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

Review of Concepts

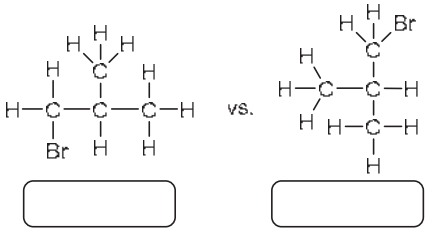
Fill in the blanks below. To verify that your answers are correct, look in your textbook at the end of Chapter 1. Each of the sentences below appears verbatim in the section entitled *Review of Concepts and Vocabulary*.

- _____ **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 _____ **hybrid orbitals** to achieve its four single bonds.
- Ethylene's planar geometry can be explained using three _____ **hybrid 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 sigma bonds and the number of _____ exhibited by each atom.
- The physical properties of compounds are determined by _____ forces, the attractive forces between molecules.
- _____ bonding is a type of attractive interaction which occurs when a lone pair of an electronegative atom interacts with an electron poor hydrogen atom.
- **London dispersion forces** result from the interaction between transient _____ and are stronger for larger molecules due to their larger surface area and ability to accommodate more interactions.
- Solubility is based on the principle of "_____ dissolves _____." Polar compounds are soluble in _____ solvents; nonpolar compounds are soluble in _____ solvents.

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 *SkillBuilder Review* section.

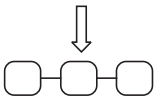
SkillBuilder 1.1 Identifying Constitutional Isomers

EXAMPLE	Step 1 Determine the molecular formula for each structure.	Step 2 Using the boxes below, draw the skeletal structure for each compound, and assign numbers to the parent chain of each, to determine if the compounds have the same connectivity of atoms
Evaluate the two drawings and determine whether they represent the same compound or constitutional isomers		

SkillBuilder 1.2 Drawing the Lewis Structure of a Small Molecule

Step 1 Connect the atoms that form more than one bond.

CH_3CH_2O



Step 2 Connect the hydrogen atoms as indicated in the condensed formula

Step 3 Determine the total number of valence electrons.

C x2 = ____ electrons

O x1 = ____ electrons

H x4 = ____ electrons

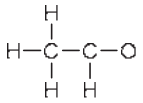
____ = total # of valence electrons

Step 4 Determine how many electrons have already been drawn

# of bonds: ____

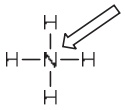
# of electrons: ____

Step 5 On the structure below, draw the remaining electrons so that each atom achieves an octet



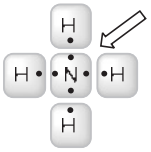
SkillBuilder 1.3 Calculating Formal Charge

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

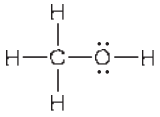


In this case, the nitrogen atom exhibits only _____ valence electrons.

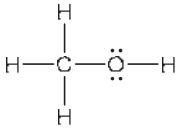
Step 3 Assign a formal charge to the nitrogen atom.

SkillBuilder 1.4 Locating Partial Charges

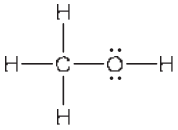
Step 1 Circle the bonds below that are polar covalent



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

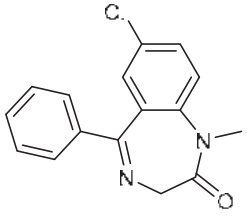


Step 3 Indicate the location of all partial charges ($\delta+$ and $\delta-$)

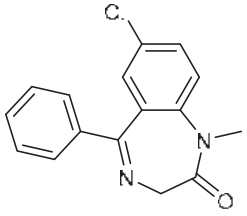


SkillBuilder 1.5 Reading Bond-Line Structures

Circle all carbon atoms in the compound below



Draw all hydrogen atoms in the compound below

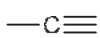


SkillBuilder 1.6 Identifying Hybridization States

A carbon atom with four σ bonds will be _____ hybridized.

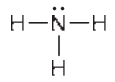
A carbon atom with three σ bonds will be _____ hybridized.

A carbon atom with two σ bonds will be _____ hybridized.



SkillBuilder 1.7 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 = ☐

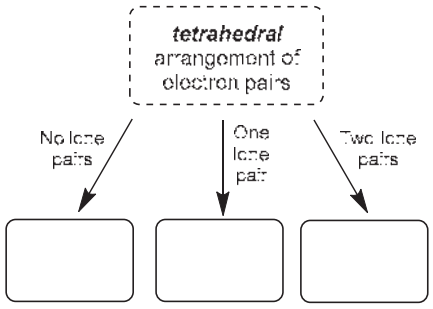
# of lone pairs = ☐

Steric Number = ☐

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

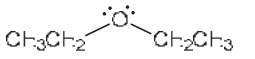
Steric #	Arrangement of Electron Pairs
4	
3	
2	

Step 3 Identify the geometry in each case below:



SkillBuilder 1.8 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).

☐

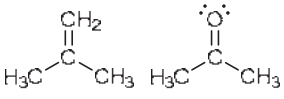
Step 3 Redraw the compound, and draw the net dipole moment:

☐

SkillBuilder 1.9 Predicting Relative Boiling Points

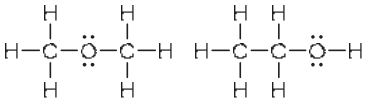
Dipole-Dipole Interactions

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



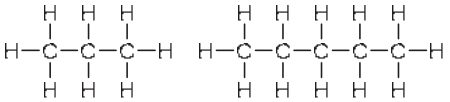
Hydrogen-Bonding Interactions

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 through 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.

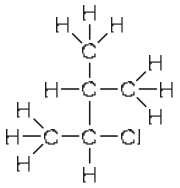
Chapter Opener

Welcome to the wonderful world of organic chemistry! Look at what we have in store…

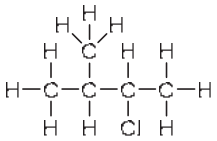
	**equipment** (snowboard, boots, goggles, backpack) that is lightweight and durable, **clothing** that is warm, waterproof, moisture-wicking and colorful, **sunscreen** that protects from sun damage and skin cancer
	**beverages** (organic compounds as sweeteners, flavors, colors), **food** (carbohydrates, proteins, fats, seasonings), **picnic accessories** (charcoal for cooking, plastic cups, insulated cooler, chair fabric), **car fluids** (gasoline, oil, antifreeze, lubricants), **car materials** (rubber tires, interior seats, dashboard), **toys** (inflatable guy, lava lamp wax, leather or vinyl basketball), **recreation** (skateboard wood, polyurethane wheels, helmet for safety, television parts)
	**carbon-based life forms** (aka all life forms: plants, animals, insects, microorganisms…all life is built with organic materials and functions with biochemistry), **Cows** (beef, milk, lactose, cheese, collagen, bone meal, leather, methane gas), **crop protection** (herbicides, pesticides, fungicides), **barn** (wood structure, fiberglass insulation, red paint), **colors in leaves** (chlorophyll, carotenoids, flavonoids, anthocyanins)
	**tennis equipment** (tennis racquet, strings, rubber ball and fuzzy cover), **tennis court** (paint for lines, nylon net, hard surface or grass), **sports clothing** (flexible, quick-drying, compression, tennis shoes, hair band), **gym bag contents** (deodorant, athletic tape, menthol cream for sore muscles, asthma inhaler, anti-inflammatories, contact lenses).
	**natural fibers** (cotton, wool, linen, hemp, cashmere, silk) and **synthetic fibers** (nylon, spandex, polyester, rayon, microfiber), **dyes** for colorful clothing, **laundromat supplies** (detergent, laundry pods, soap, fabric softener, dryer sheet, fragrance, natural gas or propane for dryer), **headphones** (plastic and foam materials)
	**tent** (canvas, nylon), **fishing gear** (pole, fishing line, lures), **hiking supplies** (bug spray, poison ivy, hydrocortisone, allergy pills), **tattoo and henna pigments**
	**school supplies** (paper, plastic and ink for pens, wood and graphite for pencils, eraser, adhesive on sticky notes), **computer** (keys, components, LCD screen), **coffee** (caffeine, coffee flavors, sugar, cream), **orange juice** (vitamin C, citric acid, orange flavors and color)
 	**personal care products** (shaving cream, mouthwash, toothbrush, dental floss, lotions, deodorants, toilet paper, Vaseline, lip balm, hair removers, wax), **cosmetics** (lipstick, mascara, makeup remover, nail polish, nail polish remover, styling gel, tanning creams), **cleansers** (shampoo, conditioner, soaps, body wash, facial cleansers), **health products** (vitamins, acne medicine, birth control pills), **cough medicine** (pain reliever, fever reducer, cough suppressant, expectorant, decongestant, grape flavor, sugar, purple color, plastic cup)

