CLASS 11 Chemistry Chapter 8 Redox Reactions

Oxidation and Reduction: Let us compare Oxidation and Reduction

	Oxidation		**Reduction**
1.	Addition of oxygen	**1.**	Removal of oxygen
2.	Removal of Hydrogen	**2.**	Addition of Hydrogen
3.	Addition of an electronegative element.	**3.**	Removal of an electronegative element.
4.	Removal of an electropositive element	**4.**	Addition of an electropositive element.
5.	Loss of electron(s)	**5.**	Gain of electron(s)
6.	Increase in oxidation number.	**6.**	Decrease in oxidation number.

- **Reducing Agent:** Donor of electrons.
- **Oxidizing Agent:** Acceptor of electrons.
- **Redox Reaction:** Reactions in which oxidation and reduction takes place simultaneously.
- **Oxidation Number:** It is charge that an atom appears to have in a givenspecies when the bonding electron are counted towards more electro- negative atom.

Calculation of Oxidation Number:

Oxidation number of all the elements in their elemental form (in standard state) is taken as zero. Oxidation number of element in a molecule Cl_2, F_2, O_2, P_4, O_3, Fe, H_2, N_2, C (graphite) is zero.

Common Oxidation number of elements of first group is +1. CommonOxidation number of elements of second group + 2.

For ions composed of only one atom, the oxidation number is equalto the charge on the ion.

The oxidation number of oxygen in most compounds is – 2. While inperoxides (*e.g.*, H_2O_2, Na_2O_2), each oxygen atom is assigned an oxidation number of – 1, in super oxides (*e.g.*, KO_2, RbO_2) each oxygen atom is assigned an oxidation number of – (½).

In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of + 2 and + 1, respectively.

The oxidation number of hydrogen is + 1 but in metal hydride itsoxidation no. is – 1.

In all its compounds, fluorine has an oxidation number of – 1.

The algebraic sum of the oxidation number of all the atoms in acompound must be zero.

In polyatomic ion, the algebraic sum of all the oxidation numbers ofatoms of the ion must equal the charge on the ion.

Calculation of average oxidation number: Solved Examples

Example-1: Calculate oxidation number of underlined element: (a) $Na_2\underline{S}_2\ _3$

(b) $N_2\underline{S}_4\ _6$

Sol. (a) Let oxidation number of S-atom is x. Now work accordingly with the rules given before. (+1) × 2 + (x) × 2 + (–2) × 3 = 0 x = + 2

(b) Let oxidation number of S-atom is x

(+1) × 2 + (x) × 4 + (–2) × 6 = 0 x = + 2.5

Types of Redox Reactions:

Combination Reaction: A combination reaction can be represented in the following manner:

$$A + B \rightarrow C$$

Either A and B or both A and B must be in the elemental form for such a reaction to be a redox reaction. All combustion reactions, which make use of elemental dioxygen, as well as other reactions involving elements other than dioxygen, are redox reactions. Some important examples of this category are:

0 0 +4 –2

C(s) + O2 (g) → CO2(g)

0 0 + 2 – 3

$$3\,Mg\,(s) + N_2\,(g) \xrightarrow{\Delta} Mg_3N_2\,(s)$$

–4+1 0 +4 –2 +1 –2

CH4(g) + 2O2(g) CO2(g) + 2H2O (l)

Decomposition Reaction: Decomposition reactions are the opposite of combination reactions. Precisely, 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. Examples of this type of reactions are

+1 +5 – 2 + 1 – 1 0

$$\mathbf{2KClO_3(s) \xrightarrow{\Delta} 2\ KC1(s) + 3O_2\ (g)}$$

Metal Displacement: A metal in a compound can be displaced by another metal in the uncombined state. We have already discussed about this class of the reactions under section 8.2.1. Metal displacement reactions find many applications in metallurgical processes in which pure metals are obtained from their compounds in ores. A few such examples are:

+ 2 + 6 – 2 0 + 2 + 6 – 2 0

$$\mathbf{CuSO_4\ (aq) + Zn(s) \rightarrow ZnSO_4\ (aq) + Cu\ (s)}$$

Non-metal displacement: The non-metal displacement redox reactions include hydrogen displacement and a rarely occurring reaction involving oxygen displacement. All alkali metals and some alkaline earth metals (Ca, Sr, and Ba) which are very good reductants, will displace hydrogen from cold water.

