Classification of Elements and Periodicity in Properties

3 Chapter

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

Ans: The primary idea of the periodic table's organisation is to organise elements into periods and groups based on their qualities. This arrangement simplifies and organises the study of elements and their compounds. Elements with comparable characteristics are grouped together in the periodic table.

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

Ans: In his periodic chart, Mendeleev organised the elements by atomic weight or mass. In order of increasing atomic weight, he divided the elements into periods and groups. He grouped items that have comparable qualities together. He did not, however, keep this arrangement for long. He discovered that if the elements were ordered strictly in order of increasing atomic weights, several of them would not fit into this classification scheme.

As a result, in certain circumstances, he neglected the atomic weight order. The atomic weight of iodine, for example, is smaller than that of tellurium. Despite this, Mendeleev ranked tellurium (Group VI) ahead of iodine (Group VII) because the characteristics of iodine are so similar to fluorine, chlorine, and bromine.

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

Ans: The physical and chemical properties of elements are periodic functions of their atomic weights, according to Mendeleev's Periodic Law. The Modern Periodic Law, on the other hand, maintains that an element's physical and chemical properties are periodic functions of its atomic number.

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

Ans: A period is the value of the primary quantum number (n) for the outermost shells in the periodic table of the elements. The filling of the primary quantum number begins each period (n). For the sixth period, the value of (n) is 6. The azimuthal quantum number (l) for $n=6$ can be $0,1,2,3$ or 4. The Aufbau principle states that electrons are introduced to distinct orbitals in order of increasing energy. The $6d$ subshell has an even higher energy than the $7s$ subshell.

Only the $6s, 4f, 5d$ and $6p$ subshells can be filled by electrons in the 6^{th} period. Now, the $6s$ has one orbital, the $4f$ has seven, the $5d$ has five, and the $6p$ has three. As a result, there are a total of sixteen orbitals available $(1+7+5+3=16)$. Each orbital can only hold two electrons, according to Pauli's exclusion principle. As a result, 16 orbitals can hold a maximum of 32 electrons. As a result, the periodic table's sixth period should have 32 elements.

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

Ans : The periodic table's 7^{th} period has elements with atomic values ranging from $Z=87-114$. As a result, the element with $Z=114$ is found in the periodic table's 7^{th} period. The first two items with $Z=87$ and $Z=88$ are s-block elements in the seventh period, while the next 14 elements with $Z=90-103$ are f-block elements. d – block elements are those with $Z=89$ and $Z=104-112$, whereas p – block elements are those with $Z=113-118$. As a result, in the 7^{th} period, the element with

$Z=114$ is the second p – block element. As a result, the element with $Z=114$ is found in the periodic table's 7^{th} period and 4^{th} group.

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

Ans: The first phase has two elements, while the second period has eight. The element with $Z=11,$ begins the third period. In the third period, there are now eight elements. As a result, the third period finishes with the element with $Z=18,$ i.e., the element in the third period's 18^{th} group. As a result, the element in the third period's 17^{th} group has the atomic number $Z=17.$

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

Ans: (i) Lawrencium (Lr) has a Z of 103 while, Berkelium (Bk) has a Z of 97.

(ii) Seaborgium (Sg) has a Z of 106.

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

Ans: The quantity of valence electrons affects the physical and chemical properties of elements. The number of valence electrons in a group of elements is the same. As a result, physical and chemical properties of elements in the same group are comparable.

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

Ans: The radius of an atom is known as its atomic radius. It is used to determine the size of an atom. If the element is a metal, the atomic radius is the metallic radius; if it is a nonmetal, the atomic radius is the covalent radius. The internuclear distance separating the metal cores in a metallic crystal is half the metallic radius. In solid copper, for example, the internuclear distance between two neighbouring copper atoms is 256 pm. As a result, the metallic radius of copper is calculated to be,

$$\begin{aligned}\text{Metallic radius of copper} &= \frac{256}{2}\\ &= 128\text{ pm}\end{aligned}$$

