CHAPTER 01

PERIODIC TABLE AND PERIODIC PROPERTIES

The Periodic Table: The Symbol of Chemistry

One of the most important turning points in the history of science was the creation of the periodic table, which led to many important innovations. It is accurate to refer to the periodic table of elements as the "Symbol of Chemistry." In the modern periodic table, 118 elements are arranged in tabular form in the current periodic table based on their atomic number.

Historical Background

Many elements, such as Gold, Silver, Iron, Phosphorus, Sulfur, Zinc, and Arsenic have been known since the pre-historic era. However, the first classification was made in the 18th century. **Antoine Lavoisier** attempted to classify known elements as metals and nonmetals.

(I) Dobereiner's Triads

In 1829, a German chemist, J.W.Dobereiner grouped the elements into **triads** (a group of three) with similar properties. According to his law of Triads:

"The atomic weight of the middle element was roughly the average of the other two elements of a triad."

Some Dobereiner's Triads

Triad	Li	Na	K	Atomic mass of Na = $\frac{7+39}{2} = 23$
Atomic mass	7	23	39	
Triad	Ca	Sr	Ba	Atomic mass of Sr = $\frac{40+137}{2} = 88.5$
Atomic mass	40	88	137	
Triad	Cℓ	Br	I	Atomic mass of Br = $\frac{35.5+127}{2} = 81.25$
Atomic mass	35.5	80	127	

(ii) Newland's Law of Octaves

In 1864, an English chemist, John Newlands first time observed periodicity in the 62 known elements. According to his law of Octaves:

"If the elements are arranged in the increasing order of their atomic masses, every eighth (8th) element had some properties in common with the first one."

He classified the elements into groups so that every eight elements resembled the first element in properties.

Newland's Octaves

Li (7)	Be (9)	B (11)	C (12)	N (14)	O (16)	F (19)
Na (23)	Mg (24)	Aℓ (27)	Si (28)	P (31)	S (32)	Cℓ (35.5)
K (39)	Ca (40)	Ti (48)	Cr (52)	Mn (55)	Fe (56)	Ni & Co

(iii) Mendeleev's Periodic Table

A Russian chemist, Dmitri Mendeleev, is considered the father of the Periodic Table.

In 1869, Dmitri Mendeleev arranged 63 known elements in order of increasing atomic mass. He organized them into eight vertical columns (groups), placing elements with similar properties together. The success of his table lay in the intentional gaps he left for undiscovered elements. He predicted their atomic masses and properties, which proved accurate when these elements were practically found.

Mendeleev's Periodic Law

"If the elements are arranged in ascending order of their atomic masses, their chemical properties repeat in a periodic manner."

(iv) Lothar Meyer Atomic Volume Curves

In 1869, **Lothar Meyer** developed his famous curves by plotting a graph between the atomic weights and atomic volumes of elements. These curves also demonstrated periodicity.

(v) Moseley's Contribution

In 1913, Moseley determined the exact atomic numbers of known elements using X-ray emission. In this way he resolved flaws and discrepancies in Mendeleev's table by arranging the elements by atomic numbers instead of atomic masses.

This significant breakthrough led Moseley to modify the Periodic Law to state that the properties of elements are periodic functions of their atomic numbers.

THE MODERN PERIODIC TABLE

In modern periodic table, all the elements are arranged in ascending order of their atomic numbers.

Modern Periodic Law

"If the elements are arranged in ascending order of their atomic numbers, their chemical properties repeat in a periodic manner."

This law was introduced by Moseley in 1911. On the basis of modern periodic law, Bohr introduced modern periodic table.

Essential Features of Modern Periodic Table

- Presently, 118 elements are grouped in the table in ascending order of their respective atomic numbers.
- There are seven horizontal rows called **periods** and eighteen vertical columns called **groups**. (In older versions of the table, there were 8 vertical groups were divided into two types of groups: Eight A Groups and Ten B-Groups.
- In the periodic table, elements within the same group exhibit similar chemical properties because they have the same number of valence electrons. However, they show a gradual change in physical properties from top to bottom in a group.
- Elements in a period show a gradual change in properties moving from left to right in periods.

Significance of Modern Periodic Table

The classification of elements in the modern periodic table helps in the easier understanding of their properties.

Classification of Elements in Modern Periodic Table

1. Groups

"Elements with similar properties are placed in vertical columns called groups."

- There are eighteen (18) groups.
  - ➢ In older versions of the table, there were 8 vertical groups which were divided into two types of groups.
  - → **Sub-group-A:** Containing representative or normal elements.
  - → **Sub-group-B:** Contain less typical elements, called transition elements and are arranged in the center of periodic table
- Elements of same group have similar chemical properties due to same valence shell electronic configuration.
- Usually number of valence electrons is same as that of group number.

2. Periods

"The horizontal rows of elements in periodic table are called periods."

- There are 7 periods in the periodic table numbered by Arabic numerals 1 to 7.
- The period number indicates the principal quantum number (n), representing the number of electron shells surrounding the nucleus.

Periodic Table of the Elements

1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8 VIII 8	9 VIII 8	10 VIII 8	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 VIIIA 8A
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.294
55 Cs Cesium 132.905	56 Ba Barium 137.328	57 La Lanthanum 138.905	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89 Ac Actinium 227.028	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [280]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

Lanthanide Series	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.243	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967
Actinide Series	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]

3. Metals, Non-Metals and Metalloids

On the basis of metallic character, elements can be broadly classified as:

(i) Metals

- Metals are elements which tend to lose electrons to form positive ions.
- Generally, elements on left-hand side, in the center and at the bottom of the periodic table are metals.

(ii) Non-metals

- Non-metals are elements which tend to gain electrons to form negative ions.
- Elements in the upper right corner of the periodic table are non-metals.
- In the periodic table elements of groups IVA to VIIIA, at the top right-hand corner above the stepped line, are non-metals.

(iii) Metalloids

- The metalloids separate the metals and nonmetals on a periodic table. The metalloids exhibit some properties of metals and some of non-metals.
- Mostly periodic tables have a "stair-step line" on the table identifying the metalloids. The line begins at boron (B) and extends down to polonium (Po) including Si, Ge, As, Sb and Te.

Metallic Elements	Nonmetallic Elements
• Distinguishing luster (shine)	• Non-lustrous, various colors
• Malleable and ductile (flexible) as solids	• Brittle, hard or soft
• Conduct heat and electricity	• Poor conductors
• Metallic oxides are basic, ionic	• Nonmetallic oxides are acidic, covalent
• Cations in aqueous solution	• Anions, oxyanions in aqueous solution

4. Blocks

Based on the subshells containing their valence electrons, elements in the periodic table can be classified into following four blocks;

(i) s-block

Elements of group 1 (alkali metals) and group 2 (alkaline earth metals) including Helium (He) have valence electrons in the 's' subshell so they are s-block elements. e.g. $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

(ii) p-block

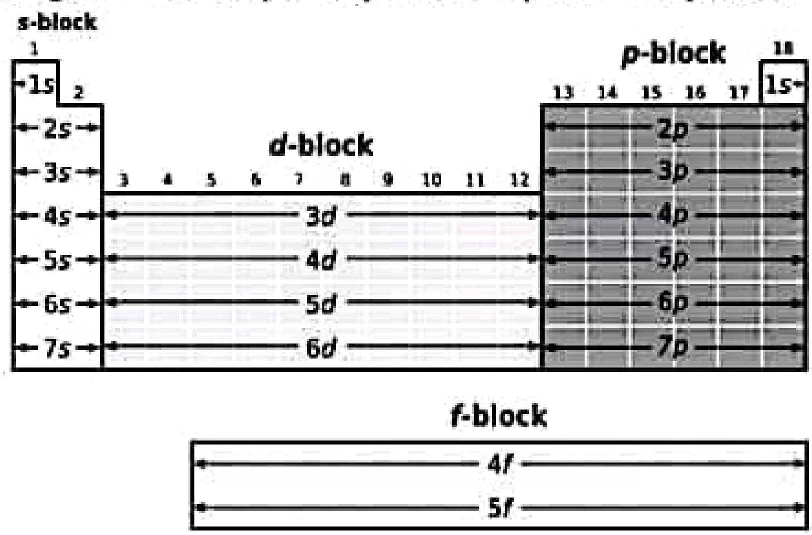
The elements of group 13 to group 18 except Helium (He) are known as p-block elements as their valence electrons are present in p-subshell. e.g. $_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

(iii) d-block

In transition elements, d-block constitutes four series of elements. In these elements the valence electrons are in the d-subshell. e.g. $_{26}Fe = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^6$

(iv) f-block

In Lanthanides and Actinides valence electrons are present in f-orbital hence these elements are called f-block elements. e.g. $_{58}Ce = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^2$



5. Families in Periodic Table

A set of elements sharing common properties is called an element family. On the basis of common properties, elements in periodic table are classified into following five famous families of elements;

(i) Alkali Metals

Elements in the group 1 of the periodic table are known as alkali metals because they produce alkalis when they react with water.

- They include the elements lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs) and francium (Fr).

Characteristics

- Alkali metals have one valence electron, making them monovalent.
- They have low densities, relatively low melting points, and low ionization energies.
- These are the most reactive metals.
- These are s-block elements.

(ii) Alkaline Earth Metals

Group 2 elements are metals primarily found in the earth and form alkalis; hence they are referred to as alkaline earth metals.

- They include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba) and radium (Ra).

Characteristics

- These elements have two electrons in their valence shell, making them divalent.
- They are metallic solids that are harder and denser than alkali metals.
- Easily oxidized, with high thermal and electrical conductivities.
- These are s-block elements.

(iii) Transition Elements

The transition metals make up the largest family of elements in the middle of periodic table. They are further classified into two groups;

(a) d-block elements (outer transition elements).

These include four series, each containing 10 elements. They are found in the center of the periodic table.

(b) f-block elements (inner transition elements).

These include the lanthanides and actinides, which are placed in two separate rows below the main body of the periodic table.

Characteristics

- They exhibit high thermal and electrical conductivities, high melting points and high density.
- They show variable oxidation states.
- They mostly form coloured compounds.

(iv) Chalcogens

The group 16 elements are called Chalcogens because most ores of copper (Greek chalkos) are oxides or sulfides.

In this group, oxygen (O) & sulphur (S) are non-metals, selenium (Se), tellurium (Te), polonium (Po) are metalloids and Livermorium (Lv).

Oxygen and sulphur are nonmetals, while selenium and tellurium are metalloids, and polonium is a metal

(v) Halogens

Elements in group 17, known as halogens. The term "halogen" means "salt-former" because these elements easily react with alkali metals and alkaline earth metals to form stable halide salts.

- They include fluorine (F), chlorine (Cℓ), bromine (Br), iodine (I) and astatine (At).

Characteristics

- Halogens are nonmetals.
- They are highly reactive nonmetals with high electron affinities.
- They can easily accept one electron to complete their outermost shell.

(VI) Noble Gases

Elements in group 18, known as noble gases, inert gases and zero group. They are present at the extreme right of the periodic table.

- They include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn).

Characteristics

- Due to their stable electron configuration (complete outermost shell), they are almost entirely unreactive under normal conditions and rarely form compounds with other elements.
- These elements are monoatomic in nature.

Exception: Although, noble gases are unreactive, however they have some compound. An example is compounds of xenon such as xenon hexafluoroplatinate ($XePtF_6$), the word inert gases was changed to noble gases.

Quick Check 1.1

a) Why are the elements in Groups 1 and 2 known as s-block elements?

