REDOX REACTIONS

Term	electron change	Oxidation number change
Oxidation	Loss of electrons	Increases
Oxidizing agent	Takes electrons	Decreases
Reduction	Gain of electrons	Decreases
Reducing agent	Gives electrons	Increases

A Classical Concept of Oxidation and Reduction

A process which involves addition of oxygen or any electronegative element or removal of hydrogen or any electropositive element is called oxidation.

A process which involves addition of hydrogen or any electropositive element or removela of oxygen or any electronegative element is called Reduction.

An Electronic Concept of Oxidation and Reduction

Oxidation is the process which involves loss of electrons.

$$\underset{\text{Sodium atom}}{Na} \xrightarrow{\text{Oxidation}} \underset{\text{Sodium ion}}{Na^+} + \underset{\text{Electron}}{e^-} \quad \underset{\text{Increase in +ve charge}}{[0 \rightarrow +1]}$$

$$\underset{\text{Chloride ion}}{Cl^-} \xrightarrow{\text{Oxidation}} \underset{\text{Chlorine atom}}{Cl} + \underset{\text{Electron}}{e^-} \quad \underset{\text{Increase in +ve charge}}{[-1 \rightarrow 0]}$$

Oxidation reaction can be known as it involves increase in positive charge or decrease in negative charge of the atom.

Reduction is the process which involves gain of electrons.

$$\underset{\text{Ferric ion}}{Fe^{3+}} + \underset{\text{Electrons}}{3e^-} \rightarrow \underset{\text{Iron}}{Fe} \quad \underset{\text{Decrease in +ve charge}}{(+3 \rightarrow 0)}$$

$$\underset{\text{Oxygen}}{O} + \underset{\text{Electrons}}{2e^-} \longrightarrow \underset{\text{Oxide ion}}{O^{2-}} \quad \underset{\text{Decrease in –ve charge}}{(0 \rightarrow -2)}$$

Reduction reaction can be known as it involves increase in negative charge or decrease in positive charge of the atom.

Oxidizing agent : Oxidizing agent oxidizes other substances on the expense of self reduction or it is the agent that gains electrons.

Reducing agent : Reducing agent reduces other substances on the expense of self oxidation or it is the agent that loses electrons.

Oxidizing agent (Gains electrons) $\xrightarrow{\text{Brings}}$ Oxidation (Loss of electrons)

Reducing agent (Loss of electrons) $\xrightarrow{\text{Brings}}$ Reduction (Gains electrons)

Redox reaction : It is the reaction in which oxidation and reduction occur simultaneously. A redox change involves the process in which a reducing agent is oxidized to liberate electrons, which are then taken up by an oxidizing agent to get itself reduced.

Oxidation

Reduction

$$\underset{\substack{\text{Cupric oxide} \\ \text{(Oxidizing agent)}}}{CuO} + \underset{\substack{\text{Hydrogen} \\ \text{(Reducing agent)}}}{H_2} \longrightarrow \underset{\text{Copper}}{Cu} + \underset{\text{Water}}{H_2O}$$

Oxidation Half Reaction and Reduction Half Reaction

Redox reaction is a combination of oxidation of half reaction and reduction of half reaction.

In oxidation of half reaction electrons are released by reducing agent.

$$\underset{\substack{\text{Zinc} \\ \text{(Reducing agent)}}}{Zn\ (s)} \longrightarrow \underset{\text{Zinc ion}}{Zn^{2+}\ (aq)} + \underset{\text{Electrons}}{2e^-}$$

In reduction of half reaction electrons are accepted by oxidizing agent.

$$\underset{\substack{\text{Cupric ion} \\ \text{(Oxidizing agent)}}}{Cu^{2+}\ (aq)} + \underset{\text{Electrons}}{2e^-} \longrightarrow \underset{\text{Cupper}}{Cu\ (s)}$$

Hence, Redox reaction is obtained by adding oxidation of half reaction and reduction of half reaction.

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

(Oxidation of half reaction)

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

(Reduction of half reaction)

$$Cu^{2+}\ (aq) + Zn\ (s) \rightarrow Cu\ (s) + Zn^{2+}\ (aq)$$

(Redox reaction)

Oxidation number is a formal charge on an atom in a compound or ion. Which is given positive sign if electrons are lost and negative sign when electrons are gained.

Oxidation state of an element is the oxidation number per atom.

Rules to calculate oxidation number of an element

I. In elementary state, the atoms are assigned an oxidation number zero.

II. In compounds, oxidation number of
   (a) Hydrogen is +1. In metallic hydrides oxidation number of hydrogen is –1.
   (b) Oxygen is –2. In peroxides it is –1 and in F_2O it is +2.
   (c) Alkali metals is +1 and alkaline earth metals is +2.
   (d) Fluorine is always –1.

III. In neutral molecules, the sum of the oxidation numbers of all the atoms is zero.

IV. The oxidation number of simple ion is equal to the charge on the ion.

V. For charged species, the sum of the oxidation numbers of all the atoms is equal to the charge on the ion.

(a) Calculate the oxidation number of Cr in $K_2Cr_2O_7$.

$K_2Cr_2O_7$: $2 \times (+1) + 2(x) + 7(-2) = 0$

Hence, $x = +6$

Oxidation number of Cr in $K_2Cr_2O_7$ will be +6.

