Class 11, Chapter- 8: Redox Reactions

1. Oxidation: In oxidation reactions there is loss of electrons or increase in positive charge or decrease in negative charge.

2. Reduction: In reduction reactions there is gain of electrons or decrease in positive charge or increase in negative charge (or a decrease in the oxidation number)

3. Oxidant or Oxidising Agent: Oxidant or oxidising agent is a chemical substance which can accept or gain one or more electrons.
Important Oxidants are Molecules of most electronegative elements such as O_2, O_3, halogens. Compounds having element in its highest oxidation state e.g., $K_2Cr_2O_7$, $KMnO_4$, $HClO_4$, H_2SO_4, $KClO_3$, $Ce(SO4)_2$, Oxides of metals and non metals like MgO, CrO_3, CO_2, etc.

4. Reductants or reducing agent: is a chemical substance which loss one or more electrons and causes reduction of some other species.
Important Reductants are all metals such as Na, AI, Zn, etc., some non metals, e.g. C, S. P, H_2, etc. and Metallic hydrides like NaH, LiH. KH, CaH_2 etc.

5. Difference b/w Oxidation and Reduction:

Sr. No.	Oxidation	Reduction
1	Loss of electron	Gain of electron
2	Addition of oxygen; $2Mg + O_2 = MgO$	Removal of oxygen; $Fe_2O_3 + 3CO = 2Fe + 3CO_2$
3	Removal of hydrogen $2HI = H_2 + I_2$	Addition of hydrogen; $H_2S + Cl_2 = 2HCl + S$
4	Addition of an electronegative element	Removal of an electronegative element
	$Zn + S = ZnS$	$H_2 + Br_2 = 2HBr$
5	Removal of an electropositive element	Addition of an electropositive element

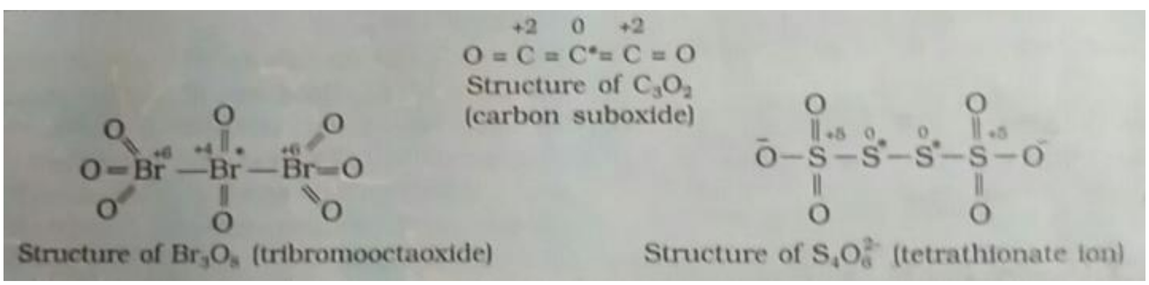
6. Oxidation Number: An oxidation number is a number that is assigned to an atom in a substance. The oxidation number could be positive, negative or zero, and it indicates if electron are loss or gained.
Rules for find out Oxidation Number: Rules for assigning oxidation number to an atom

- Oxidation number of Hydrogen is always +1 (except in hydrides, LiH, CaH_2 it is -1).
- Oxidation number of oxygen in most of compounds is -2. In peroxides it is (-1) In superoxide's, it is (-1/2). In OF_2 it is +2 and in O_2F_2 oxidation number of oxygen is +1
- Oxidation number of Fluorine is -1 in all its compounds
- Alkali metals (1A group) and alkali earth metals (2A group), oxidation number are always +1 and +2.
- For neutral molecules sum of oxidation number of all atoms is equal to zero like H_2O, NaOH.

