What Group is carbon in?  
How many valence electrons does it have?  
How many does it want to gain through covalent bonding?  
How many half-filled orbitals does it have to make these covalent bonds?

The fact that it forms 4 bonds means that 1 electron from the 2s orbital must be promoted to the empty p orbital:

This process is called excitation. The energy required to do so is compensated by the energy released when the covalent bonds are formed.

We find that a process of hybridisation occurs to form 4 hybrid orbitals of a new energy, shape and orientation.

The exact nature of these orbitals depend on the type of orbitals that have been hybridised to form them.

If a carbon is only bonded to 3 other atoms then it forms three sp² hybrid orbitals:

What geometry will be formed by 3 identical hybrid orbitals?

If we draw the example of ethene (C₂H₄) we see that after forming 3 covalent (sigma) bonds using the 3 hybrid orbitals, we coincidentally have 2 parallel unhybridised p-orbitals who also need to form a covalent bond:

What kind of covalent bond is formed by these p-orbitals?

So we find that a double bond is always formed from...

If we draw the example of ethyne (C₂H₂) we see that after forming 2 covalent (sigma) bonds using the 2 hybrid orbitals, we coincidentally have 4 parallel unhybridised p-orbitals who also need to form a covalent bond:

What kind of covalent bond is formed by these p-orbitals?

So we find that a triple bond is always formed from...

BONUS QUESTION - What kind of hybridisation might phosphorous use in PCl₅?

Lone pairs can also be involved in hybridisation.

Both ammonia and water have lone pairs that are found in hybrid orbitals. As their geometries are both based around a tetrahedron, we know they must contain four equal sp³ hybrid orbitals:

We can use this knowledge to relate hybridisation and geometry information.

**This shows the knowledge on how hybridisation and geometry interact.**
Worked example

Urea $\text{H}_2\text{N}^-\text{C}==\text{NH}_2$ is present in solution in animal urine. What is the hybridization of C and N in the molecule, and what are the approximate bond angles?

Solution

Start with the Lewis (electron dot) structure.

Because there are three electron domains around the C atom, they are arranged in a triangular planar shape with angles of 120°. The C atom must be $sp^2$ hybridized to give this shape. There are four electron domains around each N atom, so they are arranged tetrahedrally. The N atoms must be $sp^3$ hybridized to give this shape.

Exercises

43 What is the difference in spatial distribution between electrons in a pi bond and electrons in a sigma bond?

44 Give examples of sigma bonds in molecules that form from overlap of the following orbitals. The first one is completed for you as an example.

(a) $s-s$ answer $H-H$ in $H_2$
(b) $s-p$ (c) $p-p$ end-on (d) $sp^3-s$ (e) $sp^2-s$
(f) $sp-s$ (g) $sp^3-p$

45 What hybridization would you expect for the coloured atom in each of the following?

(a) $\text{H}_2\text{C}==\text{O}$ (b) $\text{BH}_4^-$ (c) $\text{SO}_3$
(d) $\text{BeCl}_2$ (e) $\text{CH}_3\text{COOH}$

46 Cyclohexane $\text{C}_6\text{H}_12$ has a puckered, non-planar shape whereas benzene $\text{C}_6\text{H}_6$ is planar.

Explain this difference by making reference to the C-C-C bond angles and the type of hybridization of carbon in each molecule.
14.2 - Hybridisation

Practice questions

1. Which bonds are arranged in order of increasing polarity?
   A. H–H < H–Cl < H–Br < H–I
   B. H–H < H–Br < H–Cl < H–I
   C. H–I < H–Br < H–Cl < H–H
   D. H–I < H–Cl < H–Br < H–H

2. Which row correctly describes the bonding type and melting point of carbon and carbon dioxide?

<table>
<thead>
<tr>
<th>Carbons</th>
<th>Carbon dioxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>covalent bonding</td>
</tr>
<tr>
<td>B</td>
<td>ionic bonding</td>
</tr>
<tr>
<td>C</td>
<td>ionic bonding</td>
</tr>
<tr>
<td>D</td>
<td>covalent bonding</td>
</tr>
</tbody>
</table>

6. Which compound forms hydrogen bonds in the liquid state?
   A. C₂H₅OH
   B. CHCl₃
   C. CH₃CH₂OH
   D. (CH₃)₂CO

7. Which molecule has a non-bonding (lone) pair of electrons around the central atom?
   A. BF₃
   B. SO₂
   C. PCl₅
   D. SF₆

8. Which species does not have delocalized electrons?
   A. NO₂⁻
   B. NO₃⁻
   C. O₂
   D. C₂H₄

9. Which species contain a dative covalent bond?
   I. HClO
   II. SO₂
   A. I and II only
   B. I only
   C. II only
   D. I, II, and III

10. Which molecule has an octahedral shape?
    A. SF₆
    B. PCl₅
    C. XeF₆
    D. BF₃

11. Which statements about σ and π bonds are correct?
    I. σ bonds result from the axial overlap of orbitals.
    II. σ bonds form only from s orbitals.
    III. π bonds result from the sideways overlap of parallel p orbitals.
    A. I and II only
    B. I and II only
    C. I and III only
    D. II and III

12. In which substance does a carbon atom have sp³ hybridization?
    A. 2-methylbutane
    B. propene
    C. CH₃CH₂OH
    D. diamond

13. The Lewis structure of XeF₆ contains two bonding pairs of electrons and three non-bonding pairs of electrons (lone pairs) around the central xenon atom. What is the shape of the XeF₆ molecule?
    A. bent
    B. trigonal bipyramidal
    C. square planar
    D. linear

18. Carbon and silicon belong to the same group of the periodic table.
   (a) Describe and compare three features of the structure and bonding in the three allotropes of carbon: diamond, graphite, and C₆₀ fullerene.  
   (b) Both silicon and carbon form oxides.
       (i) Describe the structure and bonding in SiO₂.  
       (ii) Explain why silicon dioxide is a solid and carbon dioxide is a gas at room temperature.  
       (c) Describe the bonding within the carbon monoxide molecule.  
       (d) Describe the delocalization of pi (π) electrons and explain how this can account for the structure and stability of the carbonate ion, CO₃²⁻.  
       (e) Explain the meaning of the term hybridization. State the type of hybridization shown by the carbon atoms in carbon dioxide, diamond, graphite, and the carbonate ion.  

21. SF₆, SF₅Cl, and SF₄Cl₂ have different shapes. Draw their Lewis (electron dot) structures and use the VSEPR theory to predict the name of the shape of each molecule.  

22. (a) Draw the Lewis (electron dot) structures, state the shape, and predict the bond angles for the following species.
    (i) PCl₅
    (ii) NH₄⁺
    (iii) XeF₆
    (b) (i) Compare the formation of a sigma (σ) and a pi (π) bond between two carbon atoms in a molecule.
        (ii) Identify how many sigma and pi bonds are present in propene, C₃H₆.
        (iii) Deduce all the bond angles present in propene.
        (iv) Explain how the concept of hybridization can be used to explain the bonding in the triple bond present in propene.

23. (a) Draw the Lewis (electron dot) structures, state the shapes, and predict the bond angles for the following species.
    (i) SF₅Cl²⁻  
    (ii) NO₂⁻  
    (b) Explain, using diagrams, why NO₂ is a polar molecule but CO₂ is a non-polar molecule.
    (c) Describe the structure and bonding in silicon dioxide.
    (d) Describe the molecule HCONH₂.
        (i) State the name of the compound and draw its structural formula, showing all the bonds present.
        (ii) Explain the term hybridization.
        (iii) Describe how sigma and pi bonds form.
Homework → Topic 0 Repaso de química

Q 9, 10, 18, 18

$\sigma$ bond is stronger

2 types

Covariant $\sigma$ overlap on bond axis

$\pi$: sideways

2 pairs

Double bond = 1 $\sigma$ + 1 $\pi$

Hybrid orbitals

$sp, sp^2, sp^3$

Hybridisation

Excitation of s electron

$CH_4$, methane → tetrahedral