Solutions

1.1. (a) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom and find that both structures have the same formula ($C_5H_{11}Cl$):

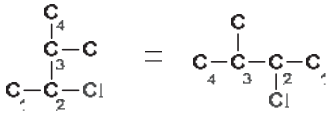


$C_5H_{11}Cl$



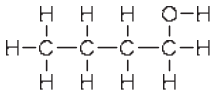
$C_5H_{11}Cl$

Next, the connectivity of the atoms must be determined. To simplify the drawings mentally, we disregard the hydrogen atoms, and we number the longest carbon chain on each skeleton from the end closest to the chlorine atom:

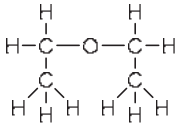


Comparison of these two drawings reveals that they represent the same compound (a four-carbon chain with a chlorine atom at C2 and a CH_3 group at C3).

1.1. (b) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom and find that both structures have the same formula ($C_4H_{10}O$):

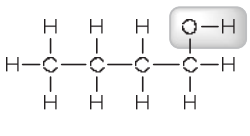


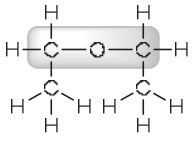
$C_4H_{10}O$



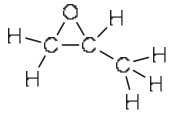
$C_4H_{10}O$

Next, the connectivity of the atoms must be determined. A comparison of the highlighted regions (below) indicates that these are different compounds (one has an OH group and the other does not). Because these two drawings represent different compounds with the same molecular formula, they are constitutional isomers.

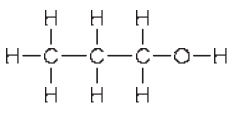




1.1. (c) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom and find that these structures do NOT have the same formula. As such, they are unrelated compounds.

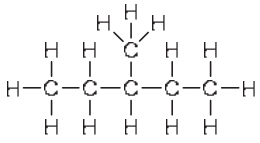


C_3H_6O

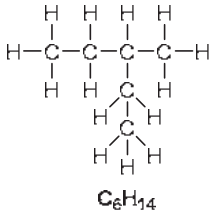


C_3H_8O

1.1. (d) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom and find that both structures have the same formula (C_6H_{14}):

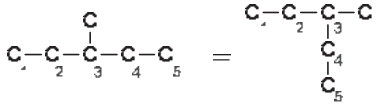


C_6H_{14}



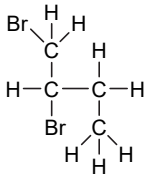
C_6H_{14}

Next, the connectivity of the atoms must be determined. To simplify the drawings mentally, we disregard the hydrogen atoms, and we number the longest carbon chain on each skeleton:

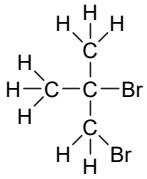


Comparison of these two drawings reveals that they represent the same compound (a five-carbon chain with a CH_3 group at C3).

1.1. (e) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom and find that both structures have the same formula ($C_4H_8Br_2$):

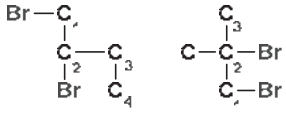


$C_4H_8Br_2$



$C_4H_8Br_2$

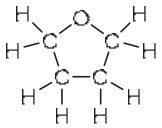
Next, the connectivity of the atoms must be determined. To simplify the drawings mentally, we disregard the hydrogen atoms, and we number the longest carbon chain on each skeleton from the end with an attached bromine atom:



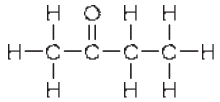
Comparison of these two drawings reveals that they do NOT represent the same compound. One has a four-carbon chain, and the other has only three carbon atoms in its longest chain. Because these two drawings represent different compounds with the same molecular formula, they are constitutional isomers.

1.1. (f) Begin by determining the molecular formula of each structure. We count the numbers of each type of atom

and find that both structures have the same formula (C_4H_8O):



C_4H_8O

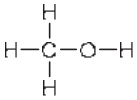


C_4H_8O

Next, the connectivity of the atoms must be determined. A quick comparison indicates that these are clearly different compounds (one has a ring and the other has a C═O group). These two drawings are constitutional isomers because they represent different compounds with the same molecular formula.

1.2. The carbon atoms are drawn in the center of the compound. The chlorine atoms and fluorine atoms are then placed at the periphery, as shown. We can tell that these are different compounds because only one has the three chlorine atoms all connected to the same carbon. In the other constitutional isomer, the chlorine atoms are distributed over both carbon atoms.

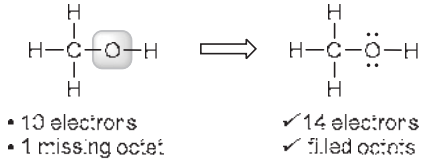
1.3. (a) First, connect the atoms that form more than one bond. Then, we connect all of the hydrogen atoms, as shown:



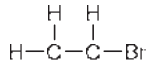
The next step is to calculate the number of electrons that will be included in the final structure. There is one carbon atom, and it contributes four valence electrons. The oxygen atom contributes six electrons, and each of the four hydrogen atoms brings one electron, for a total of 14 valence electrons.

C x1 = 4 electrons
O x1 = 6 electrons
H x4 = 4 electrons
14 = total # of valence electrons

Five bonds have already been drawn. These five bonds account for ten electrons, so four electrons still must be drawn. The oxygen atom lacks an octet, so that is where we add the electrons, as two lone pairs:



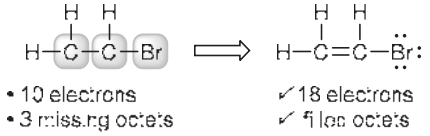
1.3. (b) First, connect the atoms that form more than one bond. In this case, the order in which the atoms are connected is implied by the given condensed formula: CH_2CHBr. Then, we connect all of the hydrogen atoms, as shown:



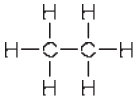
The next step is to calculate the number of electrons that will be included in the final structure. There are two carbon atoms, and they contribute a total of eight valence electrons. The bromine atom contributes seven electrons, and each of the three hydrogen atoms brings one electron, for a total of 18 valence electrons.

C x2 = 8 electrons
Br x1 = 7 electrons
H x3 = 3 electrons
18 = total # of valence electrons

Five bonds have already been drawn. These five bonds account for ten electrons, so eight electrons still must be drawn. Both carbon atoms are missing an octet. We know that carbon is tetravalent, so the octets can be filled by drawing a double bond between the two carbon atoms. The bromine atom also lacks an octet, so we add six electrons there, as three lone pairs:



1.3. (c) First, connect the atoms that form more than one bond. Then, we connect all of the hydrogen atoms, as shown:

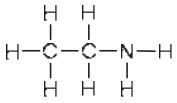


The next step is to calculate the number of electrons that will be included in the final structure. There are two carbon atoms, and they contribute a total of eight valence electrons. Each of the six hydrogen atoms brings one electron, for a total of 14 valence electrons.

