Class: 12th

Subject: Chemistry

Chapter 1: Periodic Classification Of Elements And Periodicity

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🩸 Periodic Table – Key Points Based MCQs (Exam-Focused)

1. Who gave the most comprehensive classification of elements?

(a) Newland

(b) Dobereiner

(c) Mendeleev ✅

(d) Moseley

2. In Mendeleev's periodic table, elements were arranged by:

(a) Atomic number

(b) Atomic mass ✅

(c) Valency

(d) Electronegativity

3. Modern periodic law is based on:

(a) Atomic mass

(b) Valency

(c) Atomic number ✅

(d) Group number

4. Vertical columns in the periodic table are called:

(a) Periods

(b) Blocks

(c) Groups ✅

(d) Families

5. How many periods are there in the modern periodic table?

(a) 6

(b) 7 ✅

(c) 8

(d) 9

6. Group IA elements are known as:

(a) Halogens

(b) Noble gases

(c) Alkali metals ✅

(d) Transition metals

7. Group IIA elements are called:

(a) Alkaline-earth metals ✅

(b) Halogens

(c) Noble gases

(d) Coinage metals

8. Group VIIA elements are known as:

(a) Noble gases

(b) Halogens ✅

(c) Alkali metals

(d) Metalloids

9. Group VIIIA elements are called:

(a) Halogens

(b) Noble gases ✅

(c) Transition metals

(d) Alkali metals

10. Elements are classified into blocks based on:

(a) Atomic mass

(b) Valency

(c) Valence orbital ✅

(d) Period number

11. Which block contains transition elements?

(a) s-block

(b) p-block

(c) d-block ✅

(d) f-block

12. Atomic radius increases:

(a) Across a period

(b) Down a group ✅

(c) With ionization energy

(d) With electronegativity

13. Which ion is smaller than its parent atom?

(a) Negative ion

(b) Positive ion ✅

(c) Neutral atom

(d) All ions

14. Ionization energy increases:

(a) Down a group

(b) Across a period ✅

(c) With atomic size

(d) With metallic character

15. Electron affinity decreases:

(a) Across a period

(b) With atomic number

(c) Down a group ✅

(d) With electronegativity

16. Metallic character increases:

(a) Across a period

(b) Down a group ✅

(c) With ionization energy

(d) With atomic number

17. Oxidation state of an element is related to:

(a) Period number

(b) Atomic mass

(c) Group number ✅

(d) Block type

18. Electrical conductance depends on:

(a) Atomic number

(b) Free electrons ✅

(c) Ionization energy

(d) Group number

19. Halides of group IA are:

(a) Covalent

(b) Polymeric

(c) Ionic ✅

(d) Amphoteric

20. Oxides are classified as:

(a) Ionic and covalent

(b) Acidic, basic, amphoteric ✅

(c) Metallic and non-metallic

(d) Organic and inorganic

🩸Important MCQs:

1. Who introduced the first regular periodic table in 1871?

(a) Al-Razi

(b) Dobereiner

(c) Newland

(d) Mendeleev ✅

2. Which scientist proposed the Law of Octaves?

(a) Mendeleev

(b) Newland ✅

(c) Moseley

(d) Dobereiner

3. Dobereiner's classification was based on:

(a) Atomic number

(b) Triads ✅

(c) Octaves

(d) Periods

4. Periodic Law states that properties of elements are periodic functions of their:

(a) Atomic mass

(b) Atomic number ✅

(c) Valency

(d) Density

5. Who discovered the concept of atomic number?

(a) Mendeleev

(b) Newland

(c) Moseley ✅

(d) Al-Razi

6. Which group was added in the modern periodic table to include noble gases?

(a) Group IA

(b) Group VIIA

(c) Group VIIIA ✅

(d) Group IB

7. In Mendeleev's table, elements were arranged by:

(a) Atomic number

(b) Atomic mass ✅

(c) Electron configuration

(d) Group number

8. Which element's properties were predicted by Mendeleev before its discovery?

(a) Argon

(b) Krypton

(c) Germanium ✅

(d) Neon

9. Be, Mg, Ca, Sr, Ba are placed in which group in the modern periodic table?

(a) Group IA

(b) Group IIA ✅

(c) Group IIB

(d) Group IIIA

10. Zn, Cd, Hg are placed in which group in the modern periodic table?

(a) Group IA

(b) Group IIA

(c) Group IIB ✅

(d) Group IIIA

11. Horizontal rows in the periodic table are called:

(a) Groups

(b) Periods ✅

(c) Blocks

(d) Series

12. Vertical columns in the periodic table are called:

(a) Periods

(b) Blocks

(c) Groups ✅

(d) Series

13. Al-Razi classified substances based on:

(a) Atomic number

(b) Atomic mass

(c) Physical and chemical properties ✅

(d) Valency

14. Which scientist classified 62 elements by increasing atomic mass in 1864?

(a) Dobereiner

(b) Newland ✅

(c) Mendeleev

(d) Moseley

15. The periodic table helps in studying:

(a) Electrolysis

(b) Periodic behavior of elements ✅

(c) Balancing equations

(d) Types of reactions

16. How many periods are there in the modern periodic table?

(a) 6

(b) 7 ✅

(c) 8

(d) 9

17. Which period contains only two elements?

(a) 2nd

(b) 3rd

(c) 1st ✅

(d) 7th

18. Which periods are called short periods?

(a) 1 and 2

(b) 2 and 3 ✅

(c) 4 and 5

(d) 6 and 7

19. How many elements are present in the 4th and 5th periods?

(a) 8

(b) 10

(c) 18 ✅

(d) 32

20. Lanthanides are placed in which period?

(a) 4th

(b) 5th

(c) 6th ✅

(d) 7th

21. Which group is known as Alkali Metals?

(a) Group IIA

(b) Group VIIA

(c) Group IA ✅

(d) Group VIIIA

22. Which group is called the Halogen family?

(a) Group IA

(b) Group IIA

(c) Group VIIA ✅

(d) Group VIIIA

23. Elements with valence electrons in d-orbitals are called:

(a) s-block elements

(b) p-block elements

(c) d-block elements ✅

(d) f-block elements

24. Which of the following is a metalloid?

(a) Na

(b) Si ✅

(c) Cl

(d) Ca

25. Elements with both metallic and non-metallic properties are called:

(a) Metals

(b) Non-metals

(c) Noble gases

(d) Metalloids ✅

26. Atomic radius increases down a group due to:

(a) Increase in nuclear charge

(b) Addition of electron shells ✅

(c) Decrease in shielding effect

(d) Increase in valency

27. Across a period, atomic radius decreases because:

(a) Shell number increases

(b) Electron repulsion increases

(c) Nuclear charge increases ✅

(d) Atomic number decreases

28. Lanthanide contraction refers to:

(a) Expansion of atomic size

(b) Gradual decrease in size of Lanthanides ✅

(c) Increase in ionic radius

(d) Increase in valency

29. A positive ion is always:

(a) Larger than its atom

(b) Smaller than its atom ✅

(c) Equal in size

(d) Unstable

30. A negative ion is always:

(a) Smaller than its atom

(b) Equal in size

(c) Larger than its atom ✅

(d) Neutral

31. In a group, ionic radius of similar charged ions:

(a) Decreases

(b) Increases ✅

(c) Remains constant

(d) First increases then decreases

32. Ionization energy is the energy required to:

(a) Add an electron

(b) Remove an electron ✅

(c) Break a bond

(d) Form a compound

33. Ionization energy decreases down a group due to:

(a) Increase in nuclear charge

(b) Increase in shielding effect ✅

(c) Decrease in electron number

(d) Decrease in atomic size

34. Ionization energy increases across a period because:

(a) Atomic size increases

(b) Shielding effect increases

(c) Nuclear charge increases ✅

(d) Shell number increases

35. Which element has the highest ionization energy in a period?

(a) Alkali metal

(b) Halogen

(c) Inert gas ✅

(d) Transition metal

36. Electron affinity is energy:

(a) Required to remove an electron

(b) Released or absorbed when an electron is added ✅

(c) Used to break a bond

(d) Needed to form a positive ion

37. Electron affinity is highest for:

(a) Noble gases

(b) Alkali metals

(c) Halogens ✅

(d) Transition metals

38. Second electron affinity is usually:

(a) Negative

(b) Positive ✅

(c) Zero

(d) Equal to first

39. Electron affinity increases across a period due to:

(a) Decrease in atomic number

(b) Increase in atomic size

(c) Increase in nuclear charge ✅

(d) Decrease in valency

40. Electron affinity decreases down a group because:

(a) Atomic size increases ✅

(b) Nuclear charge increases

(c) Shielding effect decreases

(d) Shell number decreases

41. Metals tend to:

(a) Gain electrons

(b) Lose electrons ✅

(c) Form acidic oxides

(d) Be poor conductors

42. Non-metals generally form:

(a) Basic oxides

(b) Neutral oxides

(c) Acidic oxides ✅

(d) Amphoteric oxides

43. Metallic character increases:

(a) Across a period

(b) Down a group ✅

(c) With ionization energy

(d) With electron affinity

44. Non-metallic character decreases:

(a) Across a period

(b) With atomic number

(c) Down a group ✅

(d) With nuclear charge

45. Which element is the most non-metallic?

(a) Oxygen

(b) Nitrogen

(c) Fluorine ✅

(d) Chlorine

46. Melting points of group IA elements are low because:

(a) They are non-metals

(b) They have large atomic size

(c) Each atom provides one bonding electron ✅

(d) They form covalent bonds

47. Which group has the highest melting point in short periods?

(a) Group IA

(b) Group IIA

(c) Group IVA ✅

(d) Group VIIA

48. Carbon has a high melting point in diamond due to:

(a) Metallic bonding

(b) Ionic bonding

(c) Giant covalent structure ✅

(d) Weak van der Waals forces

49. Melting and boiling points decrease down group IA because:

(a) Atomic size decreases

(b) Binding forces become stronger

(c) Atomic size increases ✅

(d) Number of valence electrons increases

50. In group VIIA, melting and boiling points increase down the group due to:

(a) Decrease in molecular mass

(b) Increase in polarity

(c) Increase in molecular size and polarizability ✅

(d) Decrease in nuclear charge

51. Oxidation state of sodium in NaCl is:

(a) -1

(b) 0

(c) +2

(d) +1 ✅

52. Oxidation state of an element in free state is:

(a) +1

(b) -1

(c) 0 ✅

(d) Depends on group number

53. Elements of group IIIA usually show oxidation state of:

(a) +1

(b) +2

(c) +3 ✅

(d) +4

54. Nitrogen can show oxidation states of:

(a) Only -3

(b) +3 and +5 ✅

(c) Only +5

(d) 0 and -1

55. Group VIIA elements mostly show oxidation state of:

(a) +1

(b) -1 ✅

(c) +7

(d) 0

56. Transition elements show variable oxidation states due to:

(a) Filled s-orbitals

(b) Inert pair effect

(c) Partly filled d-orbitals ✅

(d) High electronegativity

57. Which group contains coinage metals with high electrical conductance?

(a) Group IA

(b) Group IIA

(c) Group IB ✅

(d) Group VIIA

58. Which form of carbon is a good conductor of electricity?

(a) Diamond

(b) Graphite ✅

(c) Charcoal

(d) Fullerene

59. Electrical conductance of group IA metals:

(a) Decreases down the group

(b) Increases down the group ✅

(c) Remains constant

(d) Is unpredictable

60. Hydration energy depends on:

(a) Atomic number

(b) Charge-to-size ratio of ions ✅

(c) Number of neutrons

(d) Number of isotopes

61. Halides are binary compounds formed by:

(a) Hydrogen

(b) Oxygen

(c) Halogens ✅

(d) Noble gases

62. Halides of group IA elements are:

(a) Covalent

(b) Polymeric

(c) Ionic ✅

(d) Amphoteric

63. Which halide has the highest lattice energy?

(a) Chloride

(b) Bromide

(c) Iodide

(d) Fluoride ✅

64. Bonding in $SiCl_4$ is mostly:

(a) Ionic

(b) Metallic

(c) Polar covalent ✅

(d) Hydrogen bonded

65. Across a period, halides tend to become more:

(a) Ionic

(b) Covalent ✅

(c) Metallic

(d) Amphoteric

66. Which halide has the strongest van der Waals forces?

(a) Fluoride

(b) Chloride

(c) Bromide

(d) Iodide ✅

67. For a metal, the halide with highest ionic character is:

(a) Iodide

(b) Bromide

(c) Chloride

(d) Fluoride ✅

68. $PbCl_2$ is mainly:

