

WAMPEEWO NTAKKE SS

CHEMISTRY DEPARTMENT

S.5 NOTES TERM 1

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TRENDS OF ATOMIC RADIUS IN THE PERIODIC TABLE.

Across the period.

The atomic radius decreases across the period from the alkali metals to the halogens. This is because there is an increase in the nucleus charge across the period and electrons are being added in the same quantum shell which screens each other poorly from increasing nucleus charge. Hence, **the effective nuclear charge outweighs the screening effect and therefore the** atom contracts in size.

E.g. Period 2 Li –Ne

Elements	Li	Be	B	C	N	O	F	Ne
Atomic Radius/nm	0.133	0.089	0.080	0.077	0.074	0.072	0.070	0.068

However, in period 4 and subsequent periods, the steady decrease in atomic radius with increase in nuclear charge is interrupted by transition elements. For a given period, the transition elements have about the same atomic radius.

Qn.Explain why the transition elements have about the same atomic radius.

This is because across the transition elements, the nucleus charge increases while the electrons are being added to the inner 3d sub shell, the screening effect slightly increases across the series. This increase in nucleus charge is roughly

balanced by increase in the screening effect; hence the nucleus attraction for the outer electrons remains constant.

N.B: In Summary

The transition *elements have about the same atomic radius. This is because the effect of increased nuclear charge* is roughly balanced by the greater screening effect produced by adding an extra electron to the shell.

DOWN THE GROUP

Down the group, there is an increase in atomic radius; this is because the number of outer electrons remains the same while the number of full inner electron shells increases. Atomic radius increases despite increase in the nuclear charge. Adding an electron shell outweighs the increase in nuclear charge. The screening effect of full inner electron shells increases and outweighs the increase in nuclear charge. The effective nuclear charge on the outer electrons decreases and hence the atoms expand.

N.B

Screening effect has its greatest influence (effect) within the alkali metals which contain only one electron in their outermost quantum shells. Therefore; the elements have the largest atomic radii.

ELEMENT	Li	Na	K	Rb	Cs
ATOMIC RADIUS(nm)	1.33	1.57	2.03	2.16	2.35

The halogens have a smaller atomic radius.

ELEMENT	F	Cl	Br	I
ATOMIC RADIUS(nm)	0.072	0.099	0.114	0.133

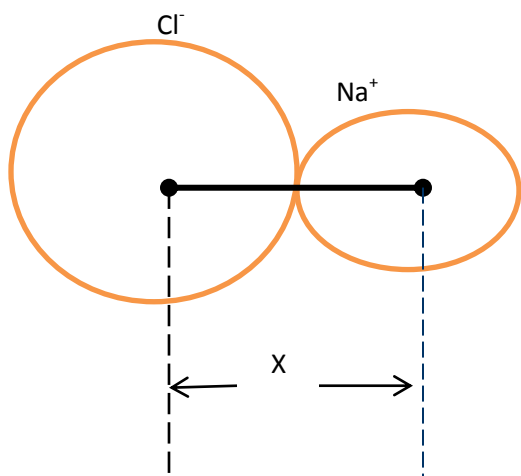
Atomic/ionic radius decreases as electrons are being removed from the atom or ion. This is because the nuclear charge remains constant while the screening effect decreases leading to increased effective nuclear charge, therefore the remaining fewer electrons are more strongly attracted by the nucleus and are drawn closer (hence the atom contracts)

IONIC RADIUS

This is half the intermolecular distance between two ions which are adjacent to each other in the lattice of an ionic compound.

OR

It is a mean distance between the centre of the nucleus of an ion and the region of its outermost shell occupied by electrons.



$$\text{Ionic radius} = x/2$$

Ionic radius is determined by the effective nuclear charge and the electronic configuration.

TRENDS ACROSS THE PERIOD.

The ionic radius decreases across the period; this is because there is an increase in the net positive nuclear charge which outweighs the screening effect as electrons are being removed. Hence, the remaining electrons of the ion experience greater nuclear charge attraction and ion contracts in size.

TRENDS DOWN THE GROUP.

The ionic radius increases down the group; this is because down the group, the effective nuclear charge attraction decreases due to the addition of electrons shells.

The increasing screening effect outweighs the increasing nuclear charge.

