Chapter – 3

Classification of Elements and Periodicity in Properties

2marks:

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

Answer:

It is to characterize the elements in periods and groups as per their properties. So, this course of action makes the investigation of elements and compounds of elements in a simple one and methodical way. In this periodic table, elements with comparative properties are set in a similar group.

2. What is the atomic number of element keeping in mind both the cases given below;

1. Element is in 3rd period of periodic table.

2. Element is in 17th group of periodic table.

Answer:

First period is having 2 elements and second period is having 8 elements. So, the third period begins with element Z = 11. Presently, third period contains 8 elements. So, 18th element is the last element of the third period and this 18th element is present in 18th group. Thus, the element in the seventeenth group of the 3rd period is having atomic number 17 i.e. Z = 17.

3.Which element are named by

a) Seaborg's group

b) Lawrence Berkeley Laboratory?

Answer:

a) Seaborgium (Sg) which has atomic number, Z = 106

b) Lawrencium (Lr) which has atomic number, Z = 103 and Berkelium (Bk) which has atomic number, Z = 97

4. Elements present in the same group are having similar chemical and physical properties. Why is it so?

Answer:

The chemical and physical properties of any elements rely on the quantity of valence electrons. In periodic table elements are in same group are having same quantity of valence electrons. This is why elements present in the same group are having similar chemical and physical properties.

5. State difference between Mendeleev's Approach for periodic law and the Modern approach for the periodic law.

Answer:

Mendeleev's Approach for periodic law	Modern approach for the periodic law
Chemical properties and Physical properties of the elements are the	Chemical properties and Physical properties of the elements are the

periodic functions of the atomic mass of the corresponding elements.	periodic functions of the atomic numbers of the corresponding elements.

6. Consider the accompanying species: N^{3-},O^{2-},F^{-},Na^{+}, Mg^{2+}, and Al^{3+}
(i) What is similar in them?
(ii) Arrange them in the according to their increasing order of ionic radii.

Answer:
The species that are given are having equal number of electrons i.e. 10 electrons. So, they are isoelectronic species.
Arrangement of the given ions according to their increasing order of nuclear charge is:
$N^{3-} < O^{2-} < F^{-} < Na^{+} < Mg^{2+} < Al^{3+}$
Arrangement of the given ions according to their increasing order of ionic radii is:
$Al^{3+} < Mg^{2+} < Na^{+} < F- < O^{2-} < N^{3-}$

7.Determine ionization enthalpy of Hydrogen atom in $Jmol^{-1}$.

Electron of hydrogen is having $-2.18*10^{-18}$J in ground state.

Answer:

Here it is given that, electron of hydrogen is having $-2.18*10^{-18}$J in ground state.

Thus, −2.18∗10−18J amount of energy will be required to expel an electron from ground state in H(hydrogen) – atom.

∴ For Hydrogen atom the Ionization of enthalpy = $-2.18*10^{-18}$J

Thus, ionization enthalpy of Hydrogen atom in $Jmol^{-1}$

$= -2.18 \times 10^{-18} \times 6.02 \times 10^{23} Jmol^{-1}$

8. What is electron affinity of O(oxygen) atoms?

1.Positive

2.More negative

3.Less negative

Justify the answer.

Answer:

O^- ion is formed when O-atom gains one electron and the energy is being released during this process. So the 1st electron affinity for O-atom is negative.

$O(g) + e^- \rightarrow O^-g$

If an electron is added in O^- ion then it forms O^{2-} ion, the energy is required to ne given to counter the strong electronic repulsions. So, the 2nd electron affinity of O-atom is positive.

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

9. State the difference between the terms electron affinity and electronegativity.

Answer:

Electron gain enthalpy	**Electronegativity**
Tendency to gain electrons for an isolated gaseous atom is its electron gain enthalpy.	Tendency to attract the shared pairs of electrons for an atom which is in chemical compound is its electronegativity.

