Class : Chemistry
Class : XI

Kannada

Classification of Elements

Why do We Need to Classify Elements ?

- The elements are the basic units of all types of matter.
- At present 118 elements are known. So, to study them and their compound individually, there was a need to classify them in a schematic way.

Genesis of Perlodic Classification:

- In 1829, Dobereiner arranged certain elements with similar properties in groups of three, known as triads, in such a way that the atomic mass of the middle element was nearly the same as the average atomic masses of the first and the third elements. A few triads proposed by him are listed.

Table : Dobereiner's Triads

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

- Dobereiner's relationship referred to as the Law of Triads.
- The Law of Triads, seemed to work only for a few elements, so it was dismissed.
- The English chemist, John Alexander Newlands in 1865 profounded the Law of Octaves in which he arranged the elements in increasing order of their atomic weights and noted that every eighth element had properties similar to the 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					

Nomenclature of Elements with Atomic Numbers > 100

The IUPAC has recommended a systematic nomenclature for the elements with atomic number more than 100 can be derived directly from the atom is number of the element using the numerical roots for 0 and numbers 1-9.

- The roots are put together in order of digits which make up the atomic number and "ium" is added at the end.

Table: Notation for IUPAC Nomenclature of Elements

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	E

As of now, elements with atomic numbers up to 118 have been discovered. Official names of all elements have been announced by IUPAC.

- 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
	Unnilennium	Une	Meitnerium	Mt

Atomic Number	Name according to IUPAC nomenclature	Symbol	IUPAC Official Name	IUPAC Symbol
109	Unniloctium	Une	Meitnerium	Mt
110	Unnilennium	Uun	Darnstadtium	Ds
111	Ununnillium	Uua	Rontgenium	Rg
112	Unununnium	Uub	Copernicium	Cn
113	Ununbium	Uut	Nihonium	Nh
114	Ununtrium	Uuq	Flerovium	Fl
115	Ununquadium	Uup	Moscovium	Mc
116	Upunpentium	Uuh	Livermorium	Lv
117	Ununhexium	Uus	Tennessine	Ts
118	Ununseptium	Uno	Oganesson	Og
	Ununoctium			

Electronic Configurations of Elements and the Periodic Table

Electronic Configurations in Periods :

Successive period in the Periodic Table is associated with the filling of the next higher principal energy level ($n = 1, n = 2$, etc.)

Period	Filling of orbital	Elements	Total no. of elements
First $(n=1)$	$1s^{1-1}$	${}_{1}\mathrm{H}, {}_{2}\mathrm{He}$	2
Second $(n-2)$	$2s^{1/2}2p^{0-6}$	${}_{3}\mathrm{Lito}_{10}\mathrm{Ne}$	8
Third $(n-3)$	$3s^{1-2}3p^{0-6}$	${}_{11}\mathrm{Nato}_{10}\mathrm{Ar}$	8
Fourth $(n-4)$	$4s^{1-2}3d^{0-10}4p^{p-6}$	${}_{10}\ \mathrm{K}-{}_{36}\mathrm{Kr}$	18
Fifth $(n=5)$	$5s^{1-2}4d^{0-10}5p^{0-6}$	${}_{32}\mathrm{Rb}-{}_{s1}\mathrm{Xe}$	18
Sixth $(n=6)$	$6s^{1-2}4f^{0-1+5d^{0-10}}6p^{0.6}$	${}_{35}\mathrm{Cs}_{-{}_{5}}\mathrm{Rn}$	32
Seventh $(n=7)$	$7s^{1-2}5f^{0-14}6d^{0-10}7p^{0-6}$	${}_{87}\ \mathrm{Fr}-{}_{110}\mathrm{Og}$	32

Groupwise Electronic Configurations :

Elements in the same vertical column or group have similar valence shell electronic configurations, the same number of electrons in the outer orbitals, and similar properties.

