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−SZ^* = Z - SZ∗=Z−S where:
  - **Z** = Actual nuclear charge (number of protons)
  - **S** = Shielding constant (depends on electron configuration)

Variation in Periodic Table:

- **Across a Period (→) → Z* increases** as protons increase, but shielding remains relatively constant.
- **Down a Group (↓) → 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	**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):

1s2→2s2→2p6→3s2→3p6→4s2→3d10→4p6→5s2→4d10→5p61s^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^61s2→2s2→2p6→3s2→3p6→4s2→3d10→4p6→5s2→4d10→5p6

Special Cases:

- **Cr (Z=24):** 4s13d54s^1 3d^54s13d5 instead of 4s23d44s^2 3d^44s23d4 (half-filled stability)
- **Cu (Z=29):** 4s13d104s^1 3d^{10}4s13d10 instead of 4s23d94s^2 3d^94s23d9 (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.**

5. Electron Affinity (EA)

Definition:

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

X(g)+e−→X−(g)+EnergyX(g) + e^- → X^-(g) + \text{Energy}X(g)+e−→X−(g)+Energy

- **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_2 & Cl_2 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.**

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.