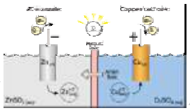


Chapter - 8

Redox Reactions

	Oxidation		Reduction
1.	Addition of oxygen	**1.**	Removal of oxygen
2.	Removal of an Hydrogen	**2.**	Addition of Hydrogen
3.	Addition of an electronegative element.	**3.**	Removal of an electronegative element.
4.	Addition of an electropositive element	**4.**	Addition of an electropositive element.
5.	Loss of electron(s)	**5.**	Gain of electron(s)
6.	Increase in oxidation number.	**6.**	Decrease in oxidation, number.

- **Reducing Agent :** Donor of electron(s).
- **Oxidising Agent :** Acceptor of electron(s).
- **Redox Reaction :** Reactions in which oxidation and reduction takes place simultaneously.
- **Oxidation Number :** It is charge that an atom appears to have in a given species when the bonding electron are counted towards more electronegative atom.
- **Calculation of Oxidation Number :**

  (a) O.S. of all the elements in their elemental form (in standard state) is taken as zero. O.S. of element of Cl_2, F_2, O_2, P_4, O_3, Fe(s), H_2, N_2, C (graphite) is zero.

  (b) Common O.S. of elements of 1st group is +1. Common O. S. of elements of 2nd group + 2.

  (c) For ions composed of only one atom, the oxidation number is equal to the charge on the ion.

(d) The oxidation number of oxygen in most compounds is – 2. While in peroxides (*e.g.*, H_2O_2, Na_2O_2), each oxygen atom is assigned an oxidation number of – 1, in super oxides (*e.g.*, KO_2, RbO_2) each oxygen atom is assigned an oxidation number of – (½).

(e) In oxygen di fluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of + 2 and + 1, respectively.

(f) The oxidation number of hydrogen is + 1 but in metal hydride its oxidation no. is – 1.

(g) In all its compounds, fluorine has an oxidation number of – 1.

(h) The algebraic sum of the oxidation number of all the atoms in a compound must be zero.

(i) In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must equal the charge on the ion.

- **Types of Redox Reactions:**

(i) Combination Reaction : **0 0 +2 –3**

$$3\,Mg\,(s) + N_2\,(g) \rightarrow Mg_3N_2\,(s)$$

(ii) Decomposition : **+1 +5 – 2 +1 – 1 0**

$$3KClO_3(s) \rightarrow 2\,KC1(s) + 3O_2\,(g)$$

(iii) Metal Displacement : **+ 2 + 6 – 2 0 + 2 + 6 – 2 0**

$$CuSO_4\,(aq) + Zn(s) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

(iv) Non-metal displacement : **0 + 1 – 2 + 2 – 2 + 1 0**

$$3\,Ca\,(s) + 2\,H_2O\,(1) \rightarrow Ca\,(OH)_2 + H_2\,(g)$$

(v) Disproportionation : It is a reaction in which same element is reduced and oxidized simultaneously.

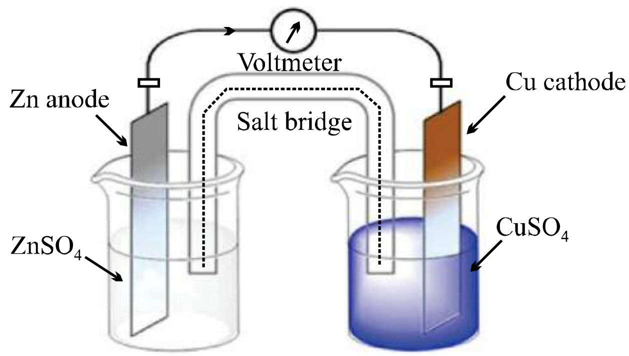
0 –1 + 1

$$C1_2\,(g) + 2\,OH^-\,(aq) \rightarrow Cl^-\,(aq) + ClO^-\,(aq) + H_2O\,(1)$$

- **Stock Notation :** Representing oxidation number of metal in Roman numerals within parenthesis after the symbol or name of metal in the molecular formula or name of a compound. *e.g.*, Stock Notation of Ferric oxide is $Fe_2(III)O_3$ or Iron (III) oxide.
- **Fractional Oxidation Number :** When two or more atoms of an element are present in different oxidation states, then calculated oxidation number may comes out as fractional due to average of all the different oxidation states.

In reality no element can have a fractional oxidation state.. .

