BRILLIANT PUBLIC SCHOOL, SITAMARHI

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X Chemistry Chapter Notes

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Chemistry
Ch.1: Chemical reactions and equations

Top definitions:

1. Chemical reaction: A chemical reaction involves a chemical change in which substances react to form new substances with entirely new properties. Substances that react or take part in the reaction are known as reactants and the substances formed are known as products.
2. Physical change: If a change involves change in colour or state but no new substance is formed, then it is a physical change.
3. Chemical change: If a change involves formation of new substances, it is a chemical change.
4. Chemical equation: The symbolic representation of a chemical reaction is called a chemical equation.
5. Exothermic and endothermic reactions: If heat is evolved during a reaction, then such a reaction is known as exothermic reaction. If heat is absorbed from the surroundings, then such a reaction is known as endothermic reaction
6. Combination reaction: Combination reaction is a reaction in which 2 or more substances combine to give a single product.
7. Decomposition reaction: In a decomposition reaction, a single reactant decomposes to give 2 or more products. Decomposition reactions require energy in the form of heat, light or electricity
8. Displacement reaction: A reaction in which a more active element displaces less active element from its salt solution.
9. Reactivity series: The Reactivity series is a list of metals arranged in the order of decreasing reactivity. The most reactive metal is placed at the top and the least reactive metal is placed at the bottom.
10. Double displacement reaction: A chemical reaction in which there is an exchange of ions between the reactants to give new substances is called a double displacement reaction.
11. Precipitation reaction: An insoluble solid known as precipitate is formed during a double displacement reaction. Such reactions are also known as precipitation reactions.
12. Redox reaction: A reaction, in which oxidation and reduction takes place simultaneously is known as redox reaction.
13. Oxidation: Oxidation is a chemical process in which a substance gains oxygen or loses hydrogen.
14. Reduction: Reduction is a chemical process in which a substance gains hydrogen or loses oxygen.

Key learnings:

1. A chemical reaction involves a chemical change in which substances react to form new substances with entirely new properties.

Substances that react or take part in the reaction are known as reactants and the substances formed are known as products**.**

2. During a chemical reaction, there is a breaking of bonds between atoms of the reacting molecules to give products.
3. A chemical reaction can be observed with the help of any of the following observations:

a) Evolution of a gas
b) Change in temperature
c) Formation of a precipitate
d) Change in colour
e) Change of state

4. Physical change: If a change involves change in colour or state but no new substance is formed, then it is a physical change.
5. Chemical change: If a change involves formation of new substances, it is a chemical change.
6. Exothermic and endothermic reactions: If heat is evolved during a reaction, then such a reaction is known as Exothermic reaction. If heat is absorbed from the surroundings, then such a reaction is known as endothermic reaction.
7. Chemical equation: The symbolic representation of a chemical reaction is called a chemical equation.
8. Features of a chemical equation:

a. The reactants are written on the left hand side with a plus sign between them.
b. The products are written on the right hand side with a plus sign between them.
c. An arrow separates the reactants from the products. The arrow head points towards the products and indicates the direction of the reaction.

9. Skeletal chemical equation: A chemical equation which simply represents the symbols and formulae of reactants and products taking part in the reaction is known as skeletal chemical equation for a reaction. For example: For the burning of Magnesium in the air, $Mg + O_2 \rightarrow MgO$ is the skeletal equation.
10. Balanced chemical equation: A balanced equation is a chemical equation in which number of atoms of each element is equal on both sides of the equation i.e. number of atoms of an element on reactant side = number of atoms of that element on the product side.
11. As per the law of conservation of mass, the total mass of the elements present in the products of a chemical reaction is equal to the total mass of the elements present in the reactants.
12. The process of equating the number of atoms on both the sides of a chemical equation is known as balancing of a chemical equation.

a. The first step in balancing a chemical equation is to write the number of atoms of each element present on the left hand side and right hand side.
b. We should always start balancing with the compound that contains maximum number of atoms. It can be reactant or a product. Then in that compound select the element which has the maximum number of atoms.
c. While balancing a chemical equation, the molecular formulae of the reactants and products should not change. The molecular formulae are simply multiplied by suitable coefficients.
d. To make a chemical equation more informative, the reaction conditions such as temperature, pressure or catalyst are written on the arrow separating the reactants and products.
e. The evolution of gas is indicated by an upward arrow.
f. The formation of precipitate is indicated by a downward arrow.
g. Heat evolved during the reaction is written as + Heat on the product side.
h. Heat absorbed during the reaction is written as + Heat on the reactant side.

13. Types of chemical reactions:
    a. Combination reaction
    b. Decomposition reaction
    c. Displacement reaction
    d. Redox reaction
14. Combination reaction is a reaction in which 2 or more substances combine to give a single product. Combination reaction can be between 2 elements, between an element and a compound or between 2 compounds.
15. Decomposition reaction: In a decomposition reaction, a single reactant decomposes to give 2 or more products. Decomposition reactions require energy in the form of heat, light or electricity
16. Types of decomposition reactions:
    a. Decomposition reactions which require heat are known as thermolytic decomposition reactions
    b. Decomposition reactions which require light are known as photolytic decomposition reactions
    c. Decomposition reactions which require electricity are known as electrolytic decomposition reactions
17. Displacement reaction: A reaction in which a more active element displaces less active element from its salt solution.

18. The reactivity series is a list of metals arranged in the order of decreasing reactivity. The most reactive metal is placed at the top and the least reactive metal is placed at the bottom.
19. Double displacement reaction: A chemical reaction in which there is an exchange of ions between the reactants to give new substances is called double displacement reaction.
20. Precipitation reaction: An insoluble solid known as precipitate is formed during a double displacement reaction. Such reactions are also known as precipitation reactions.
21. Redox reaction: A reaction in which oxidation and reduction take place simultaneously in a reaction, is known as a redox reaction.
22. Oxidation is a chemical process in which a substance gains oxygen or loses hydrogen.
23. Reduction is a chemical process in which a substance gains hydrogen or loses oxygen.
24. If a substance gains oxygen or loses hydrogen during a reaction, it is said to be oxidised.
25. If a substance gains hydrogen or loses oxygen during a reaction, it is said to be reduced.
26. A substance that loses oxygen or gains hydrogen is known as an oxidising agent.
27. A substance that loses hydrogen or gains oxygen is known as a reducing agent.
28. An oxidising agent gets reduced whereas a reducing agent gets oxidised.
29. In terms of electronic concept, Oxidation is defined as a loss of electrons while reduction is defined as a gain of electrons.
30. Corrosion is the slow eating up of metals by the action of air and moisture on their surfaces. Corrosion in case of Iron is known as Rusting.
31. Chemically, rust is hydrated ferric oxide ($Fe_2O_3.xH_2O$)
32. Advantages of corrosion: Though corrosion is undesirable, it can be advantageous in case of aluminium which on exposure to air, gets coated with a protective layer of aluminium oxide. This protects the metal underneath from further corrosion and damage.
33. Rancidity: When oils and fats or foods containing oils and fats are exposed to air, they get oxidised due to which the food becomes stale and gives a bad taste or smell. This is called Rancidity.
34. Rancidity can be prevented by:

a) Adding antioxidants i.e. the substances which prevent oxidation
b) Refrigeration
c) Storing the food in air-tight containers

