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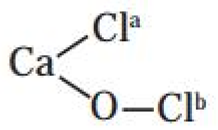
Maharashtra State Board 11th Chemistry Solutions Chapter 6 Redox Reactions

1. Choose the most correct option

Question A.
Oxidation numbers of Cl atoms marked as Cl_a and Cl_b in $CaOCl_2$ (bleaching powder) are



a. zero in each
b. -1 in Cl_a and +1 in Cl_b
c. +1 in Cl_a and -1 in Cl_b
d. 1 in each
Answer:
b. -1 in Cl_a and +1 in Cl_b

Question B.
Which of the following is not an example of redox reacton ?
a. $CuO + H_2 \rightarrow Cu + H_2O$
b. $Fe_2O_3 + 3CO_2 \rightarrow 2Fe + 3CO_2$
c. $2K + F_2 \rightarrow 2KF$
d. $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$
Answer:
d. $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

Question C.
A compound contains atoms of three elements A, B and C. If the oxidation state of A is +2, B is +5 and that of C is -2, the compound is possibly represented by
a. $A_2(BC_3)_2$
b. $A_3(BC_4)_2$
c. $A_3(B_4C)_2$
d. ABC_2
Answer:
b. $A_3(BC_4)_2$

Question D.
The coefficients p, q, r, s in the reaction
$pCr_2O_7^{2-} + q\ Fe^{2\oplus} \rightarrow r\ Cr^{3\oplus} + s\ Fe^{3\oplus} + H_2O$ respectively are :
a. 1, 2, 6, 6
b. 6, 1, 2, 4
c. 1, 6, 2, 6
d. 1, 2, 4, 6
Answer:
c. 1, 6, 2, 6

Question E.
For the following redox reactions, find the correct statement.
$Sn^{2\oplus} + 2Fe^{3\oplus} \rightarrow Sn^{4\oplus} + 2Fe^{2\oplus}$
a. $Sn^{2\oplus}$ is undergoing oxidation
b. $Fe^{3\oplus}$ is undergoing oxidation
c. It is not a redox reaction
d. Both $Sn^{2\oplus}$ and $Fe^{3\oplus}$ are oxidised
Answer:
a. $Sn^{2\oplus}$ is undergoing oxidation

Question F.
Oxidation number of carbon in H_2CO_3 is
a. +1
b. +2
c. +3
d. +4
Answer:
d. +4

Question G.
Which is the correct stock notation for magenese dioxide ?

a. Mn(I)O_2
b. Mn(II)O_2
c. Mn(III)O_2
d. Mn(IV)O_2
Answer:
d. Mn(IV)O_2

Question I.
Oxidation number of oxygen in superoxide is
a. -2
b. -1
c. −12
d. 0
Answer:
c. −12

Question J.
Which of the following halogens does always show oxidation state -1 ?
a. F
b. Cl
c. Br
d. I
Answer:
a. F

Question K.
The process $SO_2 \rightarrow S_2Cl_2$ is
a. Reduction
b. Oxidation
c. Neither oxidation nor reduction
d. Oxidation and reduction.
Answer:
a. Reduction

2. Write the formula for the following compounds :
A. Mercury(II) chloride
B. Thallium(I) sulphate
C. Tin(IV) oxide
D. Chromium(III) oxide
Answer:
i. $HgCl_2$
ii. Tl_2SO_4
iii. SnO_2
iv. Cr_2O_3

3. Answer the following questions

Question A.
In which chemical reaction does carbon exibit variation of oxidation state from -4 to +4 ? Write balanced chemical reaction.
Answer:
In combustion of methane, carbon exhibits variation from -4 to +4. The reaction is as follows:
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
In CH_4, the oxidation state of carbon is -4 while in CO_2, the oxidation state of carbon is +4.

Question B.
In which reaction does nitrogen exhibit variation of oxidation state from -3 to +5 ?
Answer:

$$\underset{\text{Ammonia}}{\overset{-3}{4NH_{3(g)}}} + \underset{\text{Oxygen}}{5O_{2(g)}} \xrightarrow[\text{500 K, 9 bar}]{\text{Pt/Rh gauge catalyst}} \underset{\text{Nitric oxide}}{\overset{+2}{4NO_{(g)}}} + 6H_2O_{(g)}$$

$$\underset{\text{Nitric oxide}}{2NO_{(g)}} + \underset{\text{Oxygen}}{O_{2(g)}} \rightleftharpoons \underset{\text{Nitrogen dioxide}}{\overset{+4}{2NO_{2(g)}}} \quad ; \quad \underset{\text{Nitrogen dioxide}}{3NO_{2(g)}} + \underset{\text{Water}}{H_2O_{(l)}} \longrightarrow \underset{\text{Nitric acid}}{\overset{+5}{2HNO_{3(l)}}} + \underset{\text{Nitric oxide}}{NO_{(g)}}$$

