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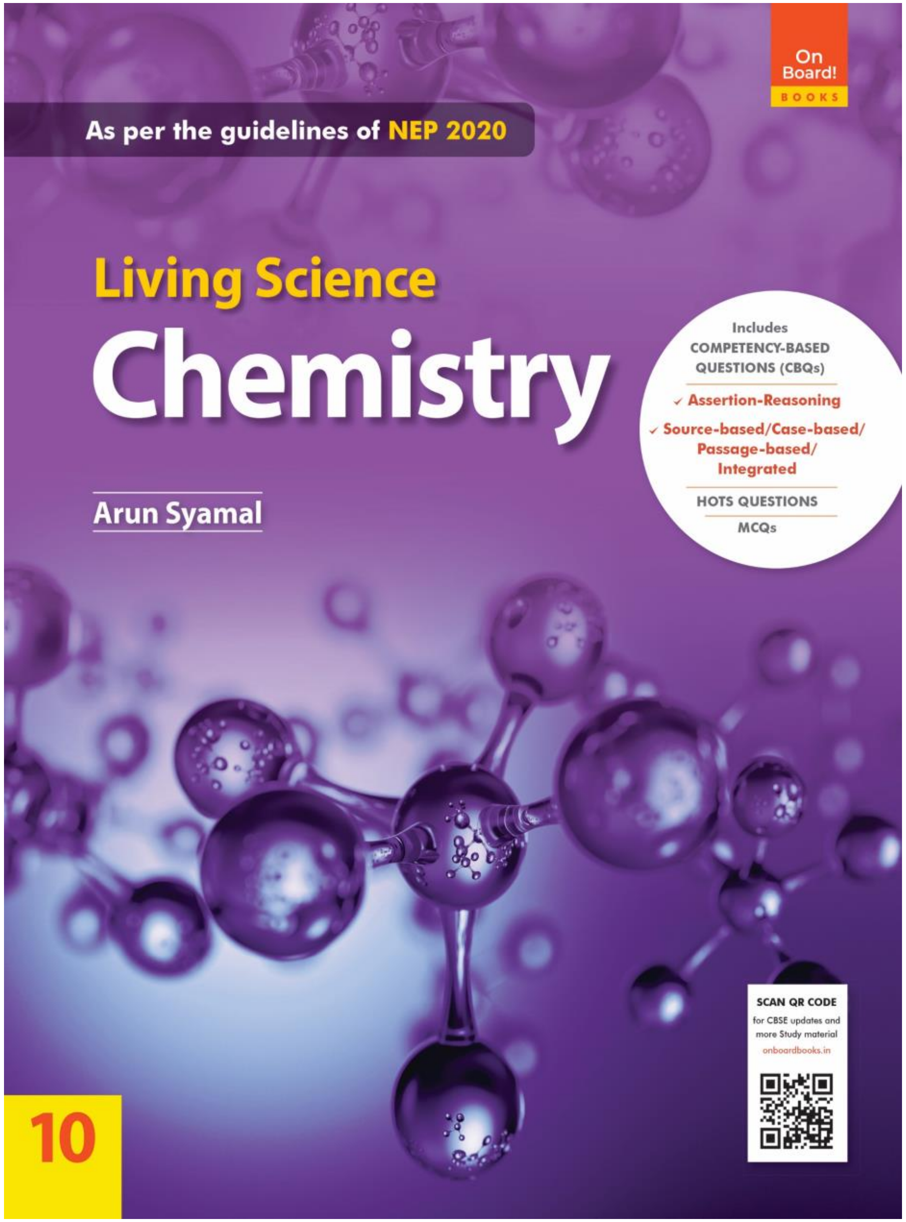

