

**MODERN'S**  
abc OF



Award for  
Excellence  
in Book  
Production

FIRST PRIZE

# Chemistry

Dr. S.P. Jauhar

CLASS-XI

PART-II



STRICTLY ACCORDING TO THE NEW SYLLABUS

**INDIA'S FIRST SMART BOOK**

According to new syllabus prescribed by Central Board of Secondary Education (CBSE), New Delhi and State Boards of Uttarakhand, Karnataka (1st Year PUC), Chhattisgarh, Jharkhand, Punjab, Haryana, Himachal, Kerala, Mizoram, Meghalaya, Nagaland, Assam, Manipur and other States following CBSE curriculum.

MODERN'S  
abc <sup>+</sup> OF

# CHEMISTRY

INDIA'S FIRST SMART BOOK

For Class XI

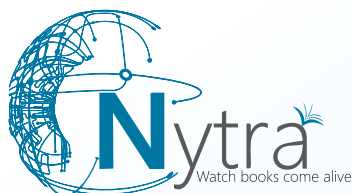
**PART-II**

Strictly  
in accordance  
with the Latest  
Guidelines and Syllabus  
Including  
Value Based Questions  
issued by  
N.C.E.R.T. / C.B.S.E.

By

**Dr. S.P. JAUHAR**

Formerly Professor of Chemistry  
Department of Chemistry  
Panjab University, Chandigarh



REVISED & UPDATED  
EDITION 2016–2017



**MODERN PUBLISHERS**  
(Producers of Quality Textbooks)

Price of Part I & II : ₹ 1050.00

## OUR ADDRESSES IN INDIA

- |  |  |
|--|--|
| <b>□ New Delhi:</b> MBD House, Gulab Bhawan, 6, Bahadur Shah Zafar Marg  | Ph. 23317931, 23318301                 |
| <b>□ Mumbai:</b> A-683, T.T.C. Industrial Area, M.I.D.C. Off. Thane-Belapur Road, Navi Mumbai                            | Ph. 32996410, 27780821, 8691053365     |
| <b>□ Chennai:</b> No. 26 B/2 SIDCO Estate, North Phase, Pataravakkam, Ambattur Industrial Estate, Ambattur               | Ph. 26359376, 26242350                 |
| <b>□ Chennai:</b> Plot No. 3018, Old Y Block, 3rd Street, 12th Main Road, Anna Nagar West                                | Ph. 23741471                           |
| <b>□ Kolkata:</b> Satyam Building, 46-D, Rafi Ahmed Kidwai Marg  | Ph. 22296863, 22161670                 |
| <b>□ Jalandhar City:</b> MBD House, Railway Road   | Ph. 2458388, 2459046, 2455663          |
| <b>□ Bengaluru:</b> 124/31, 1st Main, Industrial Town (Near Chowdeshwari Kalyan Mantap), West of Chord Road, Rajajinagar | Ph. 23103329, 23104667                 |
| <b>□ Hyderabad:</b> 3-4-492, Varun Towers, Barkatpura  | Ph. 27564788, 9985820001               |
| <b>□ Ernakulam:</b> Surabhi Building, South Janatha Road, Palarivattom   | Ph. 2338107, 2347371                   |
| <b>□ Pune:</b> Survey No. 44, Behind Matoshree Garden, Kondhwa-Khadi Machine, Pisoli Road, at Post-Pisoli                | Ph. 65271413, 65275071                 |
| <b>□ Nagpur:</b> Near N.I.T. Swimming Pool, North Ambazari Road, Ambazari Layout   | Ph. 2248104, 2248106, 2248649, 2245648 |
| <b>□ Ahmedabad:</b> Godown No. 10, Vedant Prabha Estate, Opp. ONGC Pumping Station, Sarkhej Sanand Road, Sarkhej         | Ph. 26890336, 32986505                 |
| <b>□ Cuttack:</b> Badambadi, Link Road   | Ph. 2367277, 2367279, 2313013          |
| <b>□ Guwahati:</b> Chancellor Commercial, Hem Baruah Road, Paan Bazar  | Ph. 2131476, 8822857385                |
| <b>□ Lucknow:</b> 173/15, Dr. B. N. Verma Road, Old 30 Kutchery Road   | Ph. 4010992, 4010993                   |
| <b>□ Patna:</b> 1st Floor, Annapurna Complex, Naya Tola  | Ph. 2672732, 2686994, 2662472          |
| <b>□ Bhopal:</b> Plot No. 137, 138, 139, Sector-I, Special Industrial Area, Govindpura                                   | Ph. 2581540, 2601535                   |
| <b>□ Jabalpur:</b> 840, Palash Chamber, Malviya Chowk  | Ph. 2405854                            |
| <b>□ Goa:</b> H. No. 932, Plot No. 66, Kranti Nagar (Behind Azad Bhawan), Alto Porvorim, Bardez                          | Ph. 2413982, 2414394                   |
| <b>□ Jaipur:</b> C-66A, In front of Malpani Hospital, Road No. 1, V.K. Industrial Area, Sikar Road                       | Ph. 4050309, 4020168                   |
| <b>□ Raipur:</b> Behind Kailash Provision Store, Ravi Nagar  | Ph. 2445370, 4052529                   |
| <b>□ Karnal:</b> Plot No. 203, Sector-3, HSIDC, Near Namaste Chowk, Opp. New World                                       | Ph. 2220006, 2220009                   |
| <b>□ Shimla (H.P.):</b> C-89, Sector-I, New Shimla-9   | Ph. 2670221, 2670618                   |
| <b>□ Jammu (J&amp;K):</b> MBD Office, 48 Gurjar Colony, C/o Gurjar Desh Charitable Trust, N.H. Bye Pass Road             | Ph. 2467376, 9419104035                |
| <b>□ Ranchi (Jharkhand):</b> Shivani Complex, 2nd Floor, Jyoti Sangam Lane, Upper Bazar                                  | Ph. 9431257111                         |
| <b>□ Sahibabad (U.P.):</b> B-9 & 10, Site IV, Industrial Area  | Ph. 3100045, 2896939                   |
| <b>□ Dehradun (Uttarakhand):</b> Plot No. 37, Bhagirathipuram Niranjanpur, GMS Road                                      | Ph. 2520360, 2107214                   |
- DELHI LOCAL OFFICES:**
- |   |                        |
|---|------------------------|
| <b>□ Delhi (Shakarpur):</b> MB 161, Street No. 4                    | Ph. 22546557, 22518122 |
| <b>□ Delhi (Daryaganj):</b> MBD House, 4587/15, Opp. Times of India | Ph. 23245676           |
| <b>□ Delhi (Patparganj):</b> Plot No. 225, Industrial Area          | Ph. 22149691, 22147073 |

### MODERN'S OUTSTANDING TEXTBOOKS FOR CLASS XI

- Modern's** abc of Physics
- Modern's** abc of Practical Physics
- Modern's** abc of Physics Lab Manual
- Modern's** abc + of Chemistry
- Modern's** abc of Practical Chemistry
- Modern's** abc of Chemistry Lab Manual
- Modern's** abc + of Biology
- Modern's** abc of Practical Biology
- Modern's** abc of Biology Lab Manual
- Modern's** abc + of Mathematics
- Solutions of **Modern's** abc of Mathematics
- Modern's** abc of Computer Science C++

### FOR JEE-Main, JEE-Advance & Medical (AIPMT)

- Modern's** abc of Objective Physics
- Modern's** abc of Objective Chemistry
- Modern's** abc of Objective Mathematics
- Modern's** abc of Objective Biology
- Modern's** abc of Crash Course Physics (JEE-Main)
- Modern's** abc of Crash Course Chemistry (JEE-Main)
- Modern's** abc of Crash Course Mathematics (JEE-Main)

#### KIND ATTENTION : DEAR CUSTOMER

To avoid duplicacy, a HOLOGRAM has been pasted on the cover of this book. If this hologram is missing, please do not purchase the book as it is likely to be a duplicate book. Any such duplicacy if noted may please be intimated to the Publishers.

We are committed to serve students with best of our knowledge and resources. We have taken utmost care and paid much attention while editing and printing this book but we would beg to state that Authors and Publishers should not be held responsible for unintentional mistakes that might have crept in. However, errors brought to our notice shall be gratefully acknowledged and attended to.

