



What you already know

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



What you will learn

- Types of hybridization
- Hybridization in odd electron species
- Some molecules that change hybridization in solid states
- Resonance
- Resonating structures and resonance hybrid
- Bond order



MAIN



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sp^3d^3 Hybridization

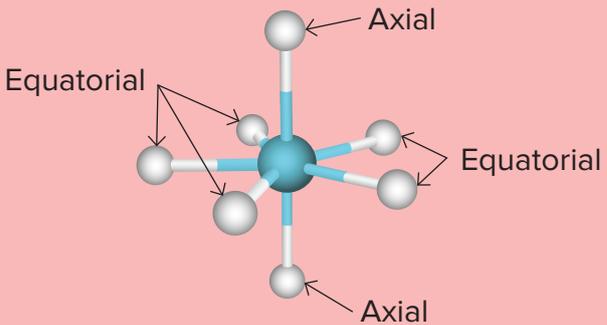
Hybridization	Number of hybrid orbitals/steric number	Geometry
sp^3d^3	7	 <p>Pentagonal bipyramidal</p>

Table 1: Geometry of sp^3d^3 hybridization

Example:

IF_7

In the ground state of iodine, there is only one unpaired electron. However, in the Lewis structure of IF_7 , there are seven sigma bonds, so the **excitation of electrons** from s and p to d-orbitals **is needed**. In the excited state, iodine has **seven unpaired electrons**, so it can form seven bonds. Hence, seven (**one s, three p and three d-orbitals**) orbitals undergo mixing to form seven sp^3d^3 hybridized orbitals, which are arranged in a **pentagonal bipyramidal geometry**.

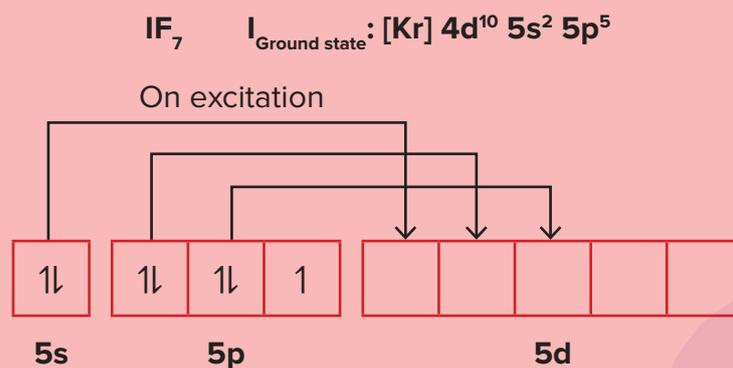


Fig. 1: Ground state electronic configuration of iodine



Fig. 2: Excited state electronic configuration of iodine in IF_7

There are two types of positions in the pentagonal bipyramidal geometry, i.e., **axial** and **equatorial** as in the trigonal bipyramidal geometry. However, in this case, the angle between the equatorial-equatorial bonds (72°) is less than that of the axial-equatorial bond (90°). Hence, the repulsion in the electron pairs is more at equatorial positions than that in axial positions. As a result, the **lone pair tends to occupy axial positions** in this case. Also, due to more repulsions, the **equatorial bonds are longer** than the axial bonds.

Bond length
(Equatorial)

>

Bond length
(Axial)

Other examples of sp^3d^3 hybridization

XeF_6

1. The ground state electronic configuration of xenon is $[\text{Kr}] 4d^{10} 5s^2 5p^6$ and it can be written as follows:



Fig. 3: Ground state electronic configuration of xenon

2. The excited state electronic configuration of xenon can be written as follows:

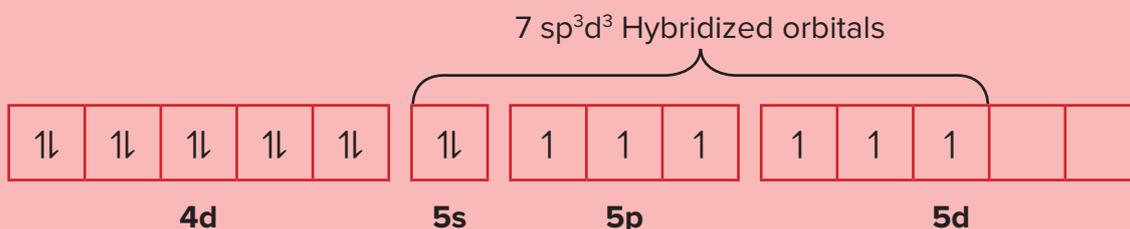


Fig. 4: Excited state electronic configuration of xenon in XeF_6

In XeF_6 , one s, three p and three d-orbitals of xenon undergo mixing to give seven hybridized orbitals. The mixing of these orbitals is known as **sp^3d^3 hybridization**. One of them has an unshared pair of electrons and the other six have one electron each, which can be used to form bonds with six monovalent F-atoms. In XeF_6 , the **lone pair is present towards any of the trigonal planes** which is formed by joining any of the three adjacent side atoms, and the side atoms are present at the six positions of an octahedron. Due to the presence of this lone pair, the geometry is distorted from the pure octahedral. However, the **distortion is not so considerable**. Hence, the shape of this molecule is **distorted octahedral**.

