

Periodic properties :-

(A)

Effective Nuclear Charge :

To explain it we should understand Screening and Penetration. Penetration is the distance of e^- from the nucleus. Screening is defined as the concept of inner e^- blocking the outer electrons from nuclear charge.

Effective Nuclear Charge
 Z_{eff}

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

S = No. of inner electrons.

Z = Total no. of electrons in an atom.

for eg $Cl_{17} = 2, 8, 7$

$$Z_{eff} = 17 - 10 = 7.$$

Inner 10 e^- will screen out the +ve charge of 10 protons.

Z_{eff} shows that nucleus is pulling the outer electrons with +7 charge and therefore the outer electrons are pulled closer to nucleus and atomic radii is smaller.

As the atomic No. Z increases the effective nuclear charge Z_{eff} increases.

So, as we move in a period the atomic no. Z increases, so does Z_{eff} increases. Therefore the nucleus has greater pull on the outer electrons and size decreases.

But in gp, the valance e^- keep the same Z_{eff} but now the orbitals are farther from the nucleus. \therefore Nucleus has lesser pull on outer electrons and atomic radii is larger.

Periodic Properties

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18 Vertical Columns — groups (same valance e'⁻)
Horizontal Rows — periods.

Ionisation Energy (IE): Energy required to remove outermost e'⁻ from an isolated gaseous atom.

Period: Ionisation Energy ↑ as we move from left to right. So the metallic behaviour disappears because there are less e'⁻ shielding out e'⁻ from the pull of nucleus.
gp: IE ↓ as we move from top to bottom as the atomic no. ↑. eg IE of Boron is 800 kJ/mol (semimetal) but other members have IE = 600 kJ/mol.

Electron affinity (ea): Energy released when an e'⁻ is added to neutral atom.
e'⁻ affinity of non-metals are higher than those of metals and in general ea of halogens are strikingly large & (large negative) so it decreases in period. Less negative in gp except N, O, F. In Noble gas ea is zero as they will not easily gain e'⁻.

Electronegativity (en): Tendency of an atom to acquire e'⁻ during bonding. The most common scale is of L. Pauling. based on bond energies. e' with low IE will have low en.

1) Increase in bond energy was a measure of difference in electronegativity

2) Pauling also used the geometrical mean bond energy as his reference for increase in bond energies. So for electronegativities X_A & X_B of two atoms A and B is

$$96 (X_A - X_B)^2 = D(A-B) - [D(A-A) D(B-B)]^{1/2}$$

D = Dissociation energy in kJ/mole.

(2)
R.S Mulliken proposed a more fundamental definition of electronegativity.

The average of an atom's IE & e_a.

Non metals are more electronegative than metals. In gp smaller atoms are more electronegative than larger one.

So, in gp electronegativity decs as we move from top to bottom as the size Inc.

And in period i.e. in rows electronegativity Inc as we move from left to right. So, Fluorine being most electronegative and Cesium being least electronegative element.

Pauling proposed en as a method for predicting the ionic character of bonds.

①

H	②	⑬	⑭	⑮	⑯	⑰
2.2						
Li	Be	B	C	N	O	F
1.0	1.6	2.0	2.6	3.0	3.4	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.3	1.6	1.9	2.2	2.6	3.2
K	Ca	Ga	Ge	As	Se	Br
0.8	1.0	1.8	2.0	2.2	2.6	3.0
Rb	Sr	In	Sn	Sb	Te	I
0.8	1.0	1.8	1.9	2.1	2.1	2.7
Cs	Ba	Tl	Pb	Bi	Po	At
0.8	0.9	1.8	2.1	2.0	2.0	2.2

Diff in electronegativity for the elements ~~serve~~ in table serve as a guide to the charge separation in a bond.

From this table we came to conclusion

that C-H bonds should not be very polar
that N-H and O-H bonds should be quite polar
and that H atom in Li-H should be more
negative than the Li atom.

Oxidation states: - is the charge which an
atom has or appears to have when all
other atoms are removed from it to acquire
a closed shell electronic configuration of
the type ns^2np^6 or nd^{10} .

This tendency is, particularly
clear in 1st and 2nd gp
and among the lighter members of gp 13.
For eg In gp 13, the valance e⁻ configuration
of the atoms is ns^2np^1 and loss of 3 e⁻ to form
+3 Oxidation state results in ions that have
 $(n-1)s^2(n-1)p^6$ or $(n-1)d^{10}$ configurations.

