Chapter 1 – Structure and Bonding

Chapter Outline

I. Atomic Structure (Sections 1.1–1.3).
   A. Introduction to atomic structure (Section 1.1).
      1. Atoms consist of a dense, positively charged nucleus surrounded by negatively charged electrons.
         a. The nucleus is made up of positively charged protons and uncharged neutrons.
         b. The nucleus contains most of the mass of the atom.
         c. Electrons move about the nucleus at a distance of about $10^{-10}$ m.
      2. The atomic number ($Z$) gives the number of protons in the nucleus.
      3. The mass number ($A$) gives the total number of protons and neutrons.
      4. All atoms of a given element have the same value of $Z$.
         a. Atoms of a given element can have different values of $A$.
         b. Atoms of the same element with different values of $A$ are called isotopes.
   B. Orbitals (Section 1.2).
      1. The distribution of electrons in an atom can be described by a wave equation.
         a. The solution to a wave equation is an orbital, represented by $\Psi$.
         b. $\Psi^2$ predicts the volume of space in which an electron is likely to be found.
      2. There are four different kinds of orbitals ($s$, $p$, $d$, $f$).
         a. The $s$ orbitals are spherical.
         b. The $p$ orbitals are dumbbell-shaped.
         c. Four of the five $d$ orbitals are cloverleaf-shaped.
      3. An atom's electrons are organized into shells.
         a. The shells differ in the numbers and kinds of orbitals they have.
         b. Electrons in different orbitals have different energies.
         c. Each orbital can hold two electrons.
      4. The two lowest-energy electrons are in the $1s$ orbital.
         a. The $2s$ orbital is the next in energy.
         b. Each $p$ orbital has a region of zero density, called a node.
      C. Electron Configuration (Section 1.3).
         1. The ground-state electron configuration of an atom is a listing of the orbitals occupied by the electrons of the atom.
         2. Rules for predicting the ground-state electron configuration of an atom:
            a. Orbitals with the lowest energy levels are filled first.
               The order of filling is $1s$, $2s$, $2p$, $3s$, $3p$, $4s$, $3d$.
            b. Only two electrons can occupy each orbital, and they must be of opposite spin.
            c. If two or more orbitals have the same energy, one electron occupies each until all are half-full (Hund's rule). Only then does a second electron occupy one of the orbitals.
               All of the electrons in a half-filled shell have the same spin.
   II. Chemical Bonding Theory (Sections 1.4–1.5).
      A. Development of chemical bonding theory (Section 1.4).
         1. Kekulé and Couper proposed that carbon has four "affinity units" – carbon is tetravalent.
         2. Other scientists suggested that carbon can form double bonds, triple bonds and rings.
3. Van't Hoff and Le Bel proposed that the 4 atoms to which carbon forms bonds sit at the corners of a regular tetrahedron.

4. In a drawing of a tetrahedral carbon, a wedged line represents a bond pointing toward the viewer, and a dashed line points behind the plane of the page.

B. Covalent bonds.
1. Atoms bond together because the resulting compound is more stable than the individual atoms.
   a. Atoms tend to achieve the electron configuration of the nearest noble gas.
   b. Atoms in groups 1A, 2A and 7A either lose electrons or gain electrons to form ionic compounds.
   c. Atoms in the middle of the periodic table share electrons by forming covalent bonds.
2. The number of covalent bonds formed by an atom depends on the number of electrons it has and on the number it needs to achieve an octet.
3. Covalent bonds can be represented two ways.
   a. In Lewis structures, bonds are represented as pairs of dots.
   b. In line-bond structures, bonds are represented as lines drawn between two atoms.
4. Valence electrons not used for bonding are called lone-pair electrons.
   a. Lone-pair electrons are represented as dots.

C. Valence bond theory (Section 1.5).
1. Covalent bonds are formed by the overlap of two atomic orbitals, each of which contains one electron. The two electrons have opposite spins.
2. Each of the bonded atoms retains its atomic orbitals, but the electron pair of the overlapping orbitals is shared by both atoms.
3. The greater the orbital overlap, the stronger the bond.
4. Bonds formed by the head-on overlap of two atomic orbitals are cylindrically symmetrical and are called \( \sigma \) bonds.
5. Bond strength is the measure of the amount of energy needed to break a bond.
6. Bond length is the optimum distance between nuclei.
7. Every bond has a characteristic bond length and bond strength.