- In the free or elementary state, the oxidation number of an atom is always zero like Cl_2, O_2, Ca
- For ions composed of only one atom, the oxidation number is equal to the charge on the ion like in $F^- = -1$, $O^+ = +1$
- The algebraic sum of the oxidation number of all the atoms in a compound must be zero
- For ions the sum of oxidation number is equal to the charge on the ion like $Cr_2O_7^{-2} = -2$, $NO_3^- = -1$
- In a polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must be equal to the charge on the ion
- Oxidation no. of element cannot be more than his no. of valence electrons.

7. Fractional Oxidation Number: The oxidation number is usually in whole numbers but in some cases they are in fractional numbers which is unconvincing to us because electrons are never share or transferred in fraction. Actually this fractional oxidation state is the average oxidation state of element like C_3O_2 (Avg. Oxidation No of C = 4/3 means 2 + 0 + 2), S_4O_6 (Oxidation No of S = 10/4 means 5 + 0 + 0 + 5), Fe_3O_4 (Oxidation No of Fe =8/3 means 3+3+2) etc

Note: When covalent bond is form b/w same atom in compound, then both the atoms acquire the oxidation state of zero



Reference Exercise: NCERT book pg- 272, Exercise No.8.2

8. Stock notation: the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula. Thus aurous chloride and auric chloride are written as Au (I) Cl and Au(III)Cl_3. Similarly, stannous chloride and stannic chloride is written as Sn (II) Cl_2 and Sn(IV)Cl_4.

Formula of the compound	Chemical name	Stock notation
Cu_2O	Cuprous oxide	Copper (I) oxide; Cu_2(I)O
Fe_2O_3	Ferric oxide	Iron (III) oxide; Fe_2(III)O_3
$HgCl_2$	Mercuric chloride	Mercury (II) chloride; Hg(II) Cl_2

Reference Exercise: NCERT book pg- 261 Example- 8.3, pg- 272, Exercise No.8.6

9. Redox Reaction: Redox reactions are those reactions in which oxidation and reduction takes place simultaneously.

Types of Redox Reactions:

a. Combination Reactions: Chemical reactions in which two or more substances (elements or compounds) combine to form a single substance like X + Y = Z

$$\overset{0}{C} + \overset{0}{O_2} \rightarrow \overset{+4\ -2}{CO_2}$$

$$\overset{0}{Cl_2} + \overset{+3\ -1}{PCl_3} \rightarrow \overset{+5\ -1}{PCl_5}$$

$$\overset{-4\ +1}{CH_4} + \overset{0}{O_2} \rightarrow \overset{+4\ -2}{CO_2} + \overset{+1\ -2}{H_2O}$$

b. Decomposition Reactions: Chemical reactions in which a compound break up into two or more simple substances

XY = X + Y

$$\overset{+1\ -2}{2H_2O} \rightarrow \overset{0}{2H_2} + \overset{0}{O_2}$$

Decomposition of CaCO3 is not a decomposition redox reaction due to oxidation number remain same in product substance combine elements like

$$\overset{+2\ +4\ -2}{CaCO_3} \rightarrow \overset{+2\ -2}{CaO} + \overset{+4\ -2}{CO_2}$$

c. Displacement Reactions: Reaction in which one ion (or atom)in a compound is replaced by an ion(or atom) of other element

X + YZ = XZ + Y

(i) Metal Displacement Reactions: Reactions in which a metal in a compound is displaced by another metal in the uncombined state

$$\overset{+5\ -2}{2V_2O_5} + \overset{0}{5Ca} \rightarrow \overset{0}{2V} + \overset{+2\ -2}{5CaO}$$

$$\overset{+2\ +6\ -2}{2CuSO_4} + \overset{0}{Zn\ (S)} \rightarrow \overset{+2\ +6\ -2}{ZnSO_4} + \overset{0}{Cu}$$

