16 (c)

(v) Why tin plated steel is used to make food cans?

Ans:

See Question 19 (b)

(vi) Explain one example from daily life which involves oxidation-reduction reaction?

Ans:

See Question 3 b (4).

Q.3: Define oxidation and reduction in terms of loss or gain of oxygen or Hydrogen.

Ans:

See Question 4 (a)

Q.4: Define oxidation and reduction in terms of loss or gain of electrons.

Ans:

See Question 4 (b)

Q.5: List the possible uses of electrolytic cell.

Ans:

See Question 12.

Q.6: Sketch a Daniel Cell, labeling the cathode, anode, and the direction of flow of the electrons.

Ans:

See Question 13 (c).

Q.7: Describe how a battery produces electrical energy.

Ans:

See Question 15 (b).

Q.8: Describe the method of recovering metal from its ores.

Ans:

See Question 8 (b)

Q.9: Explain electrolytic refining of copper.

Ans:

See Question 17.

Q.10: Compare the effects of Al_2O_3 and Fe_2O_3 formation on their parent metals and cite example from daily life.

Ans:

Fe_2O_3 is used in blast furnace method and Al_2O_3 is used in Electrolysis for the extraction of Al.metal. Both metals are used in Alloys.

$4Al + 3O_2 \longrightarrow 2Al_2O_3$, $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

Q.11: Explain how food and beverage industries deal with Corrosion?

Ans:

See Question 19 (b).

Q.12: Explain how chemistry interacts with photography?

Ans:

See Question 5.

Q.13: Electrolysis has a major role in electrochemical industries.

a) Sketch an electrolytic cell, label the anode and cathode and indicate the direction of electron transfer.

Ans:

See Question 16 (b).

a) Describe the nature of electrochemical process.

Ans:

See Question 9 (c).

c) Distinguish between electrolytic and voltaic cell.

Ans: See Question 9 (b).

Q.14; State the substances which are oxidized or reduced. Give reason for you answer.

a) $N_2 + 3H_2 \longrightarrow 2NH_3$ **b)** $CO_2 + 2Mg \longrightarrow 2MgO + C$

c) $Mg + H_2O \longrightarrow MgO + H_2$ $H_2S + Cl_2 \longrightarrow 2HCl + S$

e) $2NH_3 + 3CuO \longrightarrow 3CuO + N_2 + 3H_2O$

Ans:

Reduction

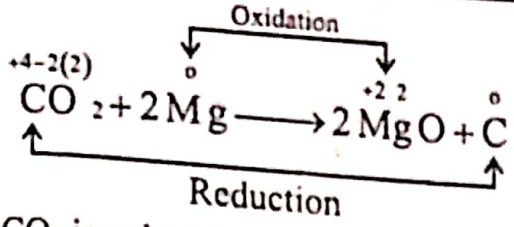
$\overset{0}{N_2} + 3\overset{0}{H_2} \longrightarrow 2\overset{-3}{N}\ \overset{+1(3)}{H_3}$

Oxidation

a)

N_2 is Reduced into NH_3 because oxidation No. of N_2 decreases from $0 \longrightarrow -3$ while H_2 is oxidized into NH_3 because its oxidation No. increases from $0 \longrightarrow +1$.

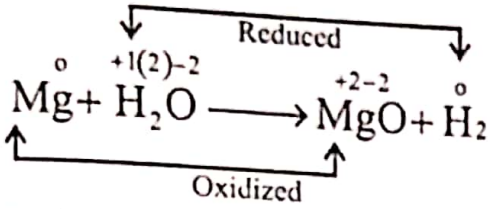
b)



CO_2 is reduced into C because oxidation No. of C in CO_2 decreases from $+4 \longrightarrow 0$

