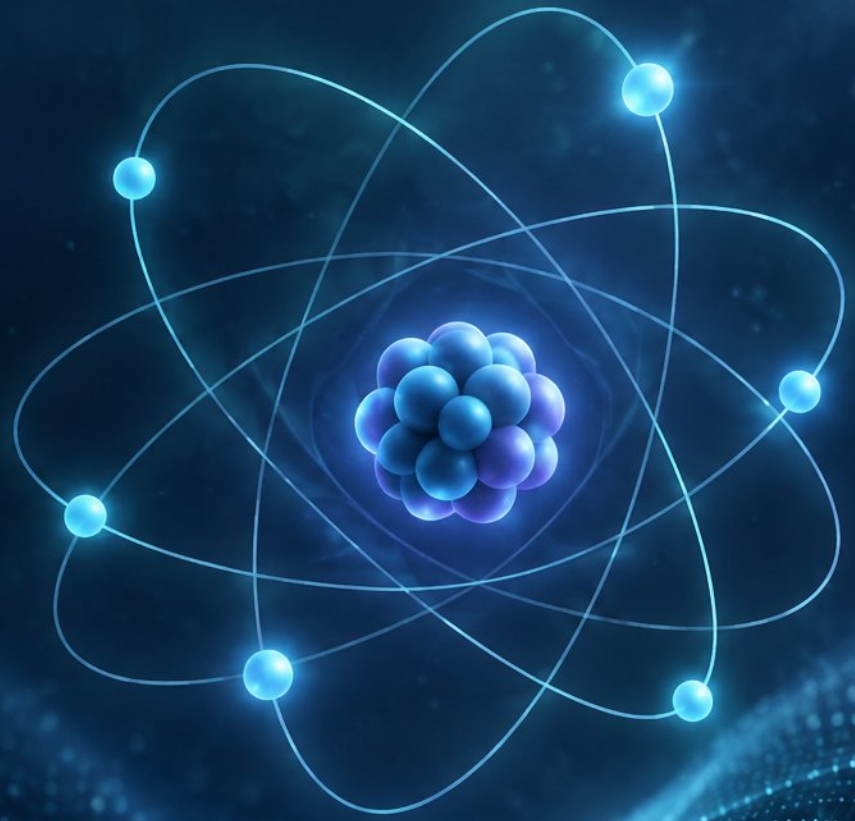
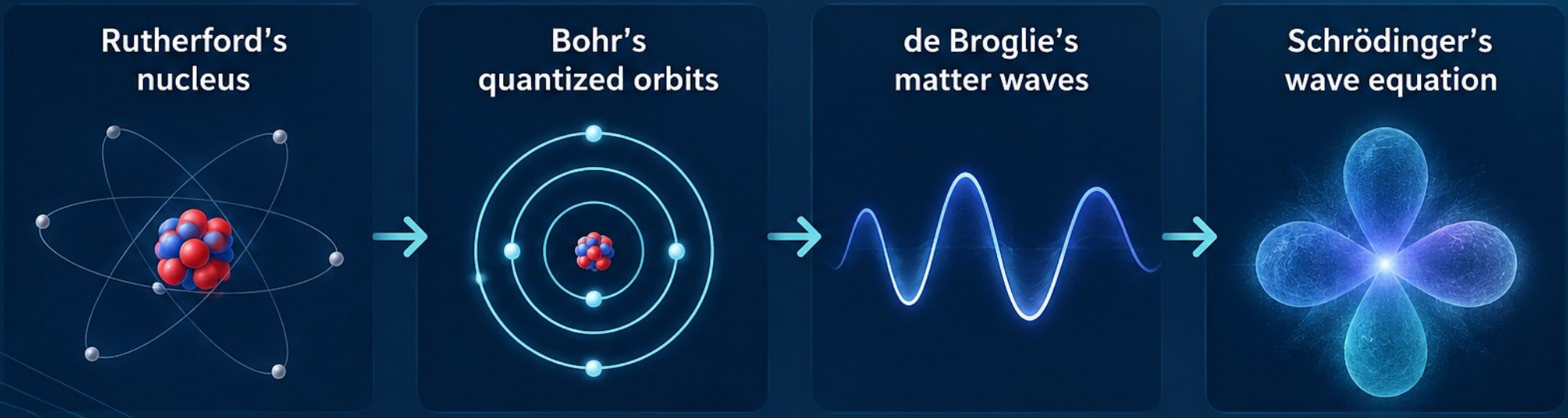
Electron Configuration:
Shells, Subshells, and Orbitals



