In 1989, researchers discovered the structure of HIV-protease, a protein synthesized by the human immunodeficiency virus (HIV), which causes AIDS.

Researchers used bonding theories—models that predict how atoms bond together to form molecules—to simulate how potential drug molecules would interact with the HIV-protease molecule.

The gold-colored structure on the computer screen is a representation of HIV-protease.

The molecule shown in the center is Indinavir, a protease inhibitor.
Without HIV-protease, HIV could not spread in the human body because the virus could not copy itself, and AIDS would not develop.

With knowledge of the HIV-protease structure, drug companies designed a molecule that would disable protease by sticking to the working part of the molecule (called the *active site*).

By the early 1990s, these companies had developed several drug molecules that inhibit the action of HIV-protease called *protease inhibitors*.

In human trials, protease inhibitors in combination with other drugs have decreased the viral count in HIV-infected individuals to undetectable levels.

Although these drugs do not cure AIDS, HIV-infected individuals who regularly take their medication can expect nearly normal life spans.
Bonding Theories

- Bonding theories predict how atoms bond together to form compounds.
- They predict what combinations of atoms form compounds and what combinations do not.
- Bonding theories predict why salt is NaCl and not NaCl$_2$ and why water is H$_2$O and not H$_3$O.
- Bonding theories explain the shapes of molecules, which in turn determine many of their physical and chemical properties.
Lewis theory is named after G. N. Lewis (1875–1946), the American chemist who developed it. We represent electrons as dots and draw what are called dot structures, or Lewis structures, to represent molecules. These structures have tremendous predictive power. Lewis theory can determine whether a particular set of atoms will form a stable molecule and what that molecule might look like. Although modern chemists also use more advanced bonding theories to better predict molecular properties, Lewis theory remains the simplest method for making quick, everyday predictions about molecules.
In Lewis theory, the valence electrons of main-group elements are represented as dots surrounding the symbol of the element.

The result is called a **Lewis structure**, or **dot structure**.

Remember, the number of valence electrons for any main-group element (except helium, which has 2 valence electrons but is in Group 8A) is equal to the group number of the element.
The electron configuration of O is as follows:

\[1s^2\, 2s^2\, 2p^4\]

Its Lewis structure is as follows:

\[O\, 6\, \text{dots representing valence electrons}\]
• Each dot represents a valence electron. The dots are placed around the element’s symbol with a maximum of two dots per side.

• Although the exact location of dots is not critical, here we fill in the dots singly first and then pair them (with the exception of helium, described shortly).

• It is important to follow this convention when doing the homework.
The Lewis Structures for All of the Period 2 Elements Are:

Li · Be · B · C · N : O : F : Ne

- Lewis structures allow us to easily see the number of valence electrons in an atom.
- Atoms with 8 valence electrons—which are particularly stable—are easily identified because they have eight dots, an octet.
Writing Lewis Structures for Elements

• Write the Lewis structure of phosphorus.
• Since phosphorus is in Group 5A in the periodic table, it has 5 valence electrons. Represent these as five dots surrounding the symbol for phosphorus.

• Solution:
Helium Is an exception. Its electron Configuration and Lewis Structure Are:

\[ 1s^2 \text{ He:} \]

- The Lewis structure of helium contains two paired dots (a **duet**).
- For helium, a duet represents a stable electron configuration.
Lewis Theory of Chemical Bonding

- In Lewis theory, a **chemical bond** involves the sharing or transfer of electrons to attain stable electron configurations for the bonding atoms.
- If the electrons are transferred, the bond is an **ionic bond**.
- If the electrons are shared, the bond is a **covalent bond**.
- In either case, the bonding atoms attain stable electron configurations.
Lewis Theory of Chemical Bonding

• A stable configuration usually consists of eight electrons in the outermost or valence shell. This observation leads to the **octet rule**:

  In chemical bonding, atoms transfer or share electrons to obtain outer shells with eight electrons.

