

Reactions of Amines and Epoxides

Preparation of Amines

Amination of alcohols: Polyalkylation is a serious problem for indirect synthesis of amines



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



Bond Types, Strength, Length, Polarity and Dipole Moment

Two different types of bonding

There are two types of bonding: ionic bonding and covalent bonding

- Ionic bonds are formed by the transfer of electrons from one element to another (ionization potential) and electron affinity (electron).



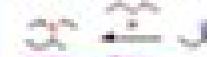
- Covalent bonds are formed by sharing of electrons between two atoms.



The Relative Electronegativity of an



Bonded and Unbonded electron pairs



Electron movement



There are two main categories

Definition of electron pair



Sigma (σ) and pi (π) bonds

Covalent bonds are classified in sigma (σ) and pi (π) bonds

- Sigma bonds are formed by head-on (frontal) overlap of orbitals.
- pi bonds are formed by side-on overlap of orbitals.
- One single bond is a sigma bond.
- One multiple bond consists of one sigma and π bonds.



π bonds are stronger than σ bonds.

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 0.5, the electrons are shared equally (nonpolar bond) for two atoms, forming a nonpolar covalent bond. Two important examples are the C-C and C-H bonds.

- If the difference in electronegativity is between 0.5 and 1.7, we have a polar covalent bond. Three important examples are shown for carbon: a bond to N, O, F and other halogens.

- A difference of 1.7 or higher can be high that the electrons are not longer shared and a ionic bond is formed.

Dipole Moment

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



- The dipole moment is shown as a vector pointing from the less electronegative to toward the more electronegative atom.
- The more electronegative atom has a partial negative charge (δ⁻) and the less electronegative atom has a partial positive charge (δ⁺).
- δ⁻ (delta minus) indicates an electron deficient atom.
- δ⁺ (delta plus) indicates an electron rich atom.

Molecular Dipole

To determine the molecular polarity, the vector sum of the dipoles is used.

The geometry of the molecule is important. Symmetrical molecules are nonpolar even if polar bonds are present. A good example is carbon dioxide.



Cancel the vector as the other pointing point in the same way.

Bond Length and Bond Strength

As the bond length increases, the bond strength decreases.

Bond	H-H	H-Cl	H-Br	H-I
Length	109 pm	127 pm	141 pm	154 pm
Strength	436 kJ/mol	431 kJ/mol	366 kJ/mol	298 kJ/mol

Double bonds are stronger than single bonds and a triple bond is stronger than a double bond.

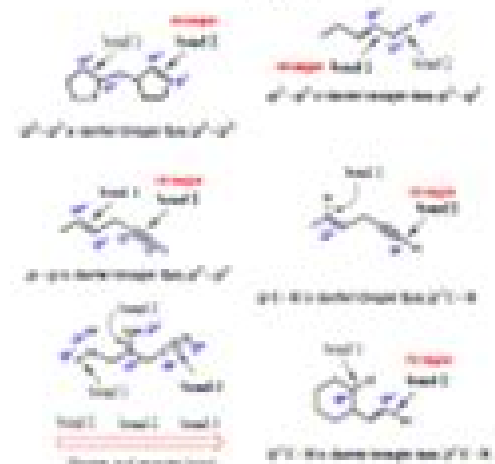


As the bond length increases, the bond strength decreases.

Average values of C-C and C-H bonds depending on hybridization			
C-C Bond	sp^3-sp^3 C-C	sp^2-sp^2 C-C	$sp-sp$ C-C
Hybridization	sp^3-sp^3	sp^2-sp^2	$sp-sp$
Length	154 pm	134 pm	121 pm
Strength	376 kJ/mol	611 kJ/mol	837 kJ/mol
Hybridization	sp^3-sp^3	sp^2-sp^2	$sp-sp$
C-H Bond	25% s	33% s	50% s
Length	109 pm	108 pm	106 pm
Strength	413 kJ/mol	505 kJ/mol	518 kJ/mol

Higher s character = stronger and shorter bond

Examples



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