

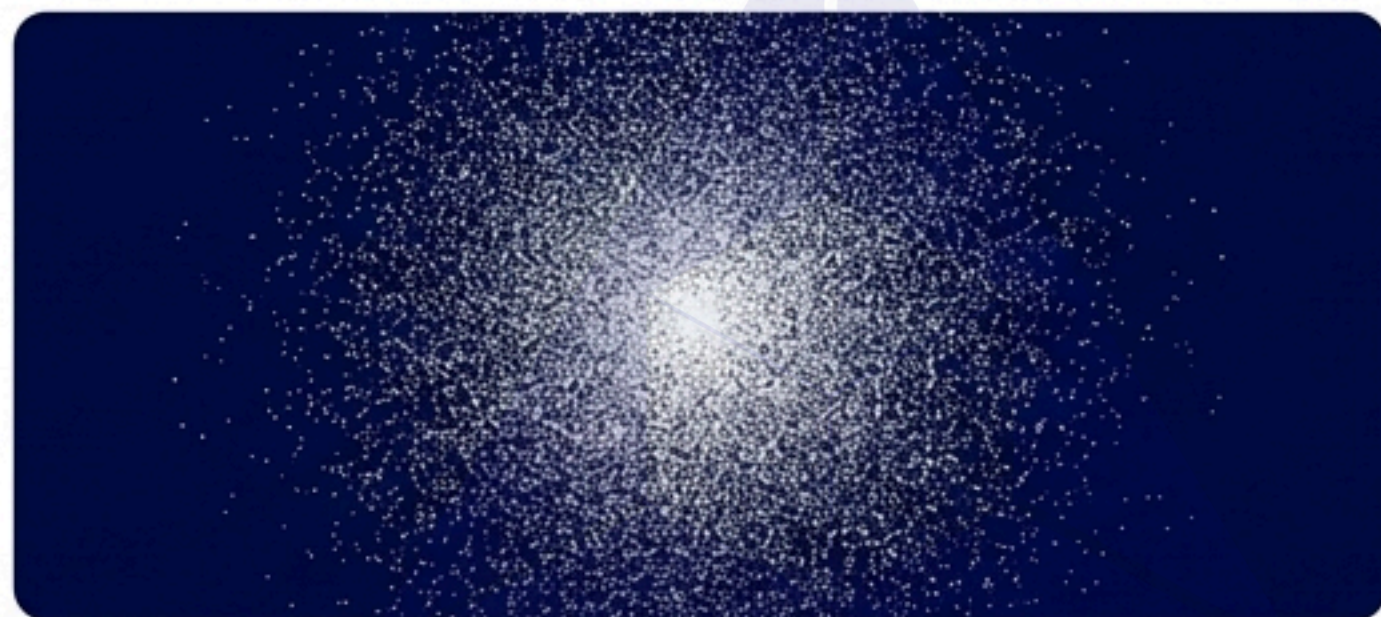
UNIT 2: STRUCTURE OF ATOM

2.6.2 Shapes of Atomic Orbitals

Visualizing the Quantum Architecture
of the Atom

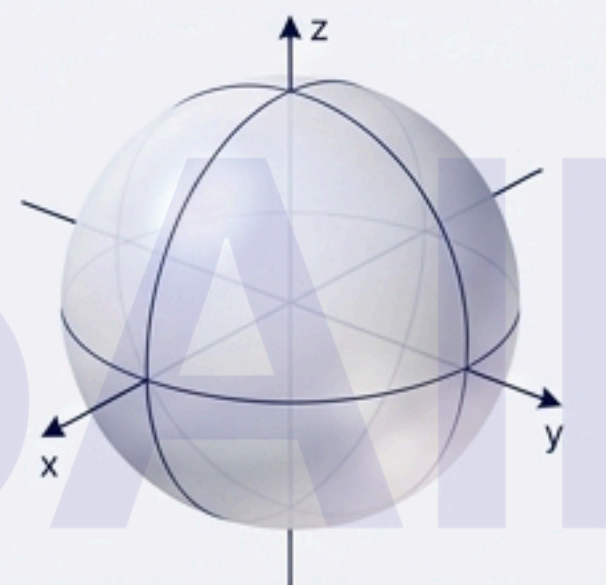
Beyond Orbits: Probability & Boundary Surfaces

Probability Density



The probability of finding an electron at a point is proportional to the square of the orbital wave function. It never truly reaches zero at any finite distance from nucleus.

