

$$\therefore 3.336 \times 10^{-29} = \delta \times 2.6 \times 10^{-10}$$

$$\therefore \delta = \frac{3.336 \times 10^{-29}}{2.6 \times 10^{-10}} = 1.283 \times 10^{-19} \text{ Coulomb}$$

$$\% \text{ ionic character} = \frac{1.283 \times 10^{-19}}{1.602 \times 10^{-19}} \times 100 = 80.09\%$$

If one unit charge, then $\delta = 1.602 \times 10^{-19} \text{ C}$

$$\therefore \mu = 1.602 \times 10^{-19} \times 2.6 \times 10^{-10} \\ = 4.1652 \times 10^{-29} \text{ Coulomb meter}$$

COORDINATE BOND

- An electron pair bond where both electrons originate from one atom and none from the other are called coordinate bonds or dative bonds. Since in coordinate bonds, two electrons are shared by two atoms, they differ from normal covalent bonds only in the way they are formed, and once formed they are identical to normal covalent bonds.



Example :

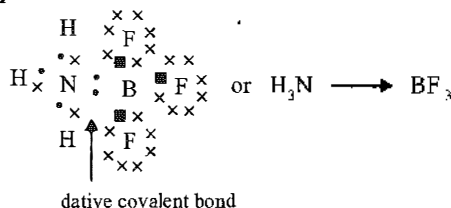
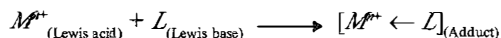


Fig : The dative covalent bond in H_3NBF_3 .

- Lewis acid-base concept of coordinate bond :** In coordination compounds, the metal ion (M^{n+}) acts as an acceptor (Lewis acid) (metal ion has an empty orbital and hence accepts an electron pair) and the ligand (L) acts as a donor (Lewis base) (ligand has a lone pair of electrons in it). The formation of a coordinate bond, $M^{n+} \leftarrow L$, in a coordination compound or adduct can be represented as :



Hence coordinate bond is also called donor-acceptor bond.

- Characteristics of coordinate compounds**
 - They exist as gases, liquids and solids.
 - Melting point and boiling point : Higher than covalent compounds and lower than ionic compounds.
 - Soluble in non-polar solvents and insoluble in polar solvents.
 - Behave as non-conductors of electricity.
 - Undergo molecular reactions.

Bond Parameters

- Bond length :** Distance between centres of bonded atoms is called bond length or bond distance.

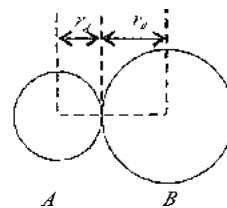


Fig : The bond length in a covalent molecule AB

- Bond length of A_2 type molecule
 $d_{A-A} = 2r_A$ [r_A = covalent radius of atom A]
- Bond length of AB type molecule
 $d_{A-B} = r_A + r_B - 0.09(X_A - X_B)$ [when r_A and r_B are in \AA]
- With the increase in the percentage of s -character in a hybrid orbital, bond length decreases, since s -orbital is smaller than p -orbital.

Molecule	Type of hybridisation	% of s -character hybrid orbitals	C-C bond length (pm)	C-H bond length (pm)
Ethane ($\text{H}_3\text{C}-\text{CH}_3$)	sp^3	25	154	109.3
Ethylene ($\text{H}_2\text{C}=\text{CH}_2$)	sp^2	33.3	134	108.7
Acetylene ($\text{HC}\equiv\text{CH}$)	sp	50	120	105.7

- For a given atom, the bond length increases with the size of the other atom bonded to it, (e.g., for hydrogen halides, bond length increases as $\text{HF} < \text{HCl} < \text{HBr} < \text{HI}$).
- Bond energy :** The amount of energy required to break a bond is called bond dissociation energy or bond energy. $A + B = AB + \text{Energy}$. The energy term corresponds to the energy released or bond energy.
 - Larger the bond energy, stronger is the bond.
 - Greater the electronegativity difference, greater is the bond polarity, greater is the bond energy, e.g., $\text{H}-\text{F} > \text{H}-\text{Cl} > \text{H}-\text{Br} > \text{H}-\text{I}$
 - More directional is a bond, higher is the bond strength.
 $sp^3 - sp^3 > sp^2 - sp^2 > sp - sp > p - p > s - p > s - s$
- Bond angles :** The angle between the bonding electron pairs in a given molecule or ion is called bond angle.
 - As the s -character of the hybrid bond increases, the bond angle increases as
 $sp^3 < sp^2 < sp$
 $109^\circ 28' \quad 120^\circ \quad 180^\circ$
 - The $B-A-B$ bond angle decreases with the decrease of the electronegativity or with the increase of the size of the central atom A .
 $\text{NF}_3 > \text{PF}_3 > \text{AsF}_3 > \text{SbF}_3$
 $102^\circ \quad 97^\circ \quad 96^\circ \quad 88^\circ$