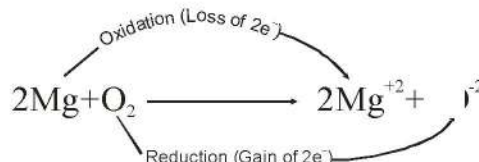


**TEST YOUR SELF****TEST YOUR SELF 7.1 OXIDATION AND REDUCTION REACTIONS**

- i. How can you justify that a reaction between magnesium and oxygen is a redox reaction, while the reaction shows only addition of oxygen (oxidation)



Ans: Assigning oxidation states



In this reaction, loss of electron also occurs instead of addition of oxygen called oxidation, while gain of electrons is called reduction. So, in redox reactions, oxidation and reduction occur simultaneously which is collectively called redox reaction.

- ii. A reaction between carbon and oxygen involved only addition of oxygen (oxidation), but, it is called a redox reaction. Comment on this.

Ans: Assigning oxidation states to the elements



In this reaction, carbon has lost 4 electrons while  $\text{O}_2$  has gained these 4 electrons. So in this reaction oxidation and reduction occur simultaneously which is collectively called redox reaction.

- iii. Oxidation and reduction proceed simultaneously. Explain, with an example.

Ans: Assigning oxidation states to the elements



In this reaction, carbon has lost 4 electrons while  $\text{O}_2$  has gained these 4 electrons. So in this reaction oxidation and reduction occur simultaneously which is collectively called redox reaction.

- iv. Identify which of the followings is oxidation or reduction reaction.

- $\text{K} \longrightarrow \text{K}^+ + 1e^-$
- $\text{Br} + 1e^- \longrightarrow \text{Br}^-$
- $\text{Cu} \longrightarrow \text{Cu}^{+2} + 2e^-$
- $\text{I}^- \longrightarrow \text{I} + 1e^-$
- $\text{Fe}^{+2} \longrightarrow \text{Fe}^{+3} + 1e^-$

Ans.

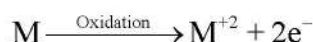
- |    |  |                               |
|----|--|-------------------------------|
| a. | $\text{K} \longrightarrow \text{K}^+ + 1e^-$           | Oxidation (Loss of electron)  |
| b. | $\text{Br} + 1e^- \longrightarrow \text{Br}^-$         | Reduction (Gain of electron)  |
| c. | $\text{Cu} \longrightarrow \text{Cu}^{+2} + 2e^-$      | Oxidation (Loss of electrons) |
| d. | $\text{I}^- \longrightarrow \text{I} + 1e^-$           | Oxidation (Loss of electron)  |
| e. | $\text{Fe}^{+2} \longrightarrow \text{Fe}^{+3} + 1e^-$ | Oxidation (Loss of electron)  |

- v. An element M reacts with another element X to form  $\text{MX}_2$ . In terms of loss or gain of electrons, identify the element which is oxidized and which is reduced.

Ans: Consider the reaction



In this reaction, "M" release two electrons and become " $\text{M}^{+2}$ " so oxidation reaction is:



While "X" gains two electrons, so reduction reaction is



So in this reaction "M" is oxidized and X is reduced.

- vi. **How can you justify that the following reaction is not only an oxidation reaction but also a complete redox reaction.**



Ans: Assigning oxidation states to the elements

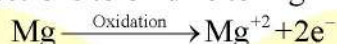


In this reaction carbon has lost 2 electrons while Fe has gained these 2 electrons. So in this reaction oxidation and reduction occur simultaneously which is collectively called redox reaction.

