Chapter: Periodicity in Properties and Classification of Elements

Exercise

Question 3.1: What is the basic theme of organisation in the periodic table?

Answer:

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

Question 3.2: Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?

Answer:

Mendeleev arranged the elements in his periodic table in the order of their atomic weight or mass. Mendeleev arranged the elements in periods and groups in order of their increasing atomic weight. He placed the elements having similar properties in the same group. But he did not stick to this arrangement for long. He noticed that if the elements were arranged in order of their increasing atomic weights, then some elements did not fit within this scheme of classification. Thus, he ignored the order of atomic weights in some cases. For example, the atomic weight of iodine is less than that of tellurium. But Mendeleev placed tellurium (in Group VI) before iodine (in Group VII) just because iodine's properties are very similar to fluorine, chlorine, and bromine.

Question 3.3: What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?

Answer:

Mendeleev's Periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic weights. And the Modern periodic Law states that the physical and chemical properties of elements are periodic functions of their atomic numbers.

Question 3.4: On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Answer:

A period in the periodic table of the elements indicates the value of the principal quantum number (n) for the outermost shells. Each period begins with the filling of principal quantum number (n).

The value of n for the sixth period is 6. For $n = 6$, azimuthal quantum number (l) will have values $0, 1, 2, 3, 4$.

According to Aufbau's principle, electrons are usually added to different orbitals in order of their increasing energies. The energy of the $6d$ subshell is higher than that of the $7s$ subshell.

In the $6th$ period, electrons can be filled in only $6s$, $4f$, $5d$, and $6p$ subshells. $6s$ has one orbital, $4f$ has seven orbitals, $5d$ has five orbitals, and $6p$ has three orbitals. Thus, there are sixteen $(1+7+5+3=16)$ orbitals available. According to Pauli's exclusion principle, each orbital can accommodate a maximum of 2 electrons.

So, 16 orbitals can accommodate a maximum of 32 electrons.

Therefore, the sixth period of the periodic table should have 32 elements.

Question 3.5: In terms of period and group where would you locate the element with Z=114?

Answer:

Elements with atomic numbers from $Z=87$ to $Z=114$ are present in the 7^{th} period of the periodic table. So, 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. Thus, the element with $Z=114$ is the second $p-$block element in the 7^{th} period. Therefore, the element with $Z=114$ is present in the 7^{th} period and 4^{th} group of the periodic table.

Question 3.6: Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Answer:

There are two elements in the 1^{st} period and eight elements in the 2^{nd} period. The third period starts with the element with $Z=11$. There are eight elements in the third period. So, the 3^{rd} period ends with the element with $Z=18$ i.e., the element in the 18^{th} group of the third period has $Z=18$. Therefore, the element in the 17^{th} group of the third period has atomic number $Z=17$.

Question 3.7: Which element do you think would have been named by

(i) Lawrence Berkeley Laboratory

(ii) Seaborg's group?

Answer:

(i) Lawrencium (Lr) with $Z=103$ and Berkelium (Bk) with $Z=97$

(ii)Seaborgium (Sg) with $Z=106$

Question 3.8: Why do elements in the same group have similar physical and chemical properties?

Answer:

The physical and chemical properties of elements usually depend on the number of valence electrons. Elements present in the same group have the same number of electrons in their outermost shell. Thus, elements present in the same group have similar physical and chemical properties.

Question 3.9: What does atomic radius and ionic radius really mean to you?

Answer:

Atomic radius is the radius of an atom which measures the size of an atom. If the element is a metal, then the atomic radius is the metallic radius. If the element is a non-metal, then it is the covalent radius. Metallic radius is calculated as half the internuclear distance in the metallic crystal. For example, the internuclear distance between two adjacent copper atoms in solid copper is $256\ pm$.

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

Covalent radius is 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 $198\ pm$. So, the covalent radius of chlorine is $\frac{198}{2}pm = 99\,pm$.

Ionic radius is the radius of an ion. The ionic radii can be found by measuring the distances between the cations and anions in ionic crystals. A cation is formed by removing an electron from an atom, so it has fewer electrons than the parent atom resulting in an increase in the effective nuclear charge.