*© All rights reserved. No part of this publication may be reproduced, stored in a retrieval system, or transmitted in any form or by any means, electronic, mechanical, photocopying, recording, or otherwise without the prior written permission of the publisher. Any breach will entail legal action and prosecution without further notice.*

**Published by :** MODERN PUBLISHERS  
MBD House, Railway Road, Jalandhar City.

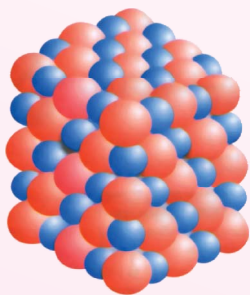
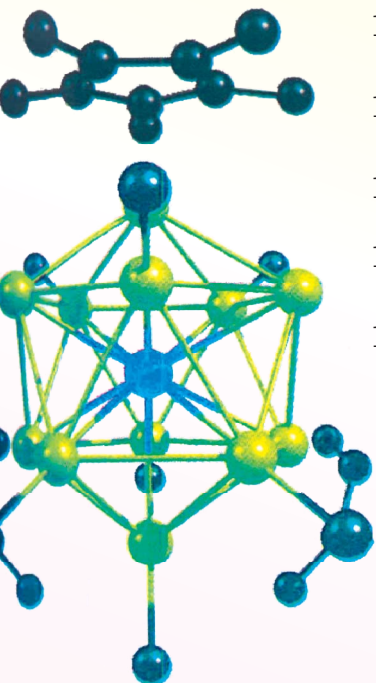
**Printed at :** HOLY FAITH INTERNATIONAL (P) LTD.  
B-9 & 10, Site IV, Industrial Area, Sahibabad (U.P.)

E info@mbdgroup.com  
W www.mbdgroup.com  
Toll Free No. : 1800 200 2233

# CONTENTS

## PART-II

8.	Redox Reactions	879 – 956
9.	Hydrogen	957 – 1014
10.	s-Block Elements (Alkali and Alkaline Earth Metals)	1015 – 1074
11.	Some p-Block Elements	1075 – 1138
12.	Organic Chemistry: Basic Principles and Techniques	1139 – 1300
13.	Hydrocarbons	1301 – 1452
14.	Environmental Pollution	1453 – 1497
	● Value Based Questions	1498 – 1502
	● MOCK TEST – 2	1503 – 1504
	● MOCK TEST – 3 (According to Completely CBSE Pattern)	1505 – 1506
	● Solution File (Unit Practice Test and Mock Test)	1507 – 1511
	● Appendices	1513 – 1525
	● Logarithm and Anti-logarithm Tables	(i) – (v)



blank

# UNIT 8

# REDOX REACTIONS

The oxidation and reduction reactions or **redox reactions** are very common in our daily life. A large number of phenomena both physical as well as chemical and biological are concerned with redox reactions. These reactions are finding wide uses in biological, pharmaceutical, industrial, metallurgical and agricultural areas. Some of the common examples are generation of electricity in dry and wet batteries, burning of different types of fuels for obtaining energy for domestic, transport or industrial purposes, digestion of food in animals, corrosion of metals and many industrial processes such as production of *caustic soda*, *potassium permanganate*, *extraction of metals like sodium, iron and aluminium*, etc. In fact, **redox reactions keep us alive**. They help to capture the energy of the sun by the photosynthesis process by plants which is used by us. Recently, environmental issues such as **hydrogen economy** i.e., use of *liquid hydrogen as fuel and development of ozone hole* are considered as redox phenomena.

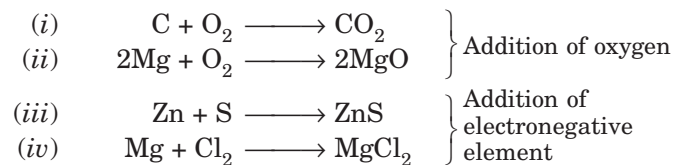
In the present unit, we shall classify the reactions into two classes, oxidation and reduction reactions and discuss electrochemical cells.

## OXIDATION AND REDUCTION REACTIONS

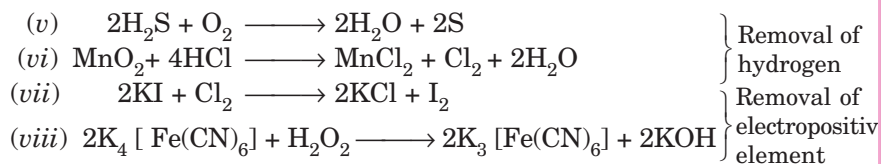
### Classical Concepts

According to classical (earlier) views, **oxidation** is defined as **the addition of oxygen or any other electronegative element or removal of hydrogen or any other electropositive element**.

For example, the oxidation reactions are :



The above reactions involve addition of oxygen or electronegative element (S and Cl<sub>2</sub>) and therefore, are oxidation reactions.



### IN THIS UNIT.....

◆ Understanding Text	879
◆ Conceptual Questions	920

<b>CHAPTER SUMMARY</b>	922
------------------------	-----

### NCERT FILE

◆ NCERT Textbook Exercises with Solutions	923
◆ NCERT Exemplar Problems (Subjective) with Answers & Solutions	931
◆ Quick Memory Test with Answers	934
◆ HOTS & Brain Twisting Qs with Answers	935

<b>REVISION EXERCISES</b>	937
---------------------------	-----

◆ Hints & Answers for Revision Exercises	939
--	-----

### COMPETITION FILE

◆ MCQs	940
◆ Competitive Examination Qs	
> AIPMT & Other State Boards for Medical Entrance	942
> JEE (Main) & Other State Boards for Engineering Entrance	943
> JEE (Advance) for IIT Entrance	946
◆ NCERT Exemplar Problems (Objective)	949
◆ Hints & Explanations for Difficult Questions	951

<b>UNIT PRACTICE TEST</b>	956
---------------------------	-----

These reactions involve removal of hydrogen or some electropositive element (potassium) and, therefore, are oxidation reactions.

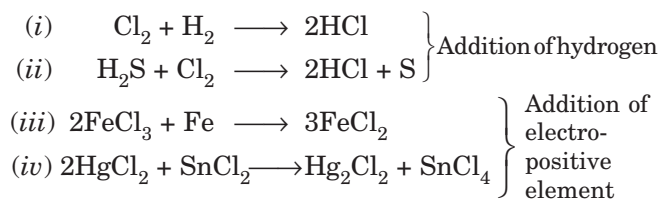
**A substance which gives oxygen or removes hydrogen is called an oxidising agent or oxidant.**

For example, in examples (i) and (ii),  $O_2$  is an oxidising agent because it causes the oxidation of C and Mg. Similarly, in reactions (iii) and (iv) S and  $Cl_2$  act as oxidising agents because they add electronegative element to Zn and Mg. Similarly,  $Cl_2$  and  $MnO_2$  act as oxidising agents in reactions (v) and (vi) because they remove hydrogen while,  $Cl_2$  and  $H_2O_2$  act as oxidising agents in reactions (vii) and (viii) because they remove electropositive element (potassium).

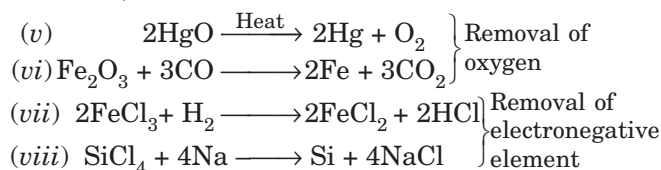
Similarly, according to classical concept, **reduction** is defined as

**the addition of hydrogen or any other electropositive element or removal of oxygen or any other electronegative element.**

These reactions are illustrated by the following examples :



In the above examples, reactions (i) and (ii) involve addition of hydrogen and reactions (iii) and (iv) involve addition of electropositive elements (Fe and Hg) and therefore, are reduction reactions.



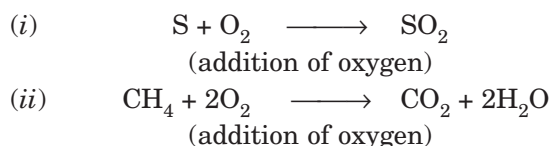
In the above examples, reactions involve removal of oxygen or electronegative element (chlorine) and therefore, are reduction reactions.

**A substance which provides hydrogen or removes oxygen is called a reducing agent or reductant.**

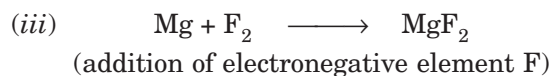
For example,  $H_2$ ,  $H_2S$ , Fe, Cu, C, CO,  $HgCl_2$ , and Na are reducing agents in the above reactions because they bring about reduction reactions.

Let us learn some more examples of oxidation and reduction.

