



What you already know

- Bond strength
- VSEPR theory



What you will learn

- Hybridization
- Shape of hybridized orbitals
- Types of hybridization
- % s-character
- Steric number
- Relation between hybridization and geometry
- Berry pseudorotation

Need for Hybridization

In order to understand the concept of hybridization, let us first look into the need for hybridization.

Consider an example of methane (CH_4) molecule.

The electronic configuration of carbon in its ground state is given as follows:

C (Z = 6, ground state) $1s^2 2s^2 2p^2$

In the given electronic configuration, two unpaired electrons are present. So, the carbon can form two bonds. In order to form four bonds, the excitation of an electron takes place, and the electronic configuration is given as follows:

C (Z = 6, excited state) $1s^2 2s^1 2p^3$

From the given configuration, it can be seen that an electron present in 2s orbital can combine with 1s orbital of hydrogen to form one type of bond, while the three electrons present in 2p subshell can combine separately with 1s orbitals of three hydrogen atoms to form three different types of bonds. In other words, the C-H bond which is made with 2s orbital electron and the C-H bond which is made with 2p subshell electrons **should be different**, and both the types of bonds should have different bond strength and bond length.

The three p-orbitals are perpendicular to one another. So, in CH_4 , three C-H bonds should be at 90° . However, it has been found experimentally that all the four H-C-H bond angles are identical with a value of 109.5° . Furthermore, all the bonds are also of a similar length and strength.

These experimental results can be explained with the concept of hybridization.



Fig. 1: Ground state electronic configuration of carbon



Fig. 2: Excited state electronic configuration of carbon



Hybridization

This has been introduced by **Linus Pauling and Slater**, and it is a **theoretical** concept used to study some structural properties like the geometry of the molecules, which cannot be explained by the simple theories of valency.

Definition

The **process of mixing atomic orbitals of nearly the same energy** to produce a **set of entirely new orbitals of equivalent energy (known as hybrid orbitals)** is known as **hybridization**.

Important conditions

- The **orbitals present in the valence shell of the atom should** take part in hybridization.
- The orbitals participating should have the **same or nearly the same energy**. Example:

(a) 1s and 2s orbitals cannot mix because the orbitals have major energy differences. In 1s and 2s orbitals, there are different principal quantum numbers . Due to this, they both have different energies and hence, they cannot mix.	1s, 2s	Can't mix	✗
(b) 2s and 2p have the same principal quantum number . So, they can combine (or mix) to give hybridized orbitals.	2s, 2p	Can mix	✓
(c) 3s and 3p can mix easily because they have nearly the same energy. However, in 3s and 3d, there is some energy difference . So, mixing of these orbitals takes place under certain conditions.	3s, 3p, 3d	Can mix	✓
(d) They cannot mix because all the orbitals have different principal quantum numbers . Due to this, the energy difference between the orbitals is significantly high and mixing becomes impossible.	3s, 4d, 5s	Can't mix	✗
- Excitation of electrons (or **promotion of electrons**) is **not a necessary condition prior to hybridization**.
- The **orbitals participating** in hybridization **can be fully filled, half-filled, or empty**.

Salient features

- The **number of hybrid orbitals formed is equal to the number of atomic orbitals that take part** in hybridization. Two atomic orbitals (s and p) combine to give two hybridized orbitals (**Fig. 3**).
- The **hybridized orbitals** formed are of **equivalent energy and shape**. It is the larger lobe of the hybridized orbital that takes part in the bond formation.
- The shape of the hybrid orbitals obtained on mixing of s and p-orbitals is as shown:

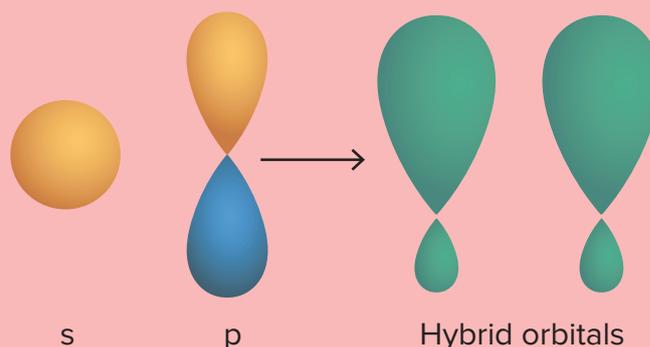


Fig. 3: Hybridized orbitals

However, the actual shape (mushroom type) of the hybridized orbital is given in **Fig. 4**. It is for the convenience that the shape drawn or used for the representation is the one shown in the right side of **Fig. 4**.

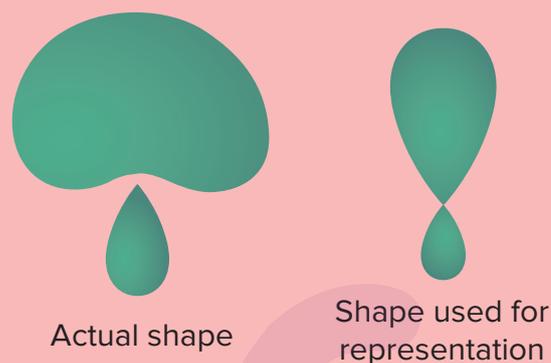
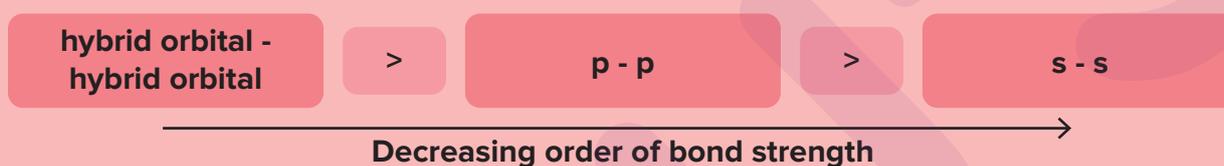


Fig. 4: Shape of hybrid orbital

- The **hybrid orbitals** are **more effective in forming stable sigma bonds** than that of the **pure atomic orbitals**. This is only valid for the same principal quantum number.

