

Stoichiometry - II

Analysis of Redox Reactions

There are two ways of analysing and solving the stoichiometric problems based on redox reactions :

- (a) One is to balance the redox reaction and then approach the problem like a simple balanced chemical equation using the concept of moles. (*Application of mole concept as done earlier*).
- (b) Second, use the concept of gram equivalents (or milli equivalents) as used in neutralisation.
Here, gmeq. (or meq) of oxidising agent = gmeq. (or meq) of reducing agent.

First we will learn how to analyse a given redox reaction and different methods used for balancing it. Then as the chapter progress, you will learn how to proceed while solving numerical problems based on the redox analysis.

OXIDATION - REDUCTION

Section - 1

Oxidation is a process that involves the loss of electrons by the atoms or ions.

Reduction is a process that involves the gain of electrons by the atoms or ions.

Any reaction, in which the electrons are exchanged between atoms or ions, represents simultaneous process of oxidation and reduction and is called as a Redox Reaction.

In a Redox Reaction, the species that loses electron (i.e., gets oxidised) is known as **reducing agent** or reductant, (since it causes reduction of other species), the species which accepts electrons from reductant (i.e., gets reduced) is known as **oxidising agent** or oxidant (as it causes oxidation of other species).

Illustrating the concept :

$\text{Mg} + \text{Zn}^{2+} \longrightarrow \text{Mg}^{2+} + \text{Zn}$ is a redox reaction.

Oxidation : $\text{Mg}^0 \longrightarrow \text{Mg}^{2+} + 2\text{e}^-$

Reduction : $\text{Zn}^{2+} + 2\text{e}^- \longrightarrow \text{Zn}^0$

Mg causes reduction of Zn^{2+} , so it is Reductant.

Zn^{2+} causes oxidation of Mg, so it is Oxidant.

Note : A redox reaction involves two half reactions: *Oxidation half reaction and Reduction half reaction*.

Now, a new concept of Oxidation Number and Oxidation State is being introduced in order to study oxidation and reduction reactions (redox reactions) in a much better way.

Oxidation Number (O.N.) : It refers to the total charge on all atoms of same kind in a compound.

Oxidation State (O.S.) : It refers to the charge per atom of all atoms of same kind in a compound.

Oxidation state, many a times, is also referred to as Oxidation Number (Refer to NCERT Book).

This means, oxidation number of an element in a compound is equal to oxidation state of that element multiplied by total atoms of that element in particular compound.

- (i) In ionic compounds, it is simply the charge on corresponding cation and anion which is expressed as oxidation state of that particular element. For example, oxidation state of potassium and chlorine in potassium chloride (KCl) is simply +1 and -1 respectively as KCl is treated as $K^+ Cl^-$.

Refer to following examples where oxidation states are written above the atoms :

+2 -1	+2 -2	+3 -1	+2 -2	+1 -1	+1 +6 -2
MgCl ₂	CaS	AlCl ₃	CaO	NaF	K ₂ SO ₄

Note : (a) In MgCl₂ and AlCl₃, -1 is the oxidation state of Cl atom and its oxidation number is -2 and -3 respectively.
 (b) In each of the cases, the sum of oxidation number of all atoms of all kinds is equal to zero (as the compound is neutral).

- (ii) In Covalent Compounds, it is not so easy to assign oxidation state of an atom. In order to simplify the concept, we are going to define a set of rules which would enable us to assign oxidation state to every element in any compound.

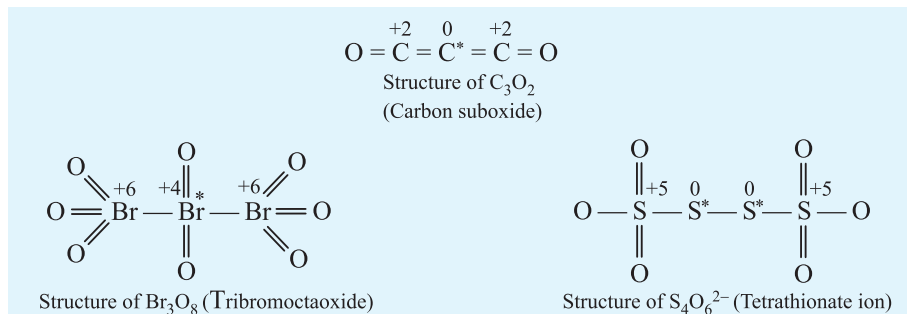
Rules for Assigning Oxidation State (O.S.) and Oxidation Number (O.N.) :

