Classification of elements and periodicity in Properties

In 1800 only 31 elements were known. By 1865 the number of elements identified more than 63. At present 118 elements are known. With such a large number of elements it is very difficult to study each element individually. To make it easy scientist proposed a systematic way to organise the elements.

Genesis of periodic table

- **Law of triads** - Johann Dobereiner first gave the idea of trends among properties of elements. In 1829 he propounded that several groups of three elements having similar physical and chemical properties which were later known as triads.
- According to him in each triad middle element's atomic weight is average of other two element's atomic weight.

Table: Dobereriner's Triads

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

- **Law of Octaves-** In 1865 John Alexander Newlands propounded the law of octaves. He arranged elements in increasing order of their atomic weights and found that every eight-element had similar properties to first element.

Table: Newlands' Octaves

Element	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
At. wt.	7	9	11	12	14	16	19
Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
At. wt.	23	24	27	29	31	32	35.5
Element	K	Ca					
At. wt.	**39**	**40**					

- **Periodic law –** Dmitri Mendeleev and Lother Meyer both chemists propounded that on arranging the elements in increasing order of atomic weight similarity appear in chemical and physical properties.
- Lother Meyer plotted the physical properties such as atomic volume, melting point and boiling point against atomic weight and obtained a periodically repeated pattern.
- Mendeleev published the periodic law first time. It states "the properties of the elements are a periodic function of their atomic weights."
- Mendeleev arranged elements in horizontal rows and vertical column of a table in order of their increasing atomic weights in such a manner that elements with similar properties occupied same vertical column.
- Mendeleev left the gap under aluminium and silicon, called these elements ka- Aluminium and Eka- Silicon. He predicted the existence of gallium, germanium and described their physical properties. These elements were discovered later.

Table: Mendeleev's Prediction for the elements Eka-aluminium (Gallium) and Eka-silicon (Germanium)

Property	**Eka-aluminium (predicted)**	**Gallium (found)**	**Eka- silicon (pridicted)**	**Germanium (found)**
Atomic weight	68	70	72	72.6
Density (g/cm^3)	5.9	5.94	5.5	5.36
Melting Point (K)	Low	302.93	High	1231
Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

Group	I	II	III	IV	V	VI	VII	VIII
Oxide : Hydride:	R_2O RH	RO RH_4	R_2O_3 RH_4	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 22.99	Al 24.31	Si 28.09	P 30.974	S 32.06	Cl 35.453	
4 First series Second series	K 39.102 Cu 63.54	Ca 40.08 Zn 65.54	Sc 44.96 Ga 69.72	Ti 47.90 Ge 72.59	V 50.94 As 74.92	Cr 50.20 Se 78.96	Mn 54.94 Br 79.909	Fe Co Ni 55.85 58.93 58.71
5 First series Second series	Rb 85.47 Ag 107.87	Sr 87.62 Cd 112.40	Y 88.91 In 114.82	Zr 91.22 Sn 118.69	Nb 92.91 Sb 121.60	Mo 95.94 Te 127.60	Tc 99 I 126.90	Ru Rh Pd 101.07 102.91 106.4
6 First series Second series	Cs 132.90 Au 196.97	Ba 137.34 Hg 200.59	La 138.91 Tl 204.37	Hf 178.40 Pb 207.19	Ta 180.95 Bi 208.98	W 183.85		Ru Rh Pd 190.2 192.2 195.09

Figure: Mendeleev's periodic table earlier

Modern periodic law

- According to this law "the physical and chemical properties of the elements are periodic functions of their atomic numbers".
- A modern version or long form of the periodic table of the element is the most convenient and used.
- Horizontal rows are called periods and vertical column are groups.
- Elements having same outer electronic configuration in their atoms are arranged in vertical column and referred as families.
- The groups are numbered from 1 to 18 replacing the older notation of groups IA… VIIA, VIII, IB and 0.
- The first period contains 2 elements. The subsequent periods consist of 8, 8, 18, 18 and 32 elements respectively.
- Seventh period is incomplete.
- 14 elements of both sixth and seventh periods are placed in separate panels at the bottom.