• The octet rule generally applies to all main-group elements except hydrogen, lithium, and beryllium. Each of these elements achieves stability when it has two electrons (a duet) in its outermost shell.
Lewis Structures of Ionic Compounds: Electrons Transferred

- When metals bond with nonmetals, electrons are transferred from the metal to the nonmetal.
- The metal becomes a cation and the nonmetal becomes an anion.
- The attraction between the cation and the anion results in an ionic compound.
- In Lewis theory, we represent this by moving electron dots from the metal to the nonmetal.
Lewis Structures of Ionic Compounds: Electrons Transferred

• For example, potassium and chlorine have the Lewis structures:

\[
\text{K} \cdot \text{Cl} : \rightarrow \text{K}^+ \text{ [Cl} :]^- \\
\]

• When potassium and chlorine bond, potassium transfers its valence electron to
The transfer of the electron gives chlorine an octet and leaves potassium with an octet in the previous principal shell, which is now the valence shell.

The potassium, because it lost an electron, becomes positively charged, while the chlorine, which gained an electron, becomes negatively charged.

The positive and negative charges attract one another, forming the compound $\text{KCl}$.

The Lewis structure of an anion is usually written within brackets with the charge in the upper right corner (outside the brackets).

The charges are shown, but the brackets are not shown in the homework drawing program online.
Write the Lewis Structure of the Compound MgO

• Draw the Lewis structures of magnesium and oxygen by drawing two dots around the symbol for magnesium and six dots around the symbol for oxygen.

• In MgO, magnesium loses its 2 valence electrons, resulting in a 2+ charge, and oxygen gains 2 electrons, attaining a 2− charge and an octet.
Lewis Theory Predicts the Correct Chemical Formulas for Ionic Compounds

- Consider the ionic compound formed between sodium and sulfur. The Lewis structures for sodium and sulfur are as follows:

![Lewis Structure: Na• S]

- Sodium must lose one valence electron to obtain an octet (in the previous principal shell), while sulfur must gain two electrons to obtain an octet.

- The compound that forms between sodium and sulfur requires two sodium atoms to every one sulfur atom. The Lewis structure is as follows:

![Lewis Structure: Na⁺ [S²⁻] Na⁺]

- The correct chemical formula is Na₂S.
Covalent Lewis Structures: Electrons Shared

• When nonmetals bond with other nonmetals, a molecular compound results.

• Molecular compounds contain covalent bonds in which electrons are shared between atoms rather than transferred.

• In Lewis theory, we represent covalent bonding by allowing neighboring atoms to share some of their valence electrons in order to attain octets (or duets for hydrogen).
Hydrogen and oxygen have the Lewis structures:

\[ \text{H} \cdot \cdot \text{O} : \]

In water, hydrogen and oxygen share their electrons so that each hydrogen atom gets a duet and the oxygen atom gets an octet.

\[ \text{H} : \text{O} : \cdot \cdot \text{H} \]
The shared electrons—those that appear in the space between the two atoms—count toward the octets (or duets) of both of the atoms.
Electrons that are shared between two atoms are called *bonding pair* electrons.

Electrons that are only on one atom are called *lone pair* (or nonbonding) electrons.

Bonding pair electrons are often represented by dashes to emphasize that they are a chemical bond. Remember that each dash represents a *pair* of shared electrons.
Lewis Theory Explains Why the Halogens Form Diatomic Molecules

• Consider the Lewis structure of chlorine.

\[ \text{Cl} \]

• If two Cl atoms pair together, they can each attain an octet.

\[ \text{Cl} : \text{Cl} \]

or

\[ \text{Cl} \text{— Cl} \]

• When we examine elemental chlorine, we find that it exists as a diatomic molecule, just as Lewis theory predicts.

• The same is true for the other halogens.
Lewis Theory Predicts That Hydrogen Should Exist as $H_2$

- Hydrogen has the Lewis structure:
  \[ \text{H} \cdot \]

- When two hydrogen atoms share their valence electrons, they each get a duet, a stable configuration for hydrogen.

\[ \text{H:H} \quad \text{or} \quad \text{H—H} \]

- Lewis theory is correct. In nature, elemental hydrogen exists as $H_2$ molecules.
Double Bonds: In Lewis Theory, Atoms Can Share More Than One Electron Pair to Attain an Octet

• We know from Chapter 5 that oxygen exists as the diatomic molecule.
• If we pair two oxygen atoms together and then try to write a Lewis structure, we do not have enough electrons to give each O atom an octet.
• We can convert a lone pair into an additional bonding pair by moving it into the bonding region.
Double Bonds: In Lewis Theory, Atoms Can Share More Than One Electron Pair to Attain an Octet

