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CHEMISTRY

INORGANIC CHEMISTRY

Level & Board	AQA (A-LEVEL)
TOPIC:	PERIODICITY
PAPER TYPE:	SOLUTION - 2
TOTAL QUESTIONS	10
TOTAL MARKS	38

Periodicity - 2

1.

(a)

The term "first ionization energy" refers to the energy (enthalpy change, ΔH) needed to remove one mole of electrons from one mole of gaseous atoms.

$Na(g) \rightarrow Na^+(g) + e^-$ $\Delta H = 496 kJ/mol$

(2)

(b)

The trend in first ionization energies from sodium to phosphorus across Period 3 increases due to the increasing nuclear charge, which results from an increasing number of protons in the nucleus.
The shielding effect remains constant, as these elements have the same number of electron shells or energy levels.
The atomic radius decreases as moving across the period, bringing the outer electrons closer to the nucleus. This decrease in distance from the nucleus, combined with the stronger attraction from the increasing nuclear charge, makes it requires more energy to remove the outermost electron.

(3)

(c)

Silicon (Si) has the highest melting point among the elements from Na to P.

Explanation:

Macromolecular Structure:

- Silicon has a macromolecular or giant molecular structure in its solid form.
- The atoms in silicon are covalently bonded in a three-dimensional network.

Strong Covalent Bonds:

- Silicon is characterized by many or strong covalent bonds between its atoms.
- The covalent bonds in silicon are strong and require a significant amount of energy to break.

Energy Required for Melting:

- *Many strong covalent bonds need to be broken in silicon to transition from a solid to a liquid state.*
- *The high energy requirement for breaking these covalent bonds contributes to the high melting point of silicon.*

(3)

(d)

The element that deviates from the general trend of increasing melting points from sodium (Na) to phosphorus (P) is aluminum (Al).

Explanation:

Electron Configuration:

- *Aluminum has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^1$.*
- *The electron being removed during ionization is from a higher energy level or a 3p orbital, not a 2p orbital.*

Energy Level of the Electron Being Removed:

- *The electron being removed in aluminum is from a higher energy level compared to the elements from Na to P.*
- *The electron is in a 3p orbital, indicating a higher principal quantum number.*

Ease of Electron Loss:

- *Aluminum requires less energy to lose this outer electron compared to the other elements.*
- *The ionization energy for aluminum is lower, indicating that the electron is more easily lost.*

(3)

2. A

(1)

3.

(a)

Aluminum has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^1$.
It belongs to p-block.

(2)

(b)

The bonding in metals has a lattice of positively charged metal cations surrounded by a sea of delocalized electrons. This shows the regular and *often close-packed arrangement of metal cations in the metallic lattice, held together by the electrostatic attraction between cations and electrons.*

(2)

(c)

The higher melting point of magnesium compared to sodium is attributed to several factors:

Greater Nuclear or Ionic Charge:

- *Magnesium has a greater nuclear or ionic charge compared to sodium.*

Smaller Atoms/Ions:

- *Magnesium ions are smaller than sodium ions.*
- *The smaller size of magnesium ions contributes to a higher charge density in the magnesium lattice.*

More Delocalized Electrons/Free Electrons:

- *In both magnesium and sodium, there are delocalized electrons forming a sea of electrons.*
- *Magnesium, being larger and having more electrons, contributes to a higher number of delocalized electrons.*

Stronger Attraction Between Ions and Delocalized Electrons:

- *The combination of greater nuclear charge, smaller ions, and more delocalized electrons in magnesium results in a stronger electrostatic attraction between the metal cations and the delocalized electrons.*

(3)

(d)

Metals conduct electricity due to delocalized electrons that Move in a particular direction under the influence applied potential difference.

(2)

4. **D**

(1)

5.

***Electronegativity** is a measure of the tendency of an atom to attract a bonding pair of electrons in a chemical bond. It is the ability of an atom to pull shared electrons towards itself when forming a chemical bond with another atom.*

Trend in Electronegativity Across Period 3 (Sodium to Chlorine):

Increase Across Period 3:

- *The electronegativity values generally increase from left to right across Period 3.*
- *This trend is due to the increase in the effective nuclear charge as the number of protons in the nucleus increases.*

Nuclear Charge and Electronegativity:

- *As moving from sodium to chlorine, the number of protons in the nucleus increases, resulting in a higher positive charge in the nucleus.*
- *The increased positive charge enhances the ability of the atom to attract electrons in a chemical bond.*

Atomic Size Influence:

- *Although electrons are added to the same principal energy level (n=3) as you move across Period 3, the increasing positive charge in the nucleus is the dominant factor.*
- *The stronger pull from the nucleus outweighs the electron-electron repulsion within the same energy level.*

(5)

6. **B**

(1)

7.