- **Electrode or Redox Couple :** A redox couple is defined as having together oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction.
- **Electrode Potential (E) :** Tendency of an electrode to gain or lose electrons.
- **Standard Electrode Potential (E°)** : Electrode Potential measured at 298 K and 1M concentration of metal ions (or 1 bar pressure of gas).
- **Electrochemical Cell :** A device in which chemical energy of a spontaneous redox reaction is converted into electrical energy.



Cell diagram,

$$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

LHS oxidation, $Zn \rightarrow Zn^{2+} + 2e^-$

RHS reduction $Cu^{2+} + 2e^- \rightarrow Cu$

Overall reaction $Zn + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu$

- Representation of an Electrochemical cell :

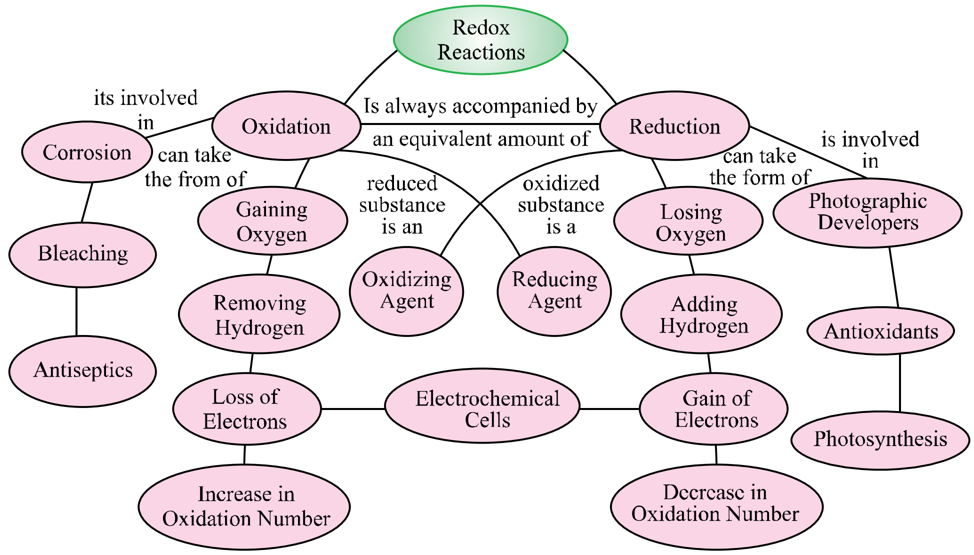
———— Flow of electrons ————→

←———— Flow of current ————

Left Electrode	Salt Bridge	**R**ight Electrode
Oxidation		**R**eduction
Anode		**C**athode
Negative		**P**ositive

- **Functions of Salt Bridge :** (i) To complete inner circuit. (ii) To maintain electrical neutrality.

Flow learning



1-Mark Questions

Oxidation-Reduction : Classical, Electronic and Oxidation Number Concept

1. Define oxidation and reduction according to electronic concept.
2. Define oxidation and reduction according to oxidation number.
3. A freshly cut apple is almost white but it turns reddish brown after some. Give reason.
4. Define oxidation number.
5. Write oxidation number of Mn in $KMnO_4$.
6. Write oxidation Number of Cr in $Cr_2O_7^{2-}$.
7. Write Stock Notation of MnO_2 and $AuCl_3$.

Redox Reactions

8. Define redox reaction with example.
9. Define disproportionation reaction. Give one example.
10. Define the term Redox tirration.
11. Name the indicator used in redox titrations involving $K_2Cr_2O_7$ as an oxidizing agent.

Redox Reactions and Electrode Processes

12. At what concentration of Cu^{2+} (aq) will electrode potential become equal to its standard electrode potential ? [**Ans.** 1 M]

13. The standard reduction potentials of three metals cations X, Y and Z are + 0.52, – 3.03 and – 1.18 V respectively. Arrange X, Y and Z in order of increasing reducing power. [**Ans.** X < Z < Y]

14. An electrochemcial cell consists of two electrodes *i.e.*, Anode and Cthode. What is the direction of flow of electrons in this cell ?

15. Why anode is negatively charged in an electrochemical cell.

16. Out of Zn and Cu vessel one will be more suitable to store 1 M HCl ? [**Ans.** Cu]

Given $E^{\circ}_{Zn^{2+}/Zn} = -0.76$ V, $E^{\circ}_{Cu^{2+}/Cu} = +0.34$ V.

