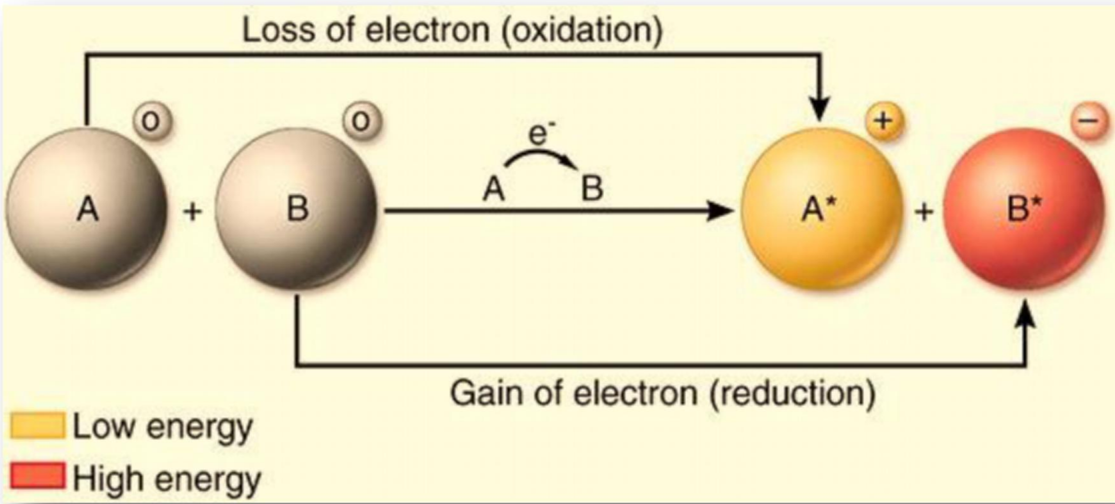
Revision Notes on Redox Reactions

Oxidation and Reduction

- Oxidation is defined as the loss of electrons by a chemical species (atom, ion or molecule).
- Reduction is the gain of electrons by a chemical species (atom, ion or molecule).
- An oxidising agent that chemical species which takes electrons thus it is an electron acceptor.
- A reducing agent is the chemical species that gives electrons and thus acts as an electron donor.

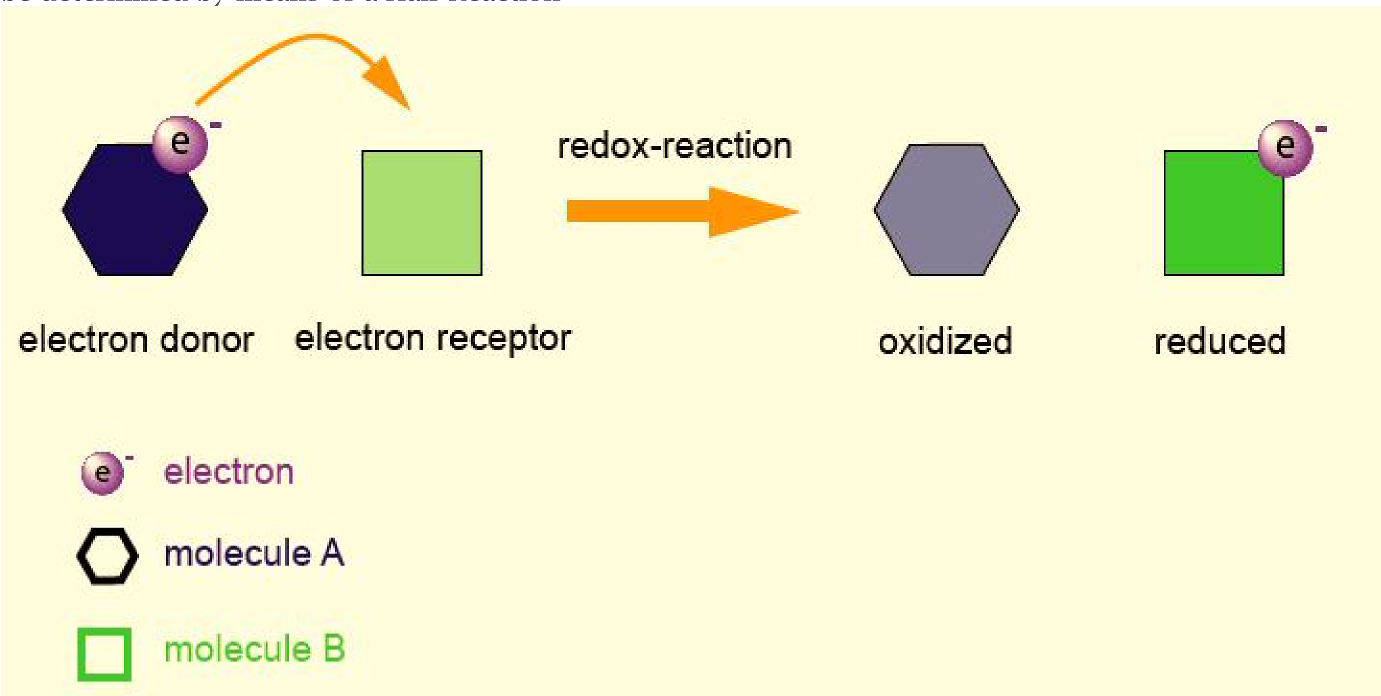


- When $Fe^{2+}(aq)$ ions are being oxidised they are acting as reducing agents, and when $Fe^{3+}(aq)$ ions are being reduced they are acting as oxidising agents. In general ;

$$\text{O.A.} + \text{electrons} \underset{\text{Oxidation}}{\overset{\text{Reduction}}{\rightleftharpoons}} \text{R.A.}$$

$$\underset{\text{(R.A.)}}{Fe^{2+}_{(aq)}} \xrightarrow{\text{Oxidation}} \underset{\text{(O.A.)}}{Fe^{3+}_{(aq)}} + e^- \qquad Fe^{3+}_{(aq)} + e^- \xrightarrow{\text{reduction}} Fe^{2+}_{(aq)}$$