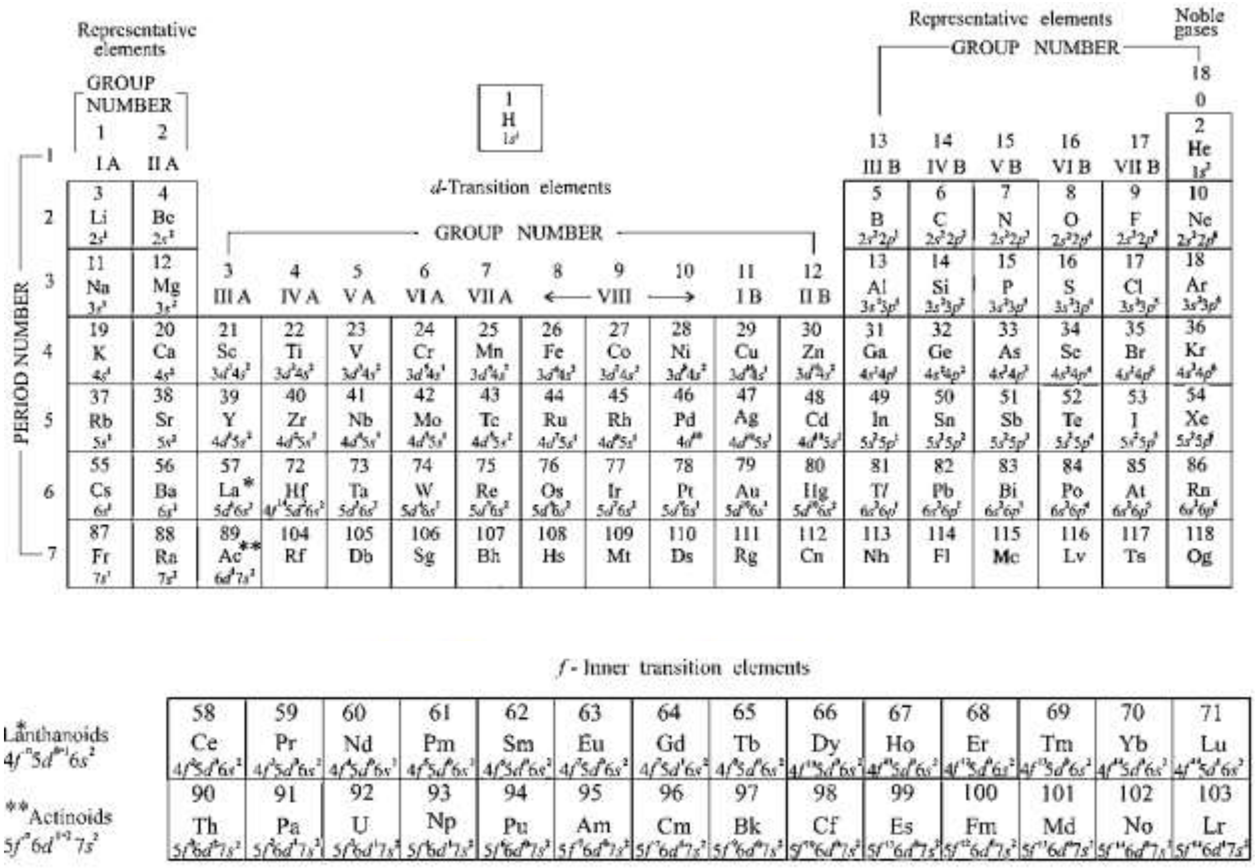


Figure: Long form of the periodic table

Nomenclature of elements with atomic numbers> 100

- Traditionally the new elements named after the name of discoverer. IUPAC (International Union of Pure and Applied Chemistry) ratified it.
- In case of element 104 both Americans and Soviet scientists claimed credit for discovering it. The Americans named it Rutherfordium and Soviets named it Kurchatovium. To avoid such problems IUPAC established.
- A systematic nomenclature derived from the numerical roots for 0-9 numbers.
- The new element first gets a temporary name with a symbol of three letters. Later permanent name and symbol are given by IUPAC representatives from each country.
- Today 118 elements with atomic numbers and official names by IUPAC are discovered.

Table: Notation for IUPAC Nomenclature of Element

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

Table: Nomenclature of Elements with atomic number above 100

Atomic Number	Name according to IUPAC nomenclature	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	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnillium	Uun	Darmstadium	Ds
111	Unununnium	Uuu	Rontgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	NIhonium	Nh
114	Ununquadium	Uuq	Flerovium	Fl
115	Ununpentium	Uup	Moscovium	Mc
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununnoctium	Uuo	Oganesson	Og

Electronic configuration of elements and the periodic table

- The characteristics of an electron in an atom are denoted by four quantum numbers n, l, s, m. The principal quantum (n) defines the main energy level known as shell. Filling of electrons in to subshell i.e., orbitals s, p, d, f in atom take place.
- The distribution of electrons into orbitals of an atom is called its electronic configuration.
- There is a direct connection between electronic configuration of an element and its position in periodic table.

A). **Electronic configuration in Periods**

- The period indicates the value of n for the outermost or valence shell.
- The first period (n=1) starts with the filling of the lowest level (1s). It has two elements hydrogen ($1s^1$) and helium ($1s^2$), the first shell (K) completed.
- The second period (n=2) starts with lithium and third electron enters the 2s orbital. The next element Be has four electrons. It's electronic configuration is $1s^2 1s^2$.
- The next element Boron (B), 2p orbital is filled by one electron. 2p orbital is filled till the L shell is completely filled at Neon ($2s^2 2p^6$). Thus, there are eight elements in second period.
- The third period begins from sodium (Na) and end at argon. The electron enters 3s orbital in case of sodium and completely filled 3p orbital in argon. Third period have eight elements from sodium to argon.
- The fourth period (n=4) starts from potassium and electron enter 4s orbital. Now 3d orbital is filled before 4s orbital because 3d is energetically more favourable. We enter 3d transition series of elements which starts from scandium (Sc) (Z=21). Sc has electronic configuration $3d^1 4s^2$.

