

## 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.

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### 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):**  $Z^* = Z - S$  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**.

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### 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	<b>Lowest energy</b>	Closest to nucleus, strong attraction
p-orbitals	Higher than s	More shielded, less penetration
d-orbitals	Higher than p	Even more shielded
f-orbitals	<b>Highest energy</b>	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):

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

#### Special Cases:

- Cr (Z=24):**  $4s^1 3d^5$  instead of  $4s^2 3d^4$  (half-filled stability)
- Cu (Z=29):**  $4s^1 3d^{10}$  instead of  $4s^2 3d^9$  (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 (→): Decreases** (due to increased  $Z^*$ ).
- Down a Group (↓): 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).

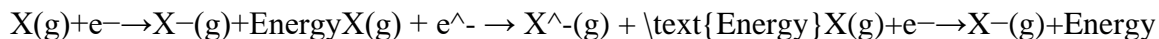
- ✔ **Significance:** Influences **bonding, solubility, and crystal structure**.

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## 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 (→): Increases** (except noble gases).
- Down a Group (↓): 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<sub>2</sub> & Cl<sub>2</sub> have high EA).

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## 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 (→): Increases** (higher Z\*).
- Down a Group (↓): Decreases** (shielding increases).

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

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## 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**.

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## 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**.

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### 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.