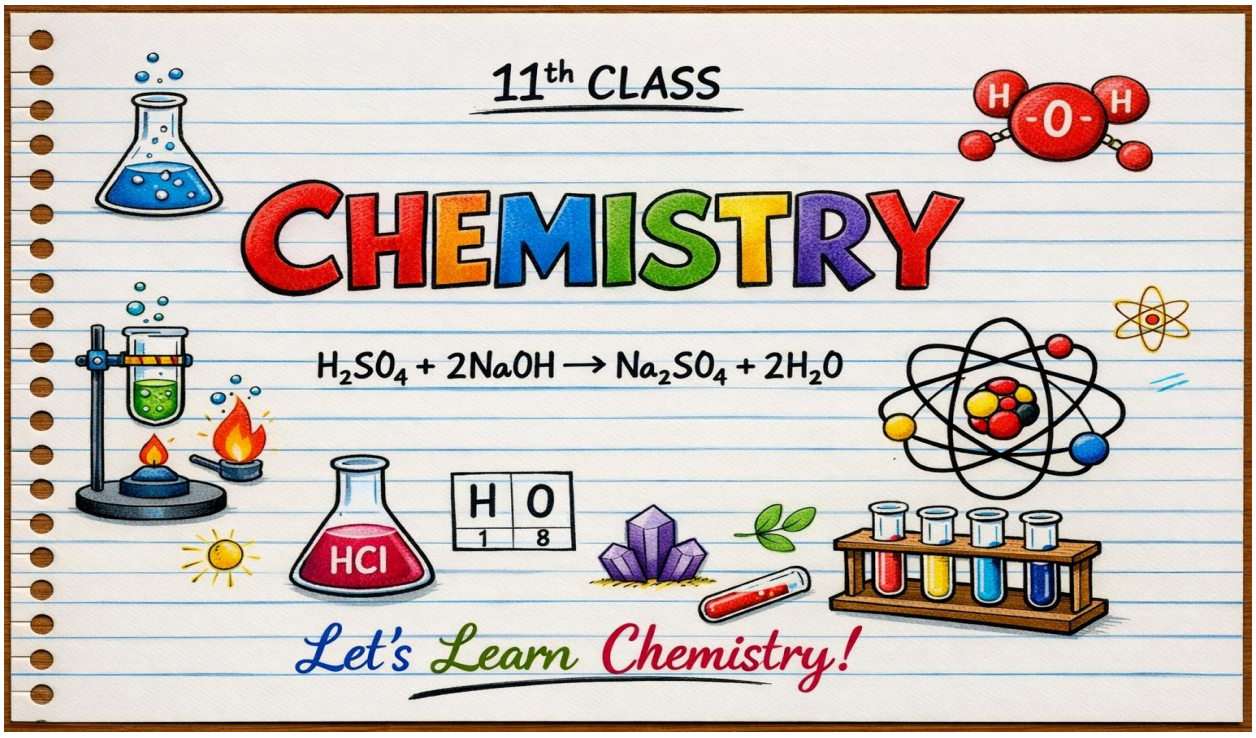


# Class: 11th

# Subject: Chemistry

# Unit 1: PERIODIC TABLE AND PERIODIC PROPERTIES

## ❖ Important MCQs:

**1. Which statement best explains the importance of the modern periodic table?**

(a) It arranges elements randomly

(b) It relates properties with atomic number ✅

(c) It is based on density

(d) It ignores chemical properties

**2. Why was Mendeleev's periodic table considered more successful than earlier classifications?**

(a) It used atomic number

(b) It included only metals

(c) It predicted undiscovered elements correctly ✅

(d) It ignored periodicity

**3. What was the major limitation of Newlands' Law of Octaves?**

(a) It worked for all elements

(b) It was based on atomic number

(c) It failed beyond calcium and grouped dissimilar elements together ✅

(d) It ignored repetition

**4. Which of the following best describes Döbereiner's triads?**

(a) Elements with equal masses

(b) Elements with similar properties and related atomic masses ✅

(c) Random grouping

(d) Based on atomic number

**5. Why did Moseley modify Mendeleev's periodic law?**

(a) To remove metals

(b) To correct arrangement based on atomic number ✅

(c) To reduce elements

(d) To change groups

**6. Which pair of elements best represents Döbereiner's triad concept?**

(a) Na, Mg, Al

(b) Li, Na, K ✅

(c) C, N, O

(d) Fe, Co, Ni

**7. What was the basis of Lothar Meyer's periodicity?**

(a) Atomic number

(b) Chemical reactions

(c) Graph between atomic weight and atomic volume ✅

(d) Electron shells

**8. Which scientist first showed periodicity through graphical representation?**

(a) Mendeleev

(b) Moseley

(c) Lothar Meyer ✅

(d) Newlands

**9. The modern periodic law states that properties of elements are periodic functions of:**

(a) Atomic mass

(b) Atomic number ✅

(c) Atomic size

(d) Valency

**10. Which statement is TRUE about Mendeleev's table?**

(a) No gaps were left

(b) Elements were arranged by atomic number

(c) Some elements were placed out of order to match properties ✅

(d) It included 118 elements

**11. Why is Moseley's work considered a turning point?**

(a) He discovered metals

(b) He introduced triads

(c) He established atomic number as fundamental property ✅

(d) He removed periodicity

**12. Which of the following is NOT a feature of Newlands' law?**

(a) Based on atomic mass

(b) Repetition after 8 elements

(c) Applicable to all known elements ✅

(d) Law of Octaves

**13. What is the key reason periodicity occurs in elements?**

(a) Atomic mass

(b) Electron configuration repetition ✅

(c) Density

(d) Volume

**14. Which classification was the earliest scientific attempt?**

(a) Mendeleev

(b) Moseley

(c) Lavoisier ✅

(d) Meyer

