NCERT solutions for class 11 chemistry chapter 3 Classification of Elements and Periodicity in Properties

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

Answer :

The basic theme of organization in the periodic table is to classify all the elements according to similar properties in periods and groups, This arrangement makes the study of elements and their compounds simple and systematic and less confusion can be generated and true information can be regenerated about the different elements present in the periodic table.

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 elements in horizontal rows and vertical columns of a table in order of their increasing atomic weights in such a way that the elements with similar properties occupied the same vertical column or group. He fully recognized the significance of periodicity and used a broader range of physical and chemical properties to classify the elements.

In particular, Mendeleev relied on the similarities in the empirical formulas and properties of the compounds formed by the elements. He realized that some of the elements did not fit in with his scheme of classification if the order of atomic weight was strictly followed. He ignored the order of atomic weights, thinking that the atomic measurements might be incorrect, and placed the elements with similar properties together.

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 and can be stated as : The physical and chemical properties of the elements are periodic functions of their **atomic weights** .

But **Modern Periodic Law** states that, The physical and chemical properties of the 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 :

The period number corresponds to the highest principal quantum number (n) of the elements in the period. The subsequent periods consists of 8, 8, 18, 18 and 32 elements, respectively. The sixth period (n = 6) contains 32 elements and successive electrons enter 6s, 4f, 5d and 6p orbitals, Now, 6s has 1 orbital, 4f has 7 orbitals, 5d has 5 orbitals, and 6p has 3 orbitals.

Therefore, there is a total of 16 orbitals available in which each orbital carry 2 electrons according to Pauli's exclusion principle. So, we have a total of 32 electrons, Hence 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 7th period of the periodic table. Thus the element with Z=114 is present in the seventh period.

In the seventh period, the first two elements are s-block elements, the next 14 elements are f-block elements, next 10 elements are d-block elements and the next 6 elements are p-block elements.

Therefore, Z=114 element is the second p-block element in the seventh period. **Thus, it is present in the 4th group of the 7th period 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 :

Given that it is present in the 3rd period and a 17th group of the periodic table.

So, the first period has 2 elements and the second period has 8 elements. The 3rd period starts with the element with Z=11. And there are 8 elements present in the third period. Thus, the 3rd period ends with the element with Z=18. The element in the 18th group of the 3rd period has Z=18. Hence, the element in the 17th group of the 3rd period has atomic number **Z=17** .

Question 3.7(i) Which element do you think would have been named by
Lawrence Berkeley Laboratory

Answer :

Lawrence Berkeley Laboratory named the elements **Lawrencium (Lr)** with $Z = 103$ and **Berkelium (Bk)** with $Z = 97$.

Question 3.7(ii) Which element do you think would have been named by Seaborg's group?

Answer :

Seaborg's group named the elements **Seaborgium (Sg)** with $Z = 106$.

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

Answer :

Elements in the same group have similar physical and chemical properties because they have the same number of valence electrons. So, their most properties are similar, as they react in the same manner as the group has the same number of valence electrons.

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

Answer :

Atomic radius is the distance between from the centre of an atom to the outer most shell containing the electrons. It is a measure of the size of an atom.

They can be of types like :

(a) Covalent radius, where it is the half of the distance between the line joining the centres of nuclei of two adjacent similar atoms.

(b) Metallic radius is half of the distance between the centres of the nuclei of two adjacent atoms in the metallic crystal.

(c) van der Waal's radius, half of the internuclear distance between 2 similar adjacent atoms.

Whereas, the Ionic radius is the distance from the centre of the nucleus of the ion up to which it exerts its influence on the electron cloud of an ion (cation or anion). Generally, cation has a smaller ionic radius than the parent atom and anions are larger than the parent atom.

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

Answer :

The atomic radius of the elements generally decreases from the left to the right in a period because the nuclear charge gradually increases by one unit and one electron is also added in the electron shell, due to this the electrons get attracted more towards the centre, as a result, the atomic radii decreases.

The atomic radius of the elements increases as we move downwards in a group because there is an increase in the principal quantum number and thus, there is an increase in the number of electrons shells. Therefore, the atomic size is expected to increase.