C x2 = 8 electrons
H x6 = 6 electrons
14 = total # of valence electrons

Seven bonds have already been drawn, accounting for the expected 14 electrons. Both carbon atoms have filled octets, so the initially drawn Lewis structure is complete.

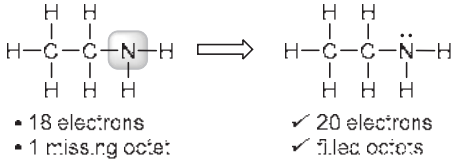
1.3. (d) First, connect the atoms that form more than one bond. In this case, the order in which the atoms are connected is implied by the given condensed formula: $CH_3CH_2NH_2$. Then, we connect all of the hydrogen atoms, as shown:



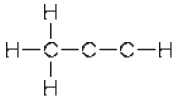
The next step is to calculate the number of electrons that will be included in the final structure. There are two carbon atoms, and they contribute a total of eight valence electrons. The nitrogen atom contributes five electrons, and each of the seven hydrogen atoms brings one electron, for a total of 20 valence electrons.

C x2 = 8 electrons
N x1 = 5 electrons
H x7 = 7 electrons
20 = total # of valence electrons

Nine bonds have already been drawn. These nine bonds account for 18 electrons, so two electrons still must be drawn. The nitrogen atom lacks an octet, so that is where we add the electrons, as a lone pair:



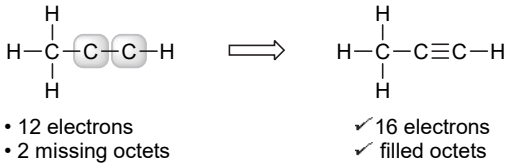
1.3. (e) First, connect the atoms that form more than one bond. In this case, the order in which the atoms are connected is implied by the given condensed formula: CH_3CCH. Then, we connect all of the hydrogen atoms, as shown:



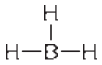
The next step is to calculate the number of electrons that will be included in the final structure. There are three carbon atoms, and they contribute a total of 12 valence electrons. Each of the four hydrogen atoms brings one electron, for a total of 16 valence electrons.

C x3 = 12 electrons
H x4 = 4 electrons
16 = total # of valence electrons

Six bonds have already been drawn. These six bonds account for 12 electrons, so four electrons still must be drawn. Two of the carbon atoms are missing octets. We know that carbon is tetravalent, so the octets can be filled by drawing a triple bond between these two carbon atoms:



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.

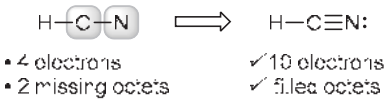


1.5. (a) First, connect the atoms that form more than one bond. Then, we connect the hydrogen atom (to the carbon atom, as implied by the given formula, HCN), as shown:

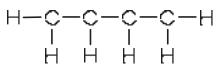
The next step is to calculate the number of electrons that will be included in the final structure. There is one carbon atom (four valence electrons), one nitrogen atom (five valence electrons), and one hydrogen atom (one electron), for a total of 10 valence electrons.

C x1 = 4 electrons
N x1 = 5 electrons
H x1 = 1 electron
10 = total # of valence electrons

Two bonds have already been drawn. These two bonds account for four electrons, so six electrons still must be drawn. The carbon atom is missing an octet. We know that carbon is tetravalent, so the octet can be filled by drawing a triple bond between the carbon atom and the nitrogen atom. The nitrogen atom still lacks an octet, so we add two electrons there, as a lone pair:



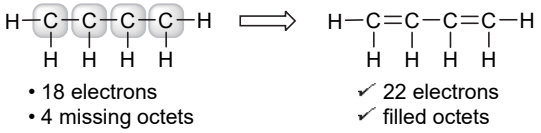
1.5. (b) First, connect the atoms that form more than one bond. In this case, the order in which the atoms are connected is implied by the given condensed formula: $CH_2CHCHCH_2$. Then, we connect all of the hydrogen atoms, as shown:



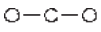
The next step is to calculate the number of electrons that will be included in the final structure. There are four carbon atoms, and they contribute a total of 16 valence electrons. Each of the six hydrogen atoms brings one electron, for a total of 22 valence electrons.

C x4 = 16 electrons
H x6 = 6 electrons
22 = total # of valence electrons

Nine bonds have already been drawn. These nine bonds account for 18 electrons, so four electrons still must be drawn. All four of the carbon atoms are missing octets. We know that carbon is tetravalent, so the octets can be filled by drawing two double bonds as shown:



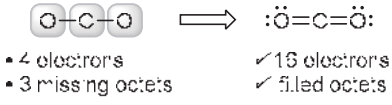
1.6. As described in the problem statement, carbon is the central atom in carbon dioxide:



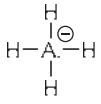
The next step is to calculate the number of electrons that will be included in the final structure. There is one carbon atom (four valence electrons) and two oxygen atoms (six valence electrons each), for a total of 16 valence electrons.

C x1 = 4 electrons
O x2 = 12 electrons
16 = total # of valence electrons

Two bonds have already been drawn. These two bonds account for four electrons, so 12 electrons still must be drawn. The carbon atom is missing an octet. We know that carbon is tetravalent, so the octet can be filled by drawing two double bonds. The oxygen atoms also lack octets, so we add four electrons to each oxygen atom, as two lone pairs:

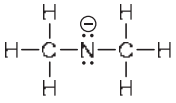


1.7. (a) Aluminum is in group 3A of the periodic table, and it should therefore have three valence electrons. In this case, the aluminum atom exhibits four valence electrons (one for each bond). With one extra electron, this aluminum atom will bear a negative charge:

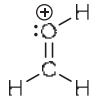


1.7. (b) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge:

1.7. (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 electrons (one for each bond and two for each lone pair). With one extra electron, this nitrogen atom will bear a negative charge:



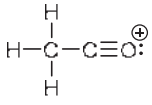
1.7. (d) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge.



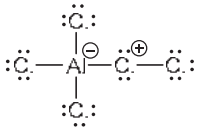
1.7. (e) Carbon is in group 4A of the periodic table, and it should therefore have four valence electrons. In this case, the carbon atom exhibits five electrons (one for each bond and two for the lone pair). With one extra electron, this carbon atom will bear a negative charge:

1.7. (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 electrons (one for each bond). This carbon atom is missing an electron, and it therefore bears a positive charge:

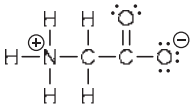
1.7. (g) Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom exhibits only five electrons (one for each bond, and two for the lone pair). This oxygen atom is missing an electron, and it therefore bears a positive charge:



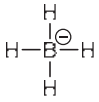
1.7. (h) Two of the atoms in this structure exhibit formal charges because each of these atoms does not exhibit the appropriate number of electrons. The aluminum atom (group 3A) should have three valence electrons, but it exhibits four (one for each bond). With one extra electron, this aluminum atom 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:



1.7. (i) Two of the atoms in this structure exhibit formal charges 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 oxygen atoms (the one on the right) exhibits seven 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:

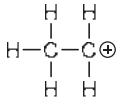


1.8. (a) The boron atom in this case exhibits four electrons (one for each bond), although boron (group 3A) should only have three valence electrons. With one extra electron, this boron atom bears a negative charge:

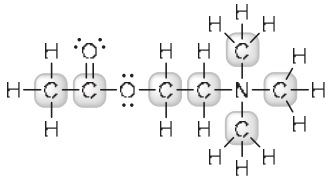


1.8. (b) Nitrogen is in group 5A of the periodic table, so a nitrogen atom should have five valence electrons. A negative charge indicates one extra electron, so this nitrogen atom must exhibit six electrons (one for each bond and two for each lone pair):

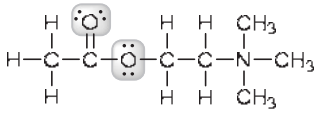
1.8. (c) One of the carbon atoms (below right) exhibits three electrons (one for each bond), but carbon (group 4A) is supposed to have four valence electrons. It is missing one electron, so this carbon atom therefore bears a positive charge:



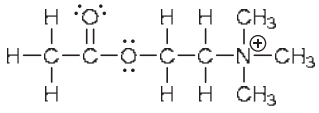
1.9. Carbon is in group 4A of the periodic table, and it should therefore have four valence electrons. Every carbon atom in acetylcholine has four bonds, thus exhibiting the correct number of electrons (four) and having no formal charge.



