Ionization Enthalpy Chemistry Questions with Solutions

Q1: Pick the correct statement.

a. The Ionization enthalpy increases along the period.
b. The Ionization enthalpy decreases along the period.
c. The electropositive nature of elements increases along the period.
d. The electronegative nature of elements decreases along the period.

Answer: (a.)

Explanation: The Ionization enthalpy depends upon the nuclear charge. As we move from left to right in a period, the nuclear charge and hence, the pull on the outermost electron increases. Hence, the Ionization enthalpy increases as we move from left to right in a period.

Q2. Tick mark the incorrect statement.

a. The first ionization enthalpy (Δ_iH_1) of Aluminium is less than that of Magnesium.
b. The first ionization enthalpy (Δ_iH_1) of Sodium is less than that of Magnesium.
c. The second ionization enthalpy (Δ_iH_2) of Magnesium is greater than the second ionization enthalpy of Sodium.
d. The third ionization enthalpy (Δ_iH_2) of Magnesium is greater than the third ionization enthalpy of Aluminium.

Answer: (c.)

Explanation: The second I.E. (Δ_iH_2) of Sodium is greater than the second I.E. of Magnesium. This is because sodium attains the noble gas core after the removal of 1 electron. Thus, the second electron from sodium would be removed from its noble gas core and hence, would require high energy. While in the case of Magnesium, the magnesium receives a noble gas core after the removal of a second electron. Hence, the Δ_iH_2 of Magnesium is lesser than that of sodium.

Q3. The second ionization enthalpy for an element is the required energy for

a. The removal of a pair of electrons
b. The removal of 1 mole of electrons from 1 mole of gaseous anions.
c. The removal of 1 mole of electrons from 1 mole of the monovalent gaseous cations of the given element.
d. The removal of 2 moles of electrons from a neutral gaseous atom of the element.

Answer: (c.)

Explanation: The second ionization enthalpy for an element is the required energy for the removal of 1 mole of electrons from 1 mole of the monovalent gaseous cations of the given element.

Q4. Which of the following atoms has the lowest first ionization enthalpy?

a. Potassium
b. Calcium
c. Strontium
d. Rubidium

Answer: (d.)

Explanation: Ionization enthalpy is the lowest towards the left within a period. Ionization enthalpy also decreases down the group. Hence, among the given elements, Rubidium has the lowest first ionization enthalpy.

Q5. Which of the following atoms has the highest first ionization enthalpy?

a. Fluorine
b. Chlorine
c. Bromine
d. Iodine

Answer: (a.)

Explanation: As the ionization enthalpy increases towards the right in a period and moving up in a group, Fluorine has the highest first ionization enthalpy.

Q6. Why is the first ionization enthalpy of the alkaline earth metals greater than the alkali metals?

Answer: This is because the alkaline earth metals have a greater nuclear charge than the alkali metals within the respective period. Hence, the ionization enthalpy increases on moving from left to right within a period.

Q7. How much energy will be required for the conversion of all of the Magnesium atoms into their respective magnesium cations present in 24 grams of Magnesium vapours? Given that the first ($\Delta_i H_1$) and second ionization enthalpies ($\Delta_i H_2$) of magnesium are 737.76 kJ mol^{-1} and 1450.73 kJ mol^{-1} respectively.

Answer: From the definition of successive ionization enthalpies:

$Mg (g) + \Delta_i H_1 \rightarrow Mg^+ (g) + e^- (g)$; $\Delta_i H_1 = 737.76$ kJ mol^{-1}

$Mg^+ (g) + \Delta_i H_2 \rightarrow Mg^{2+} (g) + e^- (g)$; $\Delta_i H_2 = 1450.73$ kJ mol^{-1}

Hence, the total energy required to create an Mg^{2+} ion from Mg (g) is ($\Delta_i H_1 + \Delta_i H_2$).