C. Calculate the oxidation number of underlined atoms.
a. $H_2\underline{S}O_4$
b. $H\underline{N}O_3$

c. $H_3\underline{P}O_3$
d. $K_2\underline{C}_2O_4$
e. $H_2\underline{S}_4O_6$
f. $\underline{Cr}_2O_7^{2-}$
g. $NaH_2\underline{P}O_4$
Answer:
i. $H_2\underline{S}O_4$
Oxidation number of H = +1
Oxidation number of O = -2
H_2SO_4 is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms of H_2SO_4 = 0
∴ 2 × (Oxidation number of H) + (Oxidation number of S) + 4 × (Oxidation number of O) = 0
∴ 2 × (+1) + (Oxidation number of S) + 4 × (-2) = 0
∴ Oxidation number of S + 2 – 8 = 0
∴ Oxidation number of S in H_2SO_4 = +6

ii. $H\underline{N}O_3$
Oxidation number of H = +1
Oxidation number of O = -2
HNO_3 is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms of HNO_3 = 0
∴ (Oxidation number of H) + (Oxidation number of N) + 3 × (Oxidation number of O) = 0
∴ (+1) + (Oxidation number of N) + 3 × (-2) = 0
∴ Oxidation number of N + 1 – 6 = 0
∴ Oxidation number of N in HNO_3 = +5

iii. $H_3\underline{P}O_3$
Oxidation number of O = -2
Oxidation number of H = +1
H_3PO_3 is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms = 0
∴ 3 × (Oxidation number of H) + (Oxidation number of P) + 3 × (Oxidation number of O) = 0
∴ 3 × (+1) + (Oxidation number of P) + 3 × (-2) = 0
∴ Oxidation number of P + 3 – 6 = 0
Oxidation number of P is H_3PO_3 = +3

iv. $K_2\underline{C}_2O_4$
Oxidation number of K = +1
Oxidation number of O = -2
$K_2C_2O_4$ is a neutral molecule.
∴ Sum of the oxidation number of all atoms = 0
∴ 2 × (Oxidation number of K) + 2 × (Oxidation number of C) + 4 × (Oxidation number of O) = 0
∴ 2 × (+1) + 2 × (Oxidation number of C) + 4 × (-2) = 0
∴ 2 × (Oxidation number of C) + 2 – 8 = 0
∴ 2 × (Oxidation number of C) = + 6
∴ Oxidation number of C = +62
∴ Oxidation number of C in $K_2C_2O_4$ = +3

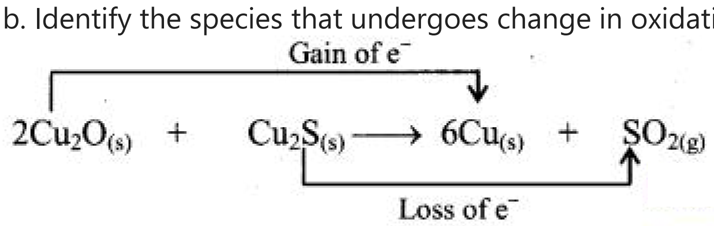
v. $H_2\underline{S}_4O_6$
Oxidation number of H = +1
Oxidation number of O = -2
$H_2S_4O_6$ is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms = 0
∴ 2 × (Oxidation number of H) + 4 × (Oxidation number of S) + 6 × (Oxidation number of O) = 0
∴ 2 × (+1) + 4 × (Oxidation number of S) + 6 × (-2) = 0
∴ 4 × (Oxidation number of S) + 2 – 12 = 0
∴ 4 × (Oxidation number of S) = + 10
∴ Oxidation number of S = +104
∴ Oxidation number of S in $H_2S_4O_6$ = +2.5

vi. $\underline{Cr}_2O_7^{2-}$
Oxidation of O = -2
$Cr_2O_7^{2-}$ is an ionic species.
∴ Sum of the oxidation numbers of all atoms = – 2
∴ 2 × (Oxidation number of Cr) + 7 × (Oxidation number of O) = -2
∴ 2 × (Oxidation number of Cr) + 7 × (-2) = – 2
∴ 2 × (Oxidation number of Cr) – 14 = – 2
∴ 2 × (Oxidation number of Cr) = – 2 + 14
∴ Oxidation number of Cr = +122
∴ Oxidation number of Cr in $Cr_2O_7^{2-}$ = +6

vii. $NaH_2\underline{P}O_4$
Oxidation number of Na = +1
Oxidation number of H = +1
Oxidation number of O = -2
NaH_2PO_4 is a neutral molecule
Sum of the oxidation numbers of all atoms = 0
(Oxidation number of Na) + 2 × (Oxidation number of H) + (Oxidation number of P) + 4 × (Oxidation number of O) = 0
(+1) + 2 × (+1) + (Oxidation number of P) + 4 × (-2) = 0
(Oxidation number of P) + 3 – 8 = 0
Oxidation number of P in NaH_2PO_4 = +5

Question D.
Justify that the following reactions are redox reaction; identify the species oxidized/reduced, which acts as an oxidant and which act as a reductant.
a. $2Cu_2O_{(s)} + Cu_2S_{(s)} \rightarrow 6Cu_{(s)} + SO_{2(g)}$
b. $HF_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + F^-_{(aq)}$
c. $I_{2(aq)} + 2\ S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + 2I^-_{(aq)}$
Answer:
i. $2Cu_2O_{(s)} + Cu_2S_{(s)} \rightarrow 6Cu_{(s)} + SO_{2(g)}$
a. Write oxidation number of all the atoms of reactants and products.