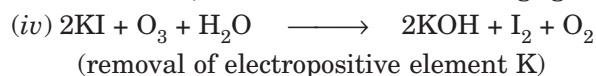
#### Oxidation reactions



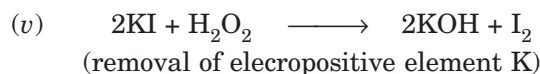
In these reactions (i) and (ii)  $O_2$  acts as an oxidising agent.



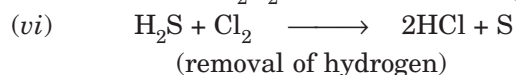
In this reaction, fluorine acts as an oxidising agent.



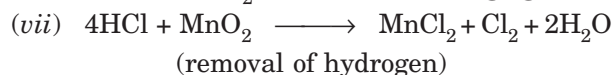
In this reaction, ozone is an oxidising agent.



In this reaction,  $H_2O_2$  acts as an oxidising agent.

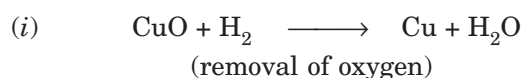


In this reaction,  $Cl_2$  acts as an oxidising agent.

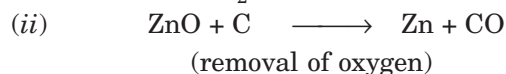


In this reaction,  $MnO_2$  acts as an oxidising agent.

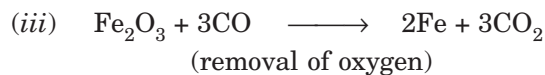
#### Reduction reactions



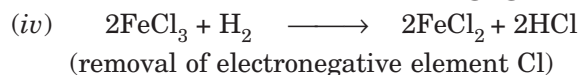
In this reaction,  $H_2$  acts as a reducing agent.



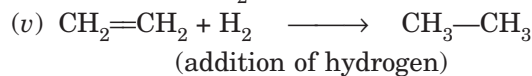
In this reaction, C acts as a reducing agent.



In this reaction, CO acts as a reducing agent.



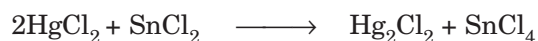
In this reaction,  $H_2$  acts as a reducing agent.



In this reaction,  $H_2$  acts as a reducing agent.

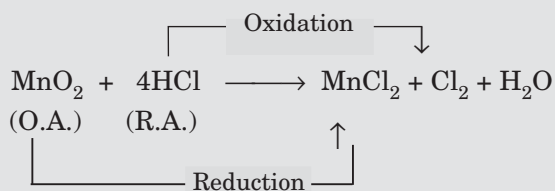
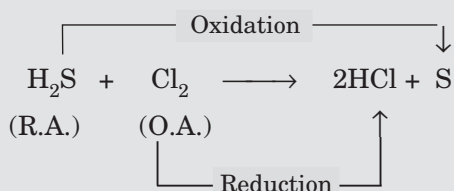
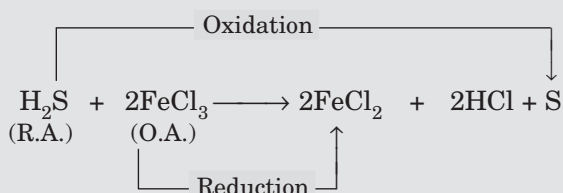
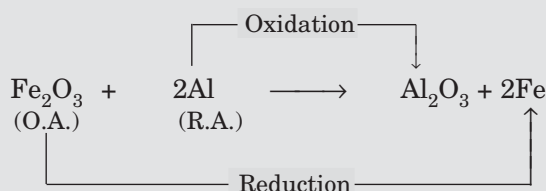
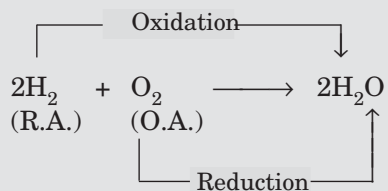
#### Oxidation-Reduction Reactions are Complementary : go side by side

Oxidation-reduction reactions are complementary reactions. Whenever any substance is oxidised, another substance is always reduced at the same time and vice versa. In other words, the oxidation reduction reactions always occur simultaneously i.e., they always go hand in hand or side by side. For example, consider the reaction :



In this reaction, stannous chloride ( $SnCl_2$ ) has been oxidised to stannic chloride ( $SnCl_4$ ) because of addition of electronegative element chlorine to it. At the same time  $HgCl_2$  is reduced to  $Hg_2Cl_2$  because of the addition of mercury to mercuric chloride. Thus,  $SnCl_2$  is oxidised by  $HgCl_2$  and  $HgCl_2$  is reduced by  $SnCl_2$ . Hence,  $HgCl_2$  acts as an oxidising agent and  $SnCl_2$  acts as a reducing agent in this reaction.

Thus, oxidation and reduction reactions are collectively called **redox reactions**. Some more examples of redox reactions are given below :



Here O.A. and R.A. stand for oxidising agent and reducing agent respectively.

### Electronic Concept of Oxidation and Reduction

Oxidation and reduction reactions are now used in a much broader and general sense. These are defined in terms of electron transfer between the reactants, known as electronic concept of oxidation and reduction.

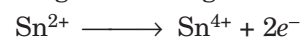
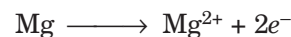
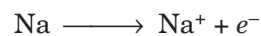
**Oxidation.** According to electronic concept, **oxidation is**

**a process which involves loss of electrons by an atom or group of atoms.**

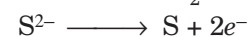
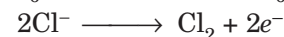
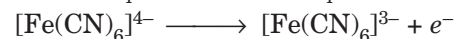
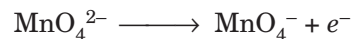
As a result of loss of electrons, there is increase in positive charge or decrease in negative charge of the atom or ion undergoing oxidation.

For example,

(i) *Loss of electrons resulting in increase in positive charge.*



(ii) *Loss of electrons resulting in decrease of negative charge.*

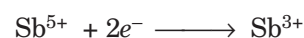
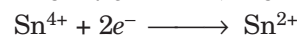


**Reduction .** According to electronic concept, **reduction is**

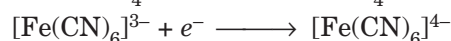
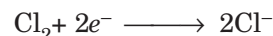
**a process which involves gain of electrons by an atom or group of atoms.**

Due to gain of electrons there is increase of negative charge or decrease of positive charge of an atom or ion undergoing reduction. For example,

(i) *Gain of electrons resulting in decrease of positive charge.*



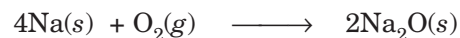
(ii) *Gain of electrons resulting in increase of negative charge.*



### OXIDATION-REDUCTION REACTIONS AS ELECTRONS TRANSFER REACTIONS

As already learnt, oxidation process involves loss of electrons whereas reduction process involves gain of electrons. During transfer of electrons in a reaction, the species which loses electrons is said to be *oxidised* and the species which gains electrons is said to be *reduced*. Since there cannot be a net loss or gain of electrons in a chemical reaction, this means that all chemical changes involving the loss and gain of electrons occur simultaneously *i.e.* one species loses electrons in oxidation while the other gains electrons during the reduction. Thus, both these changes take place side by side *i.e.*, **oxidation takes place at the cost of reduction and vice versa.**

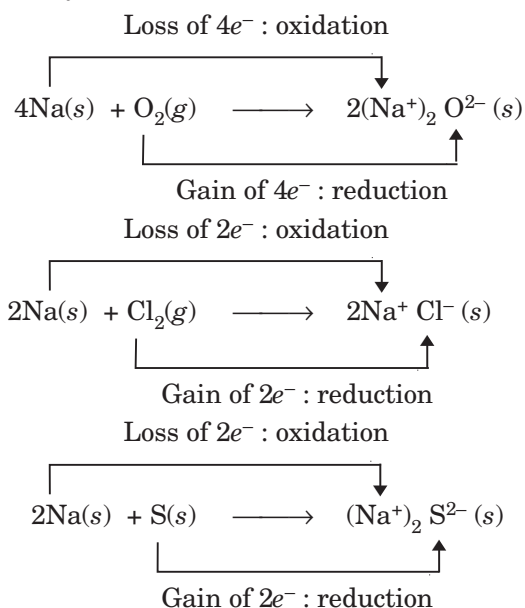
In other words, oxidation and reduction reactions go hand in hand. This may be illustrated by the following reactions:



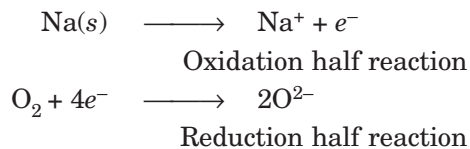
All these reactions are redox reactions because in each of these reactions sodium is oxidised due to addition of either oxygen or more electronegative elements such as chlorine or sulphur to sodium. Simultaneously oxygen, chlorine and sulphur are reduced because electropositive element sodium has been added to each of these. We have also learnt in chemical bonding that sodium oxide, sodium chloride and sodium sulphide are ionic compounds and may be written as :

Sodium oxide :  $(\text{Na}^+)_2\text{O}^{2-}$ , Sodium chloride :  $\text{Na}^+ \text{Cl}^-$ , Sodium sulphide :  $(\text{Na}^+)_2 \text{S}^{2-}$ .