III. Hybridization (Sections 1.6–1.10).
A. \( sp^3 \) Orbitals (Sections 1.6, 1.7).
1. Structure of methane (Section 1.6).
   a. When carbon forms 4 bonds with hydrogen, one 2s orbital and three 2p orbitals combine to form four equivalent atomic orbitals (\( sp^3 \) hybrid orbitals).
   b. These orbitals are tetrahedrally oriented.
   c. Because these orbitals are unsymmetrical, they can form stronger bonds than unhybridized orbitals can.
   d. These bonds have a specific geometry and a bond angle of 109.5°.
2. Structure of ethane (Section 1.7).
   a. Ethane has the same type of hybridization as occurs in methane.
   b. The C–C bond is formed by overlap of two \( sp^3 \) orbitals.
   c. Bond lengths, strengths and angles are very close to those of methane.
B. \( sp^2 \) Orbitals (Section 1.8).
1. If one carbon 2s orbital combines with two carbon 2p orbitals, three hybrid \( sp^2 \) orbitals are formed, and one p orbital remains unchanged.
2. The three \( sp^2 \) orbitals lie in a plane at angles of 120°, and the p orbital is perpendicular to them.
3. Two different types of bonds form between two carbons.
   a. A \( \sigma \) bond forms from the overlap of two \( sp^2 \) orbitals.
   b. A \( \pi \) bond forms by sideways overlap of two \( p \) orbitals.
   c. This combination is known as a carbon–carbon double bond.
4. Ethylene is composed of a carbon–carbon double bond and four \( \sigma \) bonds formed between the remaining four \( sp^2 \) orbitals of carbon and the 1s orbitals of hydrogen.
   a. The double bond of ethylene is both shorter and stronger than the C–C bond of ethane.

C. \( sp \) Orbitals (Section 1.9).
   1. If one carbon 2s orbital combines with one carbon 2p orbital, two hybrid \( sp \) orbitals are formed, and two \( p \) orbitals are unchanged.
   2. The two \( sp \) orbitals are 180° apart, and the two \( p \) orbitals are perpendicular to them and to each other.
   3. Two different types of bonds form.
      a. A \( \sigma \) bond forms from the overlap of two \( sp \) orbitals.
      b. Two \( \pi \) bonds form by sideways overlap of four \( p \) orbitals.
      c. This combination is known as a carbon–carbon triple bond.
   4. Acetylene is composed of a carbon–carbon triple bond and two \( \sigma \) bonds formed between the remaining two \( sp \) orbitals of carbon and the 1s orbitals of hydrogen.
      a. The triple bond of acetylene is the strongest carbon–carbon bond.

D. Hybridization of nitrogen, oxygen, phosphorus and sulfur (Section 1.10).
   1. Covalent bonds between other elements can be described by using hybrid orbitals.
   2. Both the nitrogen atom in ammonia and the oxygen atom in water form \( sp^3 \) hybrid orbitals.
      a. The lone-pair electrons in these compounds occupy \( sp^3 \) orbitals.
   3. The bond angles between hydrogen and the central atom is often less than 109° because the lone-pair electrons take up more room than the \( \sigma \) bond.
   4. Because of their positions in the third row, phosphorus and sulfur can form more than the typical number of covalent bonds.

IV. Molecular orbital theory (Section 1.11).
   A. Molecular orbitals arise from a mathematical combination of atomic orbitals and belong to the entire molecule.
   1. Two 1s orbitals can combine in two different ways.
      a. The additive combination is a bonding MO and is lower in energy than the two hydrogen 1s atomic orbitals.
      b. The subtractive combination is an antibonding MO and is higher in energy than the two hydrogen 1s atomic orbitals.
   2. A node is a region between nuclei in which there is zero probability of finding an electron.
      a. If a node occurs between two nuclei, the nuclei repel each other.
   3. The number of MOs in a molecule is the same as the number of atomic orbitals combined.