**15. What was the major improvement in modern periodic table over Mendeleev's table?**

(a) Removal of elements

(b) Arrangement by atomic number instead of atomic mass ✅

(c) No groups

(d) No periodicity

**16. The modern periodic table arranges elements in increasing order of:**

(a) Atomic mass

(b) Atomic number ✅

(c) Density

(d) Volume

**17. Total number of periods in the modern periodic table are:**

(a) 6

(b) 7 ✅

(c) 8

(d) 18

**18. Total number of groups in the modern periodic table are:**

(a) 8

(b) 16

(c) 18 ✅

(d) 20

**19. Elements in the same group have similar chemical properties because they have:**

(a) Same atomic mass

(b) Same number of shells

(c) Same number of valence electrons ✅

(d) Same density

**20. Moving from top to bottom in a group, elements show change in:**

(a) Chemical properties

(b) Physical properties ✅

(c) Atomic number

(d) Valence electrons

**21. Across a period (left to right), elements show:**

(a) No change

(b) Sudden change

(c) Gradual change in properties ✅

(d) Random change

**22. Metals generally tend to:**

(a) Gain electrons

(b) Share electrons

(c) Lose electrons to form positive ions ✅

(d) Remain neutral

**23. Non-metals usually:**

(a) Lose electrons

(b) Gain electrons to form negative ions ✅

(c) Form positive ions

(d) Do not react

**24. Metalloids are best described as elements that:**

(a) Are only metals

(b) Are only non-metals

(c) Show properties of both metals and non-metals ✅

(d) Are gases only

**25. The “stair-step line” in periodic table separates:**

(a) Metals and gases

(b) Metals and metalloids

(c) Metals and non-metals ✅

(d) Groups and periods

**26. Elements are classified into blocks based on:**

(a) Atomic mass

(b) Number of shells

(c) Subshell of valence electrons ✅

(d) Density

**27. Elements of group 1 and 2 belong to:**

(a) p-block

(b) d-block

(c) s-block ✅

(d) f-block

**28. Transition elements belong to:**

(a) s-block

(b) p-block

(c) d-block ✅

(d) f-block

**29. Lanthanides and Actinides are part of:**

(a) s-block

(b) p-block

(c) d-block

(d) f-block ✅

**30. Elements in groups 13 to 18 belong to:**

(a) s-block

(b) d-block

(c) p-block ✅

(d) f-block

**31. The block of an element helps to determine its:**

(a) Color only

(b) Chemical properties and reactivity ✅

(c) Size only

(d) Mass only

**32. Alkali metals belong to which group?**

(a) Group 2

(b) Group 17

(c) Group 1 ✅

(d) Group 18

**33. Alkali metals have how many valence electrons?**

(a) 1 ✅

(b) 2

(c) 3

(d) 4

**34. Alkali metals are:**

(a) Least reactive

(b) Non-metals

(c) Most reactive metals ✅

(d) Noble gases

**35. Alkaline earth metals belong to:**

(a) Group 1

(b) Group 2 ✅

(c) Group 16