For same value of 'n'

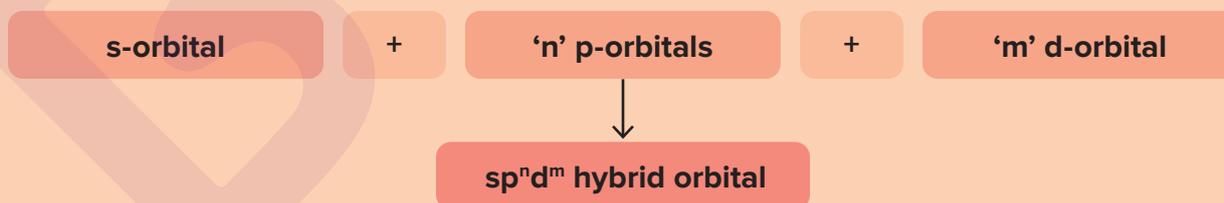


Also, the hybrid orbitals cannot form stable pi bonds because in **pi (π) bonds**, there should be a sideways overlapping, and for this, both the lobes should be of the same size. However, in a hybrid orbital, both the lobes have different sizes.

- The hybridized orbitals are directed in space in some preferred direction to have minimum repulsion between the electron pairs and thus, forms a stable arrangement. Thus, the type of hybridization indicates the geometry of the molecules.

Naming of Hybrid Orbitals

The naming of the hybrid orbitals is done on the basis of atomic orbitals that are participating in hybridization. For any given principal quantum number, there can be only one s-orbital. Thus, it can combine with 'n' p orbitals (where n lies between one to three) and 'm' d orbitals (where m lies between one to five). This is given as follows:



The types of hybridization and the number of hybrid orbitals involved can be summarised in the following table (**Table 1**):

S.No.	Number of orbitals mixed during hybridization			Type of hybridization	Number of hybrid orbitals
	s	p	d		
1.	1	1	0	sp	2
2.	1	2	0	sp ²	3

3.	1	3	0	sp^3	4
4.	1	3	1	sp^3d	5
5.	1	3	2	sp^3d^2	6
6.	1	3	3	sp^3d^3	7

Table 1: Types of hybridization



Note

- Hybridization takes place for the orbitals, not for the electrons. The orbitals participating can be **fully filled, half-filled, or empty**.
- In general, we use the term that hybridization of a molecule (say NH_3) is sp^3 , but it is to be noted that the **hybridization is a concept associated with an atom** (here, it is the central atom N, which has sp^3 hybridization).
- Excitation of electrons is not a necessary condition** prior to hybridization. For example, it is required in PCl_5 , but not needed in PCl_3 .



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Trick to Find Hybridization

- Draw the Lewis dot structure of the compound and mark the central atom.
- Count the number of sigma bonds attached to the central atom and the number of lone pairs on the central atom.
- For the central atom, apply the following formula:

$$\text{Steric number (S.N.)} = \text{Number of lone pairs} + \text{Number of sigma bonds}$$

Hybridization can be determined from the steric number. For example, if the steric number is 2, then two atomic orbitals (s and p) will participate to give two sp hybridized orbitals. Similarly, if the steric number is 5, then five atomic orbitals (one s, three p and one d) will participate to give five sp^3d hybridized orbitals.

Steric number (S.N.)	Hybridization	Number of hybrid orbitals
2	sp	2
3	sp^2	3
4	sp^3	4
5	sp^3d	5

6	sp^3d^2	6
7	sp^3d^3	7
8	sp^3d^4	8
9	sp^3d^5	9

Table 2: Steric number and hybridization

Example: NH_3

Number of sigma bonds = 3

Number of lone pairs = 1

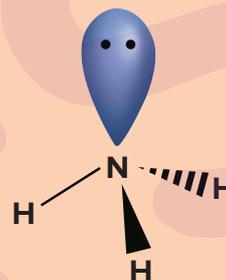


Fig. 5: NH_3 structure

Steric number = Number of sigma bonds + Number of lone pairs

$$S.N. = 3 + 1 = 4$$

From the given table, we can say that the hybridization of NH_3 (N atom) is sp^3 .