So, a cation is smaller than the parent atom. For example, the ionic radius of Na^+ ion is $95\ pm$, whereas the atomic radius of sodium atom is $186\ pm$. An anion is larger in size than its parent atom because it 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 fluorine atom is $64\ pm$.

Question 3.10: How does atomic radius vary in a period and in a group? How do you explain the variation?

Answer:

Atomic radius usually decreases from left to right across a period because within a period, the outer electrons are present in the same valence shell and the atomic number increases from left to right across a period, which results in an increased effective nuclear charge. So, the attraction of electrons to the nucleus increases.

The atomic radius usually increases down a group because down a group, the principal quantum number increases which results in an increase of the distance between the nucleus and valence electrons.

Question 3.11: 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^{+}

Answer:

Atoms and ions which have the same number of electrons are known as isoelectronic species.

(i) F^{-} ion has $9+1=10$ electrons. Therefore, 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. Therefore, 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. Therefore, 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. Therefore, 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).

Question 3.12: 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.

Answer:

(a)Each of the given species has the same number of electrons (10 electrons). Thus, the given species are isoelectronic.

(b)The ionic radii of isoelectronic species usually increases with a decrease in the magnitudes of nuclear charge.

The arrangement of the given species in order of their increasing nuclear charge is given below:

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

Nuclear charge $= +7 + 8 + 9 + 11 + 12 + 13$

Thus, the arrangement of the given species in order of their increasing ionic radii is:

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

Question 3.13: Explain why cations are smaller and anions larger in radii than their parent atoms?

Answer:

A cation has less number of electrons than its parent atom and its nuclear charge remains the same. So, the attraction of electrons to the nucleus is more in a cation than in its parent atom. Thus, a cation is usually smaller in size than its parent atom. An anion has one or more electrons than its parent atom, which results in an increased repulsion among the electrons and a decrease in the effective nuclear charge. So, the distance between the valence electrons and the nucleus is more in anions than in it's the parent atom. Therefore, an anion is larger in radius than its parent atom.

Question 3.14: What is the significance of the terms - 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy?

Answer:

The energy which is required to remove an electron from an isolated gaseous atom in its ground state is known as ionization enthalpy. Even though the atoms are widely separated in the gaseous state, some amounts of attractive forces are present among the atoms. To find the ionization enthalpy, it is not possible to isolate a single atom. But, the force of attraction can be further reduced by lowering the pressure. So, the term 'isolated gaseous atom' is used in the definition of ionization enthalpy.

Ground state of an atom means the most stable state of an atom. If an isolated gaseous atom is in its ground state, then less amount energy is required to remove an electron from it. Thus, for comparison, ionization enthalpy and electron gain enthalpy should be determined for an 'isolated gaseous atom' and its 'ground state'.

Question 3.15: 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}.

Answer:

The energy of an electron in the ground state of the hydrogen atom is $-2.18\times10^{-18}\ J$. Thus, the energy needed to remove that electron from the ground state of hydrogen atom is $2.18\times10^{-18}\ J$.

$\therefore$ Ionization enthalpy of atomic hydrogen $= 2.18\times10^{-18}\ J$

Thus, ionization enthalpy of atomic hydrogen in terms of $J\ mol^{-1}$

$= 2.18\times10^{-18}\times6.02\times10^{23}\ J\ mol^{-1}$

$= 1.31\times10^{6}\ J\ mol^{-1}$

Question 3.16: 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 **?**

Answer:

(i) During ionization, the electron to be removed from beryllium atom is a $2s-$electron, whereas the electron that is to be removed from boron atom is a $2p-$electron. $2s-$electrons are more strongly attached to the nucleus than $2p-$electrons. Thus, more energy is required to remove a $2s-$electron of beryllium than that required to remove a $2p-$electron of boron. Therefore, beryllium has higher $\Delta_i H$ than boron.