Hence, the ion expands leading to increased ionic radius

ALKALI METALS IONIC RADIUS (nm)		HALIDE IONIC RADIUS(nm)	
Li ⁺	0.68	F ⁻	1.36
Na ⁺	0.97	Cl ⁻	1.81
K ⁺	1.33	Br ⁻	1.95
Rb ⁺	1.47	I ⁻	2.16

The ionic radii for metallic ions are smaller than the atomic radii of their atoms; this is because the nuclear charge remains constant while the net positive charge increases on removal of electrons, resulting into increased effective nuclear charge which outweighs the screening effect. The remaining electrons are strongly attracted to the nucleus and drawn closer and hence ionic radius decreases.

The ionic radii for non-metallic ions are larger than the atomic radii of their atoms. Explain why?

QN: Why is sodium ion smaller than sodium atom?

This is because, during the formation of sodium ion, Na^+ , the outer most quantum shell is lost; the screening effect is reduced resulting into increased effective nuclear charge hence the ion contracts in size.

Revision questions

1. Explain why in every case, the ionic radius is smaller than the corresponding atomic radius?

2. The ions Na^+ , Mg^{2+} , and Al^{3+} have the same electronic configuration yet they have different ionic radii. Suggest reasons for this observation.

Explain why the ionic radii of Na, Mg and Al are in the order $\text{Na}^+ < \text{Mg}^{2+} < \text{Al}^{3+}$.

4(a) what is meant by the term atomic radius?

(b) Explain how atomic radii vary

(i) Across the short period of the periodic table.

(ii) Down the group in the period table.

(c) State and explain the two factors that affect the value of atomic radius.

5(a) the atomic and ionic radii of iron metal is in the order $\text{Fe} < \text{Fe}^{2+} < \text{Fe}^{3+}$. explain this observation.

(b) Across the transition elements series, the atomic radius remains almost constant. Explain this observation.

6(a) Explain what is meant by the term electronegativity?

(ii) State the factors that determine the value of electronegativity of an element.

(b) Explain how the following factors affect the value of electronegativity of an element.

(i) Nuclear charge.

(ii) Screening effect of the inner electrons.

(iii) Atomic radius.

(c) Explain the difference between electronegativity and electron affinity.

METALLIC CHARACTER (ELECTRO POSITIVITY)

Electro positivity is the relative ease by which atoms tend to give up (lose) electrons forming positive ions.

Weakly electro positive elements like group I and group ii metals are said to be strong electropositive elements and hence are strong reducing agents.

DIAGONAL RELATIONSHIP .

As electro positivity increases down the group and electronegativity increases across the period, some elements on a diagonal tend to show similarities, This is called the diagonal relationship.

Definition:

Diagonal relationship is the similarity between the chemistry of two elements and their compounds, which are in two adjacent groups and the two elements are diagonal to each other.

OR: It is a relationship that exists between two elements in adjacent groups whereby the two elements have the same chemical properties and are diagonal to each other.

TRENDS OF ELECTROPOSITIVITY IN THE PERIODIC TABLE.

Down Group.

Electro positivity increases down the group, Atomic radius increases and the nuclear charge attraction for the outer electrons decreases because the increase in the screening effect outweighs the increase in nuclear charge. The tendency for these electrons to be lost increases.

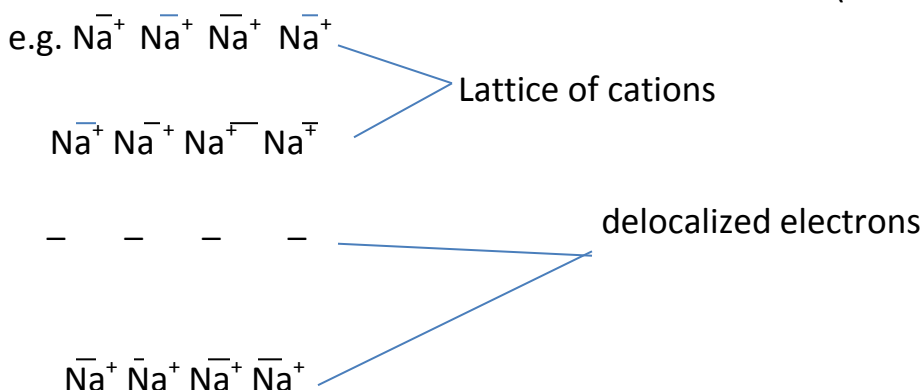
Across the Period. Electro positivity decreases across the period. The decrease in atomic radius with increase in the nuclear charge brings a stronger control of the valence electrons by the nucleus and the power to attract and hold additional electrons to form negative to form negative ions increases hence the tendency for these electrons to be lost decreases greatly.

