Classification of Elements and Periodicity in Properties

Modern Periodic Table

Mendeleev made a successful effort in grouping elements in the form of his periodic table. He had many achievements, but there were many limitations in his Periodic Table as well.

Some limitations of Mendeleev's periodic table are listed below:

The position of hydrogen was not justified in Mendeleev's periodic table.

The discovery of isotopes revealed another limitation of Mendeleev's periodic table.

Although Mendeleev arranged the elements in the increasing order of their atomic masses, there were instances where he had placed an element with a slightly higher atomic mass before an element with a slightly lower atomic mass.

The limitations of Mendeleev's periodic table forced scientists to believe that atomic mass could not be the basis for the classification of elements.

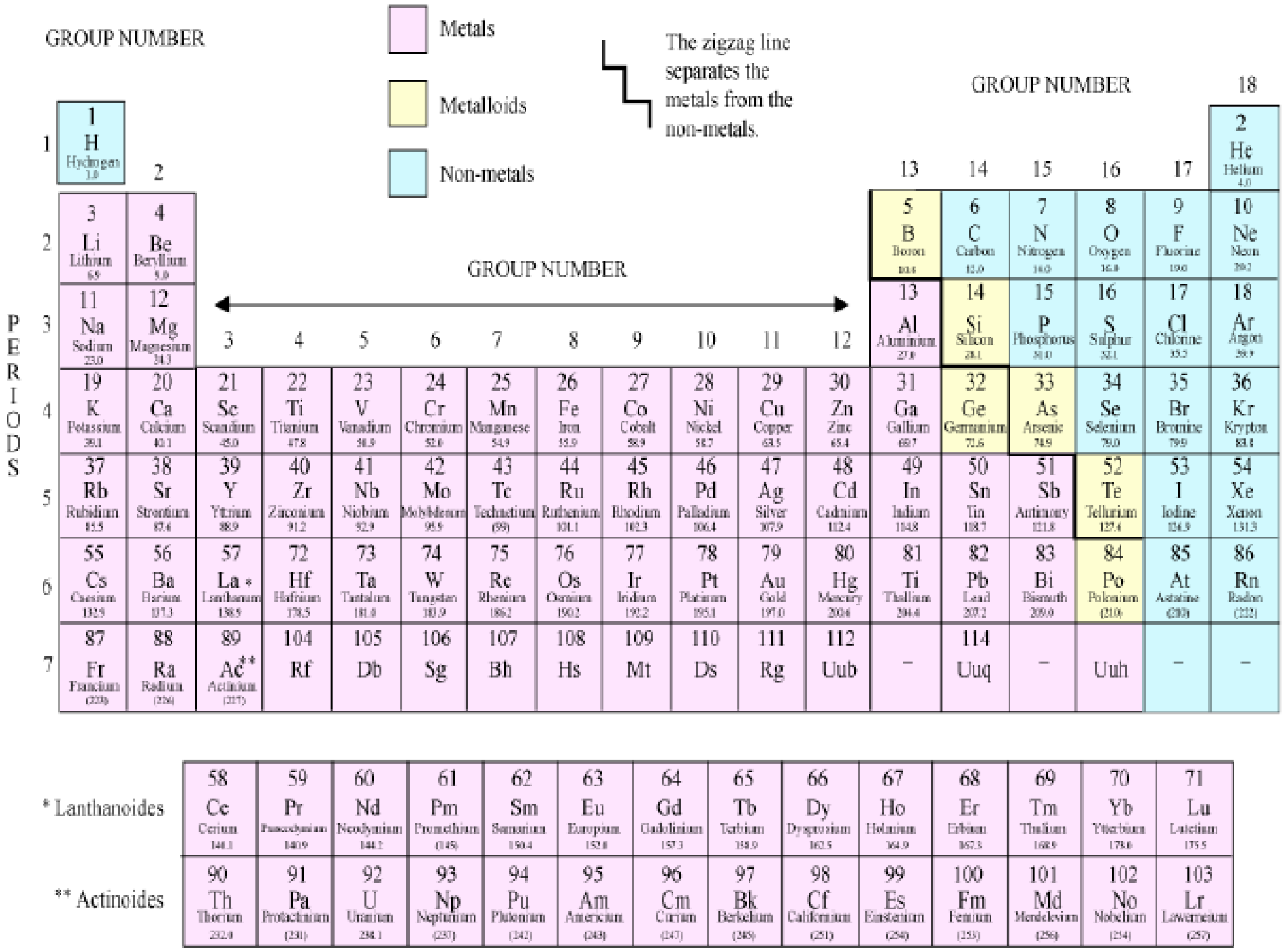
In 1913, **Henry Moseley** demonstrated that atomic number (instead of atomic mass) is a more fundamental property for classifying elements. The atomic number of an element is equal to the number of protons present in an atom of that element. Since the number of protons and electrons in an atom of an element is equal, the atomic number of an element is equal to the number of electrons present in a neutral atom.

