

Periodic Properties
Part - I

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Development of the Periodic Table

- Early model by Antoine Lavoisier

Antoine Lavoisier's 1789 classification of substances into four 'element' groups

acid-making elements	gas-like elements	metallic elements	earthy elements
sulphur	light	cobalt, mercury, tin	lime (calcium oxide)
phosphorus	caloric (heat)	copper, nickel, iron,	magnesia (magnesium oxide)
charcoal (carbon)	oxygen	gold, lead, silver, zinc	barytes (barium sulphate)
	azote (nitrogen)	manganese, tungsten	argilla (aluminium oxide)
	hydrogen	platina (platinum)	silex (silicon dioxide)

b) Dobereiner's Triads :-

Dobereiner classified elements in the increasing order of their atomic masses into groups of three elements called **triads**. In each triad the atomic mass of the middle element was approximately equal to the average atomic mass of the other two elements.

The defect in this classification was that all the then known elements could not be correctly arranged into triads.

Triad		Atomic mass	Average atomic mass of 1 st and 3 rd element
Lithium	Li	6.9	22.95
Sodium	Na	23.0	
Potassium	K	39.0	
Calcium	Ca	40.1	88.7
Strontium	Sr	87.6	
Barium	Ba	137.3	
Chlorine	Cl	35.5	81.2
Bromine	Br	79.9	
Iodine	I	126.9	

1866: John Newlands Table

Newlands ordered the elements according to atomic mass. He also put them into **groups with similar properties**. He called his table the 'law of octaves':

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co, Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce, La	Zr	Di, Mo	Ro, Ru

Problems with his table:

- More than one element in some boxes (Co, Ni)
- Metals in same groups as non-metals (Ni, Br)
- Groups only work for first 17 elements
- Elements were still being discovered

Julius Lothar Meyer (1830-1895)

- He was working on his periodic table consisting of 56 elements.
- He maintained that the properties of the elements were functions of their atomic mass and prepared a graph wherein he plotted the atomic volume of the elements against its atomic mass.

Mendeleev's periodic law

- Mendeleev's periodic law states that the physical and chemical properties of all elements are a periodic function of their atomic masses.
- Atomic weight and chemical reactivity are the two parameters he chose for classifying the elements:

Mendeleev's Periodic Table

Series	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H=1							
2	Li=7	Be=9.1	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24.4	Al=27	Si=28	P=31	S=32	Cl=35.5	Fe=56, Ni=58.5, Co 59.1, Cu 63.3.
4	K=39.1	Ca=40	— =44	Ti=48.1	V=51.2	Cr=52.3	Mn=55	
5	(Cu)=63.3	Zn=65.4	— =68	— =72	As=75	Se=79	Br=80	Rh=103, Ru=103.8, Pd=108, Ag=107.9.
6	Rb=85.4	Sr=87.5	Y=89	Zr=90.7	Nb=94.2	Mo=95.9	— =100	
7	(Ag)=107.9	Cd=112	In=113.7	Sn=118	Sb=120.3	Te=125.2	I=126.9	— — —
8	Cs=132.9	Ba=137	La=138.5	Ce=141.5	Di=145	—	—	
9	(—)	—	—	—	—	—	—	Ir=193.1, Pt=194.8, Os=200, Au=196.7.
10	—	—	Yb=173.2	—	Ta=182.8	W=184	—	
11	(Au)=196.7	Hg=200.4	Tl=204.1	Pb=206.9	Bi=208	—	—	— — —
12	—	—	—	Tb=233.4	—	U=239	—	

Mendeleev's system of classifying elements was more elaborate than that of Lothar Meyer's.

- **Recognized the significance of periodicity and used broader range of physical and chemical properties to classify the elements.**
- **Mendeleev relied on the similarities in the empirical formulas and properties of the compounds formed by the elements. He found that some of the elements did not fit in with his scheme of classification if the order of atomic weight was strictly followed.**
- **Ignored the order of atomic weights, thinking that the atomic measurements might be incorrect**

- **Placed the elements with similar properties together.**
- **For example, iodine with lower atomic weight than that of tellurium (Group VI) was placed in Group VII along with fluorine, chlorine, bromine because of similarities in properties.**
- **While arranging the elements of similar properties in the same group, he left unfilled spaces.**
- **He proposed that some of the elements were still undiscovered and, therefore, left several gaps in the table.**
- **For example, both **gallium** and **germanium** were unknown at the time.**

- Mendeleev published his Periodic Table.
- He left a gap under aluminium and a gap under silicon, and called these elements **Eka-Aluminium** and **Eka-Silicon**.
- Mendeleev predicted not only the existence of gallium and germanium, but also described some of their general physical properties. These elements were discovered later.

CONTRIBUTION OF MENDELEEV PERIODIC LAW

- ❖ **Systematic study of the element:** Simplified the study of chemistry of elements. Knowing the properties of one element in group properties of other element in a group can be predicted.

- ❖ **Correction of atomic masses:** Helped in correction of atomic mass. Atomic mass of Be was corrected from 13.5 to 9. Similarly with the help of this table, atomic masses of indium , gold , platinum etc. were corrected.
- ❖ **Prediction of new elements:** Only 56 elements were known. While arranging these elements some gaps were left which corresponded to undiscovered elements. For example, both gallium and germanium were unknown at the time.
- ❖ **Mendeleev published his Periodic Table.** He left a gap under aluminium and a gap under silicon, and called these elements Eka- Aluminium and Eka-Silicon.

Drawbacks of Mendeleev's periodic table

- ❖ **Position of Hydrogen:** Hydrogen resembled the properties of both alkali metals (like lithium) and also that of halogens (like iodine). Hence, the position of hydrogen (whether hydrogen is to be placed with halogens or alkali metals) was not specified.
- ❖ **Position for Isotopes:** Isotopes are the atoms of the same element having same atomic number but different atomic masses

Mendeleev's periodic table was based on arranging elements in increasing order of atomic masses. But isotopes were not included in his periodic table.

- ❖ **Certain elements were arranged in reverse order:** Elements having higher atomic mass were placed in front(or before) the elements with less atomic mass.
- ❖ **Example - Cobalt and Nickel.** Cobalt being more heavier than Nickel ,was placed before Nickel , in the Mendeleev's periodic table.
- ❖ **Some similar elements were separated and dissimilar elements were grouped together:** In the Mendeleev's periodic table, some similar elements were placed in different groups while some dissimilar elements had been grouped together.
- ❖ For example, **copper and mercury** resembled in their properties but they had been placed in different groups.

- ❖ At the same time, elements of group IA such as Li, Na and K were grouped with copper (Cu), silver (Ag) and gold (Au), though their properties are quite different.
- ❖ **Cause of periodicity:** Mendeleev did not explain the cause of periodicity among the elements. He could not explain why having similar properties were being repeated in different elements.
- ❖ **Position of lanthanoids (or lanthanides) and actinoids (or actinides):** The fourteen elements following lanthanum (known as lanthanoids, from atomic number 58-71) and the fourteen elements following actinium (known as actinoids, from atomic number 90 – 103) have not been given separate places in Mendeleev's table.

Modern Periodic Table

Modern Periodic Law :

Properties of elements are a periodic function of their atomic number.

- Modern periodic table contains 18 vertical columns known as groups and 7 horizontal rows known as periods.
- Elements in a group have same number of valence electrons

The structure of MODERN PERIODIC TABLE

MODERN PERIODIC TABLE is very important tool for chemists and scientists.

It is easy to predict the properties of elements by using periodic table.

In **PERIODIC TABLE** all elements *are arranged* from left to right and top to bottom in *order of increasing atomic number*

Atomic number is the order number of elements which is indicated by the **number of protons** of element.

