Periodic Table -

Periodicity in properties -

In 1869 - 5 years after John Newlands put forward his Law of Octaves, Russian Chemist called Dmitri Mendeleev published a periodic table -

Mendeleev's law: Properties of elements are periodic function of their atomic weight

Major limitations :

1. Controversial position of hydrogen

electropositive H resembled Group I alkali metals

valency -1 and also halogen → diatomic nonmetal.

2. Increase in atomic mass was not regular while moving from one element to another element.

3. Discovery of isotopes.

Properties of elements are not function of atomic weights but atomic number

Atomic weight arises due to existence of protons and neutrons within the nucleus. While physical properties of elements may be related to atomic weight. chemical properties depend on electrons that have negligible mass. It is the electrons which participate in chemical reactions. So, properties of elements will depend on atomic no. & not atomic mass.

Modern Periodic law came to picture in order to address the limitations of Mendeleev's periodic table.

Modern periodic law

1. Properties of elements are periodic function of their atomic no.

2. Periodicity in properties arises due to repetition of similar outer shell electronic configuration at certain regular intervals.

In a Periodic table,

period ≡ the no. of electronic shells present in an atom.

group ≡ the no. of electrons present in the outermost electronic shell and thus this number also corresponds to the maximum valency or oxidation state that an atom can exhibit.

Oxidation no.

H_2SO_4 : $2 + x - 8 = 0 \Rightarrow x = 6$

SO_2 : $x - 4 = 0 \Rightarrow x = 4$

(3) Modern periodic table is based on the modern periodic law in which elements are arranged in increasing order of their atomic nos.

Horizontal Rows → Periods ≡ no. of electronic shell

Vertical Column → Groups ≡ no. of electrons present in valence shell

(4) Modern periodic table consists of 7 periods and 18 groups.

(5) Period indicates the value of 'n' (principle quantum number) for the outermost or valence shell.

$2 \cdot n^2$

$2 \cdot 3^2$

(6) In a particular group, same no. of electrons are present in the outermost orbitals.

Bohr bury scheme

The outermost orbital has similar valence shell electronic configuration.

IUPAC Nomenclature for elements with atomic no. ? 100

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t

Digit	Name	Abbreviation
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

110 - un un nilium Ununnilium

Classification of elements

s block

p-block

d-block

f-block

Valence shell Electronic configuration	Nature	Position in modern periodic table
s block elements : ns^{1-2} $(n = 1-7)$	Metals Alkalie alkaline earth metals	1 & 2 group elements

Elements	Valence shell electronic configuration	Nature	Position in modern periodic table
p - block element	$ns^2 np^{1-6}$ $n = 2 - 7$	Nonmetals Metalloids & some are metals als	Group 13 - 18
d - block Elements	$(n-1)\, d^{10}$ ns^{1-2} $(n = 4 - 7)$	Metals Transition elements	3 - 12

3d series : Sc (21) – Zn (30)
4d " Y (39) – Cd (48)
5d " La (57) , Hf (72) to Hg (80)

d- block elements : Transition elements
These are element that have partially or incompletely filled d-orbital in their ground state or most stable oxidation state.

General properties of d-block elements

1. Form stable complexes.
2. Have high M.P. & B.P.
3. Large charge / radius ratio
4. Form compounds which are often paramagnetic.
5. Hard, high density
6. Form compounds with profound catalytic activity.

7. Show variable oxidation states.

8. Form colored ions and compounds

i) d-d transition
ii) Charge transfer

Color arises

$4s^2, 3d^9$ ✗

$Cu^{29}: 4s^1, 3d^{10}$

Cu^+ Cu^{2+}

All transition metals give bright coloration in flame test.

f - block

Electronic configuration.

$(n-2)f^{1-14}(n-1)s^2(n-1)p^6$

$(n-1)d^{0-1}ns^2 \quad (n = 6 \,\&\, 7)$

Radioactive.

Group - 3

4f series - Lanthanides - 14 elements

Ce (58) - Lu (71)

5f series - Actinides - 14 elements

Th (90) to Lw (103)

Why lanthanides & actinides are placed separately in periodic table?

Electronic configuration that gives rise to distinctive properties.

Most important property $\longrightarrow$ radioactivity

Application: Medical application, Nuclear power.

Periodicity in Atomic Properties

Atomic radius :

1. In a given period, atomic radius will decrease from left to right. This is because the nuclear charge increases while the electrons are being added to the same shell.

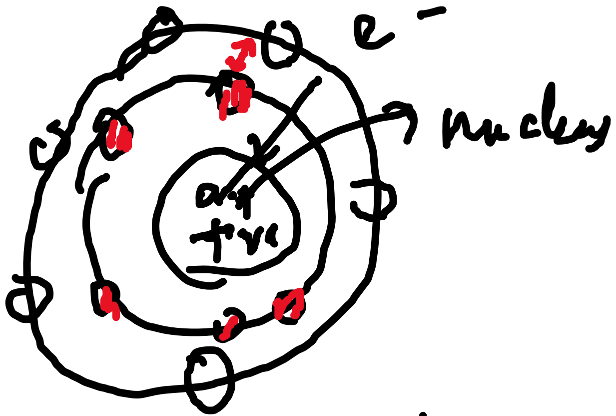
2. Down the group atomic radius increases due to the increase in no. of shells.

(3). In the first transition series the atomic size slightly decreases from Sc to Mn. Why?

When we move from Sc to Mn the effect of effective nuclear charge is stronger than the shielding effect. The atomic size from Fe to Ni remains almost the same because both the effects balance each other.

Screening effect – shielding effect

Screening effect occurs when the nucleus reduces its force of attraction on the valence electrons due to the presence of the inner-shell.



Outer electrons experience attraction from the nucleus and repulsion from the inner electrons.

As the attraction between the nucleus & outer electrons decreases, the repulsion increases between the inner electrons and outer electrons.

The atomic size from Cu to Zn slightly increases because shielding effect is more than effective nuclear charge due to d^{10} structure of Cu & Zn.

Sc – Mn	Fe – Ni	Cu $\longrightarrow$ Zn
atomic size slightly decrease	almost same	slightly increase

Inner Transition Elements –

As we move along the lanthanide series, there is a decrease in atomic as well as ionic radius.

Atomic and ionic radii are distances away from the nucleus or central atom having different periodic trends. When we talk about 'atomic radii' "atomic" refers to the distance away from the nucleus. Ionic radius is the distance away from the central atom.

The ionic radius is the distance from the nucleus to the outermost electrons in an ion whereas atomic radius is the distance from the nucleus to the normal valence electrons.

In case of lanthanides, the decrease in size is regular in ions but not so regular in atoms.

