

## Chapter 4

### Periodic Table and Periodicity of Properties

#### Periodic Table and Periodic Law

#### Multiple Choice Questions (MCQs)

1. What is the primary purpose of the periodic table?
  - a) To list all known compounds
  - b) To organize information about elements**
  - c) To calculate the atomic mass of compounds
  - d) To describe chemical reactions in detail
2. How many elements were known until the end of the 18th century?
  - a) 118
  - b) 100
  - c) 23**
  - d) 50
3. The periodic table is very useful for:
  - a) Changing atomic numbers
  - b) Predicting properties of elements**
  - c) Determining the color of an element
  - d) Calculating reaction speed
4. The modern periodic table is arranged based on:
  - a) atomic mass
  - b) Density
  - c) atomic number**
  - d) Reactivity
5. The Periodic Law states that properties are a periodic function of an element:
  - a) atomic mass
  - b) atomic number**
  - c) Number of neutrons
  - d) Boiling point
6. Who discovered the atomic number, which became the basis for the modern periodic table?

- a) Dmitri Mendeleev
  - b) John Dalton
  - c) Henry Moseley**
  - d) Linus Pauling
7. In the periodic table, elements are listed in order of increasing atomic number:
- a) From top to bottom only
  - b) From right to left only
  - c) From left to right and top to bottom**
  - d) In alphabetical order
8. Which element is in the top left corner of the periodic table?
- a) Helium (He)
  - b) Lithium (Li)
  - c) Hydrogen (H)**
  - d) Oxygen (O)
9. The horizontal rows in the periodic table are called:
- a) Groups
  - b) Families
  - c) Periods**
  - d) Blocks
10. How many periods are there in the modern periodic table?
- a) 8
  - b) 18
  - c) 7**
  - d) 10
11. Which period contains only 2 elements?
- a) Period 2
  - b) Period 6
  - c) Period 1**
  - d) Period 7
12. Periods 1, 2, and 3 are classified as:
- a) long periods
  - b) Transition periods
  - c) short periods**
  - d) Noble periods
13. As you move from left to right across a period, the properties of elements:

- a) Remain constant
  - b) Change gradually**
  - c) Change randomly
  - d) Become identical
14. The pattern of properties within a period repeats when you:
- a) Move down a group
  - b) Move across a period
  - c) Move to the next period**
  - d) Look at atomic mass
15. The international body responsible for naming new elements and placing them on the periodic table is:
- a) NASA
  - b) WHO
  - c) IUPAC**
  - d) UNESCO
16. The element with atomic number 2, located at the top right corner, is:
- a) Hydrogen (H)
  - b) Lithium (Li)
  - c) Helium (He)**
  - d) Beryllium (Be)
17. The element with atomic number 3, at the left of the second row, is:
- a) Hydrogen (H)
  - b) Helium (He)
  - c) Lithium (Li)**
  - d) Beryllium (Be)
18. The number of elements in a period:
- a) Is constant for all periods
  - b) Decreases from top to bottom
  - c) Varies, from 2 to 32**
  - d) Is always 8
19. The repetition of properties is in accordance with
- a) The Law of Conservation of Mass
  - b) The Periodic Law**
  - c) The Law of Multiple Proportions
  - d) Avogadro's Law

20. One can use the periodic table to predict the type of bond an element might form based on its:
- a) Color
  - b) atomic structure**
  - c) Discoverer
  - d) Year of discovery

## Short Questions

**Q1: Why is it impossible to remember details about all elements without a system like the periodic table?**

With 118 known elements, each with unique reactions, properties, and atomic masses, the amount of information is vast and complex. Memorizing each detail individually is impractical. The periodic table provides a logical and systematic framework that organizes this information, showing patterns and relationships that make it understandable and predictable.

**Q2: What is the key advantage of elements being placed in the same group?**

Elements in the same group have the same number of valence electrons, which dictate their chemical behavior. Therefore, if you know the properties of one element in a group, you can accurately predict the physical and chemical properties of any other element in that same group, as they will behave similarly.

**Q3: State the Modern Periodic Law.**

The Modern Periodic Law states that if the elements are arranged in the order of their increasing atomic numbers, their physical and chemical properties are repeated in a periodic manner. This means patterns in properties recur at regular intervals.

**Q4: Differentiate between a 'period' and a 'group' in the periodic table.**

**Periods** are the horizontal rows. There are 7 periods, and properties change gradually across them. **Groups** are the vertical columns. There are 18 groups, and elements in the same group share similar chemical properties due to the same number of valence electrons.

**Q5: Describe the layout of elements in the periodic table.**

Elements are listed in order of increasing atomic number from left to right and from top to bottom. Hydrogen ( $Z=1$ ) is in the top left corner, helium ( $Z=2$ ) is at the top right, and lithium ( $Z=3$ ) begins the second period on the left. This order continues sequentially.

**Q6: Why are the first three periods called 'short periods'?**

The first three periods are called short periods because they contain a relatively small number of elements. Period 1 has 2 elements, Period 2 has 8 elements, and Period 3 has 8 elements. In contrast, periods 4, 5, 6, and 7 contain 18 or more elements and are called long periods.

**Q7: What happens to the properties of elements as you move from left to right across a single period?**

As you move from left to right across a period, the properties of the elements change gradually. For example, there is a transition from metallic to non-metallic character. However, the number of electron shells remains the same, while the number of valence electrons increases.

**Q8: What is the significance of Moseley's discovery of the atomic number?**

Before Moseley, elements were arranged by atomic mass, which led to inconsistencies. Moseley's discovery that atomic number (the number of protons) is the fundamental property of an element resolved these inconsistencies and provided the correct basis for the Modern Periodic Law and the structure of the periodic table.

**Q9: What is the role of IUPAC concerning the periodic table?**

The International Union of Pure and Applied Chemistry (IUPAC) is the governing body that standardizes chemical nomenclature. It is responsible for ratifying the names of newly discovered elements, assigning their symbols, and placing them in their correct positions on the periodic table.

**Q10: How does the periodic table help in understanding the reactivity of elements?**

The periodic table allows us to relate an element's reactivity to its atomic structure. Elements on the far left (metals) tend to lose electrons and are reactive. Elements on the far right (non-metals) tend to gain electrons. Noble gases, with full valence shells, are inert. This pattern helps predict how any element will react.

