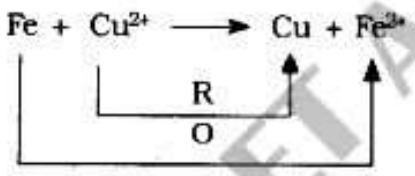
Chemistry : Redox Reaction

INTRODUCTION

- **Chemical reaction:** The process in which a single substance or many substances (reactants) interact with each other to produce one or more substances (products) is called chemical reaction.

1. Redox Reaction (Oxidation-Reduction) :

- The chemical reactions involved transfer *of* electrons from one chemical substance to another. These electron-transfer reactions are termed as Oxidation-Reduction or Redox reactions.
- These reactions involves oxidation and reduction both simultaneously.
- These reaction involves increase in oxidation number and decrease in oxidation number simultaneously.
- A complete reaction showing oxidation and reduction together is called a Redox reaction.



CLASSICAL CONCEPTS OF REDOX REACTIONS

- There are four concepts for oxidation and reduction reactions.

Classical concept :

(a) **Oxidation** : Oxidation is a process which involves.

(i) **Addition of Oxygen** : Oxidation is a chemical reaction in which oxygen is added

e.g. $2Mg + O_2 \longrightarrow 2MgO$

(Oxidation of Magnesium)

$C + O_2 \longrightarrow CO_2$ (Oxidation of Carbon).

$Na_2SO_3 + H_2O_2 \longrightarrow Na_2SO_4 + H_2O$

(Oxidation of Sodium sulphite)

(ii) **Removal of Hydrogen** : Hydrogen is removed i.e. hydrogen becomes less

e.g. $H_2S + Cl_2 \longrightarrow 2HCl + S$

(Oxidation of Hydrogen sulphide)

$4HI + O_2 \longrightarrow 2H_2O + 2I_2$

(Oxidation of Hydrogen iodide)

(iii) **Addition of Electronegative element :**

e.g. $Fe + S \longrightarrow FeS$ (Oxidation of iron)

$SnCl_2 + Cl_2 \longrightarrow SnCl_4$

(Oxidation of stannous chloride)

$2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3 + 2H_2O$

(Oxidation of ferric chloride)

(iv) **Removal of Electropositive element :**

e.g. $2NaI + H_2O_2 \longrightarrow 2NaOH + I_2$

(Oxidation of NaI)

$2KI + Cl_2 \longrightarrow 2KCl + I_2$ (Oxidation of KI)

$2K_2MnO_4 + Cl_2 \longrightarrow 2KCl + 2KMnO_4$

(Oxidation of Potassium magnate)

(v) **Increment in oxidation state of Electropositive element :**

e.g. $SnCl_2 + Cl_2 \longrightarrow SnCl_4$

Sn+2 Sn+4

(b) **Reduction** : Reduction is a process which involves

(i) **Removal of Oxygen** : Oxygen is lost.

e.g. $CuO + C \longrightarrow Cu + CO$

(Reduction of cupric oxide)

$H_2O + C \longrightarrow$ (Reduction of water)

Steam coke

$Fe_3O_4 + 4H_2 \longrightarrow 3Fe + 4H_2O$

(Reduction of Fe_3O_4)

$Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3$

$Cr_2O_3 + 2Al \longrightarrow 2Cr + Al_2O_3$

(d) **Anion** : Anion accepts electron, its negative charge increases.

$A^{-n_1} + (n_2 - n_1)\, e^- \longrightarrow A^{-n_2}$

$MnO_4^{-1} + e^- \longrightarrow MnC_4^{-2}$

$[Fe(CN)_6]^{-3} + e^- \longrightarrow [Fe(CN)_6]^{-4}$

(c) **Molecule :** Molecule accepts electron, it is a reduction reaction.