Ans: Elements of group 1 (alkali metals) and group 2 (alkaline earth metals) have valence electrons in the 's' subshell so they are s-block elements. e.g. $_{11}Na = 1s^2, 2s^2, 2p^6, 3s^1$

b) Name the elements in the chalcogen family. Give their two characteristics.

Ans: Chalcogen family (Group 16) includes oxygen (O), sulphur (S), selenium (Se), tellurium (Te), polonium (Po) and Livermorium (Lv).

Characteristics of chalcogen:

(i) All the elements of group 16 show the property of allotropy.

(ii) They have high reactivity with metals, forming oxides and sulphides.

Periodic Arrangement of Elements (Arrangement of elements in periodic table)

In periodic table, the elements are arranged in rows (periods) and columns (groups) based on their atomic numbers, chemical properties and electronic configuration. This arrangement offers valuable insight into their physical properties, such as their physical state and atomic radii, as well as their electronic structure and chemical reactivity.

Steps to right electronic configuration of an elements

(i) Electronic shell

The period number indicates the principal quantum number (n), representing the number of electron shells surrounding the nucleus. e.g, an element X in the 3rd period has three electron shells, with its valence electrons located in the 3rd shell.

(ii) Sub-shell

The specific subshell where the valence electrons are found, depends on the element's block (azimuthal quantum number). If an element X in the 3rd period is in the s-block, its valence electrons are in the 3s subshell.

(iii) Valence electrons

The group number indicates the number of valence electrons; e.g., an element X in the 3rd period and group 2 has two valence electrons in its outermost shell. Thus, the element Z in the 3rd period and group 2 (s-block) has two valence electrons in the 3s subshell, which means that X would be magnesium (Mg).

How Periods and Groups Reflect Electronic Configuration?

Suppose 'X' belongs to group 13 and period 3.

Since the element X belongs to group 13 of periodic table so it has 3 valence electrons; and it is found in period 3 so it has three shells around its nucleus. It means that the 3 valence electrons are in the 3rd shell. The configuration will be:

$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^1$
↓	↓	↓
1st shell	2nd shell	3rd shell

Periodicity Of Properties

Modern Periodic Law

"The physical and chemical properties of elements are periodic functions of their atomic numbers."

Periodicity of Properties

"The repetition of properties of elements in periodic table after regular intervals is called periodicity."

Explanation

- Modern periodic law is the "cornerstone" of the periodic table, indicating that elements with similar properties appear at certain intervals.
- When elements are arranged in order of increasing atomic number, those in the same group like lithium, sodium, potassium, and cesium show similar physical and chemical properties because they share the same valence electron configuration.
- However, due to the gradual increase in the number of protons in nucleus and the addition of new electron shells, the physical and chemical properties of elements vary systematically within a group and a period.

Quick Check 1.2

a) **X-belongs to group 14 and period 2**
(i) Write electronic configuration of the element X.
(ii) Identify block of the element. Identify this element from periodic table.

Ans: Since the element X belongs to group 14 of periodic table so it has 4 valence electrons; and it is found in period 2 so it has two shells around its nucleus. It means that the 4 valence electrons are in the 2nd shell. The configuration will be:
(i) X = $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$
(ii) Since valence electrons are in p-subshell, therefore 'X' belongs to p-block.
Periodic table shows that 'X' is silicon (Si).

b) **Identify an element that is in Period 4 and Group 17?**

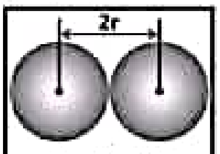
Ans: Since the element belongs to group 17 of periodic table so it has 7 valence electrons; and it is found in period 4 so it has four shells around its nucleus. It means that the 7 valence electrons are in the 4th shell. The configuration will be:
$_{35}Br = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^5$

Exams Short Answer Questions

Q. Define Dobereiner's law of triads. Also give two examples.
Q. What is Newland's law of Octaves? Q. Define Mendeleev's and modern periodic law.
Q. Write two defects of Mendeleev's periodic table.
Q. Define periodic table. How many groups and periods are present in it?
Q. d and f-Block elements are called transition elements.
Q. Lanthanide contraction controls the atomic sizes of elements of 6th and 7th period.
Q. Define Group and period. How many elements are there in period number 1?
Q. Write the names of families in periodic table.
Q. Give essential features of period four (4) in modern periodic table?
Q. How the classification of elements in different blocks helps in understanding their chemistry?

ATOMIC RADIUS

"Half of the distance between the centers of two identical atoms bonded together is called atomic radius."



Atomic radius

The atomic radius can vary depending on:
(i) Type of bond:
(van der Waals radius > metallic radius > anionic radius > covalent radius > cationic radius.)
(ii) State of the atom:
The atomic radius is typically measured in picometers (pm) or Angstroms (Å).

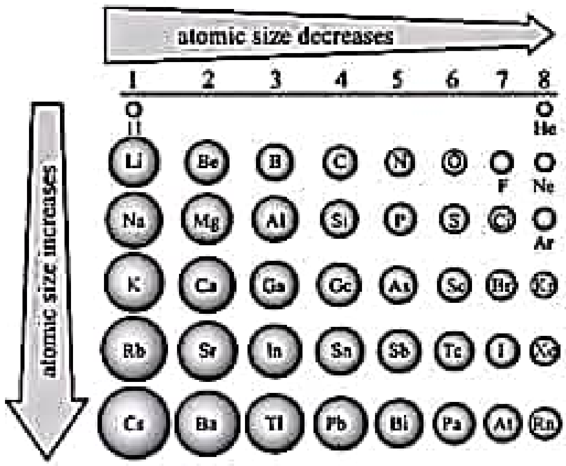
Factors affecting atomic radius
(i) Atomic number (ii) Effective nuclear charge (iii) Shielding effect of inner electrons

Periodic Trend

(a) Across the Period:
Atomic radius decreases across a period (from left to right) in the periodic table due to increasing nuclear charge, which pulls the electron cloud closer.

MCQ: Which of the following element has smaller size?
(a) Na (b) K (c) Al (d) Li
Hint: Atomic radius of Al = 143 pm
Atomic radius of Li = 152 pm

(b) Down the Group:
Atomic radius increases down a group (from top to bottom) because additional electron shells are added, so more shielding makes the atom larger despite the increase in nuclear charge (which is outweighed).



Variation in atomic radius across periods and down the groups

Ionic Radius

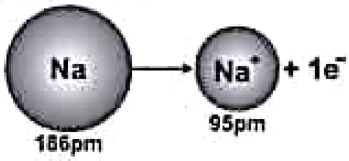
"The distance from the nucleus of an ion to the outermost electron shell is called ionic radius."

- The ionic radius is a measure of the size of an ion in a crystal lattice.
- It is measured in picometers (pm) or angstroms (Å).

Size of cation (Cationic Radius)

When a neutral atom loses one or more electrons, it becomes a positive ion (cation).

The size of cation is smaller than its parent atom due to;



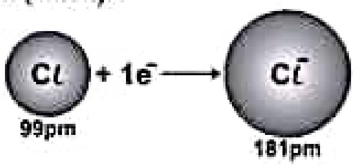
(i) Removal of one or more electrons from a neutral atom usually results in the loss of the outermost shell.

(ii) Removal of electrons causes an imbalance in proton-electron ratio. Due to the greater attraction of the nuclear charge, the remaining electrons of the ion are drawn closer to the nucleus. Thus, a positive ion is always smaller than the neutral atom from which it is derived.

Example: Na (186pm) $\longrightarrow$ Na^+ (95pm) + $1e^-$

Size of Anion (Anionic Radius)

When a neutral atom gains one or more electrons, it becomes a negative ion (anion). The size of anion is greater than its parent atom. This is because the addition of electrons increases electronic repulsion, as a result the nuclear pull on electrons decreases and the electron cloud expands.



Example: Cl (99pm) + $1e^-$ $\longrightarrow$ Cl^- (181pm)

Periodic Trend

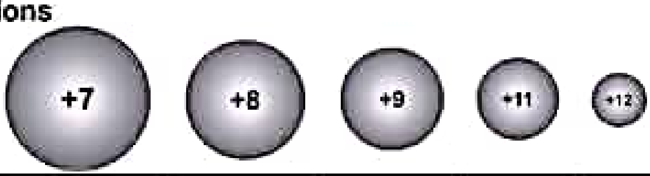
(a) Across the Period:

When we move across a period from left to right;

(i) The cationic radius decreases due to the increasing nuclear charge which pulls the electrons closer.

(ii) The anionic radius also decreases across a period because the increasing nuclear charge also pulls the electrons closer to the nucleus.

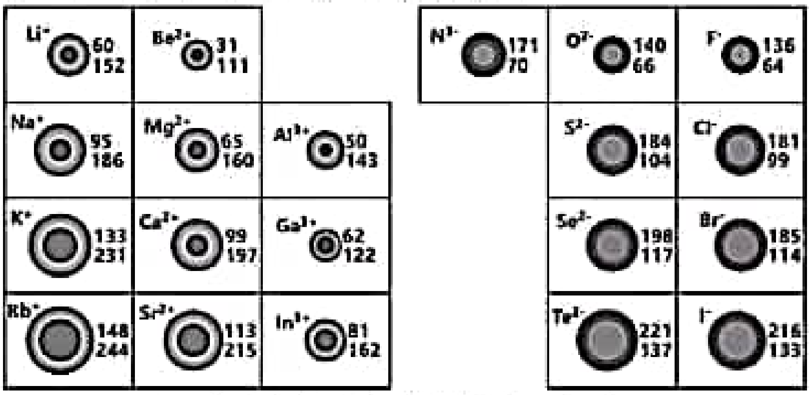
Radii of iso-electronic ions



Ion	N^{3-}	O^{2-}	F^-	Na^+	Mg^{2+}
Number of electrons	10	10	10	10	10
Charge on the nucleus	+7	+8	+9	+11	+12
Radius (pm)	171	140	136	95	60

(b) Down the Group:

Both cations and anions increase in size as we move down a group. This is because the principal quantum number (n) increases, leading to an increase in the number of electron shells. Consequently, the distance between the nucleus and the outermost electrons becomes larger, outweighing the effect of increased nuclear charge. The additional electron shells make the ions larger.



Variation in Ionic Radius (picometers)

Quick Check 1.3

a) Which factors affect atomic and ionic radii?

Ans: Factors affecting atomic radius and ionic radii

(i) Atomic number (ii) Effective nuclear charge (iii) Shielding effect of inner electrons

b) Using your knowledge of Period 3 elements, predict and explain the relative sizes of:

(i) the atomic radii of lithium and fluorine

(ii) a lithium atom and its ion, Li^+

(iii) an oxygen atom and its ion, O^{2-}

(iv) a nitride ion, N^{3-}, and a fluoride ion, F^-.

Ans: (i) Lithium has a larger atomic radius than fluorine. It is because, across a period, atomic radius decreases because more protons pull the electrons closer without adding new shells.

(ii) The Li^+ ion is much smaller than a lithium atom. When lithium becomes Li^+, it loses an electron shell. Fewer electrons and more pull from nucleus make the ion smaller.

(iii) The O^{2-} ion is larger than the oxygen atom. The addition of two extra electrons increases electron–electron repulsion and there is less effective nuclear attraction per electron, so the electron cloud spreads out more, increasing the radius.

(iv) N^{3-} is larger than F^-. Both are isoelectronic (same number of electrons), but N^{3-} has only 7 protons pulling on 10 electrons, while F^- has 9 protons pulling on the same number. So, the nuclear attraction is weaker in N^{3-}, making it larger.

