

09 Acid Base Chemistry

➤ Theory	293
➤ Rack Your Brain (Solutions)	318
➤ Exercise	319
➤ Multiple Choice Questions (MCQs)	319
➤ Short Answer Questions	320
➤ Descriptive Questions	323
➤ Numerical Problems	324
➤ Additional SLOs Based MCQs	325
➤ Additional Short Answer Questions	325
➤ Self Assessment	332

10 Electrochemistry

➤ Theory	333
➤ Rack Your Brain (Solutions)	352
➤ Exercise	353
➤ Multiple Choice Questions (MCQs)	353
➤ Short Answer Questions	355
➤ Descriptive Questions	359
➤ Numerical Problems	359
➤ Additional SLOs Based MCQs	371
➤ Additional Short Answer Questions	372
➤ Self Assessment	376

11 Hydrocarbons

➤ Theory	377
➤ Rack Your Brain (Solutions)	404
➤ Exercise	405
➤ Multiple Choice Questions (MCQs)	405
➤ Short Answer Questions	407
➤ Descriptive Questions	412
➤ Additional SLOs Based MCQs	412
➤ Additional Short Answer Questions	414
➤ Self Assessment	418

12 Nitrogen and Sulphur

➤ Theory	419
➤ Rack Your Brain (Solutions)	434
➤ Exercise	435
➤ Multiple Choice Questions (MCQs)	435
➤ Short Answer Questions	437
➤ Descriptive Questions	440
➤ Additional SLOs Based MCQs	440
➤ Additional Short Answer Questions	441
➤ Self Assessment	444

13 Halogens

➤ Theory	445
➤ Rack Your Brain (Solutions)	457
➤ Exercise	458
➤ Multiple Choice Questions (MCQs)	458
➤ Short Answer Questions	460
➤ Descriptive Questions	462
➤ Additional SLOs Based MCQs	463
➤ Additional Short Answer Questions	464
➤ Self Assessment	466

14 Atmosphere

➤ Theory	467
➤ Rack Your Brain (Solutions)	482
➤ Exercise	483
➤ Multiple Choice Questions (MCQs)	483
➤ Short Answer Questions	484
➤ Descriptive Questions	485
➤ Additional SLOs Based MCQs	486
➤ Additional Short Answer Questions	487
➤ Self Assessment	490

15 Basic Separation Techniques

➤ Theory	491
➤ Rack Your Brain (Solutions)	502
➤ Exercise	502
➤ Multiple Choice Questions (MCQs)	502
➤ Short Answer Questions	504
➤ Descriptive Questions	506
➤ Additional SLOs Based MCQs	506
➤ Additional Short Answer Questions	507
➤ Self Assessment	508

16 Lab Safety and Practical Skills

➤ Theory	509
➤ Rack Your Brain (Solutions)	520
➤ Exercise	521
➤ Multiple Choice Questions (MCQs)	521
➤ Short Answer Questions	522
➤ Descriptive Questions	524
➤ Additional SLOs Based MCQs	525
➤ Additional Short Answer Questions	526
➤ Self Assessment	528

Chapter

01

PERIODIC TABLE AND PERIODIC PROPERTIES

Student Learning Outcomes

After studying this chapter, students will be able to:

- Explain the arrangement of elements in the periodic table. (Understanding)
- Identify the positions of metals, non-metals and metalloids in the periodic table. (Understanding)
- Explain that the periodic table is arranged into four blocks associated with the four sublevels s, p, d, and f. (Understanding)
- Recognize that the period number (n) is the outer energy level that is occupied by electrons. (Understanding)
- Deduce the electronic configuration of an atom from the element's position on the periodic table, and vice versa (based on s, p, d and f subshells). (Understanding)
- State that the number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table. (Knowledge)
- Deduce the nature, possible position in the Periodic Table and identity of unknown elements from given information about physical and chemical properties. (Understanding)
- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity. (Application)
- Explain that vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity. (Understanding)
- Explain the trends in the ionization energies and electron affinities of the Group 1 and Group 17 elements. (Understanding)
- Recognize that trends in metallic and non-metallic behaviour are due to the trends in valence electrons. (Understanding)
- Suggest the types of chemical bonding present in the chlorides and oxides from observations of their physical and chemical properties. (Understanding)
- Describe (including writing equations for) the reactions, if any, of the oxides (acidic and basic) with water (including the likely pHs of the solutions obtained). (Understanding)
- Explain with the help of equations for, the acid / base behaviour of the oxides and the hydroxides NaOH, Mg(OH)₂ including, where relevant, amphoteric behavior in reactions with acids and bases (sodium hydroxide only) (Understanding)
- Explain with equations for, the reactions of the chlorides with water including the likely pHs of the solutions obtained. (Understanding)
- Explain the variation in the oxidation number of the oxides and chlorides (NaCl, MgCl₂) in terms of their outer shell (valence shell) electrons. (Understanding)
- Write equations for the reactions of Na and Mg with oxygen, chlorine and water. (Application)
- Explain the variations and trends in terms of bonding and electronegativity. (Understanding)

Introduction

- The periodic table of elements is rightly referred as the "Symbol of Chemistry."
- It is a vital and thorough source of chemical knowledge and much more than just a simple chart.
- It would be difficult to explore and comprehend the enormous area of chemistry without its systematic categorization and arrangement of elements.
- 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.
- In modern periodic table, elements are arranged in tabular form in the current periodic table based on their atomic number, electrical configuration, and recurrent chemical characteristics.
- Periodic table provides the foundation for comprehending 118 elements and their properties, making it a crucial tool in chemistry.
- The creation of the periodic table, provides a framework for researching the periodic behaviour.
- It is a significant accomplishment in scientific history.

HISTORICAL BACKGROUND

- Just three centuries ago only less than a dozen (12) elements were known to humanity.
 - By 1700 A.D., only 12 elements-Gold, Silver, Copper, Iron, Lead, Tin, Mercury, Phosphorus, Sulphur, Carbon, Zinc, and Arsenic-were recognized.
 - Over time, more elements were discovered, encouraging scientists to organize them systematically.
- (i) Lavoisier: Up to the end of 18th century, Antoine Lavoisier attempted to classify known elements as metals and nonmetals.
- (ii) Döbereiner: In 1829, Döbereiner grouped the elements into triads (a group of three) with similar properties in such a way that the atomic weight of the middle element was roughly the average of the other two. Examples of such triads include lithium, sodium, and potassium.

Element of Triads	Li	Na	K	Mean Value
Atomic masses of Triads	7	23	39	Atomic mass of Na = $\frac{7+39}{2} = 23$

- (iii) John Newlands: English chemist John Newlands, in 1864, first time observed periodicity in the 62 known elements, noticing that the properties of every eighth element were similar when arranged by the increasing order of their atomic masses.
- He classified the elements into groups so that every eight elements resembled the first element in properties.

- (iv) Mendeleev: In 1869, Russian chemist Dmitri Mendeleev is considered the father of the Periodic Table. He arranged 63 elements into eight vertical columns by increasing atomic mass, aligning elements with similar properties into vertical groups.

The great success of Mendeleev's table was hidden in leaving gaps for undiscovered elements and predicting their atomic mass and properties, which proved accurate when these elements were practically found.

- (v) Luther Meyer: In the same year 1869, Lothar Meyer developed his famous curves by plotting a graph between the atomic weights and atomic volumes of elements. These curves also showed periodicity. These curves also showed periodic periodicity.

Rock Your Brain!

1. Classification of elements in the modern periodic table is based on:
- Law of Triads
 - Law of Octaves
 - Moseley's law
 - Mendeleev's Periodic Law



Dmitri Mendeleev arranged elements according to their atomic masses and his table was the first most notable effort in the classification of elements.

Rock Your Brain!

2. What are the basis of arrangement of elements in Mendeleev's and Moseley's periodic tables?

- (vi) Moseley: In 1913, Moseley determined the exact atomic numbers of known elements using X-ray emission.
- He resolved flaws and discrepancies in Mendeleev's table by arranging the elements on the basis of atomic numbers instead of atomic masses.
 - This significant breakthrough led Moseley to modify the Periodic Law.
 - Moseley or modern periodic law states that "the properties of elements are periodic functions of their atomic numbers".

