# Periodic Table and Periodicity

## Question1

**Among the ions $Mg^{2+}, O^{2-}, Al^{3+}, F^{-}, Na^{+}$ and $N^{3-}$, the ion with largest size and ion with smallest size are respectively**

### AP EAPCET 2025 - 26th May Morning Shift

**Options:**

A.

$N^{3-}, Mg^{2+}$

B.

$O^{2-}, F^{-}$

C.

$Al^{3+}, N^{3-}$

D.

$N^{3-}, Al^{3+}$

**Answer: D**

## Solution:

1. **Determine the number of electrons for each ion:**
    - $N^{3-}$: Neutral N has 7 electrons. Gaining 3 electrons makes it 7 + 3 = 10 electrons.
    - $O^{2-}$: Neutral O has 8 electrons. Gaining 2 electrons makes it 8 + 2 = 10 electrons.
    - $F^{-}$: Neutral F has 9 electrons. Gaining 1 electron makes it 9 + 1 = 10 electrons.
    - $Na^{+}$: Neutral Na has 11 electrons. Losing 1 electron makes it 11 - 1 = 10 electrons.

- $Mg^{2+}$: Neutral Mg has 12 electrons. Losing 2 electrons makes it 12 - 2 = 10 electrons.
- $Al^{3+}$: Neutral Al has 13 electrons. Losing 3 electrons makes it 13 - 3 = 10 electrons.

All these ions are isoelectronic, meaning they all have the same electron configuration (like Neon, [Ne]).

2. **Determine the nuclear charge (number of protons) for each ion:**
   - $N^{3-}$: Z = 7
   - $O^{2-}$: Z = 8
   - $F^{-}$: Z = 9
   - $Na^{+}$: Z = 11
   - $Mg^{2+}$: Z = 12
   - $Al^{3+}$: Z = 13
3. **Apply the rule for isoelectronic species:** For isoelectronic ions, the ionic radius decreases as the nuclear charge (number of protons) increases. A higher nuclear charge means the electrons are pulled more strongly towards the nucleus, resulting in a smaller size.

   Ordering the ions by increasing nuclear charge (and thus decreasing size):
   - $N^{3-}$ (Z=7) - Largest size
   - $O^{2-}$ (Z=8)
   - $F^{-}$ (Z=9)
   - $Na^{+}$ (Z=11)
   - $Mg^{2+}$ (Z=12)
   - $Al^{3+}$ (Z=13) - Smallest size

Therefore, the ion with the largest size is $N^{3-}$ and the ion with the smallest size is $Al^{3+}$.

The final answer is $\boxed{D}$.

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# Question2

## Which of the following orders is not correct for the given property?

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**Options:**

A.

$Li < Na < K$ - metallic radius

B.

$Br < F < Cl$ - electron gain enthalpy

C.

$C < N < O$ - first ionisation enthalpy

D.

$Mg^{2+} < Na^{+} < F^{-}$- ionic radius

**Answer: C**

## Solution:

The order of first ionisation enthalpy is incorrect in option (c). The correct order is $C < O < N$

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# Question3

## The order of negative electron gain enthalpy of $Li, Na, S, Cl$ is

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**Options:**

A.

$Na > S > Cl > Li$

B.

$Cl > S > Li > Na$

C.

$Cl > Li > S > Na$

D.

$Li > Na > S > Cl$

**Answer: B**

## Solution:

The order of negative electrons gain enthalpy is

Cl > S > Li > Na

Electron gain enthalpy becomes more negative with increase in the atomic number across a period. While E.G.E. becomes less negative as we go down the group because the size of the atom increases and the added $e^-$ are farther away from the nucleus.

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# Question4

## The period and group numbers of the element having maximum electronegativity in the long form of periodic table respectively, are

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**Options:**

A.

2,17

B.

3,17

C.

1,18

D.

2,16

**Answer: A**

## Solution:

The element with the highest electronegativity in the periodic table is fluorine (F). It is located in period 2 and group 17.

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# Question5

**Electronic configurations of four elements $A, B, C, D$ are given below**

**(A)** $1s^2 2s^2 2p^6 3s^1$

**(B)** $1s^2 2s^2 2p^6 3s^2 3p^1$

**(C)** $1s^2 2s^2 2p^6 3s^2$

**(D)** $1s^2 2s^2 2p^6 3s^2 3p^2$

**The correct order of first ionisation enthalpy of these elements is**

## AP EAPCET 2025 - 23rd May Morning Shift

**Options:**

A.

D > B > C > A

B.

C > D > B > A

C.

C > A > B > D

D.

