Grade 9 Chemistry

Periodicity of Properties

A. Shielding Effect

Q1. Define shielding effect. (2 marks)

Answer: Shielding effect is the reduction in the attractive force between the nucleus and outer electrons caused by inner shell electrons.

- Inner electrons act as a shield between nucleus and outer electrons.
- It decreases the effective nuclear charge felt by outer electrons.

Q2. Explain how shielding effect changes down a group in the periodic table. (3 marks)

Answer:

- As we move down a group, number of electron shells increases.
- More inner electrons are present to shield the outer electrons.
- Therefore, shielding effect increases down the group.

Q3. Why does shielding effect remain almost constant across a period? Explain. (3 marks)

Answer:

- Across a period, electrons are added to the same outermost shell.
- The number of inner electron shells remains constant.
- Since shielding depends on inner shells only, shielding effect remains almost constant across a period.

Q4. Between sodium (Na) and magnesium (Mg), which element experiences greater shielding effect? Give reason. (3 marks)

Answer:

- Both Na and Mg experience same shielding effect.
- Both are present in Period 3 with same number of inner shells (2 shells).
- Shielding effect depends on inner shells which are equal in both.

Q5. Arrange the elements of Group 2 (Be, Mg, Ca, Sr, Ba) in order of increasing shielding effect. (2 marks)

Answer: Be < Mg < Ca < Sr < Ba

B. Atomic Size/Atomic Radius

Q1. Define atomic radius. (2 marks)

Answer: Atomic radius is the distance from the center of nucleus to the outermost shell of an atom. OR It is half the distance between nuclei of two adjacent atoms of same element.

Q2. Explain the trend of atomic radius in a period and in a group. (4 marks)

Answer: In a group:

- Atomic radius increases down the group.
- New electron shells are added.

In a period:

- Atomic radius decreases across the period.
- Nuclear charge increases while electrons enter same shell.

Q3. Atomic size increases down the group. Explain why. (3 marks)

Answer:

- New electron shell is added at each step down the group.
- Distance between nucleus and outermost electrons increases.
- Shielding effect also increases, reducing nuclear attraction.

Q4. Atomic size decreases across the period. Explain why. (3 marks)

Answer:

- Nuclear charge (protons) increases across the period.
- Electrons are added to the same shell.
- Increased nuclear charge pulls electrons closer to nucleus, decreasing atomic size.

Q5. Among sodium (Na) and potassium (K), which has the greater atomic size and why? (3 marks)

Answer:

- Potassium (K) has greater atomic size.
- K is below Na in Group 1.
- K has more electron shells (4 shells) than Na (3 shells).

Q6. Among chlorine (Cl) and fluorine (F), which has larger atomic radius? Give reason. (3 marks)

Answer:

- Chlorine (Cl) has larger atomic radius.
- Both are in Group 17, but Cl is below F.
- Cl has 3 electron shells while F has only 2 shells.

Q7. Arrange the elements of Group 1 (Li, Na, K, Rb, Cs) in order of increasing atomic size. (2 marks)

Answer: Li < Na < K < Rb < Cs

Q8. Arrange the elements of Period 3 (Na, Mg, Al, Si, P, S, Cl, Ar) in order of decreasing atomic radius. (2 marks)

Answer: Na > Mg > Al > Si > P > S > Cl > Ar

Q9. Between aluminum (Al) and sulfur (S), which has smaller atomic size? Explain. (3 marks)

Answer:

- Sulfur (S) has smaller atomic size.
- Both are in Period 3, but S is on right side with more protons.
- Higher nuclear charge in S pulls electrons closer, reducing atomic size.

Q10. Why does atomic radius decrease from left to right across a period despite the addition of electrons? (3 marks)

Answer:

- Electrons are added to the same shell, not new shells.
- Nuclear charge (protons) increases across the period.
- Stronger nuclear attraction pulls electron cloud closer, decreasing atomic radius.

C. Ionization Energy

Q1. Define ionization energy. (2 marks)

Answer: Ionization energy is the minimum energy required to remove an electron from the outermost shell of an isolated gaseous atom.

- Measured in kJ/mol or eV.

Q2. Explain the trend of ionization energy in a period and in a group. (4 marks)

Answer: In a group:

- Ionization energy decreases down the group.
- Atomic size increases, making electrons easier to remove.

In a period:

- Ionization energy increases across the period.
- Nuclear charge increases and atomic size decreases, holding electrons tightly.