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Chemistry
Chapter 2
Acids, Bases and Salts
Chapter Notes

Top concepts:

1. Definition of acids , bases and salts:

Acids	Bases	Salts
Sour in taste	Bitter in taste & soapy to touch	Acid + Base $\rightarrow$ Salt + Water
Give H^+ or H_3O^+ ions in aqueous solutions	Give OH^- ions in aqueous solutions	

2. On the basis of origin, acids are classified as:
   a. Organic acids: Acids derived from living organisms like plants and animals are called organic acids. They are weak acids and are not harmful for living organisms. For example: citric acid is present in fruits, acetic acid present in vinegar, oxalic acid present in tomato, tartaric acid present in tamarind, lactic acid present in sour milk and curd.
   b. Mineral acids: They are also called inorganic acids. They are dangerous and corrosive. Special precautions have to be taken while handling them. For example: sulphuric acid (H_2SO_4), hydrochloric acid (HCl) etc.

3. On the basis of their strength, acids are classified as:
   a. Strong acids: Strong acids are those acids which completely dissociate into its ions in aqueous solutions. Example: nitric acid (HNO_3) , sulphuric acid(H_2SO_4), hydrochloric acid(HCl)
   b. Weak acids: Weak acids are those acids which do not completely dissociate into its ions in aqueous solutions. For example: carbonic acid (H_2CO_3), acetic acid (CH_3COOH)

4. On the basis of their concentration, acids are classified as:
   a. Dilute acids: Have a low concentration of acids in aqueous solutions.
   b. Concentrated acids: Have a high concentration of acids in aqueous solutions.

5. Alkalies: Water soluble bases are called alkalies. For example: Sodium hydroxide (NaOH), potassium hydroxide(KOH)

6. On the basis of their strength, bases are classified as:

a. Strong bases: Strong bases are those bases which completely dissociate into its ions in aqueous solutions. Example: sodium hydroxide (NaOH), potassium hydroxide (KOH)
b. Weak bases: Weak bases are those bases which do not completely dissociate into its ions in aqueous solutions. For example: ammonium hydroxide (NH_4OH)

7. On the basis of their concentration, bases are classified as:
   a. Dilute bases: Have a low concentration of alkali in aqueous solutions.
   b. Concentrated bases: Have a high concentration of alkali in aqueous solutions.

8. Acids and bases conduct electricity because they produce ions in water. There is a flow of electric current through the solution by ions.

9. Indicators are those chemical substances which behave differently in acidic and basic medium and help in determining the chemical nature of the substance. Acid base indicators indicate the presence of an acid or a base by a change in their colour or smell.

10. Indicators can be natural or synthetic.

11. Olfactory indicators: These are those indicators whose odour changes in acidic or basic medium. Example: onion

12. Onion: Smell of onion diminishes in a base and remains as it is in an acid.

13. Vanilla essence: The odour of vanilla essence disappears when it is added to a base. The odour of vanilla essence persists when it is added to an acid.

14. Turmeric: In acids, yellow colour of turmeric remains yellow. In bases, yellow colour of turmeric turns red.

15. Litmus: Litmus is a natural indicator. Litmus solution is a purple dye which is extracted from lichen. Acids turn blue litmus red. Bases turn red litmus blue. Water is essential for acids and bases to change the colour of litmus paper. Remember that litmus paper will act as an indicator only if either the litmus paper is moist or the acid or base is in the form of aqueous solution. This is because acids and bases release H^+ and OH^- ions respectively in aqueous solutions.

16. Phenolphthalein: Phenolphthalein remains colourless in acids but turn pink in bases.

17. Methyl orange: Methyl orange turns pink in acids and becomes yellow in bases.

18. Reaction of acids and bases with water:

Acids	Bases
Release H^+ or H_3O^+ ions in water	Release OH^- ions in water
$HCl + H_2O \rightarrow H_3O^+ + Cl^-$	$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$

19. Reaction of acids and bases with metals:
    Metals displace hydrogen from the acids and form salt and hydrogen gas. This is a displacement reaction. So, acids react with only those metals which are placed above hydrogen in the reactivity series so that metals can displace hydrogen from acids.
    $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
    Metal Acid Salt Hydrogen gas

    Bases react with some metals to form salt and hydrogen gas.
    $Zn + 2\,NaOH \rightarrow Na_2ZnO_2 + H_2$
    Metal Base Sodium zincate (salt) Hydrogen gas
    Bases do not react with all the metals to form salt and water.

20. Reaction of acids and bases with metal carbonates:
    Acids react with metal carbonate to form salt, water and release carbon dioxide.
    $Na_2CO_3 + 2HCl \rightarrow 2\,NaCl + CO_2 + H_2O$
    Bases do not react with metal carbonates.

21. Reaction of acids and bases with metal bicarbonates:
    Acids react with metal bicarbonate to form salt, water and release carbon dioxide.
    $NaHCO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$
    Bases do not react with metal bicarbonates.

22. Reaction of acids with bases: Neutralisation reaction:
    Acids react with bases to form salt and water.
    $HCl + NaOH \rightarrow NaCl + H_2O$
    or
    $H^+ + OH^- \rightarrow H_2O$

23. Reaction of acids with metallic oxides:

Metallic oxides are basic. Therefore, acids react with metallic oxides to form salt and water.
$HCl + CuO \rightarrow CuCl_2 + H_2O$

24.Reaction of bases with non-metallic oxides: Non – metallic oxides are acidic in nature. Bases react with non- metallic oxides to form salt and water. Example: CO_2
$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$

25.Amphoteric oxides: Oxides which show acidic as well as basic properties. For example: ZnO, Al_2O_3
$HCl + ZnO \rightarrow ZnCl_2 + H_2O$
$Zn + 2\ NaOH \rightarrow Na_2ZnO_2 + H_2O$

26.Neutral oxides: Oxides which are neither acidic nor basic are called neutral oxides. Example: CO

27.pH: It is used to find out the strength of acids and bases i.e., how strong or weak the acid or a base is. p in pH stands for 'potenz' in German. The strength of acids and bases depends on the number of H^+ ions and OH^- ions produced respectively.

