NCERT Solutions for Class-XI Chemistry

Chapter-8
NCERT Chemistry Class 11

1. Assign oxidation numbers to the underlined elements in each of the following species:

(a) NaH_2PO_4

(b) $NaHSO_4$

(c) $H_4\underline{P}_2O_7$

(d) $K_2\underline{Mn}O_4$

(e) $Ca\underline{O}_2$

(f) $Na\underline{B}H_4$

(g) $H_2\underline{S}_2O_7$

(h) $KAI(\underline{S}O_4)_2.12H_2O$

1. O.N is the oxidation number

O.N of Oxygen (O) = –2 (In case of peroxide and superoxide it wil be different ON)

O.N of hydrogen (H) = +1 (In case of metalic hydride, –1)

O.N of sodium (Na) = +1

O.N of aluminium (Al) = +3

O.N of potassium (K) = +1

O.N of calcium (Ca) = +2

In neutral compounds the sum of O.N of all the atoms is zero.

(a) Let the O.N of P be x

$\therefore\ 1 * 1 + 2 * 1 + x + 4 * (-2) = 0 \Rightarrow x = +5$

(b) Let the O.N of S be x

$\therefore\ 4 * 1 + 2 * x + 7 * (-2) = 0 \Rightarrow x = +5$

(d) Let the O.N of Mn be x

$\therefore\ 2 * 1 + x + 4 * (-2) = 0 \Rightarrow x = +6$

(e) Let the O.N of O be x

Ca is an alkaline earth metal so its O.N. is +2

$\therefore\ 1 * 2 + 2 * x = 0 \Rightarrow x = -1$

(f) Let the O.N of B be x

Note that in this H exists as hydride ion H^- so its O.N. is –1

$\therefore\ 1 * 1 + x + 4 * (-1) = 0 \Rightarrow x = +3$

(g) Let the O.N of S be x

$\therefore\ 2 * 1 + 2 * x + 7 * (-2) = 0 \Rightarrow x = +6$

(h) Let the O.N of S be x

$\therefore\ 1 * 1 + 1 * 3 + [x + (-2) * 4] * 2 + 12 * [1 * 2 + (-2)] = 0 \Rightarrow x = +6$

2. What are the oxidation numbers of the underlined elements in each of the following and how do you rationalise your results?

(a) $K\underline{I}_3$

(b) $H_2\underline{S}_4O_6$

(c) $\underline{Fe}_3O_4$

(d) $\underline{C}H_3\underline{C}H_2OH$

(e) $\underline{C}H_3\underline{C}OOH$

2. (a) $K\underline{I}_3$

In KI_3, the oxidation number (O.N.) of K is +1. Hence, the average oxidation number of I is $-\frac{1}{3}$. However, O.N. cannot be fractional. Therefore, we will have to consider the structure of KI_3 to find the oxidation states.

In a KI_3 molecule, an atom of iodine forms a coordinate covalent bond with an iodine molecule.

$$\overset{+1}{K^+}\left[\overset{0}{I}-\overset{0}{I}\leftarrow\overset{-1}{I}\right]$$

Hence, in a KI_3 molecule, the O.N. of the two I atoms forming the I2 molecule is 0, whereas the O.N. of the I atom forming the coordinate bond is –1. (b) $H_2\underline{S}_4O_6$

$$\overset{+1}{H_2}\,\overset{x}{SO_4}\,\overset{-2}{O_6}$$

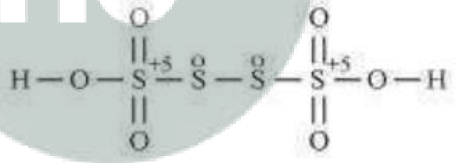
Now, 2(+1) +4(x) +6(–2) = 0

$\Rightarrow 2 + 4x - 12 = 0$

$\Rightarrow 4x = 10$

$\Rightarrow x = +2\frac{1}{2}$

However, O.N. cannot be fractional. Hence, S must be present in different oxidation states in the molecule.



The O.N. of two of the four S atoms is +5 and the O.N. of the other two S atoms is 0.

(c) $\underline{Fe}_3O_4$

On taking the O.N. of O as –2, the O.N. of Fe is found to be $+2\frac{2}{3}$. However, O.N. cannot be fractional.

Here, one of the three Fe atoms exhibits the O.N. of +2 and the other two Fe atoms exhibit the O.N. of +3.

$$\overset{+2}{Fe}O, \overset{+3}{Fe_2}O_3$$

(d) $\underline{C}H_3\underline{C}H_2OH$

$$\overset{x}{C_2}\,\overset{+1}{H_6}\,\overset{-2}{O}$$

2(x) +4(+1) +1(–2) =0

$\Rightarrow 2x + 6 - 2 = 0$

$\Rightarrow x = -2$

Hence, the O.N. of C is –2.

(e) $\underline{C}H_3\underline{C}OOH$

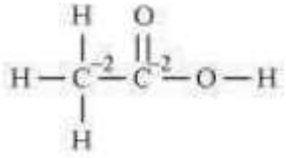
$\overset{x}{C_2}\overset{+1}{H_6}\overset{-2}{O_2}$

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

$\Rightarrow 2x + 4 - 4 = 0$

$\Rightarrow x = 0$

However, 0 is average O.N. of C. The two carbon atoms present in this molecule are present in different environments. Hence, they cannot have the same oxidation number. Thus, C exhibits the oxidation states of +2 and –2 in CH_3COOH.