- Explains bonding, periodic trends, reactivity, and spectroscopy
- Three-level vocabulary: shells (n), subshells (l), orbitals (m_l, m_s)
- Electrons live in probability distributions, not planetary orbits
- Conceptual foundations, worked examples, and analogies for all learners
- Roadmap: quantum numbers, filling rules, notation, trends, misconceptions, practice

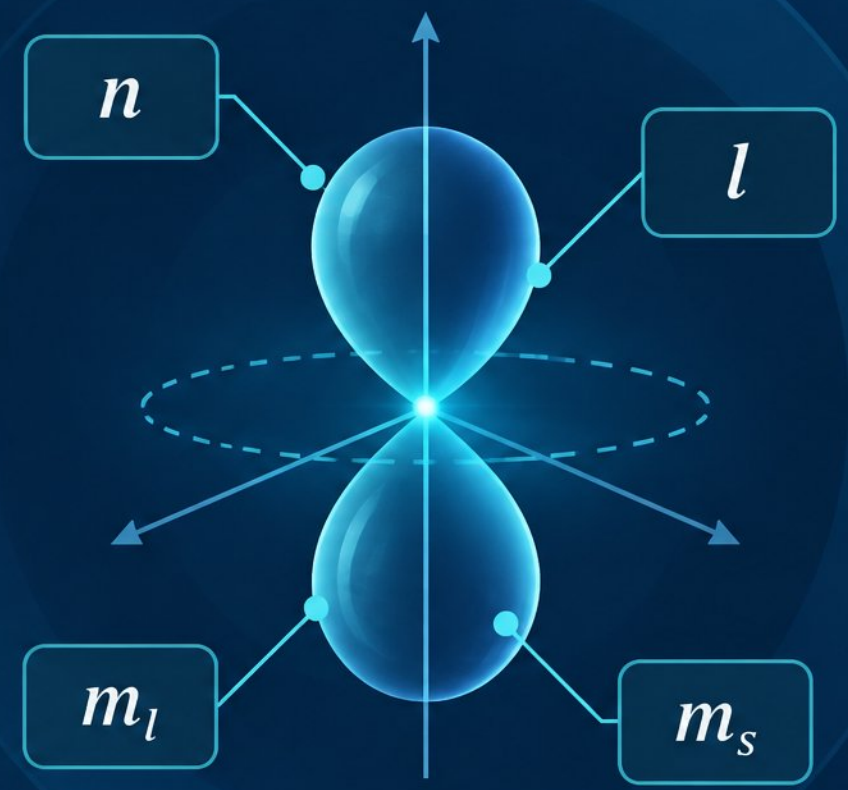
From Bohr Orbits to Quantum Mechanical Orbitals



- Rutherford's nucleus → Bohr's quantized orbits → de Broglie's matter waves → Schrödinger's wave equation
- Bohr explained hydrogen's spectrum but failed for helium and larger atoms
- Also failed: fine structure and the Zeeman effect
- Wavefunction ψ: orbitals show where electrons are most likely found
- Orbital shapes (s, p, d) are 95% boundary surfaces — not planet-like orbits