$O_2 + 4e^- \longrightarrow 2O^{-2}$

$I_2 + 2e^- \longrightarrow 2I^{-1}$

Therefore in reduction reactions-

(i) Positive charge decreases and negative charge increases

(ii) Oxidation number decreases

- **Valency concept of oxidation and reduction:**

(1) Oxidation : This process invovle the increase in (+) ve valency or decrease in (–) ve valency of a substance takes place called oxidation.

e.g. $2Mg + O_2 \longrightarrow 2MgO$

(0) (2)

(ii) **Reduction** :- This process involve decreases in (+) ve valency or increases in (–) ve valency .

e.g. $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$

(+2) (+1)

OXIDATION NUMBER CONCEPT

❑ Oxidation number of an element in a particular compound represents the number of electrons lost or gained by an element during its change from free state into that compound or Oxidation number of an element in a particular compound represents the extent of oxidation or reduction of an element during its change from free state into that compound.

(a) Oxidation number is given positive sign if electrons are lost. Oxidation number is given negative sign if electrons are gained.

(b) Oxidation number represents real charge in case of ionic compounds, however, in covalent compounds it represents for imaginary charge.

(c) It is the residual charge which an atom appears to have when other atom are withdrawn from the molecules as ions by containing electrons with more electronegative atoms.

(d) It represents the real charge in case of ionic compounds and represents the imaginary charge in case of covalent compounds.

(e) Maximum oxidation number of an element is equal to group number which be belongs in the periodic table.

(f) Minimum oxidation number of an element is equal to (group number –8)

☞ Group elements always shows +1 as oxidation number

☞ Group elements always shows +2 as oxidation number

☞ Group elements always shows +3 oxidation number

☞ Group shows -4 to +4 oxidation number

☞ Group shows -3 to +5 oxidation number

☞ Group shows -2 to +6 oxidation number

☞ Group shows –1 to +7 oxidation number (except F)

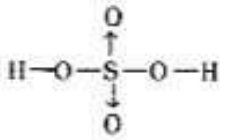
Inert gases always shows zero oxidation number.

A) Oxidation state for Coordinate bond :-

(a) Compounds having coordinate bond between low electronegative element and high electronegative element then the e- donor element show **+2** oxidation number whereas e- acceptor element show -2. Oxidation number In this type of bond of bonded compounds.

❑ Electronegativity of acceptor > electronegativity of donor.

For example in H_2SO_4-



❑ Here 'S' is low electronegative element than O therefore, oxidation number of S =+2 and oxidation number of 0 = -2 (due to single dative bond).

2. Oxidising agent or Oxidant (O. A.):

A) **Oxidant:** Oxidant is a species that accepts electron(s) and causes other substance to lose electron(s) (that is to oxidize). OR Oxidant is an electron acceptor. Oxidants (oxidizing agents) oxidizes other substances and themselves get reduced.

(a) Oxidising agents are lewis acids.

(b) Substances which shows the decrement in oxidation number.

B) **Important oxidising agents**

(a) High electronegative elements like N, O, F; Cl etc.

(b) Metallic oxides like Li_2O, Na_2O, Na_2O_2, CO_2, CaO, MgO, BaO_2 etc.

(c) Nonmetallic oxides like CO_2, SO_2, H_2O_2, O_3.

(d) Neutral compound or ion in which element shows their highest oxidaion no or state, will act as oxidant or oxidising agent like $KMnO_4$, H_2SO_4, $SnCl_4$, H_3PO_4, $K_2Cr_2O_7$, $HClO_4$, $CuCl_2$, HNO_3, H_2SO_3, $FeCl_3$, $HgCl_2$, etc.

(ii) **Addition of Hydrogen: Hydrogen is added,**

e.g. $N_2 + 3H_2 \longrightarrow 2NH_3$ (Reduction of Nitrogen)

$Cl_2 + H_2 \longrightarrow 2HCl$ (Reduction of Chlorine)

$S + H_2 \longrightarrow H_2S$ (Reduction of Sulphur)

(iii) **Removal of Electronegative element**

e.g. $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$

(Reduction of mercuric chloride)

$2FeCl_3 + H_2 \longrightarrow 2FeCl_2 + 2HCl$

(Reduction of ferric chloride)