0 + 1 – 2 + 2 – 2 + 1 0

$$Ca\ (s) + 2\ H_2O\ (1) \rightarrow Ca\ (OH)_2 + H_2\ (g)$$

Disproportionation reactions: Disproportionation reactions are a special type of redox reactions. In a disproportionation reaction an element in one oxidation state is simultaneously oxidized and reduced. One of the reacting substances in a disproportionation reaction always contains an element that can exist in at least three oxidation states. The element in the form of reacting substance is in the intermediate oxidation state; and both higher and lower oxidation states of that element are formed in the reaction. The decomposition of hydrogen peroxide is a familiar example of the reaction, were oxygen experiences disproportionation.

0 –1 + 1

$\mathbf{C1_2\ (g) + 2\ OH^- (aq) \rightarrow Cl^- (aq) + ClO^- (aq) + H_2O\ (1)}$

$\mathbf{P_4(s) + 3OH– (aq) + 3H_2O(l) \rightarrow PH_3(g) + 3H_2PO_2 – (aq)}$

Stock Notation: Representing oxidation number of metal in Roman numerals within parenthesis after the symbol or name of metal in the molecular formula or name of a compound. For *e.g.*, Stock Notation of Ferric oxide is $Fe_2(III)O_3$ or Iron (III) oxide.

Fractional Oxidation Number:

When two or more atoms of an elementare present in different oxidation states, then calculated oxidation number may come out as fractional due to average of all the different oxidationstates.

In reality no element can have a fractional oxidation state.

Balancing of Redox Reactions:

Oxidation number method: In writing equations for oxidation-reduction reactions, just as for other reactions, the compositions and formulas must be known for the reactants and the products. The oxidation number method is now best illustrated in the following steps:

Step 1: Write the correct formula for each reactant and product.

Step 2: Identify atoms which undergo change in oxidation number in the reaction by assigning the oxidation number to all elements in the reaction.

Step 3: Calculate the increase or decrease in the oxidation number per atom and for the entire molecule/ion in which it occurs. If these are not equal then multiply by suitable number so that these become equal. (If you realize that two substances are reduced and nothing is oxidized or vice-versa, something is wrong. Either the formulas of reactants or products are wrong or the oxidation numbers have not been assigned properly).

Step 4: Identify the involvement of ions if the reaction is taking place in water, add H+ or OH– ions to the expression on the required side so that the total ionic charges of reactants and products are equal. If the reaction is carried out in acidic solution, use H+ ions in the equation; if in basic solution, use OH– ions.

Step 5: Make the numbers of hydrogen atoms in the expression on the two sides equal by adding water (H_2O) molecules to the reactants or products. Now, also check the number of oxygen atoms. If same number of oxygen atoms are present in the reactants and products, the equation then becomes the balanced redox reaction.

Half reaction method:

In this method, the two half equations are balanced separately and then added together to give balanced equation. Suppose we have to 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. The following steps are involved in this task.

Step 1: Write the unbalanced equation for the reaction in ionic form:

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

Step 2: Separate the equation into half-reactions:

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

Reduction half reaction: $\mathbf{\overset{+6}{Cr_2}\overset{-2}{O_7}{}^{2-(}aq) \rightarrow \overset{+3}{Cr}{}^{3+(}aq)}$

Step 3: Balance the atoms other than O and H in each half reaction separately. Here the oxidation half reaction is already balanced with respect to Fe atoms. For the reduction half reaction, we multiply the Cr^{3+} by 2 to balance Cr atoms.

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

Step 4: For reactions occurring in acidic medium, add H_2O to balance O atoms and H^+ to balance H atoms. Thus, we get:

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

Step 5: Add electrons to one side of the half reaction to balance the charges. If required, make the number of electrons equal in the two half reactions by multiplying one or both half reactions by appropriate number.

The oxidation half reaction is thus rewritten to balance the charge:

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

Now in the reduction half reaction there are net twelve positive charges on the left - hand side and only six positive charges on the right- hand side. Therefore, we add six electrons on the left side.

