

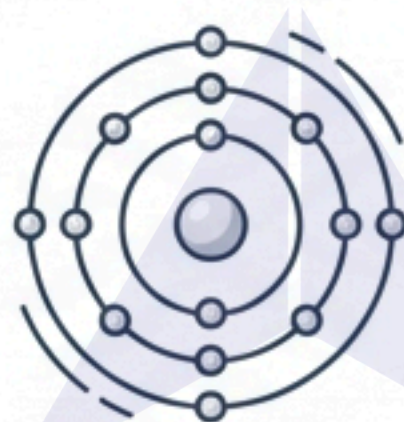
UNIT 2 • STRUCTURE OF ATOM

2.6.5 Electronic Configuration of Atoms

Mapping the Quantum Blueprint:
From Notation to Orbital Diagrams

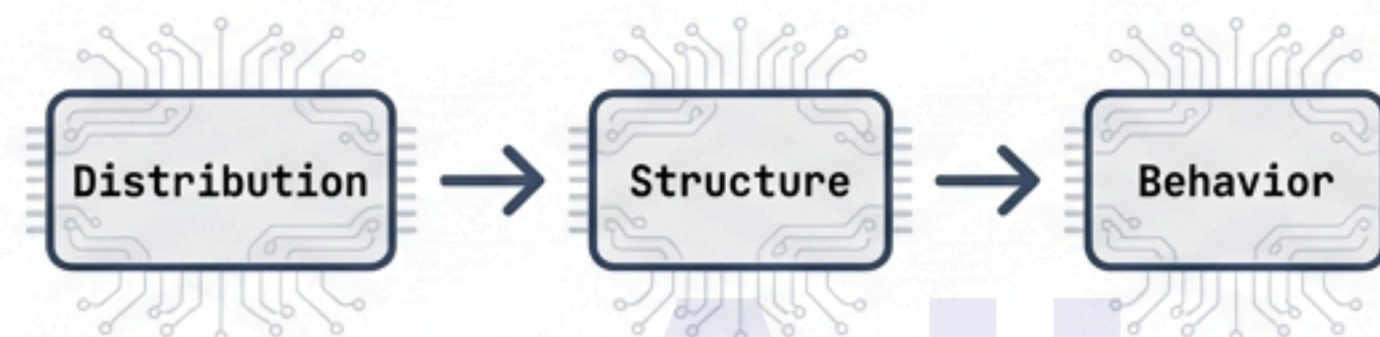


What Electronic Configuration Means



The Distribution of Electrons

The distribution of electrons into the various orbitals of an atom is called its electronic configuration. It is the ultimate map of an atom's internal structure.



Why It Matters

The rich diversity of chemical behavior across different elements—why some are reactive metals, why some are inert gases, and how molecules form—can be traced directly back to this specific arrangement of electrons.

Takeaway: If you know an atom's configuration, you can predict its chemical destiny.

REPRESENTATION METHOD 1

Subshell Notation Method

The standard $s^a p^b d^c$ notation differentiates subshells by their principal quantum number and lists the electron count as a superscript.

Subshell Letter Symbol (ℓ)

Dictates the shape of the orbital (s, p, d, f).

Principal Quantum Number (n)

Defines the main energy shell.

1s²

Electron Count

The number of electrons occupying this specific subshell.

Example Progression: $1s^2 \rightarrow 2s^2 \rightarrow 2p^6 \dots$



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REPRESENTATION METHOD 2

Orbital Diagram Method

This visual method provides a deeper look into the atom by representing all four quantum numbers, including explicitly visualizing electron spin.



(Visual depicts Carbon's valence shell configuration)

The Box

Represents a single orbital within a subshell.

The Arrow (↑ or ↓)

Represents an electron and its magnetic spin quantum number (m_s).

The Advantage

Unlike subshell notation, orbital diagrams explicitly show how electrons pair up, demonstrating Pauli and Hund's rules in action.

Building the First 10 Elements (H to Ne)

Watch how the 1s and 2s fill completely before the progressive filling of the three degenerate 2p orbitals.

Element	Symbol (Z)	Notation	Orbital Box Diagram
Hydrogen	(Z=1)	$1s^1$	$\boxed{\uparrow}$
Helium	(Z=2)	$1s^2$	$\boxed{\uparrow\downarrow}$
Lithium	(Z=3)	$1s^2 2s^1$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow}$
Beryllium	(Z=4)	$1s^2 2s^2$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}$
Boron	(Z=5)	$1s^2 2s^2 2p^1$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{}$
Carbon	(Z=6)	$1s^2 2s^2 2p^2$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{}$
Nitrogen	(Z=7)	$1s^2 2s^2 2p^3$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{\uparrow} \quad \boxed{}$
Oxygen	(Z=8)	$1s^2 2s^2 2p^4$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow} \quad \boxed{}$
Fluorine	(Z=9)	$1s^2 2s^2 2p^5$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{}$
Neon	(Z=10)	$1s^2 2s^2 2p^6$	$\boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow} \quad \boxed{\uparrow\downarrow}$

Hund's Rule:
separate boxes

Shell complete

Core Electrons vs. Valence Electrons

As atoms get larger, writing full configurations becomes cumbersome.
We simplify this using noble gas shorthand.

Sodium (Na, Z=11) Full: $1s^2 2s^2 2p^6 3s^1$

Sodium (Na) Shorthand: $[\text{Ne}] 3s^1$

The Core $[\text{Ne}]$

- Electrons in completely filled inner shells.
- In Sodium, the first 10 electrons perfectly match the configuration of Neon ($1s^2 2s^2 2p^6$).
- Highly stable, they do not participate in chemical reactions.

The Valence $3s^1$

- Electrons added to the shell with the highest principal quantum number (n).
- These dictate the chemical reactivity and bonding behavior of the atom.

The Three Rules of Orbital Filling

Every electronic configuration is written by following these three fundamental laws of quantum mechanics.



Aufbau Principle

1. Energy Staircase

Rule: Electrons first occupy the lowest energy orbital available before entering higher energy orbitals.

In Action: The 4s orbital is lower in energy than the 3d orbital, so 4s fills first (e.g., in Potassium).



Pauli Exclusion Principle

2. Maximum Capacity

Rule: No two electrons in an atom can have the same set of four quantum numbers.

In Action: A single orbital box can hold a maximum of two electrons, and they must have opposite spins ($\uparrow\downarrow$).



Hund's Rule of Multiplicity

3. Spread Out Before Pairing

Rule: Pairing of electrons in degenerate orbitals (p, d, f) does not take place until each orbital has one electron.

In Action: In Nitrogen ($2p^3$), the three electrons each take their own p-orbital box with parallel spins before any pairing occurs.



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