VARIATION OF PROPERTIES IN PERIOD 3 ELEMENTS AND COMPOUNDS

STRUCTURE AND BONDING

(a) Na ,Mg ,Al

These are metals that adopt closed packed structures or body centered structures. They adopt metallic bonding with strong force of attraction between the cations and delocalized mobile electrons (electron cloud)



The strength of metallic bond depends on the number valence electrons contributed in the electron cloud. The more the number of electrons contributed, the stronger the metallic bond and hence the higher the melting point.

Group I elements, which contribute only one electron to the electron cloud, their metallic bond is weak and thus they are soft metals which low melting and boiling points.

NOTE

Among non-metals and covalent compounds, the melting and boiling points depend on the molecular size and the structure of the molecules. The bigger the molecules (the higher the molecular size), the stronger the intermolecular forces and hence the higher the melting and boiling points.

Silicon (Si)

This has giant atomic structure (i.e. it has a tetrahedral structure similar to *that* of diamond) where the atoms are held together by very strong covalent bonds which requires a lot of energy to hence it has a high melting and boiling points.

(C) P, S, Cl, and Ar

These are non-metals consisting of simple molecules. Chlorine and Argon consisting of discrete (individuals) molecules with very weak intermolecular forces hence they are gases at ordinary temperature.

However, atoms within these molecules are held together by covalent bonds.

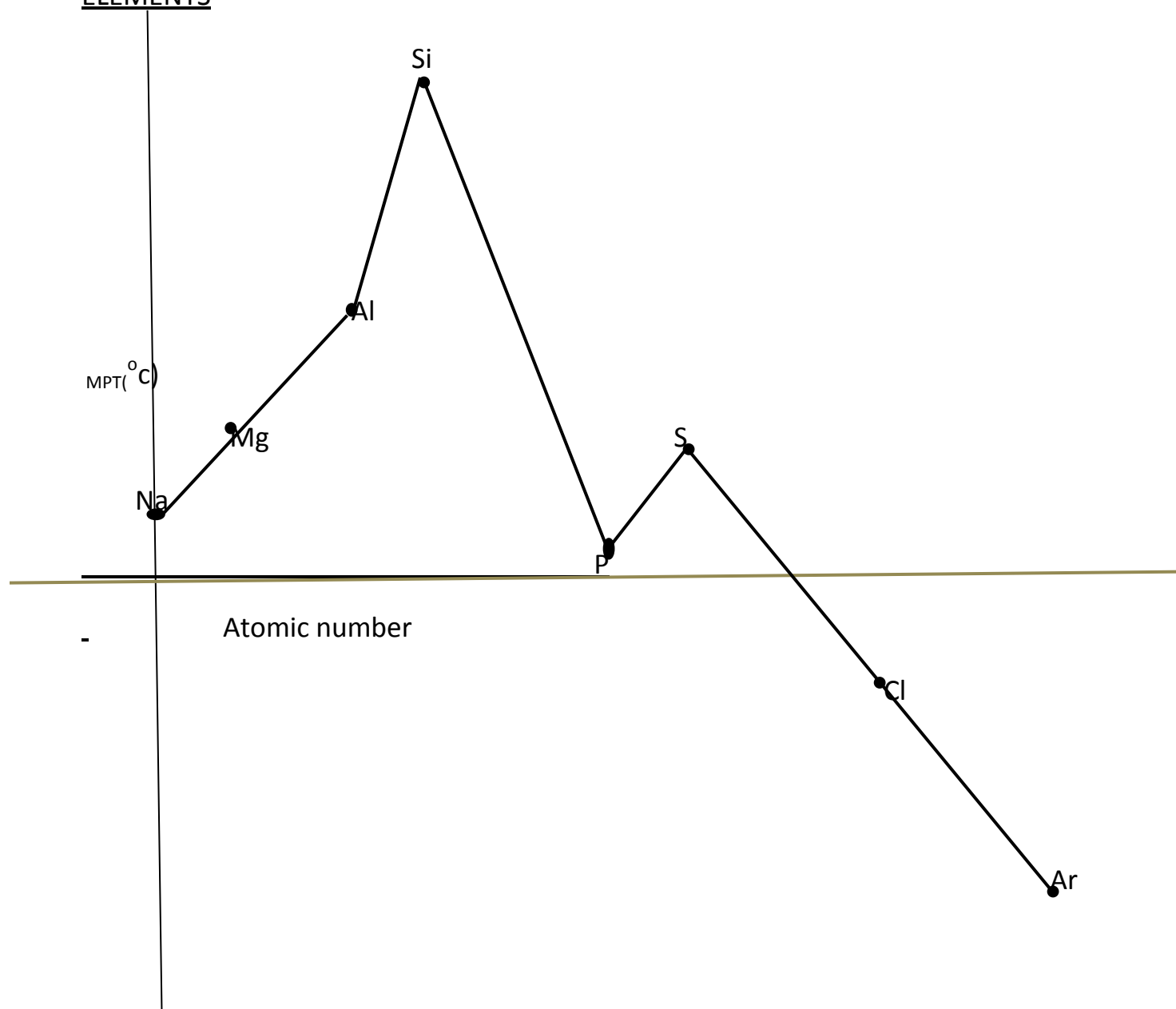
Phosphorus and Sulphur also form simple molecular structures but they are held together by covalent bonds.

