

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 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



Electron movement



There are two main categories

Definition of electron pair



Sigma (σ) and pi (π) bonds

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

- **Sigma bonds** are formed by head-on (frontal) overlap of orbitals
- **π bonds** are formed by side-on overlap of orbitals
- **Any single bond** is a **sigma bond**
- **Any multiple bond** consists of one **sigma** and **n** **pi** 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) between the 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 **dipole** is used

The geometry of the molecule is important. **symmetrical molecules** are **nonpolar** (even polar bonds for dipole moment are cancelled). A good example is carbon dioxide



Cancel the vector as the other pointing just in the other 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

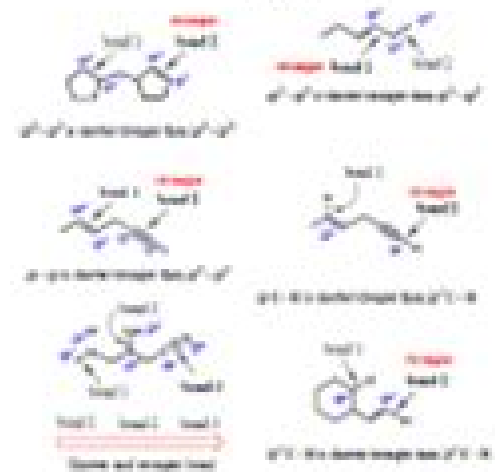


As the length of the bond increases, the bond strength decreases

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	153 pm	134 pm	120 pm
Strength	376 kJ/mol	472 kJ/mol	591 kJ/mol
Hybridization	sp^3-sp^2	sp^2-sp	$sp-sp$
C-H bonds	101 pm	106 pm	103 pm
Length	153 pm	134 pm	120 pm
Strength	413 kJ/mol	472 kJ/mol	591 kJ/mol

Higher s character = stronger and shorter bond

Examples



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1. Understanding the eBook Organic Chemistry Final Study Guide
 - The Rise of Digital Reading Organic Chemistry Final Study Guide
 - Advantages of eBooks Over Traditional Books
2. Identifying Organic Chemistry Final 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 Final Study Guide
 - User-Friendly Interface
4. Exploring eBook Recommendations from Organic Chemistry Final Study Guide
 - Personalized Recommendations
 - Organic Chemistry Final Study Guide User Reviews and Ratings
 - Organic Chemistry Final Study Guide and Bestseller Lists

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