$$2\underset{\substack{\uparrow\\+1}}{Cu_2}\underset{\substack{\uparrow\\-2}}{O_{(s)}} + \underset{\substack{\uparrow\\+1}}{Cu_2}\underset{\substack{\uparrow\\-2}}{S_{(s)}} \longrightarrow 6\underset{\substack{\uparrow\\0}}{Cu_{(s)}} + \underset{\substack{\uparrow\\+4}}{S}\underset{\substack{\uparrow\\-2}}{O_{2(g)}}$$

b. Identify the species that undergoes change in oxidation number.


c. The oxidation number of S increases from -2 to +4 and that of Cu decreases from +1 to 0. Because oxidation number of one species increases and that of the other decreases, the reaction is a redox reaction.
d. The oxidation number of S increases by loss of electrons and therefore, S is a reducing agent and it itself is oxidised. On the other hand, the oxidation number of Cu decreases by gain of electrons and therefore, Cu is an oxidising agent and itself is reduced.

Result:

1. The given reaction is a redox reaction.
2. Oxidant/oxidising agents (Reduced species): Cu_2O/ Cu_2S
3. Reductant/reducing agent (Oxidised species): Cu_2S

[Note: Cu in both Cu_2O and Cu_2S undergoes reduction. Hence, both Cu_2O and Cu_2S can be termed as oxidising agents in the given reaction.]

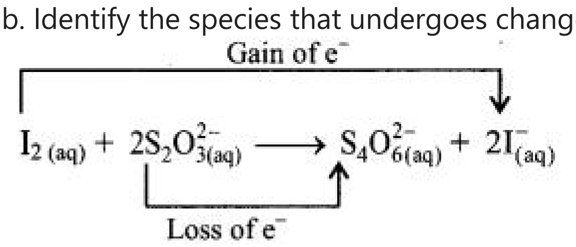
ii. $HF_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + F^-_{(aq)}$
a. Write oxidation number of all the atoms of reactants and products.

$$\underset{\substack{\uparrow\\+1}}{H}\underset{\substack{\uparrow\\-1}}{F_{(aq)}} + \underset{\substack{\uparrow\\-2}}{O}\underset{\substack{\uparrow\\+1}}{H^-_{(aq)}} \longrightarrow \underset{\substack{\uparrow\\+1}}{H_2}\underset{\substack{\uparrow\\-2}}{O_{(l)}} + \underset{\substack{\uparrow\\-1}}{F^-_{(aq)}}$$

b. Since, the oxidation numbers of all the species remain same, this is NOT a redox reaction. Result:
The given reaction is NOT a redox reaction.

iii. $I_{2(aq)} + 2\ S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + 2I^-_{(aq)}$
a. Write oxidation number of all the atoms of reactants and products.

$$\underset{\substack{\uparrow\\0}}{I_{2(aq)}} + 2\underset{\substack{\uparrow\\+2}}{S_2}\underset{\substack{\uparrow\\-2}}{O^{2-}_{3(aq)}} \longrightarrow \underset{\substack{\uparrow\\+2.5}}{S_4}\underset{\substack{\uparrow\\-2}}{O^{2-}_{6(aq)}} + 2\underset{\substack{\uparrow\\-1}}{I^-_{(aq)}}$$

b. Identify the species that undergoes change in oxidation number.


c. The oxidation number of S increases from +2 to +2.5 and that of I decreases from 0 to -1. Because oxidation number of one species increases and that of the other decreases, the reaction is a redox reaction.
d. The oxidation number of S increases by loss of electrons and therefore, S is a reducing agent and itself is oxidised. On the other hand, the oxidation number of I decreases by gain of electrons and therefore, I is an oxidising agent and itself is reduced.

Result:

1. The given reaction is a redox reaction.

2. Oxidant/oxidising agent (Reduced species): I_2
3. Reductant/reducing agent (Oxidised species): $S_2O_3^{2-}$

Question E.
What is oxidation? Which one of the following pairs of species is in its oxidized state ?
a. Mg / Mg^{2+}
b. Cu / Cu^{2+}
c. O_2 / O^{2-}
d. Cl_2 / Cl^-
Answer:
a. Mg / Mg^{2+}
Here, Mg loses two electrons to form Mg^{2+} ion.
$Mg_{(s)} \longrightarrow Mg^{2+}_{(aq)} + 2e^-$
Hence, Mg / Mg^{2+} is an oxidized state.

b. Cu/Cu^{2+}
Here, Cu loses two electrons to form Cu^{2+} ion.
$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$
Hence, Cu/Cu^{2+} is in an oxidized state.

c. O_2 / O^{2-}
Here, each O gains two electrons to form O^{2-} ion.
$O_{2(g)} + 4e^- \longrightarrow 2O^{2-}_{(aq)}$
Hence, O_2 / O^{2-} is in a reduced state.

d. Cl_2 / Cl^-
Here, each Cl gains one electron to form Cl^- ion.
$Cl_{2(g)} + 2e^- \longrightarrow 2Cl^-_{(aq)}$
Hence, Cl_2 / Cl^- is in a reduced state.