3. Justify that the following reactions are redox reactions:
(a) $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$
(b) $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$
(c) $4BCl_3(g) + 3LiAlH_4(s) \rightarrow 2B_2H_6(g) + 3LiCl(s) + 3AlCl_3$ (s)
(d) $2K(s) + F_2(g) \rightarrow 2K+F-$ (s)
(e) $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

3. (a) Let us write the O.N of each element

$\overset{2+}{Cu}\overset{-2}{O} + \overset{0}{H_2} \rightarrow \overset{0}{Cu} + \overset{+1}{H_2}\overset{-2}{O}$

Here, the O.N of Cu decreases from +2 to 0 i.e., CuO is reduced to Cu. Also, the O.N of H increases from 0 to +1 i.e., H2 is oxidized to H2O. Hence it is a redox reaction.

(b) Let us write the O.N of each element

$\overset{+3}{Fe_2}\overset{-2}{O_3} + 3\overset{+2}{C}\overset{-2}{O} \longrightarrow 2\overset{0}{Fe} + 3\overset{+4}{C}\overset{-2}{O_2}$

Here, the O.N of Fe decreases from +3 to 0. Also, the O.N of C increases from +2 to +4. Hence it is a redox reaction.

(c) Let us write the O.N of each element

Here, the O.N of Fe decreases from +3 in BCl3 to –3 in B2H6. And, the O.N of H increases from –1 in to +1 in $LiAlH_4$ to +1 in B_2H_6. Hence it is a redox reaction.

(d) We know that oxidation = loosing of e^- by atom and reduction = gaining of e^- by another atom here K lose its electron and F accept it, Hence it is a redox reaction

(e) here N(–3) → N (+2) oxidation reaction

and O(0) → O(–2) reduction reaction (oxidation state of oxygen is zero at molecular state) hence it's a redox reaction

4. Fluorine reacts with ice and results in the change:
$H_2O(s) + F_2(g) \rightarrow HF(g) + HOF(g)$
Justify that this reaction is a redox reaction.

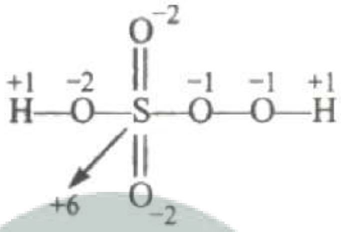
4. Let us write the oxidation number of each atom involved in the given reaction above its symbol as:

$\overset{+1}{H_2}\overset{-2}{O} + \overset{0}{F_2} \longrightarrow \overset{+1}{H}\overset{-1}{F} + \overset{+1}{H}\overset{-2}{O}\overset{+1}{F}$

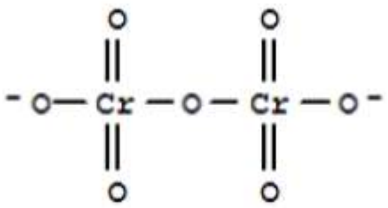
Here, we have observed that the oxidation number of F increases from 0 in F_2 to +1 in HOF. Also, the oxidation number decreases from 0 in F_2 to –1 in HF. Thus, in the above reaction, F is both oxidized and reduced. Hence, the given reaction is a redox reaction.

5. Calculate the oxidation number of sulphur, chromium and nitrogen in H_2SO_5, $Cr2O_7^{2-}$ and NO_3^-. Suggest structure of these compounds. Count for the fallacy.

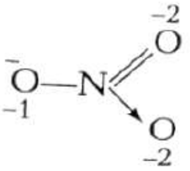
5. H_2SO_5 let the oxidation number of sulphur be x
So,
$2 * 1 + x + 5(-2) = 0$
$x = +8$
There is a fallacy Sulphur cannot have +8 oxidation state because it has maximum +6 oxidation number, not more than that.The structure of H_2SO_5 is shown as follows:



2(H) + 1(S) + 3(O) + 2(O in peroxy linkage)
$\Rightarrow 2(+1) + 1(x) + 3(-2) + 2(-1) = 0$
$\Rightarrow x = +6$
(ii) $Cr_2O_7^{2-}$
Let the oxidation number of chromium be x
Now
$2x + 7 * (-2) = -2$
$x = (-2 + 14)/2$
$x = +6$
There is no fallacy here



(iii) NO_3^-
let assume oxidation number of N is x
$x + (-2) * 3 = -1$
$(x - 6) = -1$
Now, $x = +5$



here is no fallacy about the O.N of N in NO_3^-

6. Write the formulae for the following compounds:

(a) Mercury(II) chloride

(b) Nickel(II) sulphate

(c) Tin(IV) oxide

(d) Thallium(I) sulphate

(e) Iron(III) sulphate

(f) Chromium(III) oxide

6. (a) Mercury (II) chloride:

$HgCl_2$

(b) Nickel (II) sulphate:

$NiSO_4$

(c) Tin (IV) oxide:

SnO_2

(d) Thallium (I) sulphate:

Tl_2SO_4

(e) Iron (III) sulphate:

$Fe_2(SO_4)_3$

(f) Chromium (III) oxide:

Cr_2O_3

7. Suggest a list of the substances where carbon can exhibit oxidation states from –4 to +4 and nitrogen from –3 to +5.

7. The substance of Carbon-

Substance	O.N. of C
CH_4	–4
C_2H_6	–3
C_2H_4 or CH_2Cl	–2
C_2H_2	–1
CH_2Cl_2	0
C_6Cl_6	+1
$(COOH)_2$	+3
CO_2 or CCl_4	+4

Substance for Nitrogen-

Substance	O.N. of N
NH_3	–3
N_2H_4	–2
N_2H_2	–1

N_2	0
N_2O	+1
NO	+2
N_2O_3	+3
N_2O_4	+4
N_2O_5	+5

8. While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why?

8. In sulphur dioxide (SO_2), the oxidation number (O.N.) of S is +4 and the range of the O.N. that S can have is from +6 to –2. Therefore, SO_2 can act as an oxidising as well as a reducing agent. In hydrogen peroxide (H_2O_2), the O.N. of O is –1 and the range of the O.N. that O can have is from 0 to –2. O can sometimes also attain the oxidation numbers +1 and +2. Hence, H_2O_2 can act as an oxidising as well as a reducing agent. In ozone (O_3), the O.N. of O is zero and the range of the O.N. that O can have is from 0 to –2. Therefore, the O.N. of O can only decrease in this case. Hence, O_3 acts only as an oxidant.