Fig. 6: Electronic configuration of nitrogen



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% s-character

The % s-character is calculated by the given formula. It gives the **percentage contribution of s-orbital to a hybrid orbital in a given hybridization.**

$$\% \text{ s-character} = \frac{\text{Number of s-orbitals}}{\text{Number of (s + p) orbitals}} \times 100$$

Examples:

• **sp**

Number of s-orbital = 1

Number of (s + p) orbitals = 2

$$\% \text{ s-character} = \frac{1}{2} \times 100$$

% s-character = 50%

• **sp^3**

Number of s-orbital = 1

Number of (s + p) orbitals = 4

$$\% \text{ s-character} = \frac{1}{4} \times 100$$

% s-character = 25%

Hybridization	% s-character	% p-character
sp	50	50
sp ²	33.33	66.66
sp ³	25	75

Table 3: % s and p-characters in different hybridizations

As the s-character increases, the **bond length decreases**. As the s-orbital is more penetrating, it is **closer to the nucleus**. The bond angle **increases from 109.5° in sp³ to 180° in sp**. As the s-character increases in a hybridized orbital, the **electronegativity of the concerned atom also increases**.

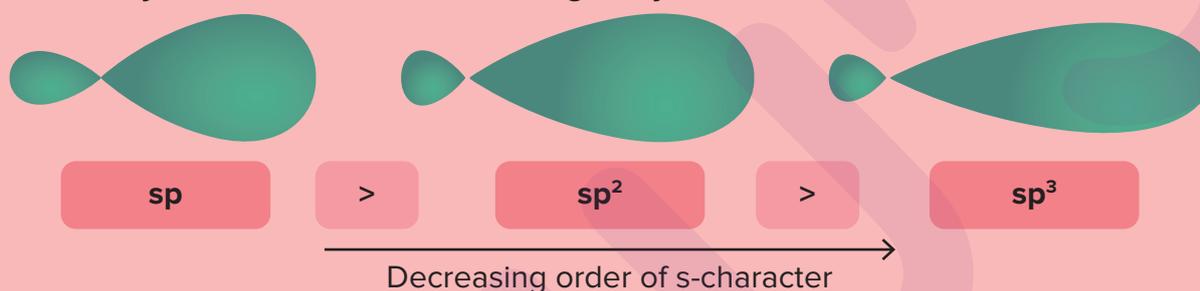


Fig. 7: % s-character in hybrid orbital

Significance of % s-character

As % s-character increases	As % p or % d-character increases
Orbital becomes bulkier and shorter	Orbital becomes thinner and longer
Energy of hybrid orbital decreases	Energy of hybrid orbital increases
Electronegativity increases	Electronegativity decreases

Table 4: Comparison between s and p-character



Energy level for different hybrid orbitals

In case of alkynes, the carbon is sp hybridized, and therefore, it has higher electronegativity as compared to the carbon present in alkanes, where carbon is sp^3 hybridized.

- As the % s-character increases, the stability of the hybrid orbitals increases and the strength of the bonding increases.

 $sp - p$

>

 $sp^2 - p$

>

 $sp^3 - p$

>

 $p - p$

Decreasing order of bond strength \rightarrow

Bond strength of different hybrid orbitals with p-orbital

Steric Number Formula

The formula for steric number is given as follows:

$$\text{S.N.} = \frac{V + M - q}{2}$$

Where, **S.N.** = Steric number

V = Number of valence electrons on the central atom

M = Number of monovalent surrounding atoms

q = Charge on the species.

Examples:

Compound	V	M	q	S.N.	Hybridization
CH_4	4	4	0	4	sp^3
CO_2	4	0	0	2	sp
CO_3^{2-}	4	0	-2	3	sp^2
O_3	6	0	0	3	sp^2
ICl_3	7	3	0	5	sp^3d
XeF_5^-	8	5	-1	7	sp^3d^3
ClO_3^+	7	0	1	3	sp^2
XeF_5^+	8	5	1	6	sp^3d^2

Table 5: Finding hybridization for the molecules using steric number formula

The geometry of the molecule can be predicted from its hybridization (or the hybridization of the central atom of the molecule). This is summarised in **Table 6**.

S.N.	Types of hybridization	Geometry	Orbitals Involved
2	sp	Linear	s, p _x /p _y /p _z
3	sp ²	Trigonal planar	s, p _x , p _y /p _y , p _z /p _z , p _x
4	sp ³	Tetrahedral	s, p _x , p _y , p _z
5	sp ³ d	Trigonal bipyramidal	s, p _x , p _y , p _z , d _{z²}
6	sp ³ d ²	Octahedral	s, p _x , p _y , p _z , d _{z²} , d _{x²-y²}
7	sp ³ d ³	Pentagonal bipyramidal	s, p _x , p _y , p _z , d _{z²} , d _{x²-y²} , d _{xy}

Table 6: Geometry for the given hybridization of a molecule

Let us take each hybridization one by one and understand the geometry aspect with the help of examples.

sp hybridization

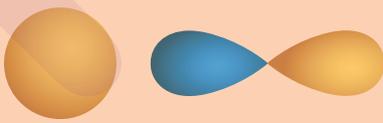
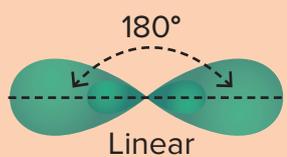
Hybridization	Number of hybrid orbitals	Orbitals involved in mixing	Geometry
sp	2	 one s + one p	 Linear

Table 7: sp hybridization

Example:

- BeCl₂

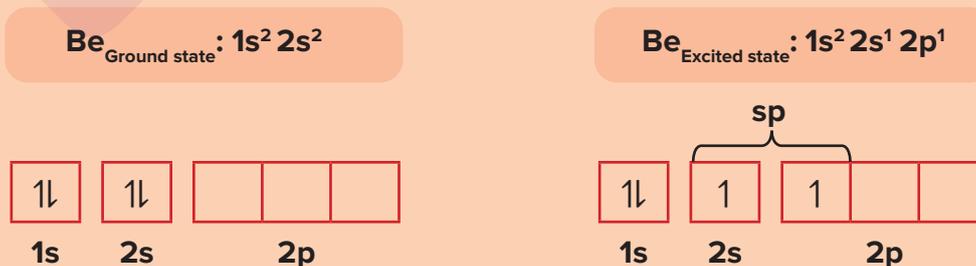


Fig. 8: Ground and excited state electronic configuration of Be

In the ground state, Be does not have any unpaired electron. So, excitation is needed. In the excited state, two unpaired electrons are present. The 2s and 2p orbitals undergo mixing and two sp hybrid orbitals are obtained. Each of the sp hybridized orbital forms a sigma bond with the p orbital of chlorine. The geometry obtained is a linear one.

sp² hybridization

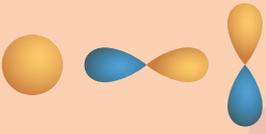
Hybridization	Number of hybrid orbitals	Orbitals involved in mixing	Geometry
sp ²	3	 one s + two p	 120° Trigonal planar

Table 8: sp² hybridization

Example:

- BF₃

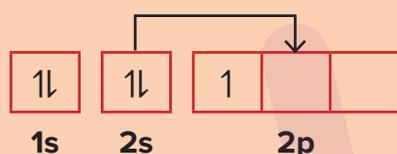
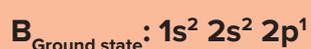


Fig. 9: Ground state configuration of boron

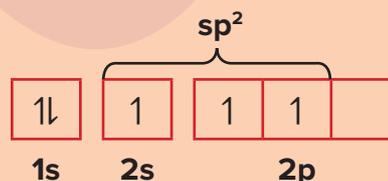


Fig. 10: Excited state configuration of boron

There is only **one unpaired electron** in the ground state of B (Fig. 9) and there are **three sigma bonds** in BF₃. So, in order to form three bonds, there is a requirement for the promotion of one electron. After the excitation, there are three unpaired electrons (Fig. 10), and the **mixing of 2s and two 2p orbitals** takes place resulting in the formation of **three sp² hybridized orbitals**. The **geometry obtained is trigonal planar**.

sp³ hybridization

Hybridization	Number of hybrid orbitals	Orbitals involved in mixing	Geometry
sp ³	4	 one s + three p	 109.5° Tetrahedral

Table 9: sp³ hybridization

Example:

- CH_4

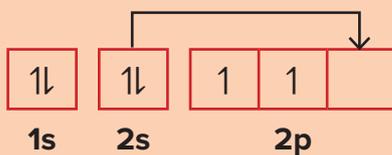


Fig. 11: Ground state electronic configuration of carbon

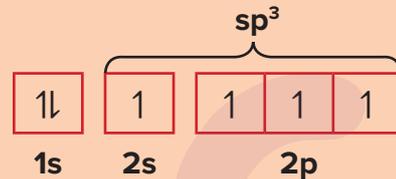


Fig. 12: Excited state electronic configuration of carbon

In the ground state (Fig. 11) of carbon, there are **two unpaired electrons**. However, in CH_4 , **four sigma bonds are formed**. So, excitation of electrons takes place. In the excited state of carbon (Fig. 12), there are **four unpaired electrons**. So, it can form four bonds. Hence, four (one s and three p-orbitals) atomic orbitals undergo mixing to obtain four sp^3 hybridized orbitals.



Hybridization Involving d-orbitals

The energies of **3s and 3p atomic orbitals** are nearly the same. They are similar for **4s and 4p orbitals**. Hence, 3s and 3p atomic orbitals (or 4s and 4p orbitals) can mix together. However, the energy of 3d orbital is higher as compared to 3p orbital, but more closer to 4s orbital. That means 3d orbital can easily mix with 4s and 4p orbitals and also with 3s and 3p orbitals (but under certain conditions).

Energy of 3d orbitals is comparable to the energy of

3s and 3p

4s and 4p

Relation between energy of 3d with 3s and 3p, and 4s and 4p

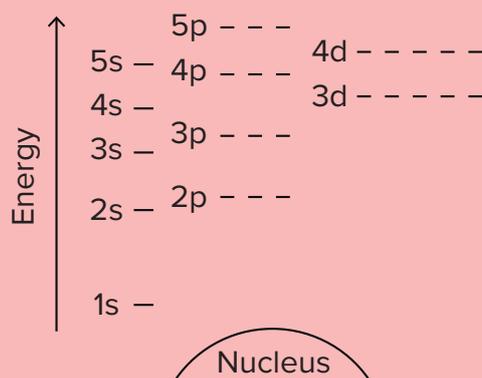


Fig. 13: Energy diagram of orbitals

sp³d hybridization

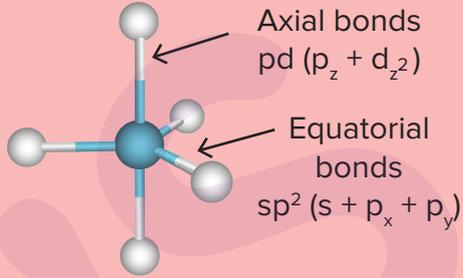
Hybridization	Number of hybrid orbitals	Orbitals involved in mixing	Geometry
sp ³ d	5	s, p _x , p _y , p _z , d _{z²}	 <p>Trigonal bipyramidal</p>

Table 10: sp³d hybridization

Example:

PCl₅

The ground state electronic configuration of phosphorus is given (Fig. 14). There are three unpaired electrons. So, it can form three bonds. However, in order to form five bonds (in PCl₅), the excitation of the electron takes place (Fig. 15). Hence, one 3s orbital, three 3p orbitals and one d-orbital (d_{z²}) combine to form five sp³d hybridized orbitals.