Lanthanide Contraction :

Lanthanide contraction is the steady decrease in the size of atoms & ions in rare earth elements with increasing atomic no from lanthanum (atomic no. 57) through lutetium (atomic no. 71)

Lanthanide Contraction arises due to the poor shielding effect of 4f electrons.

Because the elements in Row 3 / 3rd period have 4f electrons, these electrons do not shield good, causing a greater nuclear charge. This greater nuclear charge has a greater pull on the electrons.

'd' & 'f' orbitals are associated with two (d) or three (f) nodal planes through the nucleus. This means that the

F:

corresponding electrons are further away from the nucleus. Consequently, their shielding less effective than that of s & p orbitals.

- Which orbital has the highest shielding effect?

  's' orbital → because the inner electrons shield or screen the nuclear forces from reaching the outer electrons.

Consequences of Lanthanide Contraction:

1. Increase in electronegativity
2. Increase in ionisation energy.
3. Hard to separate the lanthanides.
4. Ability of complex formation decreases.

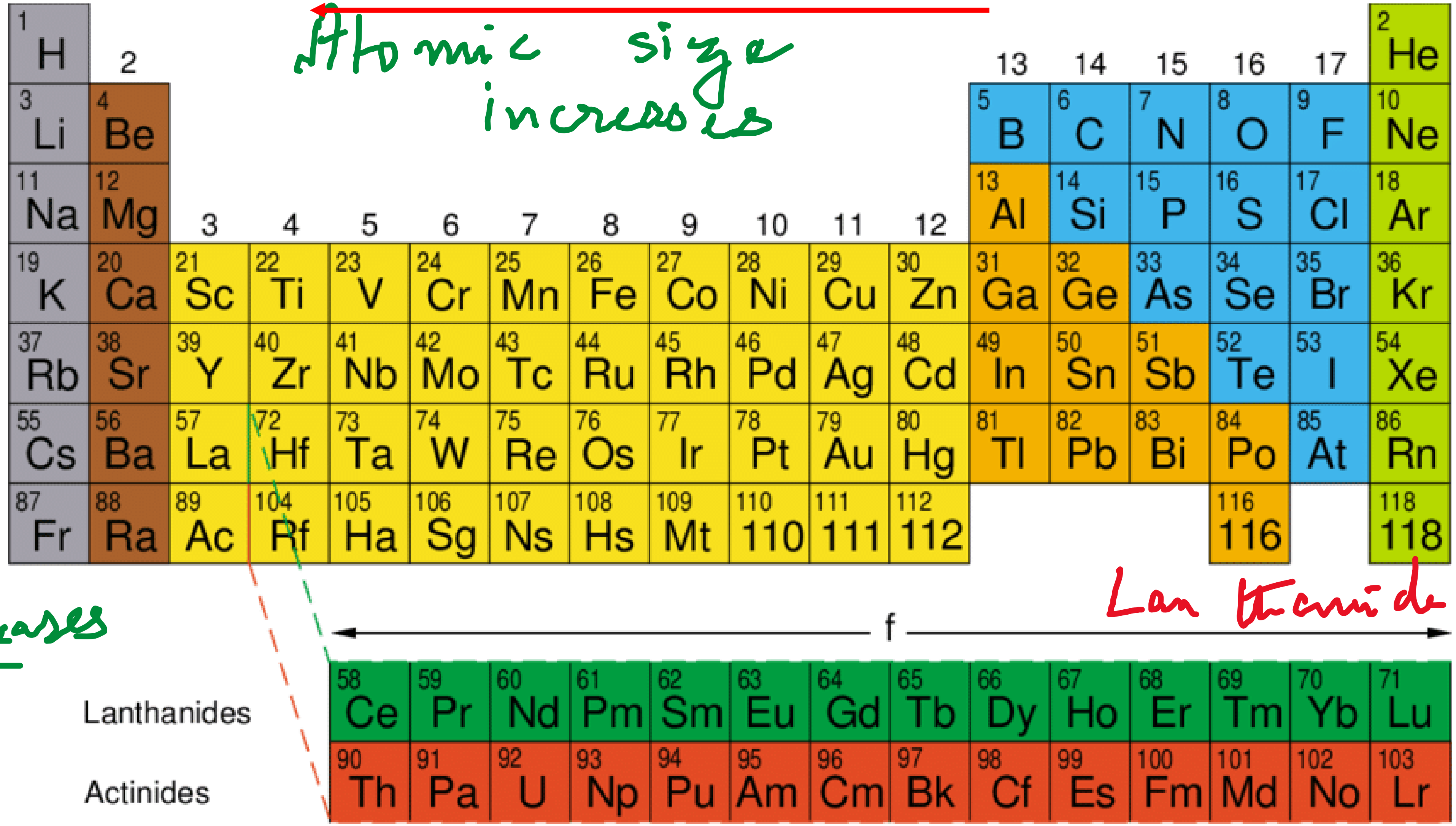
To separate lanthanides from other elements occuring with them, they are chemically combined with specific substances to form lanthanide compounds with low solubility (oxalates & fluorides, for e.g). A process known as exchange is then used to separate lanthanides from each other.

Lanthanides are difficult to separate from each other, because of similarities in their physical & chemical properties. Most separation process takes advantage of a small decrease in ionic radius that occurs across the lanthanide series.

For an ideal ligand, the decrease in ionic radius would result in steadily increasing extraction across the series. This means that the ligands would captiow more lute-tium (the lanthanide with the smallest radius) than lanthanum (with the largest radius)

Modern Periodic Table

Atomic size increases



Increases

Lanthanide Contraction

2. Ionization Potential or Ionization Energy.

Ionization energy is the minimum amount of energy required to remove the most loosely bound electron of an isolated neutral gaseous atom or molecule