MODERN PERIODIC TABLE – MAIN FEATURES

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

Main features of the modern periodic table:

- Presently, 118 elements are grouped in the table in ascending order of their respective atomic numbers.
- There are seven (07) horizontal rows called periods and eighteen (18) vertical columns called groups. (In older versions of the table, there were eight (8) vertical groups were divided into two types of groups i.e., 08 A-Groups and 10 B-Groups.
- In the periodic table, elements within the same group show 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.

Note: Other than groups and periods in the periodic table, there are different ways of grouping the elements into various blocks, families and categories in order to enhance understanding.

METALS, NON-METALS AND METALLOIDS

- The elements can be broadly classified into:
 - Metals
 - Nonmetals
 - Metalloids
- Metals: Elements which tend to lose electrons to form positive ions.

Examples: Iron, copper, gold and silver etc.
- Non-metals: Elements which tend to gain electrons to form negative ions.

Examples: Chlorine, Sulphur, phosphorous etc.
- Metalloids: Separates 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 to help in identifying the elements. The line begins at boron (B) and extends down to polonium (Po) including Si, Ge, As, Sb and Te.

NOTE:

- Elements to the left of the line are considered metals.
- Elements just to the right of the line exhibit properties of both metals and non-metals and are termed as metalloids or semimetals.
- Elements to the far right of the periodic table are non-metals. The exception is hydrogen, the first element on the periodic table.

Periodic Table of the Elements

1 H	2 He																	18 Ar	19 K	20 Ca											36 Kr	37 Rb	38 Sr																	54 Xe	55 Cs	56 Ba															86 Rn																									
3 Li	4 Be																	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe	55 Cs	56 Ba	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn	87 Fr	88 Ra	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr																																														

Fig: Modern periodic table

BLOCKS IN PERIODIC TABLE

- The elements in the periodic table can be classified into four blocks based on the subshells containing their valence electrons.

• s-block • p-block • d-block • f-block

For example:

- The valence electrons of elements in the first two groups are in the "s" subshells, placing these elements in the s-block.

Figure: Blocks in periodic table

- In d-block elements, the last electron enters in d-orbital of penultimate (inner to the outermost) shell i.e. it goes to $(n-1)$ d orbital.
- In f-block elements, last electron enters in f-orbital of antipenultimate (inner to penultimate) shell.

- Transition elements belong to the d-block, and the elements in the two series at the bottom of the table (known as Lanthanides and Actinides) are categorized as f-block elements.
- The remaining elements in groups 13 to 18, including the inert (noble) gases in the last group (18), belong to the p-block.

Note: Knowing the block to which an element belongs provides valuable information about its characteristics, chemical reactivity, oxidation states and other properties such as electronegativity and ionization energy, electron filling, etc.

So there are four blocks in the periodic table:

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

Rock Your Brain!

3. In which block of periodic table noble gases are present?
- A) s-block B) p-block
C) d-block D) f-block

FAMILIES IN PERIODIC TABLE

Elements may be categorized according to element families. An element family is a set of elements sharing common properties. There are six famous families of elements in the periodic table:

- Alkali metals (Li, Na, K, Rb, Cs, Fr) → Group-IA
- Alkaline earth metals (Be, Mg, Ca, Ba, Ra) → Group-IIA
- Transition metals (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn) → 3d-Series or 1st Transition Series
- Chalcogens (O, S, Se, Te, Po) → Group-VIA
- Halogens (F, Cl, Br, I, At, Ts) → Group-VIIA
- Noble gases (He, Ne, Ar, Kr, Xe, Rn, Og) → Group-VIIIA

Rock Your Brain!

4. d and f-Block elements are called transition elements. Explain.

i) Alkali Metals (ns^1)

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

- Sodium (Na) and potassium (K) are notable examples of these elements.
- Alkali metals are characterized by:
 - One valence electron
 - Low densities
 - Relatively low melting points
 - Low ionization energies
 - Most reactive metals.

Interesting Information!

Element Livermorium (Lv)
Atomic Number = 116
Atomic Mass = 293 (Most stable isotope)
Group = 16 (Chalcogens)
Period = 7
Electronic configuration
= $[Rn] 5f^{14}6d^{10}7s^27p^4$

ii) Alkaline Earth Metals (ns^2)

The group 2 elements are metals primarily found in the earth and form alkalis that is why they are referred to as alkaline earth metals.

Examples: Beryllium (Be), Magnesium (Mg), Calcium (Ca) etc.

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

iii) Transition Elements (d & f-block elements)

The transition metals make up the largest family of elements in the middle of periodic table. They include four series of d-block elements $(n-1)d^{1-10}ns^{1-2}$, as well as the lanthanides and actinides (f-block elements) $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$ found in the two rows below.

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

iv) Chalcogens ($ns^2 np^4$)

- The group 16 elements are called Chalcogens (Ore Forming Compounds) because most ores of copper (Greek chalkos) are oxides or sulfides.
- In this group, oxygen & Sulphur are non-metals, Se, Te, Po are metalloids and Livermorium (Lv) is a metal.

v) Halogens ($ns^2 np^5$)

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

- Examples: Fluorine, Chlorine, Bromine and Iodine.
- Halogens are highly reactive non-metals with high electron affinities.
- Halogens can easily accept one electron to complete their outermost shell.

vi) Noble Gases ($ns^2 np^6$)

The noble gases are a group of unreactive elements present at the extreme right of the periodic table in Group 18.

- Examples: Helium, Neon, Krypton, Argon etc.
- 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.



Although, noble gases are unreactive, however they have some compounds e.g., compounds of xenon such as xenon hexafluoroplatinate (XePtF_6). Due to the formation of some compounds, 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. The valence electrons of group 1 and 2 are in the s-subshell so they are called s-block elements.

Group 1: ns^1 Group 2: ns^2

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

Ans. Members of Chalcogens:

O, S, Se, Te, Po, Lv (Livermorium).

Characteristics of Chalcogens:

(i) They are ore-forming elements because many metal ores e.g., copper ores contain them in the form of oxides or sulphides e.g., Chalcopyrite (Cu FeS_2).

(ii) Chalcogens have six valence electrons and contain non-metals, metalloids and metals.

PERIODIC ARRANGEMENT AND ELECTRONIC CONFIGURATION

Understanding the periodic arrangement of elements in the periodic table gives valuable insight into their physical properties, such as their physical states and atomic radii, as well as their electronic structures and chemical reactivity.

- Period Number:** The period number indicates the principal quantum number (n), representing the number of electron shells surrounding the nucleus.

Example: an element X in the 3rd period has three electron shells, with its valence electrons located in the 3rd 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.

- Group Number:** The group number indicates the number of valence electrons;

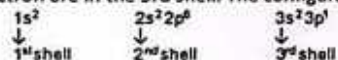
For instance, an element X in the 3rd period and group 2 has two valence electrons in its outermost shell. Thus, the element Z in the 3rd period and group 2 (s-block) has two valence electrons in the 3s subshell, which means that X would be Magnesium (Mg).

- Example to relate both period number and group number:** Example to relate period number and group number with electronic configuration and position of element in the periodic table.

X belongs to group 13 and period 3

In this example, 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:



Understanding the periodic arrangement of elements provides an explanation of an element's electronic configuration, which is essential for understanding its chemical properties and behaviour.

QUICK CHECK 1.2

a) X belongs to group 14 and period 2

i. Write electronic configuration of the element X.

Ans. $X_{(\text{Ground State})} = \frac{1s^2}{1^{\text{st}} \text{ shell}}, \frac{2s^2, 2p^2}{2^{\text{nd}} \text{ shell}}$ and period number 2, it is carbon.

ii. Identify block of the element. Identify this element from periodic table.

Ans. p-block, group 14 (Carbon Group). It is carbon having six electrons.

b) Predict the electronic configuration of an element that is in period 4 and group 17 without consulting the periodic table?