Ex. Q6 Describe the factors affecting and periodic trends of ionization energy.

IONIZATION ENERGY

First Ionization Energy (ΔH_{i1})

"The energy required to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions is called the first ionization energy (ΔH_{i1})."

Examples:

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^- \qquad \Delta H_{i1} = 494\ kJ\ mol^{-1}$$

$$Ca_{(g)} \longrightarrow Ca^+_{(g)} + e^- \qquad \Delta H_{i1} = 590\ kJ\ mol^{-1}$$

- Ionization energy is actually the qualitative measure of the stability of an isolated atom. Greater the ionization energy, more stable will be the atom.
- It is measured in $kJ\ mol^{-1}$, $kcal\ mol^{-1}$ or eV. $\boxed{1eV = 96.485\ kJ\ mol^{-1}}$
- Noble gases have the highest values of ionization energy due to complete outermost shell in them, the removal of electron is extremely difficult.
- Alkali metals have lowest values of ionization energy.

Higher Ionization Energies

An element can have several ionization energies; the exact number corresponds to its atomic number.

Second Ionization Energy (ΔH_{i2})

If a second electron is removed from each ion in a mole of gaseous 1+ ions, it is called 2nd ionization energy, ΔH_{i2}. e.g. $Ca^{+}_{(g)} \longrightarrow Ca^{2+}_{(g)} + e^{-}$ $\quad \Delta H_{i2} = 1150 kJ\ mol^{-1}$

Third Ionization Energy (ΔH_{i3})

If a third electron is removed from each ion in a mole of gaseous 2+ ions, it is called 3rd ionization energy, ΔH_{i3}. e.g. $Ca^{2+}_{(g)} \longrightarrow Ca^{3+}_{(g)} + e^{-}$ $\quad \Delta H_{i3} = 4940 kJ\ mol^{-1}$

Factors Affecting the Ionization Energies

The magnitude of the ionization energy of an element depends upon the following factors:

(i) Nuclear Charge

$\boxed{\text{I.E} \propto \text{Nuclear charge}}$

The greater the effective nuclear charge, the stronger the electrostatic force of attraction between the nucleus and the electrons. As a result, it becomes more difficult to remove an electron from the atom. Therefore, ionization energy increases with an increase in effective nuclear charge.

(ii) Size of the atom or ion

$\boxed{\text{I.E} \propto \dfrac{1}{\text{atomic radius}}}$

In bigger atoms, force of attraction between the nucleus and the outermost electrons is weaker. Therefore, the ionization energy decreases as the size of the atom increases and vice-versa.

(iii) Electronic arrangement

It is observed half-filled and completely-filled orbitals are found to be more stable. Therefore, the ionization energy is higher when an electron is to be removed from a fully filled or half-filled-shells.

(a) **Noble gases** have highest ionization energies in their respective periods. It is due to highly stable fully-filled shells ($ns^2\ np^6$).

(b) **Oxygen** has lower ionization energy than **nitrogen**. The electronic configuration of oxygen and nitrogen are:

$${}_7N = 1s^2\ 2s^2\ 2p_x^1\ 2py^1\ 2p_z^1 \qquad \Delta H_{i1} = 1403\ kJ\ mol^{-1}$$
$${}_8O = 1s^2\ 2s^2\ 2p_x^2\ 2py^1\ 2p_z^1 \qquad \Delta H_{i1} = 1365\ kJ\ mol^{-1}$$

Although, nitrogen has one unit less positive charge in its nucleus than oxygen, but due to the extra-stability of the half-filled sub-shell of nitrogen it is difficult to remove an electron from N atom.

(iv) Shielding Effect

$\boxed{\text{I.E} \propto \dfrac{1}{\text{Shielding effect}}}$

Greater the shielding, easier it is to remove the valence electrons from an atom. Larger the number of inner electrons, greater is the screening effect, therefore, lower is the ionization energy.

(v) Spin–Pair Repulsion

When electrons are spin-paired in the same orbital, the repulsion between them can lead to a slightly lower ionization energy compared to removing an unpaired electron. This is because the paired electrons experience increased repulsion, making it slightly easier to remove one of the paired electrons.

Example: Chromium (Cr) has two spin-paired electrons in its 4s orbital. The ionization energy to remove one of these paired electrons is relatively lower due to the increased repulsion between the paired electrons. In contrast, Manganese (Mn) has three unpaired electrons in its 4s orbital. Removing one of these unpaired electrons requires more energy due to the absence of spin-pairing repulsion.

${}_{24}Cr = [Ar]\ 4d_5 4s_2$ (less stable)

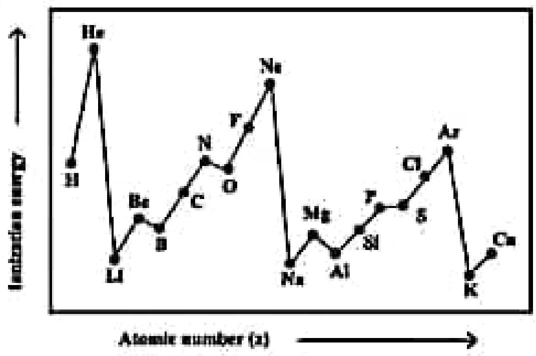
${}_{25}Mn = [Ar]\ 4d_5 4s_1$ (more stable)

Variation of Ionization Energy in the Periodic Table

Across the Period: In moving from left to right across a period, number of shells remains unchanged while the effective nuclear charge increases. Therefore, the ionization energy increases from left to right.

Down the Group: Going down in a group, ionization energy decreases from top to bottom due to increase in the atomic size and shielding effect. e.g. In Group I, the ionization energies decrease in the following order:

Li > Na > K > Rb > Cs.



Variation in Ionization Energies Across periods

Quick Check 1.4

a) Explain with reasoning following facts about ionization energy:
i. 1st ionization energy of Boron is lesser than Beryllium.
ii. 1st ionization energy of Aluminum is lower than Magnesium.

Ans: (i) Beryllium (Be) has a higher 1st ionization energy than boron (B) due to its stable, completely filled 2s orbital. On the other hand, in boron, the valence electron is in the 2p orbital, which is higher in energy and more shielded by inner electrons. This makes it easier to remove the 2p electron from boron, resulting in a lower ionization energy compared to Be.

(i) Magnesium (Mg) has a higher 1st ionization energy than aluminum (Al) due to its stable, completely filled 3s orbital. On the other hand, in aluminum, the valence electron is in the 3p orbital, which is higher in energy and more shielded by inner electrons. This makes it easier to remove the 3p electron from aluminum, resulting in a lower ionization energy compared to Mg.

b) What trend is observed in ionization energy as you go down group 3? Give reason

Ans: Down the group, increase in nuclear charge does not affect I.E considerably. These are atomic radii and shielding effect which increase down the group and cause decrease in attraction between nucleus and valence electrons. Therefore, I.E decreases as we go down group 3.

Ex.Q5 Describe the factors affecting and periodic trends of electron affinity.

ELECTRON AFFINITY

First Electron Affinity (ΔH°_{ea1})

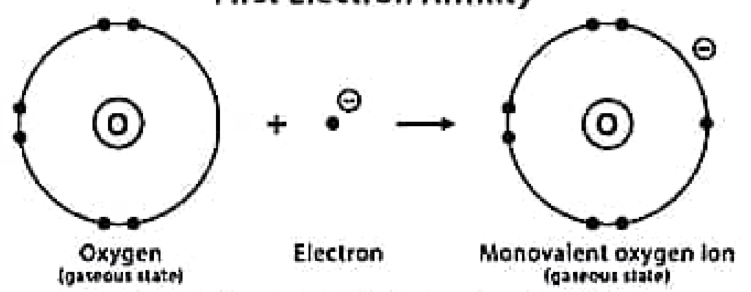
"The enthalpy change involved when 1 mole of electrons is added to 1 mole of gaseous atoms to form 1 mole of gaseous uni-negative ions under standard conditions is called first electron affinity, (ΔH^{o}_{ea1})."

Example#1: $Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)}$ $\Delta H^{\circ}_{ea1} = -348.8\ kJ\ mol^{-1}$

This is amount of energy released when 6.02×10^{23} atoms of chlorine in the gaseous state are converted into $Cl^-_{(g)}$ ions.

Example#2: $O_{(g)} + e^- \longrightarrow O^-_{(g)}$ $\Delta H^{\circ}_{ea1} = -142\ kJ\ mol^{-1}$

First Electron Affinity



Energy Released, $\Delta H^{\circ}_{ea1} = -141\ kJ\ mol^{-1}$

- Since, energy is released, so first electron affinity carries negative sign.
- Electron affinity is the measure of the attraction of the nucleus of an atom for the extra electron.

Units: Its units are $kJ\ mol^{-1}$, $k\ cal\ mol^{-1}$, electron volt.

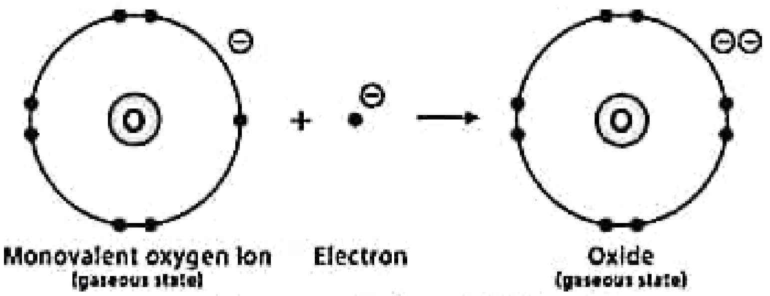
Second Electron Affinity (ΔH°_{ea2})

"The amount of energy required to add electrons to 1 mole of uni-negative gaseous ions to form 1 mole of gaseous 2– ions under standard conditions is called second electron affinity, ΔH^{o}_{ea2}" e.g.,

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \qquad \Delta H^{\circ}_{ea2} = +844\ kJ\ mol^{-1}$$

844 kJ mol^{-1} of energy is absorbed on adding second electron to a uni-negative (O^-) ion. Since, energy is absorbed, so second electron affinity carries positive sign.

Second Electron Affinity



Energy Absorbed, ΔH^o_{ea2} = +844 kJ mol^{-1}

Factors affecting Electron Affinity

Important factors affecting the magnitude of electron affinity values of elements are as follows:

(i) Size of atom

The smaller the size of an atom, the greater its electron affinity, and vice versa. This is because in smaller atoms, the nucleus exerts a stronger attraction on the incoming electron.

$$\boxed{E.A \propto \frac{1}{\text{atomic radius}}}$$

(ii) Nuclear Charge

Greater the magnitude of nuclear charge of an element stronger is the attraction of its nucleus for the incoming electron. Thus, with the increase in the magnitude of nuclear charge, electron affinity also increases.

$$\boxed{E.A \propto \text{Nuclear charge}}$$

(iii) Shielding Effect of inner electrons

Greater the shielding effect, lesser will be the electron affinity.

$$\boxed{E.A \propto \frac{1}{\text{Shielding effect}}}$$

(iv) Electronic Configuration of Atom

The electron affinity is low when the electron is added to a half-filled sub-shell than that for partially filled one. e.g. Electron affinity values of 'N' and 'P' group-15 (V-A), atoms are very low. This is because of the presence of half-filled 'np' orbitals in their valence shell (N = $2s^2\ 2p^3$, P = $3s^2\ 3p^3$). These half-filled p-subshells, being very stable, have very little tendency to accept any extra electron to be added to them.