CBSE LIVING SCIENCE CHEMISTRY

CLASS 10

Chapter 5

Periodic Classification of Elements

Learning Objectives

- ❖ Dobereiner's triads
- ❖ Newlands' law of octaves
- ❖ Mendeleev's periodic law
- ❖ Mendeleev's Periodic Table
- ❖ Modern periodic law
- ❖ Modern Periodic Table
- ❖ Properties of elements
  - a. in a group
  - b. in a period
- ❖ Electronic configuration
- ❖ Valency
- ❖ Metallic and non-metallic characters
- ❖ Electropositive nature
- ❖ Atomic radius
- ❖ Electronegative nature
- ❖ Electron affinity
- ❖ Ionisation energy
- ❖ Electronegativity

Till today 118 elements are known to us. It is difficult to study and remember the properties of each element. To overcome this problem, scientists grouped the elements of similar properties together so that if the properties of one of them is known, those of the others could be correlated. This is known as periodic classification of elements.

The periodic classification of elements has given us a systematic and immensely useful framework for correlating the properties of the elements into a few very simple and logical patterns. In this chapter, we shall study the historical development of the Modern Periodic Table in chronological order and explain the basis of periodic classification of elements.

DOBEREINER'S TRIADS

In 1817, the German chemist John Dobereiner noticed that the elements could be arranged in the order of increasing atomic mass in the groups of three in which the atomic mass of the middle element was approximately the arithmetic mean of the atomic masses of the other two elements. The middle element had properties in between those of the other two. Such groups of elements were called **Dobereiner's triads**. For example, chlorine, bromine and iodine exhibit similar chemical properties and form a triad.

Element	Chlorine	Bromine	Iodine
Atomic mass	35.5	80.0	127.0

Atomic mass of the middle element Br

$$\approx \text{Average of atomic masses of Cl and I}$$

$$= \frac{35.5 + 127.0}{2}$$

$$= 81.25$$

Limitations of Dobereiner's Triads

- It was observed that Dobereiner's classification was not applicable to all known elements but was limited to only a few elements. Dobereiner could identify only three triads from the elements known at that time. Hence, this system of classification was not found to be useful.

Dobereiner's triads

Element	Atomic mass	Element	Atomic mass	Element	Atomic mass
Li	7	Ca	40	Cl	35.5
Na	23	Sr	88	Br	80
K	39	Ba	137	I	127

- Although elements like nitrogen, phosphorus and arsenic exhibit similar chemical properties, they do not constitute a Dobereiner's triad. This is because the actual mass of the middle element phosphorus (31.0) is much lower than the average (44.45) of the atomic masses of nitrogen (14.0) and arsenic (74.9). Hence, nitrogen, phosphorus and arsenic do not constitute a Dobereiner's triad inspite of their similar chemical properties.

NEWLANDS' LAW OF OCTAVES

An English chemist, John Newlands in 1864 arranged the elements in order of increasing atomic mass and found that the properties of every eighth element are similar to the properties of the first element. Thus, he called it the 'Law of Octaves' which states that **when elements are arranged in the order of increasing atomic masses, the physical and chemical properties of every eighth element are a repetition of the first element.**

He divided the elements in the order of increasing atomic masses into horizontal rows of seven elements as shown in table below. It can be seen in the table that the eighth element starting from lithium is sodium which exhibits similarity in physical and chemical properties with lithium. Likewise, the eighth element starting from sodium is potassium which exhibits similarity in physical and chemical properties with sodium. Similarly, the eighth element starting from fluorine is chlorine which exhibits similarity in physical and chemical properties with fluorine.

Newlands' arrangement of elements

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	–	–

Drawbacks of Law of Octaves

1. The law of octaves worked well only for lighter elements and could not be extended to elements of atomic masses higher than that of calcium since after calcium every eighth element did not exhibit properties similar to that of the first.