- The 3d orbitals are filled at Zinc (Z=30) with electronic configuration $3d^{10}4s^2$. The fourth period ends at krypton with the filling up of the 4p orbitals. There are 18 elements in the fourth period.
- The fifth period (n=5) beginning with Rubidium. The fifth period contains 4d transition series. It starts from Yttrium and end at Xenon.
- The sixth period (n=6) contains 32 elements. Successive electrons enter 6s, 4f, 5d and 6p orbitals. The filling of 4f orbitals starts from Cerium (Z=58) and ends at Lutetium (Z=71) to give the 4f inner transition series which is known as Lanthanoid series.
- In seventh period (n=7) successive filling of 7s, 5f, 6d and 7p orbitals take place. Most of the elements of this series are man-made radioactive elements.
- The 5f inner transition starts from actinium (Z=89) by filling up 5f orbitals. This series is also known as actinoid series.
- The 4f and 5f inner transition series of elements are placed separately in the periodic table.

B) **Group wise electronic configurations**

- Elements which have same vertical column or group have similar valence shell electronic configuration; same number of electrons in the outer orbitals. For example group-I elements have ns^1 valence shell electronic configuration.

Table: electronic configuration of group-I

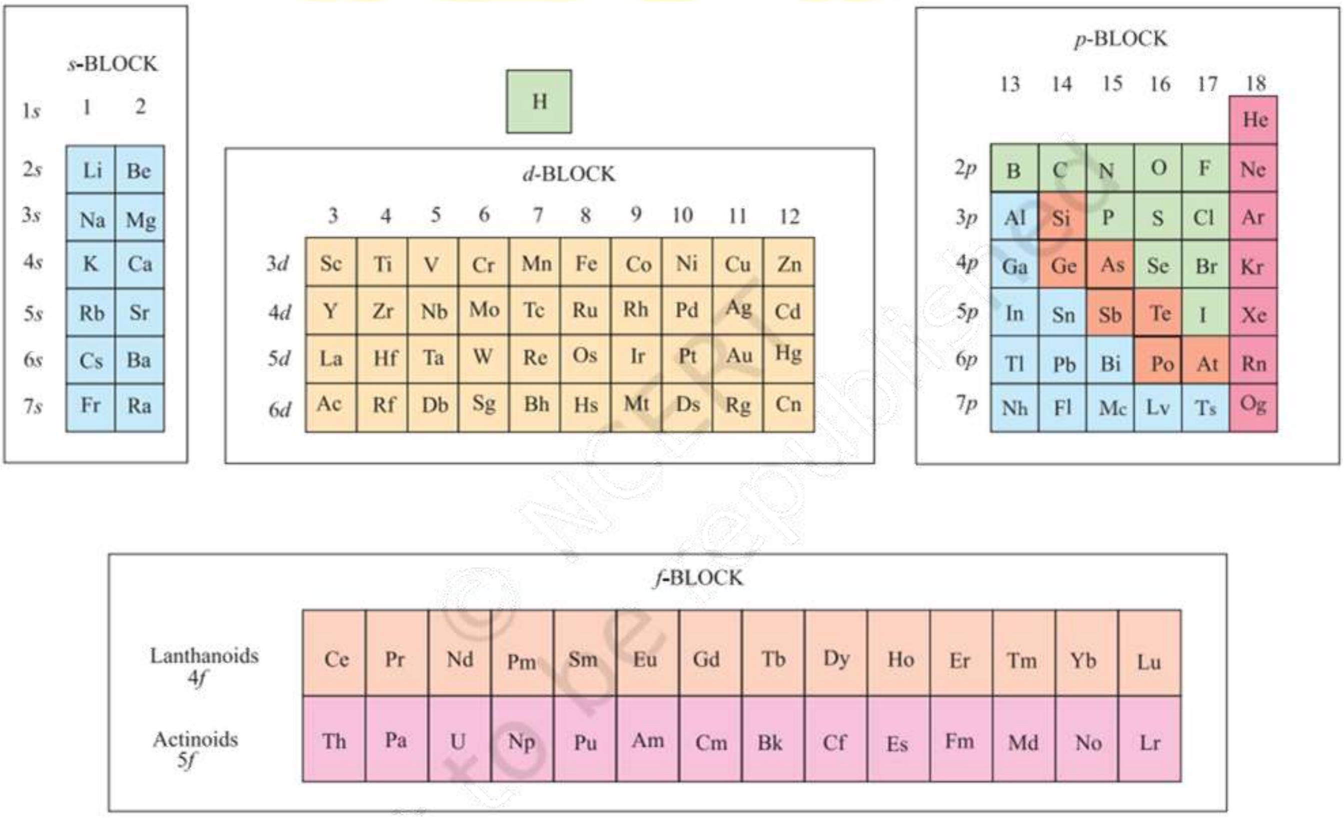
Atomic number	Symbol	Electronic configuration
3	Li	$1s^2 2s^1$ or [He] $2s^2$
11	Na	$1s^2 2s^2 2p^6 3s^1$ or [Ne]$3s^1$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or [Ar]$4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ or [Kr]$5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ or [Xe]6s1

