

[Click Here](#)





NCERT Solutions for Class 11 Chemistry Chapter 4: Chemical Bonding and Molecular Structure available online for CBSE students. Comprehensive answers can be found here, providing in-depth guidance for the board exam. BYJU'S has crafted solutions by experienced faculty to aid students preparing for the exam. The Hybridisation of Elements Involving D Orbitals - Molecular Orbital Theory Molecular Orbital Theory explains the formation of molecular orbitals through Linear Combination of Atomic Orbitals (LCAO). This theory involves the combination of atomic orbitals to form molecular orbitals, which are crucial in understanding chemical bonding. The Conditions for the Combination of Atomic Orbitals include: \* The atoms involved must have overlapping orbitals \* The orbitals must be oriented in a way that allows for effective overlap Types of Molecular Orbitals include: \* Sigma ( $\sigma$ ) bonds: formed by end-to-end overlap of atomic orbitals \* Pi ( $\pi$ ) bonds: formed by side-by-side overlap of atomic orbitals An Energy Level Diagram for Molecular Orbitals shows the relative energies of different molecular orbitals. The Electronic Configuration and Molecular Behaviour are crucial in understanding chemical bonding. The formation of a chemical bond is attributed to the tendency of a system to achieve stability, which leads to the combination of atoms to form molecules. Hydrogen Bonding is a type of intermolecular force that arises between molecules with a hydrogen atom bonded to a highly electronegative atom, such as oxygen, nitrogen, or fluorine. The cause of formation of Hydrogen Bond is the partial positive charge on the hydrogen atom and the partial negative charge on the electronegative atom. Types of H-bonds include: \* Hydrogen bonding between a hydrogen atom bonded to one molecule and an electronegative atom bonded to another molecule \* Hydrogen bonding between two molecules with hydrogen atoms bonded to electronegative atoms The formation of chemical bonds is crucial in understanding chemical reactions and properties. the Lewis dot symbol for H is  $\cdot\text{H}\cdot$ . The single negative charge on hydrogen indicates that it has gained 1 electron. So, one valence electron plus one gained electron. Thus lewis dot symbol is  $\text{H}^-$  4) Draw the Lewis structures for the following molecules and ions :  $\text{H}_2\text{S}$ ,  $\text{SiCl}_4$ ,  $\text{BeF}_2$ ,  $\text{CO}_3^{2-}$ ,  $\text{HCOOH}$  Ans.)  $\text{H}_2\text{S}$   $\text{SiCl}_4$   $\text{BeF}_2$   $\text{CO}_3^{2-}$   $\text{HCOOH}$  Q-5) Define the octet rule. Write its significance and limitations. Ans.) Octet rule says, "Atoms can combine either by transfer of valence electrons from one atom to another or by sharing their valence electrons in order to achieve the nearest inert gas configuration by having an octet in their valence shell." Octet rule explains chemical bond formation depending upon the nature of the element. Limitations: (a) Octet rule fails to predict the relative stability and shape of the molecules. (b) It is based on the inert nature of noble gases. But, some inert gases, say krypton (Kr) and xenon (Xe), form compounds like  $\text{KrF}_2$ ,  $\text{XeF}_2$  etc. (c) For elements beyond the 3rd period, the octet rule cannot be applied. Elements present beyond the 3rd period have more than 8 valence electrons surrounding the central atom. E.g.  $\text{SF}_6$ ,  $\text{PF}_6$  etc. (d) For atoms in a molecule having an odd number of electrons, the octet rule is not applied. E.g. For  $\text{NO}_2$  and  $\text{NO}$  octet rule is not applicable. (e) If a compound has less than 8 electrons surrounding the central atom, then the octet rule cannot be applied to that compound. E.g.  $\text{BeH}_2$ ,  $\text{AlCl}_3$ ,  $\text{LiCl}$  etc., does not obey the octet rule. Q-6) Write the favourable factors for the formation of an ionic bond. Ans.) The formation of an ionic bond takes place through the transfer of 1 or more electrons from one atom to another. Thus, ionic bond formation depends on the flexibility of neutral atoms to lose or gain electrons. The formation of an ionic bond also depends on the lattice energy of the compound which is formed. The factors that are favourable for ionic bond formation: (a) High electron affinity of atoms of non-metals. (b) The high lattice energy of the compound which is formed. (c) Low ionization enthalpy of an atom of metal. Q-7) Discuss the shape of the following molecules using the VSEPR model:  $\text{BeCl}_2$ ,  $\text{BCl}_3$ ,  $\text{SiCl}_4$ ,  $\text{AsF}_5$ ,  $\text{H}_2\text{S}$ ,  $\text{PH}_3$  Ans.)  $\text{BeCl}_2$  The central atom does not have any lone pair but has 2 bond pairs. Thus, its shape is  $\text{AB}_2$ , i.e., linear shape.  $\text{BCl}_3$  The central atom does not have any lone pair but has 3 bond pairs. Thus, its shape is  $\text{AB}_3$ , i.e., trigonal planar.  $\text{SiCl}_4$  The central atom does not have any lone pair but has 4 bond pairs. Thus, its shape is  $\text{AB}_4$ , i.e., tetrahedral.  $\text{AsF}_5$  The central atom does not have any lone pair but has 5 bond pairs. Thus, its shape is  $\text{AB}_5$ , i.e., trigonal bipyramidal.  $\text{H}_2\text{S}$  The central atom has 1 lone pair and has 2 bond pairs. Thus, its shape is  $\text{AB}_2\text{E}$ , i.e., bent shape.  $\text{PH}_3$  The central atom has 1 lone pair and has 3 bond pairs. Thus, its shape is  $\text{AB}_3\text{E}$ . The geometries of  $\text{NH}_3$  and  $\text{H}_2\text{O}$  molecules are distorted tetrahedral, but the bond angle in water is less than that of ammonia due to the stronger repulsion between lone pairs and bond pairs on the oxygen atom in the water molecule. This occurs because the central oxygen atom has two lone pairs and two bond pairs, causing a greater repulsion compared to the nitrogen atom in ammonia which has one lone pair and three bond pairs. Bond strength can be expressed in terms of bond order, where an increase in bond order indicates stronger bonds and increased bonding between atoms. The equilibrium distance between the nuclei of two bonded atoms is known as the bond length. This concept is essential for understanding molecular structures and interactions. Resonance plays a crucial role in determining the structure of certain molecules, such as  $\text{CO}_3^{2-}$  ion, where experimental results show that all C-O bonds are equivalent. To represent this ion accurately, it's more efficient to use resonance structures rather than a single Lewis structure with one double bond and two single bonds. In the case of  $\text{H}_3\text{PO}_3$ , represented by structures 1 and 2, these cannot be considered as canonical forms of the resonance hybrid due to the rearrangement of atoms in each structure. Resonance structures for  $\text{SO}_3$ ,  $\text{NO}_2$ , and  $\text{NO}_3^-$  can be written as follows:  $\text{SO}_3$ :  $\text{NO}_2$ :  $\text{NO}_3^-$  The Lewis symbols show electron transfer between atoms to form cations and anions in the following cases: (i) K and S: The electronic configuration of S has 2 more electrons than Ne, while K has 1 less electron. Therefore, the transfer occurs as follows. (ii) Ca and O: The electronic configuration of O has 2 less electrons than Ar, while Ca has 2 extra electrons. Thus, the transfer takes place in the following manner. (iii) Al and N: The electronic configuration of N has 3 more electrons than Ne, while Al has 3 less electrons. Hence, the transfer occurs as follows. The bond angle difference between  $\text{H}_2\text{O}$  and  $\text{NH}_3$  is due to the stronger repulsion in water molecules caused by their structure. Molecules like  $\text{H}_2\text{O}$  and  $\text{CO}_2$  are triatomic and their shapes differ significantly.  $\text{H}_2\text{O}$  has a bent shape while  $\text{CO}_2$  is linear. This distinction can be explained by considering the dipole moment of each molecule. The dipole moment of  $\text{CO}_2$  was found to be 0, indicating that its shape is linear. This occurs because the bond between C-O is equal and opposite, thus nullifying each other's effects. In contrast,  $\text{H}_2\text{O}$  has a dipole moment of 1.84 D, suggesting that its structure is bent. The unequal bond moments between O-H contribute to this non-linear shape. Understanding dipole moments is crucial for several reasons. \* It helps determine the shape of a molecule, with symmetrical shapes like linear having zero dipole moments and non-symmetrical ones displaying varied shapes. \* It indicates the polarity of molecules, with higher dipole moments indicating greater polarity. A molecule with a zero dipole moment is considered non-polar, while one with some polar character has a non-zero dipole moment. The concept of electronegativity refers to an atom's ability to attract shared electrons in a chemical compound. It varies based on the element it is bonded to and is not constant for any single element. In contrast, electron gain enthalpy measures a isolated gaseous atom's tendency to acquire additional electrons. Given text here is, the higher the ionic character of a molecule will be. So, the required order of ionic character of the given molecules is  $\text{N}_2 < \text{SO}_2 < \text{ClF}_3 < \text{K}_2\text{O}$