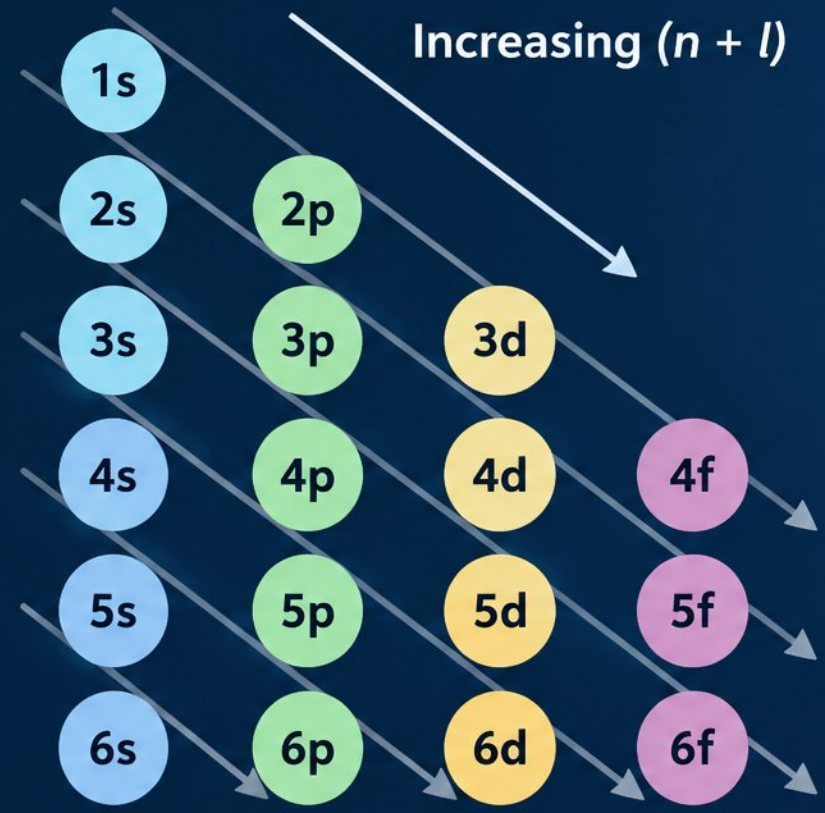
Quantum Numbers and What They Describe

- **Principal quantum number n:** shell, orbital size, and energy level; positive integers from 1
- **Angular momentum quantum number l:** subshell shape; allowed values 0 to $n - 1$
- **Magnetic quantum number m_l:** orbital orientation; $2l + 1$ values from $-l$ to $+l$
- **Spin quantum number m_s:** electron spin, $\pm 1/2$; basis of the Pauli exclusion principle
- **Worked practice:** assign n, l, m_l, and m_s to a specific electron in a given subshell



Shells and Subshells: Capacities and Energy Ordering

- Shell maximum: $2n^2$ electrons; subshells hold $s = 2$, $p = 6$, $d = 10$, $f = 14$
- Fill by $(n + l)$ rule: 1s, 2s, 2p, 3s, 3p, 4s, then 3d
- 4s fills before 3d: penetration and shielding lower its energy in early Period 4 atoms
- Penetration order $s > p > d > f$; within a shell, energies rise $s < p < d < f$,
- Aufbau ladder shows filling order, not precise energy gaps



Subshell Capacities

s = 2 p = 6 d = 10 f = 14

Filling Rules and Standard Notation

- **Aufbau principle:** lowest-energy subshell fills first (1s → 2s → 2p → 3s…)
- **Pauli exclusion:** each orbital holds at most two electrons with opposite spins
- **Hund's rule:** degenerate orbitals get one electron each, parallel spins, before pairing
- **Notation example:** $1s^2\ 2s^2\ 2p^6$; "$2p^4$" means four electrons in the n = 2 p subshell
- **Orbital box diagrams:** arrows map directly to configuration superscripts

Example: $1s^2\ 2s^2\ 2p^6$



Electron configuration:

$1s^2\ 2s^2\ 2p^6$

Worked Examples: Applying the Filling Rules

- Walk through H → C → N → O → Ne: Aufbau, Pauli, and Hund in action
- **Carbon:** $1s^2\ 2s^2\ 2p^2$ with two unpaired p electrons
- **Nitrogen:** $1s^2\ 2s^2\ 2p^3$, all three p orbitals singly occupied
- **Oxygen:** $1s^2\ 2s^2\ 2p^4$, one p orbital paired, two unpaired
- **Transition metals Sc–Zn:** 4s fills before 3d, small energy gap

