HINTS & SOLUTIONS WORKBOOK - 2

Stoiciometry-II (Redox Reactions)

Daily Tutorial Sheet	Level-0

1. Reverse the signs of oxidation potential to get the values of the reduction or electrode potentials. Thus,

$E_{Zn|Zn^{2+}} = 0763V \therefore E°_{Zn^{2+}|Zn} = -0.763V$ and $\because E°_{Cd|Cd^{2+}} = 0.403V$

$\therefore \quad E°_{Cd^{2+}|Cd} = -0.403V$

Since $Zn^{2+}|Zn$ electrode is at lower potential, therefore, it acts as the anode while $Cd^{2+}|Cd$ electrode with higher potential acts as the cathode.

In other words, Zn loses electrons and Cd^{2+} ion accepts them. Therefore, cell reaction is :

$$Zn + Cd^{2+} \longrightarrow Zn^{2+} + Cd$$

And $E^0_{cell} = E^0_{cd^{2+}|cd} - E°_{Zn^{2+}|Zn} = -0.403 - (-0.763) = +0.360V$

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

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

Subtracting equation (i) from equation (ii), we have,

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

In other words, $E°$ of the cell, $Zn \mid Zn^{2+}(M) \mid\mid Ag^{+}(M) \mid Ag$, is 1.56 V.

3. Since, $Mg^{2+}(aq) \mid Mg$ electrode $= -2.36\,V$ is at a lower potential than $Al^{3+}(aq) \mid Al$ electrode $= -1.66\,V$, therefore, $Mg^{2+}(aq) \mid Mg$ electrode acts as the anode and $Al^{3+}(aq)|Al$ acts as the cathode. In other words, Mg loses electrons and Al^{3+} ion accepts electrons. Thus, the cell reaction is

$3Mg + 2Al^{3+} \longrightarrow 3Mg^{2+} + 2Al$ and $E°_{cell} = E°_{Al^{3+}|Al} - E°_{Mg^{2+}|Mg} = -1.66 - (-2.36) = +0.70V$

4. $E°_{cell} = E°_{Cu^{2+}|Cu} - E°_{Al^{3+}|Al} = 0.34 - (-1.66) = 2.0V$

5. If the metal is to react with dil. H_2SO_4 (i.e., H^+ ions) to produce H_2 gas, the metal should have a lower electrode potential than that of standard hydrogen electrode, i.e., 0.0 V.

(i) Since, $E°_{Cu^{2+}|Cu} = 0.34V$ is higher than $E°_{H^{+}|H_2} = 0.0V$, therefore, Cu will not react with 1N H_2SO_4 to produce H_2 gas.

(ii) Since, $E°_{Pb^{2+}|Pb} = -0.13\,V$ and $E°_{H^{+}|H_2} = 0.0\,V$, therefore, lead will react with $1\,N\,H_2SO_4$ to produce H_2 gas.

(iii) Since $E°_{Fe^{2+}|Fe} = -0.44V$ and $E°_{H^{+}|H_2} = 0.0V$, therefore iron will react with $1\,N\,H_2SO_4$ to produce H_2 gas.

6. Since, $E°_{Zn^{2+}|Zn} = -E°_{Zn,|Zn^{2+}} = -0.76V$, is at a lower potential than $E°_{Cu^{2+}|Cu} = 0.34V$, therefore, Zn can only lose electrons to Cu^{2+} ions. Conversely, Zn^{2+} cannot accept electrons from Cu and hence the following reaction will not occur. $Zn^{2+} + Cu \longrightarrow Zn + Cu^{2+}$

In other words, $ZnSO_4$ solution can be safely stored in a copper vessel.

7. Since, $E^\circ_{Ag^+|Ag} = +0.80V$ is higher than $E^\circ_{Cu^{2+}|Cu} = -E^\circ_{Cu|Cu^{2+}} = -(-0.34) = +0.34V$, therefore, Ag^+ ions can easily accept electrons from Cu. In other words, the following reactions will occur.