Atomic number = proton number

Hydrogen H 1.008																	Helium He 4.002				
Lithium Li 6.941	Beryllium Be 9.012															Boron B 10.81	Carbon C 12.01	Nitrogen N 14.01	Oxygen O 16.00	Fluorine F 18.99	Neon Ne 20.18
Sodium Na 22.99	Magnesium Mg 24.31															Aluminum Al 26.98	Silicon Si 28.09	Phosphorus P 30.97	Sulfur S 32.07	Chlorine Cl 35.45	Argon Ar 39.95
Potassium K 39.10	Calcium Ca 40.08	Scandium Sc 44.96	Titanium Ti 47.87	Vanadium V 50.94	Chromium Cr 52.00	Manganese Mn 54.94	Iron Fe 55.84	Cobalt Co 58.93	Nickel Ni 58.69	Copper Cu 63.55	Zinc Zn 65.39	Gallium Ga 69.72	Germanium Ge 72.64	Arsenic As 74.92	Selenium Se 78.96	Bromine Br 79.90	Krypton Kr 83.80				
Rubidium Rb	Strontium Sr	Yttrium Y	Zirconium Zr	Niobium Nb	Molybdenum Mo	Technetium Tc	Ruthenium Ru	Rhodium Rh	Palladium Pd	Silver Ag	Cadmium Cd	Indium In	Tin Sn	Antimony Sb	Tellurium Te	Iodine I	Xenon Xe				

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 3.9 Name of elements with atomic number above 100

Atomic number	Temporary Name	Temporary Symbol	Name of the element	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

The Periodic Table of Elements

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period 1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57* La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89** Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

○ Non Metals	● Noble Gases
● Alkali Metals	● Metalloids
● Alkaline Metals	● Halogens
● Transition Metals	● Other Metals
● Rare Earth Elements	

*Lanthanides

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
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**Actinides

90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr
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s-BLOCK

1s	1	2
2s	Li	Be
3s	Na	Mg
4s	K	Ca
5s	Rb	Sr
6s	Cs	Ba
7s	Fr	Ra

H

d-BLOCK

	3	4	5	6	7	8	9	10	11	12
3d	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
4d	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
5d	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
6d	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn

p-BLOCK

	13	14	15	16	17	18
						He
2p	B	C	N	O	F	Ne
3p	Al	Si	P	S	Cl	Ar
4p	Ga	Ge	As	Se	Br	Kr
5p	In	Sn	Sb	Te	I	Xe
6p	Tl	Pb	Bi	Po	At	Rn
7p	Nh	Fl	Mc	Lv	Ts	Og

f-BLOCK

Lanthanoids 4f	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinoids 5f	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Fig. 3.3 The types of elements in the Periodic Table based on the orbitals that are being filled. Also shown is the broad division of elements into METALS (■ ■ ■ ■), NON-METALS (■ ■) and METALLOIDS (■).

s-block

- The elements of Group 1 (alkali metals) and Group 2 (alkaline earth metals) which have ns^1 and ns^2 outermost electronic configuration belong to the s-block.
- They are all reactive metals with low ionization enthalpies.
- They lose the outermost electron(s) readily to form 1+ ion (in the case of alkali metals) or 2+ ion (in the case of alkaline earth metals).
- The metallic character and the reactivity increase as we go down the group. Because of high reactivity they are never found pure in nature.
- The compounds of the s-block elements, with the exception of those of lithium and beryllium are predominantly ionic.

p-Block

- The *p*-Block Elements comprise of those elements belonging to **Group 13 to 18**.
- These together with the *s*-Block Elements are called the **Representative Elements or Main Group Elements**.
- The outermost electronic configuration varies from **ns^2np^1 to ns^2np^6** in each period.
- At the end of each period is a noble gas element with a closed valence shell **ns^2np^6** configuration.
- Preceding the noble gas family are two chemically important groups of non-metals. They are **the halogens** (Group 17) and the **chalcogens** (Group 16).

- These two groups of elements have highly negative electron gain enthalpies and readily add one or two electrons respectively to attain the stable noble gas configuration.
- The non-metallic character increases as we move from left to right across a period and metallic character increases as we go down the group.

p-BLOCK

	13	14	15	16	17	18
						He
<i>2p</i>	B	C	N	O	F	Ne
<i>3p</i>	Al	Si	P	S	Cl	Ar
<i>4p</i>	Ga	Ge	As	Se	Br	Kr
<i>5p</i>	In	Sn	Sb	Te	I	Xe
<i>6p</i>	Tl	Pb	Bi	Po	At	Rn
<i>7p</i>	Nh	Fl	Mc	Lv	Ts	Og

d-block

- These are the elements of Group **3 to 12** in the centre of the Periodic Table.
- These are characterized by the **filling of inner (n- 1) d orbitals** by electrons and are therefore referred to as *d*-Block elements.
- These elements have the general outer electronic configuration **$(n-1)d^{1-10}ns^{1-2}$** .
- They are all metals.
- They mostly form coloured ions, exhibit variable valence (oxidation states), paramagnetism and often used as catalysts.

- However, Zn, Cd and Hg which have the electronic configuration, $(n-1)d^{10} ns^2$ do not show most of the properties of transition elements.
- In a way, transition metals form a bridge between the chemically active metals of s-block elements and the less active elements of Groups 13 and 14 and thus take their familiar name “**Transition Elements**”.

		<i>d</i> -BLOCK									
		3	4	5	6	7	8	9	10	11	12
<i>3d</i>		Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
<i>4d</i>		Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
<i>5d</i>		La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
<i>6d</i>		Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn

f - block

- The two rows of elements at the bottom of the Periodic Table, are called the **Lanthanoids**, Ce($Z = 58$) – Lu($Z = 71$) and **Actinoids**, Th($Z = 90$) – Lr ($Z = 103$)
- They are characterized by the outer electronic configuration $(n-2) f^{1-14} (n-1) d^{0-1} ns^2$.
- The last electron added to each element is filled in the last but two $(n - 2) f$ - orbital.
- These two series of elements are hence called **the Inner-Transition Elements (f - Block Elements)**.
- Within each series, the properties of the elements are quite similar.

- The chemistry of the early actinoids is **more complicated** than the corresponding lanthanoids.
- This is due to the large number of oxidation states possible for these actinoid elements.
- Actinoid elements are radioactive.
- Many of the actinoid elements have been made only in nanogram quantities or even less by nuclear reactions and their chemistry is not fully studied.
- The elements after uranium are called **Transuranic Elements**.

f-BLOCK

Lanthanoids <i>4f</i>	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinoids <i>5f</i>	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

SLATER'S RULE

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Slater's Rule

INTRODUCTION:

- ❖ In 1930, a scientist *J.C. Slater* proposed a set of empirical rules to understand the concept of Effective Nuclear Charge and to calculate the Screening Constant or Shielding Constant.
- ❖ He proposed a formula for calculation of Effective Nuclear Charge

$$Z_{\text{eff}} = Z - S$$

where S is the Slater's screening constant ,
 Z is the Nuclear charge

Prior to explaining Slater's rules , certain terms like Nuclear Charge, Shielding Effect and Effective Nuclear Charge have to be Understood.

What is Nuclear Charge?

It is the charge on the nucleus with which it attracts the electrons of the atom. Basically, the Nuclear Charge is said to be equal to the Atomic Number(i.e. the number of protons) of an atom. It is denoted by the symbol **Z**.

What is Shielding Effect?

In case of Multielectron atoms, as the orbitals are filled up, the electrons in the Inner orbitals shield the electrons in the Outer orbitals from the Nucleus.

So, the electrons in the Outer orbitals do not feel the full force or charge of the nucleus. Thus, the reduction of nuclear charge on the Outermost electrons is called **Shielding Effect or Screening Effect**.

Shielding Effect is defined as a measure of the extent to which the intervening electrons shield the outer electrons from the nuclear charge. It is denoted by the symbol **S**

What is Effective Nuclear Charge ?

Effective Nuclear Charge is the actual charge felt by the outer electrons after taking into account shielding of the electrons. It is denoted by the symbol **Z^* or Z_{eff}**

Slater's Rules:

1) Write the electron configuration for the atom using the following design;

**(1s)(2s,2p)(3s,3p) (3d) (4s,4p) (4d) (4f) (5s,5p) (5d)
(5f) (6s,6p)..... etc.**

2) Any electrons to the right of the electron of interest contributes nothing towards shielding.

3) All other electrons in the same group as the electron of interest shield to an extent of 0.35 nuclear charge units irrespective of whether the electrons are in s, p, d, or f orbitals.

- 4) In case of **1s electron shielding another 1s** electron the screening constant value is taken to be **0.30**.
- 5) **If the electron of interest is an s or p electron:** All electrons with one less value i.e. **(n - 1) value** of the principal quantum number shield to an extent of **0.85 units** of nuclear charge. All electrons with two or more less values i.e. **(n - 2, n - 3, n - 4 etc.)** values of the principal quantum number shield to an extent of **1.00 units**.
- 6) **If the electron of interest is an d or f electron:** All electrons to the left shield to an extent of **1.00 unit** of nuclear charge.
- 7) **Sum** the shielding amounts from **steps 2 through 5** and **subtract** from the nuclear charge value to obtain the effective nuclear charge value.

Calculate Z^* for a valence electron in fluorine ($Z = 9$).

- Electronic configuration of fluorine is $1s^2, 2s^2, 2p^5$
- Grouping it acc. to Slater's rule : $(1s^2)(2s^2, 2p^5)$
- Rule 2 does not apply;
- Now, one electron out of the 7 valence electrons becomes the electron of interest. The other remaining 6 valence electrons will contribute 0.35 each towards shielding.
- The electrons in $(n - 1)$ orbitals i.e. 1s orbital will contribute 0.85 each towards shielding.

S = 0.35 x (No. of electrons in the same shell i.e. n orbital) + 0.85 x (No. of electrons in the (n – 1) shell)

$$\mathbf{S} = 0.35 \times 6 + 0.85 \times 2 = 3.8$$

$$\mathbf{Z}^* = \mathbf{Z} - \mathbf{S} = 9 - 3.8 = 5.2 \text{ for a valence electron.}$$

□ Calculate Z^* for a 6s electron in Platinum ($Z = 78$)

❖ The electronic configuration of Platinum ($Z = 78$) is $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^8$

.

