CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 3: Classification of Elements and Periodicity in Properties – Free PDF Download

Question: What is the basic theme of organisation in the periodic table?

Contents

Chapter 3 of NCERT Class 11 Chemistry, "Classification of Elements and Periodicity in Properties," traces the historical development of the periodic table from Döbereiner's triads to the Modern Periodic Law, and then explains the periodic trends in atomic radius, ionic radius, ionization enthalpy, electron gain enthalpy, valence and metallic/non-metallic character across periods and down groups.

NCERT Exercise Solutions

Question 3.1

Question: What is the basic theme of organisation in the periodic table?

Solution: The basic theme of organisation of elements in the periodic table is the classification of elements on the basis of similarities in their properties so that the study of their properties becomes systematic and simple. Elements are arranged in order of increasing atomic number, and elements with similar electronic configurations (and hence similar chemical properties) fall in the same vertical column (group), while elements with the same number of shells fall in the same horizontal row (period).

Question 3.2

Question: Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?

Solution: Mendeleev arranged the elements in order of increasing atomic masses, taking into account the similarities in chemical properties among the elements. He did not strictly stick to the order of atomic masses. For example, in some places he placed a heavier element before a lighter one (such as Co before Ni, and Te before I) to ensure that elements with similar properties fell into the same group.

Question 3.3

Question: What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?

Solution: Mendeleev's Periodic Law states that the physical and chemical properties of elements are a periodic function of their atomic masses. The Modern Periodic Law, given by Moseley, states that the physical and chemical properties of elements are a periodic function of their atomic numbers, not atomic masses. This resolved anomalies in Mendeleev's table (such as the position of argon and potassium) since atomic number, not atomic mass, is the fundamental property governing periodicity.

Question 3.4

Question: On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Solution: For the sixth period, the value of principal quantum number n = 6. The filling of orbitals in this period occurs in the order 6s, 4f, 5d, 6p. The maximum number of electrons that can be accommodated is: 6s (2) + 4f (14) + 5d (10) + 6p (6) = 32. Since the number of elements in a period equals the number of electrons that fill the outer/penultimate/antepenultimate shells in that period, the sixth period should contain 32 elements.

Question 3.5

Question: In terms of period and group where would you locate the element with Z = 114?

Solution: The electronic configuration of the element with Z = 114 is [Rn] $5f^{14}$ $6d^{10}$ $7s^2$ $7p^2$. Since the last electron enters the 7p orbital and the highest principal quantum number is 7, the element is located in period 7. Since it has 2 electrons in the p-orbital (p^2 configuration, analogous to carbon/silicon/germanium/tin/lead), it belongs to Group 14.

Question 3.6

Question: Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Solution: The element in the third period (n = 3) and group 17 is chlorine (Cl). Its electronic configuration is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$, so its atomic number is 17.

Question 3.7

Question: Which element do you think would have been named by (i) Lawrence Berkeley Laboratory (ii) Seaborg's group?

Solution: (i) Lawrencium (Lw/Lr, Z = 103) was named after the Lawrence Berkeley Laboratory (named for E. O. Lawrence, inventor of the cyclotron, where the element was discovered). (ii) Seaborgium (Sg, Z = 106) was named after Glenn T. Seaborg by his own research group, in recognition of his contribution to the discovery of many transuranium elements.

Question 3.8

Question: Why do elements in the same group have similar physical and chemical properties?

Solution: Elements in the same group have the same number of valence electrons and the same outer electronic configuration (differing only in the principal quantum number, n). Since chemical and many physical properties depend mainly on the number of valence electrons, elements of the same group exhibit similar chemical behaviour and gradually varying physical properties.

Question 3.9

Question: What does atomic radius and ionic radius really mean to you?

Solution: Atomic radius is the distance from the centre of the nucleus to the outermost shell containing electrons in a neutral, isolated atom. Since an electron cloud does not have a sharply defined boundary, it is generally measured indirectly as half the distance between the nuclei of two bonded/adjacent atoms (covalent radius, metallic radius, or van der Waals radius depending on the type of bonding). Ionic radius is the effective distance from the nucleus of an ion up to which it exerts influence on its electron cloud; it is measured as the distance between the nuclei of two adjacent ions in an ionic crystal, apportioned between the cation and anion.

Question 3.10

Question: How do atomic radius vary in a period and in a group? How do you explain the variation?

