CLASS 11 CHEMISTRY

CHAPTER 3 – CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

Supplementary and Advanced Notes

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1. Concept Behind Periodicity

The periodic table organizes chemical elements so that elements with similar chemical behavior appear in the same column.
This periodicity arises because the number of valence electrons (the outermost electrons) repeats in a regular pattern.

When the valence shell configuration (ns^1, ns^2, ns^2 np^1 –) repeats, the chemical properties also repeat.

Thus:

Group !’ Same outer configuration !’ Similar chemical behavior.

Period !’ Same number of shells !’ Gradual variation of properties.

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2. Basis of the Modern Periodic Table

Property	Mendeleev’s Table	Modern Table
Basis of classification	Atomic mass	Atomic number
Law	Mendeleev’s Periodic Law	Modern Periodic Law
Periods	8 short periods	7 long periods
Groups	8	18
Position of isotopes	Could not be explained	Explained (same atomic number)
Noble gases	Not included originally	Included in Group 18

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3. Quantum Explanation of Periodicity

Each period corresponds to filling of a new principal energy level (n).

Period	Principal Quantum Number	Orbitals Being Filled	No. of Elements
1st	n = 1	1s	2
2nd	n = 2	2s, 2p	8
3rd	n = 3	3s, 3p	8
4th	n = 4	4s, 3d, 4p	18
5th	n = 5	5s, 4d, 5p	18
6th	n = 6	6s, 4f, 5d, 6p	32
7th	n = 7	7s, 5f, 6d, 7p	32

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4. Groups and Their Characteristics

Group 1 – Alkali Metals

Valence configuration: ns^1

Highly electropositive, soft metals, form +1 ions.

React vigorously with water forming hydroxides.

Group 2 – Alkaline Earth Metals

Valence configuration: ns^2

Harder than alkali metals, form +2 ions, less reactive.

Group 17 – Halogens

Valence configuration: ns^2 n p u

Strong oxidising agents, form salts with metals.

Group 18 – Noble Gases

Valence configuration: ns^2 n p v

Stable octet, very low reactivity, monoatomic gases.

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5. Origin of Periodic Trends (Scientific Explanation)

(a) Effective Nuclear Charge (Zeff)

The net positive charge experienced by an electron.
As we move across a period, Zeff increases, pulling electrons closer, reducing atomic radius.

(b) Shielding or Screening Effect

Inner electrons repel outer electrons, reducing attraction from the nucleus.
Down a group, shielding increases, which counteracts the increase in nuclear charge, causing atomic size to increase.

(c) Subshell Penetration

Order of penetration: s > p > d > f.
An s-electron is more strongly attracted to the nucleus than a p-electron in the same shell.

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6. Periodic Trends – Conceptual Explanation and Exceptions

6.1 Atomic Radius Exceptions

In transition elements, atomic radius changes very little across a period due to the counteracting effects of increasing nuclear charge and addition of electrons to the (n–1)d subshell.

Lanthanide contraction: steady decrease in atomic and ionic radii among lanthanides due to poor shielding by 4f electrons.

6.2 Ionization Energy Exceptions

The general trend is disturbed between:

Be ($1s^2$ $2s^2$) and B ($1s^2$ $2s^2$ $2p^1$)
!' B has lower IE due to 2p electron being farther from nucleus.

N ($2p^3$) and O (2 p t)
!' O has lower IE because pairing of electrons causes extra repulsion.

6.3 Electron Affinity Exceptions

Fluorine has less negative electron affinity than chlorine due to small size and inter-electronic repulsion in 2p orbitals.

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7. Numerical Examples and Short Calculations

Example 1

Question: Which of the following has the largest atomic radius: Na, Mg, Al, Si?
Answer:
Across the period, atomic radius decreases.
Therefore, Na > Mg > Al > Si !' Sodium has the largest radius.

Example 2

Question: Arrange the following in increasing ionization energy: Li, Na, K.
Answer:
Down the group, ionization energy decreases.
Hence, K < Na < Li.

Example 3

Question: Among O^{2-}, F^-, Na^+, Mg^{2+}, which has the smallest radius?
Answer:
All are isoelectronic (10 electrons). The one with the highest nuclear charge (Mg^{2+}) is smallest.

Order: $O^{2-} > F^- > Na^+ > Mg^{2+}$.

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8. Chemical Reactivity Trends Explained

Metals: Reactivity increases down the group (easier electron loss).

Non-metals: Reactivity decreases down the group (harder electron gain).

Transition metals: Reactivity varies irregularly due to variable oxidation states.

