



PADMA CHEMISTRY CLASSES

CLASS 11th **IIT-JEE , NEET** **BOARDS**

LEVEL 1

Level 1

PERIODIC **PROPERTIES**

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PERIODIC PROPERTIES

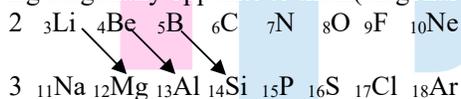
PERIODIC CLASSIFICATION OF ELEMENTS

Modern periodic Law: The Physical and Chemical properties of elements are the periodic function of their atomic number.

Modern periodic table: This is the modified form of Mendeleef 's periodic table . It was improved and modified by Moseley. The modern periodic table contains nine vertical columns known as groups. They are numbered from 1 to 8 and then zero. Except zero and eighth group, rest were divided in to sub group A and B.

CHARACTERISTICS OF PERIODS

- There are seven periods in the modern periodic table. Each period is known by the number of elements it contains.
- The seventh period will accommodate thirty two elements .
- Expect first period, each period starts with an alkali metal and finishes at an inert gas.
 - The seventh period will end at the element having atomic number 118.
 - The element having atomic number 118 will be an inert gas.
- The number of electrons increases from 1 to 8 in the outermost shell of elements in a period from left to right (exception in first period).
- Elements of second and third period are called typical or *representative elements*.
- Elements of third period only are known as *bridge elements* because a group divides in to sub groups A and B from this element.
- Elements of second period having low atomic numbers show similarities with the elements of third period lying diagonally opposite to them(diagonal relationship).



Diagonal relationship arises due to :

almost similar atomic sizes almost similar electronegativity similar polarizing power of ions

(i) *Diagonal relationship vanishes with increase in atomic number of elements*

- Valency of elements with respect to Oxygen increases from 1 to 7 in a period from left to right. With respect to Hydrogen ,the valency of elements first increase from 1 to 4 and then decreases to one .
- There is a gradation in the properties of elements when one traverses from left to right in the period

PERIOD	₁₁ Na	₁₂ Mg	₁₃ Al	₁₄ Si	₁₅ P	₁₆ S	₁₇ Cl	₁₈ Ar
3								
Valency of element w.r.t. oxygen	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇	
	1	2	3	4	5	6	7	
Valency of element w.r.t. Hydrogen	NaH	MgH ₂	(AlH ₃) _n	SiH ₄	PH ₃	H ₂ S	HCl	
	1	2	3	4	3	2	1	

CHARACTERISTICS OF GROUPS

- There are nine vertical columns in the periodic table known as groups except zero and VIII group, the rest have been divided in to sub groups A and B.
- Elements of sub group 'A' are called normal elements while those of sub group 'B' and VIII group are called *transitional elements* .
- Elements of a sub group show similarities in their properties and differ from the elements of other sub- group.
- All the elements of a sub group possess identical electronic configuration in their outermost shell .Hence they show similarities in their characteristics.

PERIODIC PROPERTIES

- Number of electrons in the outermost shell of an atom of an element is equal to its group number.
- The highest oxidation state of an element can not exceed its group number.
- The maximum valency of an element with respect to oxygen is equal to its group number.
- There is a gradation in the properties of elements from top to bottom in a group.

S-BLOCK ELEMENTS

- The elements in whose atoms the last electron (differentiating electron) enters the s- subshell of outermost energy level, are called s-Block elements.
- In the Bohr periodic table they are placed on the extreme left hand side.
- Their outer electronic configuration varies from ns^1 to ns^2 .
- They include the elements of IA and IIA sub groups of modern periodic table.
- Hydrogen and helium are also s-block elements.

IA	IIA
1	2
ns^1	ns^2
${}_1\text{H}$	${}_2\text{He}$
${}_3\text{Li}$	${}_4\text{Be}$
${}_{11}\text{Na}$	${}_{12}\text{Mg}$
${}_{19}\text{K}$	${}_{20}\text{Ca}$
${}_{37}\text{Rb}$	${}_{38}\text{Sr}$
${}_{55}\text{Cs}$	${}_{56}\text{Ba}$
${}_{87}\text{Fr}$	${}_{88}\text{Ra}$

CHARACTERISTICS OF s-BLOCK ELEMENTS

- Except hydrogen and helium, all are metals.
- Lithium is the lightest metal known.
- Metallic nature of IA (Alkali metals) > IIA(Alkaline earth metals)
- They are malleable and ductile . Ductility and malleability of IIA elements > I A elements.
- They are good conductors of heat and electricity. They have low density. Density of IA elements < IIA.
- Density increases from Li to Cs in IA(K is an exception which has lower density than Na). In IIA, Ca has lowest density . They have weak metallic bonds. Metallic bond strength of IA < IIA .
- Metallic bond strength decrease with increases in atomic number from top to bottom in IA.

