SOME BASIC CONCEPTS OF CHEMISTRY
REDOX REACTION

Oxidising And Reducing Agent

- **Oxidising Agent or Oxidant**

Oxidising agents are those compounds which can oxidize others and reduce itself during the chemical reaction. Those reagents in which for an element, oxidation number decreases or which undergoes gain of electrons in a redox reaction are termed as oxidants.

Ex. $KMnO_4$, $K_2Cr_2O_7$, HNO_3, conc. H_2SO_4 etc are powerful oxidising agents.

- **Reducing Agent or Reductant**

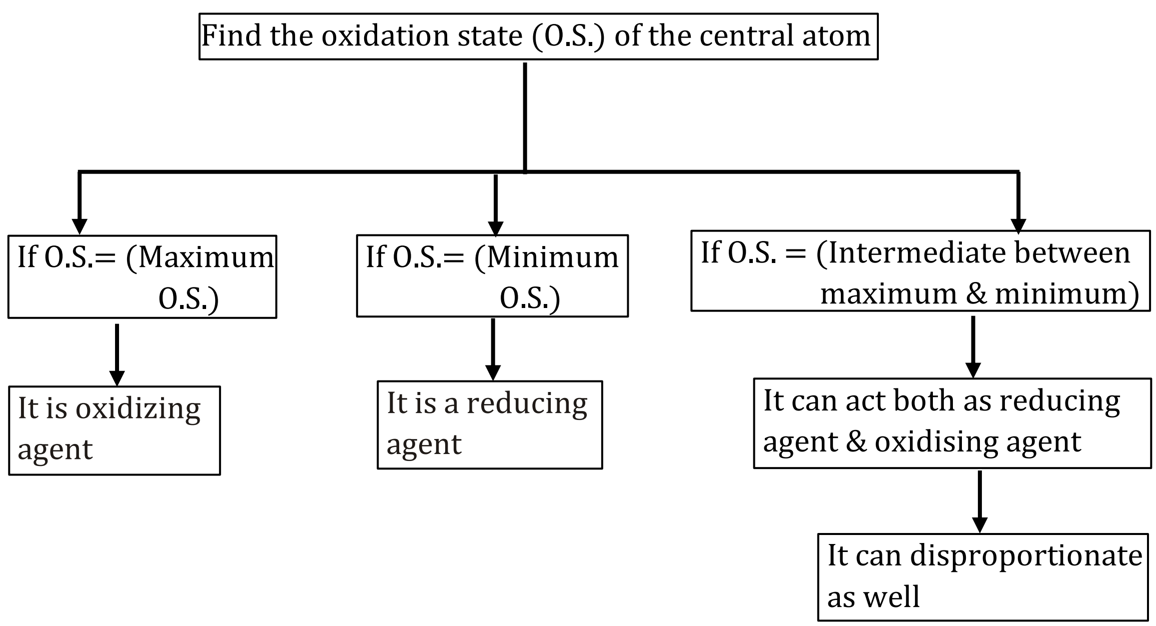
Reducing agents are those compounds which can reduce other and oxidize itself during the chemical reaction. Those reagents in which for an element, oxidation number increases or which undergoes loss of electrons in a redox reaction are termed as reductants.

Ex. KI, $Na_2S_2O_3$ etc are the powerful reducing agents.

Note: There are some compounds also which can work both as oxidising agent and reducing agent

Ex. H_2O_2, NO_2^-

HOW TO IDENTIFY WHETHER A PARTICULAR SUBSTANCE IS AN OXIDISING OR A REDUCING AGENT



- **Redox Reaction**

A reaction in which oxidation and reduction simultaneously take place is called a redox reaction in all redox reactions, the total increase in oxidation number must be equal to the total decrease in oxidation number.

Ex. $10\ \overset{+2}{FeSO_4} + 2KMnO_4^{+5} + 8H_2SO_4 \longrightarrow 5\overset{+3}{Fe_2}(SO_4)_3 + 2MnSO_4^{+2} + K_2SO_4 + 8H_2O$

- **Disproportionation Reaction**

A redox reaction in which same element present in a particular compound in a definite oxidation state is oxidized as well as reduced simultaneously is a disproportionation reaction.

Disproportionation reactions are a special type of redox reactions. One of the reactants in a disproportionation reaction always contains an element that can exist in at least three oxidation states.