D > C > B > A

**Answer: D**

## Solution:

The elements of given electronic configuration are $(A)$Na, $(B)$Al$(C)$Mg and $(D)$ Si. Thus, the correct order of first ionisation enthalpy of these elements is $D > C > B > A$

Ionisation enthalpy generally increases across a period due to increase nuclear charge and decreasing atomic radius.

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# Question6

## The correct order of the non-metallic character among the elements B, C, N, F and Si is

## AP EAPCET 2025 - 22nd May Evening Shift

**Options:**

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**

## Solution:

The correct order of non-metallic character is

$F > N > C > B > Si$

Non-metallic character increases across the period and decreases down the group.

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# Question7

## Identify the incorrect order against the stated property.

## AP EAPCET 2025 - 22nd May Evening Shift

**Options:**

A.

Ge > Sn > Pb - Ionisation enthalpy

B.

Ge > Pb > Sn - Melting point

C.

Pb > Sn > Ge– Density

D.

Ge > Pb > Sn - Electrical resistivity

**Answer: A**

## Solution:

Answer: **(A) Ge > Sn > Pb - Ionisation enthalpy**

Explanation: The general trend for ionization enthalpy down a group is a decrease. However, due to the poor shielding effect of $d$ and $f$ electrons in heavier elements like Pb, the effective nuclear charge increases, leading to an increase in ionization enthalpy from Sn to Pb. Therefore, the correct order is $\mathbf{Ge > Sn < Pb}$ (or more accurately, $\mathbf{Ge > Pb > Sn}$), making the given order in option A incorrect. The orders for melting point, density, and electrical resistivity are generally correct for these elements.

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# Question8

## The correct order of atomic radii of C, Al and S is

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**Options:**

A.

C < Al < S

B.

S < AI < C

C.

$AI < S < C$

D.

$C < S < Al$

**Answer: D**

## Solution:

The correct order of atomic radii is,

$C < S < Al$

As we go left to right in period, radii decreases. As we go down in group radii increases.

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# Question9

## Which of the following orders is not correct about the property shown against it?

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**Options:**

A.

$N > O > P > S$ - First ionisation enthalpy

B.

$F > Cl > O > S$ - Negative electron gain enthalpy

C.

$Fe^{3+} < Fe^{2+} < Fe$ - Size

D.

$O > N > S > P$ - Non-metallic character

**Answer: B**

## Solution:

The order of negative electron gain enthalpy is incorrect the correct order is,

Cl > F > S > O

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# Question10

Match the following

| | List-I (Atomic number; Z) | | List-II (Block) |
|---|---|---|---|
| A | 112 | I | s |
| B | 116 | II | p |
| C | 88 | III | d |
| D | 100 | IV | *f* |

## The correct answer is

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**Options:**

A.

A-III, B-I, C-II, D-IV

B.

A-III, B-II, C-I, D-IV

C.

A-IV, B-II, C-III, D-I

D.

A-II, B-III, C-IV, D-I

**Answer: B**

## Solution:

The match is A-III, B-II, C-I, D-IV
$112 - [\mathrm{Rn}]5f^{14}6d^{10}7s^2$
$116 - [\mathrm{Rn}]5f^{14}6d^{10}7s^27p^4$
$100 - [\mathrm{Rn}]5f^{12}7s^2$
$88 - [\mathrm{Rn}]7s^2$

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# Question11

Match the following.

| List I (Symbol of element) | | List II (Group number) | |
|---|---|---|---|
| A. | Mc | I. | 16 |
| B. | Lv | II. | 17 |
| C. | Fl | III. | 15 |
| D. | Ts | IV. | 14 |

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**Options:**

A. A -III, B - IV, C - I, D-II

B. A-IV, B-III, C-I, D-II

C. A - III, B-I, C-IV, D-II

D. A-II, B-IV, C-I, D-III

**Answer: C**

## Solution:

Here's how the elements line up by group number:

Moscovium (Mc, Z = 115) is in group 15 ⇒ III

Livermorium (Lv, Z = 116) is in group 16 ⇒ I

Flerovium (Fl, Z = 114) is in group 14 ⇒ IV

Tennessine (Ts, Z = 117) is in group 17 ⇒ II

That corresponds to Option C.

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# Question12

## Two statements are given below.

## Statement I : Nitrogen has more ionisation enthalpy and electronegativity than beryllium.

## Statement II : $CrO_3$, $B_2O_3$ are acidic oxide.

## Correct answer is

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**Options:**

A. Both Statement I and Statement II are correct.

B. Both Statement I and Statement II are not correct.

C. Statement I is correct but Statement II is not correct.

D. Statement I is not correct but Statement II is correct.

**Answer: A**

## Solution:

Statement I: "Nitrogen has more ionisation enthalpy and electronegativity than beryllium."

