



Dr. Bbosa Science

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The Periodic Table (PT)

The periodic table is a tabular arrangement such that the columns contain elements with electronically similar atoms.

The rows of the periodic table are called **periods** while the vertical columns are called **groups**. There are 7 periods and 8 groups.

The layout of the periodic table demonstrates recurring (periodic) chemical and physical properties. Elements are listed in order of increasing atomic number (i.e. the number of protons in the nucleus). Rows are arranged so that elements with similar properties fall into the same columns (group)

Some special groups

Group I (1A)

This consists of the elements, H, Li, Na, K, Rb, Cs, and Fr; except hydrogen, these elements are highly metallic in character and are known as **alkali metals**.

Group II (2A)

This consists of the elements, Be, Mg, Ca, Sr, Ba and Ra. All these elements are metals but their metallic character is less than that of group 1. These elements are known as **alkaline earth metals**.

Group VII (7B)

This consists of elements, F, Cl, Br, I and At. All these elements are nonmetals and are known as **halogens**.

Group VIII (0)

Consists of the elements, He, Ne, Ar, Kr, Xe and Rn. These elements are non-reactive and are known as the **noble elements, inert elements or zero elements** and are all gaseous elements at room temperature.

Summary

The extreme left end of the periodic table consists of metals with exception of hydrogen; the extreme right end consists of non-metals. The middle part of the periodic table consists of semi-metals or metalloids (fig. 2.1).

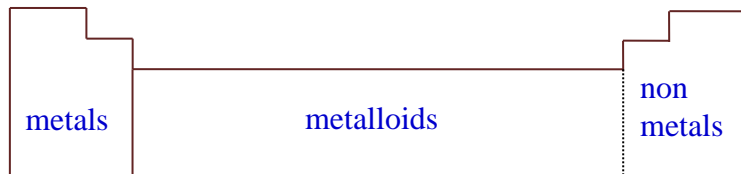


Fig. 1 Simplified PT to show the variation of properties

Therefore, the metallic characteristics of the elements decrease across the period, i.e., (left - right).

The periodic table commonly used by Uganda National Examinations Board is given in table 2.1 below.

Table 1 Periodic Table adopted by Uganda National Examinations Board

1	2											3	4	5	6	7	8
1.0 H 1	<div style="border: 1px solid red; padding: 5px; display: inline-block;"> 1.0 - Atomic mass H 1 – atomic number </div>															1.0 H 1	4.0 He 2
6.9 Li 3	9.0 Be 4											10.8 B 5	12.0 C 6	14.0 N 7	16.0 O 8	19.0 F 9	20.2 Ne 10
23.0 Na 11	24.3 Mg 12											27.0 Al 13	28.1 Si 14	31.0 P 15	32.1 S 16	35.4 Cl 17	40.0 Ar 18
39.1 K 19	40.1 Ca 20	45.0 Sc 21	47.9 Ti 22	50.9 V 23	52.0 Cr 24	54.9 Mn 25	55.8 Fe 26	58.9 Co 27	58.7 Ni 28	63.5 Cu 29	65.7 Zn 30	69.7 Ga 31	72.6 Ge 32	74.9 As 33	79.0 Se 34	79.9 Br 35	83.8 Kr 36
85.5 Rb 37	87.6 Sr 38	88.9 Y 39	91.2 Zr 40	92.9 Nb 41	95.9 Mo 42	98.9 Tc 43	101 Ru 44	103 Rh 45	106 Pd 46	108 Ag 47	112 Cd 48	115 In 49	119 Sn 50	122 Sb 51	128 Te 52	127 I 53	131 Xe 54
133 Cs 55	137 Ba 56	139 La 57	178 Hf 72	181 Ta 73	184 W 74	186 Re 75	190 Os 76	192 Ir 77	195 Pt 78	197 Au 79	201 Hg 80	207 Tl 81	207 Pb 82	209 Bi 83	209 Po 84	210 At 85	222 Rn 86
223 Fr 87	226 Ra 88	227 Ac 89															
			139 La 57	140 Ce 58	141 Pr 59	144 Nd 60	147 Pm 61	150 Sm 62	152 Eu 63	157 Gd 64	159 Tb 65	162 Dy 66	165 Ho 67	167 Er 68	169 Tm 69	173 Yb 70	175 Lu 71
			227 Ac 89	232 Th 90	231 Pa 91	238 U 92	237 Np 93	244 Pu 94	243 Am 95	247 Cm 96	247 Bk 97	251 Cf 98	254 Es 99	257 Fm 100	256 Md 101	254 No 102	260 Lw 103

According to the quantum mechanical theories of electron configuration within atoms, each row (period) in the table corresponds to the filling of quantum shell of electrons. There are progressively longer periods further down the table. And elements are grouped into: *s*, *p*, *d*- and *f*-blocks to reflect their electronic configurations.

Trends in physical properties of elements in the periodic table

The physical properties to be considered include:

1. Electronegativity/Electropositivity.
2. Metallic character.
3. Atomic sizes/atomic radii.
4. Ionic size/ionic radius.
5. Ionisation energy.
6. Electron affinity.
7. Melting points/boiling points.

Electronegativity

In a covalent bond between unlike atoms, the shared pair of electrons is not shared equally, because the atom with a greater attraction to electrons pulls the shared electrons towards itself. The measure of the power of an atom to attract bonding electrons is termed **electronegativity**. Or electronegativity of an element in a molecule is its relative tendency (or power or ability) to attract the shared/bonding pair of electrons towards it. Electronegativity of elements increases across a period as shown in table 2.2.

Table 2 Variation of electronegativity across the second period of the periodic table

Elements of 2nd period:	Li	Be	B	C	N	O	F
Electronegativity (arbitrary units):	1.0	1.5	2.0	2.5	3.0	3.5	4.0

NB: Fluorine, the most electronegative element, is given an arbitrary value of 4 and the electronegativity of the atoms of other elements are related to it.

Reason

Electronegativity increases across a particular period from alkali metal to halogen (table 2.2). Because there is a progressive increase in effective nuclear charge that exerts a contracting effect on the outer electron shells, (i.e., in the same period the additional electron resides in the same shell.)

The electronegativity **decreases down a particular group** (table 2.3). This is because the atom becomes progressively larger and the screening effect by electrons on inner shells decreases the attraction between the positive nucleus and the peripheral electrons.

Table 3 Variation of electronegativity down group 1 of the periodic table

Elements of group 1A:	Li	Na	K	Rb	Cs
Electronegativity (arbitrary units):	1.0	0.9	0.8	0.8	0.7

Trial 1

(b) Explain what is meant by the term electronegativity.

State the factors that determine the value of electronegativity of an element.

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

(i) atomic radius,

(ii) nuclear charge,

(iii) the screening effect of the inner electrons.

Electropositivity

This is the measure of the ease of an atom of an element to lose valence electrons. Electropositivity decreases across each period, due to an increase in the effective nuclear attraction on the valence electrons. Electropositivity increases down each group due to the decrease in effective nuclear attraction on the valence electrons as the atomic sizes and screening effect increase. Thus, group IA elements are the most electropositive elements.

Trial 2

Explain the variations in the electropositivity of the following elements

(i) C, Ge and Sn.

(ii) Mg, Al, P and Cl.

(8marks)

Metallic character

Since the metallic character depends on how easily an element loses its valence electrons, the metallic character **decreases across the period**, due to the increase in the effective nuclear charge on the valence electrons. Metallic character **increases down a group**. This is due to the decrease in the effective nuclear charge on the valence electrons.

Atomic radii

Atomic radius is the distance from the center of the nucleus to the furthest electron under the influence of a nuclear charge of an atom.

The radius of an atom is determined chiefly by two factors.

1. The attraction of positively charged nuclei for the valence electrons: Strong attractions result in smaller sizes/radii of the atoms as valence electrons are pulled closer to the nucleus.
2. The screening of the outer electrons from the nucleus by those from the inner shells: The greater the screening effect, the bigger the size/radius of the atom. Therefore, atomic radii decrease across a period (table 2.4) due to increase in effective nuclear charge and increase down the group (table 2.5) due to the increase in the screening effect which decreases the effective nuclear charge on the valence shell, e.g.,

Table 4 Variation of atomic radii across the second period of the periodic table

Elements of period 2A:	Li	Be	B	C	N	O	F
Atomic radii (nm)x10	1.23	0.89	0.80	0.77	0.70	0.66	0.64

Table 5 Variation of atomic radii down group 2 of the periodic table

Elements of group 2A:	Be	Mg	Ca	Sr	Ba
Atomic radii (nm) x 10:	0.89	1.36	1.74	1.91	1.98

Thus, screening has its greatest effect on the alkali metal atoms that contain only one electron in their outer most shells. Because of this, these atoms have the largest atomic radii in their respective periods.

Variation of atomic radii in transition elements

The atomic radii of elements of a particular series decrease gradually up to the midway element (from Sc to Mn) and then these values remain almost constant up to the elements of group 1B (Fe-Cu). The last element of each series i.e. Zn shows an increase in atomic radius.

Reason

For elements from Sc to Mn, the atomic radii decrease because of gradual increase in nuclear charge with increase in atomic number. From Fe to Cu, the atomic radii remain almost constant because electrons added to the *d*-orbital screen the 4*s* electron(s), and the attraction between the nucleus and 4*s* electrons decreases. The additional nuclear charge is counter balanced by the screening effect.

Towards the end of the series, there is an increase in electron-electron repulsion between the electrons added to 3*d* and 4*s* orbitals. This increase in repulsion becomes greater than the attraction between the nucleus and the 4*s* electrons, thus, because of the greater magnitude of electron-electron repulsion, the atomic size of Zn is greater than that of Cu.

Trial 3

- (a) Explain what is meant by atomic radius. (4marks)
- (b) Describe and discuss the change in radii in the following series;
- (i) Na, Mg, Al, Si, P, S and Cl (7 marks)
 - (ii) F, Cl, Br and I (4 marks)
 - (iii) Ti, V, Cr, Fe and Co (4 marks)

Trial 4

- (a) What is meant by the term 'atomic radius'? (4marks)
- (b) Explain how atomic radii vary;
- (i) Along a short period of the periodic table
 - (ii) Down the group in the periodic table (8marks)

Trial 5

The atomic numbers and atomic radii of some transition metals are given in the table 2.6 below

Table 6

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic numbers	21	22	23	24	25	26	27	28	29	30
Atomic radii (Å)	1.44	1.32	1.22	1.17	1.17	1.16	1.16	1.15	1.17	1.25

(i) Plot a graph of atomic radius versus atomic number

(4marks)

(ii) Explain the shape of your graph.

(7marks)

Ionic radii

The variations of the ionic radii across the different periods and down the groups are similar to those of the atomic radii as demonstrated by the examples in tables 2.7 and 2.8 respectively.

Table 7 Variation of ionic radii across the second period of the periodic table.

Ions	Li ⁺	Be ²⁺	B ³⁺	O ²⁻	F ⁻
Ionic radius(nm) x10	0.78	0.31	0.20	1.40	1.36

Table 8 Variation of ionic radii down group 1 of the periodic table.

Ions	Li ⁺	Na ⁺	K ⁺	Rb ⁺	Cs ⁺
Radius(nm) x10	0.78	0.095	0.133	0.148	0.165

NB: (a) Atomic radii are very much smaller than anionic radii but very much larger than cationic radii (table 2.9).

Table 9 Comparison of atomic and ionic radii of cation and anion.

Atoms	Ions
Na 0.157 nm	Na ⁺ 0.095 nm (cation)
Cl 0.099 nm	Cl ⁻ 0.181 nm (anion)

(a) The radii of iso-electronic positive ions, (i.e., cations with the same number of electrons, e.g. both Li⁺ and Be²⁺ have the same number of electrons), decrease with increasing positive charge. There is only a small change in ionic radii for iso-electronic negative ions with a change in negative charge e.g. Cl⁻ and S²⁻ in table 2.10.

Table 10 Comparison of ionic radii of iso-electronic ions.

Cations	Anions
Li^+ 0.060 nm	S^{2-} 0.184 nm
Be^{2+} 0.031 nm	Cl^- 0.181 nm

Trial 6

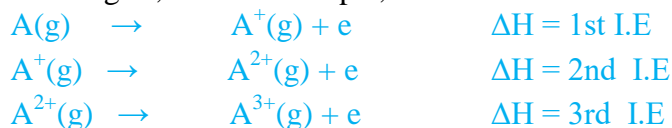
- (a) Explain how the sizes of ionic radii of group II ions affect the properties of their compounds compared to those formed by group I metals. (8marks)
- (b) Explain the following observations.
- (i) When moving from sodium (0.189 nm) to chlorine (0.099 nm) along period 3 of the periodic table, there is steady decrease in the size of covalent radii.
- (ii) Although K^+ and Cl^- ions both have 18 electrons, the ionic radius of K^+ is 0.132 nm while that of Cl^- is 0.181 nm.

Variation of ionic radii of transition metals

- (i) Since the transition metals show many oxidation states, the ionic radii of the ions of different oxidation states are different.
- (ii) Generally, the ionic radii of different cations of the same element in different oxidation states decrease with increase in oxidation state, e.g.
 $\text{Cr}^{2+} > \text{Cr}^{3+} > \text{Cr}^{4+} > \text{Cr}^{5+} > \text{Cr}^{6+}$
This is due to increase in effective nuclear charge.
- (iii) Ionic radii of cations of different elements in the same oxidation state generally decrease with the increase in nuclear charge or atomic number e.g.
 $\text{Fe}^{2+} > \text{Co}^{2+} > \text{Ni}^{2+} > \text{Cu}^{2+}$

Ionisation energy (I.E)

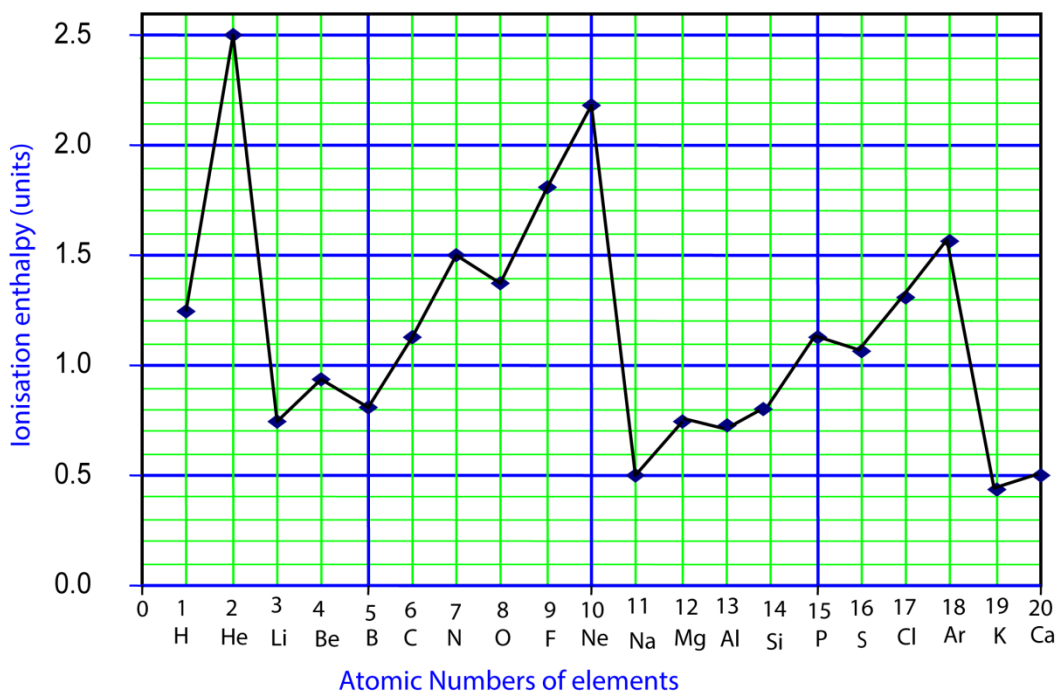
This is the energy required to remove one electron in the ground state completely from a gaseous atom or ion. The ionisation energy thus, may be classified as first, second, third ionisation energies, etc. for example,



Variation of the first ionisation energy across a period

The first ionisation energies of the first 20 elements of the PT plotted against their atomic numbers are shown in the graph below (fig. 2.2):

Fig. 1 A graph showing the first ionisation enthalpies of the first 20 elements



From lithium to neon there is a pronounced increase in ionizations energies with slight breaks occurring at positions occupied by boron and oxygen. An exactly similar trend occurs in the portion of the graph from sodium to argon.

Reasons for the two trends

1. The increase in the first ionizations energies across the period is due to increase in the effective nuclear charge on the valence electrons as the number of electrons on the outer shell increases.
2. Half filled (ns^1, np^3, nd^5) or completely filled ($ns^2np^6nd^{10}$) orbitals are comparatively more stable and hence more energy is needed to remove an electron from them. This explains why the first ionization energies of beryllium (full orbital, $1s^2 2s^2$) and nitrogen (half full orbital, $1s^2 2s^2 2p^3$) are higher than those of boron and oxygen respectively.
3. The sharp decrease in ionization energies from helium to lithium is explained by the fact that the outer electron in lithium is farther from the nucleus than in helium. Also the removal of an electron from helium leads to a less stable half full orbital while removal of an electron from lithium leads to formation of a stable full orbital $1s^2$.

Variation of the first ionization energies down a group

For elements in the same group, the ionization energies decrease down the group, due to increase in the screening effect or decrease in the effective nuclear charge because electrons are added to a new shell. For example, the first ionization energies of lithium, sodium and potassium are 520, 494, and 418 $kJ mol^{-1}$ respectively.

Trial 2.7

- (a) Study the graph (fig.2) of the first ionization energies against atomic numbers for the elements in the third short period of the periodic table given below and answer the questions that follow.

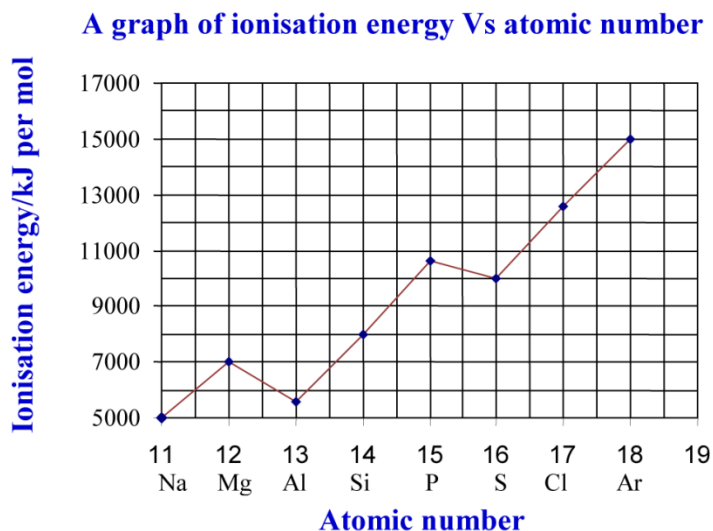


Fig. 2

- (i) What is meant by first ionisation energy? Illustrate your answer using the sodium atom. (3 marks)
- (ii) Explain the general trend in the first ionisation energies of the elements shown in the graph. (3 marks)
- (iii) The first ionisation energies of Al and S are lower than expected. Explain (5 marks)
- (b) Describe the trend in the first ionisation energies down group of the periodic table and explain your answer. (5 marks)
- (c) Show how the first ionisation energies of the elements in a group are related to their reactivities. Illustrate your answer using the reactions of group I elements with water. (4 marks)

Trial 2.8

- (a) Define the term ionisation energy.
- (b) The first ionisation energies of elements in period III of the periodic table and atomic numbers are given in the table 2.11 below

Table 11

	Na	Mg	Al	Si	P	S	Cl	Ar
Ionisation Energies (kJ mol^{-1})	495	740	580	790	1060	1000	1255	1520
Atomic Number	11	12	13	14	15	16	17	18

Plot a graph of ionizations energy against atomic number and explain the shape of the graph. (9 marks)

- (c) The melting points of magnesium, silicon and sulphur are 650°C , 1423°C and 120°C respectively. Explain the difference in melting points of the elements. (6 marks)
- (d) (i) Name the type of bonding that exists in the hydrides of Na, P and S. (1½ marks)
- (e) Write equations to show the reactions, if any, of the hydrides in (i) with water. (2½ marks)

Trial 9

The graph in figure 2.4 shows the variation of the first ionisation energies of the elements in the first row of the periodic table

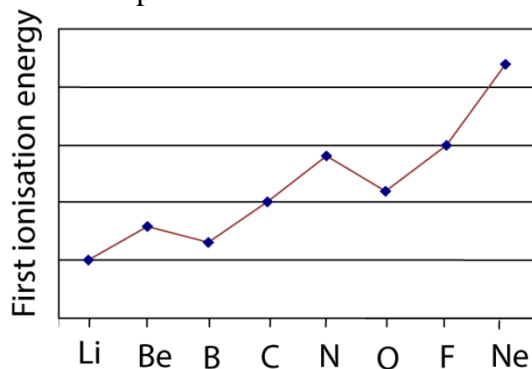


Fig.2 Elements

Explain the following observations:

- There is a general increase in first ionization energy from lithium to neon. (3marks)
- The first ionization energy of beryllium is higher than that of boron.
- The first ionization energy of oxygen is lower than that of nitrogen.

Variation of successive ionization energies of an atom

Very convincing evidence for the arrangement of electrons into definite shells is available from tabulated values of successive ionization energies of atoms; those for the potassium atom are shown in fig. 2.5. The ionization energies cover a very wide range of values and for the convenience of graphical plotting, their logarithmic values are used.

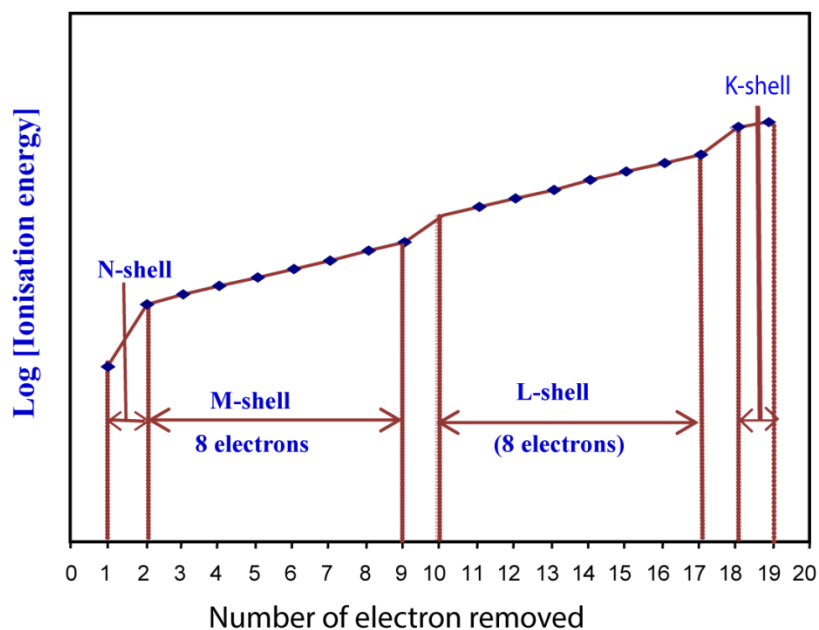


Fig. 3 Successive ionization energy of potassium atom

The clear-cut increases in ionization energy are observed when the 2nd, 10th and 18th electrons are involved, which indicates quite clearly that the nucleus of the potassium atom is surrounded by electrons grouped into a number of shells. Successive ionization of electrons becomes difficult, since every time an electron is removed; the atom carries one more unit of positive charge.

However, the large increases in ionization energy for removal of the second 2nd, 10th and 18th electrons suggest that the 2nd electron is closer to the nucleus than the 1st electron; similarly the 10th and 18th electrons are nearer to the nucleus than the 9th and 17th electrons respectively. Thus, electrons of a potassium atom can be grouped into shells as follows (table 2.12):

Table 12 Arrangement of electrons in the potassium atom

SHELL	K	L	M	N
NUCLEUS 19p, 20n	2	8	8	1

The electronic configuration of potassium is thus, written as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.

The properties of elements are closely linked to their ionization energies. The magnitude of the ionization energy is the measure of metallic character of the elements. The first ionization energies for metals are nearly all below 800 kJ mol^{-1} whereas those of non-metal are nearly all above 2000 kJ mol^{-1} .

Trial 10

Explain the following observations

- Although ionization energies generally increase across a period in the periodic table, the first ionization energy of boron is less than that of beryllium. (5marks)
- Both valence electrons of magnesium occupy the same energy level yet the second ionization energy of magnesium is greater than its ionization energy. (3marks)
- Calcium forms compounds containing Ca^{2+} ions, but none containing Ca^+ even though its first ionization energy is lower than the second ionization energy.

Trial 11

- (a) The 1st, 2nd, 3rd and 4th ionization energies of aluminum are 577, 1816, 2745, 11573 kJ mol^{-1} respectively.

Explain the trend in the ionization energies. (3marks)

- (b) Aluminium forms stable compounds in which aluminium has an oxidation state of +3 but not of +2 or +1. Explain this observation.

Trial 12

- Write the electronic configuration of strontium. (1mark)
- What is the principal oxidation state of strontium? (1mark)
- State whether the first ionisation energy of strontium is greater or less than that of barium. Explain your answers.

Trial 13

The diagram below shows successive ionisation energies for an element X, showing removal of all its electrons.

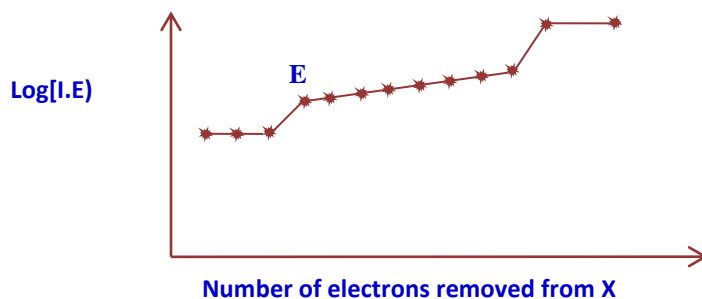


Fig. 2.6

- Giving reasons, state:
 - The group of element X.
 - The period of element X.
 - Identify element X. (4marks)
- Explain the sudden increase in the energy required to remove electron E. (2marks)
- Explain how the size of X will change as electrons are removed. (1marks)

Trial 14

(a) The first eight ionisation energies in kJ mol^{-1} of an element B are shown below (table 13).

Table 13

1 st I.E	2 nd I.E	3 rd I.E	4 th I.E	5 th I.E	6 th I.E	7 th I.E	8 th I.E
786	1580	3230	4360	16000	20000	23600	29100

- (i) Explain what is meant by the term first ionisation energy? (2 marks)
- (ii) State the factors that determine the value of the first ionisation energy and explain how they affect this value. (3 marks)
- (iii) To which group in the periodic table does element B belong. Explain your answer. (1 mark)

Trial 15

Table 2.14 below shows the ionisation energies (in kJ mole^{-1}) of five elements A, B, C, D and E.

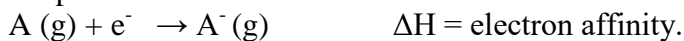
Table 14

Elements	1 st I.E	2 nd I.E	3 rd I.E	4 th I.E
A	500	4600	6900	9500
B	740	1500	7700	10500
C	630	1600	3000	4800
D	900	1800	14800	21000
E	580	1800	2700	11600

- (a) Which one of these elements is most likely to form an ion with a charge of +1? Give a reason for your answer.
- (b) State:
- Two elements which belong to the same group in the periodic table.
 - The group to which the elements you have stated in (b)(i) belong.
- (c) (i) Write the formula of the chloride of element E.
- (ii) Write an equation for the reaction between the chloride of element E and water.

Electron affinity

This is the energy involved in adding an electron to a *gaseous* atom. It is the energy change for the process.



Electron affinities of elements **increase across a period** due to increase in the effective nuclear charge (or electronegativity) and **decrease down a group** due to the decrease in the effective nuclear charge on the valence shell. Halogens have the highest values of electron affinities. However, the electron affinity of fluorine is lower than that of chlorine. This is due to the fact that the size of fluorine atoms is small, thus the incoming electron is repelled by the nine electrons that are close to each other in the valence shell. The successive electron affinities become more and more endothermic because the incoming electron is repelled by the electrons already in the atom; therefore, additional energy is required to add this electron.

Trial 16

- (a) Explain the difference between electronegativity and electron affinity.
 (b) The first electron affinities of group VII elements are given in table below.

Element	F	Cl	Br	I
1 st electron affinity / kJ mol^{-1}	-328	-349	-325	-295

- (i) Explain what is meant by the term electron affinity.
 (ii) Explain the trend in the electron affinities.
 (iii) The first and second electron affinity values for the oxygen atom are -121 kJ mol^{-1} and $+744 \text{ kJ mol}^{-1}$ respectively.
 Explain why the second electron affinity is endothermic.

Melting points of elements

The melting point of a substance or element is the measure of energy required to break down the regular arrangement of atoms, ions or molecules in the substance (solid state). Therefore, the melting point indicates the strength of the forces knitting atoms, ions or molecules together in the crystal. The strength of those forces in elements varies according to the type of crystals formed and due to this; the melting points of elements in a period do not change uniformly, (e.g table 2.15 below).

Table 2.15 The melting points of elements of the third period of the periodic table

Elements	Na	Mg	Al	Si	P	S	Cl
Configuration	2.8.1	2.8.2	2.8.3	2.8.4	2.8.5	2.8.6	2.8.7
M.pt. $^{\circ}\text{C}$	98	650	660	1423	44	120	-161
	Large increase		Large increase			Large decrease.	

From sodium to aluminium, the melting point increases due to increase in the strength of metallic bonds. The strength of metallic bonds increases as the number of electrons contributed to the formation of the metallic bond increases. Sodium contains one loosely bound electron in the valence shell, and this electron is readily contributed to the formation of the metallic bond. Magnesium contains two electrons in the valence shell and both electrons are contributed to the formation of metallic bond. The metallic bond in magnesium is thus stronger than that in sodium, which explains the large difference in the melting points.

However, the increase in melting point from Mg to Al is not as sharp as expected, probably because aluminium atoms use only two electrons of three valence electrons in the formation of the metallic bond.

Silicon has the highest melting point because it uses its four valence electrons to form an infinite three-dimensional assembly of atoms linked by a strong single covalent bond; thus, high temperatures are required to break these strong bonds.

In the structure, each silicon atom is tetrahedrally bonded to four silicon atoms with a crystal unit of five silicon atoms (ABCDE) in fig. 2.7 (the structure similar to that of diamond). Such a structure is described as a three-dimensional giant structure. As the melting of silicon involves breaking of the strong covalent bonds, the melting point of silicon is thus very high.

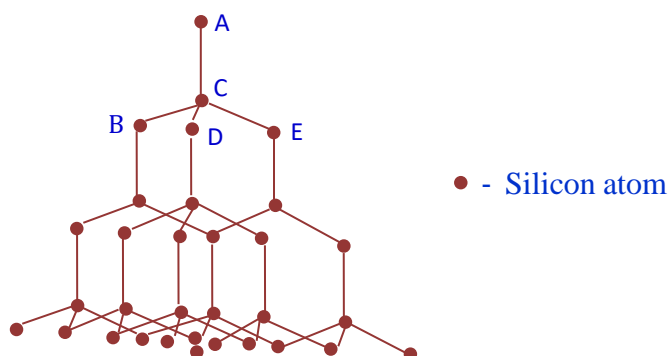


Fig. 7 Structure of silicon

The melting points of P, S, Cl are dependent on the sizes/masses of the molecules formed. The melting points of phosphorus and sulphur are relatively higher than expected because they exist as P_4 and S_8 , rather than simple individual atoms; thus, the intermolecular forces are stronger. The boiling point of chlorine is very low because it exists as discrete diatomic molecules, Cl_2 held together by weak van der Waals forces.

Trial 2.17

The atomic numbers and the melting points of some elements in period 3 of the periodic table (table 2.15) are shown below.

Table 15

Element	Na	Mg	Al	Si	P
Atomic number	11	12	13	14	15
Melting point/ $^{\circ}C$	98	650	660	1410	44

- (a) (i) Plot a graph of melting points against atomic numbers. (3 marks)
 (ii) Explain the shape of the graph in (i). (6marks)

- (b) Describe and explain how the oxides of magnesium, aluminium and silicon react with
- (i) sodium hydroxide.
 - (ii) hydrochloric acid. (9marks)
- (c) State the type of bonding in the oxides of sodium and phosphorus. (2marks)

Trial 18

- (a) The melting points of magnesium, silicon and sulphur are 650°C , 1423°C and 120°C respectively. Explain the difference in melting points of the elements. (6 marks)
- (b) Name the type of bonding that exists in the hydrides of Na, P and S. (1½ marks)
- (c) Write equations to show the reactions, if any, of the hydrides in (i) with water. (2½ marks)

Suggested answers to the trials

Trial 2.1

- (a) Electronegativity is the relative tendency of an atom to attract valence electrons.
- (b) Nuclear charge, atomic radius and screening effect.
- (c) (i) Atomic radius: Increase in atomic radii increases the distance between the bonding electrons and the nucleus of an atom. This in turn leads to a decrease in the nuclear attraction for the bonding electrons hence leading to the decrease in the electronegative value and vice versa.
- (ii) Nuclear charge: The smaller the nuclear charge experienced by the bonding electrons, the weaker they are attracted to the nucleus and the lower the electronegativity. The higher the nuclear charge experienced by the bonding electron the more strongly they are attracted to the atom and the higher the electronegativity.
- (iii) Screening effect: The higher the screening effect, the lower the nuclear attraction experienced by the outer electrons, and therefore the lower the electronegativity. The lower the screening effect, the higher the nuclear attraction experienced by the outer electron and the higher the electronegativity.

Trial 2.2

- (i) Electropositivity increases from C to Ge to Sn (down the group) because down the group atomic radii and number of electron shells increase. Screening effect outweighs increase in the nuclear attraction and the tendency to lose electron increase.
- (ii) Electropositivity decreases from Mg, to Al, P, to Cl (across the period) because from Mg to Cl additional proton is added to the nucleus and an electron to the same outer shell. Both nuclear charge and screening effect increase but increase in the nuclear charge, outweigh increase in screening effect.

Trial 3

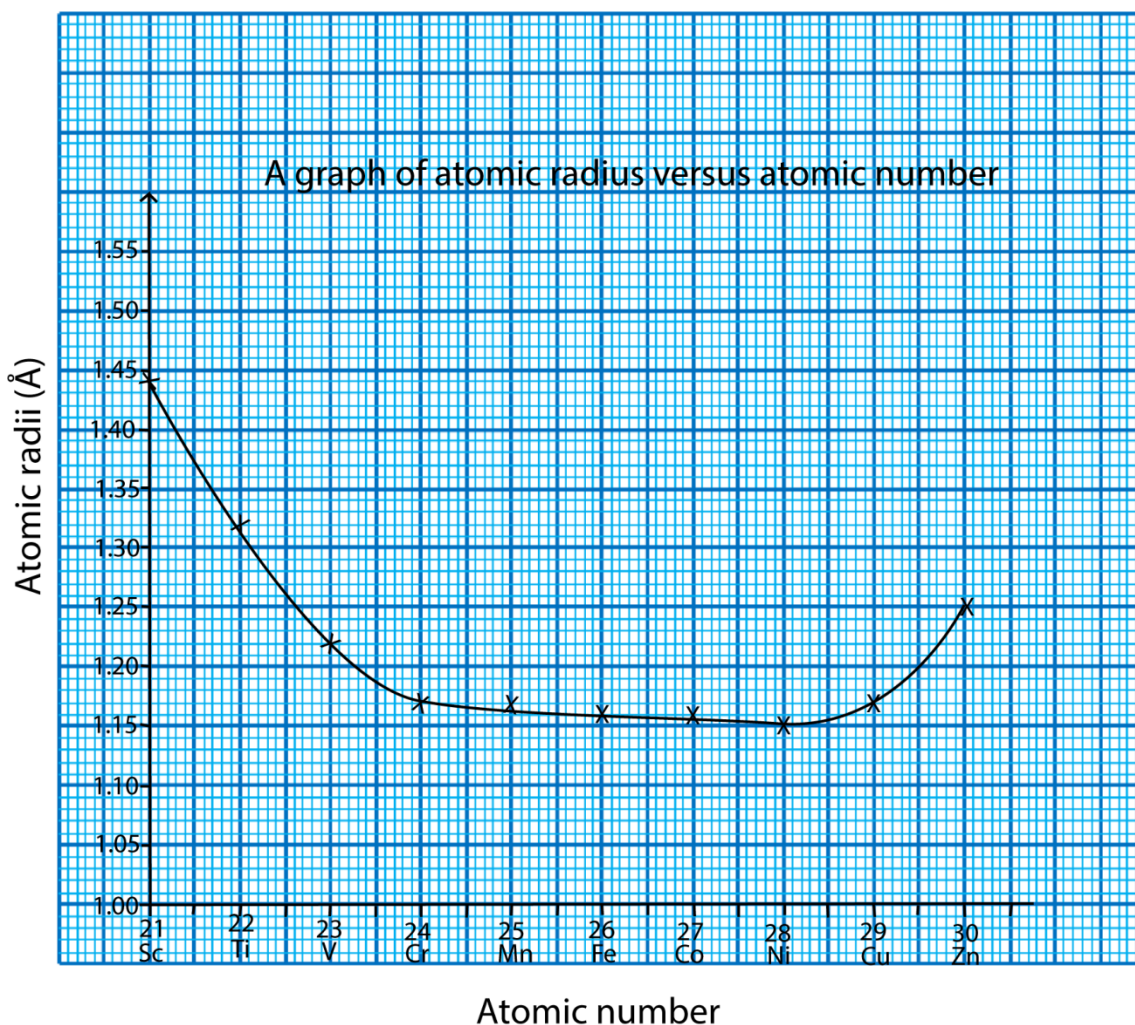
- (a) Atomic radius is the distance between the nucleus and outermost electron under the nuclear influence.
- (b) (i) Atomic radii decrease from Na>Mg>Al>Si>P>S>Cl (across the period) from Na – Cl additional proton is added to the nucleus and an electron to the same outer shell. Both nuclear charge and screening effect increase but increase in the nuclear charge, outweigh increase in screening effect.
- (ii) Atomic radii increase from F to Cl to Br to I(down the group) because from F – Cl there is increase in the number of electron shells and increase in screening effect outweigh increase in nuclear attraction the valence electrons.
- (iii) From Ti – Cr atomic radii increase to increase in the effective nuclear charge. From Fe –Co atomic radii remain unchanged because additional electron is added to inner 3d – shell; increase in screening effect is balanced by increase in nuclear charge.

Trial 4

- (a) Atomic radius is the distance between the nucleus and outermost electron under the nuclear influence.
- (b) (i) Across the period atomic radii increase due to increase in effective nuclear charge
(ii) down a group atomic radii increase due to increase in the number of electron shells and screening effect.

Trial 5

(i)



The atomic radii of elements decrease gradually from Sc to Mn and then remain almost constant from Fe-Cu and increase to Zn.

Reason

From Sc to Mn, the atomic radii decrease because of gradual increase in nuclear charge with increase in atomic number.

From Fe to Cu, the atomic radii remain almost constant because electrons added to the *d*-orbital screen the 4*s* electron(s), and the attraction between the nucleus and 4*s* electrons decreases. The

additional nuclear charge is counter balanced by the screening effect.

From Cu - Zn there is an increase in electron-electron repulsion between the electrons added to 3d and 4s orbitals. This increase in repulsion becomes greater than the attraction between the nucleus and the 4s electrons, thus, because of the greater magnitude of electron-electron repulsion, the atomic size of Zn is greater than that of Cu.

Trial 6

- (a) The ions of group II metals are smaller than those of group I metals, consequently they have higher charge densities and polarising power leading to
- Group II compound being less ionic some dissolving in organic solvents.
 - Their enthalpies of hydration being higher than those of group I.
 - Their first ionisation energies being considerably greater than those of group 1 elements.
 - Group II carbonates being unstable at high temperatures.
 - Group II metals forming complexes more readily than group I elements.
 - Group II compound generally having higher lattice energies.
- (b) (i) From Na to Cl atomic radii decrease because there is progressive increase in effective nuclear charge.
- (ii) Loss of an electron from potassium makes the remaining electrons to be strongly attracted to the nucleus reducing the size of the ion whereas a gain of an extra electron by the chlorine atom causes repulsion and increase in the size of the anion.

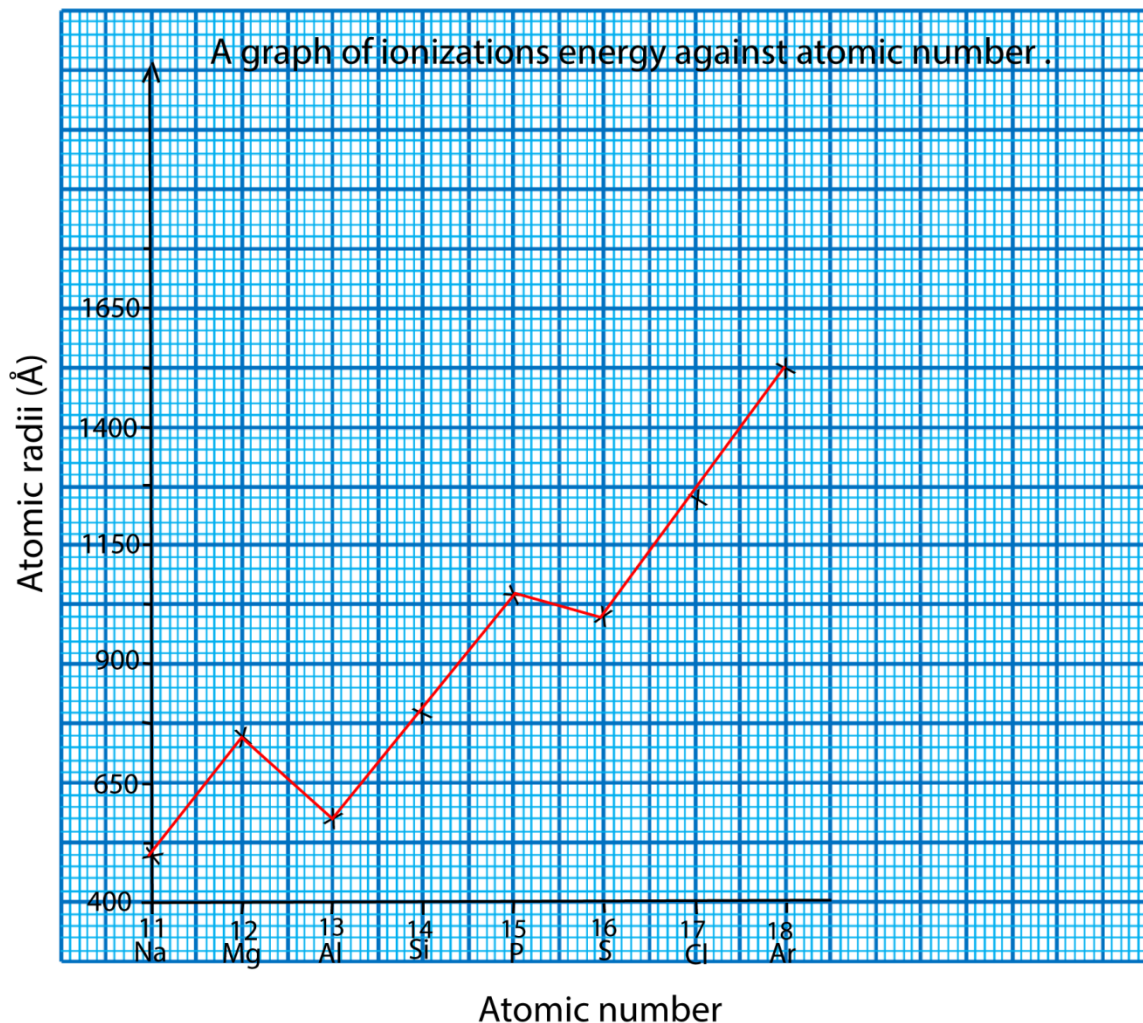
Trial 7

- (a) First ionization energy is enthalpy change when 1 mole of gaseous atoms is converted into 1 mole of gaseous ions with a single positive charge.
- $$\text{Na(g)} \rightarrow \text{Na}^+(\text{g}) + \text{e}$$
- (b) There is a general increase in ionizations energy from Na to Ar. Because electrons are being added to the same principal quantum shell. Electrons in the same principal quantum shell screen each other poorly from the nuclear attraction, thus increase in the nuclear charge outweighs the increase in the screening effect. Increase in effective nuclear charge increases the amount of energy required to remove the outer electrons.
- (c) First ionization energies of Al and S are less than expected because their first electrons are removed from unstable electron shells which are neither half full of full orbital.
- (d) First ionization energy decrease down the group because down the group both electron shells and screening effect increase reducing effective nuclear charge on valence electrons.
- (e) The lower the ionization energy of metals, the more reactive they become. For example reactivity of group 1 elements increases down the group. Li reacts slowly with water while K reacts explosively with cold water to form hydroxide.

Trial 8

(a) Ionization energy is energy change required to form gaseous ions from gaseous atoms or ions.

(b)



(c)

- There is a general increase in ionizations energy from Na to Ar. Because electrons are being added to the same principal quantum shell. Electrons in the same principal quantum shell screen each other poorly from the nuclear attraction, thus increase in the nuclear charge outweighs the increase in the screening effect. Increase in effective nuclear charge increases the amount of energy required to remove the outer electrons.
- First ionization energy of Al is lower than that of Mg because an electron from magnesium atom is removed from a stable full 3s-orbital whereas that for Al atom an electron is removed from unstable less than half full 3p-orbital.
- First ionization energy of S is lower than that of P because an electron from P is removed from a stable half full 3s-orbital whereas that for S atom an electron is removed from unstable more than half full 3p-orbital but less than full orbital.

Trial 9

(i) There is a general increase in ionisation energy from lithium to neon. Because electrons are being added to the same principal quantum shell. Electrons in the same principal quantum shell screen each other poorly from the nuclear attraction, thus increase in the nuclear charge outweighs the increase in the screening effect. Increase in effective nuclear charge increases the amount of energy required to remove the outer electrons.

(ii) Be – $1s^2 2s^2$
Li – $1s^2 2s^1$

The ionisation energy of beryllium is higher because its outermost energy level ($2s^2$) is completely filled and thermodynamically more stable. Therefore, to remove an electron from a stable sub-level requires more energy than that required to remove the $2s^1$ electron of lithium.

(iii) O – $1s^2 2s^2 3p^4$
N – $1s^2 2s^2 3p^3$

The first ionisation energy of nitrogen is higher than that of oxygen because its outer electron configuration is half filled and relatively more stable than that of oxygen that is more than half full but not full. Thus more energy is required to remove an electron from the nitrogen atom.

Trial 10

(a) There is a general increase in ionizations energy across the period. Because electrons are being added to the same principal quantum shell. Electrons in the same principal quantum shell screen each other poorly from the nuclear attraction, thus increase in the nuclear charge outweighs the increase in the screening effect. Increase in effective nuclear charge increases the amount of energy required to remove the outer electrons.

B – $1s^2 2s^2 2p^1$

The ionizations energy of beryllium ($1s^2 2s^2$) is higher because its outermost energy level ($2s^2$) is completely filled and thermodynamically more stable. Therefore, to remove an electron from a stable sub-level requires more energy than that required to remove the $2p^1$ electron of boron which is thermodynamically unstable

(b) The second ionization of Mg is higher than the first ionization energy because after removal of the first electron the remaining electron experience higher nuclear charge.

(c) Ca^{2+} is stable due to stable full electron configuration whereas Ca^+ has unstable electron configuration

Trial 11

(a) Generally ionization energies increase from 1st to 4th because removal of electrons from an atom or ion the remaining electrons experience higher nuclear attraction. But the fourth ionization energy is higher than expected, because the 4th electron is removed from inner full stable orbital.

(b) Al^{3+} forms stable compound due to stable full electron configuration; Al^{2+} and Al^{1+} have unstable configuration

Trial 12

- (a) Electron configuration of Sr = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2$
(b) +2
(c) Sr has higher first ionization energy than Ba because Sr has higher electronegativity than Ba

Trial 13

- (a) (i) Group 2
(ii) Period 1
(iii) Mg
- (b) E is removed from inner full stable orbital
- (c) The size of X decreases due to increase in effective nuclear charge.

Trial 14

- (a) First ionization energy is the energy required to change 1 mole of gaseous atoms into 1 mole of gaseous ions with single positive charge.
- (b) - atomic radii: the bigger the atomic radius the lower the ionization energy because the outer electrons experience little nuclear charge
- Electronegativity: ionization energy increases with electronegativity due to increase in effective nuclear charge on valence electrons
 - Screening effect: increase in screening effect reduces first ionization energy due to decrease in effective nuclear charge on the valence electrons.
- (b) Group 4: there is a big increase from the 4th to 5th ionization energy indicating that the 5th electron is being removed from inner electron shell

Trial 2.15

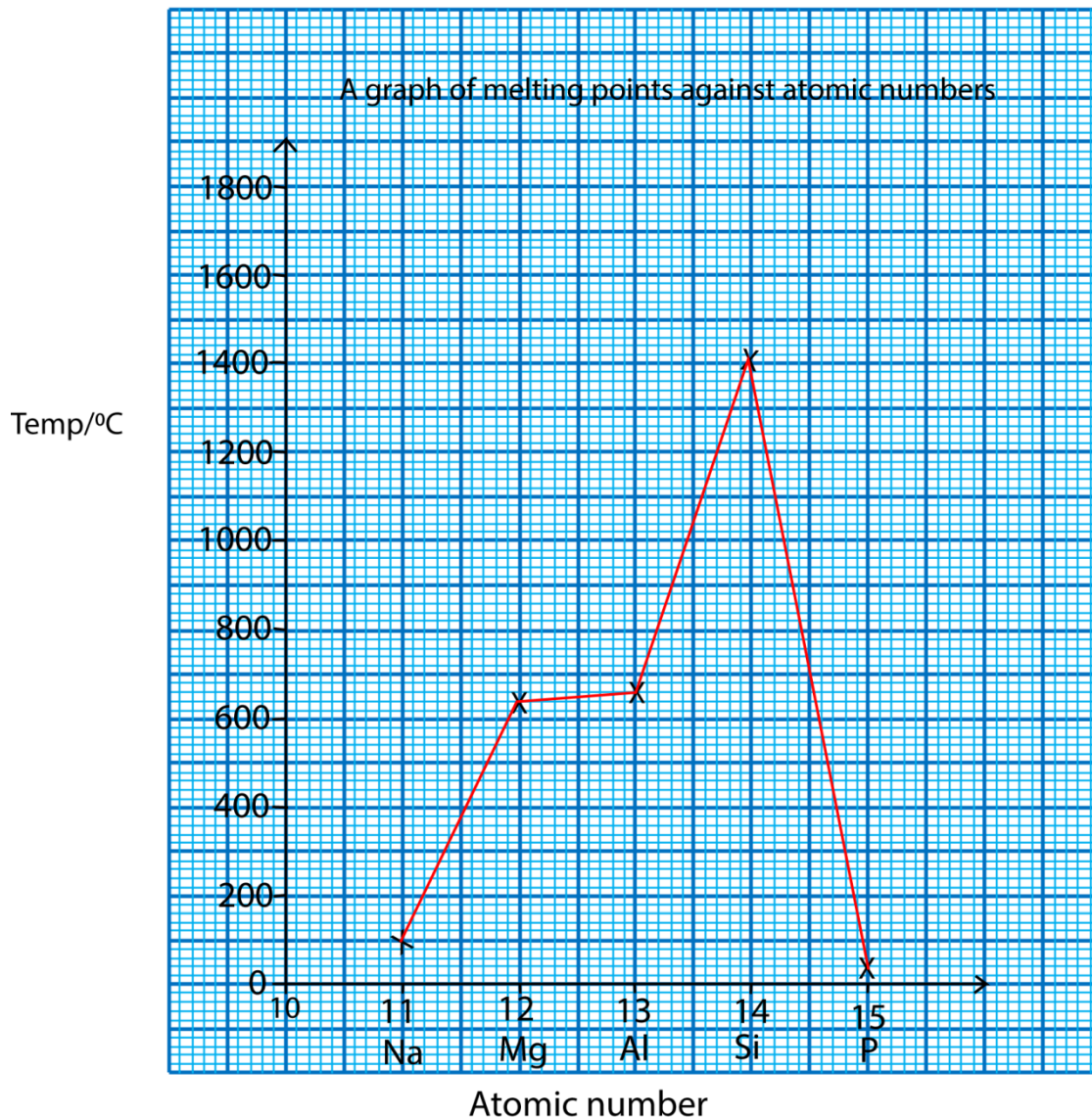
- (a) A shows a big increase in ionization energy when the 2nd electron is removed
- (b) (i) B and D (ii) from group 2
- (c) (i) ECl_3
(ii) $ECl_3(s) + 3H_2O(l) \rightarrow E(OH)_3(s) + 3HCl(aq)$

Trial 16

- (a) Electronegativity is the relative tendency of an atom to attract bonding electrons whereas electron affinity is energy change when 1 mole of electrons is added to gaseous atom or ion to form 1 mole gaseous anions
- (b) Electron affinities generally decrease down the group due to decrease in electronegativity. Fluorine has an abnormally low electron affinity because it is a small atom with high electron density that repels the incoming electron
- (c) The first ionization energy of oxygen is positive because oxygen has high electron affinity. The second electron affinity of oxygen is endothermic, because the first electron added repels the second electron.

Trial 17

(a)



- Melting point increases from Na – Al due to increase in the strength on metallic bonds as the number of electrons contributed to form metallic bonds increase, Na contributes 1 electron, Mg two and Al three electrons per atom in the formation of metallic bonds.
- The melting point of silicon has very high melting point because each silicon atom is bonded to 4 silicon atom by strong covalent bonds to form 3-dimensional structure.
- Phosphorus has low melting point because its molecules are bonded by weak van der Waal forces.

(b) (i) Magnesium does not react with sodium hydroxide

Al and Si react with hot concentrated sodium hydroxide to form aluminate and silicate respectively



(ii) Mg and Al react with hydrochloric acid to form chloride, silicon does not react



(c) Na_2O – ionic

P_2O_5 – covalent

Trial 18

(a)

- Melting point Mg is high due to strong metallic bonds
- The melting point of silicon has very high melting point because each silicon atom is bonded to 4 silicon atom by strong covalent bonds to form 3-dimensional structure.
- Sulphur has low melting point because its molecules are bonded by weak van der Waal forces. However, the melting point is relatively high because sulphur exists as S_8 molecules rather than simple sulphur atoms.

(b) NaH - ionic

PH_3 – covalent

H_2S – covalent

(c) $\text{NaH (s)} + \text{H}_2\text{O(l)} \rightarrow \text{NaOH (aq)} + \text{H}_2\text{(g)}$

PH_3 – does not react with water