Ans. Group 17 (Halogens), Period 4, it means it is Bromine (Br)

PERIODICITY OF PROPERTIES

Modern Periodic Law:

- The Modern Periodic Law states: "The physical and chemical properties of elements are periodic functions of their atomic numbers."
- The atomic number corresponds to the number of protons in an atom.
- This law is the "cornerstone" of the periodic table, indicating that elements with similar properties appear at certain intervals.
- For example, when elements are arranged by increasing atomic numbers, sodium, potassium, and cesium exhibit many physical and chemical characteristics similar to lithium, as they are all placed in the same group of the table.
- However, due to the gradual increase in the number of protons in the nucleus and the addition of new electron shells, the physical and chemical properties of elements vary systematically within a group and a period.

Atomic Radius

The atomic radius is a measure of the size of an atom. It is half of the distance between two identical atoms bonded together. The atomic radius can vary depending on the type of bond.

OR

It is half of the distance between nuclei of two identical atoms bonded together, whether by covalent, metallic or van der Waals force, depending on the state of the atom.

For example, the radius can be different in a covalent bond compared to an ionic bond.

The atomic radius is typically measured in picometers (pm) or Angstroms (Å).

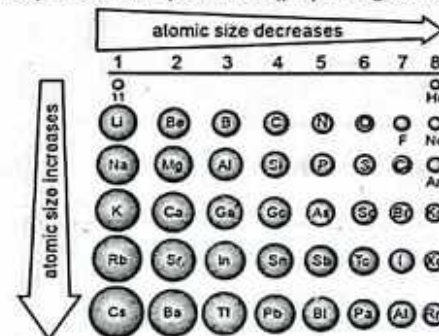


Fig: Variation in atomic radius across periods and down the groups

Periodic trends in atomic radius:

- The factors affecting the atomic radius are:
 - atomic number
 - effective nuclear charge
 - shielding effect of inner electrons
- In a Period:** Generally, 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.
- Down the Group:** The 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).

Rack Your Brain!

5. Which statement is correct?
 A) Na atom is smaller than Na^+
 B) Na atom is larger than K atom
 C) F atom is smaller than F^-
 D) F atom is larger than F^-

Ionic Radius

- The ionic radius is a measure of the size of an ion in a crystal lattice.
- It is typically defined as "the distance from the nucleus of an ion to the outermost electron shell".

- It is measured in picometers (pm) or angstroms (Å).
- When an atom loses one or more electrons to become a positive ion, it generally becomes smaller than the neutral atom. This is because the loss of electrons reduces electronic repulsion and allows the remaining electrons to be pulled closer to the nucleus.
- Contrarily when an atom gains one or more electrons to become an anion, it generally becomes larger than the neutral 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.
- Along a Period:** As we move across a period from left to right, the ionic radius of cations decreases.
- Reason:** It is due to the increasing nuclear charge which pulls the electrons closer. For anions, the ionic radius also decreases across a period because the increasing nuclear charge also pulls the electrons closer to the nucleus.
- Down the Group:** Both cations and anions increase in size as we move down a group.
- Reason:** As the principal quantum number (n) increases, there is an increase in the number of electron shells. As a result, 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.

**Rack Your Brain!**

6. Why the size of cation is always smaller than its parent atom?

In an isoelectronic ions, order of cationic radii.
 $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+}$

Li^+ 60 152	Be^{2+} 31 111		N^{3-} 171 70	O^{2-} 140 66	F^- 136 64
Na^+ 95 186	Mg^{2+} 65 160	Al^{3+} 50 143		S^{2-} 184 104	Cl^- 181 99
K^+ 133 231	Ca^{2+} 99 197	Ga^{3+} 62 122		Se^{2-} 198 117	Br^- 185 114
Rb^+ 148 244	Sr^{2+} 113 215	In^{3+} 81 162		Te^{2-} 221 137	I^- 216 133

Fig.: Variation in Ionic Radius (Ions are coloured red and blue; parent atoms brown. Radii are in picometers)

QUICK CHECK 1.3

a) Which factors affect atomic and ionic radii?

Ans. Factors

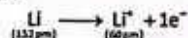
(1) Atomic number (2) Effective nuclear charge, shielding effect

b) Using your knowledge of Period 3 elements, predict and explain the relative sizes of:
 i. the atomic radii of lithium and fluorine

Ans. $\text{Li} > \text{F}$: In a period (Li and F are in period number 2) atomic radii decrease from left to right i.e. from Li to F.

ii. a lithium atom and its ion, Li^+

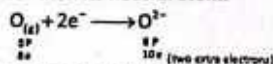
Ans. $\text{Li} > \text{Li}^+$: After the loss of one electron from Li, now Li^+ has now 3 protons but 2 electrons so control of nucleus on remaining electron increases.



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

Ans. $\text{O}^{2-} > \text{O}$

The size of O^{2-} (oxide ions) is greater than its neutral atom.



These extra electrons increase electron-electron repulsion in the atoms outer shell causing the electron cloud to expand. Moreover, the nuclear charge decreases.

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

Ans. $\text{N}^{3-} > \text{F}^-$

N^{3-} and F^- have 10 electrons each and thus are iso-electronic.

Reasons:

(a) Lower Nuclear charge of N^{3-} :

N^{3-} has 10 electrons and 7 protons so it has lower nuclear charge as compared to F^- having 9 protons and 10 electrons.

(b) More electron-electron repulsion in N^{3-} :

N^{3-} has gained three electrons thus the electron-electron repulsion increases causing electron cloud to expand more as compared to F^- .

Ionization Energy

> 1st ionization energy:

"The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1⁺ ions is known as 1st ionization energy (ΔH_{i1})."



> 2nd ionization energy:

"The minimum amount of energy required to remove a second electron from the valence shell of a positive ion is called second ionization energy."

If a second electron is removed from each ion in a mole of gaseous 1⁺ ions, we call it the 2nd ionization energy, ΔH_{i2} . Again, using calcium as an example:



> 3rd ionization energy:

"The minimum amount of energy required to remove third electron from the valence shell of a dipositive ion is called third ionization energy."

Removal of a third electron from each ion in a mole of gaseous 2⁺ ions is corresponding to the 3rd ionization energy. Again, using calcium as an example:



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

(Exercise 1.4.6)

Q. Describe the factors affecting and periodic trends of ionization energy.

> Factors affecting the ionization energy

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

i) Nuclear charge

Greater the effective nuclear charge, greater is the electrostatic force of attraction, more difficult is the removal of an electron from the atom. For this reason, ionization energy increases with an increase in the effective nuclear charge.

Effective nuclear charge is the net effect of nuclear force and screening of electrons and in fact is responsible for controlling ionization energy.

ii) Size of the atom or ion

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.

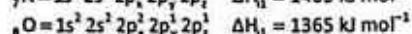
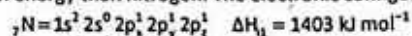
iii) Electronic arrangement

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 fulfilled 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^2np^6).

(b) Exception: Ionization energy of N > O

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



Reason: 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 (Screening) 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.

Oxygen (O) has two spin-paired electrons in its 2p orbital. The ionization energy to remove one of these paired electrons is relatively lower due to the increased repulsion between the paired electrons. Nitrogen (N) has three unpaired electrons in its 2p orbitals. Removing one of these unpaired electrons requires more energy due to the absence of spin pairing repulsion.

➤ Periodic Trends in Ionization Energy

- Down the Group: Going down in a group, the nuclear charge increases but as the size of the atom and the number of electrons causing the shielding effect also increases therefore ionization energy decreases from top to bottom.
- In Group 1, the ionization energies decrease in the following order: $\text{Li} > \text{Na} > \text{K} > \text{Rb} > \text{Cs}$.
- For example, the 6s valence electron of Cs is farther from the nucleus and thus easier to remove compared to the 5s valence electron of Rb.