Variation in the Periodic Table

Across the Period

Generally, electron affinities become more negative as we move from left to right in a period. This is firstly due to increase in the nuclear charge, which attracts additional electrons more strongly and secondly due to decreasing atomic radius.

Down the Group

As the atomic size increases down the group, the larger electron cloud causes the incoming electron to experience less attraction from the nucleus. Consequently, electron affinity generally decreases down the group. This trend is observed in the halogens (At < I < Br < F < Cl).

Electron Affinities (kJ mol^{-1}) for Group 1 and Group 17

Groups 17	E.A (kJ mol^{-1})	Group 1	E.A (kJ mol^{-1})
Fluorine	-328.0	Lithium	-60.0
Chlorine	-349.0	Sodium	-53.0
Bromine	-324.0	Potassium	-48.0
Iodine	-295.0	Rubidium	-47.0
Astatine	-270.1	Cesium	-46.0

Quick Check 1.5

Explain with reasoning following facts about electron affinity:

a) 1st electron affinity of Oxygen is –141 kJ mol⁻¹ but 2nd electron affinity is +844.0 kJ mol⁻¹.

Ans: Usually the electronegative elements release energy when first electron is added into them. But when a second electron is added in a uni-negative ion, the incoming electron is repelled by the already present anion. In order to overcome this repulsion, energy is absorbed during the process. Thus, the formation of a di-negative ion is an endothermic process and its E.A. is shown with positive sign.

$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \qquad \Delta H^o_{ea1} = -141\ kJ\ mol^{-1}$$

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \qquad \Delta H^o_{ea2} = +844\ kJ\ mol^{-1}$$

b) Which of nitrogen and phosphorus has the higher electron affinity? Justify with reason.

Ans: Phosphorus has a higher electron affinity than nitrogen because its larger size makes it easier to accept an extra electron. In nitrogen, the atom is smaller and its electrons are more tightly packed, so adding another electron causes repulsion. This makes nitrogen less willing to gain an extra electron.

$$N_{(g)} + e^- \longrightarrow N^-_{(g)} \qquad \Delta H^o_{ea1} = +7\ kJ\ mol^{-1}$$

$$P_{(g)} + e^- \longrightarrow P^-_{(g)} \qquad \Delta H^o_{ea1} = -72\ kJ\ mol^{-1}$$

c) F has lower electron affinity than Cl although its size is smaller. Explain why?

Ans: Actually, fluorine has very small size (72 pm) and seven electrons in 2s and 2p subshells have thick electronic cloud. This thick electronic cloud repels the incoming electron. Thus, fluorine has electron affinity less than that of chlorine.

$$F_{(g)} + 1e^- \longrightarrow F^{-1}_{(g)} \qquad \Delta H^o_{ea1} = -328\ kJ\ mol^{-1}$$

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \qquad \Delta H^o_{ea1} = -348.8\ kJ\ mol^{-1}$$

d) Why noble gases (group-18) have positive 1st electron affinities? Explain in terms of electronic configuration.

Ans: Noble gases have stable $ns^2\ np^6$ configuration and hence the atoms of these gases, do not accept any extra electron. This is evident from their positive 1st electron affinities.

ELECTRONEGATIVITY

"The power of an atom to attract shared pair of electrons toward itself in a molecule is called electronegativity."

- It has no unit.
- Higher electronegativity values signify a stronger attraction for electrons compared to lower values.

Measurement of electronegativity/Pauling scale

Linus Pauling, an American chemist, developed a scale of dimensionless electronegativity values for elements. According to Pauling scale;

(i) Alkali metals have minimum electronegativity (0.8).
(ii) Halogens have maximum values (F=4.0).
(iii) Normally, metals being on the left side of the periodic table, possess lower electronegativity values than those of non-metals.
(iv) Metals are electropositive and non-metals are electronegative, relatively.

Linus Pauling is the only person to have received two unshared Nobel Prizes, one for chemistry in 1954 for his work on the nature of chemical bond and one for peace in 1962 for his opposition to weapons of mass destruction.

Factors Affecting Electronegativity

(i) Atomic size

A larger atomic size will result in a lower value of electronegativity. This is because electrons being far away from the nucleus will experience a weaker force of attraction.

(ii) Effective nuclear Charge

A higher value of the effective nuclear charge will result in a greater value of electronegativity, because an increase in nuclear charge causes greater attraction to the bonded electrons. This is why the electronegativity in a period increases from left to right.

Variation in Periodic Table

Across the period

When we move from left to right along the period, the electronegativity increases. This is due to;

(i) Increase in nuclear charge (ii) Decrease in atomic size

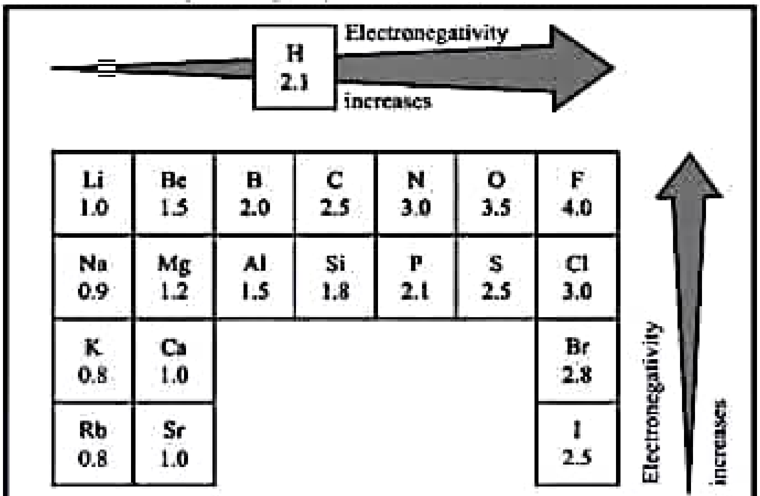
The electronegativity of Li in period 2 is 1.0 and F has a value of 4.0.

Down the Group

Electronegativity decreases down the group due to:

(i) Increase in atomic size due to addition of shells (ii) Increase in shielding effect

e.g., the electronegativities of halogens in group 17 are in the order: F > Cl > Br > I



Variation of electronegativity in groups and periods

METALLIC AND NON-METALLIC CHARACTER

Metallic Character

"The tendency of an element to lose electron and form a positive ion is called metallic character."

- Elements on the left side of the periodic table have a greater tendency to lose their outermost electrons to achieve noble gas configuration. Therefore, these elements are metals that form positive ions.
- The metallic character of an element largely depends on its valence shell electronic configuration.
- The increase in metallic character (ease of losing electron) makes the element more reactive. e.g. Among alkali metals, cesium is far more reactive and electropositive than sodium or lithium.

Non-Metallic Character

"The tendency of an element to gain electron and form a negative ion is called non-metallic character."

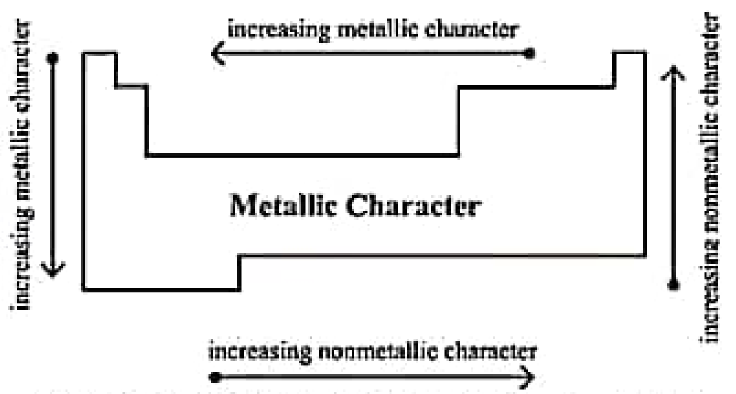
Elements on the right side of the periodic table have a greater tendency to gain their outermost electrons to achieve noble gas configuration. Therefore, these elements are non-metals that form negative ions.

Periodic Trend in Metallic Character

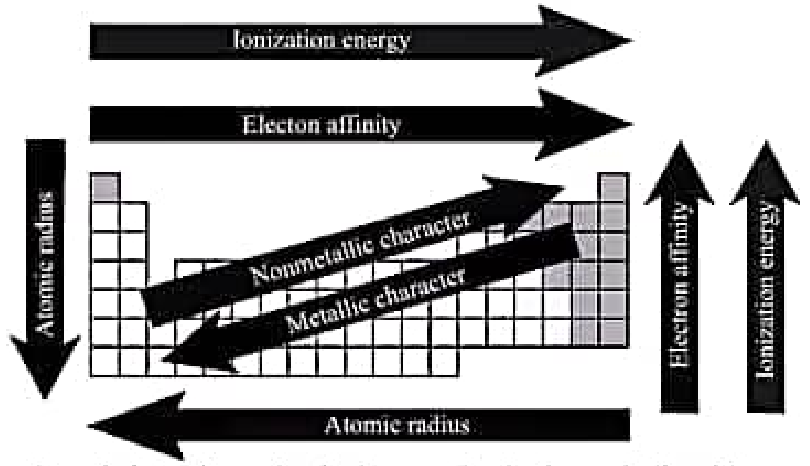
Across the period:

Metallic character decreases across a period from left to right due to an increase in nuclear charge and a decrease in atomic radius. As a result, the attraction between the nucleus and the valence electrons increases, making it more difficult for the atom to lose electrons.

Down the Group: Metallic character increases as we move down a group in the periodic table due to the increase in atomic size and the shielding effect. These factors reduce the nuclear attraction on the valence electrons, making it easier for the atom to lose electrons.



General trends for the metallic character of elements



Trends in various physical properties in the periodic table

Quick Check 1.6

a) Illustrate how does the metallic character vary in group 14?

Ans: Metallic character increases down Group 14 due to the increase in atomic size and the shielding effect. As a result, carbon (C) is a nonmetal, silicon (Si) and germanium (Ge) are metalloids, while tin (Sn) and lead (Pb) are metals.

b) Identify semi metals in groups 14, 15 and 16. Why they are semi metals?

Ans: In groups 14, 15, and 16, the semi-metals are:
(i) Group 14: silicon (Si) and germanium (Ge)
(ii) Group 15: arsenic (As) and antimony (Sb)
(iii) Group 16: tellurium (Te) and polonium (Po)
Semi-metals are elements that exhibit properties intermediate between those of metals and non-metals.

Exams Short Answer Questions

- Q. Why the atomic radius decreases from left to the right within a period and increases from top to bottom down the group? **Q.** Why the size of cation is always smaller than its parent atom? Give example also.
- Q. Why the ionic radii of negative ions are larger than the size of their parent atoms? / Why the size of anion is greater than its parent atom? Give example. / Q. Define shielding effect.
- Q. Define ionization energy. Give its trend across a period and down a group. / Define ionization energy with an example.
- Q. What is the role of shielding effect on ionization energy?
- Q. Why does the ionization energy decrease down the group and increases along a period?
- Q. Ionization energy of Al^{+3} is greater than Mg^{+2}. Give the reason.
- Q. Why second. I.E value is always greater than first I.E value?
- Q. Give variations of electron affinity down the group and across the period.
- Q. Why the second value of electron affinity of an element is usually shown with a positive sign? / Why 1st electrons affinity is negative and 2nd is positive sign?

Ex.Q.3: Write equations for the reactions of Na and Mg with oxygen, chlorine, and water. Compare the reactivity of both elements with these in terms of metallic character.