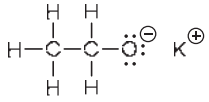
Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. Each oxygen atom in acetylcholine has two bonds and two lone pairs of electrons, so each oxygen atom exhibits six electrons (one for each bond, and two for each lone pair). With the correct number of electrons,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

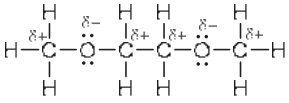


1.10. In this case, the order in which the atoms are connected is implied by the given condensed formula: CH_3CH_2OK. We expect covalent bonds between the elements that are nonmetals (carbon, hydrogen and oxygen), but the bond between oxygen (a nonmetal) and potassium (a metal) must be ionic. To become a cation, potassium must have lost its one valence electron, so no electrons are drawn on the K^+ ion:

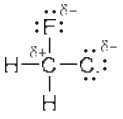


Oxygen is in group 6A of the periodic table, and it should therefore have six valence electrons. In this case, the oxygen atom bears a negative charge so it must exhibit one extra electron. Drawing three lone pairs on the oxygen atom fills its octet and attributes seven electrons to the oxygen atom (one for the bond and two for each lone pair).

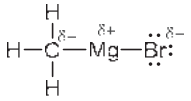
1.11. (a) Oxygen is more electronegative than carbon, and a C─O bond is polar covalent. For each C─O bond, the O will be electron-rich (δ–), and the C will be electron-deficient (δ+), as shown below:



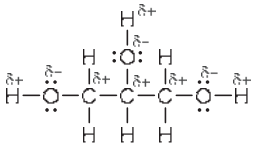
1.11. (b) Fluorine is more electronegative than carbon, and a C─F bond is polar covalent. For a C─F bond, the F will be electron-rich (δ–), and the C will be electron-deficient (δ+). Chlorine is also more electronegative than carbon, so a C─Cl bond is also polar covalent. For a C─Cl bond, the Cl will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



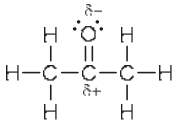
1.11. (c) Carbon is more electronegative than magnesium, so the C will be electron-rich (δ–) in a C─Mg bond, and the Mg will be electron-deficient (δ+). Also, bromine is more electronegative than magnesium. So in a Mg─Br bond, the Br will be electron-rich (δ–), and the Mg will be electron-deficient (δ+),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



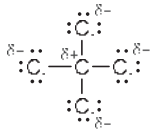
1.11. (d) Oxygen is more electronegative than carbon or hydrogen, so all C─O bonds and all O─H bond are polar covalent. For each C─O bond and each O─H bond, the O will be electron-rich (δ–), and the C or H will be electron-deficient (δ+), as shown below:



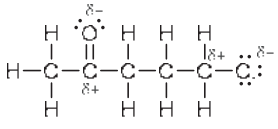
1.11. (e) Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ–) and the C will be electron-deficient (δ+) in a C═O bond, as shown below:



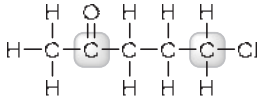
1.11. (f) Chlorine is more electronegative than carbon. As such, for each C─Cl bond, the Cl will be electron-rich (δ–) and the C will be electron-deficient (δ+), as shown below:



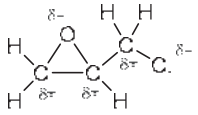
1.12. Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ–) and the C will be electron-deficient (δ+) in a C═O bond. In addition, chlorine is more electronegative than carbon. So for a C─Cl bond, the Cl will be electron-rich (δ–) and the C will be electron-deficient (δ+), as shown below:



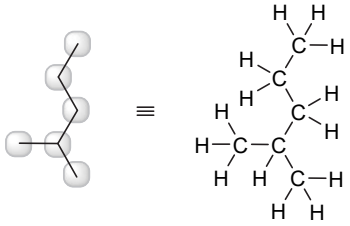
Notice that two carbon atoms are electron-deficient (δ+). These are the positions (highlighted below) that are most likely to be attacked by an electron-rich anion, such as hydroxide:



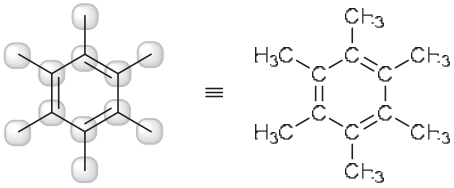
1.13. Oxygen is more electronegative than carbon. As such, the O will be electron-rich (δ−) and the C will be electron-deficient (δ+) in a C─O bond. In addition, chlorine is more electronegative than carbon. So for a C─Cl bond, the Cl will be electron-rich (δ−) and the C will be electron-deficient (δ+),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



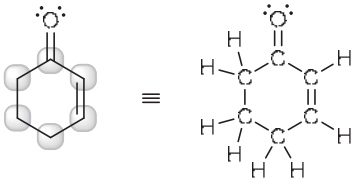
1.14. (a) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has six carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



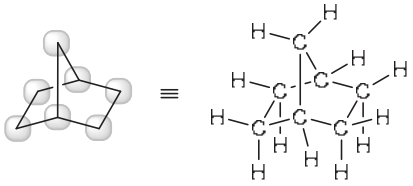
1.14. (b) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has twelve carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



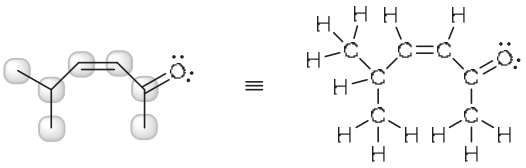
1.14. (c) Each corner represents a carbon atom (highlighted below), so this compound has six carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



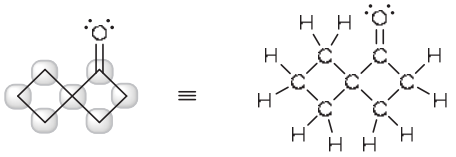
1.14. (d) Each corner represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



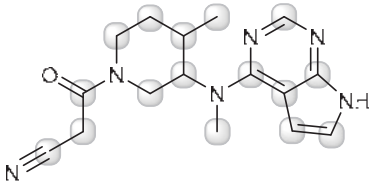
1.14. (e) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



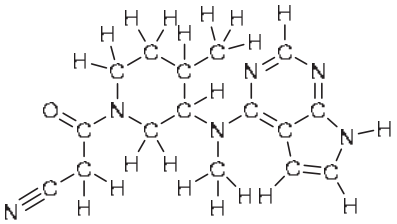
1.14. (f) Each corner represents a carbon atom (highlighted below), so this compound has seven carbon atoms. Each carbon atom is connected to enough hydrogen atoms to have a total of four bonds, as shown:



1.15. 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:



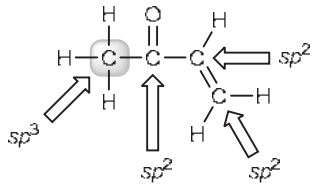
1.16. In cyclopropane, each carbon atom is sp^3 hybridized, so the ideal bond angles are 109.5º. However, the angles of an equilateral triangle are 60º (much smaller than 109.5º). Therefore, each bond angle in cyclopropane is severely strained, causing an increase in energy. This form of strain, called 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.