(d) Group 17

**36. Alkaline earth metals have:**

(a) 1 valence electron

(b) 2 valence electrons ✅

(c) 3 valence electrons

(d) 4 valence electrons

**37. Transition metals usually show:**

(a) Fixed oxidation state

(b) Variable oxidation states ✅

(c) No reactivity

(d) Only gaseous state

**38. Chalcogens belong to:**

(a) Group 15

(b) Group 16 ✅

(c) Group 17

(d) Group 18

**39. Halogens are known as:**

(a) Inert gases

(b) Salt formers ✅

(c) Metals

(d) Semiconductors

**40. Noble gases are chemically unreactive because they have:**

(a) Incomplete shells

(b) Half-filled shells

(c) Completely filled outermost shell ✅

(d) No electrons

**41. The period number of an element indicates:**

(a) Number of protons

(b) Number of valence electrons

(c) Number of electron shells ✅

(d) Atomic mass

**42. An element in period 3 will have:**

(a) 1 shell

(b) 2 shells

(c) 3 shells ✅

(d) 4 shells

**43. The group number of an element represents:**

(a) Number of shells

(b) Number of neutrons

(c) Number of valence electrons ✅

(d) Atomic size

**44. An element in period 3 and group 2 belongs to s-block. Its valence electrons are present in:**

(a) 2s subshell

(b) 3p subshell

(c) 3s subshell ✅

(d) 2p subshell

**45. An element in period 3 and group 13 will have how many valence electrons?**

(a) 1

(b) 2

(c) 3 ✅

(d) 4

**46. Modern Periodic Law states that properties of elements are periodic functions of:**

(a) Atomic mass

(b) Atomic volume

(c) Atomic number ✅

(d) Density

**47. Atomic radius is defined as:**

(a) Distance between two nuclei

(b) Half of the distance between two identical bonded atoms ✅

(c) Distance between electrons

(d) Diameter of nucleus

**48. Atomic radius is commonly measured in:**

(a) Gram

(b) Meter

(c) Picometer (pm) or Angstrom (Å) ✅

(d) Joule

**49. Atomic radius generally decreases across a period due to:**

(a) Decrease in nuclear charge

(b) Increase in shielding effect

(c) Increase in effective nuclear charge pulling electrons closer ✅

(d) Increase in number of shells

**50. Atomic radius increases down a group because:**

(a) Nuclear charge decreases

(b) Number of shells increases and shielding effect dominates ✅

(c) Electrons are removed

(d) Atoms lose protons

**51. Which factor does NOT significantly affect atomic radius?**

(a) Number of shells

(b) Effective nuclear charge

(c) Shielding effect

(d) Atomic number alone without context ✅

**52. Ionic radius refers to the size of:**

(a) Neutral atom

(b) Molecule

(c) Ion in a crystal lattice ✅

(d) Nucleus

**53. A cation is smaller than its neutral atom because:**

(a) It gains electrons

(b) It loses one or more electron shells and nucleus attracts remaining electrons strongly ✅