(b) Calculate the oxidation number of Mn in $KMnO_4$.

$KMnO_4 : 1 + x + 4(-2) = 0$

$x = +7$

Oxidation number of Mn in $KMnO_4$ will be +7.

Differences Between Oxidation Number and valency

Oxidation number	Valency
I. It is the formal charge on an atom in its combined state.	I. It is the number representing the combining capacity of an element with other elements.
II. It can be zero.	II. It is never zero except in case of noble gases.
III. It can be positive or negative.	III. It does not carry any charge, hence no plus or minus sign.
IV. It may be whole number or fractional.	IV. It is always a whole number.

Methods used to balance oxidation-reduction equations are :

I. Oxidation number method.

II. Ion electron method.

I. Oxidation number method :

(a) Identify the atoms which show change in oxidation number.

(b) Balance the oxidation number by multiplying it with a coefficient which makes the increase in oxidation number for the oxidized substances equal to the total decrease in oxidation number of reduced substances.

(c) Balance the charge by H^+.

(d) Balance the oxygen atoms by adding water and H^+ ions in acidic medium or by adding H_2O and OH^- ions in basic medium.

(e) Balance the elements which are not oxidized or reduced.

Now Balance the following redox equation :

$$CuO + NH_3 \rightarrow Cu + N_2 + H_2O$$

Step I :
$$\underset{+2}{CuO} + \underset{-3}{NH_3} \rightarrow \underset{0}{Cu} + \underset{0}{N_2} + H_2O$$

Step II: Multiply CuO by 3 and NH_3 by 2.

$$3CuO + 2NH_3 \rightarrow Cu + N_2 + H_2O$$

Step III : Balance oxygen by multiplying H_2O by 3 and put 3 before Cu.

$$3CuO + 2NH_3 \rightarrow 3Cu + N_2 + 3H_2O$$

II. Ion electron method

(a) Select the oxidant and reductant.

(b) Write the oxidation half reaction and balance the atoms of other atoms (except H and O), O atoms (by adding water), H atoms (by H^+) and also charge (by electrons).

(c) Proceed in similar manner with reduction half reaction.

(d) Multiply these reactions with suitable co-efficient so as to balance electrons appearing on either side of these reactions.

(e) Add these two half reactions.

(f) If the reaction is taking place in basic solution, add enough OH^- on both sides of the half reaction to get rid of H^+. Combine H^+ and OH^- to give H_2O and cancel any duplication.

Now, Balance fhe following redox equation.

$Cr_2O_7^{-2} + C_2O_4^{2-} \rightarrow Cr^{3+} + CO_2$ in acidic medium

Step I : Oxidation half reaction

$C_2\,O_4^{2-} \;\rightarrow\; 2\,CO_2 + 2e^-$

Step II : Reduction half reaction

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

Step III : Add the oxidation half reaction and reduction half reaction

$$
\begin{array}{rcl}
Cr_2\,O_7^{2-} + 14\,H^+ + 6e^- & \rightarrow & 2\,Cr^{3+} + 7H_2O \\
C_2O_4^{2-} & \rightarrow & 2\,CO_2 + 2e^-] \times 3 \\
\hline
Cr_2O_7^{2-} + 14\,H^+ + 3C_2O_4^{2-} & \rightarrow & 2Cr^{3+} + 7H_2O + 6CO_2
\end{array}
$$

Oxidation Number and Nomenclature

Stock system of nomenclature : In this system, oxidation number of the element is indicated by a Roman numeral written in the bracket immediately after the name of the element.

Compound	Oxidation Number	Name of the compound: According to stock system	Name of the compound: Acording to old method
$FeSO_4$	Fe–2	Iron (II) sulphate	Ferrous Sulphate
$Fe_2(SO_4)_3$	Fe–3	Iron (III) sulphate	Ferric Sulphate
Cu_2O	Cu–1	Copper (I) oxide	Cuprous oxide
Cr_2O_3	Cr–3	Chromium (III) oxide	Chromium trioxide
$K_2Cr_2O_7$	Cr–6	Potassium dichromate (VI)	Potassium dichromate

Mn_2O_7	Mn–7	Manganese (VII) oxide	Manganese heptoxide
$SnCl_2$	Sn–2	Tin (II) chloride	stannous chloride
$SnCl_4$	Sn–4	Tin (IV) chloride	stannic chloride

Types of cells : Their differences

Electrochemical cell (or galvanic cell or valtaic cell)	Electrolytic cell
I. It ia a device used to convert chemical energy into electrical energy	I. It ia device used to convert electrical energy into chemical energy.
II. Oxidation occurs at anode (negative terminal) and reduction occurs at cathode (positive terminal).	II. Oxidation occurs at cathode (negtive) and reduction occurs at anode (positive).
III. The two electrodes are placed in different containers called half cells, connected through salt bridge.	III. Both the electrodes are placed in the same container.
IV. Work is obtained from the cell.	IV. Work is done on the cell.
V. Free energy decreases as the cell reaction proceeds.	V. Free energy increases as the cell reaction proceeds.
VI. Cell reaction is spontaneous.	VI. Cell reaction is non-spontaneous.