**Ionisation Enthalpy:** Nitrogen, being a non-metal and located in group 15, has a higher ionisation energy than beryllium, a metal in group 2. For example, the first ionisation energy of nitrogen is around 1402 kJ/mol, while for beryllium it is around 900 kJ/mol.

**Electronegativity:** Nitrogen is significantly more electronegative than beryllium. This reflects in their chemical behavior, as non-metals tend to attract electrons more strongly than metals.

**Conclusion for Statement I:** This statement is correct.

Statement II: "$CrO_3$, $B_2O_3$ are acidic oxide."

**Chromium Trioxide ($CrO_3$):** This oxide is acidic. It reacts with water to form chromic acid:

$CrO_3 + H_2O \rightarrow H_2CrO_4$.

**Diboron Trioxide ($B_2O_3$):** This oxide is also acidic. It reacts with water to yield boric acid:

$B_2O_3 + 3\,H_2O \rightarrow 2\,H_3BO_3$.

**Conclusion for Statement II:** This statement is also correct.

Since both statements are correct, the answer is:

Option A

Both Statement I and Statement II are correct.

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# Question13

## In which of the following sets, elements are not correctly arranged with the property shown in brackets?

## AP EAPCET 2024 - 22th May Morning Shift

**Options:**

A. $S > Se > O$ (Electron gain enthalpy)

B. $F > O > Cl$ (Electronegativity)

C. $Na > Li > Al$ (Metallic radius)

D. $Na > K > Ba$ (Metallic nature)

**Answer: D**

## Solution:

In option D, the elements are incorrectly arranged according to their metallic nature.

**Metallic Character:** This property increases as you move down a group in the periodic table and decreases as you move from left to right across a period.

Therefore, the correct order reflecting metallic nature is: $K > Ba > Na$.

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# Question14

## The correct order of atomic radii of $N, F, Al, Si$ is

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**Options:**

A. $F > N > Si > Al$

B. $F > N > Al > Si$

C. $Al > Si > F > N$

D. $Al > Si > N > F$

**Answer: D**

## Solution:

The general trend for atomic radius in the periodic table is as follows:

As we move down a group, the atomic radius increases. This happens because additional electron shells are added, which make the atoms larger.

As we move from left to right across a period, the atomic radius decreases. This decrease is due to the increased electrostatic force of attraction between the electrons and the nucleus, which pulls the electrons closer to the nucleus and reduces the size of the atom.

Based on these trends, the correct order of atomic radii for N, F, Al, and Si is:

$Al > Si > N > F$

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# Question15

## The correct order of atomic radii of given element is

## AP EAPCET 2024 - 21th May Morning Shift

**Options:**

A. $B < Be < Mg$

B. $Mg < Be < B$

C. $Be < B < Mg$

D. $B < Mg < Be$

**Answer: A**

## Solution:

Let's analyze the atomic radii trends for the elements involved:

In the same period (period 2), atomic radii decrease from left to right due to increasing nuclear charge.

Beryllium (Be) is to the left of Boron (B), but since Boron has an extra proton with a similar shielding effect, its electrons are pulled in tighter.

Therefore, in period 2: $B < Be$.

Comparing elements in different periods:

Magnesium (Mg) is in period 3, meaning it has an additional electron shell compared to Be and B.

More electron shells result in a larger atomic radius; hence, $Mg$ is larger than both $Be$ and $B$.

Combining these observations, the order from smallest to largest is:

$B < Be < Mg$.

Thus, the correct option is Option A.

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# Question16

## Which of the following orders is not correct against the given property?

## AP EAPCET 2024 - 20th May Evening Shift

**Options:**

A. $\mathrm{Ga} < \mathrm{In} < \mathrm{Tl} < \mathrm{Al} < \mathrm{B}$ - melting point

B. $\mathrm{Al} < \mathrm{Ga} < \mathrm{In} < \mathrm{Tl} < \mathrm{B}$ - Electronegativity

C. $\mathrm{B} < \mathrm{Al} < \mathrm{Ga} < \mathrm{In} < \mathrm{Tl}$ - density

D. $\mathrm{B} < \mathrm{Al} < \mathrm{Ga} < \mathrm{In} < \mathrm{Tl}$ - atomic radius

**Answer: D**

## Solution:

Atomic radii of group 13 elements are lower than those of alkaline earth metal of group 2 due to greater nuclear charge of group 13 element.

On moving down the group, atomic radius of Ga is slightly lower than of Al . This is due to the presence of $d$ -electrons in Ga which does not sheild the nucleus effectively.