Atomic number	Symbol	Electronic configuration
3	Li	$1s^2 2s'$ (or) $[\mathrm{He}]2s'$
11	Na	$1s^2 2s^2 2p'3s'$ (or) $[\mathrm{Ne}]3s'$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 (\mathrm{or})[\mathrm{Ar}]4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ (or) $[\mathrm{Kr}]5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^7 4p^6 4d^{10} 5s^1$ $25p^4 6s^7 (or)[\mathrm{Xe}]6s^1$
87	Fr	$[Rn]7s'$

Electronic Configurations and types of Elements: $s^-, p^-, d^-, f-$ Blocks

- The elements are classify into four blocks viz., s-block, p-block, d-block and f-block depending on the type of atomic orbitals that are being filled with electrons:
- Helium belongs to the s-block but its positioning in the p-block along with other group 18 elements is justified because it has a completely filled valence shell? $(1s^2)$ and as a result, exhibits properties characteristic of other noble gases.
- Hydrogen has only one s-electron and hence can be placed in group I (alkali metals). It can also gain an electron to achieve a noble gas arrangement and hence it can behave similar to a group 17 (halogen family) elements. Because it is a special case, we shall place hydrogen separately at the top of the Periodic Table.
- Newlands's Law of Octaves seemed to be true only for elements up to calcium.
- Lothar Meyer plotted the physical properties such as Atomic volume, in pt., b.pt. against atomic weight of elements and obtained a periodically repeated pattern. Unlike Newlands, he observed a change in the length of that repeating pattern.
- By 1868, Lothar Meyer had developed a table of the elements that closely resembles the Modem Periodic Table.
- Mendeleev was responsible for publishing the Perlodic Law for the first time. It states as follows : The properties of the elements are a periodic function of their atomic weights.
- Mendeleev arranged elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group. He proposed that some of the elements were still undiscovered and, therefore, left several gaps in the table.
- He left a gap under aluminium and a gap under silicon, and called these elements Eka-Aluminium and Eka-Silicon Modern Periodic Law and the Present form of the Periodic Table
- Henry Moseley observed regularities in the characteristic X-ray spectra of the elements.
- A plot of $\sqrt{v}$ (where $\sqrt{v}$ is frequency of X-rays emitted) against atomic number (Z) gave a straight line thereby showed that the atomic number is a more fundamental property of an element than its atomic mass.
- Modern Periodic Law can be stated as the physical and chemical properties of the elements are periodic functions of their atomic numbers.
- The Periodic Law is essentially the consequence of the periodic variation in electronic configurations, which indeed determine the physical and chemical properties of elements and their compounds.
- A modem version is also called "long form" of the Periodic Table of the elements.
- It consists horizontal rows called periods and the vertical columns called groups.
- There are 18 groups and are numbered from 1 to 18 replacing the older notation of groups IA ... VIIA, VIII, IB ....VIIB and 0.
- There are seven periods corresponds to the highest principal quantum number (n) of the elements in each period. The first period contains 2 elements. The subsequent periods consists of 8,8, 18,18,32 and 32 elements, respectively.
- 14 elements of both sixth and seventh periods (lanthanoids and actinoids. respectively) are placed in separate panels at the bottom.

The s-Block Element:

- 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 elements.

Characteristics:

- They are all reactive metals with low ionization enthalpies. Thus they are never found pure in nature.
- 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 and beryllium are predominantly ionic.

The p-Block Elements:

The p-Block Elements comprise those belonging to Group 13 to 18 and these together with the 5 -Block Elements are called the Representative Elements or Main Group Elements.

Characteristics:

- 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. All the orbitals in the valence shell of the noble gases are completely filled by electrons and it is very difficult to alter this stable arrangement by the addition or removal of electrons.
  Elements of group 17 (called halogens) and elements of group 16 (called chalcogens) have highly negative electron gain enthalpies and readily add one or two electrons respectively to attain 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.

The d-Block Elements (Iransition Elements):

- Transition metals form a bridge between the chemically active metals of 5 block elements and the less active elements of groups 13 and 14.

Characteristics:

These are the elements of Group 3 to 12 in the centre of the Periodic Table.

- These are characterised by the filling of inner d-orbitals (partially) by electrons and are therefore referred to as d-Block Elements.
- These elements have the general outer electronic configuration $(n-1)d^{1-19}n5^{0-2}$.
- They are all metals, mostly form coloured ions, exhibit variable valence (oxidation states), paramagnetism and oftenly used as catalysts.
- $\mathrm{Zn, Cd}$ and Hg which have the electronic configuration, $(n-1)d^{10}ns^2$ do not show most of the properties of transition elements.

The f-Block Elements (Immer-Transition Elements):

- The two rows of elements at the bottom of the Periodic Table, called the Lanthanoids, $Ce(Z=58)-Lu(Z=71)$ and Actinoids. $Th(Z=90)-Lr(Z=103)$ are characterised by the outer electronic configuration $(n-2)\int^{1-14}(n-1)d^{0-1}-n^2$.
- The last electron added to each element is filled in f-orbital. These two series of elements are hence called the Inner-Transition Elements (f-Block Elements)
- They are all meats.
- Actinoid elements are radioactive.
- The elements after uranium are called Trans uranium Elements.