## **Blocks, Groups and Periods in Periodic table in relation to the Electronic Configuration**

### **Multiple Choice Questions (MCQs)**

1. Elements in the same vertical column of the periodic table are called a:
  - a) Period
  - b) Series

**c) Group or Family**

d) Block

2. Elements in the same group have similar properties because they have the same number of
  - a) Electron shells
  - b) Valence electrons**
  - c) Neutrons
  - d) Protons
3. The modern IUPAC system for numbering groups in the periodic table is
  - a) 1A to 8A and 1B to 8B
  - b) 1 to 8 with A and B suffixes
  - c) 1 to 18 from left to right**
  - d) By the number of valence electrons
4. In the traditional system, the groups in the middle of the table were designated as:
  - a) A groups
  - b) B groups**
  - c) Representative groups
  - d) Zero groups
5. Group A elements are also known as:
  - a) Transition elements
  - b) Noble gases
  - c) Representative or main group elements**
  - d) Inner transition elements
6. Group B elements are commonly known as:
  - a) Alkali metals
  - b) Halogens
  - c) Transition elements**
  - d) Metalloids
7. Which group is known as the alkaline earth metals?
  - a) Group 1
  - b) Group 2**
  - c) Group 16
  - d) Group 17
8. The group of elements known for their lack of reactivity is the:
  - a) Alkali metals

- b) Halogens
  - c) Noble gases**
  - d) Transition metals
9. The order in which electron orbitals are filled is determined by the:
- a) Heisenberg principle
  - b) Pauli exclusion principle
  - c) Aufbau principle**
  - d) Hund's rule
10. The block of an element is determined by the type of orbital occupied by the:
- a) First electron
  - b) Innermost electron
  - c) Last electron**
  - d) Most energetic proton
11. The period number of an element is determined by the:
- a) Number of valence electrons
  - b) Group number
  - c) Principal quantum number of the valence shell**
  - d) Block it belongs to
12. For an s-block element, the group number is equal to its:
- a) atomic number
  - b) Number of electron shells
  - c) Number of valence electrons**
  - d) Number of p orbitals
13. For a p-block element, the number of valence electrons is equal to:
- a) The group number
  - b) The period number
  - c) The group number minus 10**
  - d) The atomic number
14. Sodium (Na,  $Z=11$ ) has its valence electron in the:
- a) 2s orbital
  - b) 2p orbital
  - c) 3s orbital**
  - d) 3p orbital
15. Based on its valence electron, sodium belongs to the:
- a) p-block

- b) d-block
- c) s-block**
- d) f-block

16. The principal quantum number for sodium's valence electron is 3, so it is in period:

- a) 1
- b) 2
- c) 3**
- d) 4

17. With one valence electron in the s-block, sodium belongs to group:

- a) 1**
- b) 2
- c) 11
- d) 13

18. Elements in Groups 1 and 2 are called s-block elements because their valence electrons are in the:

- a) s sub-shell**
- b) p sub-shell
- c) d sub-shell

19. Elements in Groups 13 to 18 (except He) are known as:

- a) s-block elements
- b) p-block elements**
- c) d-block elements
- d) f-block elements

20. Lanthanides and actinides are known as:

- a) s-block elements
- b) p-block elements
- c) d-block elements
- d) f-block elements**

21. The charge on an ion formed by a Group 1 element is typically:

- a) +1**
- b) +2
- c) -1
- d) -2

22. The charge on an ion formed by a Group 2 element is typically:

- a) +1

- b) +2**
- c) -1
- d) -2

23. Elements tend to lose or gain electrons to achieve a stable configuration similar to

- a) Alkali metals
- b) Halogens
- c) Transition metals
- d) Noble gases**

24. An element with 7 valence electrons will most likely form an ion with a charge of:

- a) +7
- b) -7
- c) +1
- d) -1**

25. An element with 2 valence electrons will most likely form an ion with a charge of:

- a) +2**
- b) -2
- c) +6
- d) -6

26. Alkali metals form ions called:

- a) Divalent anions
- b) Monovalent cations**
- c) Monovalent anions
- d) Divalent cations

27. The general electron configuration for alkali metals is

- a)  $ns^1$**
- b)  $ns^2$
- c)  $ns^2 np^1$
- d)  $ns^2 np^5$

28. Carbon (C) has the electron configuration  $1s^2 2s^2 2p^2$ . It belongs to the:

- a) s-block, Group 2
- b) p-block,**

## Short Questions

**Q1: Why do elements in the same group of the periodic table have similar chemical properties?**

Elements in the same group have the same number of electrons in their outermost shell, known as valence electrons. Since chemical reactions involve the gain, loss, or sharing of these valence electrons, having the same number results in similar chemical behavior and reactivity.

**Q2: Explain the difference between the traditional and the modern (IUPAC) group numbering systems.**

The traditional system used Roman numerals (I-VIII) with the letters A and B. Groups I-A to VII-A were the representative elements, and the B groups were the transition elements. The modern IUPAC system simplifies this by numbering all groups sequentially from 1 to 18 from left to right, eliminating the A and B labels.

**Q3: Differentiate between representative elements and transition elements.**

**Representative elements** (or main group elements) are found in Groups 1, 2, and 13 to 18. Their properties are predictable as they often form ions with charges equal to their group number.

**Transition elements** are found in Groups 3 to 12. They are metals, often have variable valencies, and form colored compounds.

**Q4: How is the 'block' of an element (s, p, d, f) determined?**

The block of an element is determined by the type of atomic orbital (s, p, d, or f) that its last or valence electron occupies. For example, if the last electron is in an s-orbital, the element is in the s-block; if it's in a p-orbital, it's in the p-block, and so on.

**Q5: Using sodium (Z=11) as an example, explain how to determine its period and group.**

Sodium's electronic configuration is  $1s^2 2s^2 2p^6 3s^1$ . The highest principal quantum number is  $n=3$ , so it is in the **3rd period**. The last electron is in the s-orbital, so it's an s-block element. For s-block, the group number equals the number of valence electrons. With one valence electron ( $3s^1$ ), sodium belongs to **Group 1**.

**Q6: Identify which blocks the following groups belong to: a) Group 1, b) Group 13, c) Group 3.**

a) Group 1: s-block (valence electron in s-orbital) b) Group 13: p-block (valence electrons in s and p orbitals) c) Group 3: d-block (transition elements, last electron enters d-orbital)

**Q7: What is the relationship between the group number and the charge of the ion formed for s-block elements?**

For s-block elements (Groups 1 and 2), the group number directly indicates the number of valence electrons they lose to form positive ions. Group 1 elements lose 1 electron to form +1 ions, and Group 2 elements lose 2 electrons to form +2 ions.

**Q8: What is the relationship between the group number and the number of valence electrons for p-block elements?**

For p-block elements (Groups 13 to 18), the number of valence electrons is equal to the group number minus 10. For example, Group 17 elements have  $17 - 10 = 7$  valence electrons. Group 13 elements have  $13 - 10 = 3$  valence electrons.

**Q9: Explain why atoms lose or gain electrons.**

Atoms lose or gain electrons to achieve a stable electron configuration, typically a full outer shell of electrons resembling the nearest noble gas. This stable configuration has low energy, making the atom less reactive. Metals lose electrons to empty their outer shell, while non-metals gain electrons to complete theirs.

**Q10: Predict the charge on the ion formed by oxygen ( $Z=8$ ). Explain your reasoning.**

Oxygen has an atomic number of 8 and an electron configuration of  $1s^2 2s^2 2p^4$ . It has 6 valence electrons and is in Group 16. To achieve a stable noble gas configuration (like neon), it tends to gain 2 electrons. By gaining 2 electrons, it forms the oxide ion with a charge of -2 ( $O^{2-}$ ).