Phosphorus and Sulphur also form simple molecular structures but they are bigger in size hence higher melting points than chlorine and Argon.

Melting Point

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic Number	11	12	13	14	15	16	17	18
Mpt($^{\circ}\text{C}$)	98	650	660	1410	44	114	-34	-189

A SKETCH GRAPH OF MELTING POINTS AGAINST ATOMIC NUMBER OF PERIOD 3 ELEMENTS



TREND

The melting point increases through metals and metalloids and decreases to very low value through non-metals.

-Explanation:

Melting point depends on structure. The sudden changes in melting points across the period is a result of a sudden change in the structure

Among the metals, Na, Mg and Al, melting point depends on the number of electrons in the electron charge cloud. The greater the number of electrons in the electron cloud, the stronger the metallic bond and hence the higher is the melting point.

Among the non- metals, Melting points, depend on the size and structure of the molecules, the bigger the molecular size, the stronger are the intermolecular forces and hence the higher the melting point.

From Na to Al, stronger metallic bond has to be broken. The strength of a metallic bond increases from Na to Al because of increase in the number of electrons used per atom to make the metallic bond.

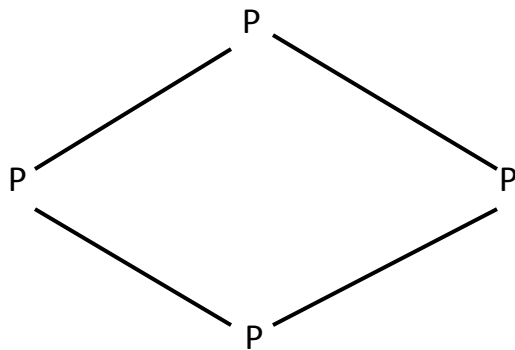
For Silicon, strong covalent bonds have to be broken against the giant atomic structure where each atom contributes four electrons to form covalent bonds

There is a sudden drop in melting point moving from Si to P due to the dramatic change of structure from metalloid to simple molecular structure

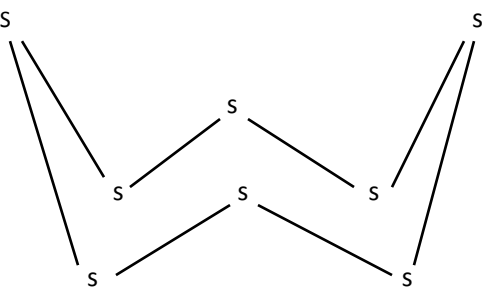
Both P and S exists as simple molecular structures. The molecules are bonded by weak van der Waals (inter molecular) forces hence a sharp decrease in melting point from Si to P.

Sulphur has a higher melting point than phosphorus because it forms a larger molecule, S₈ than phosphorus P₄. The intermolecular forces are greater because of the big size of the S₈ molecule.

P₄



S₈



.Chlorine and argon exist as discrete molecules with very weak intermolecular forces hence a very sharp decrease in the melting point from sulphur to chlorine.

Atoms of Argon cannot combine and so consists of separate atoms, the intermolecular forces are extremely very weak hence it is a gas with a very low melting point in period 3.

QN. The atomic number and melting points of some elements in period 3 of the periodic table are shown below:

ELEMENT	Na	Mg	Al	Si	P
ATOMIC NUMBER	11	12	13	14	15
MELTING POINT(°C)	98	650	660	1410	44

(a)(i) Plot a graph of melting point against atomic number

(ii) Explain the shape of the graph.

(b) State the factors that affect the melting point of

(i) metals.

(ii) Molecular substances.

(c) Explain the trend of melting points of elements in group I and group VII of the periodic table.

Qn. The melting points of elements Mg, Si and S are 650, 1410 and 114°C respectively.

(a) Explain the difference in the melting points of the elements.

(b) State the different types of bonding that exists in the hydrides of each of the elements.