- Any element in free state is assigned an oxidation state of zero (0). For example: O.S. of H, P, S, O, Fe, Br in H₂, P₄, S₈, O₂, Fe, Br₂ is equal to 0.
- The oxidation state of any cation or anion (of form A⁺ or B⁻) is equal to the magnitude of its charge. For example: O.S. of Ca in Ca²⁺ = +2, O.S. of Al in Al³⁺ = +3, O.S. of Cl in Cl⁻ = -1 and so on.
- The algebraic sum of the oxidation numbers of all atoms in a neutral compound is equal to 0.
 - The algebraic sum of the oxidation numbers of all atoms in an ion (like PO₄³⁻ etc.) is equal to the charge on the ion.
- The oxidation states of Alkali Metals (Group IA) is +1 in all of their compounds and that of Alkaline Earth elements (Group IIA) is +2 in all of their compounds.
- Hydrogen in almost all of its compounds is assigned an oxidation state of +1. The exception occurs when hydrogen forms compounds with strong metals (metallic hydrides). For example : KH, NaH, MgH₂, CaH₂ etc. In all of these, the oxidation state of hydrogen is -1.
- Oxygen in almost all of its compounds is assigned an oxidation state of -2. In a class of compounds, *Peroxides*, oxidation state of oxygen is -1, for example : H₂O₂, Na₂O₂ etc. Another exception is OF₂, where O.S. is +2. O₂F₂, where O.S. is +1 and KO₂ in which O.S. is $-\frac{1}{2}$.
- Fluorine is most electronegative element and is assigned an O.S. of -1, in all its compounds. For other halogens, O.S. is generally -1 except when they are bonded to a more electronegative halogen or oxygen. O.S. of iodine in IF₇ is +7, O.S. of chlorine in KClO₃ is +5.

Concept of Fractional Oxidation States :

Fractional oxidation state is the average oxidation state of the element under examination and it means that the element for which fractional oxidation state is realised is present in different oxidation states.

Structure of the species C_3O_2 , Br_3O_8 and $S_4O_6^{2-}$ reveal the following bonding situations :



The elements marked with asterisk in the above examples are exhibiting different oxidation state from the rest of the atoms of the same element in each of the species. This means that in C_3O_2 , two carbon atoms are present in +2 oxidation state each, whereas the third one is present in zero oxidation state and the average is $4/3$. Likewise in Br_3O_8 , each of the two terminal bromine atoms are present in +6 oxidation state and the middle bromine is present in +4 oxidation state and the average is $16/3$. In the same manner, in the species $S_4O_6^{2-}$ is 2.5, whereas the reality being +5, 0, 0 and +5 oxidation number respectively for each sulphur. Fe_3O_4 , Mn_3O_4 , Pb_3O_4 are some of the other examples of the compounds, which are mixed oxides having the metal in fractional oxidation state. However, the oxidation states may be in fraction as in O_2^+ and O_2^- where it is $+1/2$ and $-1/2$ respectively.

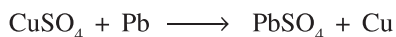
Using all these rules, we can now assign oxidation states to almost all the elements in various compounds and ions.

Try Yourself : Find the oxidation state of atoms (written in **bold**) in the following compounds.

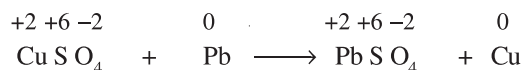
SiO_2 , OF_2 , $KClO_3$, $HClO_4$, NH_4^+ , Hg_2O , Pb_3O_4 , MnO_2 , $C_3H_6Cl_2$, BO_3^{3-} , K_2MnO_4 , Cr_2O_3 , CrO_2Cl_2 , Mg_3N_2 , N_2O_5 , $MgSiO_3$, $Na_2S_2O_3$, $S_4O_6^{2-}$, CaC_2 , $H_2C_2O_4$, NH_4NO_3

Analysis of a Redox Reaction

Consider the following Redox Reaction :



1. Assign oxidation state to each atom.



- Note :** (i) O.S. of S, O remains same, i.e., neither of the two is oxidised or reduced (by inspection).
 (ii) O.S. of Pb and Cu^{2+} has changed from left to right.

(B) Ion Electron (Half-Reaction) Method

1. Identify the oxidation and reduction half reactions and write them separately in ionic form. Note that, write whole of ion in half reaction (unlike the oxidation number method where only atoms undergoing the change of state were taken).

For example : Mn in MnO_4^- in acidic medium generally goes to MnSO_4 or Mn^{2+} .

Write it like this : $\text{MnO}_4^- \longrightarrow \text{Mn}^{2+}$ (and not as $\text{Mn}^{7+} \longrightarrow \text{Mn}^{2+}$)

2. Balance each half reaction separately. This is done accordingly to the following procedure.
3. Balance all the atoms of both reactions except O and H.
4. Now first balance O and H atoms (unlike in first method) depending upon the medium of reaction.

Acidic Medium :

To balance O atoms add same number of water (H_2O) molecules to the side deficient in O atoms.



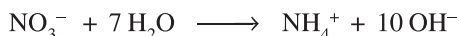
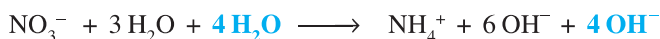
Now to balance H atoms, add as many H^+ ions required to the side deficient in H atoms (generally it is equal to double the number of water molecules added to the other side while balancing O atoms).