$PbS + H_2 \longrightarrow Pb + H_2S$

(iv) **Addition of Electropositive element :**

e.g. $HgCl_2 + Hg \longrightarrow Hg_2Cl_2$

(Reduction of mercuric chloride)

$CuCl_2 + Cu \longrightarrow Cu_2Cl_2$

(Reduction of cupric chloride)

$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$

(Reduction of mercuric chloride)

(v) **Decrement in oxidation state of Electropositive element**

e.g. $Cu_5O_4 + Fe \longrightarrow FeSO_4 + Cu$

(Reduction of $CuSO_4$)

Cu^{+2} Cu°

$FeCl_3 + H_2S \longrightarrow FeCl_2 + 2HCl + S$

(Reduction of FeCb)

Fe^{+3} F_e^{+2}

MODEM CONCEPT OR ELECTRONIC CONCEPT

(i) **Oxidation** : When element or an atom or an ion or molecule loses electron is called oxidation.de electronation is oxidation.

e.g.

(a) **Neutral atom** : Neutral atom loses electron, it gets converted to a positive ion.

$M \longrightarrow M^{+n} + ne^-$

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

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

$Al \longrightarrow Al^{+3} + 3e^-$

(b) **Cation** : Cation loses electron, there is an increase in its positive charge.

$M^{+n_1} \longrightarrow M^{+n_2} + (n_2 - n_1)e^-$

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

(ferrous) (ferric)

$Sn^{+2} \longrightarrow Sn^{+4} + 2e^-$

$Hg^{+1} \longrightarrow Hg^{+2} + e^-$

(c) **Anion** : Anion loses electron equal to its negative charge, it gets converted to a neutral atom.

$A^{-n} \longrightarrow A + ne^-$

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

$2O^{-2} \longrightarrow O_2 + 4e^-$

$2N^{-3} \longrightarrow N_2 + 6e^-$

(d) **Complex Anion** : Complex anion loses electron, its negative charge decreases.

$A^{-n_1} \longrightarrow + (n_1 - n_2)\, e^-$

$[Fe(CN)_6]_4^- \longrightarrow [Fe(CN)_6]_3^- + e^-$

$[Fe(CN)_6]_4^- \longrightarrow [Fe(CN)_6]_3^- + e^-$

(e) **Molecule** : Molecule loses electrons, it breaks up into it constituents.

$H_2O_2 \longrightarrow 2H^{+1} + O_2 + 2e^-$

Therefore in oxidation reactions-

(i) Positive charge increases and negative charge decreases

(ii) Oxidation number increases

(ii) **Reduction** When an element or an atom or an ion (positive or negative) or a molecule accepts electron,iscalled reduction. Electronation is reduction.

(a) **Neutral Atom** : Neutral element or atom accepts electron, it get converted into an anion.

$A + ne^- \longrightarrow A^{-n}$

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

$N + 3e^- \longrightarrow N^{-3}$

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

(b) **Cation** : Cation accepts electron equal to its charge, it gets converted into a neutral atom.

$Mg^{+2} + 2e^- \longrightarrow Mg^{\circ}$

$Al^{+3} + 3e^- \longrightarrow Al^{\circ}$

(c) When a cation accepts less electrons than its charge, its positive charge decreases.

For example

$M^{+2} + (n_2 - n_1)e^- \longrightarrow M^{+n_1}$

$Sn^{+4} + 2e^- \longrightarrow Sn^{++}$

(Stannic ion) (Stannous ion)

$Cu^{+2} + e^- \longrightarrow Cu^{+1}$

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

3. **Reducing agent or Reductant (R. A.):**

A) **Reductant:** Reductant is a species that donates electron(s) and causes other substance to accept electron(s) (that is to reduce). OR Reductant is an electron donor.

- Reductants (reducing agents) reduces other substances and themselves get oxidized.
- The substance which donates electrons in a chemical reaction is called reducing agent i.e., electrons donor are reducing agents.

(a) Reducting agents are lewis bases.

(b) Substances which show the increment in oxidation number.

B) Important reducing agent

(a) All metals like, K, Mg, Ca, etc.