(a) Covalent

(b) Ionic ✅

(c) Amphoteric

(d) Metallic

69. Hydrides of group IA elements are:

(a) Covalent

(b) Ionic ✅

(c) Polymeric

(d) Amphoteric

70. Hydrides of Be and Mg are:

(a) Ionic

(b) Covalent

(c) Intermediate ✅

(d) Metallic

71. Covalent hydrides are generally:

(a) Crystalline solids

(b) Non-conductors ✅

(c) Ionic compounds

(d) Amphoteric

72. Stability of covalent hydrides increases:

(a) Down a group

(b) Across a period ✅

(c) With atomic mass

(d) With metallic character

73. Most stable covalent hydride is formed by:

(a) Oxygen

(b) Fluorine ✅

(c) Nitrogen

(d) Sulphur

74. Hydrides like H_2O and HF have high boiling points due to:

(a) Ionic bonding

(b) Hydrogen bonding ✅

(c) Metallic bonding

(d) Van der Waals forces

75. Oxides of alkali metals are generally:

(a) Acidic

(b) Amphoteric

(c) Basic ✅

(d) Neutral

76. Oxides of non-metals are usually:

(a) Basic

(b) Acidic ✅

(c) Amphoteric

(d) Neutral

77. Oxides showing both acidic and basic behavior are called:

(a) Neutral oxides

(b) Amphoteric oxides ✅

(c) Metallic oxides

(d) Covalent oxides

78. Which of the following is an amphoteric oxide?

(a) Na_2O

(b) SO_3

(c) ZnO ✅

(d) CO_2

79. Hydrogen resembles alkali metals because it:

(a) Is a metal

(b) Has one electron in outer shell ✅

(c) Forms long chains

(d) Is a good conductor

80. Hydrogen resembles halogens because it:

(a) Forms positive ions

(b) Needs one electron to complete shell ✅

(c) Forms metallic bonds

(d) Is a noble gas

🩸 Periodic Table – Key Points Based Short Questions (Exam-Focused)

1. Who gave the most useful classification of elements?

Answer:

👉 Dmitri Mendeleev gave the most useful and comprehensive classification of elements.

2. On what basis did Mendeleev arrange the elements?

Answer:

👉 Mendeleev arranged the elements in ascending order of their atomic masses.

3. State the Modern Periodic Law.

Answer:

👉 "If elements are arranged in increasing order of their atomic numbers, their chemical properties repeat periodically."

4. What are groups in the periodic table?

Answer:

👉 The vertical columns in the periodic table are called groups.

5. What are periods in the periodic table?

Answer:

👉 The horizontal rows in the periodic table are called periods.

6. What are Alkali metals and Alkaline Earth metals?

Answer:

👉 Group IA elements are called Alkali metals, and Group IIA elements are called Alkaline Earth metals.

7. Why are Group VIIIA elements called noble gases?

Answer:

👉 They are called noble gases because they are least reactive and chemically stable.

8. How are elements classified into s, p, d, and f blocks?

Answer:

👉 Elements are classified into these blocks according to the type of valence orbital being filled.

9. How are elements classified based on their properties?

Answer:

👉 Elements are classified as metals, non-metals, and metalloids.

10. What is the trend of atomic radii in the periodic table?

Answer:

👉 Atomic radii increase from top to bottom in a group and decrease across a period.

11. How does the size of positive and negative ions compare with their atoms?

Answer:

👉 Positive ions are smaller, while negative ions are larger than their parent atoms.

12. What is the trend of ionization energy in the periodic table?

Answer:

👉 Ionization energy increases along a period and decreases down a group.

13. What is the trend of electron affinity in the periodic table?

Answer:

👉 Electron affinity increases across a period and decreases down a group.

14. How does metallic character change in the periodic table?

Answer:

👉 Metallic character increases down a group and decreases across a period.

15. On what factor does the electrical conductance of an element depend?

Answer:

👉 It depends on the number of free or movable electrons present in the element.

🩸 Important Short Questions:

1. What is the main purpose of the Periodic Table?

Answer:

👉 The main purpose of the Periodic Table is to systematically arrange the elements so that their physical and chemical properties can be studied easily according to their periodic behaviour.

2. Who are the main contributors to the development of the Periodic Table?

Answer:

👉 The main contributors to the development of the Periodic Table are Al-Razi, Dobereiner, Newland, and Mendeleev.

3. What is meant by Dobereiner's Triads?

Answer:

👉 Dobereiner arranged the elements into groups of three, called triads, in which all elements had similar chemical properties.

4. State Newland's Law of Octaves.

Answer:

👉 According to Newland's Law of Octaves, when elements are arranged in increasing order of their atomic masses, every eighth element has properties similar to the first one.

5. Who gave the first regular periodic table and in which year?

Answer:

👉 The first regular periodic table was given by Dmitri Mendeleev in 1871.

6. State Mendeleev's Periodic Law.

Answer:

👉 Mendeleev's Periodic Law states that "The properties of elements are a periodic function of their atomic masses."

7. What important prediction was made by Mendeleev?

Answer:

👉 Mendeleev predicted the existence and properties of undiscovered elements such as germanium, which were later discovered and matched his predictions.

8. Who discovered the atomic number and when?

Answer:

👉 The atomic number was discovered by Moseley in 1911.

9. On what basis are elements arranged in the modern periodic table?

Answer:

👉 In the modern periodic table, elements are arranged in increasing order of their atomic numbers instead of atomic masses.

10. Which new group was added in the modern periodic table?

Answer:

👉 The noble gases (Group VIIIA) were added to the modern periodic table, which were not known during Mendeleev's time.

11. How are elements arranged in the modern periodic table?

Answer:

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

12. What are groups and periods in the periodic table?

Answer:

👉 Vertical columns in the periodic table are called groups, and horizontal rows are called periods. There are 8 groups (I–VIII) and 7 periods (1–7).

13. What are short and long periods?

Answer:

👉 Short periods are the 2nd and 3rd periods containing 8 elements each, while long periods are the 4th and 5th periods containing 18 elements each.

14. What are Lanthanides and Actinides?

Answer:

👉 Lanthanides are a set of 14 elements starting after Lanthanum (La), and Actinides are 14 elements starting after Actinium (Ac).

👉 Both are shown separately at the bottom of the periodic table.

15. What are Alkali Metals and Alkaline Earth Metals?

Answer:

👉 Elements of Group IA are called Alkali Metals because they form strong alkalies with water, while elements of Group IIA are called Alkaline Earth Metals due to their alkaline nature and presence in Earth's crust.

16. What is meant by Atomic Radius?

Answer:

👉 Atomic radius is half the distance between the centres of two bonded atoms of the same element. It represents the size of the atom.

17. How does atomic radius change in a group and in a period?

Answer:

👉 In a group, atomic radius increases from top to bottom due to the addition of extra electron shells.

👉 In a period, atomic radius decreases from left to right due to increased nuclear charge.

18. What is meant by Lanthanide Contraction?

Answer:

👉 The gradual decrease in atomic size of Lanthanides with increasing atomic number is called Lanthanide Contraction.

19. What is Ionic Radius?

Answer:

👉 The Ionic Radius is the radius of an ion.

👉 A positive ion (cation) is always smaller than its parent atom, while a negative ion (anion) is larger than its parent atom.

20. Define Ionization Energy.

Answer:

👉 Ionization energy is the minimum amount of energy required to remove an electron from the outermost shell of a gaseous atom in its ground state.

21. How does ionization energy vary in a group and across a period?

Answer:

👉 Down a group, ionization energy decreases because atomic size and shielding effect increase.

👉 Across a period, ionization energy increases due to higher nuclear charge and smaller atomic size.

22. What is meant by Electron Affinity?

Answer:

👉 Electron affinity is the energy released or absorbed when an electron is added to a gaseous atom to form a negative ion.

23. How does electron affinity change in a group and across a period?

Answer:

👉 Across a period, electron affinity generally increases due to decreasing atomic size.

👉 Down a group, it decreases because atoms become larger and nuclear attraction becomes weaker.

24. What is meant by Metallic Character?

Answer:

👉 The tendency of an element to lose electrons and form positive ions is called its metallic character.

👉 It increases down a group and decreases across a period.

25. What is meant by Non-Metallic Character?

Answer:

👉 The tendency of an element to gain electrons and form negative ions is called non-metallic character.

👉 It decreases down a group and increases from left to right across a period.

26. What do melting and boiling points show about elements?

Answer:

👉 Melting and boiling points show the strength of bonds between atoms or molecules in an element.

27. How do melting and boiling points change across a period?

Answer:

👉 Across a period, they increase up to Group IVA and then decrease toward noble gases.

👉 Elements with giant covalent structures (like carbon) have very high melting points.

28. Define oxidation state.

Answer:

👉 The oxidation state is the charge an atom carries in a compound after gaining or losing electrons.

29. Why are metals good conductors of electricity?

Answer:

👉 Metals are good conductors because they have free-moving electrons in their outer shells that carry electric current easily.

30. What is hydration energy?

Answer:

👉 Hydration energy is the heat released when one mole of gaseous ions dissolves in water to form a dilute solution.

👉 It depends on the charge and size of the ions.

31. What are halides?

Answer:

👉 Halides are binary compounds formed by the combination of halogens with other elements.

32. How are halides classified based on bonding?

Answer:

👉 Halides are classified into three types:

1. Ionic halides
2. Covalent halides
3. Polymeric halides

33. What type of halides do Group IA elements form?

Answer:

👉 Group IA elements form purely ionic halides which are high melting point solids with three-dimensional lattices.