In nitric acid (HNO_3), the O.N. of N is +5 and the range of the O.N. that N can have is from +5 to –3. Therefore, the O.N. of N can only decrease in this case. Hence, HNO_3 acts only as an oxidant.

9. Consider the reactions:

(a) $6\ CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$

(b) $O_3(g) + H_2O_2(l) \rightarrow H_2O(l) + 2O_2(g)$

Why it is more appropriate to write these reactions as:

(a) $6CO_2(g) + 12H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6H_2O(l) + 6O_2(g)$

(b) $O_3(g) + H_2O_2(l) \rightarrow H_2O(l) + O_2(g) + O_2(g)$

Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

9. (a) In the photosynthesis process-

Step 1- the liberation of O_2 and $H_2 \rightarrow 2H_2O \rightarrow O_2 + 2H_2$

step-2 The H_2 produced in above reduces the CO_2 into glucose ($C_6H_{12}O_6$) and water (H_2O)

$6CO_2 + 12H_2 \longrightarrow C_6H_{12}O_6 + 6H_2O$

So, the final net reaction is

It is more appropriate to write the reaction as above because water (H_2O) molecule also produced in photosynthesis reaction.

The path of reaction can be investigated by using the radioactive H_2O^{18} instead of (H_2O)

(b) (the final net reaction)

Dioxygen is produced from both steps, one from the decomposition of ozone (O_3) and other is from the reaction of hydrogen peroxide with(O)

- The path of the reaction can be investigated by using $\frac{O_3^{18}}{H_2O^{18}}$.

10. The compound AgF_2 is an unstable compound. However, if formed, the compound acts as a very strong oxidizing agent. Why?

10. The oxidation state of Ag in AgF_2 is +2. But, +2 is an unstable oxidation state of Ag. Therefore, whenever AgF_2 is formed, silver readily accepts an electron to form Ag^+. This helps to bring the oxidation state of Ag down from +2 to a more stable state of +1. As a result, AgF_2 acts as a very strong oxidizing agent.

11. Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

11. These can be understood by the following examples-

- P_4 is reducing agent and Cl_2 is an oxidizing agent

$P_4 + 10Cl_2$ (Excess) $\rightarrow PCl_5$[O.N of phosphorus +5] $\Rightarrow$ Higher O.S of P

P_4 (Excess) $+ 6Cl_2 \rightarrow 4PCl_3$ [O.N of phosphorus +3] $\Rightarrow$ Lower O.S of P

- is an O_2 is a reducing agent and C oxidising agent

$C + O_2 \rightarrow CO_2$ (O is in excess) [O.N of C +4]

$C + O_2 \rightarrow CO$ (C is in excess [O.N of C +2]

- K is a reducing agent and O_2 is an oxidising agent

$K + O_2 \rightarrow K_2O$ (K is in excess) [O.N of O –2] (lower O.S.)

$K + O_2 \rightarrow K_2O_2$ (O is excess) [O.N of O –1] (lower O.S.)

12. How do you count for the following observations?

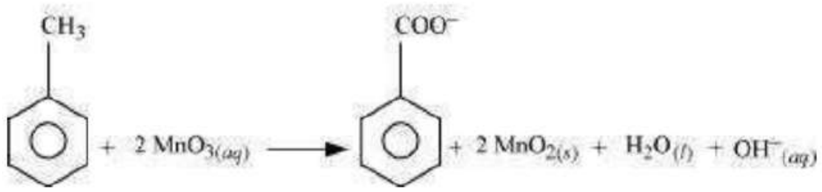
(a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why? Write a balanced redox equation for the reaction.

(b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl, but if the mixture contains bromide then we get red vapour of bromine. Why?

12. (a) In the manufacture of benzoic acid from toluene, alcoholic potassium permanganate is used as an oxidant because of the following reasons.

(i) In a neutral medium, OH^- ions are produced in the reaction itself. As a result, the cost of adding an acid or a base can be reduced.

(ii) $KMnO_4$ and alcohol are homogeneous to each other since both are polar. Toluene and alcohol are also homogeneous to each other because both are organic compounds. Reactions can proceed at a faster rate in a homogeneous medium than in a heterogeneous medium. Hence, in alcohol, $KMnO_4$ and toluene can react at a faster rate. The balanced redox equation for the reaction in a neutral medium is give as below:



(b) When conc. H_2SO_4 is added to an inorganic mixture containing bromide, initially HBr is produced. HBr, being a strong reducing agent reduces H_2SO_4 to SO_2 with the evolution of red vapour of bromine.

$2NaBr + 2H_2SO_4 \longrightarrow 2NaHSO_4 + 2HBr$

$2HBr + H_2SO_4 \longrightarrow 2HBr + H_2SO_4 \longrightarrow \underset{\text{(red vapour)}}{Br_2} + SO_2 + 2H_2O$

But, when conc. H_2SO_4 is added to an inorganic mixture containing chloride, a pungent smelling gas (HCl) is evolved. HCl, being a weak reducing agent, cannot reduce H_2SO_4 to SO_2.