The element in the form of reacting substance is in the intermediate oxidation state and both higher and lower oxidation states of that element are formed in the reaction. For example:

$$2H_2O_2^{-1}(aq) \longrightarrow 2H_2^{-2}O(\ell) + \overset{0}{O}_2(g)$$

$$S_8^0(s) + 120H^-(aq) \longrightarrow 4\ S^{-2}(aq) + 2\ S_2^{+2}O_3^{2-}(aq) + 6H_2O(\lambda)$$

$$CI_2^0\ (g) + 2OH^-(aq) \longrightarrow Cl^{+1}\ O^-(aq) + Cl^{-1-}(aq) + H_2O(\lambda)$$

- **Consider The Following Reactions**

(a) $2KClO_3\ \ 2KCl + 3O_2$

$KClO_3$ plays a role of oxidant and reductant both. Here, Cl present in $KClO_3$ is reduced and O present in $KClO_3$ is oxidized. Since same element is not oxidized and reduced, so it is not a disproportionation reaction, although it looks like one.

(b) $NH_4NO_2\ \ N_2 + 2H_2O$

Nitrogen in this compound has -3 and +3 oxidation number, which is not a definite value. So, it is not a disproportionation reaction. It is an example of comproportionating reaction, which is a class of redox reaction in which an element from two different oxidation state gets converted into a single oxidation state.

(c) $4KClO_3 \rightarrow 3 \overset{+5}{\rightarrow} 3KClO_4 + \overset{-1}{KCl}$

It is a case of disproportionation reaction and Cl atom is disproportional ting.

List of Some Important Disproportionation Reactions

1. $H_2O_2 \longrightarrow H_2O + O_2$
2. $X_2 + OH^-(\text{dil.}) \longrightarrow X^- + XO^-(X = Cl, Br, I)$
3. $x_2 + OH^-(\text{conc.}) \longrightarrow x + XO_3^-$

F_2 does not undergo disproportionation as it is the most electronegative element.

1. $F_2 + NaOH(\text{ dil.}) \longrightarrow F^- + OF_2$
2. $F_2 + NaOH(\text{ conc.}) \longrightarrow F^- + O_2$
3. $(CN)_2 + OH^- \longrightarrow CN^- + OCN^-$
4. $P_4 + OH^- \longrightarrow PH_3 + H_2PO_2^-$
5. $S_8 + OH^- \longrightarrow S^{2-} + S_2O_3^{2-}$
6. $MnO_4{}^{2-} \longrightarrow MnO_4^- + MnO_2$
7. $NH_2OH \longrightarrow N_2O + NH_3$
8. $NH_2OH \longrightarrow N_2 + NH_3$
9. Oxyacid's of Phosphorus (+1, +3 oxidation number)
10. $H_3PO_2 \longrightarrow PH_3 + H_3PO_3$
11. $H_3PO_3 \longrightarrow PH_3 + H_3PO_4$

10. Oxyacid's of Chlorine (Halogens)(+1, +3, +5 Oxidation number)
    $ClO^- \longrightarrow Cl^- + ClO_2^-$
    $ClO_2^- \longrightarrow Cl^- + ClO_3^-$
    $ClO_3^- \longrightarrow Cl^- + ClO_4^-$
11. $HNO_2 \longrightarrow NO + HNO_3$

    λ Reverse of disproportionation is called Comproportionating. In some of the disproportionation reactions, by changing the medium (from acidic to basic or reverse), the reaction goes in backward direction and can be taken as an example of Comproportionating reaction.

    $I^- + IO_3^- + H^+ \longrightarrow I_2 + H_2O$

BALANCING OF REDOX REACTIONS

All balanced equations must satisfy two criteria.

1. **Atom Balance (Mass Balance)**
   There should be the same number of atoms of each kind on reactant and product side.

2. Charge Balance

The sum of actual charges on both sides of the equation must be equal.

There are two methods for balancing the redox equations:

1. Oxidation - number change method
2. Ion electron method or half-cell method

Since First method is not very much fruitful for the balancing of redox reactions, students are advised to use second method (Ion electron method) to balance the redox reactions

Ion electron method: By this method redox equations are balanced in two different medium.

(a) Acidic medium (b) Basic medium

Balancing in Acidic Medium

Students are advised to follow the following steps to balance the redox reactions by Ion electron method in acidic medium

Ex. Balance the following redox reaction:

$$FeSO_4 + KMnO_4 + H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + MnSO_4 + H_2O + K_2SO_4$$

Sol.