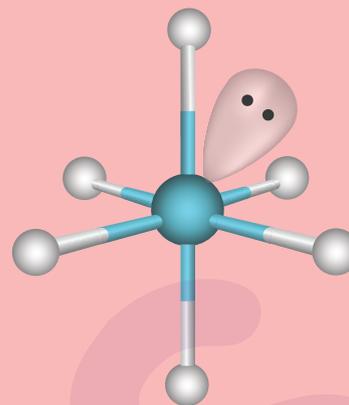


Fig. 5: Distorted octahedral structure of XeF_6

XeF_5^-

- The ground state electronic configuration of xenon is $[\text{Kr}] 4d^{10} 5s^2 5p^6$ and it can be written as follows:



Fig. 6: Ground state electronic configuration of xenon

- The excited state electronic configuration of xenon in XeF_5^- can be written as follows:



Fig. 7: Excited state electronic configuration of xenon in XeF_5^-

In XeF_5^- , one s, three p and three d-orbitals of xenon undergo mixing to give seven hybridized orbitals. The mixing of these orbitals is known as **sp^3d^3 hybridization**. After mixing, two orbitals have unshared pairs of electrons and others are singly occupied, which are used to form bonds with F-atoms. To have minimum repulsion between the lone pairs, they will be present at axial positions, whereas five bond pairs are present at equatorial positions. Hence, the shape of this molecule is **pentagonal planar**.

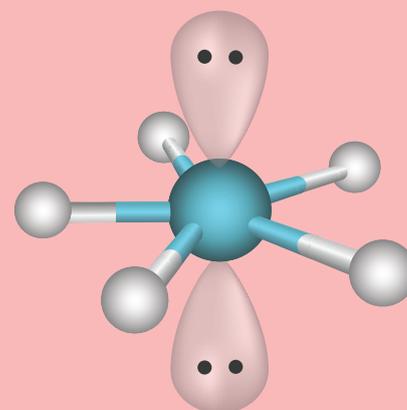


Fig. 8: Pentagonal planar

sp³d⁴ Hybridization

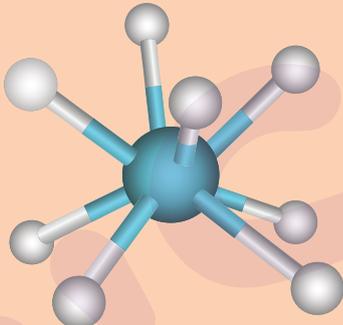
Hybridization	Number of hybrid orbitals/steric number	Geometry
sp ³ d ⁴	8	 <p>Square antiprismatic</p>

Table 2: Geometry of sp³d⁴ hybridization

Example:

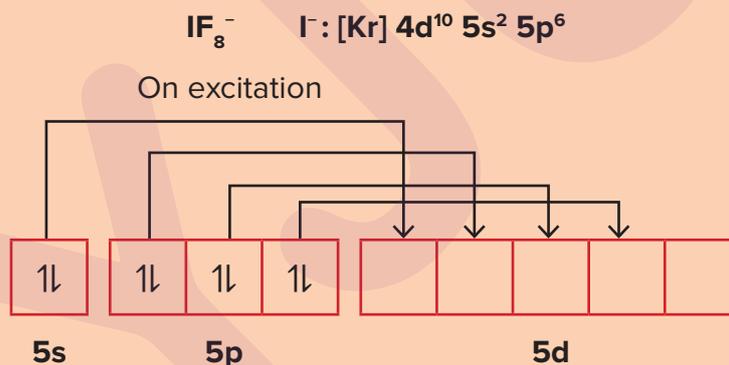


Fig. 9: Ground state electronic configuration of iodine

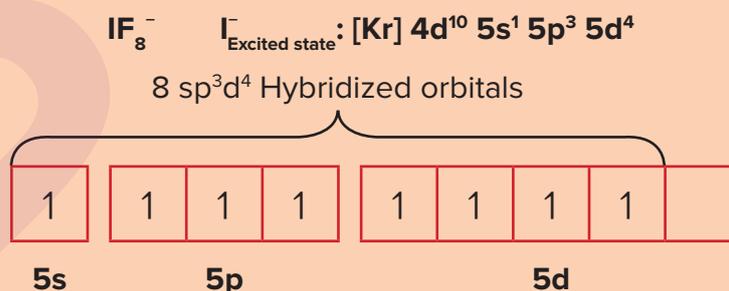


Fig. 10: Excited state electronic configuration of iodine in IF_8^-

In IF_8^- , one s, three p and four d-orbitals of iodine undergo mixing to give eight hybridized orbitals. The mixing of these orbitals is known as **sp³d⁴ hybridization**. After mixing, each orbital contains one electron each that can be used to form a single bond with F-atoms, and the shape of this species comes out to be square antiprismatic.

Square antiprismatic geometry

Consider two squares, where one square is superimposed (exactly covering) over the other. Now, rotate one of the squares by 45° (in the plane) with respect to the other, and join the vertices as shown in **Fig. 11**). The resultant shape so formed is known as a square antiprism. The molecular geometry that forms a square antiprism on joining the valence shell electron pairs of central atoms is known as square antiprismatic geometry.

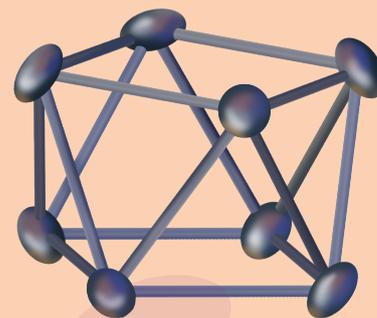


Fig. 11: Square antiprismatic geometry



Note

XeF_8 will also have a square antiprismatic geometry.

Hybridization in Odd Electron Species

The valence bond theory fails to explain hybridization in odd electron species. Generally, orbitals with odd valence electrons **do participate in hybridization**. However, experimentally, **exceptions like alkyl radicals** (CR_3^\cdot , where $\text{R} = \text{H}, \text{Me}, \text{etc.}$) are found in which the odd electron containing orbitals do not participate in hybridization.