2. Newlands assumed that only 56 elements existed in nature with no scope for discovery of new elements in future. Subsequently, quite a large number of new elements were discovered (noble gases), the properties of which did not fit into Newlands' law of octaves.
3. Newlands assigned one position to two elements in order to fit the elements in his periodic table. For example, only one position each was assigned to (Co and Ni) and (Ce and La).
4. Newlands placed Co and Ni in the same column as F, Cl and Br having quite different properties in comparison to Co and Ni.
5. Although Fe resembles Co and Ni in properties, it has been placed far away form these elements. Therefore, Newlands' Octave Law became irrelevant and worked well for lighter elements only.

MENDELEEV'S PERIODIC LAW

A breakthrough in the classification of elements came with the work of the Russian chemist Dmitri Ivanovich Mendeleev (1834–1907) who stated in 1869 the law of chemical periodicity. The law stated that **"the physical and chemical properties of elements are periodic functions of their atomic masses."**
Mendeleev's Periodic Table contained only 63 elements. Mendeleev arranged the then known elements in the order of increasing atomic masses in horizontal rows and vertical columns in such a way that the elements with similar properties became the members of the same column. The similar properties used by Mendeleev to classify elements into groups were the similar formulae of their oxides and hydrides.

This arrangement of the elements in rows and columns is known as the Periodic Table. The Periodic Table proposed by Mendeleev in 1872 is shown in the table below.

Mendeleev's periodic table of 1872 (based on atomic mass)

Group	I	II	III	IV	V	VI	VII	VIII
Oxide* Hydride+	R_2O RH	RO RH_2	R_2O_3 RH_3	RO_2 RH_4	R_2O_5 RH_3	RO_3 RH_2	R_2O_7 RH	RO_4
Periods ↓	A B	A B	A B	A B	A B	A B	A B	Transition Series
1	H 1.008							
2	Li 6.939	Be 9.012	B 10.81	C 12.011	N 14.007	O 15.999	F 18.998	
3	Na 22.99	Mg 24.31	Al 29.98	Si 28.09	P 30.974	S 32.06	Cl 35.453	
4 First series	K 39.102	Ca 40.08	* 44.96	Ti 47.90	V 50.94	Cr 50.20	Mn 54.94	Fe Co Ni 55.85 58.93 58.71
Second series	Cu 63.54	Zn 65.37	* 69.72	* 72.59	As 74.92	Se 78.96	Br 79.90	
5 First series	Rb 85.47	Sr 87.62	Y 88.91	Zr 91.22	Nb 92.91	Mo 95.94	* 99	Ru Rh Pd 101.07 102.91 106.4
Second series	Ag 107.87	Cd 112.40	In 114.82	Sn 118.69	Sb 121.75	Te 127.60	I 126.90	
6 First series	Cs 132.90	Ba 137.34	La 138.91	Hf 178.49	Ta 180.95	W 183.85		Os Ir Pt 190.2 192.2 195.09
Second series	Au 196.97	Hg 200.59	TI 204.37	Pb 207.19	Bi 208.48			

*Elements not discovered in 1872, but Mendeleev kept gap for unknown elements which were discovered subsequently.
+The letter R was used in the formula of oxides and hydrides to represent any of the elements in the group.

Merits of Mendeleev’s Periodic Table

Mendeleev’s Periodic Table was one of the outstanding contributions in the development of chemistry. It has many merits and serves many purposes. Some of the merits of Mendeleev’s Periodic Table are discussed as follows:

1. **Systematic Classification of Elements:** Mendeleev’s Periodic Table systematised the study of chemistry of elements and their compounds. Chemically similar elements are placed in a group and by knowing the property of one element in a group, the properties of other elements in the group can be predicted.
2. **Classification on the Basis of Fundamental Atomic Property:** Mendeleev’s classification of elements is based on atomic mass of elements. Atomic mass is a better fundamental property and classification based on atomic masses is a better method than classifying elements by considering atomic volume, triads, octaves, etc.
3. **Correction of Atomic Masses:** Mendeleev’s Periodic Table predicted errors in the atomic masses based on their positions in the Periodic Table and these were corrected. For example, atomic mass of berrylium (Be) was corrected from 13.9 to 9.0. The atomic masses of a few other elements such as In, Au and Pt, etc. were also corrected.
4. **Prediction of Undiscovered Elements:** Only 63 elements were known at the time of Mendeleev’s classification of elements. He kept many vacant spaces in the Periodic Table for accommodating new elements to be discovered. He even predicted the properties of these unknown elements. Subsequently, these elements were discovered and their properties were found to be similar to those predicted by Mendeleev. For example, he kept three vacant spaces below B and Al in Group III and Si in Group IV.