∴ total energy required to create an Mg^{2+} ion = (737.76 + 1450.73) kJ mol^{-1} = 2188.49 kJ mol^{-1}

Now, 24 g Magnesium contains mole = 1 mol

∴ 0.024 g Magnesium contains mole = 0.024 / 24 mol = 10^{-3} mol

Hence, the total amount of energy required to convert 10^{-3} mol of Mg (g) into Mg^{2+} (g) = 2188.49 kJ mol^{-1} x 10^{-3} mol = 2.188 kJ.

Q8. Given the first and second ionization potentials of helium (He) are 24.58 eV and 54.4 eV respectively. Calculate the energy (kJ) required to produce one mole of He^{2+} ions.

Answer: The total potential required to make an ion of H^{2+} = 24.58 eV + 54.4 eV = 78.98 eV
Hence, the total potential required to make 1 mole of H^{2+} ions = 1 x 78.98 eV = 78.98 eV
As 1 eV = 96.49 kJ mol^{-1}
Hence, 78.98 eV = 78.98 x 96.49 kJ mol^{-1} = 7620.78 kJ mol^{-1}
∴ The energy (kJ) required to produce one mole of He^{2+} ions is 7620.78 kJ.

Q9. What happens if the pressure is increased during the Haber's process?

Answer: On increasing the pressure, the equilibrium of the reaction moves to its right. Hence, the yield of ammonia will increase. However, to apply high pressures, heavy and large equipments are required which are very expensive. Hence, pressure is set at a compromisable unit of 200 atm.

Q10. Arrange the following electronic configurations in the increasing order of their ionization enthalpies.

a. $1s^2\ 2s^2\ 2p^6\ 3s^2$,
b. $1s^2\ 2s^2\ 2p^2$,
c. $1s^2\ 2s^2\ 2p^6\ 3s^1$,
d. $1s^2\ 2s^2\ 2p^6$,
e. $1s^2\ 2s^2\ 2p^3$

Answer: The ionization enthalpy decreases as the distance of the electron from the nucleus increases. The electronic configuration (a.) and (c.) have their electrons in the third principal energy level. Since, (c.) has only 1 valence electron, so the order of ionization enthalpy in between (a.) and (c.) is (a.) > (c.). (d.) and (e.) are the completely filled and half-filled electronic configurations. Hence, (d.) is the most stable configuration and has the highest ionization energy. While (e.) has the second highest ionization energy. Hence, the increasing order of their ionization enthalpies is:
(c.) < (a.) < (b.) < (e.) < (d.)

Q11. Out of Fluorine and Chlorine, which one has the lowest first ionization enthalpy?

Answer: Chlorine has the lower $\Delta_i H_1$ owing to its larger size and stronger shielding effect.

Q12. The $\Delta_i H_1$ of carbon is greater than the $\Delta_i H_1$ of boron, while the $\Delta_i H_2$ of boron is greater than that of carbon. Explain.

Answer: Since both carbon and boron belong to the 2nd period, their electronic configurations must be compared to get an explanation.

$C = 1s^2\ 2s^2\ 2p^2$ $B = 1s^2\ 2s^2\ 2p^1$

Carbon and Boron have their outermost electron in the same subshell i.e. 2p. Carbon has comparatively higher nuclear charge and hence more nuclear pull on the outermost electron than in boron. Hence, Δ_iH_1 of boron is less than that of carbon.

After the removal of an electron from carbon and boron, the electronic configuration:

$C^+ = 1s^2\ 2s^2\ 2p^1$ $B^+ = 1s^2\ 2s^2$

Since, the monovalent Carbon and Boron have their outermost electron in 2p and 2s subshell respectively. The s-subshells are more penetrated towards the nucleus, hence have more effect of nuclear attraction on its electrons. Hence, Δ_iH_2 of Boron is higher than carbon.

Q13. Out of Na, Mg, Si, and P, which one has the greatest difference in between their Δ_iH_1 and Δ_iH_2?

Answer: Na being an alkali metal has the lowest value of Δ_iH_1. However, after the removal of an electron, Na acquires a noble gas core, hence, the Δ_iH_2 of Na is very high. Hence, the difference between the Δ_iH_1 and Δ_iH_2 of Na is the highest.

