

UNIT 7: Periodic Properties in Detail

This unit focuses on various periodic properties of elements, how they vary across the periodic table, and their significance in chemistry.

1. Effective Nuclear Charge (Z^*)

Definition:

- The **effective nuclear charge (Z^*)** is the **net positive charge experienced by an electron** in a multi-electron atom.
- Due to electron shielding (inner electrons repelling outer electrons), the outer electrons **do not experience the full nuclear charge (Z)**.
- **Formula (Slater's Rule Approximation):** $Z^* = Z - S Z^* = Z - S$ where:
 - Z = Actual nuclear charge (number of protons)
 - S = Shielding constant (depends on electron configuration)

Variation in Periodic Table:

- **Across a Period (\rightarrow)** $\rightarrow Z^*$ increases as protons increase, but shielding remains relatively constant.
- **Down a Group (\downarrow)** $\rightarrow Z^*$ slightly decreases due to increased shielding by inner electrons.

Significance: Affects **atomic size, ionization energy, and electronegativity**.

2. Variation of Orbital Energies (s, p, d, f orbitals)

Key Concepts:

- **Energy of orbitals depends on:**
 - Principal quantum number (n)
 - Effective nuclear charge (Z^*)
 - Shielding and penetration effect
 - Electron-electron repulsions

Trends in Orbital Energy:

Orbital Type	Energy Trend	Reason
s-orbitals	Lowest energy	Closest to nucleus, strong attraction
p-orbitals	Higher than s	More shielded, less penetration
d-orbitals	Higher than p	Even more shielded
f-orbitals	Highest energy	Poor shielding, weak nuclear attraction

- Significance:** Explains **electron configuration stability** (e.g., why 4s fills before 3d).
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3. Electronic Configuration

Definition:

- The **arrangement of electrons** in the atomic orbitals following **Aufbau's principle**, **Hund's Rule**, and **Pauli's exclusion principle**.

General Order of Filling (Aufbau Principle):

$1s^2 \rightarrow 2s^2 \rightarrow 2p^6 \rightarrow 3s^2 \rightarrow 3p^6 \rightarrow 4s^2 \rightarrow 3d^{10} \rightarrow 4p^6 \rightarrow 5s^2 \rightarrow 4d^{10} \rightarrow 5p^6 \rightarrow 6s^2 \rightarrow 4p^6 \rightarrow 5s^2 \rightarrow 3p^6 \rightarrow 4s^2 \rightarrow 3d^{10} \rightarrow 4p^6 \rightarrow 5s^2 \rightarrow 4d^{10} \rightarrow 5p^6 \rightarrow 6p^6 \rightarrow 7s^2 \rightarrow 3p^6 \rightarrow 4s^2 \rightarrow 3d^{10} \rightarrow 4p^6 \rightarrow 5s^2 \rightarrow 4d^{10} \rightarrow 5p^6 \rightarrow 6p^6 \rightarrow 7p^6$

Special Cases:

- **Cr (Z=24):** $4s^1 3d^5 4s^1$ instead of $4s^2 3d^4 4s^2$ (half-filled d-subshell stability)
- **Cu (Z=29):** $4s^1 3d^{10} 4s^1$ instead of $4s^2 3d^9 4s^2$ (full d-subshell stability)

- Significance:** Determines **chemical reactivity and bonding**.
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4. Atomic & Ionic Sizes

Definition:

- **Atomic Radius:** Half the distance between nuclei of two identical bonded atoms.
- **Ionic Radius:** Size of an ion (cation or anion).

Trends in Atomic Size:

- **Across a Period (\rightarrow): Decreases** (due to increased Z^*).
- **Down a Group (\downarrow): Increases** (due to higher energy levels).

Trends in Ionic Size:

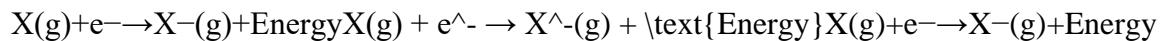
- **Cations (+ve charge): Smaller** than parent atom (loss of electrons, increased Z^*).
- **Anions (-ve charge): Larger** than parent atom (electron repulsion increases).

- Significance:** Influences **bonding, solubility, and crystal structure**.

5. Electron Affinity (EA)

Definition:

- The energy change when an electron is added to a neutral atom in the gas phase.



- First EA is exothermic (energy released), but second EA is endothermic (energy absorbed).

Trends in Electron Affinity:

- Across a Period (\rightarrow): Increases (except noble gases).
- Down a Group (\downarrow): Decreases (due to increased atomic size).
- Most negative EA: Chlorine (Cl), not Fluorine (F), due to small size of F causing electron repulsion.

Significance: Determines oxidizing power (O_2 & Cl_2 have high EA).

6. Electronegativity (χ)

Definition:

- The ability of an atom to attract electrons in a chemical bond.
- Measured using Pauling Scale (F = 4.0, most electronegative).

Trends in Electronegativity:

- Across a Period (\rightarrow): Increases (higher Z^*).
- Down a Group (\downarrow): Decreases (shielding increases).

Significance: Explains bond polarity, dipole moment, and reactivity.

7. Polarizability

Definition:

- The ease with which an electron cloud is distorted by an external electric field.

Trends:

- Larger atoms → More polarizable (more loosely held electrons).
- Smaller atoms & cations → Less polarizable (tightly held electrons).
- Highly charged anions (like I^- , S^{2-}) → Very polarizable.

Significance: Explains London forces, van der Waals interactions, and solubility.

8. Oxidation States

Definition:

- The charge an atom appears to have in a compound, based on electron transfer.

General Trends:

- **s-block (Group 1, 2): Fixed oxidation states** (+1 for alkali, +2 for alkaline earth metals).
- **p-block:** Shows **multiple oxidation states** (e.g., C = -4 to +4, N = -3 to +5).
- **d-block (Transition Metals): Variable oxidation states** (e.g., Fe^{2+} , Fe^{3+}).

Significance: Determines redox reactions and chemical bonding.

Outcomes:

- Understanding **Effective Nuclear Charge (Z^*)** and how it influences properties.
- Learning **orbital energy variations** and their role in **electronic configuration**.
- Explaining trends in **atomic size, ion size, electronegativity, and electron affinity**.
- Understanding **polarizability and oxidation states** in periodic chemistry.