I. Atomic size :-

s-block elements have biggest atomic size Atomic size of IA element > IIA element

II. Ionization potential :

- s-Block elements possess low ionization Potential.
- I.P. of IA(alkali metals) < IIA(alkaline earth metals) Ionization potential decreases from Li to Na in IA and Be to Ba in IIA.
- K, Rb and Cs are used in photoelectric cells due to their low ionization potential values.

III Electro positive nature:

All the s-block elements are highly electro-positive. Electro positive nature of IA > IIA elements

IV Valency : Elements of IA are monovalent while those IIA are divalent in their compounds.

V Reactivity: They are highly reactive metals.

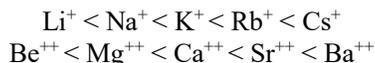
- Reactivity of IA > IIA elements.

PERIODIC PROPERTIES

- The alkali metals (IA metals) are stored under liquid paraffin.
- Reactivity increases from top to bottom in both the groups.

VI. Nature of ions :

- They form positive ions by loss of electrons from outermost shells.
- The positive ions are colourless and diamagnetic as they contain no unpaired electrons.
Ionic size increases from top to bottom ,i.e.,



VII. Nature of compounds :

- s-Block elements produce ionic compounds.
- The ionic nature of the compounds increases from top to bottom in both the groups.
- Li in IA and Be in IIA form compounds having some covalent nature. This is because of their small ionic sizes and large polarizing powers.

VIII. Reducing nature :

- All the s-Block metals are powerful reducing agents because they are highly electropositive elements.
- Reducing nature of IA elements > IIA elements
- Reducing power of alkali metals increases by dissolving them in liquid ammonia due to the presence of solvated electrons.

IX. Flame colouration : They give characteristic colours to the non-luminous flame of Bunsen burner

IA		IIA	
Li	Red	Be	No colour
Na	Golden yellow	Mg	No colour
K	Purple- blue	Ca	Brick red
Rb	violet	Sr	Crimson red
Cs	blue	Ba	Apple green
		Ra	Red

p-BLOCK ELEMENTS

The elements in the atoms of which the differentiating electron enters the p- sub-shell of outermost shell are called p-block elements. Their general outer electronic configuration varies from ns^2np^1 to ns^2np^6 . These elements are situated on-

13 (III A) ns^2ns^1	14 (IVA) ns^2np^2	15(VA) ns^2np^3	16 (VIA) ns^2np^4	17(VIIA) ns^2np^5	18(Zero group) ns^2np^6
${}_5\text{B}$	${}_6\text{C}$	${}_7\text{N}$	${}_8\text{O}$	${}_9\text{F}$	${}_{10}\text{Ne}$
${}_{13}\text{Al}$	${}_{14}\text{Si}$	${}_{15}\text{P}$	${}_{16}\text{S}$	${}_{17}\text{Cl}$	${}_{18}\text{Ar}$
${}_{31}\text{Ga}$	${}_{32}\text{Ge}$	${}_{33}\text{As}$	${}_{34}\text{Se}$	${}_{35}\text{Br}$	${}_{36}\text{Kr}$
${}_{49}\text{In}$	${}_{50}\text{Sn}$	${}_{51}\text{Sb}$	${}_{52}\text{Te}$	${}_{53}\text{I}$	${}_{54}\text{Xe}$
${}_{81}\text{Tl}$	${}_{82}\text{Pb}$	${}_{83}\text{Bi}$	${}_{84}\text{Po}$	${}_{85}\text{At}$	${}_{86}\text{Rn}$

PERIODIC PROPERTIES**CHARACTERISTICS OF p-BLOCK ELEMENTS**

- (i) Most of the p-Block elements are non-metals.
- (ii) p-Block includes metals, non-metals as well as metalloids.
- (iii) In group 13, except boron, all the elements are metals.
- (iv) In group 14, Sn and Pb metals while in group 15. Bi is a metal. Ge, As, Sb, Se and Te are metalloids.

(i) Oxidation state and valency:

1. Except oxygen, fluorine and inert gases, all the p-Block elements exhibit oxidation states from +n to (n-8) where n is the number of electrons in outer most shell.
2. Elements having 4, 5, 6, and 7 electrons in their outermost shell, show valency 4, 3, 2 and 1 in their compounds.
3. Elements of third group are trivalent and show oxidation state +1 and +3 in their compounds.
4. Many p-block elements show variable oxidation states. For example-- $\text{PCl}_3, \text{PCl}_5, \text{As}_2\text{S}_3, \text{As}_2\text{S}_5, \text{SnCl}_2, \text{SnCl}_4$ etc. There is a difference of two units in the successive valencies.
5. Stability of lower oxidation state increases from top to bottom in a group (due to inert pair effect).
6. Number of unpaired electron in the outermost shell gives the valency of an element.

