Chapter 1
A Review of General Chemistry: Electrons, Bonds and Molecular Properties

Review of Concepts

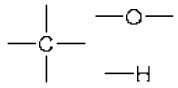
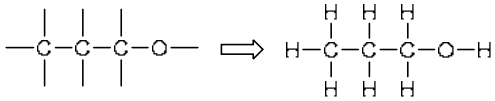
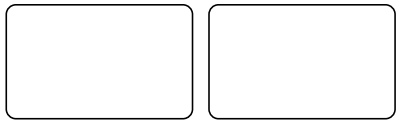
Fill in the blanks below. To verify that your answers are correct, look in your textbook at the end of Chapter 1. Each of the sentences below appears verbatim in the section entitled *Review of Concepts and Vocabulary*.

- _____ **isomers** share the same molecular formula but have different connectivity of atoms and different physical properties.
- Second-row elements generally obey the _____ **rule**, bonding to achieve noble gas electron configuration.
- A pair of unshared electrons is called a _____.
- A **formal charge** occurs when atoms do not exhibit the appropriate number of _____.
- An **atomic orbital** is a region of space associated with _____, while a **molecular orbital** is associated with _____.
- Methane's tetrahedral geometry can be explained using four degenerate _____-**hybridized orbitals** to achieve its four single bonds.
- Ethylene's planar geometry can be explained using three degenerate _____-**hybridized orbitals**.
- Acetylene's linear geometry is achieved via _____-**hybridized** carbon atoms.
- The geometry of small compounds can be predicted using valence shell electron pair repulsion (**VSEPR**) theory, which focuses on the number of _____ bonds and _____ exhibited by each atom.
- The physical properties of compounds are determined by _____ forces, the attractive forces between molecules.
- **London dispersion forces** result from the interaction between transient _____ and are stronger for larger alkanes due to their larger surface area and ability to accommodate more interactions.

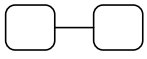
Review of Skills

Fill in the blanks and empty boxes below. To verify that your answers are correct, look in your textbook at the end of Chapter 1. The answers appear in the section entitled *SkillBuilder Review*.

SkillBuilder 1.1 Drawing Constitutional Isomers of Small Molecules

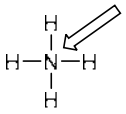
Step 1 Determine the valency of each atom in C_3H_6O.	Step 2 Connect the atoms of highest valency, and place the monovalent atoms at the periphery.	Step 3 Consider other ways to connect the atoms.
		

SkillBuilder 1.2 Drawing the Lewis Structure of a Small Molecule

Step 1 Draw the Lewis dot structure of each atom in C_2H_6O.	Step 2 First connect atoms that form more than one bond.	Step 3 Connect the hydrogen atoms.	Step 4 Pair any unpaired electrons, so that each atom achieves an octet.
			

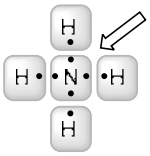
SkillBuilder 1.3 Calculating Formal Charge

Step 1 Determine the appropriate number of valence electrons



Nitrogen is in Group _____ of the periodic table, and is expected to have _____ valence electrons

Step 2 Determine the number of valence electrons in this case.

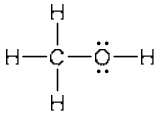


In this case, the nitrogen atom exhibits only _____ valence electrons

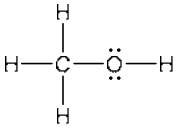
Step 3 Assign a formal charge to the nitrogen atom.

SkillBuilder 1.4 Locating Partial Charges Resulting from Induction

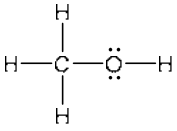
Step 1 Circle the bonds below that are polar covalent



Step 2 For each polar covalent bond, draw an arrow that shows the direction of the dipole moment.

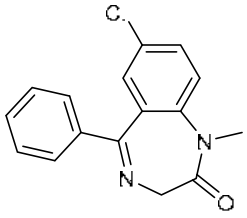


Step 3 Indicate the location of all partial charges (δ+ and δ–).

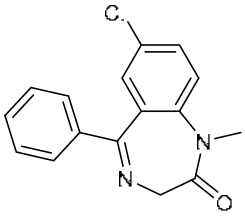


SkillBuilder 1.5 Reading Bond-Line Structures

Circle all carbon atoms in the compound below

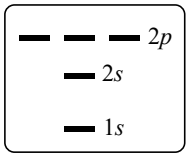


Draw all hydrogen atoms in the compound below



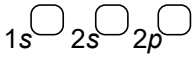
SkillBuilder 1.6 Identifying Electron Configurations

Step 1 In the energy diagram shown here, draw the electron configuration of nitrogen (using arrows to represent electrons).



Nitrogen

Step 2 Fill in the boxes below with the numbers that correctly describe the electron configuration of nitrogen.



SkillBuilder 1.7 Identifying Hybridization States

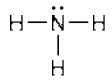
A carbon atom with four single bonds will be _____ hybridized

A carbon atom with one double bond will be _____ hybridized

A carbon atom with a triple bond will be _____ hybridized

SkillBuilder 1.8 Predicting Geometry

Step 1 Determine the steric number of the nitrogen atom here by adding the number of single bonds and lone pairs



# of single bonds = ☐

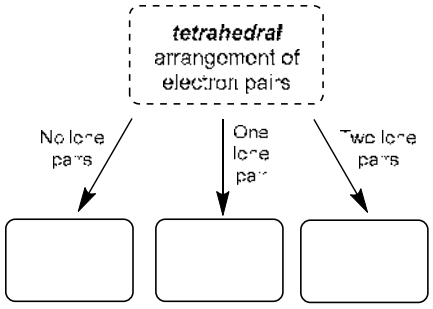
# of lone pairs = ☐

Steric Number = ☐

Step 2 The steric number determines the arrangement of electron pairs. Fill in the chart below:

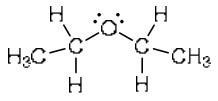
Steric #	Arrangement of Electron Pairs
4	
3	
2	

Step 3 Identify the geometry in each case below:



SkillBuilder 1.9 Identifying the Presence of Molecular Dipole Moments

Step 1 Identify the geometry of the oxygen atom below:



Geometry = ☐

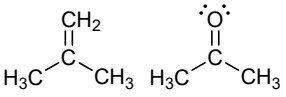
Step 2 Redraw the compound. Draw an arrow for each dipole moment (showing its direction):

Step 3 Redraw the compound, and draw the net dipole moment:

SkillBuilder 1.10 Predicting Physical Properties

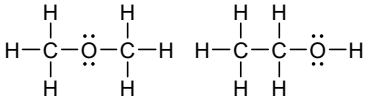
Dipole-Dipole Interactions

Circle the compound below that is expected to have the higher boiling point.



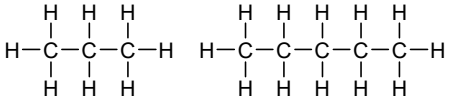
Hydrogen-Bonding Interactions

Circle the compound below that is expected to have the higher boiling point.



Carbon Skeleton

Circle the compound below that is expected to have the higher boiling point.



A Common Mistake to Avoid

When drawing a structure, don't forget to draw formal charges, as forgetting to do so is a common error. If a formal charge is present, it MUST be drawn. For example, in the following case, the nitrogen atom bears a positive charge, so the charge must be drawn:

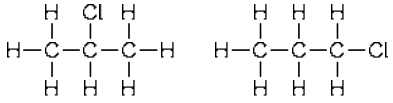
INCORRECT **CORRECT**

As we progress though the course, we will see structures of increasing complexity. If formal charges are present, failure to draw them constitutes an error, and must be scrupulously avoided. If you have trouble drawing formal charges, go back and master that skill. You can't go on without it. Don't make the mistake of underestimating the importance of being able to draw formal charges with confidence.

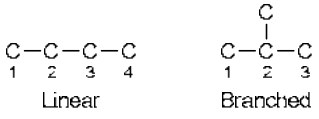
Solutions

1.1.

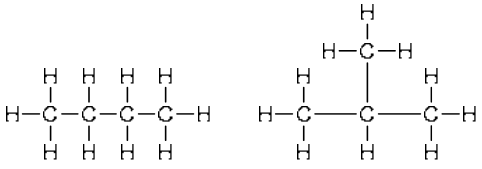
(a) Begin by determining the valency of each atom that appears in the molecular formula. The carbon atoms are tetravalent, while the chlorine atom and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the three carbon atoms) should be drawn in the center of the compound. Then, the chlorine atom can be placed in either of two locations: i) connected to the central carbon atom, or ii) connected to one of the other two (equivalent) carbon atoms. The hydrogen atoms are then placed at the periphery (ensuring that each carbon atom has a total of four bonds). The formula C_3H_7Cl has two constitutional isomers.



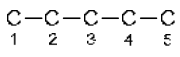
(b) Begin by determining the valency of each atom that appears in the molecular formula. The carbon atoms are tetravalent, while the hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the four carbon atoms) should be drawn in the center of the compound. There are two different ways to connect four carbon atoms. They can either be arranged in a linear fashion or in a branched fashion:



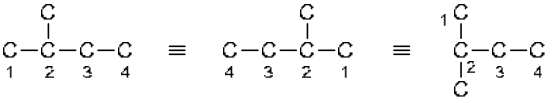
We then place the hydrogen atoms at the periphery (ensuring that each carbon atom has a total of four bonds). The formula C_4H_{10} has two constitutional isomers:



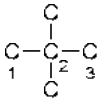
(c) Begin by determining the valency of each atom that appears in the molecular formula. The carbon atoms are tetravalent, while the hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the five carbon atoms) should be drawn in the center of the compound. So we must explore all of the different ways to connect five carbon atoms. First, we can connect all five carbon atoms in a linear fashion:



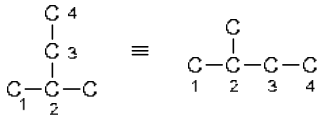
Alternatively, we can draw four carbon atoms in a linear fashion, and then draw the fifth carbon atom on a branch. There are many ways to draw this possibility:



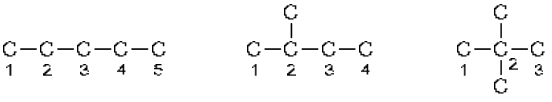
Finally, we can draw three carbon atoms in a linear fashion, and then draw the remaining two carbon atoms on separate branches.



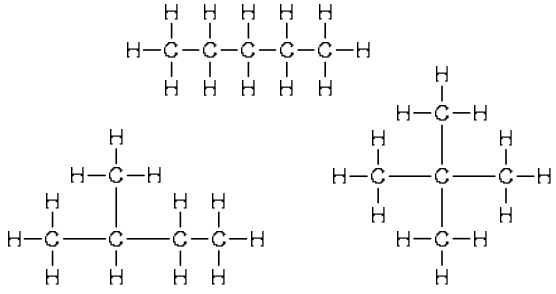
Note that we cannot draw a unique carbon skeleton (a unique arrangement of carbon atoms) simply by placing the last two carbon atoms together as one branch, because that possibility has already been drawn earlier (a linear chain of four carbon atoms with a single branch):



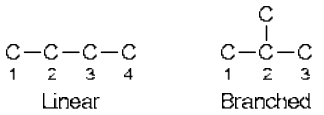
In summary, there are three different ways to connect five carbon atoms:



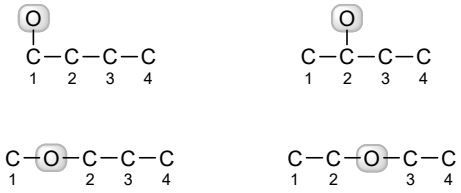
We then place the hydrogen atoms at the periphery (ensuring that each carbon atom has a total of four bonds). The formula C_5H_{12} has three constitutional isomers:



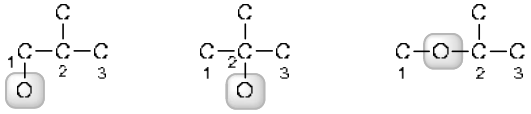
(d) Begin by determining the valency of each atom that appears in the molecular formula. The carbon atoms are tetravalent, the oxygen atom is divalent, and the hydrogen atoms are all monovalent. Any atoms with more than one bond (in this case, the four carbon atoms and the one oxygen atom) should be drawn in the center of the compound, with the hydrogen atoms at the periphery. There are several different ways to connect four carbon atoms and one oxygen atom. Let's begin with the four carbon atoms. There are two different ways to connect four carbon atoms. They can either be arranged in a linear fashion or in a branched fashion.



