

Reactions of Amines and Epoxydes

Preparation of Amines

Ammonium hydroxide: $\text{NH}_4^+ + \text{OH}^- \rightarrow \text{NH}_3 + \text{H}_2\text{O}$



Four different types of bonds



There are four types of bonding between atoms in a molecule.

- Ionic bonds are formed by the transfer of electron from one atom to another (anions and cations) (opposite charges).



Regions (p) and pi (π) bonds

- Covalent bonds are classified as sigma (σ) and pi (π).
- Sigma bonds are formed by head-on overlap of orbitals.
- pi bonds are formed by side-on overlap of orbitals.
- Very simple bond is a covalent bond.
- Non-overlapping bond consists of two regions and two pi (π) bonds.



pi bonds are stronger than sigma.

Polar and Nonpolar Covalent Bonds

Depending on the difference in the electronegativity values, covalent bonds can be polar and nonpolar.

- If the difference in electronegativity is less than 0.4, the electrons are shared equally between the two atoms. Forming a nonpolar covalent bond (like a C-C and C-H bond).
- If the difference in electronegativity is between 0.5 and 1.7, one atom has a partial negative charge and the other has a partial positive charge (like C-Cl and C-F bonds).
- A difference of 1.7 or higher one or both the atoms have a larger share and a polar bond is formed.

A Summary of Ring-Opening Reactions of Epoxydes with Strong and Weak Nucleophiles

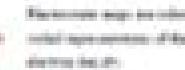
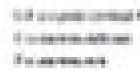


Reactions and Mechanisms

Bond Types, Strength, Length, Polarity and Dipole Moment

Dipole Moment

The polarity of a bond is denoted by the dipole moment (μ).



Electronegativity is inversely proportional to the dipole moment.

- The dipole moment is shown as a vector pointing from the less electronegative atom to the more electronegative atom.
- The vector is proportional to the partial negative charge and the partial positive charge.
- The dipole plot indicates the charge difference.
- The dipole moment is shown as a horizontal arrow.

Molecular Dipole

To measure the molecular polarity, the total sum of the dipole is used.

The polarity of the molecule is equal to the sum of individual bond dipole moments. If dipole moments are canceled, it just means a neutral dipole.



Conversely, the larger the sum of the dipole, the more polar.

Average values of C-C and C-H bond depending on hybridization			
C-C bond	$1.37 \text{ e}^- \cdot \text{Å}^{-1}$	$1.27 \text{ e}^- \cdot \text{Å}^{-1}$	$0.92 \text{ e}^- \cdot \text{Å}^{-1}$
Hybridization	sp^3 (dipole ²)	sp^2 (dipole ²)	sp (dipole)
Length	1.75 Å	1.35 Å	1.15 Å
Strength	0.15 eV	0.17 eV	0.19 eV
Hybridization			
sp^3	sp^2	sp	sp^3 (dipole ²)
Strength	0.15 eV	0.17 eV	0.19 eV
Length	1.75 Å	1.35 Å	1.15 Å
Hybridization	sp^3	sp^2	sp
Strength	0.15 eV	0.17 eV	0.19 eV
Length	1.75 Å	1.35 Å	1.15 Å

Dipole + dipole + dipole + dipole = dipole

Example



$\text{dipole}_1 + \text{dipole}_2 = \text{dipole}$



$\text{dipole}_1 + \text{dipole}_2 = \text{dipole}$



$\text{dipole}_1 + \text{dipole}_2 = \text{dipole}$



$\text{dipole}_1 + \text{dipole}_2 = \text{dipole}$

Bond Length and Bond Strength

Relative bond length



Length: $\text{C-H} > \text{C-C} > \text{C=O} > \text{C=O}$

Strength: $\text{C-H} < \text{C-C} < \text{C=O} < \text{C=O}$

Relative bond length: $\text{C-H} > \text{C-C} > \text{C=O} > \text{C=O}$

Relative bond strength: $\text{C=O} > \text{C=O} > \text{C-C} > \text{C-H}$

Relative bond length



Relative bond length: $\text{C-H} > \text{C-C} > \text{C=O} > \text{C=O}$

Relative bond strength: $\text{C=O} > \text{C=O} > \text{C-C} > \text{C-H}$

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