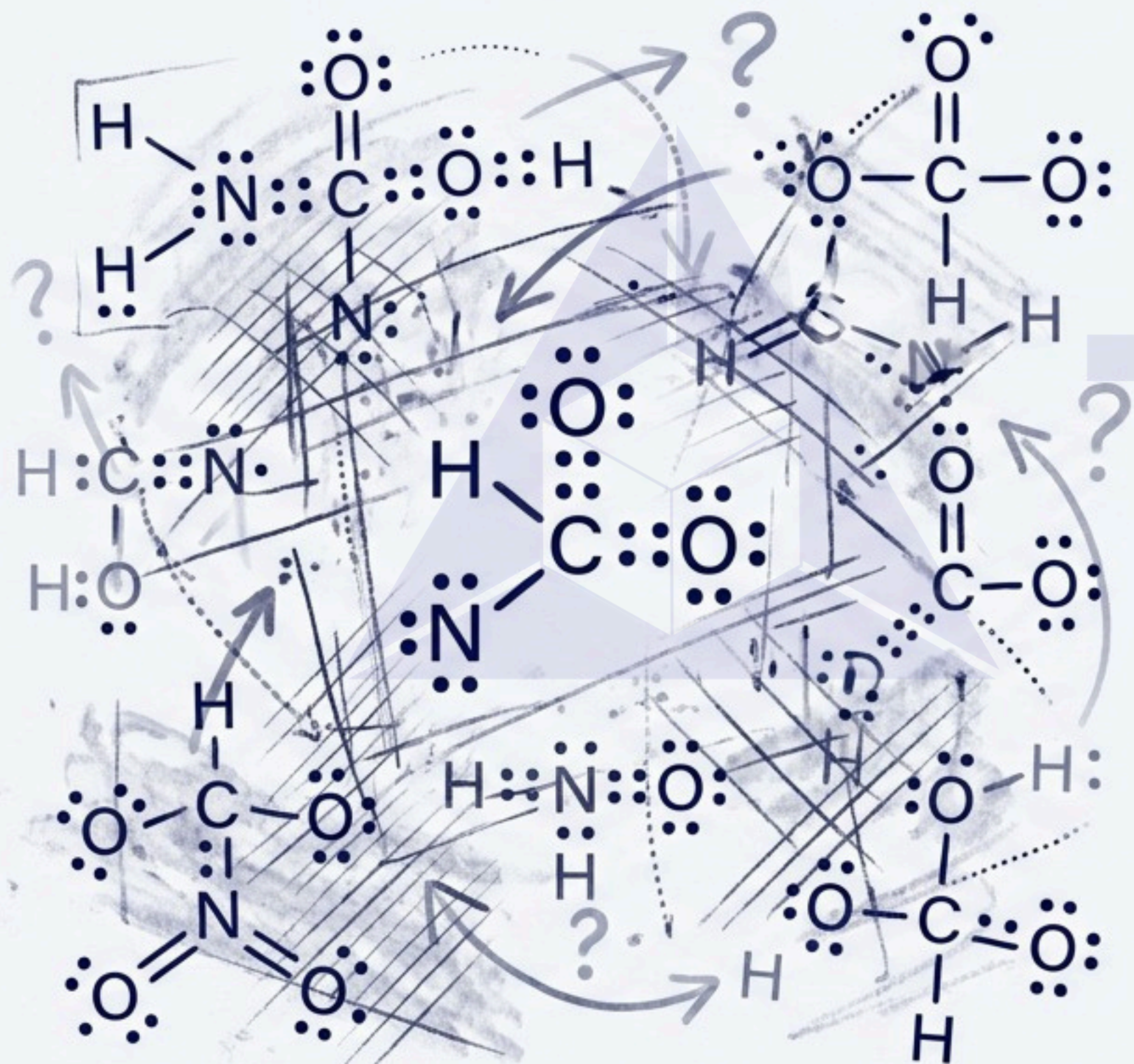


The Purpose of Formal Charge: Auditing Molecular Stability



The Ledger

Core Concept: Formal charge is a theoretical accounting system. It helps chemists keep track of valence electrons within polyatomic molecules and ions.

Primary Utility: Selection of the lowest energy structure. When multiple Lewis structures are possible, the most stable geometry is generally the one with the smallest formal charges.

The Rule: Net charge is possessed by the ion as a whole, but formal charge assigns a localized bookkeeping value to each individual atom.

