Redox Reactions

EXAM DRILL

ANSWERS

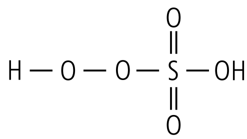
1. Decomposition reaction

2. Salt bridge

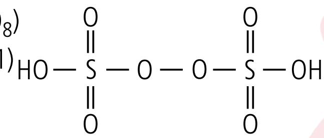
3. Oxygen difluoride (OF_2)

4. **(b) :** NH_4NO_3 may be rewritten as $NH_4^+NO_3^-$
In NH_4^+, O. N. of N is : $x + 4 = +1$ *i.e.*, $x = -3$
In NO_3^-, O. N. of N is : $x + (-6) = -1$ *i.e.*, $x = +5$

5. **(d) :**
Peroxomonosulphuric acid (H_2SO_5)
$1 + x + 2(-1) + 2(-2) + (-2) + 1 = 0$
$x = +6$



Peroxodisulphuric acid ($H_2S_2O_8$)
$1 + (-2) + x + 2(-2) + 2(-1) + x + 2(-2) + (-2) + 1 = 0$
$x = +6$



6. **(a) :** E.M.F. $= E^\circ_{cathode} - E^\circ_{anode} = +0.80 - (-0.76) = 1.56$ V
The maximum e.m.f. of the cell is obtained when the electrode potential of the anode is minimum and that of the cathode is maximum.

7. **(a) :** $\underset{(KMnO_4)}{N_1V_1} = \underset{(H_2O_2)}{N_2V_2} \Rightarrow 0.08 \times 30 = N_2 \times 20$
$N_2 = 0.12$ N
Strength of H_2O_2 = Normality × eq. wt. of H_2O_2
$= 0.12 \times 17 = 2.04\ gL^{-1}$

8. **(b) :** Zn is more reactive than Ag and hence precipitates Ag from $AgNO_3$ solution. As Ag is less reactive than Zn, $ZnSO_4$ solution can be stirred with a silver spoon.

9. **(a)**

10. **(a)**

11. **(b) :** Fe is a stronger reducing agent than H_2 or E° of Fe is less than that of H_2. Hence, liberates H_2 gas when Fe reacts with HCl.

12. **(a)**

13. $\overset{0}{Cl}_{2(g)} + 2OH^-_{(aq)} \longrightarrow \overset{+1}{Cl}O^-_{(aq)} + \overset{-1}{Cl}^{-1}_{(aq)} + H_2O_{(l)}$
Here, the Cl atom undergoes both oxidation +1 and reduction –1. Therefore, it is a disproportionation reaction.

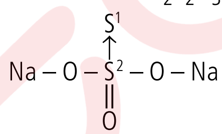
14. (i) $F_2 + 2e^- \longrightarrow 2F^-$ (Oxidant)
$BrO_3^- \longrightarrow BrO_4^- + 2e^-$ (Reductant)

(ii) $I_2 \longrightarrow NaIO_3 + 5e^-$ (Reductant)
$NaClO_3 + 5e^- \longrightarrow Cl_2$ (Oxidant)

15. Oxidation number of Cl in
$HClO \Rightarrow 1 + x - 2 = 0$ or $x = +1$
$ClO_4^- \Rightarrow x + 4(-2) = -1$ or $x = +7$
$CaOCl_2 \Rightarrow +2 + (-2) + 2x = 0$ or $x = 0$
$ClO_2 \Rightarrow x + 2(-2) = 0$ or $x = +4$

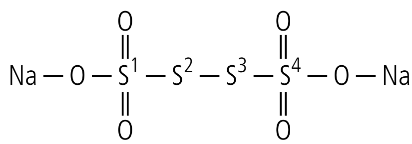
OR

(a) $Na_2S_2O_3$
Structure of $Na_2S_2O_3$ is



The oxidation state of S^1 is –2.
Let oxidation state of S^2 be x.
$\underset{\text{(For Na)}}{2 \times (+1)} + \underset{\text{(For O)}}{3(-2)} + \underset{\text{(For coordinate S)}}{x + 1 \times (-2)} = 0$
$+2 - 6 + x - 2 = 0$
$x = +6$
Thus, the oxidation states of two S atoms in $Na_2S_2O_3$ are –2 and +6.