V. Chemical structures (Section 1.12).
   A. Drawing chemical structures.
   1. Condensed structures don’t show C–H bonds and don’t show the bonds between \( \text{CH}_3 \), \( \text{CH}_2 \) and \( \text{CH} \) units.
   2. Skeletal structures are simpler still.
      a. Carbon atoms aren’t usually shown.
      b. Hydrogen atoms bonded to carbon aren’t usually shown.
      c. Other atoms are shown.
   3. Groupings such as —CH\(_3\) or —OH are sometimes written as H\(_3\)C— or HO— in order to show connectivity.
Solutions to Problems

1.1 (a) To find the ground-state electron configuration of an element, first locate its atomic number. For oxygen, the atomic number is 8; oxygen thus has 8 protons and 8 electrons. Next, assign the electrons to the proper energy levels, starting with the lowest level. Fill each level completely before assigning electrons to a higher energy level.

Notice that the 2p electrons are in different orbitals. According to *Hund’s rule*, we must place one electron into each orbital of the same energy level until all orbitals are half-filled.

\[
\begin{align*}
2p & \quad \uparrow \quad \uparrow \\
Oxygen & \quad 2s \quad \uparrow \\
1s & \quad \uparrow \\
\end{align*}
\]

Remember that only two electrons can occupy the same orbital, and that they must be of opposite spin.

A different way to represent the ground-state electron configuration is to simply write down the occupied orbitals and to indicate the number of electrons in each orbital. For example, the electron configuration for oxygen is \(1s^2 \ 2s^2 \ 2p^4\).

(b) Phosphorus, with an atomic number of 15, has 15 electrons. Assigning these to energy levels:

\[
\begin{align*}
3p & \quad \uparrow \quad \uparrow \quad \uparrow \\
3s & \quad \uparrow \\
Phosphorus & \quad 2p \quad \uparrow \quad \uparrow \quad \uparrow \quad \uparrow \\
2s & \quad \uparrow \\
1s & \quad \uparrow \\
\end{align*}
\]

The more concise way to represent ground-state electron configuration for phosphorus:

\[
1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^3
\]

(c) \(1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^4\)

\[
\begin{align*}
3p & \quad \uparrow \quad \uparrow \quad \uparrow \\
3s & \quad \uparrow \\
Sulfur & \quad 2p \quad \uparrow \quad \uparrow \quad \uparrow \quad \uparrow \\
2s & \quad \uparrow \\
1s & \quad \uparrow \\
\end{align*}
\]
The elements of the periodic table are organized into groups that are based on the number of outer-shell electrons each element has. For example, an element in group 1A has one outer-shell electron, and an element in group 5A has five outer-shell electrons. To find the number of outer-shell electrons for a given element, use the periodic table to locate its group.

(a) Magnesium (group 2A) has two electrons in its outermost shell.

(b) Cobalt is a transition metal, which has two electrons in the 4s subshell, plus seven electrons in its 3d subshell.

(c) Selenium (group 6A) has six electrons in its outermost shell.

A solid line represents a bond lying in the plane of the page, a wedged bond represents a bond pointing out of the plane of the page toward the viewer, and a dashed bond represents a bond pointing behind the plane of the page.

Identify the group of the central element to predict the number of covalent bonds the element can form.

(a) Carbon (Group 4A) has four electrons in its valence shell and forms four bonds to achieve the noble-gas configuration of neon. A likely formula is CH₂Cl₂.

(b) Sulfur belongs to Group 6A and needs to share two electrons in order to make an octet. One electron is shared with carbon and the other is shared with a hydrogen. A likely formula is CH₃SH.

(c) CH₃NH₂ (see Problem 1.6c for the structure). Nitrogen (Group 5A) shares two electrons with hydrogen and one electron with carbon in order to achieve an electron octet.
1.6 Start by drawing the electron-dot structure of the molecule.

(1) Determine the number of valence, or outer-shell electrons for each atom in the molecule. For ethanol, we know that carbon has four valence electrons, hydrogen has one valence electron, and oxygen has six valence electrons.

\[
\begin{align*}
\cdot & \cdot 4 \times 2 = 8 \\
\cdot & \cdot 1 \times 6 = 6 \\
\cdot & \cdot 6 \times 1 = 6 \\
\end{align*}
\]

\[
\text{total valence electrons} = 20
\]

(2) Next, use two electrons for each single bond.

\[
\begin{align*}
\text{CH}_3\text{CH}_2\text{OH} & \\
\text{H} & \cdot \cdot \cdot \cdot \cdot \\
\text{H}:\cdot & \cdot \cdot \cdot \cdot :\text{O}:\text{H} \\
\text{H} & \cdot \cdot \\
\end{align*}
\]

(3) Finally, use the remaining electrons to achieve an noble gas configuration for all atoms. To convert to a line-bond structure, replace the bonding electron pairs with a line.