$2NaCl + 2H_2SO_4 \longrightarrow 2NaHSO_4 + 2HCl$

13. Identify the substance oxidised, reduced, oxidising agent and reducing agent for each of the following reactions:

(a) $2AgBr (s) + C_6H_6O_2(aq) \rightarrow 2Ag(s) + 2HBr (aq) + C_6H_4O_2(aq)$

(b) $HCHO(l) + 2[Ag(NH_3)_2]^+(aq) + 3OH^-(aq) \rightarrow 2Ag(s) + HCOO^-(aq) + 4NH_3(aq) + 2H_2O(l)$

(c) $HCHO (l) + 2Cu^{2+}(aq) + 5OH^-(aq) \rightarrow Cu_2O(s) + HCOO^-(aq) + 3H_2O(l)$

(d) $N_2H_4(l) + 2H_2O_2(l) \rightarrow N_2(g) + 4H_2O(l)$

(e) $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

13. (a) Substance reduced/oxidizing agent- AgBr

Substance oxidized/reducing agent- $C_6H_6O_2$]

(b) Substance reduced/oxidising agent- $[Ag(NH_3)_2]^+$

Substance oxidised/reducing agent- HCHO

Substance oxidized/reducing agent- N_2H_4

Substance reduced/oxidizing agent- H_2O_2

(e) Substance oxidized/reducing agent- Pb

Substance reduced/oxidizing agent- PbO_2

14. Consider the reactions:

$$S_2O_3^{2-}(aq) + I_2(s) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$$

$$S_2O^2(aq) + 2Br_2(l) + 5H_2O(l) \rightarrow 2SO_4^{2-}(aq) + 4Br^-(aq) + 10H + (aq)$$

Why does the same reductant, thiosulphate react differently with iodine and bromine?

14. I_2 oxidises thiosulphate ion to tetrathionate ion i.e., from O.S. of + 2 for S $(\text{in } S_2O_3^{2-})$ to S.O. of + 5/ for S $(\text{in } S_4O_4^{2-} \text{ ion})$.

Br_2 oxidises thiosulphate ion to sulphate ion i.e., from O.S. of + 2 for S $(\text{in } S_2O_3^{2-})$ to O.S. of + 6 for S $(\text{in } SO_4^{2-})$

This is because Br_2 is a stronger
O.A. than I_2 $(E^\circ_{Br_2/2Br^-} = 1.09V,\ E^\circ_{I_2/2I^-} = 0.54V)$

15. Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant.

15. part(i)
Fluorine can oxidize other halogen ions. On the other hand Br_2, I_2, Cl_2 cannot oxidize $2F^- \rightarrow F_2$

$F_2 + 2I^- \rightarrow 2F^- + I_2$
$F_2 + 2Br^- \rightarrow 2F^- + Br_2$
$F_2 + 2Cl^- \rightarrow 2F^- + Cl_2$

And hence we say that fluorine is the better oxidant among halogen.
Part (ii)
HI & HBr are able to reduce $H_2SO_4 \rightarrow SO_2$ but HCl, HF are unable to reduce sulphuric acid.
So here we can say that HI & HBr are better reductant than HCl, HF
Again I^- can only able to reduce $Cu^{2+} \rightarrow Cu^+$ but Br^- cannot.
$4I^- + 2Cu^{2+} \rightarrow Cu_2I_2 + I_2$
Hence among hydrohalic compound hydroiodic acid is the best reductant

16. Why does the following reaction occur?
XeO_6^{4-} (aq) + $2F^-$(aq) + $6H^+$(aq) $\rightarrow$ XeO_3(g) + F_2(g) + $3H_2O$(l)
What conclusion about the compound Na_4XeO_6 (of which XeO_6^{4-} is a part) can be drawn from the reaction.

16. The given reaction occurs because XeO_6^{4-} oxidises F^- and F^-

$\overset{+8}{X}eO^{4-}_{6(aq)} + 2\overset{-1}{F}^-_{(aq)} + 6H^+_{(aq)} \longrightarrow \overset{+6}{X}eO_{3(g)} + \overset{0}{F}_{2(g)} + 3H_2O_{(l)}$ reduces XeO_6^{4-}

In this reaction, the oxidation number (O.N.) of Xe decreases from +8 in XeO_6^{4-} to +6 in XeO_3 and the O.N. of F increases from –1 in F^- to O in F_2.
Hence, we can conclude that Xa_4XeO_6 is a stronger oxidising agent than F^-.

17. Consider the reactions:
(a) H_3PO_2(aq) + $4AgNO_3$(aq) + $2H_2O$(l) $\rightarrow$ H_3PO_4(aq) + 4Ag(s) + $4HNO_3$(aq)
(b) H_3PO_2(aq) + $2CuSO_4$(aq) + $2H_2O$(l) $\rightarrow$ H_3PO_4(aq) + 2Cu(s) + H_2SO_4(aq)
(c) C_6H_5CHO(l) + $2[Ag(NH_3)_2]^+$(aq) + $3OH^-$ (aq) $\rightarrow$ $C_6H_5COO^-$(aq) + 2Ag(s) + $4NH_3$(aq) + $2H_2O$(l)
(d) $C_6H_5CHO(l)$ + $2Cu^{2+}$(aq) + $5OH^-$(aq) $\rightarrow$ No change observed.
What inference do you draw about the behaviour of Ag^+ and Cu^{2+} from these reactions?

17. In the first reaction, we can see that Ag^+ oxidizes the phosphorus from (+1 → +5) also in second, we clearly see that Cu^+ oxidize the phosphorus from (+1 → +5).

Both are oxidizing agents.

Ag^+ and Cu^{2+} act as oxidising agents in reactions (a) and (b) respectively. In reaction (c), Ag^+ oxidises C_6H_5CHO to $C_6H_5COO^-$, but in reaction (d), Cu^{2+} cannot oxidise C_6H_5CHO.