(b) All metallic hydrides like NaH, CaH_2, $LiAlH_4$, $NaBH_4$, AlH_3 etc.

(c) All hydroacids like HF; HCl, HBr, H_2S, etc.

(d) Some organic compound like Aldehyde, formic acid, oxalic acid, tartaric acid.

(e) All neutral compounds or ions, in which element shows their lowest oxidation no. or state, will act as reductant or reducing agent

(f) Some organic compounds

e.g. HCOOH, Aldehydes, Oxalic acid, Tartaric acid etc.

MnO, HClO, $HClO_2$, H_3PO_2, HNO_2, H_2SO_3, $FeCl_2$, $SnCl_2$, Hg_2Cl_2, CH_2Cl_2 etc.

C) **Compound which can acts as oxidant and reductant both**

HNO_2, SO_2, H_2O_2, O_3, Al_2O_3, CrO_2, MnO_2, ZnO, CuO

Note : Al_2O_3, CrO_2, MnO_2, ZnO, CuO are called as amphoteric oxide.

D) **Identification of oxidants and reductants form , redox reacation :**

i. $H_3PO_{4(aq)} + 3KOH_{(aq)} \longrightarrow K_3PO_{4(aq)} + 3H_2O\ (l)$

Step-I: The oxidation number is assigned to all atoms.

$H_3PO_{4(aq)} + 3KOH_{(aq)} \longrightarrow K_3PO_{4(aq)} + 3H_2O(l)$

+1+5-2 +1-2+1 +1+5-2 +1-2

Step-2: No species has undergone change in oxidation number. Therefore, the reaction is not a redox reaction.

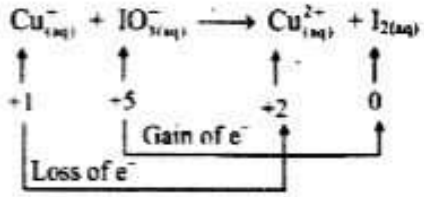
ii. $Zn_{(s)} + 2HCl(aq) \rightarrow ZnCl_{2(aq)} + H_2(g)$

Step-I: The oxidation number is assigned to all atoms.

$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_2(aq) + H_2(g)$

0 +1–1 +2–1 0

Step-2 : The species undergoing change in oxidation number are identified.



- The oxidation number of Zn increases from 0 to +2. It is therefore oxidized and is a reducing agent.
- The oxidation number of H decreases from + 1 to 0. HCl is therefore reduced and is an oxidizing agent.

iii. $6Fe^{2+}(aq) + BrO^{-3}(aq) \longrightarrow Fe^{3+}_{(aq)} + Br^-_{(aq)}$

$+ 3H_2O_{(l)} + 6H^+_{(aq)}$

Step-I:The oxidation number is assigned to all atoms.

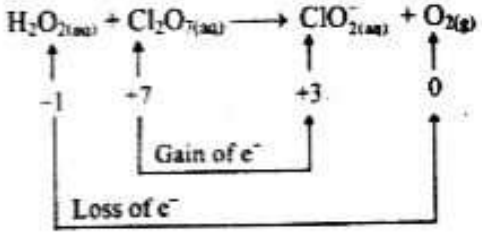
$6Fe^{2+}_{(aq)} + BrO^{3}_{(aq)} \longrightarrow$

+2 +5 –2

$Fe^{3+}_{(aq)} + Br_{(aq)} + 3H_2O_{(l)} + 6H^+_{(aq)}$

+3 –1 +1 –2 +1

Step-2: The species undergoing change in oxidation number are identified.



- The oxidation number of Fe increases from +2 to +3. Fe^{2+} is therefore oxidized to Fe^{3+} and is a reducing agent.
- The oxidation number of Br decreases from +5 to –1. BrO_3^- is therefore reduced to Br and is an oxidizing agent.