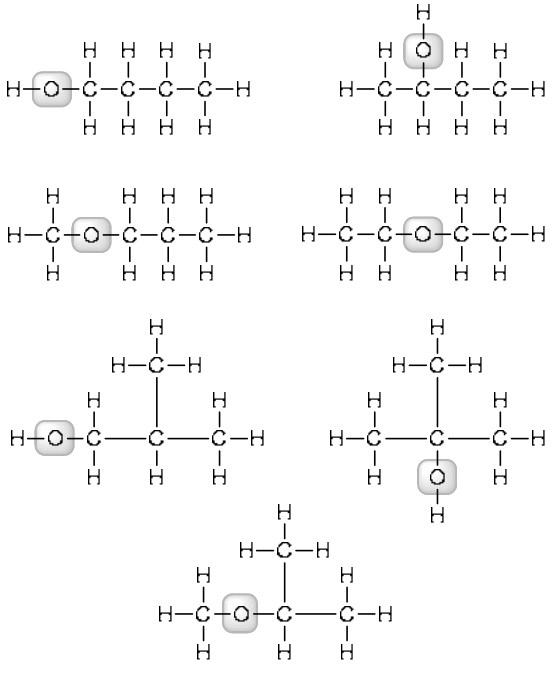
Next, the oxygen atom must be inserted. For each of the two skeletons above (linear or branched), there are several different locations to insert the oxygen atom. The linear skeleton has four possibilities, shown here:



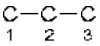
and the branched skeleton has three possibilities shown here:



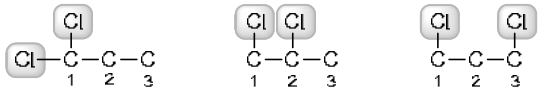
Finally, we complete all of the structures by drawing the bonds to hydrogen atoms (ensuring that each carbon atom has four bonds, and each oxygen atoms has two bonds). The formula $C_4H_{10}O$ has seven constitutional isomers:



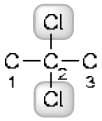
(e) Begin by determining the valency of each atom that appears in the molecular formula. The carbon atoms are tetravalent, while the chlorine atom and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the three carbon atoms) should be drawn in the center of the compound. There is only way to connect three carbon atoms:



Next, we must determine all of the different possible ways of connecting two chlorine atoms to the chain of three carbon atoms. If we place one chlorine atom at C1, then the second chlorine atom can be placed at C1, at C2 or at C3:

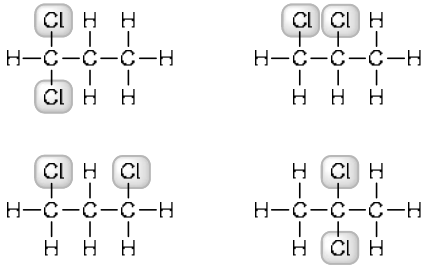


Furthermore, we can place both chlorine atoms at C2, giving a new possibility not shown above:

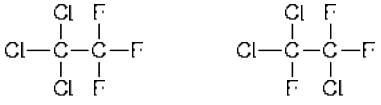


There are no other possibilities. For example, placing the two chlorine atoms at C2 and C3 is equivalent to placing them at C1 and C2:

Finally, the hydrogen atoms are placed at the periphery (ensuring that each carbon atom has a total of four bonds). The formula $C_3H_6Cl_2$ has four constitutional isomers:

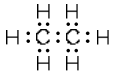


1.2. The carbon atoms are tetravalent, while the chlorine atoms and fluorine atoms are all monovalent. The atoms with more than one bond (in this case, the two carbon atoms) should be drawn in the center of the compound. The chlorine atoms and fluorine atoms are then placed at the periphery, as shown. There are only two possible constitutional isomers: one with the three chlorine atoms all connected to the same carbon, and one in which they are distributed over both carbon atoms. Any other representations that one may draw must be one of these structures drawn in a different orientation.

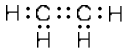


1.3.

(a) Each carbon atom has four valence electrons, and each hydrogen atom has one valence electron. Only the carbon atoms can form more than one bond, so we begin by connecting the carbon atoms to each other. Then, we connect all of the hydrogen atoms, as shown.

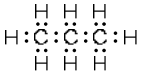


(b) Each carbon atom has four valence electrons, and each hydrogen atom has one valence electron. Only the carbon atoms can form more than one bond, so we begin by connecting the carbon atoms to each other. Then, we connect all of the hydrogen atoms, and the unpaired electrons are shared to give a double bond. In this way, each of the carbon atoms achieves an octet.

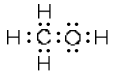


(c) Each carbon atom has four valence electrons, and each hydrogen atom has one valence electron. Only the carbon atoms can form more than one bond, so we begin by connecting the carbon atoms to each other. Then, we connect all of the hydrogen atoms, and the unpaired electrons are shared to give a triple bond. In this way, each of the carbon atoms achieves an octet.

(d) Each carbon atom has four valence electrons, and each hydrogen atom has one valence electron. Only the carbon atoms can form more than one bond, so we begin by connecting the carbon atoms to each other. Then, we connect all of the hydrogen atoms, as shown.

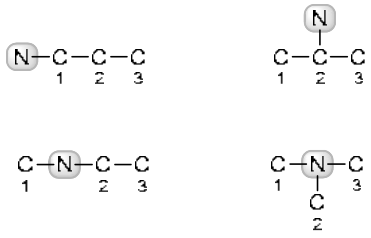


(e) The carbon atom has four valence electrons, the oxygen atom has six valence electrons, and each hydrogen atom has one valence electron. Only the carbon atom and the oxygen atom can form more than one bond, so we begin by connecting them to each other. Then, we connect all of the hydrogen atoms, as shown.

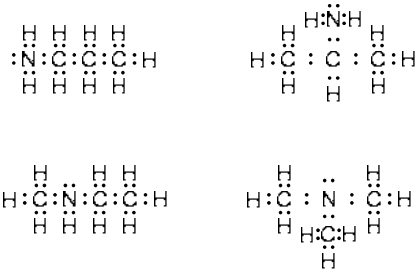


1.4. Boron is in column 3A of the periodic table, so it has three valence electrons. Each of these valence electrons is shared with a hydrogen atom, shown below. The central boron atom lacks an octet of electrons, and it is therefore very unstable and reactive.

1.5. Each of the carbon atoms has four valence electrons; the nitrogen atom has five valence electrons; and each of the hydrogen atoms has one valence electron. We begin by connecting the atoms that have more than one bond (in this case, the three carbon atoms and the nitrogen atom). There are four different ways that these four atoms can be connected to each other, shown here.



For each of these possible arrangements, we connect the hydrogen atoms, giving the following four constitutional isomers.

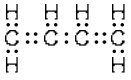


In each of these four structures, the nitrogen atom has one lone pair.

1.6.

(a) The carbon atom has four valence electrons, the nitrogen atom has five valence electrons and the hydrogen atom has one valence electron. Only the carbon atom and the nitrogen atom can form more than one bond, so we begin by connecting them to each other. Then, we connect the hydrogen atom to the carbon, as shown. The unpaired electrons are shared to give a triple bond. In this way, both the carbon atom and the nitrogen atom achieve an octet.

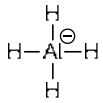
(b) Each carbon atom has four valence electrons, and each hydrogen atom has one valence electron. Only the carbon atoms can form more than one bond, so we begin by connecting the carbon atoms to each other. Then, we connect all of the hydrogen atoms as indicated in the given condensed formula ($CH_2CHCHCH_2$), and the unpaired electrons are shared to give two double bonds on the outermost carbons. In this way, each of the carbon atoms achieves an octet.



1.7.

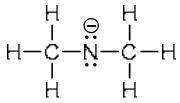
(a) Aluminum is in group 3A of the periodic table, and it should therefore have three valence electrons. In this

case, the aluminum atom exhibits four valence electrons (one for each bond). With one extra electron, this aluminum atom will bear a negative charge.

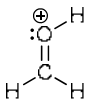


(b) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five valence electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge.

(c) Nitrogen is in group 5A of the periodic table, and it should therefore have five valence electrons. In this case, the nitrogen atom exhibits six valence electrons (one for each bond and two for each lone pair). With one extra electron, this nitrogen atom will bear a negative charge.



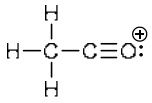
(d) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five valence electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge.



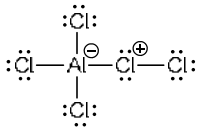
(e) Carbon is in group 4A of the periodic table, and it should therefore have four valence electrons. In this case, the carbon atom exhibits five valence electrons (one for each bond and two for the lone pair). With one extra electron, this carbon atom will bear a negative charge.

(f) Carbon is in group 4A of the periodic table, and it should therefore have four valence electrons. In this case, the carbon atom exhibits only three valence electrons (one for each bond). This carbon atom is missing an electron, and it therefore bears a positive charge.

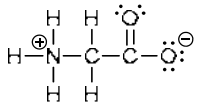
(g) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five valence electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge.



(h) Two of the atoms in this structure exhibit a formal charge because each of these atoms does not exhibit the appropriate number of valence electrons. The aluminum atom (group 3A) should have three valence electrons, but it exhibits four (one for each bond). With one extra electron, this aluminum atom will bear a negative charge. The neighboring chlorine atom (to the right) should have seven valence electrons, but it exhibits only six (one for each bond and two for each lone pair). It is missing one electron, so this chlorine atom will bear a positive charge.

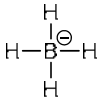


(i) Two of the atoms in this structure exhibit a formal charge because each of these atoms does not exhibit the appropriate number of valence electrons. The nitrogen atom (group 5A) should have five valence electrons, but it exhibits four (one for each bond). It is missing one electron, so this nitrogen atom will bear a positive charge. One of the two oxygen atoms (the one on the right) exhibits seven valence electrons (one for the bond, and two for each lone pair), although it should have only six. With one extra electron, this oxygen atom will bear a negative charge.

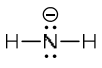


1.8.

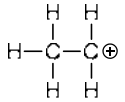
(a) The boron atom in this case exhibits four valence electrons (one for each bond), although boron (group 3A) should only have three valence electrons. With one extra electron, this boron atom bears a negative charge.



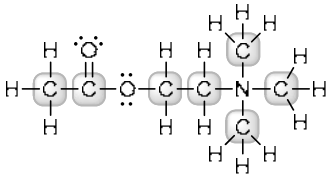
(b) Nitrogen is in group 5A of the periodic table, so a nitrogen atom should have five valence electrons. A negative charge indicates one extra electron, so this nitrogen atom must exhibit six valence electrons (one for each bond and two for each lone pair).



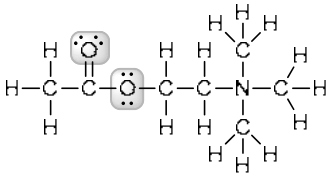
(c) One of the carbon atoms (below right) exhibits three valence electrons (one for each bond), but carbon (group 4A) is supposed to have four valence electrons. It is missing one electron, so this carbon atom therefore bears a positive charge.



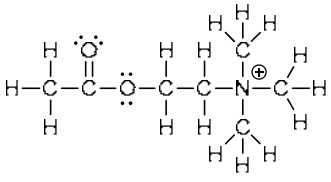
1.9. Carbon is in group 4A of the periodic table, and it should therefore have four valence electrons. Every carbon atom in acetylcholine has four bonds, thus exhibiting the correct number of valence electrons (four) and having no formal charge.



Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. Each oxygen atom in acetylcholine has two bonds and two lone pairs of electrons, so each oxygen atom exhibits six valence electrons (one for each bond, and two for each lone pair). With the correct number of valence electrons, each oxygen atom will lack a formal charge.