Q3. Ionization energy decreases down the group. Explain why. (3 marks)

Answer:

- Atomic size increases down the group.
- Outermost electrons are farther from nucleus.
- Shielding effect increases, weakening nuclear attraction on outer electrons.

Q4. Ionization energy increases across the period. Explain why. (3 marks)

Answer:

- Nuclear charge increases across the period.
- Atomic size decreases, bringing electrons closer to nucleus.
- Stronger nuclear attraction requires more energy to remove electrons.

Q5. Among sodium (Na) and potassium (K), which has the higher ionization energy and why? (3 marks)

Answer:

- Sodium (Na) has higher ionization energy.
- K has larger atomic size than Na.
- Outer electron in K is farther from nucleus and easier to remove.

Q6. Among sodium (Na) and chlorine (Cl), which element has higher ionization energy? Give reason. (3 marks)

Answer:

- Chlorine (Cl) has higher ionization energy.

- Both are in Period 3, but Cl has more protons and smaller size.
- Stronger nuclear attraction in Cl makes electron removal difficult.

Q7. Arrange the alkali metals (Li, Na, K, Rb) in order of decreasing ionization energy. (2 marks)

Answer: Li > Na > K > Rb

Q8. Arrange the halogens (F, Cl, Br, I) in order of increasing ionization energy. (2 marks)

Answer: I < Br < Cl < F

Q9. Why do noble gases have the highest ionization energies in their respective periods? (3 marks)

Answer:

- Noble gases have complete and stable outer shell configuration.
- They have smallest atomic size in their period.
- Removing electron from stable configuration requires maximum energy.

Q10. Explain why it is easier to remove an electron from potassium than from sodium. (3 marks)

Answer:

- Potassium has larger atomic size with more electron shells.
- Outer electron is farther from nucleus in K.
- Greater shielding effect in K weakens nuclear attraction, making electron removal easier.

D. Electron Affinity

Q1. Define electron affinity. (2 marks)

Answer: Electron affinity is the energy released when an electron is added to the outermost shell of an isolated gaseous atom.

- It measures the tendency of atom to accept electrons.

Q2. Explain the trend of electron affinity across a period and down a group. (4 marks)

Answer: Across a period:

- Electron affinity increases from left to right.

- Atomic size decreases and nuclear charge increases.

Down a group:

- Electron affinity decreases down the group.
- Atomic size increases, weakening attraction for incoming electrons.

Q3. Electron affinity generally increases across a period. Explain why. (3 marks)

Answer:

- Nuclear charge increases across the period.
- Atomic size decreases.
- Incoming electron is attracted more strongly due to higher nuclear charge and smaller distance.

Q4. Electron affinity decreases down the group. Explain why. (3 marks)

Answer:

- Atomic size increases down the group.
- Incoming electron is farther from nucleus.
- Shielding effect increases, reducing nuclear attraction on added electron.

Q5. Between fluorine (F) and chlorine (Cl), which has higher electron affinity and why? (3 marks)

Answer:

- Chlorine (Cl) has higher electron affinity.
- F is very small, creating strong electron-electron repulsion.
- Cl has suitable size to accept electron more easily.

Q6. Among oxygen (O) and sulfur (S), which element has greater electron affinity? Give reason. (3 marks)

Answer:

- Oxygen (O) has greater electron affinity.
- O is above S in Group 16 with smaller size.
- Higher nuclear charge and smaller distance result in stronger attraction.

Q7. Arrange the halogens (F, Cl, Br, I) in order of decreasing electron affinity. (2 marks)

Answer: Cl > F > Br > I

Q8. Why do halogens have high electron affinity values? (3 marks)

Answer:

- Halogens have 7 electrons in outermost shell.
- They need only one electron to complete their octet.
- Small size and high nuclear charge strongly attract incoming electrons.

Q9. Why do noble gases have very low or zero electron affinity? Explain. (3 marks)

Answer:

- Noble gases have complete and stable outer shell (8 electrons).
- They do not need additional electrons.
- Added electron would go to higher energy shell, which is unfavorable.

E. Electronegativity

Q1. Define electronegativity. (2 marks)

Answer: Electronegativity is the ability of an atom to attract shared pair of electrons towards itself in a chemical bond.

- Measured on Pauling scale.

Q2. Explain the trend of electronegativity in a period and in a group. (4 marks)

Answer: In a period:

- Electronegativity increases from left to right.
- Nuclear charge increases and atomic size decreases.