He named them as *Eka*-boron, *Eka*-aluminium and *Eka*-silicon, respectively. These elements were discovered subsequently and Mendeleev's prediction about the properties of these elements and their compounds was found to be precisely correct.

5. **Accomodation of Noble Gases:** Noble gases like helium, neon and argon were discovered very late because they are inert or unreactive and present in extremely low concentrations in our atmosphere. However, when it was discovered, it could be easily accommodated in Mendeleev's periodic table in a new group without disturbing the existing order.

Defects of Mendeleev's Periodic Table

Although Mendeleev's Periodic Table has many merits, it has several defects also. These defects are listed as follows:

1. **Discrepancy in periodicity:** The arrangement of elements in the order of increasing atomic masses is not maintained in the following cases and the cause of periodicity put forward by Mendeleev is not tenable.

Pair 1		**Pair 2**	
Co	Ni	Te	I
58.93	58.71	127.6	126.9

In these pairs, the elements with higher atomic masses have been placed before elements of lower atomic masses. Mendeleev's periodic law of atomic masses cannot explain these anomalies of reverse order of arrangement of elements. These anomalies occurred due to placing these elements in their correct group on the basis of their chemical properties.

2. **Position of Hydrogen:** Mendeleev placed hydrogen in Group I A. But hydrogen exhibits similarity in properties of both alkali metals and halogens. Hence, the position of hydrogen in the Mendeleev's Periodic Table is not correctly defined.

3. **Position of Isotopes:** Isotopes of the same element have similar chemical properties and different atomic masses and hence they should be placed in different groups according to the Mendeleev's periodic law based on atomic masses. But Mendeleev did not provide separate places to the isotopes of an element.

4. **Placement of Elements with Different Properties in the Same Group:** Elements with different properties have been placed in the same group. For example, the inactive coinage metals (Cu, Ag, Au) have been placed with most active metals (Li, Na, K, etc.) in Group I.

5. **Irregular Trend in Atomic Masses of Elements:** The atomic mass of elements does not increase in a regular manner on moving from one element to the next element. Due to this, it was not possible to predict the number of elements which could be discovered between two elements, specially in the case of heavier elements.

MODERN PERIODIC LAW

In 1913, Henry Moseley discovered that atomic number and not the atomic mass, is the most fundamental property which characterises an element. Consequently, Moseley modified Mendeleev's periodic law to the modern periodic law which states that **the physical and chemical properties of elements are periodic functions of their atomic numbers.**

The anomalies in the Mendeleev's Periodic Table disappear when the elements are arranged in the order of increasing atomic numbers.
Interpretations of defects of Mendeleev's periodic table on the basis of modern periodic law are given as follows:

1. **Explanation for the Positions of Some Elements:** The wrong position of certain elements like cobalt and nickel in Mendeleev's table is justified in Modern Periodic Table because atomic number of cobalt (27) is less than that of nickel (28).
2. **Explanation for the Position of Isotopes:** Since the isotopes of an element have the same number of protons, their atomic number is also the same. Hence, in spite of having different atomic masses, the isotopes of an element are assigned the same position in the Periodic Table.
3. **Position of Hydrogen:** A unique position has been assigned to hydrogen in Modern Periodic Table. It is placed at the top left corner because of its unique characteristics.

The International Union of Pure and Applied Chemistry (IUPAC) has adopted a new version of Modern Periodic Table known as **Long Form of Periodic Table**.