Hence, we can say that Ag^+ is a stronger oxidising agent than Cu^{2+}

18. Balance the following equations by ion-electron method

(a) $MnO_4^-(aq) + I^-(aq) \longrightarrow MnO_2 + I_2(s)$ (in basic medium)

(b) $MnO_4^-(aq) + SO_2(g) \longrightarrow Mn^{2+}(aq) + HSO_4^-(aq)$ (in acidic medium)

(c) $H_2O_2(aq) + Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + H_2O(l)$ (in acidic medium)

(d) $Cr_2O_7^{2-}(aq) + SO_2(g) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$ (in acidic medium)

18. (a) The skeleton equation is

$$MnO_4^-(aq) + I^-(aq) \longrightarrow MnO_2(s) + I_2(s)$$

Step 1 Separation of the equation in two half reactions

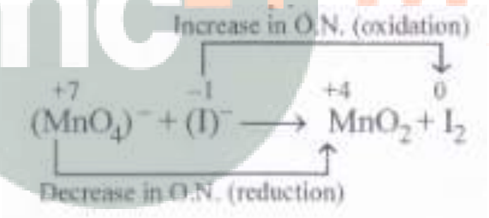
(i) Write the O.N of the atoms involved in the equation

$$(\overset{+7}{Mn}\overset{-2}{O_4})^- + (\overset{-1}{I})^- \longrightarrow (\overset{+4}{Mn}\overset{-2}{O_2})\overset{0}{I_2}$$

(ii) Identify the atoms which undergo change in O.N.

$$(\overset{+7}{Mn}\overset{-2}{O_4})^- + (\overset{-1}{I})^- \longrightarrow (\overset{+4}{Mn}\overset{-2}{O_2})\overset{0}{I_2}$$

(iii) Find out the species involved in the oxidation and reduction half reactions.



Oxidation half reaction: $I^- \longrightarrow I_2$

Reduction half reaction:

$$MnO_4^- \longrightarrow MnO_2$$

Step 2 Balancing the oxidation half reaction

$$2I^- \longrightarrow I_2 + 2e^-$$

Step 3 Balancing the reduction half reaction

The reduction half reaction is :

$$MnO_4^- \longrightarrow MnO_2$$

(i) As the decrease in O.N is 3, therefore add $3e^-$ on the reactant side

$$MnO_4^- + 3e^- \longrightarrow MnO_2$$

(ii) To balance the oxygen atoms, add two H_2O molecules on the product side

$$MnO_4^- + 3e^- \longrightarrow MnO_2 + 2H_2O$$

(iii) To balance the charges, add $4OH^-$ on the product side. Then to balance H atoms, add four H_2O molecules on the reactant side.

$$MnO_4^- + 3e^- + 4H_2O \longrightarrow MnO_2 + 4OH^- + 2H_2O$$

Thus, the reduction half reaction is balanced.

Step 4 Adding the two half reactions

In order to equate the electrons, multiply eqn. (i) by 3 and eqn. (ii) by 2. Add the two eqns.

$$2I^- \longrightarrow I_2 + 2e^-]\times 3$$

$$\underline{MnO_4^- + 3e^- + 2H_2O \longrightarrow MnO_2 + 4OH^-]\times 2}$$

$$2MnO_4^- + 6I^- + 4H_2O \longrightarrow 3I_2 + 2MnO_2 + 8OH^-$$

Or $2MnO_4^-(aq) + 6I^-(aq) + 4H_2O(l) \longrightarrow 3I_2(s) + 2MnO_2(s) + 8OH^-(aq)$

(b) The skeleton equation is :

$$MnO_4^-(aq) + SO_2(g) \longrightarrow Mn^{2+}(aq) + HSO_4^-(aq)$$

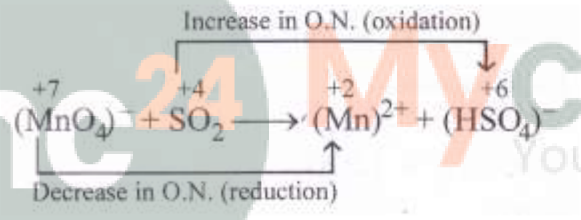
(i) Write the O.N. of the atoms involved in the equation

$$(\overset{+7}{Mn}\overset{-2}{O_4})^- + \overset{+4}{S}\overset{-2}{O_2} \longrightarrow (\overset{+2}{Mn})^{2+} + (\overset{+1}{H}\overset{+6}{S}\overset{-2}{O_4})^-$$

(ii) Identify the atoms which undergo change in O.N

$$(\overset{+7}{Mn}O_4)^- + \overset{+4}{S}O_2 \longrightarrow (\overset{+2}{Mn})^{2+} + (H\overset{+6}{S}\overset{-2}{O_4})^-$$

(iii) Find out the species involved in the oxidation and reduction half reactions.



Oxidation half reaction: $SO_2 \longrightarrow HSO_4^-$

Reduction half reaction: $MnO_4^- \longrightarrow Mn^{2+}$

Step 2: Balancing the oxidation half reaction

The oxidation half reaction is $SO_2 \longrightarrow HSO_4^-$

(i) As the increase in O.N is 2 therefore, add two electrons on the product side to balance change in O.N.

$$SO_2 \longrightarrow HSO_4^- + 2e^-$$

(ii) In order to balance the number of oxygen atoms, add two H_2O molecules on the reactant side and then to balance H atoms add $3H^+$ on the product side.

$$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-$$

Step 3 Balancing the reduction half reaction

The reduction half reaction is :

$$MnO_4^- \longrightarrow Mn^{2+}$$

(i) As the decrease in O.N. is 5, therefore add $5e^-$ on the reactant side,

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$$

(ii) In order to balance the no. of oxygen atoms, add four H_2O molecules on the product side and then to balance H atoms, add $8H^+$ on the reactant side.

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

Step 4 Adding the two half reactions

In order to equate the electrons, multiply eqn. (i) by 5 and eqn. (ii) by 2. Add the two eqns.

$$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-] \times 5$$

$$\underline{MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O] \times 2}$$

$$2MnO_4^-(aq) + 5SO_2(g) + 2H_2O(l) + H^+(aq) \longrightarrow 5HSO_4^-(aq) + 2Mn^{2+}(aq)$$

(c) Solve on similar lines.