Question F.
Justify the following reaction as redox reaction.
$2 Na_{2(s)} + S_{(s)} \rightarrow Na_2S_{(s)}$
Find out the oxidizing and reducing agents.
Answer:
i. Redox reaction can be described as electron transfer as shown below:
$2Na_{(s)} + S_{(s)} \rightarrow 2Na^+ + S^{2-}$
ii. Charge development suggests that each sodium atom loses one electron to form Na^+ and sulphur atom gains two electrons to form S^{2-}. This can be represented as follows:

Loss of e^-

$$2Na_{(s)} + S_{(s)} \longrightarrow 2Na^+ + S^{2-}$$

Gain of e^-

iii. When Na is oxidised to Na_2S, the neutral Na atom loses electrons to form Na^+ in Na_2S while the elemental sulphur gains electrons and forms S^{2-} in Na_2S.
iv. Each of the above steps represents a half reaction which involves electron transfer (loss or gain).
v. Sum of these two half reactions or the overall reaction is a redox reaction.
vi. Oxidising agent is an electron acceptor and hence, S is an oxidising agent. Reducing agent is an electron donor and hence, Na is a reducing agent.

Question G.
Provide the stock notation for the following compounds : $HAuCl_4$, Tl_2O, FeO, Fe_2O_3, MnO and CuO.
Answer:

i. $\overset{+1}{H}\overset{+3}{Au}\overset{-1}{Cl_4} \longrightarrow HAu(III)Cl_4$

ii. $\overset{+1}{Tl_2}\overset{-2}{O} \longrightarrow Tl_2(I)O$

iii. $\overset{+2}{Fe}\overset{-2}{O} \longrightarrow Fe(II)O$

iv. $\overset{+3}{Fe_2}\overset{-2}{O_3} \longrightarrow Fe_2(III)O_3$

v. $\overset{+2}{Mn}\overset{-2}{O} \longrightarrow Mn(II)O$

vi. $\overset{+2}{Cu}\overset{-2}{O} \longrightarrow Cu(II)O$

Question H.
Assign oxidation number to each atom in the following species.
a. $Cr(OH)_4^-$
b. $Na_2S_2O_3$
c. H_3BO_3
Answer:
i. $Cr(OH)_4^-$

Oxidation number of O = -2
Oxidation number of H = +1
$Cr(OH)_4^-$ is an ionic species.
∴ Sum of the oxidation numbers of all atoms = – 1
∴ Oxidation number of Cr + 4 × (Oxidation number of O) + 4 × (Oxidation number of H) = – 1
∴ Oxidation number of Cr + 4 × (-2) + 4 × (+1) = – 1
∴ Oxidation number of Cr – 8 + 4 = – 1
∴ Oxidation number of Cr – 4 = – 1 –
∴ Oxidation number of Cr = – 1 + 4
∴ Oxidation number of Cr in $Cr(OH)_4^-$ = +3

ii. $Na_2S_2O_3$
Oxidation number of Na = +1
Oxidation number of O = -2
$Na_2S_2O_3$ is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms = 0
∴ 2 × (Oxidation number of Na) + 2 × (Oxidation number of S) + 3 × (Oxidation number of O) = 0
∴ 2 × (+1) + 2 × (Oxidation number of S) + 3 × (-2) = 0
∴ 2 × (Oxidation number of S) + 2 – 6 = 0
∴ 2 × (Oxidation number of S) = + 4
∴ Oxidation number of S = +42
∴ Oxidation number of S in $Na_2S_2O_3$ = +2

iii. H_3BO_3
Oxidation number of H = +1
Oxidation number of O = -2
H_3BO_3 is a neutral molecule.
∴ Sum of the oxidation numbers of all atoms = 0
∴ 3 × (Oxidation number of H) + (Oxidation number of B) + 3 × (Oxidation number of O) = 0
∴ 3 × (+1) + (Oxidation number of B) + 3 × (-2) = 0
∴ Oxidation number of B + 3 – 6 = 0
∴ Oxidation number of B in H_3BO_3 = +3

Question I.
Which of the following redox couple is stronger oxidizing agent ?
a. Cl_2 (E_0 = 1.36 V) and Br_2 (E_0 = 1.09 V)
b. $MnO_4^{\ominus}$ (E0 = 1.51 V) and $Cr_2O_7^{2\ominus}$ (E_0 = 1.33 V)
Answer:
a. Cl_2 has a larger positive value of E_0 than Br_2. Thus, Cl_2 is a stronger oxidizing agent than Br_2.
b. $MnO_4^{\ominus}$ has larger positive value of E_0 than $Cr_2O_7^{2\ominus}$. Thus, $MnO_4^{\ominus}$ is stronger oxidizing agent than $Cr_2O_7^{2\ominus}$

Question J.
Which of the following redox couple is stronger reducing agent ?
a. Li (E_0 = – 3.05 V) and Mg(E_0 = – 2.36 V)
b. Zn(E_0 = – 0.76 V) and Fe(E_0 = – 0.44 V)
Answer:
a. Li has a larger negative value of E_0 than Mg. Thus, Li is a stronger reducing agent than Mg.
b. Zn has a larger negative value of E_0 than Fe. Thus, Zn is a stronger reducing agent than Fe.