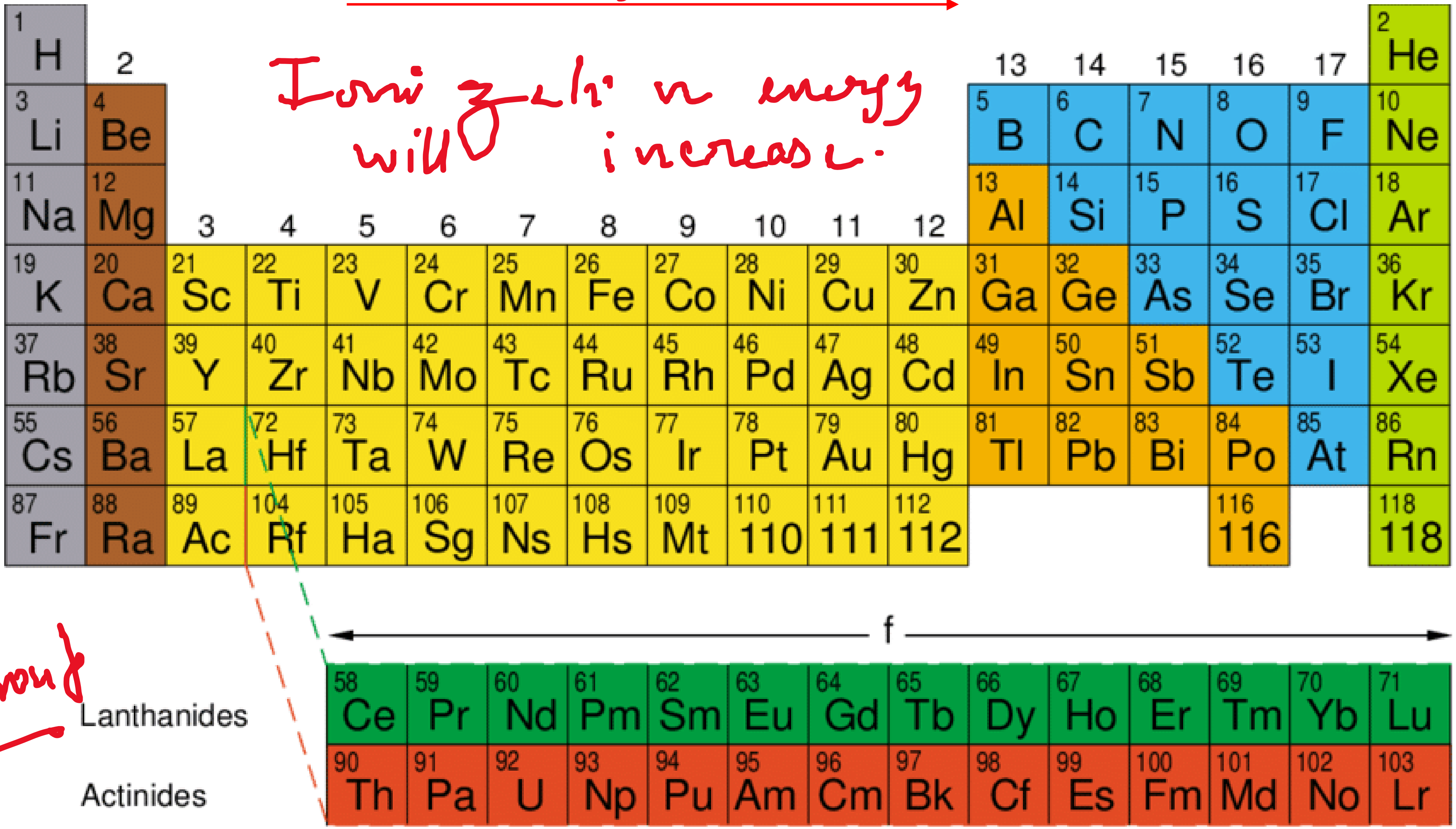
valence shell electron

In the gaseous state, atoms are widely separated, therefore interatomic forces are minimum. Due to this reason, the term "isolated gaseous atom" has been included in definition of ionisation enthalpy/energy / electron gain enthalpy

Periodic Trend:

Left to right across the period

Ionization energy will increase.



decreases down the group

Ionization energy increases along the period while decreases down the group.

Factors which influence I.E.

1. Atomic size: Larger the size of the atom, small the I.E.

2. Effective nuclear charge: Greater the effective charge on the nucleus of an atom, more difficult

would be the removal of an electron from the atom because electrostatic force of attraction between the nucleus and the outermost electron increases. So, more energy will be required to remove the electron.

③ Penetration effect of orbitals.

The order of energy required to remove electrons from s, p, d & f orbitals are s > p > d > f.

(9) Shielding or screening effect –

Screening effect results in decrease of force of attraction between the nucleus and the outermost electron and lesser energy is required to separate the electrons. Thus value of I.P. decreases.

Stability of half filled and fully filled orbitals.

According to Hund's rule, the stability of half filled or completely filled degenerate orbitals is comparatively high. So, more energy will be required to separate electrons from such atoms.

Successive Ionization energy ;

Element	IE_1	IE_2	IE_3	IE_4
Li	520.1	7297	11813	
C	1086.2	2352	4620	6225
N	1402.1	2856	4577	7474

$I.E_1 < I.E_2 < I.E_3 < \ldots\ldots$ so on

Once an electron is removed from an atom, removal of the next electron becomes even more difficult as the atoms approach closer to the inert gas configuration

3. Electron Affinity : Electron affinity is defined as the change in energy (in kJ / mole) of a neutral atom (in the gaseous phase) when an electron is added to the atom to form a negative ion.

Periodic Trend.

Electron affinity increases along the period and decreases down the group.

Factors affecting electron affinity:

1. Atomic size: Electron affinity value decreases with increasing atomic size / radius because electrostatic force / of attraction decreases between

the electron being added and the atomic nucleus due to increase of distance between them.

(2) Effective nuclear charge: Electron affinity value of an element increases as the effective nuclear charge on the atomic nucleus increases because electrostatic force of attraction between the electron being added and the

nucleus increases. As the electrostatic force of attraction increases, amount of energy released is more.

(3) Screening or shielding effect -

As the shielding effect or screening effect increases, the electron affinity value of the elements decreases. The shielding effect between the outer electrons of the nucleus increases as the number of electrons increases in the inner shells.

(4) Stability of half filled & completely filled orbitals.

The stability of half filled & completely filled orbitals of a sub-shell is comparatively more. Therefore, it is difficult to add an extra electron in such orbitals, which already possess a stable electronic configuration. So, lesser energy is released on addition of electrons; hence electron affinity values will decrease.

The electron affinity of fluorine, which is higher up in the group than chlorine is lower.