(b) $Na_2S_4O_6$



From the left, $S_{(1)} = (-2 + (-2) + (-2) + (+1)) = +5$
$S_{(2)} = 0$
$S_{(3)} = 0$
$S_{(4)} = +5$

(c) Na_2SO_3
$2 \times (+1) + x + 3(-2) = 0$
$x = +4$

(d) Na_2SO_4
$+2 + x + (-8) = 0$
$x = +6$

16. We rewrite the two equation in the reduction form
$Sn^{2+}_{(eq)} + 2e^- \longrightarrow Sn_{(s)}$; $E^\circ = -0.14$ V

$Ag^+_{(eq)} + e^- \longrightarrow Ag_{(s)}$; $E° = +0.80$ V

$E°_{cell} = E°_{cathode} - E°_{anode} = 0.80 - (-0.14) = +0.94$ V

17. $E°$ value of Fe^{2+}/Fe (– 0.44 V) is lower than that of Ag^+ / Ag (+ 0.80 V) electrode, therefore, Ag^+ gets reduced and Fe gets oxidised. As a result, concentration of Ag^+ ions decrease while that of NO_3^- ions remains unchanged.

$$2Ag^+_{(aq)} + Fe_{(s)} \longrightarrow 2Ag_{(s)} + Fe^{2+}_{(aq)}$$

18. Element (*D*) is the strongest oxidising agent since its electrode potential is the highest while element *A* is the strongest reducing agent since its electrode potential is the lowest.

19. The given representation is wrong Mg^{2+}/ Mg electrode with lower $E°$ value will act as anode and Zn^{2+}/ Zn electrode with higher $E°$ value will act as cathode.

The correct representation is: Mg | Mg^{2+} (1M) || Zn^{2+} (1M) | Zn

OR

(a) (i) $Zn_{(s)} + Cu^{2+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

Galvanic cell corresponding to the above redox reaction may be depicted as :

$$Zn|Zn^{2+}_{(aq)}\ (1\ M)\ ||Cu^{2+}_{(aq)}\ (1\ M)\ |Cu$$

(ii) The two half-cells reactions are,

$Zn \longrightarrow Zn^{2+} + 2e^-$ (oxidation-half reaction at anode)

$Cu^{2+} + 2e^- \longrightarrow Cu$ (reduction-half reaction at cathode)

Therefore, Zn acts as anode while Cu acts as cathode.

20. $2K_2Cr_2O_7 + 3Sn + 28HCl \longrightarrow 4KCl + 4CrCl_3 + 3SnCl_4 + 14H_2O$

2×294 3×118.7 (M. M. of $K_2Cr_2O_7 = 294$

$= 588$ g $= 356.1$g and At. wt. of Sn = 118.7)

356.1 g Sn react with $K_2Cr_2O_7$ = 588 g

1 g Sn react with $K_2Cr_2O_7 = \frac{588}{356.1} = 1.65$g

Eq. wt of $K_2Cr_2O_7 = \frac{\text{Mol.wt}}{6} = \frac{294}{6} = 49$

A decinormal $K_2Cr_2O_7$ solution means 1000 mL. Solution contains 4.9 g $K_2Cr_2O_7$ *i.e,.* 4.9 g $K_2Cr_2O_7$ present in 1000 mL solution

∴ 1.65 g $K_2Cr_2O_7$ will be present in

$$= \frac{1000}{4.9} \times 1.65 = 336.7 \text{ mL}$$

21. $Cr_2O_7^{2-} + 6S_2O_3^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 3S_4O_6^{2-} + 7H_2O$

$$\frac{M_1V_1}{n_1}(Cr_2O_7^{2-}) = \frac{M_2V_2}{n_2}(S_2O_3^{2-}) \Rightarrow \frac{M_1 \times 20}{1} = \frac{1\times16}{25\times6}$$

$$M_1 = \frac{16}{6\times25\times20} \text{ mol L}^{-1}$$

Conc. of $K_2Cr_2O_7$ in g/L^{-1} = $\frac{16\times294}{6\times25\times20} = 1.568$ g

22. (i) $\overset{+1\,-1}{3HCl} + \overset{+1+5-2}{HNO_3} \longrightarrow \overset{0}{Cl_2} + \overset{+3-2-1}{NOCl} + 2H_2O$

O.N. of N decreases from 5 to 3 and O.N. of Cl increases from –1 to 0.

HNO_3 is oxidising agent and HCl is reducing agent.

(ii) $\overset{+2\,-1}{HgCl_2} + \overset{+1-1}{2KI} \longrightarrow \overset{+2\,-1}{HgI_2} + \overset{+1\,-1}{2KCl}$

No change in oxidation number. So it is not a redox reaction.

(iii) $\overset{+3\,-2}{Fe_2O_3} + \overset{+2-2}{3CO} \xrightarrow{\Delta} \overset{0}{2Fe} + \overset{+4\,-2}{3CO_2}$

O.N. of Fe decreases from +3 to 0 hence, Fe_2O_3 is oxidising agent. O.N. of C increases from +2 to +4 hence, CO is reducing agent.