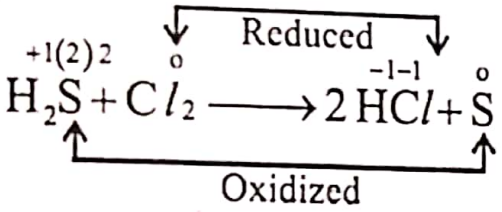
While Mg is oxidized into MgO because oxidation No. of Mg increases from $0 \longrightarrow +2$ In MgO.

c)



Mg is oxidized into MgO because oxidation No. of Mg increases from $0 \longrightarrow +2$ in MgO while H_2O reduced into H_2 because oxidation No. of H in H_2O decreases from $+1 \longrightarrow 0$

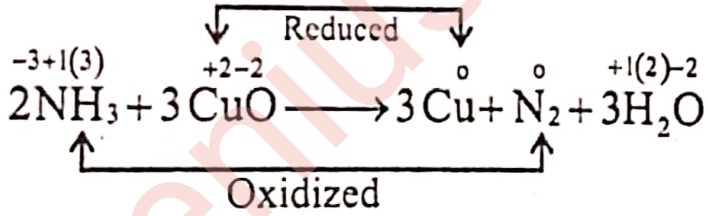
d)



H_2S is oxidized into S because oxidation No. of S in H_2S increases from $-2 \longrightarrow 0$ in S

While Cl_2 is reduced into HCl because its oxidation No. decreases from $0 \longrightarrow -1$ in HCl

e)



NH_3 is oxidized into N_2 because oxidation No. of N in NH3 increases from $-3 \longrightarrow 0$ In N_2 while CuO is reduced into Cu because oxidation No. of Cu in CuO decreases from $+2 \longrightarrow 0$ in Cu.

Q.15: a) Define oxidation number or oxidation state.

b) Find the oxidation state of nitrogen in the following compounds.

(i) NO_2 (ii) N_2O (iii) N_2O_3 (iv) HNO_3

Ans:

a) See Question 4 (c).

b) oxidation numbers of Nitrogen in

(i) $\overset{+4-2(2)}{NO_2}$ is +4 (ii) $\overset{+1(2)-2}{N_2O}$ is +1 (iii) $\overset{-3(2)-2(3)}{N_2O_3}$ is +3 (iv) $\overset{-1+5-2(3)}{HNO_3}$ is +5

Q.16: Find the oxidation state of S in the following compounds.
(i) H_2S (ii) H_2SO_3 (iii) Na_2SO_3

Ans:

Oxidation state of S in

(i) $\overset{+1(2)-2}{H_2Sis-2}$ (ii) $\overset{+1(2)+4\ 2(3)}{H_2\ SO_3}$ is +4 (iii) $\overset{+1(2)\quad 2(3)}{Na_2\,S\,O_3}$ is +2

Q.17: a) Define oxidizing and reducing agents.
b) Identify the oxidizing agents and reducing agents in the following reactions.

(i) $H_2S + Cl_2 \longrightarrow 2HCl + S$ (ii) $2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$

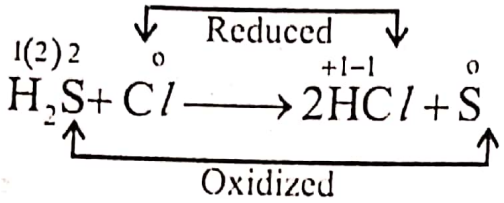
(iii) $2KI + Cl_2 \longrightarrow 2KCl + I_2$ (iv) $Mg + 2HCl \longrightarrow MgCl + H_2$

Ans:

a) See Question 7.

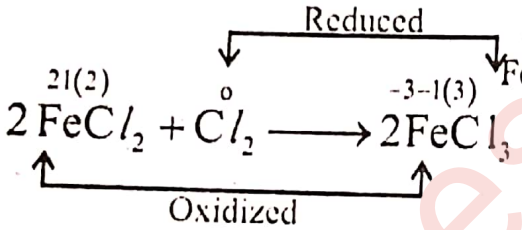
b) Identification of oxidizing and Reducing Agents in following reactions.