Anomalous behavior of F :

F has lower E.A. than Cl because the electrons in the outermost shell of a fluorine atom are closer together. Energy is required to keep the gained electrons in the shell, causing F to have a smaller electron affinity than chlorine inner shell → shielding

F : $1s^2\, 2s^2\, 2p^5$

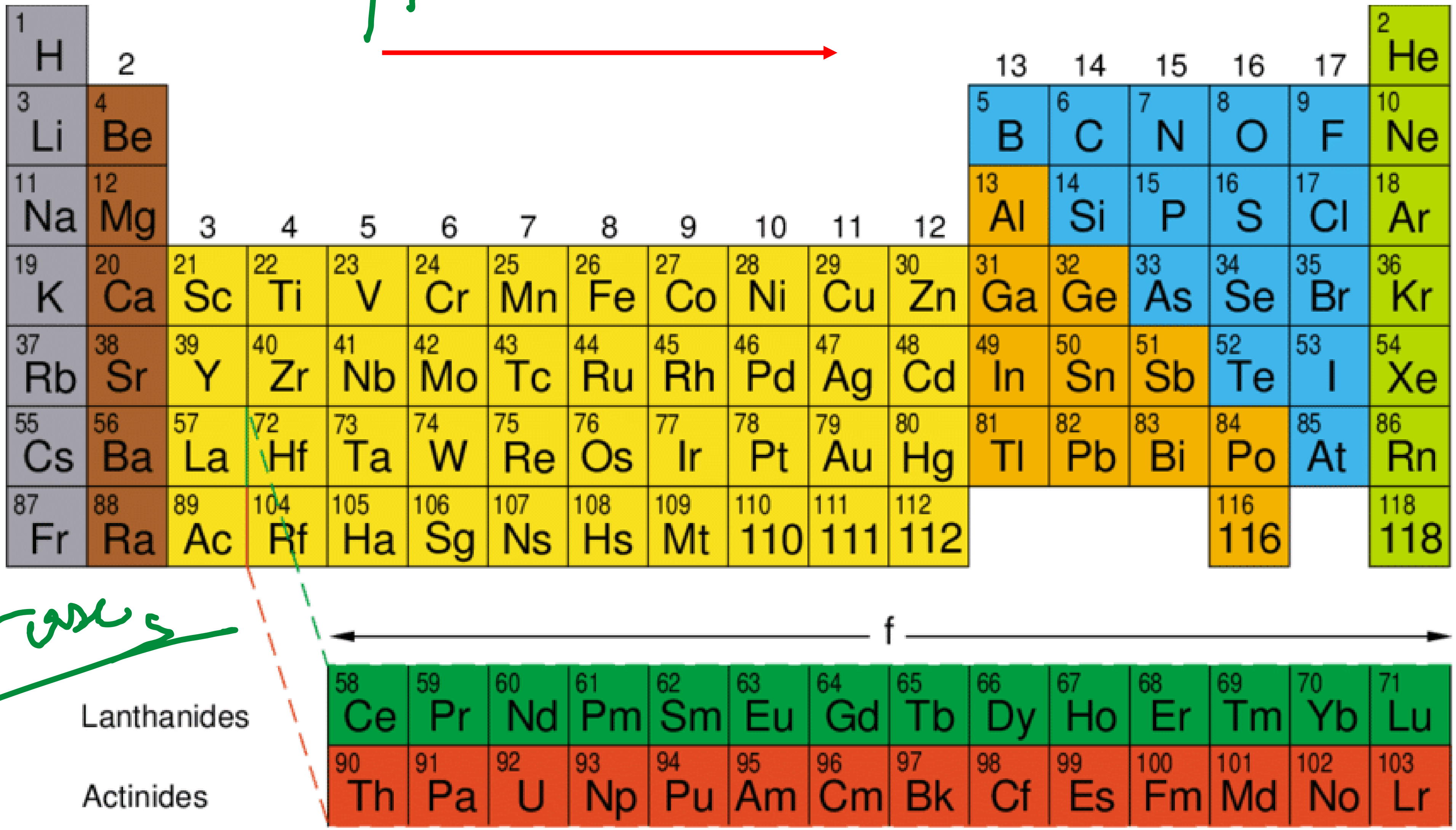
Cl : $1s^2\, 2s^2\, 2p^6,\, 3s^2\, 3p^5$

2
3

Electron - electron repulsion occurs between the electrons where n = 2

The elements in the second period have such small electron clouds that electron repulsion is greater than that of the rest of the family

Electron affinity increases



decreases

(9) Electronegativity

CaH_2 H–I I–Cl

+ve ← R – Cl R–MgCl –ve

Electronegativity (χ) measures the tendency of an atom in a molecule to attract the shared pair of electrons.

Electronegativity scales : There are some arbitrary scales that are used to measure electronegativity quantitatively.

Pauling's scale: If x_A and x_B are the electronegativity of atoms A & B respectively then,

$$\left(\frac{1}{0.208}\right)^2$$

$$0.208\sqrt{\Delta_{AB}} = x_A - x_B$$

if $x_A > x_B$

or, $\Delta_{AB} = 23.06\,(x_A - x_B)^2$

where $\Delta_{AB} = E_{A-B}$ (experimental) $- E_{A-B}$ (theoretical)

E_{A-B} is the energy of A-B bond.

In a purely covalent molecule AB, the experimental and theoretical values of bond energy AB are equal.

So, $\Delta AB = 0$

$\Rightarrow \quad 0 = 23.06\,(x_A - x_B)^2$

or, $x_A = x_B$

The factor 0.208 appears from the conversion of bond energy data in electron-volt per molecule to kilocalories per mole.

$$1 \text{ eV molecule}^{-1} \equiv 23.06 \ \frac{\text{k cal}}{\text{mol}}$$

$$\frac{1}{\sqrt{23.06}} = 0.208$$

In an ionic molecule, AB,

E_{A-B} (experimental) > E_{A-B} (theoretical)

Pauling assumed the electronegativity value of fluorine to be 4 and calculated the electronegativity values of other elements from this value.

Mulliken's electronegativity

ionization potential

$$\text{Electronegativity} = \frac{\text{Electron Affinity} - \text{I.P}}{2}$$

$$x_M = \frac{I.E + E.A}{2} \text{ (both in eV)}.$$

x_M with x_p may be related as

$$x_p = 1.35\sqrt{x_M} - 1.37$$

- <u>Alfred Rochow's electronegativity :</u>

If the distance between the circumference of outermost shell and the nucleus is r and effective nuclear charge Z_{eff} then