34. What are polymeric halides?

Answer:

👉 Polymeric halides are those in which the halogen atom acts as a bridge between two atoms of another element, forming layer or chain structures.

35. Define hydrides.

Answer:

👉 Hydrides are binary compounds of hydrogen with other elements.

36. What are the main types of hydrides?

Answer:

👉 Hydrides are classified into three main types:

1. Ionic hydrides
2. Covalent hydrides
3. Intermediate (polymeric) hydrides

37. What are amphoteric oxides?

Answer:

👉 Amphoteric oxides show both acidic and basic properties.

Example: ZnO, Al_2O_3, BeO.

38. How do oxides vary across a period?

Answer:

👉 Across a period, oxides change from strongly basic ⟶ weakly basic ⟶ amphoteric ⟶ acidic ⟶ strongly acidic.

39. Why is hydrogen placed at the top of Group IA?

Answer:

👉 Hydrogen is placed at the top of Group IA because it has one electron in its outermost shell and forms +1 ion like alkali metals.

40. Why is hydrogen considered a unique element?

Answer:

👉 Hydrogen is unique because its properties resemble alkali metals, halogens, and group IVA elements, but it does not fit perfectly in any one group.

🩸Selove Exercise

Q1. Fill in the blanks:

(i) Mendeleev in his periodic table, arranged the elements according to their atomic masses

Answer: masses

____ masses ____

(ii) Vertical columns in modern periodic table are called groups and horizontal rows are called periods

Answer: groups, periods

____ groups, periods ____

(iii) Members of group VIIA are called halogens and alkali metals is the family name of group IA

Answer: halogens, group IA

____ halogens, group IA ____

(iv) Metals form basic oxides and non-metals form acidic oxides.

Answer: basic, acidic

____ basic, acidic ____

(v) Hydrogen can be placed above the groups IA or VIIA of the periodic table.

Answer: IA or VIIA

____ IA or VIIA ____

(vi) Shielding effect is actually the repulsion due to electrons in between the nucleus and the outermost shell.

Answer: repulsion

____ repulsion ____

(vii) Noble gases have the highest values of ionization energy due to their complete outermost shells.

Answer: highest

____ highest ____

(viii) When a second electron is added to a uni-negative ion, the incoming electron is repelled by the already present negative charge.

Answer: repelled

____ repelled ____

(ix) Due to having partly filled d-orbitals, transition metals usually show variable valency.

Answer: transition metals, valency

____ transition metals, valency ____

(x) Melting and boiling points of halogens decrease down the group.

Answer: decrease

____ decrease ____

Q2. Indicate True or False

(i) In Mendeleev's periodic table elements Be, Mg, Zn and Cd are placed in the same group.

👉 False

____ False ____

(ii) The second and third periods contain eighteen elements each.

👉 False

____ False ____

(iii) Alkaline earth metals are present in Group IIA.

👉 True

____ True ____

(iv) Metals are present in the top right corner of the periodic table.

👉 False

____ False ____

(v) Metalloids are present in the lower half of Groups IVA, VA and VIA.

👉 True

____ True ____

(vi) Hydrogen forms uninegative ion like halogens.

👉 True

____ True ____

(vii) Oxidation state of an element is related to the number of period it belongs.

👉 False

____ False ____

(viii) Diamond is a good conductor of electricity.

👉 False

____ False ____

(ix) Melting points of halogens decrease down the group.

👉 True

____ True ____

(x) Zinc oxide is an example of amphoteric oxide.

👉 True

____ True ____

🌟 Q4. What are the improvements made in the Mendeleev's Periodic Table?

❖ **Introduction:**

- Dmitri Mendeleev arranged the known elements in order of increasing atomic masses and gave the first systematic classification of elements known as the Mendeleev's Periodic Table.
- However, with time, scientists discovered new facts about the atomic structure which required improvements in his table.

⚗️ Main Improvements in Mendeleev's Periodic Table:

1. Use of Atomic Number Instead of Atomic Mass

- Mendeleev arranged elements according to their atomic masses.
- **Later**, Henry Moseley (1913) discovered that atomic number is a more fundamental property than atomic mass.

- In the Modern Periodic Table, elements are arranged according to increasing atomic number, which solved the irregularities of Mendeleev's arrangement.

2. Position of Isotopes Explained

- Mendeleev could not explain the position of isotopes (atoms of same element having different atomic masses).
- In the modern table, isotopes are placed in the same position because they have same atomic number.

3. Accurate Arrangement of Elements

- In Mendeleev's table, some elements with higher atomic masses were placed before those with lower masses (e.g., Cobalt and Nickel).
- The modern arrangement by atomic number corrected this irregularity.

4. Discovery of Noble Gases

- Noble gases were unknown during Mendeleev's time.
- **Later**, they were discovered and placed in a new group (Group VIII A) without disturbing the periodic table.

5. Addition of New Elements

- After Mendeleev's table, several new elements were discovered.

- Their properties matched perfectly with the predicted ones by Mendeleev, confirming the validity of the periodic law.
- The modern table has 18 groups and 7 periods, making it more complete.

6. Division into Blocks

- The modern table classifies elements into s-block, p-block, d-block, and f-block based on the valence orbital being filled.
- This division was not present in Mendeleev's table and shows a clear understanding of atomic structure.

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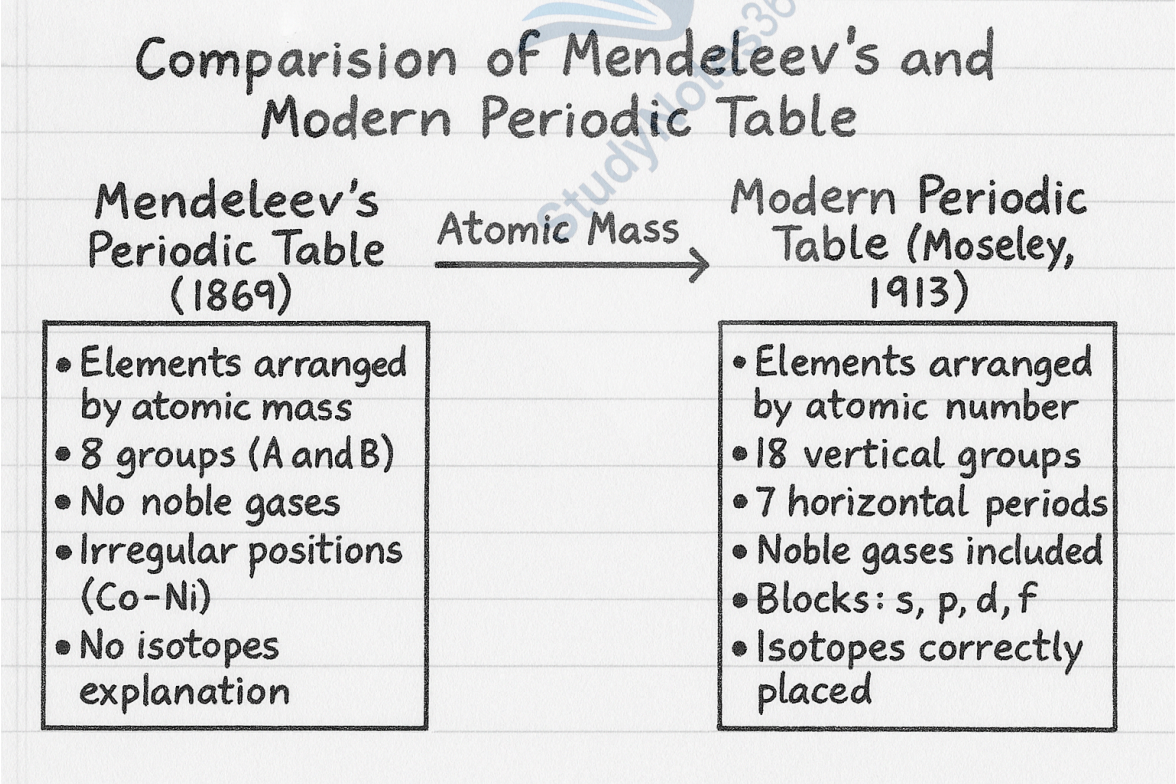

✦ **Summary:**

In short, the Modern Periodic Table is an improved version of Mendeleev's Table, as it:

- Is based on atomic number instead of mass.
- Correctly places isotopes and newly discovered elements.
- Explains periodicity more accurately.
- Divides elements into blocks and families.

Thus, these improvements made the table more logical, scientific, and complete.

🌟 **Q5. How the classification of elements in different blocks helps in understanding their chemistry?**

❖ **Introduction:**

- The Modern Periodic Table classifies elements into four major blocks — s-block, p-block, d-block, and f-block — based on the type of atomic orbital that receives the last electron.
- This classification helps chemists to predict and understand the properties and behavior of elements more easily.

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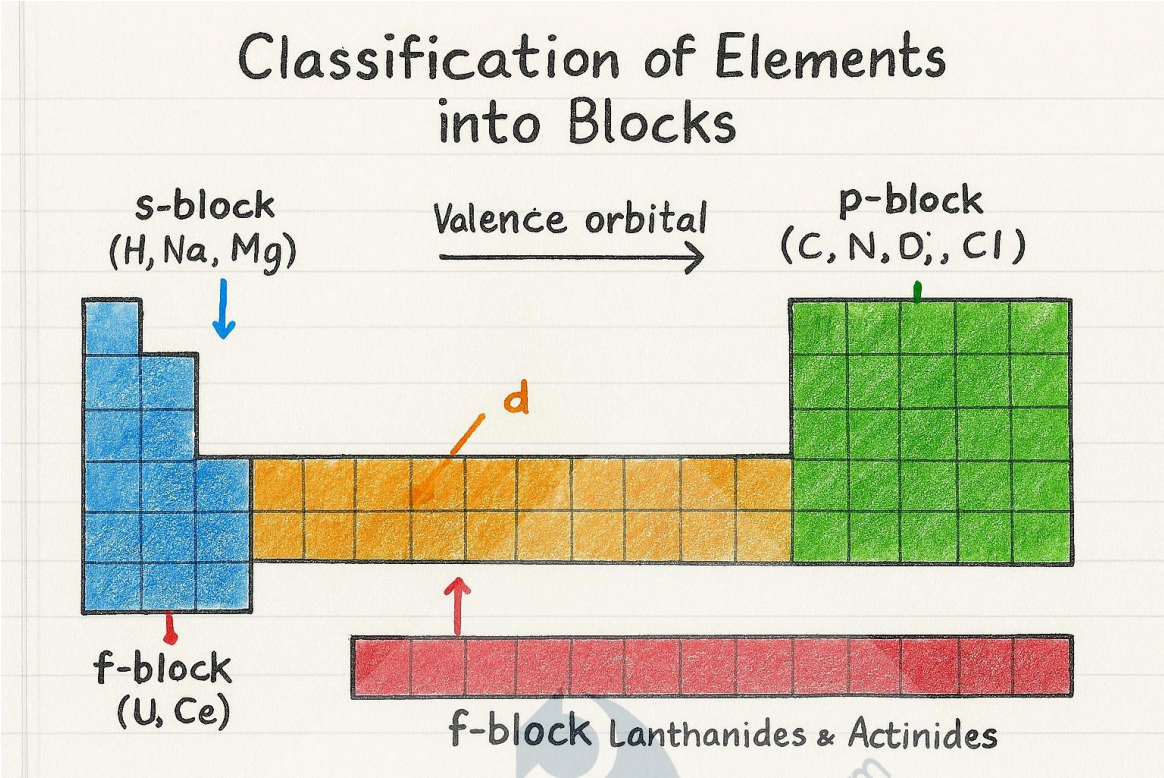

⚗️ Classification of Elements into Blocks:

1. s-block Elements

- These are elements in which the last electron enters an s-orbital.
- They include Group IA (alkali metals) and Group IIA (alkaline earth metals).
- They are highly reactive metals, have low ionization energies, and form basic oxides and ionic compounds.