(d) Solve on similar lines.

19. Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

(a) $P_4(s) + OH^+(aq) \longrightarrow PH3(g) + H2PO–2\ (aq)$

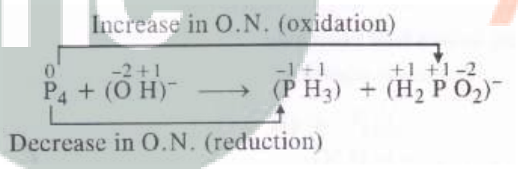
(b) $N_2H_4(7) + ClO_3^-(aq) \longrightarrow NO_{(g)} + Cl^-(g)$

(c) $C1_2O_7(g) + H_2O_2(aq) \longrightarrow ClO_2^-(aq) + O_2(g) + H^-$

19. The skeleton equation is :

$$P_4 + OH^- \longrightarrow PH_3 + H_2PO_4^{2-}$$

Step 1. Separate the equation into two half reactions :



• Here P_4 acts both as oxidising agent and reducing agent.

Step 2. Balancing the oxidation half reaction:

$$P_4 \longrightarrow (H_2PO_2)^-$$

$$P_4 + 8OH^- \longrightarrow 4H_2PO_2^- + 4e^-$$

Step 3. Balancing the reduction half reaction :

$$P_4 \longrightarrow PH_3$$

$$P_4 + 12e^- + 12H_2O \longrightarrow 4PH_3 + 12OH^-$$

Step 4. Adding the two half reactions :

In order to equation the electrons multiply eqn. (i) by 3. Add the two equations

$$P_4 + 8OH^- \longrightarrow 4H_2PO_2^- + 4e^-] \times 3$$

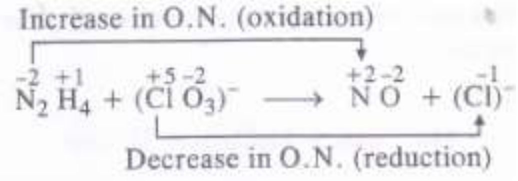
$$\underline{P_4 + 12e^- + 12H_2O \longrightarrow 4PH_3 + 12OH^-}$$

$$4P_4 + 12OH^- + 12H_2O \longrightarrow 4PH_3 + 12H_2PO_2^-$$

(b) The skeleton equation is :

$$N_2H_4 + ClO_3^- \longrightarrow NO + Cl^-$$

Step 1. Separate the equation into two half reactions :



• Here N_2H_4 is reducing agent and $(ClO_3)^-$ ions are oxidising agent.

Step 2. Balancing oxidation half reaction :

$$N_2H_4 \longrightarrow NO$$

$$N_2H_4 + 8OH^- \longrightarrow 2NO + 6H_2O + 8e^-$$

Step 3. Balancing reduction half reaction:

$$ClO_3^- + 6e^- + 3H_2O \longrightarrow Cl^- + 6OH^-$$

Step 4. Adding the two half reactions :

In order to equate electrons, multiply eqn.(ii) by 3 and eqn. (ii) by 4. Add the two equatons.

$$N_2H_4 + 8OH^- \longrightarrow 2NO + 6H_2O + 8e^-]\times 3$$

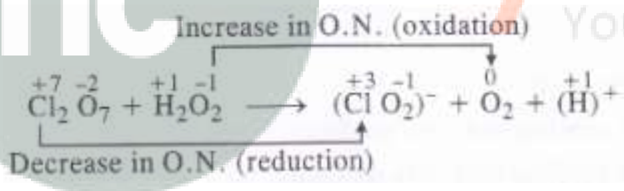
$$\underline{ClO_3^- + 6e^- + 3H_2O \longrightarrow Cl^- + 6OH^-]\times 4}$$

$$3N_2H_4 + 4ClO_3^- \longrightarrow 6NO + 6H_2O + 4Cl^-$$

(c) The skeleton equation is :

$$Cl_2O_7 + H_2O_2 \longrightarrow ClO_2^- + O_2 + H^+$$

Step 1. Separate the equation into two half reactions :



• Here H_2O_2 is reducing agent and Cl_2O_7 is oxidising agent.

Step 2. Balancing oxidation half reaction :

$$H_2O_2 \longrightarrow O_2$$

$$H_2O_2 + 2OH^- \longrightarrow O_2 + 2H_2O + 2e^-$$

Step 3. Balancing reduction half reaction :

$$Cl_2O_7 \longrightarrow ClO_2^-$$

$$Cl_2O_7 + 3H_2O + 8e^- \longrightarrow 2ClO_2^- + 6OH^-$$

Step 4. Adding the two half reactions:

In order to equate electrons multiply eqn. (i) by 4 and add to the eqn.(ii),

$$H_2O_2 + 2OH^- \longrightarrow O_2 + 2H_2O + 2e^-]\times 4$$

$$\underline{Cl_2O_7 + 3H_2O + 8e^- \longrightarrow 2ClO_2^- + 6OH^-}$$

$$Cl_2O_7 + 4H_2O_2 + 2OH^- \longrightarrow 4ClO_2^- + 4O_2 + 5H_2O$$

Note: Balance the equations by oxidation number method yourself.

20. What sorts of informations can you draw from the following reaction ?

$$(CN)_{2(g)} + 2OH^-_{(aq)} \longrightarrow CN^-_{(aq)} + CNO^-_{(aq)} + H_2O(l)$$

20. The oxidation numbers of carbon in $(CN)^2$, CN^- and CNO^- are +3, +2 and +4 respectively.

These are obtained as shown below:

Let the oxidation number of C be x.