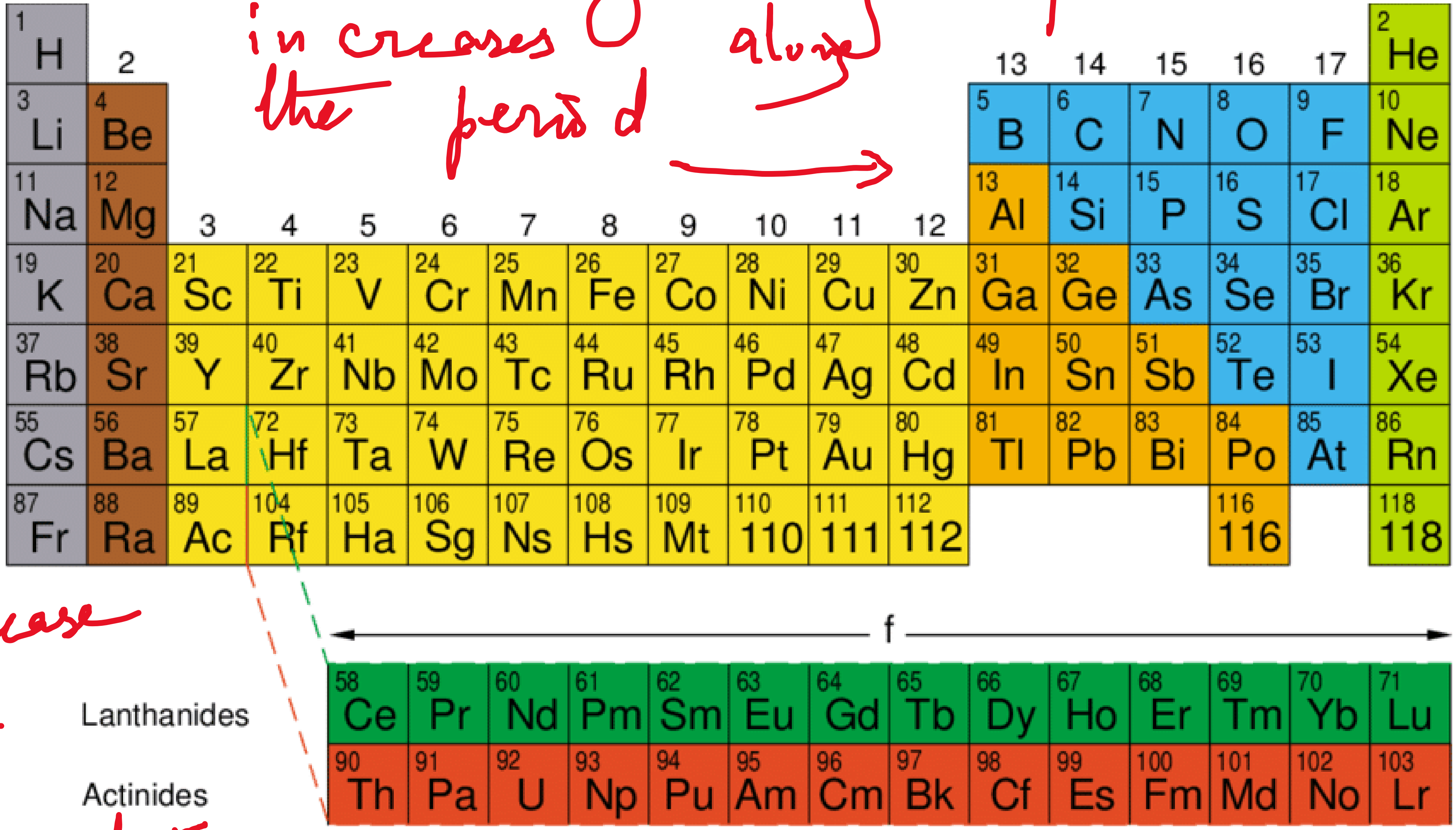
$$\frac{Z_{eff} \cdot e^2}{r^2} = \frac{0.359\, Z_{eff}}{r^2} + 0.744$$

$$Z_{eff} = Z - \sigma$$

Z = actual charge present on the nucleus i.e. no. of protons

σ = shielding constant.

Electronegativity of elements increases along the period



decrease down the group

Factors affecting electronegativity

1. **Atomic radius**: As the atomic radius of an element increases, the electronegativity value decreases.

2. **Effective nuclear charge**: The electronegativity value increases as the effective nuclear charge on the atomic nucleus increases.

(3) Oxidation state :- As the oxidation state increases, electronegativity increases.

(4) Hybridisation state of an atom in a molecule.

If the s- character in the hybridisation state of the atom increases, electronegativity also increases.

Hybridisation states	s - character	Electronegativity
sp^3	25%	2.48
sp^2	33.33%	2.75
sp	50%	3.25

CH_4 $\quad$ C_2H_4 $\quad$ C_2H_2 $\quad$ sp

$NaNH_2$/ liq. NH_3

$H \rightarrow C \equiv C \leftarrow H$

more acidic

more electronegativity

5) Valency It is the number of univalent atom which can combine with an atom of the given element

a) Valency is given by the number of electrons in outer most shell

b) If no. of valence electrons ≤ 4

valency = number of valence electrons

c) If no. of valence electrons > 4

Valency = 8 − no. of valence electrons.

O : 8 − 6 = 2

N : 8 − 5 = 3

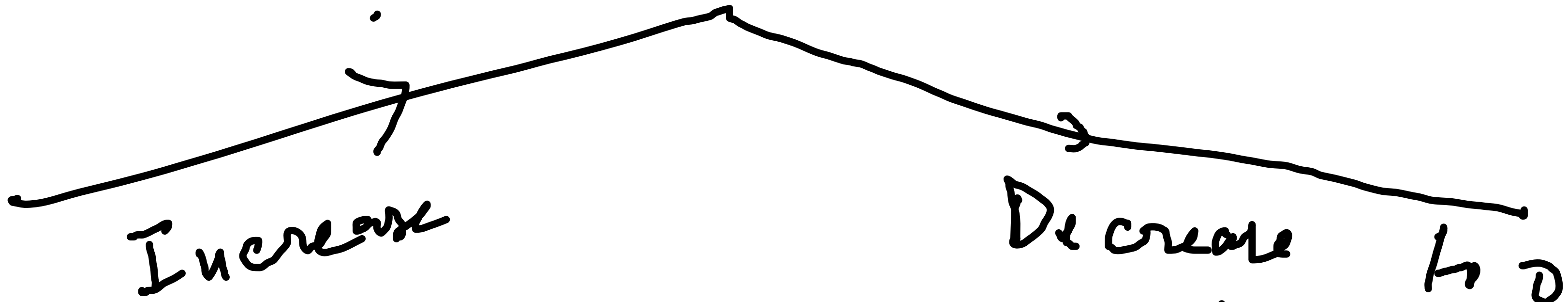
Cl : 8 − 7 = 1

d) Many elements exhibit variable valency, particularly transition metals.

e) Variation in period →

Increases from 1 → 4 and then decreases 4 → 0

1	2	3	4	3	2	1	0
Li	Be	B	C	N	O	F	Ne



f) Variation in a group. No change.

All elements belonging to a particular group will exhibit the same valency.

(E) Metallic character of an Element.

a) Non metallic elements have strong tendency to gain electrons.