❖ Grouping it acc. to Slater's rule :

$(1s^2)(2s^2, 2p^6)(3s^2, 3p^6) (3d^{10}) (4s^2, 4p^6) (4d^{10}) (4f^{14})$
 $(5s^2, 5p^6) (5d^8) (6s^2)$

❖ Rule 2 does not apply;

❖ Now, one electron out of the two valence electrons becomes the electron of interest. The other remaining valence electron will contribute 0.35 towards shielding.

❖ The electrons in $(n - 1)$ orbitals i.e. $(6 - 1) = 5^{\text{th}}$ orbital will contribute 0.85 each towards shielding.

❖ The electrons in $(n - 2)$, $(n - 3)$, $(n - 4)$ etc. orbitals { i.e. $(6 - 2)$, $(6 - 3)$, $(6 - 4)$etc.} orbitals will contribute 1.00 each towards shielding.

S = $0.35 \times (\text{No. of electrons in the same shell i.e. } n \text{ orbital}) + 0.85 \times (\text{No. of electrons in the } (n - 1) \text{ shell}) + 1.00 \times (\text{No. of electrons in the } (n - 2), (n - 3)\text{.....etc. orbitals})$

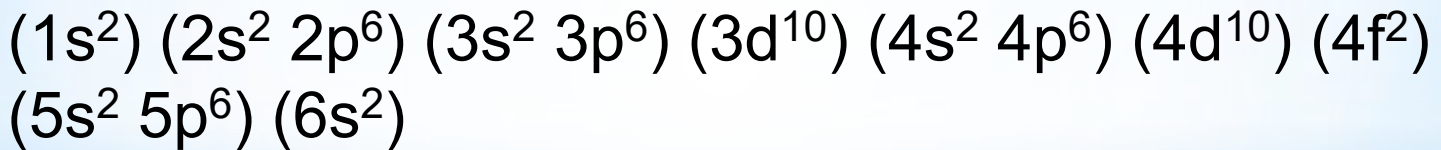
$$\mathbf{S} = 0.35 \times 1 + 0.85 \times 16 + 60 \times 1.00 = 73.95$$

$$\mathbf{Z^*} = \mathbf{Z} - \mathbf{S} = 78 - 73.95 = 4.15 \text{ for a valence electron.}$$

❑ Calculate the Effective Nuclear Charge for one of the 4f electrons of Cerium (Z = 58)

❑ The electronic configuration of Cerium is :
 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^2 5s^2 5p^6 6s^2$

❑ Grouping the orbitals acc. to Slater's :



❑ As we have to calculate the effective nuclear charge of 4f electrons of Cerium, the electrons lying after the 4f electron will not contribute to shielding

Now,

$$S = 0.35 \times (\text{No. of electrons in the same orbital} + 1.00 \times (\text{all the electrons in the lower orbitals}))$$

$$S = 0.35 \times 1 + 1.00 \times 46 = 46.35$$

$$Z_{\text{eff}} = Z - S = 58 - 46.35 = 11.65$$

□ Calculate the Effective Nuclear Charge in the periphery of Nitrogen (Z = 7)

In order to calculate the Effective Nuclear Charge in the periphery of an atom or ion, the shielding of nuclear charge by all the electrons present in the electronic configuration of the atom or ion.

□ The electronic configuration of Nitrogen is $1s^2 2s^2 2p^3$

□ Grouping acc. to Slater's rule : $(1s^2) (2s^2 2p^3)$

□ $S = 0.35 \times (\text{No. of electrons in the same orbital}) + 0.85 \times (\text{No. of electrons in the } (n - 1) \text{ orbital})$

$$S = 0.35 \times 5 + 0.85 \times 2 = 3.45$$

□ Z_{eff} in the periphery of N - atom = $Z - S = 7 - 3.45 = 3.55$

APPLICATIONS OF SLATER'S RULE

- ❖ It provides a quantitative justification for the sequence of orbitals in the energy level diagram.
- ❖ It helps to explain the filling of ns - orbital ($4s, 5s, 6s$ etc. – orbitals) prior to the filling of $(n - 1)d$ orbital ($3d, 4d, 5d \dots$ etc.).
- ❖ Let us consider the case of **Potassium ($Z = 19$), in which the last electron is supposed to enter $3d$ orbital instead it goes to $4s$ orbital**

First lets calculate the charge if the last electron enters 4s orbital

The electronic configuration of **Potassium** acc. to Slater is



- ❖ As the effective nuclear charge on electron in 4s orbital has to be calculated, the electrons in the same orbital i.e. n orbital will contribute 0.35 each, the electrons in (n – 1) orbital i.e. 3s and 3p orbitals will contribute $S = 0.85$ each and all the electrons in (n – 2, n – 3.... etc.) orbitals i.e. (2s, 2p, 1s) orbitals will contribute $S = 1.00$ each.

So,

$$\begin{aligned} S &= 0 \times 0.35 + 8 \times 0.85 + 10 \times 1.00 \\ &= 16.80 \end{aligned}$$

Therefore, Effective Nuclear Charge

$$Z^* = Z - S = 19 - 16.80 = 2.20$$

Now, Let us assume that the last electron enters the 3d orbital rather than 4s orbital,

Then the configuration acc. to Slater is



Here the d electron is under Interest, so the electrons in the same orbital i.e. 3d orbital will contribute $S = 0.35$ each, where as the electrons in all the other orbitals i.e. (3s,3p,2s, 2p, 1s) will all contribute $S = 1.00$ each.

So,

$$S = 0 \times 0.35 + 18 \times 1.00 = 18.00$$

Therefore Effective Nuclear Charge

$$Z^* = Z - S = 19.00 - 18.00 = 1.00$$

On comparing the Effective Nuclear Charge of both 4s and 3d orbitals, we see that the 4s electron is under the influence of greater Effective Nuclear charge ($Z_{\text{eff}} = 2.20$) as compared to 3d electron ($Z_{\text{eff}} = 1.00$) in Potassium atom.

□ So, the electron in 4s orbital will be more attracted by the nucleus and will have lower energy than the 3d electron.

□ Thus, the last electron will enter in the 4s orbital, rather than the 3d orbital in case of Potassium atom.

Slater's rule explain why 4s electrons are lost prior to 3d electrons during cation formation in case of Transition elements.

Let us consider the case of *Vanadium* (Z=23)

The electronic configuration of Vanadium is



- ❖ After losing 2 electrons, the electronic configuration of V^{2+} should be



And not



- ❖ The above Electronic Configuration can be explained by Slater's rules

The Effective Nuclear Charge for 4s electron is calculated as:



Now, one of the electrons of the 4s orbital becomes electron of interest. The second electron however, will contribute towards shielding effect

$$S = 1 \times 0.35 + 11 \times 0.85 + 10 \times 1.00 = 19.70$$

$$\text{So, } Z_{\text{eff}} = Z - S = 23 - 19.70 = 3.30$$

Now, lets calculate the Effective Nuclear Charge for a 3d electron

According to Slater's rule



As 4s orbital lies after the electron under interest it will contribute nothing towards shielding.

One electron of 3d orbital becomes electron of interest. The other two 3d electrons will contribute towards shielding effect.

$$S = 2 \times 0.35 + 18 \times 1.00 = 18.70$$

$$Z_{\text{eff}} = Z - S = 23 - 18.70 = 4.30$$

- Comparing the Effective Nuclear Charge of both the 3d and the 4s electron, it is seen that Effective Nuclear Charge on 3d electron is 4.30 whereas 4s electron has 3.30.
- The force of attraction experienced by the 3d electrons is more as compared to 4s electrons.
- The 3d electrons are more tightly held to the nucleus than 4s electrons.
- Thus, the 4s electrons are removed in preference to 3d electrons.

- The force of attraction experienced by the 3d electrons is more as compared to 4s electrons.
- The 3d electrons are more tightly held to the nucleus than 4s electrons.
- Thus, the 4s electrons are removed in preference to 3d electrons.

It helps to explain why size of a cation is always smaller than its neutral atom.

- Let's take the example of **Lithium atom** and **Lithium ion**
- The Electronic configuration of **Lithium atom** is $1s^2 2s^1$
- Write it according to Slater's $(1s^2) (2s^1)$
- As 2s orbital has only one electron, it becomes the electron of interest.

- Only the 1s electrons will contribute towards shielding.

$$S = 2 \times 0.85 + 0 \times 0.35 = 1.70$$

$$Z_{\text{eff}} = Z - S = 3 - 1.70 = 1.30$$

In case of Lithium (Li^+) Ion

- The Electronic Configuration of Li^+ is $1s^2$
- Grouping acc. to Slater : ($1s^2$)
- In Case of Li^+ , one of the 1s electrons becomes electron of interest and the other 1s electron contributes towards shielding.

$$S = 1 \times 0.30 = 0.30$$

$$Z_{\text{eff}} = Z - S = 3 - 0.30 = 2.70$$

- Comparison of the Effective Nuclear Charge of Li atom ($Z_{\text{eff}} = 1.30$) and Li^+ ion ($Z_{\text{eff}} = 2.70$), shows that Effective nuclear charge on Li^+ ion is more than Li atom.
- So , the size of Li^+ ion is smaller than Li atom.