Orbital box diagrams for C, N, and O

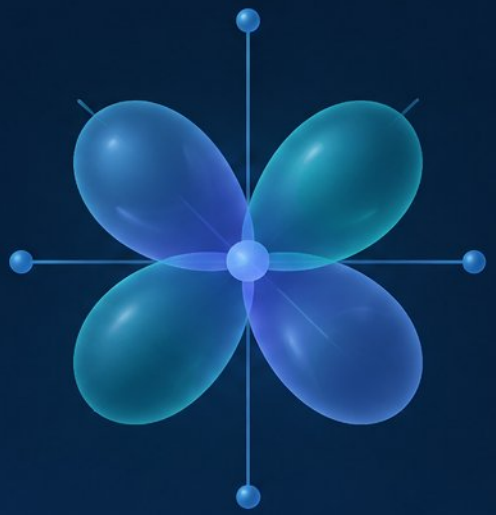


Always check: superscript sum must equal atomic number Z

Exceptions to the Predicted Filling Order

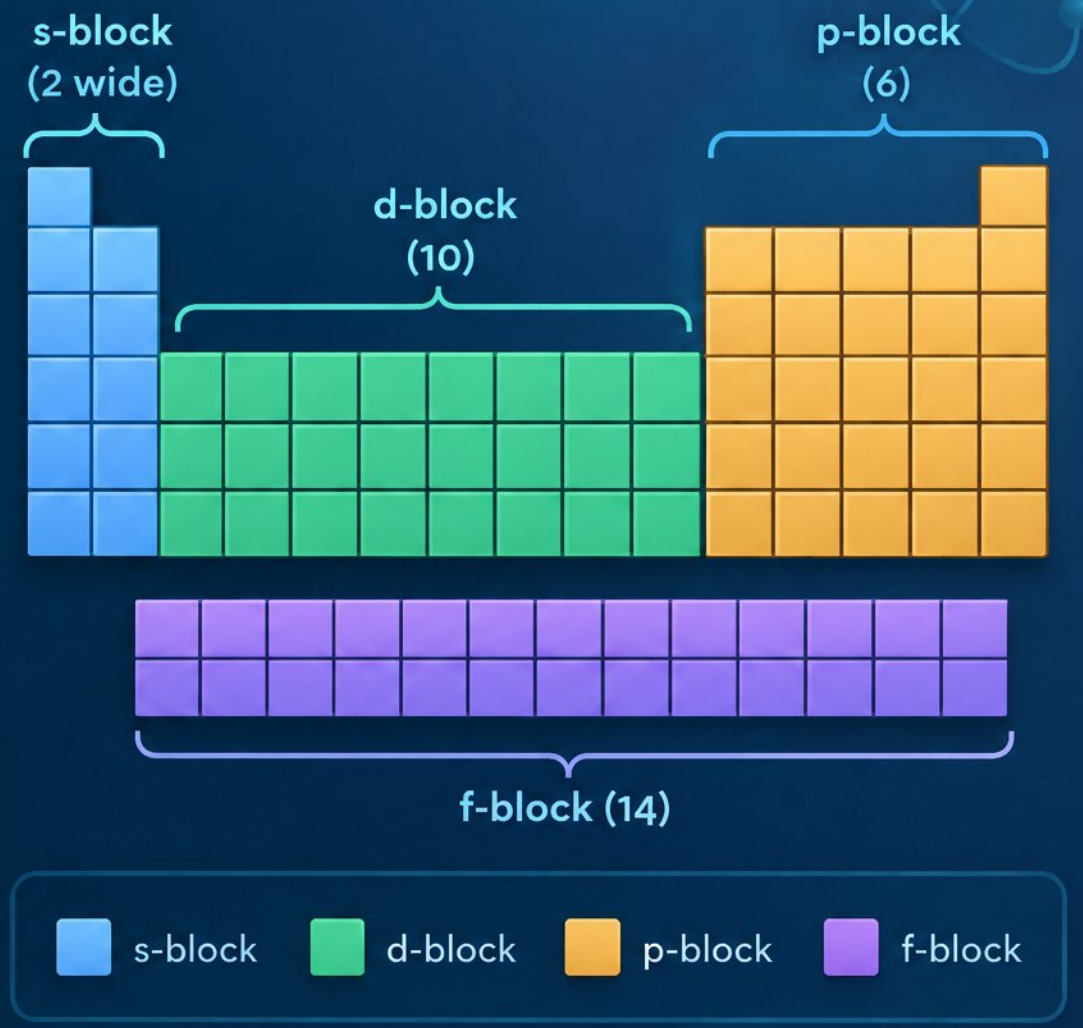
	Predicted Configuration	Observed Configuration
Chromium (Z = 24)	[Ar] $3d^4$ $4s^2$ 	[Ar] $3d^5$ $4s^1$ 
Copper (Z = 29)	[Ar] $3d^9$ $4s^2$ 	[Ar] $3d^{10}$ $4s^1$ 

- Chromium (Z = 24): predicted [Ar] $3d^4$ $4s^2$, observed [Ar] $3d^5$ $4s^1$
- Copper (Z = 29): predicted [Ar] $3d^9$ $4s^2$, observed [Ar] $3d^{10}$ $4s^1$
- Half-filled and filled d subshells gain stability; it's a balance of repulsion and a small 4s/3d gap
- Not a new rule: Nb, Mo, Ru, Rh, Pd, Ag, Pt, Au, and Lr also deviate
- Present exceptions as experimental facts; avoid saying 'half-filled is always more stable'