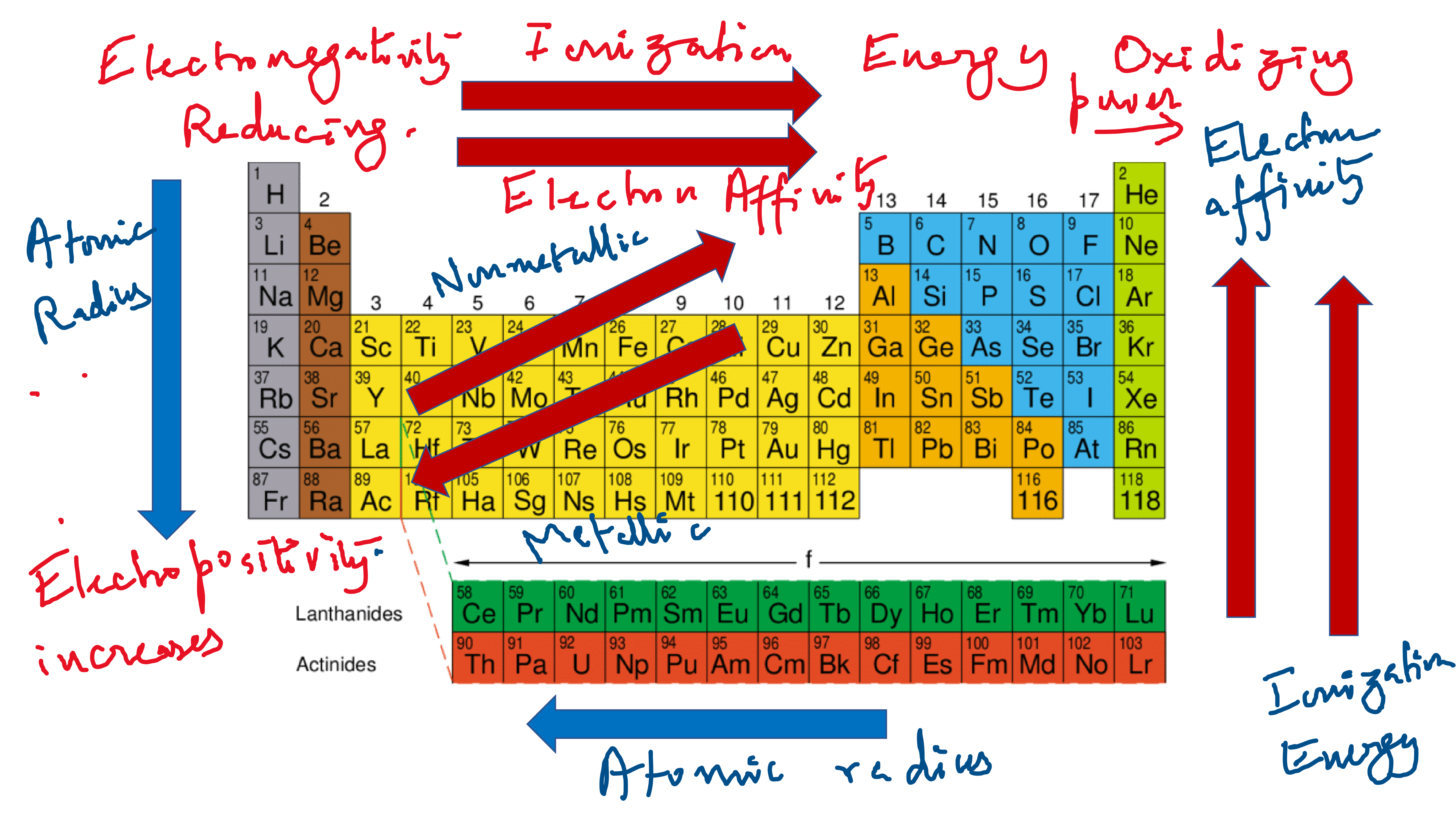
b) Non metallic character is directly related to electronegativity and metallic character is inversely related to electronegativity

c) Across a period, electronegativity increases. Therefore, metallic character will decrease.

d) Down the group, electronegativity decreases, hence non-metallic character increases.

7) Diagonal Relationship -

Diagonal relationship is said to



Electronegativity
Ionization Energy
Oxidizing power
Reducing.
Electron Affinity
Electron affinity
Atomic Radius
Nonmetallic
Metallic
Electropositivity increases
Ionization Energy
Atomic radius
Lanthanides
Actinides
f
1 H
2 He
3 Li 4 Be 5 B 6 C 7 N 8 O 9 F 10 Ne
11 Na 12 Mg 13 Al 14 Si 15 P 16 S 17 Cl 18 Ar
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
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
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
87 Fr 88 Ra 89 Ac 104 Rf 105 Ha 106 Sg 107 Ns 108 Hs 109 Mt 110 110 111 111 112 112 116 116 118 118
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
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

exist between certain pairs of diagonally adjacent elements in the second and third period of periodic table.

Diagonal pairs :-

Li – Mg
Be – Al
B – Si

exhibit similar properties

i) Semiconductors
ii) form halides that are hydrolysed in water
iii) Acidic oxides

Diagonal relationship arises due to the polarizing power i.e. ionic charge / ionic radius, being similar for the diagonally placed elements.

Cause : i) Identical size of atoms.

As we move from left to right in a periodic table, the size of atom reduces. As we move down the group, size of atom gradually increases. Similarly, from left to right electronegativity increases along the period whereas, as we move downward a group elements become more electropositive.

Therefore, diagonally placed elements possess similar properties like density, atomic size, electronegativity, polarizing power and chemical properties.

Cause of diagonal relationship -

1) Polarizing power -

2) Electronegativity -

Periodicity: Atomic size, Electron affinity, Electronegativity, Ionization Potential, Valency, Metallic character.

Bond-breaking → endothermic process.

Bond-making → exothermic

Energy is released when new bonds are formed.

Whether a reaction is endothermic or exothermic, it depends on the energy needed to break the bonds and the energy released when new bonds are formed

Except for diatomic molecules, the bond dissociation energy is different from bond energy. While bond-dissociation energy is the energy of a single chemical bond, the bond energy is the average of all the bond dissociation energies of the bonds of the same type for a given molecule.

Bond Energy	Bond Dissociation Energy -
Average amount of energy needed to break down all bonds between the same two types of atom in a compound	Amount of energy needed to break down a particular bond in homolysis

Etymology: alkaline earth metals

These are named after their oxides, alkaline earth –

Old fashioned names –

Beryllia, Magnesia, Lime, Strontia, Baryta

Basic when combined with water. "Earth" was a term applied by early chemists to nonmetallic substances that are insoluble in H_2O & resistant to heating.