## Periodicity of Properties

### Shielding Effect

#### Multiple Choice Questions (MCQs)

- The periodic recurrence of the properties of elements is due to the periodic recurrence of:
  - Atomic mass
  - electronic configuration**
  - Density
  - Melting point
- Chemical properties of elements depend primarily on the:
  - Total number of electrons
  - Number of neutrons
  - electronic configuration of the valence shell**
  - atomic size
- Physical properties within a group show a gradation because:
  - Valence configuration changes
  - The sizes of atoms change gradually**
  - The nuclear charge decreases
  - They become more metallic

4. The shielding effect is best described as:
- a) The attraction between the nucleus and electrons
  - b) The repulsion between neutrons
  - c) The reduction of nuclear attraction on valence electrons by inner-shell electrons**
  - d) The increase in atomic size across a period
5. Which atom would have a greater shielding effect, Beryllium (Be) or Magnesium (Mg)?
- a) Beryllium
  - b) Magnesium**
  - c) Both are equal
  - d) Cannot be determined
6. Down a group, the shielding effect generally:
- a) Decreases
  - b) Increases**
  - c) Remains constant
  - d) Becomes zero
7. The increase in shielding effect down a group is because:
- a) atomic number decreases
  - b) The number of inner electron shells increases**
  - c) Nuclear charge decreases
  - d) Valence electrons are added
8. Across a period, the shielding effect generally:
- a) Decreases
  - b) Increases
  - c) Remains constant**
  - d) Changes unpredictably
9. Across a period, shielding effect remains constant because:
- a) atomic size increases
  - b) The number of inner electron shells remains the same**
  - c) nuclear charge decreases
  - d) Valence electrons are lost
10. Which pair of atoms would have the same number of inner shells and thus similar shielding within their period?
- a) Li and Mg
  - b) Be and B**
  - c) Na and Cl
  - d) K and Ca

## Short Questions

**Q1: What is the fundamental reason for the periodicity of properties observed in the periodic table?**

The periodicity of properties is due to the periodic repetition of similar valence shell electron configurations as the atomic number increases. Elements with the same number of valence electrons are placed in the same group, and since chemical behavior is determined by these electrons, their properties recur at regular intervals.

**Q2: Explain why elements in the same group similar chemical properties have but show a gradation in physical properties.**

Elements in the same group have the same number of valence electrons, which dictate their chemical behavior, leading to similar chemical properties. However, physical properties depend on atomic size. Moving down a group, each element adds a new electron shell, increasing the atomic size gradually and causing a gradation in physical properties like density and melting point.

**Q3: Define the 'shielding effect'.**

The shielding effect is the phenomenon where electrons present in the inner shells reduce the attractive force between the positively charged nucleus and the valence electrons. The inner electrons "shield" or screen the valence electrons from the full pull of the nucleus.

**Q4: Explain the trend in shielding effect as we move down a group.**

The shielding effect increases as we move down a group. This is because each subsequent element has an additional inner shell of electrons compared to the one above it. These extra inner electrons provide more repulsion against the valence electrons, better shielding them from the nucleus's pull.

**Q5: Explain why the shielding effect remains constant across a period.**

Across a period, the principal quantum number ( $n$ ) remains the same, meaning all electrons are added to the same valence shell. The number of inner electron shells does not change. Since only the valence shell is being filled and the inner core remains constant, the shielding effect provided by the inner electrons remains approximately constant.

## Atomic Size

## Multiple Choice Questions (MCQs)

- Which of the following properties is NOT primarily affected by an atom's electronic configuration?
  - Atomic size
  - Ionization energy
  - Density of the element in solid state**
  - Electronegativity
- Atomic size is defined as the average distance between the nucleus and the:
  - Innermost electron
  - Outermost electron shell**
  - Proton
  - Neutron
- Moving from left to right across a period, the atomic radius generally:
  - Increases
  - Decreases**
  - Remains constant
  - Changes randomly
- The decrease in atomic radius across a period is primarily due to:
  - An increase in the number of electron shells
  - An increase in nuclear charge with electrons added to the same shell**
  - A decrease in shielding effect
  - An increase in electron-electron repulsion
- As you move across a period, the nuclear charge \_\_\_\_\_ and the shielding effect \_\_\_\_\_.
  - increases, increases
  - increases, remains constant**
  - decreases, decreases
  - remains constant, increases
- Moving down a group, the atomic radius generally:
  - Increases**
  - Decreases
  - Remains constant
  - First increases then decreases
- The increase in atomic radius down a group is primarily due to
  - An increase in nuclear charge
  - A decrease in shielding effect

**c) The addition of new electron shells**

d) A decrease in the number of protons

8. Which of the following has the largest atomic radius?

a) Li

b) Na

**c) K**

d) They are all equal

9. Which of the following has the smallest atomic radius? (All in the same period)

**a) Li**

b) Be

c) B

d) They are all equal

10. The size of an atom is ultimately determined by the:

a) Size of its nucleus

b) Number of core electrons

**c) Size and number of its electron shells**

d) Number of neutrons

## Short Questions

**Q1: List the four atomic properties mentioned that depend on electronic configuration and show periodic trends.**

The four atomic properties that depend on electronic configuration and show periodic trends are:

1. Atomic size

2. Ionization energy

3. Electron affinity

4. Electronegativity These properties increase or decrease in a predictable pattern across periods and down groups.

**Q2: Describe the trend in atomic radius across a period and explain the reason for this trend.**

Atomic radius decreases across a period from left to right. This occurs because electrons are added to the same valence shell while protons are added to the nucleus. The increasing nuclear charge pulls the electron cloud closer towards the nucleus. Since the shielding effect from inner electrons remains constant, the effective nuclear charge increases, resulting in a smaller atomic radius.

**Q3: Using lithium (Li) and beryllium (Be) as an example, explain why atomic size decreases from left to right in a period.**

Both Li ( $Z=3$ ) and Be ( $Z=4$ ) are in period 2, so they have the same number of electron shells ( $n=2$ ). However, beryllium has one more proton in its nucleus than lithium. This greater positive charge exerts a stronger pull on the same valence shell, drawing the electrons closer to the nucleus. Therefore, the atomic radius of beryllium is smaller than that of lithium.

**Q4: Describe the trend in atomic radius down a group and explain the reason for this trend.**

Atomic radius increases down a group. This is because each subsequent element has one more electron shell than the element above it. This addition of a new, outer shell increases the distance between the nucleus and the outermost electrons. Although the nuclear charge also increases, the effect of adding a new shell outweighs it, resulting in a larger atomic radius.

**Q5: Using lithium (Li) and sodium (Na) as an example, explain why atomic size increases down a group.**

Lithium ( $Z=3$ ) has two electron shells ( $n=1$  and  $n=2$ ). Sodium ( $Z=11$ ) has three electron shells ( $n=1$ ,  $n=2$ , and  $n=3$ ). This addition of a third electron shell in sodium places the valence electron much farther from the nucleus compared to lithium. The increased distance between the nucleus and the outermost electron shell is the primary reason for the larger atomic radius of sodium.