(ii) Non-metal Displacement Reactions: Such reactions in which metal displaced mainly hydrogen or oxygen during reactions like

$$\overset{0}{3Ca(s)} + \overset{+1\ -2}{2H_2O(l)} \rightarrow \overset{+2\ -2\ +1}{Ca(OH)_{2(aq)}} + \overset{0}{H_{2(g)}}$$

$$\overset{0}{2F_2} + \overset{+1\ -2}{2H_2O} \rightarrow \overset{+1\ -1}{4HF} + \overset{0}{O_2}$$

d. Disproportionation Reactions: Reactions in which an element in one oxidation state is simultaneously oxidized and reduced

$$\overset{+1\ -1}{2\ H_2O_2} \rightarrow \overset{+1\ -2}{2H_2O} + \overset{0}{O_2}$$

$$\overset{0}{Cl_2} + 2\overset{+1}{OH}(aq) \rightarrow Cl^-(aq) + \overset{+1}{ClO^-} + H_2O$$

8. Balancing of Redox reactions: There are two types method

8.1. Oxidation Number Method: Following Steps involved in balancing a Redox reaction by oxidation number method-

Step 1. Write the skeletal Redox reaction for all reactants and products of the reaction

$$Cu + NO_3^- \rightarrow NO_2 + Cu^{2+}$$

Step 2. Indicate the oxidation number of all the atoms in each compound

$$\overset{0}{Cu} + \overset{+5\ -2}{NO_3^-} \rightarrow \overset{+4\ -2}{NO_2} + \overset{+2}{Cu^{2+}}$$

Step 3. Identify the element/elements which undergo change in oxidation numbers and calculate the increase or decrease in oxidation number per atom

$$\overset{0}{Cu} + \overset{+5\ -2}{NO_3^-} \rightarrow \overset{+4\ -2}{NO_2} + \overset{+2}{Cu^{2+}}$$

--------Reduction, Gain 1 e^----→

------------Oxidation, loss 2 e^--------------→

Step 4. Equalize the increase in oxidation number with decrease in oxidation number on the reactant side by multiplying formula of oxidizing agent and reducing agents with suitable coefficients

$$Cu + \overset{\times 2}{NO_3^-} \rightarrow NO_2 + \underset{\times 1}{Cu^{2+}}$$

After multiplication $\quad Cu + \mathbf{2}NO_3^- \rightarrow NO_2 + Cu^{2+}$

Step 5. Balance the equation with respect to all other atoms except hydrogen and oxygen

$$Cu + 2NO_3^- \rightarrow \mathbf{2}NO_2 + Cu^{2+}$$

Step 6. Finally balance hydrogen and oxygen. For balancing oxygen atoms add water molecules to the side deficient in it. Balancing of hydrogen atoms depend upon the medium

a. For reactions taking place in acidic solutions add H+ ions to the side deficient in hydrogen atoms

$$Cu + 2NO_3^- \rightarrow 2NO_2 + Cu^{2+} + 2H_2O$$

$$Cu + 2NO_3^-\ \ 4H^+ \rightarrow 2NO_2 + Cu^{2+} + 2H_2O$$

b. For reactions taking place in basic solutions add H2O molecules to the side deficient in hydrogen atoms and simultaneously add equal number of OH- ions on the other side of the equation.

After than finally balance the equation by cancelling common species present on both sides of the equation

8.2. Half Reaction Method (Ion electron method): balance the equation showing the oxidation of Fe^{2+} ions to Fe^{3+} ions by dichromate ions $Cr_2O_7^{2-}$ in acidic medium, wherein, $Cr_2O_7^{2-}$ ions are reduced to Cr^{3+} ions.