**Basic Medium :**

To balance O atoms, add same number of H_2O molecules to the side having excess of O atoms and add the double the number of OH^- ions to the other side (i.e., to the side deficient in O atoms).



Now to balance H atoms (if required), add same number of OH^- ions to the side in excess of H atoms and then add same number of water molecules to the other side (i.e., the side deficient in H atoms).



5. Now add electrons to the side deficient in negative charge in order to balance the charge on both sides.



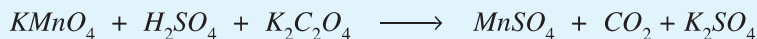
These are balanced half reactions in acidic and basic medium respectively.

6. Now add two half reactions together in such a manner that electrons from both sides cancel. So multiply by coefficients so that number of electrons produced in oxidation equals the number of electrons used in reduction (as done in oxidation number method step - 4).

Note : In this method we don't have to calculate oxidation states as done in first method.

Illustration - 1

Balance the following reaction by Oxidation number and Ion electron method.

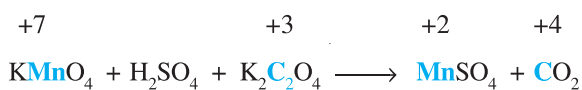


SOLUTION :

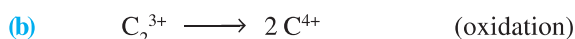
Oxidation Number Method :

Follow the steps given below to balance the reaction given above :

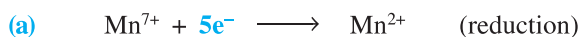
Assign O.S. to the atoms which are oxidised or reduced. Note that O.S. of O, K, H, S are same on both sides. This is done by oral inspection. We'll consider O.S. of Mn and C on both sides.



Write two half reactions and balance the atoms as follows.



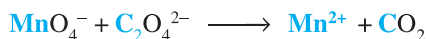
Balance the charge by adding $5e^-$ to left of (a) and $2e^-$ to the right of (b).



Add two half reactions after multiplying (a) by 2 and (b) by 5 in order that electrons cancel from both sides.

Ion Electron Method :

First write the given equation in Ionic form having ions with Central atom (which has undergone a change in Oxidation State).

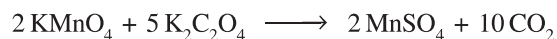


Note that O and H atoms attached to the Central atom (shown in bold), have to be retained.

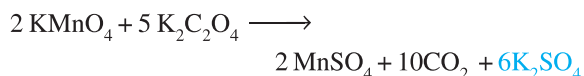
Now, write the Oxidation and Reduction half and balance them as shown :



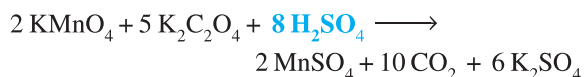
Now compare this balanced equation with the molecular unbalanced equation as follows.



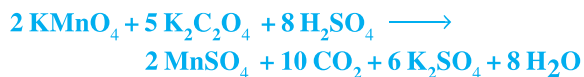
As the charge is balanced, now balance 'K' atoms on both sides by adding **6 K₂SO₄** on right.



Now to balance SO_4^{2-} ions on both sides, add **8 H₂SO₄** on left.

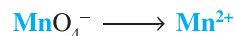


Finally add **8 H₂O** on right to balance O and H atoms to get balanced equation.



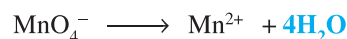
Note : You can always make a final check by counting O atoms on both sides.

Reduction :

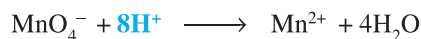


(i) First, make sure that the element undergone a change in Oxidation State is balanced.

(ii) Balance O atoms by adding **4 H₂O** on right hand side (RHS).



- (iii) Now, RHS has excess of 8 H atoms. Add 8H^+ on left hand side (LHS). Note, the medium is acidic due to the presence of H_2SO_4 .

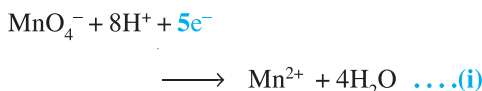


- (iv) Now O and H are balanced. Balance the charge on both sides.

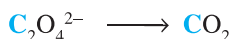
On LHS : Charge is $1 \times (-1) + 8 \times (+1) = +7$

On RHS : Charge is $1 \times (+2) + 4 \times (0) = +2$

Add 5e^- in LHS (Note : each e^- is equivalent to a charge of -1).



Oxidation :



Following the same procedure as above, we have :

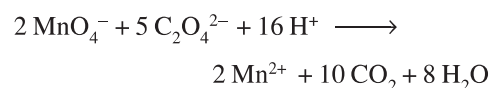
- (i) Balance C atoms : $\text{C}_2\text{O}_4^{2-} \longrightarrow 2\text{CO}_2$

- (ii) Balance O atoms : Already balanced.

