What is organic Chemistry?

• Organic Chemistry is the chemistry of compounds that contain the element carbon.

• Originally, it was thought that compounds in living things, termed organic compounds, were fundamentally different from those in nonliving things, called inorganic compounds.

Some Representative Organic Molecules

• Methane
• Ethanol
• Trifluoromethane: chlorofluorocarbon (CFC)
• Amoxicillin
• Fluoxetine: Prozac (antidepressant)
• AZT: AIDS drug
• DDT: dichlorodiphenyltrichloroethane
• Capsaicin
The nucleus contains positively charged protons and uncharged neutrons.

- The nucleus contains positively charged protons and uncharged neutrons.
- The electron cloud is composed of negatively charged electrons.
Elements in the same row are similar in size.
Elements in the same column have similar electronic and chemical properties.

The Periodic Table

- Note the location of carbon in the second row, group 4A.
• An $s$ orbital has a sphere of electron density and is lower in energy than the other orbitals of the same shell.
• A $p$ orbital has a dumbbell shape and contains a node of electron density at the nucleus. It is higher in energy than an $s$ orbital.

Since there is only one orbital in the first shell, and each shell can hold a maximum of two electrons, there are two elements in the first row, H and He.

Each element in the second row of the periodic table has four orbitals available to accept additional electrons: one 2$s$ orbital, and three 2$p$ orbitals.
Second Row Elements

• Since each of the four orbitals available in the second shell can hold two electrons, there is a maximum capacity of eight electrons for elements in the second row.

• The second row of the periodic chart consists of eight elements, obtained by adding electrons to the 2s and three 2p orbitals.

1.2 Bonding

• Bonding is the joining of two atoms in a stable arrangement.

• Through bonding, atoms attain a complete outer shell of valence electrons.

• Through bonding, atoms attain a stable noble gas configuration.

• Ionic bonds result from the transfer of electrons from one element to another.

• Covalent bonds result from the sharing of electrons between two nuclei.
• An ionic bond generally occurs when elements on the far left side of the periodic table combine with elements on the far right side, ignoring noble gases.

• A positively charged cation formed from the element on the left side attracts a negatively charged anion formed from the element on the right side. An example is sodium chloride, NaCl.

![NaCl—An ionic crystalline lattice](image)

The transfer of electrons form stable salts composed of cations and anions. LiF (lithium fluoride)
A compound may have either ionic or covalent bonds.

Bonding in Molecular Hydrogen (H₂)

- Hydrogen forms one covalent bond.
- When two hydrogen atoms are joined in a bond, each has a filled valence shell of two electrons.

![Diagram of two hydrogen atoms forming a bond](image)

• Second row elements can have no more than eight electrons around them. For neutral molecules, this has two consequences:
  - Atoms with one, two, or three valence electrons form one, two, or three bonds, respectively, in neutral molecules.
  - Atoms with four or more valence electrons form enough bonds to give an octet. This results in the following equation:

\[
predicted\ \text{number of bonds} = 8 - \text{number of valence electrons}
\]

- When second-row elements form fewer than four bonds their octets consist of both bonding (shared) and nonbonding (unshared) electrons. Unshared electrons are also called **lone pairs**.
1.3 Lewis Structures

Lewis structures are electron dot representations for molecules. There are three general rules for drawing Lewis structures:

1. Draw only the valence electrons.
2. Give every second-row element an octet of electrons, if possible.
3. Give each hydrogen two electrons.

In a Lewis structure, a solid line indicates a two-electron covalent bond.
Formal Charge

- Formal charge is the charge assigned to individual atoms in a Lewis structure.
- By calculating formal charge, we determine how the number of electrons around a particular atom compares to its number of valence electrons. Formal charge is calculated as follows:

\[
\text{formal charge} = \text{number of valence electrons} - \text{number of electrons an atom “owns”}
\]

- The number of electrons “owned” by an atom is determined by its number of bonds and lone pairs.
- An atom “owns” all of its unshared electrons and half of its shared electrons.

The number of electrons “owned” by different atoms is indicated in the following examples:

**Example 1**

\[\text{C} \]

- C shares eight electrons.
- C “owns” four electrons.

**Example 2**

\[\text{C} = \text{C} \]

- Each C shares eight electrons.
- Each C “owns” four electrons.

