

PERIODIC TABLE FOR CLASS XI CHEMISTRY

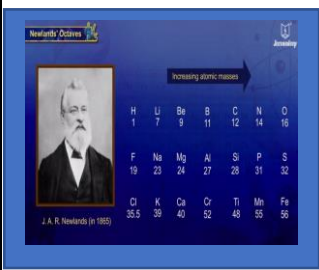
01	Total elements	=118
02	Total metals	=94
03	Total non-metals	=11
04	Total metalloids	=7
05	Total noble gas	=6

The necessity of Classification of Elements

With the discovery of a large number of elements, it became difficult to study the elements individually, so classification of elements was done to make the study easier.

Earlier Attempts to Classify Elements

Many attempts were made to classify the known elements from time to time. The earlier attempts are as follows:

(1815)	Prout's Hypothesis 	According to this theory, hydrogen atom was considered as the fundamental unit from which all other atoms were made. It is also known as unitary theory.								
(1829)	Dobereiner's Triads 	Dobereiner classified the elements into groups of three elements with similar properties in such a manner so that the atomic weight of the middle element was the arithmetic mean of the other two. For example: <table border="1"><tr><td>Elements</td><td>Li</td><td>Na</td><td>K</td></tr><tr><td>Atomic wt.</td><td>7</td><td>23</td><td>39</td></tr></table> Mean of atomic masses = $(7 + 39) / 2 = 23$. Similarly, Cl,Br,I; Ca,Sr,Ba are two more examples of such triads. Limitations Dobereiner could not arrange all the elements known at that time into triads. He could identify only three such triads that have been mentioned.	Elements	Li	Na	K	Atomic wt.	7	23	39
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(1864)	Newland's Octaves 	Newland states that when elements are arranged in order of increasing atomic masses, every eighth element has properties similar to the first just like in the musical note [Every eighth musical note is the same as the first mentioned note]. Limitations 1. This classification was successful up to the element calcium. 2. When noble gas elements were discovered at a later stage, their inclusion in these octaves disturbed the entire arrangement.								
(1869)	Lothar Meyer's Atomic Volume 	Meyer presented the classification of elements in the form of a curve between atomic volume and atomic masses and state that the properties of the elements are the periodic functions of their atomic volumes. [Here, atomic volume = molecular mass / density] He concluded that the elements with similar properties occupy similar position in the curve.								

	Mendeleev's Periodic Table 	Mendeleev's Periodic Table is based upon Mendeleev's Periodic law which states 'The physical and chemical properties of the elements are a periodic function of their atomic masses.' Only 63 elements were known during his time. This Periodic Table is divided into seven horizontal rows (periods) and eight vertical columns (groups). Zero group was added later on in the modified Mendeleev's Periodic Table.
(1913)	Modern Periodic Table 	Moseley modified Mendeleev's periodic law. He stated "Physical and chemical properties of elements are the periodic function of their atomic numbers." It is known as modern periodic law and considered as the basis of Modern Periodic Table. When the elements were arranged in increasing order of atomic numbers, it was observed that the properties of elements were repeated after certain regular intervals : 2, 8, 8, 18, 18 and 32. These numbers are called magic numbers and cause of periodicity in properties due to repetition of similar electronic configuration.

Importance of Mendeleev Periodic table:

1. Systematic study of the elements.
2. Prediction of new elements and their properties.
3. He left space for the elements yet to be discovered. Hence helped in discovering more elements after him.
4. He left spaces for Ga and Ge and named these elements as **Eka-aluminium (Ga)** and **Eka-silicon (Ge)** respectively.
5. Atomic mass correction of doubtful elements on the basis of their expected positions and properties.