Hence, the correct order of atomic radius is

$B < Ga < Al < In < Tl$

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# Question17

Match the following.

| List-I | | List-II | |
|---|---|---|---|
| A | Ionisation enthalpy | I | $P < Si < Mg < Na$ |
| B | Metallic character | II | $I < N < O < F$ |
| C | Electron gain enthalpy | III | $B < Be < C < O < N$ |
| D | Electronegativity | IV | $l < Br < F < Cl$ |

## AP EAPCET 2024 - 20th May Morning Shift

**Options:**

A. A-III, B-IV, C-I, D-II

B. A-III, B-I, C-IV, D-II

C. A-IV, B-II, C-I, D-III

D. A-IV, B-III, C-I, D-II

**Answer: B**

## Solution:

The correct match is A-III, B-I, C-IV, D-II.

(A) Electron gain enthalpy generally increases as we move left to right in the period. As we move from left to right the size of element

decreases and effective nuclear charge increase.

There is exception of Be and N .

$B < Be < C < O < N$.

(B) Metallic character decreases as We move from left to right in the period

$P < Si < Mg < Na$

(C) Generally electron gain enthalpy decreases as we go from top to bottom, but there is exception of d in group 17 element.

$I < Br < F < Cl$

(D) Electronegativity increases as we move from left to right and from top to bottom.

$I < N < O < F$

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# Question18

## Which of the following oxides is acidic in nature?

## AP EAPCET 2024 - 20th May Morning Shift

**Options:**

A. $GeO_2$

B. CO

C. $PbO_2$

D. SnO

**Answer: A**

## Solution:

| Oxide name | Nature of oxide |
|---|---|
| Germanium dioxide $(GeO_2)$ | Acidic |
| Carbon monoxide (CO) | Neutral |
| Lead dioxide $(PbO_2)$ | Amphoteric |
| Tin oxide $(SnO)$ | Amphoteric |

Hence, $GeO_2$ is acidic in nature.

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# Question19

**If the first ionisation enthalpy of $\mathrm{Li}, \mathrm{Be}$ and C respectively are $520, 899, 1086 \mathrm{~kJ~mol}^{-1}$. The first ionisation enthalpy (in $\mathrm{kJmol}^{-1}$ ) of B will be**

## AP EAPCET 2024 - 19th May Evening Shift

**Options:**

A. 487

B. 950

C. 801

D. 1402

**Answer: C**

## Solution:

Let's break down the reasoning:

The first ionisation energy is the energy needed to remove the outermost electron from an atom in its gaseous state.

The given values are:

Lithium (Li): 520 kJ/mol

Beryllium (Be): 899 kJ/mol

Carbon (C): 1086 kJ/mol

Moving across a period in the periodic table, ionisation energy generally increases. However, there is an anomaly between Be and B:

Beryllium (Be) has a $2s^2$ configuration, which is a stable, fully-filled subshell.

Boron (B), on the other hand, has a $2s^2 2p^1$ configuration. The electron in the $2p$ orbital is less strongly held (more shielded and farther from the nucleus) than the electrons in the $2s$ orbital, which leads to a lower ionisation energy than one might expect from a simple trend.

Experimentally, the first ionisation energy of boron is about 801 kJ/mol.

Therefore, the correct answer is:

Option C: 801 kJ/mol

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# Question20

## Match of following.

| List I (Element) | List II (Electron gain enthalpy (in $kJmol^{-1}$ ) |
|---|---|
| A F | I -141 |
| B Cl | II -328 |
| C O | III -200 |
| D S | IV -349 |

## The correct answer is

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**Options:**

A. A - II, B −IV,C−1,D− III

B. A=N,B−N,C−I,D−B

C. A− III, B−II,C=IN,D−1

D. A - I, B - ill, C - N, D - 1

**Answer: A**

## Solution:

The correct match is

A-II, B-IV, C-I, D-III.

Electron gain enthalpy is defined as the amount of energy released when an electron is added to an isolated gaseous atom. It increases from left to right in period. As we go down electronegativity becomes less negative.

$F \rightarrow -328\ kJ/mol$

$Cl \rightarrow -349\ kJ/mol$

$O \rightarrow -141\ kJ/mol$

$S \rightarrow -200\ kJ/mol$

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# Question21

# Identify the basic oxide from the following.

## AP EAPCET 2024 - 18th May Morning Shift

**Options:**

A. $C_2O_3$

B. $CrO_3$

C. $V_2O_5$

D. $V_2O_3$

**Answer: D**

## Solution:

We'll analyze each oxide based on its chemical behavior regarding acid–base properties:

$C_2O_3$

Compounds of carbon with oxygen are generally acidic. In this case, although $C_2O_3$ is not commonly encountered, nonmetal oxides tend to be acidic rather than basic.