28.pH scale: A scale for measuring hydrogen ion concentration in a solution is called pH scale.

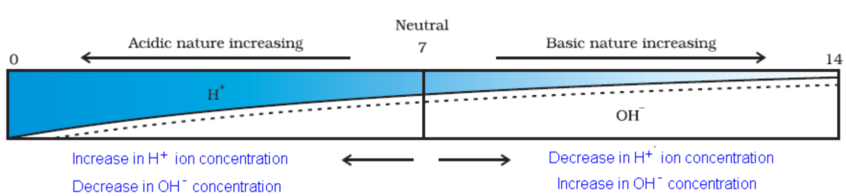
29.On pH scale, we measure pH from 0 to 14.
pH value:

pH	Type of solution
Less than 7	Acidic
Equal to 7	Neutral
More than 7	Basic

30.More the hydrogen ion (or hydronium ion) concentration, lower is the pH value.

31.More the hydroxyl ion concentration, higher is the pH value.

32.Variation in pH:



33. Acids which produce more hydrogen ions are said to be strong acids and acids which produce less hydrogen ions are said to be weak acids. In other words, strong acids have a lower pH value than weak acids.

34. Bases which produce more hydroxyl ions are said to be strong bases and bases which produce less hydroxyl ions are said to be weak bases. In other words, strong bases have a higher pH value than weak bases.

35. Living organisms are pH sensitive. Human body works within a pH range of 7.0 to 7.8.

36. Rain water with a pH less than 5.6 is called acid rain. This acid rain if it flows into river water makes the survival of aquatic life difficult.

37. Plants also require a specific pH range of soil for their healthy growth.

38. pH of our digestive system: Our stomach produces hydrochloric acid for digestion of food. But during indigestion, excess of acid is produced in the stomach and therefore, the pH decreases. This causes pain and irritation. So, to neutralise this excess acid, a mild base is used. This mild base works as an antacid. An antacid is any substance, generally a base or basic salt, which counteracts stomach acidity.

39. Tooth decay: Tooth decay starts when the pH of the mouth is lower than 5.5. Tooth enamel is made up of calcium phosphate which is the hardest substance in the body. It does not dissolve in water, but is corroded when the pH in the mouth is below 5.5. If food particles remain in the mouth after eating, bacteria present in our mouth produce acid by degradation of sugar. This decreases the pH of mouth and hence tooth decay occurs. The best way to prevent this is to clean the mouth after eating food. Using toothpastes, which are generally basic, for cleaning the teeth can neutralise the excess acid and prevent tooth decay.

40. pH is also significant as it is used in self defence by animals and plants. Bees use acids in their sting. To neutralise the effect a mild base like baking soda can be used.

41. Sodium hydroxide (NaOH)
    Preparation:
    Chlor Alkali process:
    In this process, electricity is passed through an aqueous solution of Sodium chloride (called brine). Sodium chloride decomposes to form sodium hydroxide. Chlorine gas is formed at the anode, and hydrogen gas at the cathode. Sodium hydroxide solution is formed near the cathode.
    $2NaCl(aq) + 2\ H_2O\ (l) \rightarrow 2NaOH(aq) + Cl_2(g) + H_2(g)$

42. Bleaching powder: Bleaching powder is represented as $CaOCl_2$, though the actual composition is quite complex.
Preparation: Bleaching powder is produced by the action of chlorine on dry slaked lime.
$Ca(OH)_2 + Cl_2 \rightarrow CaOCl_2 + H_2O$

43. Baking soda: Sodium hydrogen carbonate ($NaHCO_3$)
Preparation:
$NaCl + H_2O + CO_2 + NH_3 \rightarrow NH_4Cl + NaHCO_3$

44. Washing soda: Sodium carbonate $Na_2CO_3 .10H_2O$
Preparation:
In the first step, sodium carbonate is obtained by heating baking soda.
$2\ NaHCO_3 \xrightarrow{\text{Heat}} Na_2CO_3 + H_2O + CO_2$
Then washing soda is produced by recrystallisation of sodium carbonate.
$Na_2CO_3 + 10H_2O \rightarrow Na_2CO_3 .10H_2O$

45. Plaster of Paris: Calcium sulphate hemihydrate $CaSO_4.\ ½\ H_2O$
Preparation: Plaster of Paris is prepared by heating Gypsum at 373K.
$CaSO_4.\ 2H_2O \xrightarrow{\text{Heat at 373K}} CaSO_4.\ ½\ H_2O + 1½\ H_2O$

46. Water of crystallisation: It is the fixed number of water molecules present in one formula unit of a salt.

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Chemistry
Chapter 3
Metals and Non – metals

Top concepts:

1. Definition of Metals & Non metals:

Metals	Non - metals
These are the substances which are electropositive in nature i.e., they have a tendency to lose electrons	These are the substances which are electronegative in nature, i.e. they have a tendency to gain electrons.
They generally have 1, 2, or 3 electrons in their outermost shell.	They generally have 4 to 8 electrons in their outermost shell.

2. Physical properties of Metals and non- metals:

Physical Property	Metals	Non - metals
Physical state	They are generally solids. Exception: Mercury is a liquid	They are either solids or gases Exception: Bromine is a liquid
Lustre	They have a shiny lustre which is called metallic lustre	They do not have a shiny lustre Exception – graphite, iodine
Sonorous	They generally produce a sound on striking a hard surface	They are non - sonorous
Malleability	Some metals can be beaten into thin sheets. Most malleable gold and silver	Non – metals are not malleable
Ductility	The ability of metals to be drawn into thin wires is called ductility Gold is the most ductile metal	Non – metals are not ductile
Electrical conductivity	They are good conductors of electricity Best conductors: Silver	Non- metals are generally poor conductors of electricity

	and copper.	Exception: Graphite
Thermal conductivity	They are good conductors of heat Best conductors: Silver and copper. Poor conductors : Lead and mercury	They are poor conductors of heat
Hardness	They are generally hard Exception – alkali metals like sodium, potassium	They are generally soft Exception : Diamond is the hardest substance known
Melting point	They generally have high melting points Exception – gallium , alkali metals like sodium, potassium Gallium and caesium will melt if you keep them on your palm	They generally have low melting points
Densities	They generally have low densities	They generally have high densities

3. The elements which have intermediate properties between those of metals and non-metals are called metalloids.

4. Allotropes are two or more different forms of the same element.

5. Reaction of metals with oxygen: Almost all metals combine with oxygen to form metal oxides. But all metals do not react with oxygen at the same rate. Different metals show different reactivities towards oxygen.
   Metal oxides are basic in nature. But some metal oxides are amphoteric oxides.

