HYBRIDISATION

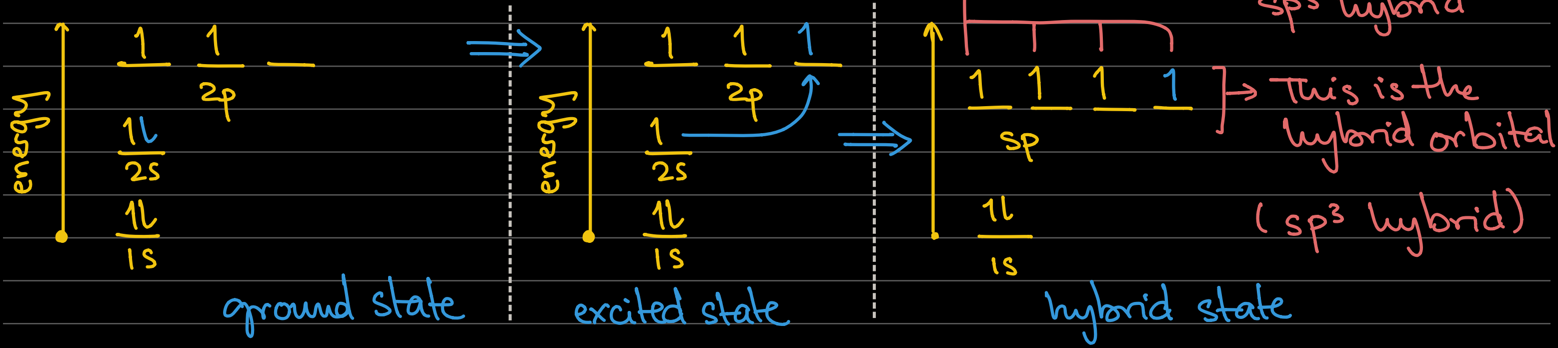
- Hybrid orbitals are formed by mixing orbitals

C−C $sp^3 \rightarrow$ 25% s and 75% p
C=C $sp^2 \rightarrow$ 33.3% s and 66.6% p
C≡C $sp \rightarrow$ 50% s and 50% p
} (in character)

$\underbrace{sp}$ → hybrids

SP^3 hybrid orbital

E.C of $_6C \rightarrow 1s^2 2s^2 2p^2$

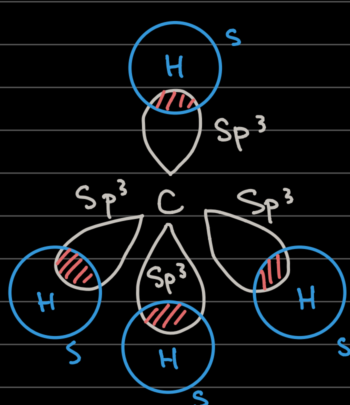


- In the sp^3 hybrid orbital − all four orbitals are at an equal energy level

- They repel each other, so that the 4 regions of electron density point towards the corners of a tetrahedron, forming a tetrahedral angle (109.5°)

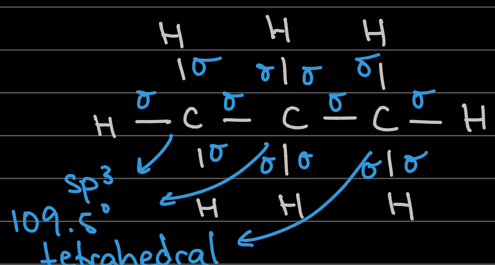
- The sp^3 hybrid orbitals form σ (sigma) bonds