Chemical Periodicity

① Na^+, Mg^+, Al^{3+}, and Si^{4+} are isoelectronic
The order of their ionic size is

a) $Na^{+} > Mg^{2+} < Al^{3+} < Si^{4+}$
b) $Na^{+} < Mg^{2+} > Al^{3+} > Si^{4+}$
c) $Na^{+} > Mg^{2+} > Al^{3+} > Si^{4+}$
d) $Na^{+} < Mg^{2+} > Al^{3+} < Si^{4+}$

Sol^n 'c' – Among the isoelectronic, the cation having more +ve charge will be smaller in size.

② Consider the following changes

$A \rightarrow A^+ + e^- ; E_1$

$A^+ \rightarrow A^{2+} + e^- : E_2$

The energy required to pull out the electrons are E_1 & E_2 respectively. The correct relationship between two energies would be

a) $E_1 < E_2$ b) $E_1 = E_2$ c) $E_1 > E_2$ d) $E_1 \geq E_2$

Correct answer

(a)

IE_1 is always less than $I.E_2$

3. The elements with zero electron affinity are –

a) Boron & carbon

b) Beryllium & helium

c) Lithium & sodium

d) Fluorine and chlorine.

Correct : b

Electronic configuration fully filled.

4. Ionization energies of atoms A & B are 400 & 300 kcal mol^{-1} respectively. The electron affinities of these atoms are 80.0 and 85.0 k.cal mol^{-1} respectively. Then which is the correct

statement regarding electronegativity χ.

a) $\chi_A < \chi_B$ b) $\chi_A > \chi_B$

c) $\chi_A = \chi_B$ d) none of these.

Correct answer : b Electronegativity

$$= \frac{I.E + E.A}{2 \times 62.5}$$

where I.E

& E.A are in k.Cal

5) The set of three elements having successive atomic numbers and having the ionization energies of 2372, 520 and 890 kJ per mole.

A) H, He, Li
b) He, Li, Be
c) Li, Be, B
d) B, C, N.

He ($Z = 2$) is a noble gas and has the highest I.E. The I.E. of Be ($Z = 4$) is more than that of Li ($Z = 3$)

(c) Similarity in chemical properties of the atoms of elements in a group of the Periodic table is most closely related to

a) atomic number
b) atomic masses
c) number of principal energy levels
d) number of valence electrons

Correct answer: 'a'

(7) In the long form of periodic table, the valence shell electronic configuration of $5s^2\ 5p^4$ corresponds to the element present in

a) Group 16 & period 6
b) " 17 & " 6
c) " 16 & " 5
d) " 17 & " 5.

Correct : 'c'

Te has $5s^2 5p^4$ valence shell configuration. It belongs to group 16 & period 5.

(8) Consider the following ionization enthalpies of two elements 'A' and 'B'.

Element	Ionization Enthalpy (kJ/mol)		
	1st	2nd	3rd
A	899	1757	14847
B	737	1450	7731

Which of the following statements is correct?

a) Both 'A' & 'B' belong to group-1 where 'B' comes below 'A'

b) Both 'A' & 'B' belong to group-I where 'A' comes below B.

c) Both 'A' & 'B' belong to group 2 where 'B' comes below 'A'

d) Both 'A' & 'B' belong to group 2 where 'A' comes below B.

'c' Correct answer. Generally

I.E increases from left to right in a period and decreases from top to down in a group.

I.E → depends on several factors

i) atomic radius 2) nuclear charge

3) shielding effect.

1st I.E. of A & B is greater than group I. C Li = 520 kJ mol^{-1} to Cs 374 kJ mol^{-1}

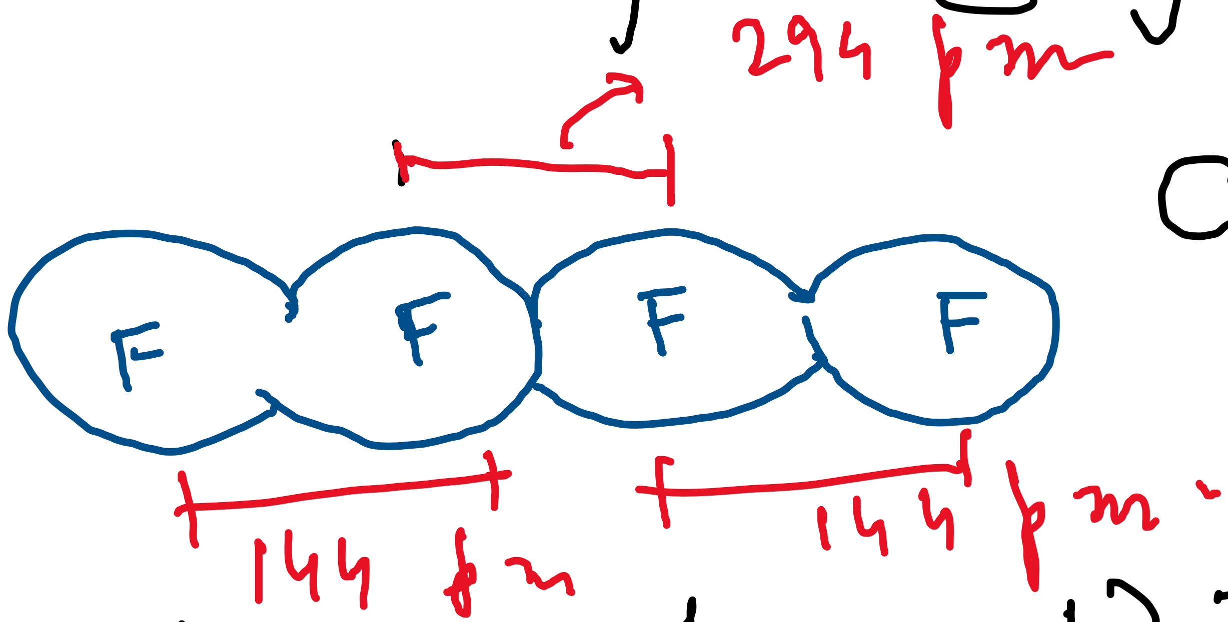
This means element A & B belong to group 2 and all 3 I.E. values are less for element B ⇒ B will come below A.