15. Is it safe to stir 1 M $AgNO_3$ solution with copper spoon ? [**Ans.** No]

Given $E^{\circ}_{Ag^{+}/Ag} = +0.80$ V, $E^{\circ}_{Cu^{2-}/Cu} = +0.34$ V.

2 Mark Questions

Oxidation-Reduction: Classical, Electronic and Oxidation Number Concept

1. Identify oxidant and reluctant in the reaction : I_2 (aq) + $2S_2O_3^{2-}$ (aq) $\rightarrow$ 2 I^- (aq) + $S_4O_6^{2-}$(aq).

2. Calculate oxidation number of Fe in Fe_3O_4 and write a suitable justification of your answer.

3. Oxidation-reduction reactions are complementary. Explain.

4. Write formula for the following compounds :

(i) Mercury (II) chloride

(ii) Nickel (II) sulphate

(iii) Iron (III) sulphate

(iv) Chromium (III) oxide

Redox Reactions

5. Justify that the reaction : $H_2O(s) + F_2 \rightarrow HF + HOF$ is a redox reaction. [NCERT]

6. A decomposition reaction may or may not be a redox reaction. Write two decomposition reactions in support of the statement.

7. Split the reaction 2 K (s) + $C1_2$ (g) $\rightarrow$ 2 KC1 (s) into oxidation and reduction half reactions.

8. Calculate the oxidation number of underlined elements in following compounds :

(i) $Ca\underline{O}_2$ (ii) $H_2\underline{S}_2O_7$ (iii) $K_2\underline{Mn}O_4$ (iv) $K\underline{I}_3$

Redox Reactions and Electrode Processes

9. Write the functions of salt bridge in an electrochemical cell.

10. Define the term redox couple. Write the practical application of redox couple.

11. The standard reduction potentials of two metals A and B are – 0.76 V and + 0.34 V respectively. An electrochemical cell is formed using electrodes of these metals.

(i) Identify the cathode and anode.

(ii) Write the direction of flow of electron.

3 Mark Questions

Oxidation-Reduction : Classical, Electronic and Oxidation Number Concept

1. Calculate oxidation number of :

(i) Cr in $Cr_2O_4^{2-}$

(ii) O in KO_2

(iii) Na in Na_2O_2.

2. Account for the following :

(i) HNO_3 acts as oxidizing agent while HNO_2 can act both as reducing and oxidizing agent.

(ii) AgF_2 is unstable compound and a strong oxidizing agent.

(iii) Ozone acts as an oxidizing agent.

Redox Reactions

3. Rermanganate ion (MnO_4^-) reacts with sulphur di oxide gas in acidic medium to produce Mn^{2+} ion and hydrogen sulphate ion. Write ionic equation and blance by ion electron method.

4. Balance the following equation by oxidation number method :

$P_4(s) + OH^- (aq) \rightarrow PH_3 + H_2PO_2^- (aq)$

5. Balance the following equation in basic medium :

$C1_2O_7 (g) + H_2O_2 (1) \rightarrow ClO_2^- (aq) + O_2 (g)$

Redox Reactions and Electrode

6. Depict the galvanic cell in which the reaction Zn (s) + 2 Ag^+ (aq) → Zn^{2+} (aq) + 2 Ag (s) takes place. Further show :

   (i) Which electrode is negatively charged ?

   (ii) The carriers of the current in the cell

   (iii) Individual reaction at each electrode. [NCERT]

7. Explain why ?

   (i) Reaction $FeSO_4$ (aq) + Cu (s) → $CuSO_4$ (aq) + Fe does not occur.

   (ii) Zinc can displace copper from aqueous $CuSO_4$ solution but Ag cannot.

   (iii) Solution of $AgNO_3$ turns blue when copper rod is immersed in it.

5 Mark Questions

Redox Reactions

1. (i) MnO_4^{2-} undergoes disproportionation reaction in acidic medium but MnO_4^- does not. Give reason.

   (ii) Give one example each of the following redox reactions:

   (a) Combination reaction

   (b) Decomposition reaction

   (c) Metal displacement reaction

Redox Reactions and Electrode Processes

2. Consider the cell reaction of an electrochemical cell : Ni(s) + 2 Ag^+(aq) → Ni^{2+} (aq) + 2 Ag (s) and answer the followmg questions :

   (i) Write anode and cathode half reactions.

   (ii) Mention the direction of flow of electrons.

   (iii) How is the electrical neutrality maintained in the solutions of the two half cells.

   (iv) Write the formula for calculating standard emf of this cell.

   (v) How does the emf change when the concentration of silver ions is decreased ?