Defects in the Mendeleev's Periodic Table

1. **Position of Hydrogen** has been placed in group IA (alkali metals), but it also resembles with halogens of group VIIA. Thus, its position in the Mendeleev's Periodic Table is controversial.
2. **Position of isotopes** As Mendeleev's classification is based on atomic weight, Isotopes would have to be placed in different positions due to their different atomic weights, e.g., 1H^1 1H^2 1H^3 would occupy different positions.
3. **Anomalous positions of some elements** Without any proper justification. In some cases, the element with higher atomic mass precedes the element with lower atomic mass. For example, Al (atomic weight = 39.9) precedes K (atomic weight = 39.1) and similarly Co (atomic weight = 58.9) has been placed ahead of Ni (atomic weight = 58.7).
4. **Position of Lanthanoids and actinoids** were not given position in the main Periodic Table.
5. Elements with similar properties are kept separate - like Gold and Platinum.
6. Elements with dis-similar properties are kept together- like Li, Na, K and Cu, Ag and Au

Structural Features of Long Form of Periodic Table

PERIOD	ELEMENTS	No. of ELEMENTS	REMARKS
1	$1\text{H} - 2\text{He}$	2	Shortest Period
2	$3\text{Li} - 10\text{Ne}$	8	Short Period
3	$11\text{Na} - 18\text{Ar}$	8	Short Period
4	$19\text{K} - 36\text{Kr}$	18	Long Period
5	$37\text{Rb} - 54\text{Xe}$	18	Long Period
6	$55\text{Cs} - 86\text{Rn}$	32	Longest Period
7	with 87Fr	19+	Incomplete Period

GROUP	ELEMENTS	GROUP NAME	REMARKS
1	7	Alkali Metals	
2	7	Alkaline Earth Metals	
11	4	Coinage Metals	
13	7	Boron Family	
14	7	Carbon Family	
15	7	Pnictogens	
16	7	Chalcogens	Ore forming elements
17	7	Halogens	Sea salt former
18	8	Noble gas	

Anomalous behaviour of the first element of a group.

The first element of a group differs considerably from its congeners

This is due to

- (i) small size
- (ii) high electronegativity and
- (iii) non availability of d-orbitals for bonding. Anomalous behaviour is observed among the second-row elements (i.e., Li to F).

- Elements with atomic number greater than 92 (U^{92}) are called the transuranium elements. All these elements are man-made through artificial nuclear reactions.
- IUPAC approved the name for the element of atomic number 110, as Darmstadtium, with symbol Ds].

Atomic No:	Name of the elements	Symbol
110	Darmstadtium	Ds
113	Nihonium	Nh
115	Moscovium	Mc
117	Tennessine	Ts
118	Oganesson	Og

The 4 main divisions of the Periodic Table

Block	Group No:		Remarks	Configuration
s	Group 1 and Group 2	Last e ⁻ enters to s orbital	Representative Elements	ns^{1-2}
p	Group 13 and Group 18	Last e ⁻ enters to p orbital	Representative Elements	$ns^2 np^{1-6}$
d	Group 3 and Group 12	Last e ⁻ enters to d orbital	Transitional Elements	$(n-1)d^{1-10} ns^{1-2}$
f	Group 3	Last e ⁻ enters to f orbital	Inner transitional Elements	$(n-2)f^{1-14} (n-1)d^{0-1} ns^2$

Limitations of Long Form of Periodic Table

In the long form of the Periodic Table:

1. The position of hydrogen still remains uncertain.
2. The inner-transition elements do not find a place in the main body of the table. They are placed separately.

Predicting the Position of an Element in the Periodic Table

The complete electronic configuration is written first

The principle quantum number of the valence shell represents the period of the element.

The subshell in which the last electron is filled corresponds to the block of the element.

Group of the element is predicted from the electrons present in the

Outermost shell (n) or penultimate (n -1) shell as follows:

Block	Calculation of Group No:
s	number of ns-electrons
p	10 + number of ns and np electrons
d	the sum of the number of (n -1) d and ns electrons.
f	Group No. is 3

IUPAC Nomenclature of Elements With Z > 100

The names are derived directly from the atomic numbers using numerical roots for 0 and numbers from 1-9 and adding the suffix - ium.