(i)



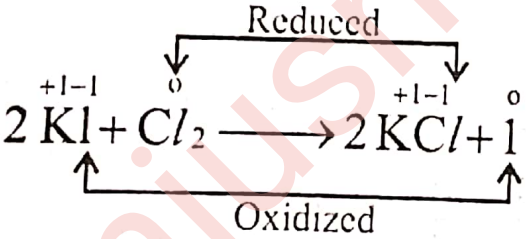
So Cl_2 oxidizing agent while H_2S is reducing agent.

(ii)



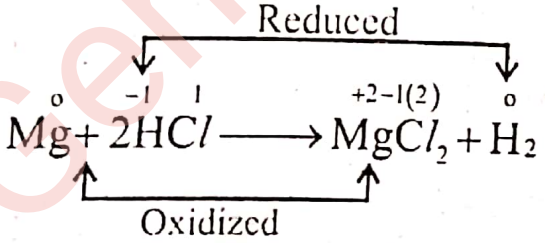
$FeCl_2$ is reducing agent while Cl_2 is oxidizing agent.

(iii)



KI is reducing agent while Cl_2 oxidizing agent.

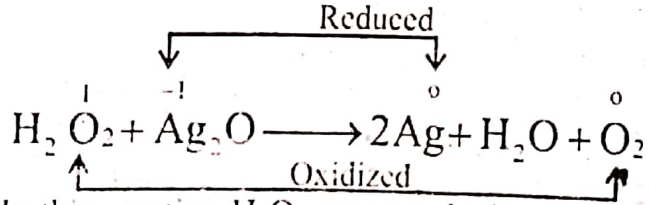
(iv)



Mg is reducing agent while HCl is oxidizing agent.

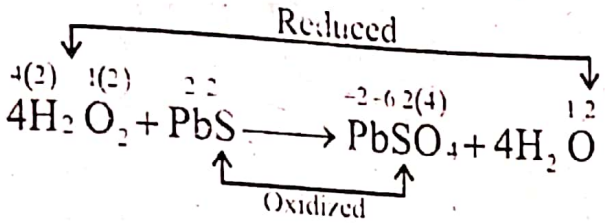
Q.18:

(i)



In this reaction H_2O_2 act as reducing agent because oxidation No. of O in H_2O_2 increases from $-1 \longrightarrow 0$ in O_2

(ii)



In this reaction, H_2O_2 act as oxidizing agent because oxidation No. of O in H_2O_2 decreases from $-1 \longrightarrow -2$ in H_2O

SELF ASSESSMENTS

SELF ASSESSMENT EXERCISE 7.1

Identify elements undergoing oxidation and reduction in the following reactions.

1) $N_2 + 3H_2 \longrightarrow 2NH_3$
2) $2H_2 + O_2 \longrightarrow 2H_2O$
3) $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
4) $4Al + 3O_2 \longrightarrow 2Al_2O_3$

Solution

1) Nitrogen undergo reduction after gaining Hydrogen and forms NH_3 (Ammonia)
2) Hydrogen undergo oxidation after gaining oxygen and forms H_2O or oxygen undergo reduction after gaining Hydrogen and forms H_2O.
3) Fe_2O_3 undergo reduction after losing oxygen and forms Iron metal or CO undergo oxidation after gaining oxygen and form CO_2.
4) Al undergo oxidation after gaining oxygen and forms Al_2O_3.