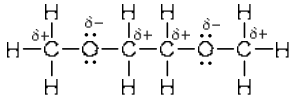


The nitrogen atom (group 5A) should have five valence electrons, but it exhibits four (one for each bond). It is missing one electron, so this nitrogen atom will bear a positive charge.

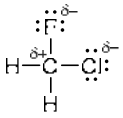


1.10.

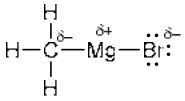
(a) Oxygen is more electronegative than carbon, and a C–O bond is polar covalent. For each C–O bond, the O will be electron-rich (δ–), and the C will be electron-poor (δ+), as shown below.



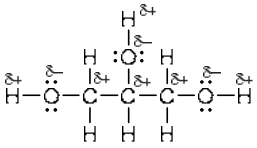
(b) Fluorine is more electronegative than carbon, and a C–F bond is polar covalent. For a C–F bond, the F will be electron-rich (δ–), and the C will be electron-poor (δ+). Chlorine is also more electronegative than carbon, so a C–Cl bond is also polar covalent. For a C–Cl bond, the Cl will be electron-rich (δ–), and the C will be electron-poor (δ+), as shown below.



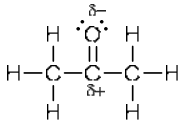
(c) Carbon is more electronegative than magnesium, so the C will be electron-rich (δ–) in a C–Mg bond, and the Mg will be electron-poor (δ+). Also, bromine is more electronegative than magnesium. So in a Mg–Br bond, the Br will be electron-rich (δ–), and the Mg will be electron-poor (δ+), as shown below.



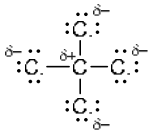
(d) Oxygen is more electronegative than carbon or hydrogen, so all C–O bonds and all O–H bond are polar covalent. For each C–O bond and each O–H bond, the O will be electron-rich (δ–), and the C or H will be electron-poor (δ+), as shown below.



(e) Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ–) and the C will be electron-poor (δ+) in a C=O bond, as shown below.

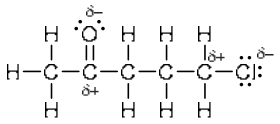


(f) Chlorine is more electronegative than carbon. As such, for each C–Cl bond, the Cl will be electron-rich (δ–) and the C will be electron-poor (δ+), as shown below.



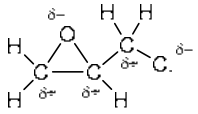
1.11. Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ–) and the C will be electron-poor (δ+) in a C=O bond. In addition, chlorine is more electronegative than carbon. So for a C–Cl bond,

the Cl will be electron-rich (δ–) and the C will be electron-poor (δ+), as shown below.



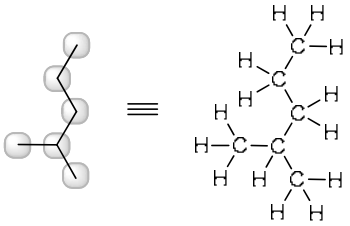
Notice that two carbon atoms are electron-poor (δ+). These are the positions that are most likely to be attacked by an electron-rich anion, such as hydroxide.

1.12. Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ−) and the C will be electron-poor (δ+) in a C─O bond. In addition, chlorine is more electronegative than carbon. So for a C─Cl bond, the Cl will be electron-rich (δ−) and the C will be electron-poor (δ+), as shown below. As you might imagine, epichlorohydrin is a very reactive molecule!

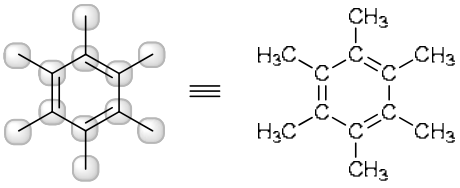


1.13.

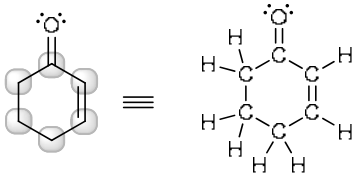
(a) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has six carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



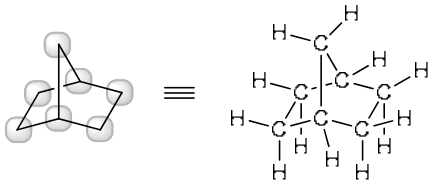
(b) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has twelve carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



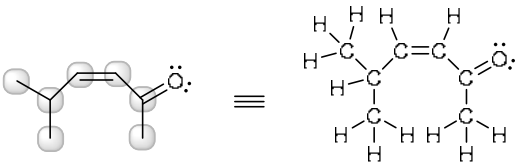
(c) Each corner represents a carbon atom (highlighted below), so this compound has six carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



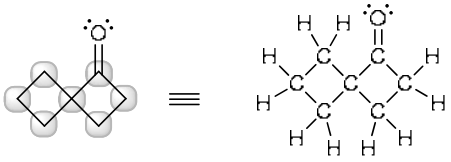
(d) Each corner represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



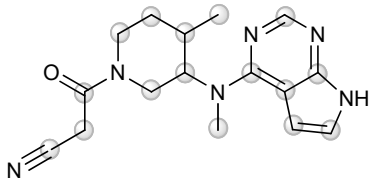
(e) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



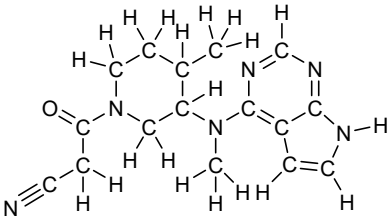
(f) Each corner represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom has enough attached hydrogen atoms to have exactly four bonds, as shown:



1.14. Remember that each corner and each endpoint represents a carbon atom. This compound therefore has 16 carbon atoms, highlighted below:



Each carbon atom should have four bonds. We therefore draw enough hydrogen atoms in order to give each carbon atom a total of four bonds. Any carbon atoms that already have four bonds will not have any hydrogen atoms:



1.15.

(a) As indicated in Figure 1.10, carbon has two *1s* electrons, two *2s* electrons, and two *2p* electrons. This

information is represented by the following electron configuration: $1s^2 2s^2 2p^2$

(b) As indicated in Figure 1.10, oxygen has two *1s* electrons, two *2s* electrons, and four *2p* electrons. This information is represented by the following electron configuration: $1s^2 2s^2 2p^4$

(c) As indicated in Figure 1.10, boron has two *1s* electrons, two *2s* electrons, and one *2p* electron. This information is represented by the following electron configuration: $1s^2 2s^2 2p^1$

(d) As indicated in Figure 1.10, fluorine has two *1s* electrons, two *2s* electrons, and five *2p* electrons. This information is represented by the following electron configuration: $1s^2 2s^2 2p^5$

(e) Sodium has two *1s* electrons, two *2s* electrons, six *2p* electrons, and one *3s* electron. This information is represented by the following electron configuration: $1s^2 2s^2 2p^6 3s^1$

(f) Aluminum has two *1s* electrons, two *2s* electrons, six *2p* electrons, two *3s* electrons, and one *3p* electron. This information is represented by the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$

1.16.
(a) The electron configuration of a carbon atom is $1s^2 2s^2 2p^2$ (see the solution to Problem 1.15a). However, if a carbon atom bears a negative charge, then it must have one extra electron, so the electron configuration should be as follows: $1s^2 2s^2 2p^3$

(b) The electron configuration of a carbon atom is $1s^2 2s^2 2p^2$ (see the solution to Problem 1.15a). However, if a carbon atom bears a positive charge, then it must be missing an electron, so the electron configuration should be as follows: $1s^2 2s^2 2p^1$

(c) As seen in SkillBuilder 1.6, the electron configuration of a nitrogen atom is $1s^2 2s^2 2p^3$**.** However, if a nitrogen atom bears a positive charge, then it must be missing an electron, so the electron configuration should be as follows: $1s^2 2s^2 2p^2$

(d) The electron configuration of an oxygen atom is $1s^2 2s^2 2p^4$ (see the solution to Problem 1.15b). However, if an oxygen atom bears a negative charge, then it must have one extra electron, so the electron configuration should be as follows: $1s^2 2s^2 2p^5$

1.17. Silicon is in the third row, or period, of the periodic table. Therefore, it has a filled second shell, like neon, and then the additional electrons are added to the third shell. As indicated in Figure 1.10, neon has two 1*s* electrons, two 2*s* electrons, and six 2*p* electrons. Silicon has an additional two 3*s* electrons and two 3*p* electrons to give a total of 14 electrons and an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^2$.

1.18. The angles of an equilateral triangle are 60º, but each bond angle of cyclopropane is supposed to be 109.5º. Therefore, each bond angle is severely strained, causing an increase in energy. This form of strain, called ring strain, will be discussed in Chapter 4. The ring strain associated with a three-membered ring is greater than the ring strain of larger rings, because larger rings do not require bond angles of 60º.

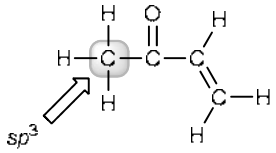
1.19.
(a) The C=O bond of formaldehyde is comprised of one σ bond and one π bond.

(b) Each C–H bond is formed from the interaction between an sp^2-hybridized orbital from carbon and an *s* orbital from hydrogen.

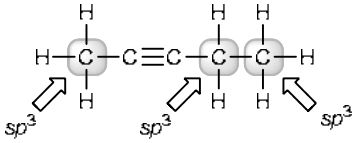
(c) The oxygen atom is sp^2 hybridized, so the lone pairs occupy sp^2-hybridized orbitals.

1.20. Rotation of a single bond does not cause a reduction in the extent of orbital overlap, because the orbital overlap occurs on the bond axis. In contrast, rotation of a π bond results in a reduction in the extent of orbital overlap between the two *p* orbitals, because the orbital overlap is NOT on the bond axis.

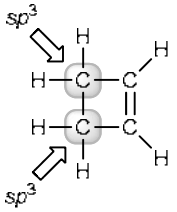
1.21.
(a) The highlighted carbon atom (below) has four σ bonds, and is therefore sp^3 hybridized. The other carbon atoms in this structure are all sp^2 hybridized, because each of them has three σ bonds and one π bond.



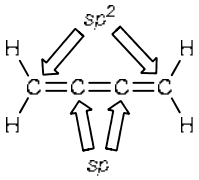
(b) Each of the highlighted carbon atoms has four σ bonds, and is therefore sp^3 hybridized. The other two carbon atoms in this structure are *sp* hybridized, because each has two σ bonds and two π bonds.



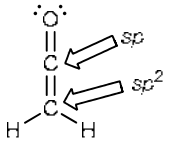
(c) Each of the highlighted carbon atoms (below) has four σ bonds, and is therefore sp^3 hybridized. The other two carbon atoms in this structure are sp^2 hybridized, because each has three σ bonds and one π bond.



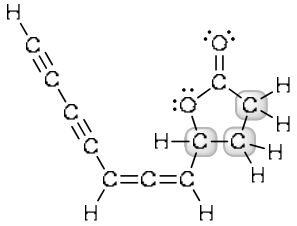
(d) Each of the two central carbon atoms has two σ bonds and two π bonds, and as such, each of these carbon atoms is *sp* hybridized. The other two carbon atoms (the outer ones) are sp^2 hybridized because each has three σ bonds and one π bond.



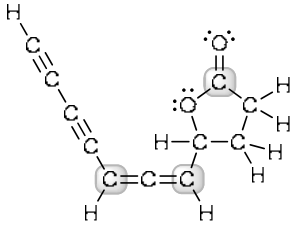
(e) One of the carbon atoms (the one connected to oxygen) has two σ bonds and two π bonds, and as such, it is *sp* hybridized. The other carbon atom is sp^2 hybridized because it has three σ bonds and one π bond.



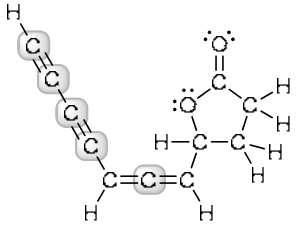
1.22. Each of the following three highlighted carbon atoms has four σ bonds, and is therefore sp^3 hybridized:



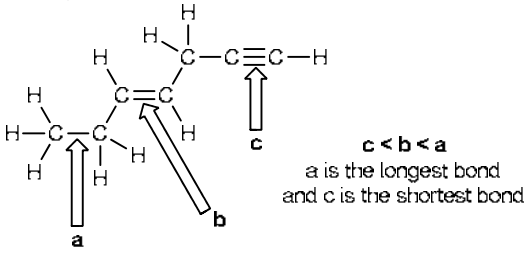
And each of the following three highlighted carbon atoms has three σ bonds and one π bond, and is therefore sp^2 hybridized:



Finally, each of the following five highlighted carbon atoms has two σ bonds and two π bonds, and is therefore *sp* hybridized.