- vii. **Explain the term oxidation on the basis of electronic concept with an example**  
Ans. **Oxidation**

*"The loss of electrons by an atom or an ion is called oxidation."*

Example: Mg loses 2 electrons to oxidize to  $\text{Mg}^{+2}$



### TEST YOUR SELF 7.2 OXIDATION STATE AND RULES FOR ASSIGNING OXIDATION STATE

- i. Find out the oxidation numbers of the following elements marked in bold in the formulae:  $\text{Ba}_3(\text{PO}_4)_2$ ,  $\text{CaSO}_4$ ,  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{Al}_2(\text{SO}_4)_3$

Ans.  $\text{Ba}_3(\text{PO}_4)_2$ :

$$2[\text{O.N. of Ba}] + 2[\text{O.N. of P}] + 8[\text{O.N. of O}] = 0$$

$$3[+2] + 2[P] + 8[-2] = 0$$

$$+6 + 2P - 16 = 0$$

$$2P - 10 = 0$$

$$2P = 10$$

$$P = \frac{10}{2}$$

$$P = +5$$

$\text{CaSO}_4$ :

$$[\text{O.N. of Ca}] + [\text{O.N. of S}] + 4[\text{O.N. of O}] = 0$$

$$[+2] + [S] + 4[-2] = 0$$

$$+2 + S - 8 = 0$$

$$S - 6 = 0$$

$$S = +6$$

$\text{Cu}(\text{NO}_3)_2$ :

$$[\text{O.N. of Cu}] + 2[\text{O.N. of N}] + 6[\text{O.N. of O}] = 0$$

$$[+2] + 2[N] + 6[-2] = 0$$

$$+2 + 2N - 12 = 0$$

$$2N - 10 = 0$$

$$2N = +10$$

$$N = \frac{10}{2}$$

$$N = +5$$



**Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub>:**

$$\begin{aligned}
 2[\text{O.N. of Al}] + 3[\text{O.N. of S}] + 12[\text{O.N. of O}] &= 0 \\
 \Rightarrow 2[+3] + 3[S] + 12[-2] &= 0 \\
 \Rightarrow +6 + 3S - 24 &= 0 \\
 \Rightarrow 3S - 18 &= 0 \\
 \Rightarrow 3S &= +18 \\
 S &= \frac{18}{3} \\
 S &= +6
 \end{aligned}$$

**ii. In a compound MX<sub>3</sub>, find out the oxidation number of M and X.****Ans. Compound = MX<sub>3</sub>**

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

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

$$(+3) + [\text{Oxidation no. of "X"}] 3 = 0$$

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

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

$$\text{Oxidation no. of X} = -1$$

Thus in a compound MX<sub>3</sub> O.N. of M is +3 and that of X is -1.**iii. Why the oxidation number of oxygen in OF<sub>2</sub> is +2?****Ans:** The oxidation number of oxygen in OF<sub>2</sub> is +2 because fluorine is more electronegative as compared to oxygen. Therefore, fluorine will carry negative (-) charge while oxygen will carry (+2) charge on it.**iv. In H<sub>2</sub>S, SO<sub>2</sub> and H<sub>2</sub>SO<sub>4</sub>, the sulphur atom has different oxidation numbers. Find out the oxidation number of sulphur in each compound.****Ans.****H<sub>2</sub>S:**

$$\begin{aligned}
 2[\text{O.N of H}] + [\text{O.N. of S}] &= 0 \\
 2[+1] + [S] &= 0 \\
 +2 + S &= 0 \\
 S &= -2
 \end{aligned}$$

**SO<sub>2</sub>:**

$$\begin{aligned}
 [\text{O.N. of S}] + 2[\text{O.N. of O}] &= 0 \\
 [S] + 2[-2] &= 0 \\
 S - 4 &= 0 \\
 S &= +4
 \end{aligned}$$

**H<sub>2</sub>SO<sub>4</sub>:**

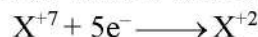
$$\begin{aligned}
 2[\text{O.N of H}] + [\text{O.N of S}] + 4[\text{O.N of O}] &= 0 \\
 2[+1] + [S] + 4[-2] &= 0 \\
 +2 + S - 8 &= 0 \\
 S - 6 &= 0 \\
 S &= +6
 \end{aligned}$$

**v. An element X has oxidation state 0. What will be its oxidation state when it gains three electrons?****Ans:** When an element X<sup>0</sup> gains the three electrons it will have oxidation state -3 such as:



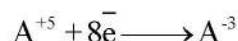
- vi. An element in oxidation state +7 gains electrons to be reduced to oxidation state +2. How many electrons did it accept?