- Each oxygen atom now has an octet because the additional bonding pair counts toward the octet of both oxygen atoms.
Triple Bonds: In Lewis Theory, Atoms Can Share More Than One Electron Pair to Attain an Octet

- Consider the Lewis structure of $\text{N}_2$.
- Since each N atom has 5 valence electrons, the Lewis structure for $\text{N}_2$ has 10 electrons.
- We do not have enough electrons to satisfy the octet rule for both N atoms.
- If we convert two lone pairs into bonding pairs, each nitrogen atom can get an octet.
Double and Triple Bonds

• When two electron pairs are shared between two atoms, the resulting bond is a double bond. In general, double bonds are shorter and stronger than single bonds. For example, the distance between oxygen nuclei in an oxygen–oxygen double bond is 121 pm. In a single bond, it is 148 pm.

• While Lewis theory does correctly predict a double bond for O₂, it is insufficient to explain the paramagnetic properties of oxygen. This problem led to the development of other bonding theories.
Double and Triple Bonds

- Triple bonds are even shorter and stronger than double bonds.
- The distance between nitrogen nuclei in a nitrogen–nitrogen triple bond is 110 pm.
- In a double bond, the distance is 124 pm.
- When we examine nitrogen in nature, we find that it exists as a diatomic molecule with a very strong, short bond between the two nitrogen atoms.
- The bond is so strong that it is difficult to break, making $\text{N}_2$ a relatively unreactive molecule.
1. Write the correct skeletal structure for the molecule.

- The skeletal structure shows the relative positions of the atoms and does not include electrons, but it must have the atoms in the correct positions.
- In nature, oxygen is the central atom, and the hydrogen atoms are terminal atoms (at the ends). The correct skeletal structure is H O H.
- The only way to absolutely know the correct skeletal structure for any molecule is to examine its structure in nature.
1. Write the correct skeletal structure for the molecule (continued).

Remember two guidelines.

- First, *hydrogen atoms will always be terminal*. Since hydrogen requires only a duet, it will never be a central atom because central atoms must be able to form at least two bonds and hydrogen can form only one.

- Second, *many molecules tend to be symmetrical*, so when a molecule contains several atoms of the same type, these tend to be in terminal positions.

- *This second guideline has many exceptions*. In cases where the skeletal structure is unclear, the correct skeletal structure is usually provided.
2. Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.

- The number of valence electrons for any main-group element is equal to its group number in the periodic table.
- If you are writing a Lewis structure for a polyatomic ion, the charge of the ion must be considered when calculating the total number of electrons.
- Add one electron for each negative charge and subtract one electron for each positive charge.
3. Distribute the electrons among the atoms, giving octets (or duets for hydrogen) to as many atoms as possible.
   - Begin by placing two electrons between each pair of atoms. These are the minimal number of bonding electrons.
   - Then distribute the remaining electrons, first to terminal atoms and then to the central atom, giving octets to as many atoms as possible.
4. If any atoms lack an octet, form double or triple bonds as necessary to give them octets. Do this by moving lone electron pairs from terminal atoms into the bonding region with the central atom.
Write the Lewis Structure for CO$_2$

Solution:
Following the symmetry guideline, write the skeletal structure:

\[
\begin{array}{cc}
\text{O} & \text{C} \\
\text{O} & & \text{O}
\end{array}
\]

- Calculate the total number of electrons for the Lewis structure by summing the valence electrons of each atom in the molecule.

\[
\text{(# valence e}^-\text{ in C) + 2(# valance e}^-\text{ in O)}
\]

\[
4\text{ e}^- + 2(6\text{ e}^-) = 16\text{ e}^- 
\]
Write the Lewis Structure for \( \text{CO}_2 \)

- Bonding electrons first:

\[
\begin{align*}
\text{O:}&\text{C:O} \\
(4 \text{ of } 16 \text{ electrons used})
\end{align*}
\]

- Lone pairs on terminal atoms next:

\[
\begin{align*}
\ddot{\text{O:}}\text{C:}\ddot{\text{O}}: \\
(16 \text{ of } 16 \text{ electrons used})
\end{align*}
\]

- Move lone pairs from the oxygen atoms to bonding regions to form double bonds.
Writing Lewis Structures for Polyatomic Ions

