

Hybridization and Sigma and Pi Bonds

How do orbitals mix to explain bond angles and double bonds?

●●○ Intermediate

Bonding & Molecular Structure

Last reviewed 5 October 2026

Before you start: it helps to understand [Molecular Shapes \(VSEPR\)](#) and [Quantum Numbers and Orbitals](#).

What is it?

Carbon's outer electrons are in one 2s and three 2p orbitals, which point in different directions at 90° to each other. Yet the four bonds in methane are **identical** and point to the corners of a tetrahedron at **109.5°**. **Hybridization** explains this: an atom's valence orbitals **mix** to form an equal number of new, identical **hybrid orbitals** that point where the bonds are.

Hybridization	Orbitals mixed	Hybrid orbitals	Electron domains	Geometry	Angle
sp^3	one s + three p	4	4	Tetrahedral	109.5°
sp^2	one s + two p	3 (+ 1 unhybridized p)	3	Trigonal planar	120°
sp	one s + one p	2 (+ 2 unhybridized p)	2	Linear	180°

Bonds come in two kinds:

- A **sigma (σ) bond** forms when orbitals overlap **head-on**, along the line between the two nuclei. **Every single bond is a σ bond.**
- A **pi (π) bond** forms when two parallel **p orbitals** overlap **side-on**, above and below the line between the nuclei.

A **double bond** is one σ + one π ; a **triple bond** is one σ + two π .

🔑 KEY IDEA

Count the **electron domains** (bonded atoms + lone pairs) around an atom, exactly as in VSEPR. **4 domains** → sp^3 , **3** → sp^2 , **2** → sp . The p orbitals that are left over are the ones that form π bonds.

Why does it matter?

- **It connects shape to orbitals.** VSEPR predicts the shape; hybridization explains which orbitals make it.
- **Double bonds don't rotate.** Twisting a π bond would break the side-on overlap, so groups on a C=C are locked in place. This is why *cis* and *trans* isomers exist.
- **Reactivity.** π electrons sit above and below the bond, exposed to attack, which is why alkenes react by addition (see [Electrophilic Addition to Alkenes](#)).

How does it work?

1. Finding the hybridization

1. Draw the Lewis structure.
2. Count electron domains around the atom: each **bonded atom** counts as one (whether the bond is single, double or triple), and each **lone pair** counts as one.
3. Match: $4 \rightarrow sp^3$, $3 \rightarrow sp^2$, $2 \rightarrow sp$.

2. Three carbon compounds

- **Ethane**, CH_3CH_3 : each C has 4 domains, so **sp^3** , 109.5° ; 7 σ bonds, free rotation about C–C.
- **Ethene**, $CH_2=CH_2$: each C has 3 domains, so **sp^2** , 120° , planar. The C=C is one σ (sp^2-sp^2) plus one π (from the leftover p orbitals).
- **Ethyne**, $HC\equiv CH$: each C has 2 domains, so **sp** , 180° , linear. The C=C is one σ plus two π bonds at right angles.

3. Atoms with lone pairs

Lone pairs occupy hybrid orbitals too. In NH_3 , N has 3 bonds + 1 lone pair = 4 domains: **sp^3** (bond angle squeezed to about 107°). In H_2O , O has 2 bonds + 2 lone pairs: **sp^3** (about 104.5°).

4. Counting σ and π bonds

Single = 1σ . Double = $1\sigma + 1\pi$. Triple = $1\sigma + 2\pi$. Count every bond in the full structure, including C–H bonds.

THINK OF IT LIKE THIS

Mixing orbitals is like mixing paint. One tin of white (s) and three tins of blue (p) make four identical tins of light blue (sp^3). The colour is “25 % s, 75 % p”, and all four tins are the same. Mix one white with two blue and you get three tins of sp^2 , with one blue tin left unmixed: that leftover p orbital is ready to make a π bond.