Question 3.11(i) What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

F^-

Answer :

Atoms and ions which contains the same number of electrons are **isoelectronic species** .

GIven F^{-} ion which has total of $9 + 1 = 10$ electrons.

Thus, the species having the same number of electrons are Na^{+} ion having $(11 - 1 = 10)$ electrons, or Ne having 10 electrons, or O_2^{-} ion $(8 + 2 = 10)$ electrons, and Al^{3+} ion having $13 - 3 = 10$electrons.

Question 3.11(ii) What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

Ar

Answer :

Atoms and ions which contains the same number of electrons are **isoelectronic species** .

GIven Ar which has a total of 18 electrons.

Thus, the species having the same number of electrons are S_2^{-} ion having $(16 + 2 = 18)$ electrons, or Cl^{-} having $(17 + 1 = 18)$ electrons, or K^{+} ion $(19 - 1 = 18)$ electrons, and Ca^{2+} ion having $(20 - 2 = 18)$ electrons.

Question 3.11(iii) What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

Mg^{2+}

Answer :

Atoms and ions which contains the same number of electrons are **isoelectronic species** .

GIven Mg^{2+} which has total of $(12-2=10)$ electrons.

Thus, the species having the same number of electrons are

F^- ion having $(9+1=10)$ electrons, or O_2^- having $(8+2=10)$ electrons,

or Ne 10 electrons, and Al^{3+} ion having $(13-3=10)$ electrons.

Question 3.11 (iv) What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

Rb^+

Answer :

Atoms and ions which contains the same number of electrons, are **isoelectronic species** .

GIven Rb^+ which has total of $(37-1=36)$ electrons.

Thus, the species having the same number of electrons are

Br^- ion having $(35+1=36)$ electrons, or Kr having 36 electrons,

or Sr^{2+} ion $(38-2=36)$ electrons.

Question 3.12(a) Consider the following species : $N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+}$ What is common in them?

Answer :

Given species $N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}$ are **isoelectronic species** as they have the same number of electrons i.e., **10 electrons** .

$N^{3-} has\ 7+3=10\ electrons.$

$O^{2-} has\ 8+2=10\ electrons.$

$F^{-} has\ 9+1=10\ electrons.$

$Na^{+} has\ 11-1=10\ electrons.$

$Mg^{2+} has\ 12-2=10\ electrons.$

$Al^{3+} has\ 13-3=10\ electrons.$

Question 3.12(b) Consider the following species : $N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}$

Arrange them in the order of increasing ionic radii.

Answer :

Given species $N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}$.

As we know that the ionic radii of isoelectric species increases with a decrease in the magnitude of nuclear charge.

So, here is the increasing ionic radii arrangement :

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

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

Answer :

A cation is smaller than its parent atom because it has fewer electrons while its nuclear charge remains the same. The size of an anion will be larger than that of the parent atom because the addition of one or more electrons would result in increased repulsion among the electrons and a decrease in effective nuclear charge.

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?
Hint : Requirements for comparison purposes.

Answer :

The significance of the term **' isolated gaseous atom'** indicates that the atoms in the gaseous phase are much far separated that there does not have any mutual attraction or repulsion interactions present which is an isolated state. Here the value of ionization enthalpy and electron gain enthalpy are not influenced by the presence of the other atoms.

The significance of the term **' ground state '** means that in an atom, electrons are present in the lowest energy state where they neither lose nor gain an electron. Ionization enthalpy and electron gain enthalpy are generally expressed with respect to the ground state of an atom only.

Question 3.15 Energy of an electron in the ground state of the hydrogen atom is $-2.18 \times 10^{-18} J$ Calculate the ionization enthalpy of atomic hydrogen in terms of $J mol^{-1}$

Answer :

Given that Energy of an electron in the ground state of the hydrogen atom is $-2.18 \times 10^{-18} J$.

So, the ionization enthalpy is for 1 mole of atoms.