It explains why a anion is always larger than its neutral atom

- Taking the example of **Chlorine atom** and **Chlorine ion**.
- In case of **Chlorine atom** ($Z = 17$), the electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$
- Grouping it acc. to Slater's rule
 $(1s^2) (2s^2 2p^6) (3s^2 3p^5)$
 $S = 6 \times 0.35 + 8 \times 0.85 + 2 \times 1.00 = 10.90$
 $Z_{\text{eff}} = Z - S = 17 - 10.90 = 6.10$

In case of Chlorine ion Cl^-

✓ The electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6$

✓ Grouping it acc. to Slater's rule

$$(1s^2) (2s^2 2p^6) (3s^2 3p^6)$$

$$\mathbf{S} = 7 \times 0.35 + 8 \times 0.85 + 2 \times 1.00 = 11.25$$

$$\mathbf{Z}_{\text{eff}} = \mathbf{Z} - \mathbf{S} = 17 - 11.25 = 5.75$$

✓ Comparison of the Effective Nuclear Charge of Cl atom ($Z_{\text{eff}} = 6.10$) and Cl^- ion ($Z_{\text{eff}} = 5.75$), shows that Effective nuclear charge on Cl atom is more than Cl^- ion.

✓ So, the size of Cl^- ion is larger than Cl atom.

* LIMITATIONS OF SLATER'S RULE

- ❑ Slater grouped both s and p orbitals together for calculating effective nuclear charge, which is incorrect. This is because radial probability distribution curves show that s orbitals are more penetrating than p orbitals. So, the s orbitals should shield to a greater extent as compared to p orbital.
- ❑ According to Slater, all the s, p, d and f electrons present in shell or energy level lower than $(n - 1)$ shell will shield the outer n electrons with equal contribution of $S=1.00$ each. This is not justified as energetically different orbitals should not contribute equally.
- ❑ Slater rules are less reliable for heavier elements

CAUSE OF PERIODICITY:

PERIODICITY: The repetition of similar properties of elements after regular intervals is known as periodicity.

Causes of periodicity:- The properties of elements are the periodic repetition of similar electronic configuration of elements as the atomic number increases.

Symbol	Atomic No.	Electronic configuration
Li	3	$1s^2 2s^1$ or $[\text{He}]2s^1$
Na	11	$1s^2 2s^2 2p^6 3s^1$ or $[\text{Ne}]3s^1$
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or $[\text{Ar}] 4s^1$
Rb	37	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ or $[\text{Kr}] 5s^1$
Cs	55	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4d^{10} 5s^2 5p^6 6s^1$ or $[\text{Xe}] 6s^1$
Fr	87	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^7 7s^1$ or $[\text{Rn}] 7s^1$

Thus it is because of similarity in electronic configuration that all the elements have similar properties

PERIODIC PROPERTIES

Physical and chemical properties of some elements show periodic variation with atomic number. Some of the properties are directly linked with electronic configuration, while some are indirectly linked.

ATOMIC RADIUS:

- The atomic radius of an element is a measure of the size of its atoms, usually the mean or typical distance from the center of the nucleus to the boundary of the surrounding shells of electrons.**
- However, absolute size of an atom is difficult to define because**

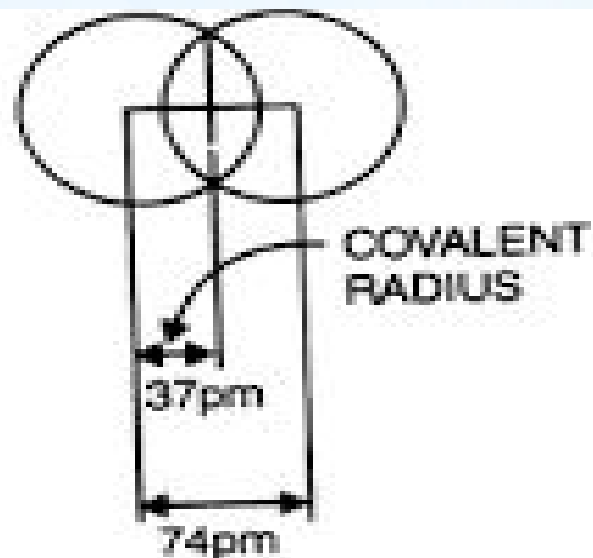
- **Since the boundary is not a well-defined physical entity, the probability of finding electron is never zero even at larger distances.**
- **It is not possible to isolate an atom and measure its radius.**
- **The probability distribution of an atom is also affected by the presence of the other atoms in its neighborhood.**
- **The size of atom may change in going from one environment to another.**
- **Thus we can only arbitrarily define atomic radius as the effective size which means the distance of closet approach of one atom from another atom in a given bonding situation.**

- The atomic radius can be defined in terms of : **Covalent radius, Ionic radius, Metallic radii and Van der Waal radius.**

Covalent radius:

In case of **homo diatomic molecule** the atomic radii is given as

“One half of the distance between the centers of two nuclei of two similarly bonded atoms in a molecule”



∴ Covalent radius of hydrogen atom

$$= \frac{74}{2} = 37 \text{ pm}$$

Similarly, covalent radius of chlorine atom

$$= \frac{1}{2} [198] = 99 \text{ pm}$$

In case of **heteroatomic or heteronuclear molecule** the covalent radii of different atoms joined by single covalent bond.

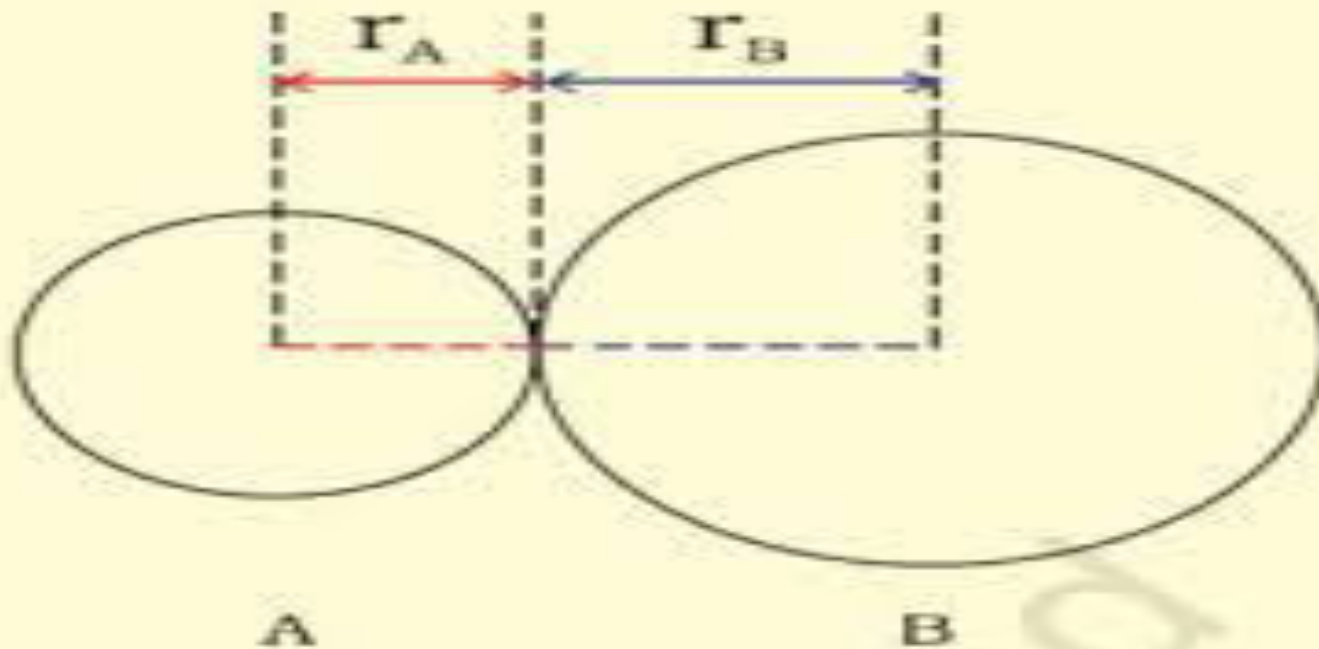


Fig. 4.1 The bond length in a covalent molecule AB.

$R = r_A + r_B$ (R is the bond length and r_A and r_B are the covalent radii of atoms A and B respectively)

Atomic radii of atom is given by

$$r_{AB} = r_A + r_B$$

if the **difference in electronegativity** of the combining atoms is **less**.

If the **difference in electronegativity** of the combining atoms is **high** then the atomic / covalent radii is given as

$$r_{AB} = r_A + r_B - |X_A - X_B|$$

Multiple Bond : The atomic radii of atoms which do not form diatomic molecules with single bonds are calculated by indirect method.