- Write Lewis structures for polyatomic ions by following the same procedure, but pay special attention to the charge of the ion when calculating the number of electrons for the Lewis structure.
- Add one electron for each negative charge and subtract one electron for each positive charge.
- Show the Lewis structure for a polyatomic ion within brackets and write the charge of the ion in the upper right corner.
Write the Lewis Structure for the CN⁻ Ion

• We begin by writing the skeletal structure:
  \[ \text{C N} \]

• Calculate the total number of electrons for the Lewis structure by summing the number of valence electrons for each atom and adding one for the negative charge.

\[
(\# \text{ valence e}^- \text{ in C}) + (\# \text{ valance e}^- \text{ in N}) + 1\text{e}^- \text{ for } -\text{ charge}
\]

\[
4 \text{ e}^- + 5 \text{ e}^- + 1 \text{ e}^- = 10 \text{ e}^-
\]
Write the Lewis Structure for the CN$^-$ Ion

• Bonding electrons first:

\[
\text{C}::\text{N} \quad \text{(2 of 10 electrons used)}
\]

\[
:\text{C}::\text{N}: \quad \text{(10 of 10 electrons used)}
\]

• Distribute the remaining electrons as lone pairs on terminal atoms.

\[
[:\text{C}::\text{N}:]^- \quad \text{or} \quad [:\text{C}=\text{N}:]^- \]

• Since neither of the atoms has octets, move two lone pairs into the bonding region to form a triple bond, giving both atoms octets.

• Enclose the Lewis structure in brackets and write the charge of the ion in the upper right corner.
Exceptions to the Octet Rule

• Lewis theory is often correct in its predictions, but exceptions exist. For example, the Lewis structure for NO, which has 11 electrons, is:

\[ \cdot\vdots\vdots\vdots \text{ or } \cdot\vdots\vdots\vdots \]

• It is impossible to write good Lewis structures for molecules with odd numbers of electrons, yet some of these molecules exist in nature. In such cases, we write the best Lewis structure that we can.

• Another significant exception to the octet rule is boron, which tends to form compounds with only 6 electrons around B, rather than 8. For example, BF\(_3\) and BH\(_3\)—both of which exist in nature—lack an octet for B.

\[ \cdot\vdots\vdots\vdots \quad \text{or} \quad \cdot\vdots\vdots\vdots \quad \cdot\vdots\vdots\vdots \]

F

B

H

H

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Exceptions to the Octet Rule

- Molecules such as SF$_6$ and PCl$_5$ have more than 8 electrons around a central atom in their Lewis structures.

- These are referred to as expanded octets.

- Beyond mentioning them, we do not cover expanded octets here.

- In spite of these exceptions, Lewis theory remains a powerful and simple way to understand chemical bonding.
Resonance: Equivalent Lewis Structures for the Same Molecule

• When writing Lewis structures for some molecules, such as SO$_2$, we can write more than one good Lewis structure.

• Begin with the skeletal structure:

\[
\text{O S O}
\]

• Sum the valence electrons.

\[
(\# \text{ valence } e^- \text{ in S}) + 2(\# \text{ valance } e^- \text{ in O})
\]

\[
6 e^- + 2(6 e^-) = 18 e^-
\]
Resonance: Equivalent Lewis Structures for the Same Molecule

• Next, we place two electrons between each pair of atoms:

\[ \text{O:} \text{S:} \text{O} \quad (4 \text{ of } 18 \text{ electrons used}) \]

and then distribute the remaining electrons, first to terminal atoms:

\[ : \text{O:} \text{S:} \text{O}: \quad (16 \text{ of } 18 \text{ electrons used}) \]

and finally to the central atom.

\[ : \text{O:} \text{S:} \text{O}: \quad (18 \text{ of } 18 \text{ electrons used}) \]

• Since the central atom lacks an octet, move one lone pair from an oxygen atom into the bonding region to form a double bond, giving all of the atoms octets.

\[ : \text{O::} \text{S:} \text{O}: \quad \text{or} \quad : \text{O} \equiv \text{S} - \text{O}: \]
Resonance: Equivalent Lewis Structures for the Same Molecule

- We could have formed the double bond with the other oxygen atom.

\[ \text{\textbullet O} - \text{\textbullet S} = \text{\textbullet O} \]

- The two structures are equally correct as Lewis structures.
- We find that the molecule exists in nature as an average or intermediate between the two Lewis structures and neither one alone represents reality.
Either one of the two Lewis structures for SO\(_2\) would predict that SO\(_2\) would contain two different kinds of bonds (one double bond and one single bond).

