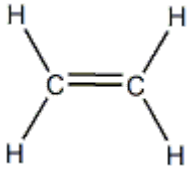
Solution

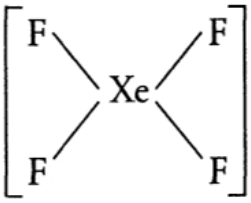
CHEMICAL BONDING

Class 11 - Chemistry

1. This is because ethyl alcohol forms H-bonds with water. The -OH group of alcohols is responsible for H-bonding.
2. The oxygen is more electronegative than sulfur. Hence, the high electronegativity of O atom pulls the bonding electron towards it while in H_2S , it is away from S atom. Thus bond pairs repulsion is more in H_2O than H_2S. So, H-O-H bond angle in H_2O is higher than H-S-H bond angle in H_2S.
3. Electrovalent or ionic bond is formed whenever there is high electronegativity difference (usually greater than 1.7) between combing atoms.
4. According to the valence bond theory, a covalent bond is formed due to the overlapping of half-filled atomic orbitals present in the valence shell of the atoms participating in bonding.
5. Bond order is defined as number of bonds between two atoms in the Lewis representation of a molecule or ion.
6. In ethene, there are 5 σ bonds (1 between C - C, and 4 between C - H) and 1 - π bond. The structure of ethene is



7. The attractive force which holds various constituents (atoms, ions etc.) together in different chemical species is called a chemical bond.
8. The electrons placed in bonding molecular orbitals tend to hold the nuclei together and hence greater stability than the corresponding antibonding molecular orbital where the mutual repulsion of the electron is more than attraction.
9. XeF_4 = 4 bp + 2 lp (This is because Xe contains 8 valence electrons) = square planar geometry.



10. The molecules given have symmetrical shapes (e.g. CO_2 is linear, BF_3 is trigonal planar and CCl_4 is tetrahedral) and thus the dipoles gets cancelled and the net dipole moment is zero.
11. Dipole moment, μ = q $\times$ r

    $\mu = 0.54 \Rightarrow D = 0.54 \times 10^{-18}$ state C cm

    $\delta = 1.14 \times 10^{-8}$ cm

    $\therefore q = \frac{0.54\times10^{-18}\text{ stat Ccm}}{1.41\times10^{-8}\text{ cm}} = 0.38 \times 10^{-10}$ stat

    Now, fraction of charge = $\frac{\text{Change present}}{\text{Electronic charge}}$

    $= \frac{0.38\times10^{-10}}{4.8\times10^{-10}} = 0.08$

    $\therefore \delta_A = 0.08, \delta_B = -0.08$
12. i. In predicting the nature of the molecules: Molecules with specific dipole moments are polar in nature and those of zero dipole moments are non-polar in nature.
    ii. In the determination of shapes of molecules.
    iii. In calculating the percentage ionic character.
13. The VSEPR theory is used to determine the geometry of the individual molecules. It is developed by Gillespie-Nyholm.The main postulates of valence shell Electron pair repulsion (VSEPR) theory are:-

    (i) The shape of a molecule depends upon the no. of electron pairs around the central atom.

    (ii) There is a repulsive force between the electron pairs, which tend to repel one another

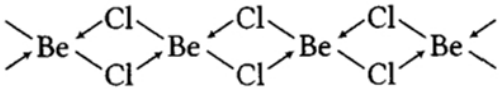
    (iii) The electron pairs in space tend to occupy such positions that they are at a maximum distance so, that the repulsive force will be minimum.

    (iv) A multiple bond is treated as if it is single bond and the remaining electron pairs which constitute the bond may be regarded as a single super pair.

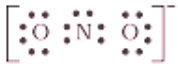
14. When two atoms with different electro-negativity are linked to each other by a covalent bond, the shared pair of electrons will not in the center because of the difference in electro-negativity. For example, in hydrogen fluoride.
molecule, fluorine has greater electronegativity than hydrogen. Thus the shared pair of electrons is displaced more towards fluorine atom, and acquire a partial negative charge (δ^-). At the same time, the hydrogen atom will have a partial positive charge (δ^+). Such a covalent bond is known as polar covalent bond or simply polar bond.
lt is represented as:
δ^+ δ^-
H - F
(2,1) (4,0)
15. In gaseous state at high temperature, $BeCl_2$ exists as linear molecule, Cl - Be - Cl, thus the hybridisation of the central atoms is sp.

$CI - Be - CI$
Structure of BeCI in gaseous state

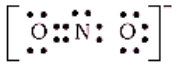
In solid state, it has a polymeric structure with chlorine bridges as follows.



Two Cl-atoms are listed to be atom by two coordination bonns and two by covalent bonds.
For these bonds to be formed, Be in the excited state with the configuration $1s^2 2s^1 2p_x^1 2p_y^0 2p_z^0$
undergoes sp^3 hybridisation.
Two half-filled hybrid orbitals will form normal covalent bonds with two Cl-atom. The other two Cl-atoms are coordinated to Be-atom by donating electron pairs into the empty hybrid orbitals.
16. **Step 1.** Count the total number of valence electrons of the nitrogen atom, the oxygen atoms and the additional one negative charge (equal to one electron).
N($2s^2$ $2p^3$), O ($2s^2$ $2p^4$)
5 + (2 × 6) +1 = 18 electrons
Step 2. The skeletal structure of NO_2^- is written as: O N O
Step 3. Draw a single bond (one shared electron pair) between the nitrogen and each of the oxygen atoms completing the octets on oxygen atoms. This, however, does not complete the octet on nitrogen if the remaining two electrons constitute lone pair on it.



Hence we have to resort to multiple bonding between nitrogen and one of the oxygen atoms (in this case a double bond). This leads to the following Lewis dot structures.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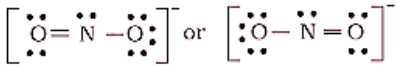


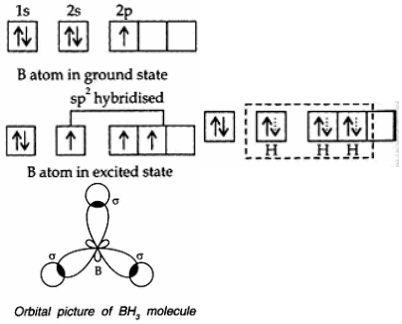
17. i. The factors for the formation of an ionic bond are:
    a. Ionization energy: It is the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom. Thus, lesser the ionization enthalpy required, easier will be the formation of a cation. e.g.Generally metals (especially alkali metals and alkaline earth metals) form cations easily due to low ionization energy.
    b. Electron gain enthalpy: It is the amount of energy released on the addition of an electron to an isolated gaseous neutral atom. Greater the negative electron gain enthalpy, easier will be the formation of an anion. e.g.Generally non-metals (specially Halogens) form anion easily due to high electron gain enthalpy.
    c. High lattice enthalpy: It is the amount of energy released when 1 mole of a crystal is formed from the gaseous state. The greater the magnitude of lattice energy greater will be the stability of the ionic bond.

    ii. Lattice energy depends on
    (a) Inter-ionic distance: As the inter-ionic distance decrease, lattice energy increases.
    (b) Product of ionic charges: As the product of ionic charges increases, lattice energy increases.
    In MgO and $MgCl_2$, the product of the magnitude of ionic charges is 4 and 2 respectively. Hence MgO has higher lattice energy than $MgCl_2$. Again the size of Mg^{2+} is smaller than Na^+ and therefore $MgCl_2$ has higher lattice energy than NaCl.
    Similarly, the product of the magnitude of ionic charges of $CaCl_2$ and NaCl is 2 and 1 respectively. So $CaCl_2$ has higher lattice

energy than NaCl. The size of Ca^{2+} is greater than Mg^{2+} and hence the lattice energy of $CaCl_2$ is lower than $MgCl_2$.Therefore the order of lattice energy is

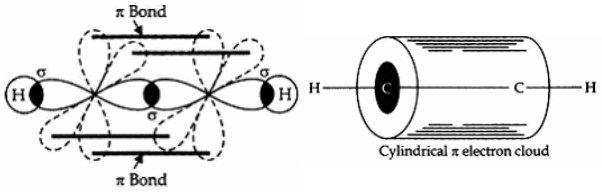
$MgO > MgCl_2 > CaCl_2 > NaCl$

18. i. Formation of BH_3 (atomic no. of B is 5)



B atom gets hybridized to form three equivalent hybrid orbitals directed towards three corners of an equivalent triangle with B atoms in the center. Bond angle = 120°.

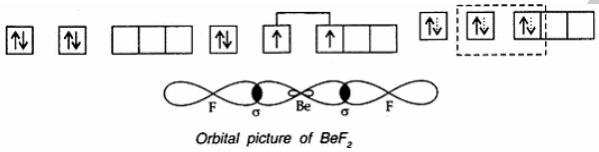
ii. C_2H_2



Orbitals picture of ethyne

Both the carbon atoms are sp hybridized. Both the carbon atoms have also two unhybridized orbitals which overlap sidewise with the similar orbitals of the other carbon atom to form two π bonds.

iii. BeF_2



The orbital picture of BeF_2 molecule is Linear.

19. The atomic orbitals mix together to generate a new set of equivalent orbitals, called the **hybrid orbitals** or **hybridised orbitals**. These orbitals are used in bond formation. The phenomenon is called **hybridisation**.

Types of hybridization in carbon atom.

i. a. **Diagonal or sp-hybridisation**- All compounds of carbon containing C ≡ C (double bond) like ethene (C_2H_2).

b. **Trigonal or sp^2-hybridisation**- All compounds of carbon containing C = C (double bond) like ethene (C_2H_4).

c. **Tetrahedral or sp^3-hybridisation**- All compounds of carbon containing C-C single bonds only like ethane (C_2H_6).

ii. a. $\overset{*}{\underset{sp^2}{C}}H_2 = CH - \overset{O}{\overset{\|}{\underset{sp^2}{C^*}}} - O - H$;

b. $CH_3 - \overset{*}{\underset{sp^3}{C}}H_2 - OH$;

c. $CH_3 - CH_2 - \overset{O}{\overset{\|}{\underset{sp^2}{C^*}}} - H$;

d. $\overset{*}{\underset{sp^3}{C}}H_3 = CH = CH - CH_3$

e. $CH_3 - \overset{*}{\underset{sp}{C}} \equiv CH$