6. Most metal oxides are insoluble in water but some of these dissolve in water to form alkalies. Example:
   Na_2O (s) + H_2O (l) → 2 NaOH (aq)

7. Amphoteric oxides: Metal oxides which show both acidic as well as basic behaviour are known as amphoteric oxides. Such metal oxides which react with both acids as well as bases to produce salts and water. For example: aluminium oxide, zinc oxide, etc.
   $Al_2O_3 + 6\ HCl \rightarrow 2\ AlCl_3 + 3\ H_2O$
   $Al_2O_3 + 2\ NaOH \rightarrow 2\ NaAlO_2 + H_2O$

8. Reaction of metals with water:

Metal	Reacts with	Products
Potassium	Violently with cold water	KOH, H_2
Sodium	Violently with cold water	$NaOH$, H_2
Calcium	Less violently with cold water	$Ca(OH)_2$, H_2
Magnesium	Hot water	$Mg(OH)_2$, H_2
Aluminium	Steam	Al_2O_3, H_2
Iron	Steam	Fe_3O_4, H_2
Zinc	Steam	ZnO, H_2
Lead	No reaction at all	
Silver	No reaction at all	
Gold	No reaction at all	
Copper	No reaction at all	

9. Reaction of metals with acids:
   Metal + Dilute acid $\rightarrow$ Salt + Hydrogen

10. Reaction of metals with nitric acid:
    Hydrogen gas is not evolved when a metal reacts with nitric acid. It is because HNO_3 is a strong oxidising agent. It oxidises the H_2 produced to water and itself gets reduced to any of the nitrogen oxides (N_2O, NO, NO_2). But magnesium (Mg) and manganese (Mn) react with very dilute HNO_3 to evolve H_2 gas.

11. Aqua regia: Aqua regia is a freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric acid in the ratio of 3:1.

12. Anodising: It is a process of forming a thick oxide layer of aluminium. During anodising, a clean aluminium article is made the anode and is electrolysed with dilute sulphuric acid. The oxygen gas evolved at the anode reacts with aluminium to make a thicker protective oxide layer. This aluminium oxide coat makes it resistant to further corrosion.

13. Reaction of metals with solution of other metal salts:

    Metal A + Salt solution of B $\rightarrow$ Salt solution of A + Metal B
    (Metal A is more reactive than metal B)

14. Reactivity series: The reactivity series is a list of metals arranged in the order of their decreasing activities.

K	Potassium	Most reactive
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
Zn	Zinc	Reactivity decreases
Fe	Iron	
Pb	Lead	
H	Hydrogen	
Cu	Copper	
Hg	Mercury	
Ag	Silver	
Au	Gold	Least reactive

15. Reaction of metals with non – metals: When a metal and a non- metal react with each other, transfer of electrons take place from metal to non-metal..

16. Ionic compounds: The compounds thus formed by complete transference of electrons from a metal to non- metal are known as ionic compounds. Ionic compounds have strong electrostatic force of attraction between the positive and negative ions.

17. Properties of ionic compounds:
a. Physical nature: Ionic compounds are solids and are somewhat hard.
b. Melting and Boiling points: Ionic compounds have high melting and boiling points
c. Solubility: Electrovalent compounds are generally soluble in water and insoluble in solvents such as kerosene, petrol, etc.
d. Conduction of Electricity: Ionic compounds do not conduct electricity in the solid state but conduct electricity in the molten state or when dissolved in water.

18. Corrosion: The process of slowly eating away of the metal due to attack of air, water, etc., on the surface of the metal is called corrosion.

19. The rusting of iron can be prevented by painting, oiling, greasing, galvanising, chrome plating, anodising or making alloys.

20. Galvanisation is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The galvanised article is protected against rusting even if the zinc coating is broken.
21. Alloys: An alloy is a homogeneous mixture of two or more metals, or a metal and a non-metal.

22. Alloys and their constituents:

Alloy	Constituents
Brass	Copper, zinc
Bronze	Copper, tin
Steel	Iron, carbon
Stainless steel	Iron, nickel, chromium
Solder	Lead, tin

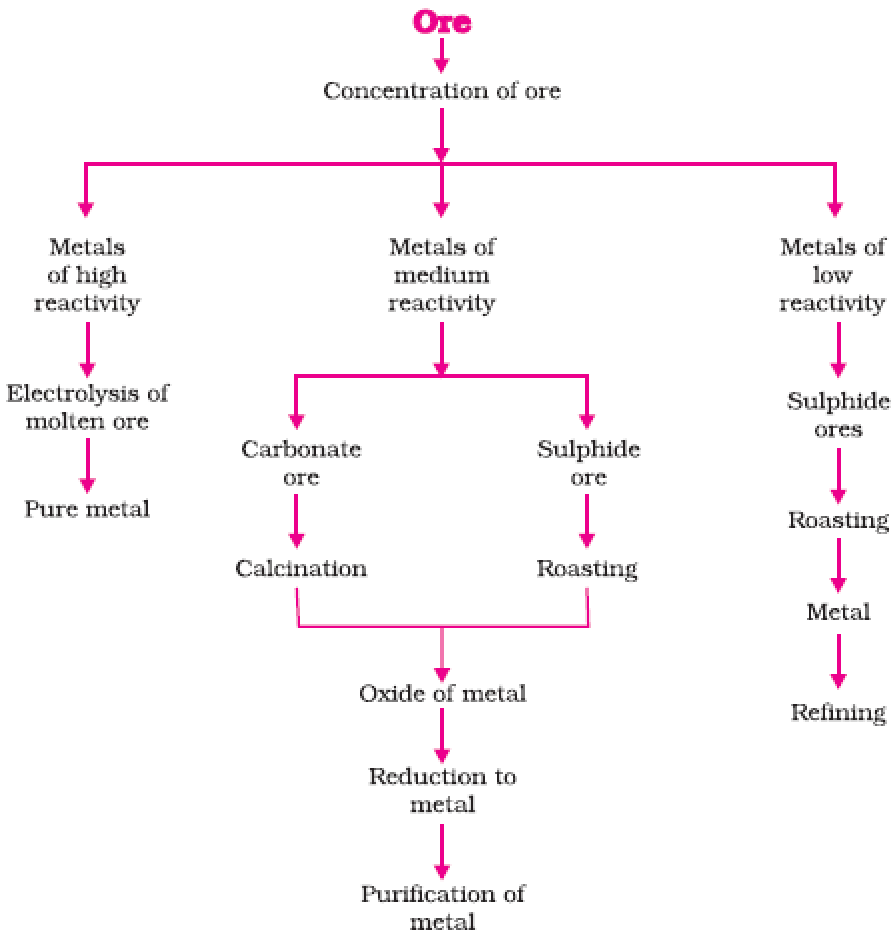
23. Mineral: The elements or compounds, which occur naturally in the earth's crust, are known as minerals.