Step 1: Write unbalanced equation in ionic form:

$$Fe^{2+} (aq) + Cr_2O_7^{2-} (aq) \rightarrow Fe^{3+} (aq) + Cr^{3+} (aq)$$

Step 2: Separate the equation into half reactions:

Oxidation half $\overset{+2}{Fe^{2+}} (aq) \rightarrow \overset{+3}{Fe^{3+}} (aq)$

Reduction half $\overset{+6\ -2}{Cr_2O_7^{2-}} (aq) \rightarrow \overset{+3}{Cr^{3+}} (aq)$

Step 3: Balance the atoms other than O and H in each half reaction individually.

$$Cr_2O_7^{2-} (aq) \rightarrow 2\ Cr^{3+} (aq)$$

Step 4: For reactions occurring in **acidic medium**, add H2O to balance O atoms and H^+ to balance H atoms.

$$Cr_2O_7^{2-} (aq) \rightarrow 2\ Cr^{3+} (aq) + 7H_2O$$

$$Cr_2O_7^{2-} (aq) + 14H^+ \rightarrow 2\ Cr^{3+} (aq) + 7H_2O$$

Note: For reaction in **alkaline solutions,** add water molecule to balance O, add OH^- to the other side to balance H.

Step 5: Add electrons to one side of the half reaction to balance the charges.

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

$$Cr_2O_7^{2-} (aq) + 14H^+ + 6e^- \rightarrow 2\ Cr^{3+} (aq) + 7H_2O \quad \text{-------(1)}$$

Step 6: Equalize the number of electrons in both the reactions by multiplying a suitable number

$$6Fe^{2+} (aq) \rightarrow 6Fe^{3+} (aq) + 6e^- \quad \text{------------- (2)}$$

Step 7: Add the two balanced half reactions (1) and (2) then cancel the common terms of opposite sides

$$6Fe^{2+} (aq) + Cr_2O_7^{2-} (aq) + 14H^+ \rightarrow 6Fe^{3+} (aq) + 2\ Cr^{3+} (aq) + 7H_2O$$

10.2 (b) Half Reaction Method (Ion electron method): Balance the following equation in **alkali medium**

$$N_2H_4\ (l) + ClO_3^-\ (aq) \rightarrow NO\ (g) + Cl^-\ (aq)$$

Step 1: Separate the equation into half reactions:

Oxidation half $N_2H_4 \rightarrow NO$

Reduction half $ClO_3^-\ (aq) \rightarrow Cl^-\ (aq)$

Step 2: Balance the atoms other than O and H in each half reaction individually.

$$N_2H_4 \rightarrow 2NO$$

Step 3: Add H_2O to balance O atoms and H^+ to balance H atoms.

$N_2H_4 + 2H_2O \rightarrow 2NO$
$N_2H_4 + 2H_2O \rightarrow 2NO + 8H^+$
Note: For reaction in **alkaline solutions,** add OH^- to both side

$N_2H_4 + 2H_2O + 8OH^- \rightarrow 2NO + 8H^+ + 8OH^-$
$N_2H_4 + 2H_2O + 8OH^- \rightarrow 2NO + 8H_2O$

Step 4: Add electrons to one side of the half reaction to balance the charges.

$N_2H_4 + 2H_2O + 8OH^- \rightarrow 2NO + 8H_2O + 8e^-$ ------- (1)

Step 5: Balance another half reaction

$ClO_3^-\ (aq) \rightarrow Cl^-\ (aq) + 3H_2O$
$ClO_3^-\ (aq) + 6H^+ \rightarrow Cl^-\ (aq) + 3H_2O$
$ClO_3^-\ (aq) + 6H^+ + 6OH^- \rightarrow Cl^-\ (aq) + 3H_2O + 6OH^-$
$ClO_3^-\ (aq) + 6H_2O \rightarrow Cl^-\ (aq) + 3H_2O + 6OH^-$
$ClO_3^-\ (aq) + 6H_2O + 6e^- \rightarrow Cl^-\ (aq) + 3H_2O + 6OH^-$ ---------(2)

Step 6: Multiplying eq 1 with 3 and eq 2 with 4 for equalize the number of electrons in both the reactions.