Note

If the central atom has an odd electron and the terminal atom is highly electronegative, then the odd electron orbital will be involved in hybridization.

Examples:

1. CF_3^\cdot

In CF_3^\cdot , the orbital containing a single electron is involved in hybridization. Therefore, the steric number, i.e., the number of sigma bond pairs + the number of lone pairs, on the central atom is 4. Hence, the hybridization is sp^3 , and the shape of CF_3^\cdot radical is **trigonal pyramidal**.

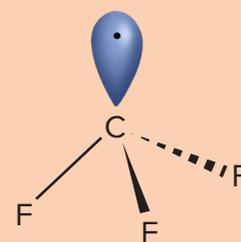


Fig. 12: Shape of CF_3^\cdot radical

2. CH_3^\cdot

In CH_3^\cdot , the orbital containing a single electron is not involved in hybridization (since it is an alkyl halide). Therefore, the steric number of the central atom is 3. Hence, the hybridization is sp^2 , and the shape of CH_3^\cdot radical is **trigonal planar**.

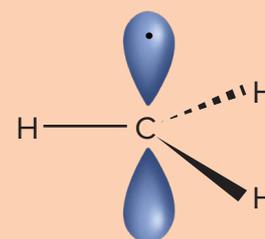


Fig. 13: Shape of CH_3^\cdot radical

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Hybridization of Phosphorus Halides in Solid State

PCl₅

In the gaseous state, PCl₅ has a trigonal bipyramidal geometry, which is not so symmetrical. Therefore, in the solid state, PCl₅ dissociates into cationic [PCl₄]⁺ and anionic [PCl₆]⁻. **Hybridizations of [PCl₄]⁺ and [PCl₆]⁻ are sp³ and sp³d², respectively, and they possess tetrahedral and octahedral shapes, which are more symmetrical than trigonal bipyramidal geometry.**

**PBr₅**

In the gaseous state, PBr₅ has a trigonal bipyramidal geometry, which is not so symmetrical. Therefore, in the solid state, PBr₅ dissociates into a cationic form [PBr₄]⁺ and an anion Br⁻. Hybridizations of [PBr₄]⁺ is sp³, and the shape is tetrahedral, which is more symmetrical than trigonal bipyramidal and is therefore more stable. Also, due to the large size of the bromine atom, phosphorus could not accommodate six Br-atoms. Hence, [PBr₆]⁻ does not exist.

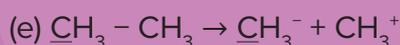
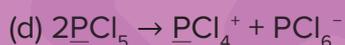
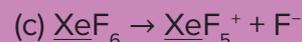
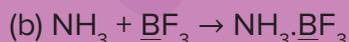
**PI₅**

PI₅ does not exist, since phosphorus could not accommodate five large iodine atoms around itself.



Based on hybridization

Find the **number of transformations** among the following that involve the **change in hybridization** of the underlined atom.



Solution

Desired species on reactant side	Number of sigma bond pairs (b.p.)	Number of lone pairs on the central atom (l.p.)	Steric number (S.N.) = b.p. + l.p.	Hybridization	Desired species on the product side	Number of sigma bond pairs (b.p.)	Number of lone pairs on the central atom (l.p.)	Steric number (S.N.) = b.p. + l.p.	Hybridization
H ₂ <u>O</u>	2	2	4	sp ³	H ₃ <u>O</u> ⁺	3	1	4	sp ³
<u>B</u> F ₃	3	0	3	sp ²	NH ₃ · <u>B</u> F ₃	4	0	4	sp ³

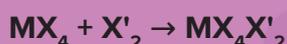
XeF_6	6	1	7	sp^3d^3	XeF_5^+	5	1	6	sp^3d^2
PCl_5	5	0	5	sp^3d	PCl_4^+	4	0	4	sp^3
$\text{H}_3\text{C}-\text{CH}_3$	4	0	4	sp^3	H_3C^-	3	1	4	sp^3

Therefore, only three transformations are bringing the change in hybridization.



Based on hybridization

Consider the following reaction:



If the atomic number of M is 52, X and X' are halogens, and X' is more electronegative than X, then choose the **incorrect** statements regarding the given information.

- Both X' atoms occupy axial positions that are formed by the overlapping of p and d-orbitals only.
- All M-X bond lengths are identical in both MX_4 and $\text{MX}_4\text{X}'_2$ compounds.
- Central atom M does not use any valence non-axial set of d-orbitals in the hybridization of the final product.
- Hybridization of the central atom M remains the same in both the reactant and the final product.

Solution

The atomic number of Xe that is present in the fifth period is 54, and the atomic number of given M is 52. Hence, it will belong to the 16th group and fifth period of the periodic table. Thus, the given **M is Te**.

Steric number (S.N.) = Number of sigma bond pairs (b.p.) + Number of lone pairs on the central atom (l.p.)

Steric number of Te in $\text{TeX}_4 = 4 + 1 = 5$

Therefore, hybridization = sp^3d , and the geometry of **TeX_4 is trigonal bipyramidal**.

Steric number of Te in $\text{TeX}_4\text{X}'_2 = 6 + 0 = 6$

Therefore, hybridization = sp^3d^2 , and the geometry of **$\text{TeX}_4\text{X}'_2$ is octahedral**.