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Electron-dot structure</th>
<th>Line-bond structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) CH$_3$CH$_2$OH</td>
<td>[Diagram]</td>
<td>[Diagram]</td>
</tr>
<tr>
<td>(b) H$_2$S</td>
<td>[Diagram]</td>
<td>[Diagram]</td>
</tr>
<tr>
<td>(c) CH$_3$NH$_2$</td>
<td>[Diagram]</td>
<td>[Diagram]</td>
</tr>
<tr>
<td>(d) N(CH$_3$)$_3$</td>
<td>[Diagram]</td>
<td>[Diagram]</td>
</tr>
</tbody>
</table>

1.7 Each of the two carbons has 4 valence electrons. Two electrons are used to form the carbon–carbon bond, and the 6 electrons that remain can form bonds with a maximum of 6 hydrogens. Thus, the formula C$_2$H$_7$ is not possible.
1.8 Connect the carbons and add hydrogens so that all carbons are bonded to four different atoms.

![Propane](attachment:image1.png)

The geometry around all carbon atoms is tetrahedral, and all bond angles are approximately 109°.

1.9

![Hexane](attachment:image2.png)

1.10

![Propene](attachment:image3.png)

The C3–H bonds are σ bonds formed by overlap of an \( sp^3 \) orbital of carbon 3 with an \( s \) orbital of hydrogen.

The C2–H and C1–H bonds are \( \sigma \) bonds formed by overlap of an \( sp^2 \) orbital of carbon with an \( s \) orbital of hydrogen.

The C2–C3 bond is a \( \sigma \) bond formed by overlap of an \( sp^3 \) orbital of carbon 3 with an \( sp^2 \) orbital of carbon 2.

There are two C1–C2 bonds. One is a \( \sigma \) bond formed by overlap of an \( sp^2 \) orbital of carbon 1 with an \( sp^2 \) orbital of carbon 2. The other is a \( \pi \) bond formed by overlap of a \( p \) orbital of carbon 1 with a \( p \) orbital of carbon 2. All four atoms connected to the carbon-carbon double bond lie in the same plane, and all bond angles between these atoms are approximately 120°. The bond angle between hydrogen and the \( sp^3 \)-hybridized carbon is approximately 109°.

1.11

![Buta-1,3-diene](attachment:image4.png)

All atoms lie in the same plane, and all bond angles are approximately 120°.
1.12

Aspirin.

All carbons are $sp^2$ hybridized, with the exception of the indicated carbon. All oxygen atoms have two lone pairs of electrons.

1.13

Propyne

The $C_3$-H bonds are $\sigma$ bonds formed by overlap of an $sp^3$ orbital of carbon 3 with an $s$ orbital of hydrogen.

The $C_1$-H bond is a $\sigma$ bond formed by overlap of an $sp$ orbital of carbon 1 with an $s$ orbital of hydrogen.

The $C_2$-$C_3$ bond is a $\sigma$ bond formed by overlap of an $sp$ orbital of carbon 2 with an $sp^3$ orbital of carbon 3.

There are three $C_1$-$C_2$ bonds. One is a $\sigma$ bond formed by overlap of an $sp$ orbital of carbon 1 with an $sp$ orbital of carbon 2. The other two bonds are $\pi$ bonds formed by overlap of two $p$ orbitals of carbon 1 with two $p$ orbitals of carbon 2.

The three carbon atoms of propyne lie in a straight line; the bond angle is 180°. The $\text{H}–\text{C}_1\equiv\text{C}_2$ bond angle is also 180°. The bond angle between hydrogen and the $sp^3$-hybridized carbon is approximately 109°.

1.14

(a) The $sp^3$-hybridized oxygen atom has tetrahedral geometry. The lone pairs occupy $sp^3$ orbitals.