4. Balance the reactions/equations :

Question A.
Balance the following reactions by oxidation number method

a. $Cr_2O_7^{2\ominus}(aq) + SO_3^{2\ominus}(aq) \longrightarrow Cr^{3\oplus}(aq) + SO_4^{2\ominus}(aq)$ (acidic)

b. $MnO_4^{\ominus}(aq) + Br^{\ominus}(aq) \longrightarrow MnO_2(s) + BrO_3^{\ominus}(aq)$ (basic)

c. $H_2SO_4(aq) + C(s) \longrightarrow CO_2(g) + SO_2(g) + H_2O(l)$ (acidic)

d. $Bi(OH)_3(g) + Sn(OH)_3^{\ominus}(aq) \longrightarrow Bi(s) + Sn(OH)_6^{2\ominus}(aq)$ (basic)

Answer:
i. $Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \longrightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$ (acidic)
Step 1: Write skeletal equation and balance the elements other than O and H.

$Cr_2O_7^{2-}{}_{(aq)} + SO_3^{2-}{}_{(a)} \longrightarrow 2Cr^{3+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$

Step 2: Assign oxidation number to Cr and S. Calculate the increase and decrease in the oxidation number and make them equal.

$$\underset{+6\ -2}{Cr_2O_7^{2-}{}_{(aq)}} + \underset{+4\ -2}{SO_{3(aq)}^{2-}} \longrightarrow \underset{+3}{2Cr_{(aq)}^{3+}} + \underset{+6}{SO_{4(aq)}^{2-}}$$

Increase in oxidation number: $\underset{+4}{SO_3^{2-}} \longrightarrow \underset{+6}{SO_4^{2-}}$

(Increase per atom = 2)

Decrease in oxidation number: $\underset{+6}{Cr_2O_7^{2-}} \longrightarrow \underset{+3}{2Cr_{(aq)}^{3+}}$

(Decrease per atom = 3)

To make the net increase and decrease equal, we must take 3 atoms of S and 2 atoms of Cr. (There are already 2 Cr atoms.)

Step 3: Balance 'O' atoms by adding $4H_2O$ to the right-hand side.

$Cr_2O_7^{2-}{}_{(aq)} + 3SO_3^{2-}{}_{(aq)} \longrightarrow 2Cr^{3+}{}_{(aq)} + 3SO_4^{2-}{}_{(aq)} + 4H_2O_{(l)}$

Step 4: The medium is acidic. To make the charges and hydrogen atoms on the two sides equal, add 8H on the left-hand side.

$Cr_2O_7^{2-}{}_{(aq)} + 3SO_3^{2-}{}_{(aq)} + 8H^+{}_{(aq)} \longrightarrow 2Cr^{3+}{}_{(aq)} + 3SO_4^{2-}{}_{(aq)} + 4H_2O_{(l)}$

Step 5: Check two sides for balance of atoms and charges.

Hence, balanced equation:

$$\mathbf{Cr_2O_{7(aq)}^{2-} + 3SO_{3(aq)}^{2-} + 8H_{(aq)}^{+} \longrightarrow 2Cr_{(aq)}^{3+} + 3SO_{4(aq)}^{2-} + 4H_2O_{(l)}}$$

ii. $MnO_4^-{}_{(aq)} + Br^-{}_{(aq)} \longrightarrow MnO_{2(s)} + BrO_3^-{}_{(aq)}$ (basic)

Step 1: Write skeletal equation and balance the elements other than O and H.

$MnO_4^-{}_{(aq)} + Br^-{}_{(aq)} \longrightarrow MnO_{2(s)} + BrO_3^-{}_{(aq)}$

Step 2: Assign oxidation number to Mn and Br. Calculate the increase and decrease in the oxidation number and make them equal.

$$\underset{+7\ -2}{MnO_{4(aq)}^-} + \underset{-1}{Br_{(aq)}^-} \longrightarrow \underset{+4}{MnO_{2(s)}} + \underset{+5}{BrO_{3(aq)}^-}$$

Increase in oxidation number: $\underset{-1}{Br^-} \longrightarrow \underset{+5}{}$

(Increase per atom = 6)

Decrease in oxidation number: $\underset{+7}{MnO_4^-} \longrightarrow \underset{+4}{MnO_2}$

(Decrease per atom = 3)

To make the net increase and decrease equal, we must take 2 atoms of Mn.

$2MnO_4^-{}_{(aq)} + Br^-{}_{(aq)} \longrightarrow 2MnO_{2(s)} + BrO_3^-{}_{(aq)}$

Step 3: Balance 'O' atoms by adding H_2O to the right-hand side.

$2MnO_4^-{}_{(aq)} + Br^-{}_{(2q)} \longrightarrow 2MnO_{2(s)} + BrO_3^-{}_{(aq)} + H_2O_{(l)}$

Step 4: The medium is basic. To make the charges and hydrogen atoms on the two sides equal, add $2H^+$ on the left-hand side.

$$2MnO_{4(aq)}^- + Br_{(aq)}^- + 2H_{(aq)}^+ \longrightarrow 2MnO_{2(s)} + BrO_{3(aq)}^- + H_2O_{(l)}$$

Add OH^- ions equal to the number of H^+ ions on both sides of the equation.

$$2MnO_{4(aq)}^- + Br_{(aq)}^- + 2H_{(aq)}^+ + 2OH_{(aq)}^- \longrightarrow 2MnO_{2(s)} + BrO_{3(aq)}^- + H_2O_{(l)} + 2OH_{(aq)}^-$$

The H^+ and OH^- ions appearing on the same side of the reaction are combined to give H_2O molecules.

$$2MnO_{4(aq)}^- + Br_{(aq)}^- + 2H_2O_{(l)} \longrightarrow 2MnO_{2(s)} + BrO_{3(aq)}^- + H_2O_{(l)} + 2OH_{(aq)}^-$$

$$2MnO_{4(aq)}^- + Br_{(aq)}^- + H_2O_{(l)} \longrightarrow 2MnO_{2(s)} + BrO_{3(aq)}^- + 2OH_{(aq)}^-$$

Step 5: Check two sides for balance of atoms and charges.