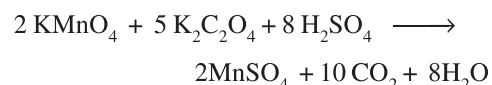
- (iii) Balance H atoms : No H atom.

- (iv) $\text{C}_2\text{O}_4^{2-} \longrightarrow 2\text{CO}_2 + 2\text{e}^- \dots\text{(ii)}$

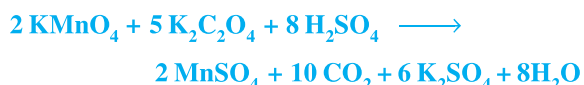
Multiply equation (i) by 2 and (ii) by 5 to balance the electrons transfer and add to get :



Using the above balanced equation, convert it to its corresponding given Neutral equation to get :

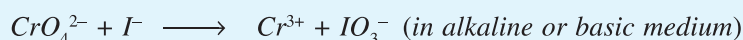


Now, K (12 atoms on LHS) and SO_4^{2-} (6 atoms on LHS) is unbalanced. Add $6\text{K}_2\text{SO}_4$ on RHS to get the final equation as :



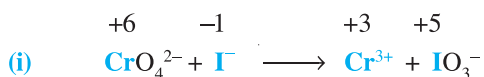
Note : At the later stages, you will find that it is Ion electron method which is more important.

Illustration - 2 Balance the following reaction by oxidation number method.

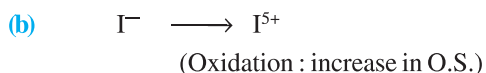


SOLUTION :

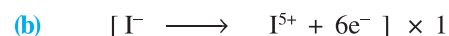
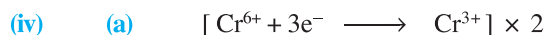
Proceeding on the pattern of last illustration



- (ii) (a) $\text{Cr}^{6+} \longrightarrow \text{Cr}^{3+}$
(Reduction : decrease in O.S.)



- (iii) (a) $\text{Cr}^{6+} + 3\text{e}^- \longrightarrow \text{Cr}^{3+}$



On adding (a) and (b) we get :



On comparing with given reaction



Note that the given reaction is ionic, so first of all charge needs to be balanced. Balance charge by adding OH^- ions (in basic medium).

A charge of -5 on left and $+5$ on right side means add 10 OH^- ions to the right side of the equation.



Finally add $5 \text{ H}_2\text{O}$ molecules on the left to balance H and O and get the balanced equation.

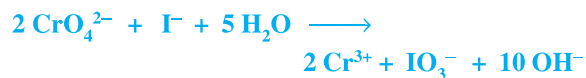
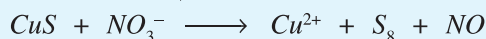
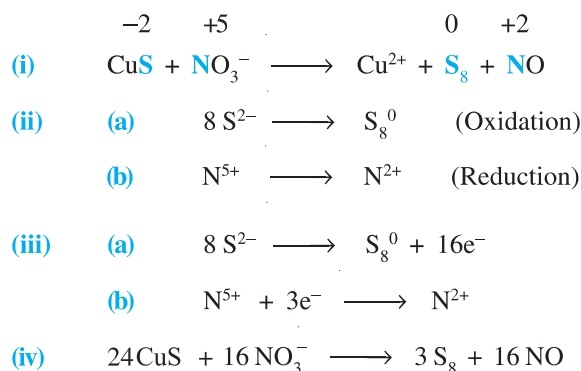


Illustration - 3 Balance the following reaction in acidic medium.

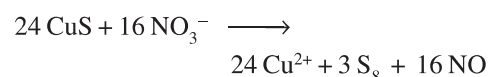


SOLUTION :

Proceeding on the same pattern of last illustration



Add 24 Cu^{2+} ions on right side to balance Cu before balancing the charge.



A charge of -16 on left and $+48$ on right means, add 64 H^+ ions to the left (acidic). Finally add $32 \text{ H}_2\text{O}$ on right to balance H and O atoms.

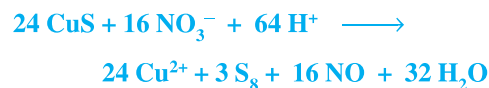
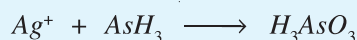
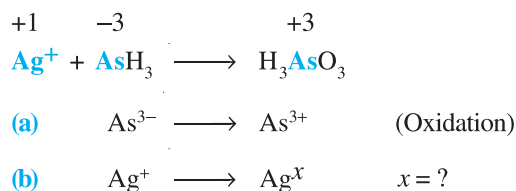


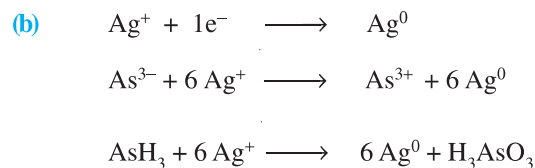
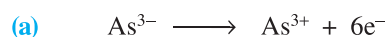
Illustration - 4 Complete and balance the following in acidic medium.