(b) Both nitrogen and carbon have tetrahedral geometry. The nitrogen lone pair lies in an $sp^3$ orbital.

(c) Like nitrogen, phosphorus has five outer-shell electrons. $\text{PH}_3$ has tetrahedral geometry, and its lone-pair electrons occupy an $sp^3$ orbital.

(d) The $sp^3$-hybridized sulfur atom has tetrahedral geometry, and the sulfur lone-pair electrons lie in an $sp^3$ orbital.
1.15 Remember that the end of a line represents a carbon atom with 3 hydrogens, a two-way intersection represents a carbon atom with 2 hydrogens, a three-way intersection represents a carbon with 1 hydrogen and a four-way intersection represents a carbon with no hydrogens.

(a) Adrenaline – C$_9$H$_{13}$NO$_3$

(b) Estrone – C$_{18}$H$_{22}$O$_2$

1.16 Several possible skeletal structures are possible for each molecular formula.

(a) C$_5$H$_{12}$

(b) C$_2$H$_7$N

(c) C$_3$H$_6$O

(d) C$_4$H$_9$Cl

1.17 PABA
Visualizing Chemistry

1.18

(a)

(b)

1.19 Citric acid ($C_6H_8O_7$) contains seven oxygen atoms, each of which has two electron lone pairs. Three of the oxygens form double bonds with carbon.

1.20 All carbons are $sp^2$ hybridized, except for the carbon indicated as $sp^3$. The two oxygen atoms and the nitrogen atom have lone pair electrons, as shown.
Additional Problems

Electron Configuration

1.22

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Number of valence electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Zinc</td>
<td>30</td>
<td>2</td>
</tr>
<tr>
<td>(b) Iodine</td>
<td>53</td>
<td>7</td>
</tr>
<tr>
<td>(c) Silicon</td>
<td>14</td>
<td>4</td>
</tr>
<tr>
<td>(d) Iron</td>
<td>26</td>
<td>2 (in 4s subshell), 6 (in 3d subshell)</td>
</tr>
</tbody>
</table>

1.23

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Ground-state Electron configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Potassium</td>
<td>19</td>
<td>1s² 2s² 2p⁴ 3s² 3p⁶ 4s¹</td>
</tr>
<tr>
<td>(b) Arsenic</td>
<td>33</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p³</td>
</tr>
<tr>
<td>(c) Aluminum</td>
<td>13</td>
<td>1s² 2s² 2p⁶ 3s² 3p¹</td>
</tr>
<tr>
<td>(d) Germanium</td>
<td>32</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p²</td>
</tr>
</tbody>
</table>

Electron-Dot and Line-Bond Structures

1.24

(a) NH₂OH (b) AlCl₃ (c) CF₂Cl₂ (d) CH₂O

1.25

(a) The 4 valence electrons of carbon can form bonds with a maximum of 4 hydrogens. Thus, it is not possible for the compound CH₅ to exist.

(b) If you try to draw a molecule with the formula C₂H₆N, you will see that it is impossible for both carbons and nitrogen to have a complete octet of electrons. Therefore, C₂H₆N is unlikely to exist.

(c) A compound with the formula C₃H₅Br₂ doesn't have filled outer shells for all atoms and is thus unlikely to exist.
1.26

\[
\begin{align*}
\text{H} & \\
\text{H:}:\text{C}:\cdot\cdot\cdot\text{N}: \\
\text{H}
\end{align*}
\]

Acetonitrile

In the compound acetonitrile, nitrogen has eight electrons in its outer electron shell. Six are used in the carbon-nitrogen triple bond, and two are a nonbonding electron pair.

1.27

\[
\begin{align*}
\text{H} & \\
\cdot\cdot\cdot\cdot\text{Cl} & \\
\text{C}=\text{C} & \\
\text{H} & \\
\text{H}
\end{align*}
\]

Vinyl chloride

Vinyl chloride has 18 valence electrons. Eight electrons are used for 4 single bonds, 4 electrons are used in the carbon-carbon double bond, and 6 electrons are in the 3 lone pairs that surround chlorine.