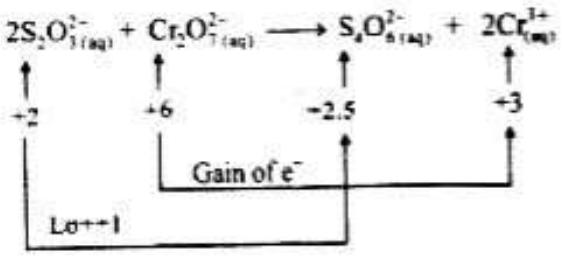
iv. $2Zn_{(s)} + O_{2(g)} \longrightarrow 2ZnO_{(s)}$

Step-I:The oxidation number is assigned to all atoms.

$2Zn_{(s)} + O_2(g) \longrightarrow 2ZnO_{(s)}$

0 0 +2 –2

Step-2: The species undergoing change in oxidation number are identified.



- The oxidation number of Zn increases from 0 to +2. Zn is therefore oxidized to ZnO and is a reducing agent.
- The oxidation number of O decreases from 0 to -2. O_2 is therefore reduced to ZnO (O^{2-}) and is an oxidizing agent.

- **Calculation of equivalent weight**

The equivalent weight of an oxidising agent is that weight which accepts one mole electron in a chemical reaction.

(a) Equivalent weight of an oxidant (get reduced)

$$= \frac{\text{Molecular weight}}{\text{Number of electrons gained by one mole}}$$

Ex. In acidic medium

$6e^- + Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$

Atoms undergoes reduction is Cr its oxidation state is decreasing from +6 to +3

Equivalent weight of $K_2Cr_2O_7$

$$= \frac{\text{Molecular weight of } K_2Cr_2O_7}{3 \times 2} = \frac{\text{Molecular weight}}{6}$$

Ex.

(i) Equivalent weight of $KMnO_4$ in acidic medium

$MnO_4^- \longrightarrow Mn_2^+$ (acidic medium)

(+7) (+2)

5 electrons are taken so equivalent weight

$$\text{Molecular weight of } KMnO_4 = \frac{158}{5} = 31.6$$

(ii) Equivalent weight of KMn04 in neutral medium

$MnO_4^- \longrightarrow Mn^{+2}$ (neutral medium)

(+7) (+2)

Only 3 electrons are gained, so equivalent weight

$$= \frac{\text{Molecular weight of } KMnO_4}{3} = \frac{158}{3} = 52.7$$

(iii) Equivalent weight of KM.n04 in alkaline medium

$MnO_4^- \longrightarrow MnO_4^{-2}$ (alkaline medium)

(+7) (+6)

Only one electron is gained, so equivalent weight

$$\frac{\text{Molecular weight of } KMnO_4}{1} = 158$$

Note It is important to note that KMn04 acts as an oxidant in every medium although with different strength which follows the order –

acidic medium > neutral medium > alkaline medium while, $K_2Cr_2O_7$ acts as an oxidant only in acidic medium as follows

$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$

$(2 \times 6) \longrightarrow (2 \times 3)$

6 electrons are gained so equivalent weight

$$= \frac{\text{Molecular weight of } K_2Cr_2O_7}{6} = \frac{294.21}{6} = 49.03$$

(d) It is clear that $KMnO_4$ is better oxidant than K_2CrO_7.

The equivalent weight of a reducing agent is the weight which donates one electron in a chemical reaction .

e.g. $2S_2O_3^- \longrightarrow S4O_6^- + er$

(R. A.)

Equivalent weight of $S_2O_3^{--} = \frac{2M}{2} = M$

(e) To determine the molecular formula of compound

e.g. Suppose that there are three atoms A, B, C and their oxidation number are 6, -1,-2, respectively. Then the molecular formula of compound will be.

Sol. Since, the charge on a free compound is zero. So

$+6 = (-1 \times 4) + (-2)$

$+6 = -6$

or $+6 = (-1 \times 2) + (-2 \times 2)$

$= -2 + (-4) = -6$

So molecular formula, AB_4C or AB_2C_2.