(c) It gains protons

(d) It becomes unstable

**54. An anion is larger than its neutral atom because:**

(a) It loses electrons

(b) It gains electrons causing increased electron-electron repulsion ✅

(c) It loses protons

(d) It loses neutrons

**55. Ionic radius of ions across a period (left to right) generally:**

(a) Increases

(b) Remains constant

(c) Decreases due to increasing nuclear charge ✅

(d) Becomes zero

**56. Ionization energy is defined as the energy required to remove:**

(a) One neutron from atom

(b) One proton from nucleus

(c) One electron from a gaseous atom forming a 1+ ion ✅

(d) Two electrons from atom

**57. Ionization energy is measured for atoms in:**

(a) Solid state

(b) Liquid state

(c) Gaseous state ✅

(d) Plasma only

**58. Ionization energy generally increases with increase in:**

(a) Atomic size

(b) Effective nuclear charge ✅

(c) Shielding effect

(d) Number of shells

**59. Larger atomic size results in:**

(a) Higher ionization energy

(b) Lower ionization energy ✅

(c) No change

(d) Infinite energy

**60. Shielding effect refers to:**

(a) Attraction between protons and neutrons

(b) Repulsion between electrons only

(c) Reduction in nuclear attraction due to inner electrons ✅

(d) Increase in nuclear charge

**61. Greater shielding effect leads to:**

(a) Higher ionization energy

(b) Lower ionization energy ✅

(c) No effect

(d) Zero energy

**62. Ionization energy is higher when electrons are removed from:**

(a) Fully-filled or half-filled subshells ✅

(b) Empty orbitals

(c) Outer shells only

(d) Nucleus

**63. Noble gases have highest ionization energies because they have:**

(a) Half-filled shells

(b) Fully-filled stable shells ($ns^2$ $np^6$) ✅

(c) No electrons

(d) Single electron

**64. Oxygen has lower ionization energy than nitrogen because:**

(a) Oxygen has fewer protons

(b) Nitrogen has half-filled stable subshell ✅

(c) Oxygen has more shells

(d) Nitrogen has more shielding

**65. Spin-pair repulsion occurs when:**

(a) Electrons are in different shells

(b) Electrons are in different orbitals

(c) Two electrons share the same orbital with opposite spins ✅

(d) No electrons are present

**66. Due to spin-pair repulsion, ionization energy:**

(a) Increases

(b) Decreases slightly ✅

(c) Becomes zero

(d) Remains constant

**67. Chromium (Cr) has configuration [Ar] $3d^5$ $4s^1$, which shows:**

(a) Less stability

(b) Greater stability due to half-filled subshell ✅

(c) No stability

(d) Maximum repulsion

**68. Manganese (Mn) has configuration [Ar] $3d^5$ $4s^2$ and shows:**

(a) More stability than Cr

(b) Less stability compared to Cr due to paired electrons ✅

(c) No electrons

(d) No subshells

**69. Ionization energy trend down a group is:**

(a) Increases

(b) Decreases ✅

(c) Remains constant

(d) Becomes irregular

**70. Ionization energy trend across a period (left to right) is:**

(a) Decreases

(b) Increases ✅

(c) Remains constant

(d) Random

**71. Electron affinity is defined as the energy change when:**

(a) An electron is removed from atom

(b) A proton is added to atom

(c) An electron is added to a gaseous atom forming an anion ✅

(d) Neutrons are added to nucleus

**72. Electron affinity is generally measured under:**

(a) Solid conditions

(b) Liquid conditions

(c) Gaseous state ✅

(d) Plasma state only

**73. Smaller atomic size results in:**

(a) Lower electron affinity

(b) Higher electron affinity due to stronger attraction of nucleus ✅

(c) No effect

(d) Zero electron affinity

**74. Electron affinity increases with increase in:**

(a) Atomic size

(b) Shielding effect

(c) Nuclear charge ✅

(d) Number of shells only

**75. Electron affinity is low for elements with:**

(a) Fully filled orbitals

(b) Half-filled subshells (stable configuration) ✅

(c) No electrons

(d) High density

**76. Elements like nitrogen (N) and phosphorus (P) have low electron affinity because:**

(a) They have large atomic size

(b) They have half-filled p-subshells which are stable and resist addition of electrons ✅