Oxidation states that corresponds to loss of np
electrons and retention of ns^2e^1 also occur
among heavier elements of gp 14, 15, 16, 17
Thus Ti & Pb which have ns^2np^2 configurations
display both +2 & +4 Oxidation states.

In gps (13 to 17) where two or more positive
Oxidation is found, the lower Oxidation
state tend to be more important further
down a column in the Table. This means
that C, Si, Ge involve +4 Oxidation state
while Sn, Pb involve +2 " " :

Negative Oxidation state appear in gp (15).
and are very frequent among the
non metals. Thus Nitrogen and phosphorus
form nitrides and phosphides i.e. N^{3-} & P^{3-}
but As and Sb ~~do not~~ chemistry
show -3 Oxidation state rarely. and
Bismuth does not acquire it virtually.

As far as gp (16) is concerned -2 Oxidation (4)
state is important for all the elements
and as we move to heavier elements
down in the gp the -2 Oxidation is
is not stable.

for gp (17) $\rightarrow -1$ Oxidation state is of
much importance.
The reason for the above behaviour is that
lighter non-metals are relatively more
electronegative.

Now, for transition elements, Variable
Oxidation states are seen. Through gp 3 to 7
Maximum Oxidation state is found for

Sc, Ti, Cr, Mn families.
for eg Sc family have configuration [noble gas]
 $(n-1)d^1 ns^2$ and show only $+3$ Oxidation
state. Mn family $(n-1)d^5 ns^2$ show $+7$ ".

Reason:- Transition elements with a d shell
no more than half filled the maximum
Oxidation state is equal to the no. of group-
except 1st transition series. for eg Fe
having configuration $3d^6 4s^2$ show $+2$ &
 $+3$ Os but $+6$ & $+8$ are very rare and
unknown. While other members Ru, Os show
 $+8$ Os.

Moreover in transition elements, the higher
Oxidation state become more important as
the atomic number increases. So this
trend of higher Os of transition elements
res down the gp opposes the trend as
observed in representative elements.

Atomic Sizes →

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Size Variation noted for the first time by Lothar Meyer in 1870. He calculated

$$\text{atomic vol} = \frac{\text{at wt}}{\text{density}}$$

Elements that exist in diff allotropic forms have more than one atomic volumes.

As the e^- cloud has no limitation the size of an atom cannot be defined but by Lennard Jones σ parameter it can be possible to solve the problem. σ represents the distance of closest approach of the nuclei of two gaseous atoms of the same element. Let the atoms be spheres and $\sigma = \text{diameter}$ so $\frac{\sigma}{2}$ is the atomic radius. If we see of noble gases.

He	Ne	Ar	Kr	Xe
1.31	1.39	1.70	1.80	2.0

It means as we move in a gp from top to bottom size increases. ^{because n (Principal Q no) increases} e^- lie at greater distances from nucleus. But in period (row) as we move from left to right size decreases as the atomic number increases. Both the trends depend on the electronic configuration changes. Because in a row principal quantum no. (n) of valance e^- is same but the nuclear charge increases. So the valance e^- tend to be drawn closer to the nucleus and as a result atoms become smaller.

Ionic Radii :- For isoelectronic sequence, the ionic radius decreases as at no. increases. Because as the nuclear charge increases, the e^- cloud is bound to contract. In gp ionic size increases as the atomic no. increases. But as we move from Li to Cs, the ionic radii of members of the gp do not increase as rapidly after

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As the interval b/w K & Rb constitutes the first transition series. And then extra elements enter the periodic table, the \uparrow in nuclear charge tends to cause the atoms and their ions to contract. So the ions that follow any transition series are smaller than ^{if only} eight elements had separated them from lighter members of their family.

Q Why size decreases along transition series?
It is due to lanthanide elements. In the 14 elements after lanthanum 4f e⁻ are added to give electronic configuration of the type $5s^2 5p^6 4f^n 6s^2$. Lanthanides form +3 ions in which the two 6s electrons and one 4f electron is lost. The size of these ions become smaller as the atomic no. increases. This \downarrow in size is called lanthanide contraction.