## Ionization Energy

### Multiple Choice Questions (MCQs)

1. Ionization energy is defined as the minimum amount of energy required to:  
Add an electron to a gaseous atom.  
**b) Remove the outermost electron from an isolated gaseous atom**  
c) Remove an electron from a gaseous ion  
d) Break a chemical bond.
  2. A high ionization energy indicates:  
a) Weak attraction between the nucleus and the outermost electron  
b) The atom is very large  
**c) Strong attraction between the nucleus and the outermost electron.**  
d) The atom is likely to form an anion.
  3. The unit commonly used for ionization energy is:  
a) Grams per mole (g/mol)  
**b) Kilojoules per mole (kJ/mol)**  
c) Joules (J)  
d) Newtons (N)
- 
1. Down a group, ionization energy generally:  
a) Increases  
**b) Decreases**

- c) Remains constant
  - d) Changes unpredictably
2. The decrease in ionization energy down a group is primarily due to
- a) Decreasing atomic size.
  - b) Increasing shielding effect.**
  - c) Decreasing nuclear charge.
  - d) Increasing electronegativity.
3. Across a period, ionization energy generally:
- a) Increases**
  - b) Decreases
  - c) Remains constant
  - d) First decreases then increases
4. Across a period, ionization energy generally:
- a) Increases**
  - b) Decreases
  - c) Remains constant
  - d) First decreases then increases
5. Which of the following would have the highest first ionization energy?
- a) Li
  - b) Be**
  - c) B
  - d) They are all equal

### Short Questions

**Q1: Define ionization energy. Write the general equation representing the first ionization energy.**

Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom in its ground state. It is represented by the equation:  $\text{M(g)} + \text{Ionization Energy} \rightarrow \text{M}^+(\text{g}) + \text{e}^-$  where M is a neutral gaseous atom.

**Q2: What does ionization energy measure about the relationship between an atom and its electrons?**

Ionization energy is a direct measure of how strongly an atom's nucleus attracts and holds onto its outermost electron(s). A high ionization energy indicates a strong attraction (the electron is hard to remove), often seen in small atoms with high nuclear charge. A low ionization energy indicates a weak attraction (the electron is easy to remove), typical of large atoms with strong shielding.

**Q3: Explain the trend in ionization energy as we move down a group.**

Ionization energy decreases down a group. This is because each subsequent element has an additional electron shell. These inner electrons shield the valence electron from the nucleus's pull more effectively. Furthermore, the increased atomic radius means the valence electron is farther away. Both factors weaken the attraction, making the electron easier to remove.

**Q4: Explain the trend in ionization energy as we move across a period.**

Ionization energy increases across a period. While electrons are added to the same shell (constant shielding), protons are added to the nucleus, increasing the nuclear charge. This stronger positive pull holds the valence electrons more tightly, increasing the energy required to remove one. The atomic radius also decreases, enhancing this effect.

**Q5: Why does beryllium (Be) have a higher ionization energy than lithium (Li)? Both elements are in the same period ( $n=2$ ). However, beryllium has 4 protons in its nucleus compared to lithium's 3. This greater nuclear charge, experienced by electrons in the same shell with similar shielding, exerts a stronger pull on the valence electrons. Therefore, more energy is required to remove an electron from beryllium than from lithium.**

## **Electron Affinity**

### **Multiple Choice Questions (MCQs)**

1. Electron affinity is defined as the energy:
  - a) Required to remove an electron from an atom.
  - b) Required to add an electron to an atom.
  - c) **Released when an electron is added to an isolated gaseous atom.**
  - d) Of an electron in its orbital.
2. Electron affinity helps to explain the formation of:
  - a) Cations
  - b) **Anions**
  - c) Free radicals
  - d) Neutrons
3. Which of the following factors does NOT affect electron affinity?
  - a) Nuclear charge
  - b) Atomic radius
  - c) Shielding effect
  - d) **Physical state (solid, liquid, gas) of the element**
4. Across a period, electron affinity generally:

- a) Decreases
  - b) **Increases**
  - c) Remains constant
  - d) Becomes zero
5. Which group of elements has the highest electron affinity values?
- a) Alkali metals (Group 1)
  - b) Alkaline earth metals (Group 2)
  - c) **Halogens (Group 17)**
  - d) Noble gases (Group 18)
6. Down a group, electron affinity generally:
- a) Increases
  - b) **Decreases**
  - c) Remains constant
  - d) First increases then decreases
7. The decrease in electron affinity down a group is due to:
- a) Decreasing atomic radius.
  - b) **Increasing shielding effect and atomic radius.**
  - c) Decreasing nuclear charge.
  - d) Increasing ionization energy.

### Short Questions

**Q1: Define electron affinity. How does it relate to the formation of anions?**

Electron affinity is the amount of energy released when an electron is added to an isolated gaseous atom to form a uni-negative ion ( $X(g) + e^- \rightarrow X^-(g) + \text{energy}$ ). A high, negative electron affinity (meaning more energy is released) indicates a strong tendency to gain an electron and form a stable anion, which is characteristic of nonmetals like halogens.

**Q2: List the three factors that affect electron affinity.**

The three main factors affecting electron affinity are:

1. **Nuclear charge:** A higher nuclear charge increases the attraction for an extra electron.
2. **Atomic radius:** A smaller radius allows the nucleus to attract an extra electron more strongly.
3. **Shielding effect:** More inner-shell electrons shield the valence electrons from the nucleus, reducing the attraction for an extra electron.

**Q3: Explain the trend in electron affinity across a period.**

Electron affinity generally becomes more negative (increases) across a period from left to right. The increasing nuclear charge and decreasing atomic radius result in a stronger attraction for an incoming electron. This means more energy is released when the electron is added, making the process more favorable for elements on the right side of the periodic table.

**Q4: Explain the trend in electron affinity down a group.**

Electron affinity generally becomes less negative (decreases) down a group. Although nuclear charge increases, the atomic radius increases significantly, and the shielding effect of inner electrons becomes more pronounced. The incoming electron is farther from the nucleus and experiences less attraction, so less energy is released upon its addition.

**Q5: Why do halogens (Group 17) have the highest electron affinities in their respective periods?**

Halogens have seven valence electrons. Adding one electron gives them a stable, full octet configuration. Combined with their high effective nuclear charge and relatively small atomic size within their period, the nucleus exerts a very strong pull on an incoming electron. This results in a large release of energy, giving them the highest electron affinity.

## **Electronegativity and Metallic Character**

### **Multiple Choice Questions (MCQs)**

1. Electronegativity is defined as the ability of an atom to:
  - a) Lose electrons.
  - b) Gain protons.
  - c) **Attract electrons toward itself in a chemical bond.**
  - d) Conduct electricity.
2. The concept of electronegativity and a scale for it was primarily devised by:
  - a) Henry Moseley
  - b) Dmitri Mendeleev
  - c) **Linus Pauling**
  - d) John Dalton
3. The reactivity of alkali metals (Group 1) \_\_\_\_\_ down the group.
  - a) **Increases**
  - b) Decreases
  - c) Remains constant
  - d) First increases then decreases
4. Which alkali metal is the least reactive?

