1 Periodic Table, Periodic Properties and Variations of Properties

SYLLABUS

Periodic properties and variations of properties — Physical and Chemical

(i) Periodic properties and their variations in groups and periods.

Definitions of following periodic properties and trends in these properties in groups and periods should be studied;

- *atomic size*
- *metallic character*
- *non-metallic character*
- *ionisation potential*
- *electron affinity*
- *electronegativity*

(ii) Periodicity on the basis of atomic number for elements.

Relation between atomic number for light elements (proton number) and atomic mass for light elements; the modern periodic table up to period 3 (students to be exposed to the complete modern periodic table but no questions will be asked on elements beyond period 3 – Argon); periodicity and other related properties to be described in terms of shells (not orbitals); special reference to the alkali metals and halogen groups.

Note : According to the recommendation of International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18 replacing the older notation of groups IA ...... VIIA, VIII, IB ... . VIIB and 0. However, for the examination both notations will be accepted.

Old notation	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
New notation	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

1.1 INTRODUCTION

Elements are pure substances made up of one type of atoms. They are the basic units of all types of matter. In order to study elements in an organised manner, they need to be classified.

You have studied in class IX contribution of Dobereiner, Newland and Mendeleev to the classification of elements.

Dobereiner, grouped the elements in three (*triads*).

Newland, observed that when elements are arranged in increasing order of their atomic mass, *every eighth element beginning from any element resembles the first* element in its physical and chemical properties.

Dmitri Mendeleev, a Russian chemist gave the first periodic table of elements base on his law which states that "*the properties of elements are the periodic functions of their atomic masses*". This arrangement enabled Mendeleev to place 63 elements known that time in vertical columns (***groups***), and in horizontal rows (***periods***). He predicted the existence of those elements which were yet to be discovered. *But this method could not explain the positions of certain elements, the rare earth metals and the isotopes.* These defects were removed when **Henry Moseley** put forward the **modern periodic table**. He stated that **"the physical and chemical properties of elements are the periodic functions of their atomic number".**

Later on Niel Bohr gave the extended form of the table known as *long form of the modern periodic table* given on page 2.

A tabular arrangement of elements in groups (vertical columns) and periods (horizontal rows) highlighting the regular trends in properties of elements is called a PERIODIC TABLE.

1.2 SALIENT FEATURES OF THE MODERN PERIODIC TABLE (SEE PAGE 2)

GROUPS

The modern periodic table has **eighteen vertical columns**. They are known as groups, arranged from left to right in the order.

TABLE 1.1 : THE LONG FORM OF PERIODIC TABLE OF ELEMENTS

GROUPS

Representative Elements (s - block) | Transition Elements (d - block) | Representative Elements (p - block) | Noble gases (p - block)

Atomic number — ^{6}C — Symbol

Solid | Liquid | Gas

Active metals (Alkali metals & Alkaline earth metals)
Heavy metals
Metalloids
Non-metals
Non-metals (Halogens)
Non-metals (Inert gas)

Periods	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 Zero
1	1 H Hydrogen																	2 He Helium
2	3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
3	11 Na Sodium	12 Mg Magnesium											13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulphur	17 Cl Chlorine	18 Ar Argon
4	19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
5	37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Tc Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
6	55 Cs Caesium	56 Ba Barium	57-71 Lanthanide Series	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Re Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
7	87 Fr Francium	88 Ra Radium	89-103 Actinide Series	104 Rf Rutherfordium	105 Db Dubnium	106 Sg Seaborguim	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Damstacium	111 Rg Roentgentium	112 Cn Copernicium	113 Uut Ununtrium	114 Fl Flerovium	115 Uup Ununpentium	116 Lv Livermorium	117 Uus Ununseptium	118 UuO Un-Un Octium

INNER TRANSITION ELEMENTS f-block															
Lanthanide Series	57 La Lanthanum	58 Ce Cerium	59 Pr Praseodymium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
Actinide Series	89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

Elements with atomic number 110 and above have been reported but not yet fully authenticated and named.
As of 2005, the table contains 116 chemical elements whose discoveries have been confirmed;
94 are found naturally on earth and the rest are synthetic elements.

Group 1 : These elements are known as **Alkali metals** as they form strong alkalis with water.

Group 2 : Alkaline earth metals – They form weaker alkalis as compared to group1 elements.

Groups 3, 4, 5, 6, 7, 8, 9, 10, 11 and 12 are known as the **transition elements**. They have their ***two*** *outermost shells incomplete*.

Group 13 : Boron family – Boron is the first member of the group.

Group 14 : Carbon family.

Group 15 : Nitrogen family.

Group 16 : Oxygen family also known as chalcogens meaning ore forming.

Group 17 : These elements form salts and so known as **Halogen** (meaning – salt former).

Group 18 : (Zero group) – Elements of this group are called the ***noble gases*** *or* ***inert gases***.

These elements have their outermost orbit complete. Due to stable electronic configuration they hardly react with other elements.

Note : The elements of groups 1, 2, 13, 14, 15, 16 and 17 are known as the **main group elements** or **representative elements** or **normal elements.**

The outermost shell of all the elements of these groups are incomplete.

PERIODS

There are **seven horizontal rows** in the modern periodic table. They are known as **periods** (see table given below).

The number of shells present in an atom determines its period. For example :

Elements of period one have one shell, elements of period two have two shells, and that of period three have three shells and so on.

Period	Type of period	Number of elements	Atomic no. of elements	No. of shell (s)	Elements in Group								
					1	2	3 - 12	13	14	15	16	17	18
1	Shortest period	2	1 – 2	1	**H** 1								**He** 2
2	Short period	8	3 – 10	2	**Li** 3	**Be** 4		**B** 5	**C** 6	**N** 7	**O** 8	**F** 9	**Ne** 10
3	Short period	8	11 –18	3	**Na** 11	**Mg** 12		**Al** 13	**Si** 14	**P** 15	**S** 16	**Cl** 17	**Ar** 18
4	Long period	18	19 – 36	4	**K** 19								**Kr** 36
5	Long period	18	37 – 54	5	**Rb** 37								**Xe** 54
6	Longest period	32	55 – 86	6	**Cs** 55		**La Hg** 57 ↔ 80						**Rn** 86
7	Longest period	32	87 – 118	7	**Fr** 87		**Ac Cn** 89 ↔ 112						**Uuo** 118

Note : Lanthanides **Group 3 of the sixth period** and actinides **Group 3 of the seventh period** have similar properties because they belong to the same Group 3. They are shown at the bottom of the periodic table because they are large in number, and if shown in the main body of the table will distort its shape.

The **third period** elements, Na, Mg, Al, Si, P, S and Cl, summarise the properties of their respective groups and are called **typical elements.**

A period is determined by the number of shells and a group is determined by the number of electrons present in the outermost shell. For example : Sodium has atomic number 11 and its electronic configuration is 2, 8, 1. It has three orbits (shells) and has one electron in the outermost orbit hence it is placed in third period and group 1. Similarly, calcium, atomic number 20, electronic configuration 2, 8, 8, 2 is placed in fourth period and group 2.

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1.3 PERIODICITY

The properties that reappear at regular intervals, or in which there is gradual variation (*i.e.* increase or decrease) at regular intervals, are called *'periodic properties'* and the phenomenon is known as the *periodicity of elements.*

Cause of periodicity

The cause of periodicity is the recurrence of similar electronic configuration i.e. having same number of electrons in the outermost orbit.