Examples: Sodium (Na), Magnesium (Mg), Calcium (Ca).

2. p-block Elements

- In these elements, the last electron enters a p-orbital.

- They include Groups IIIA to VIIIA (13–18).
- These elements show a wide range of properties: metals, non-metals, and metalloids.
- They form covalent compounds and include important elements like C, N, O, F, Cl etc.

Examples: Carbon (C), Oxygen (O), Nitrogen (N), Chlorine (Cl).

3. d-block Elements (Transition Elements)

- In these elements, the last electron enters a d-orbital.
- They are found in Groups IIIB to IIB (3–12).
- They exhibit variable oxidation states, form colored compounds, and are good conductors of heat and electricity.

Examples: Iron (Fe), Copper (Cu), Zinc (Zn).

4. f-block Elements (Inner Transition Elements)

- The last electron enters an f-orbital.
- They are placed separately at the bottom of the table and include:
- Lanthanides (atomic numbers 58–71)
- Actinides (atomic numbers 90–103)
- These elements are mostly radioactive and show similar chemical properties.

Examples: Uranium (U), Thorium (Th), Cerium (Ce).

🔬 Importance of Block Classification in Understanding Chemistry:

1. Predicting Chemical Properties

- Elements of the same block show similar properties because they have the same valence electron configuration.
- This helps in predicting their reactivity, bonding type, and compounds formed.

2. Simplifying the Study of Periodic Table

- By dividing elements into four blocks, it becomes easier to study chemical trends like ionization energy, atomic size, and metallic character.

3. Understanding Chemical Bonding

- Block classification helps in understanding how elements bond —
- s-block elements form ionic bonds,
- p-block elements form covalent bonds,
- d-block elements often form coordinate bonds.

4. Identifying Industrial and Biological Uses

- The classification helps in identifying industrial importance (like transition metals in alloys) and biological significance (like Na^+, K^+ ions in cells).

✦ **Summary:**

- In short, the classification of elements into s, p, d, and f blocks provides a systematic way to study chemistry.

It helps in understanding:

- Their electronic configuration
- Physical and chemical properties
- Reactivity trends
- Bond formation and compound types

Hence, this classification is the foundation of modern chemical understanding.

🌟 **Q6. How do you justify the position of hydrogen at the top of various groups?**

❖ **Introduction:**

- Hydrogen is the first and lightest element of the periodic table with atomic number 1 and electronic configuration $1s^1$.
- It has only one electron, which makes it similar to alkali metals, but at the same time, it can also gain one electron like halogens.
- Because of these dual properties, hydrogen has been given a unique position at the top of the periodic table.

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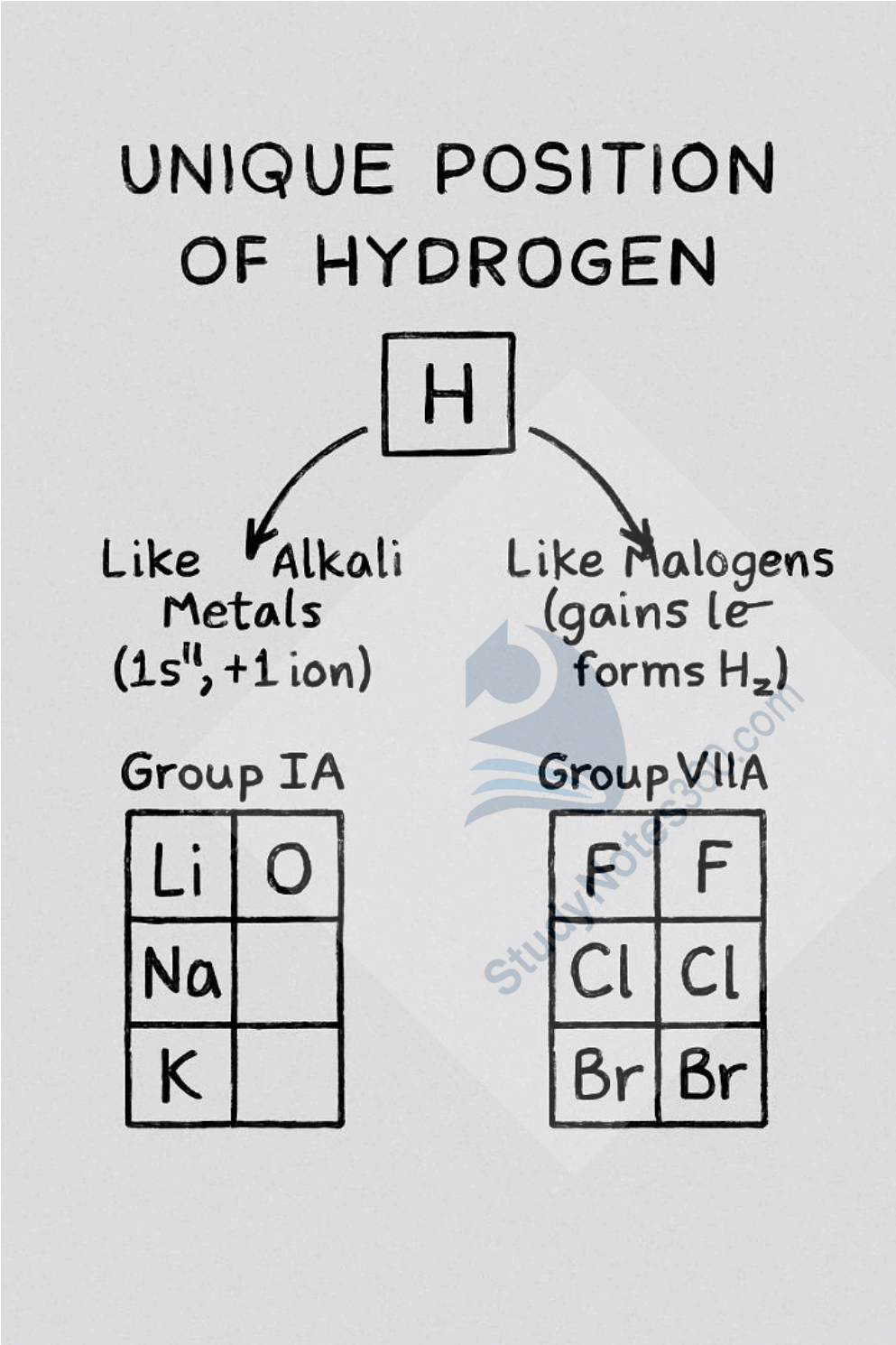

⚛️ 1. Similarity with Alkali Metals (Group IA):

Hydrogen behaves like alkali metals in many ways:

- Both contain one electron in their outermost shell (ns^1).
- Both can lose one electron to form a unipositive ion (H^+, Na^+).
- Both form ionic compounds with non-metals such as HCl, NaCl.
- Both are strong reducing agents.

👉 Therefore, hydrogen can be placed above Group IA elements.

🌫️ 2. Similarity with Halogens (Group VIIA):

Hydrogen also resembles halogens in several ways:

- Both need one electron to complete their outermost shell.
- Both can form uninegative ions (H^-, Cl^-).
- Both exist as diatomic molecules (H_2, Cl_2).
- Both form covalent compounds with non-metals.

👉 Hence, hydrogen can also be placed above Group VIIA elements.

⚖️ 3. Differences from Both Groups:

Despite these similarities, hydrogen also differs from both:

- It is a non-metal, whereas alkali metals are metals.
- It is a gas, not a solid metal.
- It does not conduct electricity or show metallic luster.
- Unlike halogens, it does not form salts in the same manner.

👉 These differences make hydrogen distinct from all other elements.

🔬 4. Unique Position of Hydrogen:

- Because hydrogen shares properties with both alkali metals and halogens — but not completely with either — it is given a separate position at the top of the periodic table.

This position shows its dual nature:

- It can lose one electron (metallic behavior).
- It can gain one electron (non-metallic behavior).

✦ **Summary:**

- Hydrogen is a unique element that cannot fit perfectly in any one group.
- It shows properties similar to alkali metals (Group IA) and halogens (Group VIIA), but also has distinct characteristics of its own.
- Because of this dual chemical behavior, hydrogen is placed separately at the top of the periodic table.
- This special position represents that hydrogen is both similar and different from all other elements — truly a class by itself.

🌟 Q7. Why the ionic radii of negative ions are larger than the size of their parent atoms?

❖ Introduction:

When an atom gains one or more electrons, it becomes a negative ion (anion).

For example:

$Cl + e^- \longrightarrow Cl^-$

This process increases the size of the atom, and the resulting anion becomes larger than its parent atom.

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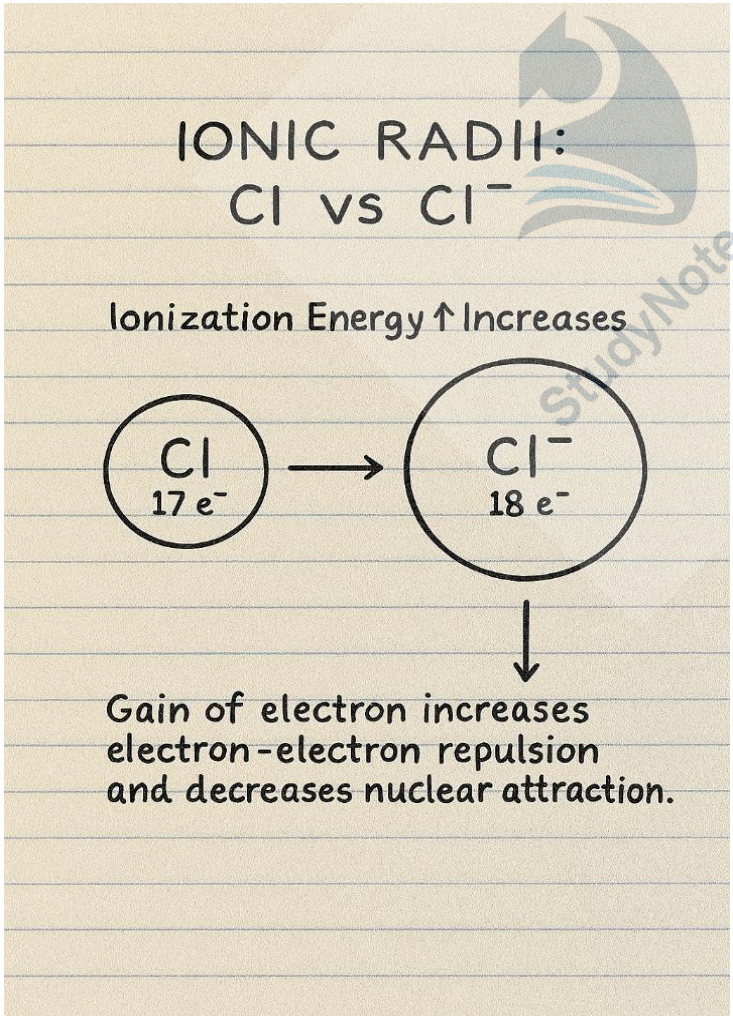

⚛️ 1. Cause of Increase in Size:

When an atom gains electrons:

- The number of electrons increases while the nuclear charge (protons) remains the same.
- The attraction between the nucleus and the electrons becomes weaker.