(c) They have more shells

(d) They have low nuclear charge

**77. Electron affinity generally becomes more negative across a period due to:**

(a) Decrease in nuclear charge

(b) Increase in atomic size

(c) Increase in nuclear charge and decrease in atomic radius ✅

(d) Increase in shielding effect only

**78. Electronegativity is defined as:**

(a) Ability to lose electrons

(b) Ability to gain neutrons

(c) Ability of an atom to attract shared pair of electrons in a molecule ✅

(d) Energy required to remove electron

**79. Electronegativity increases across a period because:**

(a) Atomic size increases

(b) Nuclear charge increases and atomic size decreases ✅

(c) Shielding effect increases

(d) Number of shells increases

**80. In the periodic table, electronegativity generally:**

(a) Increases down a group

(b) Decreases down a group due to increased size and shielding effect ✅

(c) Remains constant

(d) Becomes zero

**81. Metallic character of an element is its tendency to:**

(a) Gain electrons

(b) Lose electrons to form positive ions ✅

(c) Share electrons only

(d) Accept protons

**82. Metallic character is generally higher in elements located on the:**

(a) Right side of periodic table

(b) Left side of periodic table ✅

(c) Top right corner

(d) Middle only

**83. Across a period from left to right, metallic character:**

(a) Increases

(b) Decreases ✅

(c) Remains constant

(d) Becomes zero instantly

**84. Down a group, metallic character:**

(a) Decreases

(b) Increases ✅

(c) Remains constant

(d) Becomes irregular

**85. Cesium is more reactive than sodium because:**

(a) It has smaller atomic size

(b) It has higher electronegativity

(c) It has larger atomic size and lower ionization energy ✅

(d) It has fewer shells

**86. Sodium reacts with water to form:**

(a) Sodium oxide and oxygen

(b) Sodium hydroxide and hydrogen ✅

(c) Sodium chloride and water

(d) Sodium peroxide only

**87. Magnesium reacts with steam to form:**

(a) MgO and $H_2$ gas ✅

(b) $Mg(OH)_2$ only

(c) $MgCl_2$

(d) $MgCO_3$

**88. Sodium is stored under kerosene oil because:**

(a) It is insoluble in oil

(b) It reacts violently with air and moisture ✅

(c) It is non-reactive

(d) It is a liquid metal

**89. Sodium burns in oxygen producing:**

(a) Only sodium oxide

(b) Sodium oxide and sodium peroxide ✅

(c) Sodium chloride

(d) Sodium hydroxide

**90. Oxides of group 1, 2 and 3 are mostly:**

(a) Covalent

(b) Ionic and basic in nature ✅

(c) Acidic

(d) Neutral

**91. Oxides of group 4 to 7 are generally:**

(a) Ionic

(b) Metallic

(c) Covalent and acidic in nature ✅

(d) Basic

**92. Basic oxides react with water to form:**

(a) Acids

(b) Salts

(c) Alkalis (bases) ✅

(d) Neutral compounds

**93. Amphoteric oxides are those which:**

(a) React only with acids

(b) React only with bases

(c) React with both acids and bases ✅

(d) Do not react at all

**94. Aluminium oxide ($Al_2O_3$) is an example of:**

(a) Basic oxide

(b) Acidic oxide

(c) Amphoteric oxide ✅

(d) Neutral oxide

**95. Neutral chlorides produce solution with pH approximately:**

(a) Less than 7

(b) Greater than 7

(c) Equal to 7 ✅

(d) Zero

**96. Oxidation number is defined as:**

(a) Mass of an atom

(b) Formal charge on an atom in a molecule or ion ✅

(c) Number of neutrons

(d) Number of shells

**97. In ionic compounds, oxidation number represents:**

(a) Atomic mass

(b) Charge on the ion ✅

(c) Number of electrons shared

(d) Valence shell

**98. Oxidation numbers in oxides and chlorides are positive because:**

(a) Oxygen and chlorine are less electronegative

(b) Oxygen and chlorine are more electronegative than other elements ✅

(c) Metals are more electronegative

(d) Non-metals lose electrons easily

**99. Across the third period, oxidation number in oxides generally increases from:**

(a) 0 to +3

(b) +1 (Na) to +6 (S) ✅

(c) +2 to +8

(d) +3 to +7

**100. Phosphorus and sulfur show multiple oxidation states because they can:**

(a) Lose electrons only

(b) Gain electrons only

(c) Expand their octet by using empty 3d orbitals ✅

(d) Form only ionic bonds

## ❖ Important Short Questions:

**1. What is periodicity in the periodic table?**

**Ans:** Periodicity is the repetition of similar properties of elements at regular intervals when arranged by increasing atomic number.

**Example:** Lithium, sodium, and potassium show similar chemical properties.

**2. What is ionization energy?**

**Ans:** Ionization energy is the energy required to remove the most loosely bound electron from a neutral gaseous atom.

**Example:** $Na\ (g) \longrightarrow Na^{+}\ (g) + e^{-}$

**3. What is electronegativity?**

**Ans:** Electronegativity is the ability of an atom to attract shared electrons in a covalent bond.

**Example:** In HCl, chlorine attracts electrons more strongly than hydrogen.

**4. What is atomic radius?**

**Ans:** Atomic radius is the distance between the nucleus and the outermost electron shell of an atom.

**Example:** Sodium has a larger atomic radius than chlorine.

**5. What is electron affinity?**

**Ans:** Electron affinity is the energy released when an electron is added to a neutral gaseous atom.

**Example:** $Cl\ (g) + e^- \longrightarrow Cl^-\ (g)$

**6. What is oxidation number?**

**Ans:** Oxidation number is the formal charge assigned to an atom in a molecule or ion.

**Example:** In $H_2O$, oxygen has an oxidation number of -2.

**7. What is the difference between oxidation state and charge?**

**Ans:** Oxidation state is a theoretical value used in covalent compounds, while charge is the actual charge on ions in ionic compounds.