SOLUTION :



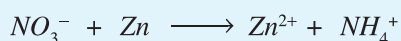
Clearly, (b) must be a reduction, so x must be 0 , as x can't be $-ve$ for Ag (Silver being metal).



Now balance charge by adding 6 H^+ on right side and finally add $3 \text{ H}_2\text{O}$ on left side.

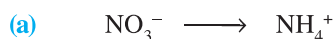


Illustration - 5 Balance the following by ion-electron method in basic medium.

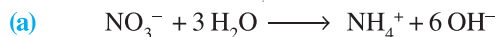


SOLUTION :

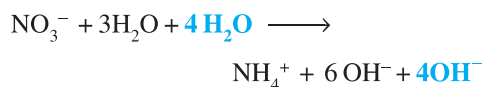
Two half reactions are :



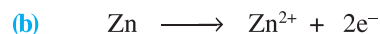
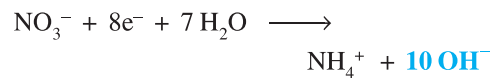
First balance 'O' atoms in basic medium.



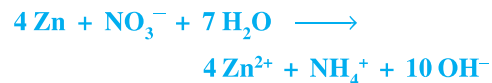
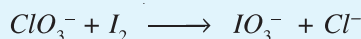
Balancing H atoms in basic medium



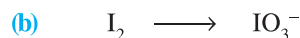
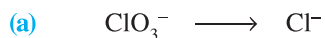
Balancing the charge :



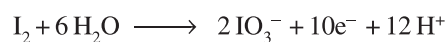
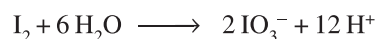
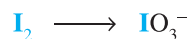
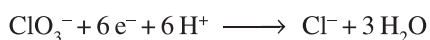
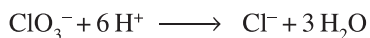
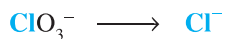
Now overall reaction is :

**Illustration - 6***Balance the following by ion-electron in acidic medium.***SOLUTION :**

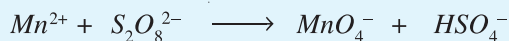
Two half reactions are :



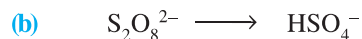
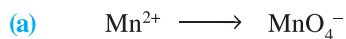
Balancing them separately :



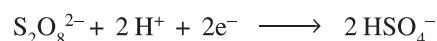
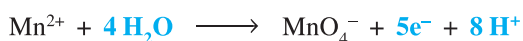
Balancing electrons in two half reactions, and adding we get :

**Illustration - 7***Balance the following by ion-electron method. (Acidic medium)***SOLUTION :**

Two half reactions are :



Balancing these two half reactions :

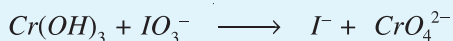


Adding the two equations, we have the balanced equation.

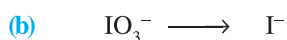
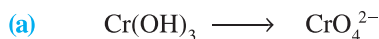


Illustration - 8

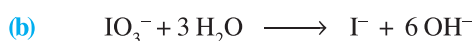
Balance the following by ion-electron method. (Basic medium)

**SOLUTION :**

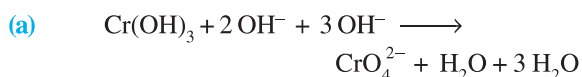
Two half reactions are :



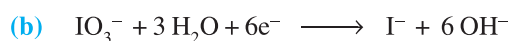
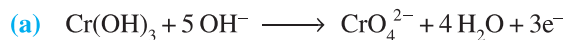
Balancing O atoms :



Balancing H atoms :



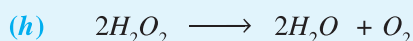
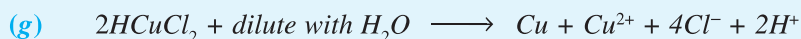
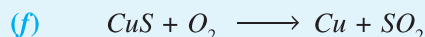
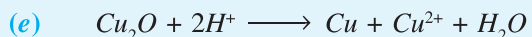
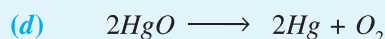
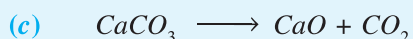
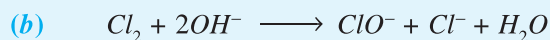
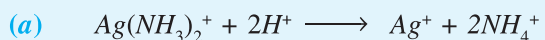
Balancing the charge :



Adding (a) and (b), we get :

**Illustration - 9**

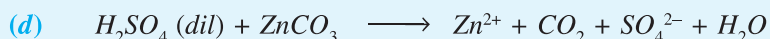
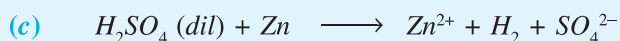
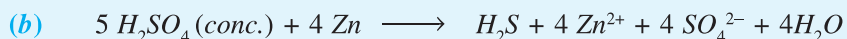
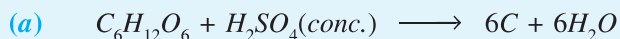
Which of the following are examples of disproportionation reactions ?

