Crash Course for NEET 2020

KEY NOTES ON Redox Reactions

Biomentors Classes Online, Mumbai

NCERT Based - Very Important Points

Oxidation

Oxidation may be defined in any of the following terms

i. Addition of oxygen.
$2\,Mg + O_2 \longrightarrow 2MgO$
ii. Removal of hydrogen
$3O_2 + 4NH_3 \longrightarrow 2N_2 + 6H_2O$
iii. Addition of electronegative element
$Cu + Cl_2 \longrightarrow CuCl_2$
iv. Removal or decrease in the electropositive element
$H_2S + Cl_2 \longrightarrow 2HCl + S$
v. De-electronation $M \longrightarrow M^{n+} + ne^-$

Reduction

Reduction may be defined in any of the following terms

i. Addition of hydrogen
$N_2 + 3H_2 \longrightarrow 2NH_3$
ii. Addition of electropositive element
$CuCl_2 + Cu \longrightarrow Cu_2Cl_2$
iii. Removal of oxygen
$CuO + H_2 \longrightarrow Cu + H_2O$
iv. Removal or decrease in the electronegative element
$2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$
v. Electronation $M + ne^- \longrightarrow M^{n-}$

Oxidant or Oxidising Agent

As stated above the oxidising agent may be defined as a substance supplying oxygen or electronegative element, removing hydrogen or electropositive element and can accept electrons. They show decrease in oxidation number examples.

$K_2Cr_2O_7,\ KMnO_4,\ H_2O_2, Cl_2, Br_2,\ KClO_3,\ FeCl_3$ etc.

Reductant or Reducing Agent

A substance supplying hydrogen or electropositive element, removing oxygen or electro negative element and can donate electrons. They show increase in oxidation number.

Examples $SnCl_2, H_2,\ H_2S,\ Mg,\ FeSO_4,\ H_2C_2O_4\ \ H_2SO_3.$

Oxidation number

It is the number of electrons lost or gained by an element during its change from free state in a particular compounds.

Or

It is defined as the formal charge present on an atom in a particular compound determined by certain arbitrary rules.

Rules for Determining Oxidation Number

i. O.N. of elements in free state is zero eg. Cl_2N_2Mg, Ca
ii. O.N. of hydrogen is always +1 except in ionic metal hydrides where it is -1.
iii. O.N. of oxygen is – 2 except in OF_2 where it is + 2 and in peroxides where it is - 1.
iv. O.N. of metals is always $+\,ve$. For IA group elements it is + 1 and for IIA group elements it is + 2
v. O.N. of halogens is – 1 in metal halides
vi. O.N of ion or radical is the number of electrons it must gain or lose to acquire neutrality i.e it is equal to the electric charge for $SO_4^{2-}\ O.N.is-2.$
vii. O.N of an atom withing compound can be $+ve, -ve$ integer, zero or fraction
viii. The algebraic sum of all the O.N. of elements is equal to zero.
ix. The algebraic sun of all the O.N of elements in an ions to equal to net charge on the ion.
x. Maximum O.N/ of an elements equal to number of valence electrons i.e. group number.
xi. Minimum O.N. of an element (except metals) $= (8 - group\ member)$.
xii. In metal carbonyl, and amalgams, O.N. of metals is zero.

COVALENCY AND OXIDATION STATE

i. **Covalency :** it is the number if hydrogen atoms which can combine with a given atom
Or
It is the number of single bonds which an atom can form.
Or
It is the number of electrons an atom can share. Valency is always a whole number.

ii. **Oxidation state :** it is defined as the O.N. per atom.