For example, the single bond radii of N_2 and O_2 **are not obtained**

by

$$r_{AB} = r_A + r_B - |X_A - X_B|$$

as both nitrogen and oxygen molecule contain multiple bonds.

The **single bond atomic radii of the nitrogen** can be calculated by indirect method in those compounds in which nitrogen is bounded to another atom by a single bond.

For eg. **C-N bond** in Methyl amine (H_3CNH_2) has been found to be 1.47 \AA . Since the atomic radius of carbon is 0.77 \AA (obtained from C-C bond distance in diamond), **the atomic radius of nitrogen is $(1.47 - 0.77) = 0.70 \text{ \AA}$**

Bond	radius	Bond	radius
C-C	0.77 \AA	N-N	0.70 \AA
C=C	0.74 \AA	N=N	0.60 \AA
C≡C	0.70 \AA	N≡N	0.55 \AA

Hybridization and atomic radius

As seen from the table, the distance / bond length between the atoms goes on decreasing with increase in the number of bond between atoms.

This can be explained on the basis of hybridization.

<i>Bond</i>	<i>Length (pm)</i>
C - C	154
C=C	134
C≡C	120
C - N	147
C=N	132
C≡N	116
C - O	143
C=O	120
C≡O	113
N - N	145
N=N	125
N≡N	110

The decrease in the radius of carbon may be attributed to **increase in s character** while moving from sp^3 , sp^2 to sp .

With increase in s-character, the electron pair between the two C atom **comes closer to the nucleus** resulting in greater attraction between the two C atoms and hence atomic radii decreases.

Hybridization	s-character	example	C-H distance	Radius of C
sp^3	25%	$CH_3 - CH_3$	1.14 Å	0.77 Å
sp^2	33%	$CH_2 = CH_2$	1.11 Å	0.74 Å
sp	50%	$CH \equiv CH$	1.07 Å	0.70 Å

❖ **The atomic radii decreases along a period (from left to right).**

- The principal energy level is the same all across the period i.e. the electrons are being added in the same energy level (n).
- Electrons being added to the same energy level are not shielded very much by the other electrons in the same energy level.
- Whereas, the nuclear charge increases from left to right due to addition of electrons.
- So, the outer electrons experience more nuclear charge and are attracted more by the nucleus, resulting in decrease in atomic radii.

❖ **The atomic radii increases on going down the group (top to bottom)**

- On going from top to bottom in a group in periodic table, the outermost electrons are placed in higher energy levels (n).
- Higher the energy level in which an electron is placed, larger will be the atomic radius of the shell.
- Nuclear charge also increases on going down the group but the increase in effective nuclear charge is not large.
- So, the increase in atomic radii/volume is much more (or predominates) than the increase in effective nuclear charge.
- As a result nuclear charge decreases down the group.

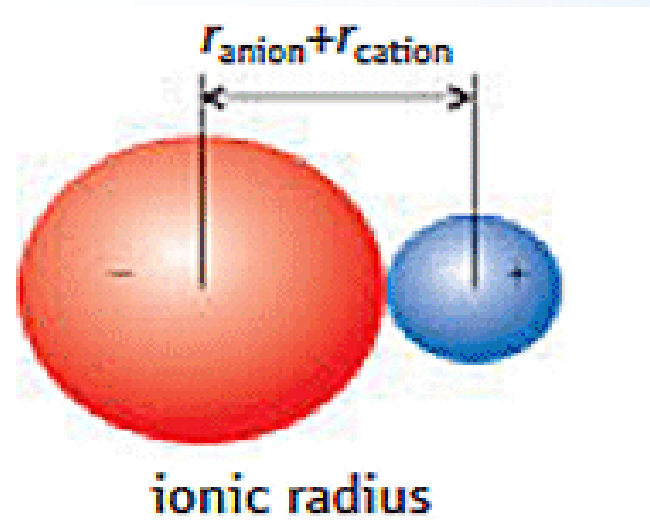
IONIC RADII:

- The ionic radii corresponds to radii of ions in ionic crystal.
- Ions are formed as a result of addition or removal of electrons from the outermost shell of atom.
- **Cations:** are formed by loss / removal of electrons
- **Anions:** Formed by addition / gain of electrons.
- **Ionic radii** is defined as “the effective distance from nucleus of the ion up to which it has an influence in the ionic bond.”
- The distance between the nuclei of a adjacent positive and negative ions in the crystal can be measured by X – ray method.

- If ions are considered spheres, the internuclear distance may be taken equal to the sum of radius of cation and anion.
- For eg. the internuclear distance of Na^+Cl^- is 2.76 \AA which corresponds to

$$r_{\text{Na}^+} + r_{\text{Cl}^-} = 2.76 \text{ \AA}$$

- Unless we know the radii of any one ion we cannot calculate the radii of the other



Measurement of ionic radius:

- Large number of methods are known but the most accepted method for calculating ionic radii is the **Pauling's method**.
- This method is limited to **univalent isoelectronic ions**.
- The method is based on the observed internuclear separation in four crystals: **NaF, KCl, RbBr and CsI**
- Pauling selected these compounds because:
 - i. They have approximately the same radius ratio ($r_+ / r_- = 0.75$)
 - ii. They have approximately equal ionic character.
 - iii. In each case the cation and the anion are isoelectronic (i.e. : they have the same number of electrons)

- Pauling also made **two more assumptions**:
 - i. The cation and anions are assumed to be in contact with each other ,so that

$$\mathbf{r}_+ + \mathbf{r}_- = \mathbf{R} \text{ (Internuclear distance)}$$
 - ii. The radius of the ion is inversely proportional to the effective nuclear charge of the ion (\mathbf{Z}_{eff})

where $\mathbf{Z}_{\text{eff}} = \mathbf{Z} - \mathbf{S}$

$$\mathbf{r}_{\text{ion}} \propto \mathbf{1} / \mathbf{Z}_{\text{eff}}$$

In other words

$$\mathbf{r}_{\text{ion}} \propto \mathbf{1} / \mathbf{Z} - \mathbf{S}$$

$$\mathbf{r}_{\text{ion}} = \mathbf{C} / \mathbf{Z} - \mathbf{S}$$

where \mathbf{C} is the constant of proportionality & has the same value for both the ions

Let us consider the example of NaF

$$r_{\text{Na}^+} + r_{\text{F}^-} = 2.31 \text{ \AA} \quad (\text{Determined by X-ray technique})$$

Both Na^+ and F^- are isoelectronic (have same no. of electrons)

Calculated Z_{eff} according to Slater rule

$$Z_{\text{eff}} \text{ for } \text{Na}^+ = 11 - 4.5 = 6.5$$

$$Z_{\text{eff}} \text{ for } \text{F}^- = 9 - 4.5 = 4.5$$

Acc. to Pauling

$$r_{\text{Na}^+} = \frac{C}{6.5} \quad \text{and} \quad r_{\text{F}^-} = \frac{C}{4.5}$$

$$\frac{C}{6.5} + \frac{C}{4.5} = 2.31 \text{ \AA}$$

$$4.5C + 6.5 C = 2.31 \times 6.5 \times 4.5$$

$$11.00C = 70.51$$

$$C = 6.41$$

Using values of C, the radii of Na^+ and F^- can be calculated

$$r_{\text{Na}^+} = \frac{C}{6.5} \quad \text{and} \quad r_{\text{F}^-} = \frac{C}{4.5}$$

$$r_{\text{Na}^+} = \frac{6.41}{6.5} \quad \text{and} \quad r_{\text{F}^-} = \frac{6.41}{4.5}$$

$$r_{\text{Na}^+} = 0.95 \text{ \AA} \quad \text{and} \quad r_{\text{F}^-} = 1.36 \text{ \AA}$$

SIZE OF CATION

❖ **Size of cation is less than its neutral /parent atom.**

- This is due to removal of electrons from the valence shell of an atom.
- The magnitude/amount of the nuclear charge remains the same where as the number of electrons decreases.
- As a result the nuclear charge now acts on lesser number of electrons.
- Due to which the electrons are more strongly attracted and pulled towards nucleus. So, the size of cation decreases.

Atomic radius of Na (Z = 11)	1.54 Å	Ionic radii of Na⁺ (Z = 11)	0.95Å
Atomic radius of K (Z = 19)	1.96Å	Ionic radii of K⁺ (Z = 19)	1.33Å

SIZE OF ANION

- ❖ **Size of anion is more than a neutral / parent atom.**
- This due to addition of electrons occur in the valence shell of an atom.
- The magnitude/amount of the nuclear charge remains the same where as the number of electrons increases.
- As a result the nuclear charge now acts on the more number of electrons.
- The effective nuclear charge decreases due to which the electrons cloud is less tightly held by the nucleus i.e. they are less strongly attracted and pulled towards nucleus. So, the size of anion is more

Atomic radii of Cl (Z = 17)	0.99Å	Ionic radii of Cl ⁻ (Z = 17)	1.81Å
Atomic radii of Br (Z = 35)	1.14Å	Ionic radii of Br ⁻ (Z = 35)	1.95Å

Periodic trends of ionic radii in period

For isoelectronic ions (ions having same number of electrons but different number of protons) the ionic radii decreases as the nuclear charge increases.