Solution: Across a period (left to right), atomic radius generally decreases because the number of shells remains the same while the nuclear charge increases, pulling the electron cloud closer to the nucleus (increasing effective nuclear charge). Down a group, atomic radius increases because a new shell of electrons is added at each successive element, and this increase in the number of shells outweighs the increase in nuclear charge, so the outermost electrons are farther from the nucleus.

Question 3.11

Question: What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.
(i) F^- (ii) Ar (iii) Mg^{2+} (iv) Rb^+

Solution: Isoelectronic species are atoms, molecules or ions that have the same number of electrons but may have different nuclear charges.
(i) F^- has 10 electrons; an isoelectronic species is Na^+ (or Ne, O^{2-}).
(ii) Ar has 18 electrons; an isoelectronic species is Ca^{2+} (or Cl^-, K^+).
(iii) Mg^{2+} has 10 electrons; an isoelectronic species is Ne (or F^-, Na^+).
(iv) Rb^+ has 36 electrons; an isoelectronic species is Kr (or Sr^{2+}, Br^-).

Question 3.12

Question: Consider the following species: N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+} and Al^{3+}
(a) What is common in them?
(b) Arrange them in the order of increasing ionic radii.

Solution: (a) All these species have 10 electrons each (electronic configuration $1s^2 2s^2 2p^6$); i.e., they are isoelectronic species.
(b) For isoelectronic species, ionic radius decreases as nuclear charge (atomic number) increases. The order of increasing atomic number is Al(13) < Mg(12) < Na(11) < F(9) < O(8) < N(7), so the order of increasing ionic radius (radius decreases with increasing nuclear charge) is: $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$.

Question 3.13

Question: Explain why cation are smaller and anions larger in radii than their parent atoms?

Solution: A cation is formed by the loss of one or more electrons from a neutral atom. This increases the effective nuclear charge per remaining electron (and may also remove an entire outer shell), so the electron cloud is pulled in more strongly, making the cation smaller than the parent atom. An anion is formed by the gain of one or more electrons. This increases electron-electron repulsion in the same shell while the nuclear charge stays the same, so the electron cloud expands, making the anion larger than the parent atom.

Question 3.14

Question: What is the significance of the terms — ‘isolated gaseous atom’ and ‘ground state’ while defining the ionization enthalpy and electron gain enthalpy?
Hint: Requirements for comparison purposes.

Solution: The term “isolated gaseous atom” ensures that the atom is free from the influence (bonding, interactions) of any other atoms, so that the energy change measured is purely due to the atom itself and not affected by neighbouring atoms/molecules as would happen in the liquid or solid state. The term “ground state” ensures that the atom is in its lowest energy state before the electron is removed or added, so that the ionization enthalpy/electron gain enthalpy values of different elements can be measured and compared under the same standard reference condition.

Question 3.15

Question: Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18}J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol^{-1}.
Hint: Apply the idea of mole concept to derive the answer.

Solution: Ionization enthalpy is the energy required to remove the electron from the ground state, i.e., it is equal in magnitude but opposite in sign to the ground state energy:
Ionization energy (per atom) = $-(-2.18\times10^{-18}$ J) = 2.18×10^{-18} J
For one mole of hydrogen atoms, multiply by Avogadro's number ($N_A = 6.022\times10^{23}$ mol^{-1}):
Ionization enthalpy = 2.18×10^{-18} J × 6.022×10^{23} mol^{-1}
= 1.3128×10^{6} J mol^{-1}
≈ 1.312×10^{6} J mol^{-1} (or 1312 kJ mol^{-1})

Question 3.16

Question: Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne.
Explain why
(i) Be has higher Δ_iH than B
(ii) O has lower Δ_iH than N and F?

Solution: (i) The electronic configuration of Be is $1s^22s^2$ and of B is $1s^22s^22p^1$. Be has a fully filled, extra-stable 2s orbital from which it is comparatively difficult to remove an electron, whereas B has to lose an electron from the higher-energy, less penetrating 2p orbital, which is easier to remove. So Be has a higher ionization enthalpy than B.
(ii) The electronic configuration of N is $1s^22s^22p^3$ (half-filled, extra stable p^3 configuration), while O is $1s^22s^22p^4$. In oxygen, one 2p orbital contains a pair of electrons, and the resulting electron-electron repulsion makes it easier to remove one of the paired electrons than to disturb the extra-stable half-filled configuration of nitrogen. Hence O has a lower ionization enthalpy than N (and, of course, than F, whose nuclear charge is even higher).