4. **Valency and Oxidation number**

- Valency of an element represents the power or capacity of the element to combine with the other element.The valency of an element is numerically equal to the number of hydrogen atoms or chlorine atoms or twice the number of oxygen atoms that combine with one atom of that element.
- It is also equal to the number of electrons lost or accepted or shared by the atoms of an element.
- In some cases (mainly in the case of electrovalent compounds), valency and oxidation number are the same but in other cases they may have different values.

5. **Redox Reactions**

- Redox reactions are the chemical reactions which involve both oxidation as well as reduction simultaneously.
- Infact, oxidation and reduction go hand in hand. Any redox reaction may be divided in two parts.

(i) Oxidation half reaction.

(ii) Reduction half reaction.

- Now, we will study some reaction.

Note :- In reaction 2 oxygens of ozone have different, oxidation state.

Structure of Ozone is

$O_3 \leftarrow O_1 = O_2$

Oxidation state of $O_1 = +2$

Oxidation state of $O_2 = 0$

Oxidation state of $O_3 = -2$

A) Types of Redox reaction

(a) **Intermolecular redox reaction :** In these reacation oxidation and reduction take place separately in the different compounds.

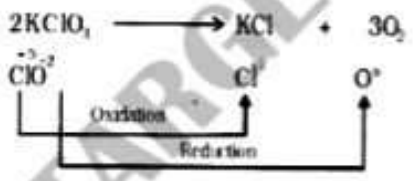
$SnCl_2 + 2FeCl_3 \longrightarrow SnCl_4 + 2FeCl_2$

$Sn^{+2} \longrightarrow Sn^{+4}$ (Oxidation)

$Fe^{+3} \longrightarrow Fe^{+2}$ (Reduction)

(b) **Intra molecular redox reaction**

In these chemical reaction oxidation and reduction take place in single compound.



(c) **Disproportionation reaction**

One and the same substance may act simultaneously as an oxidising and as a reducing agent. As a result a part of it gets oxidised to higher state and rest of it is reduced to a lower state of oxidation. Such as reaction, in which the substance undergoes simultaneous oxidation and reduction is called disproportionation.

Oxidation

$H_2O_2 + H_2O_2 \longrightarrow 2H_2O + O_2$

Reduction

7. **Balancing of redox reactions**

A) **Oxidation number method**

The general procedure involves the following steps:-

i. Unbalanced redox equation is written. The equation is balanced for all atoms (except H and O). The oxidation number is assigned to all atoms in the reactants and the products, with the help of rules used to assign the oxidation number.

ii. Atoms with change in oxidation numbers are identified and the oxidation numbers are changed by how much is also calculated. Brackets are drawn to connect atoms of the elements that are oxidized similar. Also, atoms of the elements that are reduced are connected by drawing similar brackets.

iii. Increase in oxidation number per atom of the oxidized species and the net increase in oxidation number is shown. Also, a decrease in oxidation number per atom of the reduced species and the net decrease in oxidation number is shown.

iv. Coefficients are determined so that the total increase in oxidation numbers is equal to total decrease in oxidation numbers. These coefficients are then inserted into the equation.

v. Balance oxygen atoms, H_2O is added to, the side which contains less O^- atoms (one H_2O molecule for one 0–atom is added). H^- atoms are balanced by adding W ions, to the side with less H^- atoms.

vi. The reaction which occurs in basic medium, OH^- ions, equal to the number of W ions are added on both sides of the equation. The Wand Olr ions appearing on the same side of the reaction are combined to give H_2O molecules.

vii. It is checked that the reaction is balanced with respect to both the number of atoms of each element and the charges.

a) **Balancing of reaction by the oxidation number method in acidic medium -**

1. $MnO_4^- + Fe^{+2} \rightarrow Mn^{+2} + Fe^{+3}$

Sol. Write the oxidation number of all the atoms.

$+7 \quad -2 \qquad\qquad +2 \qquad +3$

$MnO_4^- + Fe^{+2} \longrightarrow Mn^{+2} + Fe^{+3}$

Change in oxidation number has occured in Mn and Fe.