(ii)In nitrogen, the three $2p-$electrons of nitrogen occupy 3 different atomic orbitals. In oxygen, two of the four $2p-$electrons of oxygen occupy the same $2p-$orbital. This causes increased electron-electron repulsion in oxygen atom. So, the energy required to remove the fourth $2p-$electron from oxygen is less as compared to the energy required to remove one of the three $2p-$electrons from nitrogen. Thus, oxygen has lower $\Delta_i H$ than nitrogen.

Fluorine contains one electron and one proton more than oxygen. As the electron is added to the same shell, the increase in nuclear attraction is more than the increase in electronic repulsion. Thus, the valence electrons in fluorine atom face a more effective nuclear charge than that experienced by the electrons present in oxygen. So, more energy is required to remove an electron from fluorine atom than that required to remove an electron from oxygen atom. Thus, oxygen has lower $\Delta_i H$ than fluorine.

Question 3.17: 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?

Answer:

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

a. The atomic size of sodium is greater than that of magnesium

b. The effective nuclear charge of magnesium is higher than sodium

Due to these reasons, the energy needed to remove an electron from magnesium is greater than the energy required in sodium. Thus, the first ionization enthalpy of sodium is lower than that of magnesium.

But, the second ionization enthalpy of sodium is higher than that of magnesium because after losing an electron, sodium gets the stable noble gas configuration. Magnesium, after losing an electron still has one electron in the $3s-$orbital. To attain the stable noble gas configuration, it has to lose one more electron. So, the energy required to remove the second electron of sodium is much higher than that of magnesium. Thus, the second ionization enthalpy of sodium is higher than that of magnesium.

Question 3.18: What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?

Answer:

The factors that are responsible for the ionization enthalpy of the main group elements to decrease down a group are given below:

(i) Increase in the atomic size of elements: As we move down a group, the number of shells usually increases. So, the atomic size also increases gradually on moving down a group. As the distance of the valence electrons from the nucleus increases, the electrons are not held strongly. So, they can be removed easily. Thus, on moving down a group, ionization energy decreases.

(ii) Increase in the shielding effect: The number of inner shells of electrons usually increases on moving down a group. Thus, the shielding of the valence electrons from the nucleus by the inner core electrons increases down a group. So, the valence electrons are not held very tightly by the nucleus. Therefore, the energy required to remove a valence electron decreases down a group.

Question 3.19: 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?

Answer:

If we move down a group, ionization enthalpy usually decreases due to an increase in the atomic size and shielding. So, 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 don't shield the valence electrons very effectively. So, the valence electrons of Ga experience a greater effective nuclear charge than those of Al. Moving from Ga to In, the ionization enthalpy decreases because of the increase in the atomic size and shielding. But, on moving from In to Tl, the ionization enthalpy 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. Thus, the valence electron is held strongly by the nucleus. Therefore, the ionization energy of Tl is on the higher side.

Question 3.20: Which of the following pairs of elements would have a more negative electron gain enthalpy?

(i) O **or** F **(ii)** F **or** Cl

Answer:

(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 added to the same shell, the atomic size of F is smaller than the atomic size of O. As F contains one proton more than O, its nucleus can attract the incoming electron more strongly compared to the nucleus of O atom. F needs only one more electron to attain the stable noble gas configuration. Thus, 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 generally becomes less negative on moving down a group. But in this case, the value of the electron gain enthalpy of Cl is more negative than that of F as 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. Thus, there are less electron-electron repulsions in Cl and an additional electron can be accommodated easily. Therefore, the electron gain enthalpy of Cl is more negative than that of F.

Question 3.21: Would you expect the second electron gain enthalpy of O **as positive, more negative or less negative than the first? Justify your answer.**

Answer:

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

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

When an electron is added to O^- ion to form O^{2-} ion, energy should be given out to overcome the strong electronic repulsions. So, the second electron gain enthalpy of O is positive.

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

Question 3.22: What is the basic difference between the terms electron gain enthalpy and electronegativity?

Answer:

The measure of the tendency of an isolated gaseous atom to accept an electron is known as electron gain enthalpy is whereas electronegativity is the measure of the tendency of an atom in a chemical compound to attract a shared pair of electrons.

