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UNIQUE NOTES

CHEMISTRY 11

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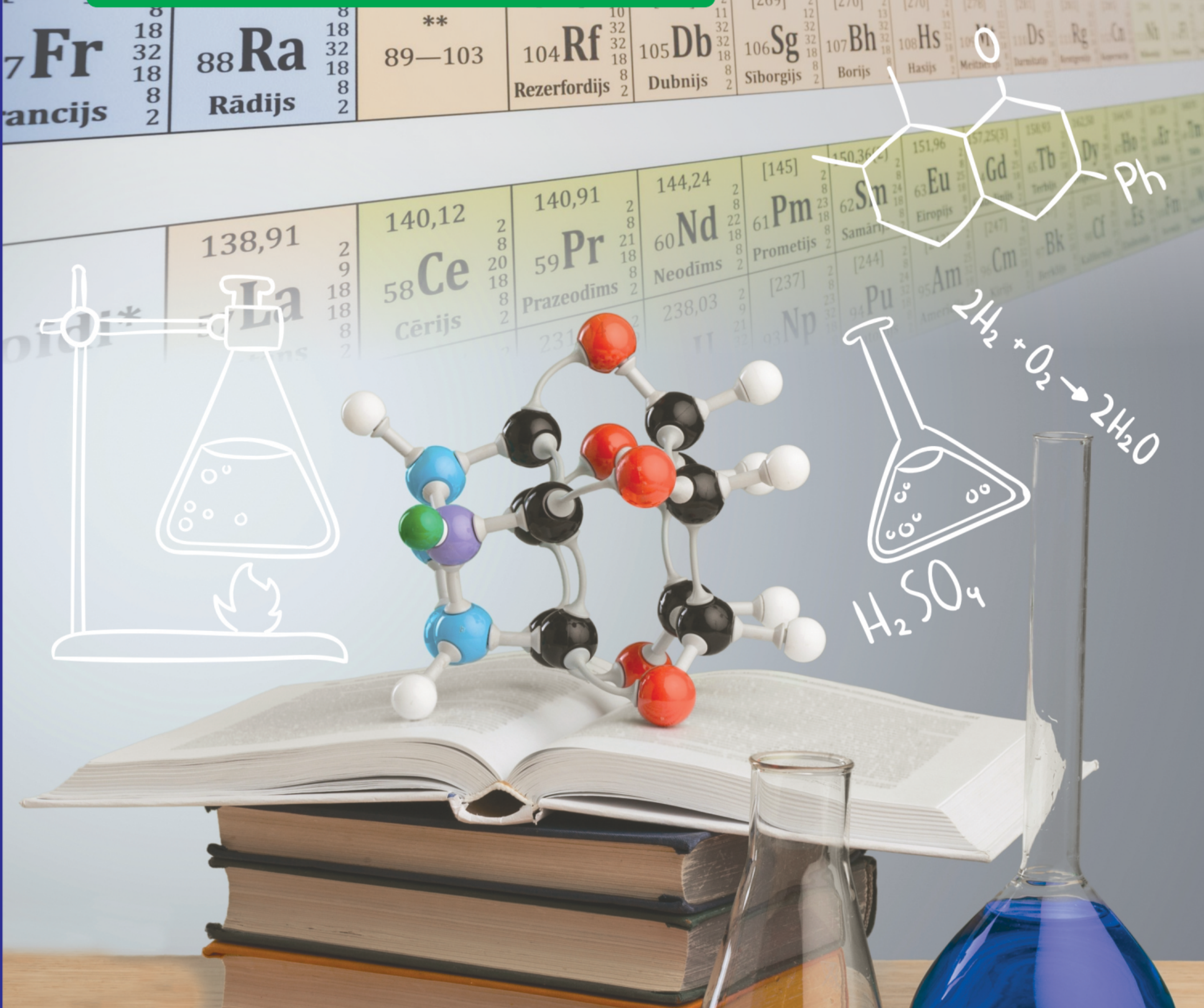

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Unit 01 PERIODIC TABLE AND PERIODIC PROPERTIES

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DESCRIPTIVE QUESTIONS

HISTORICAL BACKGROUND

Q1. Why periodic table was created? Explain historical background of periodic table. **11201001**

Ans. Creation of the periodic table

The reasons for the creation of periodic table are:

- **Symbol of chemistry:** The periodic table of elements is considered as the symbol of chemistry.
- **Systematic categorization:** The systematic categorization and arrangement of elements in periodic table provides easy way to study the wide subject of chemistry.
- **Periodicity:** The main reason for the creation of periodic table is periodicity of elements and their compounds. Periodicity means that properties of arranged elements are repeated after regular intervals.
- **Turning point:** One of the most important turning point in the history of science is the creation of periodic table.
- **Tabular form:** All the 118 elements are arranged in tabular form in the current periodic table based on their atomic number, electronic configuration, and recurrent chemical characteristics.
- **Framework:** The periodic table offers a framework for researching the periodic behaviours of elements is a significant accomplishment.

Historical background of Periodic table

(i) Dobereiner's Triads

In 1829, Dobereiner grouped the elements into triads (a group of three elements) with similar properties, noticing that the atomic weight of the middle element was roughly the average of the other two.

Examples: Lithium, sodium and potassium (^{7}Li, ^{23}Na, ^{39}K) are included in the same triad.

(ii) Newland's law of Octaves

- English chemist John Newland, in 1864, first time observed periodicity in the 62 known elements.

- He studied that the properties of every eighth element were similar when arranged by the increasing order of their atomic masses.
- He classified the elements into groups so that every eight element resembled the first element in properties.

(iii) Lother Meyer's curves

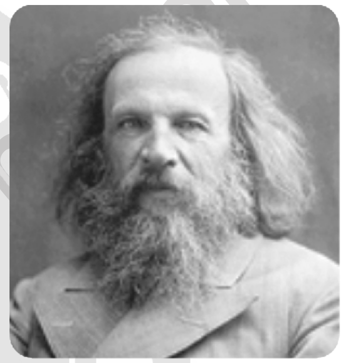
Lother Meyer developed his famous curves by plotting a graph between the atomic weights and atomic volumes of elements in 1869. Theses curves also showed periodicity.

(iv) Mendeleev's Periodic table

Russian chemist Dmitri Mendeleev considered the father of the Periodic Table. He arranged 63 elements in his periodic table named Mendeleev's Periodic table in 1869. Some features of his table are:

Mendeleev's periodic law: The properties of the elements are periodic functions of their atomic masses.

- **Groups:** The vertical columns of elements in the periodic table are called groups. He arranged elements in eight groups in his periodic table.
- **Periods:** The horizontal rows of elements in the periodic table are called periods. He arranged elements in seven periods.
- **Gaps:** The success of his table was hidden in leaving gaps for undiscovered elements and predicting their atomic masses and properties which proved accurate when these elements were practically found.



Dmitri Mendeleev arranged elements according to their atomic masses and his table was the first most notable effort in the classification of elements.

(v) Modern Periodic table

In 1913, Moseley determined the exact atomic numbers of known elements using X-ray emission, resolving flaws and discrepancies in Mendeleev's table by arranging the elements by atomic numbers instead of atomic masses.

The Modern Periodic law: The modern periodic law states that "the properties of elements are periodic functions of their atomic numbers".

MODERN PERIODIC TABLE

Q2. Explain the main features of Modern periodic table. 11201002

Ans. Main features of the Modern Periodic table

The classification of elements in the modern periodic table helps in the easier understanding of their properties.

Atomic number: All the 118 elements are arranged in groups and periods in modern period the table in ascending order of their respective atomic numbers.

Groups and Periods: There are seven horizontal rows called periods and eighteen vertical columns called groups. (In older versions of the table, there were 18 vertical groups were divided into two types of groups: Eight A-Groups and Ten B-Groups).

Properties: In the periodic table, elements within the same group exhibit similar **chemical** properties because they have the same number of valence electrons. However, they show a gradual change in **physical** properties from top to bottom in a group.

Valance electrons: Elements in a period show a gradual change in properties moving from left to right in periods, due to variation in number of valance electrons.
Representative and Transition elements: Main group elements (group 1, 2, 13 – 18) and transition elements (group 3 – 12).
Metals, Non-metals and Metalloids: Most of the elements are metals in periodic table, some of them are non-metals, only few of them are metalloids.
Block and Families: Other than groups and periods in the periodic there are different ways of grouping the elements into various blocks, families and categories just to enhance understanding.

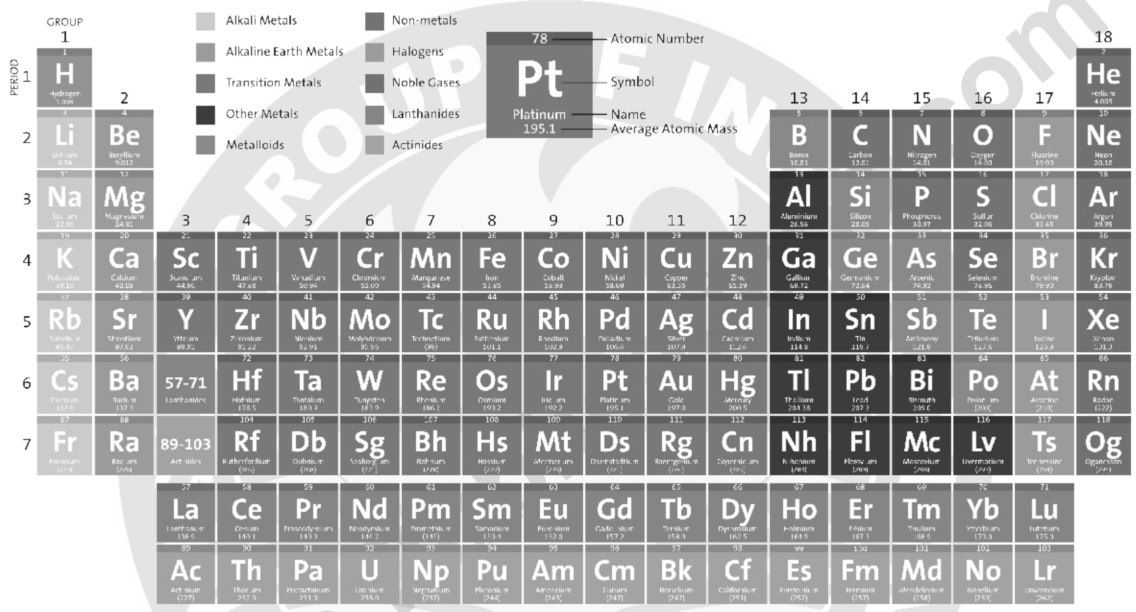


Fig: Modern Periodic Table

METALS, NON-METALS AND METALLOIDS

Q3. Explain the classification of elements as metals, non-metals and metalloids and blocks in periodic table. **11201003**

Ans. Classification of elements

Elements can be broadly classified as metals, non-metals and metalloids.