87	Fr	[Rn]$7s^1$

Electronic configurations and types of elements in s, p, d, f blocks

- The elements in a vertical column of the periodic table have similar chemical behaviour. This similar behaviour arises because these elements have same number of electrons in their outermost orbitals.
- We can classify the elements of periodic table into four blocks i.e. s-block, p-block, d-block and f-block.
- There are two exceptions to this categorisation- first one is hydrogen, it has only one s-electron and hence can be group-I. It can also gain an electron to achieve a noble gas arrangement hence it can be placed in group- 17. It is a special case so hydrogen separately place at the top of periodic table.

  Second exception is helium belongs to the s-block but it is positioned in the p-block along with other group 18 elements because it has a completely filled valence shell i.e. $1s^2$. It shows properties similar to noble gases.



The s-block elements

- The elements of group-1 (alkali metals) and group -2 (alkaline earth meals) which have ns^1 and ns^2 electronic configuration of outermost valence shell.
- S-block elements are reactive metals with low ionization enthalpies. They tend to lose one or two outer most electrons to form +1 ion or +2 ions.
- The metallic character and reactivity increase as we go down the group.

The p-block elements

- Elements of group 13 to 18 forms p-block. s-block and p-block collectively known as Representative elements or Main group elements.
- Outermost electronic configuration varies from ns^2np^1 to ns^2np^6 in each period.
- The valence shell of the noble gases is completely filled. It is very stable electronic configuration. The noble gases show very low chemical reactivity. Halogens and chalcogens are the important groups of non-metal elements.
- Halogens and chalcogens have highly negative electron gain enthalpies. They are readily accept or gain electron to attain noble gas configuration.
- As we move from left to right across a period non-metallic character increases.
- As we move down the group metallic character increases.

The d-block elements (Transition elements)

- These are the elements of group 3 to 12 situated in the centre of the periodic table.
- It is known as d-block elements because electrons are filled in d orbitals.
- The general outer electronic configuration is $(n-1)d^{1-10}ns^{0-2}$.
- They are metals and mostly form coloured ions, exhibit variable oxidation states, para-magnetism and used as catalysts.

- The transition metals form a bridge between the chemically active metals of s-block elements and less active elements of group 13 and 14 and thus take their familiar name Transition elements.

The f-block elements (Inner transition elements)

- The two rows of elements at bottom of the periodic table called Lanthanoids [Ce (Z=58) to Lu (Z=71)], Actinoids [Th (Z=90) to Lr (Z=103)]
- The general outer electronic configuration is $(n-2)f^{1-}14(n-1)d^{0-1}ns^2$
- Lanthanoids and actinoids elements are known as inner transition elements (f-block elements).
- All elements show metallic characteristics. Actinoids elements are radioactive.
- The elements after uranium are called Trans uranium elements.

Metals, non-metals and metalloids

- The elements can be divided into metals and non-metals. 78% of elements of periodic table are in metallic nature.
- Metals are solid at room temperature. Mercury is an exception.
- They are situated on the left side of periodic table. Metals usually have melting point. They are good conductor of electricity and heat. They are malleable and ductile.
- Non-metals are situated at top right-hand side of the periodic table.
- Non-metals are solid or gases at room temperature. They are poor conductor of heat and electricity.
- Most non-metal solids are brittle. They are neither ductile nor malleable.
- Metallic characters increase as we move down a group. Non-metallic characters increases as we move left to right across the periodic table.
- The elements bordering this line and running diagonally across the periodic table show properties similar to both metals and non-metals. These elements are known as semi-metals or metalloids.