**SOLUTION :**

What is the criteria for a disproportionation reaction ?

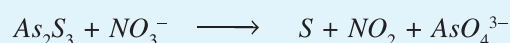
A reaction in which same atom is oxidised and reduced at the same time is called as a disproportionation reaction.

Observing very carefully among the above reactions, we note that (b), (e), (g) and (h) are disproportionation reactions.

Illustration - 10 H_2SO_4 acts as an oxidising agent, a dehydrating agent and an acid. Among the each of following reactions, which behaviour is shown by H_2SO_4 ?

SOLUTION :

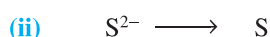
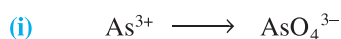
- (a) As a dehydrating agent since it simply removes water from glucose.
- (b) H_2SO_4 (S in +6 state) $\longrightarrow$ H_2S (S in -2 state), i.e., acid is reduced. So it acts as an oxidising agent.
- (c) All strong acids liberate H_2 with an active metals like Zn (i.e., it is reduced) unless SO_4^{2-} is reduced. In this case, H^+ is reduced to H_2 . So it acts as an oxidising agent. We can also call it as an acidic property.
- (d) Simply an acid.

Illustration - 11 Balance the following reaction by ion-electron method. (Acidic medium)**SOLUTION :**

This is a very special case of redox reaction.

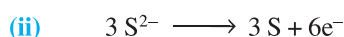
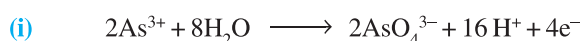
Note that, the given reaction involves two oxidations and one reduction. Proceed carefully as follows.

- (a) Two half reactions (oxidation) are :

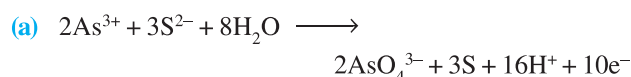


- (b) Reduction half reaction : $\text{NO}_3^- \longrightarrow \text{NO}_2$

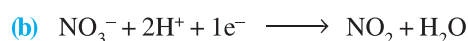
First balancing these two half reductions :



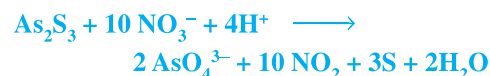
Now add the two equations to get :



Now balance oxidation half reaction and proceed as usual.



Adding the two reactions

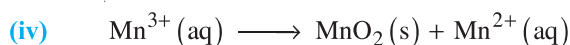
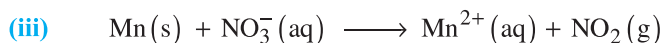
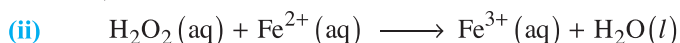
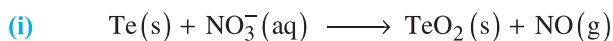


Note : In above type of cases, take total change in oxidation number (increase or decrease) for cross multiplication while adding oxidation and reduction half reactions. For example, in the given case, 10e^- 's are lost per mole of As_2S_3 .

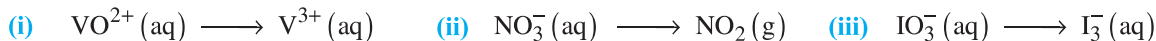
IN-CHAPTER EXERCISE - A

- What is the oxidation state of Cl in :
 (i) HClO_4 (ii) $\text{Ba}(\text{ClO}_3)_2$ (iii) CrO_2Cl_2 (iv) Cl_2O_7
- Indicate the species which are oxidised and reduced in the following reactions.
 (i) $\text{CN}^- + \text{OCl}^- \longrightarrow \text{N}_2 + \text{HCO}_3^- + \text{Cl}^-$
 (ii) $\text{C}_6\text{H}_5\text{OH} + \text{O}_3 \longrightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{O}_2$
 (iii) $\text{I}^- + \text{O}_3 \longrightarrow \text{O}_2 + \text{I}_2$
 (iv) $\text{S}_2\text{O}_3^{2-} + \text{Cl}_2 \longrightarrow \text{HSO}_4^- + \text{Cl}^-$
 (v) $\text{C} + \text{ZnO} \longrightarrow \text{Zn} + \text{CO}$

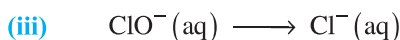
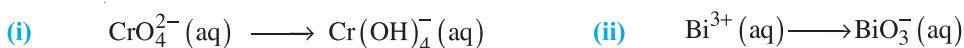
3. Write balanced ionic half equations (oxidation and reduction) for each of the following reactions.