24. Ore: The minerals contain a very high percentage of a particular metal and from which the metal can be profitably extracted is called ore.

25. Gangue: The unwanted materials or impurities present in the ores is called gangue.

26. Enrichment of ore: Ores mined from the earth are usually contaminated with gangue. The removal of gangue from the ore is called enrichment of ore. The process used for enrichment of ores is based on the differences between the physical or chemical properties of the gangue and the ore.

27. Steps involved in the extraction of metals from ores:



28. Metals low in the activity series: Metals low in the activity series are very unreactive. The oxides of these metals can be reduced to metals by heating alone.

29. Metals in the middle of the activity series: The metals in the middle of the activity series are moderately reactive. These are usually present as sulphides or carbonates in nature. They are first converted to metal oxides and then in the next step the metal oxides are reduced to metal.

31. Conversion of ore into oxide form: It can be done in two ways:

Roasting	**Calcination**
The process of heating the sulphide ore in the presence of sufficient supply of air to convert it into oxide is called roasting.	The process of heating the carbonate ore in the presence of limited supply of air to convert it into oxide is called calcination.
This process is used for sulphide ores.	This process is used for carbonate ores.
Example: $2\ ZnS\ (s) + 3O_2\ (g) \xrightarrow{Heat} 2ZnO\ (s)$	Example: $ZnCO_3\ (s) \xrightarrow{Heat} ZnO\ (s)\ +$

$+2SO_2$ (g)	CO_2 (g)

32. Metals towards the top of the reactivity series: These are highly reactive metals. Example: Sodium, calcium, magnesium, aluminium are obtained by electrolysis of molten chlorides. These metals are obtained by electrolytic reduction.

33. Electrolytic reduction of molten sodium chloride:

At cathode: $Na^+ + e^- \rightarrow Na$

At anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

34. Electrolytic refining of metal:

At anode:

$M \rightarrow M^{n+} + n\ e^-$

(Impure metal)

At cathode:

$M^{n+} + n\ e^- \rightarrow M$

(Pure metal)

X
Chemistry
Chapter 4
Carbon and its compounds

Top concepts:

1. Covalent bond: A covalent bond is a bond formed by sharing of electrons between atoms. In a covalent bond, the shared pair of electrons belongs to the valence shell of both the atoms.

2. Conditions for formation of covalent bond:
    a. The combining atoms should have 4 to 7 electrons in their valence shell.
    b. The combining atoms should not lose electrons easily.
    c. The combining atoms should gain electrons readily.
    d. The difference in electronegativities of two bonded atoms should be low.

3. Properties of covalent compounds:
    a. Physical states: They are generally liquids or gases. Some covalent compounds may exist as solids.
    b. Solubility: They are generally insoluble in water and other polar solvents but soluble in organic solvents like benzene, toluene etc.
    c. Melting and boiling point: They generally have low melting and boiling points.
    d. Electrical conductivity: They do not conduct electrical current.

4. Steps for writing the Lewis dot structures of covalent compounds:
    a. Write the electronic configuration of all the atoms present in the molecule.
    b. Identify how many electrons are needed by each atom to attain noble gas configuration.
    c. Share the electrons between atoms in such a way that all the atoms in a molecule have noble gas configuration.
    d. Keep in mind that the shared electrons are counted in the valence shell of both the atoms sharing it.

5. Electronic configuration of some non- metals:

Name of element	Symbol	Atomic No.	Electrons	Distribution of electrons	Valency	Type of element
Hydrogen	H	1	1	1	1	Non – metal
Carbon	C	6	6	2, 4	4	Non – metal
Nitrogen	N	7	7	2, 5	3	Non – metal
Oxygen	O	8	8	2, 6	2	Non – metal

Fluorine	F	9	9	2, 7	1	Non – metal
Phosphorus	P	15	15	2, 8, 5	3	Non – metal
Sulphur	S	16	16	2, 8, 6	2	Non – metal
Chlorine	Cl	17	17	2, 8, 7	1	Non – metal
Argon	Ar	18	18	2, 8, 8	0	Noble gas

6. Carbon forms covalent bonds.

7. Electronegativity – It is the ability of an atom to attract a shared pair of electrons towards itself.

8. If the atoms forming a covalent bond have different electronegativities, the atom with higher electronegativity pulls the shared pair of electron towards itself. Thus, the atom with the higher electronegativity develops a partial negative charge and the atom with the lower electronegativity develops a partial positive charge. This covalent bond with some polarity is called polar covalent bond.