Step–I Assign the oxidation number to each element present in the reaction.

$+ + + +$

Step II: Now convert the reaction in Ionic form by eliminating the elements or species, which are not undergoing either oxidation or reduction.

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

Step III: Now identify the oxidation / reduction occuring in the reaction

undergoes reduction

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

undergoes oxidation.

Step IV: Spilt the Ionic reaction in two half, one for oxidation and other for reduction.

$\boxed{Fe^{2+} \xrightarrow{\text{oxidation}} Fe^{3+}}$ | $MnO_4^- \xrightarrow{\text{Reduction}} Mn^{2+}$

Step V: Balance the atom other than oxygen and hydrogen atom in both half reactions

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

Fe & Mn atoms are balanced on both side.

Step VI: Now balance O & H atom by H_2O & H^+ respectively by the following way: For one excess oxygen atom, add one H_2O on the other side and two H^+ on the same side.

$Fe^{2+} \longrightarrow Fe^{3+}$ (no oxygen atom) ....(i)

$8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$ ....(ii)

Step VII: Equation (i) & (ii) are balanced atomwise. Now balance both equations chargewise. To balance the charge, add electrons to the electrically positive side.

$Fe^{2+} \xrightarrow{\text{oxidation}} Fe^{3+} + e^-$ ....(1)

$5e^- + 8H^+ + MnO_4^- \xrightarrow{\text{Reduction}} Mn^{2+} + 4H_2O$ ....(2)

Step VIII: The number of electrons gained and lost in each half -reaction are equalized by multiplying both the half reactions with a suitable factor and finally the half reactions are added to give the overall balanced reaction.

Here, we multiply equation (1) by 5 and (2) by 1 and add them:

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$ ....(1) $\times 5$

$5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O$ ....(2) $\times 1$

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

(Here, at his stage, you will get balanced redox reaction in Ionic form)

Step IX: Now convert the Ionic reaction into molecular form by adding the elements or species, which are removed in step (2).

Now, by some manipulation, you will get:

$$5FeSO_4 + KMnO_4 + 4H_2SO_4 \longrightarrow \frac{5}{2} Fe_2(SO_4)_3 + MnSO_4 + 4H_2O + \frac{1}{2} K_2SO_4$$

or

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 \longrightarrow 5Fe_2(SO_4)_3 + 2MnSO_4 + 8H_2O + K_2SO_4$$

Balancing in Basic Medium

In this case, except step VI, all the steps are same. We can understand it by the following example:

Ex. Balance the following redox reaction in basic medium:

$ClO^- + CrO_2^- + OH^- \longrightarrow Cl^- + CrO_4^{2-} + H_2O$

Sol. By using up to step V, we will get:

$\overset{+1}{ClO^-} \xrightarrow{\text{Reduction}} Cl^-$	$^{+3}CrO_2^{-3} \xrightarrow{\text{Oxidation}} {}^{+6}CrO_4^{2-}$

Now, students are advised to follow step VI to balance 'O' and 'H' atom.

$2H^+ + ClO^- \longrightarrow Cl^- + H_2O$ | $2H_2O + CrO_2- \longrightarrow CrO_4^{2-} + 4H^+$

Now, since we are balancing in basic medium, therefore add as many as OH^- on both side of equation as there are H^+ ions in the equation.

$2OH^- + 2H^+ + ClO^- \longrightarrow Cl^- + H_2O + 2OH^-$

$4OH^- + 2H_2O + CrO_2^- \longrightarrow CrO_4^{2-} + 4H^+ + 4OH^-$

Finally, you will get Finally you will get

$H_2O + ClO^- \longrightarrow Cl^- + 2OH^-$ ....(i)

$4OH^- + CrO_2^- \quad CrO_4{}^{2-} + 2H_2O$ .... (ii)

Now see equation (i) and (ii) in which O and H atoms are balanced by OH^- and H_2O

Now from step VIII

$2e^- + H_2O + ClO^- \longrightarrow Cl^- + 2OH^-$ .... (i) ×3

$4OH^- + CrO_2^- \longrightarrow CrO_4^{2-} + 2H_2O + 3e^-$ .... (ii) ×2

Adding: $3ClO^- + 2CrO_2^- + 2OH^- \longrightarrow 3Cl^- + 2CrO_4^{2-} + H_2O$