$Cu + 2Ag^+ \longrightarrow Cu^{2+} + 2Ag$

Hence, 1 M $AgNO_3$ solution cannot be stirred with a copper spoon.

8. Since, Cu will react with Ag^+ ions, as discussed in Ans to Q.7 above, therefore $AgNO_3$ solution cannot be stored in copper vessel.

9. Since, $E^\circ_{Zn|Zn^{2+}} = -E^\circ_{Zn^{2+}|Zn} = -0.76V$ is lower than $E^\circ_{Cu^{2+}|Cu}$, therefore, Zn will lose electrons and copper will accept them. In other words, the following reaction will occur, $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$.
Since, blue Cu^{2+} ions are consumed and colourless Zn^{2+} ions are produced during the above reaction, therefore, colour of $CuSO_4$ solution gets discharged when zinc rod is dipped in it.

10. $N \rightarrow -3$ to $+5$

$S \rightarrow -2$ to $+6$

$Cl \rightarrow -1$ to $+7$

11. In case of nitric acid, nitrogen is in maximum oxidation state of +5 and therefore, cannot be further oxidized. So, HNO_3 acts as an oxidizing agent only.
On the other hand, N is HNO_2 is in +3 oxidation state and therefore, can undergo either oxidation or reduction. So, HNO_2 acts both as an oxidizing as well as reducing agent.

12. No. In both $Cr_2O_7^{2-}$ and CrO_4^{2-}, Cr is in + 6 oxidation state and therefore, there is no change in oxidation state.

13. **(i)** $Zn \longrightarrow Zn^{2+} + 2e^-$; $Pb^{2+} + 2e^- \longrightarrow Pb$

(ii) $2I^- \longrightarrow I_2 + 2e^-$; $2Fe^{3+} + 2e^- \longrightarrow Fe^{2+}$

(iii) $2Na \longrightarrow 2Na^+ + 2e^-$; $Cl_2 + 2e^- \longrightarrow 2Cl^-$

(iv) $Mg \longrightarrow Mg^{2+} + 2e^-$; $Cl_2 + 2e^- \longrightarrow 2Cl^-$

(v) $Zn \longrightarrow Zn^{2+} + 2e^-$; $2H^+ + 2e^- \longrightarrow H_2$

14. The electrode to be chosen as the anode should have high oxidation potential.
So, out of $Zn \mid Zn^{2+}$ and $Fe \mid Fe^{2+}$, the former has higher oxidation potential.
i.e. $E^\circ_{Zn|Zn^+} = 0.76\,V$ & $E^\circ_{Fe|Fe^{2+}} = 0.44\,V$

The cathode should have higher reduction potential. So, $Cu^{2+} \mid Cu$ can act as cathode.
So, cell reaction would be

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

$E_{cell} = (0.76 + 0.34)V = 1.10V$

In the internal circuit, the direction of flow of electrons will be from cathode to anode and from anode to cathode in the external circuit.

15. $E_{Fe|Fe^{2+}} = 0.44V$; $E_{Zn|Zn^{2+}} = 0.76V$; $E^\circ_{Ni|Ni^{2+}} = 0.25V$

(i) Since oxidation potential of $Fe \mid Fe^{2+}$ is less than oxidation potential of $Zn \mid Zn^{2+}$,
$\therefore$ Fe cannot reduce Zn^{2+} ions.

(ii) Since oxidation potential of $Fe \mid Fe^{2+}$ is higher than that $Ni \mid Ni^{2+}$, so Fe can reduce Ni^{2+} ions.

16. The standard reduction potential of the four metallic elements – A, B, C and D have been given. Higher the reduction potential, lesser will be the electropositive character.