The 90% Rule

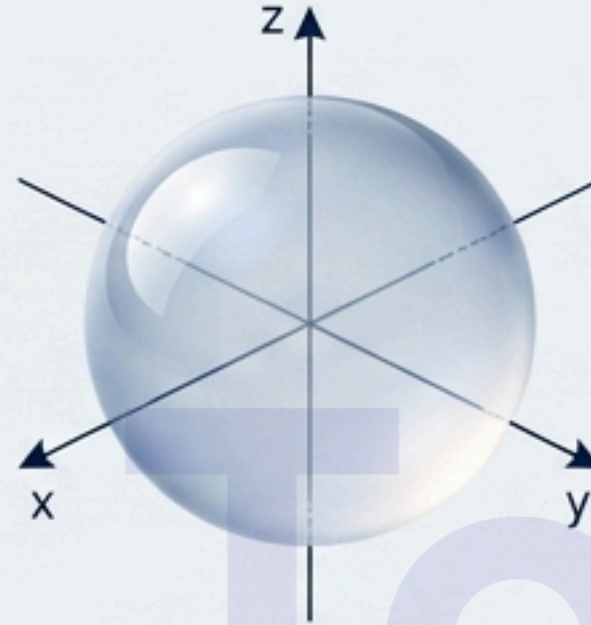


Because we cannot draw a 100% boundary, we draw a solid surface enclosing a region where the probability of finding the electron is approximately 90%.

$$|\psi|^2 > 0 \text{ at any finite distance}$$

Recap: Orbitals are not rigid paths; they are 3D regions of high probability.

The Spherical Symmetry of s Orbitals



Symmetry

Spherically symmetric. The probability of finding the electron depends strictly on the distance from the nucleus, equal in all directions.

Expansion

The size and energy of the orbital increase as the principal quantum number (n) increases, locating the electron further away from the nucleus.



Size and Energy: $4s > 3s > 2s > 1s$

Recap: The fundamental building block shape—a sphere that grows larger with higher energy levels.



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$$E=mc^2$$



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Key Highlights

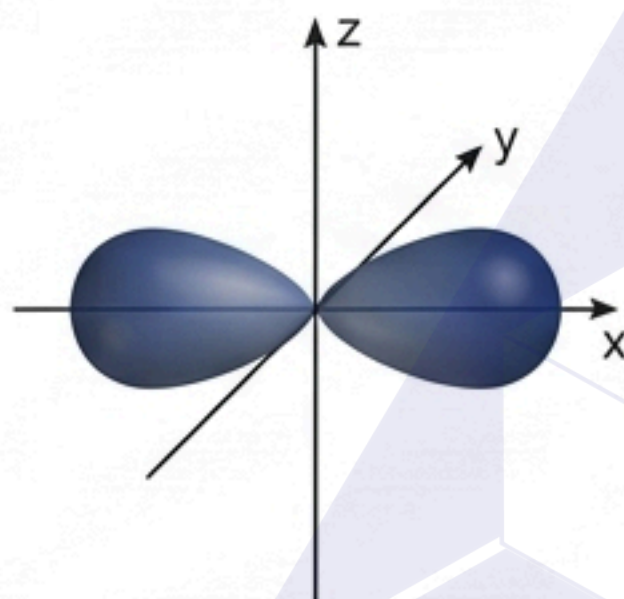
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$L = 1$ SUBSHELL