(a)

The decrease in atomic radii across Period 2 (from left to right) in the periodic table is due to an increase in effective nuclear charge. Moving from left to right across a period, the number of protons in the nucleus increases, leading to a stronger attractive force between the positively charged nucleus and the negatively charged electrons. However, the number of shells remains the same.

This stronger attractive force pulls the electrons closer to the nucleus, resulting in a decrease in atomic size. This also counteracts the shielding effect of inner electrons, making the outer electrons experience a stronger pull from the nucleus. As a result, the atomic radii of elements decrease across Period 2.

(2)

(b)

Metal oxides generally form alkaline solutions when they react with water due to the formation of hydroxide ions (OH^-).

For example:

The reaction of sodium oxide (Na_2O) with water:

$Na_2O + H_2O \rightarrow 2NaOH$

In this reaction, sodium oxide reacts with water to form sodium hydroxide. The hydroxide ions (OH^-) released during this process make the solution alkaline.

(2)

8. **C**

(1)

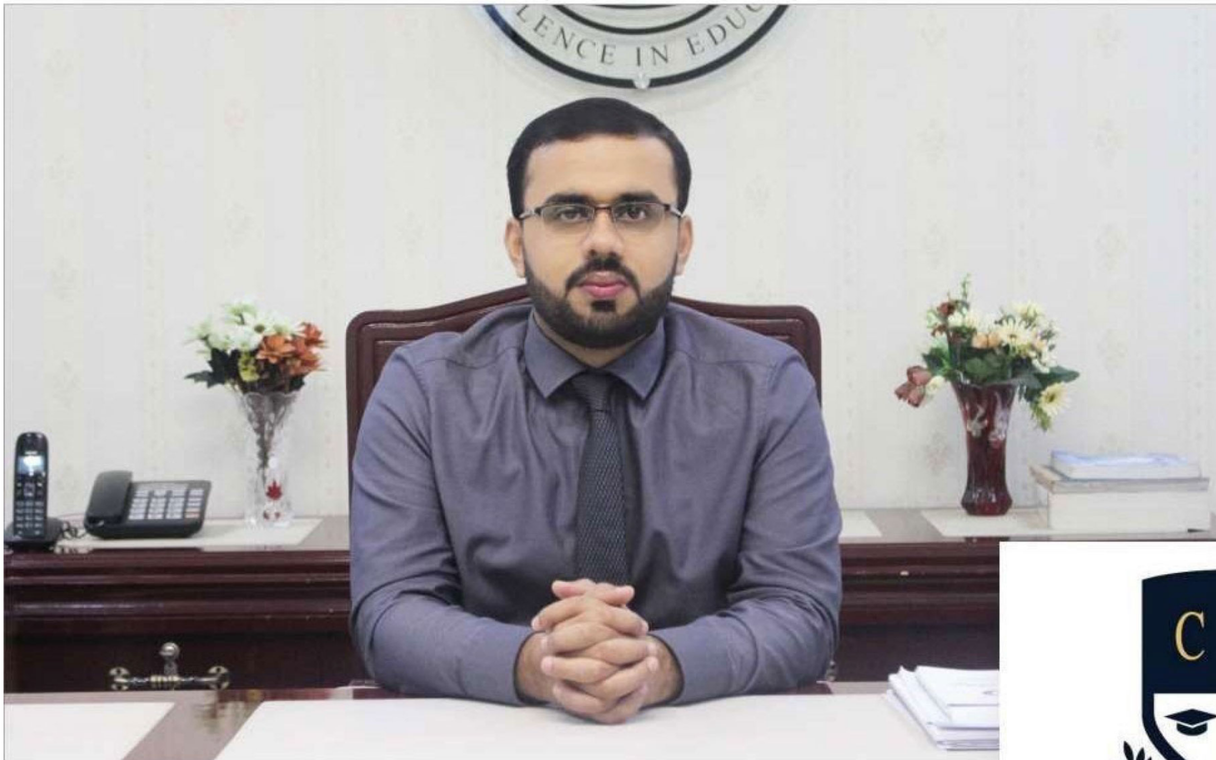
9.

Rubidium (Rb) and sodium (Na) are both alkali metals in Group 1 of the periodic table.
Moving down Group 1, the atomic size increases due to the addition of more electron shells. The outermost electrons in the alkali metals are in the same principal energy level, so the key factor influencing ionization energy is the distance of the outer electrons from the nucleus.
In this case, **rubidium is below sodium in the periodic table, and it has more electron shells. The outermost electrons in rubidium are farther away from the nucleus compared to the outermost electrons in sodium. The increased distance results in weaker attractive forces between the outer electrons and the nucleus in rubidium, making it easier to remove an electron.**
So, less energy is needed to ionize gaseous atoms of rubidium compared to gaseous atoms of sodium.

(4)

10. **C**

(1)



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