1.23. Carbon-carbon triple bonds generally have a shorter bond length than carbon-carbon double bonds, which are generally shorter than carbon-carbon single bonds (see Table 1.2).



1.24.

(a) In this structure, the boron atom has four σ bonds and no lone pairs, giving a total of four electron pairs (steric number = 4). VSEPR theory therefore predicts a tetrahedral arrangement of electron pairs. Since all of the electron pairs are bonds, the structure is expected to have tetrahedral geometry.

(b) In this structure, the boron atom has three σ bonds and no lone pairs, giving a total of three electron pairs (steric number = 3). VSEPR theory therefore predicts a trigonal planar geometry.

(c) In this structure, the nitrogen atom has four σ sigma bonds and no lone pairs, giving a total of four electron pairs (steric number = 4). VSEPR theory therefore predicts a tetrahedral arrangement of electron pairs. Since all of the electron pairs are bonds, the structure is expected to have tetrahedral geometry.

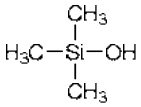
(d) The carbon atom has four σ bonds and no lone pairs, giving a total of four electron pairs (steric number = 4). VSEPR theory therefore predicts a tetrahedral arrangement of electron pairs. Since all of the electron pairs are bonds, the structure is expected to have tetrahedral geometry.

1.25. In the carbocation, the carbon atom has three bonds and no lone pairs. Since there are a total of three electron pairs (steric number = 3), and all three are bonds, VSEPR theory predicts trigonal planar geometry, with bond angles of 120°. In contrast, the carbon atom of the carbanion has three bonds and one lone pair, giving a total of four electron pairs (steric number = 4). For this ion, VSEPR theory predicts a tetrahedral arrangement of electron pairs, with a lone pair positioned at one corner of the tetrahedron, giving rise to trigonal pyramidal geometry with bond angles approximately 107°.

1.26. In ammonia, the nitrogen atom has three bonds and one lone pair. Therefore, VSEPR theory predicts trigonal pyramidal geometry, with bond angles of approximately 107°. In the ammonium ion, the nitrogen atom has four bonds and no lone pairs, so VSEPR theory predicts tetrahedral geometry, with bond angles of 109.5°. Therefore, we predict that the bond angles will increase (by approximately 2.5°) as a result of the reaction.

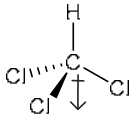
1.27. The silicon atom has four σ bonds and no lone pairs, so the steric number is 4 (sp^3 hybridization), which means

that the arrangement of electron pairs will be tetrahedral. With no lone pairs, the arrangement of the atoms (geometry) is the same as the electronic arrangement. It is tetrahedral.

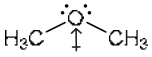


1.28.

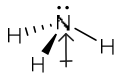
(a) This compound has three C–Cl bonds, each of which exhibits a dipole moment. To determine if these dipole moments cancel each other, we must identify the molecular geometry. The central carbon atom has four σ bonds so we expect tetrahedral geometry. As such, the three polar C–Cl bonds do not lie in the same plane, and they do not completely cancel each other out. There is a net molecular dipole moment, as shown:



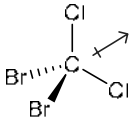
(b) The oxygen atom has two σ bonds and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. As such, the dipole moments associated with the polar C–O bonds do not fully cancel each other, and the dipole moments associated with the lone pairs also do not fully cancel each other. As a result, there is a net molecular dipole moment, as shown:



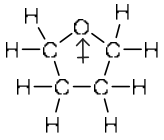
(c) The nitrogen atom has three σ bonds and one lone pair (steric number = 4), and VSEPR theory predicts trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair). As such, the dipole moments associated with the polar N–H bonds do not fully cancel each other, and there is also a dipole moment associated with the lone pair (pointing up). As a result, there is a net molecular dipole moment, as shown:



(d) The central carbon atom has four σ bonds (steric number = 4), and VSEPR theory predicts tetrahedral geometry. There are individual dipole moments associated with each of the C–Cl bonds and each of the C–Br bonds. If all four dipole moments had the same magnitude, then we would expect them to completely cancel each other to give no molecular dipole moment (as in the case of CCl_4). However, because Cl is more electronegative than Br, each C–Cl bond is more polar than each C–Br bond. Therefore, the dipole moments for the C–Cl bonds are larger than the dipole moments of the C–Br bonds, and as such, there is a net molecular dipole moment, shown here:

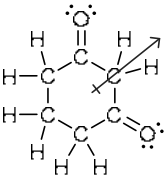


(e) The oxygen atom has two σ bonds and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. As such, the dipole moments associated with the polar C–O bonds do not fully cancel each other, and the dipole moments associated with the lone pairs also do not fully cancel each other. As a result, there is a net molecular dipole moment, as shown:



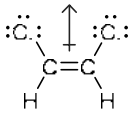
(f) There are individual dipole moments associated with each polar C–O bond and the lone pairs (as in the previous solution), but due to the symmetrical shape of the molecule in this case, they fully cancel each other to give no net molecular dipole moment.

(g) Each C=O bond has a strong dipole moment, and they do not fully cancel each other because they are not pointing in opposite directions. As such, there will be a net molecular dipole moment, as shown here:



(h) Each C=O bond has a strong dipole moment, and in this case, they are pointing in opposite directions. As such, they fully cancel each other, giving no net molecular dipole moment.

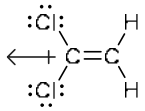
(i) Each C–Cl bond has a dipole moment, and they do not fully cancel each other because the polar bonds are not pointing in opposite directions. As such, there will be a net molecular dipole moment, as shown here:



(j) Each C–Cl bond has a dipole moment, and in this case, they are pointing in opposite directions. As such, they fully cancel each other, giving no net molecular dipole moment.

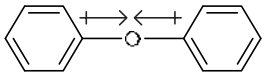
(k) Each C–Cl bond has a dipole moment, and they do not fully cancel each other because they are not pointing

in opposite directions. As such, there will be a net molecular dipole moment, as shown here:

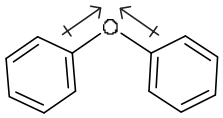


(l) Each C–Cl bond has a dipole moment, but in this case, they fully cancel each other to give no net molecular dipole moment.

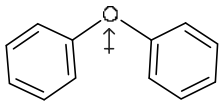
1.29. Each of the C–O bonds has an individual dipole moment, shown here:



To determine if these individual dipole moments fully cancel each other, we must determine the geometry around the oxygen atom. The oxygen atom has two σ bonds and two lone pairs, giving rise to a bent geometry. As such, the dipole moments associated with the polar C–O bonds do NOT fully cancel each other,

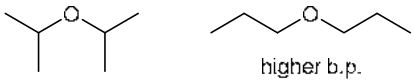


and the dipole moments associated with the lone pairs also do not fully cancel each other. As a result, there is a net molecular dipole moment, as shown:

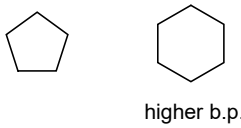


1.30.

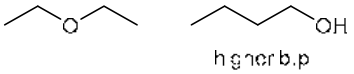
(a) Both compounds have the same molecular weight because they are isomers. The second compound is expected to have a higher boiling point, because it is less branched. The greater surface area in the second compound results in greater London dispersion forces and a higher boiling point (b.p.):



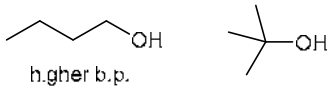
(b) The second compound is expected to have a higher boiling point, because more carbon atoms results in a higher molecular weight. Larger compounds have greater London dispersion forces and a higher boiling point (b.p.):



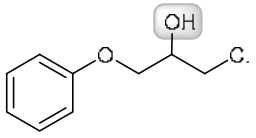
(c) The second compound is expected to have a higher boiling point, because it has an O–H bond, which will lead to hydrogen-bonding interactions between molecules.



(d) Both compounds have the same molecular weight because they are isomers, and they are both capable of forming hydrogen bonds. The first compound is expected to have a higher boiling point, however, because it is less branched. The greater surface area in the first compound results in greater London dispersion forces and a higher boiling point (b.p.):



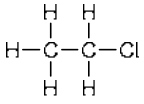
1.31. Compound **3** is expected to have a higher boiling point than compound **4**, because only compound **3** has an O-H group. Compound **4** does not form hydrogen-bonds, so it will have a lower boiling point. When this mixture is heated, the lower boiling compound (**4**) can be collected first, leaving behind compound **3**.



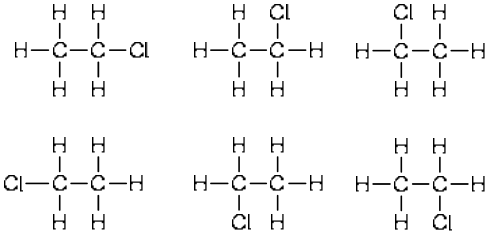
3

1.32.

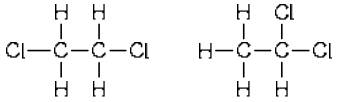
(a) The carbon atoms are tetravalent, while the chlorine atom and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the two carbon atoms) should be drawn in the center of the compound. The chlorine atom and hydrogen atoms are then placed at the periphery (ensuring that each carbon atom has a total of four bonds), as shown:



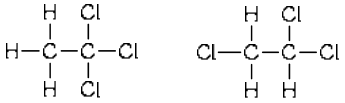
The chlorine atom can be placed in any one of the six available positions. The following six drawings all represent the same compound, in which the two carbon atoms are connected to each other, and the chlorine atom is connected to one of the carbon atoms.



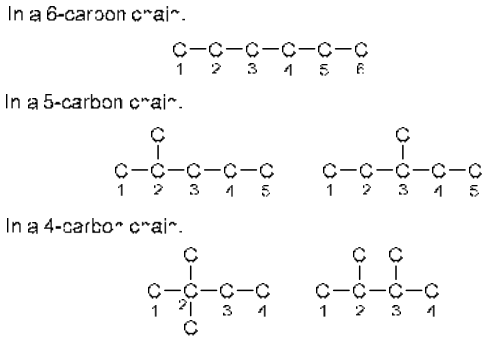
(b) The carbon atoms are tetravalent, while the chlorine atoms and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the two carbon atoms) should be drawn in the center of the compound. The chlorine atoms and hydrogen atoms are then placed at the periphery, and there are two different ways to do this. The two chlorine atoms can either be connected to the same carbon atom or to different carbon atoms, as shown.



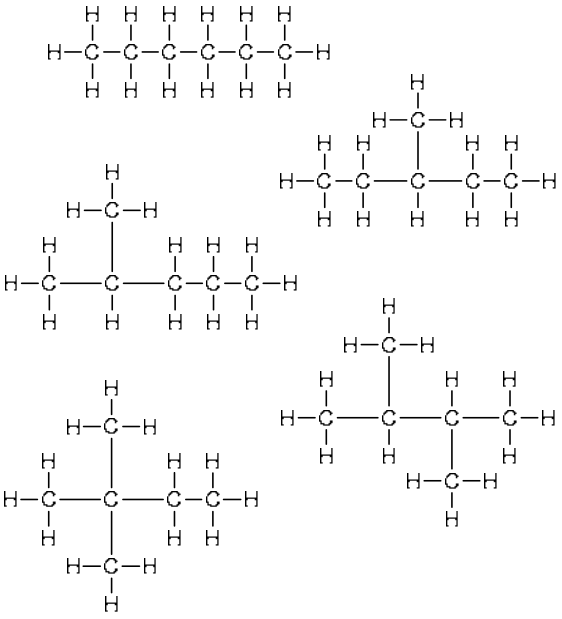
(c) The carbon atoms are tetravalent, while the chlorine atoms and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the two carbon atoms) should be drawn in the center of the compound. The chlorine atoms and hydrogen atoms are then placed at the periphery, and there are two different ways to do this. One way is to connect all three chlorine atoms to the same carbon atom. Alternatively, we can connect two chlorine atoms to one carbon atom, and then connect the third chlorine atom to the other carbon atom, as shown here:



(d) The carbon atoms are tetravalent, and the hydrogen atoms are all monovalent. Any atoms with more than one bond (in this case, the six carbon atoms) should be drawn in the center of the compound, with the hydrogen atoms at the periphery. There are five different ways to connect six carbon atoms, which we will organize based on the length of the longest chain.