Question 3.23: How would you react to the statement that the electronegativity of N **on Pauling scale is** 3.0 **in all the nitrogen compounds?**

Answer:

Electronegativity of an element is usually a variable property and is different in different compounds. So, the statement "the electronegativity of N on Pauling scale is 3.0 in all nitrogen compounds" is not correct. The electronegativity of N is different in NH_3 and NO_2.

Question 3.24: Describe the theory associated with the radius of an atom as it

(a) gains an electron

(b) loses an electron

Answer:

(a) When an atom gains an electron, its size will increase. When an electron is added, the number of electrons increase by one which results in an increase in repulsion among the electrons. But the number of protons remains the same. So, 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. Thus, the interelectronic repulsions in the atom decrease. So, the effective nuclear charge increases. Therefore, the radius of the atom decreases.

Question 3.25: Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

Answer:

The ionization enthalpy of an atom usually depends on the number of electrons and protons of that atom. The isotopes of an element have the equal protons and electrons. Thus, the first ionization enthalpy for two isotopes of the same element should be the same.

Question 3.26: What are the major differences between metals and non-metals?

Answer:

	Metals	**Non–metals**
1.	Metals can lose electrons easily.	Non-metals cannot lose electrons easily.
2.	They cannot gain electrons easily.	They can gain electrons easily.
3.	They form ionic compounds.	They form covalent compounds.
4.	Their oxides are basic in nature.	Their oxides are acidic in nature.
5.	They have low ionization enthalpies.	They have high ionization enthalpies.
6.	They have less negative electron gain enthalpies.	They have high negative electron gain enthalpies.
7.	They are less electronegative. They are rather electropositive elements.	They are electronegative.
8.	They have a high reducing power.	They have a low reducing power.

Question 3.27: 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.

Answer:

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

(b) An element having two valence electrons will lose two electrons easily to achieve the stable noble gas configuration. The general electronic configuration 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 2 electrons if it needs only 2 electrons to achieve the stable noble gas configuration. So, the general electronic configuration of such an element should be ns^2np^4. This is the electronic configuration of the oxygen family.

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

Question 3.28: 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.

Answer:

The elements of group 1 have only 1 valence electron, which they tend to lose. Group 17 elements need only one electron to achieve the noble gas configuration. On moving down group 1, the ionization enthalpies decrease i.e. the energy required to lose the valence electron decreases. So, reactivity increases on moving down a group. So, the increasing order of reactivity among group 1 elements is:

$Li < Na < K < Rb < Cs$

In group 17, as we move down the group from chlorine to iodine, the electron gain enthalpy becomes less negative i.e. its tendency to gain electrons decreases down the group 17. So, reactivity decreases down a group. The electron gain enthalpy of fluorine is less negative than chlorine. But still, it is the most reactive halogen because of its low bond dissociation energy. Hence, the decreasing order of reactivity among group 17 elements is: $F > CI > Br > I$

Question 3.29: Write the general outer electronic configuration of $s-$, $p-$, $d-$ and $f-$ block elements.

Answer:

Element	General outer electronic configuration

$s-$block	ns^{1-2} , where $n=2-7$
$p-$block	ns^2np^{1-6} , where $n=2-6$
$d-$block	$(n-1)d^{1-10}ns^{0-2}$, where $n=4-7$
$f-$block	$(n-2)f^{1-14}(n-1)d^{0-10}ns^2$, where $n=6-7$

Question 3.30: Assign the position of the element having outer electronic configuration

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

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

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

Answer:

(i) Since $n=3$, the element belongs to the 3^{rd} period. It is a $p-$block element as the last electron occupies the $p-$orbital.

There are four electrons in the $p-$orbital. So, the corresponding group of the element

$=$Number of $s-$block groups $+$ number of $d-$block groups $+$ number of $p-$electrons

$=2+10+4$

$=16$

Thus, the element belongs to the 3^{rd} period and 16^{th} group of the periodic table. So, the element is Sulphur.

(ii)Since $n=4$, the element belongs to the 4^{th} period. It is a $d-$block element as $d-$orbitals are incompletely filled.

