MIND MAP

Redox

Redox Reactions

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

Oxidation

(1) Addition of oxygen

$2Mg + O_2 \rightarrow 2MgO$

(2) Removal of Hydrogen

$H_2S + Cl_2 \rightarrow 2HCl + S$

(3) Increase in positive charge

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

(4) Increase in oxidation number

$\overset{(+2)}{SnCl_2} \longrightarrow \overset{(+4)}{SnCl_4}$

(5) Removal of electron

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

Reduction

(1) Removal of oxygen

$CuO + C \rightarrow Cu + CO$

(2) Addition of Hydrogen

$S + H_2 \rightarrow H_2S$

(3) Decrease in positive charge

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

(4) Decrease in oxidation number

$\overset{(+7)}{MnO_4^-} \longrightarrow \overset{(+2)}{Mn^{2+}}$

(5) Addition of electron

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

In Case of

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

(ii) super oxide (e.g. KO_2) is – $^1/_2$

(iii) ozonide (KO_3) is – $^1/_3$

(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.

e.g. $K\overset{+5}{Cl}O_3$ $H\overset{+5}{I}O_3$ $H\overset{+7}{Cl}O_4$ $K\overset{+5}{Br}O_3$

Oxidation number of an element in free state or in allotropic forms is always zero.

e.g. $\overset{0}{O_2}, \overset{0}{S_8}, \overset{0}{P_4}, \overset{0}{O_3}$

Oxidising Agent or Oxidant

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

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

Reducing Agent or Reductant

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

e.g. KI, $Na_2S_2O_3$ are powerful reducing agents.

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.

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

- If O.S. is = Maximum → It's an Oxidizing agent
- If O.S. = Minimum → 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.

$$2H_2O_2(aq) \longrightarrow H_2O(\ell) + O_2(g)$$

Comproportionation Reactions

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

$$NH_4NO_2 \longrightarrow N_2 + 2H_2O$$

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.

(ii) Select the atom in reducing agent whose oxidation number increases and write the loss of electrons.

(iii) Now cross multiply i.e. multiply oxidising agent by the number of loss of electrons and reducing agent by number of gain of electrons.

(iv) Balance the number of atoms on both sides whose oxidation number change in the reaction.

(v) In order to balance oxygen atoms, add H_2O molecules to the side deficient in oxygen. Then balance the number of H atoms by adding H+ ions in the hydrogen.

(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

MIND MAP

Redox

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

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

(b) Similarly equivalent wt. of a reductant

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

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	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	5	$E = \frac{M}{5}$
MnO_4^- (O.A.)	MnO_2 in neutral medium	$MnO_4^- + 3e^- + 4H^+ \rightarrow MnO_2 + 2H_2O$	3	$E = \frac{M}{5}$
MnO_4^- (O.A.)	MnO_4^{2-} in basic medium	$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	1	$E = \frac{M}{1}$

Species	Changed to	Reaction	Change in O.N.	Eq. wt.
$Cr_2O_7^{2-}$ (O.A.)	Cr^{3+} in acidic medium	$CrO_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^3 + 7H_2O$	6	$E = \frac{M}{6}$
MnO_2 (O.A.)	Mn^{2+} in acidic medium	$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^2 + 2H_2O$	2	$E = \frac{M}{2}$
Cl_2 (O.A.) in bleaching powder	Cl^-	$Cl_2 + 2e^- \rightarrow 2Cl^-$	2	$E = \frac{M}{2}$
$CuSO_4$ (O.A.) in iodometric titration	Cu^+	$Cu^{2+} + e^- \rightarrow Cu^+$	1	$E = \frac{M}{1}$

Species	Changed to	Reaction	Change in O.N.	Eq. wt.
$S_2O_3^{2-}$ (R.A.)	$S_2O_6^{2-}$	$2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$	2	$E = \frac{2M}{2} = M$
H_2O_2 (O.A.)	H_2O	$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	2	$E = \frac{M}{2}$
H_2O_2 (R.A.)	O_2	$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$	2	$E = \frac{M}{2}$
Fe^{2+} (R.A.)	Fe^{3+}	$Fe^{2+} \rightarrow Fe^{3+} + e^-$	1	$E = \frac{M}{1}$