As a result, the outermost electrons are less tightly held and move farther away from the nucleus.

👉 Hence, the ionic radius increases.

🔹 2. Electron-Electron Repulsion:

- In a negative ion, the extra electron(s) cause greater repulsion among all the electrons.
- These repulsive forces cause the electrons to spread out more.
- The electron cloud expands, resulting in a larger ionic size.

🔹 3. Example:

Let's compare chlorine atom (Cl) and its ion chloride (Cl^-):

- Atomic number of Cl = 17 ⟶ 17 protons, 17 electrons.
- In Cl^- ⟶ 17 protons, 18 electrons.

Now, 17 protons are attracting 18 electrons, so each electron feels less pull toward the nucleus.

👉 Therefore, the radius of Cl^- is greater than that of Cl atom.

⚖️ 4. General Rule:

- Positive ions (cations) ⟶ Smaller than their parent atoms because they lose electrons.
- Negative ions (anions) ⟶ Larger than their parent atoms because they gain electrons.
- This difference is due to the balance between nuclear charge and number of electrons.

✦ **Summary:**

- When an atom gains electrons, it forms a negative ion.
- Because the nuclear pull remains the same but the number of electrons increases, the repulsion among electrons increases, and the outermost shell expands.

👉 Therefore, negative ions (anions) always have larger ionic radii than their parent atoms.

🌟 Q8. Why ionization energy decreases down the group and increases along a period?

❖ **Definition of Ionization Energy:**

Ionization energy is the amount of energy required to remove one electron from a gaseous atom to form a positively charged ion.

$X(g) \longrightarrow X^+(g) + e^-$

🔸 1. Variation Down a Group

⚛️ Observation:

Ionization energy decreases from top to bottom in a group.

🧩 Reasons:

1. Increase in Atomic Size:

- As we go down a group, the number of electron shells increases, so the valence electron moves farther from the nucleus. The attraction between the nucleus and the outer electron becomes weaker, and the electron is removed more easily.

2. Increase in Shielding Effect:

- The inner electrons act as a shield between the nucleus and the outermost electron. This reduces the effective nuclear attraction, making it easier to remove the outer electron.

3. Weaker Nuclear Attraction:

- Although the number of protons increases down the group, the effect of increased distance and shielding is greater, resulting in a decrease in ionization energy.

🔸 2. Variation Across a Period

⚛️ Observation:

Ionization energy increases from left to right across a period.

🧩 Reasons:

1. Decrease in Atomic Size:

- Across a period, the number of protons in the nucleus increases while electrons are added to the same shell. This increased nuclear charge pulls the electrons closer to the nucleus, reducing the atomic size and making it harder to remove an electron.

2. Increase in Effective Nuclear Charge:

- The attraction between the nucleus and outer electrons becomes stronger, which means more energy is needed to remove an electron.

3. Stable Electronic Configuration:

- At the end of each period, noble gases have completely filled shells. Their stability makes it very difficult to remove electrons, so their ionization energies are the highest in each period.

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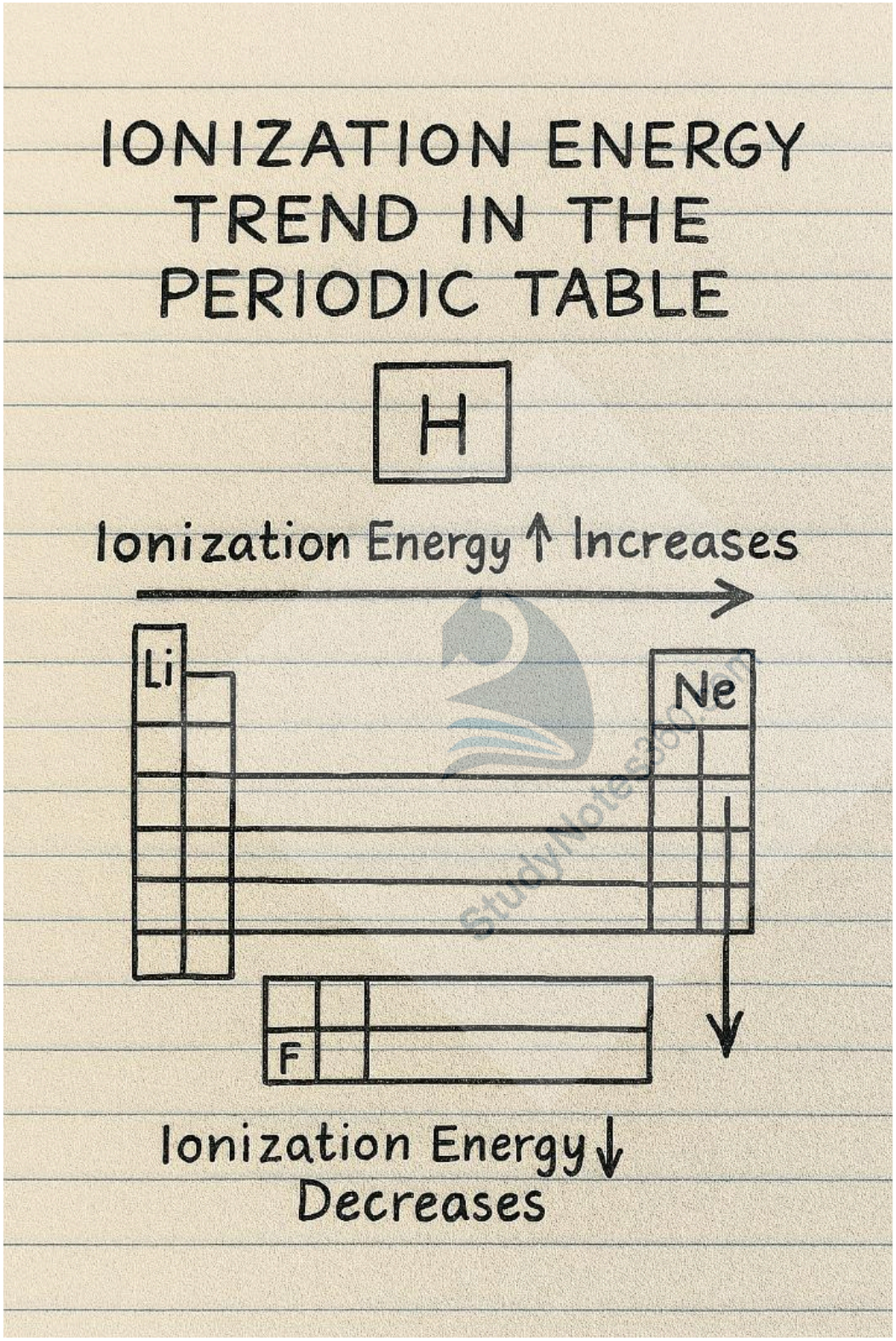
IONIZATION ENERGY TREND IN THE PERIODIC TABLE
H
Ionization Energy ↑ Increases
Li
Ne
F
Ionization Energy ↓ Decreases

✦ **Summary:**

In short, ionization energy decreases down a group because the atomic size increases and the nuclear attraction for the outer electrons becomes weaker due to the shielding effect. On the other hand, it increases across a period because the atomic size decreases, nuclear charge increases, and the electrons are held more tightly by the nucleus. Thus, the trend in ionization energy clearly shows how the structure of atoms affects their chemical behavior.

🌟 Q9. Why the second value of electron affinity of an element is usually shown with a positive sign?

❖ **Definition of Electron Affinity:**

Electron affinity is the amount of energy released or absorbed when an electron is added to a neutral gaseous atom to form a negatively charged ion.

$X(g) + e^- \longrightarrow X^-(g)$

The **energy** change in this process is called the first **electron affinity.**

If a **second electron** is added to the already negative ion, the energy change is called the second electron affinity.

$X^-(g) + e^- \longrightarrow X^{2-}(g)$

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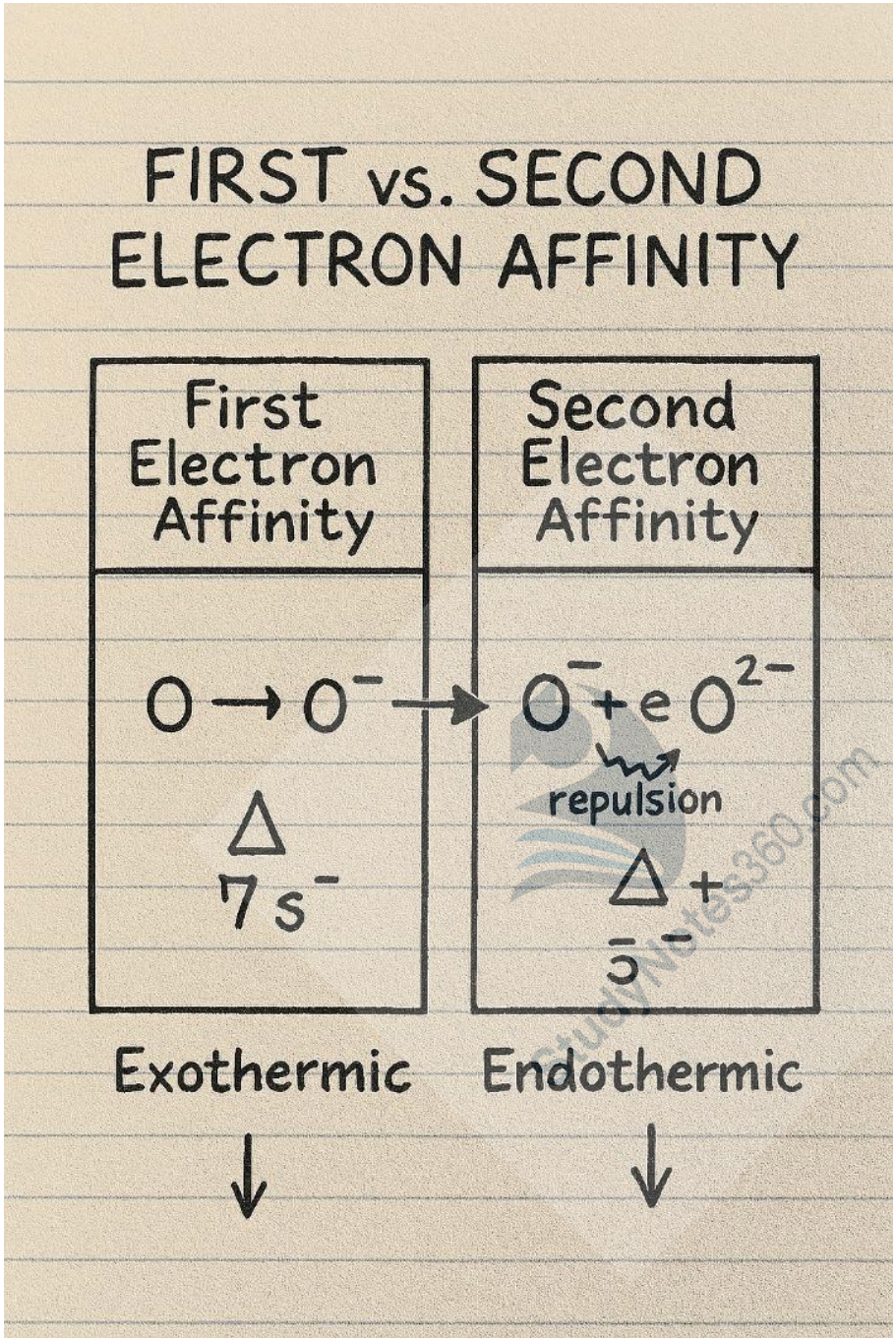