Metals, Non metals and Metalloids:

Metals: Metals comprise more than 78% of all known elements and appear on the left side of the Periodic Table.

- Metals are usually solids at room temperature except mercury which is liquid at room temperature and gallium and caesium are also liquid at slightly above room temperature.
  Metals usually have high melting (except gallium 303K and caesium - 302 K) and boiling points.
- They are good conductors of heat and electricity.
- They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires).

Non-Metals:

- Non-metals are located at the top right hand side of the Periodic Table.
  Nod-metal's are usually solids or gases at room temperature with low melting and boiling points (boron and carbon are exceptions).
  They are poor conductors of heat and electricity.
- Most nonmetallic solids are brittle and are neither malleable nor ductile.
- The elements become more metallic as we go down a group; the nonmetallic character increases as one goes from left to right across the Periodic Table.

- **Metalloids:**
- The elements e.g., silicon, germanium, arsenic, antimony and tellurium show properties that are characteristic of both metals and nonmetals. These elements are called Semi-metals or Metalloids.

Periodic Trends in Properties of Elements :

- Periodic trends in certain physical and chemical properties can be explained in terms of number of electrons and energy levels.

Trends in' Physical Properties

Atomic Radius: the term atomic radius is used to refer to both covalent and metallic radius depending on whether the element is a non metal or a metal.

- **Covalent Radius:** It is half the distance between two atoms when they are bound together by a single bond in a covalent molecule. It is generally used for non metals.

Metallic Radius: It is taken as half the internuclear distance separating the metal cores in the metallic crystal.

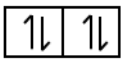
- The term Atomic radius is used to refer to both covalent and metallic radius rating the metal cores in the metallic crystal.
  The atomic size generally decreases across a period because within the period the outer electrons are in the same valence shell and the effective nuclear charge increases us the atomic number increases resulting in the increased attraction of electrons to the nucleus.

Exceptional Case:

(i) $I.E_1$, of Boron < $I.E_1$ of Beryllium

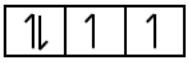
$1s^2 2s^2 2p^1$ $1s^2 2s^2$

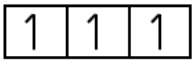




(2p electron is more shielded from nucleus thus easier to remove)

(2) $I.E_1$ of Oxygen < $I.E_1$ of Nitrogen





(more electronic repulsion hence easier to remove outer electron)

(Half filled more stable E.C)

Electron Gain Enthalpy :

- When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the Electron Gain Enthalpy $\left(\Delta_{eg}H\right)$.

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

- Group 17 elements (the halogens) have very high negative electron gain enthalpies because they can attain stable noble gas electronic configurations by picking up an electron.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.
- Electron gain enthalpy becomes more negative with increase in the atomic number across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom since the added electron on an average would be closer to the positively charged nucleus. Electron gain enthalpy become less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus.

Exceptional case : Electron gain enthalpy of O or F is less negative than that of the succeeding element. This is because when an electron is added to O or F, the added electron goes to the smaller $n = 2$ quantum level and suffers significant repulsion from the other electrons present in this level. For the $n = 3$ quantum level (S or Cl), the added electron occupies a larger region of space and the electron-electron repulsion is much less.

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. However, a number of numerical scales of electronegativity of elements viz. Pauling scale, Mulliken-Jaffe scale, Allred-Rochow scale have been developed. The one which is the most widely used is the Pauling scale.

Atomic radius increases regularly as we descend the groups, because the inner energy levels are filled with electrons, which serve to shield the outer electrons from the pull of the nucleus. Consequently the size of the atom increases as reflected in the atomic radii.

- **Ionic Radius :** The removal of an electron from an atom results in the formation of a cation, whereas pain of an electron leads to an anion.
- A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

  Some atoms and ions which contain the same number of electrons, known at isoelectronic species.
- The cation with the greater positive charge will have a smaller radius because of the greater attraction of the electrons to the nucleus. Anion with the greater negative charge will have the target radius.

- **Ionization Enthalpy :** It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.

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

  The second ionization enthalpy (IE I_2) as the energy required to remove the second most loosely bound electron;

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

  Ionization enthalpies are always positive.

  $IE_3 > IE_2 > IE_1$

Ionisation energy depends upon : (i) the attraction of electrons towards the nucleus, and (ii) the repulsion of electrons from each other.