When we examine SO\(_2\) in nature, we find that both of the bonds are equivalent and intermediate in strength and length between a double bond and single bond.
In nature, we find that the bonds are equivalent. We account for this in Lewis theory by representing the molecule with both structures, called **resonance structures**, with a double-headed arrow between them.

The true structure of $\text{SO}_2$ is intermediate between these two resonance structures.
Chemistry in the Environment: The Lewis Structure of Ozone

- Ozone is a form of oxygen in which three oxygen atoms bond together. Its Lewis structure consists of the two resonance structures:

\[
\begin{align*}
&\text{Ozone:} \\
&\text{O}_3 = \begin{array}{c}
\text{O} \\
\text{O} \\
\text{O}
\end{array} \\
&\text{O}_2 = \begin{array}{c}
\text{O} \\
\text{O}
\end{array}
\end{align*}
\]

- Compare the Lewis structure of ozone to the Lewis structure of O\textsubscript{2}:

\[
\begin{array}{c}
\text{O} = \text{O}
\end{array}
\]

- Which molecule, O\textsubscript{3} or O\textsubscript{2}, do you think has the stronger oxygen–oxygen bond?
O₂ has a stronger bond because it is a pure double bond. 
O₃ has bonds that are intermediate between single and double, so ozone has weaker bonds. 
O₃ shields us from harmful ultraviolet light entering Earth’s atmosphere. Photons at wavelengths of 280–320 nm (the most dangerous components of sunlight to humans) are just strong enough to break ozone’s bonds. 

In the process, the photons are absorbed.

\[
:O=O=O: + \text{UV light} \rightarrow :O=O: + \cdot \cdot \cdot
\]

The same wavelengths of UV light do not have sufficient energy to break the stronger double bond of O₂ so it is transparent to UV and cannot shield us from it.
Predicting the Shapes of Molecules

- Lewis theory, in combination with valence shell electron pair repulsion (VSEPR) theory, can be used to predict the shapes of molecules.
- VSEPR theory is based on the idea that electron groups—lone pairs, single bonds, or multiple bonds—repel each other.
- This repulsion between the negative charges of electron groups on the central atom determines the geometry of the molecule.
• Predict the shape of CO$_2$, which has the Lewis structure:

\[ \ddot{\text{O}} = \text{C} = \ddot{\text{O}} : \]

• The geometry of CO$_2$ is determined by the repulsion between the \textit{two} electron groups (the two double bonds) on the central carbon atom. These two electron groups get as far away from each other as possible, resulting in a bond angle of 180° and a \textit{linear} geometry for CO$_2$. 
Predicting the Shape of H$_2$CO

- H$_2$CO has the Lewis structure:

- This molecule has three electron groups around the central atom.
- These three electron groups get as far away from each other as possible, resulting in a bond angle of 120° and a trigonal planar geometry.
These angles predicted for H$_2$CO are approximate. The C═O double bond contains more electron density than do C—H single bonds, resulting in a slightly greater repulsion; thus, the HCH bond angle is actually 116° and the HCO bond angles are actually 122°.
If a molecule has four electron groups around the central atom, as does CH$_4$, it has a tetrahedral geometry with bond angles of 109.5°.

A tetrahedron is a geometric shape with four triangular faces.

CH$_4$ is shown here with both a ball-and-stick model (left) and a space-filling model (right). Space-filling models more closely portray molecules, yet ball-and-stick models are often used to clearly illustrate molecular geometries.
• The **mutual repulsion of the four electron groups** causes the tetrahedral shape—the tetrahedron allows the maximum separation among the four groups.

• When we write the structure of CH$_4$ on paper, it may seem that the molecule should be square planar, with bond angles of 90°, but it is not planar.

• In three dimensions the four electron groups can get farther away from each other by forming the tetrahedral geometry.
Predicting the Shapes of Molecules with Lone Pairs on the Central Atom

The NH₃ molecule has four electron groups (one lone pair and three bonding pairs).