Digits	Roots	Abbreviations
0	Nil	n
1	Un	u
2	Bi	b
3	Tri	t
4	Quad	q
5	Pent	p
6	Hex	h
7	Sept	s
8	Oct	o
9	enn	e

Periodic Properties

The properties which are directly or indirectly related to their electronic configuration and show gradual change as we move from left to right in a period or from top to bottom in a group are called periodic properties.

Atomic Radius

It is the distance from the centre of the nucleus to the outermost shell (valence Shell) of electrons.

Covalent radius for an atom A in a molecule A₂

$$r_A = r_A + r_A / 2 = d_{A-A} / 2$$

For heteroatomic molecule AB,

$$d_{A-B} = r_A + r_B + 0.009 (X_A - X_B)$$

where, X_A and X_B are electronegativities of A and B.

In general, the atomic size decreases on moving from left to right in a period due to increase in effective nuclear charge and increases on moving from top to bottom in a group due to addition of new shells.

van der Waals' Radius

It is defined as one-half of the distance between the nuclei of two non – bonded isolated atoms or two adjacent atoms belonging to two neighbouring molecules of an element in the solid state.

Metallic Radius

It is defined as one-half of the distance between the centres of nuclei of the two adjacent atoms in the metallic crystal.

Ionic Radius

An atom can be changed to a cation by transfer of electrons and to an anion by gain of electrons. A cation is always smaller than the parent atom because during its formation effective nuclear charge increases and sometimes a shell may also decrease. On the other hand, the size of an atom is always larger than the parent atom because during its formation effective nuclear charge decreases. In case of isoelectronic ions, the higher the nuclear charge, smaller is the size. e.g., $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^{+} < \text{F}^{-} < \text{O}^{2-} < \text{N}^{3-}$

The order of radii is: covalent radius < metallic radius < van der Waals' radius

Ionisation Enthalpy (IE)

It is the amount of energy required to remove the loosely bound electron from the isolated gaseous atom. $\text{A(g)} + \text{IE} \rightarrow \text{A}^{+}(\text{g}) + \text{e}^{-}$

Various factors with which IE varies are:

- (i) Atomic size: varies inversely
- (ii) Screening effect: varies inversely
- (iii) Nuclear charge: varies directly
 - Generally left to right in periods ionisation enthalpy increases; down the group, it decreases.
 - IE values of inert gases are exceptionally higher due to stable configuration.
 - Successive ionisation enthalpies: $\text{IE}_3 > \text{IE}_2 > \text{IE}_1$
 - IE_1 of N is greater than that of oxygen due to stable half – filled 2p-orbitals.
 - Among transition elements of 3d-series, ^{24}Cr and ^{29}Cu have higher IE_2 due to half-filled and fully-filled stable d-orbitals

Electron Gain Enthalpy (EGE or ΔH_{eg})

It is the amount of energy released when an electron is added in an isolated gaseous atom. First electron gain enthalpy is negative while the other successive electron gain enthalpy will be positive due to repulsion between the electrons already present in the anion and the electron being added.

Various factors with which electron gain enthalpy varies are:

- (i) Atomic size: varies directly
- (ii) Nuclear charge: varies directly
 - Along a period, electron gain enthalpy becomes more and more negative while on moving down the group, it becomes less negative.
 - Noble gases have positive electron gain enthalpies.
 - Halogen have maximum value of ΔH_{eg} within a period due to smallest atomic size.
 - F and O atom have small size and high charge density, therefore have lower electron gain enthalpy, than Cl and S respectively
 - $\text{Cl} > \text{F}$; $\text{S} > \text{O}$
 - Elements having half-filled and fully-filled orbitals exhibit more stability. therefore, electron gain enthalpy will be low for such elements.
 - Electron gain enthalpy can be measured by Born-Haber cycle and elements with high ΔH_{eg} are good oxidising agent.