In a particular group, **electrons in the outermost orbit** remain the same *i.e.* **electronic configuration is similar.** Since **chemical properties of elements depend upon the number of electrons in their outermost shell, thus elements of the same group have similar properties**.

For example in group 17, *i.e.,* halogens, all elements have seven electrons (see electronic configuration given in 1.4) in their respective outermost shells, therefore, they show similar properties, such as:

(i) they are coloured non-metals.

(ii) they form negative ions carrying a single charge. For example chloride ions (Cl^-).

(iii) they are very reactive and are, therefore, found in combined state.

(iv) they are only slightly soluble in water, but they dissolve much better in organic solvents like carbon disulphide, chloroform, alcohol, etc.

(v) their melting and boiling points increase regularly moving down the group.

(vi) they are good oxidising agents.

It is thus concluded that periodicity is due to the same number of electron(s) in the outermost orbit of different elements.

1.4 SHELLS (ORBITS) AND VALENCY

Orbits : Electrons revolve around the nucleus in certain definite circular paths called **orbits** or **shells.**

(1) Number of shells :

(a) ***Down a group, i.e., from top to bottom.***

The number of shells increases successively, *i.e.,* one by one, such that the number of shells that an element has, *equals the number of the period* to which that element belongs.

For example, in halogens (Group 17).

Element with atomic numbers	No. of shells	Electronic configuration K L M N O P	Period to which the element belongs
F (9)	2	2, 7	Second
Cl (17)	3	2, 8, 7	Third
Br (35)	4	2, 8, 18, 7	Fourth
I (53)	5	2, 8, 18, 18, 7	Fifth
At (85)	6	2, 8, 18, 32, 18, 7	Sixth

(b) ***Across a period, i.e., from left to right.***

On moving from left to right in a given period, the number of shells remains the same. *For example*, in the 2nd period, the number of shells remains two, *i.e.*, equal to the number of the period. Similarly in the third period the number of shells remains three and so on.

(2) Valency

Valency denotes the combining capacity of the atom of an element. It is equal to the number of electrons an atom can donate or accept or share.

On moving down a given group, the number of electrons in the outermost shell, i.e., valence electron, remains the same. Therefore, **valency,** in a group, also remains the same.

In a given period, the number of electrons in the valence (outermost) shell increases from left to right. But the valency increases only upto Group 14, where it becomes 4, and then it decreases, *i.e.*, in Group 17 it becomes 1.

Note 1 : Valency depends on the number of electrons in the outermost shell (*i.e.* valence shell). If the number of electrons present in the outermost shell are 1, 2, 3 or 4, then their valency is 1, 2, 3 or 4 respectively. If the number of electrons present in the outermost shell are 5, 6 or 7, then their valency is 8 – 5 = 3, 8 – 6 = 2 and 8 – 7 = 1 respectively.

Note 2 : Valency is the combining capacity so it is always positive.

Groups →	IA 1	IIA 2	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	Zero 18
Elements of the 2nd period	Li	Be	B	C	N	O	F	Ne
Atomic No.	3	4	5	6	7	8	9	10
Electronic configuration	K L 2, 1	K L 2, 2	K L 2, 3	K L 2, 4	K L 2, 5	K L 2, 6	K L 2, 7	K L 2, 8
No. of shells	2	2	2	2	2	2	2	2
Valency	1	2	3	4	3	2	1	0
Formula of Hydride	LiH	BeH_2	BH_3	CH_4	NH_3	H_2O	HF	-

In the periodic table
- elements are arranged in order of increasing atomic number (proton number).
- the vertical columns of elements with similar properties are called groups.
- the horizontal rows are called periods.

Intext Questions

1. (a) State modern periodic law. Name the scientist who stated the law.
   (b) What is a periodic table ? How many groups and periods does modern periodic table have?
2. Elements of group 1 and elements of group 17 both have valency 1. Explain.
3. What are horizontal rows and vertical columns in a periodic table known as ?
4. Periodicity is observed due to the similar .... (number of valence electrons/atomic number/electronic configuration).
5. How does the electronic configuration in atoms change
   (i) in a period from left to right ?
   (ii) in a group top to bottom ?
6. Correct the statements.
   (i) Elements in the same period have the same valency.
   (ii) Valency depends upon the number of shells in an atom.
   (iii) Copper and zinc are representative elements.
   (iv) Transition elements are placed at extreme right of the periodic table.
7. Name two elements in each case :
   (i) Alkali metals (ii) Alkaline earth metals
   (iii) Halogens (iv) Inert gas
   (v) Transition element (vi) Lanthanides
   (vii) Actinides
8. What do you understand by ?
   (i) Periodicity (ii) Typical elements
   (iii) Orbits
9. Name two elements that you would expect to show chemical reactions similar to calcium. What is the basis of your choice ?
10. Name the (i) metals (ii) metalloids and (iii) non-metals in the first twenty elements.
11. Fluorine, chlorine and Bromine are put in one group on the basis of their similar properties.
    (i) What are those similar properties ?
    (ii) What is the common name of this group or family ?
12. What is the main characteristic of the last element in each period of the Periodic Table ? What is the general name of such elements ?
13. According to atomic structure, what determines which element will be the first and which will be the last in a period?
14. How does the number of :
    (i) valence electrons and
    (ii) valency vary on moving from left to right in the second period of the periodic table ?
15. Name the type of elements, which have their
    (i) outermost shell complete
    (ii) outermost shell incomplete
    (iii) two outermost shell incomplete
    (iv) one electron short of octet
    (v) two electrons in the outermost orbit.
16. An element has 2 electrons in its N shell.
    (i) What is its atomic number ?
    (ii) State its position in periodic table
    (iii) Is it metal or non-metal ?
    (iv) State the name assigned to this group.
17. Answer the following in respect of element $^{32}_{16}S$.
    (i) Give its electronic configuration.
    (ii) To which group and period does it belong ?
    (iii) What is its valency ?
    (iv) Is it metal or non-metal ?
    (v) Is it a reducing agent or an oxidising agent ?
    (vi) Give its formula with chlorine.

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1.5 PERIODIC PROPERTIES

The properties of elements which are directly or indirectly related to their electronic configurations and show a regular gradation as we move across a period, from left to right or down the group from top to bottom, are called **Periodic Properties**.

Important periodic properties are :

(i) ***atomic size*** (atomic radius),

(ii) ***metallic character,***

(iii) ***non-metallic character,***

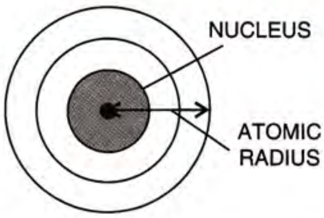
(iv) ***ionisation potential*** (ionisation energy),

(v) ***electron affinity,***

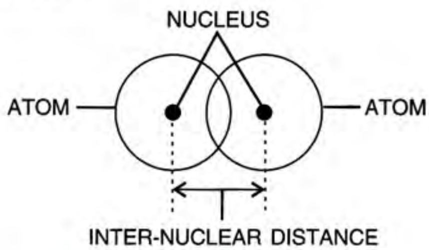
(vi) ***electronegativity.***

1.5.1 Atomic size (atomic radius)

It is the distance between the centre of the nucleus of an atom and its outermost shell.



Atomic radius can also be defined as half the inter-nuclear distance between the combined atoms in a molecule.