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

Answer:

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

11.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 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.

12.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 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.

13. 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

14. 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) Li_2O

(b) Mg_3N_2

(c) AlI_3

(d) SiO_2

(e) PF_3 or PF_5

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

15. 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 (*n*) for the outermost shell or the valence shell indicates a period in the Modern periodic table.

16. 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 because a maximum of 10 electrons can occupy all the orbitals in a *d* subshell.

17. 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

4marks:

1.Which vital property did Mendeleev use to order the elements in the periodic table that he designed and did he adhere to that?

Answer:

Mendeleev organised the components in his periodic table, according to the order of their atomic weight. Mendeleev organized the components in groups and periods according to the increasing atomic weight. Mendeleev set the elements which are having comparative properties in the similar groups.

Nonetheless, he didn't adhere to arrangement that he gave for long. He discovered that if the elements were organized according to their increasing atomic weights, then a few elements did not match within this plan of characterization.

In this manner, he overlooked the order of atomic weights now and again. For instance, the atomic mass of iodine is lower than atomic mass of tellurium.

Still Mendeleev set tellurium (in Group 6) ahead of iodine (in Group 7) essentially in light of the fact that iodine's properties are so comparable to fluorine, chlorine, and bromine

2. On the premise of the quantum numbers, verify that the 6th period of periodic table ought to have 32 components.

Answer:

In a periodic table containing elements, a period shows the value of principal quantum number (n) for the furthest shells. Every period

starts with the filling with the principal quantum number (n). And n's value for the 6th period is equal to 6. Now, for n = 6, the azimuthal quantum number (l) can have "0, 1, 2, 3, 4" values.
As indicated by Aufbau's rule, electrons will be added to various orbitals according to their increasing energies. Here, the 6d subshell is having much higher energy than the energy of 7s subshell.
In the sixth period, the electrons can occupy in just 6s, 4f, 5d, and 6 p subshells. 6s is having 1 orbital, 4f is having 7 orbitals, 5d is having 5 orbitals, and 6p is having 3 orbitals. In this way, there are a sum of 16 (1 + 7 + 5 + 3 = 16) orbitals accessible. As indicated by Pauli's exclusion, one orbital can only accommodate at max 2 electrons. Hence, sixteen orbitals can have 32 electrons.
Subsequently, the 6th period of period table ought to have 32 elements.

3. In groups and periods of periodic table where will you find the elements which is having Z =114?

Answer:

Elements whose atomic number is from Z = 87 to Z = 114 are available in the seventh period of periodic table. Therefore, the element having Z = 114 is available in the seventh period in periodic table.
In the seventh period, initial 2 elements with Z = 87 and Z= 88 are the elements of s-block and the following 14 elements except Z = 89 i.e., those from Z = 90 to Z = 103 are elements of f – block, and next 10

elements from Z = 89 and Z = 104 to Z = 112 are elements of d-block, next the elements from Z = 113 to Z = 118 are elements of p-block. In this manner, the element Z = 114 is the 2nd element of p-block in the seventh period of the periodic table.

Therefore, the element Z = 114 is available in the seventh period and fourth group in the periodic table.

4. Explain why there is variation in atomic radius in a group and period?

Answer:

Atomic radius declines as we move from left to right in a period. It happens because in a period, the external electrons are available in a similar valence shell so, the atomic number increments from left to right in a period, which results in increase in the effective nuclear charge. Therefore, the attraction of electrons towards the nucleus is increased.

Also, atomic radius declines as we move from top to bottom in the group. It happens because as we move down in a group then there is increase in principal quantum number(n) which brings about increase in the distance between nucleus and the valence electrons.

5. Cation are having smaller radii then that of their parent atom and anion are having larger radii than their parent atom. Why?

Answer:

Cations are formed by expelling an electron from outermost orbit of an atom, thus cation has less electrons compared to parent atom which results in increased effective nuclear charge but the total nuclear charge remains same which results in increased attraction of electrons towards nucleus than that of parent atom. Thus, cations are having smaller radii then that of their parent atom.