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9. Nature of Oxides Across a Period

Element	Type of Oxide	Nature
Na	Na_2O	Basic
Mg	MgO	Basic
Al	Al_2O_3	Amphoteric
Si	SiO_2	Weakly acidic
P	P_4O_{10}	Acidic
S	SO_2	Acidic
Cl	Cl_2O_7	Strongly acidic

As we move across a period, basic character decreases and acidic character increases.

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10. Concept of Isoelectronic Species

Species having the same number of electrons but different nuclear charges.
Example: O^{2} {(8 protons, 10 e {), F {(9 protons, 10 e {), N a (11 protons, 10 e {) .
In such species, size decreases with increasing atomic number because nuclear pull is greater.

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11. Applications of Periodicity

1. Predicting the chemical reactivity of elements.

2. Determining the types of bonds formed (ionic or covalent).

3. Explaining trends in acid-base properties of oxides.

4. Estimating ionization energies and electron affinities.

5. Predicting oxidation states of transition metals.

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12. Board-Style Question–Answer Notes

Short Answer (2 Marks)

Q1. Why are elements arranged according to atomic number in the modern periodic table?
Because atomic number uniquely identifies an element and defines its electronic configuration, which governs its properties.

Q2. Why is ionization enthalpy of nitrogen greater than oxygen?
Nitrogen ($2p^3$) has a half-filled p-subshell which is more stable; hence, more energy is required to remove one electron.

Q3. Explain diagonal relationship between lithium and magnesium.
Both have similar ionic sizes and polarising power, leading to similar solubility and chemical behaviour.

Long Answer (5 Marks)

Q1. Discuss the periodic trends in ionization enthalpy, electron gain enthalpy, and electronegativity.
Across a period:

Ionization enthalpy increases due to higher nuclear charge.

Electron gain enthalpy becomes more negative because atoms can gain electrons more easily.

Electronegativity increases because smaller atoms attract bonding electrons more strongly.
Down a group: opposite trends occur due to increased atomic size and shielding.

Q2. What are the defects of Mendeleev's periodic table and how are they removed in the modern table?
Defects:

1. Position of isotopes not explained.

2. Wrong order of some pairs (Co–Ni).

3. Unclear position of hydrogen.
Modern table solves these using atomic number as the basis, which is unique and avoids these anomalies.

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13. Periodic Trends in Physical Properties

Property Explanation

Density Generally increases down a group because atomic mass increases more than atomic volume.
Melting/Boiling Point For metals, usually increases down the group; for non-metals, decreases down the group.
Oxidation States Transition elements show variable oxidation states due to similar energies of (n–1)d and ns electrons.

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14. Concept of Periodicity in Bonding and Valency

Elements in the same group form compounds with analogous formulas:
$NaCl$, KCl, $RbCl$ – Group 1 halides
MgO, CaO, SrO – Group 2 oxides
This is because of the same number of valence electrons.

Valency depends on the number of electrons in the outermost shell (for main group elements).

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15. Advanced Notes – f-Block Contraction

The lanthanide contraction is the gradual decrease in atomic and ionic radii from lanthanum (La) to lutetium (Lu).
Cause: Poor shielding of 4f electrons.
Consequence:

1. Zr and Hf have almost identical radii.

2. Transition metals of 4d and 5d series exhibit similar properties.

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16. Common Conceptual Misconceptions

1. Electron affinity and electronegativity are not the same.

Electron affinity is an energy change; electronegativity is a tendency.

2. Period number does not indicate number of valence electrons but the number of shells.

3. Atomic radius is not the same as covalent or van der Waals radius.

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17. Concept Summary Chart

Property	Reason for Periodic Variation
Atomic size	Balance of nuclear charge and shielding.
Ionization enthalpy	Stability of half or fully filled orbitals.
Electron affinity	Attraction between nucleus and incoming electron.
Electronegativity	Ability to attract shared electron pair.
Metallic character	Tendency to lose electrons.

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18. Practice Questions for Board Preparation

1. Define effective nuclear charge.

2. Explain why the second ionization enthalpy is always greater than the first.

3. Arrange the following in decreasing electron gain enthalpy: O, S, Se.

4. Which element has the least electronegativity in the periodic table?

5. Why are noble gases inert?

6. Explain the term periodicity in detail.

7. State and explain the trends of metallic and non-metallic character in the periodic table.

8. Give two differences between modern and Mendeleev’s periodic tables.

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19. Chapter Wrap-Up

The periodic table is not only a list of elements but a comprehensive map of chemical behavior.
Every property—size, ionization energy, electronegativity, reactivity—can be predicted if one understands the trends and underlying principles.
Mastering periodicity builds the foundation for all of inorganic chemistry.