$CrO_3$

This is a well-known oxide of chromium in a high oxidation state. It is acidic in nature (in fact, it reacts with water to eventually form chromic acid).

$V_2O_5$

Vanadium(V) oxide is also in a high oxidation state. Like many high oxidation state transition metal oxides, it behaves as an acidic oxide (or sometimes amphoteric but predominantly leaning towards acidic behavior).

$V_2O_3$

Vanadium(III) oxide, with vanadium in a lower oxidation state, tends to be a basic oxide. In transition metals, as the oxidation state decreases, the oxide becomes more basic because the metal acts more like a typical metal ion.

Thus, the basic oxide among the options is:

$\boxed{V_2O_3}$

In summary:

Nonmetal oxides like $C_2O_3$ are acidic.

High oxidation state metal oxides such as $CrO_3$ and $V_2O_5$ are acidic.

The lower oxidation state metal oxide, $V_2O_3$, exhibits basic properties.

So, the correct answer is Option D: $V_2O_3$.

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# Question22

**In which of the following oxides of three elements $X, Y$ and $Z$ are correctly arranged in the increasing order of acidic nature. The electronic configurations of $X, Y$ and $Z$ are $[\mathrm{Ne}]3s^2 3p^1$, $[\mathrm{Ne}]3s^2 3p^5$. [Ne] $3s^2$ respectively**

## AP EAPCET 2022 - 5th July Morning Shift

**Options:**

A. $X < Y < Z$

B. $Y < Z < X$

C. $Z < X < Y$

D. $X < Z < Y$

**Answer: C**

## Solution:

| Element | Electronic configuration |
|---|---|
| $X$ | $[\mathrm{Ne}]3s^2 3p^1$ |
| $Y$ | $[\mathrm{Ne}]3s^2 3p^5$ |
| $Z$ | $[\mathrm{Ne}]3s^2$ |

According to given electronic configuration, $X$ is Al, $Y$ is Cl and $Z$ is Mg . Along a period, the basic nature of oxides decreases and acidic nature of oxides increases. Mg, Al and Cl belong to same period with Mg on left side of the period, Cl on the right side and Al in between them.

$\therefore$ Oxide of Cl is i.e. $Y$ will be most acidic oxide of Mg , i.e. Z will be basic and oxide of Al , i.e. $X$ in between the other two.

$\therefore$ Correct increasing order acidic nature will be $Z < X < Y$

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# Question23

## Assertion (A) 16 th group elements have higher ionisation enthalpy values than 15 th group elements in the corresponding periods.

## Reason (R) 15 th group elements have half-filled stable electronic configurations.

## AP EAPCET 2022 - 5th July Morning Shift

**Options:**

A. Both A and R are correct and R is the correct explanation of A.

B. Both A and R are correct but R is not the correct explanation of A.

C. A is correct but R is incorrect

D. A is incorrect but R is correct

**Answer: D**

## Solution:

**Assertion (A):** 16th group elements have higher ionisation enthalpy values than 15th group elements in the corresponding periods.

**Reason (R):** 15th group elements have half-filled stable electronic configurations.

To analyze this, let's review the concepts of ionisation enthalpy and electronic configurations:

**Ionisation Enthalpy:** Ionisation enthalpy is the energy required to remove an electron from an isolated gaseous atom. Higher ionisation enthalpy indicates it is more difficult to remove an electron.

**Electronic Configuration:** The 15th group elements (nitrogen group) have half-filled p orbitals, which are considered to be more stable due to exchange energy and symmetry. This stability can lead to slightly higher ionisation enthalpies compared to their neighboring elements.

However, the statement in Assertion (A) indicates that the 16th group elements have higher ionisation enthalpy values than the 15th group elements. Let's provide a detailed explanation:

In the 15th group, the stability due to the half-filled p orbital configuration results in relatively higher ionisation enthalpies compared to the 16th group, where such stability is not present. Thus, the 15th group elements typically have higher ionisation enthalpies than the 16th group elements.

Therefore, Assertion (A) is **incorrect**.

The Reason (R) correctly states that 15th group elements have half-filled stable electronic configurations, leading to relatively higher ionisation enthalpy.

Thus, Reason (R) is **correct**.