Electron Configuration and the Periodic Table

- Blocks map to the subshell being filled: s (2 wide), p (6), d (10), f (14)
- Group and period positions reveal valence configuration; same group, same valence count
- Derive configurations by following increasing Z through each block
- Valence electrons drive chemistry; core electrons rarely react
- Main group, transition, and inner transition elements reflect the last orbital filled



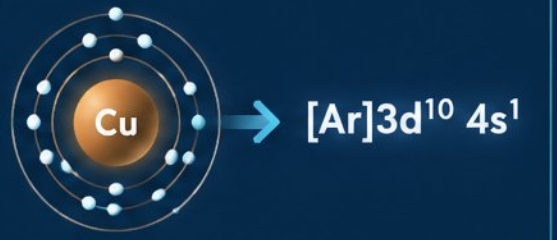
Noble-Gas Shorthand and Configurations of Ions

- **Shorthand:** bracket the preceding noble gas, then show outer electrons — Na is $[Ne]3s^1$, Cu is $[Ar]3d^{10}\ 4s^1$
- **Cations lose highest-n electrons first:** Fe^{2+} is $[Ar]3d^6$, not $[Ar]3d^4\ 4s^2$
- **Anions add electrons to the lowest available subshell:** O^{2-} is $1s^2\ 2s^2\ 2p^6$
- **Common slips:** wrong core gas, misordered superscripts, forgetting the 4s→3d switch for ions
- **Always verify:** total electrons must equal Z minus the ion charge

Sodium (Na)



Copper (Cu)



Iron(II) (Fe^{2+})

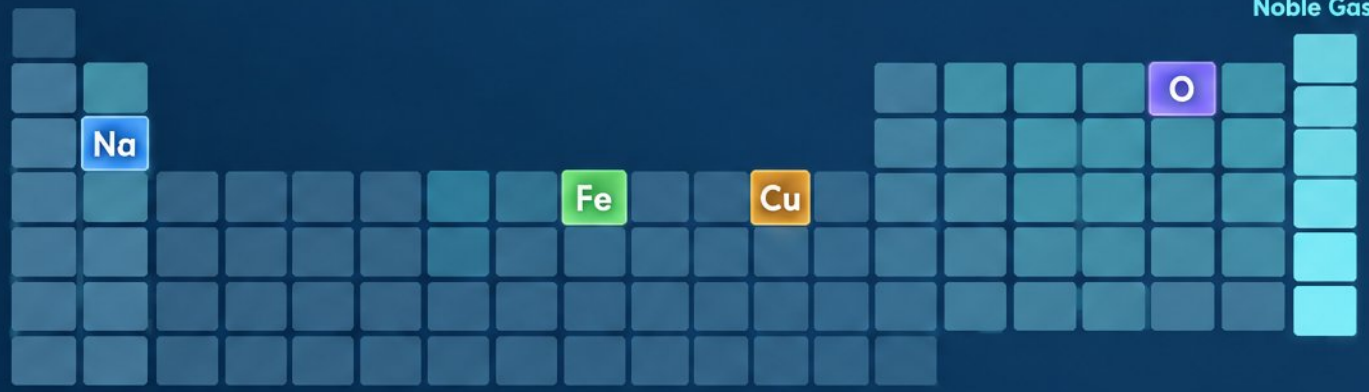


Oxide (O^{2-})