$+7 \qquad\qquad +2$

$MnO_4 \longrightarrow Mn^{+2}$ ---(1)

(Decrement in oxidation number by 5)

$Fe^{+2} \longrightarrow Fe^{+3}$ ---(2)

(Increment in oxidation number by 1)

To make increase and decrease equal, eq. (2) is multiplied by 5.

$MnO_4^- + 5Fe^{+2} \longrightarrow Mn^{+2} + 5Fe^{+3}$

To balance oxygen, $4H_2O$ are added to R.H.S. and to balance hydrogen, $8H^+$ are added to

L.H.S., $MnO_4^- + 5Fe^{+2} + 8H^+ \longrightarrow Mn^{+2} + 5Fe^{+3} + 4H_2O$

This is the balanced equation.

2. $MnCl_2(aq) + HO_2^-(aq) \longrightarrow$

$Mn(OH)_{3(s)} + Cl^-_{(aq)}$

Cl atoms are balanced and oxidation number is assigned to all atoms.

$MnCl_{2(aq)} + HO_2^-{}_{(aq)} \longrightarrow Mn(OH)_{3(s)} + 2Cl^-_{(aq)}$

↑ ↑ ↑ ↑ ↑ ↑ ↑ ↑

+2 –1 +1 –1 +3 –2 +1 –1

Increase in oxidation number:

Mn (+2) $\longrightarrow$ Mn (+3)

Increase per atom = Net increase = + 1

Decrease in oxidation number:

O (–1) $\longrightarrow$ O(–2)

Decrease per atom = – 1

Net decrease =

–2 (as there are 2 O^- atoims in HO_2^- species.)

$2MnCl_{2(aq)} + HO_2^-(aq) \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl^-_{(aq)}$

$2MnCl_{2(aq)} + HO_2^-{}_{(aq)} + 4H_2Cl(l) \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl^-_{(aq)}$

$2MnCl_{2(aq)} + HO_2^-(aq) + 4H_2Cl_{(l)} \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl^-_{(aq)} + 3H^-_{(aq)}$

The reaction occurs in basic medium, OH^- ions, equal to the number of H^+ ions, are added on both sides of the equation. Hence, $3OH^-$ ions are added on both sides.

$2MnCl_{2(aq)} + HO^-_{2(aq)} + 4H_2Cl_{(l)} \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl^-_{(aq)} + 3H^-_{(aq)} + 3OH^-$

$2MnCl_{2(aq)} + HO^-_{2(aq)} + 4H_2Cl_{(l)} \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl_{-(aq)}$

The final equation is

$2MnCl_{2(aq)} + HO^-_{2(aq)} + 4H_2Cl_{(l)} \longrightarrow$

$2Mn(OH)_{3(s)} + 4Cl^-_{(aq)}$

B) Ion-Electron method

❑ The following steps are followed while balancing redox reaction (equations) by this method.

(i) Write the equation in ionic form.

(ii) Split the redox equation into two half reactions, one representing oxidation and the other representing reduction.

(iii) Balance these half reactions separately and then add by multiplying with suitable coefficients so thqt the electrons are cancelled. Balancing is done as follows.

(iv) Add both half reaction and then balance the atoms other 'O' and H'.

(v) Then balanœ oxygen atoms by adding H_2O molecules to the side deficient in oxygen. The number of H_2O molecules added is equal to the deficiency of oxygen atgms.

(vi) Balance hydrogen atoms by adding H^+ ions equal to the deficiency in the side which is the deficient in hydrogen atoms.

(vii) If the medium of reaction is basic, OH^- ions are added to both sides of balanced equation equal in number of H+ in balanced equation.

(viii) Balance the charge by electrons to the side which is rich in +ve charges, i.e. deficiency in electrons. Number of electrons added is equal to the deficiency.

(ix) Multiply the half equations with suitable coefficients to equalize the number of electrons.

(x) Add these half equations to get an equation which is balanced with respect ot charge and atoms.

Ex.

a) Acidic Medium

1. $Cr_2O_7^{2-} + C_2O_4^{2-} \longrightarrow Cr^{3+} + CO_2$

$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$ (Reduction half reaction)

$C_2O_4^{2-} \longrightarrow CO_2$ (Oxidation half reaction)

Atoms other than H and O are balanced.