Reason:

In isoelectronic series, with increase in nuclear charge, the Z_{eff} increases. More the Z_{eff} , more will be the nuclear charge on the outermost electron & smaller will be the size.

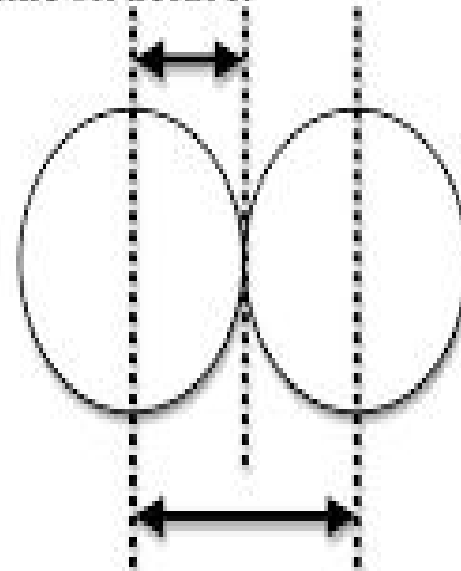
Ion	Atomic number of atom	No. of electrons	$Z_{\text{eff}} = Z - \sigma$	Ionic radii
C^{4-}	6	10	$6 - 4.15 = 1.85$	2.60
N^{3-}	7	10	$7 - 4.15 = 2.85$	1.71
O^{2-}	8	10	$8 - 4.15 = 3.85$	1.40
F^{-}	9	10	$9 - 4.15 = 4.85$	1.36

Ion	Atomic number of atom	No. of electrons	$Z_{\text{eff}} = Z - \sigma$	Ionic radii
			$10 - 4.15 = 5.85$	0.95
Na^+	11	10	7.85	0.65
Mg^{2+}	12	10	8.85	0.50
Al^{3+}	13	10	9.85	0.42
Si^{4+}	14	10	10.85	0.34
P^{5+}	15	10	11.85	0.29
S^{6+}	16	10	12.85	0.26
Cl^{7+}	17	10		

METALLIC RADIUS

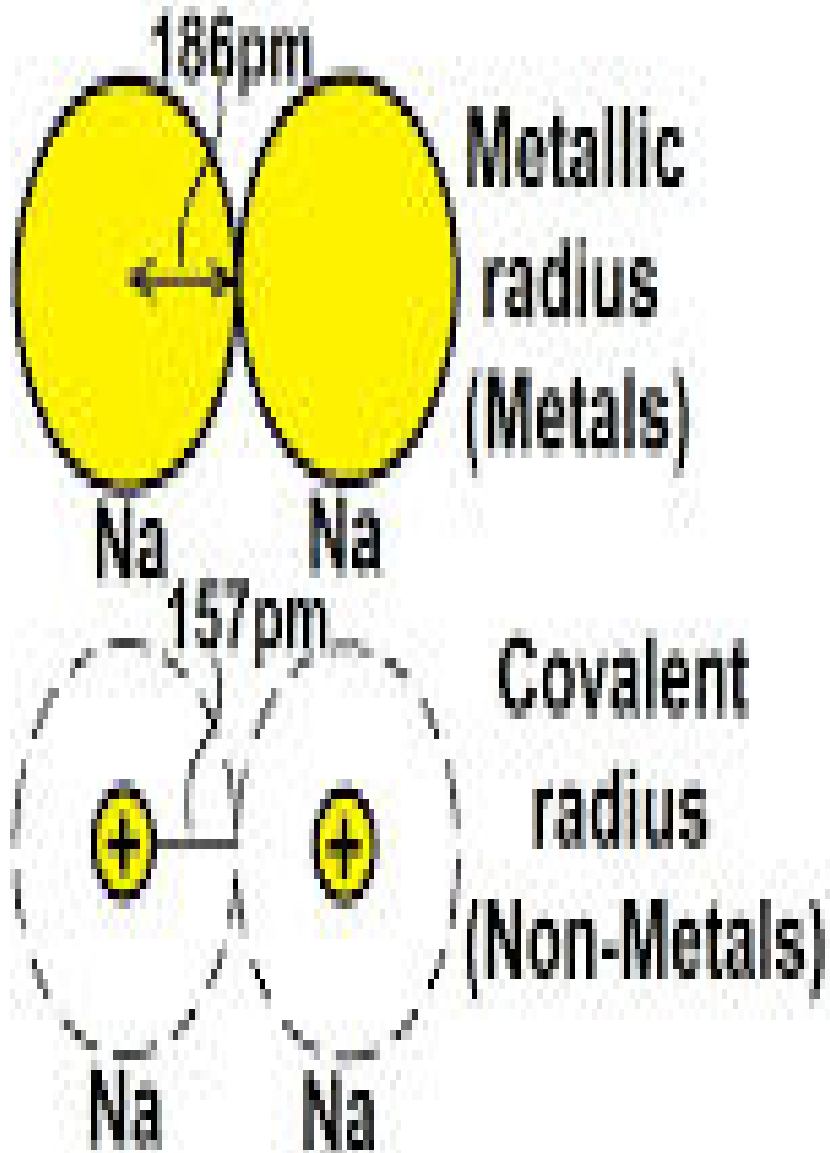
- The *metallic radius* is the *radius* of an atom joined by *metallic bond*.
- The *metallic radius* is half of the total distance between the nuclei of two adjacent atoms in a *metallic cluster*.
- Since metal is made up of group of atoms of the same element, the distance of each atom will be the same

This is the metallic radius. It is $\frac{1}{2}$ the distance between the nuclei of two atoms in a crystalline structure.



This is the distance between the nuclei of two metallic atoms.

$$r_{\text{metallic}} = \frac{\text{Internuclear distance between two adjacent atoms in metallic lattice}}{2}$$

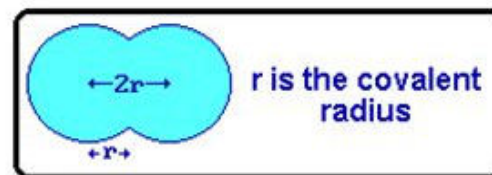
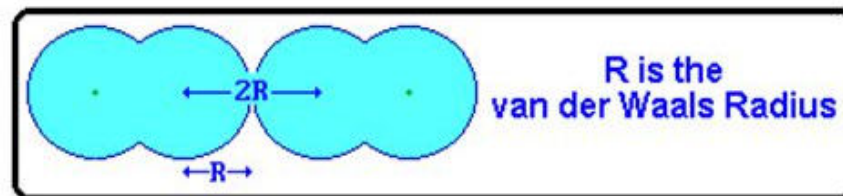


Metallic bond is weaker than covalent bond and as a result metal atoms in metallic lattice are not drawn as close to each other as in a covalent bond.

Van der Waals radius

Van der Waals Radius & Covalent Radius

- Van der Waals radius is one half of the distance between the nuclei of two atoms in adjacent molecules.
- Covalent radius is one half of the distance between two atoms in the same molecules.
- Van der Waals' radius of a non-metal is always larger than the corresponding covalent radius because the covalent bond is much stronger than van der Waals' forces.



- In covalent bond the atoms come closer due to overlapping of orbitals.
- On the other hand van der Waal forces exist between non bonded atoms or molecules.
- Van der Waal forces are weak as the atoms/molecules are at larger distances.
- Thus the internuclear distances in case of atoms held by Van der Waal forces is more than that between covalently bonded atoms.

Covalent and van der Waals radii in A⁰ of some elements

	N	O	F	Ne	P	S	Cl	Ar
Covalent	0.77	0.74	0.72	-	1.10	1.04	0.99	-
Van der waal	1.50	1.40	1.35	1.31	1.80	1.90	1.80	1.74

COMPARISON OF DIFFERENT RADII

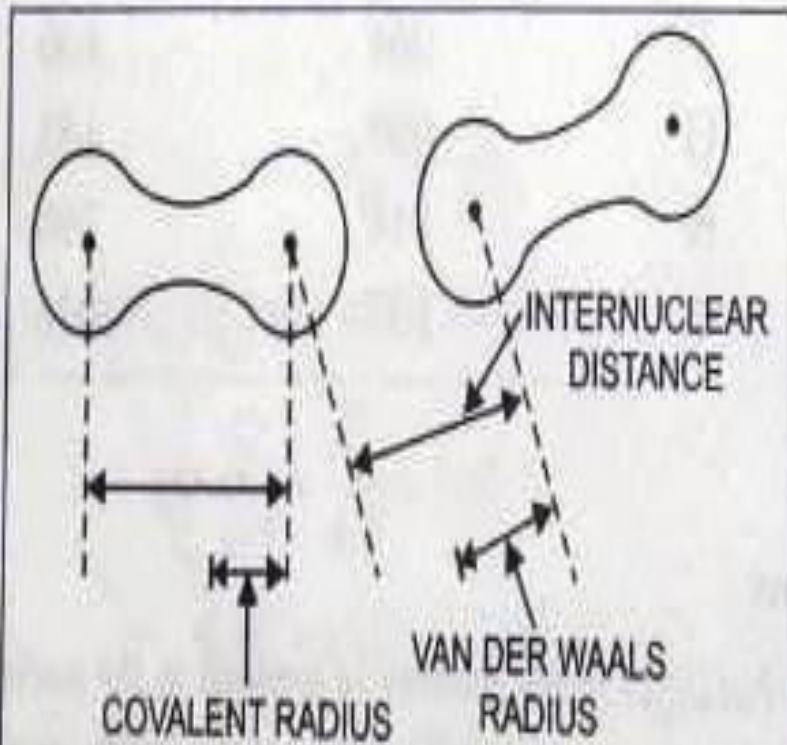


FIGURE 2.4.