- a) **Lithium (Li)**
  - b) Sodium (Na)
  - c) Potassium (K)
  - d) Francium (Fr)
5. Which is more reactive: Magnesium (Mg) or Calcium (Ca)?
- a) Magnesium
  - b) **Calcium**
  - c) They are equally reactive
  - d) Cannot be determined
6. Metallic character is defined as the tendency of an element to:
- a) Gain electrons and form anions.
  - b) **Lose electrons and form cations.**
  - c) Share electrons equally.
  - d) Attract electrons in a bond.
7. Metallic character increases down a group because:
- a) Ionization energy increases.
  - b) atomic radius decreases.
  - c) **Valence electrons are easier to lose.**
  - d) nuclear charge decreases.
8. Metallic character \_\_\_\_\_ across a period.
- a) Increases
  - b) **Decreases**
  - c) Remains constant
  - d) Varies randomly
9. The decrease in metallic character across a period is due to
- a) Decreasing atomic size and increasing effective nuclear charge.
  - b) Increasing atomic size and decreasing nuclear charge.
  - c) Constant shielding effect.
  - d) Decreasing number of valence electrons.
10. The element with the highest metallic character in a period is found on the:
- a) Far left
  - b) Far right
  - c) **Middle**
  - d) It varies

## Short Questions

### Q1: Define electronegativity. Who devised the most common scale for it?

Electronegativity is the ability of an atom to attract electrons toward itself when it is part of a chemical bond. It is a relative measure, not an absolute one. The American chemist **Linus Pauling** devised the most common and widely used scale for calculating and comparing the electronegativities of elements.

### Q2: Explain why the reactivity of alkali metals increases down Group 1.

Reactivity increases down Group 1 due to the increasing atomic radius and shielding effect. The single valence electron is farther from the nucleus and less strongly attracted. This makes it progressively easier to lose that electron to form a cation, which is the basis of their chemical reactions. Therefore, potassium is more reactive than sodium, which is more reactive than lithium.

### Q3: Define metallic character. What is the general trend for metallic character down a group and across a period?

Metallic character is the tendency of an element to lose electrons and form positive ions (cations).

- **Down a group:** Metallic character **increases** because atomic size increases and shielding effect improves, making it easier to lose electrons.
- **Across a period (left to right):** Metallic character **decreases** because effective nuclear charge increases, making it harder to lose electrons.

### Q4: Why does sodium have a greater metallic character than lithium?

Sodium is below lithium in Group 1. Sodium has more electron shells, so its valence electron is farther from the nucleus and is better shielded by inner electrons. This weaker attraction makes it easier for sodium to lose its valence electron compared to lithium, giving it a greater metallic character.

### Q5: Predict which element has higher metallic character: Aluminum (Al) or Silicon (Si). Justify your answer.

Aluminum (Al) has a higher metallic character than Silicon (Si). Both are in the same period (period 3). Aluminum is further to the left (Group 13) than silicon (Group 14). As we move left to right across a period, effective nuclear charge increases, making it harder to lose electrons. Therefore, aluminum, being on the left, loses electrons more easily and has a higher metallic character than silicon.

# Reactivity, Density and Characteristic properties of Alkali Metals

## Multiple Choice Questions (MCQs)

1. The reactivity of elements generally \_\_\_\_\_ down a group.
  - a) Decreases
  - b) **Increases**
  - c) Remains constant
  - d) Becomes zero
2. Elements on the left side of a period (e.g., Group 1) are reactive because they:
  - a) **Readily lose electrons.**
  - b) Readily gain electrons.
  - c) Are inert.
  - d) Have high ionization energy.
3. Elements on the right side of a period (e.g., Group 17) are reactive because they:
  - a) Readily lose electrons
  - b) **Readily gain electrons.**
  - c) Are gases.
  - d) Have low electron affinity.
4. The density of elements generally \_\_\_\_\_ down a group.
  - a) **Increases**
  - b) Decreases
  - c) Remains constant
  - d) Becomes zero
5. Which of the following is a characteristic property of alkali metals?
  - a) High density
  - b) **Low melting point**
  - c) Hardness
  - d) Low reactivity
6. The products of the reaction between an alkali metal and water are:
  - a) Metal oxide and hydrogen gas
  - b) **Metal hydroxide and hydrogen gas.**
  - c) Metal salt and oxygen gas.
  - d) Metal oxide and water.
7. Which alkali metal is softer: Sodium (Na) or Potassium (K)?

- a) Sodium
  - b) **Potassium**
  - c) They are equally soft
  - d) Hardness cannot be compared
8. What happens to the melting and boiling points of alkali metals down Group 1?
- a) They increase.
  - b) **They decrease.**
  - c) They remain constant.
  - d) They change unpredictably
9. The primary reason we can predict the properties of an unknown element is
- a) Its color
  - b) **Its position in the periodic table and electronic configuration**
  - c) Its state of matter
  - d) Its discoverer's name
10. If a new element in Group 1 were discovered below Francium, we would predict it to be:
- a) Harder and less reactive than lithium.
  - b) **Softer and more reactive than francium.**
  - c) A gas at room temperature.
  - d) Unable to conduct electricity.

## Short Questions

**Q1: Explain the general trend in reactivity down a group for both metals and non-metals.**

For **metals**, reactivity increases down a group because atomic size increases, making it easier to lose electrons. For **non-metals**, reactivity decreases down a group because atomic size increases, making it harder to gain electrons. In both cases, the change in atomic size (and thus the ease of losing or gaining electrons) dictates the trend.

**Q2: Explain the trend in density down a group.**

Density generally increases down a group. This occurs because the mass of the atom (due to more protons and neutrons) increases significantly with each period. Although the atomic size also increases (volume), the mass increase is more significant relative to the volume increase. This results in a higher mass being packed into a unit volume, which translates to higher density.

**Q3: List three characteristic properties of alkali metals (Group 1).**

Three characteristic properties of alkali metals are:

1. **High Reactivity:** They readily lose their one valence electron to form +1 ions.

2. **Softness and Low Density:** They are soft enough to be cut with a knife and have low densities.
3. **Low Melting and Boiling Points:** They have relatively low melting and boiling points compared to other metals.

**Q4: Predict the properties of rubidium (Rb) based on its position below potassium in Group 1.**

Based on trends, rubidium will be softer and have a lower melting and boiling point than potassium. It will be more reactive, resulting in a much more violent reaction with water to produce rubidium hydroxide and hydrogen gas. Its density will also be higher than that of potassium.

**Q5: How can the periodic table be used to predict the properties and position of an unknown element?**

By analyzing an unknown element's electron configuration, we can determine its period (from the highest principal quantum number) and its group/block (from the valence electrons and the last orbital filled). Once its position is known, we can use the established trends (e.g., atomic size, ionization energy, reactivity) for that group and period to accurately predict its physical and chemical properties.