**Example 3**

\[\text{C}^+ \]

- C shares six electrons.
- C has two unshared electrons.
- C “owns” five electrons.
1.4 A Isomers

In drawing a Lewis structure for a molecule with several atoms, sometimes more than one arrangement of atoms is possible for a given molecular formula.

Example:

Ethanol and dimethyl ether are constitutional isomers.

Isomers are different molecules having the same molecular formula.
1.4B Exceptions to the Octet Rule

Elements in Groups 2A and 3A

Two second-row elements without an octet

\[ \text{H-Be-H} \quad \text{four electrons around Be} \quad \text{six electrons around B} \]

Elements in the Third Row

10 electrons around S

\[ \text{CH}_3\text{SCH}_3 \]
dimethyl sulfoxide (abbreviated as DMSO)

12 electrons around S

\[ \text{H}_2\text{SO}_4 \]
sulfuric acid

10 electrons around each P

\[ \text{H}_3\text{PO}_4 \text{, CH}_3\text{CH}_2\text{NH}_2 \text{, and alendronic acid} \]

1.5 Resonance

Some molecules cannot be adequately represented by a single Lewis structure. For example, two valid Lewis structures can be drawn for the anion \((\text{HCONH})^-\). One structure has a negatively charged N atom and a C-O double bond; the other has a negatively charged O atom and a C-N double bond.

These structures are called resonance structures or resonance forms. A double headed arrow is used to separate the two resonance structures.

- Resonance structures are two Lewis structures having the same placement of atoms but a different arrangement of electrons.
1.5A An Introduction to Resonance Theory

Regarding the two resonance forms of (HCONH)\(^-\) shown below, it should be noted that:

- Neither resonance structure is an accurate representation for (HCONH)\(^-\). The true structure is a composite of both resonance forms and is called a resonance hybrid.
- The hybrid shows characteristics of both structures.
- Resonance allows certain electron pairs to be delocalized over two or more atoms, and this delocalization adds stability.
- A molecule with two or more resonance forms is said to be resonance stabilized.

\[\begin{array}{c}
\text{O} \\
\text{H-C=}\text{N-} \\
\text{H} \\
\end{array} \quad \text{double-headed arrow} \quad \begin{array}{c}
\text{O}^{-} \\
\text{H-C=N-H} \\
\end{array}\]

The following basic principles of resonance theory should be kept in mind:

1. **Resonance structures are not real.** An individual resonance structure does not accurately represent the structure of a molecule or ion. Only the hybrid does.

2. **Resonance structures are not in equilibrium with each other.** There is no movement of electrons from one form to another.

3. **Resonance structures are not isomers.** Two isomers differ in the arrangement of both atoms and electrons, whereas resonance structures differ only in the arrangement of electrons.
1.5B Drawing Resonance Structures

Rule [1]: Two resonance structures differ in the position of multiple bonds and nonbonded electrons. The placement of atoms and single bonds always stays the same.

Rule [2]: Two resonance structures must have the same number of unpaired electrons.

Rule [3]: Resonance structures must be valid Lewis structures. Hydrogen must have two electrons and no second-row element can have more than eight.
• **Curved arrow notation** is a convention that is used to show how electron position differs between the two resonance forms.

• **Curved arrow notation shows the movement of an electron pair.** The tail of the arrow always begins at the electron pair, either in a bond or lone pair. The head points to where the electron pair “moves.”

**Example 1:**

![Example 1 image]

**Example 2:**

![Example 2 image]

*In the two examples above, a lone pair is located on an atom directly bonded to a double bond:
Resonance Structure Examples

In the above examples, an atom bearing a (+) charge is bonded either to a double bond or an atom with a lone pair:

1.5C Resonance Hybrids

- A resonance hybrid is a composite of all possible resonance structures. In the resonance hybrid, the electron pairs drawn in different locations in individual resonance forms are delocalized.

- When two resonance structures are different, the hybrid looks more like the “better” resonance structure. The “better” resonance structure is called the major contributor to the hybrid, and all others are minor contributors.

- The hybrid is a weighted average of the contributing resonance structures.
A “better” resonance structure is one that has more bonds and fewer charges.