Finally, we complete all of the structures by drawing the bonds to hydrogen atoms (ensuring that each carbon atom has a total of four bonds). There are a total of five isomers:

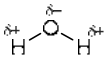


1.33.

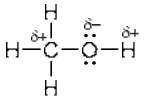
(a) According to Table 1.1, the difference in electronegativity between Br and H is 2.8 – 2.1 = 0.7, so an H–Br bond is expected to be polar covalent. Since bromine is more electronegative than hydrogen, the Br will be electron-rich (δ–), and the H will be electron-poor (δ+), as shown below:

(b) According to Table 1.1, the difference in electronegativity between Cl and H is 3.0 – 2.1 = 0.9, so an H–Cl bond is expected to be polar covalent. Since chlorine is more electronegative than hydrogen, the Cl will be electron-rich (δ–), and the H will be electron-poor (δ+), as shown below:

(c) According to Table 1.1, the difference in electronegativity between O and H is 3.5 – 2.1 = 1.4, so an O–H bond is expected to be polar covalent. Oxygen is more electronegative than hydrogen, so for each O–H bond, the O will be electron-rich (δ–) and the H will be electron-poor (δ+), as shown below:



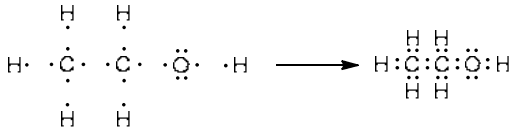
(d) According to Table 1.1, oxygen (3.5) is more electronegative than carbon (2.5) or hydrogen (2.1), and a C–O or H–O bond is polar covalent. For each C–O or H–O bond, the O will be electron-rich (δ–), and the C or H will be electron-poor (δ+), as shown below:



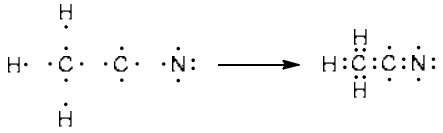
1.34.
(a) The difference in electronegativity between Na (0.9) and Br (2.8) is greater than the difference in electronegativity between H (2.1) and Br (2.8). Therefore, NaBr is expected to have more ionic character than HBr.

(b) The difference in electronegativity between F (4.0) and Cl (3.0) is greater than the difference in electronegativity between Br (2.8) and Cl (3.0). Therefore, FCl is expected to have more ionic character than BrCl.

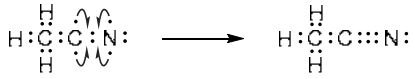
1.35.
(a) Each carbon atom has four valence electrons, the oxygen atom has six valence electrons, and each hydrogen atom has one valence electron. In this case, the information provided in the problem statement (CH_3CH_2OH) indicates how the atoms are connected to each other:



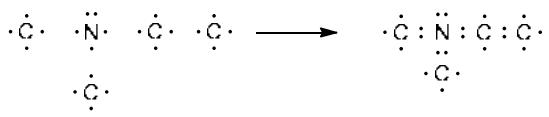
(b) Each carbon atom has four valence electrons, the nitrogen atom has five valence electrons, and each hydrogen atom has one valence electron. In this case, the information provided in the problem statement (CH_3CN) indicates how the atoms are connected to each other:



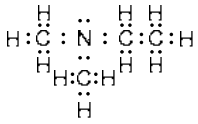
The unpaired electrons are then paired up to give a triple bond. In this way, each of the atoms achieves an octet.



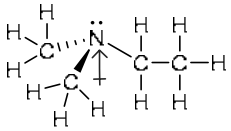
1.36. Each of the carbon atoms has four valence electrons; the nitrogen atom has five valence electrons, and each of the hydrogen atoms has one valence electron. We begin by connecting the atoms that have more than one bond (in this case, the four carbon atoms and the nitrogen atom). The problem statement indicates how we should connect them:



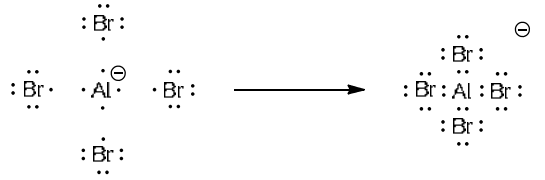
Then, we connect all of the hydrogen atoms (ensuring that each carbon atom has four bonds), as shown.



The nitrogen atom has three σ bonds and one lone pair, so the steric number is 4, which means that the arrangement of electron pairs is expected to be tetrahedral. One corner of the tetrahedron is occupied by a lone pair, so the geometry of the nitrogen atom (the arrangement of atoms around that nitrogen atom) is trigonal pyramidal. As such, the individual dipole moments associated with the C–N bonds do not fully cancel each other, and there is also a dipole moment associated with the lone pair (pointing up). As a result, there is a net molecular dipole moment, as shown:

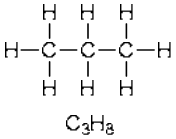


1.37. Bromine is in group 7A of the periodic table, so each bromine atom has seven valence electrons. Aluminum is in group 3A of the periodic table, so aluminum is supposed to have three valence electrons, but the structure bears a negative charge, which means that there is one extra electron. That is, the aluminum atom has four valence electrons, rather than three, which is why it has a formal negative charge. This gives the following Lewis structure:



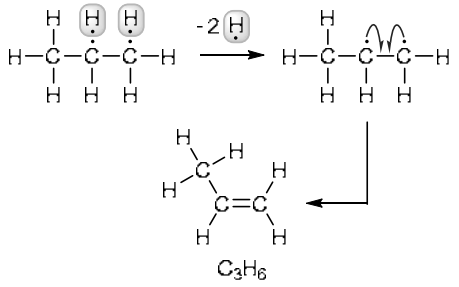
The aluminum atom has four bonds and no lone pairs, so the steric number is 4, which means that this aluminum atom will have tetrahedral geometry.

1.38. The molecular formula of cyclopropane is C_3H_6, so we are looking for a different compound that has the same molecular formula, C_3H_6. That is, we need to find another way to connect the carbon atoms, other than in a ring (there is only one way to connect three carbon atoms in a ring, so we must be looking for something other than a ring). If we connect the three carbon atoms in a linear fashion and then complete the drawing by placing hydrogen atoms at the periphery, we notice that the molecular formula (C_3H_8) is not correct:



C_3H_8

We are looking for a structure with the molecular formula C_3H_6. If we remove two hydrogen atoms from our drawing, we are left with two unpaired electrons, indicating that we should consider drawing a double bond:

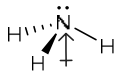


The structure of this compound (called propylene) is different from the structure of cyclopropane, but both compounds share the same molecular formula, so they are constitutional isomers.

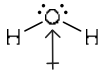
1.39.

(a) C–H bonds are considered to be nonpolar, although they do have a very small dipole moment, because there is a small difference in electronegativity between carbon (2.5) and hydrogen (2.1). With no polar bonds present, the molecule does not have a molecular dipole moment.

(b) The nitrogen atom has trigonal pyramidal geometry. As such, the dipole moments associated with the polar N–H bonds do not fully cancel each other, and there is also a dipole moment associated with the lone pair (pointing up). As a result, there is a net molecular dipole moment, as shown:



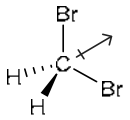
(c) The oxygen atom has two σ bonds and two lone pairs (steric number = 4), and VSEPR predicts bent geometry. As such, the dipole moments associated with the polar O–H bonds do not cancel each other, and the dipole moments associated with the lone pairs also do not fully cancel each other. As a result, there is a net molecular dipole moment, as shown:



(d) The central carbon atom of carbon dioxide (CO_2) has two σ bonds and no lone pairs, so it is *sp* hybridized and is expected to have linear geometry. Each C=O bond has a strong dipole moment, but in this case, they are pointing in opposite directions. As such, they fully cancel each other, giving no net molecular dipole moment.

(e) Carbon tetrachloride (CCl_4) has four C–Cl bonds, each of which exhibits a dipole moment. However, the central carbon atom has four σ bonds so it is expected to have tetrahedral geometry. As such, the four dipole moments completely cancel each other out, and there is no net molecular dipole moment.

(f) This compound has two C–Br bonds, each of which exhibits a dipole moment. To determine if these dipole moments cancel each other, we must identify the molecular geometry. The central carbon atom has four σ bonds so it is expected to have tetrahedral geometry. As such, the polar C–Br bonds do not completely cancel each other out. There is a net molecular dipole moment, as shown:



1.40.

(a) As indicated in Figure 1.10, a neutral oxygen atom has two *1s* electrons, two *2s* electrons, and four *2p* electrons.
(b) As indicated in Figure 1.10, a neutral fluorine atom has two *1s* electrons, two *2s* electrons, and five *2p* electrons.
(c) As indicated in Figure 1.10, a neutral carbon atom has two *1s* electrons, two *2s* electrons, and two *2p* electrons.
(d) As seen in SkillBuilder 1.6, the electron configuration of a neutral nitrogen atom is $1s^2 2s^2 2p^3$
(e) This is the electron configuration of a neutral chlorine atom.

1.41.

(a) The difference in electronegativity between sodium (0.9) and bromine (2.8) is 2.8 – 0.9 = 1.9. Since this difference is greater than 1.7, the bond is classified as ionic.

(b) The difference in electronegativity between sodium (0.9) and oxygen (3.5) is 3.5 – 0.9 = 2.6. Since this difference is greater than 1.7, the Na–O bond is classified as ionic. In contrast, the O–H bond is polar covalent, because the difference in electronegativity between oxygen (3.5) and hydrogen (2.1) is less than 1.7 but more than 0.5.

(c) Each C–H bond is considered to be covalent, because the difference in electronegativity between carbon (2.5) and hydrogen (2.1) is less than 0.5.
The C–O bond is polar covalent, because the difference in electronegativity between oxygen (3.5) and carbon (2.5) is less than 1.7 but more than 0.5.
The Na–O bond is classified as ionic, because the difference in electronegativity between oxygen (3.5) and sodium (0.9) is greater than 1.7.

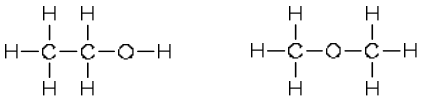
(d) Each C–H bond is considered to be covalent, because the difference in electronegativity between carbon (2.5) and hydrogen (2.1) is less than 0.5.
The C–O bond is polar covalent, because the difference in electronegativity between oxygen (3.5) and carbon (2.5) is less than 1.7 but more than 0.5.

The O–H bond is polar covalent, because the difference in electronegativity between oxygen (3.5) and hydrogen (2.1) is less than 1.7 but more than 0.5.

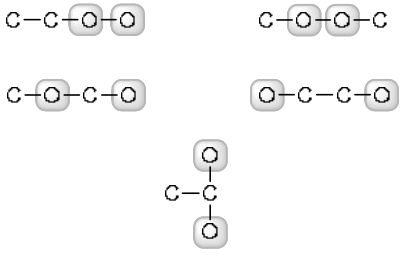
(e) Each C–H bond is considered to be covalent, because the difference in electronegativity between carbon (2.5) and hydrogen (2.1) is less than 0.5.
The C=O bond is polar covalent, because the difference in electronegativity between oxygen (3.5) and carbon (2.5) is less than 1.7 but more than 0.5.