Question 3.17

Question: How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?

Solution: Sodium ($1s^22s^22p^63s^1$) has one electron in its outermost 3s orbital, which is comparatively easy to remove, while magnesium ($1s^22s^22p^63s^2$) has a filled, more stable $3s^2$ configuration and a higher nuclear charge, so its first ionization enthalpy is higher than that of sodium. After losing one electron, Na^+ attains a stable noble-gas-like configuration ($1s^22s^22p^6$), so removing a second electron from this very stable, smaller ion requires a very large amount of energy. Mg^+, on the other hand, still has one electron in the 3s orbital ($3s^1$) after the first ionization, which is comparatively easier to remove. Hence Na's second ionization enthalpy is much higher than Mg's second ionization enthalpy.

Question 3.18

Question: What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?

Solution: Down a group: (i) the atomic radius increases due to the addition of new shells, so the valence electrons are farther from the nucleus and experience weaker nuclear attraction; (ii) the effect of increasing nuclear charge is offset/screened by the increasing number of inner shells (shielding effect), so the effective nuclear charge experienced by the outer electrons does not increase proportionately. Together, these factors make it progressively easier to remove an electron, so ionization enthalpy decreases down a group.

Question 3.19

Question: The first ionization enthalpy values (in kJ mol^{-1}) of group 13 elements are:
B: 801, Al: 577, Ga: 579, In: 558, Tl: 589
How would you explain this deviation from the general trend?

Solution: Normally ionization enthalpy is expected to decrease steadily down a group. Here it decreases from B to Al as expected (due to increased atomic size and shielding), but it does not decrease smoothly further down — Ga has a slightly higher value than Al, and Tl has a higher value than In. This is because Ga (after the 3d series) and Tl (after the 4f/lanthanoid series) have a d^{10} (and, for Tl, f^{14}) configuration in the inner shells. The d and f electrons have poor shielding ability, so they do not effectively screen the outer electrons from the increased nuclear charge. This poor shielding by the intervening d and f electrons causes the effective nuclear charge on the outer electrons of Ga and Tl to be higher than expected, leading to irregular (higher than expected) ionization enthalpy values.

Question 3.20

Question: Which of the following pairs of elements would have a more negative electron gain enthalpy?
(i) O or F
(ii) F or Cl

Solution: (i) F has a more negative electron gain enthalpy than O, because F needs only one more electron to achieve a stable noble gas configuration, and it has a smaller size and higher effective nuclear charge than O, so it has a greater tendency to accept an electron.
(ii) Cl has a more negative electron gain enthalpy than F. Although F is smaller and has higher nuclear charge, its atomic size is so small that the incoming electron experiences significant electron-electron repulsion in its compact 2p subshell, reducing the energy released. Cl, being larger, has more space to accommodate the extra electron with less repulsion, so it releases more energy (more negative electron gain enthalpy) than F.

Question 3.21

Question: Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.

Solution: The second electron gain enthalpy of oxygen is positive (i.e., energy must be supplied). The first electron gain enthalpy of oxygen ($O \rightarrow O^-$) is negative (energy is released) because a neutral atom attracts an additional electron. However, adding a second electron to the already negatively charged O^- ion ($O^- \rightarrow O^{2-}$) requires overcoming the electrostatic repulsion between the negatively charged ion and the incoming electron, so energy must be supplied — making the second electron gain enthalpy positive.

Question 3.22

Question: What is the basic difference between the terms electron gain enthalpy and electronegativity?

Solution: Electron gain enthalpy is a measurable physical quantity — it is the enthalpy change that occurs when an isolated neutral gaseous atom in its ground state gains an electron to form a gaseous anion. Electronegativity is not a measurable quantity but only a relative number/tendency — it is the tendency of an atom in a covalently bonded molecule to attract the shared pair of electrons towards itself. Electron gain enthalpy is a property of an isolated atom, while electronegativity applies to an atom that is bonded to another atom in a molecule.

Question 3.23

Question: How would you react to the statement that the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds?

Solution: This statement is not correct. The electronegativity of an atom is not a fixed, invariant property — it depends on the hybridization state of the atom and the nature of the atom(s) it is bonded to (i.e., it varies from compound to compound). For example, the electronegativity of nitrogen differs slightly depending on whether it is sp, sp^2, or sp^3 hybridized in a given compound. So 3.0 is only a representative/average value, not a constant applicable to all nitrogen compounds.