- Since there are no axial and equatorial positions in the octahedral, all the positions are equivalent positions. Thus, this is an **incorrect** statement.
- In the case of $\text{TeX}_4\text{X}'_2$, since the geometry is octahedral, all the M-X bond lengths are equal. However, in the case of TeX_4 , since there are two different types of M-X bonds, i.e., equatorial and axial, and we know that axial bond lengths are larger than equatorial bond lengths, all M-X bond lengths in TeX_4 are not equal. Hence, this is also an **incorrect** statement.
- The d-orbitals involved in octahedral geometry are $\text{d}_{x^2-y^2}$ and d_{z^2} , which are axial. Thus, no non-axial d-orbitals are involved in the hybridization in the final product. Hence, this is a **correct** statement.
- Since the hybridization of central atom in the reactant, i.e., **TeX_4 (sp^3d hybridized)** and product, i.e., **$\text{TeX}_4\text{X}'_2$ (sp^3d^2 hybridized)** are different, this is an **incorrect** statement.

Therefore, the incorrect statements are (a), (b) and (d).

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Resonance

Need of the theory

Lewis dot structure of the ozone (O_3) molecule can be represented as follows:

Single bonds are generally longer than double bonds. $O-O$ and $O=O$ bond lengths are generally observed to be equal to 148 pm and 121 pm , respectively. However, when experimental observation was made, both the bonds were found to be equal and both the lengths were between 148 pm and 121 pm , i.e., neither the single bond has 100% single bond character nor the double bond has 100% double bond character. Thus, when a **single Lewis structure cannot represent a molecule, resonance is needed** to explain the structure.

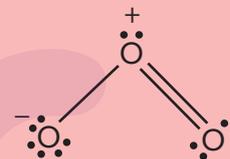


Fig. 14: Lewis structure of the ozone molecule

Definition

Resonance is the **phenomenon of delocalization of electrons**. **Delocalization** means that the electron clouds are spread over several adjacent atoms (more than two), whereas localization means that the electron clouds are situated at particular fixed positions. According to the concept of resonance, whenever a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energies, positions of nuclei and bonding and non-bonding pairs of electrons are taken as the canonical or contributing structures of the hybrid, which describes the molecule accurately.

Generally, pi bonds are weaker as compared to sigma bonds. Hence, their electron density can easily be delocalized between several atoms when an atom with an empty atomic orbital, pi bond or lone pairs are present next to it.

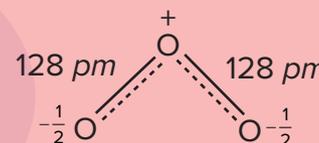


Fig. 15: Resonance hybrid of the ozone molecule

Molecules having resonance

Resonance structures (R.S.)

Resonance hybrid (R.H.)

Resonance structures or canonical structures are those hypothetical Lewis dot structures in which electrons are localized. **Resonance hybrid** is the one with delocalized electron pairs, i.e., the actual structure in the form of which the molecule exists.

Resonance structures or canonical structures are used, as the number of pi bonds or bond orders are difficult to determine in the resonance hybrid. Resonance hybrids have bond orders and bond lengths between these canonical structures.



Note

Resonance hybrids are always more stable than individual resonance structures.

BOARDS

Characteristics of Resonance Structures

1. All the resonating structures of a molecule are similar in energy.
2. Resonance structures have **identical positions of nuclei, i.e., same connectivity**. For example: The following two structures can be considered as resonance structures, since each atom is connected to the same atoms in both the structures.

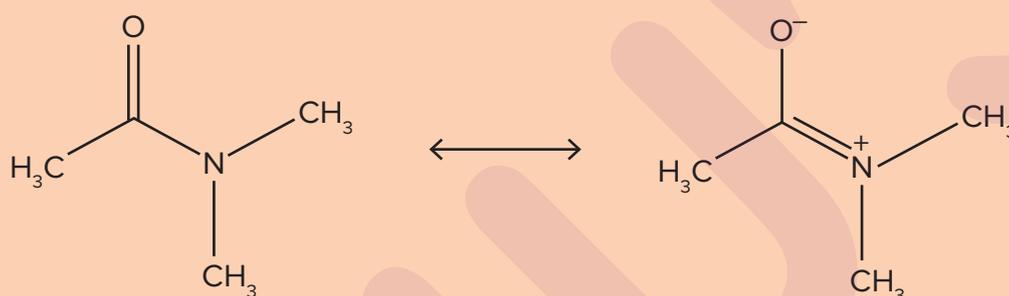


Fig. 16: Resonance structures of N,N-dimethylacetamide

3. Resonance structures have the **same number of bonding (sigma + pi bonds) and non-bonding (lone pairs) electron pairs**.

For example, in resonance structures of O_3 , there are three bond pairs and six lone pairs in both the structures.

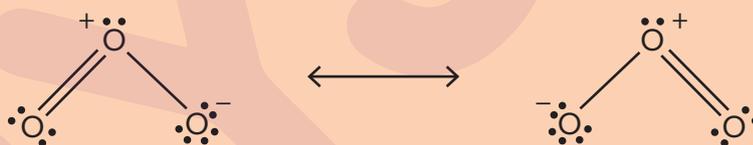


Fig. 17: Resonance structures of ozone

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Characteristics of Resonance Hybrid

1. The resonance hybrid is the **actual structure** obtained by the **combination of different possible Lewis structures** (resonance structures).
2. **More stable resonance structures** contribute more towards the resonance hybrid.
3. A resonance hybrid **does not violate the rules of covalence maxima**.