- If we look only at the electrons, the **electron geometry**—the geometrical arrangement of the electron groups—is **tetrahedral**.
- The **molecular geometry**—the geometrical arrangement of the atoms—is **trigonal pyramidal**.
- The lone pair exerts its influence on the bonding pairs.
- The bond angles in NH₃ are actually a few degrees smaller than the ideal tetrahedral angles (for reasons covered elsewhere).
The $\text{H}_2\text{O}$ molecule has four electron groups (two lone pairs and two bonding pairs).

- If we look only at the electrons, the electron geometry—the geometrical arrangement of the electron groups—is tetrahedral.
- The molecular geometry—the geometrical arrangement of the atoms—is bent.
- The lone pairs exert their influence on the bonding pairs.
- The bond angles in $\text{H}_2\text{O}$ actually are a few degrees smaller than the ideal tetrahedral angles (for reasons covered elsewhere).
### TABLE 10.1 Electron and Molecular Geometries

<table>
<thead>
<tr>
<th>Electron Groups*</th>
<th>Bonding Groups</th>
<th>Lone Pairs</th>
<th>Electron Geometry</th>
<th>Angle between Electron Groups**</th>
<th>Molecular Geometry</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>linear</td>
<td>180°</td>
<td>linear</td>
<td>:(\ddot{O} \equiv C \equiv \ddot{O}):</td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>trigonal planar</td>
<td>120°</td>
<td>trigonal planar</td>
<td>(\ddot{O}:)</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>H-C-H</td>
<td>H(-\ddot{C}-\ddot{H}):</td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>1</td>
<td>trigonal planar</td>
<td>120°</td>
<td>bent</td>
<td>:(\ddot{O} = \ddot{S} = \ddot{O}):</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>H(-\ddot{S}-\ddot{H}):</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>4</td>
<td>0</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>tetrahedral</td>
<td>H(-\ddot{C}-\ddot{H}):</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>H(-\ddot{C}-\ddot{H}):</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>3</td>
<td>1</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>trigonal pyramidal</td>
<td>H(-\ddot{N}-\ddot{H}):</td>
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<td></td>
<td></td>
<td>H(-\ddot{N}-\ddot{H}):</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>2</td>
<td>tetrahedral</td>
<td>109.5°</td>
<td>bent</td>
<td>H(-\ddot{O}-\ddot{H}):</td>
</tr>
</tbody>
</table>

* Count only electron groups around the central atom. Each of the following is considered one electron group: a lone pair, a single bond, a double bond, and a triple bond.

** Angles listed here are idealized. Actual angles in specific molecules may vary by several degrees.
1. Draw a correct Lewis structure for the molecule.

2. Determine the total number of electron groups around the central atom.

3. Determine the number of bonding groups and the number of lone pairs around the central atom.

4. Refer to Table 10.1 to determine the electron geometry and molecular geometry.
Many chemists use this notation for bonds to indicate three-dimensional structures on two-dimensional paper.

- **Straight line**
  - Bond in plane of paper
- **Hashed lines**
  - Bond projecting into the paper
- **Wedge**
  - Bond projecting out of the paper
The major molecular geometries used in this book are shown here using this notation:
Artificial sweeteners, such as aspartame, taste sweet but have few or no calories. Taste and caloric value are entirely separate properties of foods.

The caloric value of a food depends on the amount of energy released when the food is metabolized.

Sucrose (table sugar) is metabolized by oxidation to carbon dioxide and water:

$$C_{12}H_{22}O_{11} + 6 \text{O}_2 \rightarrow 12 \text{CO}_2 + 11 \text{H}_2\text{O}$$

$$\Delta H = -5644 \text{ kJ}$$

Some artificial sweeteners, such as saccharin, are not metabolized at all—they pass through the body unchanged—and therefore have no caloric value.

Other artificial sweeteners, such as aspartame, are metabolized but have a much lower caloric content (for a given amount of sweetness) than sucrose.
In the tongue, specialized cells act as highly sensitive and specific molecular detectors. The main basis for this discrimination is the molecule’s shape.

The surface of a taste cell contains specialized protein molecules called taste receptors. Each molecule that we can taste fits snugly into a special pocket on the taste receptor protein called the active site.