OR

(i) $\overset{+3-1}{PCl_3} + \overset{+1\,-2}{3H_2O} \longrightarrow \overset{+1-1}{3HCl} + \overset{+1+3\,-2}{H_3PO_3}$

No change in oxidation number. So, it is not a redox reaction.

(ii) $\overset{-3\,+1}{4NH_3} + \overset{0}{3O_2} \longrightarrow \overset{0}{2N_2} + \overset{+1\,-2}{6H_2O}$

O.N. of N increases from –3 to 0 hence, NH_3 acts as a reducing agent. O.N. of O decreases from 0 to –2 hence, O_2 acts as an oxidising agent.

(iii) $\overset{0}{Cl_2} + 2OH^- \longrightarrow \overset{+1}{ClO^-} + \overset{-1}{Cl^-} + H_2O$

Oxidation number of Cl increase from 0 to +1. Hence Cl_2 act as reducing agent oxidation number of Cl decrease from 0 to –1. Hence Cl_2 act as oxidising agent.

23. The balanced redox reactions involving oxidation of Fe^{2+} ions by $Cr_2O^{2-}_{7(aq)}$ and $MnO^-_{4(aq)}$ respectively are :

$Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

$MnO_4^- + 5Fe^{2+} + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

$$\frac{M_1V_1}{n_1}(Cr_2O_7^{2-}) = \frac{M_2V_2}{n_2}(Fe^{2+})$$

Here, we suppose V_2 cm^3 of M_2 Fe^{2+} is titrated against 24.50 cm^3 of 0.1 M $Cr_2O_7^{2-}$ and V_1 cm^3 of 0.1 M MnO_4^- solution.

$$\frac{24.5\times0.1}{1}(Cr_2O_7^{2-}) = \frac{M_2V_2}{6}(Fe^{2+}) \quad ...(i)$$

$$\frac{V_1\times0.1}{1}(MnO_4^-) = \frac{M_2V_2}{5}(Fe^{2+}) \quad ...(ii)$$

On equating eq. (i) and (ii), $V_1 = 24.5 \times \frac{6}{5} = 29.4$ cm^3

24. The Cu^+ undergoes disproportionation to form Cu^{2+} and Cu.

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

Thus, Cu^{+} or Cu_2O acts as both oxidant as well as reductant.

(i) When heated in air, Cu_2O is oxidised to CuO

$$\overset{+1\,+2}{Cu_2O} + \overset{0}{1/2O_2} \longrightarrow \overset{+2\,-2}{2CuO}$$

i.e., Cu_2O acts as a reductant and reduces O_2 to O^{2-}.

(ii) When heated with Cu_2S, it oxidises S^{2-} to SO_2 and hence Cu_2O act as an oxidant.

$$\overset{+1-2}{2Cu_2O} + \overset{+1\,-2}{Cu_2S} \longrightarrow \overset{0}{6Cu} + \overset{+4}{SO_2}$$

Cu_2O is reduced to Cu.

25. (i) $\overset{0}{Zn} + H^{+} + \overset{+5-2}{NO_3^{-}} \longrightarrow Zn^{2+} + \overset{-3\,+1}{NH_4^{+}} + \overset{+1-2}{H_2O}$

Oxidation

$$\overset{0}{Zn} + H\overset{+5}{N}O_3 \longrightarrow Zn^{2+} + \overset{-3}{N}H_4^{+} + H_2O$$

Reduction

$Zn \longrightarrow Zn^{2+} + 2e^{-}] \times 4$

$10H^{+} + NO_3^{-} + 8e^{-} \longrightarrow NH_4^{+} + 3H_2O$

$4Zn + NO_3^{-} + 10H^{+} \longrightarrow 4Zn^{2+} + NH_4^{+} + 3H_2O$

(ii) $\overset{0}{I_2} + \overset{+2}{S_2}O_3^{2-} \longrightarrow I^{-} + \overset{+2.5}{S_4}O_6^{2-}$

Total increase in O.N. = 0.5 × 4 = 2

Total decrease in O.N. = 1 × 2 = 2

To equalise O.N. multiply $S_2O_3^{2-}$ and I^{-} by 2.

$$I_2 + 2S_2O_3^{2-} \longrightarrow 2I^{-} + S_4O_6^{2-}$$

OR

(i) $H_2S + Fe^{3+} \longrightarrow Fe^{2+} + H^{+} + S$

Oxidation half reaction:

$H_2S \longrightarrow 2H^{+} + S + 2e^{-}$ ...(i)

