

Redox Reactions

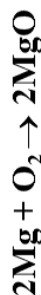
An equation having only one variable with highest power 1 is called a linear equation in one variable.

MIND MAP

Redox

Oxidation

(1) Addition of oxygen



(2) Removal of Hydrogen



(3) Increase in positive charge



(4) Increase in oxidation number



(5) Removal of electron

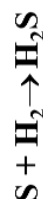


Reduction

(1) Removal of oxygen



(2) Addition of Hydrogen



(3) Decrease in positive charge



(4) Decrease in oxidation number



(5) Addition of electron



Oxidation Number

It is an imaginary or apparent charge gained by an element when it goes from its elemental free state to combined state in molecules.

Fluorine has oxidation no. equal to -1 in all its compounds.

Alkali metal (Li, Na, K, Rb,) always have oxidation number $+1$.

Alkaline earth metal (Be, Mg, Ca) always have oxidation number $+2$.

Aluminium always has $+3$ oxidation number.

In general, H atom has oxidation number equal to $+1$. But in metallic hydrides (e.g. NaH, KH) it is -1 .

In general and as well as in its oxides, oxygen atom has oxidation number equal to -2 .

In Case of

(i) peroxide (e.g. H_2O_2 , Na_2O_2) is -1

(ii) super oxide (e.g. KO_2) is $-\frac{1}{2}$

(iii) ozonide (KO_3) is $-\frac{1}{3}$

(iv) oxygen fluoride OF_2 is $+2$ & in O_2F_2 is $+1$

In general, all halogen atoms (Cl, Br, I) have oxidation

number equal to -1 . But if

halogen atom is attached with

a more electronegative atom

than halogen atom then it will

show positive oxidation

number.



Oxidation number of an

element in free state or in

allotropic forms is always

zero.

e.g. $\text{O}_2, \text{S}_8, \text{P}_4, \text{O}_3$

Oxidising Agent or Oxidant

Oxidising agents are those compound which oxidise others and reduce themselves during the chemical reaction.

e.g. $\text{KMnO}_4, \text{K}_2\text{Cr}_2\text{O}_7, \text{HNO}_3$, conc.

H_2SO_4 etc. are powerful oxidising agents.

Reducing Agent or Reductant

Reducing agents are those compound which can reduce others and oxidise itself during the chemical reaction.

e.g. $\text{KI}, \text{Na}_2\text{S}_2\text{O}_3$ are powerful reducing agents.

How To Identify Whether A Particular Substance Is An Oxidising or Reducing Agent

MIND MAP

Redox

If O.S. is = Maximum If O.S. = Minimum



It's an Oxidizing agent



It's a reducing agent

If O.S. is intermediate b/w max & minimum



It can act both as reducing agent & oxidising agent

Disproportionation Reactions

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



Comproportionation Reactions

Reaction in which a element from two different oxidation state gets converted into a single oxidation state.



Balancing of Redox Reaction

(A) Oxidation Number Change Method

- (i) Select the atom in oxidising agent whose oxidation number decreases and indicate the gain of electrons.

(B) Ion Electron Method or Half Cell Method

Balancing in Acidic medium

Assign the oxidation number to each element present in the reaction.

Now convert the reaction in ionic form by eliminating the elements or species which are not getting either oxidised or reduced.

Now identify the oxidation / reduction occurring in the reaction.

Split the ionic reaction in two half, one for oxidation and other for reduction.

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

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.

Now balance both equations charge wise.

To balance the charge, add electrons to the electrically positive side.

The number of electrons gained and lost in each half-reaction are equalised by multiplying suitable factor in both the half reaction and finally the half reactions are added to give the over all balanced reaction

Balancing In Basic Medium

Same as in Acidic Medium Add OH⁻ both sides where there is H⁺

Eq. wt. of oxidising / reducing agents in redox reaction

(a) Equivalent wt. of an oxidant

Mol. wt.

$$= \frac{\text{No. of electrons gained by one mole}}{\text{Eq. wt.}}$$

(b) Similarly equivalent wt. of a reductant

Mol. wt.

$$= \frac{\text{No. of electrons lost by one mole}}{\text{Eq. wt.}}$$

Some Oxidizing Agents/Reducing Agents With Eq. Wt.

Species	Changed to	Reaction	Change in O.N.	Eq. wt.
MnO_4^- (O.A.)	Mn^{+2} in acidic medium	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	5	$\frac{M}{5}$
MnO_4^- (O.A.)	MnO_2 in neutral medium	$\text{MnO}_4^- + 3\text{e}^- + 4\text{H}^+ \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}$	3	$\frac{M}{5}$
MnO_4^- (O.A.)	MnO_4^{2-} in basic medium	$\text{MnO}_4^- + \text{e}^- \rightarrow \text{MnO}_4^{2-}$	1	$\frac{M}{1}$

MIND MAP

Redox

Species	Changed to	Reaction	Change in O.N.	Eq. wt.
$\text{Cr}_2\text{O}_7^{2-}$ (O.A.)	Cr^{3+} in acidic medium	$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	6	$\frac{M}{6}$
MnO_2 (O.A.)	Mn^{2+} in acidic medium	$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	2	$\frac{M}{2}$
Cl_2 (O.A.) in bleaching powder	Cl^-	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	2	$\frac{M}{2}$
CuSO_4 (O.A.) in iodometric titration	Cu^+	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	1	$\frac{M}{1}$

Species	Changed to	Reaction	Change in O.N.	Eq. wt.
$\text{S}_2\text{O}_3^{2-}$ (R.A.)	$\text{S}_2\text{O}_6^{2-}$	$2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^-$	2	$\frac{2M}{2} = M$
H_2O_2 (O.A.)	H_2O	$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	2	$\frac{M}{2}$
H_2O_2 (R.A.)	O_2	$\text{H}_2\text{O}_2 \rightarrow \text{O}_2 + 2\text{H}^+ + 2\text{e}^-$	2	$\frac{M}{2}$
Fe^{2+} (R.A.)	Fe^{3+}	$\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$	1	$\frac{M}{1}$