## **Transition Elements and Lanthanides and Actinides**

### **Multiple Choice Questions (MCQs)**

1. An element with the valence shell configuration  $4s^1$  is likely to be in:  
a) **Period 4, Group 1**  
b) Period 1, Group 4  
c) Period 4, Group 2  
d) Period 4, Group 17
2. Elements located in the d-block (Groups 3-12) are called:  
a) Alkali metals  
b) Halogens  
c) **Transition elements**  
d) Noble gases
3. Which of the following is a characteristic property of transition elements?  
a) Low density  
b) Low melting points  
c) **High density**  
d) Inertness

4. Transition elements exhibit variable oxidation states because:
- a) They are gases.
  - b) They have filled d-orbitals.
  - c) Both s and d sub-shells can participate in bonding.**
  - d) They have a stable noble gas configuration.
5. The industrial process for synthesizing ammonia (Haber process) uses which transition metal as a catalyst?
- a) Platinum (Pt)
  - b) Nickel (Ni)
  - c) Iron (Fe)**
  - d) Tungsten (W)
6. Colored compounds are a characteristic of:
- a) Alkali metals
  - b) Noble gases
  - c) Transition elements**
  - d) Halogens
7. The series of 14 elements with atomic numbers 57 to 71 are called:
- a) Actinides
  - b) Alkaline earth metals
  - c) Lanthanides**
  - d) Halogens
8. The elements in Group 17 are called:
- a) Alkali metals
  - b) Noble gases
  - c) Halogens**
  - d) Transition metals
9. Which halogen is a red-brown liquid at room temperature?
- a) Fluorine
  - b) Chlorine
  - c) Bromine**
  - d) Iodine
10. Halogens tend to form ions with a charge of:
- a) +1
  - b) -1**
  - c) +7

d) -7

## Short Questions

**Q1: How can the valence electron configuration be used to predict the group and period of an unknown element? Use the example  $4s^1$ .**

The valence configuration  $4s^1$  indicates the last electron enters the s-orbital, so the element is in the s-block. The number of valence electrons is 1, so for s-block, this means **Group 1**. The highest principal quantum number is  $n=4$ , which indicates the **4th period**. Therefore, the element is Potassium (K).

**Q2: List any three characteristic properties of transition elements.**

Three characteristic properties of transition elements are:

1. **High Density and Melting Points:** Due to strong metallic bonding and closely packed atoms.
2. **Variable Oxidation States:** Due to the involvement of both  $ns$  and  $(n-1)d$  electrons in bonding.
3. **Formation of Colored Compounds:** Due to the absorption of light when electrons transition within the partially filled d-subshell.

**Q3: Why do transition elements make good industrial catalysts? Provide one example.**

Transition elements can exhibit variable oxidation states and can form unstable intermediate compounds, providing an alternative path for reactions with lower activation energy. For example, **iron (Fe)** is used as a catalyst in the **Haber process** for the synthesis of ammonia from nitrogen and hydrogen.

**Q4: What are lanthanides and actinides? Where are they located in the periodic table?**

**Lanthanides** (atomic numbers 57-71) and **Actinides** (atomic numbers 89-103) are two series of f-block elements known as inner transition metals. They are located separately at the bottom of the periodic table to keep the table compact. Lanthanides are often called rare earth elements, and most actinides are radioactive.

**Q5: Describe the trend in the physical appearance of halogens down Group 17.**

Down Group 17, the halogens become darker in color and their physical state changes at room temperature. **Fluorine** is a pale-yellow gas, **Chlorine** is a yellow-green gas, **Bromine** is a red-brown liquid, and **Iodine** is a grey-black solid that produces purple vapors upon heating. This trend is due to increasing molecular size and intermolecular forces.

# Halogens

## Multiple Choice Questions (MCQs)

1. As you move down Group 17, the density of the halogens:
  - a) Decreases
  - b) **Increases**
  - c) Remains constant
  - d) First increases then decreases
2. The primary reason for the increase in density down the halogen group is that:
  - a) atomic size decreases.
  - b) **The increase in mass is greater than the increase in volume.**
  - c) The molecules become lighter.
  - d) They change from solid to gas.
3. **Which halogen has the highest density?**
  - a) Fluorine
  - b) Chlorine
  - c) Bromine
  - d) **Iodine**
4. **The reactivity of halogens is directly related to their ability to:**
  - a) Lose electrons.
  - b) **Gain electrons.**
  - c) Lose protons.
  - d) Gain neutrons.
5. The most reactive halogen is:
  - a) **Fluorine**
  - b) Chlorine
  - c) Bromine
  - d) Iodine
6. The reactivity of halogens \_\_\_\_\_ down the group.
  - a) Increases
  - b) **Decreases**
  - c) Remains constant
  - d) Becomes zero

7. A more reactive halogen can displace a less reactive halogen from its aqueous salt solution. This is a \_\_\_\_\_ reaction.
- a) Combination
  - b) Decomposition
  - c) **Displacement**
  - d) Double displacement
8. Which of the following statements is true regarding the thermal stability of hydrogen halides?
- a) HF has the weakest bond and lowest stability.
  - b) **HF has the strongest bond and highest stability.**
  - c) HI has the strongest bond and highest stability.
  - d) All H-X bonds have equal strength.
9. The correct order of decreasing thermal stability for hydrogen halides is
- a)  $\text{HI} > \text{HBr} > \text{HCl} > \text{HF}$
  - b)  **$\text{HF} > \text{HCl} > \text{HBr} > \text{HI}$**
  - c)  $\text{HCl} > \text{HF} > \text{HBr} > \text{HI}$
  - d)  $\text{HBr} > \text{HI} > \text{HCl} > \text{HF}$
10. Which halogen CANNOT displace any other halogen from its salt solution?
- a) Fluorine
  - b) Chlorine
  - c) Bromine
  - d) **Iodine**

## Short Questions

**Q1: Explain the trend in density observed down the halogen group.**

Density increases down Group 17. Although atomic and molecular size increases, the mass of the atoms (due to more protons and neutrons) increases at a faster rate than the volume. This means more mass is packed into a unit volume. Furthermore, the change from gas ( $\text{F}_2$ ,  $\text{Cl}_2$ ) to liquid ( $\text{Br}_2$ ) to solid ( $\text{I}_2$ ) indicates stronger intermolecular forces, leading to tighter packing and higher density.

**Q2: Why does fluorine have the highest reactivity among the halogens?**

Fluorine has the smallest atomic size and the highest effective nuclear charge in its group. This gives it the highest electronegativity and the strongest tendency to gain an electron to achieve a stable octet. The small bond length of  $\text{F}_2$  also makes its bond relatively weak and easy to break, further contributing to its extreme reactivity.

**Q3: Describe a displacement reaction that demonstrates the trend in oxidizing power down Group 17.**

The trend in oxidizing power is  $F_2 > Cl_2 > Br_2 > I_2$ . This can be demonstrated by chlorine gas displacing bromide ions from a solution:  $Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq/l)$ . The pale yellow-green chlorine gas will react to form orange-brown bromine, proving chlorine is a stronger oxidizing agent than bromine.