$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

$C_2O_4^{2-} \longrightarrow 2CO_2$

Balance O-atoms by the addition of H_2O to another side

$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$

$C_2O_4^{2-} \longrightarrow 2CO_2$

Balance H-atoms by the additi on of H+ to another side

$Cr_2O_7^{2-} + 14\,H^+ \longrightarrow 2Cr^{3+} + 7H_2O$

$C_2O_4^{2-} \longrightarrow 2CO_2$

Balance the charge by the addition of electron (e^-)

$Cr_2O_7^{2-} + 14\,H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$

$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$

Multiply equations by a constant to get the same number of electrons on both side.

$Cr_2O_7^{2-} + 14\,H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$

$3C_2O_4^{2-} \longrightarrow 6CO_2 + 6e^-$

$Cr_2O_7^{2-} + 3C_2O_4^{2-} + 14\,H^+ \longrightarrow$

$2Cr^{3+} + 6CO_2 + 7H_2O$

(b) **Balancing of redox reaction in Alkaline Medium :-**

I. $H_2C_2O_4 + H_2O_2 \longrightarrow CO_2 + H_2O$

Sol.

(i) The half reaction for oxidation is,

$H_2C_2O_4 + xH_2O \longrightarrow yCO_2 + zH_2O$

Balancing carbon atoms on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2$

Balancing hydrogen atoms on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+$

Balancing the charge on both sides,

$H_2C_2O_4 \longrightarrow 2CO_2 + 2H^+ + 2e^-$ (balanced)

(ii) The half reaction for reduction is,

$H_2O_2 \longrightarrow H_2O$

Balancing oxygen atoms on both sides,

$H_2O_2 \longrightarrow 2H_2O$

Balancing hydrogen atoms,

$H_2O_2 + 2H^+ \longrightarrow 2H_2O$

Balancing the charge,

$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$ (balanced)

Now, adding both equation,

$H_2C_2O_4 + H_2O_2 \longrightarrow 2CO_2 + 2H_2O$

II. **Applications of redox reaction**

A) **Metallurgy:**

a. During roasting, sulphide minerals are converted to corresponding oxides.

$2ZnS_{(s)} + 3O_{2(g)} \xrightarrow{heat} 2ZnO_{(s)} + 2SO_{2(g)}$

↑ ↑ ↑ ↑ ↑ ↑ ↑

+2 –2 0 +2 –2 +4 –2

The oxidation number of S increases from –2 to +4 and that of oxygen decreases from 0 to –2. Thus, this is a redox reaction.

b. The reduction of ZnO by coke gives Zn.

$ZnO_{(s)} + C_{(s)} \xrightarrow{heat} Zn_{(s)} + CO_{(g)}$

↑ ↑ ↑ ↑ ↑ ↑

+2 –2 0 0 +2 –2

The oxidation number of Zn decreases from +2 to 0 and that of C increases from 0 – to +2 and therefore, this is also a redox reaction.

B) **Batteries:**

- The redox reactions occurring in batteries or galvanic cells are responsible for generation of electricity.
- In Daniel cell, Zn and Cu electrodes are connected through a wire. These electrodes are dipped in the solutions of their own ions. The electrons flow from Zn to Cu electrode through the connecting wire. Zn loses two electrons and gets oxidized to Zn^{2+}. Cu^{2+} ion gains two electrons and gets reduced to Cu.

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

↑ ↑ ↑ ↑

0 +2 +2 0

- The oxidation number of Zn increases from 0 to +2 and that of Cu decreases from +2 to 0.

  Therefore, it is a redox reaction.

C) **Bleaching:**

- Bleaching is the decolourization or lightening of coloured material. It involves a redox reaction.

a. NaOCl acts as an oxidizing agent in bleaching of clothes. It removes stains.

b. Chlorine is an oxidizing agent. It bleaches wood pulp into white paper.

c. H_2O_2 acts as an oxidizing agent. It bleaches dark hair by a redox reaction.