Ans: When an element in oxidation state +7 gains electrons to be reduced to oxidation state +2 it will accept 5 electrons in its valance shell as shown by the equation



- vii. If the oxidation state of an element changes from +5 to -3, Has it been reduced or oxidized? How many electrons are involved in this process?

Ans: If the oxidation state of an element changes from +5 to -3, then it will be reduced. In this process of reduction



### TEST YOURSELF 7.3 OXIDIZING AND REDUCING AGENTS

- i. In the following reaction, how can you justify that  $H_2S$  is oxidized and  $SO_2$  is reduced.



Ans: In this reaction hydrogen is removed from  $H_2S$ , therefore  $H_2S$  is oxidized while oxygen is removed from  $SO_2$ , therefore  $SO_2$  is reduced to S.

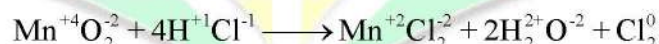
- ii. The reaction between  $MnO_2$  and  $HCl$  is a redox reaction written as balance chemical equation.



Find out:

- The substance oxidized
- The substance reduce
- The substance reduced
- The substance which acts as reducing agent.

Ans: Assigning oxidation states to the elements

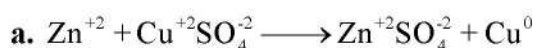


- Cl is oxidized
- Mn is reduced
- $MnO_2$  is oxidizing agent.
- $HCl$  is reducing agent

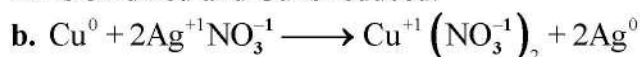
- iii. The following reactions are redox reactions. Find out the element which has been reduced and the element which has been oxidized?

- $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$
- $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$
- $H_2S + Cl_2 \longrightarrow 2HCl + S$

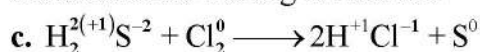
Ans:



Zn is oxidized and Cu is reduced.



Cu is oxidized and Ag is reduced.



S is oxidized and Cl is reduced.

- iv. Why the following reaction is not a redox reaction. Explain with reasons?



Ans: This is an example of neutralization reaction rather than a redox reaction because in



acid base reactions acid reacts with a base to form salt and water.



### TEST YOURSELF 7.4 ELECTRO CHEMICAL CELLS

i. **Why are the strong electrolytes termed as good conductors?**

**Ans.** Strong electrolytes are good conductor because they are completely ionized in aqueous solutions and produce more and more ions which helps in conduction.

ii. **Does non-electrolyte forms ions in solution?**

**Ans.** Non-electrolytes exist in the form of molecules. They do not form ions in solution and do not allow the current to pass through their solutions.

iii. **What is difference between a strong electrolyte and a weak electrolyte?**

**Ans:**

Strong Electrolyte	Weak Electrolyte
<i>The electrolytes which ionize completely in aqueous solutions and produce more ions, are called strong electrolytes.</i>	<i>The electrolytes which ionize to a small extent when dissolved in water and could not produce more ions are called weak electrolytes.</i>
<b>Examples:</b> Strong electrolytes are aqueous solutions of NaCl, NaOH and H <sub>2</sub> SO <sub>4</sub> . $\text{NaOH}_{(s)} \xrightarrow{\text{H}_2\text{O}} \text{Na}^{+}_{(aq)} + \text{OH}^{-}_{(aq)}$	<b>Examples:</b> Weak electrolytes are the aqueous solution of acetic acid Ca(OH) <sub>2</sub> etc. $\text{CH}_3\text{COOH}_{(l)} + \text{H}_2\text{O}_{(l)} \longrightarrow \text{CH}_3\text{COO}^{-}_{(aq)} + \text{H}_3\text{O}^{+}_{(aq)}$

iv. **Identify a strong or weak electrolyte among the following compounds.**

CUSO<sub>4</sub>, H<sub>2</sub>CO<sub>3</sub>, Ca(OH)<sub>2</sub>, HCl, AgNO<sub>3</sub>