These ionic charges suggest that the above redox reactions may also be written as :

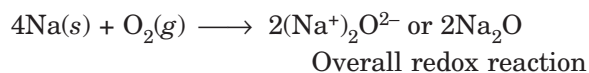
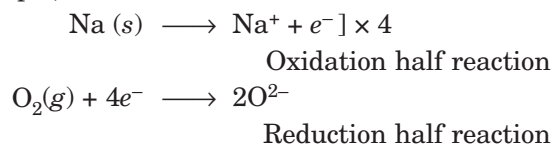


For the sake of convenience, each of the above redox reaction may be considered as the sum of two half reactions; one *involving loss of electrons i.e. oxidation* and the other *involving gain of electrons i.e. reduction* as

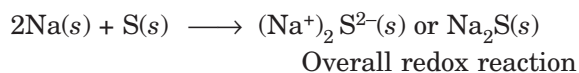
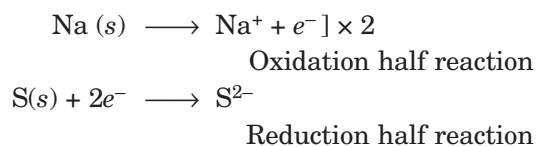
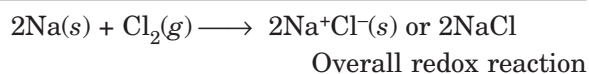
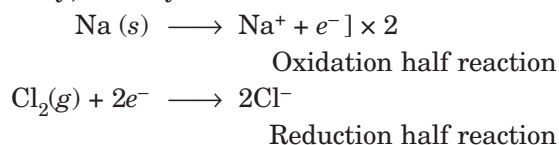


These two half reactions are called **oxidation half reaction** and **reduction half reaction** respectively. The sum of the half reactions give the overall reaction. It may be noted that the overall reaction is so written that the electrons are balanced *i.e., the number of electrons lost and gained become equal*. This means that the two half reactions are simply added if the number of electrons lost during oxidation half reaction is equal to the number of electrons gained during reduction half reaction. But if the number of electrons lost during oxidation half reaction are different from the number

of electrons gained during reduction half reaction, the two half reactions are multiplied by suitable integers so that when the two half equations are added, the electrons get cancelled out in the final redox reaction. For example,



Similarly, we may write other reactions as :



In all the above reactions, electrons are transferred from sodium to  $\text{O}_2$ ,  $\text{Cl}_2$  and S. Thus, **oxidation-reduction i.e., redox reactions are regarded as electron transfer reactions**. Moreover, we observe that  $\text{O}_2$ ,  $\text{Cl}_2$  and S have gained electrons only if there is another substance which can donate electrons. This means that **reduction can occur only if oxidation is occurring side by side and vice versa**. In these reactions, sodium loses electrons and has been oxidised to sodium ion,  $\text{Na}^+$ . This is brought about by the acceptance of electrons by oxygen, chlorine or sulphur, which get reduced. Thus, oxygen, chlorine and sulphur bring about the oxidation and are called **oxidising agents** or **oxidants**.

Thus, according to electronic concept,

**oxidising agent is a substance which can accept one or more electrons.**

Similarly, in the above example, oxygen, chlorine and sulphur are reduced to oxide ion ( $\text{O}^{2-}$ ), chloride ion ( $\text{Cl}^-$ ) and sulphide ion ( $\text{S}^{2-}$ ) respectively. This has been possible by furnishing electrons by sodium. Thus, sodium is a **reducing agent** or **reductant**. Thus,

**reducing agent is a substance which can give one or more electrons.**

**TO SUM UP**

**Oxidation** is a process in which one or more electrons are lost.

**Reduction** is a process in which one or more electrons are gained.

**Oxidising agent** or **oxidant** is a substance which can accept one or more electrons.

**Reducing agent** or **reductant** is a substance which can give one or more electrons.

In a redox reaction, **oxidising agent is reduced** by accepting electrons and **reducing agent is oxidised** by losing electrons.

**REDOX REACTIONS IN AQUEOUS SOLUTIONS**

Redox reactions are very common in our daily life. Whenever you use a battery, a redox reaction occurs. A simple example of an oxidation-reduction reaction is the reaction between zinc metal and copper (II) salt in aqueous solution :



Let us study this reaction.

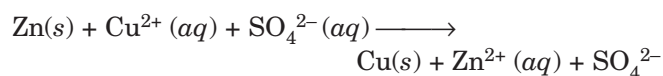
Take a strip of zinc metal and clean it with a sand paper. Place this strip in a solution of copper sulphate in a beaker [Fig. 1]. We observe that a spontaneous reaction takes place and the following observations are made :

- (i) Zinc metal starts dissolving and it loses its weight gradually.
- (ii) The copper metal starts getting either deposited on the zinc plate or settles down at the bottom of the beaker.
- (iii) The blue colour of the solution gradually fades away.
- (iv) The reaction is exothermic and proceeds with the liberation of heat energy.
- (v) The solution remains electrically neutral.

Let us try to explain the above reaction. In aqueous solution,  $\text{CuSO}_4$  dissociates into  $\text{Cu}^{2+}$  and  $\text{SO}_4^{2-}$  ions. When a strip of zinc is dipped in  $\text{CuSO}_4$  solution, zinc

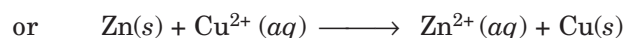
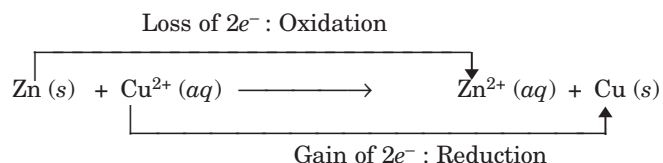
metal loses electrons and gets oxidised whereas  $\text{Cu}^{2+}$  ions accept electrons and get reduced. Since zinc is getting oxidised and converted to  $\text{Zn}^{2+}$  ions which go into the solution, the weight of zinc plate gradually decreases. On the other hand,  $\text{Cu}^{2+}$  ions present in the solution are accepting the electrons given by zinc and are getting reduced to copper. As a result, copper metal either gets deposited on the zinc plate or gets precipitated at the bottom of the beaker [Fig. 1(b)]. Since  $\text{Cu}^{2+}$  ions from the solution are changing to  $\text{Cu}(s)$ , the blue colour of the solution (which is due to  $\text{Cu}^{2+}$  ions) slowly fades.

The reaction may be written as :



Formation of  $\text{Zn}^{2+}$  ions among the products can easily be tested when the blue colour of the solution due to  $\text{Cu}^{2+}$  ions has disappeared. If we pass, hydrogen sulphide ( $\text{H}_2\text{S}$ ) gas through the solution, formation of white precipitate of zinc sulphide ( $\text{ZnS}$ ) on making the solution alkaline with ammonia indicates the presence of  $\text{Zn}^{2+}$  ions in the solution.