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

To equalize the number of electrons in both the half reactions, we multiply the oxidation half reaction by 6 and write as:

$$6Fe^{2+}\,(aq) \rightarrow 6Fe^{3+}(aq) + 6e–$$

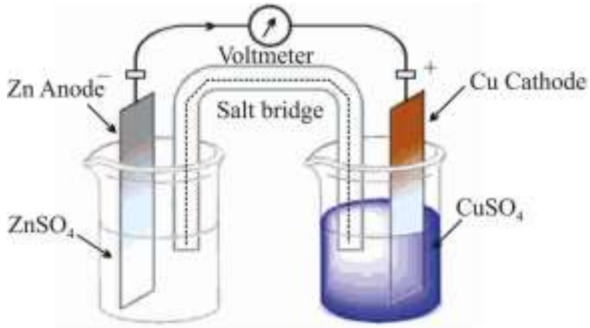
Step 6: We add the two half reactions to achieve the overall reaction and cancel the electrons on each side. This gives the net ionic equation as:

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

Electrode Potential (E): Potential difference between electrode and electrolytic solution due to charge separation.

Standard Electrode Potential (E^{θ}): Electrode Potential measured at 298 K and 1M concentration of metal ions (or 1 bar pressure of gas).

Electrochemical Cell: A device in which chemical energy of a spontaneousredox reaction is converted into electrical energy.



Cell diagram: $Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$

LHS oxidation, $Zn \rightarrow Zn^{2+} + 2e^{-}$

RHS reduction $Cu^{2+} + 2e^{-} \rightarrow Cu$

Overall reaction $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Representation of an Electrochemical cell:

———— Flow of electrons ———→

←———— Flow of current ———

$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

	Left Electrode	Salt Bridge	**R**ight Electrode
LOAN	**O**xidation		**R**eduction
	Anode		**C**athode
	Negative		**P**ositive

Functions of Salt Bridge:

(i) To complete inner circuit.

(ii) To maintainelectrical neutrality around electrodes.

MIND MAP : REDOX REACTIONS

REDOX REACTION

OXIDATION

- Addition of oxygen
  $Mg + O_2 \longrightarrow MgO$
- Addition of electronegative element
  $Mg + F_2 \longrightarrow MgF_2$
- Removal of Hydrogen
  $2H_2S + O_2 \longrightarrow 2S + 2H_2O$
- Removal of electronegative element
  $2K_4[Fe(CN)_6] + H_2O_2 \longrightarrow 2K_3[Fe(CN)_6] + 2KOH$
- Loss of electron or increase in oxidation number
  $2Na + Cl_2 \longrightarrow Na^{+}Cl^{-}$
  Na — oxidise

Reduction / Reducing Agent
Substance which get oxidised itself but help to reduce others substance is called Reducing Agent / Reductant.

Types of Redox Reaction

- Combination Reactioin
  $\overset{0}{C} + \overset{0}{O}_2 \longrightarrow C^{+4}O_2^{-2}$
- Decomposition Reaction
  $2K\overset{+5}{Cl}O_3^{-2} \longrightarrow 2KCl^{-1} + 3\overset{0}{O}_2$
- Displacement Reaction
  a) Metal Displacement
  $Cr_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Cr$
  b) Non-Metal Displacement
  $2Na + 2H_2O \longrightarrow 2NaOH + H_2$
- Disproportionation Reaction
  $H_2O_2 \longrightarrow 2H_2O + O_2$

ELECTROCHEMICAL

$Zn\,|Zn^{2+}|\,|Cu^{2+}|Cu$

$E^\circ_{Cell} = E - E^\circ_L = E^\circ_{Cathode} - E^\circ_{Anode}$

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

$E_{Cell} = E^\circ_{Cell} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$

$E^\circ_{Cell} = 0$ At equilibrium

$E_{Cell} < 0$ Non spontaneous reaction

$E_{Cell} > 0$ Spontaneous reaction

REDUCTION

- Addition of Hydrogen
  $Na + H_2 \longrightarrow NaH$
- Addition of electropositive element
  $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$
- Removal of Oxygen
  $2HgO \longrightarrow 2Hg(l) + O_2(g)$
- Removal of electronegative element
  $2FeCl_3 + H_2 \longrightarrow 2FeCl_2 + 2HCl$
- Gain of electron or decrease in oxidation number
  $2Na + Cl_2 \longrightarrow 2\overset{+}{Na}\overset{-}{Cl}$
  Cl get reduced

Oxidant or Oxidising Agents
Substance which get reduced and help to oxidise other substance.