Electronegativity (EN)

It is defined as the tendency of an atom to attract the shared electron pair towards itself in a covalent bond. Various factors with which electronegativity varies are:

(i) Atomic size: varies inversely

(ii) Charge on the ion: varies directly, e.g., $\text{Li} < \text{Li}^+$, $\text{Fe}^{2+} < \text{Fe}^{3+}$

(iii) Hybridisation: (Electronegativity; s-character in the hybrid orbital)

Electronegativity of carbon atom = $\text{C}_2\text{H}_6 < \text{C}_2\text{H}_4 < \text{C}_2\text{H}_2$

- In periods as we move from left to right electronegativity increases, while in the group's electronegativity decreases down the group.
- For noble gases, its value is taken as zero.
- Electronegativity helps to predict the polarity of bonds and dipole moment of molecules.
- Electronegativity order of some elements (on different scale) is

(i) **Mulliken scale** : Electronegativity (x) = $\text{IE} + \Delta H_{\text{eg}} / 2$

(ii) **Pauling scale** : The difference in electronegativity of two atoms A and B is given by the relationship; $x_B - x_A = 0.18 \sqrt{\Delta}$

[where, $\Delta = E_{\text{A-B}} - \sqrt{E_{\text{A-A}} * E_{\text{B-B}}}$. (Δ is known as resonance energy.) $E_{\text{A-B}}$, $E_{\text{A-A}}$ and $E_{\text{B-B}}$ represent bond dissociation energies of the bonds A – B, A – A and B – B respectively].

(iii) **Allred and Rochow method**: Electronegativity = $0.744 + 0.359 Z_{\text{eff}} / r^2$

[where, Z_{eff} is the effective nuclear charge = $Z - \sigma$

where, σ is screening constant. It's value can be determined by Slater's rule].

Valency

It is defined as the combining capacity of the element. The valency of an element is related to the electronic configuration of its atom and usually determined by electrons present in the valence shell,

- On moving along a period from left to right, valency increases from 1 to 4 and then decreases to zero (for noble gases) while on moving down a group the valency remains the same.
- Transition metals exhibit variable valency because they can use electron (from outer as well as penultimate shell).

Chemical Reactivity

- Reactivity of metal increases with decrease in IE, electronegativity and increase in atomic size as well as electropositive character.
- Reactivity of non-metals increases with increase in electronegativity as well as electron gain enthalpy and decrease in atomic radii.

Melting and Boiling Points

- On moving down the group, the melting point and boiling point for metallic elements go on decreasing due to the decreasing forces of attraction. However, for non-metals, melting point and boiling point generally increase down the group.

[Along a period from left to right, melting point and boiling point increases and reaches a maximum value in the middle of the period and then start decreasing].

1. Tungsten (W) has highest m.p. (3683K) among metals,
2. carbon (diamond) has the highest m.p., (4000 K) among non-metals.
3. Li metal has minimum density while
4. Iridium (Ir) metal has maximum density.

Electropositivity or Metallic Character

- The tendency of an atom of the element to lose valence electrons and form positive ion is called electropositivity.

- Greater the electropositive character, greater is the metallic character.
- Electropositive character decreases on moving across the period and increases on moving down the group.
- Alkali metals are the most electropositive and halogens are the least electropositive element in their respective period.
- Basic nature of oxides of metallic character, i.e., it also decreases along a period and increases down the group.

Diagonal Relationship

Certain elements of 2nd period show similarity in properties with their diagonal elements in the 3rd period. Thus, (Li resembles Mg), (Be resembles Al) and (B resembles Si).

This is called diagonal relationship and is due to the reason that these pairs of elements have almost

1. identical ionic radii and
2. polarizing power (i.e., charge/size ratio).

Elements of third period, i.e., Mg, Al and Si are known as bridge elements.