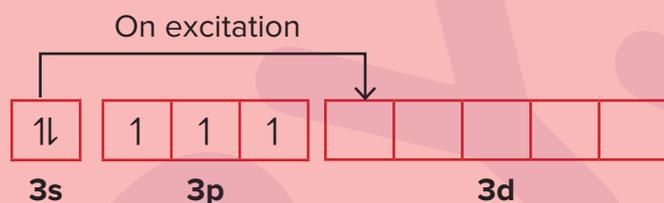


Fig. 14: Ground state electronic configuration of phosphorus

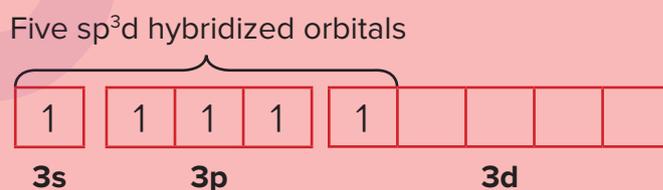


Fig. 15: Excited state electronic configuration of phosphorus



Note

In PCl₃, no excitation is required as three unpaired electrons are present and thus, three bonds can be formed.

BOARDS

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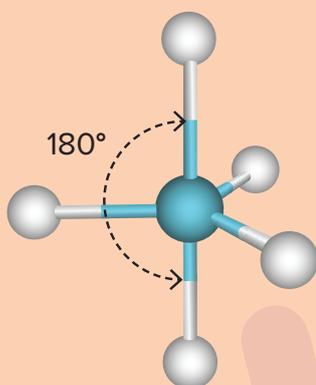
Trigonal Bipyramidal Geometry (TBP Geometry)

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In TBP, the hybridization is sp^3d , and it is not a pure hybridization but a **mixed hybridization**. It is considered to be a **combination of sp^2 and pd hybridization**. In pd (p_z, d_{z^2}) hybridization, there is zero % s-character and these hybridized orbitals are considered to be involved in forming bonds at the axial positions, while sp^2 (s, p_x, p_y) hybridized orbitals are considered to be involved in forming bonds at the equatorial positions.

Since there is **zero % s-character involved in pd hybridized orbitals**, the **axial bonds are longer and weaker**. In sp^2 hybridization, the **% s-character is 33.33%**, and these hybridized orbitals are present in equatorial positions. Thus, the **equatorial bonds are shorter and stronger**. The axial bonds have a **bond angle of 180°** to each other and the equatorial bonds have a **bond angle of 120°** between them.

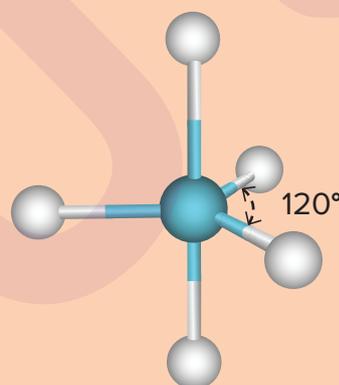
Axial bonds



$$\% s = 0$$

Longer and weaker bond

Equatorial bonds



$$\% s = 33.33$$

Shorter and stronger bond

Fig. 16: Axial and equatorial bonds in trigonal bipyramidal geometry



Note

pd hybridization is not an independent hybridization.

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Bent's Rule

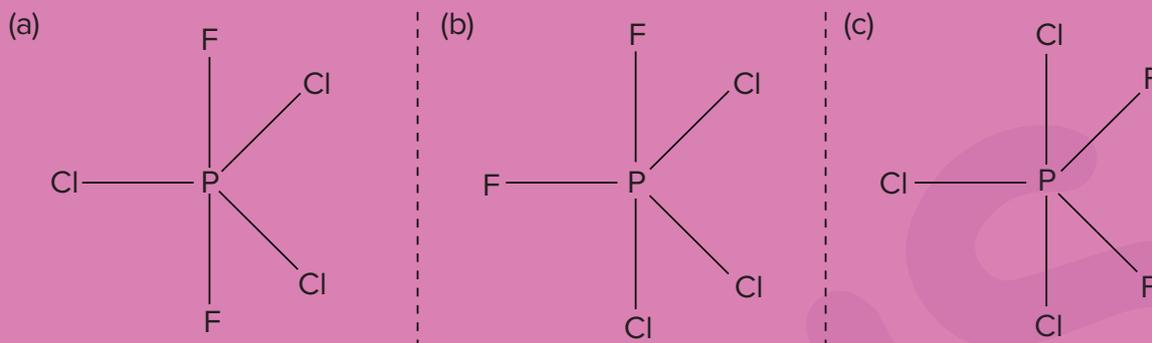
According to this rule,

- **Axial bonds** are made with **pd hybridization**. So, they have **zero % s-character**. Hence, these positions are occupied by the **most electronegative atoms** present in a compound.
- **Equatorial bonds** are made with **sp^2 hybridization**. They have **33.33% s-character**. Hence, these positions are occupied by the **lone pairs, multiple bonds, or less electronegative atoms**.