REACTIONS OF SODIUM AND MAGNESIUM

(i) Reactions with Water

Reaction of Sodium with Water:

Sodium is more reactive than magnesium towards water. Na reacts vigorously with water to form sodium hydroxide and hydrogen. The reaction is highly exothermic and sometimes hydrogen gas catches fire.

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

Reaction of Magnesium with Water:

Mg reacts with water more slowly than Na and forms magnesium hydroxide and hydrogen. However, magnesium reacts with steam more vigorously to make magnesium oxide and hydrogen gas.

$$Mg_{(s)} + 2H_2O_{(l)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$$

Magnesium powder burns very rapidly with an intense white flame. This has led to its use in fireworks and S.O.S. flares.

(ii) Reactions with Oxygen

Reaction of Sodium with Oxygen:

Sodium burns in oxygen with a golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air.

- Na reacts vigorously with oxygen in open air to form peroxide.

$$2Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{2(s)}$$

- Under special conditions like limited O_2 or high temperature, sodium oxide is formed.

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)}$$

Reaction of Magnesium with Oxygen:

Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(g)}$$

(iii) Reactions with Chlorine

Chlorine reacts with both metals to give soluble salts.

Reaction of Sodium with Chlorine:

Chlorine reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed. $2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$

Reaction of Magnesium with Chlorine:

Magnesium reacts with chlorine to give white solid, magnesium chloride.

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

Quick Check 1.7

(a) What is the nature of oxides and hydroxides of Na and Mg?

Ans: The oxides and hydroxides of sodium (Na) and magnesium (Mg) are basic in nature.

- Sodium forms strongly basic compounds like sodium oxide (Na_2O) and sodium hydroxide (NaOH), both of which readily dissolve in water to produce highly alkaline solutions.
- Magnesium forms weakly basic compounds such as magnesium oxide (MgO) and magnesium hydroxide ($Mg(OH)_2$), which are only slightly soluble in water and produce weakly alkaline solutions.

(b) What could you predict about the reactivity of Ca, a group 2 element, when reacted with water and oxygen?

Ans: **Reaction of Ca with Water:** Calcium reacts with water to produce calcium hydroxide and hydrogen gas. This reaction is relatively slow and a white precipitate of calcium hydroxide is formed. $Ca_{(s)} + 2H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + H_{2(g)}$

Reaction of Ca with Oxygen:

Calcium reacts with oxygen to form calcium oxide (CaO), also known as quicklime.

$$2Ca_{(s)} + O_{2(g)} \longrightarrow 2CaO_{(s)}$$

Trends in Bonding in Oxides of Period 3

As we move from left to right across Period 3, the electronegativity of elements increases, which leads to a decrease in the ionic character and an increase in the covalent character of their oxides.

(i) Oxides of **Groups 1, 2, and 3** are more ionic in nature. These oxides form giant ionic lattices with strong electrostatic forces between oppositely charged ions. Examples include Na_2O and MgO.

(ii) Oxides of **Groups 4, 5, 6, and 7** are more covalent. These exist as molecular covalent compounds with relatively weak intermolecular forces. Examples include SO_3 and Cl_2O_7.

Oxides of Period 3

Oxides of Period 3	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Bonding	Ionic	Ionic	Partly Covalent	Covalent	Covalent	Covalent	Covalent
Structure	Giant Ionic	Giant Ionic	Giant Ionic	Giant Covalent	Molecular	Molecular	Molecular

Trends in Bonding in Chlorides of Period 3

As we move from left to right across Period 3, the covalent character in chlorides increases due to decrease in difference of electronegativity between the halogen and the other atom.

(i) Chlorides of Groups 1, 2, and 3 are predominately ionic. e.g., NaCl etc.

(ii) Chlorides of Groups 4, 5, 6 and 7 are covalent e.g., CCl_4, PCl_5, S_2Cl_2 etc.

Chlorides of Period 3	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5	S_2Cl_2
Bonding	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Structure	Giant Ionic	Giant Ionic	Polymeric	Molecular	Molecular	Molecular

Ex.Q.4: Explain with the help of equations acidic and basic behavior of oxides and chlorides.

OXIDES

"Binary compounds formed by the reaction of oxygen with other elements are called oxides."

Oxides have quite unusual properties & there is an extensive and varied chemistry of these compounds.

Classification of Oxides: Oxides can be classified on the basis of:

(A) Nature of bonding (B) Oxidation state of oxygen (C) Acidic and basic character

(A) On the basis of oxidation state of oxygen

Oxide	Oxidation State of Oxygen	Examples
Normal oxide	−2	Li_2O, MgO, ZnO, Al_2O_3, Cl_2O_7
Peroxide	−1	H_2O_2, Na_2O_2, BaO_2
Super oxide	−1/2	KO_2, RbO_2, CsO_2
Sub oxide	—	C_3O_2 (Average O.S of Carbon = 4/3)
Ozonide	−1/3	KO_3

(B) On the basis of acidic or basic character

On this basis there are following three types of oxides;

(i) Basic oxides

"An oxide that when combined with water gives off an alkali is called basic oxide."

- Metals of group I and 2 (except Be) form basic oxides when react with oxygen.
- These oxides are usually ionic in nature.
- On adding in water; they give basic solutions. e.g. Na_2O, MgO, CaO, BaO etc.

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$
$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)}$$

- Group 2 hydroxides solubility increases down the group so alkalinity also increases down the group.

(ii) Amphoteric oxides

"Amphoteric oxides are oxides that can react with both acids and bases."

- They have the ability to behave as either an acid or a base, depending on the conditions.
- They behave as bases when react with strong acids and behave as acids with strong bases.
- The less electropositive elements (Be, Al, Zn, Ga, In) form amphoteric oxides. e.g. BeO, Al_2O_3, ZnO, Bi_2O_3 etc.

Aluminum oxide (Al_2O_3) is insoluble in water but reacts with hydrochloric acid to form aluminium chloride and water, and with sodium hydroxide to form sodium aluminate and water.

$$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(s)} + 3H_2O_{(l)}$$
$$Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow Na_2Al_2O_{4(s)} + H_2O_{(l)}$$

(iii) Acidic oxides

"An oxide that when combined with water gives off an acid is called an acidic oxide."

- Non-metals (C, N, O, P etc.) react with oxygen to form acidic oxides which are held together by covalent bonds.
- On adding in water, they give acidic solutions. e.g. CO_2, P_4O_{10}, SO_3, Cl_2O_7 etc.

$$P_2O_{3(s)} + H_2O_{(l)} \longrightarrow H_3PO_{3(aq)}$$
$$SO_{2(g)} + H_2O_{(l)} \longrightarrow H_2SO_{3(aq)}$$

Silicon dioxide (SiO_2) is acidic oxide as it can react with bases.

- Reactions of theses oxides with bases are given below:

$$SiO_{2(s)} + 2NaOH_{(aq)} \longrightarrow Na_2SiO_{3(aq)} + H_2O_{(l)}$$
$$P_2O_{3(s)} + 6NaOH_{(aq)} \longrightarrow 2Na_3PO_{3(aq)} + 3H_2O_{(l)}$$
$$SO_{2(g)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{3(aq)} + H_2O_{(l)}$$

CHLORIDES

"Binary compounds of chlorine with other elements are called chlorides."

These chlorides show characteristic behavior when we add them into water, resulting in solutions that can be acidic or neutral.

Classification of Chlorides

(i) Neutral Chlorides:

"The chloride that, when dissolved in water, produce a neutral solution with a pH close to 7 are called neutral chlorides."

- At the start of period 3, chlorides of sodium and magnesium do not react with water. The solutions formed contain the positive metal ions and negative chloride ions surrounded by water molecules.
- These ions are now known as hydrated ions and this process is known as hydration.

Examples:

$$NaCl_{(s)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} \qquad (pH=7)$$
$$MgCl_{2(s)} \longrightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)} \qquad (pH=6.5)$$

Group 1 and group 2 chlorides are also neutral with few exceptions.

(ii) Acidic Chlorides:

"The chlorides that, when dissolved in water, produce an acidic solution (with a pH less than 7) due to hydrolysis, are called acidic chlorides."

If we move in period 3, from aluminum to sulphur all chlorides undergo hydrolysis to make acidic solution.

Hydrolysis of $AlCl_3$:

When $AlCl_3$ is added to water, it dissociates into aluminum ions and chloride ions. Al^{3+} ion is hydrolyzed to produce H^+ ions. This turns the solution acidic.

$$AlCl_{3(s)} + 3H_2O_{(l)} \longrightarrow Al(OH)_{3(s)} + 3HCl_{(aq)}$$

Hydrolysis of $SiCl_4$ and PCl_3:

$$SiCl_{4(l)} + 2H_2O_{(l)} \longrightarrow SiO_{2(s)} + 4HCl_{(aq)}$$
$$PCl_{3(l)} + 3H_2O_{(l)} \longrightarrow H_3PO_{3(s)} + 3HCl_{(aq)}$$

Quick Check 1.8

(a) ZnO reacts with HCl to give $ZnCl_2$ and with NaOH to give Na_2ZnO_2. Give equations and also predict the type of this oxide?

Ans: $ZnO + HCl \longrightarrow ZnCl_2 + H_2O$

$ZnO + 2NaOH \longrightarrow Na_2ZnO_2 + H_2O$

Zinc oxide (ZnO) behaves as a base when it reacts with a strong acid like HCl, and as an acid when it reacts with a strong base like NaOH. Therefore, ZnO is classified as an amphoteric oxide.

(b) Why $AlCl_3$ is an acidic halide, but NaCl not?

Ans: $AlCl_3$ is an acidic halide because it reacts with water to produce H^+ ions due to its covalent nature and acts as a Lewis acid. In contrast, NaCl is ionic and neutral in water, so it does not show acidic behavior.

(c) Predict whether the chlorides PCl_5, NCl_3 would be acidic or basic, give reason.

Ans: PCl_5 can act as a Lewis acid because it has an empty d orbital that can accept an electron pair from a Lewis base. NCl_3, on the other hand, can act as a Lewis base because the lone pair on nitrogen can be donated to a Lewis acid.

(d) Would SO_2 and P_2O_5 react with HCl and H_2SO_4 or with NaOH?

Ans: SO_2 and P_2O_5 are acidic oxides, so they do not react with acids like HCl or H_2SO_4.
Instead, they react with bases like NaOH to form salts:

$$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$$
$$P_2O_5 + 6NaOH \rightarrow 2Na_3PO_4 + 3H_2O$$

OXIDATION NUMBER

"The formal charge on an atom in a molecule or ion is called the oxidation number of that atom."

- It may be zero, positive, negative or in fraction.
- The oxidation number is also referred to as the oxidation state.

Oxidation Number and Nature of Species

- **In ionic compounds**, the oxidation number of an atom is the charge which appears on the, ions. In sodium chloride oxidation state of Na is +1 and chlorine is –1 i.e. Na^{+1}, Cl^{-1}.
- **In covalent compounds**, it is decided on the basis of difference in their relative electronegativities. e.g. $SnCl_4$ is a covalent compound in which oxidation state of Sn is +4 and of Cl is –1. Similarly, oxidation states of oxygen are $H_2\overset{-2}{O}$, $\overset{+2}{O}F_2$.
- **In free (Native) state:** (Cl_2, Br_2, O_2, H_2 etc.) oxidation state of elements is zero.