**Ans:**

- i. CuSO<sub>4</sub>: Strong electrolyte
- ii. H<sub>2</sub>CO<sub>3</sub>: Weak electrolyte
- iii. Ca(OH)<sub>2</sub>: Weak electrolyte.
- iv. HCl: Strong electrolyte
- v. AgNO<sub>3</sub>: Strong electrolyte

v. **Which force drives the non-spontaneous reaction to take place?**

**Ans:** Non-spontaneous reactions take place in the presence of an external agent. The external agent are electrons that cause electricity. So, electric energy helps the non-spontaneous reactions to proceed.

vi. **Which type of chemical reaction takes place in electrolytic cell?**

**Ans:** Non-spontaneous chemical reaction takes place in electrolytic cell with the help of electricity



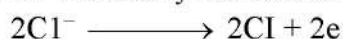
vii. **What type of reaction takes place at anode in electrolytic cell?**

**Ans:** Oxidation takes place at anode in electrolytic cell. Anode is a positive charge electrode.



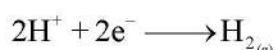
viii. **Why the positively charged electrode is called anode in electrolytic cell?**

**Ans:** The positive charged electrode is called anode in electrolytic cell because it is connected to the positive (+) terminal of the battery and anions move towards it for oxidation.



ix. **In the electrolysis of water, towards which terminal H<sup>+</sup> ions move?**

**Ans:** In the electrolysis of water, H<sup>+</sup> ions move toward the cathode and reduced to hydrogen gas which is liberated.



x. **In the electrolysis of water, where is the oxygen produced?**



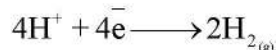
**Ans:** In the electrolysis of water oxygen is produced at anode.



**xi. Towards which electrode of the electrolytic cell moves the cations and what do they do there?**

**Ans:** in the electrolytic cell cations carry (+) charge, they move towards the cathode and are reduced there.

Such as:



**xii. How the half cells of a galvanic cell are connected? What is function of salt bridge?**

**Ans:** The two half cells of the galvanic cell are connected by salt bridge.

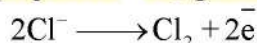
**Function of salt bridge:**

The function of salt bridge is to maintain the electric neutrality between the two half cells of a galvanic cell and thus maintain the flow of ionic current.

### TEST YOURSELF 7.5 (ELECTRO CHEMICAL INDUSTRIES)

**i. Anode of Down's cell is made of a non-metal, what is its name? What is the function of this anode?**

**Ans:** In Down's cell, the anode is made of graphite.  $\text{Cl}^-$  ions get oxidized to give chlorine gas ( $\text{Cl}_2$ ) gas at the anode. So, its function is to oxidize  $\text{Cl}^-$  to  $\text{Cl}_2$  gas. The  $\text{Cl}^-$  ions are oxidized there produce  $\text{Cl}_2$  gas at anode.

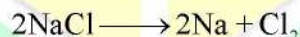


**ii. Where does the sodium metal is collected in Down's cell?**

**Ans:** In Down's cell  $\text{Na}^+$  are reduced at cathode and molten Na-metal floats on the denser molten salt mixture from where it is collected in a side tube.