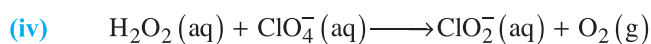
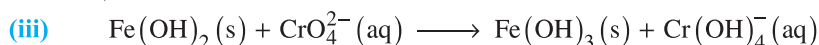
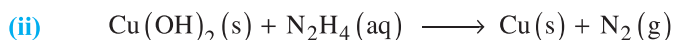
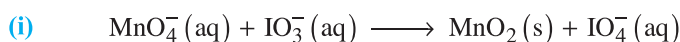
4. Balance the following half-reactions in acidic medium.



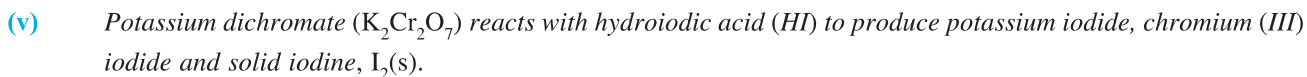
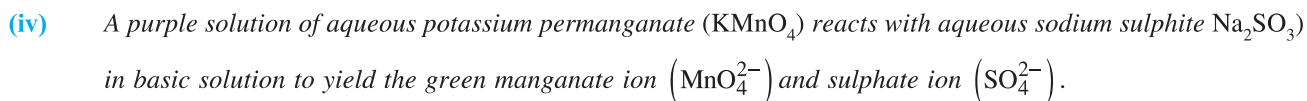
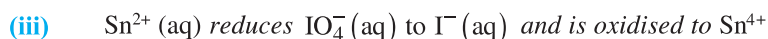
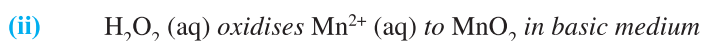
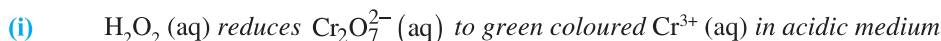
5. Balance the following half-reactions in basic medium.



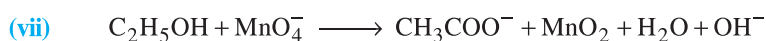
6. Write balanced net ionic equations for the following reactions in basic solution.



7. Write balanced redox reactions for each of the following reactions.



8. Balance the following chemical reactions. (It is better to use ion electron method)



- (viii) $\text{P}_4 + \text{H}^+ + \text{NO}_3^- + \text{H}_2\text{O} \xrightarrow{\text{(dil)}} \text{H}_2\text{PO}_4^- + \text{NO}$
- (ix) $\text{HS}^- + \text{HSO}_3^- \longrightarrow \text{S}_2\text{O}_3^{2-} + \text{H}_2\text{O}$
- (xi) $\text{O}_2^- + \text{H}_2\text{O} \longrightarrow \text{OH}^- + \text{O}_2$
(superoxide ion)
- (xiii) $\text{I}^- + \text{IO}_3^- + \text{H}^+ \longrightarrow \text{H}_2\text{O} + \text{I}_2$
- (xv) $\text{IBr} + \text{BrO}_3^- + \text{H}_2\text{O} \longrightarrow \text{IO}_3^- + \text{Br}^- + \text{H}^+$
- (xvii) $\text{H}_2\text{O}_2 + \text{H}^+ + \text{Fe}^{2+} \longrightarrow \text{H}_2\text{O} + \text{Fe}^{3+}$
- (x) $\text{Fe}^{3+} + \text{NH}_2\text{OH} \longrightarrow \text{Fe}^{2+} + \text{N}_2\text{O} + \text{H}^+$
- (xii) $\text{IO}_3^- + \text{HSO}_3^- \longrightarrow \text{I}^- + \text{H}^+ + \text{SO}_4^{2-}$
- (xiv) $\text{Fe}_2\text{S}_3 + \text{H}_2\text{O} + \text{O}_2 \longrightarrow \text{Fe}(\text{OH})_3 + \text{S}$
- (xvi) $\text{CrI}_3 + \text{H}_2\text{O}_2 + \text{OH}^- \longrightarrow \text{CrO}_4^{2-} + \text{IO}_4^- + \text{H}_2\text{O}$
- (xviii) $\text{As}_2\text{S}_3 + \text{MnO}_4^- + \text{H}^+ \longrightarrow \text{H}_3\text{AsO}_4 + \text{Mn}^{2+} + \text{S}$