Periodic trends in properties

There are many observable unique patterns in the chemical and physical properties of elements as we move across a period or down in a group in the periodic table. There are some trends in chemical and physical properties of atoms which are discussed below:

Trends in physical properties

- There are various physical properties of elements such as boiling point, melting point, heats of fusion and vaporization, energy of atomization, etc.
- Periodic trends with respect to ionic and atomic radii, ionization enthalpy, electron gain enthalpy and electronegativity discussed one by one below-

1. Atomic radius

- The size of an atom is ~ 1.2 Å, which is very small. There is not any practical way to measure the size of an individual atom.
- There is a practical approach to measure radius is first measure the distance between two atoms when they are bound together by a single bond in a covalent molecule and from this value the covalent radius of element can be calculated.
- For example in Cl_2, the bond distance between two chlorine molecule is 198pm and half of this distance i.e. 99pm ids atomic radius of chlorine.
- In case of metals we called it metallic radius. Metallic radius is half of the internuclear distance separating the metal cores in the metallic crystal. For example the internuclear distance between two copper atoms in solid copper is 265pm. The metallic radius of copper will be 265/2 i.e., 128pm.
- We can refer atomic radius for both metallic and covalent radius.
- Generally, the atomic size decreases across a period because with in a period the outer electrons are in the same valence shell and the effective nuclear charge (Z_{eff}) increases as the atomic number increases, resulting in the increased attraction between nucleus and electrons.
- In case of family or vertical column of a periodic table, atomic radius increases regularly with atomic number

Table: Atomic radii (pm) across the periods

Atom (period II)	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
Atomic radius	152	111	88	77	74	66	64
Atom (period III)	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**
Atomic radius	186	160	143	117	110	104	99

Table: Atomic Radii (pm) down a family

Atom (Group I)	Atomic Radius	**Atom (Group 17)**	Atomic Radius
Li	152	**F**	64
Na	186	**Cl**	99
K	231	**Br**	114
Rb	244	**I**	133
Cs	262	**At**	140

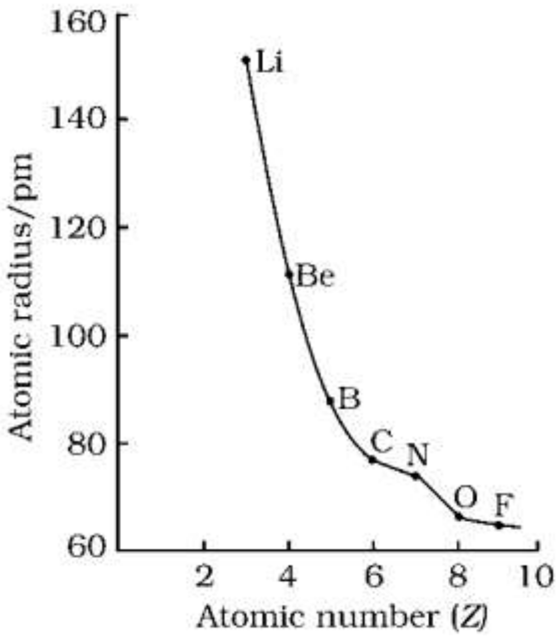


Figure: variation of atomic radius with atomic number across the second period

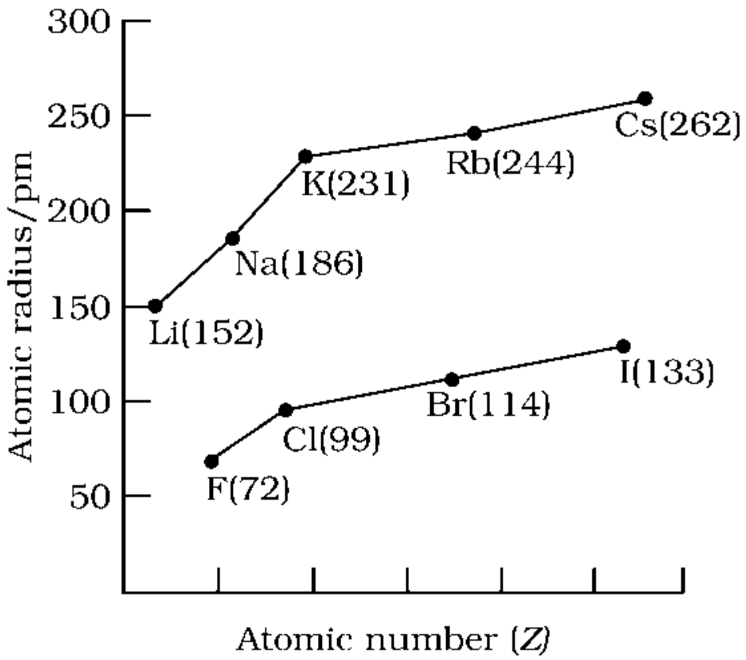


Figure: Variation of atomic radius with atomic number for alkali metals and halogens

2. Ionic Radius

- The removal of an electron from an atom result in the formation of a cation and gain of an electron leads to an anion.
- The ionic radii can be measured by calculating the distance between cation and anion in an ionic crystal.
- The size of cation is smaller than its parent atom as it has less electrons but nuclear charge remains same.
- The size of an anion will be larger than that of parental atom because the addition of one or more electrons increases repulsion among electrons and decrease in Z_{eff}.
- Example: Ionic radius of fluoride ion (F^-) = 136 pm; atomic radius of fluorine = 64 pm

  The atomic radius of sodium = 186 pm; ionic radius of Na^+= 95 pm.