Anions are formed by gaining an electron in the outermost orbit of an atom, thus anion has more electrons compared to parent atom which results in decreased effective nuclear charge but the total nuclear charge remains same which results in increased distance the nucleus and the valence electrons as the attraction of electrons towards nucleus decreases than that of parent atom. Thus, anions are having larger radii then that of their parent atom.

6. State significance of following terms:

1. “isolated gaseous atom”

2. “ground state”

in the definition of ionization enthalpy and electron gain enthalpy?

Answer:

“Ionization enthalpy is the energy that is required to expel an electron from an isolated gaseous atom in ground state”. Despite the fact that in gaseous state the atoms are generally widely separated, there are a few measures of attractive forces between the atoms. To find the ionization enthalpy of any ion, it is difficult to isolate a solitary atom. This attractive force can be further diminished by bringing down the pressure. Hence, the term “isolated gaseous atom” is utilized as a part of the meaning of ionization enthalpy.

An atom’s ground state is the most stable state. Less energy is required to expel an electron if isolated gaseous atom is present in the ground state. In this way, for the purpose of comparison, electron gain enthalpy and ionization enthalpy must be calculated for an “isolated gaseous atom” and its “ground stat.

7. Explain why the 1^{st} ionization enthalpy of magnesium is higher than 1st ionization enthalpy of sodium but the 2nd ionization enthalpy of magnesium is lower than 2^{nd} ionization enthalpy of sodium?

Answer:

The 1^{st} ionization enthalpy of magnesium is higher than 1^{st} ionization enthalpy of sodium because,

1. Magnesium is having greater atomic size than sodium.
2. Magnesium is having higher effective nuclear charge than sodium.

Thus, energy required to expel an electron from sodium is lower than that in magnesium. Thus, the 1^{st} ionization enthalpy of magnesium is higher than 1^{st} ionization enthalpy of sodium.

The 2^{nd} ionization enthalpy of magnesium is lower than 2^{nd} ionization enthalpy of sodium is because after expelling an electron, there is still 1 electron remaining in the 3s-orbital of magnesium, whereas sodium achieves stable inert gas configuration after expelling an electron. So, magnesium still requires to expel 1 electron to achieve stable inert gas configuration.

Thus, energy required to expel 2nd electron from magnesium is lower than that in sodium. Thus, the 2nd ionization enthalpy of magnesium is lower than 2nd ionization enthalpy of sodium.

8. State the factors because of which in elements of main group the ionization enthalpy decreases when we move down the group.

Answer:

The factors because of which in elements of main group the ionization enthalpy decreases when we move down the group are given below:

1. "Increase in the shielding effect": Inner shells increases as we move down the group. Thus, the shielding effect of valence electrons increases by inner core electrons from nucleus. Thus, the attractive force on electrons towards nucleus is very strong. So, energy required to expel a valence electron decreases as we move down the group.

2. "Increase in atomic size of elements": Inner shells increases as we move down the group. Thus, the atomic size increases as we move down the group. Also, the distance between the valence electron and nucleus of an atom, as a result the electrons are not strongly bounded. So, valence electrons can be expelled easily. Thus, energy required to expel a valence electron decreases as we move down the group.

9. Find out which of the pair given below will have high negative electron affinity?

a)F or Cl b) O or F

Answer:

1. F and Cl are the elements of the same group in periodic table. On

moving down the group the electron affinity becomes less negative. Here, the value of electron affinity of F is less negative than that of Cl. It is because atomic size of Cl is larger than that of F. In Cl, the electron will be added to n = 3 quantum level, whereas in F, the electron will be added to n = 2 quantum level. Thus, as the electron-electron repulsion is reduced in Cl so an extra electron can easily be accommodated. So, electron affinity of Cl is more negative compared to that of F.