(ii) Ionization potential:

1. p-Block elements have higher ionization potential than s-Block elements.
2. It is due to decreased atomic size and increased nuclear charge.
3. Hence, p-Block elements do not form positive ions. They possess a tendency to form negative ions,

Element	:	B	C	N	O	F
Ionization Potential(eV)	:	8.3	11.26	14.5	13.6	17.4

(iii) Atomic radii:

1. p-Block elements have smaller atomic radii than s-Block elements.
2. The atomic radius decrease from left to right in A, the right hand side of the Bohr's Periodic table.
3. This Block contains the elements of group 13(IIIA), 14(IVA), 15(VA), 16(VIA), 17(VIIA) and 18(zero group).

Period due to increasing nuclear charge.

Element	:	B	C	N	O	F
Atomic radius (A)	:	0.8	0.77	0.74	0.74	0.72

- (iv) **Electro negativity:** They have high electro negativities it is their small atomic size and increased number of electron in outer most shell.

Element	:	B	C	N	O	F
E.N.	:	2.0	2.5	3.0	3.5	4.0

(Pauling scale)

- (v) **Electron Affinity:** p-Block elements possess higher electron affinities than s-Block elements. Oxygen, sulphur and halogens possess high electron affinities. Halogens have highest electron affinity to their respective periods.

▪ **Chlorine has maximum electron affinity in the periodic table.**

- (vi) **Reactivity:** Halogens, oxygen, sulphur and phosphorus are elements of p-Block rest of elements are less reactive.

• Reactivity of non-metal (p-Block elements) \propto electronegative character

- (vii) **Catenation:** It is property of forming long chain by combination of many identical non metallic atoms. Boron, carbon, silicon, nitrogen, phosphorus, oxygen, sulphur etc. Elements form straight or closed chain by uniting together.

• **Carbon has maximum catenation property.**

PERIODIC PROPERTIES

- (viii) **Nature of compounds:** The p-Block elements gain or share electron to complete the octet. Thus they form ionic compounds when they accept electrons and covalent compounds when they share electrons.

- $$\begin{array}{ccc} \text{p-Block element} + \text{p-Block element} & \xrightarrow[\text{electrons}]{\text{Sharing of}} & \text{covalent compound} \end{array}$$
- $$\begin{array}{ccc} \text{s-Block element} + \text{p-Block element} & \xrightarrow[\text{electrons}]{\text{transference of}} & \text{Ionic or electrovalent compound} \end{array}$$

Tendency to form polyatomic molecules :

1. Some of the p-block elements form polyatomic molecules.
2. For example, Diamond (C₄)_n; Sulphur, (S₈); selenium, (se₈); phosphorus, (P₄); Red phosphorus, (P₄)_n; Ozone (O₃), etc.
3. **Nature of oxides :** Generally, the oxides of p-Block elements are acidic in nature.
4. Actually they form neutral, acidic, basic, amphoteric and mixed (compound) oxides for example, Al₂O₃ is amphoteric in nature, Bi₂O₃ is basic oxide.

The acidic nature of non-metallic oxide increases with increasing oxidation number of the non-metal.

d-BLOCK ELEMENTS

1. In their atoms, the differentiating electrons enters the d-sub-shell.
2. The penultimate shell, i.e (n-1). d –sub shell. Their general outer electronic configuration is (n-1) d¹⁻¹⁰ns¹⁻².
3. Their outer two shells are incompletely filled with electrons. They are also known as *transition elements*.
4. In the long form, they are placed in the middle between s-and p-Block elements.
5. All the elements of 'B' sub groups and VIII group are transitional elements, i.e., elements of 3,4,5,6,7,8,9,10, 11 and 12 groups are transitional elements.
6. In the periodic table, there are four d- series which are describe as follows-
 - i. 3d-series (1st transition series): This series is present in the 4th period. It comprises of ten elements.
3d-series : ₂₁Sc.....₃₀Zn.
 - ii. 4d-series : The elements of this series are situated in the 5th period. It also contains ten elements.
4d-series : ₃₉Y.....₄₈Cd,
(Yttrium)
 - iii. 5d-series : Ten elements are present in this series which are included in the sixth period.
This series starts with Lanthanum(57) and ends at Mercury (80). 5d-series : ₅₇La.....₈₀Hg .
 - iv. 6d-series : This series is incomplete, i.e. all the elements of this series are not known. It starts with Actinium(89) and ends at Examercury (112) Its elements are situated in 7th period.

GENERAL CHARACTERISTICS OF d-BLOCK ELEMENTS**1. All the d-Block elements are metals :**

- a. Metals having (n-1) d¹⁰ configuration are characterized by the presence of metallic bond while other possess metallic as well as covalent bonding .
- b. Their metallic bond strength is greater than s-Block metals.
- c. Mercury has weakest metallic bond.
- d. They are crystalline metals. They crystallize in b.c.c., f.c.c. or h.c.p. lattice.
For example, copper crystallizes in f.c.c. lattice. Its co-ordination no. is 12.