The Formal Charge Equation

$$FC = V - N - \frac{1}{2} B$$

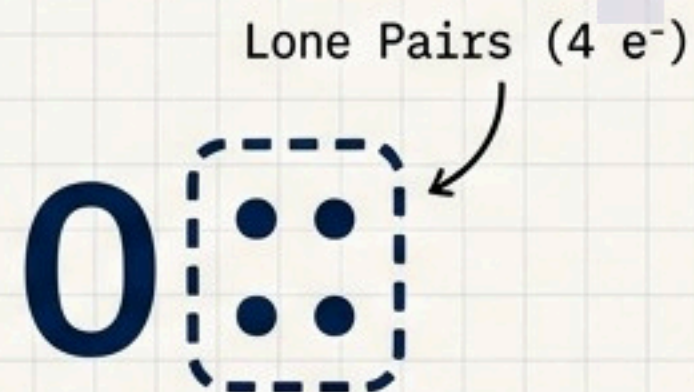
V = Valence Electrons

Total valence electrons in the isolated free atom.



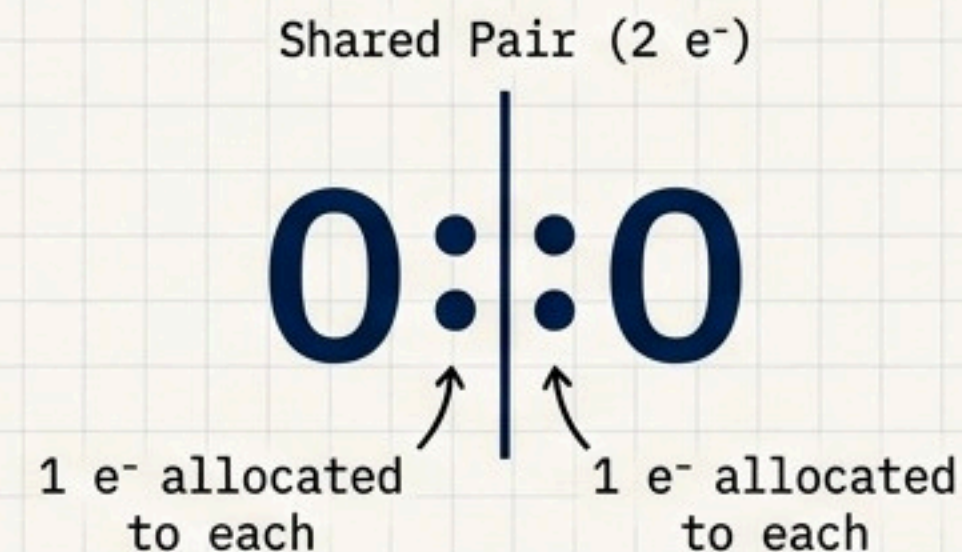
N = Non-Bonding Electrons

Total electrons in lone pairs.