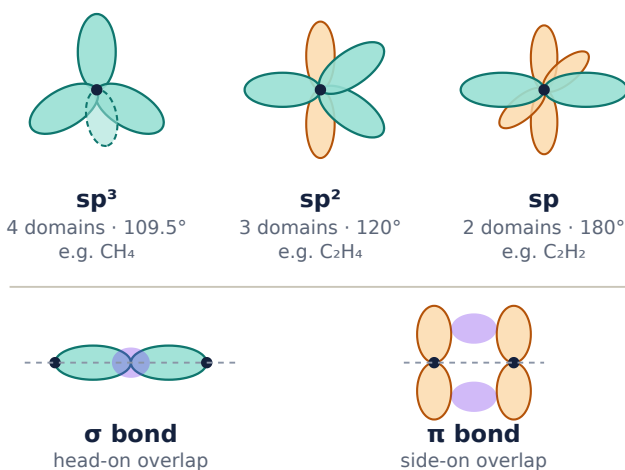
Q MORE PRECISELY

Hybridization is a model within **valence bond theory**; it is a way of describing the bonding, not a step that atoms “do”. **Molecular orbital theory** is a more complete description that explains, for example, why O_2 is magnetic. A π bond is weaker than a σ bond, because side-on overlap is smaller than head-on overlap: the $\text{C}=\text{C}$ bond (614 kJ/mol) is less than twice as strong as $\text{C}-\text{C}$ (348 kJ/mol). Atoms with more than four electron domains, such as S in SF_6 , are better described without d-orbital hybrids, so they are left out here.

Visualise it

Hybrid orbitals and bond types

Teal: hybrid orbitals · amber: unhybridized p orbitals



sp^3 , sp^2 and sp hybrids, and the difference between σ (head-on) and π (side-on) overlap.

Worked example

✎ WORKED EXAMPLE: Hybridization of carbon

Question: Give the hybridization of carbon in (a) CH_4 (b) $\text{H}_2\text{C}=\text{O}$ (methanal) (c) CO_2 .

- Step 1** (a) 4 bonded atoms, no lone pairs: 4 domains, **sp^3** .
- Step 2** (b) 3 bonded atoms (2 H, 1 O): 3 domains, **sp^2** , trigonal planar.
- Step 3** (c) 2 bonded atoms ($\text{O}=\text{C}=\text{O}$): 2 domains, **sp** , linear.

 WORKED EXAMPLE: Counting σ and π bonds

Question: How many σ and π bonds are in ethanoic acid, CH_3COOH ?

Step 1 Bonds: $3 \times \text{C-H}$, $1 \times \text{C-C}$, $1 \times \text{C=O}$, $1 \times \text{C-O}$, $1 \times \text{O-H}$.

Step 2 σ bonds: every bond has one, so $3 + 1 + 1 + 1 + 1 = 7 \sigma$.

Step 3 π bonds: only the C=O has one: **1 π** .

 WORKED EXAMPLE: A triple bond

Question: Describe the bonding in hydrogen cyanide, $\text{H-C}\equiv\text{N}$.

Step 1 Carbon has 2 domains (H and N): **sp**, linear, 180° .

Step 2 C-H is a σ bond (sp-s). $\text{C}\equiv\text{N}$ is $1 \sigma + 2 \pi$.

Step 3 Total: **2 σ** and **2 π** . Nitrogen also has 2 domains (C and its lone pair), so it is **sp** too.

 WORKED EXAMPLE: An atom with lone pairs

Question: What is the hybridization of the oxygen atom in water, and why is the bond angle less than 109.5° ?

O has 2 bonds + 2 lone pairs = 4 domains: **sp³**. The lone pairs repel more strongly than bonding pairs, squeezing the H-O-H angle to about **104.5°** .

Common mistake

 COMMON MISTAKE: Counting a double bond as two domains

A double or triple bond counts as **one** electron domain, because it points in one direction. CO_2 has two domains on carbon (sp), not four.

 COMMON MISTAKE: Forgetting that every bond has a σ

A double bond is $\sigma + \pi$, not two π bonds. The first bond between two atoms is always σ .

 COMMON MISTAKE: Forgetting lone pairs

Lone pairs count as domains: N in NH_3 is sp^3 , not sp^2 , even though it has only three bonds.

NOTATION NOTE

- sp^3 is read “s-p-three”; the superscript counts the p orbitals mixed, not electrons.
- σ is the Greek letter sigma; π is pi.

Remember this

REMEMBER THIS

- Electron domains = bonded atoms + lone pairs. 4 $\rightarrow$ sp^3 (109.5°), 3 $\rightarrow$ sp^2 (120°), 2 $\rightarrow$ sp (180°).
- σ : head-on overlap, along the bond axis. π : side-on overlap of parallel p orbitals.
- Single = σ ; double = $\sigma + \pi$; triple = $\sigma + 2\pi$.
- Leftover (unhybridized) p orbitals make the π bonds.
- π bonds prevent rotation about double bonds.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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