**Q4: Explain the trend in the thermal stability of the hydrogen halides (H-X bonds).**

Thermal stability decreases down the group:  $HF > HCl > HBr > HI$ . This is because the bond strength between hydrogen and the halogen decreases. The bond length increases down the group, and the electronegativity difference decreases, leading to a weaker, longer bond that is easier to break with heat.

**Q5: Predict the physical state and likely color of the halogen astatine (At) at room temperature.**

Following the trend down Group 17, astatine, being below iodine, should be a solid at room temperature. The color of halogens darkens down the group (pale yellow  $\rightarrow$  greenish yellow  $\rightarrow$  red brown  $\rightarrow$  dark grey/black). Therefore, astatine is predicted to be a very dark solid, likely black.

## Noble Gases and Metals, Non-Metals

### Multiple Choice Questions (MCQs)

- The elements in Group 18 are known as:
  - Halogens
  - Alkali metals
  - Noble gases**
  - Chalcogens
- The general electron configuration for noble gases (except He) is
  - $ns^2$
  - $ns^2 np^6$**
  - $ns^2 np^5$
  - $ns^1$
- The key reason for the low reactivity of noble gases is:
  - High ionization energy.
  - A complete valence shell electron configuration.**
  - Large atomic size.

- d) Low density.
4. Which noble gas has the electron configuration  $1s^2$ ?
- a) Neon
  - b) **Helium**
  - c) Argon
  - d) Krypton
5. Which noble gas is used in neon signs?
- a) Helium
  - b) Argon
  - c) **Neon**
  - d) Krypton
6. Which of the following is a good conductor of electricity?
- a) Sulfur (non-metal)
  - b) **Copper (metal)**
  - c) Bromine (halogen)
  - d) Argon (noble gas)
7. Metals are malleable because:
- a) They have free electrons.
  - b) **Metallic bonds allow layers of atoms to slide.**
  - c) They have high melting points.
  - d) They form acidic oxides.
8. Non-metals are generally:
- a) Malleable and ductile
  - b) **Brittle**
  - c) Good conductors of heat
  - d) Shiny
9. The high melting points of metals are primarily due to
- a) Weak van der Waals forces.
  - b) **Strong metallic bonding.**
  - c) Covalent bonding.
  - d) Hydrogen bonding.
10. A common use of noble gases like argon is
- a) Making acids.
  - b) As a fuel source.
  - c) **As an inert shielding gas in welding.**
  - d) Making fertilizers.

## Short Questions

**Q1: Why are noble gases chemically inert? What is the exception to their electron configuration?**

Noble gases are inert because they have a stable electron configuration with a completely filled valence shell (octet rule). This makes them energetically stable with no tendency to gain, lose, or share electrons. The exception is helium, which has a complete valence shell with only 2 electrons (duet rule:  $1s^2$ ).

**Q2: List two important uses of noble gases and state the property that makes them suitable for each use.**

1. **Neon Signs:** Neon gas is used in advertising signs because it emits a characteristic bright red-orange glow when an electric current is passed through it (electrical discharge).
2. **Inert Shielding Gas:** Argon is used in welding to shield the hot metal from oxygen and nitrogen in the air, preventing oxidation. This use relies on its extreme inertness and inability to react.

**Q3: Compare metals and non-metals based on their electrical conductivity and explain the reason for the difference.**

**Metals** are good conductors of electricity due to the presence of a "sea of delocalized electrons" that are free to move and carry an electric current. **Non-metals** are poor conductors (insulators) because their electrons are tightly bound in covalent bonds or localized, leaving no free charges to move and conduct electricity.

**Q4: Explain why metals are malleable but non-metals are brittle.**

Metals are malleable because layers of positive ions can slide over one another without breaking the structure, as the sea of delocalized electrons readily redistributes itself to maintain the metallic bonds. Non-metals are brittle because they are held by rigid, directional covalent bonds. Applying force shatters this rigid structure instead of bending it.

**Q5: What is the difference between thermal and electrical conductivity in metals? Why do metals generally possess both?**

**Electrical conductivity** is the ability to conduct an electric current, while **thermal conductivity** is the ability to conduct heat. Metals possess both properties for the same fundamental reason: the presence of mobile, delocalized electrons. These free electrons can transfer kinetic energy (heat) and also carry charge (electricity) efficiently through the metal lattice.

## Exercise

### 1. Encircle the correct answer.

(i) Number of periods in the periodic table are:

(a) 8

**(b) 7**

(c) 16

(d) 5

(ii) Which of the following groups contain alkaline earth metals?

(a) 1A

**(b) IIA**

(c) VIIA

(d) VIIIA

(iii) Which of the following elements belongs to VIIIA?

(a) Na

(b) Mg

(c) Br

**(d) Xe**

(iv) Main group elements are arranged in \_\_\_\_\_ groups.

(a) 6

(b) 7

**(c) 8**

(d) 10

(v) Period number of Al is:

(a) 1

(b) 2

**(c) 3**

(d) 4

(vi) Valence shell electronic configuration of an element M (atomic no. 14) is:

(a)  $(2s^2, 2p^1)$

(b)  $(2s^2, 2p^2)$

(c)  $(2s^2, 2p^3)$

**(d)  $(3s^2, 3p^2)$**

(vii) Which of the following elements you expect to have greater shielding effect?

(a) Li

b) Na

(c) K

**(d) Rb**

(viii) As you move from right to left across a period, which of the following does not increase:

(a) electron affinity

(b) ionization energy

(c) nuclear charge

**(d) shielding effect**

(ix) All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:

**(a) high ionization energies**

(b) relatively greater atomic size

(c) similar electronic configuration

(d) decreased nuclear charge

## 2. Give short answer.

**(i) Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.**

The element is in period 3, so its outermost electrons are in the  $n=3$  shell. As a Group IIIA (or Group 13) element, it has 3 valence electrons. Therefore, its valence shell electronic configuration is  $3s^2 3p^1$ .

**(ii) Define halogens.**

Halogens are the elements found in **Group VIIA (or Group 17)** of the periodic table. They are highly reactive non-metals that have **seven valence electrons ( $ns^2 np^5$ )** and readily form  $-1$  ions. They are known for forming salts with metals (e.g., NaCl).

**(iii) Which atom has higher shielding effect, Li or Na?**

**Na has a higher shielding effect** than Li. This is because sodium has three electron shells ( $n=1, 2, 3$ ), while lithium only has two ( $n=1, 2$ ). The inner shells of electrons in sodium more effectively shield the valence electron from the pull of the nucleus.

**(iv) Explain why Na has higher ionization energy than K?**

Ionization energy decreases down a group. **Sodium has a higher ionization energy than potassium** because potassium has an extra electron shell. This increases the atomic radius and the shielding effect, making its valence electron easier to remove than the valence electron in the smaller sodium atom.

**(v) Alkali metals belong to S-block in the periodic table, why?**

Alkali metals belong to the s-block because their **last or differentiating electron enters the s-orbital**. Their general electronic configuration is  $ns^1$ , meaning they have one electron in their outermost s orbital, which defines their position in the s-block.