**Correct option:** D. A is incorrect but R is correct.

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# Question24

## Assertion (A) Fluorine has smaller negative electron gain enthalpy than chlorine.

## Reason (R) The electron-electron repulsion is higher in chlorine than in fluorine.

## AP EAPCET 2022 - 5th July Morning Shift

**Options:**

A. Both A and R are correct and R is the correct explanation of A .

B. Both A and R are correct but R is not the correct explanation of A .

C. A is correct but R is incorrect

D. A is incorrect but R is correct

**Answer: C**

## Solution:

To evaluate the Assertion (A) and Reason (R) and determine the correct option, we need to understand the concepts of electron gain enthalpy and the factors affecting it.

**Assertion (A): Fluorine has a smaller negative electron gain enthalpy than chlorine.**

Electron gain enthalpy refers to the enthalpy change when an electron is added to a neutral atom to form a negative ion. A more negative electron gain enthalpy indicates a greater tendency to gain an electron. Fluorine has a smaller atomic radius than chlorine, which means that the added electron in fluorine is closer to the nucleus and experiences a stronger attraction. Therefore, technically, one would expect fluorine to have a more negative electron gain enthalpy. However, due to the very small size of the fluorine atom, the inter-electronic repulsion in the compact 2p orbital is quite high. On the other hand, these repulsions are slightly lesser in the 3p orbitals of chlorine, making the overall electron gain enthalpy of chlorine more negative than that of fluorine. Hence, Assertion (A) is correct.

**Reason (R): The electron-electron repulsion is higher in chlorine than in fluorine.**

As discussed, the electron-electron repulsion is actually higher in fluorine due to its smaller size and the compact nature of its orbitals. Chlorine, with its larger size, experiences lesser electron-electron repulsion when an additional electron is added. Hence, Reason (R) is incorrect.

Based on this discussion, the correct evaluation of Assertion (A) and Reason (R) is given by:

**Option C**

A is correct but R is incorrect

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# Question25

## In which of the following, elements are arranged in the correct order of their electron gain enthalpies?

## AP EAPCET 2022 - 4th July Evening Shift

**Options:**

A. $F > S > O > N$

B. $F > O > S > N$

C. $F > O > N > S$

D. $F > N > O > S$

**Answer: A**

## Solution:

Electron gain enthalpy increases across a period on moving from left to right. So, F has maximum electron gain enthalpy whereas N has minimum electron gain enthalpy.On moving down the group, electron gain enthalpy decreases but from 2 nd period to 3 rd period, it increases because of high electronic repulsions in 2nd period. Hence, the correct order is

$F > S > O > N$

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# Question26

**In second period of the long form of the periodic table an element $X$ has second lowest first ionisation enthalpy and element $Y$ has second highest first ionisation enthalpy values. What are $X$ and $Y$ ?**

## AP EAPCET 2022 - 4th July Evening Shift

**Options:**

A. B, F

B. Be, Ne

C. Be, O

D. C, O

**Answer: A**

## Solution:

The order of first ionisation enthalpy across period 2 is

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

So, $X$ element having second lowest first ionisation enthalpy is boron (B).

$Y$ element having second highest first ionisation enthalpy is fluorine ( F ).

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# Question27

**Lithium shows diagonal relationship with element '$X$' and aluminium with $Y$. $X$ and $Y$ respectively are**

## AP EAPCET 2022 - 4th July Morning Shift

**Options:**

A. Mg, Be

B. Be, Mg

C. Na, Si

D. B, Be

**Answer: A**

## Solution:

Diagonal relationship exists between certain pairs of diagonally adjacent elements in the second and third period of the periodic table. It occurs because of the directions of trends in various properties as we move across the period or down the periodic table.

Lithium shows diagonal relationship with magnesium whereas aluminium with beryllium. Hence, the element $X$ and $Y$ are Mg and Be respectively.

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# Question28

## The correct order of the metallic character of the elements Be, Al, Na, K is

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**Options:**

A. $K > Na > Al > Be$

B. $K > Al > Na > Be$

C. $Al > K > Na > Be$

D. $Na > K > Be > Al$

**Answer: A**

## Solution:

Metallic character decreases on moving from left to right across a period. So, sodium ( Na ) is more metallic than aluminium ( Al ).

Also, metallic character increases on moving down the group. Thus, potassium (K) is more metallic than sodium (Na).

As Be lies in 2nd period, so it is least metallic.

So, the correct order of the metallic character is

$K > Na > Al > Be$

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# Question29

## Which of the following tetrahalides does not exist?

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**Options:**

A. $CCl_4$

B. $SiCl_4$

C. $PbCl_4$

D. $Pbl_4$

**Answer: D**

## Solution:

Tetrahalide, $PbI_4$ does not exist because iodine is a good reducing agent and it reduces $Pb^{4+}$ to $Pb^{2+}$. So, the compound formed is $PbI_2$ not $PbI_4$.