Covalent and van der Waals radius.

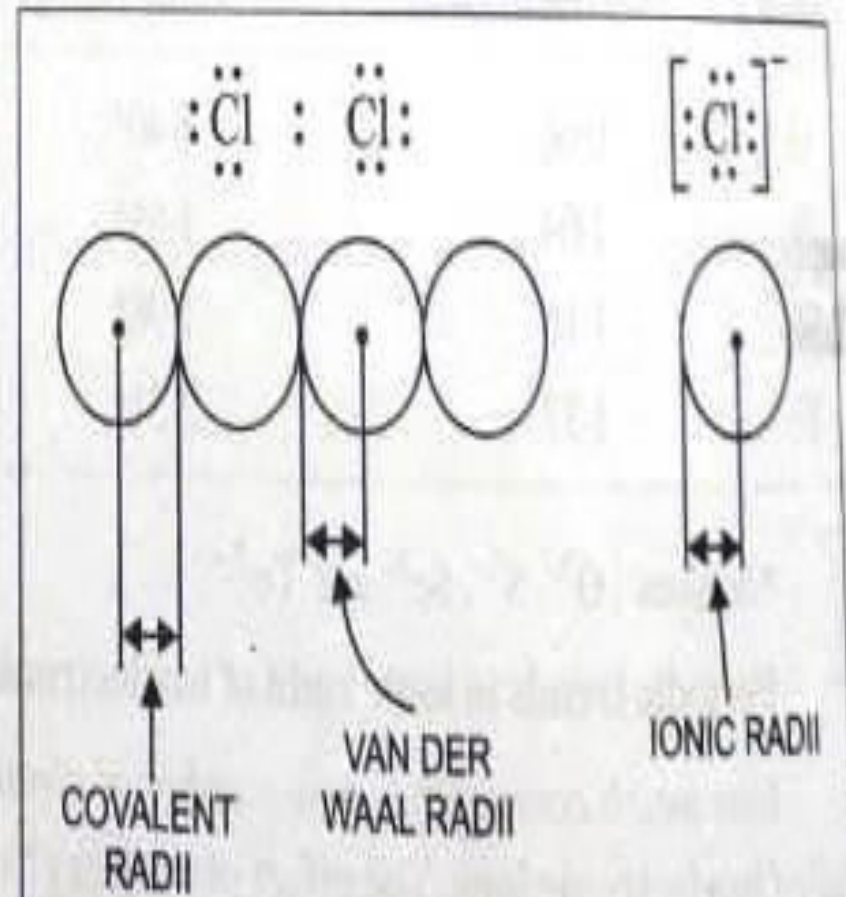


FIGURE 2.5.

Comparisons of covalent, van der waal and ionic radii.

IONIZATION ENERGY

- It is a quantitative measure of the tendency of an element to lose electron.
- It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.



- Ionization energy is also called **ionization potential** because it is measured as the amount of potential needed to remove the most loosely held electron from gaseous atom.
- Smaller the value of ionization energy, the easier it is to remove electron.

Factors on which ionization energy depends

1. Size of an atom:

- Ionization energy depends upon the distance between the nucleus and electron that has to be removed.
- The attractive force between the electron and the nucleus is inversely proportional to the distance between them.
- As size increases, the outer most electrons are less tightly held by the nucleus. So it becomes easier to remove the electron.
- Hence ionization energy decreases with increase in atomic size

2. Charge on the nucleus:

- The attractive force between the nucleus and the electrons increases with increase in nuclear charge.
- Force of attraction is directly proportional to the product of charges on nucleus and that of electrons.
- With increase in nuclear charge it becomes more difficult to remove an electron and IE increases.

3. Screening effect of the inner electrons:

- In multielectron atoms, the outermost electrons are shielded or screened from the nucleus by the inner electrons. This is called screening or shielding effect

- As the result the outermost electron does not feel the full nuclear charge.
- The actual nuclear charge felt is $Z_{\text{eff}} = Z - S$
- If no. of electrons in the inner shell is more, the screening effect will be more.
- As a result the nucleus-electron attraction will be less and Ionization energy will be less.

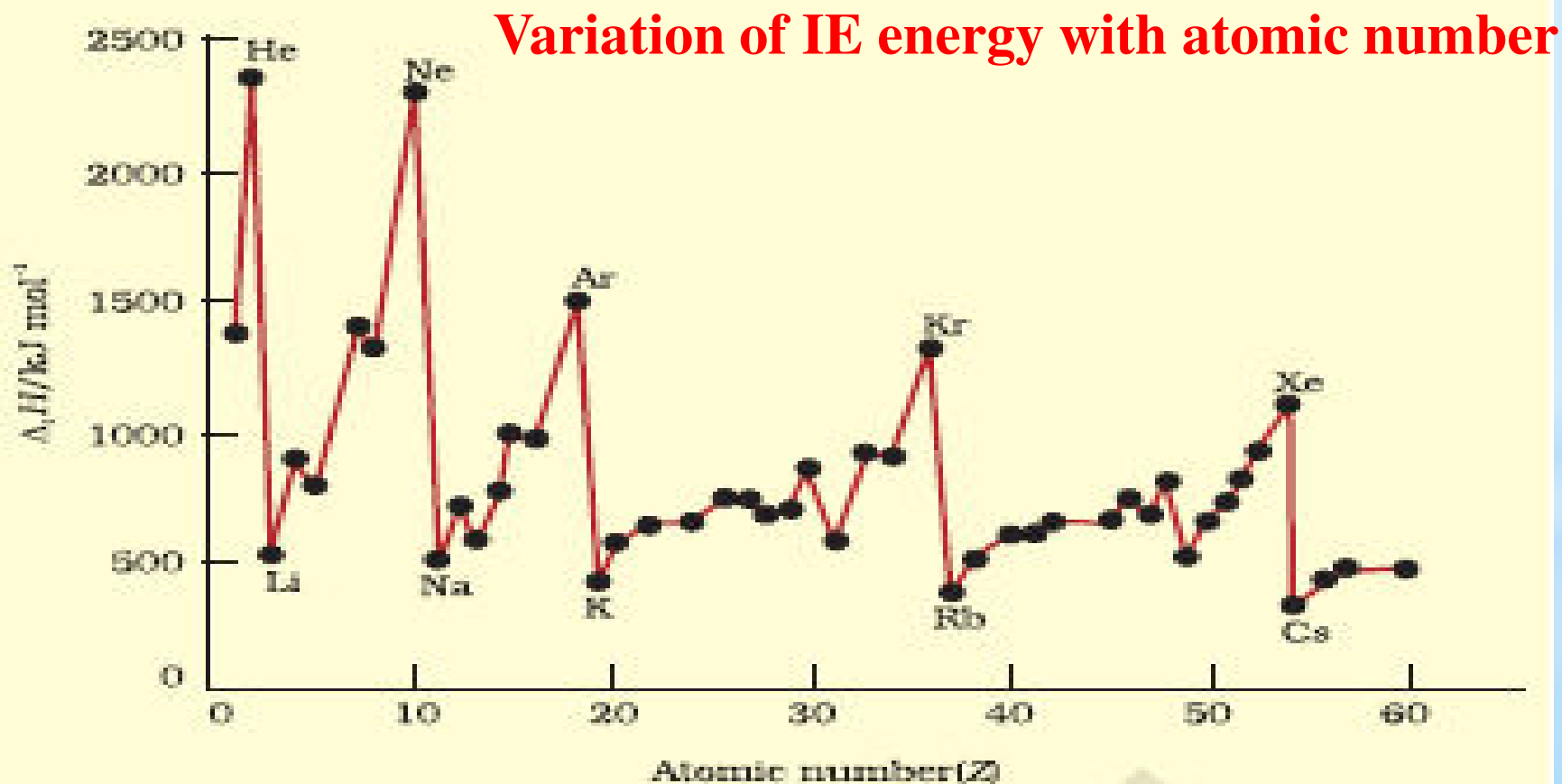
4. Penetration effect:

- In multielectron atoms, the penetration of a $2s$ electron to the nucleus is more than that of a $2p$ -electron.