2. O and F are the elements of the same period in periodic table. An F-atom is having 1 electron and 1 proton more than that of O-atom as electron is added in the same shell, Thus the atomic size of O-atom is larger than F-atom. As O-atom is having 1 proton less than F-atom. So, the nucleus of O-atom cannot attract an incoming electron that strongly as that of an F-atom. Also F-atom requires only 1 electron to achieve stable inert gas configuration. So, the electron affinity of F(Fluorine) is more negative than that of O(oxygen).

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

Answer:

	Metals		**Non–metals**
1.	Metals can lose electrons easily.	**1.**	Non-metals cannot lose electrons easily.

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

11. 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 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.

12(i). 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 as follows:

$F^- < Ne < Na^+$

Z 9 10 11

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

$Na^+ < Ne < F^-$

(ii). 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. 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.

7marks:

1. What do you understand by the term 'Ionic radius' and 'atomic radius'?

Answer:

Radius of an atom is known as atomic radius. It quantifies the size of an atom. On chance that the element is a metal, then its radius is termed as metallic radius, and if element is a non-metal, then its radius is termed as covalent radius. The metallic radius can be calculated as inter-nuclear distance between two molecules divided by 2. For instance, the inter-nuclear distance between two adjoining copper atoms is 256 pm in solid copper.

Metallic radius of copper = 256/2 pm=128pm

Covalent radius can be measured as the interatomic distance between 2 atoms when they are together by a solitary bond in a covalent atom. For instance, the interatomic distance between 2 chlorine atoms of chlorine molecule = 198 pm.

Covalent radius of copper = 198/2 pm=99pm

Radius of an ion (cation or anion) is known as ionic radius. Ionic radius is computed by measuring the inter-ionic distance between the cation and anion in an ionic crystal. Since cations are created by expelling an electron from outermost orbit of an atom, thus cation has less electrons compared to parent atom which results in increased effective nuclear charge.

In this way, a cation is small in size than parent atom. For instance, the ionic radius of Na^+ ion (sodium ion) = 95 pm, while the atomic radius of Na (sodium) atom = 186 pm. An anion is bigger in size than the parent atom. It is because an anion is having the same nuclear charge, yet more number of electrons compared to the parent atom which results in increased repulsion within atom among the electrons which also results in decreased effective nuclear charge. For instance, ionic radius of F^- (fluorine ion) = 136 pm, while the atomic radius of F(fluorine) atom = 64 pm.

2. Explain what is isoelectronic species? Give names of the species which will be isoelectronic species with each ion or atom given below.

1. Ar

2. Rb^+

3. F^-

4. Mg^+

Answer:

Ions and atoms which are having equal numbers of the electrons are called the isoelectronic species.

1. Ar (Argon) is having 18 electrons. Hence, the species which is isoelectronic with Ar must also have 18 electrons.

It's some isoelectronic species are

i) S^{2-} ion it is also having 18 electrons (16 + 2 = 18).

ii) Cl^- ion it is also having 18 electrons (17 + 1 = 18).

iii) K^+ ion it is also having 18 electrons (19 – 1 = 18).

2. Rb^+ (Rubidium) is having 36 electrons (37 – 1 = 36). Hence, the species which is isoelectronic with Rb+ must also have 36 electrons.

It's some isoelectronic species are

i)Br^- ion it is also having 36 electrons (35 + 1 = 36).

ii)Kr ion it is also having 36 electrons.

iii)Sr^{2+} ion it is also having 36 electrons (38 – 2 = 36).

3. F− (Fluorine) ion is having 10 electrons (9 + 1 = 10). Hence, the species which is isoelectronic with F− must also have 10 electrons.

It's some isoelectronic species are

i)Na^+ ion it is also having 10 electrons (11 – 1 = 10).

ii)Ne ion it is also having 10 electrons.

iii)Al^{3+} ion it is also having 10 electrons (13 – 3 = 10).