1.42.
(a) Begin by determining the valency of each atom in the compound. The carbon atoms are tetravalent, the oxygen atom is divalent, and the hydrogen atoms are all monovalent. Any atoms with more than one bond (in this case, the two carbon atoms and the oxygen atom) should be drawn in the center of the compound, with the hydrogen atoms at the periphery. There are two different ways to connect two carbon atoms and an oxygen atom, shown here:

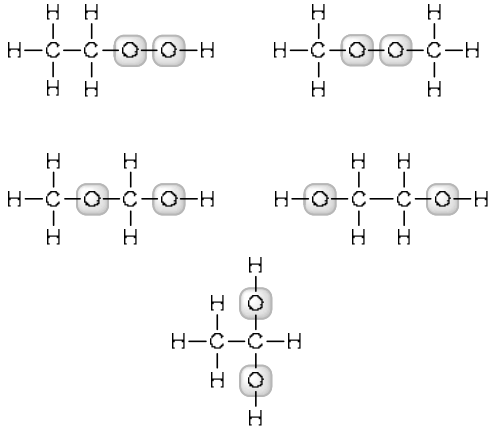
We then complete both structures by drawing the remaining bonds to hydrogen atoms (ensuring that each carbon atom has four bonds, and each oxygen atom has two bonds):



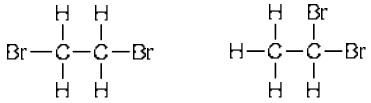
(b) Begin by determining the valency of each atom in the compound. The carbon atoms are tetravalent, the oxygen atoms are divalent, and the hydrogen atoms are all monovalent. Any atoms with more than one bond (in this case, the two carbon atoms and the two oxygen atoms) should be drawn in the center of the compound, with the hydrogen atoms at the periphery. There are several different ways to connect two carbon atoms and two oxygen atoms (highlighted, for clarity of comparison), shown here:



We then complete all of these structures by drawing the remaining bonds to hydrogen atoms:



(c) The carbon atoms are tetravalent, while the bromine atoms and hydrogen atoms are all monovalent. The atoms with more than one bond (in this case, the two carbon atoms) should be drawn in the center of the compound. The bromine atoms and hydrogen atoms are then placed at the periphery, and there are two different ways to do this. The two bromine atoms can either be connected to the same carbon atom or to different carbon atoms, as shown.



1.43.
(a) Oxygen is more electronegative than carbon, and the withdrawal of electron density toward oxygen can be indicated with the following arrow:

(b) Carbon is more electronegative than magnesium, and the withdrawal of electron density toward carbon can be indicated with the following arrow:

(c) Nitrogen is more electronegative than carbon, and the withdrawal of electron density toward nitrogen can be indicated with the following arrow:

(d) Carbon is more electronegative than lithium, and the withdrawal of electron density toward carbon can be indicated with the following arrow:

(e) Chlorine is more electronegative than carbon, and the withdrawal of electron density toward chlorine can be indicated with the following arrow:

(f) Carbon is more electronegative than silicon, and the withdrawal of electron density toward carbon can be indicated with the following arrow:

(g) Oxygen is more electronegative than hydrogen, and the withdrawal of electron density toward oxygen can be indicated with the following arrow:

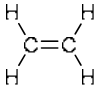
(h) Nitrogen is more electronegative than hydrogen, and the withdrawal of electron density toward nitrogen can be indicated with the following arrow:

1.44.

(a) The oxygen atom has two σ bonds and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. The C-O-H bond angle is expected to be approximately 105º, and all other bonds angles are expected to be approximately 109.5º (because each carbon atom has four σ bonds and tetrahedral geometry).

(b) The central carbon atom has three σ bonds and no lone pairs (steric number = 3), and VSEPR theory predicts trigonal planar geometry. As such, all bond angles are approximately 120º.

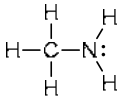
(c) Each of the carbon atoms has three σ bonds and no lone pairs (steric number = 3), and VSEPR theory predicts trigonal planar geometry. As such, all bond angles are approximately 120º.



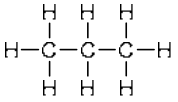
(d) Each of the carbon atoms has two σ bonds and no lone pairs (steric number = 2), and VSEPR theory predicts linear geometry. As such, all bond angles are 180º.

(e) The oxygen atom has two σ bonds and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. Therefore, the C-O-C bond angle is expected to be around 105º. The remaining bond angles are all expected to be approximately 109.5º (because each carbon atom has four σ bonds and tetrahedral geometry).

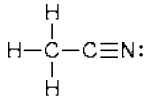
(f) The nitrogen atom has three σ bonds and one lone pair (steric number = 4), and VSEPR theory predicts trigonal pyramidal geometry, with bond angles of 107º. The carbon atom is also tetrahedral (because it has four σ bonds), although the bond angles around the carbon atom are expected to be approximately 109.5º.



(g) Each of the carbon atoms has four σ bonds (steric number = 4), so each of these carbon atoms has tetrahedral geometry. Therefore, all bond angles are expected to be approximately 109.5º.



(h) The structure of acetonitrile (CH_3CN) is shown below (see the solution to Problem **1.35b**).



One of the carbon atoms has four σ bonds (steric number = 4), and is expected to have tetrahedral geometry. The other carbon atom (connected to nitrogen) has two σ bonds and no lone pairs (steric number = 2), so we expect linear geometry.
As such, the C–C≡N bond angle is 180º, and all other bond angles are approximately 109.5º.

1.45.

(a) The nitrogen atom has three σ bonds and one lone pair (steric number = 4). It is sp^3 hybridized (electronically tetrahedral), with trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair).

(b) The boron atom has three σ bonds and no lone pairs (steric number = 3). It is sp^2 hybridized, with trigonal planar geometry.

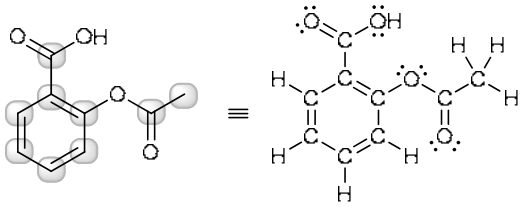
(c) This carbon atom has three σ bonds and no lone pairs (steric number = 3). It is sp^2 hybridized, with trigonal planar geometry.

(d) This carbon atom has three σ bonds and one lone pair (steric number = 4). It is sp^3 hybridized (electronically

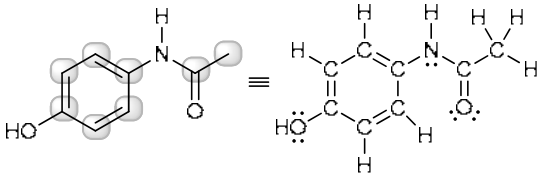
tetrahedral), with trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair).

1.46.

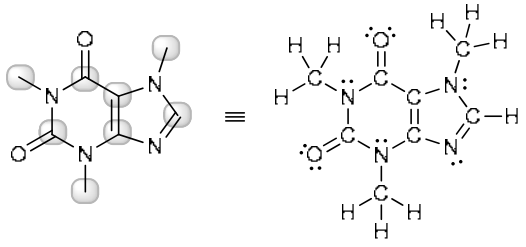
(a) Each corner and each endpoint represents a carbon atom (highlighted), so this compound has nine carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, as shown.



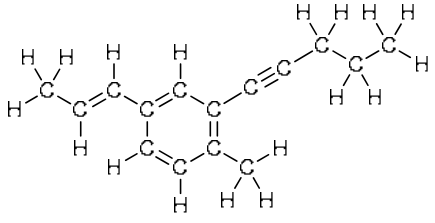
(b) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has eight carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, as shown.



(c) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has eight carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, as shown.

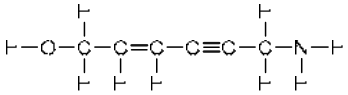


1.47. Each corner and each endpoint represents a carbon atom, so this compound has fifteen carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, giving a total of eighteen hydrogen atoms, as shown here:



$C_{15}H_{18}$

1.48. The double bond represents one σ bond and one π bond, while the triple bond represents one σ bond and two π bonds. All single bonds are σ bonds. Therefore, this compound has sixteen σ bonds and three π bonds.



1.49.

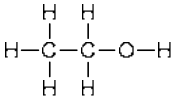
(a) The second compound is expected to have a higher boiling point, because it has an O–H bond, which will lead to hydrogen bonding interactions.

(b) The second compound is expected to have a higher boiling point, because it has more carbon atoms, and thus a higher molecular weight and more opportunity for London dispersion forces.

(c) The first compound has a C=O bond, which has a strong dipole moment, while the second compound is nonpolar. The first compound is therefore expected to exhibit strong dipole-dipole interactions and to have a higher boiling point than the second compound.

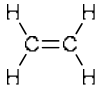
1.50.

(a) This compound possesses an O–H bond, so it is expected to exhibit hydrogen bonding interactions.



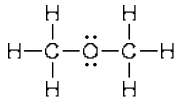
(b) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pairs can serve as hydrogen-bond acceptors, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect there to be any hydrogen bonding interactions.

(c) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound will not exhibit hydrogen bonding interactions.

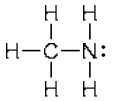


(d) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom . Therefore, this compound will not exhibit hydrogen bonding interactions.

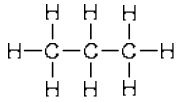
(e) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pairs can serve as hydrogen-bond acceptors, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect there to be any hydrogen bonding interactions.



(f) This compound possesses two N–H bonds, so it is expected to exhibit hydrogen bonding interactions.



(g) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound will not exhibit hydrogen bonding interactions.



(h) This compound possesses N–H bonds, so it is expected to exhibit hydrogen bonding interactions.

1.51.

(a) Boron is in group 3A of the periodic table, and therefore has three valence electrons. It can use each of its valence electrons to form a bond, so we expect the molecular formula to be BH_3 (x=3).

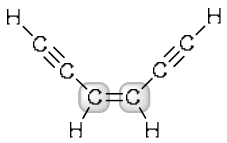
(b) Carbon is in group 4A of the periodic table, and therefore has four valence electrons. It can use each of its valence electrons to form a bond, so we expect the molecular formula to be CH_4 (x=4).

(c) Nitrogen is in group 5A of the periodic table, and therefore has five valence electrons. But it cannot form five bonds, because it only has four orbitals with which to form bonds. One of those orbitals must be occupied by a lone pair (two electrons), and each of the remaining three electrons is available to form a bond. Nitrogen is therefore trivalent, and we expect the molecular formula to be NH_3 (x=3).

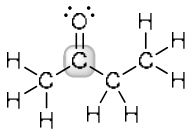
(d) Carbon is in group 4A of the periodic table, and therefore has four valence electrons. It can use each of its valence electrons to form a bond, and indeed, we expect the carbon atom to have four bonds. Two of the bonds are with hydrogen atoms, so the other two bonds must be with chlorine atoms. The molecular formula is CH_2Cl_2 (x=2).

1.52.

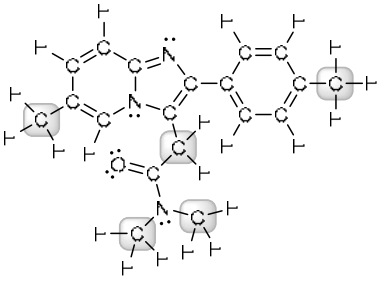
(a) Each of the highlighted carbon atoms has three σ bonds and no lone pairs (steric number = 3). Each of these carbon atoms is sp^2 hybridized, with trigonal planar geometry. Each of the other four carbon atoms has two σ bonds and no lone pairs (steric number = 2). Those four carbon atoms are all *sp* hybridized, with linear geometry.



(b) The highlighted carbon atom has three σ bonds and no lone pairs (steric number = 3). This carbon atom is sp^2 hybridized, with trigonal planar geometry. Each of the other three carbon atoms has four σ bonds (steric number = 4). Those three carbon atoms are all sp^3 hybridized, with tetrahedral geometry.



1.53. Each of the highlighted carbon atoms has four σ bonds (steric number = 4), and is sp^3 hybridized, with tetrahedral geometry. Each of the other fourteen carbon atoms in this structure has three σ bonds and no lone pairs (steric number = 3). Each of these fourteen carbon atoms is sp^2 hybridized, with trigonal planar geometry.



1.54.

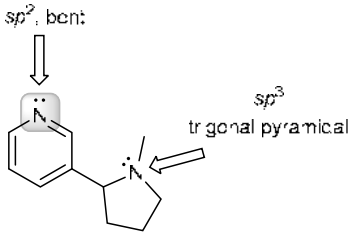
(a) Oxygen is the most electronegative atom in this compound. See Table 1.1 for electronegativity values.

