Chapter-8 : Redox Reactions

1. (a) O.N. of Mn in MnO_4^- is +7 and in Mn^{2+} it is +2. The difference is of 5 electrons.
2. (a) In a redox reaction, one molecule is oxidised and other molecule is reduced i.e. oxidation number of reactants are changed.

$$\overset{0}{H_2} + \overset{0}{Br_2} \longrightarrow 2\,\overset{+1-1}{HBr}$$

Here, H_2 is oxidised and Br_2 is reduced, thus it is oxidation-reduction reaction.
3. (b) In the given reaction oxidation state of Mg is changing from 0 to +2 while in nitrogen it is changing from 0 to – 3. So, oxidation of Mg and reduction of nitrogen takes place.
4. (a) In this reaction oxidation occurs.
5. (a) O.N. of P in H_3PO_3 (phosphorous acid)
$3 \times (+1) + x + 3 \times (-2) = 0$ or $x = +3$
In orthophosphoric acid (H_3PO_4) O.N. of P is + 5, in hypophosphorous acid (H_3PO_2) it is + 1 while in metaphosphoric acid (HPO_3), it is + 5,
6. (b) Sum of oxidation state of all atoms in neutral compound is zero.
Let the oxidation state of iron in the complex ion
$[Fe(H_2O)_5(NO)]^{2+}SO_4^{2-}$ be x; then $x + 5 \times 0 + 0 = +2$.
$\therefore\ x = +2$
7. (a) (i) Oxidation state of an element in its free state is zero.
(ii) Sum of oxidation states of all atoms in compound is zero.
O.N. of S in $S_8 = 0$; O.N. of S in $S_2F_2 = +1$;
O.N. of S in $H_2S = -2$
8. (d) $(NH_4)_2SO_4$ split into ions, NH_4^+ and SO_4^{2-}. Let O.N. of N in NH_4^+ be x then, $1 \times (x) + 4 \times (+1) = 1$
$\therefore\ x = -3$
9. (b) $Na_2S_4O_6$
Let O.N. of S be x then $2 \times (+1) + 4 \times (x) + 6 \times (-2) = 0$
$\therefore\ x = 2.5.$
10. (c) In metal carbonyls metal always has O.N. zero.
11. (a) $O{=}\overset{+2}{C}{=}\overset{0}{C}{=}\overset{+2}{C}{=}O$
In C_3O_2, two C atoms linked with oxygen atoms are present in +2 oxidation state and central carbon has zero oxidation state. So, the average oxidation state of C is + 4/3.
12. (d) In disproportionation reaction, one element of a compound will simultaneously get reduced and oxidised. In ClO_4^-, oxidation number of Cl is + 7 and it can not increase further. So, ClO_4^- will not get oxidised and so will not undergo disporportionation reaction.
13. (d) In Na_2O, $SnCl_2$ and Na_2O_2 central atom is either in lowest or highest oxidation state, so it can function either as an oxidising or a reducing agent but not both. However, the oxidation state of N in $NaNO_2$ is +3 which lies between its highest (+5) and lowest (–3) values.
14. (a) More is E°_{RP}, more is the tendency to get itself reduced or more is oxidising power.
15. (a) Fluorine has highest E° value and more reactive than MnO_2.
16. (a) Zinc rod dipped in blue copper sulphate solution is oxidised to Zn^{2+} and Cu^{2+} are reduced to Cu and get deposited on zinc rod.
17. (b) Oxygen has oxidation number – 1 in H_2O_2 and +2 in OF_2. Fluorine has oxidation number – 1 in all its compounds. Hydrogen shows + 1 in most of its compounds while – 1 in binary metallic hydrides. Sulphur shows – 2 in all sulphides.
18. (b) Oxidation number of a compound must be 0. Using the values for A, B and C in the four options, we find that $A_3(BC_4)_2$ is the answer.
Check : $(+2)3 + [(+5)+4(-2)]2 = 6 + (5-8)2 = 0$
19. (c) $CaOCl_2$ or $Ca(OCl)Cl$ is the mixed salt of $Ca(OH)_2$ with HCl and HOCl.
20. (a) In acidic medium MnO_4^- changes to Mn^{2+}, hence O.N. changes from +7 to +2.
21. (d) O.N. of carbon in CH_3CHO is –1; in other cases it is zero.
22. (c) $K_2Cr_2O_7 + 3SO_2 + 4H_2SO_4 \rightarrow$
$$K_2SO_4 + Cr_2(SO_4)_3 + 3SO_3 + 4H_2O$$
O.N. of chromium changes from +6 to +3.
23. (b) O.N. of iodine in I_3^- is –1/3.