The hybrid looks more like this Lewis structure.

more bonds
fewer charges

major contributor

Drawing Resonance Hybrids

individual resonance structures

resonance hybrid
1.6 Determining Molecular Shape

Two variables define a molecule’s structure: bond length and bond angle.

1.6A Bond Length

- Bond length decreases across a row of the periodic table as the size of the atom decreases.
- Bond length increases down a column of the periodic table as the size of an atom increases.

### Structure and Bonding

**Table 1.2 Average Bond Lengths**

<table>
<thead>
<tr>
<th>Bond Length (Å)</th>
<th>Bond Length (Å)</th>
<th>Bond Length (Å)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H – H</td>
<td>0.74</td>
<td>H – F</td>
</tr>
<tr>
<td>C – H</td>
<td>1.09</td>
<td>H – Cl</td>
</tr>
<tr>
<td>N – H</td>
<td>1.01</td>
<td>H – Br</td>
</tr>
<tr>
<td>O – H</td>
<td>0.96</td>
<td>H – I</td>
</tr>
</tbody>
</table>
1.6B Bond Angle

Bond angle determines the shape around any atom bonded to two other atoms.

- The number of groups surrounding a particular atom determines its geometry. A group is either an atom or a lone pair of electrons.

- The most stable arrangement keeps these groups as far away from each other as possible. This is exemplified by Valence Shell Electron Pair Repulsion (VSEPR) theory.

<table>
<thead>
<tr>
<th>Number of groups</th>
<th>Geometry</th>
<th>Bond angle</th>
</tr>
</thead>
<tbody>
<tr>
<td>two groups</td>
<td>linear</td>
<td>180°</td>
</tr>
<tr>
<td>three groups</td>
<td>trigonal planar</td>
<td>120°</td>
</tr>
<tr>
<td>four groups</td>
<td>tetrahedral</td>
<td>109.5°</td>
</tr>
</tbody>
</table>

Two groups around an atom—
Three groups around an atom—

Two trigonal planar molecules

\[
\begin{align*}
120^\circ & \\
\text{F} & \text{B} \quad \text{F} \\
\text{F} & \\
\text{three atoms around B} & \\
\text{three groups} & \\
\end{align*}
\]

All three B–F bonds lie in one plane.

All six atoms lie in one plane.

ethylene

Four groups around an atom—

Tetrahedral arrangement

\[
\begin{align*}
109.5^\circ & \\
\text{H} & \text{C} \quad \text{H} \\
\text{H} & \\
\text{preferred geometry} & \\
\text{larger H–C–H bond angle} & \\
\end{align*}
\]

Square planar arrangement

\[
\begin{align*}
90^\circ & \\
\text{H} & \text{C} \quad \text{H} \\
\text{H} & \\
\text{This geometry does not occur.} & \\
\end{align*}
\]
Drawing Three Dimensional Structures

- A solid line is used for a bond in the plane.
- A wedge is used for a bond in front of the plane.
- A dashed line is used for a bond behind the plane.

The molecule can be turned in many different ways, generating many equivalent representations. All of the following are acceptable drawings for \( \text{CH}_4 \).

Each drawing has two solid lines, one wedge, and one dashed line.
Note that wedges and dashes are used for groups that are really aligned one behind another. It does not matter in the following two drawings whether the wedge or dash is skewed to the left or right, because the two H atoms are really aligned.

The position of the wedge and dash does not matter. The two H atoms are really aligned.

A Nonbonded Pair of Electrons is Counted as a “Group”

In ammonia (NH₃), one of the four groups attached to the central N atom is a lone pair. The three H atoms and the lone pair point to the corners of a tetrahedron. The H-N-H bond angle of 107° is close to the theoretical tetrahedral bond angle of 109.5°. This shape is referred to as a trigonal pyramid.
In water (H₂O), two of the four groups attached to the central O atom are lone pairs. The two H atoms and two lone pairs around O point to the corners of a tetrahedron. The H-O-H bond angle of 105° is close to the theoretical tetrahedral bond angle of 109.5°. Water has a bent shape, because the two groups around oxygen are lone pairs of electrons.