Unit : Angstrom : 1A = 10^{-10} m
Picometre : 1 pm = 10^{-12} m.

Atomic size depends upon :

(i) number of shells and (ii) nuclear charge.

(i) *Number of shells*

An increase in the number of shells increases the size of an atom because the distance between the outermost shell and the nucleus increases.

(ii) *Nuclear charge*

It is the positive charge present in the nucleus of an atom, which is equal to the number of protons in the nucleus, *i.e.*, the atomic number.

An increase in nuclear charge decreases the size of the atom because the electrons are then attracted towards the nucleus with a greater force, thereby bringing the outermost shell closer to the nucleus.

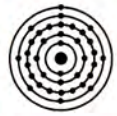
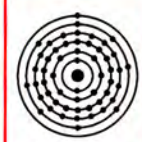
Trends in atomic size (atomic radius)

(a) Down a group :

In a group, the size of an atom increases as one proceeds from top to bottom. This is due to the successive addition of shells (which overweighs the increased nuclear charge) as one moves from one period to the next in a group. For example :

***Elements in Group* 1**

In group 1 the size of hydrogen is smallest.

Elements	Atomic No.	Atomic radius	Electronic configuration K L M N O P	Size
Hydrogen	1	37 pm	1	
Lithium	3	152 pm	2, 1	
Sodium	11	186 pm	2, 8, 1	
Potassium	19	231 pm	2, 8, 8, 1	
Rubidium	37	244 pm	2, 8, 18, 8, 1	
Caesium	55	262 pm	2, 8, 18, 18, 8, 1	

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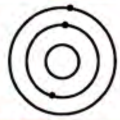
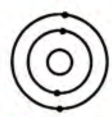
In group 17 the size of fluorine is smallest.

F	<	Cl	<	Br	<	I	<	At
64 pm		99 pm		114 pm		133 pm		140 pm

(b) Across a period :

In a period, the size of an atom decreases from left to right. This is because the nuclear charge, *i.e.*, the atomic number increases from left to right in the same period, thereby bringing the outermost shell closer to the nucleus. Therefore, considering the second period given above, it has been found that lithium (Li) has the largest atomic size while fluorine (F) has the smallest.

Trends in atomic size across a period.

Elements	Lithium	Beryllium	Boron	Carbon	Nitrogen	Oxygen	Fluorine
Atomic No.	3	4	5	6	7	8	9
Atomic Radius	152 pm	112 pm	88 pm	77 pm	70 pm	66 pm	64 pm
Electronic configuration	K L 2, 1	K L 2, 2	K L 2, 3	K L 2, 4	K L 2, 5	K L 2, 6	K L 2, 7
Size							

In third period, sodium is largest in size

Na	>	Mg	>	Al	>	Si	>	P	>	S	>	Cl
186 pm		160 pm		143 pm		117 pm		110 pm		104 pm		99 pm

Note : As an exception the size of the atoms of inert gases are bigger. This is because the outer shell of inert gases is complete They have the maximum number of electrons in their outer most orbit thus the electronic repulsions are maximum. Hence the size of the atom of an inert gas is bigger.

Note : **1. Cation is always smaller than the parent atom,** from which it is formed.

Reason :
Cation is formed by the loss of electron(s), hence proton(s) are more than electron(s) in a cation. So electrons are strongly attracted by the nucleus and are pulled inward. Hence the size decreases.

$$\underset{(2,\,8,\,1)}{Na - e^-} \rightarrow \underset{(2,\,8)}{Na^+}$$

2. Anion is larger than the parent atom.

Reason : Anion is formed by the gain of electron(s). Thus, the number of electron(s) are more than proton(s). The effective positive charge in the nucleus is less, so less inward pull is experienced. Hence the size expands.

$$\underset{(2,\,8,\,7)}{Cl + e^-} \rightarrow \underset{(2,\,8,\,8)}{Cl^-}$$

3. Size of isoelectronic ions *i.e.* the ions having the same number of electrons depends upon the nuclear charge (no. of protons). Greater is the nuclear charge smaller is the size .

Example :

Isoelectronic ions	Mg^{2+}	Na^+	F^-	O^{2-}
No. of electrons	10	10	10	10
No. of protons	12	11	9	8
Size in A°	0·65	0·95	1·26	1·42

1.5.2 Metallic character

Those elements, which have a tendency to lose their valence electrons (electrons of the outermost orbit) *and form a positive ion, are considered metals.*

$$\underset{(2,\,8,\,1)}{Na - e^-} \rightarrow \underset{(2,\,8)}{Na^+}$$

$$\underset{(2,\,8,\,2)}{Mg - 2e^-} \rightarrow \underset{(2,\,8)}{Mg^{2+}}$$

Note : Hydrogen is an element (non metal) which does not have a neutron, it has only one electron and one proton. On losing this electron it forms its cation which has only one proton, therefore its cation can also be called a **proton**. $H - e^- \rightarrow H^+$

The metallic character of elements depends on:

(i) atomic size and (ii) nuclear charge.

(i) ***Atomic size*** **:** The greater the **atomic size**, the farther the outermost orbit, and thus, lesser the nuclear pull exerted on it. As a result, electron(s) can be removed more easily from the valence shell, thus making the elements more metallic.

(ii) ***Nuclear charge*** **:** The greater the **nuclear charge**, the greater is the force exerted by the nucleus on the electron(s) of the outermost orbit. This makes it difficult to remove the electron(s) from the outermost orbit. Thus metallic nature decreases.

Elements which lose electron(s) to complete their octet are **reducing agents**. Metals are good reducing agents. Greater the tendency to lose electron(s) stronger is the reducing agent.

Trends in metallic character

Down a group

On moving ***down a group***, the atomic size increases and the nuclear charge also increases. The effect of an increased atomic size is greater as compared to the increased nuclear charge. Therefore, ***metallic nature increases as one moves down a group,*** *i.e.,* they can lose electrons easily.

Li
Na
K
Rb
Cs
↓ Metallic nature increases

Example : **In group 1,** lithium is the least metallic element.

Note : Francium is radioactive element so its properties are not known.

Table 1.2

Group	*GROUP 1*	*GROUP 2*	*GROUP 13*	*GROUP 14*	*GROUP 15*	*GROUP 16*	*GROUP 17*	*GROUP 18*
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic number	11	12	13	14	15	16	17	18
Electronic configuration	K, L, M 2, 8, 1	K, L, M 2, 8, 2	K, L, M 2, 8, 3	K, L, M 2, 8, 4	K, L, M 2, 8, 5	K, L, M 2, 8, 6	K, L, M 2, 8, 7	K, L, M 2, 8, 8
Metallic property	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal	Noble gas
Valency	1	2	3	4	3	2	1	0

Across a period

On moving ***across a period***, nuclear pull increases due to the increase in atomic number, and thus the atomic size decreases. Hence, elements cannot lose electrons easily.

Therefore, ***the metallic nature decreases across a period, moving from left to right.***

Example : **In the 2nd period,** lithium is the most metallic.

Li Be	B	C N O
Metals	Metalloid	Non-metals

In the 3rd period, *sodium is the most metallic element.*

1.5.3 Non-metallic character

Those elements, which have a tendency to gain electrons, in order to attain octet in their outermost orbit, are considered as ***non-metals.***

Non-metals usually have 5, 6 or 7 electrons in their outermost orbits. They can gain 3, 2 or 1 electron respectively, in order to attain octet or to complete their valence shells (outermost orbits) and form an anion.

