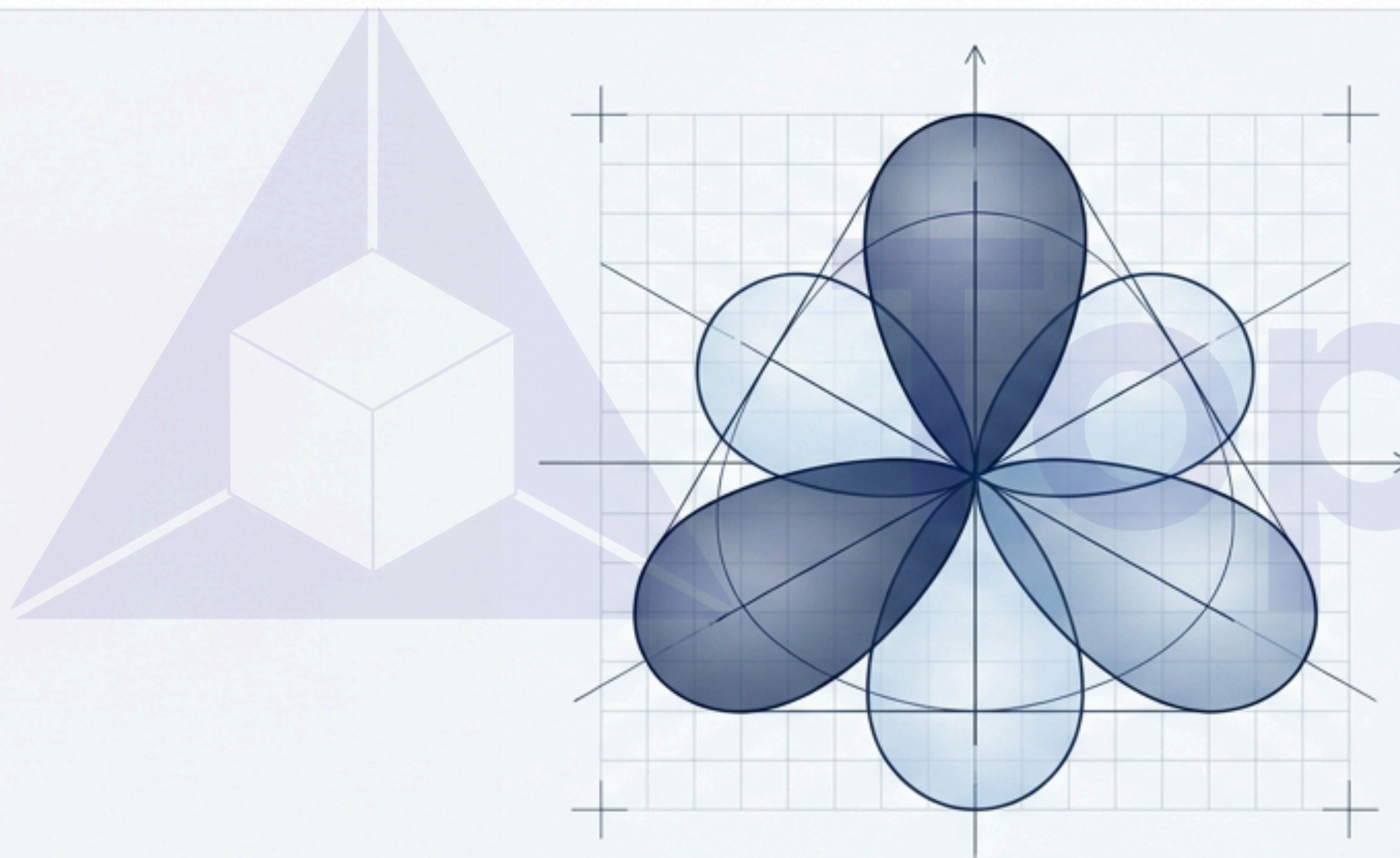


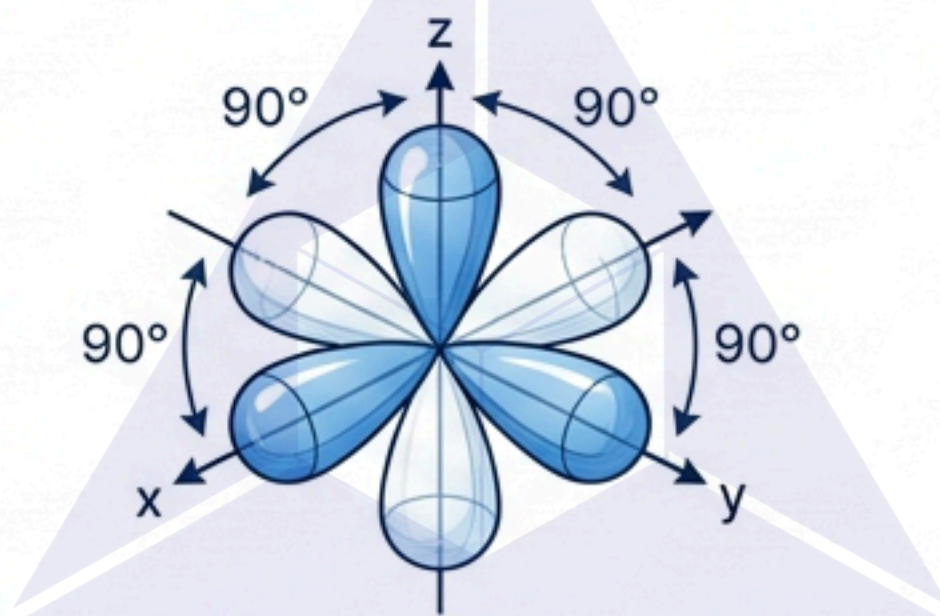
4.6 Hybridisation



THE PROBLEM

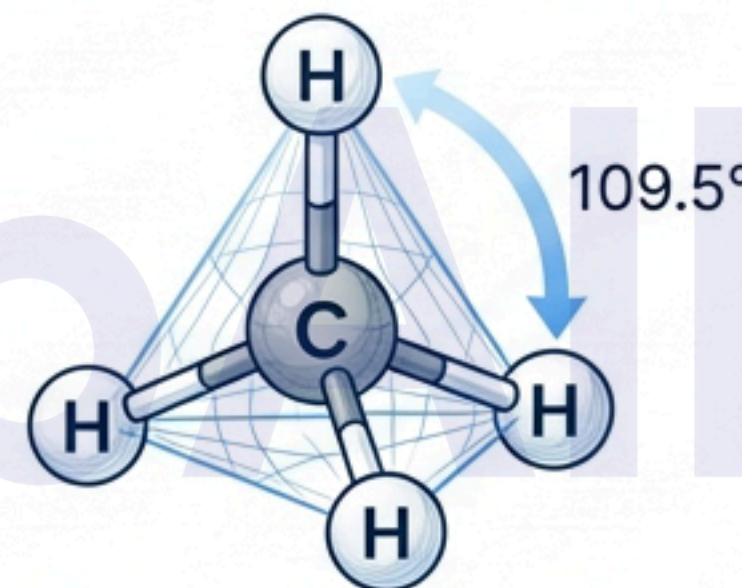
Why Hybridisation Was Introduced

Pure Orbital Overlap Fails



Simple overlap of pure atomic orbitals (like s and p) predicts incorrect, constrained bond angles (e.g., exactly 90°).

Experimental Geometry



Real polyatomic molecules (CH₄, NH₃, H₂O) possess highly specific, symmetrical geometric shapes that pure orbitals cannot explain.

Linus Pauling introduced the concept of hybridisation to explain the characteristic geometrical shapes of polyatomic molecules.

The Definition Of Hybridisation



The process of intermixing of the orbitals of slightly different energies...
...so as to redistribute their energies...

...resulting in the formation of a new set of orbitals of equivalent energies and shape.



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$$E=mc^2$$



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SALIENT FEATURES

Equivalent Hybrid Orbitals



Conservation

The number of hybrid orbitals formed is exactly equal to the number of atomic orbitals that get hybridised.



Equivalence

The hybridised orbitals are always strictly equivalent in both energy and shape.



Effectiveness

Hybrid orbitals are significantly more effective in forming stable bonds than pure atomic orbitals.



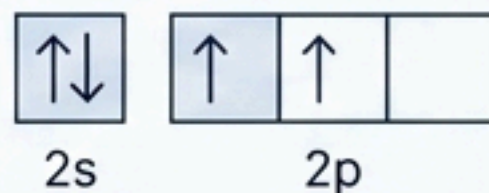
Directionality

They are directed in space to minimize electron pair repulsion, dictating a stable molecular geometry.