1.28

(a) 
\[
\begin{align*}
\text{H}_3\text{C} & \\
\text{S} & \\
\text{S} & \\
\text{CH}_3
\end{align*}
\]
(b) 
\[
\begin{align*}
\text{H}_3\text{C} & \\
\text{C} & \\
\text{N} & \\
\text{H}_2
\end{align*}
\]
(c) 
\[
\begin{align*}
\text{H}_3\text{C} & \\
\text{C} & \\
\text{O} & \\
\text{O}
\end{align*}
\]

1.29

In molecular formulas of organic molecules, carbon is listed first, followed by hydrogen. All other elements are listed in alphabetical order.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Molecular Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) Aspirin</td>
<td>C$_9$H$_8$O$_4$</td>
</tr>
<tr>
<td>(b) Vitamin C</td>
<td>C$_6$H$_8$O$_6$</td>
</tr>
<tr>
<td>(c) Nicotine</td>
<td>C$<em>{10}$H$</em>{14}$N$_2$</td>
</tr>
<tr>
<td>(d) Glucose</td>
<td>C$<em>6$H$</em>{12}$O$_6$</td>
</tr>
</tbody>
</table>

1.30

To work this type of problem, you must list all possible structures consistent with the rules of valence. You must systematically consider all possible attachments, including those that have branches, rings and multiple bonds.

(a) 
\[
\begin{align*}
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{C} & \\
\text{C} & \\
\text{C} & \\
\text{H}
\end{align*}
\]
(b) 
\[
\begin{align*}
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{C} & \\
\text{N} & \\
\text{H}
\end{align*}
\]
(c) 
\[
\begin{align*}
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{C} & \\
\text{C} & \\
\text{O} & \\
\text{H}
\end{align*}
\]
(d) 
\[
\begin{align*}
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{C} & \\
\text{C} & \\
\text{C} & \\
\text{Br}
\end{align*}
\]

(d) 
\[
\begin{align*}
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{H} & \\
\text{C} & \\
\text{C} & \\
\text{C} & \\
\text{Br}
\end{align*}
\]
Hybridization

1.35 The H$_3$C– carbon is $sp^3$ hybridized, and the –CN carbon is $sp$ hybridized.
1.36

(a) \( sp^3 \) \( sp^3 \) \( sp^3 \) 
\( \text{CH}_3\text{CH}_2\text{CH}_3 \)

(b) \( sp^3 \) \( sp^2 \) \( sp^2 \) 
\( C\equiv\text{CH} \)

(c) \( sp^2 \) \( sp^2 \) \( sp \) \( sp \) 
\( \text{H}_2\text{C}\equiv\text{CH} \equiv \text{C} \equiv \text{CH} \)

(d) \( sp^3 \) \( sp^2 \) \( sp^2 \)
\( \text{CH}_3\text{CO}_2\text{H} \)

1.37

\( \text{Benzene} \)

All carbon atoms of benzene are \( sp^2 \) hybridized, and all bond angles of benzene are 120°. Benzene is a planar molecule.

1.38

(a) \( sp^2 \) \( \text{H}_2\text{C} \) \( \text{C} \) \( \text{O} \) \( \text{O} \) \( \text{H} \)
\( \text{NH}_2 \)

(b) \( sp^2 \) \( \text{H}_2\text{C} \) \( \text{C} \) \( \text{N} \) \( \text{H} \)

(c) \( sp^2 \) \( \text{H}_3\text{C} \)
\( \text{O} \) \( \text{O} \) \( \text{OH} \)

\( \text{Glycine} \) \( \text{Pyridine} \) \( \text{Lactic acid} \)

1.39 Examples:

(a) \( \text{CH}_3\text{CH}_2\text{CH} \equiv \text{CH}_2 \) (b) \( \text{H}_2\text{C} \equiv \text{CH} \) \( \text{CH} \equiv \text{CH}_2 \) (c) \( \text{H}_2\text{C} \equiv \text{CH} \) \( \text{C} \equiv \text{CH} \)

1.40

(a) \( sp^2 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^2 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) 
\( \text{H}_2\text{N} \) \( \text{C} \) \( \text{C} \) \( \text{C} \) \( \text{C} \) \( \text{H}_2\text{N} \)

(b) \( sp^3 \) \( sp^2 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) \( sp^3 \) 
\( \text{CH}_2\text{OH} \) \( \text{HO} \) \( \text{O} \) \( \text{C} \)

\( \text{Procaine} \) \( \text{Vitamin C} \)
The bond angles formed by atoms having $sp^3$ hybridization are approximately 109°.