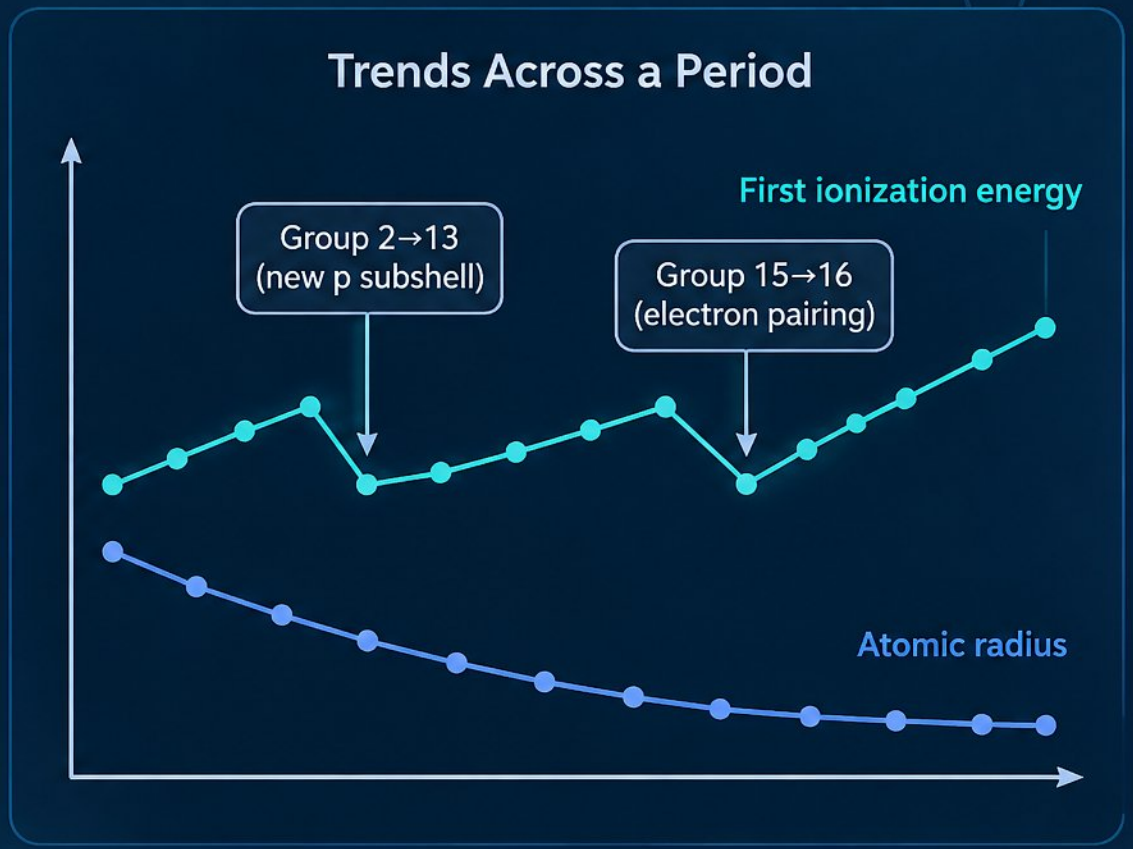
The Big Picture

Noble gases provide a stable core for shorthand notation.



Configuration and Periodic Trends

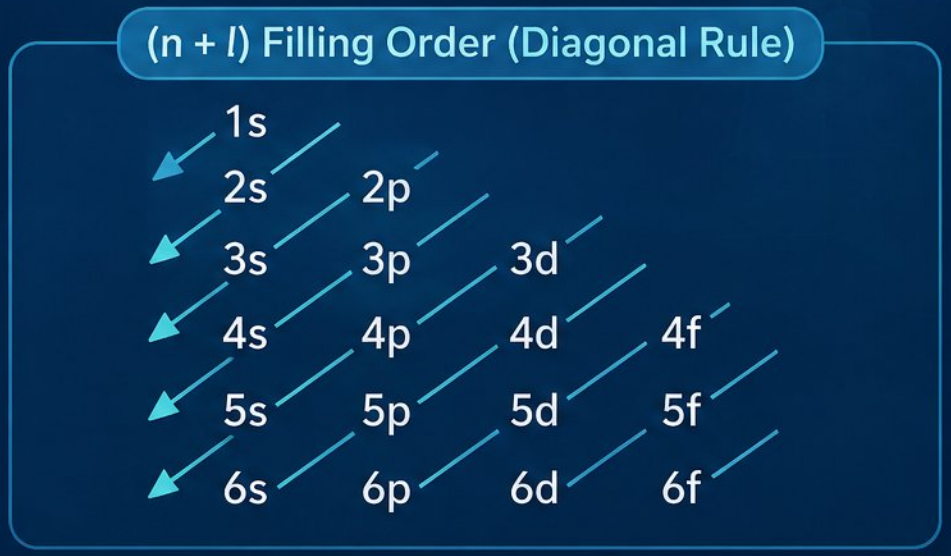
- Atomic radius shrinks across a period, grows down a group (effective nuclear charge, added shells)
- Ionization energy rises across a period, falls down a group (distance, shielding)
- Two dips: Group 2→13 (new p subshell) and Group 15→16 (electron pairing)
- Successive ionization jumps confirm the number of valence electrons
- Electronegativity and metallic character follow the same configuration logic