**Example:** In $SO_2$, sulfur has oxidation state +4.

**8. Why are elements arranged in periods and groups?**

**Ans:** Elements are arranged based on increasing atomic number and similar valence electrons to show periodic trends.

**Example:** Elements in Group 1 (Li, Na, K) have similar properties.

**9. What are s-block elements?**

**Ans:** Elements whose last electron enters the s-orbital are called s-block elements.

**Example:** Lithium (Li), Sodium (Na).

**10. What are p-block elements?**

**Ans:** Elements whose last electron enters the p-orbital are called p-block elements.

**Example:** Carbon (C), Oxygen (O).

**11. Why are Group 13–17 elements called p-block elements?**

**Ans:** Because their valence electrons enter the p-subshell.

**Example:** Nitrogen (Group 15) has p-orbital electrons.

**12. What is metallic character?**

**Ans:** Metallic character is the ability of an element to lose electrons and form positive ions.

**Example:** Sodium easily loses one electron to form $Na^+$.

**13. How does metallic character change across a period?**

**Ans:** It decreases from left to right.

**Example:** Sodium is more metallic than chlorine.

**14. How does metallic character change down a group?**

**Ans:** It increases from top to bottom.

**Example:** Potassium is more metallic than lithium.

**15. Why does atomic radius decrease across a period?**

**Ans:** Due to increasing nuclear charge pulling electrons closer.

**Example:** Na is larger than Cl.

**16. Why does atomic radius increase down a group?**

**Ans:** Due to addition of new electron shells.

**Example:** K is larger than Na.

**17. Why does ionization energy increase across a period?**

**Ans:** Because nuclear charge increases and atomic size decreases.

**Example:** Na has lower ionization energy than Cl.

**18. Why does ionization energy decrease down a group?**

**Ans:** Due to increased atomic size and shielding effect.

**Example:** Li has higher ionization energy than K.

**19. What is the trend of electronegativity across a period?**

**Ans:** It increases from left to right.

**Example:** Fluorine is more electronegative than carbon.

**20. What is the trend of electron affinity down a group?**

**Ans:** It generally decreases down a group.

**Example:** Chlorine has higher electron affinity than iodine.

**21. Why do elements in the same group have similar chemical properties?**

**Ans:** Because they have the same number of valence electrons.

**Example:** Na and K both have one valence electron.

**22. Why do noble gases have high ionization energy?**

**Ans:** Because they have stable complete valence shells.

**Example:** Helium and neon are very stable.

**23. Why do alkali metals show high reactivity?**

**Ans:** Because they have one valence electron and low ionization energy.

**Example:** Sodium reacts vigorously with water.

**24. Why do halogens show high electronegativity?**

**Ans:** Because they need one electron to complete their octet.

**Example:** Chlorine readily gains one electron to form $Cl^-$.

**25. Why do some elements show variable oxidation states?**

**Ans:** Because they can use different numbers of electrons for bonding due to d-orbitals.

**Example:** Sulfur shows oxidation states of +4 and +6 in $SO_2$ and $SO_3$.

# 📘 EXERCISE

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

**I. Which scientist first time observed the periodicity in the elements?**

a) J. Newlands ✅

b) L. Meyer

c) J.W. Döbereiner

d) D. I. Mendeleev

**II. Recognize the element (3 shells, s-block, 2 outer electrons)**

a) Calcium

b) Sodium

c) Magnesium ✅

d) Potassium

**III. Which one 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) Phosphorus ✅

**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?**

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) Aluminium (Al)

d) Argon (Ar) ✅

## Q.2 Attempt the following short-answer questions:

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

**Ans:** The first ionization energy is the amount of energy required to remove the most loosely bound electron from a neutral gaseous atom. For example, the first ionization energy of sodium is the energy needed to remove one electron from Na(g) to form $Na^+$(g).

**b) Explain why sulfur has a lower first ionization energy than phosphorus.**

**Ans:** Sulfur has a lower first ionization energy than phosphorus because sulfur has paired electrons in one of its 3p orbitals, which

creates extra electron–electron repulsion. Phosphorus has a half-filled stable $3p^3$ configuration, which is more stable, so it requires more energy to remove an electron from phosphorus than from sulfur.

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

**Ans:** Elements in Groups 13 to 17 are called p-block elements because their last (valence) electrons enter the p-subshell. Their outermost electrons are in p orbitals, which defines their block in the periodic table.

**d) What are the factors that affect electronegativity?**

**Answer:**

Electronegativity is affected by:

- Atomic size (smaller atoms have higher electronegativity)
- Nuclear charge (greater nuclear charge increases electronegativity)
- Shielding effect (greater shielding reduces electronegativity)
- Position in the periodic table (increases across a period and decreases down a group)

**e) What factors are responsible for the increasing reactivity of alkali metals as we move down the group?**

**Ans:** Reactivity increases down the group because atomic size increases and the outer electron is farther from the nucleus. The nuclear

attraction decreases and ionization energy becomes lower, so the outer electron is lost more easily.