ie. $CH_4 \rightarrow$ C in methane is sp^3 hybridized

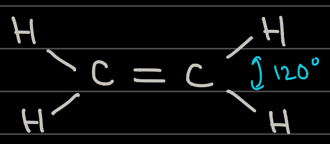


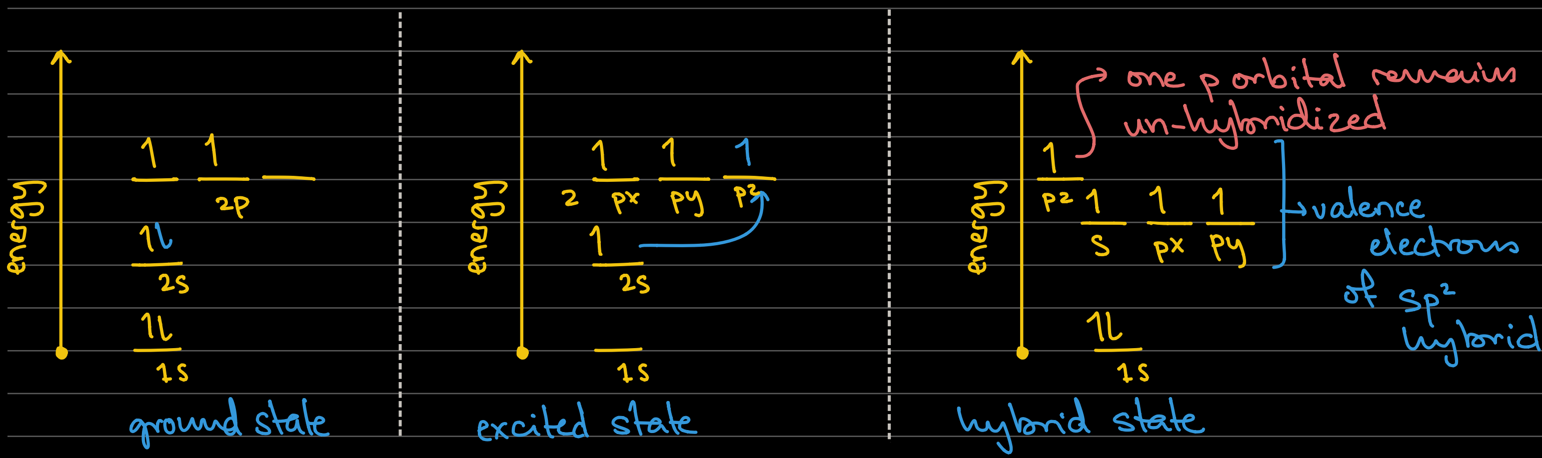
▨ = sigma (σ) bond

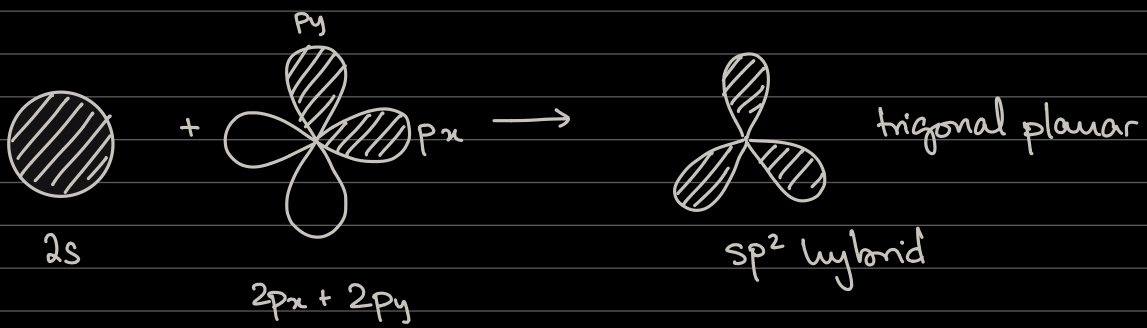
→ Methane has 4σ bonds



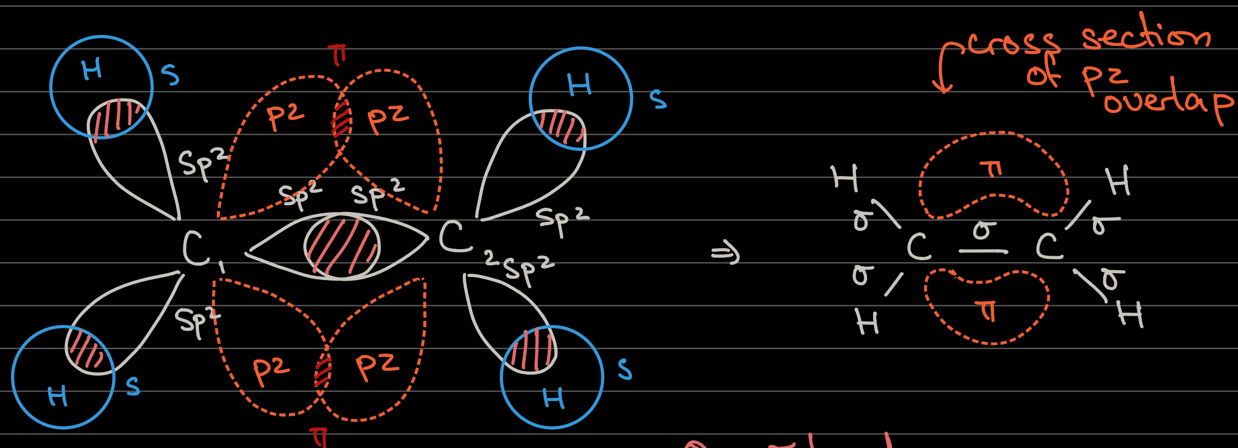
SP^2 hybrid orbital → present in double bonded carbons
ie. Ethene (C_2H_4)







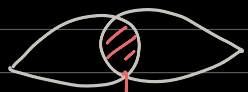
- There are three orbitals in the sp^2 hybrid orbital of equal energy
- The remaining p orbital (ie. pz) orbital is unhybridized