The Long Form of Periodic Table was constructed by Bohr (1920). The classification of elements in the Long Form of Periodic Table is based on electronic configuration of elements. The chemical periodicity of elements depends upon electronic configuration of elements.

Modern Periodic Table (Long form of the periodic table)

Atomic number → 6
Symbol → C
Atomic mass → 12.011

GROUP NUMBER

PERIOD NUMBER / IUPAC (1984)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.0079																	2 **He** 4.0026
2	3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 16.0	9 **F** 19.0	10 **Ne** 20.179
3	11 **Na** 23.0	12 **Mg** 24.305											13 **Al** 27.0	14 **Si** 28.0855	15 **P** 31.0	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.08	21 **Sc** 45.0	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 52.0	25 **Mn** 54.9380	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.9216	34 **Se** 79.0	35 **Br** 79.909	36 **Kr** 83.80
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.22	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (99)	44 **Ru** 101.07	45 **Rh** 102.305	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.40	49 **In** 114.82	50 **Sn** 118.69	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.904	54 **Xe** 131.29
6	55 **Cs** 132.9054	56 **Ba** 137.34	57 **La*** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.206	76 **Os** 190.2	77 **Ir** 192.22	78 **Pt** 195.09	79 **Au** 197.0	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.19	83 **Bi** 208.9804	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** 226.0254	89 **Ac**** 227.028	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (269)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (278)	110 **Ds** (281)	111 **Rg** (282)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (290)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (295)

*LANTHANOIDES	58 **Ce** 140.12	59 **Pr** 140.907	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 152.0	64 **Gd** 157.25	65 **Tb** 158.92	66 **Dy** 162.50	67 **Ho** 164.9304	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 175.5
**ACTINOIDES	90 **Th** 232.0381	91 **Pa** 231.0359	92 **U** 238.0289	93 **Np** 237.0482	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (245)	98 **Cf** (251)	99 **Es** (254)	100 **Fm** (253)	101 **Md** (256)	102 **No** (254)	103 **Lr** (257)

Description of Modern Periodic Table (Long Form of Periodic Table)

Periods

The horizontal rows of elements in the Periodic Table are called periods. There are seven periods in the Modern Periodic Table. The number of elements in each period is as follows:

Period 1 is the shortest period with only two elements (H, He).

Period 2 is a short period with eight elements (Li, Be, B, C, N, O, F, Ne).

Period 3 is a short period with eight elements (Na, Mg, Al, Si, P, S, Cl, Ar).

Period 4 is a long period with eighteen elements.

Period 5 is a long period with eighteen elements.

Period 6 is a long period with thirty-two elements.

Period 7 is incomplete and is a long-period with thirty-two elements.

The number of elements accommodated in a period depends upon the maximum number of electrons which can be accommodated in various shells (*K*, *L*, *M*, *N*, etc.) of the atoms. The maximum number of electrons which can be accommodated in an energy shell (orbit) is given by $2n^2$ where *n* stands for the number of the orbit.

Period 1 can accommodate two elements because the first electron shell (K shell) has the capacity to accommodate only two electrons ($n = 1$, $2n^2 = 2 \times (1)^2 = 2$).

Period 2 can accommodate eight elements because the second electron shell (L shell) has the capacity to accommodate eight electrons [$n = 2$, $2n^2 = 2 \times (2)^2 = 8$].

Period 3 can theoretically contain eighteen elements [n = 3, $2n^2 = 2 \times (3)^2 = 18$], yet only eight elements are accommodated since the outermost shell (*M*) cannot hold more than eight electrons.

Period 4 can accommodate eighteen elements. Additional ten transition elements are present in Period 4.

Period 5 can accommodate eighteen elements. Additional ten transition elements are present in Period 5.