For example, the resonance hybrid of O_3 can be drawn as follows:

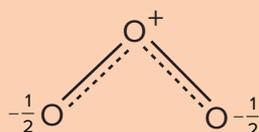


Fig. 18: Resonance hybrid of the ozone molecule

Characteristics of Resonance

1. Resonance structures are **hypothetical**. It is the resonance hybrid that exists in **reality**.
2. A resonance hybrid has its **individual identity**.
3. The **bond lengths** in resonance hybrids are **intermediate** to those of resonance structures.

Rules for Resonance Structures

1. The resonance structure should have a **proper or a valid Lewis structure**.
2. In the resonance structure, **only pi electrons or lone pairs can be moved**.
3. The **overall charge** on the structure must remain the **same in all the structures**.
4. The **bonding framework** must remain **intact**, i.e., all the structures should have the same connectivity. In simpler words, atoms cannot move to different locations in different resonance structures.

For example, the following (I) and (II) structures of nitromethane are not resonance structures, as they have different connectivity. In structure (I), N is bonded to two O atoms as well as the C of CH_3 , while in structure (II), N is only bonded to two O-atoms and not to the C of CH_3 .

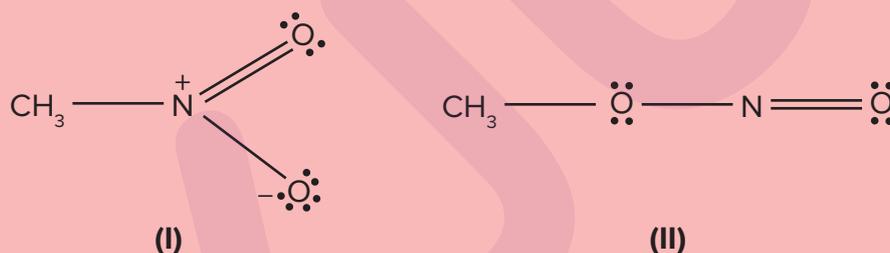


Fig. 19: Possible Lewis structure of nitromethane

Some other examples of resonance structures are as follows:

Resonance in the carbonate ion

The resonance structures of the carbonate ion (CO_3^{2-}) can be written as follows:

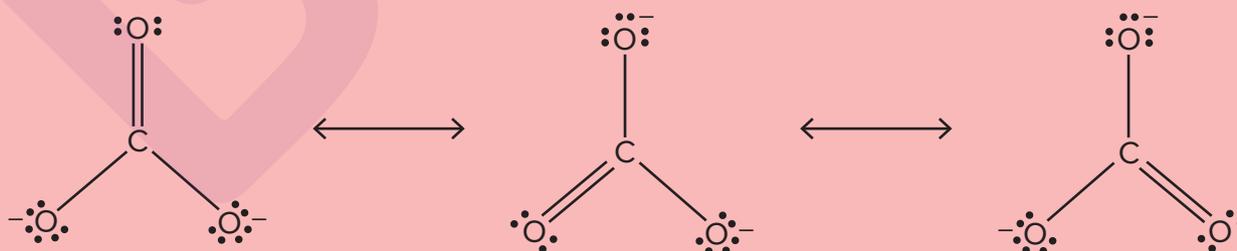


Fig. 20: Resonance structures of carbonate ion

All three structures have the same connectivity of atoms. Also, all atoms have the same charge, and one pi bond is present in each structure. Hence, all the resonating structures of the carbonate ion are equivalent or degenerate. The resonance hybrid of the carbonate ion can be shown as follows:

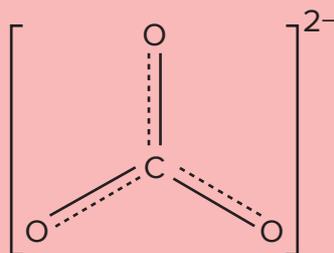


Fig. 21: Resonance hybrid of the carbonate ion

Resonance in the acetate ion

The resonance structures of the acetate ion (CH_3COO^-) can be written as follows:

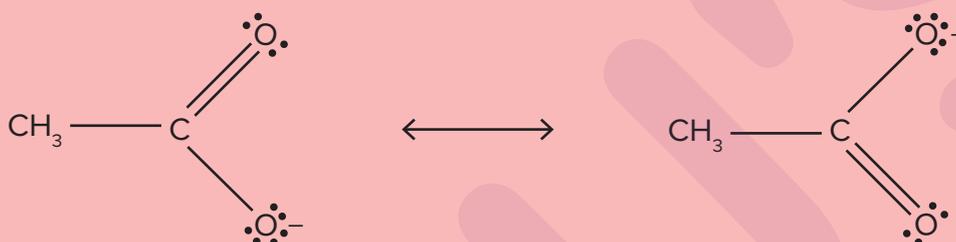


Fig. 22: Resonance structures of acetate ion

Both the structures have the same connectivity of atoms. Also, all the atoms have the same charge, and one pi bond is present in each structure. Hence, both the resonating structures of the acetate ion are equivalent or degenerate. The resonance hybrid of the acetate ion can be shown as:

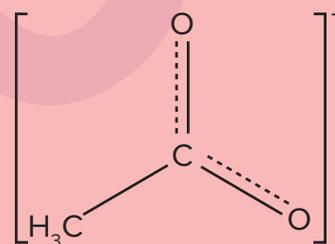


Fig. 23: Resonance hybrid of the acetate ion

Resonance in benzene

In the ground state of carbon, there are two unpaired electrons. However, since the covalency of carbon is **four**, **excitation of an electron** from s to p-orbital is **needed**. In the excited state of carbon, there are four unpaired electrons. Thus, it can form four bonds. In benzene, each carbon is connected to two adjacent carbon atoms and one hydrogen atom by sigma bonds. Hence, three (one s and two p-orbitals) orbitals undergo mixing to form **three sp^2 hybridized** orbitals, and the third p-orbital having one unpaired electron is used for pi bond formation.