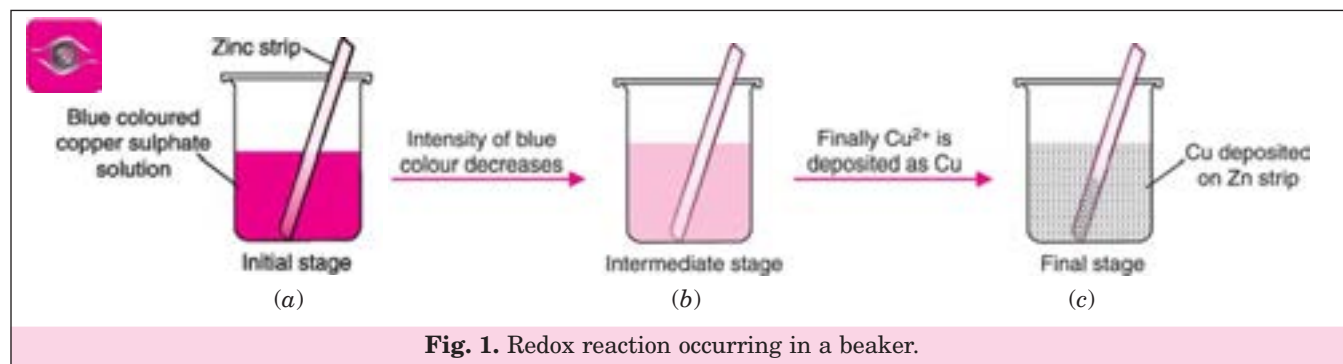
In this reaction, sulphate ions,  $\text{SO}_4^{2-}$  do not participate and we may write the reaction as :



The reaction is commonly described as zinc having displaced  $\text{Cu}^{2+}$  ions from the  $\text{CuSO}_4$  solution.

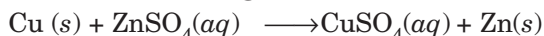
In this case **zinc acts as a reducing agent** while  **$\text{Cu}^{2+}$  ions act as oxidising agent**.

In this redox reaction, the transference of electrons from Zn metal to  $\text{Cu}^{2+}$  ions is occurring directly because the two are in direct contact. The energy released appears as heat and therefore, it is an exothermic reaction.



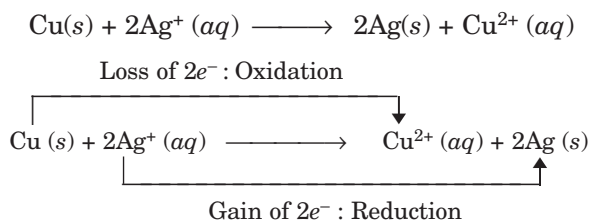
**Fig. 1.** Redox reaction occurring in a beaker.

Let us study the state of equilibrium for the above reaction. For this purpose, place a rod of copper in zinc sulphate ( $\text{ZnSO}_4$ ) solution. You will notice that no visible reaction occurs. If we pass  $\text{H}_2\text{S}$  gas through the solution, there will be no black precipitate of copper sulphide indicating that no reaction has occurred. In other words, the following reaction :



**does not occur.** Thus, we conclude that the state of equilibrium for the above reaction between zinc and copper sulphate greatly favours the products over the reactants.

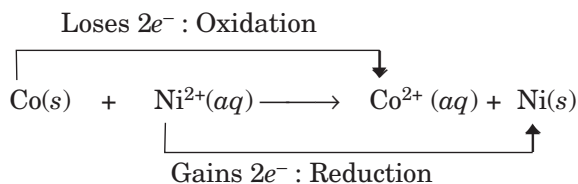
Similarly, when we dip a copper strip in a silver nitrate solution, copper gets oxidised and goes into the solution whereas  $\text{Ag}^+$  ions accept electrons and get reduced (Fig. 2). The reaction may be written as :



Thus, copper is oxidised to  $\text{Cu}^{2+}$  and  $\text{Ag}^+$  is reduced to  $\text{Ag(s)}$ . Therefore, copper acts as an oxidising agent or oxidant while silver acts as a reducing agent or reductant.

The equilibrium greatly favours the products  $\text{Cu}^{2+}(\text{aq})$  and  $\text{Ag(s)}$ .

For the purpose of comparison, let us study the reaction between cobalt metal and nickel sulphate solution. For this, place a rod of cobalt metal in nickel sulphate solution. The following reaction occurs :



At equilibrium, chemical tests show that both  $\text{Ni}^{2+}(\text{aq})$  and  $\text{Co}^{2+}(\text{aq})$  are present in moderate concentration. This means that in this case, neither the reactants [ $\text{Co(s)}$  and  $\text{Ni}^{2+}(\text{aq})$ ] nor the products [ $\text{Co}^{2+}(\text{aq})$  and  $\text{Ni(s)}$ ] are greatly favoured.

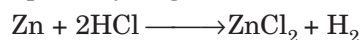
Thus, we observe that there is a competition of loss or gain of electrons by metals. This is similar to the competition between various acids to lose a proton ( $\text{H}^+$ ) in water. This similarity suggests that we can develop a table in which the metals and their ions are listed on the basis of their tendency to lose or gain electrons similar to the table in which various acids are arranged according to their acid strengths. For example, we have learnt in above experiments that zinc loses electrons to copper and copper releases electrons to silver. Therefore, the electron releasing tendency of these three metals is in the order :



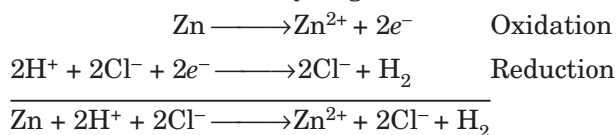
Similarly, by comparing the relative tendencies of other metals, we can build a table to arrange the metals in their decreasing or increasing order of losing electrons. This series is called **activity series** or **electrochemical series** and is discussed later.

#### Some other examples of common redox reactions

When a metal reacts with acids, hydrogen is liberated. For example metals like Zn, Mg, Ni etc. can easily displace hydrogen from dil. HCl.



This reaction occurs as redox reaction in which metal is oxidised and hydrogen is reduced as :



In electrolytic process also, redox reaction takes place. For example, during electrolysis of aqueous solution of copper chloride, oxidation and reduction occur as :

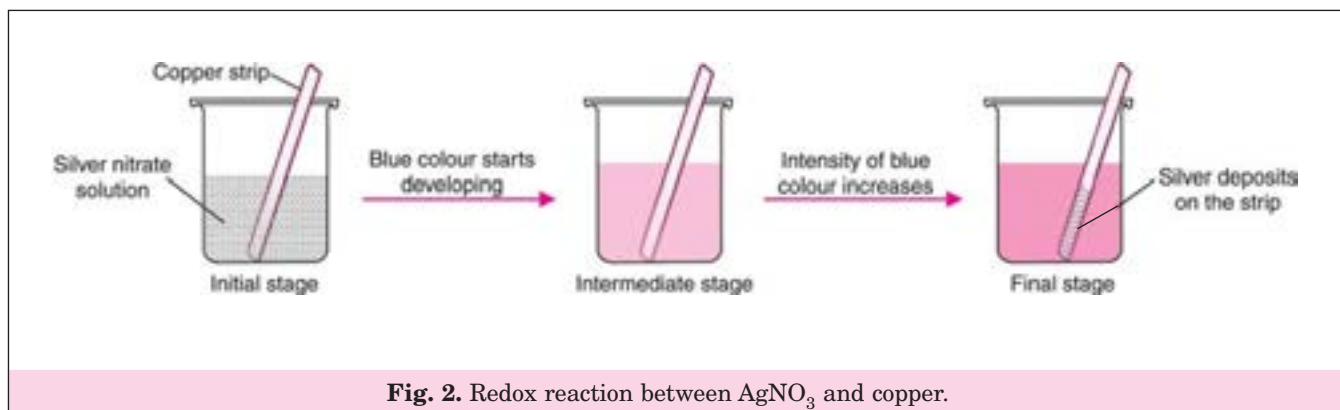
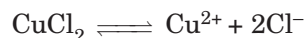
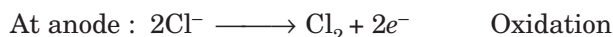
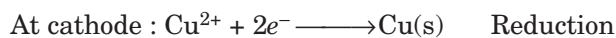
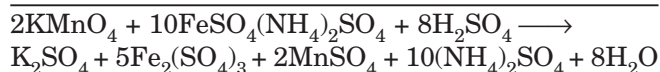
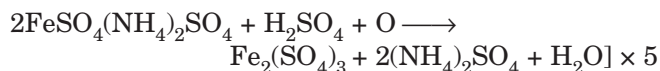


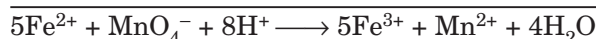
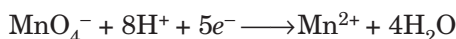
Fig. 2. Redox reaction between  $\text{AgNO}_3$  and copper.