9) Which of the following has largest ionic radius?

a) Li^+ b) O_2^{2-} c) B^{3+} d) F^-

Li^+ : On moving along a period ionic radii decreases due to increase in effective nuclear charge

(10) The van der Waal and covalent radii of fluorine atom respectively from the following figure are :-



Correct answer :- 'c'

a) 219 pm, 72 pm
b) 75 pm, 72 pm
c) 147 pm, 72 pm
d) 147 pm, 144 pm

Covalent radius is radius of an atom in its bound state i.e - in F_2 it is $\frac{1}{2} \times$ distance b/w 2 covalently bonded F atoms.

van der Waal radii is one-half of the distance between the nuclei of 2 identical non-bonded isolated atoms

(11) The element $Z = 114$ has been discovered recently. It will belong to which of the following family/group and its electronic configuration?

a) Halogen family $[Rn]\,5f^{14}\,6d^{10}\,7s^{2}\,7p^{5}$

b) Carbon family $[Rn]\,5f^{14}\,6d^{10}\,7s^{2}\,7p^{2}$

c) Oxygen family $[Rn]\,5f^{14}\,6d^{10}\,7s^{2}\,7p^{4}$

d) Nitrogen family $[Rn]\,5f^{14}\,6d^{10}\,7s^{2}\,7p^{6}$

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1. b) $Z = 114$ Flerovium

Super heavy artificial element

Transactinide element in p block.

7th period → Carbon family.

$Z = 114$.

$[Rn]_{86}\ 5f^{14}, 6d^{10}, 7s^2, 7p^2$

(14) In which of the following options the order of arrangement does not agree with variation of property indicated against it?

a) $B < C < N < O$ (increasing first I.E.)

b) $I < Br < Cl < F$ (increasing electron gain enthalpy)

c) $Li < Na < K < Rb$ (increasing metallic radius)

d) $Al^{3+} < Mg^{2+} < Na^{+} < F^{-}$ (increasing ionic size)

(13) The species Ar, K^+, and Ca^{+2} contain same number of electrons. In which order do their radii increase?

A) $Ar < K^+ < Ca^{2+}$ B) $Ca^{2+} < Ar < K^+$

C) $Ca^{2+} < K^+ < Ar$ D) $K^+ < Ar < Ca^{2+}$

Ar, K^+, Ca^{2+} → isoelectronic.

For isoelectronic species ionic radii decrease with increase in effective +ve charge

(14) The correct order of the decreasing ionic radii among the following isoelectronic species is

a) $Ca^{2+} > K^{+} > S^{2-} > Cl^{-}$

b) $Cl^{-} > S^{2-} > Ca^{2+} > K^{+}$

c) $S^{2-} > Cl^{-} > K^{+} > Ca^{2+}$

d) $K^{+} > Ca^{2+} > Cl^{-} > S^{2-}$

Correct option : c

Ionic radii $\propto$ charge on anion $\propto \dfrac{1}{\text{charge of cation}}$

(15) The correct order of acidic strength

a) $Cl_2O_7 > SO_2 > P_4O_{10}$

b) $K_2O > CaO > MgO$

c) $CO_2 > N_2O_5 > SO_3$

d) $Na_2O > MgO > Al_2O_3$

Correct option

c)

Acidic character of oxide $\propto$ Non metallic nature of element

Non metallic character increases along the period. Hence the order of acidic character is -

$Cl_2O_7 > SO_2 > P_4O_{10}$

(16) Elements X, Y and Z have atomic numbers 19, 37 and 55 respectively. Which of the following statement is true about them?

x y z : 19, 37, 55

a) Their I.P would increase with increasing atomic no.

b) y would have an intermediate potential between those of x & z.

c) z would have the highest I.P.

d) y would have the highest I.P.

(b) Correct option

17 Consider the following information about elements P & Q:

Trivalent

Divalent

	Period no.	Group no.
P	2	15
Q	3	2

$P - 3$

$Q - 2$

Q_3P_2

The formula of the compound formed by P & Q is

a) QP b) Q_2P_3 c) Q_3P_2 d) Q_2P

Correct answer: c

(16) Electron affinity is positive for

a) $O(g) + e^- \longrightarrow O^-(g)$

b) $S(g) + e^- \longrightarrow S^-(g)$

c) $O^+(g) + e^- \longrightarrow O(g)$

d) $O^-(g) + e^- \longrightarrow O^{2-}(g)$

'd' Correct of him. O^- ion repel incoming electron thus energy is reqd. to add more electron

(19) Which of the following ionic species has the greatest proton affinity to form stable compound.

a) Cl^{-} b) F^{-} Correct c) I^{-} d) Br^{-}

Proton affinity decreases in moving across the period from L → R due to increase in charge, within a group the P.A. decreases from top to bottom

(20) The electronegativity follows the order:

a) $F > O > Cl > Br$ b) $F > Cl > Br > O$

c) $O > F > Cl > Br$ d) $Cl > F > O > Br$

Correct : "a"

F & O belong to 2nd period whereas Cl and Br - 3rd & 4th periods respectively.

Hence sequence of E.N. is $F > O > Cl > Br$

(21) The increasing order of E.A. of N, P & As is -

a) $N < P < As$

b) $As < P < N$

c) $As < N < P$

d) $As < N > P$.

Correct option : 'c'

Phosphorus has vacant 'd' orbitals due to which it has higher E.A. than nitrogen. As ↓ vacant d - orbitals but its atomic size is very big, also has E.A. is less than N.

(22) Which transition involves maximum amount of energy?

a) $M^{\ominus}(g) \longrightarrow M(g) + e^{\ominus}$

b) $M^{\ominus}(g) \longrightarrow M^{+}(g) + 2e^{\ominus}$

c) $M^{\oplus}(g) \longrightarrow M^{2+}(g) + e^{\ominus}$

d) $M^{2+}(g) \longrightarrow M^{3+}(g) + e^{\ominus}$

Correct option: d) The energy involved is I.E. 3^{rd} ionisation energy will be greater than 2^{nd} and 1^{st} ✓

(22)

The chemistry of Li is very similar to that of Mg even though they are placed in different groups. Reason –

a) Both are found together in nature

b) " have nearly same size

c) Both " similar electronic configuration

d) The ratio of their charge & size is nearly same

Correct answer

(24) The first ionization potential of aluminium is smaller than Mg because :

a) Atomic size of Al > Atomic size of Mg.

b) " " " " < " " " Mg.

Correct

c) Al has one electron in p-orbital

d) None of these.

Al ($3s^2 3p^1$) Mg ($3s^2$). Lower energy is reqd. to remove $3p^1$ electron than $3s^1$ electron (penetrating effect s > p > d > f).
Secondly Mg has stable electronic configuration than Al.

25 Which of the following order is wrong?

a) $NH_3 < PH_3 < AsH_3$ — Acidic

b) $Li < Be < B < C$ — IE,

c) $Al_2O_3 < MgO < Na_2O < K_2O$ — Basic

d) $Li^+ < Na^+ < K^{\oplus} < Cs^+$ — Ionic radius.

Correct order :

$Li < B < Be < C$