In a covalent molecule, the covalent radius is the distance between two atoms when they are connected united by a single bond. The distance between two chlorine atoms in a chlorine molecule, for example, is 198 pm. Thus, the covalent radius of chlorine is,

$$\begin{aligned}\text{Covalent radius of chlorine} &= \frac{198}{2}\\ &= 99\text{ pm}\end{aligned}$$

The radius of an ion is referred to as its ionic radius (cation or anion). The distances between the cations and anions in ionic crystals can be used to compute the ionic radii. Because a cation is generated by removing an electron from an atom, it has less electrons than the parent atom and hence has a higher effective nuclear charge. As a result, a cation is smaller than its parent. The ionic radius of Na^{+} is 95 pm, while the atomic radius of the Na atom is 186 pm. An anion, on the other hand, is larger than its parent atom. This is due to the fact that an anion has the same nuclear charge as the parent atom but contains more electrons, resulting in higher electron repulsion and a drop in the effective nuclear charge. The ionic radius of the F^{-} ion, for example, is 136 pm, while the atomic radius of the F atom is 64 pm.

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

Ans: Across a period, the atomic radius decreases from left to right. Because the outer electrons are present in the same valence shell during a period and the atomic number increases from left to right across a period, the effective nuclear charge increases. As a result, electrons become more attracted to the nucleus.

The atomic radius, on the other hand, tends to rise as you progress through the groups. This is due to the fact that as the primary quantum number (n) decreases, the distance between the nucleus and valence electrons rises.

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^+

Ans: Isoelectronic species refer to atoms and ions that have the same number of electrons.

(i) The F^- ion possesses a total of 10 electrons. As a result, the species that isoelectronic with it will contain 10 electrons as well. Na^+ ion $(11-1=10)$ electrons, Ne (10 electrons), O^{2-} ion $(8+2=10)$ electrons, are some of its isoelectronic species.

(ii) Ar possesses a total of 18 electrons. As a result, the species that isoelectronic with it will contain 18 electrons as well. S^{2-} ion $(16+2=18)$ electrons, Cl^- ion $(17+1=18)$ electrons, and Ca^{2+} ion $(20-2=18)$ electrons are some of its isoelectronic species.

(iii) The electron count of the Mg^{2+} ion is 10 electrons. As a result, the species that isoelectronic with it will contain 10 electrons as well. Na^{+} ion $(11-1=10)$ electrons, Ne (10 electrons), O^{2-} ion $(8+2=10)$ electrons, are some of its isoelectronic species.

(iv) The electron count of the Rb^{+} ion is 36 electrons. As a result, the species that isoelectronic with it will contain 36 electrons as well. Br^{-} ion $(35+1=36)$ electrons, Kr (36 electrons), and Sr^{2+} ion $(38-2=36)$ electrons) are some of its isoelectronic species.

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.**

Ans: (a) The number of electrons in each of the provided species (ions) is the same (10 electrons). As a result, the species in question are isoelectronic.

(b) As the magnitudes of nuclear charge fall, the ionic radii of isoelectronic species increases. The following is the order in which the supplied species are arranged in order of increasing nuclear charge.

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

As a result, the following is the order of the supplied species in terms 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?

Ans: A cation possesses fewer electrons than its parent atom while maintaining the same nuclear charge. As a result, a cation's nucleus attracts electrons more than its parent atom's nucleus. As a result, a cation is smaller than its parent atom in size. An anion, on the other hand, possesses one or more electrons than its parent atom, resulting in higher electron repulsion and a reduction in effective nuclear charge. As a result, anions have a greater distance between their valence electrons and the nucleus than their parent atom. As a result, the radius of an anion is bigger than that of 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?

Ans: The energy required to remove an electron from an isolated gaseous atom in its ground state is known as ionisation enthalpy. Despite the fact that the atoms are far separated in the gaseous state, there are some attraction forces between them. It is impossible to isolate a single atom to calculate the ionisation enthalpy. However, by lowering the pressure, the power of attraction can be reduced much more. As a result, in the definition of ionisation enthalpy, the term "isolated gaseous atom" is employed.