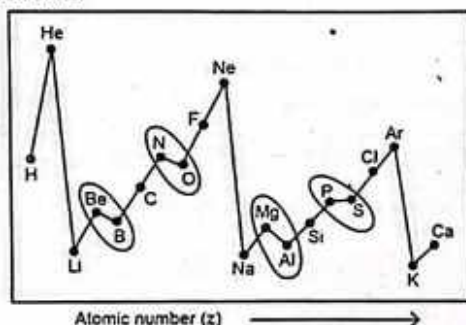


Figure: Variation in Ionization Energies across periods

- Along a Period: As we move from left to right across a period, no. of shells remains unchanged while the effective nuclear charge increases, making it more difficult to remove an electron. Although the number of electrons also increases across a period, the shielding effect within the same shell is same so not considered. Consequently, the ionization energy increases. So, the ionization energy increases.
- The trend of ionization energies of period (1-3) is shown in Figure. The figure shows reveals that noble gases have the highest values of ionization energy due to complete outermost shell in them, the removal of electron is extremely difficult, whereas alkali metals have lowest values of ionization energy.

Rock Your Brain!

7. Ionization energy of calcium is lower than _____ element.
A) Strontium B) Magnesium
C) Barium D) Sodium

Interesting Information!

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

- $24\text{Cr} = [\text{Ar}] 3d^5 4s^1$ (more stable)
- $25\text{Mn} = [\text{Ar}] 3d^5 4s^2$ (less stable)

Interesting Information!

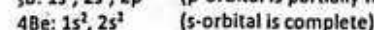
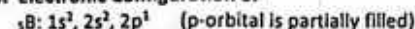
Helium (He) has the highest first ionization energy among all elements in the periodic table.

QUICK CHECK 1.4

a) Explain with reasoning following facts about ionization energy:

i. 1st Ionization energy of Boron is lesser than Beryllium.

Ans. Electronic Configuration of



I.E. of Be > B

Be has greater effective nuclear charge due to its stable 2s² electronic configuration and smaller size making it harder to remove an electron. Boron has lesser ionization energy due to a single electron in 2p-orbital and removal of 2p electron is easier due to higher energy level and shielding effect.

ii. 1st Ionization energy of Aluminum is lower than Magnesium.

Ans. Electronic configuration of



I.E. of Al < Mg

Mg has a stable 3s² electronic configuration making it harder to remove an electron. 3p electron in Al is shielded by the 3s electrons making it easier to remove.

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

Ans. First ionization energy generally decreases from top to bottom in group 3 (B, Al, Ga, In, Tl)

This is because atomic radius increases down the group due to addition of an extra shell successively. So valence electrons are further away from the nucleus, leading to a weaker attraction and lower ionization energy. Moreover, shielding effect increases. In group 3, trend is not smooth.

I.E. of Ga (579 kJ/mol) > I.E. Al (577 kJ/mol) due to poor shielding effect of 3d electrons of Ga.

I.E. of Tl (589 kJ/mol) > I.E. of In (558 kJ/mol) due to poor shielding effect of 4f and 5d in Thallium.

➤ Electron Affinity (H_{EA}°)

First Electron Affinity:

- The first electron affinity, ($\Delta H_{\text{EA}}^\circ$) is 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. Electron affinity of chlorine atom.



- This is amount of energy released when 6.02×10^{23} atoms of chlorine in the gaseous state are converted into $\text{Cl}_{(\text{g})}^-$ ions.
- Since, energy is released, so first electron affinity carries negative sign.

➤ Second Electron Affinity:

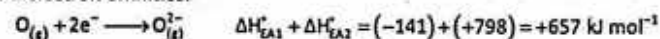
- The second electron affinity, $\Delta H_{\text{EA}}^\circ$ is the amount of energy required to add electrons to 1 mole of uni-negative gaseous ions to form 1 mole of gaseous 2-ions under standard conditions.
- Example: 1st and 2nd Electron Affinities Oxygen: when first electron is added to a neutral oxygen atom, 141 kJ mol⁻¹ energy is released.



- But 798 kJ mol⁻¹ of energy is absorbed on adding second electron to a uni-negative (O^-) ion.



- The net enthalpy change for the formation of the oxide ion (O^{2-}) can be calculated by adding the first and second electron affinities.



Rock Your Brain!

8. Which Group 17 element has the most negative electron affinity?
A) Fluorine B) Chlorine
C) Bromine D) Iodine

Q. Describe the factors affecting and periodic trends of electron affinity.**> Factors affecting electron affinity**

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

i) Size of atom

For small sized atoms the attraction of the nucleus for the incoming electron is stronger.

Thus, smaller is the size of the atom, greater is its electron affinity.

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.

iii) 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.

- **Half-filled p-subshells:** 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.

- **Noble gases:** Noble gases group-18 (VIII-A) 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.

> Periodic trends in electron affinity

- **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 ($At < I < Br < F < Cl$).
- **Along a Period:** Generally, electron affinities become more negative as we move from left to right period. This is firstly due to increase in the nuclear charge, which attracts additional electrons more strongly and secondly due to decreasing atomic radius.

Table: Electron Affinities (kJ/Mol) for Group 1 and Group 17

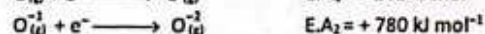
ELEMENT	ELECTRON AFFINITY (kJ/Mol)	ELEMENT	ELECTRON AFFINITY (kJ/Mol)
Fluorine	-328.0	Lithium	-60.0
Chlorine	-349.0	Sodium	-53.0
Bromine	-324.0	Potassium	-48.0
Iodine	-295.0	Rubidium	-47.0
Astatine	-270.1	Cesium	-46.0

**QUICK CHECK 1.5**

Explain with reasoning following facts about electron affinity:

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

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 -ive ion (charge). 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.

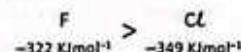


b) Which of nitrogen and phosphorus has the higher electron affinity? Justify with reason.

Ans. The electron affinities of both Nitrogen and Phosphorus are low with value -7 kJ/mol (or zero E.A.) and -72 kJ/mol respectively. So, phosphorus has a more negative electron affinity or in other words greater tendency to accept electrons.

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

Ans. Electron affinity of



Explanation:

Fluorine has unexpectedly lower electron affinity than Chlorine. It may be due to small size of Fluorine atom and its thick electronic cloud. The addition of an extra electron produces high electron density which increases strong electron-electron repulsion. The repulsive forces between electrons result in low electron affinity.

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

Ans. Noble gases have positive electron affinity because their outermost shells are completely filled (e.g., Ne: $1s^2 2s^2 2p^6$). Adding an extra electron forces it into a higher energy level (new shell), which is unfavorable. This process requires energy, so electron affinity is positive (endothermic), unlike most elements where it is negative.

> Electronegativity

Electronegativity is the power of an atom to attract shared pair of electrons toward itself in a molecule. Linus Pauling, an American chemist, developed a scale of dimensionless electronegativity values, which range from just below one for alkali metals to a maximum of four for fluorine. Higher electronegativity values signify a stronger attraction for electrons compared to lower values.

> 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. For example, the electronegativities of halogens in group 17 are in the order:

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

> Periodic Trends in Electronegativity

When we move from left to right along the period, the electronegativity increases this is due to increasing nuclear charge and decreasing size. In the groups, it decreases from top to bottom this is due to the increase in size due to the addition of shells and increasing shielding effect. For example, in the halogen group, the electronegativity value decreases from fluorine (4.0) to iodine (2.5) (as shown in a part of the periodic table in Fig).

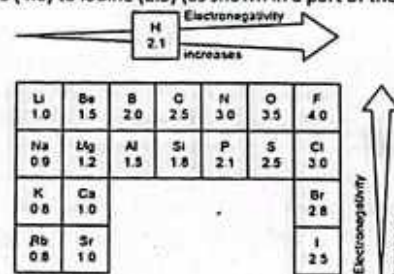


Figure: Variation of electronegativity in groups and periods

Did You Know?

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.

Rack Your Mind!

9. Which element shows the most non-metallic behavior? (Understanding)
A) Lithium B) Magnesium
C) Sulfur D) Fluorine

Generally, metals being on the left side of the periodic table, possess lower electronegativity values than those of non-metals. Hence, metals are electropositive and non-metals are electronegative, relatively. Figure provides a summary of all the variation trends in various physical properties of elements in the periodic table.