(i) Metals

Definition: Metals are elements which tend to lose electrons to form positive ions.
Location: Elements to the left of the stair – step line in periodic table are considered as metals.
Examples: Iron, Copper, Gold and Silver etc. are metal elements in periodic table.

(ii) Non-Metals

Definition: Non-metals are elements which tend to gain electrons to form negative ions.
Location: Elements to the far right side of the stair – step line in the periodic table are non-metals. The exception is hydrogen, the first element of the periodic table.
Examples: Chlorine, Sulfur and Phosphorous, Carbon and Hydrogen etc. are non-metal elements in periodic table.

(iii) Metalloids

Definition: The metalloids exhibit some properties of metals and some of non-metals.

Location: The metalloids separate the metals and non-metals on a periodic table. Mostly periodic tables have a "stair-step line" on the table identifying the element groups. The line begins at boron (B) and extends down to polonium (Po) including Si, Ge, As, Sb and Te.
Examples: Si, Ge, As and Sb etc. are metalloid elements in the periodic table.

BLOCKS IN PERIODIC TABLE

Blocks in Periodic table

Definition: Elements in the periodic table can be classified based on the subshells containing their valence electrons.
Example: The blocks of periodic table are given in the following:

s-Block

- The valence electrons of elements in the first two groups (1 and 2) are in the "s" subshells, so these elements are placed in the s-block.
- It is present at extreme left side of the periodic table.

p-Block

- The elements in groups 13 to 18 having their valence electrons in their "p" subshells so they are included in p-block.
- It is located at the right side of the periodic table.

d-Block

- The elements in groups 3 to 12 having "d" subshells in the process of filling so they are included in d-block, and they are called transition elements.
- Transition means in between two things, as these elements are located in between s and p blocks so they are given the name transition elements.

f-Block

- The elements in the two series at the bottom of the table (known as Lanthanides and Actinides) are categorized as f-block elements.
- Their "f" subshells are in the process of filling.

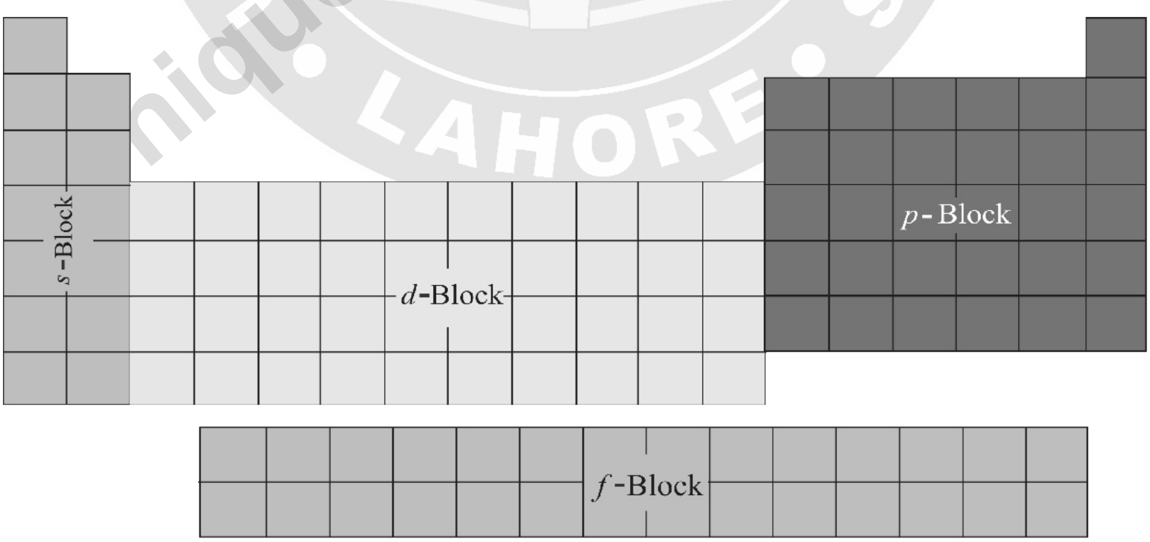


Fig: Blocks in periodic table

FAMILIES IN PERIODIC TABLE

Q4. What types of families are present in the periodic table? 11201004

Ans. Families in Periodic table

Definition: An element family is a set of elements sharing common properties are placed in the same group of periodic table.

There are six famous families of elements in the periodic table:

- Alkali metals (Li, Na, K, Rb, Cs, Fr)
- Alkaline earth metals (Be, Mg, Ca, Sr, Ba, Ra)
- Transition metals (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn etc.)
- Chalcogens (O, S, Se, Te, Po, Lv)
- Halogens (F, Cl, Br, I, At, Ts)
- Noble gases (He, Ne, Ar, Kr, Xe, Rn, Og)

(i) Alkali Metals

Definition: Elements in the group 1 of the periodic table are known as alkali metals because they produce alkalis when they react with water.

Example: Sodium and potassium are notable examples of these elements.

Characteristics

- Alkali metals are characterized by one valence electron.
- Alkali metals show low densities, relatively low melting points, and low ionization energies.
- These are the most reactive metals.
- They are soft metals.

(ii) Alkaline Earth Metals

Definition: Group 2 elements are metals primarily found in the earth and form alkalis, hence they are referred to as alkaline earth metals.

Example: Calcium and magnesium are notable the examples of alkaline earth metals.

Characteristics

- These elements have two electrons in their valence shell, making them divalent.
- They are metallic solids that are harder and denser than alkali metals.
- Easily oxidized with high thermal and electrical conductivities.
- They are soft metals but harder than alkali metals.

(iii) Transition Elements

Definition: The transition metals make up the largest family of elements in the middle of periodic table.

Example: They include four series of d-block elements (groups 3 to 12), as well as the lanthanides and actinides (f-block elements) found in the two rows at the bottom of periodic table.

Characteristics

- They exhibit high thermal and electrical conductivities, high melting points, high density, and variable oxidation states.
- They mostly form coloured compounds.
- They form complex compounds.

(iv) Chalcogens

Definition: The group 16 elements are called chalcogens because most ores of copper (Greek Chalkos = ores, gen = forming) are oxides or sulphides.

Example: In this group, oxygen & sulphur are non-metals, Se, Te, Po are metalloids and Livermorium is a metal.

Characteristics

- They have six valance electrons.
- They mostly show – 2 oxidation state.

(v) Halogens

Definition: Elements in group 17, known as halogens are non-metallic. The term "halogen" means "salt-former" because these elements easily react with alkali metals and alkaline earth metals to form stable halide salts.

Example: Fluorine, chlorine, bromine and iodine etc. are halogens.

Characteristics

- Halogens are highly reactive non-metals with high electron affinities.
- Halogens can easily accept one electron to complete their outermost shell.

(vi) Noble gases

Definition: The noble gases are a group of unreactive elements present at the extreme right of the periodic table in Group 18.

Example: Helium, argon, krypton etc. are noble gases.

Characteristics

- Stable electronic configuration (complete outermost shell).
- They are almost entirely unreactive under normal conditions and rarely form compounds with other elements.
- These elements are mono-atomic in nature.

Keep in Mind

S.Q. How the word inert gases changes into noble gases?

Ans. Although, noble gases are unreactive, however they have some compounds. An example is compounds of xenon such as xenon hexafluoroplatinate ($XePtF_6$), the word inert gases was changed to noble gases.

Quick Check 1.1

(a) Why are the elements in Groups 1 and 2 known as s-block elements?

Ans. Elements of groups 1 & 2 have valance electrons in "s" sub-shell, so they are called s-block elements.

Examples: Na, Mg are s-block elements

(b) Name the elements in the chalcogen family. Give their two characteristics.

Ans. The elements present in chalcogen family are: O, S, Se, Te, Po and Lv.

Characteristics

- All they have 6 electrons in their valance shells.
- All they can accept 2 electrons and form -2 ions.

PERIODIC ARRANGEMENTS AND ELECTRONIC CONFIGURATION

Q5. Explain how elements are arranged in periodic table according to their electronic configuration? 11201005

Ans. Electronic configuration

Definition: The arrangement of electrons in sub-shells or orbitals around the nucleus in an atom is called electronic configuration.

Periodic arrangement: The periodic arrangement of elements in the periodic table provides information about the physical properties, such as their physical state and atomic radii as well as their electronic structure and chemical reactivity.

Electronic configuration helps in the following ways in arrangement of periodic table.

- **Period Number:** The period number indicates the principal quantum number (n), representing the number of electronic shells surrounding the nucleus.
  Example: An element **X** in the 3rd period have three shells, with its valence electrons located in the 3rd shell.
- **Sub-shell and block:** The specific subshell where the valence electrons are found, depends on the element's block (azimuthal quantum number). If an element **X** in the 3rd period is in the s-block, its valence electrons are in the 3s subshell.
- **Group Number:** The group number indicates the number of valence electrons.
  Example of Mg: An element **X** in the 3rd period and group 2 has two valence electrons in its outermost shell. Thus, the element Z in the 3rd period and group 2 (s-block) has two valence electrons in the 3s subshell, which means that **X** would be magnesium (Mg).
  Example of Al: Another example to relate period number and group number with electronic configuration and position of element in period table.

X belongs to group 13 and period 3
The element X belongs to group 13 of periodic table so it has 3 valence electrons; and it is found in period 3 so it has three shells around its nucleus.
It means that the 3 valence electron are in the 3rd shell which means element is Al.
The configuration will be: $1s^2$ $2s^2\ 2p^6$ $3s^2\ 3p^1$

$$\underset{1^{st}\text{ shell}}{\underset{\downarrow}{1s^2}}\quad \underset{2^{nd}\text{ shell}}{\underset{\downarrow}{2s^2\,2p^6}}\quad \underset{3^{rd}\text{ shell}}{\underset{\downarrow}{3s^2\,3p^1}}$$

- **Chemical Properties:** Understanding the periodic arrangement of elements provides an explanation of an element's electronic configuration, which is essential for understanding its chemical properties and behavior.

Quick Check 1.2

(a) X belongs to group 14 and period 2?
(i) Write electronic configuration of the element X.
(ii) Identify block of the element. Identify this element from periodic table.

Ans. X belongs to 14^{th} group and 2^{nd} period

- Group 14 shows that it is in IV-A group
- IV-A group present in p-block
- An element present in IV-A of period 2 means, it has two shells only
- So, X = $1s^2\quad 2s^2\quad 2p^2$
  2^{nd} period
- $2p^2$ shows that X is in 2^{nd} group of p-block
- X is C = $1s^2\quad 2s^2\quad 2p^2$ So the element X is carbon ($_6C$)

(b) Identify an element that is in period 4 Group 17?

Ans.