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

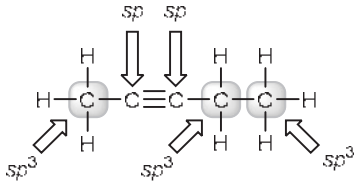
1.17. (b) Each C─H bond is formed from the interaction between an sp^2 hybrid orbital from carbon and an *s* orbital from hydrogen.

1.18. 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.

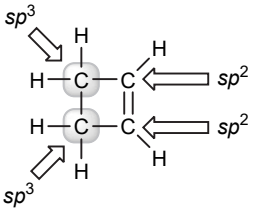
1.19. (a) The highlighted carbon atom (below) has four σ bonds and is therefore sp^3 hybridized. The other carbon atoms in this structure are all sp^2 hybridized, because each of them has three attached atoms (three σ bonds):



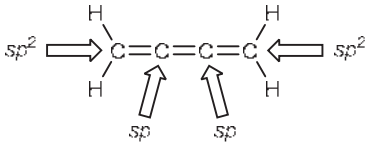
1.19. (b) Each of the highlighted carbon atoms has four σ bonds and is therefore sp^3 hybridized. The other two carbon atoms in this structure are *sp* hybridized, because each has two attached atoms (two σ bonds):



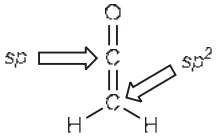
1.19. (c) Each of the highlighted carbon atoms (below) has four σ bonds and is therefore sp^3 hybridized. The other two carbon atoms in this structure are sp^2 hybridized, because each has three attached atoms (three σ bonds):



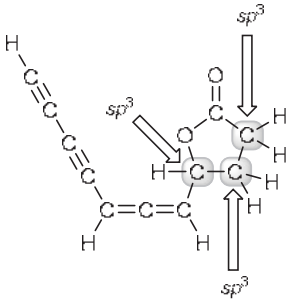
1.19. (d) Each of the two central carbon atoms has two attached atoms (two σ bonds), and as such, each of these carbon atoms is *sp* hybridized. The other two carbon atoms (the outer ones) are sp^2 hybridized because each has three attached atoms (three σ bonds):



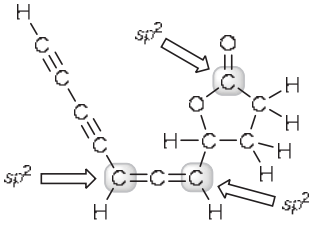
1.19. (e) One of the carbon atoms (the one connected to oxygen) has two attached atoms (two σ bonds), and as such, it is *sp* hybridized. The other carbon atom is sp^2 hybridized because it has three attached atoms (three σ bonds):



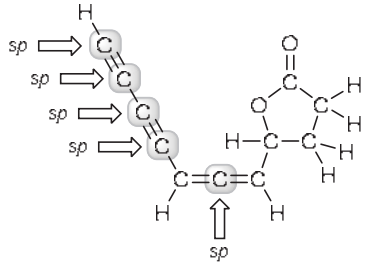
1.20. Each of the following three highlighted carbon atoms has four σ bonds, and is therefore sp^3 hybridized:



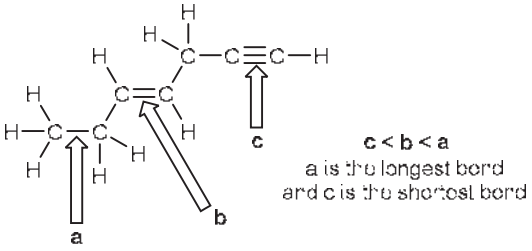
And each of the following three highlighted carbon atoms has three attached atoms (three σ bonds) and is therefore sp^2 hybridized:



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



1.21. 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 1.3).



1.22. (a) In this structure, the boron atom has four attached atoms 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.

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

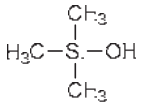
1.22. (c) In this structure, the nitrogen atom has four attached atoms and no lone pairs, giving a total of four electron pairs (steric number = 4). VSEPR theory therefore predicts a tetrahedral arrangement of electron pairs. Since all of the electron pairs are bonds, the structure is expected to have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

1.22. (d) The carbon atom has four attached atoms 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.

1.23. In the carbocation, the carbon atom has three attached atoms (three σ 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 attached atoms (three σ 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 of approximately 107°.

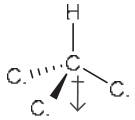
1.24. In ammonia, the nitrogen atom has three attached atoms and one lone pair. Therefore, VSEPR theory predicts trigonal pyramidal geometry, with bond angles of approximately 107°. In the ammonium ion, the nitrogen atom has four attached atoms and no lone pairs, so VSEPR theory predicts tetrahedral geometry, with bond angles of 109.5°. Therefore, when ammonia is converted into an ammonium ion, we predict that the bond angles will increase (by approximately 2.5°).

1.25. The silicon atom has four attached atoms and no lone pairs:

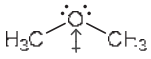


So the steric number is 4 (sp^3 hybridization), which means that the arrangement of electron pairs will be tetrahedral. With no lone pairs, the arrangement of the atoms (geometry) is the same as the electronic arrangement. It is tetrahedral.

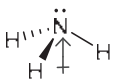
1.26. (a) This compound has three C─Cl bonds, each of which exhibits a dipole moment. To determine if these dipole moments cancel each other, we must identify the molecular geometry. The central carbon atom has four attached atoms so we expect tetrahedral geometry. As such, the three polar C─Cl bonds 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:



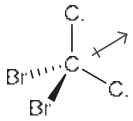
1.26. (b) The oxygen atom has two attached atoms and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. As such, the dipole moments associated with the polar C─O bonds do not fully cancel each other, and the dipole moments associated with the lone pairs also do not fully cancel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



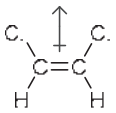
1.26. (c) The nitrogen atom has three attached atoms and one lone pair (steric number = 4), and VSEPR theory predicts trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair). As such, the dipole moments associated with the polar N─H bonds do not fully cancel each 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:



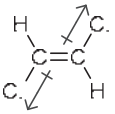
1.26. (d) The central carbon atom has four attached atoms (steric number = 4), and VSEPR theory predicts tetrahedral geometry. There are individual dipole moments associated with each of the C─Cl bonds and each of the C─Br bonds. If all four dipole moments had the same magnitude, then we would expect 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:



1.26. (e) 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:

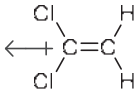


1.26. (f) Each C─Cl bond has a dipole moment, and in this case, they are pointing in opposite directions. As such, they fully cancel each other, giving no net molecular dipole moment.

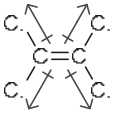


No net molecular dipole moment

1.26. (g) 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:



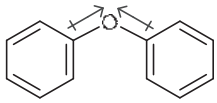
1.26. (h) Each C─Cl bond has a dipole moment, but in this case, they fully cancel each other to give no net molecular dipole moment.



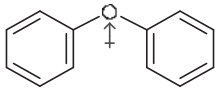
No net molecular dipole moment

1.27. Each of the C─O bonds has an individual dipole moment. 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.

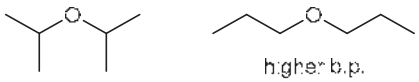
As such, the dipole moments associated with the polar C─O bonds do NOT fully cancel each other,



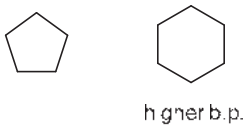
and the dipole moments associated with the lone pairs also do not fully cancel each other. As a result, there is a net molecular dipole moment, as shown:



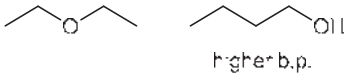
1.28. (a) Both compounds have the same molecular weight because they are isomers (they both have the molecular formula $C_6H_{14}O$). The second compound is expected to have a higher boiling point, because it is less branched. The greater surface area in the second compound results in greater London dispersion forces and a higher boiling point (b.p.):



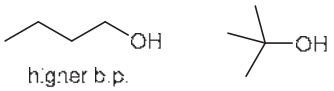
1.28. (b) The second compound is expected to have a higher boiling point, because it has more carbon atoms and therefore a higher molecular weight. Larger compounds have greater London dispersion forces and a higher boiling point (b.p.):



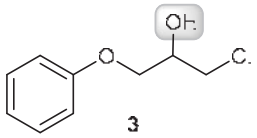
1.28. (c) The second compound is expected to have a higher boiling point, because it has an O─H bond, which will lead to hydrogen-bonding interactions between molecules.