PERIODIC PROPERTIES**2. Melting and boiling points :**

- Transitional metals possess much higher melting point and boiling point, than s-Block elements.
- It is due to strong inter atomic attractions in these metals.
- Higher melting and boiling points are of tungsten (W).
- Zn, Cd, and mercury has lowest M.P. and B.P. in their respective series due to (n-1) d¹⁰ configuration.
- Lowest melting and boiling points of mercury as it has weakest metallic bond.

3. Conductivity:

- All the d-Block metals are good conductors of heat and electricity, due to the presence of free and mobile electrons.
- Silver is the best electrical conductor, copper comes next to silver in electrical conductance.

4. Density :

- Densities of d-Block elements are very high as compared with s-Block metals.
- It is due to their small radii and close packed structures.
- In a given transition series density increase on moving from left to right.
- It becomes maximum on reaching group VIII and decrease further to group IB and IIB.
- Density increases on moving down a group.
- The reason is that the atomic size of elements in the same group but in second and third transition series are practically the same, but their atomic weights are almost doubled. So densities are increased.
- ²¹Scandium has the lowest while Os has the highest density in d-Block elements .

5. Variable valency :

- d-Block elements having partially filled (n-1)d sub-shell show variable valencies (or oxidation states).
- In these elements, not only the ns-electrons take part in bond formation but the (n-1) d electrons also because there is little difference in the energy of n and (n-1)d orbitals.
- Zn and Cd do not exhibit variable valency.
- Maximum variation in oxidation state of a transitional metal is shown when the c is (n-1)d⁵ ns² i.e. from +2 to +7.

6. Formation of coloured and paramagnetic ions :

- d-Block elements form coloured ions. Colour arises due to d-d transition.
- This transition is possible in ion having partially filled (n-1) d sub shell.
- Thus transitional metal ion with (n-1) d¹⁻⁹ configuration are coloured and paramagnetic.
- If the configuration of the ion is (n-1) d¹⁰ or (n-1) d⁰ it will be *colour less and diamagnetic*.
- Atoms ions or molecules having unpaired electrons show paramagnetic character. They possess permanent magnetic moment. They are attracted in magnetic field.
- Paramagnetic character \propto number of unpaired electrons.
- Magnetic moment is measured in Bohr Magneton (B.M.).
- When there are 'n' unpaired electrons in a transitional metal ion, magnetic moment in Bohr Magneton is given by, $\mu = \sqrt{n(n+2)}$.

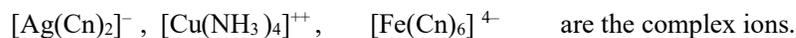
7. Catalytic Properties :

- d-block metals and their compounds possess excellent catalytic properties.
- This property of these metals is due to their variable valencies on their surface atoms.
- Pt / PtO is Adam's catalyst.
- A mixture of R₃Al and TiCl₄ is called Zeigler-Natta catalyst.
- Pt is used as a catalyst for oxidation, hydrogenation and dehydrogenation reactions.
- Transitional metals occlude hydrogen gas. Hence, they are good catalyst for hydrogenation reactions.

8. Formation of co-ordination complexes:

- The cations of transition metals have a strong tendency to form complexes.
- The cations have ionic radius, high positive charge density and vacant (n-1) d- atomic orbitals of suitable energy to accept electron pair from ligand molecules.

For example,



Charge

- The stability of a complex \propto Ionic radius

PERIODIC PROPERTIES

- **Reactivity** : Comparatively these elements are less reactive than s and p-Block elements. The reason being -
 - (i) Large heat of sublimation.
 - (ii) High ionization potential
 - (iii) High heat of hydration of ions

8. Formation of Non-stoichiometric compounds :

- a. Non-stoichiometric compounds are the compounds of indefinite composition.
- b. These compounds do not obey the valency rules.
- c. For example, FeO is a Non-stoichiometric compound as its composition varies between $\text{Fe}_{0.84}\text{O}$ and $\text{Fe}_{0.94}\text{O}$. Similarly $\text{LaH}_{2.72}$, $\text{VSe}_{0.98}$, $\text{VSe}_{1.2}$ and $\text{VSe}_{1.6}$ are Non-stoichiometric compounds.
- d. Transitional metals form Non-stoichiometric compounds due to their variable valencies.
- e. These metals exist in mixed oxidation state in Non-stoichiometric compounds.

Nature of compounds : The d-Block metals form ionic, covalent as well as dative compounds.

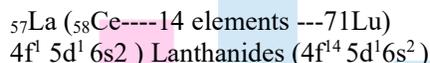
f-BLOCK ELEMENTS

- The elements in whose atoms, the last electron enters the $(n-2)f$ sub shell are called f-Block elements, i.e. in these elements the f-sub shell of ante-penultimate shell expands by filling with electrons.
- The general outer electronic configuration of these elements is $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$.
- Their outer three shells are incompletely filled with electrons.
- They are also called inner transition elements. There are two f-series in the periodic table.