9. Choose the correct option for each of the following questions. Only one option is correct.

- (i) Which of the following changes requires a reducing agent ?
- (A) $\text{CrO}_4^{2-} \longrightarrow \text{Cr}_2\text{O}_7^{2-}$ (B) $\text{BrO}_3^- \longrightarrow \text{BrO}^-$
- (C) $\text{H}_3\text{AsO}_3 \longrightarrow \text{HAsO}_4^{2-}$ (D) $\text{Al}(\text{OH})_3 \longrightarrow \text{Al}(\text{OH})_4^-$
- (ii) In balancing the half-reaction :
- $\text{CN}^- \longrightarrow \text{CNO}^-$ (skeletal)
- The number of electrons that must be added is
- (A) 0 (B) 1 on the right (C) 1 on the left (D) 2 on the right
- (iii) Which of the following is a disproportionation reaction ?
- (A) $\text{CaCO}_3 + 2\text{H}^+ \longrightarrow \text{Ca}^{2+} + \text{H}_2\text{O} + \text{CO}_2$ (B) $2\text{CrO}_4^{2-} + 2\text{H}^+ \longrightarrow \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$
- (C) $\text{Cr}_2\text{O}_7^{2-} + 2\text{OH}^- \longrightarrow 2\text{CrO}_4^{2-} + \text{H}_2\text{O}$ (D) $\text{Cu}_2\text{O} + 2\text{H}^+ \longrightarrow \text{Cu} + \text{Cu}^{2+} + \text{H}_2\text{O}$
- (iv) Oxidation number of P in $\text{Ba}(\text{H}_2\text{PO}_2)_2$ is :
- (A) +3 (B) +2 (C) +1 (D) -1
- (v) The number of electrons involved in the reduction of nitrate ion (NO_3^-) to hydrazine (N_2H_4) is :
- (A) 8 (B) 7 (C) 5 (D) 3
- (vi) Sulphur has highest oxidation state in :
- (A) SO_2 (B) H_2SO_4 (C) $\text{Na}_2\text{S}_2\text{O}_3$ (D) $\text{Na}_2\text{S}_4\text{O}_6$
- (vii) Which of the following can act as an oxidising agent as well as a reducing agent ?
1. H_2O_2 2. H_2S 3. SO_2 4. HNO_2
- (A) 2, 3, 4 (B) 1, 2, 3 (C) 1, 3, 4 (D) All
- (viii) SO_2 under atmospheric condition changes to SO_x^{2-} . If oxidation number of S in SO_x^{2-} is +6, what is the value of x in SO_x^{2-} ?
- (A) 1 (B) 2 (C) 3 (D) 4

(ix) For the redox reaction :



the correct coefficients of the reactants for the balanced reaction are :

	$\text{Cr}_2\text{O}_7^{2-}$	Ni	H^+
(A)	1	3	14
(B)	2	3	14
(C)	1	1	16
(D)	3	3	12

ANALYSIS BASED ON MOLE CONCEPT

Section - 2

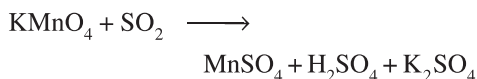
Let us now solve some numerical problems based on the analysis of redox reactions. In such type of applications based problems, first we have to create a chemical equation and then balance it. And then proceed as per the given problem, using Mole concept.

Illustration - 12 Find the amount of Iron pyrites (FeS_2) which is sufficient to produce enough SO_2 on roasting (heating in excess of O_2) such that it (SO_2) completely decolourises a 1L solution of KMnO_4 containing 15.8 g/L of it.

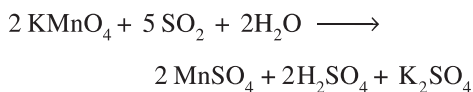
The equations are : $\text{FeS}_2 + \text{O}_2 \longrightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2$ and $\text{KMnO}_4 + \text{SO}_2 \longrightarrow \text{MnSO}_4 + \text{H}_2\text{SO}_4 + \text{K}_2\text{SO}_4$

SOLUTION :

First calculate the amount of SO_2 required to decolourise 15.8 g/L of KMnO_4 solution. For this you need to balance the following chemical reaction.

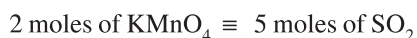


Proceeding in the manner as given in above illustrations, we have the balanced equation as :



Now, using our understanding to solve a typical stoichiometric problems, we can solve this problem as follows :

From stoichiometry of balanced equation, we have:



Calculate moles in 15.8 g/L of KMnO_4 :

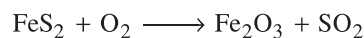
Using : Strength (g/L) = $M M_0$

$$\text{Molarity} = \frac{15.8}{158} = 0.1 \text{ M}$$

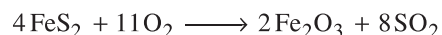
$\Rightarrow$ 1.0 L of KMnO_4 contains 0.1 moles

$$\text{Hence moles of } \text{SO}_2 \text{ required} = \frac{5}{2} (0.1) = 0.25$$

To calculate the amount of pyrites, we have to balance the reaction :



Balancing the reaction (using Hit and Trial Method) we have



From stoichiometry of roasting, we have :

$$8 \text{ moles of } \text{SO}_2 \equiv 4 \text{ moles of } \text{FeS}_2$$

$$\begin{aligned} 0.25 \text{ moles of } \text{SO}_2 &\equiv \frac{4}{8} (0.25) \text{ moles of } \text{FeS}_2 \\ &= 0.125 \text{ moles of } \text{FeS}_2 \end{aligned}$$

$$\Rightarrow \text{Mass of } \text{FeS}_2 = 0.125 \times 120 = 15 \text{ gm}$$