

Chapter 3: Classification of Elements and Periodicity in Properties

NCERT · Class 11 · Chemistry

Textbook exercises

Question 1

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

Solution

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.

Question 2

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

Solution

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 elements 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.

Question 3

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

Solution

Ans:

Mendeleev's Periodic Law	Modern Periodic Law
Based on atomic mass and chemical properties.	Based on atomic number and chemical properties.
Arranged elements in increasing order of atomic mass.	Arranged elements in increasing order of atomic number.
Left gaps for undiscovered elements and predicted their properties.	No provision for gaps; elements placed strictly according to atomic number.
Isotopes are placed in the same slot based on average atomic mass.	Each isotope is considered as a separate entity and placed accordingly.
Elements with similar chemical properties placed in the same group.	Similar chemical behaviour attributed to elements in the same group.
Initially faced scepticism but gained acceptance over time.	Widely accepted and forms the basis of the modern periodic table.

Question 4

On the basis of quantum numbers, justify that the sixth period of the Periodic Table should have 32 elements.

Solution

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.

Question 5

In terms of period and group, where would you locate the element with

$$Z = 114?$$

Solution

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 elements 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.

Question 6

Write the atomic number of the element present in the third period and the seventeenth group of the Periodic Table.

Solution

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 18th group. As a result, the element in the third period's 17th group has the atomic number

$$Z = 17.$$

Question 7

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

Solution

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.

Question 8

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

Solution

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.

Question 9

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

Solution

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 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 fewer 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.

Question 10

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

Solution

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.

Question 11

What do you understand about isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

(i) F^-

(ii) Ar

(iii) Mg^{2+}

(iv) Rb^+

Solution

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 is 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 is 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 is 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 is 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.

Question 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.

Solution

Ans:

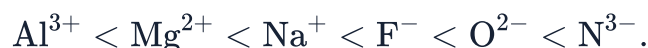
The number of electrons in each of the provided species (ions) is the same (10 electrons). As a result, the species in question are isoelectronic.

Ans :

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.



As a result, the following is the order of the supplied species in terms of increasing ionic radii:



Question 13

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

Solution

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.

Question 14

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

Solution

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 an 'isolated gaseous atom' and its 'ground state' must be established for comparison reasons.

Question 15

Energy of an electron in the ground state of the hydrogen atom is $-2.18 \times 10^{-18} \text{ J}$. Calculate the Ionization Enthalpy of Atomic Hydrogen in Terms of J mol^{-1} .

Solution

Ans :

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

Thus, the ionization enthalpy of atomic hydrogen is $2.18 \times 10^{-18} \text{ J}$.

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

$$\text{Ionization enthalpy} = 2.18 \times 10^{-18} \times 6.02 \times 10^{23} \text{ J mol}^{-1}$$

$$= 1.31 \times 10^6 \text{ J mol}^{-1}$$

Question 16

Among the 2nd 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?

Solution

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.

Ans:

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.

Question 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?

Solution

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. 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.

Question 18

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

Solution

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.

Question 19

The first ionization enthalpy values kJ mol^{-1} of Group 13 elements are:

Elements	Ionization Enthalpy (kJ mol ⁻¹)
Boron	801
Aluminum	577
Gallium	579
Indium	558
Thallium	589

How would you explain this deviation from the general trend?

Solution

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.

Question 20

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

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

Solution

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.

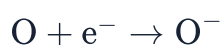
Question 21

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

Solution

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.



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.



Question 22

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

Solution

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.

Question 23

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

Solution

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.

Question 24

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

Solution

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.

Question 25

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

Solution

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.

Question 26

What are the major differences between metals and nonmetals?

Solution

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.

Question 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.

Solution

Ans :

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.

Ans:

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.

Ans:

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.

Ans:

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

Question 28

The increasing order of reactivity among group 1 elements is $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$ Whereas That Among Group 17 Elements Is $\text{F} > \text{Cl} > \text{Br} > \text{I}$. Explain.

Solution

Ans :

Group 1 elements have 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.



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: $F > Cl > Br > I$.

Question 29

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

Solution

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, $n = 6 - 7$

Question 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)

$$(n - 2)f^7(n - 1)d^1ns^2, \text{ for } n = 6$$

in the Periodic Table.

Solution

Ans:

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.

Ans:

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.

corresponding group of element = number of elements in (d - block + s - block)

$$= 2 + 2$$

$$= 4$$

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

Ans :

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,

$$\text{atomic number} = 54 + 7 + 2 + 1$$

$$= 64$$

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

Question 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.

Solution

Ans:

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 $(\Delta_{eg} H)$.

Ans:

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

Ans:

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

Ans:

Because it has a very high first ionization enthalpy ($\Delta_i H_1$) and a positive electron gain enthalpy ($\Delta_{eg} H$). element V is predicted to be the least reactive nonmetal.

(e) The metal which can form a stable binary halide of the formula MX_2 (X =halogen).

Ans:

The negative electron gain enthalpy of element VI is low ($\Delta_{eg} H$). 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 MX_2 (X =halogen) can be formed.

(f) The metal which can form a predominantly stable covalent halide of the formula MX (X =halogen)?

Ans :

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.

Question 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

Solution

Ans: Li_2O

Ans : Mg_3N_2

Ans: AlI_3

Ans: SiO_2

(e) Phosphorus and fluorine

Ans: PF_3 or PF_5

(f) Element 71 and fluorine

Ans: 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 .

Question 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

Solution

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.

Question 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 the Value of Azimuthal Quantum Number (l) For the Last Subshell That Received Electrons in Building up the Electronic Configuration.

Solution

Ans :

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

Question 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.

Solution

Ans :

The valence electrons are unaffected by nuclear mass.

Question 36

Anything that influences the valence electrons will affect the chemistry of the element. which one of the following factors does not affect the valence shell?

A.

Valence Principal Quantum Number (n)

B.

Nuclear Charge (Z)

C.

Nuclear Mass

D.

Number of Core Electrons.

Solution

Ans :

The valence electrons are unaffected by nuclear mass.

Question 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 electrons from core noble gas configuration.

C.

End of valence electrons is marked by a big jump in ionization enthalpy.

D.

Removal of electrons from orbitals bearing lower n value is easier than from an orbital having higher n value.

Solution

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.

Question 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.

$K > Mg > Al > B$

Solution

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 $K > Mg > Al > B$

Question 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.

$F > N > C > Si > B$

Solution

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 $B > Si$.

Question 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.

$O > F > N > Cl$

Solution

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 $F > O > Cl > N$.