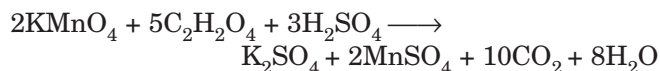
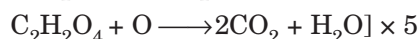
Potassium permanganate is a strong oxidising agent in the presence of dil.  $\text{H}_2\text{SO}_4$ . In the presence of dil.  $\text{H}_2\text{SO}_4$ , it gives nascent oxygen which oxidises ferrous ammonium sulphate  $[\text{FeSO}_4(\text{NH}_4)_2\text{SO}_4]$  to ferric sulphate as :



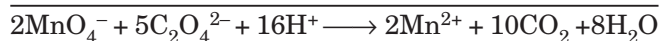
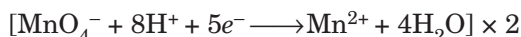
The ionic equation for the reaction may be written as :



Similarly, acidified  $\text{KMnO}_4$  oxidises oxalic acid ( $\text{C}_2\text{H}_2\text{O}_4$ ) solution. In this reaction, acidified  $\text{KMnO}_4$  liberates nascent oxygen which oxidises oxalic acid to carbon dioxide as:



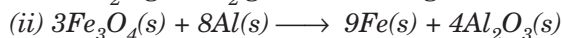
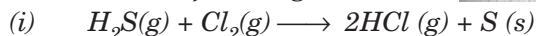
The ionic equation for the reaction may be written as :



### SOLVED EXAMPLES

#### Example 1.

Identify the species undergoing oxidation and reduction in the following reaction : **N.C.E.R.T.**

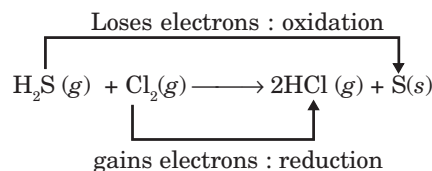


**Solution:**

(i)  $\text{H}_2\text{S}$  is oxidised because a more electronegative element, chlorine is added to hydrogen (or a more electropositive element hydrogen has been removed from  $\text{H}_2\text{S}$ ).

Chlorine has been reduced because hydrogen is added to it.

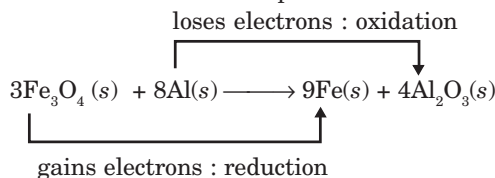
In terms of electronic concept, we can say,



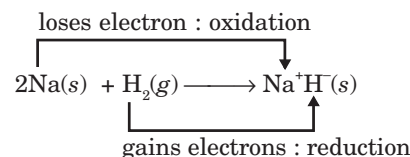
(ii) Aluminium is oxidised because oxygen is added to it.

Ferrous ferric oxide ( $\text{Fe}_3\text{O}_4$ ) is reduced because oxygen has been removed from it.

In terms of electronic concept :



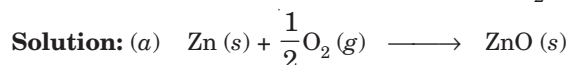
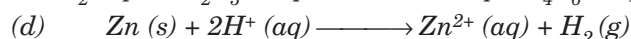
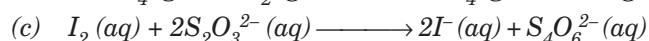
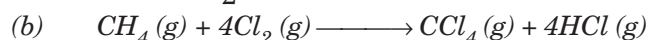
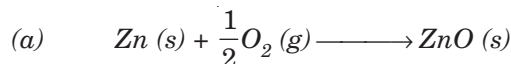
(iii) In this case, sodium has been oxidised and hydrogen has been reduced. This becomes more clear in terms of electronic concept :



$\text{NaH}$  is ionic compound and exists as  $\text{Na}^+\text{H}^-$

#### Example 2.

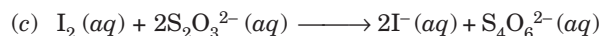
Identify the oxidant and reductant in the following reactions :



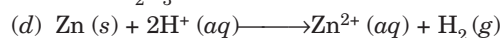
In this reaction, zinc donates electrons to O to give zinc ions and oxide ions. Thus, Zn acts as reductant while oxygen acts as **oxidant**.



$\text{CH}_4$  is oxidised and acts as **reductant** while  $\text{Cl}_2$  is reduced and acts as **oxidant**.



$\text{I}_2$  gains electrons and is reduced. Therefore, it acts as **oxidant**.  $\text{S}_2\text{O}_3^{2-}$  acts as **reductant**.



Zinc loses electrons and gets oxidised. Therefore, it acts as **reductant**. Hydrogen gets reduced and acts as an **oxidant**.

#### Example 3.

Justify that the reaction :

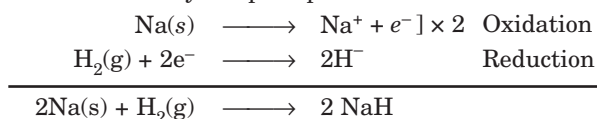


**N.C.E.R.T.**

**Solution:** Sodium hydride is an ionic compound and it exists as  $\text{Na}^+\text{H}^-$ .



The reaction may be split up into two half reactions:



In this reaction Na is oxidised to  $\text{Na}^+$  and hydrogen is reduced from  $\text{H}_2$  to  $2\text{H}^-$  ion. Therefore, it is a redox reaction.

## OXIDATION NUMBER

The concept of electron transfer can easily explain the redox reactions in case of ionic substances. However, we cannot easily explain the redox changes in terms of electron transfer in case of covalent compounds. Now-a-days, to understand oxidation-reduction reactions, systematically, scientists assign an *oxidation number* or *oxidation state* to each chemical species (ionic or covalent) according to a set of rules. The concept of oxidation number has been very useful in balancing of redox reactions. The **oxidation number** is defined as

*the charge which an atom of the element has in its ion or appears to have when present in the combined state with other atoms.*

**Oxidation number** also gives the charge which an atom appears to have when all other atoms are removed from it as ions. During the removal of atoms, the electrons on the atoms are counted according to the following two arbitrary rules :

(i) *Electrons shared between two like atoms are divided equally between the sharing atoms.* In case of hydrogen molecule, for example, the electron pair is equally shared between the two hydrogen atoms. Therefore, one electron is counted with each hydrogen atom as shown below :



Now, each hydrogen atom is having one electron while its nucleus has one proton. Therefore, there is no net charge on each atom of hydrogen. In other words, oxidation number of hydrogen in hydrogen molecule is zero.

(ii) *Electrons shared between two unlike atoms are counted with more electronegative atom.* For example, in hydrogen chloride molecule, chlorine is more electronegative than hydrogen. Therefore, the shared pair is counted towards chlorine atom as shown below :



As a result of this, chlorine gets one extra electron and acquires a unit negative charge. Hence, oxidation number of chlorine is  $-1$ . On the other hand, hydrogen

atom without electron has a unit positive charge. Hence, oxidation number of hydrogen in hydrogen chloride is  $+1$ .

The oxidation number is a **fictitious charge** in case of covalent species. The oxidation number can have positive, zero or negative values depending upon their state of combination.

## Rules for the Determination of Oxidation Number of an Atom

The following general rules are used for the calculation of oxidation number of an atom in a molecule :

1. *The oxidation number of an element in the free or elementary state or in any of its allotropic forms is always zero.* For example, oxidation numbers of helium in He, hydrogen in  $\text{H}_2$ , oxygen in  $\text{O}_2$  or  $\text{O}_3$ , iron in Fe, carbon in diamond, helium atom in He, bromine in  $\text{Br}_2$ , phosphorus in  $\text{P}_4$ , sulphur in  $\text{S}_8$ , are zero.
2. *The oxidation number of an element in a single (monoatomic) ion is the same as the charge on the ion.* For example, oxidation number of  $\text{K}^+$  is  $+1$ , of  $\text{Ca}^{2+}$  is  $+2$ , of  $\text{Al}^{3+}$  is  $+3$ . Similarly, the oxidation numbers of  $\text{Cl}^-$ ,  $\text{SO}_4^{2-}$  and  $\text{PO}_4^{3-}$  are  $-1$ ,  $-2$  and  $-3$  respectively.
3. *In binary compounds of metal and non-metal, the oxidation number of metal is always positive while that of the non-metal is negative.* For example, in NaCl, the oxidation number of sodium is  $+1$  and that of chlorine is  $-1$ .
4. *In compounds formed by the combination of non-metallic atoms, the atom with higher electronegativity is given negative oxidation number.* For example, in HCl, the oxidation number of chlorine in HCl is  $-1$  because of its high electronegativity.