In a group:

- Electronegativity decreases from top to bottom.
- Atomic size increases, weakening attraction on bonding electrons.

Q3. Electronegativity increases across a period. Explain why. (3 marks)

Answer:

- Nuclear charge increases across the period.

- Atomic size decreases.
- Atoms can attract shared electrons more strongly due to stronger nuclear pull.

Q4. Electronegativity decreases down the group. Explain why. (3 marks)

Answer:

- Atomic size increases down the group.
- Bonding electrons are farther from nucleus.
- Shielding effect increases, reducing ability to attract shared electrons.

Q5. Among sodium (Na) and chlorine (Cl), which element is more electronegative and why? (3 marks)

Answer:

- Chlorine (Cl) is more electronegative.
- Cl has higher nuclear charge and smaller size than Na.
- Cl attracts electrons much more strongly than Na.

Q6. Between fluorine (F) and iodine (I), which has higher electronegativity? Give reason. (3 marks)

Answer:

- Fluorine (F) has higher electronegativity.
- F is at top of Group 17 with smallest atomic size.
- Bonding electrons are very close to F's nucleus, experiencing maximum attraction.

Q7. Arrange the elements nitrogen (N), oxygen (O), and fluorine (F) in order of increasing electronegativity. (2 marks)

Answer: N < O < F

Q8. Arrange the alkali metals (Li, Na, K) in order of decreasing electronegativity. (2 marks)

Answer: Li > Na > K

Q9. Why is fluorine the most electronegative element? Explain. (3 marks)

Answer:

- Fluorine has smallest atomic size among all elements (2 shells only).
- It has high nuclear charge relative to its size.
- Bonding electrons experience maximum nuclear attraction in F.

Q10. Among carbon (C) and silicon (Si), which is more electronegative? Give reason. (3 marks)

Answer:

- Carbon (C) is more electronegative than silicon (Si).
- C is above Si in Group 14 with smaller atomic size.
- Bonding electrons are closer to C's nucleus.

F. Metallic Character

Q1. Define metallic character. (2 marks)

Answer: Metallic character is the tendency of an element to lose electrons and form positive ions.

- It indicates how easily an atom behaves like a metal.

Q2. Explain the trend of metallic character across a period and down a group. (4 marks)

Answer: Down a group:

- Metallic character increases down the group.
- Ionization energy decreases, making electron loss easier.

Across a period:

- Metallic character decreases from left to right.
- Ionization energy increases, making electron loss difficult.

Q3. Metallic character increases down the group. Explain why. (3 marks)

Answer:

- Atomic size increases down the group.
- Outer electrons are farther from nucleus.
- Shielding effect increases, making electrons easier to lose.

Q4. Metallic character decreases across a period. Explain why. (3 marks)

Answer:

- Nuclear charge increases across the period.
- Atomic size decreases.
- Electrons are held more tightly, making them difficult to lose.

Q5. Among sodium (Na) and magnesium (Mg), which element has greater metallic character and why? (3 marks)

Answer:

- Sodium (Na) has greater metallic character.
- Na has lower ionization energy than Mg.
- Na can lose its outer electron more easily.

Q6. Between lithium (Li) and cesium (Cs), which is more metallic? Give reason. (3 marks)

Answer:

- Cesium (Cs) is more metallic.
- Cs is at bottom of Group 1 with larger atomic size.
- Its outer electron is very loosely held and easily lost.

Q7. Arrange the elements of Period 3 (Na, Mg, Al) in order of decreasing metallic character. (2 marks)

Answer: Na > Mg > Al

Q8. Arrange the elements of Group 1 (Li, Na, K, Rb) in order of increasing metallic character. (2 marks)

Answer: Li < Na < K < Rb

Q9. Why are elements on the left side of the periodic table more metallic in nature? (3 marks)

Answer:

- Left side elements have fewer outer electrons.
- They have low ionization energy and large atomic size.
- They can easily lose electrons to form positive ions.

G. Reactivity

Q1. Explain the trend of reactivity of metals down a group. (3 marks)

Answer:

- Reactivity of metals increases down the group.
- Atomic size increases and ionization energy decreases.
- Outer electrons become easier to lose, increasing reactivity.

Q2. Explain the trend of reactivity of non-metals down a group. (3 marks)

Answer:

- Reactivity of non-metals decreases down the group.
- Atomic size increases, making electron gain difficult.
- Nuclear attraction on incoming electrons becomes weaker.