Oxidation Numbers in Oxides and Chlorides of 3rd Period elements

- The oxidation number of an element of 3rd Period in its oxide and chlorides corresponds to the number of electrons used for bonding.
- It is always positive because oxygen and chlorine are more electronegative than any of these elements.
- The oxidation number matches the group number, reflecting the total number of valence electrons.
- Phosphorus and sulfur exhibit several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals. e.g. In SO_2, sulfur has an oxidation number of +4 because only four electrons are used for bonding, while in SO_3, sulfur has an oxidation number of +6 because all six electrons are used for bonding.

(i) <u>In oxides</u>, the oxidation number increases from +1 in Na to +6 in S.

Period 3 Elements	Na	Mg	Al	Si	P	S	S
Oxides	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10} / P_4O_6	SO_2	SO_3
Oxidation Number	+1	+2	+3	+4	+5 / +3	+4	+6

(ii) <u>In chlorides</u>, the oxidation number increases from +1 in Na to +5 in P.

Period 3 Elements	Na	Mg	Al	Si	P	P	S
Chlorides	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	PCl_5	SCl_2
Oxidation Number	+1	+2	+3	+4	+3	+5	+2

Quick Check 1.9

(a) Calculate the oxidation number of Sulphur in SO_2 and SO_3.

Ans:

Oxidation Number of Sulphur in SO_2	Oxidation Number of Sulphur in SO_3
O.N of O = –2	O.N of O = –2
O.N of S = ?	O.N of S = ?
(O.N of S) + 2 (O.N of O) = 0	(O.N of S) + 3 (O.N of O) = 0
O.N of S + 2 (–2) = 0	O.N of S + 3 (–2) = 0
O.N of S – 4 = 0	O.N of S – 6 = 0
O.N of S = +4	O.N of S = +6

(b) Why some p block elements show variable oxidation state?

Ans: Some p-block elements exhibit variable oxidation states because they can expand their octet by promoting electrons into empty d-orbitals. For example, In SO_2, sulfur has an oxidation state of +4, as it forms bonds using four of its valence electrons. In SO_3, sulfur has an oxidation state of +6, as all six valence electrons are involved in bonding.

Exams Short Answer Questions

Q. What is the difference between acidic and basic oxides? Give one example of each.
Q. What are amphoteric oxides? Give one example.
Q. Name different classes of oxides and mention trend of oxides across the period?
Q. Although both sodium and phosphorous are present in the same period of the periodic table yet their oxides are different in nature. Na_2O is basic while P_2O_5 is acidic in character.
Q. Why the metals are good conductors?
Q. Why Na_2O is basic while SO_3 is acidic in nature.
Q. Why the basicity of IIA group metal oxides increase on descending a group of periodic table?

MULTIPLE CHOICE QUESTIONS

EXERCISE MCQs

(i) Which scientist first time observed the periodicity in the elements?
(a) J. Newlands (b) L. Meyer
(c) J.W. Dobereiner (d) D. I. Mendeleev

(ii) Recognize the element if it has 3 electron shells, belongs to "s" block and has 2 electrons in its outer most shell.
(a) Calcium (b) Sodium
(c) Magnesium (d) Potassium

(iii) Which one do you think is correct about metallic character?
(a) It decreases from top to bottom in a group.
(b) It increases from top to bottom in a group.
(c) It remains constant from left to right in a period.
(d) It increases from left to right in a period.

(iv) Which property increases as you go down a group in the periodic table?
(a) Atomic radius (b) Electron Affinity
(c) Electronegativity (d) Ionization energy

(v) Which set of the following conditions results in higher ionization energy?
(a) Smaller atom and greater nuclear charge.
(b) Smaller atom and smaller nuclear charge
(c) larger atom and greater nuclear charge
(d) larger atom and the smaller nuclear charge

(vi) Which of the following atoms show more than one (variable) oxidation states?
(a) Sodium (b) Magnesium
(c) Aluminum (d) Phosphorous

(vii) Which is the correct general trend in the variation of electron affinity in a group?
(a) It becomes less negative from top to bottom.
(b) It becomes more negative from top to bottom.
(c) It remains the same.
(d) It has no definite trend and changes irregularly.

(viii) What is the oxidation state of sulfur in the sulfate ion (SO_4^{-2}).
(a) +4 (b) +2
(c) +6 (d) 0

(ix) Which is the correct trend in variation of electronegativity along a period of the periodic table?
(a) It decreases from left to right across a period.
(b) It increases from left to right across a period.
(c) It remains constant.
(d) It has no definite trend

(x) The atomic radius generally_____ across a period in the periodic table.
(a) Increases
(b) Decreases
(c) Remains constant
(d) First increases then decreases

(xi) Which one of the following elements has the highest ionization energy?
(a) Sodium (Na) (b) Magnesium (Mg)
(c) Aluminum (Al) (d) Argon (Ar)

ADDITIONAL PRACTICE MCQs

1. Maximum electron affinity values of N are:
(A) 1 (B) 2
(C) 3 (D) None

2. The oxides of electropositive elements are:
(A) Acidic (B) Basic
(C) Amphoteric (D) Neutral

3. Which of the following elements form polymeric halides?
(A) Be (B) Ga
(C) Al (D) All of these

4. Zinc oxide is:
(A) basic (B) Amphoteric
(C) Acidic (D) Neutral

5. The elements of group II-A are called:
(A) Alkali metals
(B) Alkaline earth metals
(C) Coinage metals
(D) Metalloids

6. In modern periodic table, the elements are arranged in the ascending order of:
(A) Atomic masses (B) Valency
(C) Valence electrons (D) Atomic number

7. Memorable contributions in the classification of elements were made by:
(A) Al-Razi (B) Dobereiner
(C) Newlands (D) All of them

8. The longest period in the periodic table is:
(A) 7th (B) 6^{th}
(C) 5th (D) 4^{th}

9. f-block elements are also called:
(A) Non-typical transition elements
(B) Outer transition elements
(C) Inner transition elements
(D) Normal transition elements

10. The number of elements in fourth period of periodic table is:
(A) 32 (B) 18
(C) 10 (D) 8

11. Hydrogen can be placed with the elements of group IV A because both:
(A) Act as strong oxidizing agent
(B) Act as strong reducing agent
(C) Possess the property of catenation
(D) Form neutral oxides

12. Most of the elements are:
(A) Crystalloids (B) Metals
(C) Metalloids (D) Non-metals

13. The decrease in atomic sizes is not much prominent across rows containing elements of:
(A) s-Block (B) p-Block
(C) d-Block (D) f-Block

14. Elements of the periodic table are classified into blocks:
(A) Four (B) Three
(C) Five (D) Six

15. How many elements are present in 5th period of the periodic table?
(A) 32 (B) 8
(C) 18 (D) 28

16. Which of the following oxide is different in nature?
(A) MgO (B) BeO
(C) SrO (D) CaO

17. Both hydrogen and alkali metals have:
(A) One electron in valence shell
(B) Valence shell is completely filled
(C) Lack one electron in valence shell
(D) Valence shell is partially filled

18. Lithium and beryllium in 2nd period resemble in properties with:
(A) Ca & Be (B) Mg & Al
(C) O_2, H_2 (D) Na & O_2

19. Non-metallic character is tendency to _____ electron.
(A) Gain (B) Lose
(C) a and b (D) None of these

20. The inner transition elements of 7th period are called:
(A) Lanthanides (B) Actinides
(C) Rare earth elements (D) Both b & c

21. Which of the following sets of elements are metalloids?
(A) C & Si (B) As & Sb
(C) Mg & Na (D) Both a and b

22. From which we can remove electron easily?
(A) Cesium (B) Lithium
(C) Aluminum (D) Magnesium

23. Which of the following order is correct for the first ionization energies of following elements?
(A) B < Be < N < O (B) B < Be < O < N
(C) Be < B < N < O (D) B < O < Be < N

24. Which of the following species has the highest ionization potential?
(A) Li^+ (B) Al^+
(C) Mg^+ (D) Ne

25. For which element is the gaining of an electron most exothermic?
(A) Li (B) N
(C) F (D) B

26. Which property decreases as you move down a column in the periodic table?
(A) atomic size (B) ionization energy
(C) metallic character (D) nuclear charge

27. The outermost configuration of most electronegative element is:
(A) $ns^2\ np^5$ (B) $ns^2\ np^6$
(C) $ns^2\ np^4$ (D) ns^2

28. Which of following will show maximum penetration effect?
(A) 4f (B) 4d
(C) 4p (D) 4s

29. The first four ionization energies of an element Z are 738, 1451, 7730 and 10541 kJ mol^{-1}. Which one of following ions is most likely to be formed when Z reacts with fluorine?
(A) Z^{2+} (B) Z^{3+}
(C) Z^{4+} (D) Z^+

30. Magnitude of ionization energy depends upon:
(A) number of positive charges
(B) shielding effect increases
(C) spin pair repulsion
(D) All of the above

31. Ionization energy between last element of one period and first element of next period receives a rapid:
(A) increase (B) decrease
(C) constancy (D) neutral

32. Force of attraction between nucleus and electrons increases across periods because:
(A) nuclear charge increases
(B) distance remains constant
(C) shielding effect almost constant
(D) All of the above

33. Electrons always reside in certain energy level outside:
(A) nucleus (B) axis
(C) zone (D) lobe

34. Which of the following would have the lowest first ionization energies?
(A) Alkali metals
(B) Transition metals
(C) Halogens
(D) Alkaline Earth metals

35. Which equation correctly represents the first ionization of aluminum?

(A) $Al_{(s)} \longrightarrow Al^{+1}_{(g)} + e^-$

(B) $Al_{(g)} + e^- \longrightarrow Al^{-1}_{(g)}$

(C) $Al_{(g)} \longrightarrow Al^{-1}_{(g)} + e^-$

(D) $Al_{(g)} \longrightarrow Al^{+1}_{(g)} + e^-$

36. Which of the following atoms would have the largest second ionization energy?

(A) Mg (B) Cl

(C) S (D) Na

37. Removing an electron from sodium is an _____ process and adding an electron to chlorine is an _____ process:

(A) Endothermic, exothermic

(B) Exothermic, endothermic

(C) Endothermic, endothermic

(D) Exothermic, exothermic

38. When atomic radius increases electron affinity:

(A) decreases (B) increases

(C) remains constant (D) unpredictable

39. For electron affinity of halogens, which of the following is correct?

(A) Br < I (B) F > Cl

(C) Br > Cl (D) F > I

40. The ionization energy is _________ to the size of an atom.

(A) directly proportional (B) equal to

(C) inversely proportional (D) independent

ANSWERS TO MULTIPLE CHOICE QUESTIONS

Hints and Explanations

EXERCISE MCQs

(i). a Newlands was the first to recognize and propose periodicity clearly and systematically

(ii). c $_{12}Mg = 1s^2\ 2s^2\ 2p^6\ 3s^2$

(iii). b Metallic character increases down the group due to increasing atomic size and decreasing ionization energy.

(iv). a Atomic size increases due to the addition of new electron shells.

(v). a A small atom with a high nuclear charge holds electrons tightly, requiring more energy to remove an electron.

(vi). d Phosphorus can show multiple oxidation states: –3, +3, +5, etc.

(vii). a As atomic size increases down a group, atoms attract electrons less strongly, resulting in a less exothermic (less negative) electron affinity.

(viii). c Let oxidation state of S = x.
Oxygen = –2, and there are 4 O atoms:
$x + 4(-2) = -2 \rightarrow x - 8 = -2 \rightarrow x = +6$

(ix). b Effective nuclear charge increases, attracting bonding electrons more strongly.