Atomic number = Number of protons = Number of electrons

The number of protons or electrons in an element is fixed. No two elements can have the same atomic number. Hence, elements can be easily classified in the increasing order of their atomic numbers. In the light of this fact, Mendeleev's Periodic Law was done away with. As a result, the modern periodic law came into the picture.

The modern periodic law states that the properties of elements are a periodic function of their atomic numbers, not their atomic masses.

The table that is obtained when elements are arranged in the increasing order of their atomic numbers is called the **Modern Periodic Table** or **Long Form of the Periodic Table** as shown in the figure.



The Modern periodic table

In the modern periodic table, the elements are arranged in rows and columns. These rows and columns are known as **periods** and **groups** respectively. The table consists of 7 periods and 18 groups.

Do You Know:

In the modern periodic table, hydrogen is placed above alkali metals because of resemblance with their electronic configurations. However, it is never regarded as an alkali metal. This makes hydrogen a unique element.

If you look at the modern periodic table, you will find that all elements in the same group contain the same number of valence electrons. Let us see the following activity to understand better.

Activity 1: Look at group two of the modern periodic table. Write the name of the first three elements followed by their electronic configurations.

What similarity do you observe in their electronic configurations? How many valence electrons are present in these elements?

The first three elements of group two are beryllium, magnesium, and calcium. All these elements contain the same number of valence electrons. The number of valence electrons present in these elements is 2. On the other hand, the number of shells increases as we go down the group.

Again, if you look at periods in the modern periodic table, you will find that all elements in the same period contain the same valence shell. Let us see the following activity to understand better.

Activity 2: Look at the elements of the third period of the modern periodic table. Write the electronic configuration of each element and calculate the number of valence electrons present in these elements.

What do you observe from the given activity? Do these elements contain the same number of shells? How many valence electrons are present in these elements?

You will find that elements such as sodium, magnesium, aluminium, silicon, phosphorus, sulphur, chlorine, and argon are present in that period. The valence shell in all these elements is the same, but they do not have the same number of valence electrons.

Name of the element	**Electronic configuration (K, L, M)**
Sodium	2, 8, 1
Magnesium	2, 8, 2
Aluminium	2, 8, 3
Silicon	2, 8, 4
Phosphorus	2, 8, 5
Sulphur	2, 8, 6
Chlorine	2, 8, 7
Argon	2, 8, 8

Thus, the number of electrons in the valence shell increases by one unit as the atomic number increases by one unit on moving from left to right in a period.

Let us calculate the number of elements that are present in the first, second, third, and fourth periods.

The maximum number of electrons that a shell can hold can be calculated using the formula $2n^2$. Here, n represents the number of shells from the nucleus. For example, n is equal to 1, 2, and 3 for K, L, and M shells respectively. Hence, the maximum number of electrons that each of these shells can hold can be calculated by substituting the value of n in the given formula.

Number of electrons that K shell can accommodate = $2n^2$

= 2×1^2

= 2

Hence, K shell can accommodate only 2 electrons and only two elements are present in the first period.

Similarly, the second and third shell (L and M respectively) can accommodate 8 and 18 electrons respectively. Since the outermost shell can contain only 8 electrons, there are only 8 elements in both the periods.

Important Note:

The position of an element in the Modern Periodic Table tells us about its chemical reactivity. The valence electrons determine the kind and the number of bonds formed by an element.

IUPAC Nomenclature for Elements with Atomic Number $>$ 100

- Latin word roots for various digits are listed in the given table.

Notation for IUPAC Nomenclature of Elements

Digit	Name	Abbreviation
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

- Latin words for various digits of the atomic number are written together in the order of digits, which make up the atomic number, and at the end, 'ium' is added.
- Nomenclature of elements with the atomic number above 100 is listed below.