**iii. What is the name of the by-product produced in the Downs cell?**

**Ans:** The name of by-product produced in the Downs cell is chlorine gas.



**iv. Are anodes of Down's cell and Nelson cell made of same element? If yes, what is its name?**

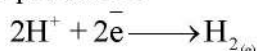
**Ans:** Yes, anodes of Down's cell and Nelson cell are made of same element named as graphite.

**v. What is the shape of cathode in Nelson's cell? Why is it perforated?**

**Ans:** In Nelson's cell the cathode is internally lined with asbestos diaphragm. It is U shaped and perforated. It is made perforated because sodium hydroxide slowly percolates through it into a catch basin.

**vi. Which ions are discharged at cathode in Nelson's cell and what is produced at cathode?**

**Ans:** The  $\text{H}^+$  ions are discharged at cathode in Nelson's cell. At cathode, sodium hydroxide ( $\text{NaOH}$ ) and hydrogen ( $\text{H}_2$ ) is produced.



### TEST YOURSELF 7.6 (CORROSION AND ITS PREVENTION)

**i. What is the difference between corrosion and rusting?**

**Ans:**

Corrosion	Rusting
<b>i.</b> Corrosion is the general term used for all the metals	<b>i.</b> Corrosion of iron is called rusting
<b>ii.</b> Corrosion of some metals may be stopped	<b>ii.</b> Rusting is the continuous process.
<b>iii.</b> It is redox reaction	<b>iii.</b> It is also redox reaction.

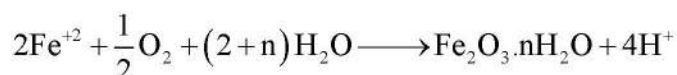
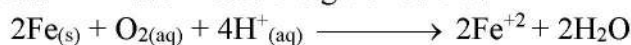
**ii. What happens to iron in the rusting process?**



**Ans:** During rusting Fe is oxidized to  $\text{Fe}^{+2}$ , that spread throughout the surrounding water and reacts with  $\text{O}_2$  to form the salt  $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$  called rust. Due to being porous, the rusting process continues until all the iron is eaten up.

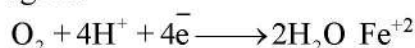
**iii. Rusting completes in how many redox reactions?**

**Ans:** Rusting completes in 2 redox reactions given below.

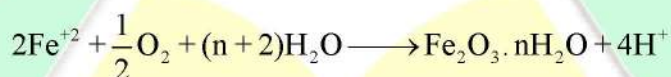


**iv. Explain the role of  $\text{O}_2$  in rusting?**

**Ans:**  $\text{O}_2$  is necessary for rusting because in the cathode region, the electrons released by iron reduce oxygen in the presence of  $\text{H}^+$  ions to form water. Oxygen acts as an oxidizing agent.



The  $\text{Fe}^{+2}$  ions are further oxidized to  $\text{Fe}^{+3}$  which then combine with oxygen to form rust.



**v. 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 metal to protect them from corrosion.

**vi. What do you mean by galvanizing?**

**Ans:** The process of coating a thin layer of zinc on iron is called galvanizing. It has advantage that it resists corrosion even if zinc coating is broken.

**vii. 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.

**viii. Why tin plated iron is rusted rapidly when tin layer is broken?**

**Ans:** When tin layer is broken the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

**ix. Name the metal which is used for galvanizing iron?**

**Ans:** Zinc metal is used for galvanizing iron.

### TEST YOURSELF 7.7 (ELECTROPLATING)

**i. Define electroplating?**

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

**Objective:**

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

**ii. How electroplating of zinc is carried out?**

**Ans:** The target metal is cleaned in alkaline detergent solution and it is treated with acid in order to remove any rust or surface scales.

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



**iii. Which material is used to make cathode electroplating?**

You are using the  
**UNREGISTERED**  
version now.

- Ans.** Some sheet of iron is used to make cathode in electroplating.
- iv. Why is the anode made up of a metal to be deposited during electrolysis**
- Ans.** The anode is made up of a metal to be deposited during electrolysis, because, when the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and deposit on it.



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