$\therefore$ Decreasing order of electropositive character is – B > C > D > A

17. Since the reduction potential of Br_2 $(E^\circ_{Br_2|Br^-} = 1.09\,V)$ is higher than that of $I_2(E^\circ_{I_2|I^-} = 0.54\,V)$, so, Br_2 will get reduced while I^- ions will undergo oxidation. So, the reaction would be

$$Br_2 + 2I^- \longrightarrow 2Br^- + I_2 \quad (E^\circ = 0.55V)$$

18. **(i)** $E_{Zn^{2+}|Zn} = -0.76\,V$ and $E^\circ_{Cu^{2+}|Cu} = 0.34\,V$

Zinc will displace copper from its salt solution, copper sulphate solution should not be stored in a zinc vessel.

(ii) Since reduction potential of $Ag^+ \mid Ag$ is higher than $Cu^{2+} \mid Cu$, so, copper ion will not be reduced. So, copper sulphate solution can be stored in silver vessel.

(iii) Similar to the case of silver vessel, copper sulphate solution can be stored in gold vessel.

19. In case of Cu_2O, Cu is in +1 oxidation state. So, it can undergo both oxidation as well as reduction. That is why it can act both as an oxidant and reductant.

As an oxidant

$$Cu_2O + H_2 \longrightarrow 2Cu + H_2O$$

0.5 +1 0 As a reductant

20. Reduction $2MnO_4^- + 6e^- \longrightarrow 2MnO_2$

Oxidation $Br^- \longrightarrow BrO_3^- + 6e^-$

Complete redox reaction would be $2MnO_4^- + Br^- + H_2O \longrightarrow 2MnO_2 + BrO_3^- + 2OH^-$

21. $2MnO_4^- + 6I^- + 4H_2O \longrightarrow 2MnO_4 + 3I_2 + 8OH^-$

22. **(a)** $KI_3 \longrightarrow +1 + 3(x) = 0$

$\therefore \quad x = +1/3$

(b) $H_2S_4O_6 \longrightarrow 2(+1) + 4(x) + 6(-2) = 0$

$\Rightarrow \quad 4x - 10 = 0 \quad \Rightarrow \quad x = 10/4$

(c) $Fe_3O_4 \longrightarrow 3(x) + 4(-2) = 0$

$\therefore \quad x = +8/3$

(d) $\overset{1}{C}H_3\overset{2}{C}H_2OH \longrightarrow 2(x) + 6(+1) + 1(-2) = 0$

$\Rightarrow \quad 2x + 4 = 0 \quad \Rightarrow \quad 2x = -4$

So, sum of 0.5. of both carbon atoms is –4.

Oxidation number of $C-1 = -3$

Oxidation number of $C-2 = -1$

(e) $\overset{1}{C}H_3\overset{2}{C}OOH \longrightarrow 2(x) + 4(+1) + 2(-2) = 0$

$\Rightarrow \quad 2x = 0 \quad \Rightarrow \quad x = 0$

So, sum of oxidation states of both carbon atoms is zero.

Oxidation number of $C-1 = -3$

Oxidation number of $C-2 = +3$

23. Compounds in which elements are in higher oxidation states are unstable. So, they act as oxidizing agents. Here, Ag in AgF_2 is in unstable + 2 oxidation states.

24. Alcohol will act as a solvent for toluene which is insoluble in aqueous medium.

(a) $C_6H_5CH_3 + 2KMnO_4 \longrightarrow C_6H_5COOH + 2MnO_2 + 2KOH$

(b) **(i)** $H_2SO_4 + Cl^- \longrightarrow \underset{\text{(pungent smell)}}{HCl \uparrow} + HSO_4^-$

(ii) $H_2SO_4 + Br^- \longrightarrow HBr + HSO_4^-$

$\underset{\text{(OA)}}{H_2SO_4} + 2HBr \longrightarrow Br_2 + SO_2 + 2H_2O$

25. Higher the standard electrode potential, higher will be the reducing power.

$\therefore$ Increasing order of reducing power is $K^+ < Mg^{2+} < Cr^{3+} < Hg_2^{2+} < Ag^+$