$(CN)_2$

$2(x - 3) = 0$

$x = 3$

$CN– x - 3 = -1$

$x = 2$

CNO^-

$x - 3 - 2 = -1$

$x = 4$

The oxidation number of carbon in the various species is:

$$\left(\overset{+3}{C}N\right)_{2(g)} + 2OH^-_{(aq)} \longrightarrow \overset{+2}{C}N^-_{(aq)} + \overset{+4}{C}NO^-_{(aq)} + H_2O_{(l)}$$

It can be easily observed that the same compound is being reduced and oxidized simultaneously in the given equation. Reactions in which the same compound is reduced and oxidised is known as disproportionation reactions. Thus, it can be said that the alkaline decomposition of cyanogen is an example of disproportionation reaction.

21. The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+}, MnO_2, and H^+ ion. Write a balanced ionic equation for the reaction.

21. $Mn^{3+} \to Mn^{2+} + MnO_2 + H^+$

write oxidation half with their oxidation state

$Mn^{3+} (+3) \to MnO_2(+4)$

Balance the charge on Mn by adding $1e^-$ on RHS side. To balance charge add H^+ ions on RHS side and then for oxygen balance add H_2O molecule on LHS side.

$Mn^{3+} + 2H_2O \to MnO_2 + 1e^- + 4H^+$

reduction half

$Mn^{3+}(+3) \to Mn^{2+}(+2)$

balancing the reduction half by adding $1e^-$ on LHS side

$Mn^{3+} + 1e^- \to Mn^{2+}$

Add both balanced reduction half and oxidation half

$2Mn^{3+} + 2H_2O \to Mn^{2+} + MnO_2 + 4H^+$

22. Consider the elements:

Cs, Ne, I and F

(a) Identify the element that exhibits only negative oxidation state.

(b) Identify the element that exhibits only positive oxidation state.

(c) Identify the element that exhibits both positive and negative oxidation states. (d)

Identify the element which exhibits neither the negative nor does the positive oxidation state.

22. (a) F exhibits only negative oxidation state of –1.

(b) Cs exhibits positive oxidation state of +1.

(c) I exhibits both positive and negative oxidation states. It exhibits oxidation states of –1, + 1, + 3, + 5, and + 7.

(d) The oxidation state of Ne is zero. It exhibits neither negative nor positive oxidation states.

23. Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

23. Base equation- $SO_2 + Cl_2 \rightarrow Cl^- + SO_4^{2-}$ (have to remember)

Now we have to balance the oxidation half and reduction half.

Oxidation half = $SO_2 \rightarrow SO_4^{2-}$

balancing - Oxygen is balanced by adding water molecule, Hydrogen is balanced by H^+ ion and for charge add e^- (electron) $SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$

Reduction hallf = $Cl_2 \rightarrow Cl^-$

Balancing - to balance charge add an electron $Cl_2 + 2e^- \rightarrow 2Cl$

Now add both balanced oxidation half and reduction half, we get

$Cl_2 + SO_4 + 2H_2O \rightarrow 2Cl^- + SO_4^{2-} + 4H^+$

24. Refer to the periodic table given in your book and now answer the following questions:

(a) Select the possible non metals that can show disproportionation reaction.

(b) Select three metals that can show disproportionation reaction.

24. In disproportionation reactions, one of the reacting substances always contains an element that can exist in at least three oxidation states.

(a) P, Cl, and S can show disproportionation reactions as these elements can exist in three or more oxidation states.

(b) Mn, Cu, and Ga can show disproportionation reactions as these elements can exist in three or more oxidation states.

25. In Ostwald's process for the manufacture of nitric acid, the first step involves the oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen?

25. we have,

number of moles(n) = given mass/ molecular mass ------------------------------(eq.1)

No. of moles of ammonia (NH_3) = 10/17 = 0.588

No. of moles of oxygen (O_2) = 20/32= 0.625

Balanced Reaction $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

Here we see that 4 moles of ammonia required 5 moles of oxygen. So

0.588 moles of ammonia = (5/4) + 0.588 = 0.735 moles of (O_2). But we have only 0.625 moles of (O_2).

It means oxygen is a limiting reagent and the maximum weight of nitric oxide (NO) can be produced by 0.635 moles of (O_2)

So, 5 moles of (O_2) produced 4 moles of C.

therefore 0.625 moles of (O_2) = (4/5) + 0.625 = 0.5 moles of (NO).

from Eq. 1

mass of (NO) = number of moles (NO) * molecular weight (NO)

= 0.5 * 30

= 15 g

Alternate Method

directly consider the molecular weight

(17*4) g of NH_3 required (5*32) g of O to produce (30*4) g of NO

So, 10g of NH_3 required = (5*32/17*4)*10 = 23.5g of O. But we have only 20g (means O is limiting reagent) whatever the max. NO produce is from 20g of O.

and we know that 5*32g of O produce 30*4 g of NO

So, 20g of O produce = (30*4/5*32)*20 g of NO = 15g of NO

26. Using the standard electrode potentials given in the Table 8.1, predict if the reaction between the following is feasible:

(a) Fe^{3+}(aq) and I^-(aq)

(b) Ag^+(aq) and Cu(s)

(c) Fe^{3+} (aq) and Cu(s)

(d) Ag(s) and Fe^{3+}(aq)

(e) Br_2(aq) and Fe^{2+}(aq)

26. (a) The possible reaction between $Fe^{3+}_{(aq)} + I^-_{(aq)}$

$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + I_{2(s)}$ is given by,

Oxidation half equation $2I^-_{(aq)} \longrightarrow I_{2(s)} + 2e^-; E^\circ = -0.54V$

Reduction half equation $[Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(aq)}] \times 2;\ E^\circ = +0.77V$

$$2Fe^{3+}_{(aq)} + 2I^-_{(aq)} \longrightarrow Fe^{2+}_{(aq)} + I_{2(s)}; E^\circ = +0.23V$$

E° for the overall reaction is positive. Thus, the reaction between $Fe^{3+}_{(aq)}$ and $I^-_{(aq)}$ is feasible.