Nomenclature of Elements with Atomic Number Above 100

Atomic number	Name	Symbol	IUPAC Official Name	IUPAC Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassnium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium	Rg
112	Ununbium	Uub		
113	Ununtrium	Uut		
114	Ununquadium	Uuq		
115	Ununpentium	Uup		
116	Ununhexium	Uuh		
117	Ununseptium	Uus		
118	Ununoctium	Uuo		

Electronic Configuration and the Periodic Table

Electronic Configuration in Periods

- Period indicates the value of 'n' (principal quantum number) for the outermost or valence shell.
- Successive periods in the periodic table are associated with the filling of the next higher principal energy level (n = 2, n = 3, etc).
- First period (n = 1) → hydrogen ($1s^1$) and helium ($1s^2$) [2 elements]
- Second period (n = 2) → Li ($1s^2\ 2s^1$), Be ($1s^2\ 2s^2$), B ($1s^2\ 2s^2\ 2p^1$) to Ne ($2s^2\ 2p^6$) [8 elements]
- Third period (n = 3) → filling to $3s$ and $3p$ orbitals gives rise to 8 elements (Na to Ar)
- Fourth period (n = 4) → 18 elements (K to Kr) – filling of the $4s$ and $4p$ orbitals

  $3d$ orbital is filled up before $4p$ orbitals ($3d$ orbitals → energetically favourable)
- $3d$-transition series → Sc ($3d^1\ 4s^2$) to Zn ($3d^{10}\ 4s^2$)
- Fifth period (n = 5) → 18 elements (Rb to Xe)
- $4d$-transition series starts at Ytterbium and ends at Cadmium.
- Sixth period (n = 6) → 32 elements; electrons enter $6s$, $4f$, $5d$, and $6p$ orbitals successively. Elements from Z = 58 to Z = 71 are called $4f$-inner transition series or lanthanoid series (filling up of the $4f$ orbitals).
- Seventh period (n = 7) → electrons enter at $7s$, $5f$, $6d$, and $7p$ orbitals successively. Filling up of $5f$ orbitals after Ac (Z = 89) gives $5f$-inner transition series or the actinoid series.

Electronic Configuration in Groups

- Same number of electrons is present in the outer orbitals (that is, similar valence shell electronic configuration).
- Electronic configuration of group 1 elements is given in the following table.

Atomic number	**Symbol**	**Electronic configuration**
3	Li	$1s^2\ 2s^1$(or) [He]$2s^1$
11	Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$(or) [Ne]$3s^1$

19	K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$(or) $[Ar]4s^1$
37	Rb	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^1$(or) $[Kr]5s^1$
55	Cs	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^6\ 6s^1(or)[Xe]6s^1$
87	Fr	$[Rn]7s^1$

Electronic Configurations and Types of Elements

- ***s*- Block Elements**

- Group 1 (alkali metals) – ns^1 (outermost electronic configuration)
- Group 2 (alkaline earth metals) – ns^2 (outermost electronic configuration)
- Alkali metals form +1 ion and alkaline earth metals form +2 ion.
- Reactivity increases as we move down the group.
- They are never found in the pure state in nature. (Reason – they are highly reactive)

- ***p* - Block Elements**

- Elements belonging to Groups 13 to 18
- Outermost electronic configuration varies from ns^2np^1 to ns^2np^6 in each period.
- Group 18 (ns^2np^6) – noble gases
- Group 17 – halogen
- Group 16 – chalcogens
- Non-metallic character increases from left to right across a period.

- ***d*- Block Elements (Transition Elements)**

- Elements of group 3 to group 12
- General electronic configuration is $(n-1)\ d^{1-10}\ ns^{0\text{-}2}$.
- Called transition elements
- Zn, Cd, and Hg with $(n-1)\ d^{10}\ ns^2$ configuration do not show properties of transition elements.
- All are metals. They form coloured ions, exhibit variable oxidation states, paramagnetism, and are used as catalysts.

- ***f*- Block Elements**

- Lanthanoids → Ce (Z = 58) to Lu (Z = 71)
- Actinoids → Th (Z = 90) to Lr (Z = 103)
- Outer electronic configuration → $(n-2)\ f^{1-14}\ (n-1)\ d^{0-1}\ ns^2$
- They are called inner-transition elements.

- All are metals.
- Actinoid elements are radioactive.
- Elements after uranium are called **Transuranium** elements.