  Isoelectronic species: some atoms and ions which contain the same number of electrons are known as isoelectronic species. O^{-2}, F^-, Na^+ and

Mg^{+2} have same number of electrons i.e. 10. But the radius would be different due to different nuclear charges.

3. Ionization Enthalpy

- The tendency of an element to lose electron is known as ionization enthalpy.
- The energy required to remove an electron from an isolated gaseous atom in its ground state is called ionization enthalpy.
- The first ionization enthalpy for an element X:

$$\mathbf{X(g) \rightarrow X^{+}(g) + e^{-}}$$

- The ionization enthalpy ($\Delta_i H$) is expressed in unit of kJ mol^{-1}.
- The second ionization enthalpy is energy required to remove the second most loosely bound electron.

$$\mathbf{X^{+}(g) \rightarrow X^{2}+(g) + e^{-}}$$

- It is an endothermic process. Ionization enthalpies are always positive.

$$\mathbf{\Delta_i H = (+ive)}$$

- Successive ionization enthalpies are increase.

$$X(g) + (IE)_1 \rightarrow X^{+}(g) + 1e^{-}$$
$$X^{+}(g) + (IE)_2 \rightarrow X^{2+}(g) + 1e^{-}$$
$$X^{+2}(g) + (IE)_3 \rightarrow X^{3+}(g) + 1e^{-}$$

Here IE = Ionization enthalpy

$$\mathbf{(IE)_3 > (IE)_2 > (IE)_1}$$

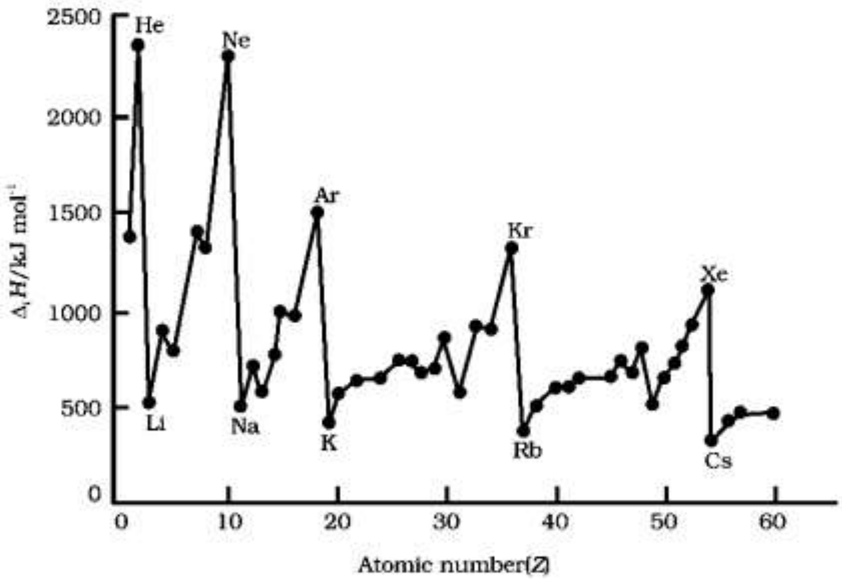


Figure: variation of first ionization enthalpies ($\Delta_i H$) with atomic number of elements with Z=1 to 60

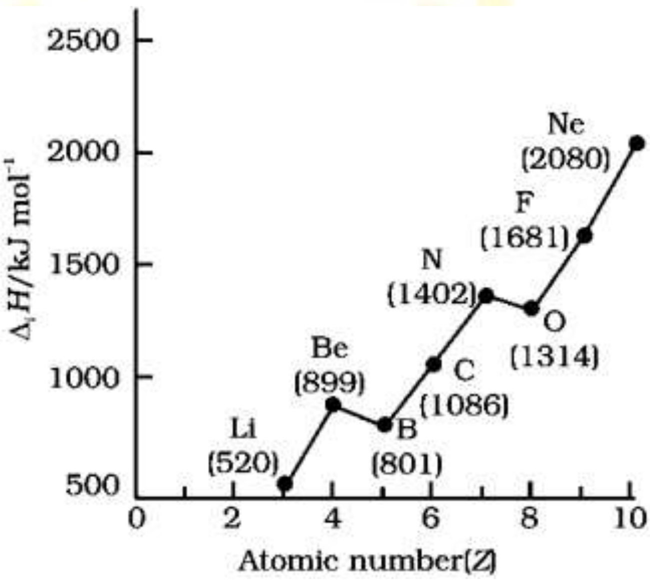


Figure: First Ionization enthalpies ($\Delta_i H$) of elements of the second period as a function of atomic number (Z)