Application of Bent's rule

Which of the following is the **correct structure for PCl_3F_2** ?



Solution

In PCl_3F_2 ,

Number of sigma bonds = 5

Number of lone pairs = 0

S.N. = Number of sigma bonds + Number of lone pairs

S.N. = 5 + 0 = 5

So, hybridization = sp^3d (TBP geometry)

According to **Bent's rule**, in TBP geometry, the axial bonds are occupied by the **most electronegative** atoms due to **zero % s-character**. F has more electronegativity than Cl. Hence, the axial positions are occupied by **F**.

The equatorial positions are occupied by the less electronegative atoms and Cl has **less electronegativity than F**. Hence, the equatorial positions are occupied **by Cl**.

So, the correct representation is (a).

Berry Pseudorotation

Berry pseudorotation is the **isomerism phenomenon** in which **two axial bonds** are interconverted with the two **equatorial bonds** in TBP geometry. For steric number five, there are two geometries that are possible: **trigonal bipyramidal and square pyramidal**. In pseudorotation, when the interconversion of axial and equatorial bonds takes place, a transition state having square pyramidal geometry is observed.

Example:

In PF_5 , all the bonds are **identical** and they have the same **average bond length** and strength due to pseudorotation. This mechanism only happens when the energy difference between the TBP and square pyramidal geometry is less than the thermal energy at the room temperature.

PF_5 is also known as a **fluxional molecule** [Fluxional (or non-rigid) molecules are the molecules that undergo dynamics such that some or all of their atoms interchange between symmetry-equivalent positions].

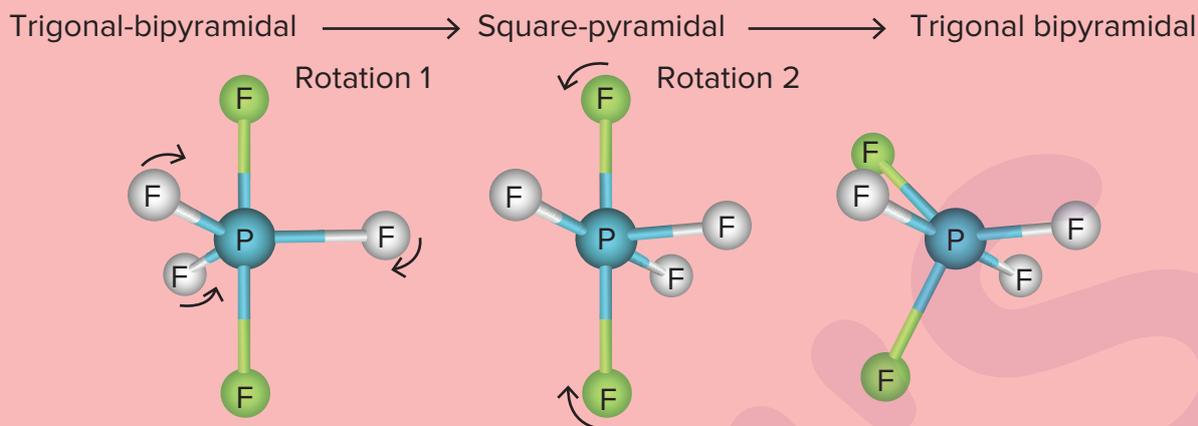


Fig. 17: Pseudorotation in PF_5

Other examples of sp^3d hybridization are as follows:

- SF_4

It has sp^3d hybridization with **four bonds and one lone pair**. The geometry is trigonal bipyramidal and the **shape is see-saw**.

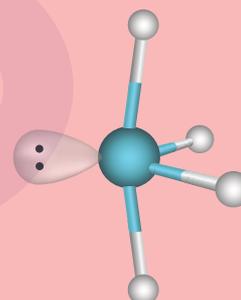


Fig. 18: SF_4 molecule

- XeF_2

It has sp^3d hybridization with **two bonds and three lone pairs**. The geometry is trigonal bipyramidal and the **shape is linear**.

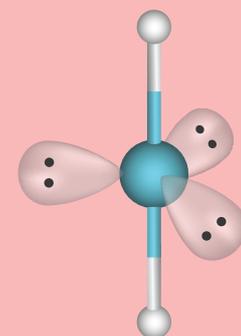


Fig. 19: XeF_2 molecule

- XeOF_2

It has sp^3d hybridization with **three sigma bonds and two lone pairs**. The geometry is trigonal bipyramidal and the shape is **T-shape**. According to Bent's rule, **lone pairs and double bond** will occupy **equatorial positions** because they feel **less repulsion** whereas fluorine occupy axial positions.

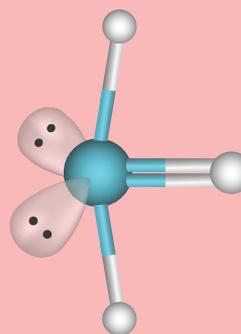


Fig. 20: XeOF_2 molecule

sp³d² hybridization

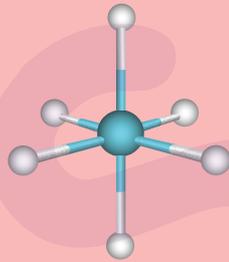
Hybridization	Number of hybrid orbitals	Orbitals involved in mixing	Geometry
sp ³ d ²	6	s, p _x , p _y , p _z , d _{z²} , d _{x²-y²}	 Octahedral

Table 11: sp³d² hybridization

Example:

SF₆

The ground state electronic configuration of sulphur is given (Fig. 21). There are two unpaired electrons. So, it can form two bonds. However, in order to form six bonds (in SF₆), the excitation of the electrons takes place (Fig. 22). Hence, one 3s orbital, three 3p orbitals, and two d-orbitals (d_{z²}, d_{x²-y²}) combine to form six sp³d² hybridized orbitals.

