Balancing of oxidation-reduction reactions.

Though there are a number of methods for balancing oxidation – reduction reactions, two methods are very important.

These are:

(1) Oxidation number method

(2) Ion – electron method

(1) **Oxidation number method**: Themethod: for balancing redox reactions by oxidation number change method was developed by **Johnson.** In a balanced redox reaction, total increase in oxidation number must be equal to the total decrease in oxidation number. This equivalence provides the basis for balancing redox reactions. This method is applicable to both molecular and ionic equations. The general procedure involves the following steps,

(i) Write the skeleton equation (if not given, frame it) representing the chemical change.

(ii) Assign oxidation numbers to the atoms in the equation and find out which atoms are undergoing oxidation and reduction. Write separate equations for the atoms undergoing oxidation and reduction.

(iii) Find the change in oxidation number in each equation. Make the change equal in both the equations by multiplying with suitable integers. Add both the equations.

(iv) Complete the balancing by inspection. First balance those substances which have undergone change in oxidation number and then other atoms except hydrogen and oxygen. Finally balance hydrogen and oxygen by putting H_2O molecules wherever needed. The final balanced equation should be checked to ensure that there are as many atoms of each element on the right as there are on the left.

(v) In ionic equations the net charges on both sides of the equation must be exactly the same. Use H^+ ion/ions in acidic reactions and OH^- ion/ions in basic reactions to balance the charge and number of hydrogen and oxygen atoms.

The following example illustrate the above rules:

Step:I $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$ (Skeleton equation)

Step: II writing the oxidation number of all the atoms.

$$\overset{0}{Cu} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O}_3 \rightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O}_3)_2 + \overset{+4}{N}\overset{-2}{O}_2 + \overset{+1}{H}_2\overset{-2}{O}$$

Step: III Change in oxidation number has occurred in copper and nitrogen.

$$\overset{0}{Cu} \rightarrow \overset{+2}{Cu}(NO_3)_2 \qquad \text{......(i)}$$

$$H\overset{+5}{N}O_3 \rightarrow \overset{+4}{N}O_2 \qquad \text{......(ii)}$$

Increase in oxidation number of copper = 2 units per molecule Cu

Decrease in oxidation number of nitrogen = 1 unit per molecule HNO_3

Step: IV To make increase and decrease equal, equation (ii) is multiplied by 2.

$$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

Step: V Balancing nitrate ions, hydrogen and oxygen, the following equation is obtained.

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

This is the balanced equation.

(2) **Ion-electron method (half reaction method)**

The method for balancing redox-reactions by ion electron method was developed by **Jette** and **LaMev** in 1927. It involves the following steps

(i) Write down the redox reaction in ionic form.

(ii) Split the redox reaction into two half reactions, one for oxidation and other for reduction.

(iii) Balance each half reaction for the number of atoms of each element. For this purpose,

(a) Balance the atoms other than H and O for each half reaction using simple multiples.

(b) Add water molecules to the side deficient in oxygen and H^+ to the side deficient in hydrogen. This is done in acidic or neutral solutions.

(c) In alkaline solution, for each excess of oxygen, add one water molecule to the same side and $2OH^-$ ions to the other side. If hydrogen is still unbalanced, add one OH^- ion for each excess hydrogen on the same side and one water molecule to the other side.

(iv) Add electrons to the side deficient in electrons as to equalize the charge on both sides.

(v) Multiply one or both the half reactions by a suitable number so that number of electrons become equal in both the equations.

(vi) Add the two balanced half reactions and cancel any term common to both sides.

The following example illustrate the above rules

Step: I $I_2 + OH^- \rightarrow IO_3^- + I^- + H_2O$ (Ionic equation)

Step: II Splitting into two half reactions, $I_2 + OH^- \rightarrow IO_3^- + H_2O$; $I_2 \rightarrow I^-$

(Oxidation half reaction) (Reduction half reaction)

Step: III Adding OH^- ions, $I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O$

Step: IV Adding electrons to the sides deficient in electrons,

$$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-;\ I_2 + 2e^- \rightarrow 2I^-$$

Step: V Balancing electrons in both the half reactions.

$$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-;\ 5[I_2 + 2e^- \rightarrow 2I^-]$$

Step: VI Adding both the half reactions.

$6I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10I^-$; Dividing by 2, `

$3I_2 + 6OH^- \rightarrow IO_3^- + 5I^- + 3H_2O$