

Reactions of Amines and Epoxides

Preparation of Amines

Amination of alcohols: Polyalkylation is a serious problem for tertiary hydroxyl group



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



Bond Types, Strength, Length, Polarity and Dipole Moment

Two different bond types

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. Examples:



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



The Relative Electronegativity of an



Bonded and Unbonded electron



Bonding electrons



There are two main categories

1. Definition of electron pair



Sigma (σ) and pi (π) bonds

Covalent bonds are classified as **sigma (σ) bonds** and **pi (π) bonds**.

- **Sigma bonds** are formed by head-on (end-to-end) overlap of orbitals.
- **pi bonds** are formed by side-on overlap of orbitals.
- **Any single bond** is a **sigma bond**.
- **Any multiple bond** consists of one **sigma** and **π** bond(s).



σ bonds are stronger than π 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).
- If the difference in electronegativity is **between 0.5 and 1.7**, we have a **polar covalent bond**. More electronegative atom attracts the electron pair more.
- 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 defined as a vector pointing from the less electronegative to 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 the dipoles are present. A good example is carbon dioxide.



Consider the vector sum of the dipoles pointing from the positive to the negative end.

Bond Length and Bond Strength

As the bond order increases, the bond length decreases and the bond strength increases.

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 triple bonds are stronger than double bonds.

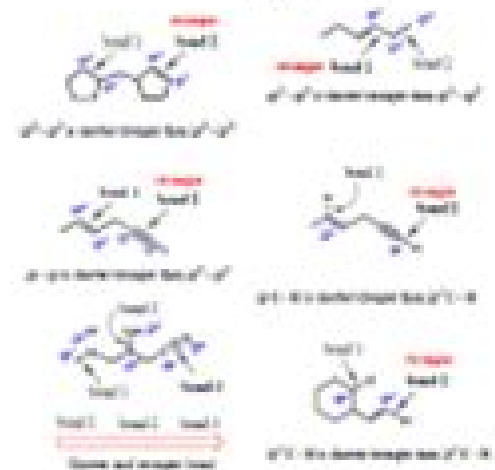


As the bond order increases, the bond length decreases and the bond strength increases.

Average values of C-C and C-H bonds depending on hybridization			
C-C Bond	sp ³ -C-C	sp ² -C-C	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	591 kJ/mol
Hybridization	sp ³ -sp ³	sp ² -sp ²	sp-sp
C-H Bond	sp ³ -H	sp ² -H	sp-H
Length	109 pm	108 pm	106 pm
Strength	413 kJ/mol	506 kJ/mol	548 kJ/mol

Higher s character = stronger and shorter bond

Examples



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Table of Contents Organic Chemistry Study Guide

1. Understanding the eBook Organic Chemistry Study Guide
 - The Rise of Digital Reading Organic Chemistry Study Guide
 - Advantages of eBooks Over Traditional Books
2. Identifying Organic Chemistry Study Guide
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Organic Chemistry Study Guide
 - User-Friendly Interface
4. Exploring eBook Recommendations from Organic Chemistry Study Guide
 - Personalized Recommendations

- Organic Chemistry Study Guide User Reviews and Ratings
- Organic Chemistry Study Guide and Bestseller Lists
- 5. Accessing Organic Chemistry Study Guide Free and Paid eBooks
 - Organic Chemistry Study Guide Public Domain eBooks
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- 6. Navigating Organic Chemistry Study Guide eBook Formats
 - ePub, PDF, MOBI, and More
 - Organic Chemistry Study Guide Compatibility with Devices
 - Organic Chemistry Study Guide Enhanced eBook Features
- 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Organic Chemistry Study Guide
 - Highlighting and Note-Taking Organic Chemistry Study Guide
 - Interactive Elements Organic Chemistry Study Guide
- 8. Staying Engaged with Organic Chemistry Study Guide
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Organic Chemistry Study Guide
- 9. Balancing eBooks and Physical Books Organic Chemistry Study Guide
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Organic Chemistry Study Guide
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Organic Chemistry Study Guide
 - Setting Reading Goals Organic Chemistry Study Guide
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Organic Chemistry Study Guide
 - Fact-Checking eBook Content of Organic Chemistry Study Guide

- Distinguishing Credible Sources
- 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
- 14. Embracing eBook Trends
 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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