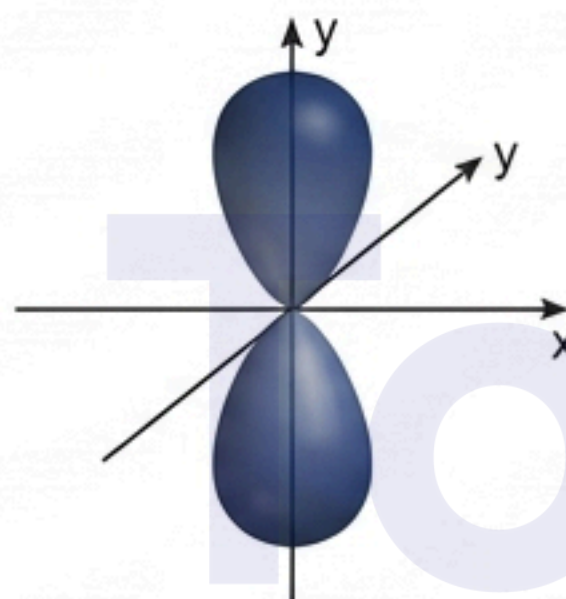
The Directional Lobes of p Orbitals

$2p_x$



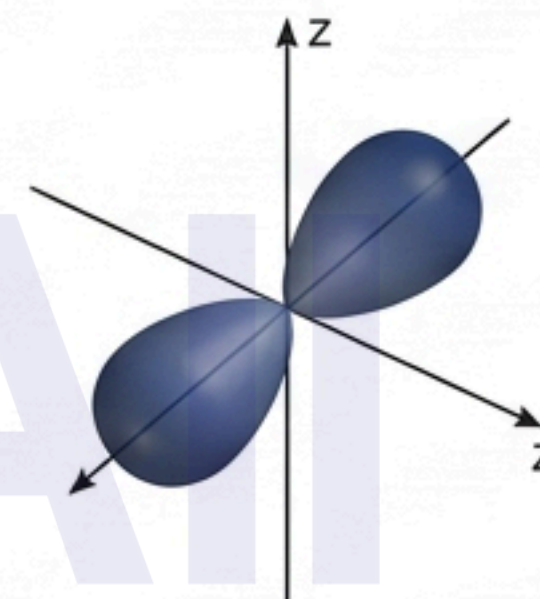
Two lobes lying horizontally along the x-axis.

$2p_y$



Two lobes lying vertically along the y-axis.

$2p_z$



Two lobes lying depth-wise along the z-axis.

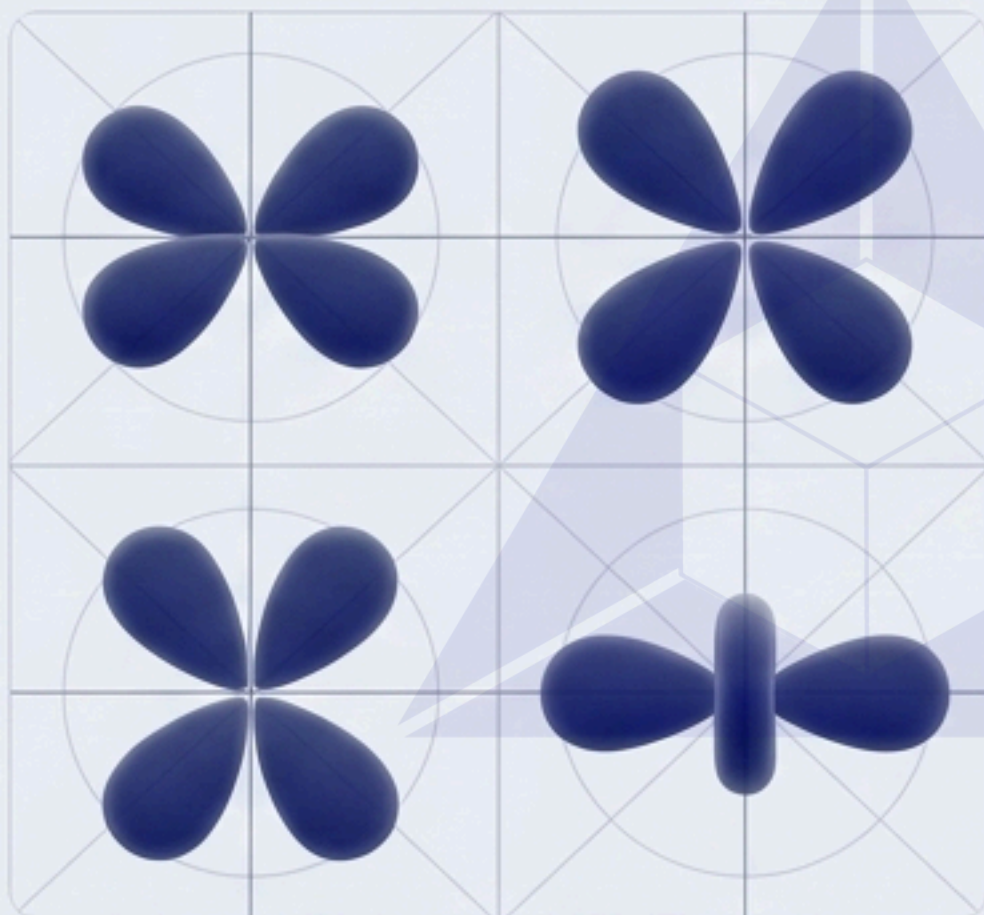
Each p orbital has two lobes separated by a plane with zero probability density.
All three orientations are degenerate (equal in energy, size, and shape).