Artificial sweeteners taste sweet because they fit into the receptor pocket that normally binds sucrose.

Both aspartame and saccharin bind to the active site in the protein more strongly than sugar does.

For this reason, artificial sweeteners are “sweeter than sugar.”

It takes 200 times as much sucrose as aspartame to trigger the same amount of nerve signal transmission from taste cells.

The ability of scientists to determine the shapes of key biological molecules is largely responsible for the revolution in biology that has occurred over the last 50 years.
Electronegativity and Polarity: Why Oil and Water Don’t Mix

• If you combine oil and water in a container, they separate into distinct regions. Why? Something about water molecules causes them to bunch together into one region, expelling the oil molecules into a separate region.
Electronegativity and Polarity: Why Oil and Water Don’t Mix

- The two bonds between O and H each consist of an electron pair—two electrons shared between the oxygen atom and the hydrogen atom.
- The oxygen and hydrogen atoms each donate one electron to this electron pair; however, they don’t share them equally.
- The oxygen atom takes more than its fair share of the electron pair.
Electronegativity

- The ability of an element to attract electrons within a covalent bond is called electronegativity.
- Oxygen is more electronegative than hydrogen, which means that, on average, the shared electrons are more likely to be found near the oxygen atom than near the hydrogen atom.
• Consider this representation of one of the two OH bonds:

![Dipole moment diagram]

• The oxygen atom (getting the larger share) has a partial negative charge, symbolized by $\delta^-$ (delta minus).

• The hydrogen atom (getting the smaller share) has a partial positive charge, symbolized by $\delta^+$ (delta plus).

• The result of this uneven electron sharing is a **dipole moment**, a separation of charge within the bond.
Polar Covalent Bonds

• Covalent bonds that have a dipole moment are called **polar covalent bonds**.
• The magnitude of the dipole moment, and the polarity of the bond, depend on the electronegativity difference between the two elements in the bond and the length of the bond.
• For a fixed bond length, the greater the electronegativity difference, the greater the dipole moment and the more polar the bond.
Electronegativity

- The value of electronegativity is assigned using a relative scale on which fluorine, the most electronegative element, has an electronegativity of 4.0.
- Linus Pauling introduced the electronegativity scale used here. He arbitrarily set the electronegativity of fluorine at 4.0 and computed all other values relative to fluorine.
Identical Electronegativities

- If two elements with identical electronegativities form a covalent bond, they share the electrons equally, and there is no dipole moment.
- In Cl\(_2\), the two Cl atoms share the electrons evenly. This is a pure covalent bond. The bond has no dipole moment; the molecule is nonpolar.
Large Electronegativity Difference

• If there is a large electronegativity difference between the two elements in a bond, such as what normally occurs between a metal and a nonmetal, the electron is completely transferred and the bond is ionic.

• In NaCl, Na completely transfers an electron to Cl. This is an ionic bond.
Intermediate Electronegativity Difference

- If there is an intermediate electronegativity difference between the two elements, such as between two different nonmetals, then the bond is polar covalent.
- In HF, the electrons are shared, but the shared electrons are more likely to be found on F than on H. The bond is **polar covalent**.
The degree of bond polarity is a continuous function. The guidelines given here are approximate.
Does the presence of one or more polar bonds in a molecule always result in a polar molecule? The answer is no.

A **polar molecule** is one with polar bonds that add together—they do not cancel each other—to form a net dipole moment.

If a diatomic molecule contains a polar bond, then the molecule is polar.

For molecules with more than two atoms, it is more difficult to tell polar molecules from nonpolar ones because two or more polar bonds may cancel one another.
Consider carbon dioxide:

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

- Each bond is polar because the difference in electronegativity between oxygen and carbon is 1.0.
- CO\(_2\) has a linear geometry, the dipole moment of one bond completely cancels the dipole moment of the other, and the molecule is nonpolar.
Vector Notation for Dipole Moments

• We can represent polar bonds with arrows (vectors) that point in the direction of the negative pole and have a plus sign at the positive pole (as just shown for carbon dioxide).

• If the arrows (vectors) point in exactly opposing directions as in carbon dioxide, the dipole moments cancel.