Period 6 can accommodate thirty-two elements. This period starts with alkali metal Cs and ends with Rn. This period accommodates additional ten transition elements [La (57), Hf (72) to Hg (80)] and fourteen inner transition elements [Ce (58) to Lu (71)]. The fourteen inner transition elements are placed at the bottom of the Periodic Table in a separate block. The inclusion of these fourteen elements in a separate block is justified in order to maintain recurrence of periodicity in properties. The recurrence of periodicity occurs after thirty-two elements in the periods 6 and 7.

Period 7 can accommodate thirty-two elements. This period also contains fourteen inner transition elements and ten transition elements.

Groups

The vertical columns of elements in the Periodic Table are called groups. There are eighteen groups in the Modern Periodic Table and these groups are numbered as 1, 2, 3, … 18 from left to right across the Periodic Table.

Valence Electrons

The electrons present in the outermost shell of an atom of an element are called its valence electrons. They are so called because the electrons in the outermost shell determine the valency of an element.
Elements containing one valence electron are placed in group 1. Elements containing two valence electrons are placed in group 2.
The group number of elements having more than two valence electrons can be calculated as:
Group number of element = valence electrons of element + 10
For example, elements containing 3 valence electrons are placed in Group 13.
Elements containing 4 valence electrons are placed in Group 14. Elements containing 5 valence electrons are placed in Group 15.
Elements containing 6 valence electrons are placed in Group 16. Elements containing 7 valence electrons are placed in Group 17.
Elements containing 8 valence electrons are placed in Group 18.

Trends in the Modern Periodic Table

Valency

Valency is the combining capacity of an atom of an element to acquire the inert gas configuration. It depends upon the number of valence electrons.

- **Down a group:** In a group, outer electronic configuration is the same for all the elements. So all the elements in a group have same valence electrons and valency.
- **Along a period:** As one moves from left to right in each short period (period number 2 and 3), the valency of elements increases from 1 to 4 and then decreases to zero in the last elements of the period.

Atomic Radius

The atomic radius can be visualized as the distance from the centre of the nucleus to the outermost shell of electrons in an atom.

- **Variation of atomic radius in a group:** The atomic radii of elements increase with increase in atomic number on moving from top to bottom in a group. On moving down a group, a new shell of electrons is added in each succeeding element. Hence, the electrons in the valence shell of each succeeding element lie farther and farther away from the nucleus. Consequently, the attraction of the nucleus for the electrons decreases and as a result, the atomic radius increases with increase in atomic number. But the nuclear charge of elements also increases. The increase in the atomic radius due to the addition of new shell is so large that it outweighs the contractive effect of the increased nuclear charge. As a result, the atomic radius increases on moving down a group from top to bottom.

- **Variation of atomic radius in a period:** The atomic radii of elements decrease with increase in atomic number on moving from left to right in a period. On moving from left to right along a period, the nuclear charge increases by one unit in each succeeding element, while the added electron enters in the same shell. As a result, the electrons of all the shells are strongly pulled towards the nucleus. Hence, there is a steady contraction in atomic radius from left to right in a period. Thus, in any period, the alkali metal atom such as Li, Na, K, etc., has the largest atomic radius while the halogen atom has the smallest atomic radius. The atomic radius suddenly increases for the noble gas atom due to the inter-electronic repulsion within a shell.

Ionisation Energy

In an atom, the electrons are held by electrostatic force of attraction to the positively charged nucleus. In order to remove an electron from an atom, energy must be supplied. The amount of energy required to remove the most loosely held electron from an isolated gaseous atom is called **ionisation energy**.

- **Variation of ionisation energy (IE) in a group:** On moving down a group, the IE of elements decreases with increase in atomic number. On moving down a group, the atomic radius increases due to the addition of a new electron shell at each succeeding element and the nuclear charge increases with increase in atomic number. Hence, the electrostatic force of attraction between the nucleus and electrons increases and accordingly the IE should increase.

But the effect of the increase in the atomic radius more than compensates the effect of the increased nuclear charge in increasing the IE. Hence, the IE decreases on moving down a group.