- Period 4 shows that it has four shells
- Group 17 means it is present in 5^{th} group of p-block
- 5^{th} group of p-block is halogen
- A halogen atom present in 4^{th} period is "Br".
- Br = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5 \longrightarrow$ 5^{th} group of p-block

  $\downarrow$
  4^{th} period of periodic table

ATOMIC RADIUS

Q6. Define atomic radius and ionic radius. Give their factors and trends in the periodic table. **11201006**

Ans. Atomic radius

Definition: The atomic radius is a measure of the size of an atom, it is half of the distance between two identical atoms bonded together.

- The atomic radius can vary depending on the type of bond (covalent, Metallic & van der Waal's force) or the state of the atom.
- The radius can be different in a covalent bond compared to an ionic bond.

Example: The atomic radius of sodium atom is 186pm.

Unit: The atomic radius is typically measured in picometers (pm) or Angstroms (Å).

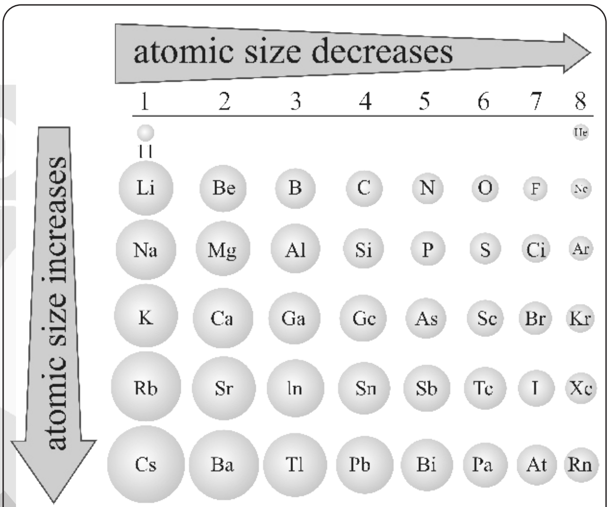


Fig: Variation in atomic radius across periods and down the groups

Factors affecting atomic radius

The factors affecting the atomic radius are

(i) Atomic number
(ii) Effective nuclear charge
(iii) Shielding effect of inner electrons.

Periodic trends in atomic radius:

Variation in periods

Atomic radius decreases across a period (from left to right) in the periodic table.

Reason: Increasing nuclear charge in period from left to right, which pulls the electronic cloud closer, results in decrease in atomic radius.

Variation in group

Atomic radius increases down a group (from top to bottom) in the periodic table.

Reason: The additional electronic shells are added moving down the group, so more shielding makes the atom larger despite the increase in nuclear charge (which is outweighed), as a result atomic radius increases.

IONIC RADIUS

Ionic radius

Definition: It is defined as the distance form the nucleus of an ion to the outermost electron shell.

Unit: It is measured in picometers (pm) or angstroms (Å).

Example: The cationic radius of K^{+1} is 133 pm.

Cationic radius: When an atom loses one or more electrons to become a positive ion, it generally becomes smaller than the neutral atom.

Reason: The loss of electrons reduces electronic repulsion and allows the remaining electrons to be pulled closer to the nucleus, so cationic radius is generally smaller than the parent atom.

Anionic radius: When an atom gains one or more electrons, it forms an anion, it generally becomes larger than the neutral atom.

Reason: The addition of electrons increases electronic repulsion, as a result the nuclear pull on electrons decreases and the electron cloud expands, so the anionic radius is larger than its parent atom.

Periodic trends in ionic radius:

Variation in period

When we move across a period from left to right, the ionic radius of cations and anions decreases.

Reason: Increasing nuclear charge which pulls the electrons closer for cations. The anionic radius also decreases across a period because the increasing nuclear charge also pulls the electrons closer to the nucleus.

Variation in group

Both cations and anions increase in size as we move down a group.

Reason: The principal quantum number (n) increases, leading to an increase in the number of electron shells. Consequently, the distance between the nucleus and the outermost electrons becomes larger, outweighing the effect of increased nuclear charge. The additional electron shells make the ions larger.

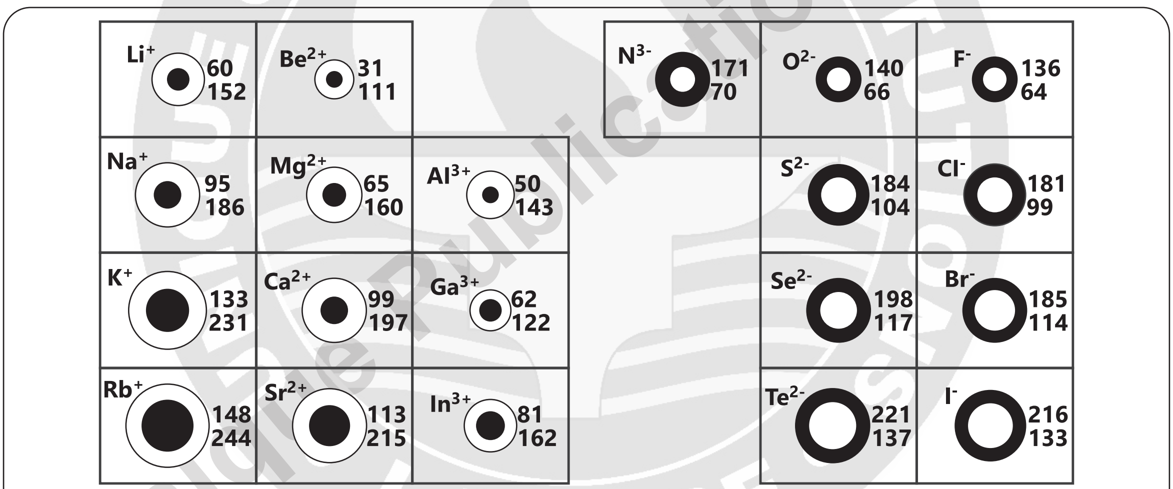


Fig: Variation in Ionic Radius

Quick Check 1.3

(a) Which factors affect atomic and ionic radii?

Ans. Factor affecting an atomic and ionic radii are:

- Atomic number
- Effective nuclear charge
- Shielding effect

Increase in atomic number is responsible in increase in effective nuclear charge. So, atomic and ionic radii decrease in period and increase in groups.

Shielding effect

Shielding effect of inner electrons in period remain same while in groups it increases from top to bottom. So, radii increase in groups and decrease in periods.

(b) Using your knowledge of Period 2 elements, predict and explain the relative sizes of:

(i) The atomic radii of lithium and fluorine

Ans. Both Li & F are present in the same period and we know that the sizes are decreased from left to right. So, Li has greater size than F

Li = 152 pm F = 64 pm

(ii) A lithium atom and its ion, Li^+

Ans. As we know that the size of cation is always less than its parental atomic size.

Li = 152 pm Li^+ = 60 pm

(iii) An oxygen atom and its ion, O^{2-}

Ans. The size of anion will always be greater than its parental atomic size.

O = 66 pm O^{2-} = 140 pm

(iv) A Nitride ion, N^{3-}, and a fluoride ion, F^-

Ans. Nitride ion (N^{3-}) has more negative charge than fluoride ion (F^-). So, N^{3-} ion has larger size than F^- ion.

N^{3-} = 171 pm F^- = 136 pm

IONIZATION ENERGY

Q7. Define and explain ionization energy. Write its factors and trends in the periodic table. **11201007**

Ans. Definition of 1st Ionization Energy: The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1+ ions is known as 1st ionization energy (ΔH_{i1}).

Example: 1st Ionization Energy

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^- \qquad \Delta H_{i1} = 496\ kJ\ mol^{-1}$$
$$Ca_{(g)} \longrightarrow Ca^+_{(g)} + e^- \qquad \Delta H_{i1} = 590\ kJ\ mol^{-1}$$

Definition of 2nd Ionization Energy: If a second electron is removed from each ion in a mole of gaseous 1+ ions, it is called the 2nd ionization energy (ΔH_{i2}).

Example: 2nd Ionization Energy

$$Ca^+_{(g)} \longrightarrow Ca^{2+}_{(g)} + e^- \qquad \Delta H_{i2} = 1150\ kJ\ mol^{-1}.$$

Definition of 3rd Ionization Energy: If third electron is removed from each ion in a mole of gaseous 2+ ions, it is called the 3rd ionization energy (ΔH_{i3}).

Example: 3rd ionization energy

$$Ca^{2+}_{(g)} \longrightarrow Ca^{3+}_{(g)} + e^- \qquad \Delta H_{i3} = 4940\ kJ\ mol^{-1}$$

An element can have several ionization energies; the exact number corresponds to its atomic number.

Factors affecting the ionization energy

The magnitude of the ionization energy of an element depends upon the following factors;

(i) Nuclear charge

Greater the effective nuclear charge, greater is the electrostatic force of attraction, more difficult is the removal of an electron from the atom. So, ionization energy increases with an increase in the effective nuclear charge.

(ii) Size of the atom or ion

In bigger atoms force of attraction between the nucleus and the outermost electrons is weaker. Therefore, the ionization energy decreases as the size of the atom increases and vice-versa.

(iii) Electronic arrangement

It is observed half-filled and completely-filled orbitals are found to be more stable. Therefore, the ionization energy is higher when an electron is to be removed from a fully-filled or half-filled shells.

- Noble gases have highest ionization energies in their respective periods. It is due to highly stable fully-filled shells (ns^2np^6).
- Oxygen has lower ionization energy than nitrogen due to the half filled sub-shell of nitrogen.

The electronic configurations of oxygen and nitrogen are:

$$_7N = 1s^2\,2s^2\,2p_x^1\,2p_y^1\,2p_z^1 \quad \text{(more stable)} \quad \Delta H_{i1} = 1403\ kJmol^{-1}$$

$$_8O = 1s^2\,2s^2\,2p_x^2\,2p_y^1\,2p_z^1 \quad \text{(less stable)} \quad \Delta H_{i1} = 1365\ kJmol^{-1}$$

(iv) Shielding effect

Greater the shielding effect, easier it is to remove the valence electrons from an atom. Larger the number of inner electrons, greater is the shielding effect, therefore, lower is the ionization energy.

(v) Spin-Pair Repulsion

When electrons are spin-paired in the same orbital, the repulsion between them can lead to a slightly lower ionization energy compared to removing unpaired electron.

Reason: The paired electrons experience increased repulsion, making it slightly easier to remove one of the paired electrons.

Manganese (Mn) has two spin-paired electrons in its 4s orbital. The ionization energy to remove one of these paired electrons is relatively lower due to increased repulsion between the electron pairs. In contrast chromium (Cr) has one unpaired electron in its 4s orbitals. Removing one of these unpaired electrons requires more energy due to the absence of spin-pairing repulsion.

$_{24}Cr$ = [Ar] $3d^5 4s^1$ (more stable)

$_{25}Mn$ = [Ar] $3d^5 4s^2$ (less stable)

Periodic trends in ionization energy

Variation in groups

Going down in a group ionization energy decreases.