9. Carbon forms a large number of compounds because of two unique properties:
    a. Tetravalency
    b. Catenation

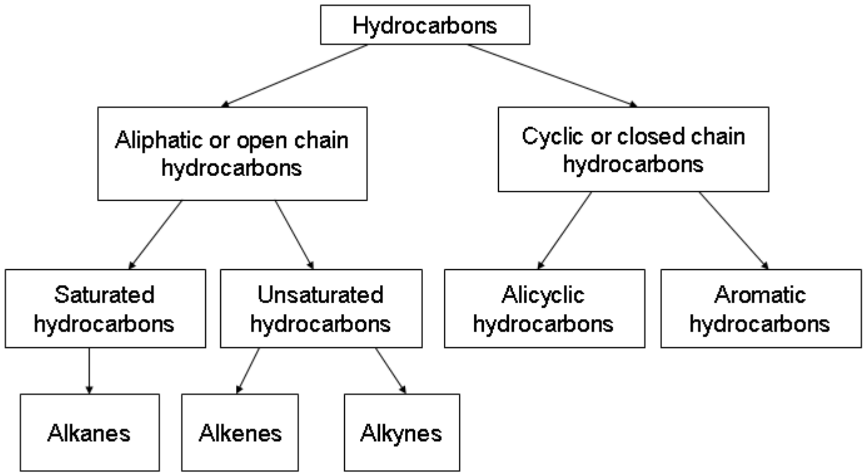
10. Tetravalency of carbon:
    Atomic number = 6
    Electronic configuration: 2, 4
    Valence electrons = 4
    Valency = 4

    So, carbon needs four electrons to attain noble gas configuration. Or in other words, carbon has the ability to form four bonds with carbon or atoms of other mono-valent elements.

11. Catenation: Carbon has the unique ability to form bonds with other atoms of carbon, giving rise to large molecules. This property is called catenation.

12. Steps for writing the Lewis dot structures of Hydrocarbons:
    a. Write the electronic configuration of all the atoms present in the molecule.
    b. Identify how many electrons are needed by each atom to attain noble gas configuration.
    c. First complete the noble gas configuration of all the hydrogen atoms by bonding each hydrogen atom with a carbon atom by a single bond.
    d. The remaining valency of each carbon is completed by forming carbon – carbon single, double or triple bonds.
    e. Keep in mind that the shared electrons are counted in the valence shell of both the atoms sharing it.

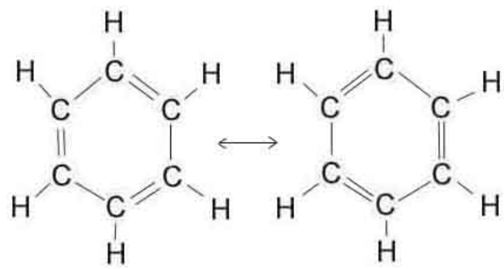
13.Classification of hydrocarbons:



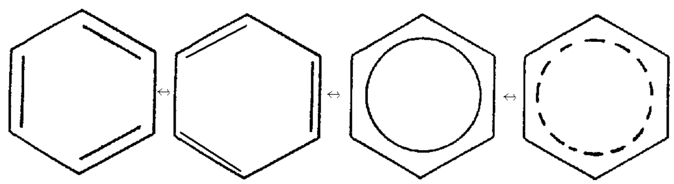
a. Aliphatic or open chain hydrocarbons: These are the carbon compounds which have carbon carbon long open chains. They are classified as:
   i. Saturated hydrocarbons: These hydrocarbons have all carbon – carbon single bonds.
   ii. Unsaturated hydrocarbons: These hydrocarbons have at least one carbon – carbon double or triple bonds.
      - Hydrocarbons with at least one carbon-carbon double bond are called alkenes.
        General formula = C_nH_{2n} where n = number of carbon atoms
      - Hydrocarbons with at least one carbon-carbon triple bond are called alkynes.
        General formula = C_nH_{2n-2} where n = number of carbon atoms

b. Cyclic or closed chain hydrocarbons: These are the hydrocarbons which have carbon carbon closed chain.
   They are classified as:
   i. Alicyclic hydrocarbons: These are the hydrocarbons which do not have benzene ring in their structure.
   ii. Aromatic hydrocarbons: These are the hydrocarbons which have benzene ring in their structure. When hydrogen bonded to carbon of benzene is substituted with halogens, radicals or other functional groups, the derivatives are called aromatic compounds.

14.Benzene: It is an aromatic hydrocarbon which has the molecular formula C_6H_6. It has alternating carbon - carbon single and double bonds.



Benzene can also be represented as:



15. IUPAC name of hydrocarbon consists of two parts:
    a. Word root: Number of carbons in the longest carbon chain

Number of carbon atoms	Word root (Greek name)
1	Meth
2	Eth
3	Prop
4	But
5	Pent
6	Hex
7	Hept
8	Oct
9	Non
10	Dec

    b. Suffix: Depends on the type of carbon – carbon bond: for single bond, suffix is –ane ; for double bond, suffix is –ene, and for triple bond suffix is –yne

16. Steps to write the IUPAC nomenclature of hydrocarbons:
    a. Select the parent carbon chain:
        i. Select the longest carbon chain as the parent chain.
        ii. If a double or a triple bond is present in the carbon chain, it should be included in the parent chain.
    b. Number the parent carbon chain from that carbon end such that the double bond, triple bond or side chain gets the lowest number.

c. Identify and name the side chain if any: -CH_3 is named as methyl, -C_2H_5 is named as ethyl etc. Also identify the position of the side chain.
d. Write the name of the hydrocarbon as:
Position number-name of the side chain word root – Position number- suffix
Example: 2-Methyl but-1-ene
e. Remember if the hydrocarbon is an alkane, the position number of suffix is not written.

17. Types of formula for writing hydrocarbons:
  a. Molecular formula: The actual number of each type of atom present in the compound.
  b. Structural formula: The actual arrangement of atoms is written
  c. Condensed formula: It is the shortened form of the structural formula

18. Conditions for Isomerism:
  a. Only alkanes with more than three carbon atoms can have isomers.
  b. The side chains cannot be present on the terminal carbons.

19. How to write different chain isomers of hydrocarbons:
  a. First draw the different carbon chains keeping in mind the conditions for isomerism.
  b. Complete the tetravalency of carbon by forming single covalent bonds with hydrogens.
  c. In the end, check that the molecular formula of each isomer should be same.

20. How to write different position isomers of unsaturated hydrocarbons:
  a. First draw the different carbon chains keeping in mind the conditions for isomerism.
  b. If it is an alkene draw the first isomer always by drawing a double bond between C_1 and C_2 or if it is an alkyne draw the first isomer always by drawing a triple bond between C_1 and C_2
  c. The next isomers will be dawn by drawing the same chain and changing the positions of the double and triple bonds in alkenes and alkynes respectively.
  d. Complete the tetravalency of carbon by forming single covalent bonds with hydrogens.
  e. In the end, check that the molecular formula of each isomer should be same.

21. Homologous Series: A series of organic compounds in which every succeeding member differs from the previous one by $-CH_2$ or 14 u. The molecular formula of all the members of a homologous series can be derived from a general formula.