Reduction half reaction:

$Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$ ...(ii)

Multiplying equation (ii) by 2 and adding it to equation (i), we have

$H_2S + 2Fe^{3+} \longrightarrow 2Fe^{2+} + 2H^{+} + S$

(ii) $\overset{0}{Cl_2} \longrightarrow \overset{-1}{Cl^{-}}$ (Reduction half-reaction)

$Cl_2 + 2e^{-} \longrightarrow 2Cl^{-}$ ...(i)

$\overset{+5}{BrO_3^{-}} \longrightarrow \overset{+7}{BrO_4^{-}}$ (Oxidation half-reaction)

$BrO_3^{-} \longrightarrow BrO_4^{-} + 2e^{-}$

$BrO_3^{-} + 2OH^{-} \longrightarrow BrO_4^{-} + 2e^{-} + H_2O$ ...(ii)

Adding (i) and (ii)

$Cl_2 + BrO_3^{-} + 2OH^{-} \longrightarrow BrO_4^{-} + 2Cl^{-} + H_2O$

26. (i) In this redox reaction, H in $LiAlH_4$ gets oxidised because of the addition of oxygen atom that leads to the formation of OH^{-}. Propan-2-ol (CH_3COCH_3) gets reduced because of addition of hydrogen atom to propan-2-ol ($CH_3CH(OH)CH_3$).

(ii) This is not a redox reaction as neither hydrogen or oxygen or e^{-} is removed or added.

(iii) This is not a redox reaction as neither hydrogen or oxygen or e^{-} is removed or added.

27. (a) $Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$

(b) $Zn \mid ZnSO_4(1M) \parallel CuSO_4(1M) \mid Cu$

(c) $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode} = 0.34 - (-0.76) = 1.1$ V

(d) Zn electrode will act as anode as oxidation takes place in the electrode. Cu electrode will act as the cathode as reduction takes place in this electrode.

(e) The direction of electrons flow is from anode to cathode *i.e.*, from Zn electrode to copper electrode. The direction of current flow is from cathode to anode *i.e.*, from Cu electrode to Zn electrode.

OR

(a) A reaction in which a particular species simultaneously gets oxidised and reduced, is known as disproportionation reaction.

e.g., $\overset{+1}{Hg_2Cl_2} \longrightarrow \overset{0}{Hg} + \overset{+2}{HgCl_2}$

In this reaction, Hg_2Cl_2 is getting oxidised to $HgCl_2$ and also getting reduced to Hg.

(b) **Step 1 :** The skeletal ionic equation is

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

Step 2 : Assign oxidation numbers for Cr and S

$\overset{+6}{Cr_2}O_7^{2-}{}_{(aq)} + \overset{+4}{S}O_3^{2-}{}_{(aq)} \longrightarrow Cr^{+3}_{(aq)} + 3\overset{+6}{S}O_4^{2-}{}_{(aq)}$

This indicates that the dichromate ion is the oxidant and the sulphite ion is the reductant.

Step 3 : Calculate the increase and decrease of oxidation number, and make them equal :

$\overset{+6\,-2}{Cr_2O_7^{2-}}{}_{(aq)} + \overset{+4\,-2}{3SO_3^{2-}}{}_{(aq)} \longrightarrow 2Cr^{3+}_{(aq)} + \overset{+6\,-2}{3SO_4^{2-}}{}_{(aq)}$

Step 4 : As the reaction occurs in the acidic medium, and further the ionic charges are not equal on both the sides, add $8H^{+}$ on the left to make ionic charges equal.

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

Step 5 : Finally, count the hydrogen atoms, and add appropriate number of water molecules (*i.e.*, $4H_2O$) on the right to achieve balanced redox change.

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

(c) Skeleton equation :

$$\overset{+7\,-2}{MnO_4^{-}} + \overset{+4\,-2}{SO_2} \longrightarrow \overset{+2}{Mn^{2+}} + \overset{+1+6-2}{HSO_4^{-}}$$

Oxidation half : $SO_2 \longrightarrow HSO_4^{-}$

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

Oxidation half :

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

(Add $2H_2O$ molecules to balance O atoms)

$SO_2 + 2H_2O \longrightarrow HSO_4^- + 3H^+ + 2e^-$ ...(i)

Reduction half :

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

(Add $4H_2O$ molecules to balance O atoms and H atoms)

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

Overall reaction, adding eq. (i) × 5, eq. (ii) × 2, Now

$$2MnO_4^- + 5SO_2 + 2H_2O + H^+ \rightarrow 5HSO_4^- + 2Mn^{2+}$$

28. (a) The potential difference set up between the metal and its own ions in the solution is called electrode potential. In general, it is the tendency of an electrode to gain or lose electrons.