(x). b As, we move across a period, electrons are added to the same shell while nuclear charge increases, pulling electrons closer.

(xi). d Argon is a noble gas with a full outer shell, making it very stable. High energy is required to remove an electron.

ADDITIONAL PRACTICE MCQs

(1). D Electron affinity of Nitrogen is -7kJ/mol. Which is very small because of its stable electronic configuration?

(2). B Electropositive elements are metals (Na, Li, K) and metals form basic oxides.

(3). D $BeCl_2$, $AlCl_3$, $GaCl_3$ are polymeric halides.

(4). B Reacts as acid with base and reacts as base with acid.

$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$
(ZnO as a base)

$ZnO + 2NaOH + H_2O \longrightarrow Na_2[Zn(OH)_4]$
(ZnO as an acid)

(5). B (i) They give alkalies in water.
(ii) Exist abundantly on earth crust in form of carbonates and silicates.

(6). D In modern periodic table the base of arrangement is number of protons in nucleus (atomic number).

(7). D All these have contribution in classifying elements.

(8). B It has 32 elements and is a longest period.

(9). C In these elements, three outermost shells are not completely filled. The last electron in them enters in the (n – 2) subshells.

(10). B Fourth period is called long period.

(11). B $H_2 + CuO \longrightarrow Cu + H_2O$
$C + SnO_2 \longrightarrow CO_2 + Sn$

(12). B Out of 118 elements, about 83 elements are metals.

(13). C Due to diffused d sub - shell.

(14). A (i) s-block (ii) p-block (iii) d-block (iv) f-block

(15). C 5^{th} period is also called long period.

(16). B BeO is amphoteric in nature.

$BeO + H_2SO_4 \longrightarrow BeSO_4 + H_2O$

$BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O$

(17). A Valence shell electronic configuration of
$_1H = 1s^1$ & $_{11}Na = 1s^2\ 2s^2\ 2p^6\ 3s^1$

(18). B Li & Be resemble in properties with Mg & Al respectively due to diagonal relationship.

(19). A Greater the tendency of non-metals to gain electron, greater is the non-metallic character.

(20). D Actinides are elements of 7^{th} period which are rare earth elements.

(21). B Metallic character increases down the group. N & P are non – metals. As & Sb are metalloid & Bi is metal.

(22). A It is easy to remove an electron from cesium due to its large size.

(23). B These elements belong to same period (2^{nd}) of periodic table. Atoms with half-filled and completely filled orbitals in the outer shell are more stable and possess higher ionization energies. s-orbitals have greater penetration power than p-orbitals. Hence the removal of electrons from s-orbitals requires more energy.

Boron, B is smaller than beryllium, Be atom. Hence, we expect increase in ionization energy from Be to B.

However, Be atom has greater ionization energy than B atom. It is due to stable $2s^2$ configuration and the presence of valence electron in the s-orbital. Removal of electron from s-orbital requires more energy than from p-orbital as stated above.

$2s^22p^3$ configuration is more stable than $2s^22p^4$ due to half-filled p-sublevel. Hence nitrogen, N atom has greater ionization energy than oxygen, O atom.

(24). A Li^+ has $1s^2$ configuration, which is the configuration of He atom. Hence it should possess highest IP value.

(25). C Fluorine (F) has the highest electron affinity among the given elements. This means it releases the most energy when it gains an electron, making the process highly exothermic

(26). B I.E decreases down the group (column) due to increase in atomic radius and shielding effect.

(27). A Halogens (VII A) are most electronegative elements having ns^2np^5 outermost configuration.

(28). D Closer to orbital to a nucleus, more will be its penetration effect. Order of penetration effect is;

$$4s > 4p > 4d > 4f$$

(29). A A huge gap between 2^{nd} and 3^{rd} I.E values shows that it exists usually as Z^{2+}.

(30). D Magnitude of ionization energy depends upon Number of positive charges Shielding effect increases Spin pair repulsion.

(31). B I.E of Ne (Last element of period 2) = 2081 kJ mol^{-1}

I.E of Na (First element of period 3) = 496 kJ mol^{-1}

(32). A As we more across the period, nuclear charge increases while atomic radius decreases. This increase force of attraction between nucleus and electrons.

(33). A Electrons are located in energy levels or shells that surround the nucleus of an atom. These energy levels are fixed and quantized, not random or found in zones, lobes, or axes.

(34). A Alkali metals have one loosely held electron in their outer shell, making them the easiest to ionize. Their low effective nuclear charge and large atomic size mean they give up their outer electron easily.

(35). D Necessary conditions for I.E:

(i) Isolated gaseous atom

(ii) Loss of valence electron

(iii) Formation of cation

(36). D Na loses its one electron in valence shell to have stable electronic configuration.

$$Na^{+1} = 1s^2\,2s^2\,2p^6$$

Now it becomes difficult to remove one more electron (2^{nd} I.E) from Na^{+1}.

(37). A $Na_{(g)} \longrightarrow Na^{+1}_{(g)} + 1\bar{e}$ I.E = +496 kJ mol^{-1}

$Cl_{(g)} + 1\bar{e} \longrightarrow Cl^{-1}_{(g)}$ E.A = –349 kJ mol^{-1}

Removing an electron (I.E) is an endothermic (+ive) while adding an electron (E.A) is an exothermic (–ive).

(38). A With the increase in atomic radius, the force of attraction between valence electrons and the nucleus decreases. As a result, the electron affinities usually decrease.

(39). D Fluorine has the highest electron affinity among the halogens. As we move down the group from F to I, the atomic size increases, and the incoming electron is less strongly attracted, so the electron affinity decreases.

(40). C $I.E \propto \dfrac{1}{\text{atomic radius}}$

SHORT ANSWER QUESTIONS

EXERCISE SHORT ANSWER QUESTIONS

Q.2a. What is 1st ionization energy? Give an example.

Ans: 1st Ionization Energy (ΔH_{i1})

The energy required to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous positive ions is called the first ionization energy (ΔH_{i1}).

Example: $Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$ $\Delta H_{i1} = 494\ kJ\ mol^{-1}$

Q.2b. Explain why sulphur has a lower first ionization energy than phosphorus.

Ans: Sulphur has a slightly lower first ionization energy than phosphorus because of the presence of paired electrons in its 3p orbitals, which increases electron-electron repulsion and makes it easier to remove an electron. In contrast, phosphorus has half-filled 3p orbitals, which are relatively more stable due to exchange energy, thereby requiring more energy to remove an electron.

$$S_{(g)} \longrightarrow S^+_{(g)} + e^- \quad \Delta H_{i1} = 999.6\ kJ\ mol^{-1}$$
$$P_{(g)} \longrightarrow P^+_{(g)} + e^- \quad \Delta H_{i1} = 1012\ kJ\ mol^{-1}$$

Q.2c. Why the elements in Group 13 to 17 are called p-block elements?

Ans: The elements of group 13 to group 17 are known as p-block elements as their valence electrons are present in p-subshell. e.g. $_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

Q.2d. What are the factors that affect electronegativity?

Ans: Factors affecting Electronegativity

(i) Atomic size: A larger atomic size will result in a lower value of electronegativity. This is because electrons being far away from the nucleus will experience a weaker force of attraction.

(ii) Effective nuclear Charge: A higher value of the effective nuclear charge will result in a greater value of electronegativity, because an increase in nuclear charge causes greater attraction to the bonded electrons. This is why the electronegativity in a period increases from left to right.

Q.2e. What factors are responsible for the increasing reactivity of alkali metals as you move down the group?

Ans: The reactivity of alkali metals increases as we move down the group due to the following reasons:

(i) Increase in atomic size

(ii) Decrease in ionization energy

Q.2f. Why some of the elements show variable oxidation numbers while others do not?

Ans: Some elements exhibit variable oxidation numbers by utilizing empty d-orbitals to accommodate additional electrons after promoting outer-shell electrons. As a result, they can form more bonds than allowed by the octet rule, leading to variable oxidation numbers.

Q.2g. Identify the element which is in period 5 and group 15?

Ans: Since the element belongs to group 15 of periodic table so it has 5 valence electrons; and it is found in period 5 so it has five shells around its nucleus. It means that the 5 valence electrons are in the 5th shell. The configuration will be:

$$_{51}Sb = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^3$$

Q.2h. Why oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorous?

Ans: Sodium and magnesium oxides are more ionic because of the large electronegativity difference between the metals and oxygen. In contrast, nitrogen and phosphorus have higher electronegativities, so their oxides are more covalent in nature and exhibit weaker ionic character.

Q.2i. Give reason for the different chemical reactivities of Na and Mg toward oxygen and chlorine.

Ans: Sodium (Na) and magnesium (Mg) exhibit different reactivities toward oxygen and chlorine due to differences in their electronic configurations and ionization energies. Sodium has one valence electron and a lower ionization energy, making it more reactive than magnesium, which has two valence electrons and a comparatively higher ionization energy.

Q.2j. Why the ionization energy of lithium is much lower than that of helium despite the fact that the nuclear charge of lithium is +3 and that of helium is +2?

Ans: The ionization energy of lithium is lower than helium because lithium's outermost electron is in the second shell (2s), farther from the nucleus and more shielded by inner electrons. In contrast, helium's electrons are in the first shell (1s), closer to the nucleus with minimal shielding, leading to a stronger electrostatic attraction and a higher ionization energy.

Q.2k. The ionization energy of Be (atomic no. 4) is higher than that of B (atomic no. 5), despite the fact that the nuclear charge of Be is +4 and that of B is +5.

Ans: In boron (B), the electron is removed from a p-orbital, which is more extended than the s-orbital in beryllium (Be). Additionally, in beryllium, the electron is removed from a paired s-orbital, which requires more energy to remove. In contrast, the unpaired electron in boron's p-orbital results in a lower ionization energy for boron.

Q.2l. What is common in Na^+, Mg^{2+}, Al^{3+}, Ne^0 and F^-? Arrange them in increasing order of sizes.

Ans: Na^+, Mg^{2+}, Al^{3+}, Ne^0 and F^- have 10 electrons each so they are isoelectronic species.

Ion	F^-	Na^+	Mg^{2+}	Ne^0	Al^{3+}
Number of electrons	10	10	10	10	10
Radius (pm)	136	95	72	68	53

Increasing order of sizes: $F^- > Na^+ > Mg^{2+} > Ne^0 > Al^{3+}$

Q.2m. Consider the chlorides of sodium, magnesium, and phosphorus (V): NaCl, $MgCl_2$, and PCl_5.
(i) Classify each of these chlorides as acidic, basic, or neutral.
(ii) For each chloride, briefly explain the reason for your classification, referring to their behavior when dissolved in water.

Ans: Classification

(i) NaCl – Neutral (ii) $MgCl_2$ – Slightly acidic (iii) PCl_5 – Acidic

(i) Sodium chloride (NaCl):
Reason: NaCl is a salt of a strong acid (HCl) and a strong base (NaOH). When dissolved in water, it dissociates completely into Na^+ and Cl^- ions, neither of which hydrolyzes water.
In water:

$$NaCl \longrightarrow Na^+ + Cl^-$$

No effect on pH → neutral solution.

(ii) Magnesium chloride ($MgCl_2$):
Reason: $MgCl_2$ comes from a strong acid (HCl) and a weak base ($Mg(OH)_2$). The Mg^{2+} ion is small and highly charged, so it can hydrolyze water to a small extent, releasing H^+ ions.
In water:

$$Mg^{2+} + 6H_2O \rightleftharpoons [Mg(H_2O)_6]^{2+} \rightleftharpoons [Mg(H_2O)_5(OH)]^+ + H^+$$

Slightly lowers pH → slightly acidic solution.