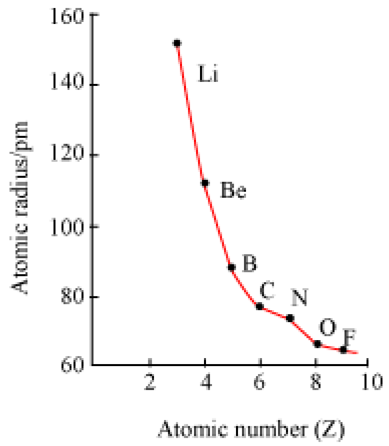
Metals, Non-metals, and Metalloids

- Metals → Appear on the left side of the periodic table
- Non-metals → Located at the top right-hand side of the periodic table
- Elements change from metallic to non-metallic from left to right.
- Elements such as Si, Ge, As, Sb, Te show the characteristic properties of both metals and non-metals. They are called semi-metals or metalloids.

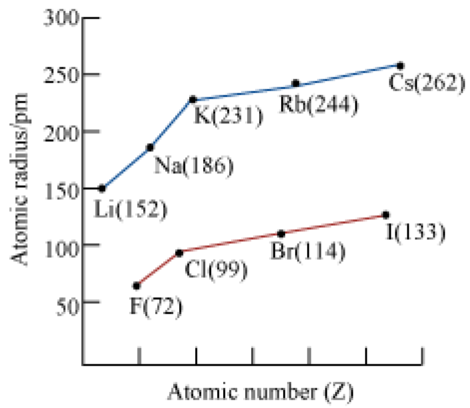
Periodic Trends in Physical Properties

Atomic Radius

- Atomic radii decrease with the increase in the atomic number in a period.
- For example, atomic radii decrease from Li to F in the second period.



- Nuclear charge increases progressively by one unit on moving from left to right across the period. As a result, the electron cloud is pulled closer to the nucleus by the increased effective nuclear charge, which causes decrease in atomic size.
- Atomic radii increase from top to bottom within a group of the periodic table.
- Variation of atomic radii with atomic number among alkali metals and halogen:



Ionic Radius

- Cation is smaller than its parent atom.
- The size of the anion is larger than its parent atom.

Ionization Enthalpy

- Defined as the amount of energy required to remove the most loosely bound electron from the isolated gaseous atom in its ground state

$$A(g) \longrightarrow A^{+}(g) + e^{-} \qquad \Delta H = \Delta_i H_1$$

$$A^{+}(g) \longrightarrow A^{2+}(g) + e^{-} \qquad \Delta H = \Delta_i H_2$$

Unipositive
ion

$$A^{2+}(g) \longrightarrow A^{3+}(g) + e^{-} \qquad \Delta H = \Delta_i H_3$$

Dipositive
ion

$$\Delta_i H \longrightarrow \Delta_i H_3 > \Delta_i H_2 > \Delta_i H_1$$

- Decreases with the increase in atomic size
- Increases with the increase in nuclear charge
- Decreases with the increase in the number of inner electrons
- Increases with the increase in penetration power of electrons
- Atom having a more stable configuration has high value of enthalpy.

- Variation across a period: Increases with the increase in atomic number across the period.

- Variation in a group: Decreases regularly with the increase in atomic number within a group.

Electron Gain Enthalpy

- Defined as the enthalpy change taking place when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion

$$X(g) + e^- \longrightarrow X^-(g)$$

- Larger the value of electron gain enthalpy, greater is the tendency of an atom to accept electron.

- Greater the magnitude of nuclear charge, larger will be the negative value of electron gain enthalpy.

- Larger the size of the atom, smaller will be the negative value of electron gain enthalpy.

- More stable the electronic configuration of the atom, more positive will be the value of its electron gain enthalpy.

- Variation across a period – Tends to become more negative as we go from left to right across a period

- Variation down a group – Becomes less negative on going down the group

Electronegativity

- Defined as the tendency of an atom in a molecule to attract the shared pair of electrons towards itself

- Greater the effective nuclear charge, greater is the electronegativity.

- Smaller the atomic radius, greater is the electronegativity.

- In a period – Increases on moving from left to right

- In a group – Decreases on moving down a group

Valency

- It is defined as the number of univalent atoms which can combine with an atom of the given element.
- Valency is given by the number of electrons in outermost shell.
- If the number of valence electrons ≤4: valency = number of valence electrons
- If the number of valence electrons >4: valency = (8 - number of valence electrons)

- In a period – Increases from 1 to 4 and then decreases from 4 to zero on moving from left to right

- In a group – No change in the valency of elements on moving down a group. All elements belonging to a particular group exhibit same valency.