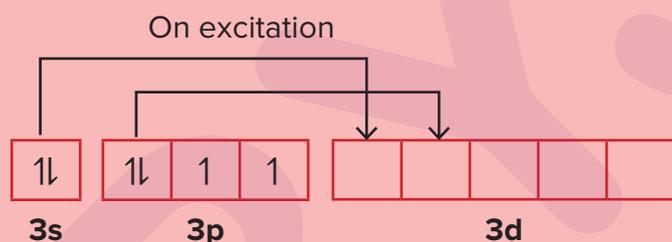


Fig. 21: Ground state configuration of sulphur

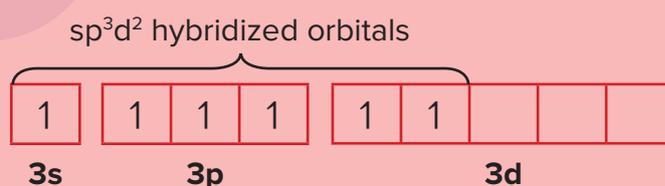


Fig. 22: Excited state configuration of sulphur

Other examples of sp³d² are as follows:

- XeOF₄

Number of sigma bonds = 5

Number of lone pairs = 1

S.N. = 5 + 1 = 6 (sp³d² hybridization)

The geometry of XeOF₄ is octahedral and the shape is square pyramidal.

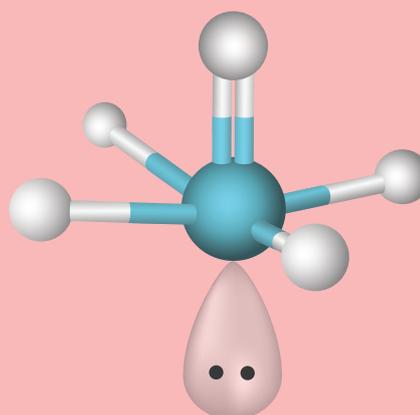


Fig. 23: XeOF₄ molecule

- XeF_4

Number of lone pairs = 2

Number of sigma bonds = 4

S.N. = 2 + 4 = 6 (sp^3d^2 hybridization)

It has an octahedral geometry and a **square planar shape**.

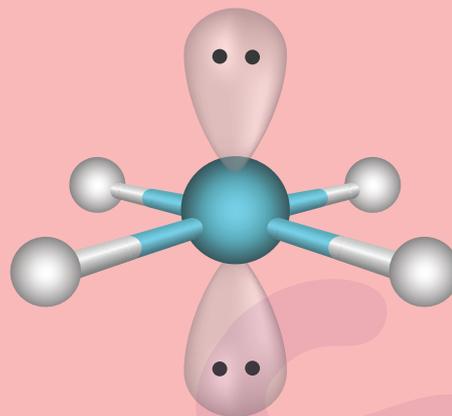


Fig. 24: XeF_4 molecule



In sp^3d^2 hybridization (octahedral geometry), there are no axial and equatorial bonds. All the bonds are identical in nature.



Application of hybridization in finding $\text{d}\pi\text{-p}\pi$ bonds

What is the **number of $\text{d}\pi\text{-p}\pi$ bonds** present in SO_2 , SO_3 and ClO_4^- , respectively?

(a) 0, 1, 2

(b) 1, 2, 3

(c) 2, 3, 4

(d) 2, 3, 3

Solution

Step 1: Calculation of hybridization

Compounds	SO_2	SO_3	ClO_4^-
Lewis structure			
Number of sigma bonds	2	3	4
Number of lone pairs on the central atom	1	0	0
Steric number	3	3	4
Hybridization	sp^2	sp^2	sp^3

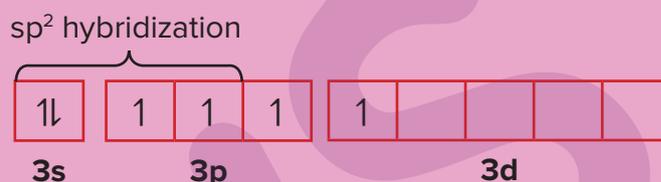
Step 2: Finding the $d\pi-p\pi$ bonds in SO_2

The ground state electronic configuration of S is $[\text{Ne}] 3s^2 3p^4$.

In SO_2 , the covalency of S is four. So, we need excitation of only 3p orbital electron before mixing during hybridization.



In the excited state, there are four unpaired electrons. Two of the three sp^2 hybridized orbitals participate in the sigma bond formation with two oxygen atoms (third hybridized orbital contains the lone pair), while the unhybridized orbitals having unpaired electrons participate in the pi bond formation.



Since there is only one d-orbital having an unpaired electron, it can get involved in the **$d\pi-p\pi$ bond formation, where the p-orbital is of the O atom**. Also, there will be a formation of **$p\pi-p\pi$ bond** (as one of the unhybridized orbitals containing an unpaired electron is a p-orbital).