For example, chlorine (Cl) and oxygen (O) gain 1 and 2 electron(s) respectively to form anions.

$$\underset{(2,\,8,\,7)}{Cl} + e^- \rightarrow \underset{(2,\,8,\,8)}{Cl^-}$$

$$\underset{(2,\,6)}{O} + 2e^- \rightarrow \underset{(2,\,8)}{O^{2-}}$$

Non-metallic character also depends on :

(i) atomic size and

(ii) nuclear charge.

(i) ***Atomic size*** : The smaller the **atomic size,** the greater is the nuclear pull and the tendency to gain electrons. Therefore, the element is more non-metallic by nature.

(ii) ***Nuclear charge :*** The greater the nuclear charge, the greater is the tendency to gain electron(s), hence more non-metallic is the element.

Trends in non-metallic character

Down a group

The atomic size increases, due to the addition of new shells over successive periods. Though the nuclear charge increases, due to an increase in atomic number, the effect of an increasing atomic size is greater. Therefore, ***non-metallic nature decreases down the group.***

Example : **Group 14**

C	Non metal
Si	Metalloid
Ge	Metalloid
Sn	Metal
Pb	Metal

Across the period (left to right)

On moving across a period, the tendency to gain electron(s) increases due to an increase in nuclear pull and a decrease in the atomic size. Therefore, *non-metallic character increases across a period, i.e., from left to right.*

For example, in the 3rd period,

$$\underbrace{Na < Mg < Al}_{\text{Metals}} < \underbrace{Si}_{\text{Metalloid}} < \underbrace{P < S < Cl}_{\text{Non-metals}}$$

Non-metallic character increases, or metallic character decreases. →

Note : Non-metals are oxidising agents.

The nature of oxides also show periodicity.

Across a period : The oxides of elements in a particular period *show decreasing basic nature* and finally become acidic. *For example*, the oxides of the elements of the third period.

Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Strongly basic	Basic	Amphoteric	Feebly acidic	Acidic	More acidic	Most acidic

Down a group : The basic nature of oxides of metals increases.

Chemical reactivity

The reactivity of elements depend upon their tendency to lose or gain electrons to complete their outermost orbit. Greater the tendency to lose electron(s), greater is the reactivity in the case of metals. Similarly greater the tendency to gain electron(s), greater is the reactivity of non-metals.

Trends in chemical reactivity

Across a period :

On moving from left to right in a period, the chemical reactivity of elements first decreases and then increases.

The group 1 element can lose electrons easily as compared to group 2. Similarly, group 2 elements can also lose electrons easily but only in comparison to group 13 and not at all in comparison to group 1. As the tendency to lose electrons decreases, reactivity also decreases and thus silicon is the least reactive element in third period. As we move from phosphorus to chlorine, the tendency to gain electrons increases, hence reactivity again increases., *e.g.,*

Third period :

Na	Mg	Al	Si	P	S	Cl
Most reactive metal	reactivity decreases →		least reactive	← reactivity decreases		Most reactive Non-metal

Down a group

The tendency of loosing electrons increases down the group. Since chemical reactivity in metals depends upon the tendency to lose electrons, thus reactivity increases on going down the group.

The most reactive metal is at the bottom of group 1.

The chemical reactivity of non-metals decreases on going down the group as it depends upon the tendency to gain electrons, which decreases down the group. The most reactive non-metal is at the top of group 17 *i.e.*, Fluorine.

Gradation in physical properties

The melting and boiling points of metals **decrease** on going down the group, *e.g.*,

Metals	m.p.	b.p.
Li	180·5°C	1347°C
Na	94·5°C	883°C
K	63·5°C	774°C

↓ decreases

The melting and boiling points of non-metals **increase** on going down the group.

Non-Metals	m.p.	b.p.	Physical state
Fluorine	–219·6°C	–187°C	gas
Chlorine	–101°C	–34·6°C	gas
Bromine	–7·2°C	+58·8°C	liquid
Iodine	+113·6°C	+183°C	solid

↓ increases

Note : **Across a period**, left to right, melting point and boiling point usually increases upto group 14 (IV A) and then decreases.

Element of 3rd period	Na	Mg	Al	Si	P	S
M.P. (°C)	98	650	660	1410	44·2	115·2

Density of elements across a period increases gradually to maximum and then slight decrease may be noticed.

Element	Na	Mg	Al	Si	P	S
Density g/cc	1·0	1·7	2·7	2·3	1·8	2·1

Down a group density of elements increases gradually.

Element	Li	Na	K	Rb	Cs
Density g/cc	0·54	0·97	0·86	1·53	1·87

Intext Questions

1. What do you understand by atomic size ? State its unit.
2. Give the trends in atomic size on moving :
   (i) down the group,
   (ii) across the period left to right.
3. Arrange the elements of second and third period in increasing order of their atomic size.
4. Why is the size of (i) neon greater than fluorine? (ii) sodium is greater than magnesium ?
5. Which is greater in size ?
   (i) an atom or a cation (ii) an atom or an anion
   (iii) Fe^{2+} or Fe^{3+}
6. Arrange :
   (i) Be, Li, C, B, N, O, F (in increasing metallic character).
   (ii) Si, Na, Al, Mg, Cl, P, S (in decreasing non-metallic character).
7. State the trend in chemical reactivity :
   (i) across the period left to right,
   (ii) down the group.
8. A metal M forms an oxide having the formula M_2O_3. It belongs to third period. Write the atomic number and valency of the metal.
9. An element X belong to 4th period and 17th group, state
   (i) no. of valence electrons in it.
   (ii) name of the element.
   (iii) name the family to which it belong.
   (iv) write the formula of the compound formed when X reacts with $^{27}_{13}Y$.
10. The given table shows elements with same number of electrons in its valence shell.

Elements	A	B	C
m.p.	63·0	180·0	97·0

   State :
   (i) Whether these elements belong to same group or period.
   (ii) Arrange them in order of increasing metallic character.
11. Which one of the following has the largest atomic radius ?
   (i) Sodium (ii) Potassium
   (iii) Magnesium (iv) Aluminium
12. Which one has the largest size ?
   (i) Br (ii) I (iii) I^- (ii) Cl

13. The metals of group 2 from top to bottom are Be, Mg, Ca, Sr and Ba
   (i) Which one of these elements will form ions most readily and why ?
   (ii) State the common feature in their electronic configuration.
14. Write the number of protons, neutrons and electronic configuration of ${}^{39}_{19}K$, ${}^{31}_{15}P$. Also state their position in periodic table.
15. The electronic configuration of an element T is 2, 8, 8, 1.
   (i) What is the group number of T ?
   (ii) What is the period number of T ?
   (iii) How many valence electrons are there in an atom of T ?
   (iv) What is the valency of T ?
   (v) Is it a metal or a non-metal ?
16. Complete the following sentences choosing the correct word or words from those given in brackets at the end of each sentence :
   (i) The properties of the elements are a periodic function of their ________ (atomic number, mass number, relative atomic mass).
   (ii) Moving across a ______ of the Periodic Table the elements show increasing _____ character (group, period, metallic, non-metallic).
   (iii) The elements at the bottom of a group would be expected to show ______ metallic character than the element at the top (less, more).
   (iv) The similarities in the properties of a group of elements are because they have the same ________ (electronic configuration, number of outer electrons, atomic numbers).
17. Give reasons for the following :
   (i) The size of a Cl^- ion is greater than the size of a Cl atom.
   (ii) Argon atom is bigger than chlorine atom.
   (iii) Ionisation potential of the element increases across a period.
   (iv) Alkali metals are good reducing agents.
18. Name the element which has :
   (i) two shells, both of which are completely filled with electrons ?
   (ii) the electronic configuration 2, 8, 3 ?
   (iii) a total of three shells with five electrons in its valence shell ?
   (iv) a total of four shells with two electrons in its valence shell ?
   (v) twice as many electrons in its second shell as in its first shell ?
19. (i) State the number of elements in Period 1, Period 2, and Period 3 of the periodic table. Name them.
   (ii) What is the common feature of the electronic configuration of the elements at the end of Period 2 and Period 3 ?
   (iii) If an element is in Group 17, it is likely to be ________ [metallic/non-metallic] in character, while with one electron in its outermost energy level (shell), then it is likely to be ______ [metallic/non-metallic].