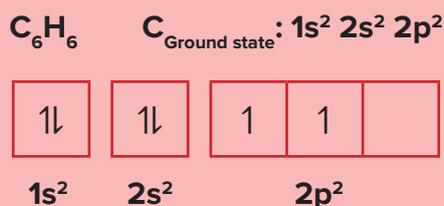


Fig. 24: Ground state electronic configuration of carbon

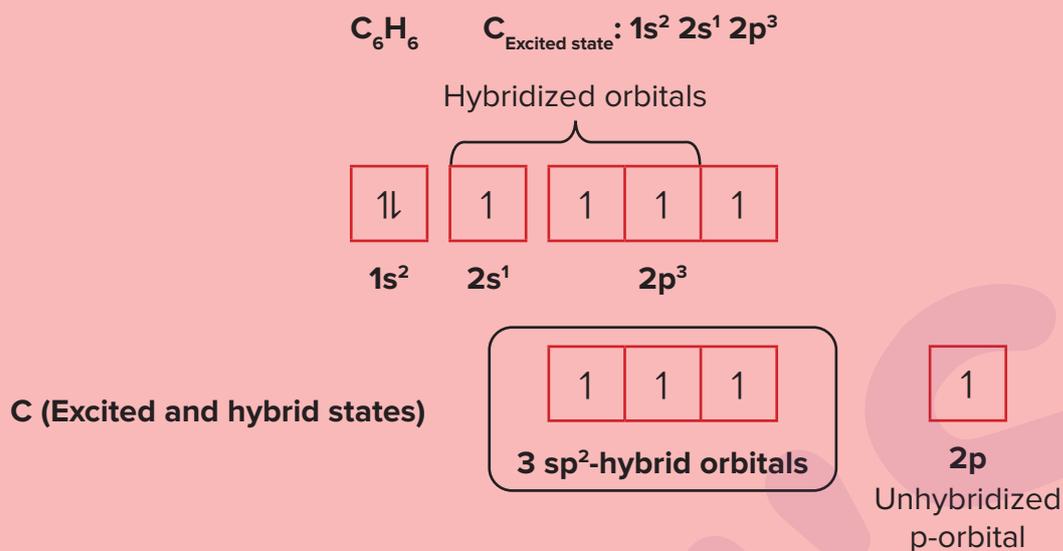


Fig. 25: Excited state electronic configuration of carbon in benzene

The resonance structures of benzene (C_6H_6) can be written as follows:

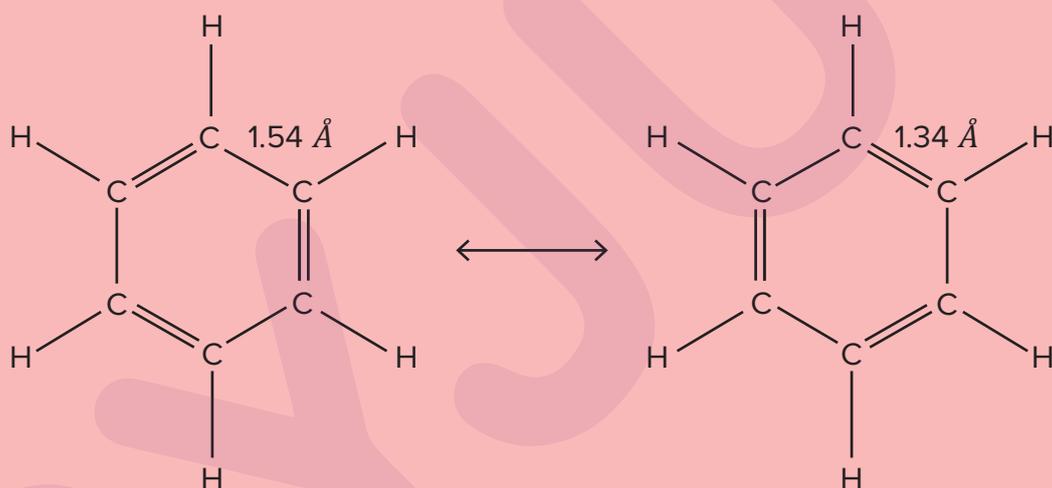


Fig. 26: Resonance structures of benzene molecule

There are two types of C-C bonds that are expected, one with the bond length equivalent to the single C-C bond (1.54 \AA) and another with the bond length equal to the double C-C bond (1.34 \AA). However, experimentally, each C-C bond length is found to be equal to 1.39 \AA . This can be easily explained by the delocalization of the pi electron clouds. Generally, we have two resonating structures of benzene, which are equivalent or degenerate.