Therefore, the ground state energy of the atoms may be expressed as :

E (ground state) = $-2.18 \times 10^{-18} J \times (6.022 \times 10^{23} mol^{-1})$

$= -1.312 \times 10^{6} J\ mol^{-1}$

Whereas, ionization enthalpy is,

$= E_{\infty} - E_{ground\ \ state}$

$= 0 - (-1.312 \times 10^{6} mol^{-1}) = 1.312 \times 10^{6} J\ mol^{-1}$

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 Δ_i H than B

Answer :

The electronic configuration of Be is $(1s^2 2s^2)$. The outermost electron is present in the s-orbital. Whereas, in B the electronic configuration is $(1s^2 2s^2 2p^1)$ here outermost electron is present in p-orbital.

The electrons in s-orbital are more tightly attracted by the nucleus than the p-orbital, as a result, more amount of energy is required to knock out a s-orbital electron.

Hence Be has higher Δ_i H than B.

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 **(ii)** O has lower $\Delta_i H$ than N and F?

Answer :

In oxygen, two of the four 2p-electrons of oxygen occupy the same 2p-orbital and in nitrogen, the three 2p-electrons of nitrogen occupy three different atomic orbitals.

This results in increased electron-electron repulsion in oxygen atom as a result, 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 sp-electrons from nitrogen.

Hence, oxygen has lower $\Delta_i H$ than nitrogen.

In case of fluorine the electron is added to the same shell, the increase in nuclear attraction dominates over the increase in electronic repulsion.

Therefore, the valence electrons in fluorine atom experience a more effective nuclear charge than that experienced by the electrons present in oxygen. As a result, more energy is required to remove an electron from a fluorine atom than that required to remove an electron from oxygen atom.

Hence, 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 electronic configuration of Na and Mg are:

$Na = 1s^2 2s^2 2p^6 3s^1$ and $Mg = 1s^2 2s^2 2p^6 3s^2$

So, the first electron in case of both has to be removed from 3s-orbital, but the nuclear charge of Sodium has +11 charge and Magnesium has +12 charge, which causes electrons to held more tightly in case of Mg, therefore, **the first ionization energy of sodium is lower than that of magnesium.**

After the first ionization happens, the second electron has to be removed from the p-orbital in case of sodium which has already attained its stable noble gas configuration and from the s-orbital in case of magnesium.

Therefore, 2nd 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 :

Ionization enthalpy of the main group elements tends to decrease down a group because of the following reasons:

(a) **Atomic Size:** moving down the group there is an increase in the atomic size, increasing number of electron shells, which results in the weaker binding force with the nucleus. Hence the ionization enthalpy decreases.

(b) **Screening or shielding effect :** moving down the group increases the new shells, the number of inner electron shells which shield the valence electrons increases. As a result, the force of attraction of the nucleus for the valence electrons further decreases and hence the ionization enthalpy decreases.

Question 3.19 The first ionization enthalpy values (in $K\,Jmol^{-1}$) 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 :

The given trend can be explained by the following steps:

(i) Moving from B to Al, there is an increase in the size of the atom as a result decrease in the value of ionization enthalpy.

(ii) Moving from Al to Ga, there are 10 electrons in Ga which do not screen as is done by Sulphur and Phosphorus. Therefore, there is an unexpected increase in the value of effective nuclear charge resulting in increased ionization energy value.

(iii) Moving from Ga to In and Tl, there are 14 electrons in Tl with very poor shielding effect, which increases the effective nuclear charge thus the value of ionization energy increases.

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

O or F

Answer :

Oxygen and Fluorine both are of the second group elements and as we move from O to F, there is a decrease in atomic size, as a result, there is an increase in nuclear charge. Further, by a gain of one electron, $F \rightarrow F^{-}$ ion has an inert gas configuration and $O \rightarrow O^{-}$ gives ion which does not have a stable inert gas configuration.

The energy released is much higher in going from fluorine to oxygen or the electron gain enthalpy value of F is much more negative than that of oxygen.

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

F or Cl

Answer :

The electron gain enthalpy value of Cl $\triangle_{eg}H = -349kJ\ mol^{-1}$ is more negative than that of F $\triangle_{eg}H = -328kJ\ mol^{-1}$.

This is because of the reason for the smaller size of F due to its small size, the electron repulsions in the relatively compact 2p-subshell are comparatively large and hence the attraction for an incoming electron is less as in the case of Cl.