(b) Smaller the reduction potential, more easily the substance is oxidised and hence acts as a strong reducing agent. Therefore, order of increasing reducing power will be Ag < Mg < K.

(c) $E°_{Cu^{2+}/Cu} - E°_{Ag^+/Ag} = -0.46$ ...(i)

$E°_{Cu^{2+}/Cu} - E°_{Zn^{2+}/Zn} = 1.10$ ...(ii)

Subtracting eq. (i) from eq. (ii) we get

$$E°_{cell} = E°_{Ag^+, Ag} - E°_{Zn^{2+}, Zn} = 1.10 - (-0.46) = 1.56 \text{ V}$$

The $E°$ value for the cell $Zn/Zn^{2+}(1M) || Ag^+(1M) | Ag$ is 1.56 V.

OR

(a) EMF of a cell is used to predict the spontaneity of a redox reaction. If the EMF comes out to be positive, the reaction takes place, and if the EMF comes out to be negative, the reaction will not take place.

(b) $Zn_{(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_{2(g)}$

The cell is represented as :

$$Zn_{(s)} | Zn^{2+}_{(aq)} || H^+_{(aq)} | H_{2(g)} (Pt_{(s)})$$

$E°_{cell} = E°_{H^+/H_2} - E°_{Zn^{2+}/Zn} = 0 - (-0.76) = +0.76$ V

Now, the $E°_{cell}$ is positive, hence the reaction takes place. Yes, zinc on reaction with 1M sulphuric acid will liberate hydrogen gas.

(c) $E°_{cell} = E°_{Cu^{2+}/Cu} - E°_{Ni^{+2}/Ni} = 0.34 \text{ V} - (-0.25 \text{ V}) = 0.59$ V

Hence, $E°_{cell}$ comes out to be positive. The reaction will take place. Therefore, $CuSO_4$ solution cannot be stored in a nickel vessel.

29. (a) The arrangement of various elements in order of their increasing reduction potential (with respect to standard hydrogen electrode) is called electrochemical series. We can compare the relative activity of elements by this series because activity can be compared in terms of oxidation potential.

The redox reaction of an electrochemical cell is feasible if EMF of the cell is positive. If the EMF comes out to be negative then the reverse reaction would be feasible.

(b) Reduction half : $Zn^{2+} + 2e^- \longrightarrow Zn$

Oxidation half : $Cd \longrightarrow Cd^{2+} + 2e^-$

$E°_{cell} = E°_{Zn/Zn^{2+}} - E°_{Cd^{2+}/Cd} = 0.763 - 0.403 = 0.36$ V

OR

(a) Electrode potential is defined as the tendency of an electrode to either lose or gain electrons *i.e.*, tendency to get either oxidised or reduced when it is in contact with a solution of its own ions is known as electrode potential.

We can determine the reducing power of an element by electrode potential. If a substance has

(i) a negative $E°$, it is stronger reducing agent than H_2.

(ii) a positive $E°$, it is stronger oxidising agent than H_2.

(b) Mg^{2+}/Mg electrode has lower potential it act as anode and Al^{3+}/Al electrode has higher potential it act as cathode.

The cell reaction is : $3Mg + 2Al^{3+} \longrightarrow 3Mg^{2+} + 2Al$

$E°_{cell} = E°_{cathode} - E°_{anode} = -1.66 - (-2.36) = +0.70$ V

30. **(i)** The electrode with higher reduction potential has a strong tendency to gain electrons and hence acts as the cathode while the electrode with lower reduction potential has a strong tendency to set oxidised and hence acts as the anode. Cr with lower electrode potential ($E°$ = –0.74 V) acts as the anode while Fe with higher electrode potential ($E°$ = –0.44 V) acts as the cathode.

(ii) $E°_{cell} = E°_{cathode} - E°_{anode}$

$= -0.44 - (-0.74) = +0.30$ V

(iii) $2Cr_{(s)} + 3Fe^{2+}_{(aq)} \longrightarrow 2Cr^{3+}_{(aq)} + 3Fe_{(s)}$

$Cr_{(s)} | Cr^{3+}_{(1M)} || Fe^{2+}_{(1M)} | Fe_{(s)}$

(iv) $Cr_{(s)} \longrightarrow Cr^{3+}_{(aq)} + 3e^-$ (Oxidation)

$Fe^{2+}_{(aq)} + 2e^- \longrightarrow Fe_{(s)}$ (Reduction)

(v) $E°_{cell} = -0.036 - (-0.74) = +0.704$ V