- s-orbitals have the maximum probability of being near the nucleus.
- The penetrating power decreases as $s > p > d > f$.
- If the penetration of the electron is more, it will be closer to the nucleus and will be firmly held which will increase the ionization energy.
- For the same shell the required ionization energy to remove an electron from an s orbital is more than that of a p-orbital, which in turn will be more than that for a d orbital.
- **5. Electronic Configuration:**
- Ionization energy depends upon the electronic configuration of the atom.

- It has been observed that certain electronic configuration are more stable than others.
- For example half filled and fully filled orbitals are more stable than the others.
- So, it will be difficult to remove electron from these stable configuration and ionization energy required will be high.
- **GENERALISATIONS**
 - i) The noble gas have the most stable electronic configuration in each period. They have the highest I.E.
 - ii) The elements like Be($1s^2, 2s^2$) and Mg ($1s^2 2s^2 2p^6 3s^2$) have completely filled orbitals. Hence their I.E. will be high.

iii) The elements like N($1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$) and P($1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$) have half filled orbital and hence more stable. Hence they require more energy to remove electron. So, their ionization energy will be large.

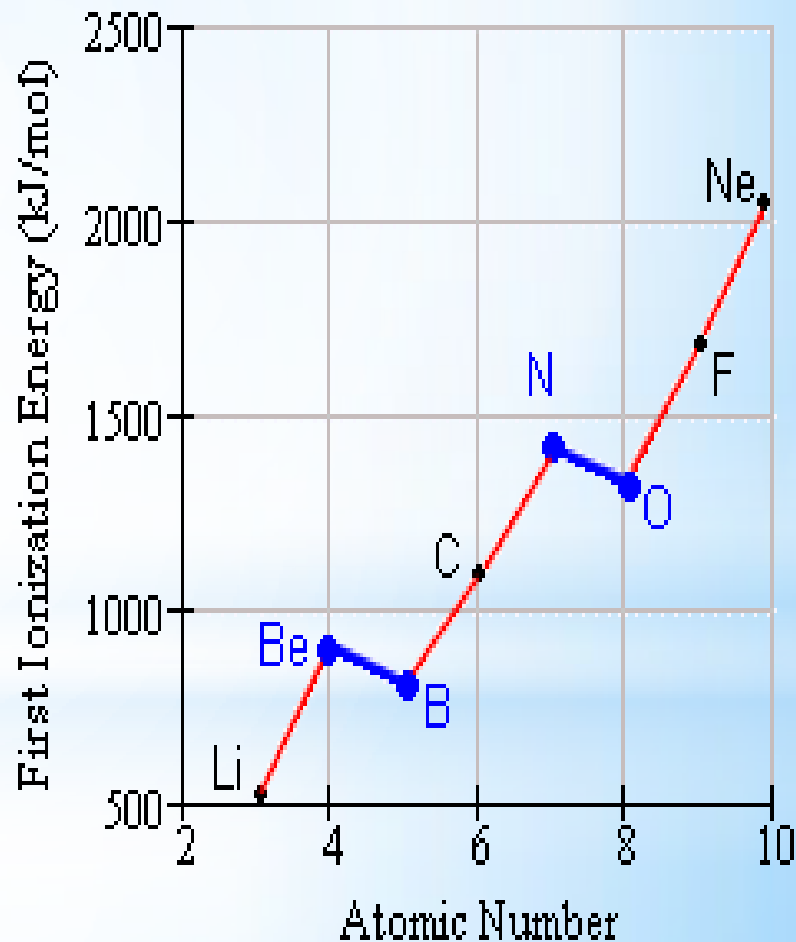


Variation of Ionization Energy In Periodic Table:

Ionization energy increases with increase in atomic number in period.

- i) On moving across a period from left to right, the nuclear charge increases.
- ii) The atomic size decreases along a period though the energy level remains same.

As a result of increasing nuclear charge and simultaneous decrease in atomic size, the valence electrons are more tightly held by nucleus.

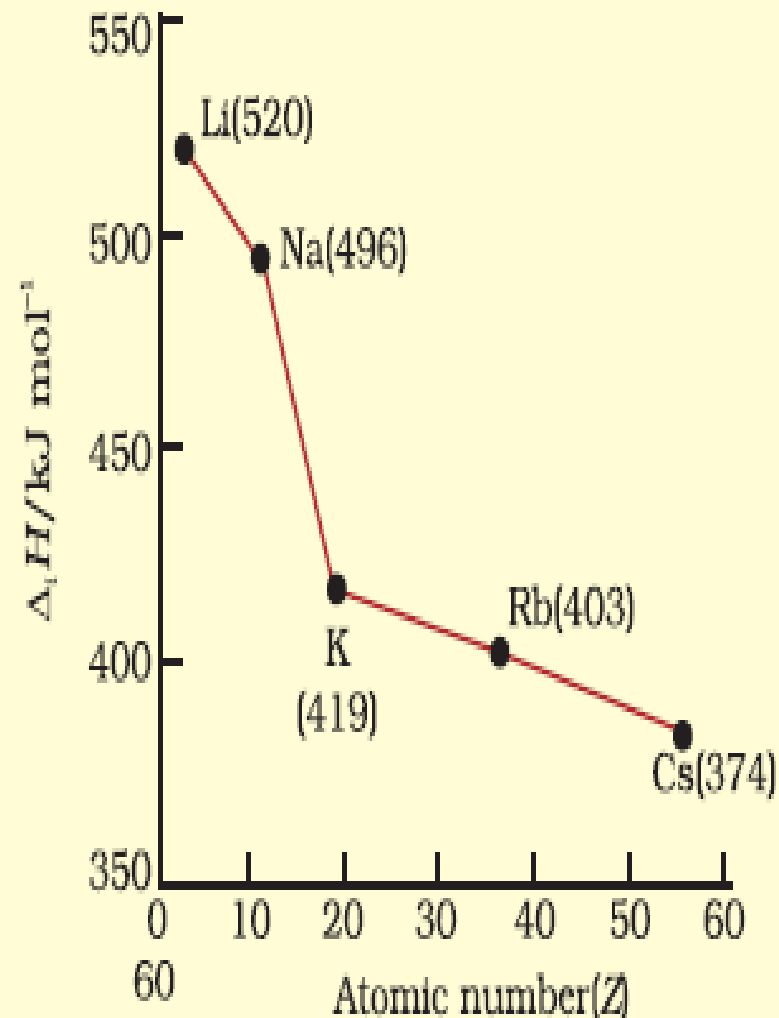


Ionization energy in group:

Ionization energy decreases gradually in moving from top to bottom in a group .

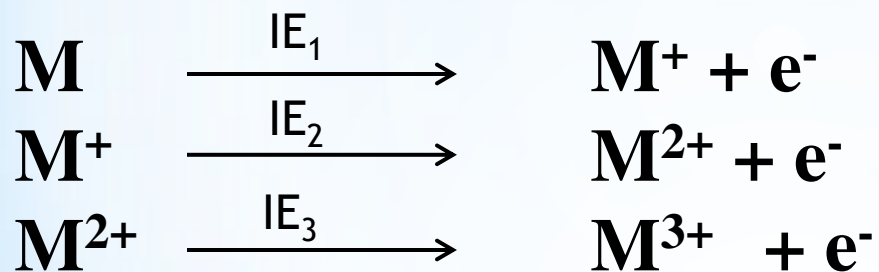
On moving down a group from top to bottom ,

- i) the nuclear charge increases
- ii) the atomic size increases along a group due to addition of shell
- iii) There is increase in shielding and screening effect on the outermost electron due to an increase in number of inner electron.



The effect in atomic size and shielding effect is more than the effect of increase in nuclear charge. As a result electron becomes less and less firmly held to the nucleus as we move down in the group.

Successive Ionization energy: Atom may not lose just one electron but more than one electron. Therefore it is essential to specify the Ionization energy for each electron removed. The energy required to lose or remove subsequent electrons from atom in gaseous state are known as successive ionization energy.



For example, the first and second ionization of Li may be given as



Ionization Energy



Table 8.5 Successive Ionization Energies of the Elements Lithium Through Sodium

Z	Element	Number of Valence Electrons	Ionization Energy (MJ/mol)*												
			IE ₁	IE ₂	IE ₃	IE ₄	IE ₅	IE ₆	IE ₇	IE ₈	IE ₉	IE ₁₀			
3	Li	1	0.52	7.30	11.81										
4	Be	2	0.90	1.76	14.85	21.01									Core electrons
5	B	3	0.80	2.43	3.66	25.02	32.82								
6	C	4	1.09	2.35	4.62	6.22	37.83	47.28							
7	N	5	1.40	2.86	4.58	7.48	9.44	53.27	64.36						
8	O	6	1.31	3.39	5.30	7.47	10.98	13.33	71.33	84.08					
9	F	7	1.68	3.37	6.05	8.41	11.02	15.16	17.87	92.04	106.43				
10	Ne	8	2.08	3.95	6.12	9.37	12.18	15.24	20.00	23.07	115.38	131.43			
11	Na	1	0.50	4.56	6.91	9.54	13.35	16.61	20.11	25.49	28.93	141.37			



THE END

THANK YOU