SELF ASSESSMENT EXERCISE 7.2

In the following reactions, identify which element is oxidized and which element is reduced.

i) $Mg + S \longrightarrow MgS$
ii) $4Na + O_2 \longrightarrow 2Na_2O$
iii) $2Al + 3Cl_2 \longrightarrow 2AlCl_3$
iv) $Mg + H_2 \longrightarrow MgH_2$

Solution

(i) $Mg^{\circ} + S^{\circ} \longrightarrow Mg^{+2}S^{-2}$

So Mg° is oxidized into Mg^{+2} in MgS (oxidation no increases)
While S° is reduced into S^{-2} in MgS (oxidation no decreases)

(ii) $4Na^{\circ} + O^{\circ}_2 \longrightarrow \overset{+1(2)}{2\,Na_2\,O^{-2}}$

Na° is oxidized into Na^{+1} in Na_2O (oxidation no increases)
While O°_2 is reduced into O^{-2} in Na_2O (oxidation no decreases)

(iii) $2Al^{\circ} + 3Cl^{\circ}_2 \longrightarrow \overset{+3\ -1(3)}{2\,Al\,Cl_3}$

Al° is oxidized into Al^{+3} in $AlCl_3$ (oxidation no increases)
While Cl_2 is reduced into Cl^{-1} in $AlCl_3$ (oxidation no decreases)

(iv) $Mg^{\circ} + H^{\circ}_2 \longrightarrow Mg^{+2}H_2^{-1(2)}$

Mg° is oxidized into Mg^{+2} in MgH_2 (oxidation no increases)
While H°_2 is reduced into H^{-1} in MgH_2 (oxidation no decreases)

SELF ASSESSMENT EXERCISE 7.3 or

What is acid rains. How it is formed?

One major problem of air pollution is the formation of acid rain. Air pollutants such as SO_2 and NO_2 combine with oxygen and water vapours in the air to form H_2SO_4 and HNO_3. These acids fall to the ground with the rain, making the rain acidic. Clouds can also absorb the acids and carry them hundreds of kilometer away from where the pollutants are released. Determine the oxidation number of N in NO_2 and HNO_3, S in SO_2 and H_2SO_4.

Solution.

Acid rains is formed by the following chemical reactions.

$$N_2 + O_2 \longrightarrow 2NO$$

$$2NO + O_2 \longrightarrow 2NO_2$$

$$2NO_2 + \frac{1}{2}O_2 + H_2O \longrightarrow 2HNO_3,$$

thus or $4NO_2 + O_2 + 2H_2O \longrightarrow 4HNO_3$ acid fall in the form of rain called Acid Rain

Similarly

$$S + O_2 \longrightarrow SO_2$$

$$2SO_2 + O_2 \longrightarrow 2SO_3$$

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

So acid formed fall in the form of rain so called Acid Rain.

Determination of Oxidation Number

O. N of N in NO_2 is

O. N of N + 2 (O) = 0

O. N of N + 2 (-2) = 0

O. N of N – 4 = 0

O. N of N = +4

In HNO_3

H + N + 3 (O) = 0

+1 + N + 3 (-2) = 0

+1 + N – 6 = 0

N – 5 = 0

N = + 5

0. N of S in SO_2

S + 2 (O) = 0

S + 2 (-2) = 0

S – 4 = 0

S = + 4

0.N of S in H_2SO_4

2H + S + 4(O) = 0

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

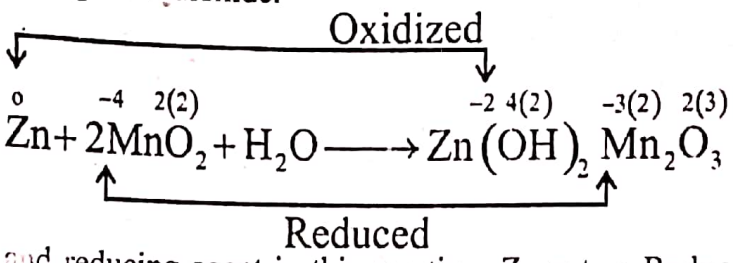
+2 + S – 8 = 0

S – 6 = 0

S = + 6

SELF ASSESSMENT EXERCISE 7.4

1) The torch cell discharges electricity because of an oxidation-reduction reaction that takes place between zinc and manganese dioxide.



Identify the oxidizing and reducing agent in this reaction. Zn act as Reducing Agent as itself oxidized while MnO_2 act as oxidizing agent as itself being reduced.