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# Question30

## The electronegativity of the given elementsincreases in the order.

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**Options:**

A. C, N, Si, P

B. N, Si, C, P

C. Si, P, C, N

D. P, Si, N, C

**Answer: C**

## Solution:

Electronegativity is the tendency to attractshared pair of electrons.

Electronegativity of elements generally increasesacross the period and decreases down the group.Si = 1.8, P = 2.1, C = 2.5 and N = 3.0

So, the correct order is

Si < P < C < N

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# Question31

## The first ionisation enthalpies of $\mathrm{Mg}$ and $\mathrm{Al}$ can be expected to be ..... .

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**Options:**

A. $\mathrm{IE}_1(\mathrm{Mg}) = 577.5\ \mathrm{kJ\ mol}^{-1}, \mathrm{IE}_1(\mathrm{Al}) = 577.5\ \mathrm{kJ\ mol}^{-1}$

B. $\mathrm{IE}_1(\mathrm{Mg}) = 577.5\ \mathrm{kJ\ mol}^{-1}, \mathrm{IE}_1(\mathrm{Al}) = 737.5\ \mathrm{kJ\ mol}^{-1}$

C. $\mathrm{IE}_1(\mathrm{Mg}) = 737.7\ \mathrm{kJ\ mol}^{-1}, \mathrm{IE}_1(\mathrm{Al}) = 737.7\ \mathrm{kJ\ mol}^{-1}$

D. $\mathrm{IE}_1(\mathrm{Mg}) = 737.7\ \mathrm{kJ\ mol}^{-1}, \mathrm{IE}_1(\mathrm{Al}) = 577.5\ \mathrm{kJ\ mol}^{-1}$

**Answer: D**

## Solution:

Magnesium ( Mg ) has the first ionisation enthalpy higher than that of Al. Due to fulfilled *s*-orbital of Mg it requires higher energy to remove the electron from outermost shell. On the other hand, Al can easily lose one electron as its *p*-orbital has only one electron.

Therefore,

$IE_1(Mg) = 737.7 kJmol^{-1}$
$IE_1(Al) = 577.5\ kJ\ mol^{-1}$

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# Question32

## Which of the following statements is/are correct?

## 1. Mercury is the only metal that exists as liquid at room temperature.

## 2. Among non-metals, carbon has the highest melting point.

## 3. Hydrogen is the most abundant element in the universe.

## 4. Oxygen is the most abundant element in the Earth's crust.

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**Options:**

A. 1

B. 2

C. 3

D. 1, 2, 3, 4

**Answer: D**

## Solution:

**1. Mercury is the only metal that exists as liquid at room temperature.**

✅ **Correct.**

Mercury is indeed the only metal that is liquid at room temperature (~25°C). Gallium also melts just above room temperature (at about 29.7°C), but it is solid under normal room conditions.

**2. Among non-metals, carbon has the highest melting point.**

✅ **Correct.**

Carbon (in the form of diamond) has the highest melting (or sublimation) point among non-metals — around 3550°C.

**3. Hydrogen is the most abundant element in the universe.**

✅ **Correct.**

Hydrogen accounts for about 75% of the elemental mass of the universe.

**4. Oxygen is the most abundant element in the Earth's crust.**

✅ **Correct.**

Oxygen constitutes about 46% by mass of the Earth's crust, making it the most abundant element there.

✅ **Answer: Option D (1, 2, 3, 4) — All are correct.**

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# Question33

## The element with outer electronic configuration $(n-1)d^2ns^2$, where $n=4$, would belong to

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**Options:**

A. 2nd period, 2nd group

B. 4th period, 4th group

C. 4th period, 2nd group

D. 2nd period, 4th group

**Answer: B**

## Solution:

For electronic configuration $=(n-1)d^2ns^2$

Sum of electrons present in outer most orbital ($d$ and $s$ ) denoted group and period position is represented by principal quantum number.

e.g., $(n-1)d^2ns^2$

Group = number of valence electrons in $d$ and $s$-orbital.

Sum of electrons in $d$ and $s$ orbital $= 2 + 2 = 4$

$\therefore$ Number of electrons in $d$-orbital and valence shell $= 4$

$\therefore$ 4th group and 4th period.