Hence, balanced equation: $\mathbf{2MnO_{4(aq)}^- + Br_{(aq)}^- + H_2O_{(l)} \longrightarrow 2MnO_{2(s)} + BrO_{3(aq)}^- + 2OH_{(aq)}^-}$

iii. $H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_{2(g)} + SO_{2(g)} + H_2O_{(l)}$ (acidic)

Step 1: Write skeletal equation and balance the elements other than O and H.

$H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_{2(g)} + SO_{2(g)} + H_2O_{(l)}$

Step 2: Assign oxidation number to S and C. Calculate the increase and decrease in the oxidation number and make them equal.

$$\underset{+1\ +6\ -2}{H_2SO_{4(aq)}} + \underset{0}{C_{(s)}} \longrightarrow \underset{+4}{CO_{2(g)}} + \underset{+4}{SO_{2(g)}} + H_2O_{(l)}$$

Increase in oxidation number: $\underset{0}{C} \longrightarrow \underset{+4}{CO_2}$

(Increase per atom = 4)

Decrease in oxidation number: $\underset{+6}{H_2SO_4} \longrightarrow \underset{+4}{SO_2}$

(Decrease per atom = 2)

To make the net increase and decrease equal, we must take 2 atoms of S.

$2H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_{2(g)} + 2SO_{2(g)} + H_2O_{(l)}$

Step 3: Balance 'O' atoms by adding H2O to the right-hand side.

$2H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_{2(g)} + 2SO_{2(g)} + H_2O_{(l)} + H_2O_{(l)}$

Step 4: The medium is acidic. There is no charge on either side. Hydrogen atoms are equal on both side.

$2H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_2 + 2SO_{2(g)} + H_2O_{(l)}$

Step 5: Check two sides for balance of atoms and charges.

Hence, balanced equation: $2H_2SO_{4(aq)} + C_{(s)} \rightarrow CO_{2(g)} + 2SO_{2(g)} + H_2O_{(l)}$

iv. $Bi(OH)_{3(s)} + Sn(OH)^-_{3(aq)} — Bi_{(s)} + Sn(OH)^{2-}_{6(aq)}$ (basic)

Step 1: Write skeletal equation and balance the elements other than O and H.

$Bi(OH)_{3(s)} + Sn(OH)^-_{3(aq)} — Bi_{(s)} + Sn(OH)^{2-}_{6(aq)}$

Step 2: Assign oxidation numbers to Bi and Sn. Calculate the increase and decrease in the oxidation number and make them equal.

$$\underset{+3\ -2\ +1}{Bi(OH)_{3(s)}} + \underset{+2\ -2\ +1}{Sn(OH)^-_{3(aq)}} \longrightarrow \underset{0}{Bi_{(s)}} + \underset{+4}{Sn(OH)^{2-}_{6(aq)}}$$

Increase in oxidation number: $\underset{+2}{Sn(OH)_3^-} \longrightarrow \underset{+4}{Sn(OH)_6^{2-}}$

(Increase per atom = 2)

Decrease in oxidation number: $\underset{+3}{Bi(OH)_3} \longrightarrow \underset{0}{Bi}$

(Decrease per atom = 3)

To make the net increase and decrease equal, we must take 3 atoms of Sn and 2 atoms of Bi.

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)}$$

Step 3: Balance 'O' atoms by adding $3H_2O$ to the left-hand side.

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3H_2O_{(l)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)}$$

Step 4: The medium is basic. To make hydrogen atoms on the two sides equal, add 3W on the right-hand side.

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3H_2O_{(l)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)} + 3H^+_{(aq)}$$

Add OH^- ions equal to the number of H^+ ions on both sides of the equation.

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3H_2O_{(l)} + 3OH^-_{(aq)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)} + 3H^+_{(aq)} + 3OH^-_{(aq)}$$

The H^+ and OH^- ions appearing on the same side of the reaction are combined to give H_2O molecules.

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3H_2O_{(l)} + 3OH^-_{(aq)} \longrightarrow 4Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)} + 3H_2O_{(l)}$$

$$2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3OH^-_{(aq)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)}$$

Step 5: Check two sides for balance of atoms and charges.

Hence, balanced equation: $\mathbf{2Bi(OH)_{3(s)} + 3Sn(OH)^-_{3(aq)} + 3OH^-_{(aq)} \longrightarrow 2Bi_{(s)} + 3Sn(OH)^{2-}_{6(aq)}}$

Question B.

Balance the following redox equation by half reaction method

a. $H_2C_2O_4\ (aq) + MnO_4^{\ominus}\ (aq) \longrightarrow CO_2(g) + Mn^{2\oplus}\ (aq)$ (acidic)

b. $Bi\ (OH)_3\ (s) + SnO_2^{2\ominus}\ (aq) \rightarrow SnO_3^{2\ominus}\ (aq) + Bi\ (s)$ (basic)

Answer:

i. $H_2C_2O_{4(aq)} + MnO^-_{4(aq)} \rightarrow CO_{2(g)} + Mn^{2+}_{(aq)}$

Step 1: Write unbalanced equation for the redox reaction. Assign oxidation number to all the atoms in reactants and products. Divide the equation into two half equations.

$$\underset{+1\ +3\ -2}{H_2C_2O_{4(aq)}} + \underset{+7\ -2}{MnO^-_{4(aq)}} \longrightarrow \underset{+4\ -2}{CO_{2(g)}} + \underset{+2}{Mn^{2+}_{(aq)}}$$

Gain of e^-

Loss of e^-

Oxidation half reaction: $H_2C_2O_{4(aq)} \longrightarrow CO_{2(g)}$

Reduction half reaction: $MnO^-_{4(aq)} \longrightarrow Mn^{2+}_{(aq)}$

Step 2: Balance the atoms except O and H in each half equation. Balance half equation for O atoms by adding $4H_2O$ to the right side of reduction half equation.

Oxidation: $H_2C_2O_{4(aq)} \longrightarrow 2CO_{2(g)}$

Reduction: $MnO^-_{4(aq)} \longrightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

Step 3: Balance H atoms by adding H^+ ions to the side with less H. Hence, add $2H^+$ ions to the right side of oxidation half equation and $8H^+$ ions to the left side of reduction half equation.

Oxidation: $H_2C_2O_{4(aq)} \longrightarrow 2CO_{2(g)} + 2H^+_{(aq)}$

Reduction: $MnO^-_{4(aq)} + 8H^+_{(aq)} \longrightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

Step 4: Now add 2 electrons to the right side of oxidation half equation and 5 electrons to the left side of reduction half equation to balance the charges.

Oxidation: $H_2C_2O_{4(aq)} \longrightarrow 2CO_{2(g)} + 2H^+_{(aq)} + 2e^-$

Reduction: $MnO^-_{4(aq)} + 8H^+_{(aq)} + 5e^- \longrightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$

Step 5: Multiply oxidation half equation by 5 and reduction half equation by 2 to equalize number of electrons in two half equations. Then add two half equation.

Oxidation: $5H_2C_2O_{4(aq)} \longrightarrow 10CO_{2(g)} + 10H^+_{(aq)} + 10e^-$

Reduction: $2MnO^-_{4(aq)} + 16H^+_{(aq)} + 10e^- \longrightarrow 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$

Add two half equations:

$5H_2C_2O_{4(aq)} + 2MnO^-_{4(aq)} + 6H^+_{(aq)} \longrightarrow 10CO_2 + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}$

The equation is balanced in terms of number of atoms and the charges.

Hence, balanced equation: $\mathbf{5H_2C_2O_{4(aq)} + 2MnO^-_{4(aq)} + 6H^+_{(aq)} \longrightarrow 10CO_2 + 2Mn^{2+}_{(aq)} + 8H_2O_{(l)}}$

ii. $Bi(OH)_{3(s)} + SnO_2^{2-}{}_{(aq)} \longrightarrow SnO_3^{2-}{}_{(aq)} + Bi_{(s)}$

Step 1: Write unbalanced equation for the redox reaction. Assign oxidation number to all the atoms in reactants and products. Divide the equation into two half equations.

$$\underset{+3\ -2\ +1}{Bi(OH)_{3(s)}} + \underset{+2\ -2}{SnO^{2-}_{2(aq)}} \longrightarrow \underset{+4\ -2}{SnO^{2-}_{3(aq)}} \quad \underset{0}{}$$

Gain of e^-

Loss of e^-

Oxidation half reaction: $SnO^{2-}_{2(aq)} \longrightarrow SnO^{2-}_{3(aq)}$

Reduction half reaction: $Bi(OH)_{3(s)} \longrightarrow Bi_{(s)}$

Step 2: Balance half equations for O atoms by adding H_2O to the side with less O atoms. Add $1H_2O$ to left side of oxidation half equation and $3H_2O$ to the right side of reduction half equation.

Oxidation: $SnO^{2-}_{2(aq)} + H_2O_{(l)} \longrightarrow SnO^{2-}_{3(aq)}$

Reduction: $Bi(OH)_{3(s)} \longrightarrow Bi_{(s)} + 3H_2O_{(l)}$

Step 3: Balance H atoms by adding H^+ ions to the side with less H. Hence, add $2H^+$ ions to the right side of oxidation half equation and $3H^+$ ions to the left side of reduction half equation.

Oxidation: $SnO^{2-}_{2(aq)} + H_2O_{(l)} \longrightarrow SnO^{2-}_{3(aq)} + 2H^+_{(aq)}$

Reduction: $Bi(OH)_{3(s)} + 3H^+_{(aq)} \longrightarrow Bi_{(s)} + 3H_2O_{(l)}$

Step 4: Now add 2 electrons to the right side of oxidation half equation and 3 electrons to the left side of reduction half equation to balance the charges.

Oxidation: $SnO^{2-}_{2(aq)} + H_2O_{(l)} \longrightarrow SnO^{2-}_{3(aq)} + 2H^+_{(aq)} + 2e^-$

Reduction: $Bi(OH)_{3(s)} + 3H^+_{(aq)} + 3e^- \longrightarrow Bi_{(s)} + 3H_2O_{(l)}$

Step 5: Multiply oxidation half equation by 3 reduction half equation by 2 to equalize number of electrons in two half equations. Then add two half equation.

Oxidation: $3SnO^{2-}_{2(aq)} + 3H_2O_{(l)} \longrightarrow 3SnO^{2-}_{3(aq)} + 6H^+_{(aq)} + 6e^-$

Reduction: $2Bi(OH)_{3(s)} + 6H^+_{(aq)} + 6e^- \longrightarrow 2Bi_{(s)} + 6H_2O_{(l)}$

Add two half equations:

$2Bi(OH)_{3(s)} + 3SnO^{2-}_{2(aq)} \longrightarrow 3SnO^{2-}_{3(aq)} + 2Bi_{(s)} + 3H_2O_{(l)}$

Reaction occurs in basic medium. However, H^+ ions cancel out and the reaction is balanced. Hence, no need to add OH^- ions. The equation is balanced in terms of number of atoms and the charges.

Hence, balanced equation:

$\mathbf{2Bi(OH)_{3(s)} + 3SnO^{2-}_{2(aq)} \longrightarrow 3SnO^{2-}_{3(aq)} + 2Bi_{(s)} + 3H_2O_{(l)}}$

5. Complete the following table :

Assign oxidation number to the underlined species and write Stock notation of compound

Compound	Oxidation number	Stock notation
$AuCl_3$	…………….	…………….
$SnCl_2$	…………….	…………….
$V--_{2}O_{4-7}$	…………….	…………….
$Pt---Cl_{2-6}$	…………….	…………….
H_3AsO_3	…………….	…………….

Answer:

Compound	Oxidation number	Stock notation
$AuCl_3$	+3	$Au(III)Cl_3$
$SnCl_2$	+2	$Sn(II)Cl_2$
$V--_{2}O_{4-7}$	+5	$V_2(V)O_{4-7}$
$Pt---Cl_{2-6}$	+4	$Pt(IV)Cl_{2-6}$
H_3AsO_3	+3	$H_3As(III)O_3$

11th Chemistry Digest Chapter 6 Redox Reactions Intext Questions and Answers

Can you tell? (Textbook Page No. 81)

Question i.
Why does cut apple turn brown when exposed to air?
Answer:
Cut apple turns brown when exposed to air because polyphenols are released. These polyphenols undergo oxidation in the presence of air and impart brown colour.

Question ii.
Why does old car bumper change colour?
Answer:
Car bumper is made of iron which undergoes rusting over a period of time. Hence, old car bumper changes colour.

Question iii.
Why do new batteries become useless after some days?
Answer:
Batteries generate electricity by redox reactions. Once the chemicals taking part in redox reaction are used up, the battery cannot generate power. Hence, new batteries become useless after some days.

Can you recall? (Textbook Page No. 81)

Question i.
What is combustion reaction?
Answer:
Combustion is a process in which a substance combines with oxygen.

Question ii.
Write an equation for combustion of methane.
Answer:
Combustion of methane: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ + Heat + Light

Question iii.
What is the driving force behind reactions of elements?
Answer:
The ability of element to combine with other element or the ability of element to replace other element in compound is the driving force behind the reactions. This may involve formation of precipitates, formation of water, release of gas, etc.

Try this. (Textbook Page No. 82)

Question 1.
Complete the following table of displacement reactions. Identify oxidising and reducing agents involved.

Reactants	Products
$Zn_{(s)}$ + ————$_{(aq)}$	————$_{-(aq)}$ + $Cu_{(s)}$
$Cu_{(s)}$ + $2Ag^{+}_{(aq)}$	———— + ————
———— + ————-	$Co^{2+}_{(aq)}$ + $Ni_{(s)}$

Answer:

Reactants	Products	Oxidising agent	Reducing agent
$Zn_{(s)} + Cu^{2+}_{(aq)}$	$Zn^{2+}_{(aq)} + Cu_{(s)}$	Cu^{2+}	Zn
$Cu_{(s)} + 2Ag^{+}_{(aq)}$	$Cu^{2+}_{(aq)} + 2Ag_{(s)}$	Ag^{+}	Cu
$Co_{(s)} + Ni^{2+}_{(aq)}$	$Co^{2+}_{(aq)} + Ni_{(s)}$	Ni^{2+}	Co

Try this (Textbook Page No. 88)

Question 1.
Classify the following unbalanced half equations as oxidation and reduction.

Example	Type
$Cl^{-}_{(aq)} \longrightarrow Cl_{2(g)}$	Oxidation
$OCl^{-}_{(aq)} \longrightarrow Cl^{-}_{(g)}$	----------
$Fe(OH)_2 \longrightarrow Fe(OH)_3$	----------
$VO^{2+}_{(aq)} \longrightarrow V^{3+}_{(aq)}$	

Answer:

Example	Type
$Cl^{-}_{(aq)} \longrightarrow Cl_{2(g)}$	Oxidation
$OCl^{-}_{(aq)} \longrightarrow Cl^{-}_{(g)}$	Reduction
$Fe(OH)_2 \longrightarrow Fe(OH)_3$	Oxidation
$VO^{2+}_{(aq)} \longrightarrow V^{3+}_{(aq)}$	Reduction