22. Properties of a homologous series: As the molecular mass increases in a series, so physical properties of the compounds show a variation, but chemical properties which are determined by a functional group remain the same within a series.

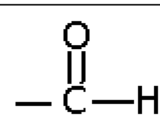
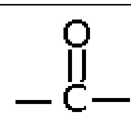
23. Homologous series of alkanes: General formula: C_nH_{2n+2} where n = number of carbon atoms

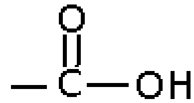
24. Homologous series of alkenes: General formula: C_nH_{2n} where n = number of carbon atoms

25. Homologous series of alkynes: General formula: = C_nH_{2n-2} where n = number of carbon atoms

26. Functional group: An atom or a group of atoms which when present in a compound gives specific properties to it, regardless of the length and nature of the carbon chain is called functional group.
    a. Free valency or valencies of the group are shown by the single line.
    b. The functional group is attached to the carbon chain through this valency by replacing one hydrogen atom or atoms.
    c. Replacement of hydrogen atom by a functional group is always in such a manner that valency of carbon remains satisfied.
    d. The functional group, replacing the hydrogen is also called as hetero atom because it is different from carbon, and can be nitrogen, sulphur, or halogen etc.

27. Some functional groups in carbon compounds:

Heteroatom	**Functional group**	**Formula of the functional group**	**Suffix**
Cl/ Br	Halo- Chloro Bromo	 -Cl -Br	Named as prefix Chloro – Bromo -
Oxygen	Alcohol	-OH	- ol
	Aldehyde		-al
	Ketone		-one

	Carboxylic acid		-oic acid

28. Steps to write the IUPAC name of organic compounds:

i. Select the parent carbon chain:
   1. Select the longest carbon chain as the parent chain.
   2. If a double or a triple bond is present in the carbon chain, it should be included in the parent chain.
   3. If a functional group is present, the carbon chain should include the functional group.

ii. Number the parent carbon chain from that carbon end such that the functional group, double bond, triple bond or side chain gets the lowest number.
Remember here that the aldehyde and carboxylic acid functional group are present on the terminal carbon atom.

iii. Identify the name and position of the functional group, double bond, triple bond or side chain.

iv. The name of the functional group is written with either a prefix or a suffix as given in the above table.

v. If the name of the functional group is to be given as a suffix, the name of the carbon chain is modified by deleting the final 'e' and adding the appropriate suffix. For example, a three-carbon chain with a ketone group would be named in the following manner – Propane – 'e' = propan + 'one' = propanone.

vi. Remember that in the compounds which have carbon containing functional groups, the name of the word root includes the functional group carbon atom also.

vii. If the carbon chain is unsaturated, then the final 'ane' in the name of the carbon chain is substituted by 'ene' or 'yne' as given in the table above. For example, a three-carbon chain with a double bond would be called propene and if it has a triple bond, it would be called propyne.

29. Difference between chemical properties of saturated and unsaturated hydrocarbons:

Saturated hydrocarbons	Unsaturated hydrocarbons
Give a clean blue flame on complete combustion and on incomplete combustion give a yellow sooty flame.	Give a yellow sooty flame on combustion.
Undergo substitution reaction like chlorination	Undergo addition reaction like hydrogenation, addition reaction with bromine in carbon tetrachloride

Are fairly unreactive and inert in the presence of most reagents	Are reactive

30. Catalysts are substances that cause a reaction to occur or proceed at a different rate without the reaction being affected.

31. Oxidizing agents are substances which are capable of providing oxygen to other compounds for their oxygen. Example: Alkaline $KMnO_4$, acidified $K_2Cr_2O_7$ etc.

32. Reactions of ethanol:

$$2\ CH_3CH_2OH + 2\ Na \rightarrow 2\ CH_3CH_2ONa + H_2$$

$$2\ CH_3CH_2OH \xrightarrow{\text{Hot conc. } H_2SO_4} CH_2{=}CH_2 + H_2O$$

$$2\ CH_3CH_2OH \xrightarrow[\text{Or acidified } K_2Cr_2O_7 + \text{Heat}]{\text{Alkaline } KMnO_4 + \text{Heat}} CH_3COOH$$

33. Reactions of ethanoic acid:

$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$

$$CH_3COOH + CH_3CH_2OH \xrightarrow{\text{acid}} \underset{\text{ester}}{CH_3COOCH_2CH_3}$$

$$CH_3COOC_2H_5 + NaOH \longrightarrow CH_3COONa + C_2H_5OH$$

$$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$$

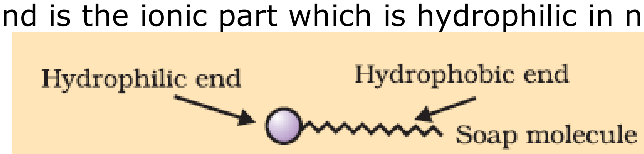
$$2CH_3COOH + Na_2CO_3 \longrightarrow 2CH_3COONa + H_2O + CO_2$$

$$CH_3COOH + NaHCO_3 \longrightarrow CH_3COONa + H_2O + CO_2$$

34. Catalysts are substances that cause a reaction to occur or proceed at a different rate without the reaction being affected.

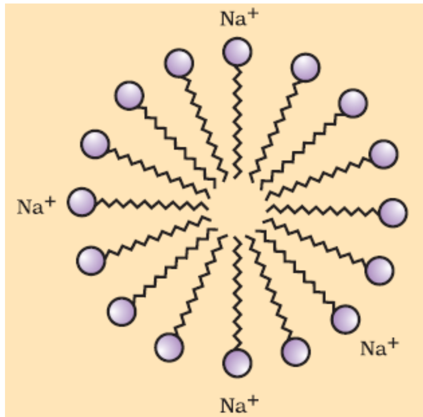
35. Soaps are sodium or potassium salts of long chain carboxylic acids.

36. Structure of soap molecule: The structure of soap molecule consists of a long hydrocarbon tail at one end which is hydrophobic in nature. The other end is the ionic part which is hydrophilic in nature.



37. Cleansing action of soap: When soap is at the surface of water, the ionic end of soap orients itself towards water and the hydrocarbon 'tail' orients itself aligns itself along the dirt. Thus, clusters of molecules are

formed in which the hydrophobic tails are in the interior of the cluster and the ionic ends are on the surface of the cluster. This formation is called a micelle.