• In the vector representation of a dipole moment, the vector points in the direction of the atom with the partial negative charge.
Consider water (H$_2$O):

- Each bond is polar because the difference in electronegativity between oxygen and hydrogen is 1.4.
- Water has two dipole moments that do not cancel, and the molecule is polar.
TABLE 10.3 Common Cases of Adding Dipole Moments to Determine Whether a Molecule Is Polar

**Nonpolar**
Two identical polar bonds pointing in opposite directions will cancel. The molecule is nonpolar.

**Polar**
Three polar bonds in a trigonal pyramidal arrangement (109.5°) will not cancel. The molecule is polar.

**Nonpolar**
Three identical polar bonds at 120° from each other will cancel. The molecule is nonpolar.

**Polar**
Two polar bonds with an angle of less than 180° between them will not cancel. The molecule is polar.

**Nonpolar**
Four identical polar bonds in a tetrahedral arrangement (109.5° from each other) will cancel. The molecule is nonpolar.

Note: In all cases where the polar bonds cancel, the bonds are assumed to be identical. If one or more of the bonds are different than the other(s), the bonds will not cancel and the molecule is polar.
Electronegativity and Polarity: Why Oil and Water Don’t Mix

• Water molecules are polar; the molecules that compose oil are generally nonpolar.
• Polar molecules interact strongly with other polar molecules because the positive end of one molecule is attracted to the negative end of another, just as the south pole of a magnet is attracted to the north pole of another magnet.
• A mixture of polar and nonpolar molecules is similar to a mixture of small magnetic and nonmagnetic particles. The magnetic particles clump together, excluding the nonmagnetic ones and separating into distinct regions. Similarly, the polar water molecules attract one another, forming regions from which the nonpolar oil molecules are excluded.
Attraction
After eating a greasy cheeseburger, your hands are coated with grease and oil. If you try to wash them with only water, they remain greasy. However, if you add a little soap, the grease washes away. Why?

Water molecules are polar, and the molecules that compose grease and oil are nonpolar. As a result, water and grease repel each other.

One end of a soap molecule is polar, while the other end is nonpolar.

The polar head of a soap molecule strongly attracts water molecules, while the nonpolar tail strongly attracts grease and oil molecules.

Soap allows water and grease to mix, removing the grease from your hands and washing it down the drain.
• Lewis theory: Chemical bonds are formed when atoms transfer valence electrons (ionic bonding) or share valence electrons (covalent bonding) to attain noble gas electron configurations.

• Molecular shapes: The shapes of molecules can be predicted by combining Lewis theory with valence shell electron pair repulsion (VSEPR) theory, where electron groups around the central atom repel one another and determine the geometry of the molecule.
Electronegativity:

- Electronegativity refers to the relative ability of elements to attract electrons within a chemical bond.
- Electronegativity increases as you move to the right across a period in the periodic table and decreases as you move down a column.
Electronegativity and polarity:

• When two nonmetal atoms of different electronegativities form a covalent bond, the electrons in the bond are not evenly shared and the bond is polar.

• In diatomic molecules, a polar bond results in a polar molecule.

• In molecules with more than two atoms, polar bonds may cancel, forming a nonpolar molecule, or they may sum, forming a polar molecule.
Chemical Skills Learning Objectives

1. LO: Write Lewis structures for elements.
2. LO: Write Lewis structures of ionic compounds.
3. LO: Use Lewis theory to predict the chemical formula of an ionic compound.
4. LO: Write Lewis structures for covalent compounds.
5. LO: Write resonance structures.
6. LO: Predict the shapes of molecules.
7. LO: Determine whether a molecule is polar.
Free radicals are molecules that contain an odd number of valence electrons and therefore contain an unpaired electron in their Lewis structure. As you know from Lewis theory, such molecules are not chemically stable and quickly react with other molecules.

Some theories on aging suggest that free radicals cause a variety of diseases and aging.

Free radicals may attack molecules within the cell, such as DNA, changing them and causing cancer or other diseases.

Free radicals may also attack molecules on the surfaces of cells, making them appear foreign to the body’s immune system. The immune system then attacks the cell and destroys it, weakening the body.
Highlight Problem 10.113

• Draw the Lewis structure for each of these free radicals, which have been implicated in theories of aging.
  • $O_2^-$
  • $O^-$
  • OH
  • CH$_3$OO (unpaired electron on terminal oxygen)