Question 3.24

Question: Describe the theory associated with the radius of an atom as it
(a) gains an electron
(b) loses an electron

Solution: (a) When an atom gains an electron, it becomes an anion. The nuclear charge remains the same, but the number of electrons increases, resulting in increased electron-electron repulsion in the same shell. This causes the electron cloud to expand, so the radius increases (anion > parent atom).
(b) When an atom loses an electron, it becomes a cation. The remaining electrons experience a higher effective nuclear charge per electron (and an outer shell may be completely removed), so the electron cloud is pulled in closer to the nucleus. This causes the radius to decrease (cation < parent atom).

Question 3.25

Question: Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

Solution: The first ionization enthalpies of two isotopes of the same element are expected to be almost the same (nearly identical). This is because ionization enthalpy depends on the electronic configuration and the nuclear charge (number of protons), both of which are identical for isotopes of the same element. Isotopes differ only in the number of neutrons, which has a negligible direct effect on the force of attraction between the nucleus and the electrons, so the ionization enthalpy remains essentially unchanged.

Question 3.26

Question: What are the major differences between metals and non-metals?

Solution: Metals are electropositive elements: they have low ionization enthalpies, tend to lose electrons to form cations, have low electronegativity, form basic oxides, are good conductors of heat and electricity, are malleable and ductile, and are usually lustrous solids (except mercury). Non-metals are electronegative elements: they have high ionization enthalpies, tend to gain electrons to form anions, have high electronegativity, form acidic (or neutral) oxides, are generally poor conductors of heat and electricity, and are typically brittle solids, liquids, or gases.

Question 3.27

Question: Use the periodic table to answer the following questions.
(a) Identify an element with five electrons in the outer subshell.
(b) Identify an element that would tend to lose two electrons.
(c) Identify an element that would tend to gain two electrons.
(d) Identify the group having metal, non-metal, liquid as well as gas at the room temperature.

Solution: (a) Nitrogen (N, $2s^22p^3$) is an element with five electrons in its outer subshell ($2p^3 + 2s^2$, i.e., group 15 elements in general have ns^2np^3, five electrons in the outer shell).
(b) Magnesium (Mg, $3s^2$) tends to lose two electrons to attain a stable noble gas configuration, forming Mg^{2+}.
(c) Oxygen (O, $2s^22p^4$) tends to gain two electrons to attain a stable noble gas configuration, forming O^{2-}.
(d) Group 17 (halogens) has all three states at room temperature: fluorine and chlorine are gases, bromine is a liquid, and iodine and astatine are solids (with iodine being a non-metal and astatine having some metallic character).

Question 3.28

Question: The increasing order of reactivity among group 1 elements is Li < Na < K < Rb <Cs whereas that among group 17 elements is F > Cl > Br > I. Explain.

Solution: Group 1 (alkali metals) elements react by losing their single outermost electron; reactivity therefore depends on how easily this electron can be lost, i.e., on ionization enthalpy. Down the group, atomic size increases and ionization enthalpy decreases, so the electron is lost more easily, and reactivity increases from Li to Cs.
Group 17 (halogens) elements react by gaining one electron; reactivity therefore depends on the ease of gaining an electron, i.e., on electron gain enthalpy (and atomic size). Down the group, atomic size increases, so the attraction for an incoming electron decreases and the negative electron gain enthalpy magnitude generally decreases, making it harder to gain an electron. Hence reactivity decreases from F to I.

Question 3.29

Question: Write the general outer electronic configuration of s-, p-, d- and f- block elements.

Solution: s-block elements: ns^{1-2}
p-block elements: ns^2np^{1-6}
d-block elements: $(n-1)d^{1-10}ns^{0-2}$
f-block elements: $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$

Question 3.30

Question: Assign the position of the element having outer electronic configuration
(i) ns^2np^4 for n=3
(ii) $(n-1)d^2ns^2$ for n=4, and
(iii) $(n-2)f^7(n-1)d^1ns^2$ for n=6, in the periodic table.