Calculation / Determination of Oxidation Number of Underlined Element in Some Compounds

i. $K_2\underline{Cr_2}\,O_7$

Let the O.N. of Cr be x then

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

$2 + 2x - 14 = 0 \qquad \therefore x = +6$

ii. $K\underline{Mn}O_4$

Let the O.N. of Mn be x then

$1 \times (+1) + 1 \times (x) + 4 \times (-2) = 0$

$1 + 1x - 8 = 0 \qquad \therefore x = +7$

iii. $H_2\underline{S}O_4$

Let the O.N. of S be x then

$2 \times (+1) + 1 \times (x) + 4 \times (-2) = 0$

$2 + x - 8 = 0 \qquad \therefore x = +6$

Types of redox reactions

1. Combination reactions:

A combination reaction may be denoted in the manner: $A + B \rightarrow C$

All combustion reactions, which make use of elemental dioxygen, as well as other reactions involving elements other than dioxygen, are redox reactions.

$$\overset{0}{C}(s) + \overset{0}{O_2}(g) \xrightarrow{\Delta} \overset{+4\ -2}{CO_2}(g)$$

$$3\overset{0}{Mg}(s) + \overset{0}{N_2}(g) \xrightarrow{\Delta} \overset{+2\ -3}{Mg_3N_2}(s)$$

2. Decomposition reactions

Decomposition reactions are the opposite of combination reactions. A decomposition reaction leads to the breakdown of a compound into two or more components at least one of which must be in the elemental state

$$2\overset{+1\ -2}{H_2O}\,(l) \xrightarrow{\Delta} 2\overset{0}{H_2}\,(g) + \overset{0}{O_2}(g)$$

$$2\overset{+1\ -1}{NaH}\,(s) \xrightarrow{\Delta} 2\overset{0}{Na}\,(s) + \overset{0}{H_2}(g)$$

3. Displacement reactions

In a displacement reaction, an ion (or an atom) in a compound is replaced by an ion (or an atom) of another element. It may be denoted as: $X + YZ \rightarrow XZ + Y$

4. Disproportionation reactions:

In a disproportionation reaction an element in one oxidation state is simultaneously oxidised and reduced.

$$\overset{+1\ -1}{2H_2O_2}\,(aq) \rightarrow \overset{+1\ -2}{2H_2O}(l) + \overset{0}{O_2}(g)$$

BALANCING OF REDOX REACTIONS

For balancing a chemical equation the two important methods are

Oxidation Number Method – The certain rules are as follows

(i) Assign oxidation number to the atoms showing a change in oxidation state.
(ii) Balance the total number of atoms undergoing change in oxidation state.
(iii) Balance the number of electrons gained and lost.
(iv) Balance [O] on both sides by adding H_2O.
(v) Balance H atoms by adding H^+ ions
(vi) If the reaction proceeds in basic solution and sufficient number of OH^- ions on both sides.

Example 1 :

Balance the equation involving oxidation of ammonia by copper oxide to give Cu, N_2 and H_2O

Solution : $\underset{\substack{\uparrow \\ 3\times 2 = 6e^-}}{\overset{+2}{CuO}} \quad + \quad \underset{\substack{\downarrow \\ 6e^-}}{\overset{-3}{2NH_3}} \longrightarrow Cu^0 + N_2^0 + H_2O$

$\therefore 3CuO + 2NH_3 \longrightarrow 3Cu + N_2 + H_2O$

Balance O by adding $H_2O\ to\ RHS$

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

Example 2 :

The reduction of permanganate ion by ferrous ion in presence of a dilute acid

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

Solution :

$$\underset{\substack{\uparrow \\ 5\bar{e}}}{\overset{+7}{MnO_4^-}} \quad + \quad \underset{\substack{\downarrow \\ 5\times 1\bar{e}}}{Fe^{2+}} \quad + H^+ \longrightarrow Mn^2 + Fe^{3+} + H_2O$$

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

Balance O and H by adding H_2O and H^+ ; $3H_2O$ on RHS and $7H^+$ on LHS

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

ION ELECTRON METHOD – The rules are as follows :

i. Split up the reaction into two half reactions showing oxidation and reduction separately.
ii. Balance number of atoms undergoing the change oxidation state.
iii. Balance O on both sides by adding H_2O.
iv. Balance H atoms by adding H^+ ions.
v. Balance charge by adding required number if electrons
vi. Make the number of electrons equal in two half reactions by multiplying with suitable coefficient.
vii. Add the two half reactions

Example 1

Oxidation of ferrous salt by potassium dichromate in acid solution.

$$Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + Fe^{3+} + H_2$$

Solution – (i) Reduction half reactions

$$Cr_2O_7^{2-} + H^+ \longrightarrow Cr^{3+} + H_2O$$

Equalize Cr atoms

$$Cr_2O_7^{2-} + H^+ \longrightarrow 2Cr^{3+} + H_2$$

Balance O and H atoms on both sides by adding H_2O and H^+ ions.

$$Cr_2O_7^{2-} + H^+ \longrightarrow 2Cr^{3+} + H_2O$$

$$+13H^+ \qquad\qquad 6H_2O$$

Balance charge on both sides by adding electrons.