1.28. (d) Both compounds have the same molecular weight because they are isomers (they both have the molecular formula $C_4H_{10}O$), 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.):



1.29. Compound **3** is expected to have a higher boiling point than compound **4**, because only compound **3** has an O─H group. The O─H group serves as both a hydrogen bond donor and a hydrogen bond acceptor, so molecules of **3** will exhibit hydrogen bonding. In contrast, molecules of **4** cannot form hydrogen bonds between themselves, so compound **4** will have a lower boiling point. When this mixture is heated, the lower boiling compound (**4**) can be collected first, leaving behind compound **3**.

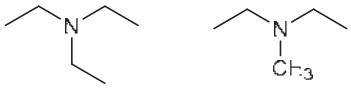


3

1.30. (a) Both compounds are polar and can form hydrogen bonds with water. However, the first compound is expected to be more soluble in water, because it has O─H groups that can serve as both hydrogen bond donors and hydrogen bond acceptors, so this compound will exhibit more extensive hydrogen bonding with water. For your information, the solubility of each compound in room temperature water is provided:

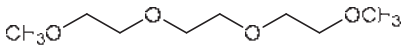
more soluble (miscible with water) **less soluble** (20 g per 100 mL)

1.30. (b) Both compounds are polar and can form hydrogen bonds with water. However, the second compound is expected to be more soluble in water, because the CH_3 group serves as a smaller nonpolar region than a CH_2CH_3 group. For your information, the solubility of each compound in room temperature water is provided:



less soluble (11 g per 100 mL) **more soluble** (31 g per 100 mL)

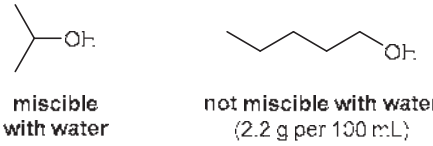
1.31. (a) The first compound (shown below) is polar and can form hydrogen bonds with water, so it must be the one that is miscible in water. The second compound is nonpolar and is expected to have very little water solubility.



1.31. (b) The second compound (shown below) is polar and can form hydrogen bonds with water, so it must be the one that is miscible in water. The first compound is

nonpolar and is expected to have very little water solubility.

1.31. (c) Both compounds are polar and can form hydrogen bonds with water, but the first compound has a shorter nonpolar carbon chain, so it must be the one that is miscible in water. For your information, the solubility of the second compound in room temperature water is provided:

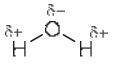


1.32. Polymer A has numerous O─H groups. Each O─H group serves as both a hydrogen bond donor and a hydrogen bond acceptor, so Polymer A will exhibit extensive hydrogen bonding with water (making it sticky). Polymer A represents starch, and it can be used to make biodegradable packing peanuts (they readily decompose when exposed to moisture). On the other hand, Polymer B has only carbon and hydrogen atoms, so it is nonpolar and will not dissolve in water. Polymer B is polystyrene, which is the plastic used to make traditional packing peanuts.

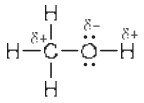
1.33. (a) According to Table 1.1, the difference in electronegativity between Br and H is 2.8 – 2.1 = 0.7, so an H─Br bond is expected to be polar covalent. Since bromine is more electronegative than hydrogen, the Br will be electron-rich (δ–), and the H will be electron-deficient (δ+), as shown below:

1.33. (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-deficient (δ+), as shown below:

1.33. (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-deficient (δ+), as shown below:



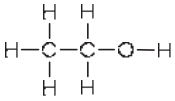
1.33. (d) According to Table 1.1, oxygen (3.5) is more electronegative than carbon (2.5) or hydrogen (2.1), and a C─O or H─O bond is polar covalent. For each C─O or H─O bond, the O will be electron-rich (δ–), and the C or H will be electron-deficient (δ+), as shown below:



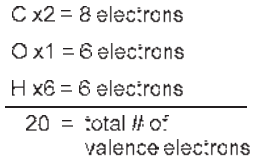
1.34. (a) According to Table 1.1, the difference in electronegativity between Na (0.9) and Br (2.8) is greater than the difference in electronegativity between H (2.1) and Br (2.8). Therefore, NaBr is expected to have more ionic character than HBr.

1.34. (b) According to Table 1.1, 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.

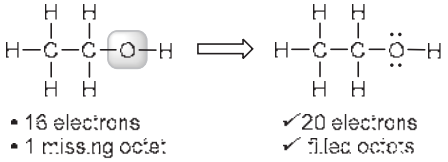
1.35. (a) The connectivity of atoms is implied by the given condensed formula (CH_3CH_2OH):



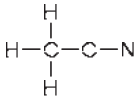
The next step is to calculate the number of electrons that will be included in the final structure. There are two carbon atoms (eight valence electrons), one oxygen atom (six valence electrons), and each of the six hydrogen atoms brings one electron, for a total of 20 valence electrons.



Eight bonds have already been drawn. These eight bonds account for 16 electrons, so four electrons still must be drawn. The oxygen atom lacks an octet, so that is where we add the electrons, as two lone pairs:



1.35. (b) The connectivity of atoms is implied by the given condensed formula (CH_3CN):

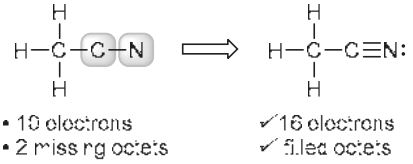


The next step is to calculate the number of electrons that will be included in the final structure. There are two

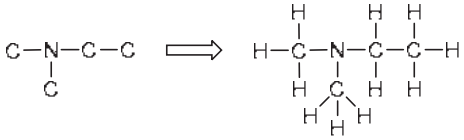
carbon atoms (eight valence electrons), one nitrogen atom (five valence electrons), and each of the three hydrogen atoms brings one electron, for a total of 16 valence electrons.

C x2 = 8 electrons
N x1 = 5 electrons
H x3 = 3 electron
16 = total # of valence electrons

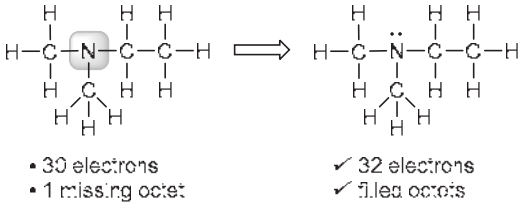
Five bonds have already been drawn. These five bonds account for ten electrons, so six electrons still must be drawn. The carbon atom is missing an octet. We know that carbon is tetravalent, so the octet can be filled by drawing a triple bond between the highlighted carbon and nitrogen atoms. The nitrogen atom still lacks an octet, so we add two electrons there, as a lone pair:



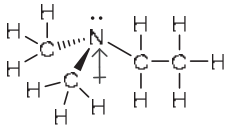
1.36. We begin by connecting the atoms that have more than one bond (the problem statement indicates how we should connect them). Then, we connect all of the hydrogen atoms (ensuring that each carbon atom has four bonds), as shown:



Usually, the next step would be the calculation of the number of valence electrons, but this becomes very time-consuming for larger molecules! Instead, we can simply complete the Lewis structure by adding lone pairs and/or pi bonds to complete any missing octets. In this case, we note that only the nitrogen atom has an incomplete octet, and that adding a lone pair would solve the problem. By being familiar with typical bonding patterns for uncharged atoms, we can complete Lewis structures by inspection. By the way, the number of total valence electrons is 32, if you would like to check your work.