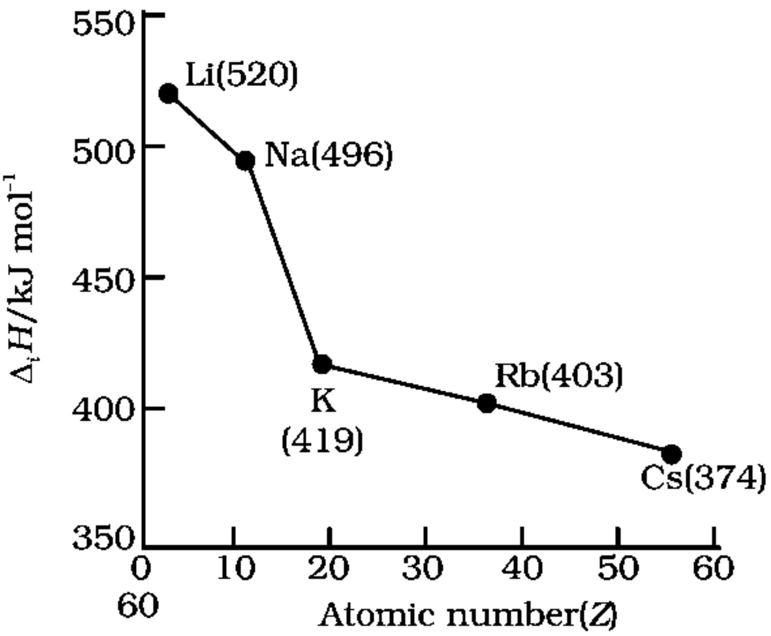


Figure: $\Delta_i H$ of alkali metals as a function of Z

- There are two factors responsible for the various trends of ionization enthalpy in the periodic table – (1) the attraction of electrons towards the nucleus (2) the repulsion of electrons from each other.
- The Z_{eff} experienced by a valence electron in an atom will be less than the actual charge on the nucleus because of **shielding** or **screening** of valence electron from the nucleus by the superseding core electrons. In case of 2s electron in Li is shielded from the nucleus by the inner 1s electrons. The valence electron experiences a net positive charge which is less than the actual charge of +3.

4. Electron gain enthalpy

- Electron gain enthalpy is the amount of energy that is released when an electron is added in valence shell of an isolated gaseous atom.
- It is denoted as ($\Delta_{eg}H$). It can be either exothermic or endothermic process.

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

Table: Electron gain enthalpies (kJ mol^{-1}) of some main group elements

Group I	$\Delta_{eg}H$	**Group 16**	$\Delta_{eg}H$	**Group 17**	$\Delta_{eg}H$	**Group 0**	$\Delta_{eg}H$
H	-73					**He**	+48
Li	-60	**O**	-141	**F**	-328	**Ne**	+116
Na	-53	**S**	-200	**Cl**	-349	**Ar**	+96
K	-48	**Se**	-195	**Br**	-325	**Kr**	+96
Rb	-47	**Te**	-190	**I**	-295	**Xe**	+77
Cs	-46	**Po**	-174	**At**	-270	**Rn**	+68

- Electron gain enthalpy becomes more negative with increase in the atomic number across a period. The Z_{eff} increases from left to right across a period and consequently it would be easier to add an electron to a smaller atom.

5. Electronegativity

- A qualitative measure of the ability of an atom in a chemical compound to attract shared electrons to itself, is called electronegativity.
- It is not a measurable quantity. Pauling scale is mostly used to measure the electronegativity.
- Electronegativity of any given element is not constant. It varies depending upon the element to which it is bound.
- Electronegativity generally increases across a period from left to right in the period.