(b) Fluorine is the most electronegative atom. See Table 1.1 for electronegativity values.

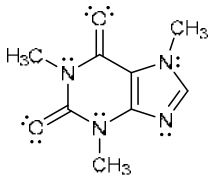
(c) Carbon is the most electronegative atom in this compound. See Table 1.1 for electronegativity values.

1.55. The highlighted nitrogen atom has two σ bonds and one lone pair (steric number = 3). This nitrogen atom is sp^2 hybridized. It is electronically trigonal planar, but one of the sp^2 hybridized orbitals is occupied by a lone pair, so the geometry (arrangement of atoms) is bent. The other nitrogen atom (not highlighted) has three σ bonds and a lone pair (steric number = 4). That nitrogen atom is sp^3

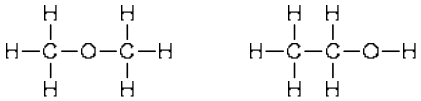
hybridized and electronically tetrahedral. One corner of the tetrahedron is occupied by a lone pair, so the geometry (arrangement of atoms) is trigonal pyramidal.



1.56. Each of the nitrogen atoms in caffeine achieves an octet with three bonds and one lone pair, while each oxygen atom in this structure achieves an octet with two bonds and two lone pairs, as shown:



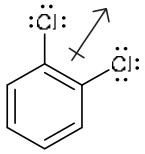
1.57. As seen in Section 1.2, the following two compounds have the molecular formula C_2H_6O.



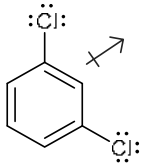
The second compound will have a higher boiling point because it possesses an OH group which can exhibit hydrogen bonding interactions.

1.58.

(a) Each C–Cl bond has a dipole moment, and the two dipole moments do not fully cancel each other because the polar bonds are not pointing in opposite directions. As such, there will be a net molecular dipole moment, as shown here:

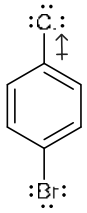


(b) Each C–Cl bond has a dipole moment, and the two dipole moments do not fully cancel each other because the polar bonds are not pointing in opposite directions. As such, there will be a net molecular dipole moment, as shown here:



(c) Each C–Cl bond has a dipole moment, and in this case, the two dipole moments are pointing in opposite directions. As such, they fully cancel each other, giving no net molecular dipole moment.

(d) The C–Cl bond has a dipole moment, and the C–Br bond also has a dipole moment. These two dipole moments are in opposite directions, but they do not have the same magnitude. The C–Cl bond has a larger dipole moment than the C–Br bond, because chlorine is more electronegative than bromine. Therefore, there will be a net molecular dipole moment, as shown here:

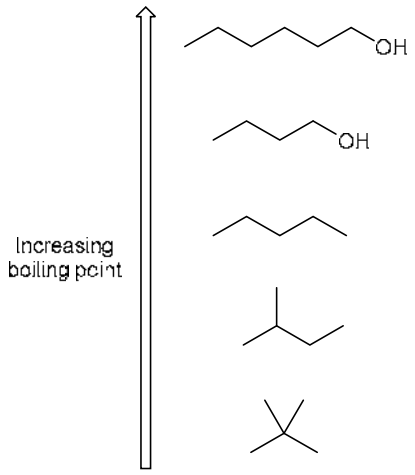


1.59. The C–Cl bond in chloroform partially cancels the dipole moments of the other two C–Cl bonds, thereby reducing the molecular dipole moment relative to methylene chloride.

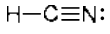
1.60. $CHCl_3$ is expected to have a larger molecular dipole moment than $CBrCl_3$, because the bromine atom in $CBrCl_3$ serves to partially cancel out the dipole moments of the three C–Cl bonds (as is the case for CCl_4).

1.61. The carbon atom of O=C=O has two σ bonds and no lone pairs (steric number = 2) and VSEPR theory predicts linear geometry. As a result, the individual dipole moments of each C=O bond cancel each other completely to give no overall molecular dipole moment. In contrast, the sulfur atom in SO_2 has a steric number of three (because it also has a lone pair, in addition to the two S=O bonds), which means that it has bent geometry. As a result, the individual dipole moments of each S=O bond do NOT cancel each other completely, and the molecule does have a molecular dipole moment.

1.62. Two compounds possess OH groups. These two compounds will have the highest boiling points, because they can form hydrogen bonds. Among these two compounds, the one with more carbon atoms (six) will be higher boiling than the one with fewer carbon atoms (four), because the higher molecular weight results in greater London dispersion forces. The remaining three compounds all have equal molecular weights (five carbon atoms) and lack an OH group. The difference between these three compounds is the extent of branching. Among these three compounds, the compound with the greatest extent of branching has smallest surface area and therefore the lowest boiling point, and the one with the least branching has the highest boiling point.



1.63. The correct answer is (a). We must first draw the structure of HCN. To draw a Lewis structure, we begin by counting the valence electrons (hydrogen has one valence electron, carbon has four valence electrons, and nitrogen has five valence electrons, for a total of ten valence electrons). The structure must have ten valence electrons (no more and no less). Carbon should have four bonds, and it can only form a single bond with the hydrogen atom, so there must be a triple bond between carbon and nitrogen:

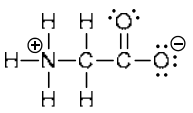


The single bond accounts for two electrons, and the triple bonds accounts for another six electrons. The remaining two electrons must be a lone pair on nitrogen. This accounts for all ten valence electrons, and it gives all atoms an octet.
Since the carbon atom has a triple bond, it must be *sp* hybridized, with linear geometry.

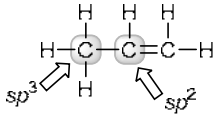
1.64. The molecular formula of cyclobutane is C_4H_8. Of the four structures shown, only structure (c) has the same molecular formula (C_4H_8).

1.65. The correct answer is (b). Each of the structures has two carbon atoms and one oxygen atom. However, only the second structure has an OH group. This compound will have an elevated boiling point, relative to the other three structures, because of hydrogen bonding.

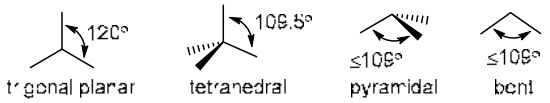
1.66. The first statement (a) is the correct answer, because an oxygen atom has a negative charge, and the nitrogen atom has a positive charge, as shown here:



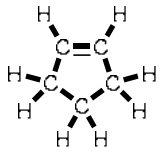
1.67. The correct answer is (d). Carbon-carbon σ bonds are formed by overlapping hybridized orbitals, so we must determine the hybridization of both carbon atoms (highlighted below). The carbon on the left has four σ bonds, so it is sp^3-hybridized, and the carbon on the right has three σ bonds, so it is sp^2-hybridized. The indicated σ bond results from the overlap of one sp^3 hybridized orbital (from the carbon atom on the left), and one sp^2 hybridized orbital (from the carbon atom on the right).



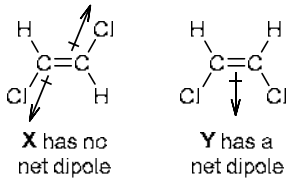
1.68. The correct answer is (c). An atom with a trigonal planar arrangement of bonds has a bond angle of 120°, the largest bond angle among the choices listed. A tetrahedral arrangement of bonds corresponds to an angle of 109.5°, while trigonal pyramidal and bent geometries typically have bond angles that are slightly smaller than 109.5°.



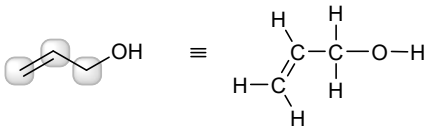
1.69. The correct answer is (c). There are 13 σ bonds in the given compound (highlighted in bold below). All of the single bonds are σ bonds, and the double bond is comprised of one σ bond and one π bond.



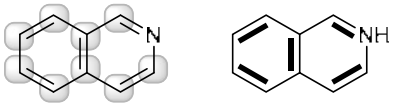
1.70. The correct answer is (b). Each C–Cl bond in **Y** has a dipole moment, and compound **Y** has a molecular dipole moment. Each C–Cl bond in **X** also has a dipole moment, but in this case, the two polar bonds are pointing in opposite directions. As such, the dipole moments fully cancel each other, giving no net molecular dipole moment for compound **X**. The polar isomer (**Y**) has more significant dipole-dipole attractions between molecules and, therefore, has a higher boiling point than the nonpolar isomer (**X**).



1.71. The correct answer is (b). In the bond-line drawing, each corner and each endpoint represents a carbon atom, so this compound has three carbon atoms (highlighted below). Each carbon atom will have enough hydrogen atoms to have exactly four bonds. Together with the hydrogen atom that is connected to the oxygen atom, there are a total of six hydrogen atoms, as shown below:



1.72. The correct answer is (d). In the bond-line drawing, each corner represents a carbon atom (highlighted below), so this compound has nine carbon atoms. Each double bond is comprised of one σ bond and one π bond, so there are five π bonds (in bold below).



1.73.

(a) Compounds A and B share the same molecular formula (C_4H_9N) but differ in their constitution (connectivity of atoms), and they are therefore constitutional isomers.

(b) The nitrogen atom in compound B has three σ bonds and one lone pair (steric number = 4). It is sp^3 hybridized (electronically tetrahedral), with trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair).

(c) A double bond represents one σ bond and one π bond, while a triple bond represents one σ bond and two π bonds. A single bond represents a σ bond. With this in mind, compound B has 14 σ bonds, as compared with compounds A and C, which have 13 and 11 σ bonds, respectively.

(d) As explained in the solution to part (c), compound C has the fewest σ bonds.

(e) A double bond represents one σ bond and one π bond, while a triple bond represents one σ bond and two π bonds. As such, compound C exhibits two π bonds.

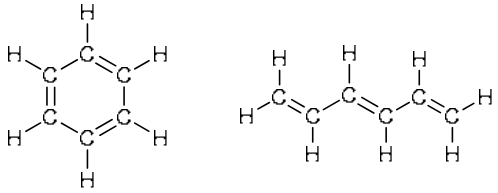
(f) Compound A has a C=N bond, in which the carbon atom has three σ bonds and no lone pairs (steric number = 3). It is sp^2 hybridized.

(g) Each of the carbon atoms in compound B is sp^3 hybridized with four σ bonds (steric number = 4). Similarly, the nitrogen atom in compound B has three σ bonds and one lone pair (steric number = 4). This nitrogen atom is also sp^3 hybridized.

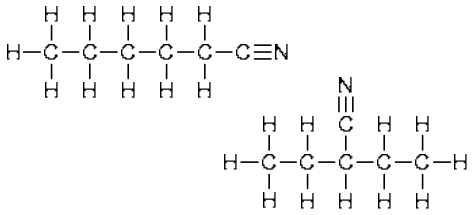
(h) Compound A has an N–H bond, and is therefore expected to form hydrogen bonding interactions. Compounds B and C do not contain an N–H bond, so compound A is expected to have the highest boiling point.

1.74.

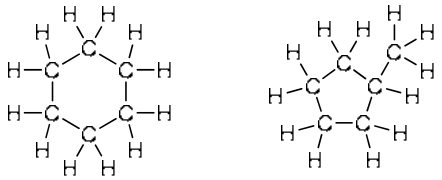
(a) In each of the following two compounds, all of the carbon atoms are sp^2 hybridized (each carbon atom has three σ bonds and one π bond). There are certainly many other possible compounds for which all of the carbon atoms are sp^2 hybridized.



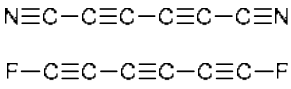
(b) In each of the following two compounds, all of the carbon atoms are sp^3 hybridized (because each carbon atom has four σ bonds) with the exception of the carbon atom connected to the nitrogen atom. That carbon atom has two σ bonds and is therefore *sp* hybridized. There are certainly many other acceptable answers.