- Neither reduction nor oxidation occurs alone. Both of them occur simultaneously. Since both these reactions must occur at the same time they are often termed as "redox reactions". The oxidation or reduction portion of a redox reaction, including the electrons gained or lost can be determined by means of a Half-Reaction



Oxidation Number

- The oxidation number for an element in a covalent compound is by taking the oxidation number to be equal to the charge that the element would carry, if all the bonds in the compound were regarded as ionic instead of covalent.
- The sum of the oxidation numbers of all the atoms in an uncharged compound is zero. In case of an ion, the algebraic sum of the oxidation numbers of all the atoms is equal to the charge on the ion.
- Oxidation number of any element in its elementary state is zero.
- Fluorine is the most electronegative element. Its oxidation number is always -1 .
- Oxygen after fluorine is the second most electronegative element. It shows an oxidation state of -2 in almost all the compounds excepts peroxides and superoxides,
- In all compounds, except ionic metallic hydrides, the oxidation number of hydrogen is +1.
- In any compounds has more than two elements, the oxidation number of any one of them may have to be obtained by first assigning reasonable oxidation numbers to the other elements.

Oxidation Number / State Method For Balancing Redox Reactions:

This method is based on the principle that the number of electrons lost in oxidation must be equal to the number of electrons gained in reduction. The steps to be followed are :

- Write the equation (if it is not complete, then complete it) representing the chemical changes.
- By knowing oxidation numbers of elements, identify which atom(s) is(are) undergoing oxidation and reduction. Write down separate equations for oxidation and reduction.
- Add respective electrons on the right of oxidation reaction and on the left of reduction reaction. Care must be taken to ensure that the net charge on both the sides of the equation is same.
- Multiply the oxidation and reduction reactions by suitable numbers to make the number of electrons lost in oxidation reactions equal to the number of electrons gained in reduction reactions.
- Transfer the coefficient of the oxidizing and reducing agents and their products to the main equation.

- By inspection, arrive at the co-efficients of the species not undergoing oxidation or reduction.
- Some elements like manganese, chromium, nitrogen and chlorine show variable oxidation states.
- When an element is oxidised its oxidation number gets increased while reduction on any element decreases its oxidation number. Change in oxidation number can be used to decide whether an oxidation or a reduction has taken place. In the change from chloromethane to dichloromethane, for example,
 C H$_3$ Cl C H$_2$ Cl$_2$
 -2 +1 -1 0 +1 -1
 The oxidation number of carbon is increased from -2 to 0. The carbon is therefore being oxidised.

Half-Reaction or Ion-Electron Method For Balancing Redox Reactions

This method involves the following steps :

- Divide the complete equation into two half reactions, one representing oxidation and the other reduction.
- Balance the atoms in each half reaction separately according to the following steps:
  - First of all balance the atoms other than H and O.
  - In a reaction taking place in acidic or neutral medium, oxygen atoms are balanced by adding molecules of water to the side deficient in oxygen atoms while hydrogen atoms are balanced by adding H^+ ions to the other side deficient in hydrogen atoms. On the other hand, in alkaline medium (OH^-), every excess of oxygen atom on one side is balanced by adding one H_2O to the same side and $2OH^-$ to the other side. In case hydrogen is still unbalanced, then balance by adding one OH^-, for every excess of H atom on the same side as the excess and one H_2O on the other side.
  - Equalize the charge on both sides by adding a suitable number of electrons to the side deficient in negative charge.
  - Multiply the two half reactions by suitable integers so that the total number of electrons gained in one half reaction is equal to the number of electrons lost in the other half reaction.

Common Oxidising and Reducing Agents

Oxidising agent	Effective Change	Decrease in Oxidation Number
$KMnO_4$ in acid solution	$MnO_4^- \rightarrow Mn^{2+}$	5
$KMnO_4$ in alkaline solution	$MnO_4^- \rightarrow MnO_2$	3
$K_2Cr_2O_7$ in acid solution	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$	3
dilute HNO_3	$NO_3^- \rightarrow NO$	3
concentrated HNO_3	$NO_3^- \rightarrow NO_2$	1
concentrated H_2SO_4	$SO_4^{2-} \rightarrow SO_2$	2
manganese (IV) oxide	$MnO_2 \rightarrow Mn^{2+}$	2
chlorine	$Cl \rightarrow Cl^-$	1
chloric (I) acid	ClO^-	2

	$\rightarrow Cl^-$	
KlO_3 in dilute acid	$IO_3^- \rightarrow I$	5
KlO_3 in concentrated acid	$IO_3^- \rightarrow I^-$	4
Reducing agent	**Effective Change**	**Increase in Oxidation Number**
iron (II) salts (acid)	$Fe^{2+} \rightarrow Fe^{3+}$	1
tin (II) salts (acid)`	$Sn^{2+} \rightarrow Sn^{4+}$	2
ethanedioates (acid)	$C_2O_4^{2-} \rightarrow CO_2$	1
sulphites (acid)	$SO_3^{2-} \rightarrow SO_4^{2-}$	2
hydrogen sulphide	$S^{2-} \rightarrow S$	2
iodides (dilute acid)	$I^- \rightarrow I$	1
iodides (concentrated acid)	$I^- \rightarrow I^+$	2
metals, e.g. Zn	$Zn \rightarrow Zn^{2+}$	2
hydrogen		