Thus, **one $d\pi-p\pi$ bond is present in the SO_2 molecule.**

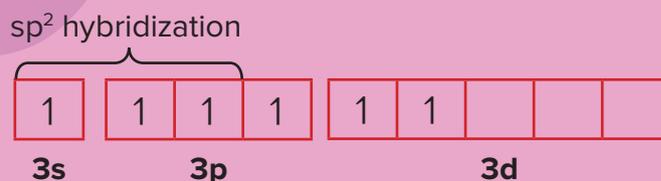
Step 3: Finding the $d\pi-p\pi$ bonds in SO_3

The ground state configuration of S is $[\text{Ne}] 3s^2 3p^4$.

In SO_3 , the covalency of S is six. So, the excitation of both 3s and 3p orbitals takes place.



In the excited state, there are six unpaired electrons. The three sp^2 hybridized orbitals participate in the sigma bond formation with the three O atoms, while the unhybridized orbitals having unpaired electrons participate in the pi bond formation. Since there are two d-orbitals having unpaired electrons, only these can get involved in the **$d\pi-p\pi$ bond formation, where the p-orbital is of the O atom**.



Thus, **two $d\pi-p\pi$ bonds are present in the SO_3 molecule.**

Step 4: Finding the $d\pi-p\pi$ bonds in ClO_4^-

The ground state electronic configuration of Cl is $[\text{Ne}] 3s^2 3p^5$. In ClO_4^- , the covalency of Cl is seven. So, the excitation of electrons takes place.

The excited state electronic configuration of Cl is $[\text{Ne}] 3s^1 3p^3 3d^3$. It has sp^3 hybridization.

In the excited state, there are seven unpaired electrons. The four sp^3 hybridized orbitals participate in the sigma bond formation with the four O atoms, while the unhybridized orbitals having unpaired electrons participate in the pi bond formation. Since there are three unhybridized d-orbitals having unpaired electrons, only these can get involved in the **$d\pi-p\pi$ bond formation, where the p-orbital is of the O atom**.

So, in ClO_4^- , there are three $d\pi-p\pi$ bonds.

Hence, option (b) is the correct answer.



Summary

- The **process of mixing of atomic orbitals of nearly the same energy** to produce a **set of entirely new orbitals of equivalent energy (known as hybrid orbitals)** is known as **hybridization**.
- Important conditions of hybridization are as follows:**
 - The **orbitals present in the valence shell of the atom** should take part in hybridization.
 - The orbitals participating should have the **same or nearly the same energy**.
 - Excitation of electrons (or **promotion of electrons**) is **not a necessary condition prior to hybridization**.
 - The **orbitals participating** in hybridization **can be fully filled, half-filled, or empty**.
- Salient features of hybridization are as follows:**
 - The **number of hybrid orbitals formed** is **equal to the number of atomic orbitals that take part** in hybridization.
 - The **hybridized orbitals** formed are of **equivalent energy and shape**.
 - It is the **larger lobe of the hybridized orbital** that **takes part in the bond formation**. The **hybrid orbitals** are **more effective** in **forming stable sigma bonds** than the **pure atomic orbitals**. The hybridized orbitals are not involved in the pi bonding.
 - The hybridization / hybridized orbitals are directed in space in some preferred direction to have a minimum repulsion between the electron pairs and thus, a stable arrangement. The type of hybridization / hybridized indicates the geometry of the molecules.
- Steric number**

$$\text{S.N.} = \frac{V + M - q}{2}$$

Where,

S.N. = steric number.

V = number of valence electrons on the central atom.

M = number of monovalent surrounding atoms.

q = charge on the species.

- s-character and p-character**

Hybridization	% s-character	% p-character
sp	50	50
sp ²	33.33	66.67
sp ³	25	75

• **Orbitals involved and geometry for the given hybridization**

S.N.	Types of hybridization	Geometry	Orbitals Involved
2	sp	Linear	$s, p_x/p_y/p_z$
3	sp^2	Trigonal planar	$s, p_x, p_y/p_y, p_z/p_z, p_x$
4	sp^3	Tetrahedral	s, p_x, p_y, p_z
5	sp^3d	Trigonal bipyramidal	$s, p_x, p_y, p_z, d_{z^2}$
6	sp^3d^2	Octahedral	$s, p_x, p_y, p_z, d_{z^2}, d_{x^2-y^2}$
7	sp^3d^3	Pentagonal bipyramidal	$s, p_x, p_y, p_z, d_{z^2}, d_{x^2-y^2}, d_{xy}$

- **Bent's rule:** According to this, axial bonds are made with pd hybridization. So, they have **zero %s-character**. Hence, these positions are occupied by the **most electronegative atoms** present in a compound. Equatorial bonds are made with sp^2 hybridization. They have **33.33% s-character**. Hence, these positions are occupied by the **lone pairs, multiple bonds, or less electronegative atoms**.
- **Berry pseudorotation:** It is the **isomerism phenomenon** in which **two axial bonds** are interconverted with two **equatorial bonds** in TBP geometry. For steric number five, there are two geometries that are possible: **trigonal bipyramidal and square pyramidal**.

In pseudorotation, when the interconversion of axial and equatorial bonds takes place, a transition state having a square pyramidal geometry is observed. For example, PF_5 shows Berry pseudorotation.