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# Question34

## Choose the correct option regarding the following statements

## Statement 1 Nitrogen has lesser electron gain enthalpy than oxygen.

## Statement 2 Oxygen has lesser ionisation enthalpy than nitrogen.

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**Options:**

A. Statement 1 is correct but statement 2 is incorrect.

B. Both statements 1 and 2 are incorrect.

C. Both statements 1 and 2 are correct.

D. Statement 1 is incorrect but statement 2 is correct.

**Answer: C**

## Solution:

The outermost electronic configuration of nitrogen $\left(2s^2 2p_x^1 2p_y^2 2p_z^1\right)$ is very stable because $p$-orbital is half-filled. The addition of an extra electrons to any of the $2p$-orbitals require energy as it breaks the stability of N. Oxygen has 4 electrons in $2p$-orbital and acquires stable configuration i.e. $2p^3$ after removing one electron and complete octet after gaining two electrons. A nitrogen has positive electron gain enthalpy, whereas oxygen has negative, however oxygen has lower ionisation enthalpy than nitrogen.Hence, both statements 1 and 2 are correct.

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# Question35

## Among the given configurations, identify the element which does not belong to the same family as the others?

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**Options:**

A. $[\mathrm{Ne}]3s^2 3p^5$

B. $[\mathrm{Ar}]3d^{10}4s^2$

C. $[\mathrm{Kr}]4d^{10}5\ \mathrm{s}^2$

D. $[\mathrm{Xe}]4f^{14}5d^{10}4s^2$

**Answer: A**

## Solution:

In contrast element neon [Ne] has seven electrons in the valence shell and hence, does not i.e. in the same group as the other 3 elements i.e. [Ar], [Kr], [Xe].

In a family all elements have same outermost electronic configuration. Since, $[\mathrm{Ne}]3s^2 3p^5$, chlorine belongs to halogen family while the remaining three are in same group i.e. group 12 (noble gas ).

$$\left.\begin{aligned} {}_{80}\mathrm{Hg} &= [\mathrm{Xe}]4f^{14}5d^{10}6s^2 \\ {}_{48}\mathrm{Cd} &= [\mathrm{Kr}]4d^{10}5s^2 \\ {}_{30}\mathrm{Zn} &= [\mathrm{Ar}]3d^{10}4s^2 \end{aligned}\right\} \text{Group-12}$$

$${}_{17}\mathrm{Cl} = [\mathrm{Ne}]3s^2 3p^5$$

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# Question36

## Which statements among the following are correct about helium?

## (i) Liquid helium is used to sustain powerful superconducting magnets.

## (ii) Liquid helium is useful to carry low temperature experiments.

**(iii) It is a heavy gas.**

**(iv) It is a flammable gas.**

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**Options:**

A. (i) and (ii)

B. (ii) and (iii)

C. (i) and (iv)

D. (iii) and (iv)

**Answer: A**

## Solution:

Helium is as a safe, non-flammable gas to fillin balloons and parade balloons. It is lighter inweight. Liquid helium is used for MRI system. Itis needed as are refrigerant and superconductingappliances. It in useful to carry low temperatureexperiments in research.

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# Question37

## Which of the following represents the correctorder of increasing electron gain enthalpywith negative sign for the elements?

## 1. Nitrogen (N) 2. Phosphorus (P) 3. Chlorine (Cl) 4. Fluorine (F)

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**Options:**

A. P < N < F < Cl

B. N < P < F < Cl

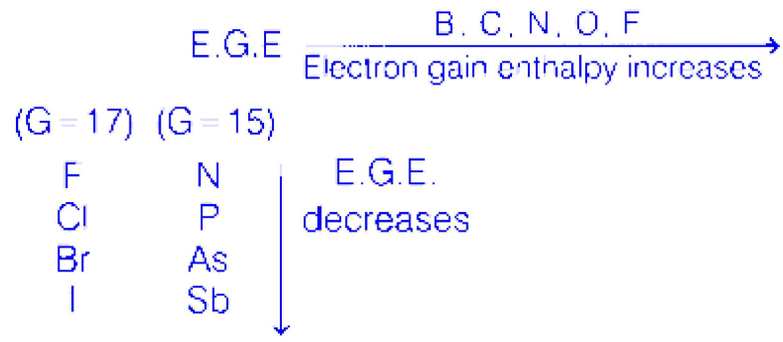
C. Cl < F < P < N

D. F < Cl < N < P

**Answer: B**

## Solution:

Electron gain enthalpy, (EGE) increases inperiod from left to right and decreases in group onmoving downward.




Here, in group 17, chlorine having very highelectron gain enthalpy (349 kJ/mol). Due to highelectronegativity of fluorine, it has second highelectron gain enthalpy (328 kJ/mol).

Order of increasing electron gain enthalpy (in kJ/mol)

$$\underset{(7)}{N} < \underset{(72)}{P} < \underset{(328)}{F} < \underset{(349)}{Cl}$$

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