Speed Test-36

1. **(a)** $[Fe(H_2O)_5NO]SO_4$
Let O.N. of Fe be x then,
$1\times(x)+5\times(0)+1\times(+1)+1\times(-2)=0 \quad \therefore x=+1$

2. **(c)** In $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$ all atoms are present in the same O.S. in reactants and products.

3. **(a)** $Cr_2O_7^{2-} + 6I^- + 14H^+ \longrightarrow 2Cr^{3+} + 3I_2 + 7H_2O$
+ 3 oxidation state of Cr.

4. **(d)** O.N. of N in NO_2 and N_2O_4 is +4
$\therefore$ difference is zero.
O.N. of P in P_2O_5 and P_4O_{10} is +5
$\therefore$ difference is zero
O.N. of N in N_2O is +1 and in NO is +2. The difference is 1
O.N. of S in SO_2 is +4 and in SO_3 is +6. The difference is +2.

5. **(d)** Xe = 53.5 % $\therefore$ F = 46.5%
Relative number of atoms Xe
$$=\frac{53.5}{131.2}=0.4 \text{ and } F=\frac{46.5}{19}=2.4$$
Simple ratio Xe = 1 and F = 6
$\therefore$ Molecular formula is XeF_6
O.N. of Xe is +6.

6. **(d)** The element is Ti (At. no. 22). Electronic configuration is $1s^2, 2s^2p^6, 3s^2p^6d^2, 4s^2$, the energy level of 3d and 4s is very close. It can exhibit +4 oxidation state (Ti^{4+}).

7. **(d)** In Ag_2O, O.N. of Ag is +1 and in Ag the O.N. is 0. There is gain of electrons, hence H_2O_2 act as a reducing agent.

8. **(d)** In redox reaction oxidation and reduction take place simultaneoushy. $Cu_2S + 2FeO \rightarrow 2Cu + 2Fe + SO_2$.
O.N. of Cu changes from +1 to 0 (reduction) and O.N. of S changes from –2 to +4 (oxidation).

9. **(d)** Order of decreasing electrode potentials of Mg, K, Ba and Ca is
Mg > Ca > Ba > K
It can be explained by their standard reduction potentials.
$$E^\circ_{K^+|K} = -2.925$$
$$E_{Ba^{2+}|Ba} = -2.90$$
$$E^\circ_{Ca^{2+}|Ca} = -2.87$$
$$E^\circ_{Mg^{2+}|Mg} = -2.37$$
Highly negative value of E°_{red} shows the least value of electrode potential.

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

Stronger reducing agent ⇒ Easy to oxidise
⇓
Lower reduction potential ⇐ higher oxidation potential

(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.

11. **(d)** Negative E° ⇒ Stronger reducing agent or weaker oxidising agent
Positive E° ⇒ Weaker reducing agent or stronger oxidising agent.

12. **(a)** Both Fe(II) and S(IV) in SO_3^{2-} can be oxidised to Fe (III) and $(SO_4)^{2-}$ respectively hence (3/5)×0.5 = 0.3 mol/litre.
$$\left[MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O\right]\times\frac{3}{5}$$
$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$
$$SO_3^{2-} \longrightarrow SO_4^{2-} + 2e^-$$
$$\frac{3}{5}MnO_4^- + \frac{24}{5}H^+ + Fe^{2+} + SO_3^{2-} \longrightarrow 3/5\,Mn^{2+} + 12/5\,H_2O + Fe^{3+} + SO_4^{2-}$$

13. **(d)** Reduction potential of Cu(II) is greater than that of Zn(II) and Al(III) thus can be easily replaced by these ions. Moreover solution of copper is blue in color.

14. **(b)** $3SO_2 + O_3 \rightarrow 3SO_3$
O.N. of S changes from +4 to +6. Two electron change
$\therefore$ Eq. Wt = M/2. (molecular wt. = M)

15. **(a)** $$2K\overset{+7}{Mn}O_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2\overset{+2}{Mn}SO_4 + 3H_2O + 5O$$
$$2\overset{+2}{Fe}SO_4 + H_2SO_4 + O \rightarrow \overset{+3}{Fe_2}(SO_4)_3 + H_2O$$

O.N. of Mn changes from +7 to +2 (Reduction)

O.N. of Fe changes from +2 to +3 (Oxidation)

16. **(c)** On balancing the given equations, we get

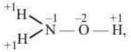
$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \longrightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

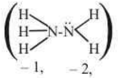
So, $x = 2, y = 5$ & $z = 16$

17. **(b)** On reaction with hot and concentrated alkali a mixture of chloride and chlorate is formed

$$3Cl_2 + 6\,NaOH(\text{excess}) \xrightarrow{\text{Hot}} 5Na\overset{-1}{Cl} + Na\overset{+5}{Cl}O_3 + 3H_2O$$

18. **(a)** $\underset{-3}{[Co(NH_3)_5Cl.]Cl_2}$,





, $\underset{-3}{Mg_3N_2}$

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

20. **(a)** $3MnO_4^{2-} + 2H_2O \rightarrow MnO_2 + 2MnO_4^- + 4OH^-$

or $MnO_4^{2-} + \frac{2}{3}H_2O \rightarrow \frac{1}{3}MnO_2 + \frac{2}{3}MnO_4^- + \frac{4}{3}OH^-$