There are two electrons in the $d-$orbital.

So, the corresponding group of the element

$=$Number of $s-$block groups $+$ number of $d-$block groups

$=2+2$

$=4$

Thus, it is a 4^{th} period and 4^{th} group element. So, the element is Titanium.

(iii) Since $n=6$, the element is present in the 6^{th} period. It is an $f-$block element as the last electron occupies the $f-$orbital. It belongs to group 3 of the periodic table as all $f-$block elements belong to group 3. Its electronic configuration is $[Xe]4f^7 5d^1 6s^2$. So, its atomic number is $54+7+2+1=64$. Therefore, the element is Gadolinium.

Question 3.31: The first $(\Delta_i H1)$ and the second $(\Delta_i H)$ ionization enthalpies $\left(in\ kJ\ mol^{-1}\right)$ and the $(\Delta_{eg} H)$ electron gain enthalpy $\left(in\ kJ\ mol^{-1}\right)$ of a few elements are given below:

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

Answer:

(a) Element V is most likely to be the least reactive element because it has the highest first ionization enthalpy $(\Delta_i H1)$ and a positive electron gain enthalpy $(\Delta_{eg} H)$.

(b) Element II is most likely to be the most reactive metal because it has the lowest first ionization enthalpy $(\Delta_i H1)$ and a low negative electron gain enthalpy $(\Delta_{eg} H)$.

(c) Element III is most likely to be the most reactive non–metal because it has a high first ionization enthalpy $(\Delta_i H1)$ and the highest negative electron gain enthalpy $(\Delta_{eg} H)$.

(d) Element V is most likely to be the least reactive non–metal because it has a very high first ionization enthalpy $(\Delta_i H2)$ and a positive electron gain enthalpy $(\Delta_{eg} H)$.

(e)Element VI has a low negative electron gain enthalpy $(\Delta_{eg} H)$. So, it is a metal. And it has the lowest second ionization enthalpy $(\Delta_i H2)$. Thus, 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. Thus, it can form a stable covalent halide of the formula MX (X = halogen).

Question 3.32: 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**

Answer:

(a) LiO_2

(b) Mg_3N_2

(c) AlI_3

(d) SiO_2

(e) PF_3 or PF_5

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

Question 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.

Answer:

The value of the principal quantum number for the outermost shell or the valence shell indicates a period in the modern periodic table.

Question 3.34: 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.

Answer:

The $d-$block has 10 columns as a maximum of 10 electrons can occupy all the orbitals in a $d-$ subshell.

Question 3.35: 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.

Answer:

Nuclear mass does not affect the valence electrons.

Question 3.36: 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.

Answer:

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:

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

Z 9 10 11

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

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

Question 3.37: 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.

Answer:

Electrons in orbitals having a lower n value are more attracted to the nucleus than electrons in orbitals bearing a higher n value. So, 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.

Question 3.38: 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)** $K > Mg > Al > B$

Answer:

The metallic character of elements decreases from left to right across a period. So, the metallic character of Mg is more than that of Al.

The metallic character of elements increases down a group. So, the metallic character of Al is more than that of B.

So, we get $K > Mg$.

So, the correct order of metallic character is $K > Mg > Al > B$.

Question 3.39: 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

Answer:

The non-metallic character of elements increases from left to right across a period. So, the decreasing order of non-metallic character is $F > N$>C> B.

The non-metallic character of elements decreases down a group. So, the decreasing order of non-metallic characters of C and Si are $C > Si$. Si is less non-metallic than B i.e. $B > Si$.

Thus, the correct order of their non-metallic characters is $F > N$>C> B>Si.

Question 3.40: 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)** $O > F > N > Cl$

Answer:

The oxidizing character of elements increases from left to right across a period. So, the decreasing order of oxidizing property is $F > O > N$.

The oxidizing character of elements decreases down a group. So, we get $F > Cl$.

The oxidizing character of O is more than that of Cl i.e. $O > Cl$.

So, the correct order of chemical reactivity of F, Cl, O and N in terms of their oxidizing property is $F > O > Cl > N$.