In both NH₃ and H₂O, the bond angle is smaller than the theoretical tetrahedral bond angle because of repulsion of the lone pairs of electrons. The bonded atoms are compressed into a smaller space with a smaller bond angle.
## Predicting Geometry Based on Counting of Groups Around the Central Atom

**Figure 1.4**
Summary: Determining geometry based on the number of groups

<table>
<thead>
<tr>
<th>Number of groups around an atom</th>
<th>Geometry</th>
<th>Bond angle</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>linear</td>
<td>180°</td>
<td>BeH₂, HC≡CH</td>
</tr>
<tr>
<td>3</td>
<td>trigonal planar</td>
<td>120°</td>
<td>BF₃, CH₂=CH₂</td>
</tr>
<tr>
<td>4</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>CH₄, NH₃, H₂O</td>
</tr>
</tbody>
</table>

## 1.7 Drawing Organic Structures

### 1.7A Condensed Structures

- All atoms are drawn in, but the two-electron bond lines are generally omitted.
- Atoms are usually drawn next to the atoms to which they are bonded.
- Parentheses are used around similar groups bonded to the same atom.
- Lone pairs are omitted.

```
H     H     H
H--C--C--C--H = CH₃CH₂CH₂CH₃ or CH₃(CH₂)₂CH₃  
H     H     H

H
H--C--H
H     H
H--C--C--H = (CH₃)₃CH
H     H     H

2 CH₂ groups bonded together
```
Examples of Condensed Structures (from Figure 1.5)

Examples of Condensed Structures Containing a C-O Double Bond (from Figure 1.6)
1.7B Skeletal Structures

• Assume there is a carbon atom at the junction of any two lines or at the end of any line.
• Assume there are enough hydrogens around each carbon to make it tetravalent.
• Draw in all heteroatoms and hydrogens directly bonded to them.

Examples of Skeletal Structures (from Figure 1.7)

- This C needs 1 H.
- This C needs 1 H.
- This C needs 3 H's.
- Each of these C's needs 3 H's.
- There are 2 C's on the triple bond.
Words of Caution Regarding Interpretation of Skeletal Structures...

- A charge on a carbon atom takes the place of one hydrogen atom.
- The charge determines the number of lone pairs. Negatively charged carbon atoms have one lone pair and positively charged carbon atoms have none.

1.8 Hybridization

1.8A Hydrogen

When the 1s orbital of one H atom overlaps with the 1s orbital of another H atom, a sigma ($\sigma$) bond that concentrates electron density between the two nuclei is formed.

This bond is cylindrically symmetrical because the electrons forming the bond are distributed symmetrically about an imaginary line connecting the two nuclei.
1.8B Bonding in Methane

To account for the bonding patterns observed in more complex molecules, we must take a closer look at the $2s$ and $2p$ orbitals of atoms in the second row.

Carbon has two core electrons, plus four valence electrons. To fill atomic orbitals in the most stable arrangement, electrons are placed in the orbitals of lowest energy. For carbon, this places two in the $2s$ orbital and one each in $2p$ orbitals.

\[
\begin{array}{c}
\text{C (1s}^2\text{) +} \\
\text{4 valence electrons}
\end{array}
\]

Note: The lowest energy arrangement of electrons for an atom is called its ground state.

In this description, carbon should form only two bonds because it has only two unpaired valence electrons, and CH$_2$ should be a stable molecule. However, CH$_2$ is a very unstable species that cannot be isolated under typical laboratory conditions. Note that in CH$_2$, carbon would not have an octet of electrons.

Two bonds from two unpaired electrons

\[\text{H} - \overset{\text{\scriptsize \ldots}}{\text{C}} - \text{H}\]

no octet (unstable)
There is a second possibility. Promotion of an electron from a 2s to a vacant 2p orbital would form four unpaired electrons for bonding. This process requires energy because it moves an electron to a higher energy orbital. This higher energy electron configuration is called an electronically excited state.

![Diagram of ground state and excited state for carbon]

But this description is still not adequate. Carbon would form two different types of bonds: three with 2p orbitals and one with a 2s orbital. However, experimental evidence points to carbon forming four identical bonds in methane.