- The effective nuclear charge experienced by a valence electron in an atom will be less than the actual charge on the nucleus because of "shielding" or "screening" of the valence electron from the nucleus by the intervening core electron. Shielding is effective when the orbitals in the inner shells are completely filled.
- Shielding effect order, s electrons > p electrons > d electrons > f electrons, Across a period, increasing nuclear charge outweighs the shielding. Consequently, the outermost electrons are held more and more tightly and the ionisation enthalpy increases across a period.
- As we go down a group, the outermost electron being increasingly farther from the nucleus, there is an increased shielding of the nuclear charge by the electrons in the inner levels. In this case, increase in shielding outweighs the increasing nuclear charge and the removal of the outermost electron requires less energy and hence I.E decreases down a group.
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Ionic Radius :

- The removal of an electron from an atom results in the formation of a cation, whereas gain of an electron leads to an anion.
- A cation is smaller than its parent atom because it has fewer electron while its nuclear charge remains the same. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.
- Some atoms and ions which contain the same number of electrons, known as isoelectronic species.
- The cation with the greater positive charge will have a smaller radius because of the greater attraction of the electrons to the nucleus. Anion with the greater negative charge will have the larger radius.

- **Ionization Enthalpy :** It represents the energy required to remove an electron from an isolated gaseous atom (X) in its ground state.

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

  The ionization enthalpy is expressed in units of $kJmol^{-1}$

  The second ionization enthalpy (IE 2_2) as the energy required to remove the second most loosely bound electron;

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- $\therefore I IE_3 > IE_2 > IE_1$ Ionisation energy depends upon : (i) the attraction of electrons towards the nucleus, and (ii) the repulsion of electrons from each other.
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- Shielding effect order: s electrons > p electrons > d electrons > f electrons.
- Across a period, increasing nuclear charge outweighs the shielding. Consequently, the outermost electrons are held more and more tightly and the ionization enthalpy increases across a period.
- As we go down a group, the outermost electron being increasingly farther from the nucleus, there is an increased shielding of the nuclear charge by the electrons in the inner levels. In this case, increase in shielding outweighs the increasing nuclear charge and the removal of the outermost electron requires less energy and hence I.E decreases down a group.
- Electronegativity generally increases across a period from left to right (say from lithium to fluorine) and decrease down a group (say from facts to astatine) in the periodic table.

Table: Electronegativity Values (on Pauling scale) Across the Periods

Atom (period II)	4+	He	R	6	N .	θ	E.
Electronegativity	1.0	1.5	2	2.5	3.0	3.5	4.0
Atom (Period III)	Na	Mg	N	SI	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

- The attraction between the outer (or valence) electrons and the nuclear increases as the atomic radius decreases in a period. The electronegativity also increases. On the same account electronegativity values decrease with the increase in atomic radii down a group. The trend is similar to that d ionization enthalpy.
- The electronegativity is inversely related to the metallic properties of elements

Periodic Trends in Chemical Properties

Periodicity of Valence or Oxidation States :

- The valence of representative elements is usually equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons as shown below:

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

- The term oxidation state is frequently used for valence. The oxidation state of an element in a particular compound can be defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule.
- Transition elements and actinoids, exhibit variable valance

- **Anomalous properties of Second Period Elements :** The first element of each of the groups 1 (lithium) and 2 (beryllium) and groups 13-17 (boron to fluorine) differs in many respects from the other members of their respective group.
- The behaviour of lithium and beryllium is more similar with the second element of the following group i.e, magnesium and aluminium, respectively. This sort of similarity is commonly referred to as diagonal relationship in the periodic properties.
- The anomalous behaviour is attributed to their small size, large charge/ radius ratio and high electronegativity of the elements.
-

Periodic Trends and Chemical Reactivity:

- All chemical and physical properties are a manifestation of the electronic configuration of elements.
- The maximum chemical reactivity at the extreme left (among alkali metals) is exhibited by the loss of an electron leading to the formation of a cation and at the extreme right (among halogens) shown by the gain of an electron forming an anion. This property can be related with the reducing and oxidizing behaviour of the elements.
- The metallic character of an element, which is highest at the extreme left decreases and the non-metallic character increases while moving from left to right across the period.