The resonance hybrid of benzene can be shown as:

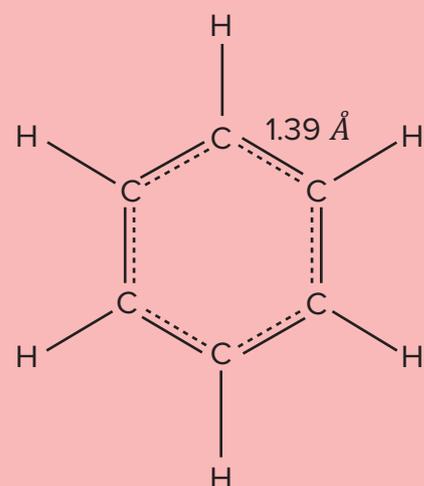


Fig. 27: Resonance hybrid of the benzene molecule

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Bond Order in Resonating Structures

For covalent bonds, the bond order is the **number of electron pairs shared between two atoms**. For localized bonds, it can be simply predicted using Lewis structures. For resonance structures, the bond order is calculated by using the following formula:

Bond order (B.O.)
between two atoms

=

Total number of
bonds between two
atoms in all structures

÷

Total number of
resonance structures

Examples:

Bond order of C–O in CO_3^{2-}

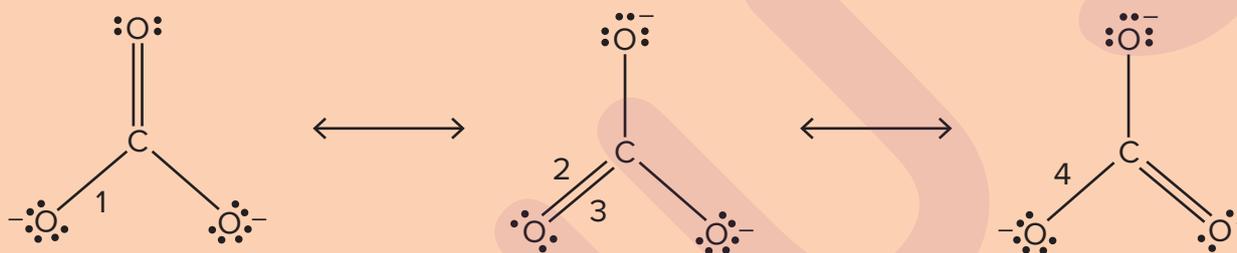


Fig. 28: Resonance structures of CO_3^{2-}

Total number of bonds between two atoms in all resonating structures = 4

Number of resonance structures = 3

Hence, bond order of C–O = $\frac{4}{3} = 1.33$

Bond order of O–O in O_3

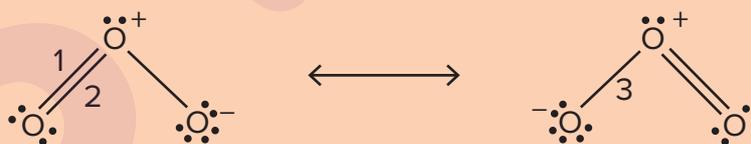


Fig. 29: Resonance structures of ozone

Total number of bonds between two atoms in all resonating structures = 3

Number of resonance structures = 2

Hence, bond order of O–O = $\frac{3}{2} = 1.5$

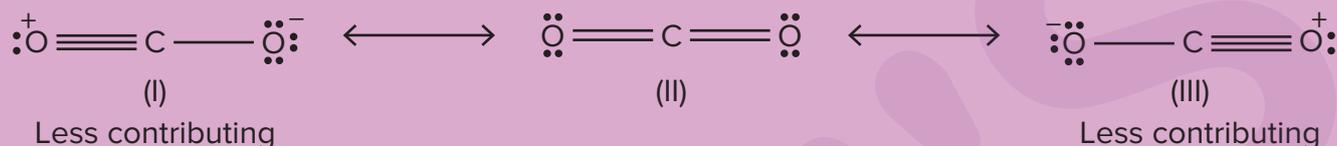


Based on the bond order in resonance structures

The experimentally determined **carbon to oxygen bond length in CO_2 is 115 pm**. The lengths of a **normal carbon to oxygen double bond ($\text{C}=\text{O}$)** and a **carbon to oxygen triple bond ($\text{C}\equiv\text{O}$)** are **121 pm and 110 pm**, respectively. Explain.

Solution

CO_2 has three resonating structures as shown in the following:



The resonance hybrid will have contributions from all three resonating structures, as there are triple, double, and single bonds in the resonating structures for both the C-O bonds. Thus, the overall bond order remains 2 for each bond. However, due to the contribution of (I) and (III) structures, the % s-character increases, as in the (I) and (III) structures, triple-bonded O-atoms are sp hybridized, whereas in (II), O-atoms are sp^2 hybridized. Hence, due to the **increase in the % s-character**, the bond **length decreases** and is **less** than that of a regular **C=O bond**.



Note

Resonance is one of the most important parameters to explain the stability of a certain molecule.

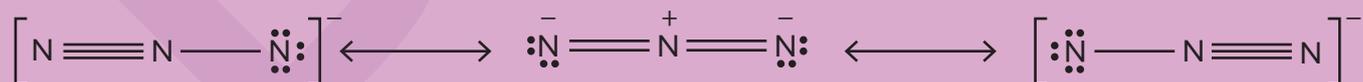


Based on bond order

The **numbers of types of N-N bond lengths** in azide ion and hydrazoic acid are p and q , respectively. Find the **value of $p + q$** .

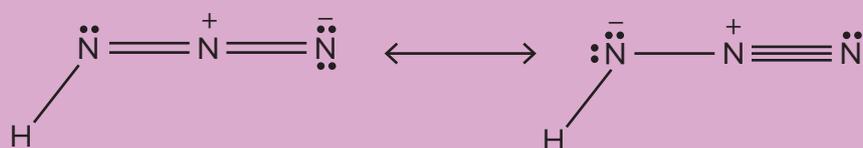
Solution

Azide ion (N_3^-) has three resonance structures as shown in the following:



Due to resonance, both the bonds have an identical bond order, i.e., 2. Thus, there is only one type of bond, i.e., $p = 1$.

Hydrazoic acid has two resonance structures as shown in the following:



There are two types of N-N bonds, one with bond order 1.5 and other with bond order 2.5. Hence, $q = 2$

Therefore, the value of $p + q = 1 + 2 = 3$



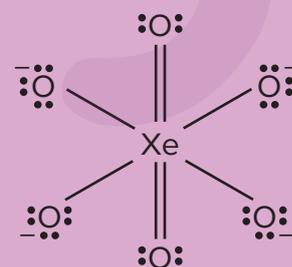
Based on bond order

Choose the correct statement(s) regarding the **perxenate (XeO_6^{4-}) ion**.

- The bond order of the Xe-O bond is 1.5.
- The bond order of the Xe-O bond is $8/6$.
- Total number of possible resonating structures = 15
- Total number of possible resonating structures = 10

Solution

The Lewis structure of the perxenate ion can be drawn as follows:



(a) We know,

Bond order = Total number of bonds in a resonating structure / Number of resonating structures = $20/15 = 4/3$

Hence, the bond order of the Xe-O bond is $8/6$. Thus, this is an incorrect option.

(b) Since the bond order of the Xe-O bond is $8/6$, this is a correct option.

(c) The number of resonating structures can be calculated by using permutation and combination. Since there are two pi bonds and six O-atoms, the number of resonating structures =

$${}^6C_2 = \frac{6!}{2! \times 4!} = \frac{6 \times 5 \times 4!}{2 \times 1 \times 4!} = 15$$

Hence, total number of resonating structures = 15

Thus, this is a correct option.

(d) As the total number of resonating structures = 15, this is an incorrect option.

Therefore, correct answers are options (b) and (c).



Summary

Molecules with steric number 7 show sp^3d^3 hybridization and pentagonal bipyramidal geometry.

- There are two types of positions in pentagonal bipyramidal geometry, i.e., **axial** and **equatorial** as in trigonal bipyramidal geometry. However, in this case, the angle between the equatorial-equatorial bonds (72°) is less than that of the axial-equatorial bond (90°). Hence, the repulsion in the electron pairs is more at equatorial positions than in axial positions. As a result, in this case, the **lone pair tends to occupy the axial positions**.
- Due to more repulsion, **equatorial bonds are longer** than axial bonds for pentagonal bipyramidal geometry.
- In the case of **pentagonal bipyramidal** geometry having six sigma bond pairs and one lone pair, the shape of such a molecule is that of a **distorted octahedral**.
- Molecules with five sigma bond pairs and two lone pairs have a pentagonal planar shape.
- Molecules with steric number 8 show sp^3d^4 hybridization and square antiprismatic geometry.
- The valence bond theory fails to explain the hybridization in odd electron species.

- Generally, orbitals with odd valence electrons **do participate in hybridization**. However, experimentally, **exceptions like alkyl radicals** ($\cdot\text{CR}_3$, where R = H, Me, etc.) are also found in which odd electron containing orbitals do not participate in hybridization.
- In the gaseous state, PCl_5 and PBr_5 have a trigonal bipyramidal geometry, which is not symmetrical. Therefore, in the solid states, PCl_5 dissociates into a cationic form $[\text{PCl}_4]^+$ and an anionic form $[\text{PCl}_6]^-$, whereas in solid the state, PBr_5 dissociates into a cationic form $[\text{PBr}_4]^+$ and an anion Br^- . **The hybridizations of $[\text{PCl}_4]^+$ and $[\text{PCl}_6]^-$ are sp^3 and sp^3d^2** , respectively, and they possess tetrahedral and octahedral shapes, which are **more symmetrical** than trigonal bipyramidal geometry. The hybridizations of $[\text{PBr}_4]^+$ is sp^3 , and the shape is tetrahedral, which is more symmetrical than the trigonal bipyramidal and is therefore more stable. Also, due to the large size of the Br-atom, phosphorus could not accommodate six Br-atoms. Hence, $[\text{PBr}_6]^-$ does not exist.
- **PI_5 does not exist**, since phosphorus could not accommodate five large iodine atoms around itself.
- Resonance is the **phenomenon of delocalization of electrons**. **Delocalization** means that the electron clouds are spread over several adjacent atoms (more than two), whereas **localization** means that the electron clouds are between two atoms or at particular positions around an atom. According to the concept of resonance, whenever a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energy, positions of nuclei, and bonding and non-bonding pairs of electrons are taken, as the canonical or contributing structures of the hybrid which describes the molecule accurately.
- **Resonance structures or canonical structures** are hypothetical Lewis dot structures in which electrons are localized. Resonance hybrid is the one with delocalized electron pairs, i.e., the actual structure in which the molecule exists.
- The characteristics for resonance structures are as follows:
 1. Resonance structures have similar or degenerate energy.
 2. Resonance structures have identical positions of nuclei, i.e., the same connectivity.
 3. Resonance structures have the same number of bonding (sigma + pi bonds) and non-bonding (lone pairs) electron pairs.
- More stable resonance structures contribute more towards the resonance hybrid.
- A resonance hybrid does not violate the rules of covalence maxima.
- For a valid resonating structure, it must follow the given rules:
 1. All the resonating structures of a molecule are similar in energy.
 2. In the resonance structure, only pi electrons or lone pairs can be moved.
 3. The overall charge on the structure must remain the same in all the structures.
 4. The bonding framework must remain intact, i.e., all the structures should have the same connectivity. In simpler words, atoms cannot move in resonance structures.
- For covalent bonds, the bond order is the **number of electron pairs shared between two atoms**.

The bond order of resonating structures

B.O. between two atoms

=

Total number of **bonds** between two atoms in all structures

÷

Total number of **resonating structures**