Solution: (i) For n = 3, configuration is $3s^23p^4$. This is a p-block element in period 3, group 16 (p^4 means the 4th element of the p-block, i.e., group 12+4=16). This corresponds to sulphur (S).
(ii) For n = 4, configuration is $3d^24s^2$. This is a d-block element in period 4, group 4 (d^2 corresponds to group 3+2 = 4... following the d-block group numbering where d^n corresponds to group (n+2) for the first

few, i.e., group 4). This corresponds to titanium (Ti).
(iii) For n = 6, configuration is $4f^7 5d^1 6s^2$. This is an f-block element (inner transition element), period 6, and belongs to the actinoid/lanthanoid series — specifically it corresponds to gadolinium (Gd), placed in the lanthanoid series of period 6.

Question 3.31

Question: The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{eg}H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below:

Element I: $\Delta_i H_1$=520, $\Delta_i H_2$=7300, $\Delta_{eg}H$=–60
Element II: $\Delta_i H_1$=419, $\Delta_i H_2$=3051, $\Delta_{eg}H$=–48
Element III: $\Delta_i H_1$=1681, $\Delta_i H_2$=3374, $\Delta_{eg}H$=–328
Element IV: $\Delta_i H_1$=1008, $\Delta_i H_2$=1846, $\Delta_{eg}H$=–295
Element V: $\Delta_i H_1$=2372, $\Delta_i H_2$=5251, $\Delta_{eg}H$=+48
Element VI: $\Delta_i H_1$=738, $\Delta_i H_2$=1451, $\Delta_{eg}H$=–40

Which of the above elements is likely to be:
(a) the least reactive element
(b) the most reactive metal
(c) the most reactive non-metal
(d) the least reactive non-metal
(e) the metal which can form a stable binary halide of the formula MX_2 (X=halogen)
(f) the metal which can form a predominantly stable covalent halide of the formula MX (X=halogen)?

Solution:
(a) **Element V** — it has the highest first ionization enthalpy (2372) and a positive electron gain enthalpy, both characteristic of a noble gas, which is the least reactive.
(b) **Element II** — it has the lowest first ionization enthalpy (419) among the metals, so it loses an electron most easily, making it the most reactive metal (resembles Na).
(c) **Element III** — it has a high first ionization enthalpy and the most negative electron gain enthalpy (–328), so it has the strongest tendency to gain an electron, making it the most reactive non-metal (resembles Cl).
(d) **Element IV** — it has a fairly high ionization enthalpy and a negative but smaller electron gain enthalpy than element III, making it a non-metal but the least reactive of the non-metals present (resembles O).
(e) **Element VI** — its first and second ionization enthalpies are both moderate and comparable in magnitude (738, 1451), suggesting it readily loses two electrons to form a stable M^{2+} ion and hence an MX_2 halide (resembles Mg).
(f) **Element I** — it has a low first ionization enthalpy (520) but an extremely high second ionization enthalpy (7300, a huge jump), indicating it readily loses only one electron; because of its small size and high polarizing power (as in Li), the halide MX formed tends to have significant covalent character (resembles Li).

Question 3.32

Question: Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.
(a) Lithium and oxygen
(b) Magnesium and nitrogen
(c) Aluminium and iodine

(d) Silicon and oxygen
(e) Phosphorus and fluorine
(f) Element 71 and fluorine

Solution: (a) Li has valence +1 and O has valence −2, so the compound is Li_2O.
(b) Mg has valence +2 and N has valence −3, so the compound is Mg_3N_2.
(c) Al has valence +3 and I has valence −1, so the compound is AlI_3.
(d) Si has valence +4 and O has valence −2, so the compound is SiO_2.
(e) P has valence +5 (or +3) and F has valence −1, so the compound is PF_5 (or PF_3).
(f) Element 71 is lutetium (Lu), which has valence +3, and F has valence −1, so the compound is LuF_3.

Question 3.33

Question: In the modern periodic table, the period indicates the value of :
(a) atomic number
(b) atomic mass
(c) principal quantum number
(d) azimuthal quantum number

Solution: The correct answer is **(c) principal quantum number**. The period number corresponds to the highest principal quantum number (n) of the shell occupied by the outermost/valence electrons of the elements in that period.

Question 3.34

Question: Which of the following statements related to the modern periodic table is incorrect?
(a) The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-shell.
(b) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.
(c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.
(d) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Solution: The correct answer is **(b)**, which is the incorrect statement. This statement is wrong because the d-subshell can hold a maximum of 10 electrons (not 8), so the d-block actually has 10 columns, not 8.