4. Mg^+ (Magnesium) ion is having 11 electrons (12 – 1 = 11). Hence, the species which is isoelectronic with Mg+ must also have 11 electrons.

It's some isoelectronic species are

i)Al^{2+} ion it is also having 11 electrons (13 – 2 = 11).

ii)Na ion it is also having 11electrons.

iii)Si^{3+} ion it is also having 11 electrons (14 – 3 = 11).

3. For some elements of the 2nd period the arrangement according to their ionization enthalpy is given as follows

Li<B<C<O<N<F<Ne

Explain Why?

1. ΔiH for O is lower than ΔiH of N and F.

2. ΔiH for Be is higher than ΔiH than B?

Answer:

1. In nitrogen, there are three 2p-electrons and all of these 3 occupy 3 distinct atomic orbital. While in oxygen 2 out of 4, 2p – electrons occupy same 2p-orbital, so the repulsion between the electrons in the oxygen atom increases.

Thus, the energy required to expel 2^{nd} 2p –electron in oxygen atom is higher than the energy required to expel 4th 2p –electron in nitrogen atom.

Thus, ΔiH for O is lower than ΔiH of N.

Fluorine atom is having one proton and one electron more than that in oxygen atom. As electron is added to a similar shell, the increment in attractive force between nucleus and electron (as proton is added) is higher than the increment in the repulsive force between electron-electron(as electron is added). Thus, valence electrons in the fluorine atom experiences higher effective nuclear charge compared to that, which is experienced by electron of oxygen atom. Thus, the energy required to expel an electron from fluorine is higher than the energy required to expel an electron from oxygen.

Thus, ΔiH for O is lower than ΔiH of O.

2. During ionization process, the electron that can be expelled from Be(beryllium) – atom is 2s – electron, but the electron that can be expelled from boron is 2p – electron.

The attractive force between a 2s – electron and nucleus is higher than between a 2s – electron and nucleus.

Thus, the energy required to expel 2s –electron is higher than the energy required to expel 2p –electron.
Thus, ΔiH for Be is higher than ΔiH than B

4.For the elements of group 13 the values of 1st ionization enthalpy is given below:

B	Al	Ga	In	Tl
801	577	579	558	589

Explain the 'deviation from the general trend'?
Answer:

Inner shells increases as we move down the group. Thus, the shielding effect of valence electrons increases by inner core electrons from nucleus. Thus, the attractive force on electrons towards nucleus is very strong. So, ionization enthalpy decreases as we move down the group. Hence for elements of group 13 the ionization enthalpy decreases as we move down from B to Al.
Here, Ga is having high ionization enthalpy than that of Al. This is because Al comes after the s-blocks elements, while Ga comes after the d-blocks elements. The shielding that is provided by electrons of d-block elements is not effective. So, the valence electrons are not shielded effectively. Thus, valence electrons in Ga atom experience higher effective nuclear charge compared to Al.
Further on moving down from Ga to In, the value of ionization enthalpy is decreased because of the increase in the shielding effect

and increase in atomic size.
But, Tl is having high ionization enthalpy than that of In. This is because Tl comes after the ‘4f and 5d electrons’. The shielding that is provided by these ‘4f and 5d electrons’ is not effective. So, the valence electrons are not shielded effectively. Thus, valence electrons in Tl atom experience higher effective nuclear charge compared to In.

5. 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 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

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

(i) $ns^2\ np^4$ for n = 3 (ii) $(n - 1)d^2\ ns^2$ for n = 4, and (iii) $(n - 2)\ f^7\ (n - 1)d^1\ ns^2$ for n = 6, in the periodic table.

Answer:

(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

= Number of *s*–block groups + number of *d*–block groups + number of *p*–electrons

= 2 + 10 + 4

= 16

Therefore, 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

= Number of *s*–block groups + number of *d*–block groups

= 2 + 2

= 4

Therefore, it is a 4th period and 4th group element. Hence, 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.