2) Identify oxidizing and reducing agents in the following reactions:

a) $2\overset{0}{S} + \overset{0}{Cl_2} \longrightarrow \overset{4(2)\ -1(2)}{S_2Cl_2}$ S is Reducing Agent and Cl_2 is oxidizing Agent.

b) $2\overset{0}{Na} + \overset{0}{Br_2} \longrightarrow \overset{4\ 1}{2NaBr}$ Na is Reducing Agent and Br_2 is oxidizing Agent.

c) $\overset{0}{H_2} + \overset{0}{S} \longrightarrow \overset{4(2)\ 2}{H_2S}$ H_2 is Reducing Agent and S is oxidizing Agent.

SEFL ASSESSMENT EXERCISE 7.5

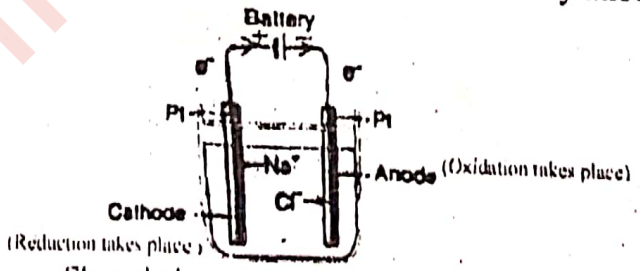
1) Sketch an electrolysis cell for the electrolysis of fused KCl

Solution

Fig. Migration of ions in electrolytic cell
During the electrolysis of molten KCl

During the electrolysis of molten KCl

2) Electrolytic cell shown in figure below is used for the electrolysis of fused sodium chloride. Indicate the direction of flow of electrons. Identify anode an cathode.



Electrolytic cell for the electrolysis of fused NaCl

The direction of flow of electrons will be from Anode to cathode

SELF ASSESSMENT EXERCISE 7.5(A)

Sketch a voltaic cell labeling the cathode, the anode and the direction of flow of electrons. Use the following chemicals:
Silver, Zinc, Silver Nitrate ($AgNO_3$) and zinc sulphate ($ZnSO_4$)
(Hint: Zn is more active than Ag)

SELF ASSESSMENT EXERCISE 7.6

Identify the half-cell in which oxidation occurs and the half cell in which reduction occurs in the following voltaic cell according to the following figure.

Following reactions occur in the cell.

At Anode $Zn(s) \longrightarrow Zn^{+2}_{(aq)} + 2e^-$ oxidation takes place

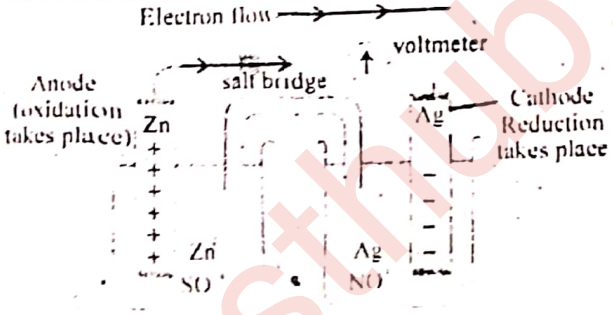
At Cathode $H_{2\,(g)} + 2e^- \longrightarrow 2H^-_{(aq)}$ Reduction takes place

THINK TANK

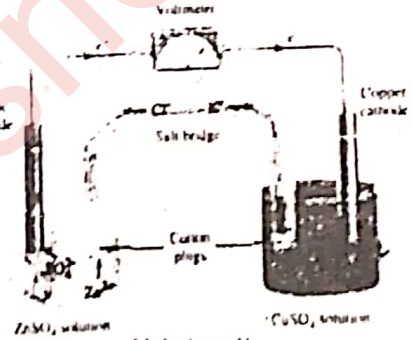
Q.1: What materials do you need to electroplate copper onto an iron nail. Make a diagram showing how these materials should be arranged.