$3N_2H_4 + 6H_2O + 24OH^- \rightarrow 6NO + 24H_2O + 24e^-$
$4ClO_3^-\ (aq) + 24H_2O + 24e^- \rightarrow 4Cl^-\ (aq) + 12H_2O + 24OH^-$

Step 7: Add the two balanced half reactions (1) and (2) then cancel the common terms of opposite sides

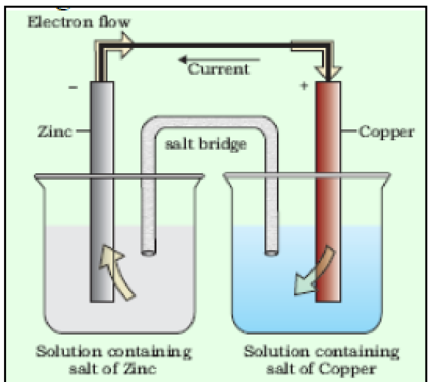
$3N_2H_4 + 4ClO_3^-\ (aq) \rightarrow 6NO + 4Cl^-\ (aq) + 6H_2O$

11. Redox reactions and Electrode Processes (Electrochemical cells and Electrode Potential):

1) An electrochemical cell is a device in which chemical energy of a redox reaction is converted into electrical energy.

2) The simplest electrochemical cell is Daniel cell or Galvanic cell or voltaic cell in which a zinc rod is placed in a solution of Zn^{2+} ions (say $ZnSO_4$) in the left container and a bar of copper metal is immersed in a solution of Cu^{2+} ions (say $CuSO_4$) in the right container. The two metals which act as electrodes are connected by a metallic wire through a voltmeter.

3) The electrode at which oxidation occurs is called **anode** and is negatively charged. The electrode at which reduction takes place is called **cathode** and is positively charged.
Note: LOAN concept for understanding Electrochemical cell- Left Oxidation Anode Negative.
4) In an electrochemical cell the transfer of electrons takes place from anode to cathode so the flow of current is from cathode to anode.



Daniell cell having electrodes of zinc and copper dipping in the solutions of their respective salts.

$Zn \rightarrow Zn^{+2} + 2e^-$ (Oxidation) $Cu^{+2} + 2e^- \rightarrow Cu$ (Reduction)

5) Salt bridge: In the electrochemical cell, the electrical circuit is completed with a salt bridge. Salt bridge also maintains the electrical neutrality of the two half cells. A salt bridge is a U shaped tube filled with solution of inert electrolyte like sodium chloride or sodium sulphate which will not interfere in the redox reaction. The ions are set in a gel or agar agar jelly so that only ions flow when inverted.
6) Electrode potential: Electrical potential difference developed between the metal and its solution is called electrode potential. It can also be defined as tendency of an electrode in a half cell to gain or lose electrons.
7) Oxidation potential is the tendency of an electrode to lose electrons or to get oxidized and Reduction potential is the tendency of an electrode to gain electrons or get reduced.

8) The electrode (half cell) having a higher reduction potential has a higher tendency to gain electrons. So, it acts as a cathode while the electrode having a lower reduction potential acts as an anode and vice versa.

9) Redox couple: A redox couple is defined as having together oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction.

10) Cell potential (E_{cell}): The difference between the electrode potentials of two half cell is called cell potential and it measured in Volts.

$$E_{cell} = E_{cathode} - E_{anode}$$

$$E_{cell} = E_{right} - E_{left}$$

11) EMF (Electromotive force) of a cell: Cell potential is called the EMF of the cell when no current is drawn through the cell.

12) Standard electrode potential (E°): It may be defined as the electrode potential of an electrode (half cell) determined relative to standard hydrogen electrode under standard conditions. It is denoted as E°. The standard electrode potential of hydrogen electrode is 0.00 volts. The standard conditions taken are:

- 1 M concentration of each ion in the solution.
- 298 K temperature and 1 bar pressure of each gas.

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

13) Standard Cell potential ($E°_{cell}$):

$$E^o_{cell} = E^o_{cathode} - E^o_{anode}$$

Reference Exercise: NCERT book pg- 274 Exercise No. 8.26 to 8.30