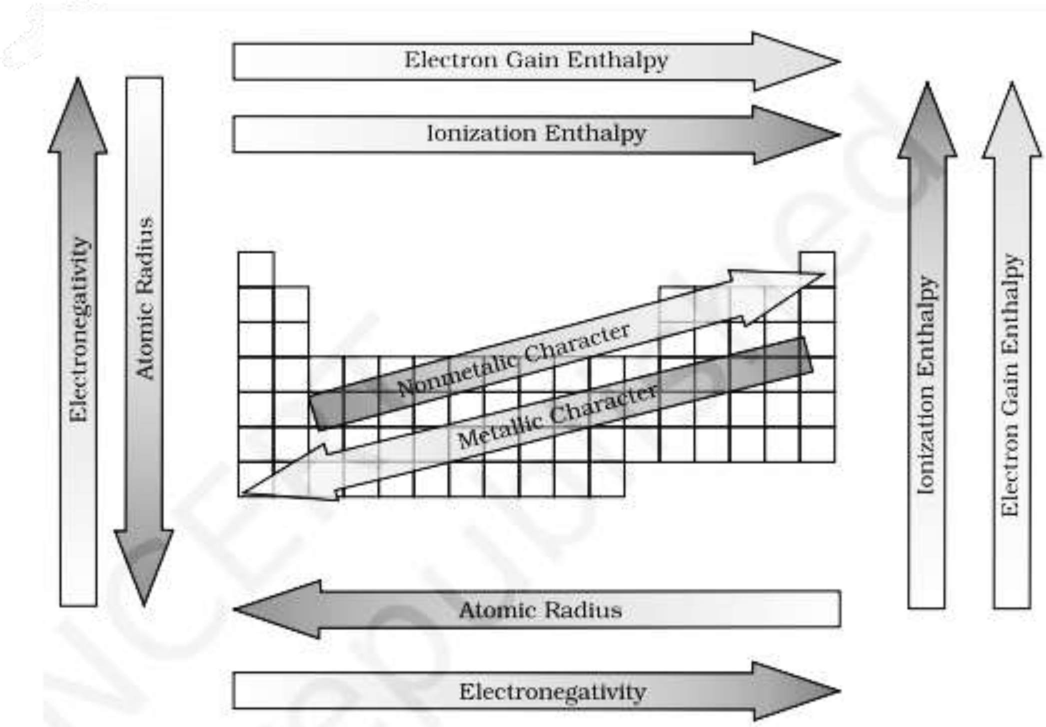


Figure: The periodic trends of elements in the periodic table

Table: Electronegativity values (on Pauling scale) across the periods

Atom (Period II)	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Atom (period III)	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Table: Electronegativity values (on Pauling scale) down a family

Atom (Group I)	Electronegativity value	Atom (Group 17)	Electronegativity value
Li	1.0	F	4.0
Na	0.9	Cl	3.0
K	0.8	Br	2.8
Rb	0.8	I	2.5
CS	0.7	At	2.2

- Electronegativity is directly related to non-metallic properties of elements. Electronegativity is inversely related to the metallic properties of elements. The increase in electronegativity across a period is accompanied by an increase in non-metallic properties of elements. The decrease in non-metallic properties of elements.

Periodic trends in chemical properties

1. Periodicity of valence or oxidation states

- The valence is the most characteristic property of the elements.
- Valence of electron = number of electrons in outermost orbital

  Or

  Valence of electron = 8- number of outermost electron
- The term oxidation state is used valence. For example: In OF_2 and Na_2O. Order of electronegativity is F > O > Na. Electronic configuration of fluorine is $2s^2 2p^5$, fluorine shares one electron with oxygen in the OF_2 molecule. Being highest electronegative element, here fluorine's oxidation state is -1. Oxidation state of oxygen is +2 as it shares 2 electrons with fluorine. In Na_2O, oxygen is more electronegative and accepts two electrons; one from each of the two sodium atoms thus shows oxidation state -2. Sodium oxidation state is +1.

  The oxidation state of an element in a particular compound can be defined as the charge acquired by it's atom on the basis of electronegative consideration from other atoms in the molecule.

Group	1	2	13	14	15	16	17	18
Number of valence electron	1	2	3	4	5	6	7	8
Valence	1	2	3	4	3.5	2.6	1.7	0.8

Table: Periodic trends in Valence of elements as shown by the formulas of their compounds

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 –	–	

2. Anomalous properties of second elements

- The first element of each of the group-I (lithium) and group-II (beryllium) and group 13-17 (boron to fluorine) different in many respects from the other members of their respective group.
- Behaviour of Li and Be is more similar with the second element of the group i.e., Mg and Al respectively. This type of similarity is known as **diagonal relationship** in the periodic properties.

Property	Element		
Metallic radius M (pm)	Li	Be	B
	152	111	88
	Na	Mg	Al
	186	160	143
Ionic radius M+ (pm)	Li	Be	
	76	31	
	Na	Mg	
	102	72	

- The anomalous behaviour is attributed to the small size, large charge/ radius ratio and high electronegativity of the element.
- In addition to it, the first member of group has only four valence orbitals (2s and 2p) available for bonding while the second member of the groups have nine valence orbitals (3s, 3p, 3d).
- The maximum covalency of the first member of each group is 4, $[BF_4]^-$ whereas the other members can expand their valence shell to accommodate more than four pairs of electrons $[AlF_6]^{3-}$.

Periodic trends and Chemical reactivity

- The entire chemical and the physical properties are an expression of the electronic configuration of elements.
- There is high chemical reactivity at the two extremes (i.e the right and left one in the periodic table). It is exhibited by the loss of an electron and formation of cation take place at the left extreme. At the extreme right exhibited by gain of an electron and formation of an anion take place.
- The chemical reactivity of an element is described well by its reaction with O_2 gas and halides.
- Elements positioned at the extreme gives oxide by reacting with oxygen, elements on extreme left mostly gives basic oxide (example- Na_2O) while the elements of extreme right gives acidic oxide (example-Cl_2O_7). Elements in the centre of periodic table gives amphoteric oxides like Al_2O_3, As_2O_3 or neutral oxides like CO, NO, N_2O.
- In case of transition elements, the change in atomic radius is much smaller than representative elements across the period. The change in atomic radius is also small in case of inner transition metals i.e. 4f series.