To solve this dilemma, chemists have proposed that atoms like carbon do not use pure s and pure p orbitals in forming bonds. Instead, atoms use a set of new orbitals called hybrid orbitals. **Hybridization** is the combination of two or more atomic orbitals to form the same number of hybrid orbitals, each having the same shape and energy.
Shape and Orientation of $sp^3$ Hybrid Orbitals

The mixing of a spherical $2s$ orbital and three dumbbell shaped $2p$ orbitals together produces four hybrid orbitals, each having one large lobe and one small lobe (Figure 1.8).

The four hybrid orbitals are oriented towards the corners of a tetrahedron, and form four equivalent bonds.

Bonding Using $sp^3$ Hybrid Orbitals

Each bond in CH$_4$ is formed by overlap of an $sp^3$ hybrid orbital of carbon with a $1s$ orbital of hydrogen. These four bonds point to the corners of a tetrahedron.

Figure 1.9
Bonding in CH$_4$ using $sp^3$ hybrid orbitals
1.8C Other Hybridization Patterns

- One 2s orbital and three 2p orbitals form four sp³ hybrid orbitals.
- One 2s orbital and two 2p orbitals form three sp² hybrid orbitals.
- One 2s orbital and one 2p orbital form two sp hybrid orbitals.

To determine the hybridization of an atom in a molecule, we count the number of groups around the atom. The number of groups (atoms and nonbonded electron pairs) corresponds to the number of atomic orbitals that must be hybridized to form the hybrid orbitals.
Hybridization Examples

**BeH₂**
- 1s Be
- 1s H
- sp hybrid orbitals
- two Be–H bonds

**BF₃ — Two views**
- From above: 2p orbitals on B and F
- From the side: 120° B–F bond angle
- The three B–F bonds all lie in a plane, 120° apart.
- The unhybridized p orbital extends above and below the plane.

**Figure 1.12**
Hybrid orbitals of NH₃ and H₂O

NH₃ = \([\text{sp}^3] \text{N} \quad \text{H} \quad \text{H} \quad \text{H}\)

H₂O = \([\text{sp}^3] \text{O} \quad \text{H} \quad \text{H} \quad \text{H}\)

1.9 Ethane, Ethylene, and Acetylene

1.9A Ethane

- ethane
- tetrahedral C atoms
- Two sp³ hybrid orbitals overlap to form the C–C bond.
- Each C–H bond is formed by overlap of an sp³ hybrid on C with a 1s orbital on H.
Making a model of ethane illustrates one additional feature about its structure. Rotation occurs around the central C—C $\sigma$ bond.

1.9B Ethylene

Each carbon is trigonal and planar. Each carbon is $sp^2$ hybridized

Forming an $sp^2$ hybridized carbon atom

$2s$ unhybridized C

$2p$ $2p$ $2p$

hybridize $sp^2$ $sp^2$ $sp^2$

$sp^2$ hybridized C

This $2p$ orbital has one electron.
Unlike the C—C bond in ethane, rotation about the C—C double bond in ethylene is restricted. It can only occur if the $\pi$ bond first breaks and then reforms, a process that requires considerable energy.
1.9C Acetylene

From Figure 1.13

- Forming an sp hybridized carbon atom
- Two 2p orbitals have one electron each.
- Two groups around C

Acetylene

- Overlap of the two sp^2 hybrid orbitals forms the C–C σ bond.
- Overlap of the two 2p orbitals forms the C–C π bond.

- Each C has two sp hybrid orbitals.
- The C–H bonds and C–C bond are σ bonds.
Each carbon atom has two unhybridized $2p$ orbitals that are perpendicular to each other and to the $sp$ hybrid orbitals. The side-by-side overlap of two $2p$ orbitals on one carbon with two $2p$ orbitals on the other carbon creates the second and third bonds of the triple bond. All triple bonds are composed of one sigma and two pi bonds.
1.10 Bond Length and Bond Strength

1.10A A Comparison of Carbon-Carbon Bonds

- As the number of electrons between two nuclei increases, bonds become shorter and stronger.
- Thus, triple bonds are shorter and stronger than double bonds, which are shorter and stronger than single bonds.
1.10B A Comparison of Carbon-Hydrogen Bonds

- The length and strength of C—H bonds vary depending on the hybridization of the carbon atom.