Similarly, N is given an oxidation number of  $-3$ , when it is bonded to less electronegative atom as in  $\text{NH}_3$  and  $\text{NI}_3$ . However, it is given an oxidation number of  $+3$  when it is bonded to more electronegative atom as in  $\text{NCl}_3$ .

**Note.** Rules (3) and (4) are based on the fact that atoms having high electronegativity have a tendency to form negative ions (anions) while atoms having low electronegativity have great tendency to form positive ions (cations).

5. *In all compounds of hydrogen, the oxidation number of hydrogen is  $+1$  except in hydrides of active metals such as LiH, NaH, KH,  $\text{MgH}_2$ ,  $\text{CaH}_2$ , etc., where hydrogen has the oxidation number of  $-1$ .*
6. *The oxidation number of oxygen is  $-2$  in most of the compounds. However, there are two exceptions. The first exception is **peroxides***

and **superoxides** in which oxygen atoms are directly linked to each other. In peroxides (e.g.,  $H_2O_2$ ,  $Na_2O_2$ ), each oxygen atom is assigned an oxidation number of  $-1$  and in superoxides (e.g.,  $KO_2$ ,  $RbO_2$ ) each oxygen atom is assigned an oxidation number of  $-1/2$ . The second exception is found in compounds in which **oxygen is bonded to fluorine**. For example, in  $OF_2$  (oxygen difluoride) the oxidation number of oxygen is  $+2$  and in  $O_2F_2$  (dioxygen difluoride) the oxidation number of each oxygen is  $+1$ . This is because of the fact that fluorine being most electronegative element known has always an oxidation number of  $-1$ .

7. The most electronegative element, fluorine has oxidation number  $-1$ . For other halogens, the oxidation number is generally  $-1$ , but there are exceptions when these are bonded to a more electronegative halogen atom or oxygen. For example, in HI, the oxidation number of I is  $-1$  but in  $IF_5$ , it is  $+5$  and in  $IF_7$ , it is  $+7$ .

8. For neutral molecule, the sum of the oxidation numbers of all the atoms is equal to zero. For example, in  $NH_3$  the sum of the oxidation numbers of nitrogen atom and 3 hydrogen atoms is equal to zero.

For a polyatomic ion, the sum of the oxidation numbers of all the atoms is equal to charge on the ion. For example, in  $SO_4^{2-}$  ion, the sum of the oxidation numbers of sulphur atom and 4 oxygen atoms must be equal to  $-2$ .

It may be noted that oxidation number is also frequently called as **oxidation state**. For example, in  $H_2O$ , the oxidation state of hydrogen is  $+1$  and the oxidation state of oxygen is  $-2$ . Similarly, in  $CO_2$ , the oxidation number of carbon is  $+4$ , which is also its oxidation state and the oxidation number of oxygen is  $-2$ , which is also its oxidation state. This means that *oxidation number gives the oxidation state of an element in a compound*.

**Procedure for calculation of oxidation numbers.** By applying the above rules, we can calculate the oxidation numbers of elements in the molecules/ions by the following steps :

- Write down the formula of the given molecule/ion leaving some space between the atoms.
- Write oxidation number on the top of each atom. In case of the atom whose oxidation number has to be calculated write  $x$ .
- Beneath the formula, write down the total oxidation numbers of each element. For this purpose, multiply the oxidation numbers of each atom with the number of atoms of that kind in the molecule/ion. Write the product in a bracket.
- Equate the sum of the oxidation numbers to zero for neutral molecule and equal to charge on the ion.
- Solve for the value of  $x$ .

## SOLVED EXAMPLES

**Example 4.**

Calculate the oxidation number of sulphur in the following molecules/ions :

- (a)  $H_2S$                       (b)  $H_2SO_3$                       (c)  $SO_4^{2-}$   
 (d)  $Na_2S_2O_3$                       (e)  $S_2O_7^{2-}$                       (f)  $H_2SO_4$   
 (g)  $S_2O_4^{2-}$

**Solution:** (a)  $H_2S$ . The oxidation number of hydrogen is  $+1$ . Let the oxidation number of sulphur be  $x$ .

$$\begin{array}{rcl} & +1 & x \\ & H_2 & S \\ & (1 \times 2) & (x) \\ 1 \times 2 + x & = & 0 \end{array} \quad \therefore \quad x = -2$$

Oxidation number of S in  $H_2S$  is  $-2$ .

(b)  $H_2SO_3$ . Oxidation number of hydrogen is  $+1$  and that of oxygen is  $-2$ . Let the oxidation number of S be  $x$ .

$$\begin{array}{rcl} & +1 & x & -2 \\ & H_2 & S & O_3 \\ & (+1 \times 2) & x & (-2 \times 3) \end{array}$$

$$(+1 \times 2) + x + (-2 \times 3) = 0$$

$$\text{or} \quad +2 + x - 6 = 0$$

$$\text{or} \quad x = +4$$

Oxidation number of S in  $H_2SO_3$  is  $+4$ .

(c)  $SO_4^{2-}$ . The oxidation number of oxygen is  $-2$  and let the oxidation number of S be  $x$ .

$$\left[ \begin{array}{r} x \\ S \\ O_4 \end{array} \right]^{2-}$$

$$\begin{array}{rcl} & x & (-2 \times 4) \\ x + (-2 \times 4) & = & -2 \end{array} \quad \text{or} \quad \begin{array}{rcl} & x & -2 + 8 \\ & x & = +6 \end{array}$$

Oxidation number of S in  $SO_4^{2-}$  is  $+6$ .

(d)  $Na_2S_2O_3$ . The oxidation number of Na is  $+1$  and that of oxygen is  $-2$ . Let the oxidation number of S be  $x$ .

$$\begin{array}{rcl} & +1 & x & -2 \\ & Na_2 & S_2 & O_3 \\ & (+1 \times 2) & (x \times 2) & (-2 \times 3) \\ (+1 \times 2) + (x \times 2) + (-2 \times 3) & = & 0 \\ +2 + 2x - 6 & = & 0 \end{array} \quad \text{or} \quad \begin{array}{rcl} & 2x & = & 4 \\ & x & = & +2 \end{array}$$

Oxidation number of S in  $Na_2S_2O_3$  is  $+2$ .

(e)  $S_2O_7^{2-}$ . The oxidation number of oxygen is  $-2$  and let the oxidation number of sulphur be  $x$ .

$$\left[ \begin{array}{r} x & -2 \\ S_2 & O_7 \end{array} \right]^{2-}$$

$$\begin{array}{rcl} & (x \times 2) & (-2 \times 7) \\ x \times 2 + (-2 \times 7) & = & -2 \\ 2x & = & -2 + 14 \end{array}$$

$$\therefore \quad x = 6$$

Oxidation number of S in  $S_2O_7^{2-}$  is  $+6$ .

(f)  $\text{H}_2\text{SO}_4$ . The oxidation number of hydrogen is +1 and that of oxygen is -2. Let the oxidation number of S be  $x$

$$\begin{array}{ccc} +1 & x & -2 \\ \text{H}_2 & \text{S} & \text{O}_4 \\ (+1 \times 2) & x & (-2 \times 4) \\ \hline (+1 \times 2) + x + (-2 \times 4) = 0 \end{array}$$

or  $2 + x - 8 = 0$

or  $x = +6$

Oxidation number of S in  $\text{H}_2\text{SO}_4$  is +6.

(g)  $\text{S}_2\text{O}_4^{2-}$ . The oxidation number of oxygen is -2 and let the oxidation number of sulphur be  $x$ .

$$\begin{array}{cc} \left[ \begin{array}{cc} x & -2 \\ \text{S}_2 & \text{O}_4 \end{array} \right]^{2-} \\ (x \times 2) + (-2 \times 4) \\ x \times 2 + (-2 \times 4) = -2 \\ 2x - 8 = -2 \therefore x = +3 \end{array}$$

$\therefore$  Oxidation number of S in  $\text{S}_2\text{O}_4^{2-}$  is +3.

#### □ Example 5.

What is the oxidation number of the underlined atoms in each of the following molecules / ions ?