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

Elements	$\Delta_i H$	$\Delta_i H$	$\Delta_{eg} H$
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 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 V has the highest first ionization energy and high second ionization energy. Therefore, it can form a predominantly stable covalent halide of the formula MX (X=halogen).

8(i). 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 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.

(ii). 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. 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.

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

(iii). 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. 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.

Fill in the blanks:

1.The modern periodic table is arranged based on the increasing order of ___________ of elements.

Answer: atomic number

2.Elements in the same ___________ have similar chemical properties.

Answer: group

3.The horizontal rows in the periodic table are called __________.

Answer: periods

4.The first ionization energy generally __________ across a period and __________ down a group.

Answer: increases, decreases

5.Elements in the periodic table are classified into s, p, d, and f __________ based on their electron configurations.

Answer: blocks

6.The element with the highest electronegativity is found in the __________ of the periodic table.

Answer: top right corner

7.The atomic radius __________ across a period and __________ down a group.

Answer: decreases, increases

8.Elements in the same group have the same number of __________ electrons.

Answer: valence

9.The metalloids are found along the __________ between metals and non-metals.

Answer: staircase

10.Elements with similar electronic configurations are placed in the same ___________ of the periodic table.

Answer: column/group

Multiple choice:

1.Which property is the basis for the modern periodic table's arrangement of elements?

a) Atomic mass

b) Atomic number

c) Number of protons

d) Number of neutrons

Answer:

b) Atomic number

2.In the periodic table, elements in the same group share:

a) Atomic mass

b) Atomic radius

c) Electronic configuration

d) Number of isotopes

Answer:

c) Electronic configuration

3.The horizontal rows in the periodic table are called:

a) Sections

b) Periods

c) Blocks

d) Groups

Answer:

b) Periods

4.Which of the following trends is observed across a period in the periodic table?

a) Decreasing ionization energy

b) Increasing atomic radius

c) Increasing electronegativity

d) Decreasing metallic character

Answer:

c) Increasing electronegativity

5.Elements in the p-block of the periodic table are characterized by having electrons in their:

a) s subshell

b) p subshell

c) d subshell

d) f subshell

Answer:

b) p subshell

6.The element with the highest ionization energy is typically found in the:

a) Bottom left corner of the periodic table

b) Bottom right corner of the periodic table

c) Top left corner of the periodic table

d) Top right corner of the periodic table

Answer:

d) Top right corner of the periodic table

7.As you move down a group in the periodic table, the atomic radius generally:

a) Increases

b) Decreases

c) Remains constant

d) Becomes unpredictable

Answer:

a)Increases

8.Elements in the same column of the periodic table have the same number of:

a) Electrons

b) Protons

c) Neutrons

d) Valence electrons

Answer:

d) Valence electrons

9.Metalloids are found in which region of the periodic table?

a) Group 1

b) Group 14

c) Group 17

d) Along the staircase between metals and non-metals

Answer:

d) Along the staircase between metals and non-metals

10.Elements with similar properties are found in the same:

a) Period

b) Block

c) Group

d) Family

Answer:

c) Group

Summary:

The "Classification of Elements and Periodicity in Properties" chapter explores the systematic arrangement of elements in the periodic table and the patterns in their properties. The modern periodic table, organized based on the increasing atomic number, serves as the backbone for understanding the behaviour of elements.

Periods, the horizontal rows in the periodic table, and groups, the vertical columns, play pivotal roles in categorizing elements. Elements within the same group share similar chemical properties, as they have identical valence electron configurations. The chapter elucidates the significance of electronic configuration in determining the periodic trends in atomic size, ionization energy, electronegativity, and metallic character.

The p-block, d-block, and f-block elements are introduced, each contributing to the diversity of properties within the periodic table. The concept of metalloids, positioned along the staircase between metals and non-metals, highlights the intermediate nature of their properties.