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

(ii) Oxidation half-reaction

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

Balance electrons of two half reactions

$$6Fe^{2+} - 6e^- \longrightarrow 6Fe^{3+}$$

Adding two half reaction (electrons are cancelled)

$$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

Electrochemical Series

When elements are arranged in increasing order of standard electrode reduction potential (or decreasing order of standard electrode oxidation potential) as compared to that of standard hydrogen electrode, it is called electrochemical series (or electromotive force series)

Reaction (Oxidised form + ne^-	$\rightarrow$ Reduced form)	$E^{\ominus}$ / V
$F_2(g) + 2e^-$	$\rightarrow 2F^-$	2.87
$Co^{3+} + e^-$	$\rightarrow Co^{2+}$	1.81
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow 2H_2O$	1.78
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow Mn^{2+} + 4H_2O$	1.51
$Au^{3+} + 3e^-$	$\rightarrow Au(s)$	1.40
$Cl_2(g) + 2e^-$	$\rightarrow 2Cl^-$	1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2(g) + 4H^+ + 4e^-$	$\rightarrow 2H_2O$	1.23
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow Mn^{2+} + 2H_2O$	1.23
$Br_2 + 2e^-$	$\rightarrow 2Br^-$	1.09
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow NO(g) + 2H_2O$	0.97
$2Hg^{2+} + 2e^-$	$\rightarrow Hg_2^{2+}$	0.92
$Ag^+ + e^-$	$\rightarrow Ag(s)$	0.80
$Fe^{3+} + e^-$	$\rightarrow Fe^{2+}$	0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow H_2O_2$	0.68
$I_2(s) + 2e^-$	$\rightarrow 2I$	0.54
$Cu^+ + e^-$	$\rightarrow Cu(s)$	0.52
$Cu^{2+} + 2e^-$	$\rightarrow Cu(s)$	0.34
$AgCl(s) + e^-$	$\rightarrow Ag(s) + Cl^-$	0.22
$AgBr(s) + e^-$	$\rightarrow Ag(s) + Br^-$	0.10
$\mathbf{2H^+ + 2e^-}$	$\rightarrow \mathbf{H_2(g)}$	**0.00**
$Pb^{2+} + 2e^-$	$\rightarrow Pb(s)$	–0.13
$Sn^{2+} + 2e^-$	$\rightarrow Sn(s)$	–0.14
$Ni^{2+} + 2e^-$	$\rightarrow Ni(s)$	–0.25
$Fe^{2+} + 2e$	$\rightarrow Fe(s)$	–0.44
$Cr^{3+} + 3e^-$	$\rightarrow Cr(s)$	–0.74
$Zn^{2+} + 2e^-$	$\rightarrow Zn(s)$	–0.76
$2H_2O + 2e^-$	$\rightarrow H_2(g) + 2OH^-$	–0.83
$Al^{3+} + 3e^-$	$\rightarrow Al(s)$	–1.66
$Mg^{2+} + 2e^-$	$\rightarrow Mg(s)$	–2.36
$Na^+ + e^-$	$\rightarrow Na(s)$	–2.71
$Ca^{2+} + 2e^-$	$\rightarrow Ca(s)$	–2.87
$K^+ + e^-$	$\rightarrow K(s)$	–2.93
$Li^+ + e^-$	$\rightarrow Li(s)$	–3.05

Increasing strength of oxidising agent

Increasing strength of reducing agent

A negative E° means that the redox couple is a stronger reducing agent than the H^+/H_2 couple. A positive E° means that the redox couple is a weaker reducing agent than the H^+/H_2 couple.