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 neutral O atom, a monovalent anion (O^{-}) is formed and there is release in energy, i.e., the first electron gain enthalpy is negative. After that when the second electron is added to anion (O^{-}) to form (O^{2-}) anion, there is a lot of electrostatic repulsions is there as both (O^{-}) anion and electron have a negative charge, more difficult to add the second electron. To overcome the force of repulsion, we have to give some energy. Thus, the second electron gain enthalpy of Oxygen is positive. i.e., $O_g \rightarrow O^{-} \rightarrow O^{2-}$

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

Answer :

The basic differences between the terms electron gain enthalpy and electronegativity are:

Electron gain enthalpy refers to the tendency of an isolated gaseous atom to accept an additional electron from a negative ion, whereas **electronegativity** refers to the tendency of an atom of an element to attract a shared pair of electrons towards it in a covalent bond.

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 :

The Given statement "the electronegativity of N on Pauling scale is 3.0 in all the nitrogen compounds" is **wrong** , because the electronegativity is a variable property of any element and is different for different kinds of compounds. For example electronegativity in NO_2 compound is different from NH_3 compound.

Question 3.24(a) Describe the theory associated with the radius of an atom as it

gains an electron

Answer :

When an atom gains an electron it forms an anion. And the size of an anion will be larger than the parent atom because of the addition of one or more electrons would result in increased repulsion among electrons and a decrease in the effective nuclear charge. This results in the increase in the atomic radius of an atom

Question 3.24(b) Describe the theory associated with the radius of an atom as it

loses an electron

Answer :

When an atom loses an electron it becomes positively charged and now for the same nuclear charge, there are lesser electrons present than the parent atom. Hence there will be an overall increase in the attraction between electrons and nucleus. As a result, there will be a reduction in the atomic radius of an atom and it is smaller than the parent atom.

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 :

As the ionization enthalpy, depends upon the number of electrons (electronic configuration) and nuclear charge (number of protons) in the nucleus, And there are the same electronic configuration of isotopes and also the nuclear charge.

Hence they will have the same ionization enthalpy.

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

Answer :

Metals are usually solids at room temperature [mercury is an exception; gallium and caesium also have very low melting points]. Metals usually have high melting and boiling points. They are good conductors of heat and electricity because they have the tendency to lose electrons easily. They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires).

Non-metals are usually solids or gases at room temperature with low melting and boiling points (boron and carbon are exceptions). They are poor conductors of heat and electricity. Most nonmetallic solids are brittle and are neither malleable nor ductile hence forming mainly covalent compounds.

Question 3.27(a) Use the periodic table to answer the following questions.

Identify an element with five electrons in the outer subshell.

Answer :

The element having five electrons in the outer subshell belongs to the nitrogen family or group 15, Example **Nitrogen** .

Question 3.27(b) Use the periodic table to answer the following questions.

Identify an element that would tend to lose two electrons.

Answer :

The element that would tend to lose two electrons should belong to alkaline earth family or group 2. For example **Magnesium** .

Question 3.27(c) Use the periodic table to answer the following questions.

Identify an element that would tend to gain two electrons.

Answer :

The element that would tend to gain two electrons should belong to the oxygen family or group 16. For example **Oxygen** .

Question 3.27(d) Use the periodic table to answer the following questions.

Identify the group having metal, non-metal, liquid as well as gas at the room temperature.

Answer :

Group 17 has metals, non-metals, liquid as well as gas at room temperature, i.e., chlorine, bromine are non-metals while iodine is a metal.

Fluorides are liquids in the state while chlorine, Br and Iodine are gases 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 :

In the case of group 1, we have ordered $Li < Na < K < Rb < Cs$ because there is only one valence electron. and thus have a strong tendency to lose this electron. The tendency to lose electrons in turn also depends upon the ionization enthalpy. Moving down the group there is a decrease in the ionization enthalpy, therefore, increasing reactivity down the group.

In the case of group 17, we have the order $F > CI > Br > I$ because there are seven electrons present in the valence shells and thus have a strong tendency to accept electrons to make it a stable noble gas electronic configuration. So, moving down we have a decrease in both electron gain enthalpy and electronegativity. Hence there is a decrease in the reactivity also.