1.5.4 Ionisation potential or Ionisation energy (I.E.) or ionisation enthalpy

We know that the electrons in an atom are attracted by the positively charged nucleus. So, if we want to remove an electron from an atom, some energy has to be supplied to overcome the strong attraction of the nucleus. And after the electron gets removed from the atom, the atom then acquires one unit positive charge and becomes a positive ion (or cation).

The energy required to remove an electron from a neutral isolated gaseous atom and convert it into a positively charged gaseous ion is called ionisation potential (I.P.) or ***ionisation energy (I.E.)*** or **first ionisation energy** (IE_1).

$M(g) + I.E. \rightarrow M^+(g) + e^-$

M can be any element.

Unit : *I.E. is measured in electron volts per atom (eV/atom) and its S.I. unit is kilojoule per mole (kJ mol^{-1}).*

Ionisation energy depends on :

(i) atomic size (ii) nuclear charge

(i) ***Atomic size :*** The greater the atomic size, the lesser the force of attraction. Since the electrons of the outermost shell lie further away from the nucleus, thus make their removal easier, *i.e.*, the ionisation energy required is less.

(ii) ***Nuclear charge :*** The greater the nuclear charge, greater is the attraction for the electrons of the outermost shell. Therefore, the electrons in the outermost shell are more firmly held because of which greater energy is required to remove the electron(s).

Trends in ionisation energy

Across a period

The ionisation energy tends to increase as one moves from left to right across a period (with exceptions), because the atomic size decreases due to an increase in the nuclear charge, and thus, more energy is required to remove the electron(s).

The elements of the 2nd period	Li	Be	B	C	N	O	F	Ne
Ionisation energy in kJ mol^{-1}	520	899*	801	1088	1402*	1314	1681	2080

The elements of the 3rd period	Na	Mg	Al	Si	P	S	Cl	Ar
I.E. in kJ mol^{-1}	496	737*	577	786	1011*	999	1256	1520

Down a group

There is an increase in atomic number (nuclear charge) and atomic size down the group due to the addition of extra shells. This increase in the atomic size overcomes the effect of an increase in the nuclear charge.

Therefore, *ionisation energy decreases with an increase in the atomic size, i.e., it decreases as one moves down a group.*

The elements of Group 1	Ionisation energy in kJ mol^{-1}
H	1312
Li	520
Na	498
K	419
Rb	403
Cs	375

The elements of Group 17	Ionisation energy in kJ mol^{-1}
F	1681
Cl	1256
Br	1143
I	1008

* Exceptions

Note : • Helium will have the highest ionisation energy 2372·0 kJ mol^{-1} while caesium (Cs) will have the lowest ionisation energy of 375·0 KJ mol^{-1} (I.E. of Fr is not determined correctly as it is radioactive).

• Metals usually have low I.E. whereas non-metals have high I.E.

1.5.5 Electron Affinity (E.A.) or Electron Gain Enthalpy

We have just studied that to remove an electron from an atom, energy is required. Similarly, when an extra electron is added to an atom, energy is released.

The amount of energy released while converting a neutral gaseous isolated atom into a negatively charged gaseous ion (anion) by the addition of electron is called Electron Affinity (E.A.)

$$X(g) + e^- \rightarrow X^-(g) + E.A.$$

X is any element taken in its gaseous state.

Unit : *electron volts per atom (eV/atom) or kJ* mol^{-1}*. Electron affinity is represented by negative sign [–] e.g.,*

$$Cl(g) + e^- \rightarrow Cl^-(g) + 349 \text{ KJ/mol.}$$

Therefore, electron affinity of chlorine is – 349 KJ/mol.

Electron affinity depends on :

(i) atomic size and (ii) nuclear charge.

(i) Atomic size : *The smaller the atomic size, the greater the electron affinity,* because the effective attractive force between the nucleus and the valence electrons is greater in smaller atoms, and so the electrons are held firmly.

(ii) Nuclear charge : *The greater the nuclear charge, greater is the electron affinity,* because increase in nuclear charge increases the effective attractive force on the valence electrons.

Variation (Trends) in electron affinity

In a period, i.e., from left to right in a horizontal row of the periodic table, the atomic size decreases and the nuclear charge increases, so the electron affinity increases. Thus electron affinity (E.A.) is the highest for group 17 elements (halogens) and the least for group 1 (alkali metals).

Electron Affinity Values in KJ mol^{-1}

	IA (1)	IIA (2)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	Zero (18)
I	H −72·8							He*
II	Li −59·8	Be*	B −88·0	C −122	N*	O −140·9	F −327·9	Ne*
III	Na −53·1	Mg*	Al −50·0	Si −119	P −74	S −200	Cl −349	Ar*

Down a group

Group I	
Li	−59·8
Na	−53·1
K	−48·4
Rb	−47·0

Moving from the top to the bottom in a group, the atomic size increases more than the nuclear charge, thereby causing a net decrease in E.A.

> *Note :* In group 17 fluorine has lower E.A. than chlorine and in group 16 oxygen has lower E.A. than sulphur. This is because the size of Fluorine and oxygen atom is very small. As a result, there are strong inter electronic repulsions and thus incoming electron does not feel much attraction.

1.5.6 Electronegativity (E.N.)

> The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is called its ***electronegativity.***

Electronegativity is a *dimensionless property,* since it is only a tendency. It only indicates the net result of the tendencies of different elements to attract the bond-forming electron pair.

Electronegativity is measured on several scales. The most widely used scale of electronegativity was devised by ***Linus Pauling*** (1932) in which the highest value of electronegativity [for fluorine] is taken as 4·0 and the lowest one [for caesium] as 0·7.

ELECTRONEGATIVITY VALUES

DECREASING ↓	1	2	13	14	15	16	17
	H 2·1						
	Li 1·0	Be 1·5	B 2·0	C 2·5	N 3·0	O 3·5	F 4
	Na 0·9	Mg 1·2	Al 1·5	Si 1·8	P 2·1	S 2·5	Cl 3·0
	K 0·8	Ca 1·0	Ga 1·6	Ge 1·8	As 2·0	Se 2·4	Br 2·8
	Rb 0·8	Sr 1·0	In 1·7	Sn 1·8	Sb 1·9	Te 2·1	I 2·5
	Cs 0·7	Ba 0·9	Tl 1·8	Pb 1·9	Bi 1·9	Po 2·0	At 2·2

I N C R E A S I N G →

Electronegativity values too depend on :

(i) size of atom and

(ii) nuclear charge.