Reason: The nuclear charge increases but as the size of the atom and the number of electrons causing the shielding effect also increase therefore ionization energy decreases from top to bottom.

Group 1: In Group 1, the ionization energies decrease in the following order: Li> Na> K>Rb>Cs.

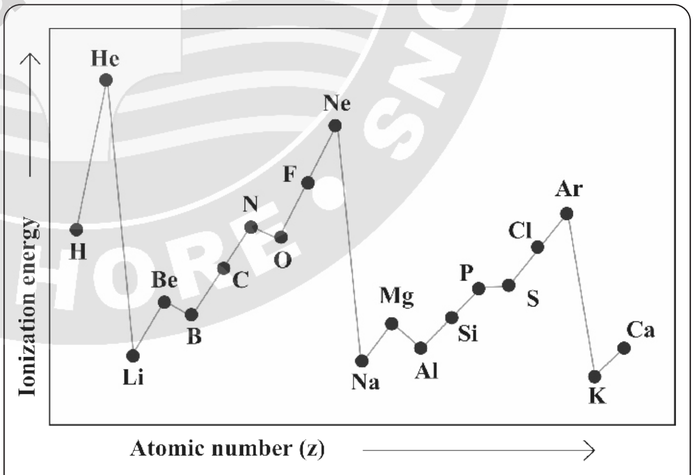


Fig: Variation in ionization energies across periods

Example: The 6s valence electron of Cs is farther from the nucleus and thus easier to remove compared to the 5s valence electron of Rb.

Variation in period

When we move from left to right across a period the ionization energy generally increases.

Reason: Number of shells remains unchanged while the effective nuclear charge increases, making it more difficult to remove an electron. Although the number of electrons also increases across a period, the shielding effect within the same shell is same so not considered. So, the ionization energy increases.

Graph

- The trend of ionization energies of period (1-3) is shown in Figure.
- The figure also reveals that noble gases have the highest values of ionization energy due to complete outermost shell in them, the removal of electron is extremely difficult.
- Alkali metals have lowest values of ionization energy.

Abnormality: Group number 13, 16 elements show abnormality in general trend. Actual trend of ionization energy of second period elements is Li < B < Be < C < O < N < F < Ne.

Quick Check 1.4

(a) Explain with reasoning following facts about ionization energy:

(i) 1st ionization energy of Boron is lesser than Beryllium.

Ans. Beryllium has two electrons in its s-sub shell which is stable from, while Boron has one electron in its p-sub shell which makes it relatively unstable. That's why 1st ionization energy of Boron is less than Beryllium. Because of full filled rule of Beryllium's 2s subshell.

(ii) 1st ionization energy of Aluminium is lower than Magnesium.

Ans. Aluminium has one electron in its 3p sub shell which make it unstable as compared to Magnesium, that has filled 3s sub shell. That's why 1st ionization energy of Aluminium is lower than that of Magnesium.

(b) What trend is observed in ionization energy as you go down group 3? Give reason.

Ans. The trend of ionization energy as we go down in the group 3 decreases due to the following reasons:

- Atomic size increases down the group.
- No. of shells increase down the group.
- Effective nuclear charge decreases down the group.
- Shielding effect increases down the group.

ELECTRON AFFINITY

Q8. Define and explain electron affinity. Give its factors and trends in periodic table. **11201008**

Ans. Definition of 1st Electron Affinity: The first electron affinity, (ΔH^o_{ea1}), is the enthalpy change involved when 1 mole of electrons is added to 1 mole of gaseous atoms to form 1 mole of gaseous uni-negative ions under standard conditions.

Example of Cℓ (Chlorine)

Electron affinity of chlorine atom.

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \qquad \Delta H^o_{eal} = -348.8 kJmol^{-1}$$

This is amount of energy released when 6.02×10^{23} atoms of chlorine in the gaseous state are converted into $Cl^-_{(g)}$ ions. Since, energy is released, so first electron affinity carries negative sign.

Definition of 2nd Electron Affinity: The second electron affinity, ΔH^o_{ea2} is the amount of energy required to add electrons to 1 mole of uni-negative gaseous ions to form 1 mole of gaseous 2- ions under standard conditions.

Example of oxygen: When first electron is added to a neutral oxygen atom, 141kJ mol^{-1} energy is released.

$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \qquad \Delta H^o_{ea1} = -141 kJmol^{-1}$$

But 844 $kJmol^{-1}$ of energy is absorbed on adding second electron to a uni-negative (O^-) ion.

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \qquad \Delta H^o_{ea2} = +844 kJmol^{-1}$$

The net enthalpy changes for the formation of the oxide ion (O^{2-}) can be calculated by adding the first and second electron affinities

$$\Delta H^o_{ea1} + \Delta H^o_{ea2} = (-141) + (844) = +701 kJmol^{-1}$$

Factors affecting electron affinity

Important factors affecting the magnitude of electron affinity values of elements are as follows:

(i) Size of atom

For small sized atoms the attraction of the nucleus for the incoming electron is stronger. Thus, smaller is the size of the atom, greater is its electron affinity.

(ii) Nuclear charge

Greater the magnitude of nuclear charge of an element stronger is the attraction of its nucleus for the incoming electron. Thus, with the increase in the magnitude of nuclear charge, electron affinity also increases.

(iii) Electronic configuration of atom

Group 15: The electron affinity is low when the electron is added to a half filled sub-shell than that for partially filled one. Electron affinity values of 'N' and 'P' group-15 (V-A), atoms are very low. This is because of the presence of half-filled 'np' orbitals in their valence shell ($N = 2s^2\ 2p^3$, $P = 3s^2\ 3p^3$). These half-filled p-subshells, being very stable, have very little tendency to accept any extra electron to be added to them.

Group 18: Noble gases group-18 (VIII-A) have stable $ns^2\ np^6$ configuration and hence the atoms of these gases, do not accept any extra electron, This is evident from their positive 1st electron affinities.

Periodic trends in electron affinity

Variation in groups

Electron affinity generally decreases down the group.

Reason: As the atomic size increases down the group, the larger electron cloud causes the incoming electron to experience less attraction from the nucleus.

This trend is observed in the halogens (At<I<Br<F<Cl).

Variation in periods

Electron Affinity generally increases in period from left to right.

Reason: This is firstly due to increase in the nuclear charge, which attracts additional electrons more strongly and secondly due to decreasing atomic radius.

Abnormality: Group number 2, 15 and 18 of periodic table show abnormality in general trend moving left to right. Actual order of electron affinity of 2nd period elements.

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

Table 1.1: Electron affinities (kJ/Mol) for group 1 and group 17

Element	Electron Affinity (kJ/mol)	Element	Electron Affinity (kJ/mol)
Flourine	-328.0	Lithium	-60.0
Chlorine	-349.0	Sodium	-53.0
Bromine	-324.0	Potassium	-48.0
Iodine	-295.0	Rubidium	-47.0
Astatine	-270.1	Cesium	-46.0

Quick Check 1.5

Explain with reasoning following facts about electron affinity:

(a) 1st electron affinity of Oxygen is -141kJ/mol but 2nd affinity is +844.0kJ/mol.

Ans. When the 1st electron is added in an atom to form uni-negative ion then energy is released, so it will be negative enthalpy. While 2nd electron is added to uni-negative ion to form di-negative ion, energy is required for this addition, which is shown with +ve sign. That's why 1st electron affinity of oxygen is -141kJ/mol while 2nd electron affinity is +844kJ/mol.

$$O + e^- \longrightarrow O^{1-} \qquad \Delta H_{e1} = -141 kJ/mol$$

$$O^{1-} + e^- \longrightarrow O^{2-} \qquad \Delta H_{e2} = +844 kJ/mol$$

(b) Which of nitrogen and phosphorus has the higher electron affinity? Justify with reason.

Ans. Electron affinity decreases in the group downward. It is because of the following reasons:

(i) Atomic size increases
(ii) No. of shells increases
(iii) Shielding effect increases
(iv) Effective nuclear charge decreases

So, Nitrogen has higher electron affinity than phosphorous, because Nitrogen has smaller size than phosphorous.

(c) F has lower electron affinity than Cl although its size is smaller. Explain why?

Ans. The value of electron affinity of F is less than Cl, because of the following reasons:

(i) Small size of "2p" orbital of high electron density of F.
(ii) Highest electronegativity of F, but has two shells only.
(iii) Greater size of "3p" orbital of Cl with low charge density of electrons.

$$F + e^- \longrightarrow F^- \qquad \Delta H_{ea1} = -328 kJ/mol$$

$$Cl + e^- \longrightarrow Cl^- \qquad \Delta H_{ea1} = -349kJ/mol$$

(d) Why noble gases (group-18) have positive 1st electron affinities? Explain in terms of electronic configuration.

Ans. Noble gases have ns^2 np^6 (except He) configuration so they never lose or gain electron at ordinary conditions. Their first electron affinity is always positive because their duplets and octets are complete.

ELECTRONEGATIVITY

Q9. Write a note on electronegativity in detail. 11201009

Ans. Definition: Electronegativity is the power of an atom to attract shared pair of electrons toward itself in a molecule.

Example: The electronegativity of F is 4.0 on Pauling scale.

Unit: It has no unit.

Scale: Linus Pauling, an American chemist, developed a scale of dimensionless electronegativity values, which range from just below one for alkali metals to a maximum of four for fluorine. Higher electronegativity values signify a stronger attraction for electrons compared to lower values.

> **Did you know?**
>
> **S.Q. Give achievements of Linus Pauling.**
>
> **Ans.** Linus Pauling is the only person to have received two unshared Nobel Prizes, one for chemistry in 1954 for his work on the nature of chemical bond and one for peace in 1962 for his opposition to weapons of mass destruction.

Factors affecting Electronegativity

(i) Atomic size

A larger atomic size will result in a lower value of electronegativity. This is because electrons being far away from the nucleus will experience a weaker force of attraction.

Example: The electronegativities of halogens in group 17 are in the order:

F > Cl > Br > I

(ii) Effective nuclear Charge

A higher value of the effective nuclear charge will result in a greater value of electronegativity, because an increase in nuclear charge causes greater attraction to the bonded electrons. This is why the electronegativity in a period increases from left to right. The electronegativity of Li in period 2 is 1.0 and F has a value of 4.0.

Periodic Trends in Electronegativity

Variation in periods

When we move from left to right along the period, the electronegativity increases,

Reason: This is due to increasing nuclear charge and decrease in atomic size.

Variation in groups

In the groups, it decreases from top to bottom in periodic table.

Reason: This is due to the increase in size by the addition of shells and increasing shielding effect.

Example: The halogen group, the electronegativity value decreases from fluorine (4.0) to iodine (2.5) as shown in a part of the periodic table in figure.

- Normally metals being on the left side of the periodic table, possess lower electronegativity.
- Non-metals have higher values of electronegativity than those of metals.
- Metals are electropositive and non-metals are electronegative, relatively.