![Increasing bond strength and bond length diagram]

### Structure and Bonding

**Table 1.3** Bond Lengths and Bond Strengths for Ethane, Ethylene, and Acetylene

<table>
<thead>
<tr>
<th>Compound</th>
<th>C—C bond length (Å)</th>
<th>Bond strength kcal/mol (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃CH₃</td>
<td>1.53</td>
<td>88 (368)</td>
</tr>
<tr>
<td>CH₂=CH₂</td>
<td>1.34</td>
<td>152 (635)</td>
</tr>
<tr>
<td>HC≡CH</td>
<td>1.21</td>
<td>200 (837)</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Compound</th>
<th>C—H bond length (Å)</th>
<th>Bond strength kcal/mol (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃CH₂=H</td>
<td>1.11</td>
<td>98 (410)</td>
</tr>
<tr>
<td>CH₂=C=CH₂</td>
<td>1.10</td>
<td>104 (435)</td>
</tr>
<tr>
<td>HC≡C=CH</td>
<td>1.09</td>
<td>125 (523)</td>
</tr>
</tbody>
</table>
Note:

- As the percent s-character increases, a hybrid orbital holds its electrons closer to the nucleus, and the bond becomes shorter and stronger.
- Although $sp^3$, $sp^2$ and $sp$ hybrid orbitals are similar in shape, they are different in size.
1.11 Electronegativity and Bond Polarity

Electronegativity is a measure of an atom’s attraction for electrons in a bond.

Electronegativity values for some common elements:

Electronegativity values are used as a guideline to indicate whether the electrons in a bond are equally shared or unequally shared between two atoms. When electrons are equally shared, the bond is nonpolar. When differences in electronegativity result in unequal sharing of electrons, the bond is polar, and is said to have a “separation of charge” or a “dipole”.

- A carbon—carbon bond is nonpolar. The same is true whenever two different atoms having similar electronegativities are bonded together.
- C—H bonds are considered to be nonpolar because the electronegativity difference between C and H is small.
Bonding between atoms of different electronegativity values results in unequal sharing of electrons.

Example: In the C—O bond, the electrons are pulled away from C (2.5) toward O (3.4), the element of higher electronegativity. The bond is polar, or polar covalent. The bond is said to have dipole; that is, separation of charge.

$\delta^+$ means the indicated atom is electron deficient.

$\delta^-$ means the indicated atom is electron rich.

The direction of polarity in a bond is indicated by an arrow with the head of the arrow pointing towards the more electronegative element. The tail of the arrow, with a perpendicular line drawn through it, is drawn at the less electronegative element.

1.12 Polarity of Molecules

Use the following two-step procedure to determine if a molecule has a net dipole:

1. Use electronegativity differences to identify all of the polar bonds and the directions of the bond dipoles.
2. Determine the geometry around individual atoms by counting groups, and decide if individual dipoles cancel or reinforce each other in space.
A polar molecule has either one polar bond, or two or more bond dipoles that reinforce each other. An example is water:

The two individual bond dipoles reinforce.

\[
\begin{array}{c}
\text{H}^+ \\
\delta^+
\end{array}
\quad \quad \quad \quad
\begin{array}{c}
\text{O}^-
\quad \text{H}^+ \\
\delta^- \quad \delta^+
\end{array}
\]

The net dipole bisects the H–O–H bond angle. The bent representation shows that the dipoles reinforce.

Do *not* draw H₂O as

\[
\begin{array}{c}
\text{H} \\
\delta^+
\end{array}
\quad \quad \quad \quad
\begin{array}{c}
\text{O} \\
\delta^-
\end{array}
\quad \quad \quad \quad
\begin{array}{c}
\text{H} \\
\delta^-
\end{array}
\]

Answer: H₂O is a polar molecule.

A nonpolar molecule has either no polar bonds, or two or more bond dipoles that cancel. An example is carbon dioxide:

The two dipoles cancel.

\[
\begin{array}{c}
\delta^- \quad \text{O}^= \quad \delta^+
\end{array}
\quad \quad \quad \quad
\begin{array}{c}
\delta^+ \\
\text{C} \\
\delta^-
\end{array}
\quad \quad \quad \quad
\begin{array}{c}
\delta^- \\
\text{O}^= \\
\delta^+
\end{array}
\]

Answer: CO₂ is a nonpolar molecule.