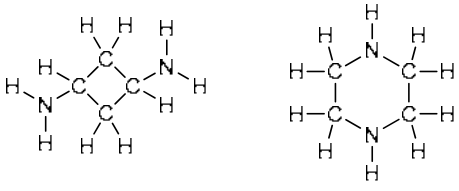
(c) In each of the following two compounds, there is a ring, and all of the carbon atoms are sp^3 hybridized (because each carbon atom has four σ bonds). There are certainly many other acceptable answers.



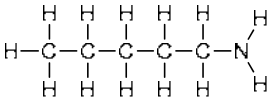
(d) In each of the following two compounds, all of the carbon atoms are *sp* hybridized (because each carbon atom has two σ bonds). There are certainly many other acceptable answers.



1.75. In each of the following two compounds, the molecular formula is $C_4H_{10}N_2$, there is a ring (as suggested in the hint given in the problem statement), there are no π bonds, there is no net dipole moment, and there is an N-H bond, which enables hydrogen bonding interactions.

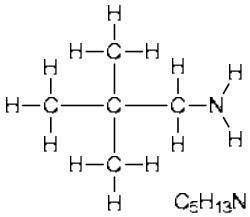


1.76. If we try to draw a linear skeleton with five carbon atoms and one nitrogen atom, we find that the number of hydrogen atoms is not correct (there are thirteen, rather than eleven):



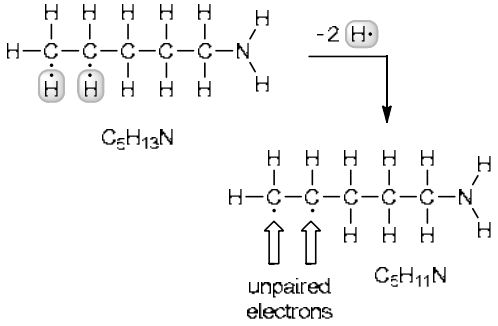
$C_5H_{13}N$

This will be the case even if try to draw a branched skeleton:



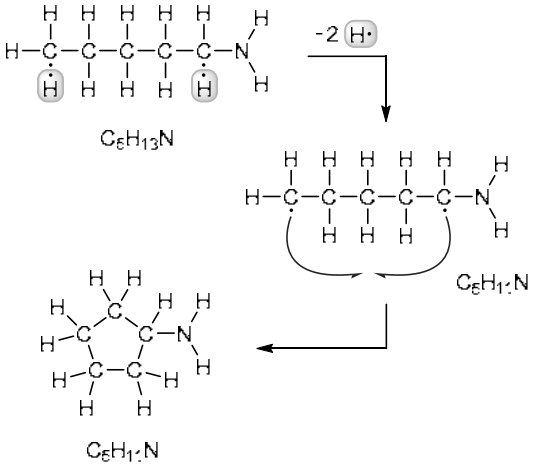
$C_5H_{13}N$

In fact, regardless of how the skeleton is branched, it will still have thirteen hydrogen atoms. But we need to draw a structure with only eleven hydrogen atoms ($C_5H_{11}N$). So we must remove two hydrogen atoms, which gives two unpaired electrons:



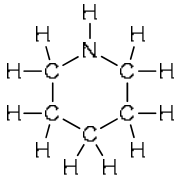
$C_5H_{13}N$ $C_5H_{11}N$

This indicates that we should consider pairing these electrons as a double bond. However, the problem statement specifically indicates that the structure cannot contain a double bond. So, we must find another way to pair the unpaired electrons. Let's consider forming a ring instead of a double bond:



$C_5H_{13}N$ $C_5H_{11}N$ $C_5H_{11}N$

Now we have the correct number of hydrogen atoms (eleven), which means that our structure must indeed contain a ring. But this particular cyclic structure (cyclic = containing a ring) does not meet all of the criteria described in the problem statement. Specifically, each carbon atom must be connected to exactly two hydrogen atoms. This is not the case in the structure above. This issue can be remedied in the following structure, which has a ring, and each of the carbon atoms is connected to exactly two hydrogen atoms, as required by the problem statement.



1.77.

(a) In compound **A**, the nitrogen atom has two σ bonds and no lone pairs (steric number = 2). It is *sp* hybridized. The highlighted carbon atom has one σ bond and one lone pair (steric number = 2), so that carbon atom is also *sp* hybridized.

(b) The highlighted carbon atom is *sp* hybridized, so the lone pair occupies an *sp*-hybridized orbital.

(c) The nitrogen atom is *sp* hybridized and therefore has linear geometry. As such, the C-N-C bond angle in **A** is expected to be 180°.

(d) The nitrogen atom in **B** has two σ bonds and one lone pair (steric number = 3). It is sp^2 hybridized. The highlighted carbon atom has three σ bonds and no lone pairs (steric number = 3), and that carbon atom is sp^2 hybridized. Each of the chlorine atoms has three lone pairs and one bond (steric number = 4), and the chlorine atoms are sp^3 hybridized.

(e) The nitrogen atom is sp^2 hybridized, so the lone pair occupies an sp^2-hybridized orbital.

(f) The nitrogen atom is sp^2 hybridized so the C-N-C bond angle in **B** is expected to be approximately 120°.

1.78. By analyzing the data, we can see that C(sp^2)–Cl must be shorter than 1.79Å [compare with C(sp^3)–Cl], while C(*sp*)–I must be longer than 1.79Å [compare with C(*sp*)–Br]. Therefore, C(*sp*)–I must be longer than C(sp^2)–Cl.

1.79.

(a) In the first compound, the fluorine isotope (^{18}F) has no formal charge. Therefore, it must have three lone pairs (see Section 1.4 for a review of how formal charges are calculated). Since it has one σ bond and three lone pairs, it must have a steric number of 4, and is sp^3 hybridized. The bromine atom also has no formal charge. So, it too, like the fluorine isotope, must have three lone pairs. Once again, one σ bond and three lone pairs give a steric number of 4, so the bromine atom is sp^3 hybridized.

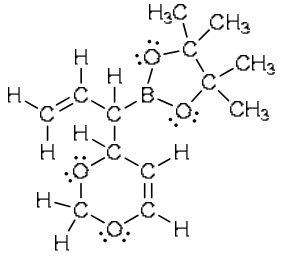
In the second compound, the nitrogen atom has no formal charge. Therefore, it must have one lone pair. Since the nitrogen atom has three σ bonds and one lone pair, it must have a steric number of 4, and is sp^3 hybridized.
In the product, the fluorine isotope (^{18}F) has no formal charge. Therefore, it must have three lone pairs. Since it has one σ bond and three lone pairs, it must have a steric number of 4, and is sp^3 hybridized. The nitrogen atom does have a positive formal charge. Therefore, it must have no lone pairs. Since it has four σ bonds and no lone pairs, it must have a steric number of 4, and is sp^3 hybridized. Finally, the bromine atom has a negative charge and no bonds. So it must have four lone pairs. With four lone pairs and no bonds, it will have a steric number of 4, and is expected to be sp^3 hybridized.
In summary, all of the atoms that we analyzed are sp^3 hybridized.

(b) The nitrogen atom is sp^3 hybridized. With four bonds, we expect the geometry around the nitrogen atom to be tetrahedral. So, the bond angle for each C-N-C bond is expected to be approximately 109.5°.

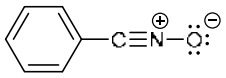
1.80.
(a) Boron is in group 3A of the periodic table and is therefore expected to be trivalent. That is, it has three valence electrons, and it uses each one of those valence electrons to form a bond, giving rise to three bonds. It does not have any electrons left over for a lone pair (as in the case of nitrogen). With three σ bonds and no lone pairs, the boron atom has a steric number of three, and is sp^2 hybridized.

(b) Since the boron atom is sp^2 hybridized with three σ bonds, we expect the geometry to be trigonal planar and the bond angles to be approximately 120°. However, in this case, the O-B-O system is part of a five-membered ring. That is, there are five different bond angles (of which the O-B-O angle is one of them) that together must form a closed loop. That requirement could conceivably force some of the bond angles (including the O-B-O bond angle) to deviate from the predicted value. In fact, we will explore this very phenomenon, called ring strain, in Chapter 4, and we will see that five-membered rings actually possess very little ring strain compared with smaller rings.

(c) Each of the oxygen atoms has no formal charge, and must therefore have two bonds and two lone pairs. The boron atom has no lone pairs, as explained in the solution to part (a) of this problem.



1.81.
(a) If we analyze each atom (in both **1** and **2**) using the procedure outlined in Section 1.4, we find that none of the atoms in compound **1** have a formal charge, while compound **2** possesses two formal charges:



The nitrogen atom has a positive charge (it should have five valence electrons, but it is exhibiting only four), and the oxygen atom has a negative charge (it should have six valence electrons, but it is exhibiting seven).

(b) Compound **1** possesses polar bonds, as a result of the presence of partial charges (δ+ and δ-). The associated dipole moments can form favorable interactions with the dipole moments present in the polar solvent molecules (dipole-dipole interactions). However, compound **2** has formal charges (negative on O and positive on N), so the dipole moment of the N-O bond is expected to be much more significant than the dipole moments in compound **1**. The dipole moment of the N-O bond in compound **2** is the result of full charges, rather than partial charges. As such, compound **2** is expected to experience much stronger interactions with the solvent molecules, and therefore, **2** should be more soluble than **1** in a polar solvent.

(c) In compound **1**, the carbon atom (attached to nitrogen) has three σ bonds and no lone pairs (steric number = 3). That carbon atom is sp^2 hybridized, with trigonal planar geometry. As such, the C-C-N bond angle in compound **1** is expected to be approximately 120°. However, in compound **2**, the same carbon atom has two σ bonds and no lone pairs (steric number = 2). This carbon atom is *sp* hybridized, with linear geometry. As such, the C-C-N bond angle in **2** is expected to be 180°. The conversion of **1** to **2** therefore involves an increase in the C-C-N bond angle of approximately 60°.

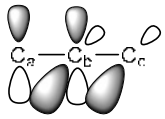
1.82.
(a) C_a has three σ bonds and no lone pairs, so it has a steric number of 3, and is sp^2 hybridized. The same is true for C_c. In contrast, C_b has two σ bonds and no lone pairs, so it has a steric number of 2, and is therefore *sp* hybridized.

(b) Since C_a is sp^2 hybridized, we expect its geometry to be trigonal planar, so the bond angle should be approximately 120°.

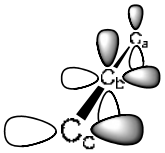
(c) Since C_b is *sp* hybridized, we expect its geometry to be linear, so the bond angle should be approximately 180°.

(d) The central carbon atom (C_b) is *sp* hybridized, so it is using two *sp* hybridized orbitals to form its two σ bonds, which will be arranged in a linear fashion. The remaining two *p* orbitals of C_b used for π bonding will be 90° apart

from one another (just as we saw for the carbon atoms of a triple bond; see Figure 1.33).

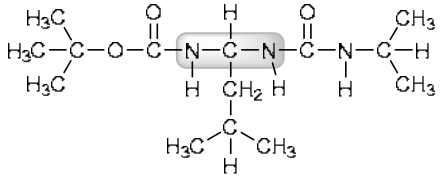


As a result, the two π systems are orthogonal (or 90°) to each other. Therefore, the *p* orbitals on C_a and C_c are orthogonal. The following is another drawing from a different perspective (looking down the axis of the linear C_a-C_b-C_c system.

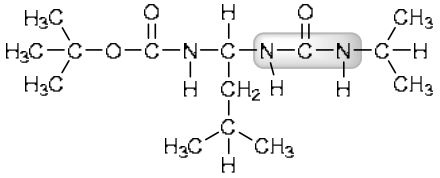


1.83.

(a) The N-C-N unit highlighted below exhibits a central carbon atom that is sp^3 hybridized and is therefore expected to have tetrahedral geometry. Accordingly, the bond angles about that carbon atom are expected to be approximately 109.5°.



The other N-C-N unit (highlighted below) exhibits a central carbon atom that is sp^2 hybridized and is therefore expected to have trigonal planar geometry. Accordingly, the bond angles about that carbon atom are expected to be approximately 120°.



(b) The interaction is an intramolecular hydrogen bond that forms between the δ+ H (connected to the highlighted nitrogen atom) and the lone pair of the δ– oxygen atom: