Redox Reactions

TRY YOURSELF ANSWERS

1. The term "oxidation" is defined as the addition of oxygen/electronegative element to a substance or removal of hydrogen/electropositive element from a substance.

Example : $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$ (Oxidation of CH_4)

$2H_2S_{(g)} + O_{2(g)} \longrightarrow 2S_{(s)} + 2H_2O_{(l)}$ (Oxidation of H_2S)

2. The term "reduction" is defined as the removal of oxygen/electronegative element from a substance or addition of hydrogen/electropositive element to a substance.

Example : $2HgO_{(s)} \xrightarrow{\Delta} 2Hg_{(l)} + O_{2(g)}$ (Reduction of HgO)

$CH_2{=}CH_{2(g)} + H_{2(g)} \rightarrow H_3C{-}CH_{3(g)}$ (Reduction of ethylene)

3. $BaSO_4$ is reduced to BaS because of the removal of oxygen from $BaSO_4$.

4. A substance which readily accept electrons to oxidise other substance, as a result itself gets reduced is called oxidising agent or an oxidant.

Example : $\underset{\text{Oxidant}}{\overset{+7}{Mn}O_4^-} + SO_3^{2-} \longrightarrow Mn^{2+} + SO_4^{2-}$

In the above reaction, the Mn reduce from +7 (in MnO_4^-) to +2 (in Mn^{2+}). Therefore, MnO_4^- is the oxidising agent in this reaction.

5. A substance which readily lose electrons to reduce other substance, as a result itself gets oxidised is called reducing agent or reductant.

Example : $Cl_{2(aq)} + 2Br^-_{(aq)} \longrightarrow 2Cl^-_{(aq)} + Br_{2(aq)}$

Br^- is reducing agent as in the above reaction oxidation state of Br changes from –1 (in Br^-) to 0 (in Br_2). This means that Br^- oxidises to Br_2.

6. $\underset{\text{(Oxidising agent)}}{Br_2} + 2e^- \longrightarrow 2Br^-$ (Reduction half reaction)

$\underset{\text{(Reducing agent)}}{2Na} \longrightarrow 2Na^+ + 2e^-$ (Oxidation half reaction)

7. (c) : (a) $\overset{+2}{C}O \rightarrow \overset{+4}{C}O_2$ (b) $\overset{+2}{Cu}O \rightarrow \overset{+2}{Cu}Cl_2$

(c) $\overset{+1}{H_2}O \rightarrow \overset{0}{H_2}$ (d) $\overset{0}{C} \rightarrow \overset{+4}{C}O_2$

Only in (c), O.N. of hydrogen decreases from + 1 to 0 and hence H_2O gets reduced to H_2.

8. The decreasing order of electron-releasing tendency or reactivity is Zn > Cu > Ag. Such order is found from experiments performed by comparing the reactivity of metals and their aqueous solutions. It is found out that Zinc (Zn) reduces both Cu^{2+} to copper (Cu) and $Ag^+_{(aq)}$ to silver $(Ag)_{(s)}$. While copper only reduces $Ag^+_{(aq)}$ to silver, $Ag_{(s)}$.

9. (i) $KMnO_4$: $+1 + x - 8 = 0$; $x = +7$; Ox. no. of Mn = +7

K_2MnO_4: $+2 + x - 8 = 0$; $x = +6$; Ox. no. of Mn = +6

MnO_2: $x - 4 = 0$; $x = +4$; Ox. no. of Mn = +4

Mn_2O_3: $2x - 6 = 0$; $x = +3$; Ox. no. of Mn = +3

The highest oxidation number for Mn is in $KMnO_4$ (+7).

(ii) $K_3[Fe(OH)_6]$: $+3 + x - 6 = 0$; $x = +3$ Ox. no. of Fe = +3

$K_2[FeO_4]$: $+2 + x - 8 = 0$; $x = +6$ Ox. no. of Fe = +6

$FeSO_4{\cdot}(NH_4)_2SO_4{\cdot}6H_2O$: In $FeSO_4$ Ox. no. of Fe = +2

$[Fe(CN)_6]^{3-}$: $x - 6 = -3$; $x = +3$

The least oxidation number for Fe is in mohr salt $FeSO_4{\cdot}(NH_4)_2SO_4{\cdot}6H_2O$ *i.e.*, +2.

10. (i) Oxidation number of Cr in $Cr_2O_7^{2-} = +6$

Oxidation number of Cr in $CrO_4^{2-} = +6$

Since, the oxidation number of chromium neither increased nor decreased, therefore, this reaction cannot be regarded as a redox reaction. The oxidation number of other elements, *i.e.*, O and H also remains unchanged.

(ii) $2\overset{+1}{Cu_2}\overset{-2}{O}_{(s)} + \overset{+1}{Cu_2}\overset{-2}{S}_{(s)} \longrightarrow 6\overset{0}{Cu}_{(s)} + \overset{+4}{S}\overset{-2}{O}_{2(g)}$

Oxidation state of sulphur changes from –2 in Cu_2S to +4 in SO_2 *i.e.*, it is getting oxidised and hence, Cu_2S acts as a reductant or reducing agent.

Oxidation state of copper changes from +1 in Cu_2O to zero in elemental copper *i.e.*, It is getting reduced and hence, Cu_2O acts as an oxidant or oxidising agent.

11. Out of the given, ClO_4^- does not undergo disproportionation reaction because in this oxidation state of +7 is highest for chlorine.

$3\overset{+1}{Cl}O^- \longrightarrow 2\overset{-1}{Cl^-} + \overset{+5}{Cl}O_3^-$

$$6\overset{+3}{Cl}O_2^- \longrightarrow 4\overset{+5}{Cl}O_3^- + 2\overset{-1}{Cl^-}$$

$$4\overset{+5}{Cl}O_3^- \longrightarrow \overset{-1}{Cl^-} + 3\overset{+7}{Cl}O_4^-$$

12. Writing oxidation numbers of all the atoms

$$\overset{+2\,-2}{PbS} + \overset{+1\,-1}{H_2O_2} \longrightarrow \overset{+2+6-2}{PbSO_4} + \overset{+1\,-2}{H_2O}$$

The O.N. of S is increased and that of O is decreased.

$$Pb\overset{-2}{S} \longrightarrow Pb\overset{+6}{S}O_4 \quad ...(i)\ ; \quad H_2\overset{-1}{O_2} \longrightarrow H_2\overset{-2}{O} \quad ...(ii)$$

Increase in O.N. of S = 8 units per PbS molecule

Decrease in O.N. of O = 1 unit per 1/2 H_2O_2 molecule

= 2 units per H_2O_2 molecule

Multiply eqn. (ii) by 4 as to make increase in O.N.

$$PbS + 4H_2O_2 \longrightarrow PbSO_4 + 4H_2O$$

The number of moles of water required in the balanced chemical equation is 4.

13. $\overset{0}{P}_{4(s)} + OH^-_{(aq)} \longrightarrow \overset{-3}{P}H_{3(g)} + H_2\overset{+1}{P}O^-_{2(aq)}$

Oxidation half-reaction,

$$8H_2O_{(l)} + P_{4(s)} \longrightarrow 4H_2PO^-_{2(aq)} + 8H^+_{(aq)} + 4e^- \quad ...(i)$$

Reduction half-reaction,

$$12e^- + 12H^+_{(aq)} + P_{4(s)} \longrightarrow 4PH_{3(g)} \quad ...(ii)$$

On multiplying equation (i) by 3 and adding it to equation (ii), we have

$$24H_2O_{(l)} + 3P_{4(s)} + P_{4(s)} \longrightarrow 12H_2PO^-_{2(aq)} + 12H^+_{(aq)} + 4PH_{3(g)}$$

Adding OH^- ions on both sides, we have

$$12OH^-_{(aq)} + 24H_2O_{(l)} + 4P_{4(s)} \longrightarrow 12H_2PO^-_{2(aq)} + (12H^+_{(aq)} + 12OH^-_{(aq)}) + 4PH_{3(g)}$$

$$12OH^-_{(aq)} + 12H_2O_{(l)} + 4P_{4(s)} \longrightarrow 12H_2PO^-_{2(aq)} + 4PH_{3(g)}$$

or $\quad P_{4(s)} + 3OH^-_{(aq)} + 3H_2O_{(l)} \longrightarrow PH_{3(g)} + 3H_2PO^-_{2(aq)}$

14. No. of millimoles of Mohr's salt present in 72 mL of 0.5 M solution = 72 × 0.5 = 36

The balanced chemical equation for the redox reaction is

$$K_2Cr_2O_7 + 6(NH_4)_2SO_4\cdot FeSO_4\cdot 6H_2O + 7H_2SO_4 \longrightarrow K_2SO_4 + 6(NH_4)_2SO_4 + 3Fe_2(SO_4)_3 + Cr_2(SO_4)_3 + 43H_2O$$

6 moles of Mohr's salt are oxidised by 1 mole of $K_2Cr_2O_7$

36 millimoles of Mohr's salt are oxidised by $\frac{1}{6} \times 36$

= 6 millimoles of $K_2Cr_2O_7$

15. Cl_2 is a stronger oxidising agent than I_2, therefore, when Cl_2 is passed through KI solution, Cl_2 gets reduced to colourless Cl^- ions while I^- ions gets oxidised to violet coloured iodine.

$$Cl_{2(aq)} + 2I^-_{(aq)} \longrightarrow 2Cl^-_{(aq)} + I_{2(s)}$$

The I_2 produced forms a blue coloured complex with starch which is responsible to impart blue colour to the solution.

16. Meq. of $Na_2CO_3 . x\, H_2O$ in 20 mL = $19.8 \times \frac{1}{10}$

Meq. of $Na_2CO_3 . xH_2O$ in 100 mL = $19.8 \times \frac{1}{10} \times 5$

$$\therefore \quad \frac{W}{E} \times 1000 = 19.8 \times \frac{1}{10} \times 5$$

$$\frac{0.7}{M/2} \times 1000 = \frac{19.8}{2}$$

$\therefore \quad M = 141.41$

$23 \times 2 + 12 + 3 \times 16 + 18x = 141.41, \quad \therefore \quad x = 2$

17. KCl cannot be used as an electrolyte in the salt bridge of the given cell because Cl^- ion will combine with Ag^+ ions to form white precipitate of AgCl.

18. Anodic reaction : $2H_2O_{(l)} \longrightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

Cathodic reaction : $2H_2O_{(l)} + 2e^- \longrightarrow H_{2(g)} + 2OH^-_{(aq)}$

19. $E°$ of H_2O is greater than that of Na^+ ions, reduction of H_2O takes place preferentially to evolve H_2 gas at the cathode.

20. $E°$ of Br_2, is higher than that of I_2. Therefore, Br_2 has a higher tendency to accept electrons than I_2. Conversely I^- ion has a higher tendency to lose electrons than Br^- ions.

Therefore, the following reaction will occur

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

$$Br_2 + 2e^- \longrightarrow 2Br^-$$

$$\overline{2I^- + Br_2 \longrightarrow I_2 + 2Br^-}$$

I^- ion will get oxidised to I_2 while Br_2 will be reduced to Br^- ions.