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

Answer :

The general electronic configuration of s-,p-,d- and f- block elements are shown below:

For the s-block elements: ns^{1-2} , where $n = 2 - 7$.

For the p-block elements: ns^2np^{1-6} , where $n = 2 - 6$.

For the d-block elements, $(n-1)d^{1-10}ns^{0-2}$, where $n = 4 - 7$.

For the f- block elements: $(n-2)f^{0-14}(n-1)d^{0-1}ns^2$, where $n = 6 - 7$.

Question 3.30(i) Assign the position of the element having the outer electronic configuration

$ns^2np^4 \;\; for \;\; n=3$

Answer :

Given the outer electronic configuration of the element:

$ns^2np^4 \;\; for \;\; n=3$

For n=3 hence the element belongs to the third period, p-block element.

Since the valence shell contains 6 electrons and Group number = 10+6 =16

configuration $= 1s^22s^22p^63s^23p^4$ element name is Sulphur.

Question 3.30(ii) Assign the position of the element having outer electronic configuration

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

Answer :

Given the outer electronic configuration of the element:

$(n-1)d^2ns^2 \;\; for \; n=4$.

For n=4 hence the element belongs to the fourth period.

Since the valence shell contains (2+2) electrons and Group number = 4

configuration $= 1s^22s^22p^63s^23p^63d^24s^2$ element name is **Titanium (Ti)** .

Question 3.30(iii) Assign the position of the element having outer electronic configuration

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

Answer :

Given the outer electronic configuration of the element:

$(n-2)f^7(n-1)d^1ns^2 \ \ for \ n=6$.

For n=6 hence the element belongs to the **sixth period** and the last electron goes to the f-orbital, the element is from f-block.

We have the group number =3. Electronic configuration: $[X_e]4f^7 5d^1 6s^2$.

Hence the element is **Gadolinium** with Z=64.

Question 3.31(a) The first $\Delta_i H_1$ and the second $\Delta_i H_2$ ionization enthalpies (in $KJmol^{-1}$) $\Delta_{eg} H$electron gain enthalpy (in $KJmol^{-1}$) of a few elements are given below:

Elements $\Delta H_1 \quad \Delta H_2 \quad \Delta_{eg} H$

I 520 7300 –60

II 419 3051 –48

III 1681 3374 –328

IV 1008 1846 –295

V 2372 5251 +48

VI 738 1451 –40

the least reactive element.

Answer :

The element which has highest first ionization enthalpy $(\triangle_i H_1)$ and positive electron gain enthalpy $(\triangle_{eg} H)$ is $element\ V$.

And also element V shows similar behaviour like inert gases because of positive electron gain enthalpy.

Question 3.31(b) The first $\Delta_i H_1$ and the second $\Delta_i H_2$ ionization enthalpies (in $KJmol^{-1}$) $\Delta_{eg} H$electron gain enthalpy (in $KJmol^{-1}$) of a few elements are given below:

Elements ΔH_1 ΔH_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

the most reactive metal.

Answer :

The **element II** which has the least first ionization enthalpy value $(\triangle_i H_1)$ and a low negative electron gain enthalpy value $(\triangle_{eg} H)$ is the most reactive metal.

Question 3.31(c) The first $\Delta_i H_1$ and the second $\Delta_i H_2$ ionization enthalpies (in $KJmol^{-1}$) $\Delta_{eg} H$electron gain enthalpy (in $KJmol^{-1}$) of a few elements are given below:

Elements ΔH_1 ΔH_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

the most reactive non-metal.

Answer :

The **element III** which has high first ionization enthalpy $(\triangle_i H_1)$ and a very high negative electron gain enthalpy $(\triangle_{eg} H)$ is the most reactive non-metal.

Question 3.31(d) The first $\Delta_i H_1$ and the second $\Delta_i H_2$ ionization enthalpies (in $KJmol^{-1}$) $\Delta_{eg} H$electron gain enthalpy (in $KJmol^{-1}$) of a few elements are given below:

Elements ΔH_1 ΔH_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

the least reactive non-metal.

Answer :

The **element IV** has a high negative electron gain enthalpy $(\triangle_{eg} H)$ but not so high that the first ionization enthalpy $(\triangle_i H_1)$ and is the least reactive non-metal.