(i) **Atomic size :** *The greater the size of the atom, the lesser the electronegativity,* since the electrons *being farther away from the nucleus, experience a lesser force of attraction.*

(ii) **Nuclear charge :** *The greater the nuclear charge, the greater the electronegativity,* because increase in nuclear charge causes electron attraction with a greater force.

Trends in electronegativity

Across a period

Since the nuclear charge increases due to an increase in atomic number, ***electronegativity increases from left to right in a period.***

For example, in the second period, electronegativity increases from lithium to fluorine.

Down a group

There is an increase in atomic number down a group, *i.e.*, nuclear charge increases, but due to the addition of extra shells, the atomic size increases. The effect of an increase in the atomic size overcomes the effect of an increase in the nuclear charge, hence ***electronegativity decreases down a group.***

For example, in the first group, it significantly decreases from lithium downwards to francium.

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* *Inert gases have zero electron affinity due to their stable electronic configuration.*

Groups 2 and 15 do not show negative values. They are exceptions.

Note : *1. Generally, metals show lower electronegativity as compared to non-metals. Thus metals are electropositive and non-metals are electronegative.*

2. The elements of second period differs in properties from their respective groups. This is due to small size of the atom and their high electronegativity.

Diagonal relationship : The elements of the **second period** show resemblance in properties with the elements of the next group of the **third period**, due to very less electronegativity difference. This leads to a **diagonal relationship**, *viz.* Li & Mg, Be & Al, B & Si. These elements are called **bridge elements**.

Group →	**1**	**2**	**13**	**14**
Period 2	Li	Be	B	C
Period 3	Na	Mg	Al	Si

Noble gases have complete octet so they do not have tendency to attract electrons.

Note : Greater the value of electron affinity easier it is to gain electron(s) and more non metallic or more electronegative or more oxidising is the element.

Intext Questions

1. (a) Define the term 'ionisation potential'. **(2010)**
   (b) Represent it in the form of an equation. In which unit it is measured?
2. Ionisation Potential values depends on (a) atomic size (b) nuclear pull. Explain.
3. State the trends in ionisation energy :
   (a) across the period, (b) down the group.
4. Name the elements with highest and lowest ionisation energies in first three periods.
5. Arrange the elements of second and third period in increasing order of ionisation energy.
6. (a) Define the term 'electron affinity'. State its unit. **(2010)**
   (b) Arrange the elements of second period in increasing order of their electron affinity. Name the elements which do not follow the trend in this period.
7. Electron affinity values generally .... across the period left to right and .... down the group top to bottom.
8. (a) Define the term 'Electronegativity'. State its unit.
   (b) Among the elements given below, the element with least Electronegativity is :
   (i) Lithium (ii) Boron (iii) Carbon (iv) Fluorine
9. Explain the following :
   (a) Group 17 elements are strong non-metals, while group I elements are strong metals.
   (b) Metallic character of elements decreases from left to right in a period while it increases in moving down a group.
   (c) Halogens have a high electron affinity.
   (d) The reducing power of element increases down in the group while decreases in a period.
   (e) Size of atoms progressively becomes smaller when we move from sodium (Na) to chlorine (Cl) in the third period of the Periodic Table.
10. Name the periodic property which relates to the :
    (a) amount of energy required to remove an electron from an isolated gaseous atom,
    (b) character of element which loses one or more electrons when supplied with energy,
    (c) tendency of an atom to attract the shared pair of electron.
11. This question refers to the elements of the Periodic Table with atomic numbers from 3 to 18. Some of the elements are shown by letters, but the letters are not the usual symbols of the elements.

3	4	5	6	7	8	9	10
A	B	C	D	E	F	G	H
11	12	13	14	15	16	17	18
I	J	K	L	M	N	O	P

Which of these :
(a) are most electronegative element.
(b) is a halogen ?
(c) is an alkali metal ?
(d) is an element with valency 4 ?
(e) have least Ionisation Energy.
(f) have least atomic size in period 3.

1.6 RELATION BETWEEN ATOMIC NUMBER AND MASS NUMBER OF LIGHT ELEMENTS

Atomic number (Z)

The atomic number of an element is equal to the number of protons in the nucleus. ***Atomic number is the unique property of an element, because no two elements have the same atomic number.***

The significance of atomic number (Z)

1. Atomic number = Number of protons.
   = Number of electrons.
2. It distinguishes an element from other elements, because no two elements have the same atomic number.
3. It gives the electronic configuration of an element, *e.g.*, an element with atomic number 13 will have electronic configuration 2, 8, 3.
4. It helps us in finding the position of an element in the periodic table.

   For example, the element with atomic number 17 will have electronic configuration 2, 8, 7. This element will be placed in the 3rd period of Group 17 *i.e.*, VIIA, because it has : (i) three energy shells and (ii) seven electrons in its outermost shell.

Table 1.4 : Arrangement of electrons, protons and neutrons of elements from atomic number 1 to 20

Elements	*Symbol*	*Atomic No. (Z) No. of protons (or electrons)*	*Mass No. (A) or Protons + neutrons*	*No. of neutrons (A – Z)*
Hydrogen	H	1	1	—
Helium	He	2	4	2
Lithium	Li	3	7	4
Beryllium	Be	4	9	5
Boron	B	5	11	6
Carbon	C	6	12	6
Nitrogen	N	7	14	7
Oxygen	O	8	16	8
Fluorine	F	9	19	10
Neon	Ne	10	20	10
Sodium	Na	11	23	12
Magnesium	Mg	12	24	12
Aluminium	Al	13	27	14
Silicon	Si	14	28	14
Phosphorus	P	15	31	16
Sulphur	S	16	32	16
Chlorine	Cl	17	35, 37	18, 20
Argon	Ar	18	40	22
Potassium	K	19	39	20
Calcium	Ca	20	40	20

Mass number (A)

The mass number of an element is the sum of the number of protons and neutrons in the nucleus of the atom of that element.

Mass number (A) = No. of protons (p)
+ No. of neutrons (n)

A close look at the electronic configurations of lighter elements reveals that those elements which have an ***even*** *number of protons, for example, atomic numbers like* $^{4}_{2}$He, $^{12}_{6}$C, etc., *have their mass numbers twice the atomic numbers except* $^{9}_{4}$Be and $^{40}_{18}$Ar.

Elements which have an **odd** *number of protons like* $^{7}_{3}$Li, $^{11}_{5}$B, $^{19}_{9}$F, $^{23}_{11}$Na, *etc., have their mass numbers twice the atomic numbers* +1 (A = 2Z + 1) *except* $^{14}_{7}$N and $^{1}_{1}$H.

Note : (i) Elements with *n/p* (neutron/proton) ratio around 1 are stable, *e.g.*, light metals like sodium, potassium, calcium, etc.

(ii) Elements with *n/p* ratio 1·5 and above are radioactive, *i.e.*, they emit radiations. They are unstable elements, *e.g.*, heavy metals like uranium.