Figure provides a summary of all the variation trends in various physical properties of elements in the periodic table.

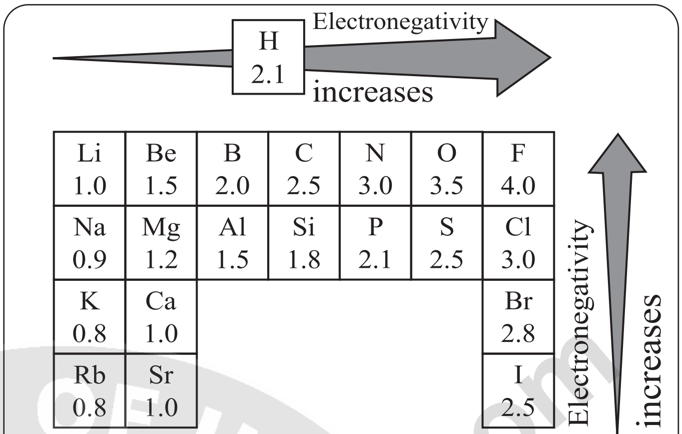


Fig: Variation of electronegativity in groups and periods

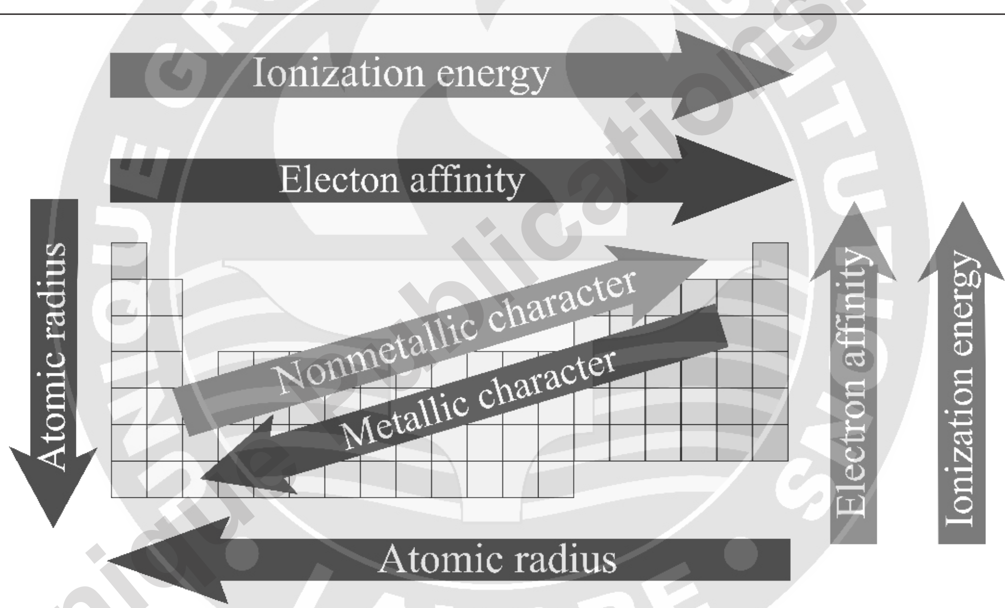


Fig: Trends in various physical properties in the periodic table

VARIATION IN METALLIC CHARACTER

Q10. Explain metallic character with its trend in the periodic table. **11201010**

Ans. Definition

The metallic character of elements is typically their tendency to lose electrons.

Explanation

- The elements on the left side of the periodic table have a greater tendency to lose their outermost electrons to achieve noble gas configuration.
- Elements on the right side of the table tend to gain electrons.
- Elements on the left side of the periodic table are metals that form positive ions.
- Elements on the right side, particularly in the right corner are nonmetals that form negative ions.
- Metallic character of an element largely depends on its valence shell electronic configuration.

Periodic Trends in metallic character

Variation in periods: The metallic character of the elements decreases.

Reason: The increase in nuclear charge pulls the electron could closer to the nucleus, making it more difficult for the atom to lose electrons and thereby decreasing across the period, the nuclear charge increases while the size deceases, which results in stronger attraction to the valence electrons making its difficult. So, the metallic character decrease in period.

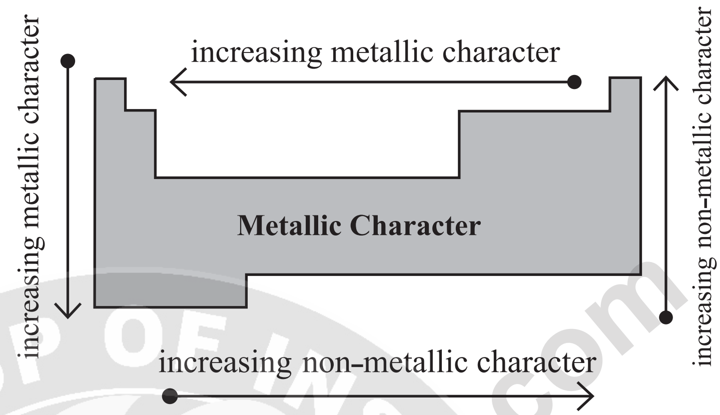


Fig : General trends for the metallic character of elements

Variation in groups: Metallic character increases as one moves down in a group of the periodic table.

Reason: The increases in atomic size and the shielding effect, which reduce the nuclear attraction on the valence electrons.

The increase in metallic character (ease of losing electron) makes the element more reactive.

Example: Cesium is far more reactive and electropositive than sodium or lithium.

Quick Check 1.6

(a) Illustrate how does the metallic character vary in group 14

Ans. Metallic character increases down the group because of greater size and high shielding effect. So, electrons can easily be removed due to poor attraction between nucleus and outermost electrons. That's why metallic character of group 14(IVA) increases down the group.

(b) Identify semi metals in groups 14, 15 and 16. Why they are semi metals?

Ans. Group 14 semi metal are Si, Ge

Group 15 semi metal are As, Sb

Group 16 semi metal are Te and Po

They are semi metals because their properties are in between metals and non-metals. They are also named as semi-conductors.

REACTIONS OF SODIUM AND MAGNESIUM

Q11. Describe the reactions of Na and Mg with oxygen, chlorine and water. 11201011

Ans. Reactions of Na and Mg

(i) With water

Reaction of Sodium: Sodium is more reactive than magnesium towards water.

Na reacts vigorously with water to form sodium hydroxide and hydrogen gas.

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

Reaction of Magnesium: Mg reacts more slowly in forming magnesium hydroxide and hydrogen gas. However, magnesium reacts with steam more vigorously to make magnesium oxide and hydrogen gas.

$$Mg_{(s)} + 2H_2O_{(1)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

$$Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$$

(ii) With Oxygen

Reaction of Sodium: Sodium burns in oxygen with a golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air. It reacts vigorously with oxygen in open air to form peroxide.

$$2Na_{(s)} + O_{2(g)} \longrightarrow Na_2O_{2(s)}$$

Under special conditions like limited O_2 or high temperature, sodium oxide is formed.

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)}$$

Reaction of Magnesium: Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide. Its reaction with oxygen is slow in comparison with sodium.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

Magnesium powder burns very rapidly with an intense white flame. This has led to its use in fireworks and S.O.S (save our ship) flares.

(iii) With Chlorine: Chlorine reacts with both metals to give soluble salts. It reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed. Magnesium also reacts with chlorine to give white solid, magnesium chloride.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

$$Mg_{(s)} + Cl_{2(g)} \longrightarrow MgCl_{2(s)}$$

Quick Check 1.7

(a) What is the nature of oxides and hydroxides of Na and Mg?

Ans. Na can react with oxygen more quickly as compared to Mg. Na gives golden yellow flame to form Na_2O_2 or Na_2O in limited oxygen. The nature of oxides and hydroxides of Na and Mg is strongly alkaline.

$$2Na_{(s)} + O_{2(g)} \longrightarrow Na_2O_{2(s)} \qquad (excess\ oxygen)$$

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)} \qquad (Limited\ oxygen)$$

Mg burns with intense white flame with oxygen.

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$$

Na and Mg react with water to give hydroxide.

$$2Na_{(s)} + 2H_2O_{(\ell)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2H_2O_{(\ell)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

$$Mg + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$$
(Steam)

(b) What could you predict about the reactivity of Ca. a group 2 element, when reacted with water and oxygen?

Ans. (i) With water: Ca reacts with water at 100°C to give calcium hydroxide and hydrogen gas.

$$Ca + 2H_2O \xrightarrow{100°C} Ca(OH)_2 + H_2$$

(ii) With Oxygen: Ca reacts with excess with oxygen to give calcium oxide.

$$2Ca + O_2 \longrightarrow 2CaO$$

TRENDS IN BONDING IN OXIDES AND CHLORIDES OF PERIOD 3

Q12. Explain the trends of bonding and classification of chlorides and oxides of period 3. **11201012**

Ans. Oxides

Definition: Oxides are binary compounds formed by the reaction of oxygen with other elements.

Ionic and covalent nature of oxides of period 3

- Oxides of group 1, 2 & 3 (Na_2O) have more ionic character.
- These oxides exist as giant ionic lattices with strong electrostatic forces between oppositely charged ions.
- Oxides of group 4, 5, 6 & 7 (SO_2) are more covalent.
- These oxides exist as covalent molecules with weak intermolecular forces.
- This transition is a result of the increasing electronegativity and decreasing ionic character.

Classification of Oxides

The classification of oxides is done into neutral, amphoteric and basic or acidic based on their characteristics.

(i) Basic oxides

Definition: A basic oxide is an oxide that when combined with water gives off an alkali.

Example: Na_2O, CaO, BaO.

Characteristics: Metals react with oxygen to give basic oxides. These oxides are usually ionic in nature. Group 1 and 2 form basic oxides when react with oxygen. Group 2 hydroxides solubility increases down the group so alkalinity also increases down the group.

$$Na_2O_{(s)} + H_2O_{(1)} \longrightarrow 2NaOH_{(aq)}$$
$$CaO_{(s)} + H_2O_{(1)} \longrightarrow Ca(OH)_{2(aq)}$$

(ii) Amphoteric oxides

Amphoteric oxides are oxides that can react with both acids and bases. This means they have the ability to behave as either an acid or a base, depending on the conditions. Aluminium oxide ($A\ell_2O_3$) is insoluble in water but reacts with hydrochloric acid to form aluminium chloride and water, and with sodium hydroxide to form sodium aluminate and water.

$$A\ell_2O_{3(s)} + 6HC\ell_{(aq)} \longrightarrow 2A\ell C\ell_{3(s)} + 3H_2O_{(\ell)}$$
$$A\ell_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaA\ell O_{2(s)} + H_2O_{(\ell)}$$

(iii) Acidic oxides

An acidic oxide is an oxide that when combined with water gives off an acid. Non-metals react with oxygen to form acidic oxides which are held together by covalent bonds. Silicon dioxide is acidic oxide as it can react with bases.