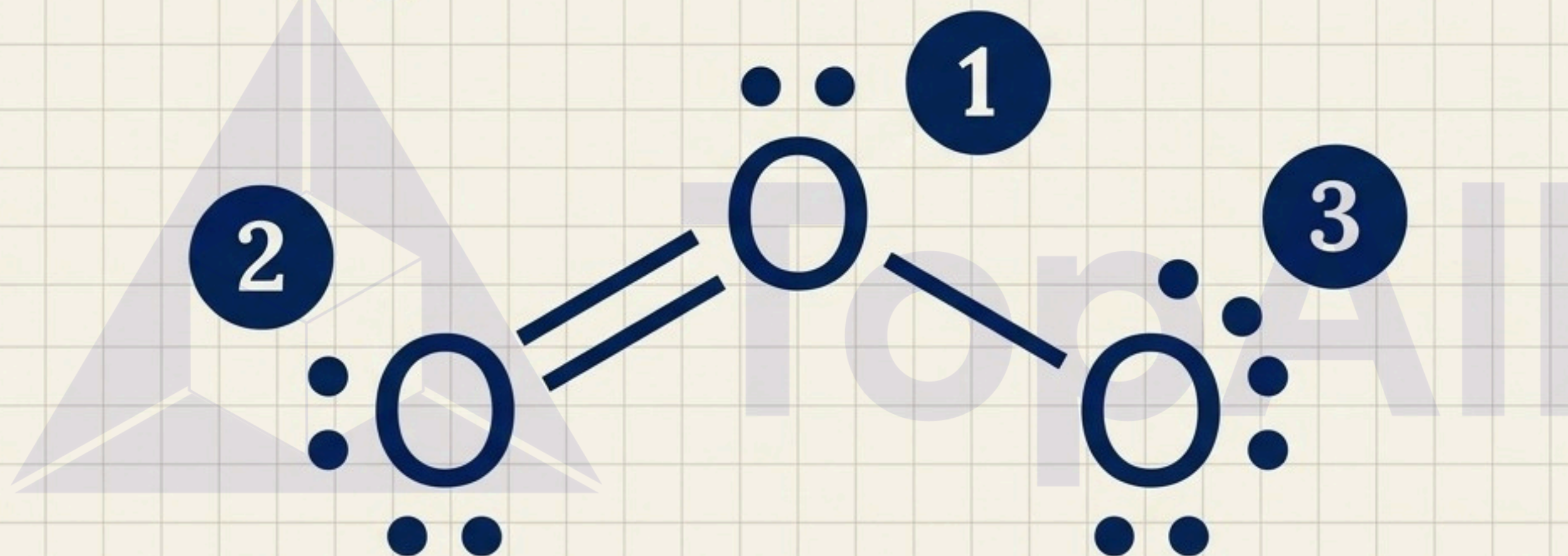


$\frac{1}{2} B$ = Bonding Electrons

Half of the shared electrons.



The Test Subject: Auditing the Ozone Molecule (O_3)



All three atoms are Oxygen, yet their unique bonding environments mean they 'own' different amounts of electrons. We must audit each one individually to determine the molecule's internal charge distribution.



**You've Learned the Theory.
Let's Test Them Now.**

$$E=mc^2$$



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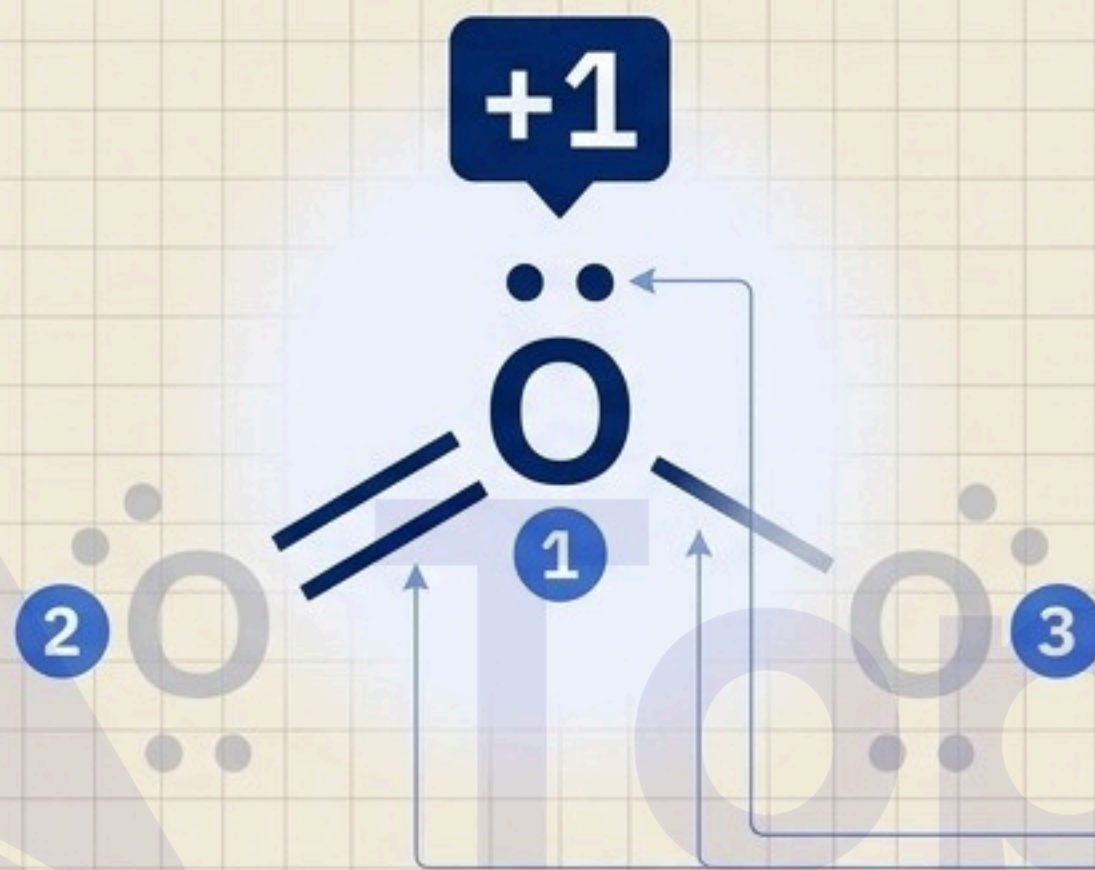
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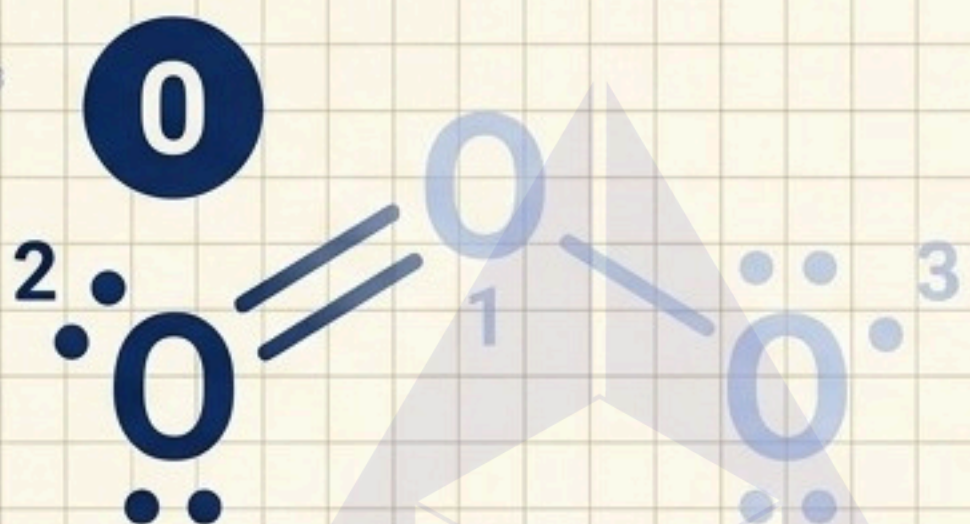