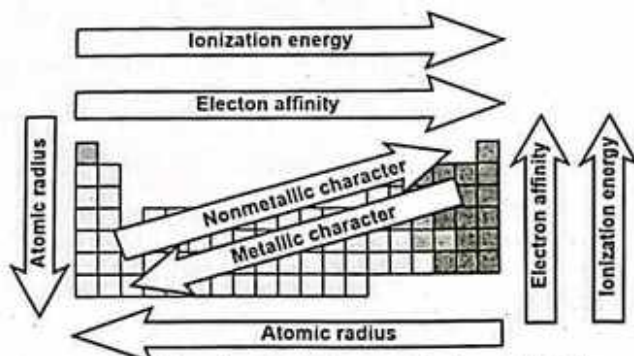


Figure: Trends in various physical properties in the periodic table

Variation in Metallic Character

- The metallic character of elements is typically their tendency to lose electrons.
- It is observed that the elements on the left side of the periodic table have a greater tendency to lose their outermost electrons to achieve noble gas configuration.
- In contrast, elements on the right side of the table tend to gain electrons.
- Thus, elements on the left side of the periodic table are metals that form positive ions, while elements on the right side, particularly in the right corner, are nonmetals that form negative ions.
- Thus the metallic character of an element largely depends on its valence shell electronic configuration.

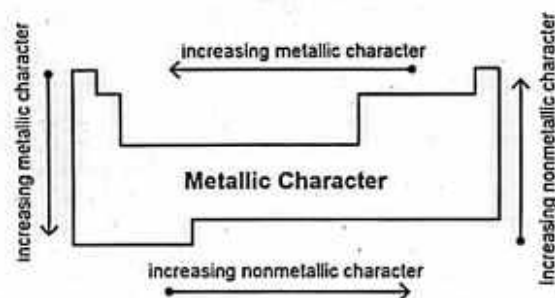


Figure: General trends for the metallic character of elements

- As a result, the metallic character of the elements decreases.
- In other words, the increase in nuclear charge pulls the electron cloud closer to the nucleus, making it more difficult for the atom to lose electrons and thereby decreasing across the period, the nuclear charge increases while the atomic size decreases, which results in stronger attraction to the valence electrons making it difficult. Therefore, metallic character (m.c) decreases from left to right. Its metallic character. Thus, metallic character decreases across a period from left to right.
- On the other hand, metallic character increases as one moves down in a group of the periodic table. This is due to the increase in atomic size and the shielding effect which reduce the nuclear attraction on the valence electrons. The increase in metallic character (ease of losing electron) makes the element more reactive.
- Example: Cesium is far more reactive and electropositive than sodium or lithium.

QUICK CHECK 1.6

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

Ans. In Group 14 (Carbon group) of the periodic table, metallic character increases, down the group.

C = Non-metal
Si = Metalloid
Ge = Metalloid
Sn = Metal
Pb = Metal

Reason: As the atomic size increases and ionization energy decreases, making it easier for atoms to lose electrons. Tendency to lose electrons is one of the key traits of metallic character.

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

Ans. Group 14 = Si and Ge

Group 15 = As and Sb

Group 16 = Te and Po

Semi-metals (Metalloids) have intermediate properties between metals and non-metals.

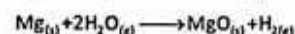
REACTIONS OF Na AND Mg WITH WATER, OXYGEN AND CHLORINE

(Exercise L.O.3)

Q. 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.

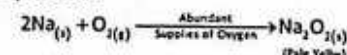
With water

Sodium is more reactive than magnesium towards water. Na reacts vigorously with water to form sodium hydroxide and hydrogen while Mg reacts more slowly in forming magnesium hydroxide and hydrogen. However magnesium reacts with steam more vigorously to make magnesium oxide and hydrogen gas.



With Oxygen

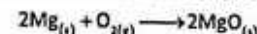
Sodium burns in oxygen with a golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air. It reacts vigorously with oxygen in open air to form peroxide.



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



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



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

With Chlorine

Chlorine reacts with both metals to give soluble salts. It reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed. Magnesium also reacts with chlorine to give white solid, magnesium chloride.



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



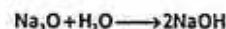
QUICK CHECK 1.7

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

Ans. The oxides and hydroxide of sodium (Na) and Magnesium (Mg) are basic in water.

Sodium (Group IA-Alkali metal)

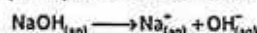
• Oxide: Na_2O



Na_2O reacts strongly with water to form basic oxide.

• Hydroxide: NaOH

NaOH is strong base and it completely dissociates as follows:



pH of solution when dissolved in water = 13 – 14

Magnesium (Group IIA-Alkaline Earth Metal)

• Oxide: MgO



MgO is basic.



Mg(OH)_2 is sparingly soluble in water. pH of solution when dissolved in water = 10

(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: $\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$

Calcium reacts relatively slowly with water to produce calcium hydroxide and hydrogen gas and a white precipitate of calcium hydroxide is formed.

Reaction of Ca with Oxygen: $2\text{Ca} + \text{O}_2 \longrightarrow 2\text{CaO}$

Calcium reacts with oxygen to form calcium oxide.

Rock Your Brain!

10. Write down the Reactions of Na with water, oxygen and chlorine.

TRENDS IN BONDING IN OXIDES AND CHLORIDES OF PERIOD 3

• Oxides of group 1, 2 & 3 (e.g., Na_2O) have more ionic character. These oxides exist as giant ionic lattices with strong electrostatic forces between oppositely charged ions.

• Oxides of group 4, 5, 6 & 7 (e.g., SO_2) are more covalent. These oxides exist as covalent molecules with weak intermolecular forces. This transition is a result of the increasing electronegativity and decreasing ionic character.

• Chlorides of group 1, 2 and 3 (e.g., NaCl) are predominately ionic.

• Chlorides of elements from group 4, 5, 6 and 7 (e.g., PCl_5) are covalent. The covalent character in chlorides increases due to decrease in difference of electronegativity between the halogen and the other atom.

Acidic Strength
 $\text{SO}_3 > \text{SO}_2$
 $\text{N}_2\text{O} > \text{NO} > \text{N}_2\text{O}_5$
 $\text{N}_2\text{O}_4 > \text{N}_2\text{O}_5$
 Increase in oxidation state increases acidic strength.

Q. Explain with the help of equations acidic and basic behaviour of oxides and chlorides.

> Classification of Oxides

i) Oxides

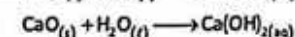
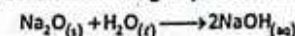
Oxides are binary compounds formed by the reaction of oxygen with other elements. Oxides are classified into neutral, amphoteric and basic or acidic based on their characteristics.

ii) Basic Oxides

A basic oxide is an oxide that when combined with water gives off an alkali. Metals react with oxygen to give basic oxides. These oxides are usually ionic in nature. The elements of group 1 and 2 form basic oxides when react with oxygen.

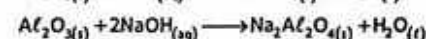
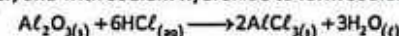
Examples: Na_2O , CaO , BaO .

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



iii) Amphoteric Oxides

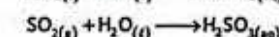
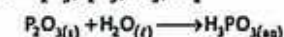
Amphoteric oxides are oxides that can react with both acids and bases. So they have the ability to behave as either an acid or a base, depending on the conditions. Aluminum oxide (Al_2O_3) is insoluble in water but reacts with hydrochloric acid to form aluminium chloride and water, and with sodium hydroxide to form sodium aluminate and water.



iv) Acidic Oxides:

An acidic oxide is an oxide that when combined with water gives off an acid. Non-metals react with oxygen to form acidic oxides which are held together by covalent bonds. Silicon dioxide is acidic oxide as it can react with bases.