Carbon As The Main Example

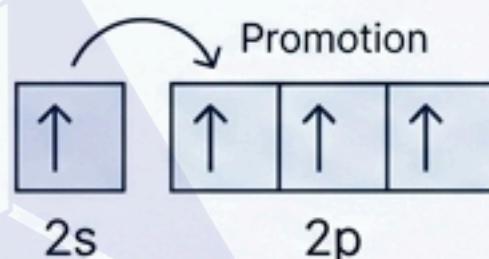
[He] 2s² 2p²

Ground State



[He] 2s¹ 2p_x¹ 2p_y¹ 2p_z¹

Excited State

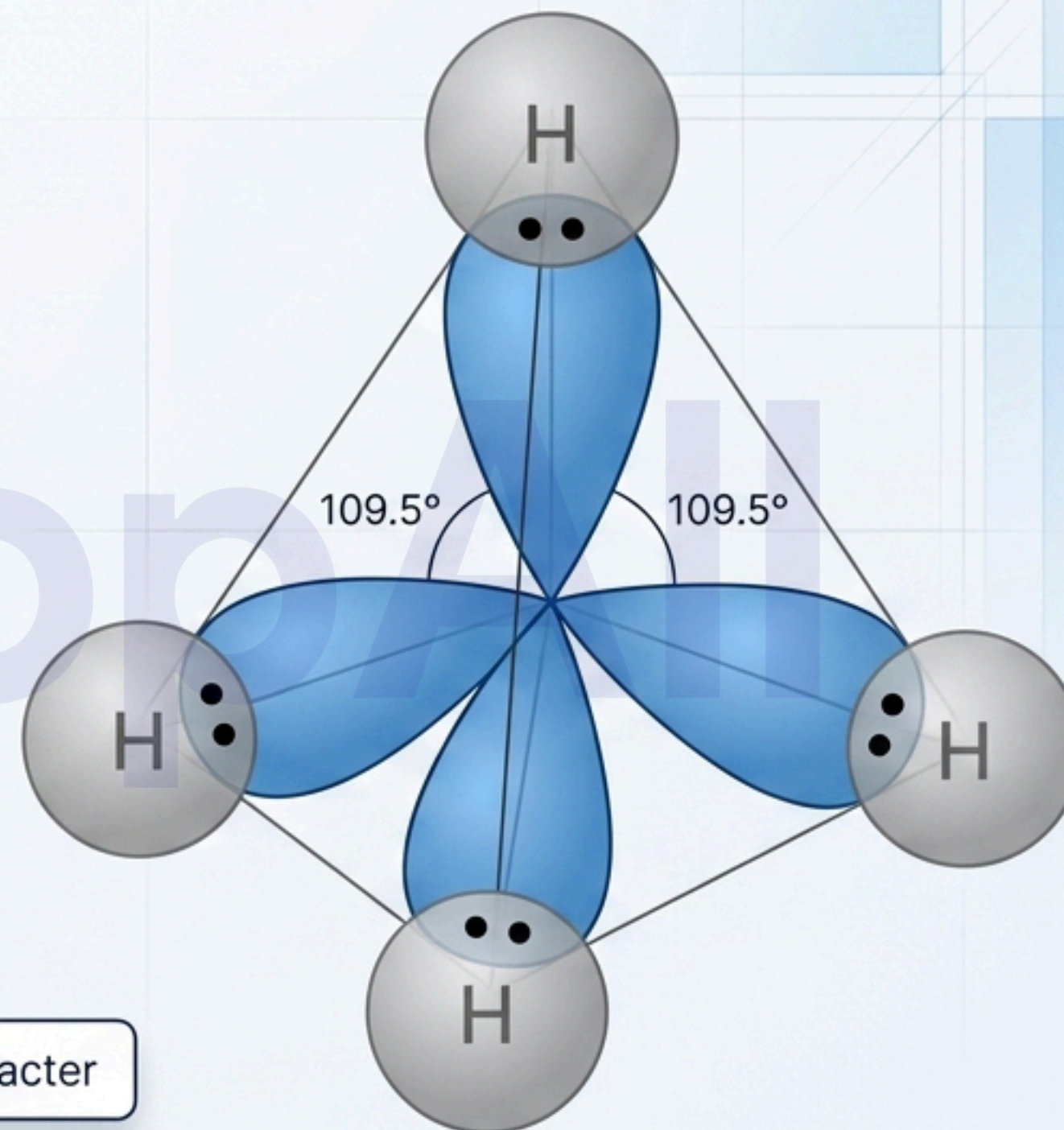


4 × sp³

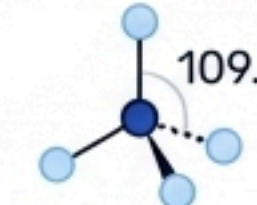
1 s-orbital + 3 p-orbitals → 4 equivalent hybrid orbitals.

25% s-character

75% p-character



Final Recap: Connecting Orbitals With Real Geometry

Orbital	Geometry	Angle	Example
sp	Linear	180°	$BeCl_2$ 
sp^2	Trigonal Planar	120°	BCl_3 
sp^3	Tetrahedral	109.5°	CH_4 

Hybridisation is the ultimate bridge: it connects theoretical quantum orbitals with the real, physical geometry of molecules.



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