🔸 1. First Electron Affinity (Exothermic Process):

- When a neutral atom gains its first electron, energy is usually released because the electron is attracted towards the positively charged nucleus.

Hence, the first electron affinity has a negative sign (energy is given out).

✅ **Example:**

$Cl(g) + e^- \longrightarrow Cl^-(g) + 349\ kJ/mol$

🔸 **2. Second Electron Affinity (Endothermic Process):**

When a second electron is added to the already negatively charged ion, the process becomes unfavorable because:

- The incoming electron faces repulsion from the already present negative charge.
- Extra energy is required to overcome this repulsion and force the second electron into the ion.

Therefore, the second electron affinity is positive, meaning energy is absorbed.

✅ **Example:**

$O^-(g) + e^- + Energy \longrightarrow O^{2-}(g)$

🔸 **3. Explanation of the Positive Sign:**

- The positive sign indicates that the process is endothermic — energy must be supplied to add the second electron.
- It also shows that forming a doubly negative ion is energetically less favorable than forming a singly negative ion.

✦ **Summary:**

The first electron affinity of an element is negative because energy is released when a neutral atom gains an electron. However, the second electron affinity is positive because the addition of another electron to a negatively charged ion requires extra energy to overcome the strong repulsion between like charges. Thus, energy is absorbed, and the value is shown with a positive sign.

🌟 **Q10. Why metallic character increases from top to bottom in a group of metals?**

❖ **Definition of Metallic Character:**

- Metallic character is the tendency of an atom to lose electrons and form positive ions (cations).
- The more easily an element loses its valence electrons, the more metallic it is.

Metal $\longrightarrow$ Metal^{n+} + n e^-

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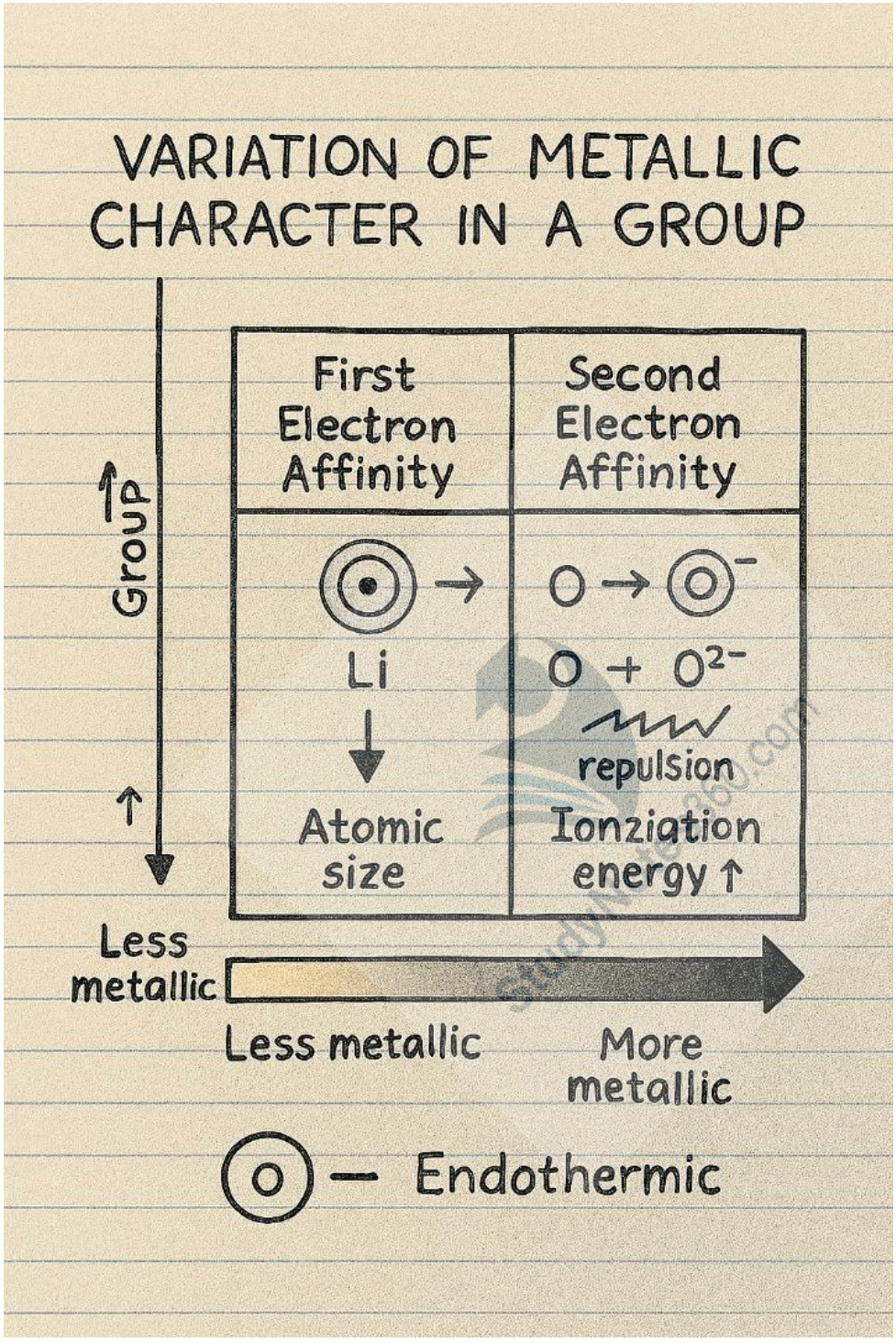

🔹 1. Increase in Atomic Size:

- When we move down a group in the periodic table, the number of electron shells increases.

- Each new shell increases the distance between the nucleus and the outermost electrons.
- Due to this increase in distance, the attraction between the nucleus and valence electrons decreases.

➡️ **Result**: It becomes easier for atoms to lose electrons, thus metallic character increases.

🔹 **2. Increase in Shielding Effect:**

- As we go down the group, the inner electrons shield the outermost electrons from the pull of the nucleus.
- This shielding effect weakens the effective nuclear attraction on valence electrons.

➡️ **Therefore**, the outer electrons can be removed more easily, enhancing metallic behavior.

🔹 **3. Decrease in Ionization Energy:**

- Ionization energy means the amount of energy required to remove an electron from an atom.

Down the group:

- Atomic size increases
- Shielding effect increases
- Attraction between nucleus and valence electrons decreases

➡️ **Hence**, ionization energy decreases, and metals lose electrons more easily.

Thus, the metallic character increases down the group.

🔹 **4. Example:**

Let's take Group I (Alkali metals):

Li < Na < K < Rb < Cs

✦ **Summary:**

As we move from top to bottom in a group, the atomic size increases, shielding effect becomes stronger, and ionization energy decreases. These factors make it easier for atoms to lose electrons, resulting in an increase in metallic character down the group.

🌟 **Q11. Explain the variation in melting points along the short periods.**

❖ **Definition:**

- The melting point of an element is the temperature at which it changes from a solid to a liquid.
- It depends on the type of bonding and strength of attractive forces between atoms or molecules in the element.

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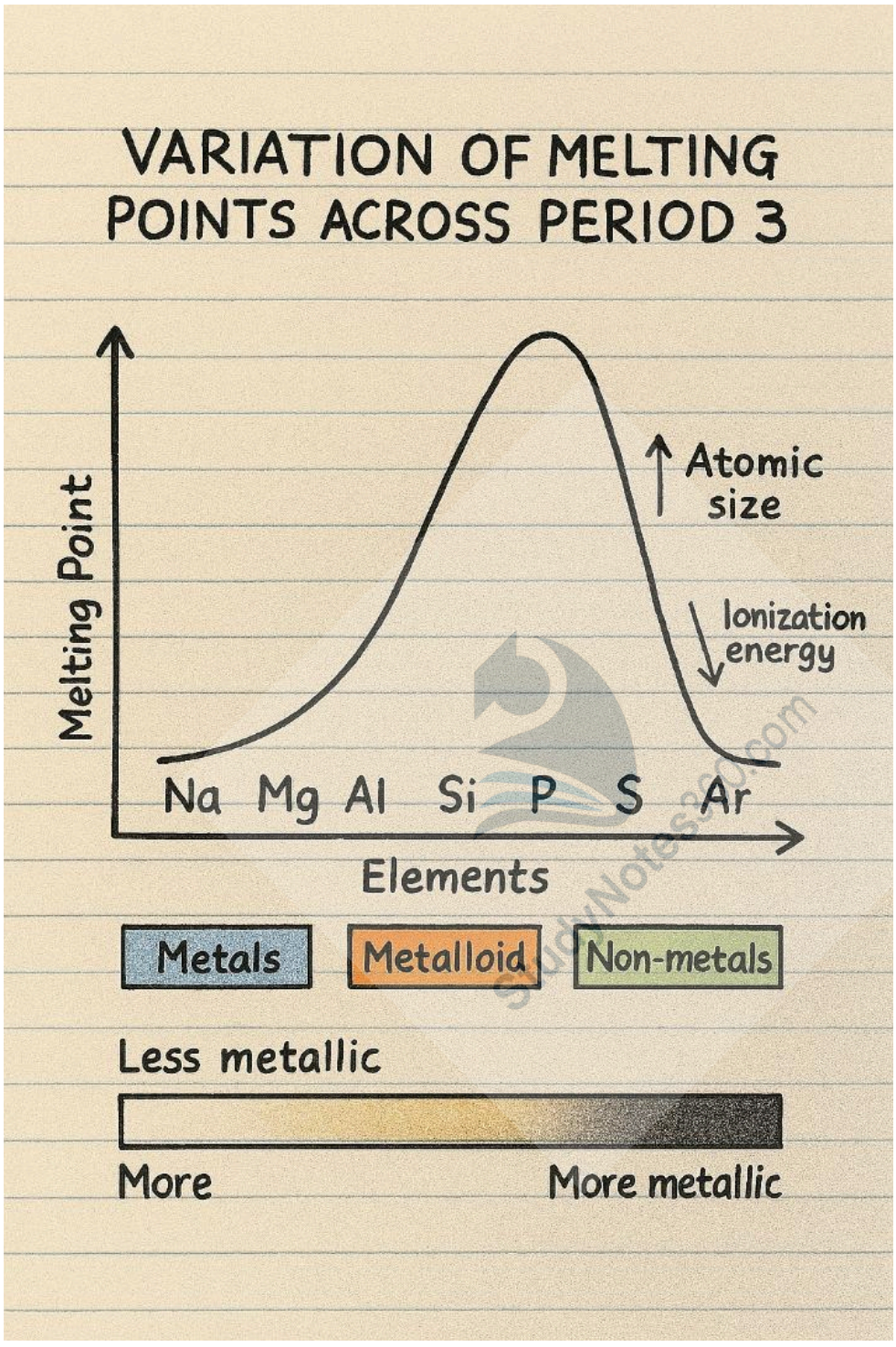

