Studybeesofficial

Where Minds Pollinate

Class 11
Chemistry

Chapter 10
Periodic Table
Handwritten Notes

Ch 10: PERIODIC TABLE

Periodic Table

A table which shows the systematic arrangement of elements is called periodic table.

-Periodic Law

If the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.

Periods

The horizontal rows of the periodic table are called periods. There are 7 periods. Properties of elements within a period change gradually as we move from left to right. When we move from one period to next period, the pattern of properties repeats.

-Number of elements per period

Period I has two elements. Period 2 and 3 has eight elements. Period 4 and 5 has eighteen elements. Period 6 and 7 has thirty-two elements. The first three periods are called short periods and rest are called long periods.

Groups

The vertical columns of the periodic table are known as groups. Elements having similar properties lie in the same group. Elements with similar valence shell electronic configuration are placed in the same group. There are two types of group i.e. group A and group B. Group A elements are called as normal or representative elements while Group B elements are called transition elements.

-Group number and electronic configuration

For normal group.

Group No.	No. of Electrons in Valence Shell	Valence Shell electronic configuration
1	1	ns^1
2	2	ns^2
13	3	ns^2, np^1
14	4	ns^2, np^2
15	5	ns^2, np^3
16	6	ns^2, np^4
17	7	ns^2, np^5
18	8	ns^2, np^6

-Modern representation of groups

The international Union of Pure and Applied Chemistry (IUPAC) decided that the groups would be 1-18 from left to right.

- Name of Some groups

Group IA are called alkali metals.

Group IIA are called alkaline earth metals.

Group VIA are called chalcogens

Group VIIA are called halogens

Group VIIIA are called nobel gases

Blocks of Elements

- S Block

The elements present in Groups 1 and 2 and Hellium (Group 18) are called s-block elements. They contain their valence electrons in s sub-shells. Alkali and alkaline earth metals are present in s-block.

- p block

The elements present in group 13 to 18 are called p-block elements. They contain their valence electrons in p-sub shell. This block contains both metals and non-metals but majority are non-metals.

- d block

Elements present in Groups 3 to 12 (or sub-group B) are called d-block elements. They contain their valence electrons in d-sub shell. These elements are also called outer transition elements.

- f block

Two rows of elements located at the bottom of the periodic table consist of f-block elements. They contain their valence electron in f-sub-shell. These elements are also called inner transition metals and also classified as lanthanides and actinides.

Charateristic Properties of elements in a Group

- Atomic radius gradually increases from top to bottom.
- Ionization energy decreases from top to bottom in a group.
- Electronegativity decreases down the group.
- Metallic character increases down the group.
- Elements in the same group generally have similar chemical reactivity because they have similar electronic configuration in their valence shell.

Metal, Non-Metals and Metalloides

The elements are generally classified into three categoreies on the basis of their properties and position in periodic table.

-Metals

- Left hand side of periodic table.
- Good conductor, Malleable and Ductile.

- Non-Metals

- Right hand side of period table.
- Poor conductor, Non-malleable, Non-ductile.

-Metalloids

- A ladder type separation b/w metals and non metals.
- Properties of both metals and non-metals.

Chemical Periodicity / Periodicity of Properties

A chemical periodicity refers to the repeating pattern of the properties of elements in periodic table. Electronic configuration of elements shows a periodic fluctuation with increasing atomic number. Therefore, the physical and chemical properties of the elements vary in a periodic manner. Elements within the same group generally have similar chemical properties because they have similar electronic configuration. Physical properties depend on the size of atom. Therefore, the elements show a gradation of physical properties within the same group.

-Electron Affinity

Electron Affinity explains the anion formation. The amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uni-negative gaseous ion.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)} + \text{electron affinity.}$$

• Factors affecting electron affinity

- Nuclear charge: Electron affinity tends to increase from left to right due to increase in nuclear charge. This is mainly due to the increase in nuclear charge. As the no. of protons in nucleus increases, the attraction to electrons also increases, making it eaeir for the atom to accept additional electron.
- Atomic size: Electron affinity generally decreases from top to bottom. This is due to the increase in atomic size. Larger atoms have electrons further from nucleus and the attraction b/w the nucleus and outer electrons is weaker. Therefore it is more difficult for larger atoms to accept extra electrons.

Sheilding effect: Sheilding effect can lead to a decrease in electron affinity.

- **Trend in periodic table**

  The electron affinity decreases down the group. The electronic affinity increases in a period as we move left to right.