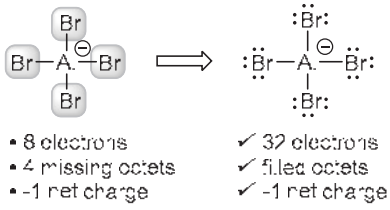
The nitrogen atom has three attached atoms (three σ bonds) and one lone pair, so the steric number is 4, which means that the arrangement of electron pairs is expected to be tetrahedral. One corner of the tetrahedron is occupied by a lone pair, so the geometry of the nitrogen atom (the arrangement of 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. To calculate the number of electrons that will be included in the final structure, we must account for the overall charge. A net charge of −1 indicates that there is one *extra* electron, so that is added to the total. There is one aluminum atom (three valence electrons) and each of the four bromine atoms brings seven valence electrons. After adding one additional electron to account for the charge, we find that there will be a total of 32 electrons shown in the Lewis structure.

Al x1 = 3 electrons
Br x4 = 28 electrons
-1 charge = 1 extra electron
32 = total # of valence electrons

After attaching the four bromine atoms to aluminum, we find that the aluminum atom has four electrons attributed to it, rather than three, which gives it a negative formal charge. Each bromine atom is missing an octet and requires the addition of three lone pairs to complete the octets and to complete the Lewis structure:



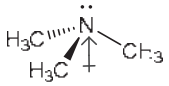
The aluminum atom has four bonds and no lone pairs, so the steric number is 4, which means that this aluminum atom will have tetrahedral geometry.

1.38. Beta-carotene is composed of only carbon atoms and hydrogen atoms, making it a completely nonpolar molecule. There are no polar bonds that would contribute to a molecular dipole moment. Using the "like dissolves like" principle, we expect β-carotene to have no solubility in a highly polar solvent such as water (Figure 1.53) and good solubility in a nonpolar solvent such as benzene (as illustrated with another nonpolar solvent in Figure 1.56).

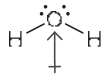
1.39. (a) C─H bonds are considered to be nonpolar, although they do have a very small dipole moment, because there is a small difference in electronegativity between carbon (2.5) and hydrogen (2.1). Nevertheless, even these small dipole moments fully cancel each other

out because of the tetrahedral geometry of CH_4, so the molecule does not have a molecular dipole moment.

1.39. (b) The nitrogen atom has trigonal pyramidal geometry. As such, the dipole moments associated with the polar N─C 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:



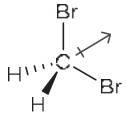
1.39. (c) The oxygen atom has two attached atoms (two σ bonds) and two lone pairs (steric number = 4), and VSEPR predicts bent geometry. As such, the dipole moments associated with the polar O─H bonds do not cancel each other, and the dipole moments associated with the lone pairs also do not fully cancel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



1.39. (d) The central carbon atom of carbon dioxide (CO_2) has two attached atoms (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, so there is no net molecular dipole moment.

1.39. (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.

1.39. (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 (bond angles of 109.5º). 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.

1.40. (b) As indicated in Figure 1.10, a neutral fluorine atom has two *1s* electrons, two *2s* electrons, and five *2p* electrons.

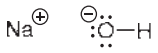
1.40. (c) As indicated in Figure 1.10, a neutral carbon atom has two *1s* electrons, two *2s* electrons, and two *2p* electrons.

1.40. (d) As indicated in Figure 1.10, a neutral nitrogen atom has two *1s* electrons, two *2s* electrons, and three *2p* electrons.

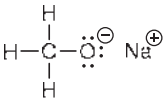
1.40. (e) In this electron configuration, the second shell is completely full, and the third shell is the valence shell. The third shell has seven valence electrons (two 3*s* electrons, and five 3*p* electrons), so this is the electron configuration of a neutral chlorine atom.

1.41. (a) There is a large difference in electronegativity between sodium (a metal) and bromine (a nonmetal). Therefore, the NaBr bond is classified as ionic.

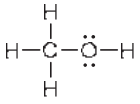
1.41. (b) There is a large difference in electronegativity between sodium (a metal) and oxygen (a nonmetal). Therefore, the bond between these two atoms is classified as ionic. In contrast, the O─H bond is polar covalent, because oxygen and hydrogen are both nonmetals, and oxygen is more electronegative. Note the different representations of ionic and covalent bonds in the Lewis structure of NaOH:



1.41. (c) There is a large difference in electronegativity between sodium (a metal) and oxygen (a nonmetal). Therefore, the bond between these two atoms is classified as ionic. The C─O bond is polar covalent, because oxygen and carbon are both nonmetals, and oxygen is more electronegative. Each of the C─H bonds is nonpolar covalent, because carbon and hydrogen are both nonmetals, and they have very similar electronegativity values (carbon = 2.5 and hydrogen = 2.1). Note the different representations of ionic and covalent bonds in the Lewis structure of $NaOCH_3$:



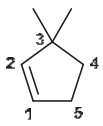
1.41. (d) All bonds in CH_3OH are covalent, because only nonmetals are present. Each of the C─H bonds is nonpolar covalent, because the electronegativity values are very similar for carbon (2.5) and hydrogen (2.1). On the other hand, oxygen is more electronegative than both carbon and hydrogen, so the O─H and C─O bonds are both described as polar covalent.



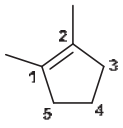
1.41. (e) All bonds in CH_2O are covalent, because only nonmetals are present. Each of the C─H bonds is nonpolar covalent, because the electronegativity values are very similar for carbon (2.5) and hydrogen (2.1). On

the other hand, oxygen is more electronegative than carbon, so the C═O double bond is described as polar covalent.

1.42. (a) These compounds both have the same molecular formula (C_7H_{12}), but they differ in their connectivity of atoms, or constitution. Therefore, they are constitutional isomers.

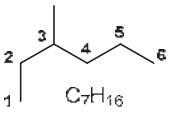


C_7H_{12}

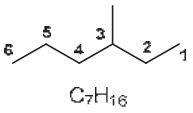


C_7H_{12}

1.42. (b) These structures have the same molecular formula (C_7H_{16}), AND they have the same constitution (connectivity of atoms), so they represent the same compound.

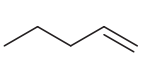


C_7H_{16}

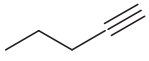


C_7H_{16}

1.42. (c) The first compound has the molecular formula C_5H_{10}, while the second compound has the molecular formula C_5H_8. As such, they are different compounds that are not isomeric.



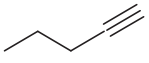
C_5H_{10}



C_5H_8

1.42. (d) These compounds both have the same molecular formula (C_5H_8), but they differ in their connectivity of atoms, or constitution. Therefore, they are constitutional isomers.

C_5H_8



C_5H_8

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

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

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

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

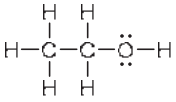
1.43. (e) Chlorine is more electronegative than carbon, and the withdrawal of electron density toward chlorine can be indicated with the following arrow:

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

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

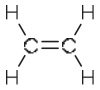
1.43. (h) Nitrogen is more electronegative than hydrogen, and the withdrawal of electron density toward nitrogen can be indicated with the following arrow:

1.44. (a) The oxygen atom has two attached atoms (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 (like the bond angles in H_2O), and all other bonds angles are expected to be approximately 109.5º (because each carbon atom has four σ bonds and tetrahedral geometry).

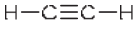


1.44. (b) The central carbon atom has three attached atoms (three σ bonds) and no lone pairs (steric number = 3), and VSEPR theory predicts trigonal planar geometry. As such, all bond angles are approximately 120º.