🔹 1. General Trend Along a Period:

When we move from left to right across a short period in the periodic table:

- Metallic elements are present on the left,
- Non-metals are present on the right,
- Metalloids and transition points occur in the middle.

➡️ Therefore, melting points first increase, reach a maximum, and then decrease toward the end of the period.

🔹 2. Explanation of Trend:

(a) Metals (Left side of the period):

- Elements like Na, Mg, and Al have metallic bonds.

The number of valence electrons and bond strength increase from Na ⟶ Mg ⟶ Al.

➡️ **Hence**, their melting points increase.

(b) Metalloids (Middle of the period):

- Silicon (Si) forms a giant covalent structure similar to diamond.
- Each silicon atom is strongly bonded to four others, creating a rigid 3D network.

➡️ Melting point is very high.

(c) Non-metals (Right side of the period):

- Elements like P, S, Cl, Ar exist as simple molecules held together by weak van der Waals forces.
- These weak forces require less energy to break,

➡️ **Hence**, melting points decrease sharply at the end of the period.

🔹 3. Example from Period 3:

Elements (Left to Right):

Na ⟶ Mg ⟶ Al ⟶ Si ⟶ P ⟶ S ⟶ Cl ⟶ Ar

Type of Elements:

- Na, Mg, Al ⟶ Metals
- Si ⟶ Metalloid
- P, S, Cl ⟶ Non-metals
- Ar ⟶ Noble gas

Melting Point Trend:

Increases from Na ⟶ Mg ⟶ Al ⟶ Si (Highest)

Then decreases from Si ⟶ P ⟶ S ⟶ Cl ⟶ Ar (Lowest)

✦ **Summary:**

- In short periods, melting points increase from Group 1 to Group 14, then decrease from Group 15 to 18.
- The reason is the change in bonding nature — from strong metallic and covalent bonds to weak molecular forces.

🧩 In short:

➡️ Metallic ⟶ Strong covalent ⟶ Weak molecular

➡️ Melting points: Low ⟶ High ⟶ Low

🌟 Q12. Why the oxidation state of noble gases is usually zero?

❖ **Answer:**

- Noble gases are present in Group VIII A (Group 18) of the periodic table.
- They are also known as inert gases because they do not readily react with other elements under ordinary conditions.
- This chemical inertness is the main reason why their oxidation state is usually zero.

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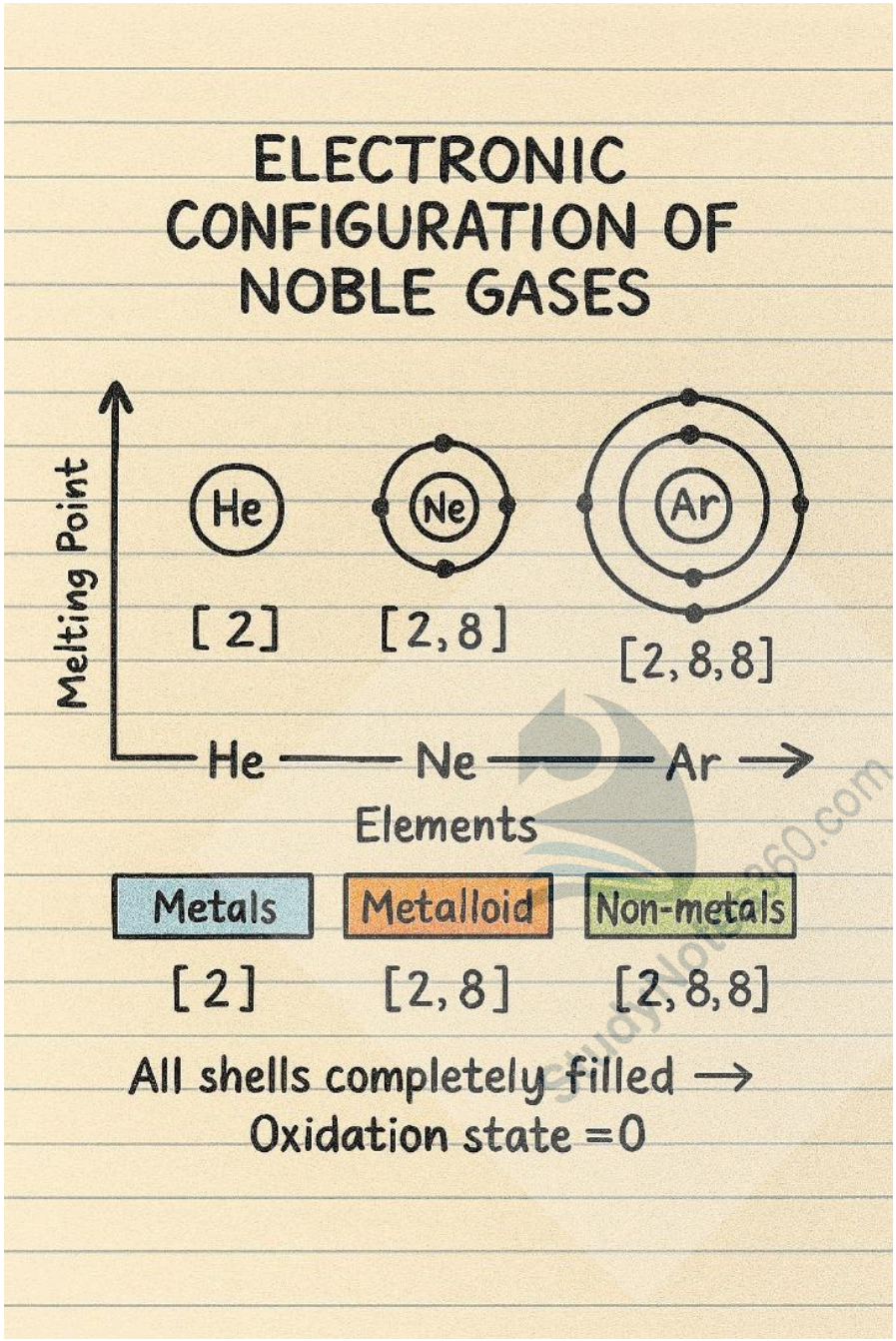

✦ Explanation:

1. Complete Outer Shell:

- Each noble gas atom has a completely filled outermost shell (e.g., He = 2, Ne = 8, Ar = 8 electrons).

Therefore, they do not need to gain, lose, or share electrons to achieve stability.

2. High Ionization Energy:

- Noble gases have very high ionization energies, meaning it requires a large amount of energy to remove an electron.

Hence, they cannot easily form positive ions.

3. Zero Electron Affinity:

- They also have zero or almost zero electron affinity, meaning they do not attract extra electrons — so negative ion formation is also not possible.

4. Stable Electronic Configuration:

- Their octet or duplet configuration is already stable, so they have no tendency to undergo chemical reactions or form compounds.

5. Exception (Xenon and Krypton):

- Under extreme conditions, heavier noble gases like Xenon (Xe) and Krypton (Kr) can form compounds (e.g., XeF_2, XeF_4), but generally, their oxidation state remains zero.

✦ **Summary:**

👉 Noble gases usually show zero oxidation state because they have a completely filled outermost shell, high ionization energy, and no tendency to gain or lose electrons.

Thus, they remain chemically inert under normal conditions.

🌟 Q13. Why diamond is a non-conductor and graphite is fairly a good conductor?

❖ **Answer:**

Diamond and graphite are two different forms (allotropes) of the same element, carbon. Although both are made of carbon atoms only, their structures and bonding patterns are completely different. This difference in structure changes their physical and electrical properties — especially conductivity.

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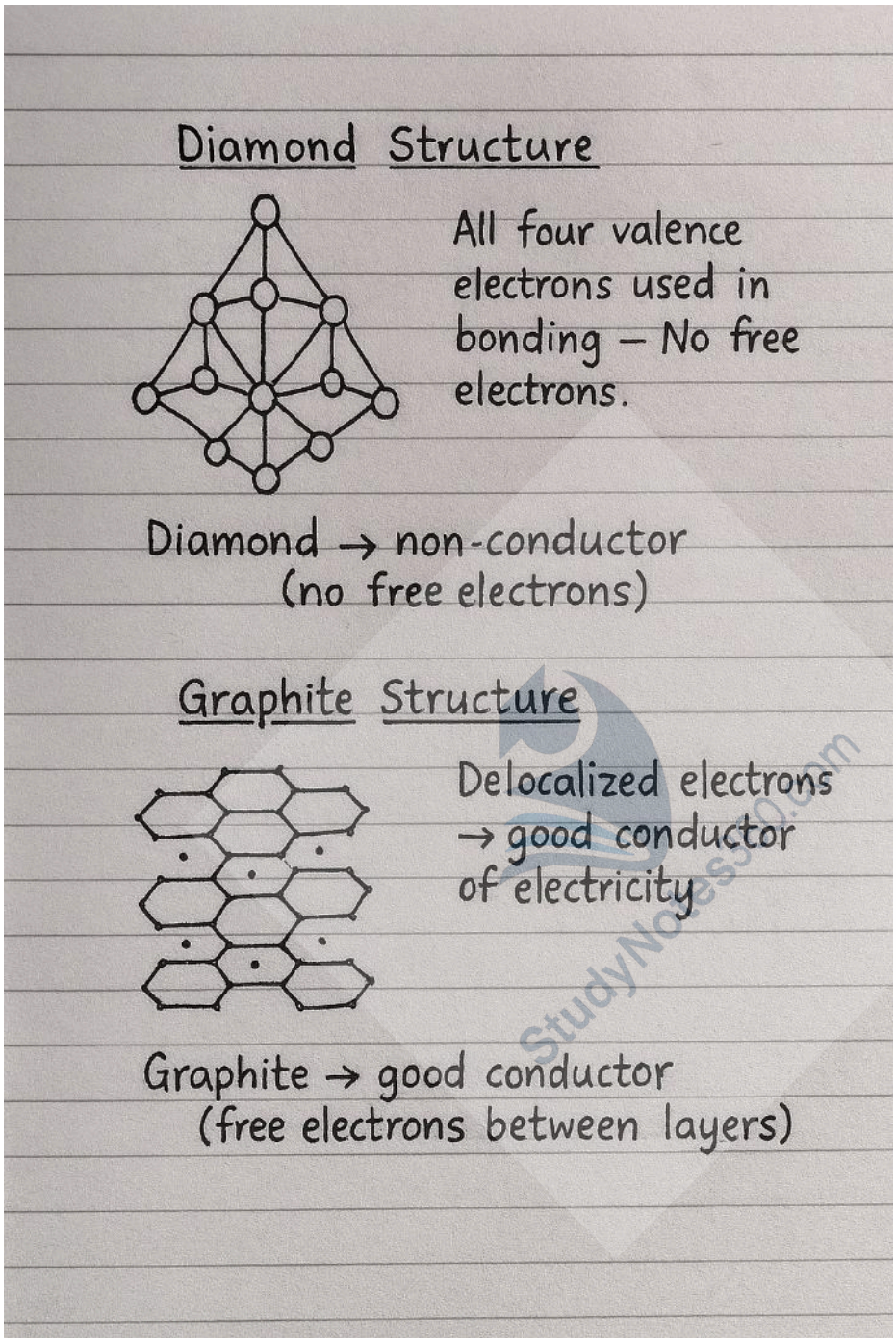