Trends in Metalic and Non-Metalic Behaviour

- Metallic Behaviour

Metals tend to easily lose valence electrons and form cations to achieve stable electron configurations. Metals usually have 1-3 valence electrons. Their ionization energy is usually low. Therefore, metals can easily lose electrons. Metals have low electronegativety and do not tend to attract electrons.

- Non-Metallic Behaviour

Non-metallic behaviour is due to their tendency to gain or ~~ath~~ share valence electrons to achieve a stable structure. Non-metals usually have 5-7 valence electrons. They have high ionization energy. This makes them less likely to loose electrons. They also have high electronegativity, which allows them to easily gain or share electrons to form covelent bonds.

- Trends in Group

- **Group 1**

  Alkali metals have strong metallic character. As you move down the group, the metallicity of the group 1 elements increases. This is due to their atomic radii and ability to lose electrons.

- **Group 17**

  Group 17 halogens have a strong non-metallic character. The non-metallic behaviour decreases. This is due to the increase in atomic radii and decrease in their ability to attract electrons.

- **Group 14**

  In group 14, which consists of carbon (C), silicon (Si), gemanium (Ge), tin (Sn) and lead (Pb), the non-metallic character decreases and metalic character increases down the group.

- Electronegativity and type of Bond

The electronnegativity difference (ΔEN) of the bond b/w two bonded atoms give a rough indication of the expected nature of the bond.

$\Delta EN > 1.8$ Bond is ionic

$\Delta EN = 0.4 - 1.8$ Bond is polar covelent

$\Delta EN < 0.4$ Bond is non-polar covelent.

- Trend in Chemical Properties

Reactions of typical metals of group 1 and 2, such as sodium and magnesium.

- **Reaction with Oxygen**

Sodium is a silvery white soft metal. It is an extremely reactive element. It readily combines with oxygen, chlorine and water. In a limited supply of oxygen, sodium burns to form sodium oxide. ($2Na + 1/2 O_2 \longrightarrow Na_2O$). In excess of oxygen, it forms pale yellow sodium peroxide ($2Na + O_2 \longrightarrow Na_2O_2$).

Magnesium is relatively hard and is also an extremely reactive element. When ignited, Mg burns in oxygen with an intense white flame to give solid magnesium oxide. ($2Mg + O_2 \rightarrow 2MgO$).

- **Reaction with Chlorine**

Sodium burns in chlorine with a bright yellow flame, producing a white solid, sodium chloride. ($2Na + Cl_2 \longrightarrow 2NaCl$).

Magnesium burns in chlorine with intense white flame, producing a white solid, magnesium chloride. ($Mg + Cl_2 \longrightarrow MgCl_2$).

- **Reaction with water**

Sodium reacts voilently with cold water producing hydrogen gas. The reaction occurs with a light explosion. ($2Na + 2H_2O \longrightarrow 2NaOH + H_2$).

Magnesium slowly reacts with cold water to form a thin layer of magnesium hydroxide on its surface, which forms a proptective coating that stops furthur reaction.

In steam, magnesium burns with a white flame, forming magnesium oxide and hydrogen. ($Mg + 2H_2O \longrightarrow MgO + 2H_2$)

Reaction of Chlorides

- Ionic Chlorides

Metal chlorides, such as sodium chloride and magnesium chloride are ionic in nature.

- **Solubility in water** Ionic chlorides dissolve in water forming a neutral solution. These solutions have pH = 7.

- Covelent Chlorides

Aluminium chloride ($AlCl_3$), silicon tetra chloride ($SiCl_4$) and phosphorous penta chloride (PCl_5) are covelent compounds.

- **Solubility in water** They are soluble in water, but they react vigoursly with water forming acidic solution. They have low pH.

- Reactions

$AlCl_3 + 3H_2O \xrightarrow[pH=3]{} Al(OH)_3 + 3HCl$ $\quad$ $SiCl_4 + 4H_2O \xrightarrow[pH=0]{} Si(OH)_4 + 4HCl$

Ionization Energy

The minimum amount of energy required to remove the outer most electron from an isolated gaseous atom.

-Factors effecting I.E

- Nuclear force of attraction ↑ I.E ↑
- Sheilding effect ↑ I.E ↓
- Atomic size ↑ I.E ↓

- Trends in periodic table

The ionization energy decreases in any given main group. as we move down. Cause of electron shells increase, leading to increase in sheilding effect.