21. **(b)** In iodometry, $K_2Cr_2O_7$ liberates I_2 from iodides (NaI or KI) which is titrated with $Na_2S_2O_3$ solution.

$$K_2Cr_2O_7 + I^- + H^+ \longrightarrow Cr^{3+} + I_2$$

Here, one mole of $K_2Cr_2O_7$ accepts 6 mole of electrons.

$$\therefore \text{Equivalent weight} = \frac{\text{molecular weight}}{6}$$

22. **(d)** $\overset{0}{Cl_2} + 2NaOH \rightarrow Na\overset{-1}{Cl} + Na\overset{+1}{Cl}O + H_2O$

$$3\overset{+6}{Mn}O_4^{2-} + 2H_2O \rightarrow 2\overset{+7}{Mn}O_4^- + \overset{+4}{Mn}O_2 + 4OH^-$$

$$2\overset{+4}{N}O_2 + H_2O \rightarrow H\overset{+5}{N}O_3 + H\overset{+3}{N}O_2$$

All undergo disproportionation.

23. **(c)** $\overset{-4\ +4}{N_2H_4} \xrightarrow{\text{loss of }10e^-} N_2^{+6}(y)$;

O.N. of N changes from -2 to +3

24. **(b)** Phosphorus, sulphur and chlorine disproportionate in the alkaline medium.

25. **(c)** Oxidation number of oxygen in $OF_2 = +2$ and in $KO_2 = \frac{-1}{2}$.

26. **(d)** In H_2SO_4, sulphur is in highest oxidation state (+6). Hence H_2SO_4 will be strongest oxidising agent.

27. **(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

28. **(a)** Higher the value of reduction potential higher will be the oxidising power whereas the lower the value of reduction potential higher will be the reducing power.

29. **(a)** More the reduction potential, more will be the oxidising power.

30. **(c)** The violent reaction between sodium and water is an example of redox reaction :

$$\text{(Oxd. state)}\quad \underset{0}{2Na} + \underset{+1}{2H_2O} \longrightarrow \underset{+1}{2NaOH} + \underset{0}{H_2}$$

In this reaction, sodium (Na) is oxidised to NaOH while H_2O is reduced to H_2.

31. **(a)** $FeSO_4$ is oxidised to $Fe_2(SO_4)_3$, change in O.N. of Fe is by 1. Hence equivalent weight of Mohr's salt is M/1 = M.

32. **(d)** $2K_2CrO_4 + 2HCl \rightarrow K_2Cr_2O_7 + 2KCl + H_2O$

Coefficients are 2, 2, 1, 2, 1

33. **(a)**

34. **(c)** In Cannizzaro's reaction

$$2HCHO + KOH \rightarrow CH_3OH + HCOOK$$

formaldehyde is getting reduced as well as oxidised.

35. **(a)**

36. **(d)** Zinc gives H_2 gas with dil H_2SO_4/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)

$$[Zn + \underset{\text{(nearly 6\%)}}{2HNO_3} \longrightarrow Zn(NO_3)_2 + 2H] \times 4$$

$$HNO_3 + 8H \longrightarrow NH_3 + 3H_2O$$

$$NH_3 + HNO_3 \longrightarrow NH_4NO_3$$

$$4Zn + 10HNO_3 \longrightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$$

Zn is on the top position of hydrogen in electrochemical series. So Zn displaces H_2 from dilute H_2SO_4 and HCl with liberation of H_2.

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

37. **(c)**

38. **(a)** ON of S in $S_8 = 0$

ON of S in $S_2F_2 = +1$

ON of S in $H_2S = -2$

39. **(a)** More is E°_{RP}, more is the tendency to get itself reduced or more is oxidising power.

40. **(a)** For statement (iii), $HgCl_2$ is reduced to Hg_2Cl_2

41. **(d)** More the negative reduction potential, more is the tendency to lose electron. The reducing power increases as the standard reduction potential becomes more and more negative.

Thus, Li is the strongest reducing agent as the standard reduction potential of Li^+/Li is most negative, –3.05 V.

42. (a) $\overset{+5}{Cl}O_3^- \longrightarrow \overset{0}{Cl_2}$

$x - 6 = -1 \quad x = 0$

$x = +5 \quad x = 0\ (x = \text{oxidation number})$

$$\text{Equivalent mass} = \frac{\text{Molecular mass}}{\text{Oxidation number}} = \frac{84.45}{5} = 16.89$$

43. (b) In SO_2 the O.N. of S can increase and decrease. Hence can behave as reducing and oxidising agent. Oxidation state of S varies from –2 to 6.

44. (a) –(4/3) is the average oxidation state of C in C_3H_4.

45. (c) The redox reaction involve loss or gain of electron(s) i.e. change in oxidation state. Given reaction is not a redox reaction as this reaction involves no change in oxidation state of reactant or product.