CBSE Class 11 Chemistry

Quick Revision Notes

Chapter 8

Redox Reactions

1. Redox reactions are those reactions in which oxidation and reduction takes place simultaneously
2. **Classical view of redox reactions**
    - Oxidation is addition of oxygen / electronegative element to a substance or removal of hydrogen / electropositive element from a substance
    - Reduction is removal of oxygen / electronegative element from a substance or addition of hydrogen / electropositive element to a substance
3. **Redox reactions in terms of Electron transfer**
    - Oxidation is defined as loss of electrons by any species.
    - Reduction is defined as gain of electrons by any species.
4. In oxidation reactions there is loss of electrons or increase in positive charge or decrease in negative charge
5. In reduction reactions there is gain of electrons or decrease in positive charge or increase in negative charge
6. Oxidising agents are species which gain one or more electrons and get reduced themselves
7. Reducing agents are the species which lose one or more electrons and gets oxidized themselves
8. Oxidation number denotes the oxidation state of an element in a compound ascertained according to a set of rules. These rules are formulated on the basis that electron in a covalent bond belongs entirely to the more electronegative element.
9. **Rules for assigning oxidation number to an atom**
    - Oxidation number of Hydrogen is always +1 (except in hydrides, it is -1).
    - Oxidation number of oxygen in most of compounds is -2. In peroxides it is (-1). In superoxides, it is (-1/2). In OF_2 oxidation number of oxygen is +2. In O_2F_2 oxidation number of oxygen is +1
    - Oxidation number of Fluorine is -1 in all its compounds

- For neutral molecules sum of oxidation number of all atoms is equal to zero
- In the free or elementary state, the oxidation number of an atom is always zero. This is irrespective of its allotropic form
- For ions composed of only one atom, the oxidation number is equal to the charge on the ion
- The algebraic sum of the oxidation number of all the atoms in a compound must be zero
- For ions the sum of oxidation number is equal to the charge on the ion
- In a polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must be equal to the charge on the ion

10. Oxidation state and oxidation number are often used interchangeably
11. According to Stock notation the oxidation number is expressed by putting a Roman numeral representing the oxidation number in parenthesis after the symbol of the metal in the molecular formula
12. **Types of Redox Reactions**
    - **Combination Reactions:** Chemical reactions in which two or more substances (elements or compounds) combine to form a single substance
    - **Decomposition Reactions:** Chemical reactions in which a compound break up into two or more simple substances
    - **Displacement Reactions:** Reaction in which one ion(or atom)in a compound is replaced by an ion(or atom) of other element
13. **Metal Displacement Reactions:** Reactions in which a metal in a compound is displaced by another metal in the uncombined state
14. **Non-metal Displacement Reactions:** Such reactions are mainly hydrogen displacement or oxygen displacement reactions
    - **Disproportionation Reactions:** Reactions in which an element in one oxidation state is simultaneously oxidized and reduced
15. Steps involved in balancing a Redox reaction by oxidation number method
    - Write the skeletal redox reaction for all reactants and products of the reaction
    - Indicate the oxidation number of all the atoms in each compound above the symbol of element
    - Identify the element/elements which undergo change in oxidation numbers
    - Calculate the increase or decrease in oxidation number per atom

- Equate the increase in oxidation number with decrease in oxidation number on the reactant side by multiplying formula of oxidizing agent and reducing agents with suitable coefficients
- Balance the equation with respect to all other atoms except hydrogen and oxygen
- Finally balance hydrogen and oxygen. For balancing oxygen atoms add water molecules to the side deficient in it. Balancing of hydrogen atoms depend upon the medium

16. For reactions taking place in acidic solutions add H+ ions to the side deficient in hydrogen atoms
17. For reactions taking place in basic solutions add H2O molecules to the side deficient in hydrogen atoms and simultaneously add equal number of OH^- ions on the other side of the equation
    - Finally balance the equation by cancelling common species present on both sides of the equation
18. Steps involved in balancing a Redox by Ion-Electron Method (Half reaction method)
    - Find the elements whose oxidation numbers are changed. Identify the substance that acts as an oxidizing agent and reducing agent
    - Separate the complete equation into oxidation half reaction and reduction half reaction
    - Balance the half equations by following steps
19. Balance all atoms other than H and O
20. Calculate the oxidation number on both sides of equation. Add electrons to whichever side is necessary to make up the difference
21. Balance the half equation so that both sides get the same charge
22. Add water molecules to complete the balancing of the equation
    - Add the two balanced equations. Multiply one or both half equations by suitable numbers so that on adding two equations the electrons are balanced
23. **Application of Redox reactions:** Redox Titrations
    - Potassium permanganate in redox reactions: Potassium permanganate ($KMnO_4$) is very strong oxidizing agent and is used in determination of many reducing agents like Fe^{2+}, oxalate ions etc. It acts as self indicator in redox reactions.
      Equation showing $KMnO_4$ as an oxidising agent in acidic medium is:

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

- Acidified Potassium dichromate ($K_2Cr_2O_7$) in redox reactions: $K_2Cr_2O_7$ is used as an oxidizing agent in redox reactions. Titrations involving $K_2Cr_2O_7$ uses diphenylamine and potassium ferricyanide (external indicator).
  Equation showing $K_2Cr_2O_7$ as an oxidising agent in acidic medium is:

  $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
- Iodine (I_2) in redox reactions: I_2 acts as mild oxidising agent in solution according to equation

  $I_2 + 2e^- \rightarrow 2I^-$

24. **Direct redox reaction:** Redox reactions in which reduction and oxidation occurs in same solution (i.e. same reaction vessel). In these reactions transference of electrons is limited to very small distance.
25. **Indirect redox reactions:** Redox reactions in which oxidation and reduction reactions take place in different reactions vessels and thus transfer of electrons from one species to another does not take place directly
    - Electrochemical cell is a device that converts chemical energy produced in a redox reaction into electrical energy. These cells are also called Galvanic cells or Voltaic cells
    - The electrode at which oxidation occurs is called anode and is negatively charged
    - The electrode at which reduction takes place is called cathode and is positively charged
26. In an electrochemical cell the transfer of electrons takes place from anode to cathode
27. In an electrochemical cell the flow of current is from cathode to anode
28. In the electrochemical cell, the electrical circuit is completed with a salt bridge. Salt bridge also maintains the electrical neutrality of the two half cells
29. A salt bridge is a U shaped tube filled with solution of inert electrolyte like sodium chloride or sodium sulphate which will not interfere in the redox reaction. The ions are set in a gel or agar agar so that only ions flow when inverted
30. Electrical potential difference developed between the metal and its solution is called electrode potential. It can also be defined as tendency of an electrode in a half cell to gain or lose electrons

31. Oxidation potential is the tendency of an electrode to lose electrons or to get oxidized
32. Reduction potential is the tendency of an electrode to gain electrons or get reduced
33. In an electrochemical cell, by the present convention, the electrode potentials are represented as reduction potential
34. The electrode having a higher reduction potential will have a higher tendency to gain electrons
35. By convention, the standard electrode potential of hydrogen electrode is 0.00 volts
36. A redox couple is defined as having together oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction
37. The difference between the electrode potentials of eth two electrodes constituting the electrochemical cell is called EMF (Electromotive force) or the cell potential

    EMF = $E^{\ominus}$ cathode - $E^{\ominus}$ anode
38. A negative $E^{\ominus}$ means that the redox couple is a stronger reducing agent than the H^+/H_2 couple
39. A positive $E^{\ominus}$ means that the redox couple is a weaker reducing agent than the H^+/H_2 couple