Step 1: Calculating the Central Oxygen



Target:	Central O atom marked 1
Valence:	6 (Standard for Oxygen)
Non-Bonding:	2 (One lone pair)
Bonding (Shared):	6 (Three bonds total × 2 electrons)
The Ledger: $6 - 2 - \frac{1}{2}(6) = +1$	

Step 2: Comparing the Terminal Atoms



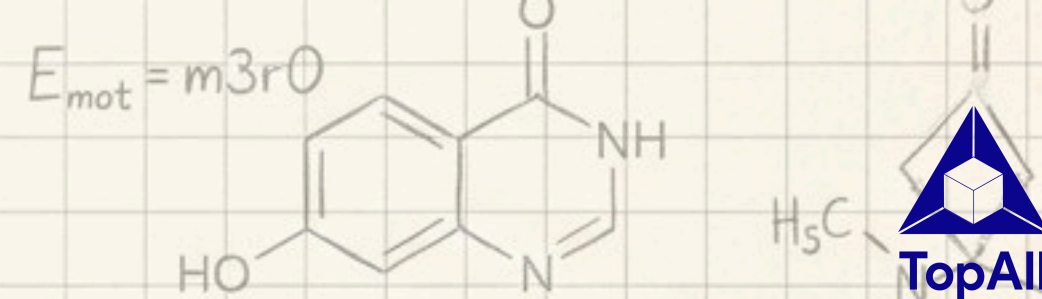
Target:	End 0	atom 2
Valence:		6
Non-Bonding:		-4
½ Shared:		-2
Formal Charge:		0



Target:	End 0	atom 3
Valence:		6
Non-Bonding:		-6
½ Shared:		-1
Formal Charge:		-1

Insight: Despite both being Oxygen, the single-bonded atom holds a negative formal charge, while the double-bonded atom is neutral. The fully audited molecule has charges of [0, +1, -1], summing to the neutral net charge of O_3 .

Revision Frame: The Limits of the Ledger



The Formula

$$\text{FC} = \text{Valence} - \text{Lone Pair } e^- - \frac{1}{2} \text{ Shared } e^-$$

Always calculate **atom by atom**. The sum of all **formal charges** in a species must equal its **net charge**.



The Primary Utility

Select the **Lowest Energy Structure**.

- The most stable structural **geometry** for a given species is generally the one with the **smallest formal charges** on its atoms.



The Crucial Caveat

Bookkeeping, Not Reality.

Formal charges do **not** indicate **real physical charge separation** within a molecule. It is a theoretical factor based on a pure covalent view where electron pairs are assumed to be shared perfectly equally.



You've Learned the Theory.

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