Question 3.35

Question: Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell?
(a) Valence principal quantum number (n)
(b) Nuclear charge (Z)
(c) Nuclear mass
(d) Number of core electrons.

Solution: The correct answer is **(c) Nuclear mass**. The chemistry of an element (governed by its valence electrons) depends on the principal quantum number of the valence shell, the effective nuclear charge, and the shielding by core electrons — but the mass of the nucleus (determined by the number of neutrons and protons) has no significant effect on the valence shell or chemical behaviour.

Question 3.36

Question: The size of isoelectronic species — F^-, Ne and Na^+ is affected by
(a) nuclear charge (Z)
(b) valence principal quantum number (n)
(c) electron-electron interaction in the outer orbitals
(d) none of the factors because their size is the same.

Solution: The correct answer is **(a) nuclear charge (Z)**. F^-, Ne and Na^+ are isoelectronic (all have 10 electrons), so the number of shells (n) and the electron-electron interactions are essentially the same for all three. The only factor that differs among them is the nuclear charge (Z = 9, 10, 11 respectively); as Z increases, the electron cloud is pulled in more strongly, so the ionic/atomic radius decreases: F^- > Ne > Na^+.

Question 3.37

Question: Which one of the following statements is incorrect in relation to ionization enthalpy?
(a) Ionization enthalpy increases for each successive electron.
(b) The greatest increase in ionization enthalpy is experienced on removal of electron from core noble gas configuration.
(c) End of valence electrons is marked by a big jump in ionization enthalpy.
(d) Removal of electron from orbitals bearing lower n value is easier than from orbital having higher n value.

Solution: The correct answer is **(d)**, which is the incorrect statement. This statement is wrong because electrons in orbitals with a lower principal quantum number (n) are closer to the nucleus and experience greater nuclear attraction, so they are actually harder (not easier) to remove than electrons in orbitals with a higher n value.

Question 3.38

Question: Considering the elements B, Al, Mg, and K, the correct order of their metallic character is :
(a) B > Al > Mg > K
(b) Al > Mg > B > K
(c) Mg > Al > K > B
(d) K > Mg > Al > B

Solution: The correct answer is **(d) K > Mg > Al > B**. Metallic character increases down a group and decreases across a period (left to right). K (group 1) is the most metallic; Mg (group 2) is next; Al (group 13) is less metallic than Mg; and B (group 13, but a metalloid at the top of the group) is the least metallic among these four.

Question 3.39

Question: Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is :
(a) B > C > Si > N > F
(b) Si > C > B > N > F
(c) F > N > C > B > Si
(d) F > N > C > Si > B

Solution: The correct answer is **(c) F > N > C > B > Si**. Non-metallic character increases across a period (left to right) and decreases down a group. Comparing period 2 elements: F > N > C > B (non-metallic character decreases left to right in period 2, i.e., increases right to left, so F is most non-metallic, followed by N, then C, then B). Si, being in period 3 (below C), is less non-metallic than its period-2 counterpart C, and in fact less non-metallic than B as well, placing it last in the sequence.

Question 3.40

Question: Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidizing property is :
(a) F > Cl > O > N
(b) F > O > Cl > N
(c) Cl > F > O > N
(d) O > F > N > Cl

Solution: The correct answer is **(b) F > O > Cl > N.** Oxidizing power depends on the tendency of an element to accept electrons, which is related to electronegativity. F, being the most electronegative element, is the strongest oxidizer, followed by O, then Cl, and finally N, which is the weakest oxidizer among these four due to the extra stability of its half-filled p^3 configuration.

Notes and Extra Questions

This chapter builds the conceptual foundation for the entire periodic table: it moves from the historical attempts at classification (Döbereiner, Newlands, Mendeleev) to the Modern Periodic Law based on atomic number, and finally to a detailed, quantitative treatment of periodic trends — atomic/ionic radius, ionization enthalpy, electron gain enthalpy, electronegativity, and metallic/non-metallic character — explained in terms of nuclear charge, shielding, and electronic configuration. Students should focus on being able to explain (not just memorise) each trend using effective nuclear charge and shielding, since a large share of the 2026-27 exercise questions (including several new assertion-style multiple-choice questions, 3.33 to 3.40) test this reasoning directly. Diagonal relationships, anomalous properties of second-period elements, and the s-, p-, d-, f-block classification are also important recurring exam themes from this chapter.