Dumbbell shapes aligned to specific 3D axes, bringing directionality to atomic bonding.

L = 2 SUBSHELL

The Complex Geometry of d Orbitals

d_{xy} , d_{yz} , d_{xz} , $d_{\{x^2-y^2\}}$



d_{xy} , d_{yz} , d_{xz}
(Lobes project between the principal axes)

$d_{\{x^2-y^2\}}$
(Lobes align directly on the x and y axes)



$d_{\{z^2\}}$

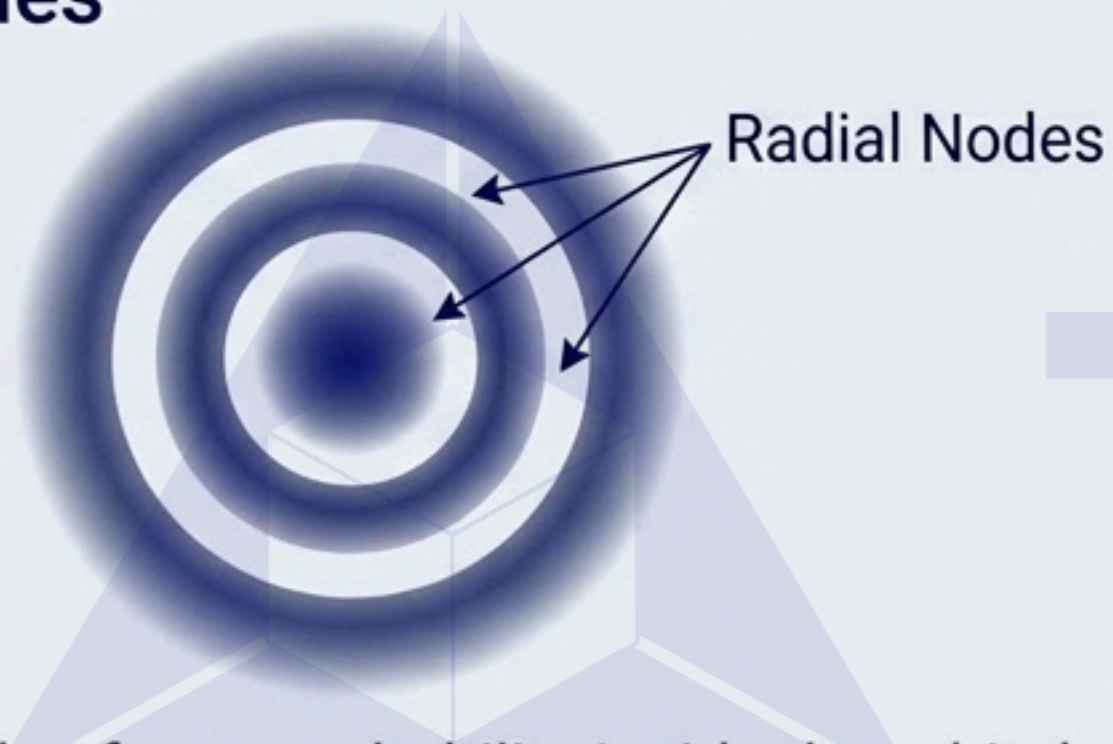
A dumbbell shape along the z-axis encircled by a donut-shaped ring of electron density at the center.