Ans:

Apparatus for electroplating copper. Copper dissolves at the anode where it is oxidized to Cu^{+2}. Copper is deposited on the cathode (the nail) where Cu^{+2} is reduced.



Q.2: Describe the process that is occurring in the following illustration. Shoe has steel strips.



Voltaic cell

Ans:

The above sketch indicated the electroplating of copper over shoe steel strips. In this experiment, copper bar act as anode releases two electrons (oxidation takes place) while shoe strip connected with cathode where reduction takes place i.e., pickup electron. As a whole, Redox Reaction is carried out as shown in the following reactions.

At Anode, oxidation takes place which means loss of electrons.

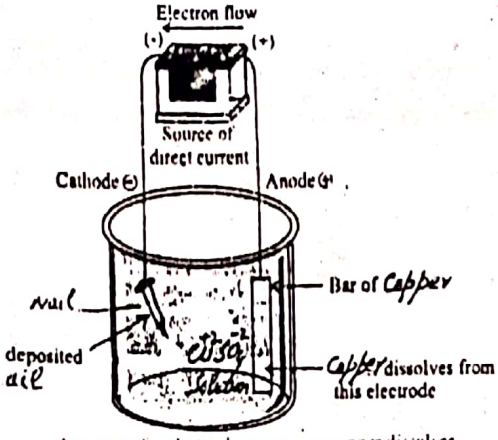
$$Cu - 2e^- \longrightarrow Cu^{+2}$$

At Cathode, reduction takes place which means gain of electrons.

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

Q.3: Following redox reaction occurs in the voltaic cell illustrated below.

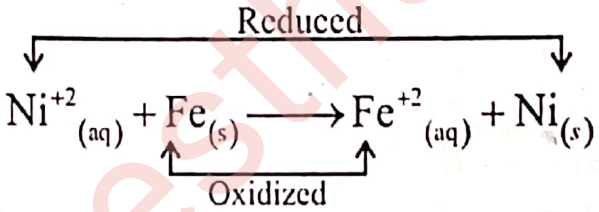
$$Ni^{+2}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{+2}_{(aq)} + Ni_{(s)}$$



Apparatus for electorplating copper copper dissolves at the anode where it is oxidized to Cu^{+2} is deposited on the cathode (the nail) where Cu^{+2} is reduced

Identify that anode, cathode and indicate the direction of flow of electrons

Ans:



$$Ni^{+2}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{+2}_{(aq)} + Ni_{(s)}$$

In the given Redox Reaction.

Ni^{+2} is reduces to Ni so act as oxidizing agent and the electrode of Ni^{+2} act as cathode while Fe is oxidizes into Fe^{+2} so act as reducing agent and the electrode of Fe act as Anode. The reactions carried out as shown below.

At Anode (oxidation) loss of electrons

$$Fe_{(s)} - 2e^- \longrightarrow Fe^{+2}_{(aq)}$$

At Cathode (reduction) gain of electrons.

$$Ni^{+2}_{(aq)} + 2e^- \longrightarrow Ni_{(s)}$$

Hence the direction of flow of electrons will be from Anode to cathode.

Q.4: Design an experiment to demonstrate cathodic protection from corrosion.

Ans:

The process in which a thin layer of tin is coated on metal (Iron) sheet to stop corrosion is called Tin Plating. It is called Coating Iron with Tin or Cathode Coating. It is done by dipping a clean sheet of iron in molten hot tin e.g., Tin cans, oil container are coated by Tin. Tin itself is very stable and protects other metals from corrosion. If tin coating is damaged, then corrosion of Iron takes, place more rapidly and Galvanic Cell is formed in which tin acts as cathode and Iron act as anode. Thus Iron flows from iron to Tin. Therefore H^+, OH^- and Fe^{+3} ion are produced in solution. Hence Fe is eaten away in the form of $Fe(OH)_3$ rapidly.

For experimentation explanation See Question 18 (c) ii.