Examples: Acidic oxides in period 3 = P_2O_3 , P_2O_5 , SO_3 , SO_2



Reactions of these oxides with bases are given below:

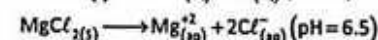
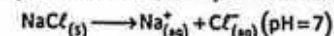


> Classification of Chlorides

Chlorine forms compounds with other elements known as chlorides. These chlorides show characteristic behaviour when we add them into water, resulting in solutions that can be acidic or neutral.

i) Neutral Chlorides

Neutral chlorides are salts that, when dissolved in water, produce a neutral solution with a pH close to 7. At the start of period 3, chloride 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. For example,



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

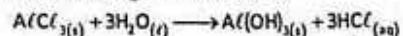
Rock Your Brain!

11. BeO is amphoteric oxide. Justify it by giving reactions.

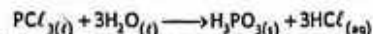
Examples of Amphoteric Oxides:
 BeO , ZnO , Bi_2O_3 etc.

ii) Acidic Chlorides:

If we move in period 3, from Aluminium to Sulphur, all chlorides react with water to make acidic solution with pH less than 7. This process is called hydrolysis, when $AlCl_3$ is added to water, aluminium and chloride ions in solution. Al^{3+} ion is hydrated and causes a water molecule to lose an H^+ ion, this process is hydrolysis. This turns the solution acidic. The following reaction occurs:



Other examples of acidic chlorides are given below.

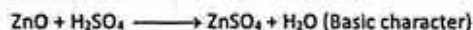


QUICK CHECK 1.8

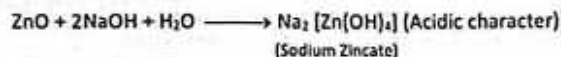
- (a) ZnO reacts with HCl to give $ZnCl_2$ and with NaOH to give Na_2ZnO_2 . Give equations and also predict the type of this oxide?

Ans. ZnO is amphoteric in nature. It is explained by following two reactions.

When ZnO reacts with an acid:



When ZnO reacts with a base:



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

Ans. when $AlCl_3$ is added to water, aluminium and chloride ions in solution. Al^{3+} ion is hydrated and causes a water molecule to lose an H^+ ion, this process is hydrolysis.

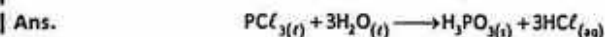


$NaCl$ is a neutral salt with pH about 7.



Strong acid and base react here to form neutral salt and water.

- c) Predict whether the chlorides PCl_5 , PCl_3 would be acidic or basic, give reason.



Both PCl_3 and PCl_5 in water show acidic in nature.

- 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 will react with NaOH.



VARIATION IN OXIDATION NUMBER IN OXIDES AND CHLORIDES

Oxidation Number:

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

- The oxidation number is also referred to as the oxidation state.
 - In ionic compounds the oxidation number of an atom is defined as the charge which appears on the ions.
- Let's examine the oxidation numbers in oxides and chlorides of the third period.
- The oxidation number of an element of 3rd Period in its oxide or chloride corresponds to the number of electrons used for bonding and 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. Consider the following table for oxidation states of various elements of the periodic table.
 - In the oxides, the oxidation number increases from +1 in Na to +6 in S. In chlorides, the oxidation number increases from +1 in Na to +5 in P.
 - Phosphorus and sulphur show several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals. For instance, in SO_2 , sulphur has an oxidation number of +4 because only four electrons are used for bonding, while in SO_3 , sulphur has an oxidation number of +6 because all six electrons are used for bonding.

Table: Oxidation Numbers in Oxides and Chlorides of 3rd Period elements

Oxide	Oxidation Number	Chloride	Oxidation Number
Na in Na_2O	+1	Na in $NaCl$	+1
Mg in MgO	+2	Mg in $MgCl_2$	+2
Al in Al_2O_3	+3	Al in $AlCl_3$	+3
Si in SiO_2	+4	Si in $SiCl_4$	+4
P in P_4O_{10} / P in P_4O_6	+5/+3	P in PCl_5	+5
S in SO_3	+6	P in PCl_3	+3
S in SO_2	+4	S in SCl_2	-2

QUICK CHECK 1.9

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

Ans. SO_2

$$\{\text{Oxidation Number of S}\} + 2\{\text{Oxidation number of O}\} = 0$$

$$S + 2(-2) = 0$$

$$S - 4 = 0$$

$$S = +4$$

SO_3

$$\{\text{Oxidation Number of S}\} + 3\{\text{Oxidation number of O}\} = 0$$

$$S + 3(-2) = 0$$

$$S - 6 = 0$$

$$S = +6$$

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

Ans. Some p-block elements with higher atomic number show variable valence due to inert pair effect in p-block elements.

- Group IIIA: $ns^2 np^1$
- Group IIIA: $ns^2 np^1$
- Sn and Pb shows +2 and +4 oxidation states.



Sr. #	Option	Explanation
1.	C	Classification of elements in the modern periodic table is based on Moseley's law
2.	S.Q	Mendeleev's periodic table was arranged on the basis of increasing atomic mass. Moseley's periodic table was arranged on the basis of increasing atomic number.
3.	B	Noble gases are present on the extreme right side of periodic table in p-block.
4.	S.Q	The d-block and the f-block elements are called transition elements because they are placed 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.
5.	C	F atom is smaller than F^-
6.	S.Q	A cation is formed by the loss of one or more electrons from a neutral atom. The size of cation is always smaller than its parent atom. Reasons: (i) There is imbalance of electron-proton ratio (No. of Protons are greater than no. of electrons). (ii) Nucleus holds the remaining electrons with a stronger force. (iii) Usually loss of shell occurs during removal of electron/s. Examples: $\begin{array}{ccc} \text{Li} & \longrightarrow & \text{Li}^+ + e^- \\ 123 \text{ pm} & & 60 \text{ pm} \\ \text{Na} & \longrightarrow & \text{Na}^+ + e^- \\ 186 \text{ pm} & & 95 \text{ pm} \end{array}$
7.	B	The 1 st ionization energy of Mg = 738 kJ/mol, while Calcium = 590 kJ/mol approx.
8.	B	$\text{Cl}_{(g)} + e^- \longrightarrow \text{Cl}_{(g)}^- \quad \Delta H_{\text{ex}}^\circ = -348.8 \text{ kJ mol}^{-1}$
9.	D	Fluorine is highly electronegative with strong non-metallic character.
10.	S.Q	$2\text{Na}_{(s)} + 2\text{H}_2\text{O}_{(l)} \longrightarrow 2\text{NaOH}_{(aq)} + \text{H}_{2(g)}$ $2\text{Na}_{(s)} + \text{O}_{2(g)} \xrightarrow[\text{Supplies of Oxygen}]{\text{Abundant}} \text{Na}_2\text{O}_{2(s)} \quad (\text{Pale Yellow})$ $2\text{Na}_{(s)} + \text{Cl}_{2(g)} \longrightarrow 2\text{NaCl}_{(s)}$
11.	S.Q	BeO – An Amphoteric Oxide: "An oxide which acts as a base when reacts with strong acid and as an acid when reacts with strong base is called an amphoteric oxide." BeO is an amphoteric oxide. - With acids, BeO behaves as a base and gives salt and water. $\text{BeO} + \text{H}_2\text{SO}_4 \longrightarrow \text{BeSO}_4 + \text{H}_2\text{O}$ - With bases, BeO just like acids. $\text{BeO} + 2\text{NaOH} \longrightarrow \text{Na}_2\text{BeO}_2 + \text{H}_2\text{O}$

Exercise

MULTIPLE CHOICE QUESTIONS (MCQs)

Q.1 Four choices are given for each question. Select the correct choice.