(iii) Phosphorus(V) chloride (PCl_5):
Reason: PCl_5 is a covalent chloride that reacts with water in a hydrolysis reaction to form phosphoric acid (H_3PO_4) and hydrogen chloride (HCl), both of which are acids.
In water: $PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$

ADDITIONAL SHORT ANSWER QUESTIONS

1. d and f–Block elements are called transition elements.

Ans: The d and f-block elements are located between the 's' and p-block elements and their properties are in transition between the metallic elements of the s-block and non-metallic elements of the p-block. Thus, they are called transition elements.

2. Lanthanide contraction controls the atomic sizes of elements of 6th and 7th period.

Ans: The lanthanide contraction is the gradual decrease in the atomic radii of elements as we move from left to right across the lanthanide series. This contraction occurs due to the poor shielding effect of the 4f electrons so that the effective nuclear charge attracting each electron steadily increases through the lanthanide series.

3. Define Dobereiner's law of triads. Also give two examples.

Ans: Dobereiner's law of Triads

The atomic mass of middle element is the average of atomic masses of the other two elements of triads.

Triad	Li	Na	K	
Atomic mass	7	23	39	Atomic mass of Na = $\frac{7+39}{2} = 23$
Triad	Ca	Sr	Ba	
Atomic mass	40	88	137	Atomic mass of Sr = $\frac{40+137}{2} = 88.5$

4. What is Newland's law of Octaves?

Ans: Newland's law of Octaves:

If the elements are arranged in the increasing order of their atomic masses, every eighth (8^{th}) element had some properties in common with the first one.

5. Define Mendeleev's and modern periodic law.

Ans: Mendeleev's Periodic Law: If the elements are arranged in ascending order of their atomic masses, their chemical properties repeat in a periodic manner.

Modern Periodic Law: If the elements are arranged in ascending order of their atomic numbers, their chemical properties repeat in a periodic manner.

6. Write two defects of Mendeleev's periodic table.

Ans: Defects of Mendeleev's Periodic Table

(i) Similar elements Separated: In Mendeleev's periodic table, certain chemically similar elements such as copper and mercury; gold and platinum have been placed in different groups.

(ii) Position of Hydrogen: Hydrogen is placed in Group IA. However, it actually resembles the elements of Group-IA (alkali metals) as well as the elements of Group VII-A (halogens). Thus, the position of hydrogen in the periodic table is not clear.

7. Define periodic table. How many groups and periods are present in it?

Ans: Periodic Table

A table obtained by the arrangement of elements into periods and groups is called periodic table.

- There are eight (8) groups, which are usually numbered by Roman numerals I to VIII.
- There are 7 periods in the periodic table numbered by Arabic numerals 1 to 7.

8. Give essential features of period four (4) in modern periodic table?

Ans: Essential Features of Period 4

The period 4 contains eighteen (18) elements and is called long period. Out of these, 8 are representative elements which belong to sub-group A. Whereas the other 10 elements, placed in the center of the table, belong to subgroup B and are known as outer transition elements.

9. Define Group and period. How many elements are there in period number 1?

Ans: Group: The vertical column of elements having similar properties is called group.

There are eight (8) groups, which are further classified in subgroup A and B. They are usually numbered by Roman numerals, I to VIII.

Period: The horizontal row of elements in periodic table is called period.

There are 7 periods in modern periodic table numbered by Arabic numerals 1 to 7.

Elements in Period 1: There are two elements in period number 1 i.e., hydrogen (H) and helium (He).

10. Write the names of families in periodic table.

Ans: Families in periods: Lanthanides and actinides

Families in groups: (i) Alkali Metals (ii) Alkaline Earth metals (iii) Halogens (iv) Noble gases

11. How classification of elements in different blocks helps in understanding their chemistry?

Ans: Classification of Elements in Blocks

Elements in the periodic table can be classified into four blocks.

(i) s-block (ii) p-block (iii) d-block (iv) f-block

This classification is quite useful in understanding the chemistry of elements and predicting their properties especially the concept of valency or oxidation state.

12. Why ionization energy decreases down the group and increases along a period?

Ans: (A) Across the period: The I.E increases from left to right in a period. It is due to;

(i) decrease in atomic radius (ii) increase in nuclear charge

(b) Down the Group: The I.E decreases from top to bottom in a group. It is due to;
(i) Increase in atomic radius (ii) Increase in shielding effect

13. Why the second value of electron affinity of an element is usually shown with a positive sign?

Ans: When an atom gains a second electron, the added electron experiences repulsion from the negatively charged anion already formed. To overcome this repulsion, energy must be supplied. Thus, the formation of a di-negative ion is an endothermic process and its electron affinity is shown with positive sign.

$$O_{(g)} + e^- \longrightarrow O^{-1}_{(g)} \quad E.A_1 = -141\ kJ\ mol^{-1}$$
$$O^{-1}_{(g)} + e^- \longrightarrow O^{-2}_{(g)} \quad E.A_2 = +780\ kJ\ mol^{-1}$$

14. Why the atomic radius decreases from left to the right within a period and increases from top to bottom down the group?

Ans: Trend of Atomic Radius

(A) Across the period: Atomic radius decreases from left to right in a period due to;
(i) Increase in atomic number (ii) Increase in nuclear charge
(b) Down the Group: Atomic radius increases down the group due to;
(i) Increase in number of shells
(ii) Increase in shielding effect down the group due to increase in intervening electrons

15. Why the size of cation is always smaller than its parent atom? Give example also.

Ans: The size of a cation is smaller than its parent atom due to imbalance of electron-proton ratio. Nucleus holds the remaining electrons with a stronger force. Usually there is loss of shell occurs during removal of electron. e.g., Na (186pm) $\longrightarrow$ Na^+ (102pm) + 1e

16. Give variation of ionic radius across the period.

Ans: Variation of Ionic Radius across the Period

Within a period, isoelectronic positive ions show decrease in ionic radius from left to right because of increasing nuclear charge. $Na^{+1} > Mg^{+2} > Al^{+3}$. Similar is in the case with isoelectronic negative ions. i.e. $P^{-3} > S^{-2} > Cl^{-1}$.

17. Define shielding effect.

Ans: Shielding Effect/Screening Effect

The decrease in attraction of nucleus and valence electrons due to presence of intervening electrons is called shielding effect or screening effect.

- It remains almost constant across the period.
- It increases down the group due to increase in number of intervening electrons.

18. Define ionization energy. Give its trend across a period and down a group. or Define ionization energy with an example.

Ans: Ionization Energy

The minimum amount of energy which is required to remove an electron from the outermost shell of its isolated gaseous atom is called ionization energy." e.g.

$$Na_{(g)} \longrightarrow Na^{+1}_{(g)} + 1e^- \qquad I.E = 496\ kJ\ mol^{-1}$$

(A) Across the Period: The value of I.E increases from left to right in a period. It is due to;
(i) decrease in atomic radius (ii) increase in nuclear charge
(b) Down the Group: The value of I.E decreases from top to bottom in a group. It is due to;
(i) Increase in atomic radius (ii) Increase in shielding effect

19. What is the role of shielding effect on ionization energy?

Ans: Role of Shielding Effect on Ionization Energy

An increase in shielding effect decreases the ionization energy and vice versa.

- Down the group, due to increase in number of intervening electrons, the shielding effect also increases, therefore ionization energy decreases.
- Across the period, although the number of electrons increases but shielding effect is not very effective within the same shell.

20. Ionization energy of Al^{+3} is greater than Mg^{+2}. Give the reason.

Ans: Both Al^{3+} and Mg^{2+} are isoelectronic species. The effective nuclear charge of Al^{3+} is greater and size is smaller than Mg^{2+}. So, it costs more energy to remove the outermost electron from Al^{3+} than from Mg^{2+}. Therefore, ionization energy of Al^{+3} is greater than Mg^{+2}.

21. Why second, I.E value is always greater than first I.E value?

Ans: Energy required to remove electron from a uni-positive ion is called second ionization energy.

$$Mg_{(g)} \longrightarrow Mg^{+1}_{(g)} + 1e^- \quad I.E_1 = +738 \text{ kJ mol}^{-1}$$
$$Mg^{+1}_{(g)} \longrightarrow Mg^{+2}_{(g)} + 1e^- \quad I.E_2 = +1450 \text{ kJ mol}^{-1}$$

$I.E_2 > I.E_1$ due to;

(i) Small cationic radius than parent atom **(ii)** Greater nuclear attraction on the remaining electrons

22. Give variations of electron affinity down the group and across the period.

Ans: Periodic Trend of Electron Affinity

Across the period, electron affinity increases due to;

(i) Increase in nuclear charge **(ii)** Decrease in atomic radius

Down the Group, electron affinity decreases due to;

(i) Increase in atomic radius **(ii)** Increase in shielding effect

23. Although both sodium and phosphorous are present in the same period of the periodic table yet their oxides are different in nature, Na_2O is basic while P_2O_5 is acidic in character.

Ans: A metal forms basic oxide while a non-metal forms acidic oxide. Sodium (Na) being a metal forms basic oxide as it yields base in water.

$$Na_2O + H_2O \longrightarrow 2NaOH \text{ (Base)}$$

While phosphorous (P) being a non-metal gives acidic oxide as it yields acid in water.

$$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4$$

24. What is the difference between acidic and basic oxides? Give one example of each.

Ans:

Basic Oxides	Acidic Oxides
These oxides when dissolve in water give the basic solution.	These oxides when dissolve in water give the acidic solution.
Metals form basic oxides.	Non-metals form acidic oxides.
Examples: Li_2O, MgO, CaO etc.	**Examples:** CO_2, P_4O_{10}, SO_3, Cl_2O_7 etc.

25. What are amphoteric oxides? Give one example.

Ans: Amphoteric Oxides

Those oxides which have both acidic and basic character are called amphoteric oxides.

The less electropositive elements (Be, Al, Zn, Ga, In) form amphoteric oxides.

e.g. BeO, Al_2O_3, ZnO, etc.

26. Name different classes of oxides and mention trend of oxides across the period?

Ans: Classes of Oxides

(i) Acidic oxides **(ii)** Basic oxides **(iii)** Amphoteric oxides

Across the Period: In a given period, the oxides progress from strongly basic through weakly basic, amphoteric and weakly acidic to strongly acidic.

Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Strongly Basic	Less basic	Amphoteric	Weakley Acidic	Acidic	Strongly acidic	Even more Acidic

27. Why the metals are good conductors?

Ans: Metals are good conductor of electricity due to:

(i) Presence of relatively loose electrons in their valence shell.

(ii) Ease of movement of electron in the metal lattice

28. Why the basicity of IIA group metal oxides increase on descending a group of periodic table?

Ans: The basicity of IIA group metal oxides increases down the group due to;

(i) Increase in atomic radius **(ii)** Increase in metallic characteristics **(iii)** Decrease in ionization energy

The order of basicity is $BeO < MgO < CaO < SrO < BaO$

29. Why Na_2O is basic while SO_3 is acidic in nature?

Ans: Generally, metals form basic oxides while non-metals form acidic oxides. Sodium (Na) is a metal, so, Na_2O is a basic oxide. On the other hand, sulphur (S) is a non-metal, thus SO_3 is an acidic oxide.