Q3. Why does the reactivity of alkali metals increase down the group? (3 marks)

Answer:

- Alkali metals react by losing one outer electron.
- Atomic size increases and shielding increases down the group.
- Outer electron becomes easier to remove, increasing reactivity.

Q4. Why does the reactivity of halogens decrease down the group? (3 marks)

Answer:

- Halogens react by gaining one electron.
- Atomic size increases down the group.
- Nuclear attraction on incoming electron weakens, decreasing reactivity.

Q5. Among sodium (Na) and potassium (K), which metal is more reactive and why? (3 marks)

Answer:

- Potassium (K) is more reactive.
- K has larger size and lower ionization energy than Na.
- Its outer electron is lost more easily during reactions.

Q6. Between fluorine (F) and chlorine (Cl), which non-metal is more reactive? Give reason. (3 marks)

Answer:

- Fluorine (F) is more reactive.
- F has smaller size and higher electronegativity.
- It can gain electron more easily than Cl.

Q7. Arrange the alkali metals (Li, Na, K) in order of increasing reactivity. (2 marks)

Answer: Li < Na < K

Q8. Arrange the halogens (F, Cl, Br, I) in order of decreasing reactivity. (2 marks)

Answer: F > Cl > Br > I

Q9. Why does reactivity of metals decrease across a period from left to right? (3 marks)

Answer:

- Atomic size decreases across the period.
- Ionization energy increases.
- Metals find it harder to lose electrons, so reactivity decreases.

Q10. Why does reactivity of non-metals increase across a period from left to right? (3 marks)

Answer:

- Atomic size decreases across the period.
- Electronegativity and nuclear charge increase.
- Non-metals can gain electrons more easily, so reactivity increases.

H. Density

Q1. Define density in relation to elements. (2 marks)

Answer: Density is the mass per unit volume of an element.

- Formula: Density = Mass/Volume
- Unit: g/cm^3 or kg/m^3

Q2. Explain the general trend of density down a group. (3 marks)

Answer:

- Density generally increases down a group.
- Both atomic mass and atomic size increase.
- Increase in mass is greater than increase in volume.

Q3. Why does density generally increase down the group? Explain. (3 marks)

Answer:

- Atomic mass increases significantly down the group.
- Atomic size also increases but to lesser extent.
- Since density = mass/volume, greater mass increase results in higher density.

Q4. Among sodium (Na) and potassium (K), which element has greater density and why? (3 marks)

Answer:

- Potassium (K) has greater density.
- K is below Na in Group 1 with higher atomic mass.
- Increase in mass is more than increase in volume.

Q5. Arrange the alkali metals (Li, Na, K, Rb) in order of increasing density. (2 marks)

Answer: Li < Na < K < Rb

Q6. Explain why density first increases and then decreases across a period. (3 marks)

Answer:

- Initially, mass increases while size decreases, increasing density.
- Towards right side, non-metals have open molecular structures.
- This causes density to decrease despite higher atomic mass.

Mixed Conceptual Questions

Q1. An element has low ionization energy and high atomic radius. Where would you expect to find this element in the periodic table? (3 marks)

Answer:

- This element is located in bottom-left corner of periodic table.
- These are alkali metals in lower periods (e.g., Cs, Rb).

- They have large size and lose electrons easily.

Q2. An element X has high electronegativity and high electron affinity. Predict whether X is a metal or non-metal and give reason. (3 marks)

Answer:

- Element X is a non-metal.
- High electronegativity and electron affinity indicate strong tendency to gain electrons.
- Metals have low values of these properties.

Q3. Element A has atomic number 11 and element B has atomic number 17. Which element has smaller atomic size and why? (3 marks)

Answer:

- Element B (Cl) has smaller atomic size.
- Both are in Period 3, but Cl has more protons (17 vs 11).
- Higher nuclear charge in Cl pulls electrons closer.

Q4. Compare the ionization energy of sodium (Na) and magnesium (Mg). Which is higher and why? (3 marks)

Answer:

- Magnesium (Mg) has higher ionization energy.
- Mg has more protons (12) than Na (11).
- Higher nuclear charge and smaller size hold electrons more tightly.

Q5. Why do elements in the same group show similar chemical properties? (3 marks)

Answer:

- Elements in same group have same number of valence electrons.
- Chemical properties depend on valence electrons.
- Same outer electron configuration gives similar chemical behavior.
