NCERT Solutions for Class-XI Chemistry

Chapter-3

NCERT Chemistry Class 11

1. What is the basic theme of organisation in the periodic table?

1. The basic theme of organisation of elements in the periodic table is to classify the elements in periods and groups according to their properties. This arrangement makes the study of elements and their compounds simple and systematic. In the periodic table, elements with similar properties are placed in the same group.

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

2. The first meaningful and remarkable contribution in the field of classification of elements was done by Russian Chemist by the name Dmitri I. Mendeleev in the year of 1869 and he proposed a law called Mendeleev periodic law which says that the physical & the chemical properties of the elements are a periodic function of their atomic masses.

On the foundation of this law, he developed a periodic table which was named as Mendeleev Periodic table, where he arranged all the elements in his periodic table in terms of their atomic weight or mass. In the table he made vertical columns called groups and vertical rows called as periods for the arranging the elements. He placed the elements with similar physical and chemical properties in the same group.

Nevertheless, he did not adhere to this system for long. He observed that if the elements were arranged strictly in terms of their increasing atomic weights, then some elements could not fit within this system of categorization.

Therefore, the order of atomic weights was ignored in some cases. For instance, the atomic weight of iodine is lesser than that of tellurium. Then to tellurium [in Group VI] was placed before iodine [in Group VII] just because iodine's properties are so similar to that of fluorine, chlorine, and bromine. Also argon was placed before potassium.

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

3. Mendeleev's Periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic weights. On the other hand, the Modern periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic numbers.

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

4. The principal quantum number is a number which determines the main energy level or principal shell to which the electron of the atom belongs. It also gives the average distance of the electron from the nucleus and also specifies the energy value of the electron.

Azimuthal quantum number is the figure which determines the sub shell in a principal shell to which an electron belongs and it also signifies the number of subshell present in the principal shell.

In the periodic table, a period indicates the value of the principal quantum number [n] denotes the total number of shells present in the atom of the element. Each period starts with the filling of principal quantum number [n]. The value of principal quantum number for the sixth period is 6. For n = 6, azimuthal quantum number [L] can have values of 0, 1,2,3,4. According to Aufbau's principle, electrons are added to different orbitals in order of their increasing energies. Thus the 6d shell has a greater energy as compared to the 7s subshell.

In the 6^{th} period, electrons can be filled in only 6s, 4f, 5d, and 6 p subshells. So the 6s subshell consists of 1 orbital, 4f subshell consists of 7 orbitals, 5d subshell has 5 orbitals and 6p subshell has three orbitals. As a result the total number of orbitals present is 16. According to Pauli's exclusion principle, only 2 electrons can be accommodated in each orbital. So, 32 electrons can be accommodated at a maximum level in 16 orbitals.

Therefore 32 elements would be present in the 6^{th} period of the periodic table.

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

5. Elements with atomic numbers from Z = 87 to Z = 114 are present in the 7^{th} period of the periodic table. Thus, the element with Z = 114 is present in the 7^{th} period of the periodic table.

In the 7^{th} period, first two elements with Z = 87 and Z= 88 are s-block elements, the next 14 elements excluding Z = 89 i.e., those with Z = 90 – 103 are f – block elements, ten elements with Z = 89 and Z = 104 – 112 are d – block elements, and the elements with Z = 113 – 118 are p – block elements. Therefore, the element with Z = 114 is the second p – block element in the 7^{th} period. Thus, the element with Z = 114 is present in the 7^{th} period and 4^{th} group of the periodic table.

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

6. 1^{st} period of the periodic table has only 2 elements and 2^{nd} period consists of only eight elements. The first element of the third period is sodium with atomic number 11. Now, there are nine elements in the third period. The last element of the third period is Argon which has atomic number of 18. So the 17^{th} group element of third period is chlorine with atomic number = 17 and atomic mass = 35.453g. It belongs to p-block.

7. Which element do you think would have been named by

(i) Lawrence Berkeley Laboratory

(ii) Seaborg's group?

7. (i) Lawrencium (Lr) with Z = 103 and Berkelium (Bk) with Z = 97
(ii) Seaborgium (Sg) with Z = 106

8. Why do elements in the same group have similar physical and chemical properties?

8. The total number of groups in the modern periodic table is 18 and each group id distinct from every other group. The elements in the same group contains same number of valence electrons and hence the same group elements possess similar physical and chemical properties as the physical and chemical properties is dependent on the number of valence electrons in the element.

9. What does atomic radius and ionic radius really mean to you?

9. Atomic radius is the radius of an atom. It measures the size of an atom. If the element is a metal, then the atomic radius refers to the metallic radius, and if the element is a nonmetal, then it refers to the covalent radius. Metallic radius is calculated as half the internuclear distance separating the metal cores in the metallic crystal. For example, the internuclear distance between two adjacent copper atoms in solid copper is 256 pm.

Thus, the metallic radius of copper is taken as $\frac{256}{2}\,pm = 128\,pm$.

Covalent radius is measured as the distance between two atoms when they are found together by a single bond in a covalent molecule. For example, the distance between two chlorine atoms in chlorine molecule is 198 pm. Thus, the covalent radius of chlorine is taken as $\frac{198}{2}\,pm = 99\,pm$.

Ionic radius means the radius of an ion (cation or anion). The ionic radii can be calculated by measuring the distances between the cations and anions in ionic crystals. Since a cation is formed by removing an electron from an atom, the cation has fewer electrons than the parent atom resulting in an increase in the effective nuclear charge. Thus, a cation is smaller than the parent atom. For example, the ionic radius of Na^+ ion is 95 pm, whereas the atomic radius of Na atom is 186 pm. On the other hand, an anion is larger in size than its parent atom. This is because an anion has the same nuclear charge, but more electrons than the parent atom resulting in an increased repulsion among the electrons and a decrease in the effective nuclear charge. For example, the ionic radius of F^- ion is 136 pm, whereas the atomic radius of F atom is 64 pm.

10. How does atomic radius vary in a period and in a group? How do you explain the variation?

10. Atomic radius of the elements generally decreases from left to the right in a period because on moving from left to right in a period the number of shells remains the same but more and more electrons get added to the atoms and hence the effective nuclear charge increases and thereby pulling the shells closer to the nucleus and hence atomic radius decreases on moving from left to right in periodic table.

Atomic radius of the elements increases on going down the group because on going down the group there is an increase in the number of shells. As the number of shells increases, the valence shell gets farther and farther away from the nucleus resulting in reduction of overall effective nuclear charge. Therefore the atomic size is obviously expected to increase.

11. 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^+

11. Atoms and ions having the same number of electrons are called isoelectronic species.
(i) F^- ion has 9 + 1 = 10 electrons. Thus, the species isoelectronic with it will also have 10 electrons. Some of its isoelectronic species are Na^+ ion (11 – 1 = 10 electrons), Ne (10 electrons), O^{2-} ion (8 + 2 = 10 electrons), and Al^{3+} ion (13 – 3 = 10 electrons).
(ii) Ar has 18 electrons. Thus, the species isoelectronic with it will also have 18 electrons. Some of its isoelectronic species are S^{2-} ion (16 + 2 = 18 electrons), Cl^- ion (17 + 1 = 18 electrons), K^+ ion (19 – 1 = 18 electrons), and Ca^{2+} ion (20 – 2 = 18 electrons).
(iii)Mg^{2+} ion has 12 – 2 = 10 electrons. Thus, the species isoelectronic with it will also have 10 electrons. Some of its isoelectronic species are F^- ion (9 + 1 = 10 electrons), Ne (10 electrons), O^{2-} ion (8 + 2 = 10 electrons), and Al^{3+} ion (13 – 3 = 10 electrons).
(iv)Rb^+ ion has 37 – 1 = 36 electrons. Thus, the species isoelectronic with it will also have 36 electrons. Some of its isoelectronic species are Br^- ion (35 + 1 = 36 electrons), Kr (36 electrons), and Sr^{2+} ion (38 – 2 = 36 electrons).

12. 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.

12. A. What is common in them?
In the given examples of isoelectronic species have a same number of electrons which is equal to 10. Hence, the given species are isoelectronic, i.e.
N^{3-} has 7+3 = 10 electrons
O^{2-} has 8+2 = 10 electrons
F^- has 9+1 = 10 electrons
Na^+ has 11-1 = 10 electrons
Mg^{2+} has 12-2 = 10 electrons
AL^{3+} has 13-3 = 10 electrons
B. Arrange them in the order of increasing ionic radii.
As the magnitude of the effective nuclear decreases, the ionic radii of the ions increase.
The following arrangement denotes the elements arranged in the order of increasing nuclear charge:
$N^{3-} < O^{2-} < F^- < Na^+ < Mg^{2+} < Al^{3+}$
Nuclear charge = +7 +8 +9 +11 +12 +13
The following arrangement denotes the elements arranged in the order of increasing ionic radii:

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

13. Explain why cations are smaller and anions larger in radii than their parent atoms?

13. A cation has a fewer number of electrons than its parent atom, while its nuclear charge remains the same. As a result, the attraction of electrons to the nucleus is more in a cation than in its parent atom. Therefore, a cation is smaller in size than its parent atom. On the other hand, an anion has one or more electrons than its parent atom, resulting in an increased repulsion among the electrons and a decrease in the effective nuclear charge. As a result, the distance between the valence electrons and the nucleus is more in anions than in it's the parent atom. Hence, an anion is larger in radius than its parent atom.

14. 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.

14. Ionization enthalpy is the minimal quantity of energy which is demanded to get rid of the most loosely bound electron from a neutral isolated gaseous atom to form a cation. A cation is formed when an atom loses electrons. But for the removal electrons the atoms should be in isolated gaseous form. Though a gaseous atom is randomly separated but there exists some attractive forces between the atoms. To calculate the ionization enthalpy, it is practically impossible to isolate a single atom. By lowering the pressure it is possible to minimize the force of attraction between the atoms. This is the reason why the term 'isolated gaseous atom' is used in the definition of the ionization enthalpy.
The most stable state of a gaseous atom is its ground state and in the ground state of the atom, it is very easy to remove the electron from the atom. Hence the term ground state is used in the definition.

15. 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.

15. The energy of an electron in the ground state of the hydrogen atom is –2.18 × 10–18 J. Therefore, the energy required to remove that electron from the ground state of hydrogen atom is
2.18×10^{-18} J.
Ionization enthalpy of atomic hydrogen = 2.18×10^{-18} J
Hence, ionization enthalpy of atomic hydrogen in terms of J mol^{-1}
$= 2.18 \times 10^{-18} \times 6.02 \times 10^{23}$ J mol^{-1} = 1.31×10^{6} J mol^{-1}

16. 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 $\Delta_i H$ than B
(ii) O has lower $\Delta_i H$ than N and F?

16. (i) symmetric nature as compared to boron. The electron of beryllium is places in 2s subshell and the 2s subshell is more tightly held to the nucleus as compared to 2p subshell of boron. Thus during ionization more energy is required to remove an electron from beryllium atom as compared to boron atom. Hence, beryllium has higher Δ_i H than boron.

(ii) In nitrogen atom all three orbitals of the 2p subshell are fully filled and in case of oxygen one orbital is fully filled and the two orbitals are partially filled of 2p subshell. So nitrogen has greater symmetricity as compared to oxygen and hence it is difficult to remove an electron from nitrogen atom as compared to oxygen atom. Thus nitrogen has higher value of ionization enthalpy as compared to oxygen.

Fluorine has one electron and proton more as compared to oxygen atom and hence in case of fluorine atom the electron is more tightly held as compared to oxygen atom. Thus it is difficult to remove an electron from fluorine atom as compared to oxygen atom. Thus fluorine has higher value of ionization enthalpy as compared to oxygen.

17. 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?

17. The first ionization enthalpy of sodium is more than that of magnesium. This is primarily because of two reasons:

(a) The atomic size of sodium is greater than that of magnesium

(b) The effective nuclear charge of magnesium is higher than that of sodium

For these reasons, the energy required to remove an electron from magnesium is more than the energy required in sodium. Hence, the first ionization enthalpy of sodium is lower than that of magnesium. However, the second ionization enthalpy of sodium is higher than that of magnesium. This is because after losing an electron, sodium attains the stable noble gas configuration. On the other hand, magnesium, after losing an electron still has one electron in the 3s-orbital. In order to attain the stable noble gas configuration, it still has to lose one more electron. Thus, the energy required to remove the second electron in case of sodium is much higher than that required in case of magnesium. Hence, the second ionization enthalpy of sodium is higher than that of magnesium.

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

18. Ionization enthalpy is the minimal quantity of energy which is demanded to get rid of the most loosely bound electron from a neutral isolated gaseous atom to form a cation. The factors responsible for the decrease in ionization enthalpy in moving down the group are given as follows:

[i] Increase in the atomic size of elements: On going down the group the number of shells in the atom also increases. As a result, the size of the atom also increases gradually on going down the group. As a result the valence electrons get farther away from the nucleus and thus weakening the effective nuclear charge of the atom and hence

electrons can be easily removed from the atoms. Hence, on going down the group, ionization energy decreases.
[ii] Increase in the shielding effect: The number of inner shells of electrons increases on moving down a group. The inner shells or orbits of the atom acts as a shield for the outer or valence shell electrons. The inner shell shields the nuclear charge and as a result only a small amount of nuclear charge is able to attract the valence shell electrons. Hence the energy required to remove the electron reduces on moving down the group.

19. The first ionization enthalpy values (in $kJmol^{-1}$) of group 13 elements are:

B	Al	Ga	In	Tl
801	577	579	558	589

How would you explain this deviation from the general trend?

19. On moving down a group, ionization enthalpy generally decreases due to an increase in the atomic size and shielding. Thus, on moving down group 13, ionization enthalpy decreases from B to Al. But, Ga has higher ionization enthalpy than Al. Al follows immediately after s–block elements, whereas Ga follows after d–block elements. The shielding provided by d-electrons is not very effective. These electrons do not shield the valence electrons very effectively. As a result, the valence electrons of Ga experience a greater effective nuclear charge than those of Al. Further, moving from Ga to In, the ionization enthalpy decreases due to an increase in the atomic size and shielding. But, on moving from In to Tl, the ionization enthalpy again increases. In the periodic table, Tl follows after 4f and 5d electrons. The shielding provided by the electrons in both these orbitals is not very effective. Therefore, the valence electron is held quite strongly by the nucleus. Hence, the ionization energy of Tl is on the higher side.

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

20. (i) O and F are present in the same period of the periodic table. An F atom has one proton and one electron more than O and as an electron is being added to the same shell, the atomic size of F is smaller than that of O. As F contains one proton more than O, its nucleus can attract the incoming electron more strongly in comparison to the nucleus of O atom. Also, F needs only one more electron to attain the stable noble gas configuration. Hence, the electron gain enthalpy of F is more negative than that of O.
(ii) F and Cl belong to the same group of the periodic table. The electron gain enthalpy usually becomes less negative on moving down a group. However, in this case, the value of the electron gain enthalpy of Cl is more negative than that of F. This is because the atomic size of F is smaller than that of Cl. In F, the electron will be added to quantum level n = 2, but in Cl, the electron is added to quantum level n = 3. Therefore, there are less electron- electron repulsions in Cl and an additional electron can be accommodated easily. Hence, the electron gain enthalpy of Cl is more negative than that of F.

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

21. When an electron is added to O atom to form O^- ion, energy is released. Thus, the first electron gain enthalpy of O is negative.

$$O_{(g)} + e^- \rightarrow O^-_{(g)}$$

On the other hand, when an electron is added to O^- ion to form O^{2-} ion, energy has to be given out in order to overcome the strong electronic repulsions. Thus, the second electron gain enthalpy of O is positive.

$$O^-_{(g)} + e^- \rightarrow O^{2-}_{(g)}$$

22. What is the basic difference between the terms electron gain enthalpy and electronegativity?

22.

Sl. No	Electron Gain Enthalpy	Electronegativity
1	It is the tendency of an atom to attract outside electrons.	It is the tendency of an atom to attract shared pair of electrons.
2	It is the property of an isolated atom.	It is the property of bonded atom.
3	It is the absolute electron attracting tendency of the atom.	It is the relative electron attracting tendency of an atom.
4	Electron gain enthalpy can be either positive or negative depending on the electron configuration of the atom that is going to gain an electron.	Electronegativity is always a positive value.
5	It has units like KJ/mol & Ev/atom.	It has no units.

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

23. Electronegativity of an element is a variable property. It is different in different compounds. Hence, the statement which says that the electronegativity of N on Pauling scale is 3.0 in all nitrogen compounds is incorrect. The electronegativity of N is different in NH_3 and NO_2.

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

24. (a) When an atom gains an electron, its size increases. When an electron is added, the number of electrons goes up by one. This results in an increase in repulsion among the electrons. However, the number of protons remains the same. As a result, the effective nuclear charge of the atom decreases and the radius of the atom increases.
(b) When an atom loses an electron, the number of electrons decreases by one while the nuclear charge remains the same. Therefore, the interelectronic repulsions in the atom decrease. As a result, the effective nuclear charge increases. Hence, the radius of the atom decreases.

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

25. The ionization enthalpy of an atom depends on the number of electrons and protons (nuclear charge) of that atom. Now, the isotopes of an element have the same number of protons and electrons. Therefore, the first ionization enthalpy for two isotopes of the same element should be the same.

26. What are the major differences between metals and non-metals?

26.

Sl. No	Metals	Non-Metals
1	Metals can lose electrons easily.	Non-metals cannot lose electrons easily.
2	Metals cannot gain electrons easily.	Non-metals can gain electrons easily.
3	Metals generally form ionic compounds.	Non–metals generally form covalent compounds.
4	Metals oxides are basic in nature.	Non–metallic oxides are acidic in nature
5	Metals have low ionization enthalpies.	Non–metals have high ionization enthalpies.
6	Metals have less negative electron gain enthalpies.	Non–metals have high negative electron gain enthalpies.
7	Metals are less electronegative. They are rather electropositive elements.	Non–metals are comparatively more electronegative.
8	Metals have a high reducing power.	Non–metals have a low reducing power.

27. 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.

27. (a) The electronic configuration of an element having 5 electrons in its outermost subshell should be ns^2 np^5. This is the electronic configuration of the halogen group. Thus, the element can be F, CL, Br, I, or At.

(b) An element having two valence electrons will lose two electrons easily to attain the stable noble gas configuration. The general electronic configuration of such an element will be ns^2. This is the electronic configuration of group 2 elements. The elements present in group 2 are Be, Mg, Ca, Sr, Ba.

(c) An element is likely to gain two electrons if it needs only two electrons to attain the stable noble gas configuration. Thus, the general electronic configuration of such an element should be ns^2 np^4. This is the electronic configuration of the oxygen family.

(d) Group 17 has metal, non–metal, liquid as well as gas at room temperature.

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

28. The elements present in group 1 have only 1 valence electron, which they tend to lose. Group 17 elements, on the other hand, need only one electron to attain the noble gas configuration. On moving down group 1, the ionization enthalpies decrease. This means that the energy required to lose the valence electron decreases. Thus, reactivity increases on moving down a group. Thus, the increasing order of reactivity among group 1 elements is as follows:

Li < Na < K < Rb < Cs

In group 17, as we move down the group from Cl to I, the electron gain enthalpy becomes less negative i.e., its tendency to gain electrons decreases down group 17. Thus, reactivity decreases down a group. The electron gain enthalpy of F is less negative than Cl. Still, it is the most reactive halogen. This is because of its low bond dissociation energy. Thus, the decreasing order of reactivity among group 17 elements is as follows: F > Cl > Br > I.

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

29.

Element	General outer electronic configuration
s–block	ns^{1-2}, where n = 2 – 7
p–block	ns^2np^{1-6}, where n = 2 – 6
p–block	$(n-1)\ d^{1-10}\ ns^{0-2}$, where n = 4 – 7
p–block	$(n-2)f^{1-14}(n-1)d^{0-10}ns^2$, where n = 6 – 7

30. Assign the position of the element having outer electronic configuration
(a) $ns^2\ np^4$ for n = 3
(b) $(n - 1)d^2\ ns^2$ for n = 4, and
(c) $(n - 2)\ f^7\ (n - 1)d^1\ ns^2$ for n = 6, in the periodic table.

30. (i) Since n = 3, the element belongs to the 3rd period. It is a p–block element since the last electron occupies the p–orbital. There are four electrons in the p–orbital.
Thus, the corresponding group of the element is :
= Number of s–block groups + number of d–block groups + number of p–electrons
= 2 + 10 + 4 = 16
Hence, the element belongs to the 3rd period and 16th group of the periodic table. Hence, the element is Sulphur.
(ii) Since n = 4, the element belongs to the 4th period. It is a d–block element as d–orbitals are incompletely filled. There are 2 electrons in the d–orbital.
Thus, the corresponding group of the element is :
= Number of s–block groups + number of d–block groups
= 2 + 2 = 4
Hence, it is a 4th period and 4th group element. Therefore, the element is Titanium.
(iii) Since n = 6, the element is present in the 6th period. It is an f –block element as the last electron occupies the f–orbital. It belongs to group 3 of the periodic table since all f-block elements belong to group 3. Its electronic configuration is [Xe] $4f^7\ 5d^1\ 6s^2$. Thus, its atomic number is 54 + 7 + 2 + 1 = 64. Hence, the element is Gadolinium.

31. The first (Δ_iH1) and the second (Δ_iH) 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:

Elements	Δ_iH	Δ_iH	$\Delta_{eg}H$
I	520	7300	–60
II	419	3051	–48
III	1681	3374	–328

IV	1008	1846	–295
V	2372	5251	+48
VI	738	1451	–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)?

31. (a) Element V is likely to be the least reactive element. This is because it has the highest first ionization enthalpy ($\Delta_i H_1$) and a positive electron gain enthalpy ($\Delta_{eg}H$).

(b) Element II is likely to be the most reactive metal as it has the lowest first ionization enthalpy ($\Delta_i H_1$) and a low negative electron gain enthalpy ($\Delta_{eg}H$).

(c) Element III is likely to be the most reactive non–metal as it has a high first ionization enthalpy ($\Delta_i H_1$) and the highest negative electron gain enthalpy ($\Delta_{eg}H$).

(d) Element V is likely to be the least reactive non–metal since it has a very high first ionization enthalpy ($\Delta_i H_2$) and a positive electron gain enthalpy ($\Delta_{eg}H$).

(e) Element VI has a low negative electron gain enthalpy ($\Delta_{eg}H$). Thus, it is a metal. Further, it has the lowest second ionization enthalpy ($\Delta_i H_2$). Hence, it can form a stable binary halide of the formula MX_2 (X = halogen).

(f) Element I has low first ionization energy and high second ionization energy. Therefore, it can form a predominantly stable covalent halide of the formula MX (X = halogen).

32. Predict the formula 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

32. (a) LiO_2

(b) Mg_3N_2

(c) AlI_3

(d) SiO_2

(e) PF_3 or PF_5

The element with the atomic number 71 is Lutetium (Lu). It has valency 3. Hence, the formula of the compound is LuF_3.

33. 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.

33. The value of the principal quantum number (n) for the outermost shell or the valence shell indicates a period in the Modern periodic table.

34. 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.

34. The d-block has 10 columns because a maximum of 10 electrons can occupy all the orbitals in a d subshell.
Hence, the statement related to the modern periodic table which is incorrect is - B. The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.

35. 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.

35. Nuclear mass does not affect the valence electrons

36. 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.

36. The size of an isoelectronic species increases with a decrease in the nuclear charge (Z).
For example, the order of the increasing nuclear charge of F^-, Ne, and Na^+ is as follows:

$$F^-(Z=9) < Ne(Z=10) < Na^+(Z=11)$$

Therefore, the order of the increasing size of F^-, Ne and Na^+ is as follows:

$$Na^+ < Ne < F^-$$

Hence, the correct option is - (a) nuclear charge (Z).

37. 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.

37. Electrons in orbitals bearing a lower n value are more attracted to the nucleus than electrons in orbitals bearing a higher n value. Hence, the removal of electrons from orbitals bearing a higher n value is easier than the removal of electrons from orbitals having a lower n value.

38. 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

38. The metallic character of elements decreases from left to right across a period. Thus, the metallic character of Mg is more than that of Al. The metallic character of elements increases down a group. Thus, the metallic character of Al is more than that of B. Considering the above statements, we get K > Mg and the only relation out of the given relations which holds the above condition is - (d) K > Mg > Al > B.
Hence, the correct order of metallic character of the elements B, Al, Mg, and K is – (d) K > Mg > Al > B.

39. 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

39. The non-metallic character of elements increases from left to right across a period. Thus, the decreasing order of non-metallic character is F > N > C > B.
Again, the non-metallic character of elements decreases down a group. Thus, the decreasing order of non-metallic characters of C and Si are C > Si. However, Si is less non-metallic than B i.e., B > Si. Hence, the correct order of their non-metallic characters is F > N > C > B > Si.

40. 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) Si > C > B > N
(c) Cl > F > O > N
(d) O > F > N > Cl

40. The oxidizing character of elements increases from left to right across a period. Thus, we get the decreasing order of oxidizing property as F > O > N. Again, the oxidizing character of elements decreases down a group. Thus, we get F > Cl. However, the oxidizing character of O is more than that of Cl i.e., O > Cl. Hence, the correct order of chemical reactivity of F, Cl, O, and N in terms of their oxidizing property is F > O > Cl > N.

Therefore, the correct option is - (b) F > O > Cl > N