Non –Metallic (and Metallic Character) of an Element

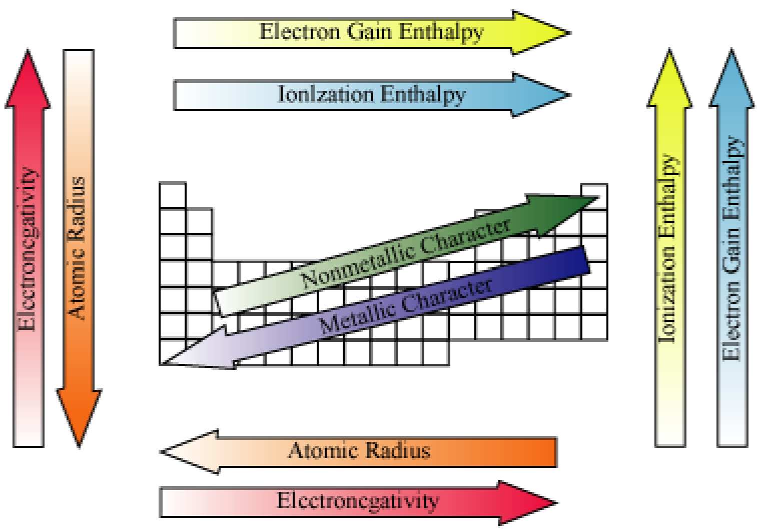
- Non-metallic elements have strong tendency to gain electrons.

- Non-metallic character is directly related to electronegativity and metallic character is inversely related to electronegativity.

- Across a period, electronegativity increases. Hence, non-metallic character increases (and metallic character decreases).

- Down a group, electronegativity decreases. Hence, non-metallic character decreases (and metallic character increases).

The periodic trends of various properties of elements in the periodic table are shown in figure.



Periodic Trends in Chemical Properties

Periodicity of Valence or Oxidation States

- Valence of the elements = Number of electrons in the outermost orbitals (if valence electrons ≤ 4)
- Or, valency of the element = 8 − Number of outermost electrons (if valence electrons > 4)

Group	1	2	13	14	15	16	17	18
Number of valence electrons	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3,5	2,6	1,7	0,8

- The given table shows the periodic trends observed in the valence of elements (hydrides and oxides).

Group	1	2	13	14	15	16	17
Formula of hydride	LiH		B_2H_6	CH_4	NH_3	H_2O	HF
	NaH	CaH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
	KH			GeH_4	AsH_3	H_2Se	HBr
				SnH_4	SbH_3	H_2Te	HI
Formula of oxide	Li_2O	MgO	B_2O_3	CO_2	N_2O_3, N_2O_5		
	Na_2O	CaO	Al_2O_3	SiO_2	P_4O_6, P_4O_{10}	SO_3	Cl_2O_7
	K_2O	SrO	Ga_2O_3	GeO_2	As_2O_3, As_2O_5	SeO_3	
		BaO	In_2O_3	SnO_2	Sb_2O_3,Sb_2O_5	TeO_3	
				PbO_2	Bi_2O_3		

- Many elements exhibit variable valence (particularly transition elements and actinoids).

Anomalous Properties of Second Period Elements

- First member of each group (the element in the second period from lithium to fluorine) differs in many respects from the rest of the members of the same group.
- For example, the behaviour of Li and Be is more similar with the second element of the following group i.e., Mg and Al respectively.

- Such sort of similarity is commonly known as **diagonal relationship** in periodic properties.

- Reasons for the different chemical behaviour of the first member of a group of elements in the *s*-and *p*-blocks as compared to the other members in the same group are as follows.

- Small atomic size of the first element
- Large charge/radius ratio
- High electronegativity
- Absence of *d*-orbitals in the valence shell
- Ability of form $p\pi - p\pi$ multiple bonds

- First member of each group of *p*-block element has the tendency to form $p\pi - p\pi$ multiple bonds to itself and to the other second period elements. For example, C = C, C ≡ C, C = O, C = N

- Reason – This property of the elements is due to their small size.

- Higher members of the group have little tendency to form $p\pi - p\pi$ bonds.

Periodic Trends and Chemical Reactivity

- High chemical reactivity at the two extremes of a period and the lowest in the centre

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- Maximum chemical reactivity is at the extreme left of a period because of the ease of electron loss (or low ionization enthalpy).

- Elements at the extreme left exhibit strong reducing behaviour and elements at the extreme right exhibit strong oxidizing behaviour.

- Oxides formed by the elements on the left are basic and by the elements on the right are acidic in nature.

- Oxides of elements in the centre are amphoteric or neutral.

- The electron gain enthalpy increases across a period and decreases down the group.