Reading and Writing Configurations: Guided Practice

1. Find Z
2. Fill by (n + *l*)
3. Apply Pauli and Hund
4. Abbreviate with noble-gas core

- Four steps: find Z, fill by (*n* + *l*), apply Pauli and Hund, abbreviate with noble-gas core
- Practice: sulfur atom, Fe^{3+} cation, and Cl^- anion, with full worked answers
- Watch for slips: filling 3d before 4s, wrong noble-gas core, ignoring 4s removal in cations
- Orbital diagrams: count unpaired electrons to confirm Hund's rule and predict magnetism
- Self-check: total electrons equals Z, and valence shell matches the periodic-table block









Common Misconceptions About Electrons and Orbitals

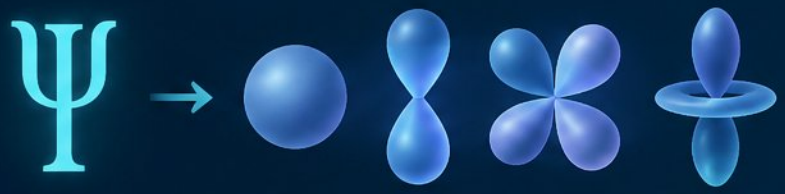
Misconception: electrons orbit like planets.
Correction: orbitals are probability distributions from a wavefunction.



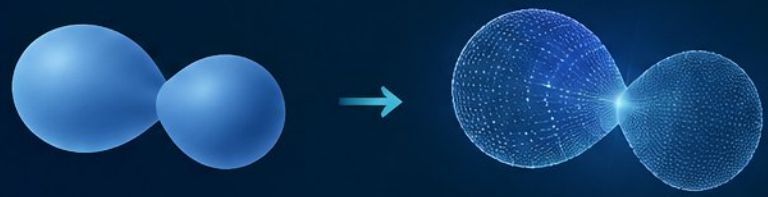
Misconception: orbitals fill in strict shell order.
Correction: ($n + l$) order and observed exceptions apply.



Misconception: quantum numbers are arbitrary labels.
Correction: each reflects energy, shape, orientation, or spin.



Misconception: orbitals are concrete objects.
Correction: they are mathematical models with boundary surfaces.



Misconception: ions fill like neutral atoms.
Correction: cations lose 4s electrons before 3d.



Teaching Strategies and Analogies That Support Learning

Clear Models
Deeper Understanding
Lasting Learning

Analogies with explicit limits:
"hotel with floors and rooms"
state where the model breaks down

Sequence deliberately:
shells and subshells → quantum numbers → filling rules → notation → periodic trends

Formative assessment targeting concepts:
ask learners to explain why 4s fills before 3d

Adult learners:
connect to prior real-world knowledge; scale worked examples to build confidence

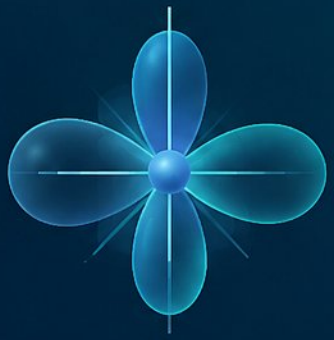
Degree-track students:
add formal quantum-number problem sets

Scaffold formal-thinking readiness
with concrete examples and visual models first

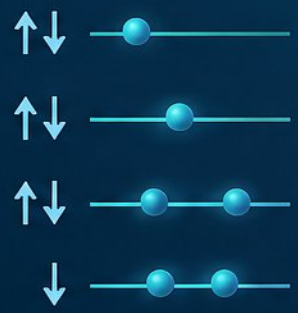


Recap, Self-Assessment, and Next Steps

Quantum Numbers define orbitals



Filling Rules define configurations



Configurations explain periodic trends



Self-check quiz:
notation, ions, exceptions, and trends, with worked answers

Next:
molecular orbital theory, hybridization, and spectroscopy

Resources:
OpenStax Chemistry 2e, RSC Education guides, NIST atomic spectroscopy

Educators:
ready-to-use sequence from shells to periodic trends

PersonWise.

PersonWise • AI digital-human course platform



Scan the code or click anywhere to watch this course live, with voice Q&A

personwise.ai/t/d6da9548/public