In case of Mg, Si and P, the value of Δ_iH_1 is higher than Na, but their Δ_iH_2 is not very high. Hence, the difference between the Δ_iH_1 and Δ_iH_2 of these elements is not very high.

Q14. Arrange the following ions in decreasing order of ionic radii: Li^{2+}, He^+, and Be^{3+}.

Answer: Li^{2+}, He^+, and Be^{3+} are isoelectronic ions. The number of protons present in Li^{2+}, He^+, and Be^{3+} are 3, 2 and 4 respectively. Hence, the decreasing order of ionic radii is $He^+ > Li^{2+} > Be^{3+}$.

Q15. Which among the following have the largest and the smallest sizes?

Mg^{2+}, Mg, Al^{3+}, Al

Answer: Both Mg and Al belong to the third period. As the atomic radii decreases on moving from left to right in a period due to increasing nuclear charge. Hence, Mg has a larger atomic radius than Al. Now, as the cations are smaller in size than their parent atoms, Mg^{2+} is smaller than Mg and Al^{3+} is smaller than Al. As Mg^{2+} and Al^{3+} are isoelectronic, Al^{3+} is smaller than Mg^{2+}.

Hence, Mg has the largest and Al^{3+} has the smallest size.

Practise Questions on Ionization Enthalpy

Q1. For which of the following electronic configurations, the value of ionization enthalpy is the lowest?

a. $1s^2\ 2s^2\ 2p^6\ 3s^2$
b. $1s^2\ 2s^2\ 2p^6\ 3s^1$
c. $1s^2\ 2s^2\ 2p^4$
d. $1s^2\ 2s^2\ 2p^6$

Answer: (b.)

Explanation: Since the outermost subshell has only 1 electron. After the removal of 1 electron, the element will acquire a noble gas configuration. Hence, the ionization enthalpy for the electronic configuration (b.) is the lowest.

Q2. Given are the first and second ionization enthalpies of some elements in kJ mol^{-1}.

Element	$\Delta_i H_1$	$\Delta_i H_2$
A	2372	5251
B	520	7300
C	900	1760
D	1680	3380

a. Which of the above is the most reactive metal?
b. Which of the above is the most reactive non-metal?
c. Which of the above is a noble gas?

Answer: (a.) Since the element B has a very less $\Delta_i H_1$, this implies that the element B is very reactive. However, the $\Delta_i H_2$, for B is very high, this implies that B must be a very reactive metal which leaves its second electrons at a very high energy provided. Hence, element B must be a very reactive alkali metal.
(b.) The $\Delta_i H_1$ of the element D is very high but its $\Delta_i H_2$ is not as considerably high. Hence, the element D must be a reactive non- metal, preferably a halogen.
(c.) Element A has the highest value of $\Delta_i H_1$ but $\Delta_i H_2$ is not so high. Hence, A must be a noble gas.

Q3. Explain why cations are smaller and anions are larger in size?

Answer: The size of a cation is always smaller than its parent atom. This is because after the removal of 1 or more electrons, the effective nuclear charge experienced by the rest of the electrons increases. As a result, the force of attraction by the nucleus on the remaining electrons increases and hence, the size of the ion decreases.
While in case of anions, the anionic size is always greater than the parent atom. This is because the addition of 1 or more electrons to the neutral atom decreases the effective nuclear charge, hence, the force of attraction experienced by the electrons decreases. Hence, the anionic size increases.

Q4. Would you expect the first ionization enthalpies of the two isotopes of hydrogen to be the same or different?

Answer: The ionization energy depends upon various factors including the electronic configuration of the atom and the nuclear charge. The electronic configuration and the nuclear charge (number of protons) of the two isotopes are the same. Hence, the two isotopes must have the same ionization energies.

Q5. Chlorine can be converted into the chloride ion more easily than the conversion of Fluorine into fluoride ion. Explain.

Answer: This is because the electron gain enthalpy of Chlorine is more negative than that of Fluorine.