- **Variation of ionisation energy (IE) in a period:** In any period in the Periodic Table, the ionisation energies of elements increase from left to right. This is due to the combined effect of the increased nuclear charge and decrease of atomic radius with increase in atomic number.

Electron Affinity

Electron affinity (EA) of an element is the energy released when an isolated gaseous atom accepts an electron to form the gaseous negative ion. Thus, a quantitative measure of the tendency of an element to gain electron is given by electron affinity.

- **Variation of electron affinity in a group:** The electron affinity of elements decreases on moving down a group. There is some exception: electron affinities of N, O and F (second period) are lower than those of P, S and Cl (third period), respectively. The elements of second period have the smallest atomic radius among the elements in their respective groups. Due to the smaller atomic radius, there occurs considerable electron–electron repulsion within the atom. As a result, the added electron is not accepted with the same ease like the other elements in the group. Thus, the electron affinity of nitrogen (-0.1 kJ mol^{-1}) is lower than that of phosphorus (-72 kJ mol^{-1}). The electron affinity of oxygen (-141 kJ mol^{-1}) is lower than that of sulphur (-200 kJ mol^{-1}).

- **Variation of electron affinity in a period:** The electron affinity of elements increases on moving from left to right in a period but not in graded manner. On moving in a period from left to right, the atomic radius decreases and nuclear charge increases. Both these factors increase the attraction of added electron by the nucleus. Hence, electron affinity increases in a period from left to right.

Metallic and Non-metallic Properties

The tendency of an element to lose electrons and form positive ions (cations) is called as electropositive character. Since, metals exhibit such a tendency, the electropositive character is also known as metallic character. Hence, metals are electropositive in nature. The tendency of an element to gain electrons and form negative ions (anions) is called as electronegative character. Since, non-metals have the tendency to accept electrons and form anions, they are electronegative in nature.

- **Variation of metallic and non-metallic character in a group:** On moving down a group, the metallic character of elements increases. This is due to the increase in atomic radius on moving down the group. Since the distance between the nucleus and outermost electrons increases on moving down a group, the force of attraction between nucleus and electrons decreases and less energy is required to lose electrons. The non-metallic character decreases on moving down a group since the tendency to gain electrons decreases on moving down a group.

- **Variation of metallic and non-metallic character in a period:** On moving from left to right in a period, the metallic character decreases. As the effective nuclear charge acting on the valence shell electrons increases in a period, the tendency to lose electrons decreases. The non-metallic character increases across a period from left to right since the tendency to gain electrons increases across a period from left to right.

Electronegativity

The tendency of an atom in a molecule to attract a shared pair of electrons towards itself is called as electronegativity.

- **Variation of electronegativity in a group:** Electronegativity of elements decreases on moving down a group. Due to the increase in size with increase in atomic number, the tendency of an atom to attract the shared pair of electrons decreases. Hence, electronegativity decreases on moving down a group.

- **Variation of electronegativity in a period:** Electronegativity of elements increases across a period. On moving from left to right in a period, the atomic radius decreases and the nuclear charge increases. Due to the combined effect of decreasing atomic radius and increasing nuclear charge, the shared pair of electrons is more strongly attracted towards the nucleus. As a result, the electronegativity increases in a period from left to right.