- The sp^2 hybrid orbitals repel each other equally and point towards the corners of an equilateral triangle - making an angle of 120°.



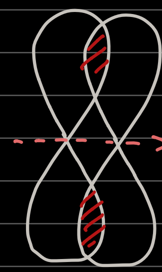
⊘ = σ bond
⊘ = π bond

σ bond : is made from the end-on / head-on overlap of p-orbitals or s-p orbitals

π bond : is made from the sideways overlap of p-orbitals.



σ : head on/end on overlap



σ bond in the middle

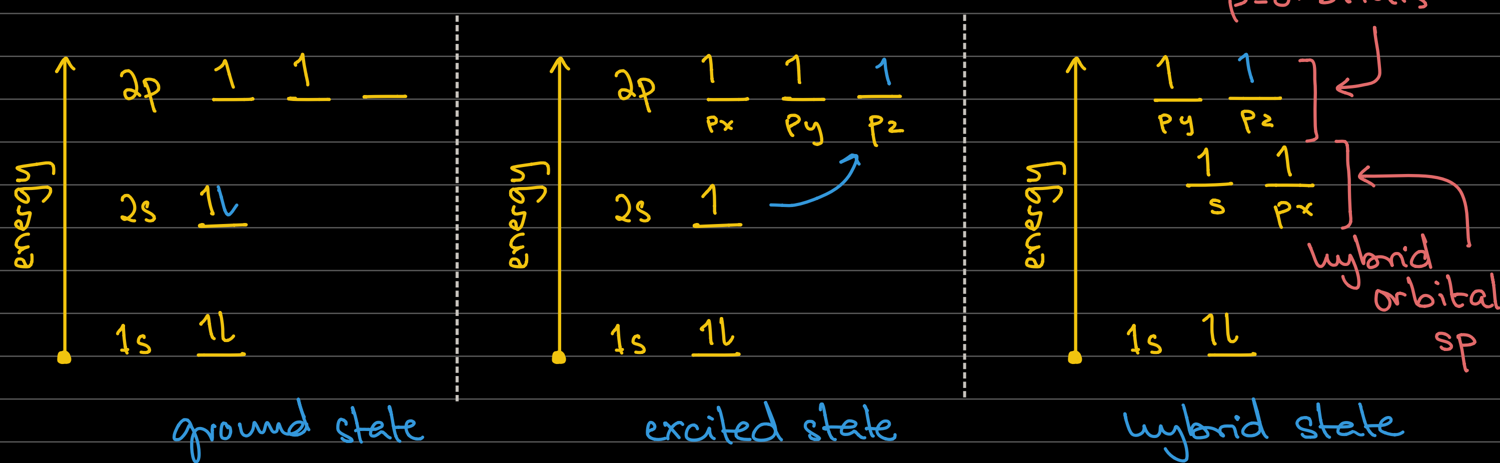
π : sideways overlap

- The remaining p-orbital (unhybridised) orbitals overlap sideways with another unhybridised p-orbital to form the pi-bond.
  - The carbon-carbon double bond is made of 1 σ plus 1 π bond.

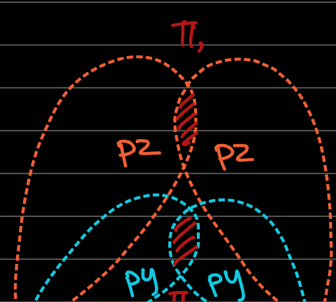
SP hybrid orbital

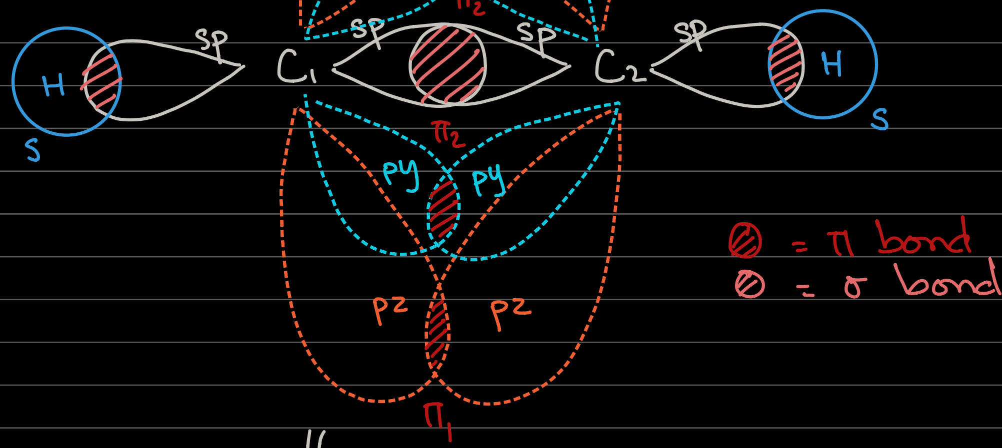
ie. ethyne H–C≡C–H

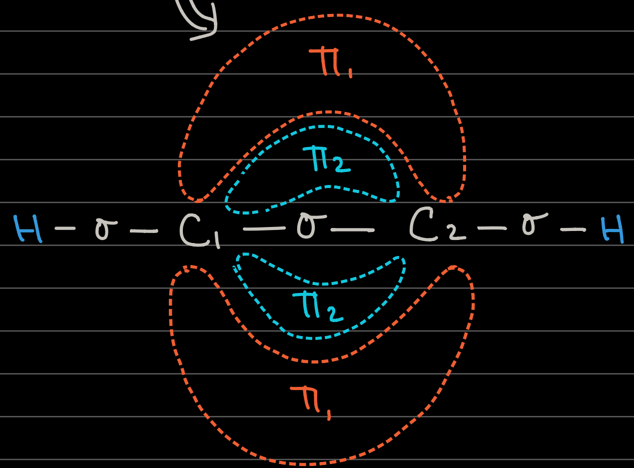
sp hybridized



- Hybrid sp orbital electrons → form the σ bonds
- Unhybridised p orbitals (electrons) → form the π bonds







⇒ H — C ≡ C — H

How do you determine the hybridization on an atom?
Count the regions of electron density around that atom, including lone pairs, single bonds, double bonds, triple bonds.

↳ But all of these regions are treated equally, ie. lone pair is considered just as much of a single region as a triple bond.

4 regions → sp^3 ie. CH_4 H–C–H (with H above and below C)

3 regions → sp^2 ie. ethene $H_2C=CH_2$

2 regions → sp ie. ethyne H − C ≡ C − H

$CH_3 - CH = CH - CH_2 - C \equiv C - H$

sp^3 sp^2 sp^2 sp^3 sp sp

Bond Energies

	kJ/mol	
C = C	610	σ + π
C − C	− 350	− σ
	260	→ π → pi bond energy

- Alot less energy is required to break a π bond than a σ bond

- π bonds are very reactive as they are regions of high electron and also because less energy is required to break a π bond than a σ bond.

—— END OF : CHEMICAL BONDING ——