24. (b) $YBa_2Cu_3O_7$
$3 + 2 \times (+2) + 3x + (-2) \times 7 = 0$
$3 + 4 + 3x - 14 = 0$
$3x = 7 \Rightarrow x = \frac{7}{3}.$

25. (c) It has four O atoms as peroxide with oxidation number –1 and one O atom with oxidation number –2. Hence, $x + 4(-1) + 1(-2) = 0$ or $x = +6$

26. (d) $3Br_2 + 6CO_3^{2-} + 3H_2O \longrightarrow 5\overset{-1}{Br} + \overset{+5}{Br}O_3^- + 6HCO_3^-$
Bromine is getting oxidised as well as reduced in this reaction.

27. (a) $2MnO_4^- + 5H_2O_2 + 6H^+ \rightarrow 2Mn^{2+} + 5O_2 + 8H_2O.$

28. (d) $8KMnO_4 + 3NH_3 \longrightarrow 8MnO_2 + 3KNO_3 + 5KOH + 2H_2O$

29. (b) $$2\overset{\text{Reduction}}{\overbrace{Fe^{3+} + Sn^{2+} \rightarrow 2Fe^{2+}}} + Sn^{4+}$$
$$2Fe^{3+} + \underbrace{Sn^{2+} \rightarrow 2Fe^{2+} + Sn^{4+}}_{\text{Oxidation}}$$

30. (c) The balanced reaction is
$H_2SO_4 + 8HI \rightarrow H_2S + 4I_2 + 4H_2O$
Hence the value of x, y, z are 8, 4, 4 respectively.

31. (b) The compound which undergo oxidation itself and reduces others is known as reducing agent. In this reaction O.N. of Ni changes from 0 to + 2 and hence Ni acts as a reducing agent.

32. (a) (i) $M^{n+} + ne^- \rightleftharpoons M$, for this reaction, high negative value of $E°$ indicates lower reduction potential, that means M will be a good reducing agent.

$$\left[\begin{array}{c}\text{Stronger reducing agent} \Rightarrow \text{Easy to oxidise} \\ \Downarrow \\ \text{Lower reduction potential} \Leftarrow \text{higher oxidation potential}\end{array}\right]$$

(ii)

Element	F	Cl	Br	I
Reduction potential ($E°$ volt)	+2.87	+1.36	+1.06	+0.54

As reduction potential decreases from fluorine to iodine, oxidising nature also decreases from fluorine to iodine.

(iii) The size of halide ions increases from F^- to I^-. The bigger ion can loose electron easily. Hence, the reducing nature increases from HF to HI.

33. (a) More negative or lower is the reduction potential, more is the reducing property. Thus the reducing power of the corresponding metal will follow the reverse order, i.e. Y > Z > X.

34. (c) Since X displaces Ni from $NiSO_4$ solution, it means X has higher oxidation potential or lower reduction potential than Ni^{2+}. Since X cannot displace Mn from $MnSO_4$ solution, it means it has higher reduction potential than Mn^{2+}. Lower the reduction potential, stronger is the reducing agent hence reducing power is Mn > X > Ni.

35. (d) All the statements are correct.

36. (c) Zinc can reduce Ag^+ and Cu^{2+} due to lowest reduction potential.

37. (c) The redox couple with maximum reduction potential will be best oxidising agent and with minimum reduction potential will be best reducing agent.

38. (d) Zinc gives H_2 gas with dil H_2SO_4 and HCl but not with HNO_3 because in HNO_3, NO_3^- ion is reduced and give NH_4NO_3, N_2O, NO and NO_2 (based upon the concentration of HNO_3)
$4Zn + 10HNO_3 \longrightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$

39. (b) $2KMnO_4 + H_2O + KI \rightarrow 2MnO_2 + 2KOH + KIO_3$
One mole of KI reduce 2 moles of $KMnO_4$.