- (a)  $\text{Cl}\underline{\text{O}}_3^-$  (b)  $\underline{\text{Br}}\text{F}_3$  (c)  $\underline{\text{C}}\text{H}_4$  (d)  $\underline{\text{C}}_6\underline{\text{H}}_{12}\underline{\text{O}}_6$  (e)  $\text{Na}_2\underline{\text{B}}_4\underline{\text{O}}_7$   
(f)  $\text{Na}_4[\underline{\text{Fe}}(\underline{\text{C}}\underline{\text{N}})_6]$  (g)  $\underline{\text{N}}_2\underline{\text{H}}_4$

**Solution:** Suppose  $x$  be the oxidation number of the underlined atom :

(a)  $\text{ClO}_3^-$   $x + (-2 \times 3) = -1$

$x - 6 = -1$  or  $x = +5$

(b)  $\text{BrF}_3$   $x + (-1 \times 3) = 0$

$x - 3 = 0$  or  $x = +3$

(c)  $\text{CH}_4$   $x + (1 \times 4) = 0$

$x + 4 = 0$  or  $x = -4$

(d)  $\text{C}_6\text{H}_{12}\text{O}_6$   $6 \times x + 1 \times 12 + (-2 \times 6) = 0$

$6x + 12 - 12 = 0$  or  $x = 0$

(e)  $\text{Na}_2\underline{\text{B}}_4\underline{\text{O}}_7$   $(+1 \times 2) + (x \times 4) + (-2 \times 7) = 0$

$+2 + 4x - 14 = 0$  or  $x = +3$

(f)  $\text{Na}_4[\underline{\text{Fe}}(\underline{\text{C}}\underline{\text{N}})_6]$   $(+1 \times 4) + x + (-1 \times 6) = 0$

$+4 + x - 6 = 0$  or  $x = +2$

(g)  $\underline{\text{N}}_2\underline{\text{H}}_4$   $2 \times x + (1 \times 4) = 0$

$2x + 4 = 0$  or  $x = -2$

#### □ Example 6.

Calculate the oxidation number of (i) Fe in  $\text{Fe}_3\text{O}_4$  (ii) S in  $\text{Na}_2\text{S}_4\text{O}_6$  (iii) Pb in  $\text{Pb}_3\text{O}_4$  (iv) N in  $(\text{NH}_4)_2\text{SO}_4$ .

**Solution:**

(i) **Fe in  $\text{Fe}_3\text{O}_4$ .** Let the oxidation number of Fe be  $x$ .

$$\begin{array}{ccc} x & -2 & 3x - 8 = 0 \\ \text{Fe}_3 & \text{O}_4 & 3x = 8 \end{array}$$

$$\begin{array}{ccc} (x \times 3) & (-2 \times 4) & x = 8/3 = \mathbf{2.67} \end{array}$$

(ii) **S in  $\text{Na}_2\text{S}_4\text{O}_6$ .** Let the oxidation number of S be  $x$ .

$$\begin{array}{ccc} +1 & x & -2 & 2 + 4x - 12 = 0 \\ \text{Na}_2 & \text{S}_4 & \text{O}_6 & 4x = 10 \end{array}$$

$$\begin{array}{ccc} (x \times 4) & (-2 \times 6) & 4x = 10 \end{array}$$

$$\begin{array}{ccc} (+1 \times 2) & (x \times 4) & (-2 \times 6) & x = \frac{10}{4} = \mathbf{2.5} \end{array}$$

(iii) **Pb in  $\text{Pb}_3\text{O}_4$ .** Let the oxidation no. of Pb be  $x$ .

$$\begin{array}{ccc} x & -2 & 3x - 8 = 0 \\ \text{Pb}_3 & \text{O}_4 & 3x = 8 \end{array}$$

$$\begin{array}{ccc} (x \times 3) & (-2 \times 4) & 3x = 8 \end{array}$$

$$\begin{array}{ccc} (x \times 3) & (-2 \times 4) & \text{or} & x = \frac{8}{3} = \mathbf{2.67} \end{array}$$

(iv) **N in  $(\text{NH}_4)_2\text{SO}_4$ .** Let the oxidation no. of N be  $x$ .

$$\begin{array}{ccc} \left( \begin{array}{cc} x+1 & -2 \\ \text{NH}_4 & \text{SO}_4 \end{array} \right)_2 & 2x + 2(+1 \times 4) + (-2) = 0 \\ & 2x + 6 = 0 \end{array}$$

$$\therefore 2x = -6 \text{ or } x = -\frac{6}{2} = \mathbf{-3}$$

### KEY NOTE

#### PARADOX OF FRACTIONAL OXIDATION NUMBER

► For some compounds, the general rules give *fractional oxidation number* to some atoms. For example, the oxidation number of Fe in  $\text{Fe}_3\text{O}_4$  comes out to be  $+\frac{8}{3}$ . This is **strange** because electrons are never shared or transferred in fractions.

Actually, this fractional oxidation number is the **average** oxidation number of the element in the compound. This can be explained by the fact that  $\text{Fe}_3\text{O}_4$  contains Fe atoms of both +2 and +3 oxidation number. It is a stoichiometric mixture of ferrous ( $\text{FeO}$ ) and ferric ( $\text{Fe}_2\text{O}_3$ ) oxides combined as  $\text{FeO} \cdot \text{Fe}_2\text{O}_3$ . Therefore, the oxidation number found in such cases is **average oxidation number**.

Similarly,  $\text{Mn}_3\text{O}_4$  is regarded as a mixture of  $\text{MnO}$  and  $\text{MnO}_2$  having the composition  $2\text{MnO} \cdot \text{MnO}_2$ . Therefore, the oxidation number of Mn in  $\text{MnO}$  is +2 and that in  $\text{MnO}_2$  is +4. Hence, average comes out to be  $[2(+2) + 1(+4)]/3 = 8/3$ .

Red lead,  $\text{Pb}_3\text{O}_4$  is regarded as a mixture with composition of  $2\text{PbO} \cdot \text{PbO}_2$ . In  $\text{PbO}$ , the oxidation number of Pb is +2 while in  $\text{PbO}_2$ , the oxidation number of Pb is +4. The average value is  $[2(+2) + 1(+4)]/3 = 8/3$ .

Similarly, in ferriferrocyanide,  $\text{Fe}_4[\text{Fe}(\text{CN})_6]_3$  molecule, there are four iron atoms with oxidation number +3 and three iron atoms with oxidation number +2. The oxidation number of iron in this molecule is the average of all these atoms equal to  $18/7$  (strange number). Thus, *the average oxidation number indicates average of atoms in different bonding situations in a compound.*

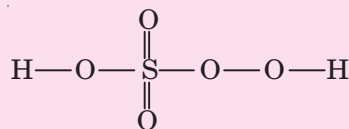
► In some compounds, the knowledge of chemical bonds or structures is essential to calculate oxidation numbers. For example

(a) **Caro's acid,  $\text{H}_2\text{SO}_5$  (peroxomonosulphuric acid)**

The O.N. can be calculated as

$$2 \times (+1) + x + 5(-2) = 0, \quad x = +8$$

But this is wrong because the maximum oxidation number of S cannot be more than +6 because it has only six electrons in its valence shell. In this structure, two oxygen atoms are bonded by a peroxide bond as



In the peroxide bond each O atom has -1 oxidation state

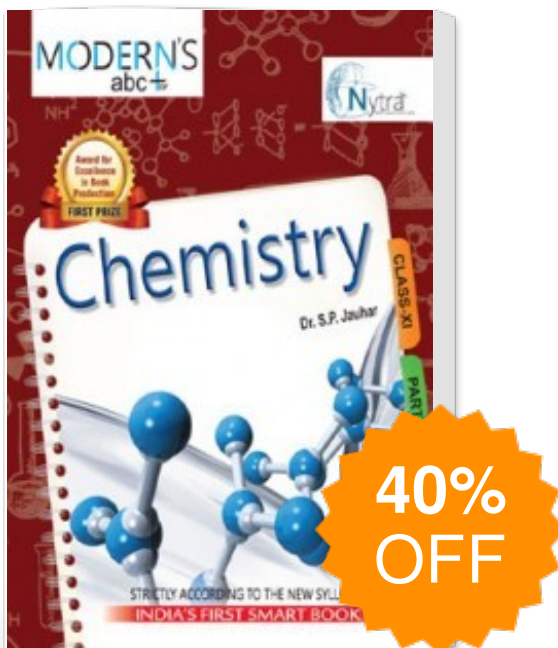
$\text{H}_2\text{S}(\text{O}_2)\text{O}_3$

$$2(+1) + x + 2(-1) + 3(-2) = 0 \quad \text{or} \quad x - 6 = 0$$

$$\therefore x = +6$$

$\therefore$  O.N. of S in  $\text{H}_2\text{SO}_5$  is +6

# Modern ABC Plus Of Chemistry Class 11 Part-II



Publisher : MBD Group  
Publishers

ISBN : 9789351841210

Author : Dr. S. P. Jauhar

Type the URL : <http://www.kopykitab.com/product/12162>



## Get this eBook