4f-series (Lanthanides, Lanthanones, rare earth elements) :

This series starts with Cerium (58) and ends with Lutetium(71).

These 14 elements are known as Lanthanides after the element Lanthanum, although it is not the member of the lanthanide family.

**5f-series (Actinides, or Actinones):**

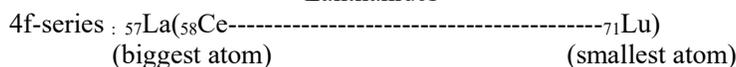
This series also contain 14 elements. In these elements 5f-subshell is progressively filled with electrons.

It starts with thorium (90) and ends at lawrencium (103). 14elements ${}_{89}\text{Ac}({}_{90}\text{Th} \text{-----} 103\text{Lr})$

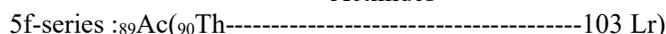
CHARACTERISTICS OF f-BLOCK ELEMENTS

- All the f-Block elements are metals. They are highly reactive elements. *All the actinides are radioactive.*
- F-Block elements within each of the series are very similar in their chemical properties and hence they are grouped separately.
- Atomic size : In a series, the atomic radius slightly decreases with increase in atomic number.
- This decrease in atomic radii is known as lanthanide contraction for lanthanides and actinide contraction for actinides.
- The cause of decrease in atomic size is that the last electron enters the $(n-2)$ f sub-shell whose electron do not shield the outer shell electrons against the nuclear attraction which increases progressively. So due to increased nuclear attraction the atomic size goes on decreasing.

Lanthanides



Actinides



All the f-Block elements are highly electro-positive metals.

6. Oxidation state:

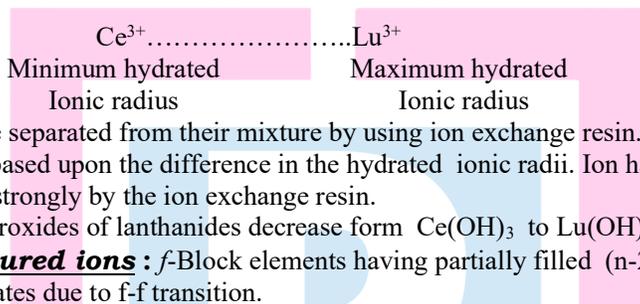
PERIODIC PROPERTIES

- a. The electrons present in (n-2)f-subshell are not easily available for taking part in bond formation.
- b. Hence these elements do not show as many oxidation states as the d-block elements.
- c. **The most common oxidation state of Lanthanides is +3.** This oxidation state is attained by the loss of 6s² and the single 5d electron or one of the 4f electrons.
- d. Some lanthanides exhibit +2 and +4 oxidation states also. The common oxidation state of actinides is also +3, but some show even higher oxidation states of +4, +5 and +6.
- e. The ionic radii of lanthanides and actinides also decreases with increasing atomic number.

Lanthanides: $_{57}\text{La}^{+3}$ (Ce³⁺.....Lu³⁺)
 (biggest ion) (smallest ion)

Actinides: $_{89}\text{Ac}^{3+}$ (90Th³⁺.....¹⁰³Lr³⁺)
 (biggest ion) (smallest ion)

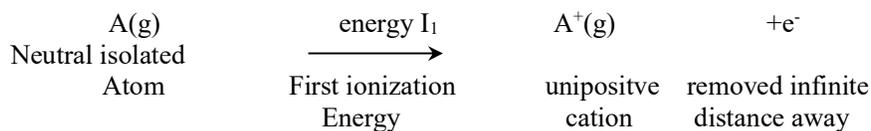
- f. The hydrated ionic radii increases from Ce³⁺ to Lu³⁺ because degree of hydration increases with decreasing ionic size.



- g. The lanthanides are separated from their mixture by using ion exchange resin.
- h. The separation is based upon the difference in the hydrated ionic radii. Ion having last hydrated ionic radius is absorbed strongly by the ion exchange resin.
 The basicity of hydroxides of lanthanides decrease from Ce(OH)₃ to Lu(OH)₃.
- 7. **Formation of coloured ions :** f-Block elements having partially filled (n-2)f sub-shells form colored ions. The colour originates due to f-f transition.
 - Ions having unpaired electrons possess paramagnetic character also.
- 8. **Tendency to form complex ion :** f-block elements also produce co-ordination complexes but the stability of these complexes is less than complex formed by d-Block elements.
- 9. **Nature of compounds :** they generally form ionic compounds.
- 10. Elements after uranium (92) are known as Trans-uranic, post-uranic, synthetic, man-made or artificial elements they do not occur in the earth's crust. The maximum contribution in their preparation is of Prof. G.T. Seaborg.

IONIZATION POTENTIAL (I.P.)