Despite their drastically different spatial geometries, all five 3d orbitals are perfectly equivalent in energy (degenerate).

Five orientations providing complex 3D spacing, crucial for transition metal chemistry.

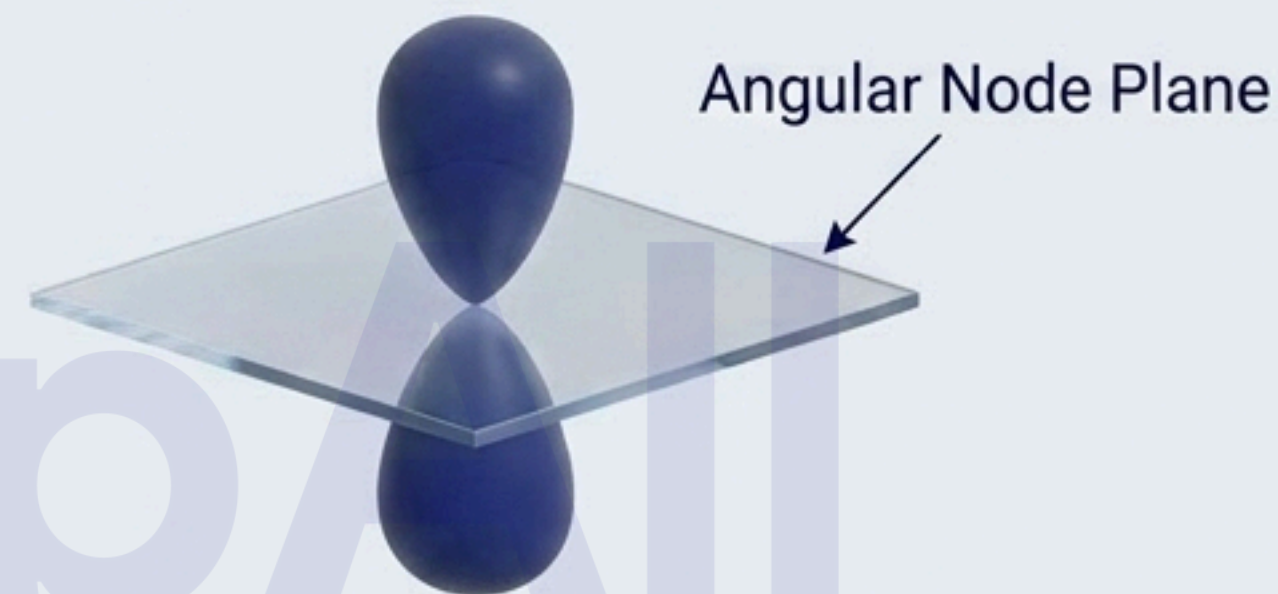
Nodes: Where the Electron Cannot Be

Radial Nodes



Spherical shells of zero probability inside the orbital. The density decreases to zero, then increases again at a further distance.

Angular Nodes



Flat planes (or cones) of zero probability separating orbital lobes.

$$\text{Total Nodes} = n - 1 \quad | \quad \text{Angular Nodes} = \ell \quad | \quad \text{Radial Nodes} = n - \ell - 1$$

Nodes act as quantum boundary lines, increasing in number as the principal energy level (n) increases.

The Atomic Architecture Matrix

Subshell	Quantum No. (l)	Shape	No. of Orbitals	Orientations (m_l)
s	0	Spherical	1	0
p	1	Dumbbell	3	-1, 0, +1
d	2	Clover & Torus	5	-2, -1, 0, +1, +2

Recap

Key Takeaway: Quantum numbers strictly dictate the size, shape, and spatial orientation of the regions where electrons reside.



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