## 3. Arrange the elements in order of increasing ionization energy:

**(a) Li, Na, K** Ionization energy decreases down a group. **Order:  $K < Na < Li$**

**(b) Cl, Br, I** Ionization energy decreases down a group. **Order:  $I < Br < Cl$**

**4. Arrange the elements in order of decreasing shielding effect:**

Shielding effect increases with the number of inner electron shells.

**(a) Li, Na, K Order:  $K > Na > Li$**

**(b) Cl, Br, I Order:  $I > Br > Cl$**

**(c) Cl, Br Order:  $Br > Cl$**

**5. Specify which of the following elements you would expect to have the greatest electron affinity:**

**S, P, Cl** Electron affinity generally increases across a period. Among these three (all in period 3), **Chlorine (Cl)** has the highest electron affinity. It has a strong effective nuclear charge to attract an electron, and adding one electron gives it a stable noble gas configuration.

**6. Group the elements in pairs that would represent similar chemical properties.**

Elements with the same number of valence electrons have similar chemical properties.

- **A ( $1s^2 2s^2$ )** and **F ( $1s^2 2s^2$ )**: Both are Group II elements (Be and Be? Likely a typo, but same configuration).
- **B ( $1s^2 2s^2 2p^6$ )** and **H ( $1s^2 2s^2 2p^6 3s^2$ )**: Both have a full valence shell configuration (Ne and  $Mg^{2+}$ , but similar inertness).
- **C ( $1s^2 2s^2 2p^3$ )** and **E ( $1s^2 2s^2 2p^6 3s^2 3p^3$ )**: Both have  $ns^2 np^3$  configuration (Group VA: N and P).
- **D ( $1s^2$ )** is Helium, a noble gas. It is paired with other noble gases like B, but B is Neon. So D is unique.
- **G ( $1s^2 2s^2 2p^6 3s^2$ )** is Magnesium, a Group II element. It could be paired with A and F, but they are in  $n=2$ .

**Better pairing based on groups:**

- **Pair 1:** C (N, Group VA) and E (P, Group VA)
- **Pair 2:** A (Be, Group II) and G (Mg, Group II)
- **Pair 3:** B (Ne, Noble Gas) and H (Mg, but has noble gas core of Ne)
- D (He, Noble Gas) is similar to B and H in being chemically inert.
- F is a duplicate of A.

**7. Arrange the elements in groups and periods in Q. No. 6.**

- **A ( $1s^2 2s^2$ ):** Period 2, Group IIA
- **B ( $1s^2 2s^2 2p^6$ ):** Period 2, Group VIIIA (Noble Gas)
- **C ( $1s^2 2s^2 2p^3$ ):** Period 2, Group VA
- **D ( $1s^2$ ):** Period 1, Group VIIIA (Noble Gas)
- **E ( $1s^2 2s^2 2p^6 3s^2 3p^3$ ):** Period 3, Group VA
- **F ( $1s^2 2s^2$ ):** Period 2, Group IIA (Duplicate of A)
- **G ( $1s^2 2s^2 2p^6 3s^2$ ):** Period 3, Group IIA
- **H ( $1s^2 2s^2 2p^6 3s^2$ ):** Period 3, Group IIA (Duplicate of G)

**8. Find the group number.**

The rule states: Group Number = Number of Valence Electrons

- **$_{27}\text{Co}$ :** Electronic config:  $[\text{Ar}] 4s^2 3d^7$ . Valence electrons = 2. **Group IIB** (It's a transition metal, so the rule doesn't hold perfectly. Valence  $e^-$  are 2).
- **$_{32}\text{S}$ :** Electronic config:  $1s^2 2s^2 2p^6 3s^2 3p^4$ . Valence electrons = 6. **Group VIA**
- **$_{39}\text{K}$ :** Electronic config:  $[\text{Ar}] 4s^1$ . Valence electrons = 1. **Group IA**
- **$_8\text{O}$ :** Electronic config:  $1s^2 2s^2 2p^4$ . Valence electrons = 6. **Group VIA**
- **$_{13}\text{Al}$ :** Electronic config:  $1s^2 2s^2 2p^6 3s^2 3p^1$ . Valence electrons = 3. **Group IIIA**
- **$_{16}\text{S}$ :** Electronic config:  $1s^2 2s^2 2p^6 3s^2 3p^4$ . Valence electrons = 6. **Group VIA**
- **$_{19}\text{K}$ :** Electronic config:  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ . Valence electrons = 1. **Group IA**
- **$_6\text{C}$ :** Electronic config:  $1s^2 2s^2 2p^2$ . Valence electrons = 4. **Group IVA**

**9. Write the valence shell electronic configuration:**

- Alkali metals (Group IA):**  $ns^1$
- Alkaline earth metals (Group IIA):**  $ns^2$
- Halogens (Group VIIA):**  $ns^2 np^5$
- Noble gases (Group VIIIA):**  $ns^2 np^6$  (He is  $1s^2$ )

**10. Write electron dot symbols:**

- Be** (2 valence  $e^-$ ): Be:
- K** (1 valence  $e^-$ ): K•

(c) **N** (5 valence  $e^-$ ):  $\cdot\text{N}\cdot$  (with 3 dots, one on each side)

(d) **I** (7 valence  $e^-$ ):  $\cdot\dot{\text{I}}\cdot$  (with 5 dots around it)

**3. Write the valence shell electronic configuration:**

(a) Period 3, Group VA:  **$3s^2 3p^3$**

(b) Period 2, Group VIA:  **$2s^2 2p^4$**

**4. Complete the table:**

Atomic number	Mass number	No. of protons	No. of neutrons	No. of electrons
11	23	11	12	11
15	29	14	15	14
22	47	22	25	22
13	27	13	14	13
<i>Explanation: Atomic Number = Protons = Electrons. Neutrons = Mass Number - Atomic Number.</i>				

**13. Identify block, group, and period:**

(a)  **$1s^2 2s^1$**

- Last electron in s-orbital  $\rightarrow$  **s-block**
- **Valence** configuration is  $2s^1 \rightarrow$  **Group 1 (IA)**
- Highest n is 2  $\rightarrow$  **Period 2** (Element: Lithium)

(b)  **$1s^2 2s^2 2p^1$**

- Last electron in p-orbital  $\rightarrow$  **p-block**
- Valence electrons = 2 (s) + 1 (p) = 3  $\rightarrow$  **Group 13 (IIIA)**

- Highest  $n$  is 2  $\rightarrow$  **Period 2** (Element: Boron)

(c)  $1s^2 2s^2 2p^6 3s^2$

- Last electron in s-orbital  $\rightarrow$  **s-block**
- Valence configuration is  $3s^2 \rightarrow$  **Group 2 (IIA)**
- Highest  $n$  is 3  $\rightarrow$  **Period 3** (Element: Magnesium)

(d)  $1s^2$

- Last electron in s-orbital  $\rightarrow$  **s-block**
- Valence shell is full of 2 electrons  $\rightarrow$  **Group 18 (VIIIA)**
- Highest  $n$  is 1  $\rightarrow$  **Period 1** (Element: Helium)