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

**Ans:** Elements that have access to d-orbitals can use different numbers of electrons for bonding by exciting electrons, so they show variable oxidation numbers. Elements without available d-orbitals usually have fixed oxidation numbers because they cannot expand their octet.

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

**Ans:** The element is Antimony (Sb).

**h) Why oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorus?**

**Ans:** Sodium and magnesium are metals with low electronegativity, so they tend to lose electrons and form ionic bonds with oxygen. Nitrogen and phosphorus are non-metals with higher electronegativity, so they share electrons with oxygen and form covalent oxides. Hence, $Na_2O$ and MgO are more ionic than oxides of N and P.

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

**Ans:** Sodium has one valence electron while magnesium has two valence electrons and higher nuclear charge. Sodium loses its single electron more easily (lower ionization energy), so it is more reactive

than magnesium. Magnesium requires more energy to lose two electrons, so its reactivity is comparatively lower toward oxygen and chlorine.

**j) Why the ionization energy of lithium is much lower than that of helium despite lithium having higher nuclear charge?**

**Ans:** Lithium has a larger atomic radius with electrons in the second shell, and inner electrons shield the outer electron from the nucleus. Helium has a very small atomic size with electrons in the first shell and stronger nuclear attraction. Therefore, helium has much higher ionization energy than lithium.

**k) Why is the ionization energy of Be higher than that of B despite higher nuclear charge of B?**

**Ans:** Beryllium has a stable fully filled $2s^2$ subshell, whereas boron has its outermost electron in the 2p orbital, which is higher in energy and less tightly held. The 2p electron in boron is easier to remove than the 2s electrons in beryllium, so Be has higher ionization energy than B.

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

**Answer:**

All these species are isoelectronic (they have 10 electrons each).

**Increasing order of size:**

$Al^{3+} < Mg^{2+} < Na^{+} < Ne < F^{-}$

**m) Consider NaCl, $MgCl_2$, and $PCl_5$**

**Answer:**

**(i) Classification:**

- NaCl ⟶ Basic
- $MgCl_2$ ⟶ Basic (slightly acidic/neutral in solution, but generally treated as basic due to metal chloride nature)
- $PCl_5$ ⟶ Acidic

**(ii) Reason:**

**NaCl:** Formed from a strong base (NaOH) and strong acid (HCl), gives a neutral/basic solution in water.

**$MgCl_2$:** Formed from a metal and reacts slightly with water but generally behaves like a salt of a metal, showing basic character.

**$PCl_5$:** A covalent compound of a non-metal; it reacts with water to produce phosphoric acid and hydrochloric acid, hence acidic in nature.

## DESCRIPTIVE QUESTIONS

**🌟 Q3: Write equations for the reactions of Na and Mg with oxygen, chlorine, and water. Compare the reactivity of both elements in terms of metallic character.**

**Reactions of Sodium (Na):**

**With oxygen:**

$4Na + O_2 \longrightarrow 2Na_2O$

**With chlorine:**

$2Na + Cl_2 \longrightarrow 2NaCl$

**With water:**

$2Na + 2H_2O \longrightarrow 2NaOH + H_2$

## Reactions of Magnesium (Mg):

**With oxygen:**

$2Mg + O_2 \longrightarrow 2MgO$

**With chlorine:**

$Mg + Cl_2 \longrightarrow MgCl_2$

**With water (steam):**

$Mg + H_2O \longrightarrow MgO + H_2$

## Comparison in terms of reactivity and metallic character:

Sodium is more reactive than magnesium because it has a larger atomic size and lower ionization energy. Sodium has one valence electron which it loses easily to form $Na^+$, whereas magnesium has two valence electrons and requires more energy to remove both electrons.

Due to easier loss of electrons, sodium shows stronger metallic character and reacts more vigorously with oxygen, chlorine, and water. Magnesium is less reactive in comparison because its electrons are held more strongly by the nucleus.

**🌟 Q4: Explain with the help of equations, acidic and basic behavior of oxides and chlorides.**

## Basic behavior of metal oxides:

Metal oxides are basic in nature because they react with water to form metal hydroxides, which release hydroxide ions ($OH^-$) in solution.

**Examples:**

$Na_2O + H_2O \longrightarrow 2NaOH$

$MgO + H_2O \longrightarrow Mg(OH)_2$

These hydroxides dissociate in water to produce $OH^-$ ions, which confirms their basic nature.