Examples of acidic oxides in period 3 are P_2O_3, P_2O_5, SO_3, SO_2

$$P_2O_{3(s)} + H_2O_{(\ell)} \longrightarrow H_3PO_{3(aq)}$$

$$SO_{2(g)} + H_2O_{(1)} \longrightarrow H_2SO_{3(aq)}$$

Reactions of these oxides with base are given below:

$$SiO_{2(s)} + 2NaOH_{(aq)} \longrightarrow Na_2SiO_{3(aq)} + H_2O_{(\ell)}$$

$$P_2O_{3(s)} + 6NaOH_{(aq)} \longrightarrow 2Na_3PO_{3(aq)} + 3H_2O_{(\ell)}$$

$$SO_{2(g)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{3(aq)} + H_2O_{(1)}$$

Chlorides

Definition: Chlorine forms binary compounds with other elements known as chlorides.

Ionic and covalent of chlorides

Chlorides of group 1, 2 and 3 (NaCl) are predominately ionic. Chlorides of elements of group 4, 5, 6 and 7 (PCl_5) are covalent. The covalent character in chlorides increases due to decrease in difference of electronegativity between the halogen and the other atom.

Classification of chlorides on the bases of pH

These chlorides show characteristic behaviour when we add them into water, resulting in solutions that can be acidic or neutral.

(i) Neutral chlorides

Neutral chlorides are salts, when dissolved in water, produce a neutral solution with a pH close to 7. At the start of period 3 chloride, sodium and magnesium do not react with water. The solutions formed contain the positive metal ions and negative chloride ions surrounded by water molecules. These ions are now known as hydrated ions and this process is known as hydration.

The following reaction occurs:

$$NaC\ell_{(s)} \longrightarrow Na^+_{(aq)} + C\ell^-_{(aq)} (pH = 7)$$

$$MgC\ell_{2(s)} \longrightarrow Mg^{+2}_{(aq)} + 2C\ell^-_{(aq)} (pH = 6.5)$$

Group 1 and group 2 chlorides are also neutral with few exceptions.

(ii) Acidic chlorides

If we move in period 3, from aluminium to sulphur all chlorides react with water to make acidic solution with pH less than 7 this process is called hydrolysis. When $AlCl_3$ is added to water, aluminium and chloride ions in solution. Al^{3+} ion is hydrated and cause a water molecule to lose an H^+ ion, this process is hydrolysis. This turns the solution acidic.

The following reaction occurs:

$$A\ell C\ell_{3(s)} + 3H_2O_{(\ell)} \longrightarrow A\ell(OH)_{3(s)} + 3HC\ell_{(aq)}$$

Other examples of acidic chlorides are given below.

$$SiC\ell_{4(\ell)} + 2H_2O_{(\ell)} \longrightarrow SiO_{2(s)} + 4HC\ell_{(aq)}$$

$$PC\ell_{3(\ell)} + 3H_2O_{(\ell)} \longrightarrow H_3PO_{3(s)} + 3HC\ell_{(aq)}$$

Quick Check 1.8

(a) ZnO reacts with HCl to give $ZnCl_2$ and with NaOH to give Na_2ZnO_2. Give equation and also predict the type of this oxide.

Ans.

$$\underset{\text{(base)}}{ZnO} + \underset{\text{(acid)}}{2HC\ell} \longrightarrow \underset{\text{(salt)}}{ZnC\ell_2} + \underset{\text{(water)}}{H_2O}$$

$$\underset{\text{(acid)}}{ZnO} + \underset{\text{(base)}}{2NaOH} \longrightarrow \underset{\text{(salt)}}{Na_2ZnO_2} + \underset{\text{(water)}}{H_2O}$$

ZnO, is amphoteric in nature, because when it reacts with acid it acts as base and when it reacts with base it acts as acid as in given reactions.

(b) Why $A\ell Cl_3$ is an acidic halide, but NaCl not?

Ans. The reaction of $AlCl_3$ is called hydrolysis with water, because it gives H^+ ions and form acidic medium having pH less than 7 while the reaction of NaCl with water is hydration process not hydrolysis and solution's pH is 7.

$$AlCl_3 + 3H_2O_{(\ell)} \longrightarrow Al(OH)_{3(s)} + 3HCl_{(aq)}$$

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

(c) Predict whether the chlorides PCl_5, NCl_3 would be acidic or basic, given reason.

Ans. The chlorides of P, S, N, C of group IIIA to VIA are acidic in nature because when they react with water they give H^+ ions in the solution, making pH less than 7 by the process called hydrolysis.

$$PCl_{5(s)} + 4H_2O_{(\ell)} \longrightarrow H_3PO_{4(aq)} + 5HCl_{(aq)}$$

$$NCl_3 + 3H_3O_{(\ell)} \longrightarrow NH_{3(aq)} + 3HOCl_{(aq)}$$

(d) Would SO_2 and P_2O_5 react with HCl and H_2SO_4 or with NaOH?

Ans. Yes, both SO_2 and P_2O_5 can react with acid or base.

$$P_2O_5 + 6HCl \longrightarrow 2POCl_3 + 3H_2O$$

$$P_2O_5 + 6NaOH \longrightarrow 2Na_3PO_4 + 3H_2O$$

$$SO_2 + 2HCl \longrightarrow SOCl_2 + H_2O$$

$$SO_2 + 2NaOH \longrightarrow Na_2SO_3 + H_2O$$

Q13. Discuss variation in oxidation number in oxides and chlorides. 11201013

Ans. Oxidation number

Definition: The oxidation number of an atom is the formal charge on that atom in a molecule or ion.

- The oxidation number is also referred to as the oxidation state.
- In ionic compounds the oxidation number of an atom is defined as the charge which appears on the, ions.

The oxidation numbers in oxides and chlorides of the third period.

The oxidation number of an element of 3rd Period in its oxide or chloride corresponds to the number of electrons used for bonding and is always positive because oxygen and chlorine are more electronegative than any of these elements.

The oxidation number matches the group number, reflecting the total number of valence electrons.

Examples: Following are some examples of oxidation state of oxide and chlorides of third period:

- In the oxides of third period, the oxidation number increases from +1 in Na to + 6 in S.
- In chlorides of third period, the oxidation number increases from +1 in Na to + 5 in P.
- Phosphorus and sulfur exhibit several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals.
- In SO_2, sulfur has an oxidation number of + 4 because only four electrons are used for bonding.
- In SO_3, sulfur has an oxidation number of + 6 because all six electrons are used for bonding.

Consider table 1.2 for oxidation states of various elements of the periodic table.

Table 1.2: Oxidation Numbers in Oxides and Chlorides of 3rd Period elements

Oxide	Oxidation Number	Chloride	Oxidation Number
Na in Na_2O	+1	Na in $NaC\ell$	+1
Mg in MgO	+2	Mg in $MgC\ell_2$	+2
$A\ell$ in $A\ell_2O_3$	+3	$A\ell$ in $A\ell C\ell_3$	+3
Si in SiO_2	+4	Si in $SiC\ell_4$	+4
P in P_4O_{10}/ P in P_4O_6	+5/+3	P in $PC\ell_5$	+5
S in SO_3	+6	P in $PC\ell_3$	+3
S in SO_2	+4	S in $SC\ell_2$	+2

Quick Check 1.9

(a) Calculate the oxidation number of sulphur in SO_2 and SO_3.

Ans. **SO_2 :**

$(O.N\ of\ S) + 2(O.N\ of\ O) = 0$
$(O.N\ of\ S) + 2(-2) = 0$
$(O.N\ of\ S) - 4 = 0$
$(O.N\ of\ S) = +4$

SO_3 :

$(O.N\ of\ S) + 3(O.N\ of\ O) = 0$
$(O.N\ of\ S) + 3(-2) = 0$
$(O.N\ of\ S) - 6 = 0$
$(O.N\ of\ S) = +6$

(b) Why some p-block elements show variable oxidation state?

Ans. Some of the elements present in p-block can show variable oxidation states because of the presence of empty d-sub shell. Like P, S show variable oxidation states.

Example: Oxidation states of S

(i) S in H_2S is – 2 (ii) S in SO_2 is + 4 (iii) S in SO_3 is + 6

Oxidation states of P

(i) P in $PC\ell_3$ is + 3 (ii) P in $PC\ell_5$ is + 5

MULTIPLE CHOICE QUESTIONS (EXERCISE)

Select the correct answer.

1. Which scientist first time observed the periodicity in the elements? **11201014**

(a) J. Newlands (b) L. Meyer (c) Dobereiner (d) D. Mendeleev

2. Recognize the element if it has 3 electron shells, belongs to "s" block and has 2 electrons in its outer most shell. 11201015
(a) Calcium (b) Sodium
(c) Magnesium (d) Potassium

3. Which one do you think is correct about metallic character? 11201016
(a) It decreases from top to bottom in a group
(b) It increases from top to bottom in a group
(c) It remains constant from left to right in a period
(d) It increases from left to right in a period

4. Which property increases as you go down a group in the periodic table?
(a) Atomic radius 11201017
(b) Electron Affinity
(c) Electronegativity
(d) Ionization Energy

5. Which set of the following conditions results in higher ionization energy? 11201018
(a) Smaller atom and greater nuclear charge
(b) Smaller atom and smaller nuclear charge
(c) Larger atom and greater nuclear charge
(d) Larger atom and the smaller nuclear charge

6. Which of the following atoms show more than one (variable) oxidation states? 11201019
(a) Sodium (b) Magnesium
(c) Aluminium (d) Phosphorous

7. Which is the correct general trend in the variation of electron affinity in a group? 11201020
(a) It becomes less negative from top to bottom
(b) It becomes more negative from top to bottom
(c) It remains the same
(d) It has no definite trend and changes irregularly

8. What is the oxidation state of sulfur in the sulfate ion SO_4^{2-}? 11201021
(a) + 4 (b) + 2
(c) + 6 (d) 0

9. Which is the correct trend in variation of electronegativity along a period of the periodic table? 11201022
(a) It decreases from left to right across a period
(b) It increases from left to right across a period
(c) It remains constant
(d) It has no definite trend

10. The atomic radius generally ___ across a period in the periodic table. 11201023
(a) Increases
(b) Decreases
(c) Remains constant
(d) First increases then decreases

11. Which one of the following elements has the highest ionization energy? 11201024
(a) Sodium (Na) (b) Magnesium (Mg)
(c) Aluminium(Al) (d) Argon (Ar)

SLO BASED MULTIPLE CHOICE QUESTIONS

Modern Periodic Table and Its History

12. Which one is the correct statement among the following? 11201025
(a) Anionic radius is generally smaller than atomic radius
(b) Cationic radius is generally bigger than atomic radius
(c) Cationic radius is generally smaller than atomic radius
(d) Both anionic and cationic radii are smaller than atomic radius