SUMMARY

1. The elements are grouped together on the basis of similarities in their properties.
2. **Dobereiner's triads:** Dobereiner's triad consists of a set of three elements (triad) in the order of increasing atomic mass in which the atomic mass of the middle element is approximately the arithmetic mean of the atomic masses of the other two elements.
3. **Newlands' law of octaves:** When elements are arranged in the order of increasing atomic masses, the similarities in physical and chemical properties reappear after each interval of eight elements just like the eighth musical note resembles the first.
4. **Mendeleev's periodic law:** The properties of the elements are periodic functions of their atomic masses.
5. **Modern periodic law:** The physical and chemical properties of elements are periodic functions of their atomic numbers.
6. **Long Form of Periodic Table:** It consists of seven periods and eighteen groups. The groups are numbered as 1, 2, 3, ... 18 from left to right across the Periodic Table. The earlier notation of designating sub-groups A and B is dropped.
7. **Representative elements:** The elements of the Groups 1, 2, 13–18 are called **representative elements.**
8. **Periodicity:** The repetition of similar properties after regular intervals is called **periodicity**. The periodicity of chemical properties of elements depends upon the periodicity of electronic configuration of elements.
9. **Valence electrons:** The electrons present in the outermost shell of an atom of an element are called **valence electrons.**
10. On moving down a group, the number of electron orbits increases by one at each successive element while the number of electrons in the outermost shell remains the same.
11. On moving from left to right in a period, the electrons are added to the same orbit.
12. The chemical properties of the elements depend on the number of valence electrons.
13. The elements of the same group show the same valency in their compounds.
14. On moving from left to right in each short period, the valency of elements increases from 1 to 4 and then decreases to 0 in the last element of the period.
15. **Atomic properties:** The physical characteristics of the atom of an element are called **atomic properties**. The physical properties such as atomic radius, ionisation energy, electron affinity, electronegativity, density, valency, M.P., B.P. are called **atomic properties**. The elements exhibit periodicity in atomic properties.

16. **Atomic radius:** The distance from the centre of the nucleus to the outermost shell of electrons of atom of an element is called **atomic radius**. On moving down a group, the atomic radii of elements increase in a regular fashion. On moving from left to right in a period, the atomic radii of elements decrease. The first element of any period has the largest atomic radius and the last element of any period has the smallest atomic radius.

17. **Metallic or electropositive character:** The tendency of an element to lose electrons and form positive ion is called **metallic** or **electropositive character**. On moving down a group, the metallic character of elements increases. On moving from left to right in a period, the metallic character decreases and non-metallic character increases. The reactivity of metals increases with increase in metallic character.

18. **Electronegative character:** The tendency of a non-metal to gain electrons and form negative ion is called **electronegative character**. The reactivity of non-metals increases with increase in electronegative character.

19. **Ionisation energy:** The amount of energy required to remove the most loosely held electron from an isolated gaseous atom is called **ionisation energy**. The magnitude of ionisation energy depends upon atomic radius and nuclear charge of element. On moving down a group, the ionisation energies of elements decrease with increase in atomic number. The ionisation energies of elements increase in a period on moving from left to right. The elements having completely filled valence shell have higher ionisation energy. Metals have comparatively lower ionisation energy than non-metals.

20. **Electron affinity:** Electron affinity of an element is the energy released when an isolated gaseous atom accepts an electron to form gaseous negative ion. Electron affinity of elements decreases on moving from top to bottom in a group. Electron affinity of elements increases on moving from left to right in a period. Halogens have the highest electron affinity among all the elements. Electron affinity trends in the Periodic Table are parallel to those of ionisation energies.

21. **Electronegativity:** Electronegativity of an element is the tendency of an atom in a molecule to attract a shared pair of electrons towards itself. Electronegativity of elements decreases on moving down a group. Electronegativity of elements increases in a period on moving from left to right. The most electronegative element is fluorine and the least electronegative element is francium.

MIND MAP

PERIODIC CLASSIFICATION OF ELEMENTS

DOBEREINER'S TRIADS

- three elements form a triad written in the order of increasing atomic masses
- atomic mass of the middle element is roughly the average of the atomic mass of other two elements

NEWLANDS' LAW OF OCTAVES

- elements are arranged in the order of increasing atomic masses
- physical and chemical properties of every 8th element are a repetition of the first element

MENDELEEV'S PERIODIC LAW

- physical and chemical properties of elements are periodic functions of their atomic masses

MODERN PERIODIC TABLE

- Modern Periodic Law : Physical and chemical properties of elements are periodic functions of their atomic numbers.
- 7 periods and 18 groups

PERIODIC TRENDS