- Which scientist first time observed the periodicity in the elements?
 a) J. Newlands b) L. Meyer c) J.W. Döbereiner d) D. I. Mendeleev
- 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
- 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.
- Which property increases as you go down a group in the periodic table?
 a) Atomic radius b) Electron Affinity c) Electronegativity d) Ionization energy
- 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
- Which of the following atoms show more than one (variable) oxidation states?
 a) Sodium b) Magnesium c) Aluminum d) Phosphorous
- 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.
- What is the oxidation state of sulphur in the sulfate ion (SO_4^{2-}).
 a) +4 b) +2 c) +6 d) 0
- 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.
- The atomic radius generally across a period in the periodic table.
 a) Increases b) Decreases c) Remains constant d) First increases then decreases
- Which one of the following elements has the highest ionization energy?
 a) Sodium (Na) b) Magnesium (Mg) c) Aluminium (Al) d) Argon (Ar)

Answer Key with Explanations

Sr.No.	Option	Answer	Explanation
I.	a	J. Newlands	- John Newlands proposed the Law of Octaves (1864), noting periodicity in properties every 8th element. - Mendeleev later formalized the periodic table (1869).
II.	a	Magnesium	3 electron shells ($n = 3$) + 2 outer electrons (s^2) → Group 2 (alkaline earth metals). - Calcium: [Ar] $4s^2$ ($n = 4$) - Magnesium [Ne] $3s^2$. The correct answer is (c) Magnesium ($n = 3$, 2 valence electrons).

III.	b	It increases from top to bottom in a group	• Metallic character increases down a group as atomic size grows, reducing ionization energy.
IV.	a	Atomic radius	• Down a group, atomic radius increases due to additional electron shells.
V.	a	Smaller atom and greater nuclear charge	• Ionization energy increases with smaller size (stronger attraction) and higher nuclear charge.
VI.	d	Phosphorous	• P shows variable oxidation states (-3, +3, +5) due to vacant d-orbitals. • Na/Mg/Al have fixed states (+1, +2, +3).
VII.	a	It becomes less negative from top to bottom	• Electron affinity decreases (less negative) down a group as atomic size increases, reducing attraction for added electrons.
VIII.	c	+6	• SO_4^{2-}: Let oxidation state of S = x • $x + 4(-2) = -2 \rightarrow x = +6$
IX.	b	It increases from left to right across a period	• Electronegativity increases left to right due to higher nuclear charge and smaller atomic size.
X.	b	Decreases	• Across a period, atomic radius decreases due to increasing nuclear charge pulling electrons closer.
XI.	d	Argon (Ar)	• Noble gases (e.g., Ar) have the highest ionization energy in a period due to stable full valence shells.

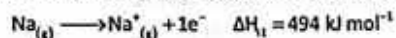
SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

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

Ans. 1st ionization energy:

"The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1⁺ ions is known as 1st ionization energy (ΔH_1)."



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

Ans. Sulphur has a lower first ionization energy than phosphorus due to electron-electron repulsion in its 3p orbitals, which makes it easier to remove an electron.

Electronic Configuration:

- Phosphorus: $1s^2 2s^2 2p^6 3s^2 3p^3$
- Sulphur: $1s^2 2s^2 2p^6 3s^2 3p^4$

Half-filled stability (Phosphorus):

- Phosphorus has a half-filled 3p subshell ($3p^3$), which is relatively stable due to symmetrical distribution and minimized electron repulsion.
- This stability makes it harder to remove one of its electrons.

Increased electron repulsion in Sulphur:

- Sulphur has one more electron than phosphorus, giving it a $3p^4$ configuration.
- In this case, one of the 3p orbitals contains a pair of electrons, which experience increased repulsion.
- This repulsion makes it easier to remove one of the paired electrons, hence lowering the ionization energy.

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

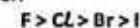
Ans. In Group 13 to 17 elements, the last electron enters in p-orbital of their respective shells. That is why they are called p-block elements e.g., B, C, N, O, F.

d. 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. For example, the electronegativities of halogens in group 17 are in the order:



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

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

Ans. Down the group (in alkali metals), Atomic size increases and ionization energy decreases. It results in greater tendency to form positive ions and they react more readily.

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

Ans. The variability in oxidation numbers among elements is primarily due to differences in electronic configurations, atomic structure, and position in the periodic table. Here's a breakdown of why some elements show variable oxidation states while others do not:

1. Transition Metals and Variable Oxidation States

- Transition metals (like iron, copper, manganese) show variable oxidation states because they have partially filled d-orbitals.
- Their outer electron configuration typically ends in $(n-1)d^{1-10}ns^{0-2}$, so both s- and d-electrons can be involved in bonding.
- Example: Iron (Fe) can lose 2 electrons to become Fe^{2+} or 3 electrons to become Fe^{3+} .

2. Relativistic and Inert Pair Effects (In heavier elements)

- Heavier p-block elements (like Sn, Pb) show variable oxidation states due to the inert pair effect, where the s-electrons are less available for bonding due to relativistic stabilization.
- Example: Lead (Pb) shows both +2 and +4 oxidation states.

3. Main Group Elements and Fixed Oxidation States

- Group 1 and Group 2 metals (like Na, Mg) almost always show a fixed oxidation number (+1 and +2 respectively) because they lose their valence s-electrons easily, and further loss would require breaking into a stable inner shell, which is energetically unfavorable.
- Similarly, nonmetals like fluorine always have an oxidation number of -1 because it is the most electronegative element and can only gain one electron to complete its octet.

4. Stability of Electronic Configurations

- Some elements achieve stable noble gas configurations after gaining or losing a specific number of electrons. Once this stability is reached, further changes are energetically unfavorable.
- For example, oxygen typically forms -2 oxidation state to complete its octet.

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

Ans. Antimony (Sb)

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

Ans. Oxides of sodium (Na) and magnesium (Mg) are more ionic than the oxides of nitrogen (N) and phosphorus (P) due to differences in electronegativity, metallic character, and bonding tendencies of these elements.

1. Nature of Elements (Metals and Non-metals)

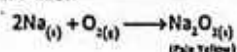
- Sodium and magnesium are metals (from Groups 1 and 2). Metals tend to lose electrons easily and form cations.

- Nitrogen and phosphorus are non-metals (Group 15). Non-metals prefer to share electrons rather than form ions, leading to covalent bonding.
2. **Electronegativity Difference**
- Ionic character is stronger when there is a large electronegativity difference between the metal and oxygen.
 - Sodium (0.93) and magnesium (1.31) have much lower electronegativities than oxygen (3.44), resulting in highly polar/ionic bonds.
 - Nitrogen (3.04) and phosphorus (2.19) have closer electronegativity values to oxygen, leading to covalent bonds.
3. **Bond Formation Tendencies**
- Na_2O and MgO form ionic lattices, with strong electrostatic attraction between ions (e.g., Na^+ and O^{2-}).
 - N_2O_5 and P_4O_{10} , oxides of nitrogen and phosphorus, exist as molecular compounds with covalent bonds (electron sharing).

I. Give reason for the different chemical reactivities of Na and Mg toward oxygen and chlorine.

Ans. With Oxygen

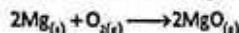
Sodium burns in oxygen with a golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air. It reacts vigorously with oxygen in open air to form sodium peroxide.



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

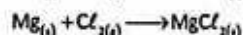


Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide.



With Chlorine

Chlorine reacts with both metals to give soluble salts. It reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed. Magnesium also reacts with chlorine to give white solid, magnesium chloride.



J. 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 much lower than that of helium because of the following key factors:

- 1. Electron Configuration and Shell Structure**
Helium has the electron configuration: $1s^2$
Both electrons are in the first energy level, very close to the nucleus and strongly held.
Lithium has the electron configuration: $1s^2 2s^1$
The outermost electron (2s) is in the second energy level, which is farther from the nucleus and shielded by the inner 1s electrons.
- 2. Electron Shielding Effect**
In lithium, the two inner 1s electrons shield the outer 2s electron from the full +3 nuclear charge. As a result, the effective nuclear charge felt by the 2s electron is significantly less than +3, it is roughly around +1.
In helium, there's no shielding, both electrons are strongly attracted to the nucleus.
- 3. Distance from the Nucleus**
The 2s electron in lithium is farther from the nucleus compared to helium's 1s electrons. Greater distance = weaker attraction = easier to remove = lower ionization energy.

k. 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. Beryllium (Be) has a higher ionization energy than boron (B) despite having a lower nuclear charge.