## Acidic behavior of non-metal oxides:

Non-metal oxides are acidic because they react with water to form acids, which release hydrogen ions ($H^+$) in solution.

**Examples:**

$SO_3 + H_2O \longrightarrow H_2SO_4$

$P_2O_5 + 3H_2O \longrightarrow 2H_3PO_4$

These acids ionize in water and produce $H^+$ ions, showing acidic behavior.

## Behavior of chlorides:

### Metal chlorides:

Metal chlorides such as NaCl and MgCl2 are generally neutral because they are formed from a strong acid (HCl) and a strong base (NaOH or Mg(OH)2). They dissolve in water without significant hydrolysis.

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

$MgCl2 \longrightarrow Mg^{2+} + 2Cl^-$

These ions do not react with water to produce $H^+$ or $OH^-$, so the solution remains neutral.

### Non-metal chlorides:

Non-metal chlorides, such as phosphorus pentachloride (PCl5), react with water and undergo hydrolysis to form acids.

$PCl5 + 4H2O \longrightarrow H3PO4 + 5HCl$

The formation of acids (H3PO4 and HCl) makes the solution acidic.

### Conclusion:

Metal oxides and metal chlorides show basic or neutral behavior due to the formation of hydroxides or non-reactive ions in water, whereas

non-metal oxides and chlorides show acidic behavior because they react with water to produce acids that release $H^+$ ions.

**🌟 Q5: Describe the factors affecting and periodic trends of electron affinity.**

**Electron Affinity:**

Electron affinity is the energy released when an electron is added to a neutral gaseous atom.

**✦ Factors Affecting Electron Affinity:**

**1. Atomic size:**

Smaller atoms have stronger attraction for incoming electrons, so electron affinity is higher (more negative). Larger atoms have lower electron affinity due to greater distance of the outer shell from the nucleus.

**2. Nuclear charge:**

Higher nuclear charge increases the attraction for electrons, resulting in higher electron affinity.

**3. Shielding effect:**

Inner electrons shield the nucleus from the incoming electron. Greater shielding reduces the effective nuclear attraction, thus decreasing electron affinity.

**4. Electronic configuration (stability):**

Atoms with stable configurations (e.g., noble gases or half-filled subshells) have low or even positive electron affinity because they do not easily accept additional electrons.

### Periodic Trends of Electron Affinity:

**Across a period (left to right):**

Electron affinity generally increases (becomes more negative) because atomic size decreases and nuclear charge increases, which strengthens the attraction for added electrons.

**Down a group (top to bottom):**

Electron affinity generally decreases (becomes less negative) due to increasing atomic size and shielding effect, which reduces the attraction for incoming electrons.

✦ **Summary:**

Electron affinity depends mainly on atomic size, nuclear charge, shielding effect, and electronic configuration. It increases across a period and decreases down a group, although some exceptions exist due to stable electronic configurations.

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

❖ **Definition:**

Ionization energy is the amount of energy required to remove the most loosely bound electron from a neutral gaseous atom to form a positive ion.

**Example:**

$Na\ (g) \longrightarrow Na^{+}\ (g) + e^{-}$

### ✦ Factors Affecting Ionization Energy:

**1. Atomic size:**

Larger atoms have electrons farther from the nucleus, so the attraction is weaker and ionization energy is low. Smaller atoms have higher ionization energy due to stronger nuclear attraction.

**2. Nuclear charge:**

Higher nuclear charge increases the attraction between the nucleus and electrons, making it more difficult to remove an electron, thus increasing ionization energy.

**3. Shielding effect:**

Inner-shell electrons reduce the effective nuclear attraction on valence electrons. Greater shielding lowers ionization energy.

**4. Electronic configuration:**

Atoms with stable configurations (full or half-filled subshells) have higher ionization energy because their electrons are more stable and harder to remove.

### Periodic Trends of Ionization Energy:

**Across a period (left to right):**

Ionization energy increases due to increasing nuclear charge and decreasing atomic size, which strengthens the attraction on electrons.

**Down a group (top to bottom):**

Ionization energy decreases because atomic size increases and shielding effect increases, reducing the effective nuclear attraction on outer electrons.

✦ **Summary:**

Ionization energy depends on atomic size, nuclear charge, shielding effect, and electronic configuration. It increases across a period and decreases down a group due to changes in nuclear attraction and distance of valence electrons from the nucleus.