40. (c) Let oxidation state of oxygen in $OF_2 = x$
$\therefore \quad x + (-1 \times 2) = 0$
$\therefore \quad x = +2$

41. (d) $Zn \rightarrow Zn^{2+} + 2e^-$ ....(i)
$8e^- + 10H^+ + NO_3^- \rightarrow NH_4^+ + 3H_2O$ ....(ii)
operate eq. (i) × 4 + eq. (ii) × 1
$4Zn + 10H^+ + NO_3^- \rightarrow 4Zn^{2+} + NH_4^+ + 3H_2O$

42. (c) $2\overset{-1}{HI} + H_2\overset{+6}{S}O_4 \longrightarrow \overset{0}{I_2} + \overset{+4}{S}O_2 + 2H_2O$ in this reaction oxidation number of S is decreasing from + 6 to +4 hence undergoing reduction and for HI oxidation number of I is increasing from –1 to 0 hence undergoing oxidation, therefore H_2SO_4 is acting as oxidising agent.

43. (c) $\overset{-4}{N_2}\overset{+4}{H_4} \xrightarrow{\text{loss of } 10e^-} Y;$
Since, O.N. of 'H' remains same. So, in new compound Y, (all nitrogen retained) N will observe loss of $10e^-$. Hence, O.N. of N changes from -2 to +3.

44. (b) $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$
Oxygen is the most electronegative element in the reaction and has the oxidation states of –1 (in H_2O_2) and –2 (in $BaSO_4$). In H_2O_2, peroxo ion is present.

45. (d) Since a metal with lower electrode potential is a stronger reducing agent, Mg can displace all the given metals, Al can displace all metals except Mg. Zn can displace all metals except Mg and Al. Fe can displace only Cu. The order in which they can displace each other from their salt solutions is Mg, Al, Zn, Fe, Cu.

46. (b) $KMnO_4 + FeC_2O_4 \longrightarrow Fe^{3+} + 2CO_2 + Mn^{2+}$
So half reaction,
$\overset{7+}{Mn}O_4^- \longrightarrow Mn^{2+}$ (decrease in O.N. = 5)
$\overset{2+}{Fe}C_2O_4 \longrightarrow Fe^{3+} + 2CO_2$ (increase in O.N. = 1)
Now equating the change in O.N. and then by adding both half reactions we get
$5FeC_2O_4 + MnO_4^- \longrightarrow 5Fe^{3+} + 10CO_2 + Mn^{2+}$
On balancing equation,

$$3MnO_4^- + 5FeC_2O_4 + 24H^+ \longrightarrow 5Fe^{3+} + 3Mn^{2+} + 10CO_2 + 12H_2O$$

$\therefore$ 3 moles of $KMnO_4$ = 5 moles of FeC_2O_4

1 mole $FeC_2O_4 = \frac{3}{5}$ moles of $KMnO_4$

47. (d) Fe^{2+} is converting into Fe^{3+} and sulphur is changing from –1 oxidation state to +4 oxidation state. There are two S^- and one Fe^{2+} in FeS_2. Thus total no. of electrons lost in the given reaction are 11.

48. (c) $\overset{2+}{Mn}O + NaOH + KNO_3 \longrightarrow K_2\overset{6+}{Mn}O_4 + H_2O$

49. (b)

$\overset{+5}{HClO_3} \rightarrow \overset{+7}{HClO_4}$

$\overset{-3}{NH_4^+} \rightarrow \overset{-3}{NH_3}$

$\overset{+4}{NO_2} \rightarrow \overset{+4}{N_2O_4}$

$\overset{+4}{HSO_3^-} \rightarrow \overset{+6}{SO_4^{2-}}$

$\overset{-1}{H_2O_2} \rightarrow \overset{-2}{H_2O}$

50. (a) $Na_2[\underline{Fe}(CN)_5NO] : +2 + x + 5(-1) + 0 = 0$

$\Rightarrow x = +3$

$K_2\underline{Ta}F_7 : +2 + x + (-7) = 0 \Rightarrow x = +5$

$Mg_2\underline{P_2}O_7 : +4 + 2x + (-14) = 0 \Rightarrow x = +5$

$Na_2\underline{S_4}O_6 : +2 + 4x + (-12) = 0$ or $4x = 10 \Rightarrow x = +2.5$

$\underline{N_3}H : 3x + 1 = 0 \Rightarrow x = -\frac{1}{3}$