1. The process of separation of an electron from an atom, or molecule is called ionization .
2. It results in the production of a positive ion . **Ionization potential (I.P)** may be defined as-**'The amount of energy required to take out the most loosely held electron from a neutral isolated atom'** .
3. First ionization potential (I₁) removes first electron from a neutral isolated atom and converts it into a unipositive ion.



4. Energy is supplied to the atom for ionization to overcome the forces of attraction between the positively charged nucleus and negatively charged electrons.
5. Energy in the form of electricity, heat or radiation may be given to atom for ionization .
6. Ionization potential is expressed in eV/atom kcal/mol or K jule/mole.
7. The size of unipositive cation is smaller than neutral atom and the effective nuclear charge is greater, so the second ionization potential (I₂) of an element is always greater than first ionization potential(I₁)

PERIODIC PROPERTIES

5. Moreover, this electron experiences less attraction from the nucleus because it is farther than 2s-electrons. Hence, less energy is needed to expel this electron, thus decreasing the first Ionization potential.
6. Similarly, it can be explained that I_1 of Al is lower than that of Mg.
7. The first Ionization potential (I_1) of Nitrogen is greater than I_1 of oxygen
 ${}_8\text{O}: 1s^2, 2s^2, 2p^4$
 ${}_7\text{N}: 1s^2, 2s^2, 2p^3$
8. In nitrogen atom, the outermost electron is present in half-filled 2p-subshell which is a stable electronic arrangement. So, a greater amount of energy is required to pull out one electron from neutral gaseous nitrogen atom. Hence its I_1 is greater than Oxygen.
9. It should be noted that I_2 of oxygen will be higher than that of nitrogen because O^+ ion will be having exactly half filled 2p-subshell. In a similar way, it can be explained that the I_1 of Phosphorus is higher than Sulphur.

(ii) Down a group :

1. On moving down a group, the ionization energy (or potential) of elements decreases.
2. It is due to increase in atomic size, increases in screening (or shielding) effect of inner electrons and decrease in effective nuclear charge. For example, I.P of halogens decreases from top (Fluorine) to bottom (Iodine).

ELECTRONEGATIVITY

1. The tendency of an atom in its molecule to attract the shared electron pair towards it self is called **Electronegativity of the element**.
 2. Thus, electronegativity is the electron attracting property of an atom in its combined state.
 3. Electro negativity increases with increasing number of electrons in the outermost shell of elements, but decreases with increasing atomic size.
- (a) **Across a period** : In a period, electronegativity increases from left to right.
- Alkali metals possess minimum and the halogens have the maximum electronegativity in a period.
 - Noble gas element have zero electronegativity.
- (a) **In a group** : Electronegativity decreases from top to bottom in a group due to increase in atomic size.

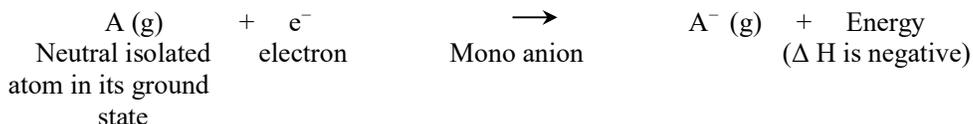
Importance of concept of Electro negativity :

- (i) It indicates the polarity of a bond between two atoms.
- Bond polarity \propto Electronegativity difference of bonded atoms.**
- When the Electronegativity difference between bonded atoms is 1.7, the bond is 50% ionic.
 - When the electronegativity difference between bonded atoms is zero, the bond is Non-polar covalent bond.
- (ii) Bond strength \propto electronegativity difference
 example: H-F is the strongest bond, while H-I is the weakest one among hydrogen halides.
- (iii) Electronegativity of non metals is higher than metals.
- (iv) Electronegativity of an element in its higher oxidation state is greater than in lower oxidation state.
- (v) Electronegativity of an element also depends upon its state of hybridization. Greater the s-character in a hybrid orbital, higher will be the electronegativity of the element. Thus electronegativity of carbon decreases in the order :

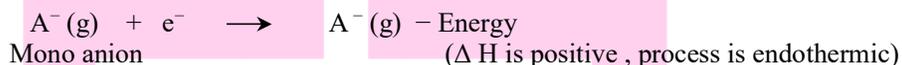


PERIODIC PROPERTIES**ELECTRON AFFINITY**

“The amount of energy liberated when an electron is added to a neutral, isolated atom in its ground state”



1. Greater the amount of energy released, greater the electron affinity. The magnitude of electron affinity is a measure of the tightness with which an extra electron can be held by an atom.
2. Electron affinity is expressed in eV/atom or kcal/mole. Like ionization energy, there are also first electron affinity, second electron affinity etc.
3. The first electron affinity is the energy released when first electron is added to neutral isolated atom, i.e. in the conversion of neutral atom in to mono anion.
4. The second electron affinity is not the energy released but the energy supplied, i.e. second electron affinity is the energy absorbed when a mono anion is converted in to a divalent anion.

**It can be explained as follows:**

When second electron is introduced in a mono anion, it experiences repulsive forces, so energy will have to be given to the system to overcome these forces of repulsion, secondly the size of mono anion is bigger than the neutral atom, i.e. the force of attraction of nucleus on outer electrons becomes less hence the energy will have to be given to the mono anion to enable to accept the second electron. Similarly, third electron affinity will also be the energy supplied to the system.

- Electron affinity is expressed in eV/atom or kcal or k joule per mole.

Factors affecting electron affinity :

(i) **Atomic size** : Electron affinity decreases with increasing atomic size.

$$\text{Electron affinity} \propto \frac{1}{\text{Atomic size}}$$

(ii) **Nuclear charge** : With increasing number, i.e. nuclear charge, the electron affinity of elements increases.

$$\text{Electron affinity} \propto \text{Nuclear charge}$$

(iii) **Half or fully filled sub-shells** : These are stable electronic configuration. Hence it is difficult to introduce electron in to a stable electronic arrangement, i.e. electron affinity become very low.

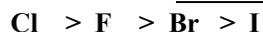
Variation of electron affinity :

(b) **Across a period** :

1. On traversing from left to right in a period, electron affinity of elements generally increase. Thus halogens possess maximum electron affinity while alkali metals have least electron affinity values.
2. Electron affinity values of elements of IIA and Noble gas elements are zero
3. Electron affinity of nitrogen and phosphorus are exceptionally low because of half p sub-shell (extra stable configuration)

(c) **In a group** :

1. Electron affinity of elements decreases on moving down a group because atomic size increases.
2. In halogen fluorine has lower electron affinity than chlorine because fluorine has very small atomic size.
3. Hence the electrons already present in fluorine atom exert a repulsive force on incoming extra electron.
4. It results in the liberation of less energy in the formation of $F^{-}(g)$ ion.
 - Decreasing order of EA of halogens:

PERIODIC PROPERTIES

- Chlorine has maximum electron affinity in the periodic table .
- Elements of third period have higher values of electron affinities than second period elements.

PERIODIC PROPERTIES

1. **Atomic size or radius** : Due to the wave nature of revolving electron, the size of an atom can not be taken as the distance between the centre of the nucleus and the electron in the outermost shell. Atomic size is defined differently for different types of elements.
2. **Metallic radius or size** : Size of radius in such a crystal is defined as “*The half of the nuclear distance between adjacently placed metallic atoms.*”
3. **Single bond covalent radius (SBCR)**: “Half of the single covalent bond length between identical non-metallic atoms.”

$$\text{Thus covalent radius} = \frac{\text{Single covalent bond length}}{2}$$

For example ,

$$\text{Atomic size of hydrogen} = \frac{0.74\text{A}^0}{2} = 0.37\text{A}^0$$

4. **Vander Waal's radius** : It is taken as the atomic size of noble gas atoms. It may be defined as “*The half of the distance between the nuclei of adjacently placed atoms in the solid state of noble gas.*”

**PERIODIC VARIATION OF ATOMIC SIZE**

1. **Across a period** : In a given period, atomic radius generally decreases from left to right because nuclear charge increases. For example, in second period the atomic radius are in the order:



- (i) Alkali metals from biggest atoms in their period .
- (ii) Halogens from smallest atom in their periods.
- (iii) Atomic size of noble gas atom is greater than its neighbouring halogen atom because the former has Vander Waal's radius .

$$r_{\text{Ne}} > r_{\text{F}}$$

$$(1.60 \text{ A}^0) > (0.72 \text{ A}^0)$$

2. **In a group** : The atomic radius increases from top to bottom in group because number of shells increases, shielding or screening effect of inner electrons increases and effective nuclear charge decreases. **For example**, the group of alkali metals, the atomic size increases from Li to Cs,



3. **Ionic radius** : “*The effective distance from the nucleus of an ion up to which it has its influence on the electron cloud .*”
4. **size of cation** : The size of cation is always smaller than its neutral parent atom because effective nuclear charge (Z_{eff}) is increased. The cation size goes on decreasing as more number of electrons are removed. Thus,

$$\text{cation size} \propto \frac{1}{\text{Magnitude of +ve charge}}$$

$$\text{For example : } r_{\text{Sn}} > r_{\text{Sn}^{2+}} > r_{\text{Sn}^{4+}}$$

PERIODIC PROPERTIES

5. **Size of anion (negative ion)**: Size of anion is always greater than neutral parent atom because effective nuclear charge (Z_{eff}) decrease. The anion size goes on increasing as more electrons are added to the atom because the effective nuclear charge also goes decreasing

Thus, anion size \propto Magnitude of charge on anion

For example, $r_o < r_o^- < r_o^{--}$

- o If A is a neutral isolated atom, the size of various species is in the order



6. **Variation of ionic size**: In any particular group, size of ion increases from top to bottom.

For example :



7. **Size of iso-electronic ions**: Species having the same number of electrons, but different nuclear charge are known as iso-electronic. The size of such species decreases with increasing nuclear charge because the attraction for the electrons increases.