13. Modern periodic table provides information about_________ elements. 11201026
(a) 100 (b) 110
(c) 118 (d) 130

14. By 1700 A.D. only_________ elements were recognized 11201027
(a) 6 (b) 12
(c) 10 (d) 3

15. Concept of Triads was given by: 11201028
(a) Mendeleev (b) L.Meyer
(c) Newlands (d) Dobereiner

16. John Newlands gave the idea of: 11201029
(a) Law of triads
(b) Law of Octaves
(c) Modern periodic law
(d) Curves between weight and volume

17. Who considered as the father of the periodic table? 11201030
(a) Moseley (b) Newland
(c) Mendeleev (d) Dobereiner

18. Select metalloid from the following:
(a) Po (b) P 11201031
(c) Cl (d) Fe

19. Group 16 elements are also called: 11201032
(a) Pnictogens (b) Chalcogens
(c) Halogens (d) Alkali metals

20. Select which is not alkaline earth metal?
(a) Be (b) Co 11201033
(c) Sr (d) Ba

21. The term "Halogen" means: 11201034
(a) Ash former (b) Salt former
(c) Copper giver (d) Sulphar giver

22. Noble gases are present in 18 group, also called: 11201035
(a) IA (b) IIB
(c) VIIIA (d) VIIIB

23. Which ion will have greater size? 11201036
(a) Na^+ (b) Mg^{2+}
(c) Al^{3+} (d) Si^{4+}

Ionization Energy, Electron Affinity And Electronegativity

24. Which of the following has highest ionization energy? 11201037
(a) P (b) S
(c) Cl (d) Ar

25. Select the element which has highest 1st ionization energy: 11201038
(a) Boron (b) Carbon
(c) Nitrogen (d) Oxygen

26. Cl has more electron affinity value than F and its value is: 11201039
(a) +328kJ/mol (b) +349kJ/mol
(c) -328kJ/mol (d) -349kJ/mol

27. 2nd value of electron affinity: 11201040
(a) will be negative always
(b) will always be positive
(c) equals to 1st value
(d) double than 1st value

28. L.Pauling developed a scale for:
(a) Ionization energy 11201041
(b) Electron affinity
(c) Electronegativity
(d) None

29. Highest electronegative element is____:
(a) F (b) Cl 11201042
(c) Br (d) I

30. Least electronegative element is_____:
(a) H (b) He 11201043
(c) Cs (d) Li

31. Electronegativity value of "Cs" is:
(a) 4 (b) 2.1 11201044
(c) 0.1 (d) 0.79

Metallic and Non-metallic Character

32. Tendency to lose electrons called:
(a) Electronegativity 11201045
(b) Ionization energy
(c) Electron affinity
(d) Metallic character

33. Metallic character increases: 11201046
(a) Downward and forward from left to right
(b) Upward and forward from left to right
(c) Downward and backward from right to lef
(d) Upward and backward from right to left

34. The most reactive metals of the periodic table are: 11201047
(a) Coinage metals
(b) Rare earth metals
(c) Alkali metals
(d) Alkaline metals

Oxides and Hydroxides

35. Which of the following will not give hydroxide upon treating with water:
(a) $Na_{(s)} \& H_2O_{(\ell)}$ 11201048
(b) $Mg_{(s)} \& H_2O_{(g)}$
(c) $Mg_{(s)} \& H_2O_{(\ell)}$ (d) $Ca_{(s)} \& H_2O_{(\ell)}$

36. Peroxides are those in which oxidation state of oxygen is: 11201049
(a) -2 (b) -1
(c) -1/2 (d) +2

37. Which element is used in S.O.S flares and in firework? 11201050
(a) Na powder (b) $CaCO_3$
(c) Mg powder (d) KO_2

38. Na burns with oxygen to give: 11201051
(a) Golden yellow flame
(b) Golden brown flame
(c) Crimson red flame
(d) Intense white flame

39. Select the one which is acidic oxide: 11201052
(a) Na_2O (b) CaO
(c) Al_2O_3 (d) CO_2

40. Metallic oxides are _______ in nature: 11201053
(a) Acidic
(b) Basic
(c) Amphoteric
(d) Depend upon oxide

41. Select basic oxide: 11201054
(a) CO_2 (b) SiO_2
(c) Al_2O_3 (d) MgO

42. Which one is amphoteric oxide: 11201055
(a) CaO (b) MgO
(c) Al_2O_3 (d) Na_2O

43. Acidic oxides are: 11201056
(a) Metallic oxides
(b) Non-metallic oxides
(c) Metalloid oxides
(d) All

44. Acidic chlorides have pH? 11201057
(a) more than 7
(b) less than 7
(c) equals to 7
(d) may be any

45. Oxidation state of P in P_4O_{10} is: 11201058
(a) +2 (b) +3
(c) +4 (d) +5

46. Element that can show variable oxidation states: 11201059
(a) Na (b) Mg
(c) S (d) Ca

47. Which oxidation state will not shown by S: 11201060
(a) -2 (b) +4
(c) +2 (d) +8

ANSWER KEY

1.	c	2.	c	3.	b	4.	a	5.	a	6.	d	7.	a
8.	c	9.	b	10.	b	11.	d	12.	c	13.	c	14.	b
15.	d	16.	b	17.	c	18.	a	19.	b	20.	d	21.	b
22.	c	23.	a	24.	d	25.	c	26.	d	27.	b	28.	c
29.	a	30.	c	31.	d	32.	d	33.	c	34.	c	35.	b
36.	b	37.	c	38.	a	39.	d	40.	b	41.	d	42.	c
43.	b	44.	b	45.	d	46.	c	47.	d				

SHORT QUESTION ANSWERS (EXERCISE)

Attempt the following short-answer questions.

Q1. What is 1st ionization energy? Give an example. 11201061

Ans. The amount of energy which is required to remove one valance electron from one mole of isolated gaseous atoms to form mono-positive gaseous ions is called 1st ionization energy.

Example:

$$Na_{(g)} \longrightarrow e^- + Na^+_{(g)} \quad \Delta H_{i1} = +496 \text{ kJ/mol}$$

Q2. Explain why sulfur has a lower first ionization energy than phosphorous? 11201062

Ans. Sulfur has four electrons in its 3p sub-shell while phosphorous has three electrons in its 3p sub-shell which makes phosphorous more stable than sulfur due to half filled stability rule, as shown in given example.

$$_{15}P = [Ne]\ 3s^2\ 3p_x^1\ 3p_y^1\ 3p_z^1$$

$$_{16}S = [Ne]\ 3s^2\ 3p_x^2\ 3p_y^1\ 3p_z^1$$

So, it is easy to remove one electron from sulfur but difficult to ionize phosphorous.

Q3. Why the elements in group 13 to 17 are called p-block elements? 11201063

Ans. The elements in group 13 to 17 are called p-block elements because their valance electrons are ended up in p-sub shell. According to valance electrons the elements are classified as s, p, d and f-block elements.

Q4. What are the factors that affect electronegativity? 11201064

Ans. Factors affecting the electronegativity are:

(i) Atomic size:

Greater the atomic size, lesser will be the electronegativity.

$$E.N \propto \frac{1}{Atomic\ Size}$$

(ii) Effective nuclear charge (Z-effect)

Greater the value of effective nuclear charge, larger will be the electronegativity.

E.N $\propto$ Effective nuclear charge

Q5. What factors are responsible for the increasing reactivity of alkali metals as you move down the group? 11201065

Ans. The main factors which are responsible for increasing the reactivity of alkali metals are "increase in atomic size" and "low ionization energy" down the group. As we know alkali metals have only one valance electron in their shells, a low ionization energy is required to remove valance electron and hence their reactivity increases. Increase in atomic size down the group also makes them reactive.

Q6. Why some of the elements show variable oxidation numbers while others do not? 11201066

Ans. Some of the elements can show variable oxidation numbers because they can expand their octet by exciting their electrons to the empty orbitals if they have. If the elements do not have empty orbitals, then they cannot show variable oxidation number. Examples: Oxidation number of P is +3 in PCl_3 and +5 in PCl_5, while Na, Mg, Ca cannot show variable oxidation states.

Q7. Identity the element which is in period 5 and group 15? 11201067

Ans. X is the element present in 5th period. So, its valance shell configuration ended up in 5th shell having principal quantum number 5. And its electron is showing its group number "15".

$$X = 1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^3$$

$$n = 5, \quad 5s^2 \quad 5p^3$$

5th period

15th group $5s^2\ 4d^{10}\ 5p^3$

$$2 + 10 + 3 = 15$$

$$X = Sb = 51$$

Q8. Why the oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorous? 11201068

Ans. Na and Mg are present at the left side of the periodic table and they are metals as well. So, that's why their oxides are more ionic in nature. While moving to the right side of the periodic table ionic character is decreased and the elements are non-metals as well. So, their ionic character decreased like in nitrogen and phosphorous oxides.

Q9. Give reasons for the different chemical reactivities of Na and Mg toward oxygen and chlorine. 11201069

Ans. Both Na and Mg can react with oxygen. Na gives golden yellow flame and forms peroxide while in limiting amount of oxygen Na can form normal oxide.

$$2Na_{(s)} + O_{2(g)} \longrightarrow Na_2O_{2(s)} \quad \text{(white solid)}$$

(excess) (peroxide)

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)} \quad \text{(white solid)}$$

(limited) (normal oxide)

Mg gives intense white flame with oxygen

$$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO \quad \text{(white solid)}$$

(normal oxide)

Both Na and Mg give chlorides with chlorine.

$$2Na_{(s)} + C\ell_{2(g)} \longrightarrow 2NaC\ell_{(s)} \quad \text{(white solid)}$$

$$Mg_{(s)} + C\ell_{2(g)} \longrightarrow MgC\ell_{2(s)}$$

(white to colourless solid)

Q10. Why the ionization energy of lithium is much lower than that of helium despite the fact that the nuclear charge of lithium is +3 and that of helium is +2? 11201070

Ans. Lithium is metal and has the tendency to lose electron easily. While helium is a noble gas and has filled valance shell, so its ionization energy is very high as compared to lithium despite the fact that helium has +2 charge while lithium has +3 nuclear charge.

Q11. The ionization energy of Be (atomic no. 4) is higher than that of B (atomic no. 5) despite the fact that the nuclear charge of Be is +4 and that of B is +5. 11201071

Ans. The reason is lies in the stability of electronic configuration of Be which has filled 2s sub-shell and unstable configuration of B which has one electron in its 2p sub-shell.

$$_4Be = 1s^2 \quad 2s^2$$

$$_5B = 1s^2 \quad 2s^2 \quad 2p_x^1$$

Q12. What is common in Na^+, Mg^{2+}, Al^{3+}, Ne^o, and F^-? Arrange them in increasing order of sizes. 11201072

Ans. The common thing in Na^+, Mg^{2+}, $A\ell^{3+}$, Ne^o and F^- is that all these are iso-electronic species and have 10 electrons in them. Their increasing sizes order is

$$A\ell^{3+} < Mg^{2+} < Na^+ < Ne < F^-$$

Q13. Consider the chlorides of sodium, magnesium and phosphorous: $NaC\ell$, $MgC\ell_2$ and $PC\ell_5$

(i) Classify each of these chlorides as acidic, basic or neutral.