✦ Structure of Diamond:

- In diamond, each carbon atom is sp^3 hybridized. Every carbon atom is bonded covalently with four other carbon atoms in a tetrahedral arrangement. This creates a

three-dimensional crystal lattice that is extremely hard and rigid.

- Because all four valence electrons of each carbon atom are used in bonding, there are no free electrons available to move or carry electric current. As a result, diamond cannot conduct electricity and acts as a non-conductor (insulator).
- This strong network of covalent bonds also gives diamond its exceptionally high melting point and hardness.

⚡ Structure of Graphite:

- In graphite, each carbon atom is sp^2 hybridized and bonded to three other carbon atoms forming flat hexagonal layers (like a honeycomb pattern). The fourth electron of each carbon atom remains free and delocalized.
- These free electrons are not tied to any specific atom — instead, they move freely between the layers. This movement of delocalized electrons allows graphite to conduct electricity efficiently.
- The layers of graphite are held together by weak van der Waals forces, which makes them slide easily over each other — this is why graphite feels soft and slippery.

💡 Main Concept:

- The ability of a substance to conduct electricity depends on the presence of free or mobile electrons.
- **In diamond,** there are no free electrons, so it is a non-conductor.
- **In graphite,** there are free electrons that move between layers, so it is a good conductor of electricity.

✦ **Summary:**

👉 Diamond does not conduct electricity because all its valence electrons are involved in covalent bonds, leaving no free electrons.

👉 Graphite, on the other hand, conducts electricity because one electron per carbon atom remains free and can move between layers.

👉 Therefore, the difference in their electrical behaviour is entirely due to the difference in bonding and atomic arrangement of carbon atoms in diamond and graphite.

🌟 **Q14. Give brief reason for the following:**

🅐 **d- and f-block elements are called transition elements.**

Answer:

- The d- and f-block elements are called transition elements because they represent a transition (or change) between the highly reactive s-block metals and the less reactive p-block elements.

➤ In these elements, d or f orbitals are being filled with electrons, leading to gradual change in properties such as:

- Variable oxidation states
- Formation of colored compounds
- Magnetic behavior and catalytic activity

👉 **Hence**, they are known as transition elements because their properties "transition" between metallic and non-metallic elements.

🅑 Lanthanide contraction controls the atomic sizes of elements of 6th and 7th periods.

Answer:

- As electrons are added to the 4f-subshell in lanthanides, these 4f-electrons poorly shield the outer electrons from the increasing nuclear charge.
- As a result, the effective nuclear attraction increases, and the atomic size gradually decreases across the lanthanide series.
- This phenomenon is called lanthanide contraction, and it affects the atomic sizes of elements in the 6th and 7th periods, making them smaller than expected and chemically similar (for example, Zr and Hf have almost same size).

Ⓒ The melting and boiling points of the elements increase from left to right up to the middle of s- and p-block elements and decrease onward.

Answer:

- Across a period, metallic bonding becomes stronger from left to right because atoms have more valence electrons and smaller sizes.
- This stronger bonding increases melting and boiling points until the middle of the period (e.g., silicon has the highest).
- After the middle, elements become non-metals, which exist as molecules held by weak van der Waals forces, causing a sharp decrease in melting and boiling points.

Ⓓ The oxidation states vary in a period but remain almost constant in a group.

Answer:

- Across a period, the number of valence electrons increases, allowing elements to show different oxidation states.
- **However**, in a group, all elements have the same valence electron configuration, so their oxidation states remain nearly constant (e.g., all Group II elements show +2).

Ⓔ The hydration energies of the ions are in the following order: $Al^{3+} > Mg^{2+} > Na^{+}$

Answer:

- Hydration energy depends on the charge and size of the ion.
- Smaller and highly charged ions attract water molecules more strongly, releasing greater hydration energy.
- Since Al^{3+} has the highest charge (+3) and smallest radius, it has the highest hydration energy, followed by Mg^{2+}, and then Na^{+} which has the lowest.

👉 **Therefore**, the order of hydration energy is:

$Al^{3+} > Mg^{2+} > Na^{+}$

Ⓕ Ionic character of halides decreases from left to right in a period.

Answer:

- As we move from left to right in a period, the electronegativity of the elements increases and their ability to form covalent bonds also increases.
- **Therefore**, the difference in electronegativity between the metal and halogen decreases, resulting in less ionic and more covalent character in the halides.

👉 **Thus**, ionic character decreases from left to right across a period.

Ⓖ Alkali metals give ionic hydrides.

Answer:

- Alkali metals (like Na, K, Li) are highly electropositive, meaning they can easily lose one electron to form a positive ion (M^+).
- **Hydrogen**, being slightly electronegative, gains this electron to form a hydride ion (H^-).
- The compound formed (like NaH, KH) consists of ions (M^+ and H^-) held by strong ionic bonds, making them ionic hydrides.

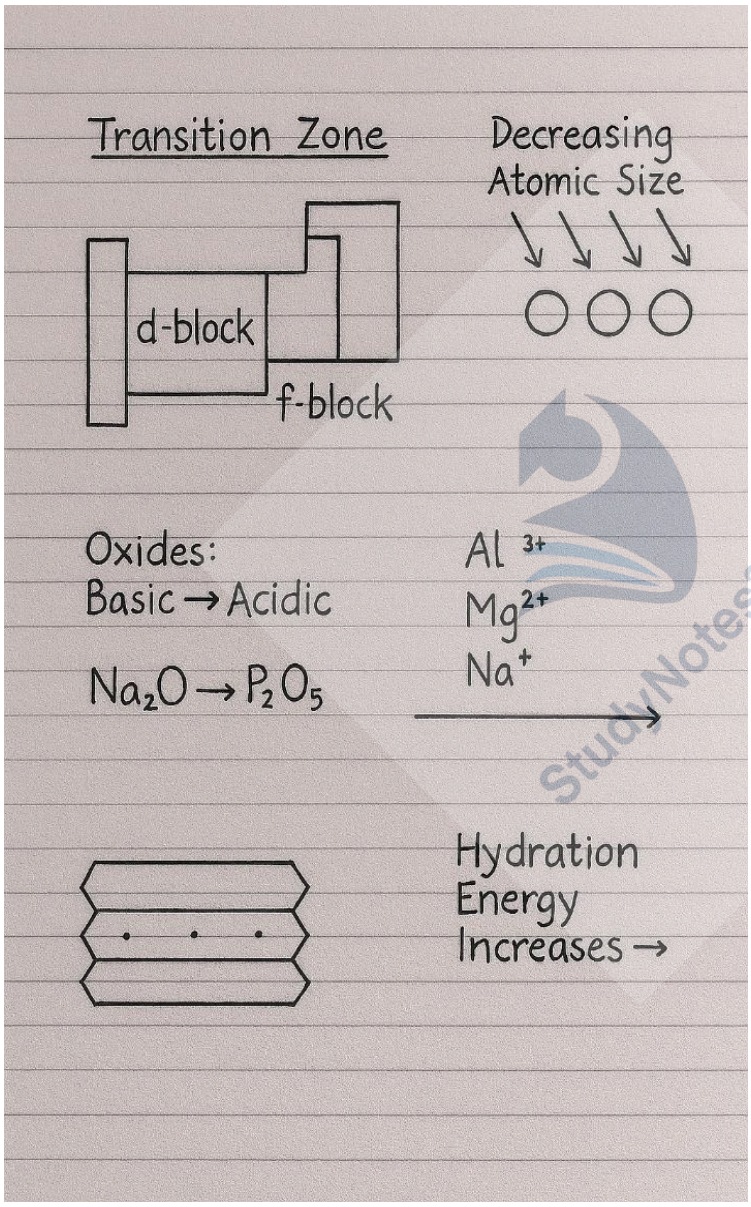
Ⓗ Although both sodium and phosphorus are present in the same period of the periodic table yet their oxides are different in nature — Na_2O is basic while P_2O_5 is acidic in character.

Answer:

- Across a period, metallic character decreases while non-metallic character increases.
- **Sodium**, being a metal, forms basic oxide (Na_2O) that reacts with water to produce a base (NaOH).
- **Phosphorus**, a non-metal, forms acidic oxide (P_2O_5) that reacts with water to produce an acid (H_3PO_4).

👉 **Thus**, as we move from left (Na) to right (P) across a period, the nature of oxides changes from basic ⟶ amphoteric ⟶ acidic.

✦ **Digram:**



✦ **Summary:**

(a) d & f-block ⟶ Transition elements due to gradual property change.

(b) Lanthanide contraction ⟶ Reduces atomic size in 6th and 7th periods.

(c) Melting/boiling points ⬆️ then ↓ across a period.

(d) Oxidation states vary in period, constant in group.

(e) Hydration energy: $Al^{3+} > Mg^{2+} > Na^{+}$ (due to charge & size).

(f) Ionic character ↓ across period due to increasing electronegativity.

(g) Alkali metals form ionic hydrides ($M^{+}H^{-}$).

(h) Oxides change from basic (Na_2O) to acidic (P_2O_5) across a period.

Note:

This chapter is designed to provide a solid foundation of knowledge, with the goal of deepening understanding and encouraging further exploration of the subject. The content has been carefully selected to support effective learning and inspire students to engage with the topic more deeply.

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Purpose: To contribute to education by offering insightful, valuable content that enhances learning and understanding.