- o Size of iso-electronic species $\propto \frac{1}{\text{Nuclear charge}}$

Iso-electronic Species with $18e^-$ may be arranged in the decreasing order of size as:

**IMPORTANT FACTS TO REMEMBER**

1. First element in the periodic table : *Hydrogen*
2. Lightest element in the periodic table : *Hydrogen*
3. Element with no neutron in its nucleus : *Hydrogen*
4. Most abundant element in the universe : *Hydrogen*
5. Element with highest specific heat : *Hydrogen*
6. Rogue element : *Hydrogen*
7. Lightest inert gas : **Helium**
8. Lightest metal : Lithium
9. Most abundant element (non-metal) in the earth's crust : OXYGEN
10. Most abundant metal in the earth's crust : *Aluminium*
11. Smallest atomic size : *Hydrogen*
12. Largest atomic size : Caesium
13. Highest electro-negativity : Fluorine
14. Lowest electro-negativity : Caesium
15. Highest ionization potential : **Helium**
16. Lowest ionization potential : Caesium
17. Lowest electron affinity : Noble gases (Zero)
18. Highest electron affinity : *Chlorine*
19. Least electro positive element : Fluorine
20. Metal with lowest melting point : MERCURY

PERIODIC PROPERTIES

21. Element with lowest melting and boiling point : **Helium**
22. Highest melting and boiling point metal : Tungston
23. Most reactive gaseous element : Fluorinet
24. Most reactive liquid element (metal) : Caesium
25. Most stable element : *Tellurium* ($t_{1/2} = 2 \times 10^{21}$ yrs)
26. Largest anion : **Astatide ion (At⁻)**
27. Largest atomic size : Caesium
28. Smallest anion : F⁻
29. Element with electro negativity next to fluorine : OXYGEN
30. Group heaving maximum number of gaseous elements in the periodic table : **Zero group**
31. Total number of gaseous elements in the periodic table : **11** (H₂, N₂, O₂, F₂, Cl₂, He, Ne, Ar, Kr, Xe,
32. Total number of liquid elements in the periodic table : **5** (Ga, Br, Cs, Hg and Fr)
33. Total number of solid elements in the periodic table : **89**
34. Liquid radioactive element : Francium
35. Total number of radioactive element in the periodic table : **25**
36. Largest group in the modern periodic table : **III B**
37. Volatile d-block elements : **Zinc, Cadmium and Mercury**
38. Rarest element in the earth's crust : **Astatine**
39. Element which forms maximum number of compounds : *Hydrogen*
40. Element having maximum catenation property : *Carbon*
41. Non-metal with highest melting and boiling point : *Diamond (Carbon)*
42. Most poisonous metal : **Plutonium**
43. Most stable carbonate : **Caesium carbonate** (Cs₂CO₃)
44. Strongest base (alkali) : CsOH
45. Strongest basic oxide : *Cs₂O*
46. Best conductor among non-metals : **Graphite**
47. The only liquid non-metal : Bromine (Br₂)
48. Elements stored under water : **Phosphorus**
49. Elements stored under kerosene (liquid paraffins) : **Na, K , Rb, Se**
50. Elements which sublimes on heating : **IODINE**
51. Noble metals : **Gold and platinum**
52. Amphoteric metals : *Zinc, Aluminium, Tin and lead .*
53. Amphoteric non-metal : **Silicon**
54. Metalloid elements : **Boron, Silicon, Germanium, Arsenic, Antimony, Selenium**
55. Diagonally related element pairs : **Li-Mg, Be-Al and B-Si**
56. Non metals with metallic luster : **Graphite (C), IODINE (I₂)**
57. Heaviest naturally occurring element : U²³⁸
58. Heaviest d-block element : **Osmoim**
59. Lightest d-block element : *Scandium*
60. Elements with least electrical conductance : **Lead (metal), Sulphur (non-metal)**
61. Lightest Metalloid (semimetal) : **Boron**
62. Heaviest Metalloid (semimetal) : *Tellurium*
63. Heaviest solid non-metal : **Astatine**
64. The element which constitutes 90% mass of the Sun : *Hydrogen*
65. Decreasing order of most abundant elements :
Oxygen > Silicon > Aluminium > Iron > Calcium > Magnesium > Sodium > Potassium
66. Amphoteric oxides : **ZnO, Al₂O₃, PbO, SnO₂, BeO, Sb₂O₃, TeO₂ .**
67. Netural oxides : **H₂O, N₂O, NO, CO, F₂O**
68. Dry bleacher : **H₂O₂**
69. Dry ice, drikold : **Solid CO₂**

PERIODIC PROPERTIES

- 70. The lanthanide which is not found in the earth's crust : **Promethium**
- 71. Radioactive lanthanides : **Promethium and Samarium**
- 72. First Man Made Element : **Technitium (43)**