(ii) For each chloride, briefly explain the reason for your classification, referring to their behaviour when dissolved in water. 11201073

Ans. Neutral chlorides: These chlorides in which chloride ion and metal ions are surrounded by water (hydrates) are called neutral chlorides. In these chlorides, they simply ionizes and do not react with water. Their pH near to 7.

$$NaC\ell_{(s)} \longrightarrow Na^+_{(aq)} + C\ell^-_{(aq)} \quad (pH = 7)$$

$$MgC\ell_{2(s)} \longrightarrow Mg^{2+}_{(aq)} + 2C\ell^-_{(aq)} \, (pH = 6.5)$$

Acidic chlorides

These chlorides which can react with water and produced H+ ion to give acidic solutions. Chlorides of Al to S from IIIA to VIA give acidic chlorides. This process is called hydrolysis.

$$PC\ell_{5_{(s)}} + 4H_2O_{(\ell)} \longrightarrow H_3PO_{4_{(aq)}} + 5HC\ell_{(aq)} \, (pH = 0)$$

SLO BASED SHORT QUESTION ANSWERS

History of Periodic Table

Q14. Which elements were discovered by the 1700AD? 11201074

Ans. By 1700 A.D., only 12 elements were recognized Gold (Au). Silver (Ag), Copper (Cu), Iron (Fe), Lead(Pb), Tin(Sn), Mercury (Hg), Phosphorous (P), Sulfur (S), Carbon (C), Zinc (Zn) and Arsenic (As).

Q15. What is Dobereiner's Triads? 11201075

Ans. In 1829, Dobereiner grouped three elements in such a way that the atomic mass of the middle element is the average of atomic masses of first and third elements. In this way he noticed some periodicity in the elements.

Example

$$^7Li, \ ^{23}Na, \ ^{39}K$$

$$Na = \frac{7+39}{2} = 23$$

Q16. Explain Newland's Octaves. 11201076

Ans. After every eight elements the next element show periodicity with first element if they are arranged on the basis of their atomic masses. It is called Newland's Octaves.

Q17. State Mendeleev's periodic law. **11201077**

Ans. In 1869 Dimitri Mendeleev stated his periodic law, it is stated as: Properties of the elements are the periodic functions of their atomic masses.

Q18. What do you know about Lother Meyer's curves law? **11201078**

Ans. In 1869, Lother Meyer developed the famous curves between atomic weight and atomic volume of the elements, These curves also showed periodicity.

Q19. State Moseley's Periodic Law/Modern periodic. **11201079**

Ans. Properties of the elements are the periodic functions of their atomic numbers. This Moseley's law also called modern periodic law.

Q20. How many groups and periods are present in modern periodic table? **11201080**

Ans. There are 7 periods and 18 groups, vertical columns are groups while periods are horizontal row. 18 groups are further divided into 8-A and 10-B groups.

Metal, Non-metal and Metalloids

Q21. What are metals and non-metals? Give examples. **11201081**

Ans. Elements which have the tendency to lose electron easily are called metals.

e.g: Li, K, Cs etc

Elements which have the tendency to accept electrons easily are called non-metals. e.g: N, O, F etc.

Q22. What are metalloids? Give their positions in the periodic table. **11201082**

Ans. Elements which have the properties some of the metals and some of the non metals are called metalloids. They are also referred as semi metals. Position: They are present in the P-block in "Stair-step line" arrangement including B from IIIA and to the Po of VIA.

Blocks in Periodic Table

Q23. Why some of the elements are called s-block elements and some are called p-block? **11201083**

Ans. Elements can be classified as s, p, d and f-block elements, because of their valance electrons are ended up in respective sub shell. If the valance electron are ended up in s-sub shell, then they are called s-block elements. And if their valance electrons are present in p-sub shell then they are called p-block elements.

Q24. What are lanthanides and Actinides? **11201084**

Ans. The series of 14 elements after the element Lanthanum -57 are called Lanthanides and the 14 elements after the element Actinium -89 are called Actinides. Both these series are also called f-block elements.

Families in Periodic Table

Q25. Name some families present in the periodic table. **11201085**

Ans. Some of the families in the periodic table are:

- Alkali metals (Li, Na, K, Rb, Cs, Fr)
- Alkaline earth metals (Be, Mg, Ca, Sr, Ba, Ra)
- Transition metals (Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn)
- Chalcogens (O, S, Se, Te, Po, Lv)
- Halogens (F, Cl, Br, I, At, Ts)
- Noble gases (He, Ne, Ar, Kr, Xe, Rn, Og)

Q26. Why 16 group elements are called Chalcogens? **11201086**

Ans. Chalogens word derived from two Greek words "Chalcos" means "copper" and "gen" means "form". So, group 16 elements are called Chalcogens are also called copper giver. Elements of this group are O, S, Se, Te, Po, Lv.

Q27. Why 17 group elements are called "Halogens"? **11201087**

Ans. The word "halogen" means "salt-former" because these elements easily react with alkali metals and alkaline earth metals to form stable halide salts. Elements include in this group are F, Cl, Br, I, At, Ts.

Q28. Why the noble gases are called unreactive elements? **11201088**

Ans. Due to their stable electronic configuration (complete outer most shell), they are almost entirely unreactive under normal conditions and rarely form compounds. Elements in this group are He, Ne, Ar, Kr, Xe, Rn, Og.

Q29. Differentiate between Cations and Anions. **11201089**

Ans.

Cations	Anions
When an atom carries positive charge, it forms cation.	When an atom carries negative charge, it forms anion.
It forms by losing electron.	It forms by accepting electron.
It's size is smaller than its parental size. eg: Li^+, Na^+	It's size is greater than its parental size. eg: F^-, Cl^-

Q30. Explain "Diagonal relationship" of the elements with suitable example. **11201090**

Ans. Elements diagonally positioned in the periodic table show some similarities in "diagonal relationship" despite the fact that they are present in different groups. For example lithium and magnesium, sodium and calcium have some common physical and chemical properties like sizes of their atoms and ions, charge densities, polarizing powers etc.

	IA	IIA	IIIA	IVA
Period 2	Li	Be	B	C
Period 3	Na	Mg	Al	Si

Ionization Energy, Electron Affinity

Q31. Define ionization energy. Give example. **11201091**

Ans. It is the amount of energy which is required to remove an electron from outermost shell of an isolated gaseous atom to form positive ion.

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^- \qquad \Delta H_{i1} = +496\ kJ/mol$$

Q32. What is "Spin-pair repulsion"? **11201092**

Ans. Those electrons which are present in same orbital experience some repulsion. It is easy to remove electron that is present in the paired form as compared to those electrons which are present in unpaired form due to spinning of electrons. It is called "Spin-paired repulsion".

$$Na_{(g)} \longrightarrow Na^+_{(g)} + e^- \qquad \Delta H_{i1} = +1403kJ/mol$$

$$O_{(g)} \longrightarrow O^+_{(g)} + e^- \qquad \Delta H_{i1} = +1365kJ/mol$$

Q33. Why the ionization energy of Be is 899kJ/mol while that of B is 801kJ/mol although the size of B is smaller than Be? **11201093**

Ans. He has 4 electrons as $1s^2$, $2s^2$ which makes it stable configuration of filled s-sub shell while B has five electron as $1s^2 \quad 2s^2 \quad 2p_x^1$ and it's configuration shows that it is relatively unstable due to one electron in its p-sub shell. That's why its value is 801kJ/mol while that of Be is 899kJ/mol even that it has smaller size than Be.

Q34. Give trend of electron affinity in a period. **11201094**

Ans. Electron affinity value increases in a period as we move from left to right, because of smaller size and increase in effective nuclear charge (z-effect or proton effect).

Electronegativity

Q35. Define electronegativity. Give trend in 2nd period. **11201095**

Ans. It is the power of an atom to attract the shared pair of electrons towards itself in a molecule trend of electronegativity in the 2nd period is increased from left to right.

2nd Period	Li	Be	B	C	N	O	F
E.N	1	1.5	2	2.5	3	3.5	4

Metallic Character

Q36. Define metallic character? Give its trend. **11201096**

Ans. Metallic character of an element is the tendency to lose electron. It increases down the group and decreases in a period.

Oxide

Q37. Na can form normal and peroxides. Give reactions. **11201097**

Ans. Sodium can react with O_2 vigorously in air to give peroxide while in limited amount of oxygen it forms normal oxide.

$$2Na + O_2 \longrightarrow Na_2O_2 \qquad (Peroxide)$$

$$4Na + O_2 \longrightarrow 2Na_2O \qquad (Normal\ oxide)$$

Q38. What are acidic oxides? **11201098**

Ans. Non metallic oxides are acidic in nature because when they react with water, they form acids.

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

Q39. What are basic oxides? **11201099**

Ans. Metallic oxides are basic in nature because when they react with water, they form base.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Q40. What are amphoteric oxides? **11201100**

Ans. These oxides which act as an acid and as well as a base are called amphoteric oxides.

Example: ZnO, Al_2O_3 and BeO

$$A\ell_2O_3 + 6HC\ell \longrightarrow 2A\ell C\ell_3 + 3H_2O$$

$$A\ell_2O_3 + 2NaOH \longrightarrow 2NaA\ell O_2 + H_2O$$

Oxidation State

Q41. Calculate the oxidation state of P in P_4O_6 and in P_4O_{10}. **11201101**

Ans. P_4O_6:

4(O.N of P) +6(O.N of O) = 0

4(O.N of P) +6(-2) =0

4(O.N of P)-12 = 0

4(O.N of P) = 12

$$P = \frac{12}{4} = +3$$

P_4O_{10}:

$$4(O.N\ of\ P) + 10(O.N\ of\ O) = 0$$

$$4(O.N\ of\ P) + 10(-2) = 0$$

$$4(O.N\ of\ P) - 20 = 0$$

$$4P - 20 = 0$$

$$P = \frac{20}{4}$$

$$P = +5$$

DESCRIPTIVE QUESTIONS (EXERCISE)

Q.3 Write equations for the reactions of Na and Mg with oxygen, chlorine and water. Compare the reactivity of both elements with these in terms of metallic character. **11201102**

Ans. See Q11. of theory.

Q.4 Explain with the help of equations, acidic and basic behaviour of oxides and chlorides. **11201103**

Ans. See Q12. and Q13. of theory.

Q.5 Describe the factors affecting and periodic trends of electron affinity. **11201104**

Ans. See Q8. of theory.

Q.6 Define ionization energy. Discuss the factors affecting and periodic trends of ionization energy. **11201105**

Ans. See Q7. of theory.