Question 3.31(e) The first $\Delta_i H_1$ and the second $\Delta_i H_2$ ionization enthalpies (in $KJmol^{-1}$) $\Delta_{eg} H$electron gain enthalpy (in $KJmol^{-1}$) of a few elements are given below:

Elements ΔH_1 ΔH_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

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

Answer :

The metal which can form a stable binary halide of the formula: MX_2 (X=halogen) is the element VI which has low first ionization enthalpy $(\triangle_i H_1)$ but higher than that of alkali metals.Hence it is an alkaline earth metal and form binary halide.

Question 3.32(a) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Lithium and oxygen

Answer :

For lithium and oxygen, the formula of the stable binary compound is: LiO_2 (Lithium oxide)

Question 3.32(b) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Magnesium and nitrogen

Answer :

For Magnesium and nitrogen, the formula of the stable binary compound is: Mg_3N_2 (Magnesium nitride) .

Question 3.32(c) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Aluminium and iodine

Answer :

For Aluminium and iodine, the formula of the stable binary compound is: AlI_3 (Aluminium iodide) .

Question 3.32(d) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Silicon and oxygen

Answer :

The stable binary compound that would be formed by the combination of Silicon and oxygen is: Silicon dioxide SiO_2 .

Question 3.32(e) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Phosphorus and fluorine

Answer :

The stable binary compound that would be formed by the combination of Phosphorus and fluorine are: Phosphorus trifluoride PF_3 or Phosphorus pentafluoride PF_5 .

Question 3.32(f) Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

Element 71 and fluorine

Answer :

The element with the atomic number 71 is Luteteium (Lu) having valency of 3.

Hence, the formula of the compound is LuF_3 (Lutetium trifluoride).

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 :

Answer - (C) Principal quantum number

In the modern periodic table, the period indicates the value of, **principal quantum number** (n) for the outermost shell or the valence shell.

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 theelectronic configuration

.

Answer :

Answer - (b) because d-block has a maximum of 10 electrons in all orbitals in a d-subshell and there are 10 columns. As each subshell can have a maximum of 5-orbitals and 10 electrons, therefore, there are 10 vertical columns in a d-block.

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 :

Answer - (c) Nuclear Mass.

Valence shell electrons are present in the outermost shell of an atom and nuclear mass does not affect the valence shell because the nuclear mass has so small value that it is considered as negligible.

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 :

Answer - (a) Nuclear charge (Z)

Because as the nuclear charge increases the size of an isoelectronic species decreases.

Here we have the isoelectronic species: F^{-} , Ne and Na^{+}

F^{-} has $Z=9$

Ne has $Z = 10$

Na^+ has $Z = 11$

Therefore, the order of the increasing size is as follows:

$Na^+ < Ne < F^-$

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 :

Answer - (d)

Because 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.

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 - (d) K > Mg > Al > B

Because the metallic character of an element is the tendency of an element to lose electrons easily. So, moving from left to right across the period metallic character of element decreases.

Here, K and Mg are s-block elements while B and Al belong to p-block elements.

Hence the order is as follows:

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 - (c) F > N > C > B > Si

Because non-metallic character of an elements increases from left to right across the period. And here given position of elements:

Boron (B) - 2^{nd} period and 13^{th} group.

Carbon (C) - 2^{nd} period and 14^{th} group.

Nitrogen (N) - 2^{nd} period and 15^{th} group.

Fluorine (F) - 2^{nd} period and 17^{th} group.

And the Silicon (Si) is present in the 3^{rd} period and 14^{th} group.

Hence the correct order of non-metallic character is as follows:

F > N > C > B > S

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 :

Answer - (b) $F > O > Cl > N$

The oxidizing property of elements increases from left to right across a period because of the presence of vacant d-orbitals in their valence shells.

Thus, we get the decreasing order of oxidizing property as $F > O > N$.

And the oxidizing character of elements decreases down a group. Thus we get $F > Cl$.

However, the oxidizing character of Oxygen is more than that of Chlorine.

So, $O > Cl$.

Hence the correct order their chemical reactivity in terms of oxidizing property is :

$F > O > Cl > N$.