The bond angles formed by atoms having $sp^2$ hybridization are approximately 120°.

Skeletal Structures

1.42

(a)  
(b)  
(c)  
(d)  

1.43

(a)  
(b)  
(c)  

$C_{10}H_{11}N$  
$C_{11}H_{11}BrO_2$  
$C_{9}H_{12}O$
1.44

Quetiapine (Seroquel) \( \text{C}_{21}\text{H}_{25}\text{N}_{3}\text{O}_{2}\text{S} \)

1.45

Oseltamivir (Tamiflu) \( \text{C}_{16}\text{H}_{28}\text{N}_{2}\text{O}_{4} \)

Clopidogrel (Plavix) \( \text{C}_{16}\text{H}_{16}\text{ClNO}_{2}\text{S} \)

General Problems

1.46 In a compound containing a carbon–carbon triple bond, atoms bonded to the \( sp \)-hybridized carbons must lie in a straight line. It is not possible to form a five-membered ring if four carbons must have a linear relationship.
The central carbon of allene forms two σ bonds and two π bonds. The central carbon is \( sp \)-hybridized, and the two terminal carbons are \( sp^2 \)-hybridized. The bond angle formed by the three carbons is 180°, indicating linear geometry for the carbons of allene.

Carbon dioxide is a linear molecule.

All of the indicated atoms are \( sp^2 \)-hybridized.

(a) The positively charged carbon atom is surrounded by six valence shell electrons; carbon has three valence electrons, and each hydrogen brings three valence electrons.
(b) The positively charged carbon is \( sp^2 \)-hybridized.
(c) A carbocation is planar about the positively charged carbon.
1.51

(a) A carbanion is isoelectronic with (has the same number of electrons as) a trivalent nitrogen compound.
(b) The negatively charged carbanion carbon is surrounded by eight valence electrons.
(c) The carbon atom is $sp^3$-hybridized.
(d) A carbanion is tetrahedral.

1.52 According to the Pauli Exclusion Principle, two electrons in the same orbital must have opposite spins. Thus, the two electrons of triplet (spin-unpaired) methylene must occupy different orbitals. In singlet (spin-paired) methylene the two electrons can occupy the same orbital because they have opposite spins. Including the two C–H bonds, there are a total of three occupied orbitals. We predict $sp^2$ hybridization for singlet methylene. In triplet methylene, there are two possibilities. In the first, $sp$-hybridized carbon forms one bond to each of two hydrogens. Each of the two unpaired electrons occupies a $p$ orbital. In the other possibility, carbon is $sp^2$ hybridized. One electron lies in an $sp^2$ orbital and the other electron occupies a $p$ orbital. Triplet methylene does not show $sp^3$ hybridization.

1.53

The two compounds differ in the way that the carbon atoms are connected.

1.54

One compound has a double bond, and the other has a ring.
1.55

\[ \text{CH}_3\text{CH}_2\text{OH} \quad \text{CH}_3\text{OCH}_3 \]

The two compounds differ in the location of the oxygen atom.

1.56

\[ \text{CH}_3\text{CH}_2\text{CH}==\text{CH}_2 \quad \text{CH}_3\text{CH}==\text{CH}_3 \]

\[ \text{H}_2\text{C}==\text{C} \]

The compounds differ in the way that the carbon atoms are connected and in the location of the double bond.

1.57

\[ \begin{array}{c}
\text{Ibuprofen} \\
\text{Naproxen} \\
\text{Acetaminophen}
\end{array} \]

* = \( \text{sp}^3\)-hybridized carbon
# = \( \text{sp}^2\)-hybridized carbon

(a), (b)

<table>
<thead>
<tr>
<th>Compound</th>
<th>( \text{sp}^3)-Hybridized carbons</th>
<th>( \text{sp}^2)-Hybridized carbons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ibuprofen</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>Naproxen</td>
<td>3</td>
<td>11</td>
</tr>
<tr>
<td>Acetaminophen</td>
<td>1</td>
<td>7</td>
</tr>
</tbody>
</table>

(c) Each of the structures has at least one six-membered ring containing three double bonds, and each has a C=O group.