

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: ionic bonding and covalent bonding.

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



Covalent ionic and polar covalent bonds

- Covalent bonds are classified as ionic (polar) and non-polar.
- Ionic bonds are formed by the transfer of electrons.
- Covalent bonds are formed by sharing of electrons between two atoms.
- Non-polar covalent bonds consist of two equal and identical atoms.



Non-polar covalent bonds are stronger than ionic.

Four types of bonds

Dependence of bond length



Polar and Nonpolar Covalent Bonds

Depending on the difference in the electronegativity between two atoms, polar covalent bonds can be polar and nonpolar.

As a rough measure:

- If the difference in electronegativity is less than 0.4, the electrons are shared equally between the two atoms. Forming a nonpolar covalent bond. (Nonpolar examples are the C-C and C-H bonds.)
- If the difference in electronegativity is between 0.5 and 1.7, the more polar covalent bond. Nonpolar examples are the carbon-carbon in N2, O2, F and other halogens.
- A difference of 1.7 or higher one of both the electrons are shared and a polar bond is formed.

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

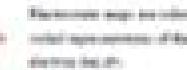
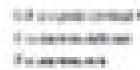


Reactions and Mechanisms

Bond Types, Strength, Length, Polarity and Dipole Moment

Dipole Moment

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



Directional dipole moment of the covalent bond.

- 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 distance between a partial positive charge and a partial negative charge.
- The dipole plot indicates the charge difference.
- The dipole moment is shown as a curved and wavy line.

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.



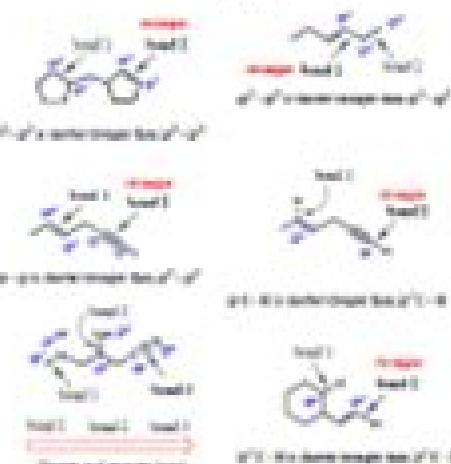
Sum of vector of the bond dipole moment for the molecule.

Average values of C-C and C-H bond depending on electronegativity

C-C Bond	$\text{H}_2\text{C}-\text{CH}_3$	$\text{H}_2\text{C}=\text{CH}_2$	$\text{CH}_3\text{CH}_2\text{CH}_3$
Electronegativity	1.10	1.10	1.10
Length	1.57 Å	1.35 Å	1.53 Å
Strength	1.01 kcal/mol	1.11 kcal/mol	1.01 kcal/mol
Hybridization	sp^3	sp^2	sp^3
Efficiency	55.7%	55.7%	55.7%
Length	1.10 Å	1.10 Å	1.10 Å
Strength	0.91 kcal/mol	0.91 kcal/mol	0.91 kcal/mol

Dipole moment and bond length

Example



Bond Length and Bond Strength

Relative bond length



Length: $\text{C-C} > \text{C-O} > \text{C=O} > \text{C=O}_2$

Strength: $\text{C-C} < \text{C-O} < \text{C=O} < \text{C=O}_2$

Efficiency: $\text{C-C} < \text{C-O} < \text{C=O} < \text{C=O}_2$

Pauli exclusion principle: $\text{C-C} < \text{C-O} < \text{C=O} < \text{C=O}_2$

London dispersion forces: $\text{C-C} < \text{C-O} < \text{C=O} < \text{C=O}_2$

Relative bond strength



Relative bond to heteroatom



$\text{C-C} < \text{C-O} < \text{C=O} < \text{C=O}_2$

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