(b)The possible reaction between $Ag^+_{(aq)} + Cu_{(s)}$ is given by,

Oxidation half equation : $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$; $E^\circ = -0.34V$

Reduction half equation : $[Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}] \times 2$; $E^\circ = +0.80V$

$$2Ag^+_{(aq)} + Cu_{(s)} \longrightarrow 2Ag_{(s)} + Cu^{2+}; E^\circ = +0.46V$$

E° positive for the overall reaction is positive. Hence, the reaction between $Ag^+_{(aq)}$ and Cu(s) is feasible.

(c) The possible reaction between $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$ is given by,

$2Fe^{3+}_{(aq)} + Cu_{(s)} \longrightarrow 2Fe^{2x}_{(s)} + Cu^{2+}_{(aq)}$

Oxidation half equation : $Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$; $E° = -0.34V$

Reduction half equation : $[Fe^{3+}_{(aq)} + e^- \longrightarrow Fe^{2+}_{(s)}] \times 2$; $E° = +0.77V$

$$2Fe^{3+}_{(aq)} + Cu_{(s)} \longrightarrow 2Fe^{2+}_{(s)} + Cu^{2+}_{(s)}; E° = +0.43V$$

E° positive for the overall reaction is positive. Hence, the reaction between $Fe^{3+}_{(aq)}$ and $Cu_{(s)}$ is feasible.

(d) The possible reaction between $Ag_{(s)}$ and $Fe^{3+}_{(aq)}$

$Ag_{(s)} + 2Fe^{3+}_{(aq)} \longrightarrow Ag^{+}_{(aq)} + Fe^{2+}_{(aq)}$ is given by,

Oxidation half equation : $Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^-] \times 2$; $E° = -0.77V$

Reduction half equation : $Br_{2(aq)} + 2e^- \longrightarrow 2Br^-_{(aq)}$; $E° = +1.09V$

$$Br_{2(aq)} + 2Fe^{2+}_{(aq)} \longrightarrow 2Br^-_{(aq)} + 2Fe^{3+}_{(aq)}; E° = -0.32V$$

Here, E° for the overall reaction is positive. Hence, the reaction between $Br_{2(aq)}$ and $Fe^{2+}_{(aq)}$ is feasible.

27. Predict the products of electrolysis in each of the following:

(i) An aqueous solution of $AgNO_3$ with silver electrodes

(ii) An aqueous solution $AgNO_3$ with platinum electrodes

(iii) A dilute solution of H_2SO_4 with platinum electrodes

(iv) An aqueous solution of $CuCl_2$ with platinum electrodes.

27. (i) $AgNO_3$ dissociate into Ag^+ and NO_3^-

@ Cathode:- $Ag^+(aq) + e^- \rightarrow Ag(s)$ (reduction potential of silver is higher than water molecule. So silver electrode oxidized) $E^0 = -0.83V$

(ii) since platinum (Pt) electrode cannot easily oxidize. So at the anode H_2O will oxidize and liberate oxygen and at cathode Ag will be deposited.

At cathode- $Ag^+ + e^+ \rightarrow Ag$

At anode- $H_2O \rightarrow O_2 + 4H^+ + 4e^-$

(iii) given sulphuric acid is dilute.

ionize into $H_2SO_2 \rightarrow 2H^+ + SO_4^{2-}$

At cathode $H^+ + e^- \rightarrow 1/2H_2(g)$

At anode, There will be -(liberation of oxygen gas)

$H_2O \rightarrow 1/2O_2 + 4H^+ + 4e^-$

Predict the products of electrolysis in each of the following:

An aqueous solution of $CuCl_2$ with platinum electrodes

(iv) In aqueous solution $CuCl_2$ ionise into Cu^{2+} and $2Cl^-$

At the cathode , the copper ion will be deposited because it has a higher reduction potential than the water molecule

At the anode , the lower electrode potential value will be preferred but due to overpotential of oxygen, chloride ion gets oxidized at the anode.

@ Anode- $2Cl^- \rightarrow Cl_2 + 2e^-$

@ cathode- $Cu^{2+} + 2e^- \rightarrow Cu$

28. Arrange the following metals in the order in which they displace each other from the solution of their salts.
Al, Cu, Fe, Mg and Zn.

28. A metal of stronger reducing power displaces another metal of weaker reducing power from its solution of salt.
The order of the increasing reducing power of the given metals is Cu < Fe < Zn < Al < Mg.
Hence, we can say that Mg can displace Al from its salt solution, but Al cannot displace Mg.
Thus, the order in which the given metals displace each other from the solution of their salts is given below:
Mg > Al > Zn > Fe, > Cu

29. aGiven the standard electrode potentials,
$K^+/K = -2.93V$, $Ag^+/Ag = 0.80V$, $Hg^{2+}/Hg = 0.79V$
$Mg^{2+}/Mg = -2.37V$. $Cr^{3+}/Cr = -0.74V$
Arrange these metals in their increasing order of reducing power.

29. A negative electrode potential (E^0) means redox couple is a stronger reducing agent. So as per data the increasing order of the following is–
Ag < Hg < Cr < Mg < K

30. Depict the galvanic cell in which the reaction Zn(s) + 2Ag+(aq) → Zn2+(aq) + 2Ag(s) takes place, further show:
(i) which of the electrode is negatively charged,
(ii) the carriers of the current in the cell, and
(iii) individual reaction at each electrode.

30. The galvanic cell for the reaction is
$Zn(s)|Zn^{2+}(aq)\,||\,2Ag^+(aq)|2Ag(s)$
Zinc electrode (anode) is negatively charged.
Current flows from silver to zinc in outer circuit
Anode : $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (oxidation)
Cathode : $2Ag^+(ag) + 2e^- \rightarrow 2Ag(s)$ (reduction)