Past Years ONE-LINERS

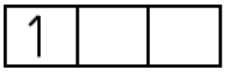
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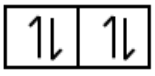
- IUPAC official name for elements having atomic number greater than 100.
  Unnilunium – Mendelevium
  Unniltrium – Lawrencium
  Unnilhexium – Seaborgium
  Ununnnium – Roentgenium
- The normal oxide formed by the element on extreme left are the most basic. (BaO), whereas that formed by the element on extreme right is the most acidic (Cl_2O_7). Oxides of the element in the centre are amphoteric (Al_2O_3). Neutral oxide (CO) have no acidic or basic properties.
- First ionisation enthalpy increases across the period. However anomaly arises in case of stability of electronic configuration due to half filled or completely filled atomic orbital.
- First ionisation energy of Be is greater than B and that of nitrogen is greater than oxygen. Periodic trend in ionic size, ionisation energy, electron gain enthalpy and metallic radius.
- For isoelectronic species, the radius decreases with increase in nuclear charge.
- Order of electron gain enthalpy of 16^{th} group element: $S > Se > Te > 0$
- Due to interelectronic repulsion electron gain enthalpy of oxygen become less than Te.
- Atomic radius decreases along a period while electronegativity, electron gain enthalpy and ionisation enthalpy increases along a period.

Exceptional Case:

(i) $I.E_1$, of Boron < $I.E_1$ of Beryllium

$1s^2 2s^2 2p^1$ $1s^2 2s^2$





(2p electron is more shielded from nucleus thus easier to remove)

(2) $I.E_1$ of Oxygen < $I.E_1$ of Nitrogen





(more electronic repulsion hence easier to remove outer electron)

(Half filled more stable E.C)

Electron Gain Enthalpy :

- When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the Electron Gain Enthalpy $(\Delta_{eg}H)$.

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

- Group 17 elements (the halogens) have very high negative electron gain enthalpies because they can attain stable noble gas electronic configurations by picking up an electron.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to a very unstable electronic configuration.
- Electron gain enthalpy becomes more negative with increase in the atomic number across a period. The effective nuclear charge increases from left to right across a period and consequently it will be easier to add an electron to a smaller atom since the added electron on an average would be closer to the positively charged nucleus. Electron gain enthalpy become less negative as we go down a group because the size of the atom increases and the added electron would be farther from the nucleus.

- **Exceptional case :** Electron gain enthalpy of O or F is less negative than that of the succeeding element. This is because when an electron is added to 0 or F, the added electron goes to the smaller $n = 2$ quantum level and suffers significant repulsion from the other electrons present in this level. For the $n = 3$ quantum level (S or Cl), the added electron occupies a larger region of space and the electron-electron repulsion is much less.

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. However, a number of numerical scales of electronegativity of elements viz., Pauling scale, Mulliken-Jaffe scale. Allred-Rochow scale have been developed. The one which is the most widely used is the Pauling scale.
- Electronegativity decreases while size increases down the group.
  Electron affinity becomes more negative across a period and become by negative down the group.
- Isoelectronic species are the atoms or ions which contain same number electrons.
- Ionisation energy increases along a period. Hence, alkali metals have the lowest ionisation energy.
- For isoelectronic species, size of anion increases as negative charge increases.

- **Tips/Tricks/Techniques ONE-LINERS**
  (Exam Special)
- Period no. is equal to the valence shell present in the configuration. Also first s and p -block elements.
  Period no. = valence shell, for d -block = shell +1, for f block = shell +2 and soon
- Heaviest solid metal : Os (highest density $22.6\ g/cm^3$)
- The basic character of hydroxides increases in the order:
  $CsOH > RbOH > KOH > NaOH > LiOH$
  The acidic character of oxides and hydroxides decreases in the order:
  $HClO_4 > HBrO_4 > HIO_4$
- The non-metallic character, oxidising character and acidic nature of oxides of the elements increases from left to right in a period and decrease from top to bottom in a group.

Na_2O Strongly basic	MgO Besic acidic	Al_2O_3 Amphoteric	SiO_2 Weakly acidic	P_4O_{10} acidic	SO_2 Strongly acidic	Cl_2O_7 Very Strongly acidic

- 2nd period elements shows diagonal relationship with 3rd period elements.
- All transuranic elements are radioactive and artificial.
- Electronegativity percent -5-character of hybridised atom.
- Mulliken scale According to Mulliken, electronegativity is average value of ionisation potential and electron affinity of an element.

$$X = \frac{\text{Ionisation potential} + \text{Electron affinity}}{2}$$

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PW Web/App - https://smart.link/7wwosivoicgd4

Library- https://smart.link/sdfez8ejd80if