1.7 COMPARISON OF ALKALI METALS AND HALOGENS

	Alkali metals [Group 1]	**Halogens [Group 17]**
Elements	Li, Na, K, Rb, Cs, Fr	F, Cl, Br, I, At
Occurrence	*Combined state.* Due to their reactive nature.	Combined state as *salts.* Due to their reactive nature.
Physical state	Shining white solid metals. Soft, and can be cut with a knife. Lithium is the hardest. They are shiny when freshly cut but soon becomes dull as they react with air.	Non-metals; diatomic in the gaseous state. Flourine (very reactive poisonous yellow gas) chlorine (poisonous yellow green gas) Bromine (poisonous red brown volatile liquid) Iodine (dark grey crystalline solid).

Valence electrons	Possess *one* valence electron and therefore show similar properties.	Possess *seven* valence electrons each and therefore show similar properties.
Conduction	*Good conductors* of electricity.	*Non-conductors* of electricity.
Nature	Highly reactive, *electropositive* metals. Metallic character, *increases* from lithium (Li) to francium (Fr).	Highly reactive, *electronegative* non-metals. Non-metallic character, decreases from fluorine (F) to iodine (I).
Melting point and boiling point	Decreases down the group.	Increases down the group.
Atomic size	They have the largest atomic size in their period (except inert gases). The atomic size further increases down the group.	They have the smallest atomic size in their period. Atomic size increases down the group.
Ionisation energy	They have lowest I.E. in their period. It decreases down the group.	They have high I.E. (lower than noble gases) in their period.
Electron affinity	They have low E.A. values which further decrease down the group.	They have high E.A. values. They too decrease down the group.
Electronegativity	They have the lowest E.N. in their period. E.N. decreases down the group.	They have high E.N. highest in their period.
Reactivity	They are reactive metals. Reactivity further *increases* down the group.	They are reactive non-metals. The reactivity *decreases* down the group.
Reaction with water and acids	They react vigorously with water and acids liberating hydrogen. Reactivity further increases down the group.	Generally they do not react with dil. acids and water.
Reducing/ oxidising agents	Strong *reducing agents* as they lose electrons to complete their octet.	Strong *oxidising agents* as they accept electrons to complete their octet.
Compound formation	Form *electrovalent* compounds with non-metals. Example : NaCl, KBr.	Form *electrovalent compounds* with metals. *e.g.,* KCl, $CaCl_2$. Form ***covalent compounds*** with hydrogen and other non-metals, *e.g.,* HBr, HCl, HI,CCl_4.

Period No.	Type of period	No. of elements	Atomic number	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIII	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0
1	Shortest	2	1 and 2	H_1 1																	He_2 2
2.	Short	8	3 to 10	Li_3 2,1	Be_4 2,2											B_5 2,3	C_6 2,4	N_7 2,5	O_8 2,6	F_9 2,7	Ne_{10} 2,8
3.	Short	8	11 to 18	Na_{11} 2,8,1	Mg_{12} 2,8,2											Al_{13} 2,8,3	Si_{14} 2,8,4	P_{15} 2,8,5	S_{16} 2,8,6	Cl_{17} 2,8,7	Ar_{18} 2,8,8
4.	Long	18	19 to 36	K_{19} 2,8,8,1	Ca_{20} 2,8,8,2				TRANSITION		ELEMENTS									Br_{35}	Kr_{36}
5.	Long	18	37 to 54	Rb_{37}																I_{53}	Xe_{54}
6.	Longest	32	55 to 86	Cs_{55}		La_{57}														At_{85}	Rn_{86}
7.	Incomplete		87 to	Fr_{87}		AC_{89}														Uus_{117}	Uuo_{118}

Lanthanide series	Ce_{58}	INNER TRANSITION ELEMENTS	Lu_{71}
Actinide series	Th_{90}		Lr_{103}

Group 1 are Alkali metals (except hydrogen)
Group 2 Alkaline earth metals
Group 3 to 12 Transition elements
Group 17 Halogens
Group 18 Inert gases
Group IIIB 6th period Lanthanides
Group IIIB 7th period Actinides
— Inner Transition Elements

VARIATIONS IN PERIODIC PROPERTIES

Properties	Across the period (Left to right)	Down the group (Top to bottom)
No. of valence electrons	Increases	Remains same
Atomic size (radius)	Decreases	Increases
Ionisation energy	Increases	Decreases
Metallic character	Decreases	Increases
Non-metallic character	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases
Basic nature of oxides	Decreases	Increases
Melting point	Increases from group I to group IV and then decreases	Group I and II decreases Group III and IV decreases Group V to VII increases
Boiling point	Increases from group I to group IV and then decreases	Group I and II decreases Group III and IV decreases Group V to VII increases
Oxidising nature	increases	Decreases
Reducing nature	decreases	Increases

EXERCISE

1. An element Barium has atomic number 56. Look up its position in the Periodic Table and answer the following questions.

(a) Is it a metal or a non-metal ?

(b) Is it more or less reactive than calcium ?

(c) What is its valency ?

(d) What will be the formula of its phosphate ?

(e) Is it larger or smaller than caesium (Cs) in size ?

2. Choose the most appropriate answer from

[SO_2, SiO_2, Al_2O_3, CO, MgO, Na_2O]

(a) A covalent oxide of a metalloid.

(b) An oxide which when dissolved in water form acid.

(c) A basic oxide.

(d) An amphoteric oxide.

3. In group I of the Periodic Table, three elements X, Y and Z have ionic radii 1.33 Å, 0.95 Å and 0.60 Å respectively. Giving a reason, arrange them in the order of increasing atomic numbers in the group.

4. Arrange the following as per **instructions** given in the brackets.

(a) Mg, Cl, Na, S, Si (increasing order of atomic size)

(b) K, Na, Cl, S, Si (increasing non-metallic character)

(c) Na, K, Cl, S, Si (increasing ionisation potential)

(d) Cl, F, Br, I (increasing electron affinity)

(e) Cs, Na, Li, K, Rb (decreasing electronegativity)

5. Name

(a) An alkali metal in period 3 and halogen in period 2.

(b) The noble gas with 3 shells.

(c) The non-metals present in period 2 and metals in period 3.

(d) The element of period 3 with valency 4.

(e) The element in period 3 which does not form oxide.

(f) The element of lower nuclear charge out of Be and Mg.

(g) Which has higher E.A., Fluorine or Neon.

(h) Which has maximum metallic character Na, Li or K.

6. Chlorine in the Periodic Table is surrounded by the elements with atomic number 9, 16, 18 and 35.

(a) Which of these have Physical and Chemical properties resembling chlorine.

(b) Which is more electronegative than chlorine.

7. First ionisation enthalpy of two elements X and Y are 500 kJ mol^{-1} and 375 kJ mol^{-1} respectively. Comment about their relative position in a group as well as in a period.

8. Explain why are the following statements not correct :

(a) All groups contain metals and non metals.

(b) Atoms of elements in the same group have the same number of electron(s).

(c) Non-metallic character decreases across a period with increase in atomic number.

(d) Reactivity increases with atomic number in a group as well as in a period.

9. Arrange the following in order of increasing radii :

(a) Cl^-, Cl (b) Mg^{2+}, Mg, Mg^+ (c) N, O, P

10. Which element from the following has the highest ionisation energy ?

(a) P, Na, Cl (b) F, O, Ne (c) Ne, He, Ar

Explain your choice.

11. The electronegativities (according to Pauling) of the elements in Period 3 of the Periodic Table are as follows with the elements arranged in alphabetical order:

Al	Cl	Mg	Na	P	S	Si
1·5	3·0	1·2	0·9	2·1	2·5	1·8

Arrange the elements in the order in which they occur in the Periodic Table from left to right.

(The group 1 element first, followed by the group 2 element and so on, up to group 7).

12. Choose the word or phrase from the brackets which correctly completes each of the following statements :-

(a) The element below sodium in the same group would be expected to have a .... (lower/higher) electro-negativity than sodium and the element above chlorine would be expected to have a .... (lower/higher) ionization potential than chlorine.

(b) On moving from left to right in a given period, the number of shells .... (remains the same/ increases decreases).

(iii) On moving down a group, the number of valence electrons .... (remains the same/increases/ decreases).

(iv) Metals are good .... (oxidising agent/ reducing agent) because they are electron .... (acceptors/donors). **(2016)**

13. Parts (a) to (e) refer to changes in the properties of elements on moving from left to right across a period of the Periodic Table. For each property, choose the correct answer.

(a) The non-metallic character of the elements :

(i) decreases, (ii) increases,

(iii) remains the same, (iv) depends on the period

(b) The electronegativity :

(i) depends on the number of valence electrons,

(ii) remains the same,

(iii) decreases,

(iv) increases.

(c) The ionization potential :

(i) goes up and down (ii) decreases

(iii) increases (iv) remains the same

(d) The atomic size :

(i) decreases,

(ii) increases,

(iii) remains the same,

(iv) sometimes increases and sometimes decreases.

(e) The electron affinity of the elements in groups 1 to 7 :

(i) goes up and then down.

(ii) decreases and then increases,

(iii) increases,

(iv) decreases.

14. The elements of one short period of the Periodic Table are given below in order from left to right :

Li Be B C O F Ne

(a) To which period do these elements belong ?

(b) One element of this period is missing. Which is the missing element and where should it be placed ?

(c) Which one of the elements in this period shows the property of catenation ?

(d) Place the three elements fluorine, beryllium and nitrogen in the order of increasing electronegativity.

(e) Which one of the above elements belongs to the halogen series ?

2007

A group of elements in the Periodic Table are given below (boron is the first member of the group and Thallium is the last).

Boron, Aluminium, Gallium, Indium, Thallium

Answer the following questions in relation to the above group of elements :

(a) Which element has the most metallic character ?

(b) Which element would be expected to have the highest electronegativity ?

(c) If the electronic configuration of aluminium is 2, 8, 3, how many electrons are there in the outer shell of thallium ?

(d) The atomic number of boron is 5. Write the chemical formula of the compound formed when boron reacts with chlorine.

(e) Will the elements in the group to the right of this boron group be more metallic or less metallic in character ? Justify your answer.

2008

Select the correct answer from the choices A, B, C, D which are given. Write down only the letter corresponding to the correct answer.

With reference to the variation of properties in the Periodic Table, which of the following is generally true?

(a) Atomic size increases from left to right across a period.

(b) Ionization potential increases from left to right across a period.

(c) Electron affinity increases going down a group.

(d) Electro-negativity increases going down a group.

2009

(a) Among period 2 elements A, B, C and D, the one which has high electron affinity is

A. Lithium B. Carbon

C. Fluorine D. Neon

(b) Group No.'s

IA	IIA	IIIA	IVA	VA	VIA	VIIA	O
1	2	13	14	15	16	17	18
Li		D			O	J	Ne
A	Mg	E	Si		H	K	
B	C		F	G			L

Select from the table

(i) Which is most electronegative.

(ii) How many valence electrons are present is G.

(iii) Write the formula of the compound between B and H.

(iv) In the compound between F and J what type of bond will be formed.

(v) Draw the electron dot structure for the compound formed between C and K.

2010

(a) The number of electrons in the valence shell of a halogen is .... A – 1, B – 3, C – 5, D – 7.

(b) Electronegativity across the period .... [increases/ decreases].

(c) Non-metallic character down the group .... [increase/decreases].

(d) Atomic number of an element is 16. State

(i) to which period it belongs

(ii) the number of valence electron in the element

(iii) is the element metal or non-metal.

2011

(a) Give reasons – The oxidising power of elements increases from left to right along a period.

(b) Select the correct answer –

(i) Across a period, the ionization potential .... [increase, decreases, remains same]

(ii) Down the group, electron affinity .... [increases, decreases, remains same].

(c) Choose the correct answer from the choice given :

(i) In the periodic table alkali metals are placed in the group – A : 1, B : 11, C : 17, D : 18.

(ii) Which of the following properties do not match with elements of the halogen family –

A. They have seven electrons in their valence shell.

B. They are highly reactive chemically.

C. They are metallic in nature.

D. They are diatomic in their molecular form.

(d) State the group and period, of the element having three shells with three electron in valence shell.

2012

(a) Choose the correct answer from the option : An element in period 3 whose electrons affinity is zero.

A. Neon B. Sulphur

C. Sodium D. Agron

(b) Give reason :

(i) Ionisation potential of the element increases across a period.

(ii) Alkali metals are good reducing agents.

(c) There are three elements E, F, G with atomic numbers 19, 8 and 17 respectively –

Classifiy the above elements as metals and non-metals.

(d) Name : A metal present in period 3, group I of the periodic table.

2013

(a) Among the period 2 elements, the element which has high electron affinity is

A. Lithium B. Carbon

C. Chlorine D. Fluorine

(b)

Group No.	1-IA	2-IIA	13-IIIA	14-IVA	15-VA	16-VIA	17-VIIA	18-O
2nd period	Li		D			O	I	Ne
3rd period	A	Mg	E	Si		H	M	
4th period	R	T	I		Q	U		Y

In the above tabe H does not represent hydrogen. Some elements are given in their own symbol and position in the periodic table while others are shown with a letter. Answer the following questions.

(i) Identify the most electronegative element.

(ii) Identify the most reactive element of group I.

(iii) Identify the element from period 3 with least atomic size.

(iv) How many valence electrons are present in Q.

(v) Which element from group 2 would have the least ionization energy.

(vi) Identify the noble gas of the fourth period.

(vii) In the compound between A and H what type of bond is formed and give its molecular formula.

(c) Identify : The element which has the highest ionization potential.

2014

(a) Choose the correct answer from the choice given :

(i) Ionisation potential increases over a period from left to right because the :

A. Atomic radius and nuclear charge increases

B. Atomic radius and nuclear charge decreases

C. Atomic radius increases and nuclear charge decreases

D. Atomic radius decreases and nuclear charge increases

(ii) An element A belonging to period 3 and group II will have,

A. 3 shells and 2 valence electrons

B. 2 shells and 3 valence electrons

C. 3 shells and 3 valence electrons

D. 2 shells and 2 valence electrons

(b) Atomic number of an element Z is 16. Answer the following :

(i) State the period and group to which Z belongs.

(ii) Is Z a metal or a non-metal.

(c) State the formula of the compound between Z and hydrogen. What kind of compound is this formula M_2O when dissolved in water forms the corresponding hydroxide which is a good conductor of element in the above context answer the following :

(i) What kind of combination exists between M and O.

(ii) State the number of electrons in the outermost shell of M.

(iii) Name the group to which M belongs.

(d) Give one word or phrase for : The amount of energy *released* when an atom in the gaseous state accepts an electron to from an anion.

(e) Match the option A and B with the statements (i) and (ii) :

A. metal	(i) The metal that forms two types of ion
B. iron	(ii) An element with electronic configuration 2, 8, 8, 3