Beryllium (Be): Atomic number = 4, Electronic configuration: $1s^2 2s^2$

Boron (B): Atomic number = 5, Electronic configuration: $1s^2 2s^2 2p^1$

- In Be, the electron is removed from a 2s orbital, which is lower in energy, more penetrating, and closer to the nucleus.
- In B, the electron is removed from a 2p orbital, which is higher in energy, farther from the nucleus, less penetrating and more shielded by the inner 1s and 2s electrons. So, even though B has a higher nuclear charge (+5 vs. +4), the 2p electron is easier to remove than the 2s electron in Be.
- So, 2p orbitals are less tightly bound than 2s orbitals, despite being in the same principal energy level ($n = 2$). This makes the first ionization energy of B lower, because you're removing an electron that's already less stable.

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

Ans. $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{Ne} < \text{F}^-$

They all have 10 electrons each so they are iso-electronic species.

As the nuclear charge increases (with constant number of electrons), the electron cloud is pulled in tighter, and the radius decreases.

Al^{3+} : Highest nuclear charge with 10 electrons $\rightarrow$ most tightly held $\rightarrow$ smallest

F^- : Lowest nuclear charge with 10 electrons $\rightarrow$ weakest pull on electrons $\rightarrow$ largest

m. 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 behaviour when dissolved in water.

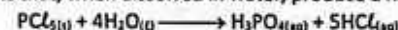
Ans.

(i) Neutral = NaCl , Basic = MgCl_2 , Acidic = PCl_5

(ii) $\text{NaCl}_{(s)} \longrightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} \quad (\text{pH} = 7)$

$\text{MgCl}_{2(s)} \longrightarrow \text{Mg}^{2+}_{(aq)} + 2\text{Cl}^-_{(aq)} \quad (\text{pH} = 6.5)$

- Neutral chlorides are salts that, when dissolved in water, produce a neutral solution with a pH close to 7.



- PCl_5 reacts with water to produce phosphoric acid and hydrochloric acid i.e. produces acidic solution.

DESCRIPTIVE QUESTIONS

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.

Ans. See Page No. (15)

Q.4 Explain with the help of equations acidic and basic behaviour of oxides and chlorides.

Ans. See Page No. (17)

Q.5 Describe the factors affecting and periodic trends of electron affinity.

Ans. See Page No. (12)

Q.6 Define ionization energy. Discuss the factors affecting and periodic trends of ionization energy.

Ans. See Page No. (09)

ADDITIONAL SLOs BASED MCQs

- Livermorium (Lv) is an element which is present in:
 - Group 1, Period 2
 - Group 2, Period 3
 - Group 16, Period 7
 - Group 17, Period 7
- The period number of an element shows:
 - Group number
 - Number of valence electrons
 - Type of block
 - Number of occupied shells
- Which of the following has the configuration $1s^2 2s^2 2p^6 3s^2 3p^1$?
 - Na
 - Mg
 - Al
 - Si
- The element in group 2 and period 4 has which configuration?
 - [Ar] $3d^1$
 - [Ar] $4s^2$
 - [Ar] $4s^1$
 - [Ne] $3s^2$
- An unknown gas forms a -1 ion and is highly reactive. It likely belongs to:
 - Group 1
 - Group 14
 - Group 17
 - Group 18
- An element reacts with water to form a strong base and hydrogen gas. It is likely:
 - A noble gas
 - An alkali metal
 - A halogen
 - A metalloid
- Which property increases down Group 1 elements?
 - Electronegativity
 - Reactivity
 - Ionization energy
 - Non-metallic character
- Why do Group 1 elements have low ionization energy?
 - Small size
 - High nuclear charge
 - Large atomic radius
 - High shielding
- Which factor mainly determines an element's metallic character?
 - Number of neutrons
 - Atomic mass
 - Valence electrons
 - Density
- Which oxide forms a basic solution in water?
 - SO_2
 - CO_2
 - MgO
 - P_2O_5
- Which of the following oxides is acidic in nature?
 - CaO
 - Na_2O
 - SO_2
 - MgO
- Which hydroxide is slightly soluble and weakly basic?
 - NaOH
 - $Mg(OH)_2$
 - $Ca(OH)_2$
 - KOH
- Which chloride forms an acidic solution in water due to hydrolysis?
 - NaCl
 - $AlCl_3$
 - KCl
 - $MgCl_2$
- Which equation correctly shows magnesium reacting with chlorine?
 - $Mg + Cl \rightarrow MgCl$
 - $Mg + Cl_2 \rightarrow MgCl_2$
 - $Mg + 2Cl \rightarrow MgCl_2$
 - $Mg + 2Cl_2 \rightarrow MgCl_4$
- Why is MgO more ionic than Al_2O_3 ?
 - Al is more reactive
 - Mg has fewer electrons
 - Mg is less electronegative than Al
 - Al has more orbitals

Answer Key with Explanations

Sr.No.	Option	Explanations
1.	C	Livermorium has atomic no.116 and is present in group 16 and period 7.
2.	D	Period number equals the number of electron shells.
3.	C	Aluminum has 13 electrons; the configuration matches.
4.	B	Group 2 elements end with s^2 ; $4s^2$ indicates period 4.
5.	C	Group 17 elements (halogens) form -1 ions and are reactive.
6.	B	Group 1 metals react with water to form strong bases and H_2 .
7.	B	Alkali metals become more reactive down the group.
8.	C	Larger atoms with shielding make electron removal easier.
9.	C	Fewer valence electrons = more metallic behavior.
10.	C	$MgO + H_2O \rightarrow Mg(OH)_2$, a basic solution.

11.	C	SO_2 is an acidic oxide that forms H_2SO_4 in water.
12.	B	$Mg(OH)_2$ is sparingly soluble and weakly basic.
13.	B	Al^{3+} causes hydrolysis, lowering pH of solution.
14.	B	Magnesium forms $MgCl_2$ with Cl_2 gas.
15.	C	Greater difference in electronegativity makes MgO more ionic.

ADDITIONAL SHORT ANSWER QUESTIONS

Q.1 How are elements arranged in the modern periodic table?

Ans. Elements in the modern periodic table are arranged in order of increasing atomic number. Elements with similar chemical properties are placed in the same group, and the pattern of properties repeats periodically.

Q.2 What do the s, p, d, and f blocks in the periodic table represent?

Ans. These blocks (s, p, d, and f) represent the type of atomic orbital being filled by the valence electrons. The s-block includes Groups 1 and 2, the p-block includes Groups 13 to 18, the d-block contains transition metals, and the f-block includes lanthanides and actinides.

Q.3 What does the period number of an element indicate?

Ans. The period number indicates the number of occupied electron shells in an atom of the element. For example, an element in Period 3 has electrons in three energy levels.

Q.4 We determine the electron configuration of an element using its position on the periodic table. How?

Ans. The element of the period number represents the highest principal energy level, and the block shows the type of orbital. The group number of the element helps to determine the number of electrons in the outermost shell.

Q.5 How can we identify an unknown element based on its physical and chemical properties?

Ans. We can identify an unknown element based on its physical and chemical properties by comparing the properties of an element such as reactivity, ion formation, and bonding behavior with known trends in the periodic table, we can estimate its possible group and period and identify it.

Q.6 How can we predict the properties of elements in the same group?

Ans. Elements in the same group have the same number of valence electrons, so they exhibit similar chemical properties and reactions. For example, Group 1 elements all react with water to form hydroxides and release hydrogen gas.

Q.7 How do atomic radius and ionization energy change across a period and down a group?

Ans. Across a period: The atomic radius decreases and ionization energy increases across a period due to greater nuclear attraction on the valence shell electrons.

Down a group: The atomic radius increases and ionization energy down the group decreases due to added electron shells and increased shielding. So nuclear attractions decreases.

Q.8 How does the number of valence electrons affect metallic and nonmetallic character?

Ans. An element with fewer valence electrons (like 1 to 3) tend to lose them easily and show metallic character, on the other hand, those elements with more valence electrons (like 5 to 7) gain electrons and show nonmetallic character.

Q.9 How do oxides of metals and non-metals behave in water?

Ans. The metal oxides usually react with water to form basic solutions, on the other hand, non-metal oxides form acidic solutions. For example, Na_2O forms NaOH, and SO_3 forms H_2SO_4 .