The most stable state of an atom is called the ground state. If an isolated gaseous atom is in its ground state, removing an electron from it will need less energy. As a result, ionisation enthalpy and electron gain enthalpy for a 'isolated gaseous atom' and its 'ground state' must be established for comparison reasons.

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** $J\ mol^{-1}$.

Ans: The energy of an electron in the hydrogen atom's ground state is 2.18×10^{-18} J. As a result, the energy required to remove that electron from the hydrogen atom's ground state is 2.18×10^{-18} J.

Thus, the ionization enthalpy of atomic hydrogen is 2.18×10^{-18} J.

As a result, the ionization enthalpy of atomic hydrogen in terms of J mol^{-1} can be calculated as,

$$\text{Ionization enthalpy} = 2.18\times10^{-18}\times6.02\times10^{23}\,\text{J mol}^{-1}$$
$$=1.31\times10^{6}\,\text{J mol}^{-1}$$

16. Among the second period elements the actual ionization enthalpies are in the order $\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$. **Explain why**
(i) Be **has higher ionization enthalpy than** B.
(ii) O **has lower ionization enthalpy than** N and F?

Ans: (i) The electron to be removed from a beryllium atom during the ionisation process is a $2s$ electron, whereas the electron to be removed from a boron atom is a $2p$-electron. $2s$-electrons are now stronger than $2p$-electrons in their attachment to the nucleus. As a result, removing a $2s$ electron from beryllium requires more energy than removing a $2p$ electron from boron.

(ii) The three $2p$-electrons of nitrogen occupy three separate atomic orbitals in nitrogen. In oxygen, however, two of the four $2p$-electrons share the same $2p$-orbital. Increased electron-electron repulsion occurs in the oxygen atom as a result of this. As a result, removing the fourth $2p$-electron from oxygen requires less energy than removing one of the three $2p$-electrons from nitrogen. As a result, oxygen has a lower ionization enthalpy than nitrogen. Fluorine has one extra electron and one more proton than oxygen. The increase in nuclear attraction (due to the addition of a proton) is greater than the rise in electronic repulsion as the electron is added to the same shell (due to the addition of an electron). As a result, the valence electrons in fluorine atoms have a greater effective nuclear charge than the electrons in oxygen atoms. As a result, removing an electron from a fluorine atom requires more energy than removing an electron from an oxygen atom. As a result, oxygen has a lower ionization enthalpy than fluorine.

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?

Ans: Sodium has a higher initial ionisation enthalpy than magnesium. This is mostly due to the following two factors:

(a) Sodium's atomic size is larger than that of magnesium.

(b) Magnesium has a higher effective nuclear charge than sodium.

As a result, the energy required to take an electron from magnesium is more than that necessary to remove an electron from sodium. As a result, sodium's first ionisation enthalpy is smaller than magnesium's. The second ionisation enthalpy of sodium, on the other hand, is larger than that of magnesium. Because sodium loses an electron, it achieves the stable noble gas structure. Magnesium, on the other hand, retains one electron in the 3s orbital after losing an electron. It still has to lose one more electron to achieve the stable noble gas structure. As a result, the energy required to remove the second electron in sodium is significantly more than that required in magnesium. As a result, sodium's second ionisation enthalpy is greater than magnesium's.

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

Ans: The following are the causes that cause the ionisation enthalpy of the primary group elements to fall along a group:

(i) Elements' atomic sizes grow larger: The number of shells grows as we proceed down a group. As a result, as you move down a group, the atomic size gradually increases. The valence electrons are not held as tightly when the distance between them and the nucleus grows. As a result, they can be simply removed. As a result, as you move down a group, the ionisation energy diminishes.

(ii) Shielding effect: As you move down a group, the number of electrons in their inner shells grows. As a result, the inner core electrons protect the valence electrons from the nucleus as they go along a group. As a result, the nucleus does not hold the valence electrons particularly tightly. As a result, the energy required to remove a valence electron reduces as the group number increases.

19. The first ionization enthalpy values $\left(\mathrm{kJ\ mol^{-1}}\right)$ of group 13 elements are :

Elements	**Ionization Enthalpy $\left(\mathrm{kJ\ mol^{-1}}\right)$**
Boron	**801**
Aluminum	**577**
Gallium	**579**
Indium	**558**
Thallium	**589**

How would you explain this deviation from the general trend?

Ans: The ionization enthalpy reduces with the atomic size and electron shielding lowers as you move down a group. Ionization enthalpy reduces from B to Al, as you move down group 13. Ga, on the other hand, has a larger ionization enthalpy than Al. Al comes after the s–block elements, while Ga comes after the d–block elements. d-electrons do not provide very good shielding. The valence electrons are not well shielded by these electrons. As a result, Ga valence electrons have a higher effective nuclear charge than Al valence electrons. Furthermore, when the atomic size and shielding drop from Ga to In, the ionization enthalpy reduces. However, as the ionization enthalpy increases from In to Tl, the ionization enthalpy decreases. Tl is found after the $4f$ and $5d$ electrons in the periodic table. The electron shielding offered by each of these orbitals is ineffective.

As a result, the nucleus tightly holds the valence electron. As a result, Tl's ionization energy is on the upper 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

Ans: The elements O and F are found in the same periodic table period. F atom has one more proton and one more electron than an O atom, and because an electron is added to the same shell, F atomic size is lower than O's. Because F has one extra proton than O its nucleus may attract incoming electrons more strongly than the nucleus of the O atom. Furthermore, F just requires one extra electron to achieve the stable noble gas structure. As a result, the electron gain enthalpy of F is lower than that of O.

(ii) F and Cl belong to the same periodic table group. Moving down a group, the electron gain enthalpy usually gets less negative. However, the electron gain enthalpy of Cl is less negative than that of F in this scenario. This is due to the fact that F atomic size is smaller than Cl. The electron will be added to quantum level $n=2$ in F but it will be added to quantum level $n=3$ in Cl. As a result, there are fewer electron-electron repulsions in Cl and an extra electron can be easily tolerated. As a result, Cl electron gain enthalpy is lower 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.

Ans: Energy is released when an electron is added to the O atom to form the O^- ion. As a result, O first electron gain enthalpy is negative.

$O + e^- \rightarrow O^-$

When an electron is added to an O^- ion to generate an O^{2-} ion, however, energy must be given out to overcome the strong electronic repulsions. As a result, O's second electron gain enthalpy is positive.

$O^- + e^- \rightarrow O^{2-}$

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

Ans: The tendency of an isolated gaseous atom to accept an electron is measured by electron gain enthalpy, whereas the tendency of an atom in a chemical compound to attract a shared pair of electrons is measured by electronegativity.

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

Ans: An element's electronegativity is a changeable property. It varies depending on the chemical. As a result, the statement that all nitrogen compounds have an electronegativity of 3.0 on the Pauling scale is erroneous. The electronegativity of N in NH_3 and NO_2 is different.

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

Ans: (a) An atom's size increases when it gains an electron. The number of electrons increases by one when one is added. The electrons' repulsion increases as a result of this. The number of protons, however, remains constant. As a result, the atom's effective nuclear charge drops while its radius increases.

(b) When an atom loses an electron, the number of electrons in the atom reduces by one, but the nuclear charge does not change. As a result, the atom's interelectronic repulsions decrease. The effective nuclear charge rises as a result. As a result, the atom's radius shrinks.

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

Ans: The number of electrons and protons (nuclear charge) in an atom determines its ionisation enthalpy. The protons and electrons in an element's isotopes are now the same. As a result, for two isotopes of the same element, the first ionisation enthalpy should be the same.

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

Ans:

	Metals	Non-metals
1.	Metals are prone to losing electrons.	Non-metals have a difficult time losing electrons.
2.	Metals have a difficult time gaining electrons.	Non-metals have an easy time gaining electrons.
3.	Ionic compounds are formed by metals in general.	Covalent compounds are formed by nonmetals in general.
4.	Metal oxides are inherently basic.	Acidity is a property of nonmetallic oxides.

5.	Ionization enthalpies of metals are low.	Ionization enthalpies of nonmetals are high.
6.	Metals have a lower electronegative charge. They are a group of elements that are electropositive.	Non–metals are electronegative in nature.

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.

Ans: (a) An element with 5 electrons in its outermost subshell should have the electrical configuration ns^2np^5. The electronic configuration of the halogen group is as follows. The element can therefore be F, Cl, Br or I.

(b) To achieve the stable noble gas configuration, an element with two valence electrons can simply lose two electrons. The element's overall electrical configuration will be ns^2. The electrical configuration of group 2 components is shown here. Be, Mg, Ca, Sr and Ba are the elements found in group 2.

(c) If an element only requires two electrons to achieve the stable noble gas state, it is likely to gain two electrons. As a result, such an element's general electrical configuration should be ns^2np^4. The electrical configuration of the oxygen family is as follows.

(d) At room temperature, Group 17 contains metals, nonmetals, liquids, and gases.

28. The increasing order of reactivity among group 1 elements is $\mathrm{Li} < \mathrm{Na} < \mathrm{K} < \mathrm{Rb} < \mathrm{Cs}$ whereas that among group 17 elements is $\mathrm{F} > \mathrm{Cl} > \mathrm{Br} > \mathrm{I}$. Explain.

Ans: Group 1 elements having only one valence electron, which they frequently lose. The noble gas structure, on the other hand, requires only one electron for Group 17 elements. The ionization enthalpies drop as you move along group 1. This reduces the energy required to remove the valence electron. As a result, on moving down a group, the reactivity increases.

$\mathrm{Li} < \mathrm{Na} < \mathrm{K} < \mathrm{Rb} < \mathrm{Cs}$

As a result, the following is the increasing order of reactivity among group 1 elements:

The electron gain enthalpy decreases as we proceed down the group from Cl to I in group 17. Its tendency to gain electrons reduces as it moves along group 17. As a result, responsiveness reduces the size of a group. F has a lower negative electron gain enthalpy than Cl. Even yet, it is the most reactive of the halogens. This is due to the fact that it has a low bond dissociation energy. As a result, the following is the decreasing order of reactivity among group 17 elements: $\mathrm{F} > \mathrm{Cl} > \mathrm{Br} > \mathrm{I}$.

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

Ans: s block element: ns^{1-2}, where, $n = 2-7$

p block element: ns^2np^{1-6}, where, $n = 2-6$

d block element: $(n-1)d^{1-10}ns^{0-2}$, where, $n = 4-7$

f block element: $(n-2)f^{1-14}(n-1)d^{1-10}ns^2$, where, $n = 6-7$

30. Assign the position of the element having outer electronic configuration

(i) ns^2np^4, for $n=3$

(ii) $(n-1)d^2ns^2$, for $n=4$

(iii) $(n-2)f^7(n-1)d^1ns^2$, for $n=6$ **in the periodic table.**

Ans: (i) The element belongs to the 3rd period since $n=3$. Because the last electron occupies the p orbital, it is a p–block element. In the p–orbital, there are four electrons. As a result, the element's matching group.

$$\begin{aligned}\text{corresponding group of element} &= \text{number of elements in}\left(d-\text{block}+p-\text{block}+s-\text{block}\right)\\ &=10+4+2\\ &=16\end{aligned}$$

As a result, the element belongs to the periodic table's 3rd period and 16th group. As a result, Sulfur is the element.

(ii) The element belongs to the 4th period since $n=4$. Because the d– orbitals are not completely filled, it is a d–block element. In the d–orbital, there are two electrons. As a result, the element's matching group.

$$\begin{aligned}\text{corresponding group of element} &= \text{number of elements in}\left(d-\text{block}+s-\text{block}\right)\\ &=2+2\\ &=4\end{aligned}$$

As a result, it belongs to the fourth period and fourth group. As a result, the element's name is Titanium.

(iii) The element is present in the 6th period because $n=6$. The last electron fills the f–orbital, making it a f–block element. It belongs to group 3 of the periodic table, which includes all f-block elements. The electronic configuration is as follows: $[\text{Xe}]4f^7 5d^1 6s^2$. As a result,

$$\begin{aligned}\text{atomic number} &= 54+7+2+1\\ &=64\end{aligned}$$

As a result, the element's name is Gadolinium.

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

Elements	$(\Delta_i H_1)$	$(\Delta_i H_2)$	$(\Delta_{eg} H)$
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)?

Ans: (a) The least reactive element is likely to be element V. This is due to the fact that it has the highest initial ionization enthalpy $(\Delta_i H_1)$ and the highest positive electron gain enthalpy $\left(\Delta_{eg} H\right)$.

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

(c) Element III, with a high first ionization enthalpy $(\Delta_i H_1)$ and the largest negative electron gain enthalpy $\left(\Delta_{eg} H\right)$. is anticipated to be the most reactive nonmetal.

(d) Because it has a very high first ionization enthalpy $(\Delta_i H_2)$ and a positive electron gain enthalpy $\left(\Delta_{eg} H\right)$. element V is predicted to be the least reactive nonmetal.

(e) The negative electron gain enthalpy of element VI is low $\left(\Delta_{eg} H\right)$. As a result, it is a metal. It also has the smallest second ionization enthalpy, $(\Delta_i H_2)$. As a result, a stable binary halide with the formula $\mathrm{MX_2}$ (X =halogen) can be formed.

(f) The first ionization energy of element I is low, whereas the second ionization energy is large. As a result, a primarily stable covalent halide with the formula MX (X =halogen) can be formed.

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

Ans : (a) $\mathrm{Li_2O}$

(b) $\mathrm{Mg_3N_2}$

(c) $\mathrm{AlI_3}$

(d) $\mathrm{SiO_2}$

(e) PF_3 or PF_5

(f) Lutetium is an element with the atomic number 71. (Lu). It has a valency of three. As a result, the compound's formula 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

Ans: In the Modern periodic table, the value of the primary quantum number (n) for the outermost shell or the valence shell denotes a period.

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 d subshell.
(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.

Ans: Because a maximum of 10 electrons can fill all of the orbitals in a d subshell, the d-block contains ten columns.

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

(b) Nuclear charge (Z)

(c) Nuclear mass

(d) Number of core electrons.

Ans: The valence electrons are unaffected by nuclear mass.

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.

Ans: 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:
$F^- < Ne < Na^+$
Z 9 10 11
Therefore, the order of the increasing size of F^-, Ne and Na^+ is as follows:
$Na^+ < Ne < F^-$

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.

Ans: The nucleus is more attracted to electrons in orbitals with a lower n value than electrons in orbitals with a higher n value. As a result, removing electrons from orbitals with a larger n value is easier than removing electrons from orbitals with a smaller 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) $K > Mg > Al > B$

Ans: Over time, the metallic character of components reduces from left to right. As a result, Mg has a more metallic quality than Al. Elements get more metallic as they progress through the group. As a result, Al has a more metallic character than B. $K > Mg$ is the result of the foregoing statements.
As a result, the right metallic character order is $K > Mg > Al > B$

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) $F > N > C > Si > B$

Ans: Over time, the non-metallic nature of elements grows from left to right. $F > N > C > B$ is the decreasing order of non-metallic nature.
The non-metallic nature of elements reduces with each successive group. As a result, the non-metallic characteristics of C and Si are in decreasing order: $C > Si$. Si, on the other hand, is less non-metallic than B, hence $B > Si$.

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) $F > O > Cl > N$

(c) $Cl > F > O > N$

(d) $O > F > N > Cl$

Ans: Over time, the oxidizing property of elements rises from left to right. As a result, we have $F > O > N$ as the decreasing order of oxidizing properties.
The ability of elements to oxidize reduces as they progress through a group. As a result, $F > Cl$. However, O has a higher oxidizing character than Cl, i.e., $O > Cl$. As a result, in terms of their oxidizing property, the right sequence of chemical reactivity of F, Cl, O and N, is $F > O > Cl > N$.