Soap in the form of a micelle is able to clean, since the oily dirt will be collected in the centre of the micelle. The micelles stay in solution as a colloid and will not come together to precipitate because of ion-ion repulsion. Now, when water is agitated, the dirt suspended in the micelles is also easily rinsed away.

38. When hard water is treated with soap, scum is formed. This is caused by the reaction of soap with the calcium and magnesium salts, which cause the hardness of water.

39. Detergents are generally ammonium or sulphonate salts of long chain carboxylic acids.

40. Detergents do not form scum with hard water. This is because the charged ends of these compounds do not form insoluble precipitates with the calcium and magnesium ions in hard water. Thus, they remain effective in hard water.

41.

Soaps	Detergents
Soaps are sodium or potassium salts of long chain carboxylic acids.	Detergents are generally ammonium or sulphonate salts of long chain carboxylic acids.
Soaps are not effective for cleaning in hard water.	Detergents are effective for cleaning in hard as well as soft water.
Soaps are biodegradable.	Detergents are non - biodegradable.

Chapter notes

Class: X
Chapter Name: Periodic classification of elements

Top concepts

1. Dobereiner's triads: Johann Wolfgang Dobereiner, a German chemist, classified the known elements in groups of three elements on the basis of similarities in their properties. These groups were called triads.

- Characteristic of Dobereiner's Triads:

a. Properties of elements in each triad were similar.

b. Atomic mass of the middle elements was roughly the average of the atomic masses of the other two elements.

- Example of Dobereiner's Triads :

Element	Atomic mass	Element	Atomic mass	Element	Atomic mass
Lithium (Li)	6.9	Calcium (Ca)	40.1	Chlorine (Cl)	35.5
Sodium (Na)	23.0	Strontium (Sr)	87.6	Bromine (Br)	79.9
Potassium (K)	39.0	Barium (Ba)	137.3	Iodine (I)	126.9

- Limitations: Dobereiner could identify only three triads. He was not able to prepare triads of all the known elements

2. Newlands' Law of Octaves: John Newlands, an English scientist, arranged the known elements in the order of increasing atomic masses and called it the 'Law of Octaves'. It is known as 'Newlands' Law of Octaves'

- Characteristics of Newlands' Law of Octaves

a. It contained the elements from hydrogen to thorium

b. Properties of every eighth element were similar to that of the first element

- Table showing Newlands' Octaves:

sa (do)	re (re)	ga (mi)	ma (fa)	pa (so)	da (la)	ni (ti)
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

Co and Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce and La	Zr	-	-

- Limitations of Newlands' law of Octaves:

a. The law was applicable to elements upto calcium (Ca) only
b. It contained only 56 elements. Further it was assumed by Newlands that only 56 elements existed in nature and no more elements would be discovered in the future.

c. In order to fit elements into the table. Newlands' adjusted two elements in the same slot and also put some unlike elements under same note. For example cobalt and nickel are in the same slot and these are placed in the same column as fluorine, chlorine and bromine which have very different properties than these elements. Iron, which resembles cobalt and nickel in properties, has been placed differently away from these elements

3. Mendeleev's Periodic Table: Dmitri Ivanovich Mendeleev, a Russian chemist, was the most important contributor to the early development of a periodic table of elements wherein the elements were arranged on the basis of their atomic mass and chemical properties..

- Characteristic of Mendeleev's Period Table:

a. Mendeleev arranged all the 63 known elements in an increasing order of their atomic masses.

b. The table contained vertical columns called 'groups' and horizontal rows called 'periods'.

c. The elements with similar physical and chemical properties came under same groups.

- Mendeleev's Periodic Law: The properties of elements are the periodic function of their atomic masses.

- Achievements of Mendeleev's Periodic Table:

a. Through this table, it was very easy to study the physical and chemical properties of various elements.

b. Mendeleev adjusted few elements with a slightly greater atomic mass before the elements with slightly lower atomic mass, so that elements with similar properties could be grouped together. For example, aluminum appeared before silicon, cobalt appeared before nickel.

c. Mendeleev left some gaps in his periodic table. He predicted the existence of some elements that had not been discovered at that time. His predictions were quite true as elements like scandium; gallium and germanium were discovered later

d. The gases like helium, neon and argon, which were discovered later, were placed in a new group without disturbing the existing order

- Limitations of Mendeleev's Periodic Table :

a. He could not assign a correct position to hydrogen in the periodic table

b. Positions of isotopes of all elements was not certain according to Mendeleev's periodic table
c. Atomic masses did not increase in a regular manner in going from one element to the next. So it was not possible to predict how many elements could be discovered between two elements

4. Modern Periodic Table: Henry Moseley gave a new property of elements, 'atomic number' and this was adopted as the basis of Modern Periodic Table'.

- Modern Periodic Law: Properties of elements are a periodic function of their atomic number

- Position of elements in Modern Periodic Table:

a. The modern periodic table consists of 18 groups and 7 periods

b. Elements present in any one group have the same number of valence electrons. Also, the number of shells increases as we go down the group.

c. Elements present in any one period, contain the same number of shells. Also, with increase in atomic number by one unit on moving from left to right, the valence shell electron increases by one unit

d. Each period marks a new electronic shell getting filled

- Trends in the Modern Periodic Table:

(i) Valency: Valency of an element is determined by the number of valence electrons present in the outermost shell of its atom

- Valency of elements in a particular group is same

- Valency of elements in a particular period increases by one unit from left to right with the increase in atomic number by one unit

(ii) Atomic Size: Atomic size refers to the radius of an atom.

- In a period, atomic size and radii decreases from left to right. This is due to increase in nuclear charge which tends to pull the electrons closer to nucleus and reduces size of atom

- In a group, atomic size and radii increases from top to bottom. This is because on moving down, new shells are added. This increases distance between outermost electrons and nucleus which increases the size of atom

(iii) Metallic and Non- metallic Properties:

- The tendency to lose electrons from the outermost shell of an atom, is called metallic character of an element

- Metallic character decreases across a period and increases down the group

- The tendency to gain electron in the outermost shell of an atom, is called non- metallic character of an element

- Non-metallic character increases across a period and decreases down the group

- Elements intermediate between metal and non-metals that show characteristic of both metals and non-metals are called as semi-metals or metalloids

5. Metals have a tendency to loose electrons while forming bond. Hence they are electropositive in nature

6. Non-metals have a tendency to gain electrons while forming bond. Hence they are electronegative in nature

7. Oxides formed by metals are generally basic and oxides formed by non-metals are generally acidic