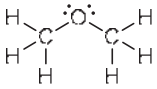
1.44. (c) Each of the carbon atoms has three attached atoms (three σ bonds) and no lone pairs (steric number = 3), and VSEPR theory predicts trigonal planar geometry. As such, all bond angles are approximately 120º.



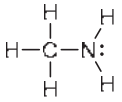
1.44. (d) Each of the carbon atoms has two attached atoms (two σ bonds) and no lone pairs (steric number = 2), and VSEPR theory predicts linear geometry. As such, all bond angles are 180º.



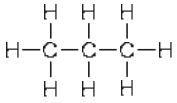
1.44. (e) The oxygen atom has two attached atoms (two σ bonds) and two lone pairs (steric number = 4), and VSEPR theory predicts bent geometry. Therefore, the C─O─C bond angle is expected to be around 105º (like the bond angles in H_2O). The remaining bond angles are all expected to be approximately 109.5º (because each carbon atom has four σ bonds and tetrahedral geometry).



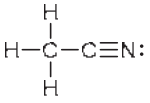
1.44. (f) The nitrogen atom has three attached atoms (three σ bonds) and one lone pair (steric number = 4), and VSEPR theory predicts trigonal pyramidal geometry, with bond angles of 107º. The carbon atom is also tetrahedral (because it has four σ bonds), although the bond angles around the carbon atom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



1.44. (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.



1.44. (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 attached atoms (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.

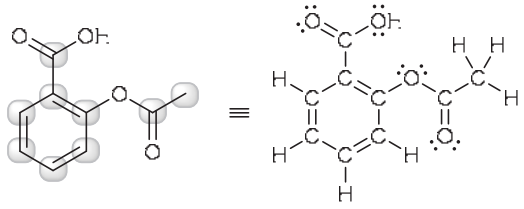
1.45. (a) The nitrogen atom has three attached atoms (three σ bonds) and one lone pair (steric number = 4). It is sp^3 hybridized (electronically tetrahedral), with trigonal pyramidal geometry (because one corner of the tetrahedron is occupied by a lone pair).

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

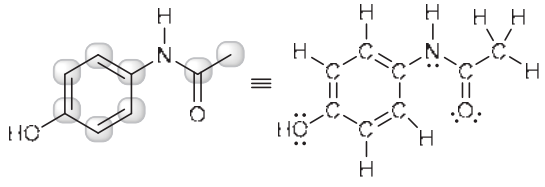
1.45. (c) This carbon atom has three attached atoms (three σ bonds) and no lone pairs (steric number = 3). It is sp^2 hybridized, with trigonal planar geometry.

1.45. (d) This carbon atom has three attached atoms (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).

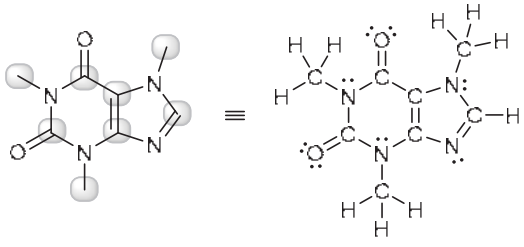
1.46. (a) Each corner and each endpoint represents a carbon atom (highlighted), so this compound has nine carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, as shown.



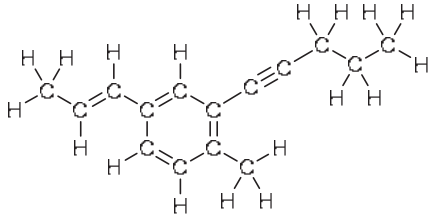
1.46. (b) Each corner and each endpoint represents a carbon atom (highlighted below), so this compound has eight carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, as shown.



1.46. (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.

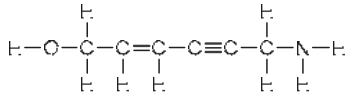


1.47. Each corner and each endpoint represents a carbon atom, so this compound has fifteen carbon atoms. Each carbon atom will have enough hydrogen atoms to have exactly four bonds, giving a total of eighteen hydrogen atoms, as shown here:



$C_{15}H_{18}$

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.

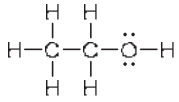


1.49. (a) The second compound is expected to have a higher boiling point, because it has an O─H bond, which will lead to hydrogen bonding interactions between molecules.

1.49. (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.

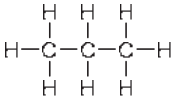
1.49. (c) The first compound has a C═O bond, which has a strong dipole moment, while the second compound is nonpolar. The first compound is therefore expected to exhibit strong dipole-dipole interactions and to have a higher boiling point than the second compound.

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

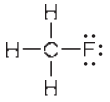


1.50. (b) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pairs can serve as hydrogen-bond acceptors, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect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

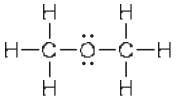
1.50. (c) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound will not exhibit hydrogen bonding interactions.



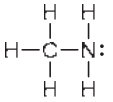
1.50. (d) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pairs can serve as hydrogen-bond acceptors, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect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



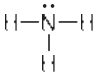
1.50. (e) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pairs can serve as hydrogen-bond acceptors, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect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



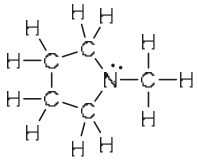
1.50. (f) This compound possesses two N─H bonds, so it is expected to exhibit hydrogen bonding interactions.



1.50. (g) This compound possesses N─H bonds, so it is expected to exhibit hydrogen bonding interactions.

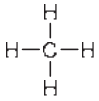


1.50. (h) This compound lacks a hydrogen atom that is connected to an oxygen or nitrogen atom. Therefore, this compound cannot serve as a hydrogen-bond donor (although the lone pair can serve as hydrogen-bond acceptor, in the presence of a hydrogen-bond donor). As a pure compound, we do not expect there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



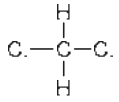
1.51. (a) Boron is in group 3A of the periodic table, and therefore has three valence electrons. It can use each of its valence electrons to form a bond, so we expect the molecular formula to be BH_3 (x=3).

1.51. (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).

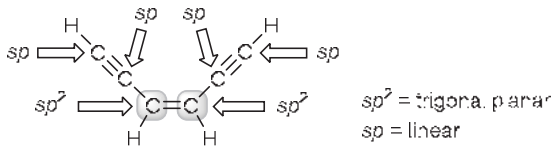


1.51. (c) Nitrogen is in group 5A of the periodic table, and therefore has five valence electrons. But it cannot form five bonds, because it only has four orbitals with which to form bonds. One of those orbitals must be occupied by a lone pair (two electrons), and each of the remaining three electrons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

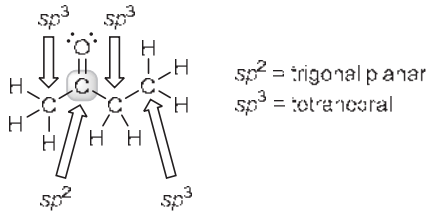
1.51. (d) Carbon is in group 4A of the periodic table, and therefore has four valence electrons. It can use each of its valence electrons to form a bond, and indeed, we expect the carbon atom to have four bonds. Two of the bonds are with hydrogen atoms, so the other two bonds must be with chlorine atoms.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



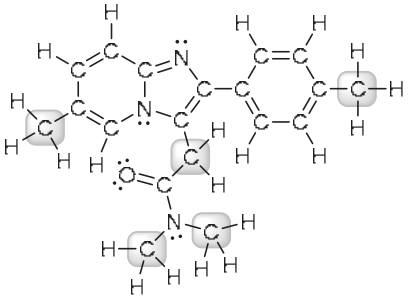
1.52. (a) Each of the highlighted carbon atoms has three attached atoms (three σ bonds) and no lone pairs (steric number = 3). Each of these carbon atoms is sp^2 hybridized, with trigonal planar geometry. Each of the other four carbon atoms has two attached atoms (two σ bonds) and no lone pairs (steric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



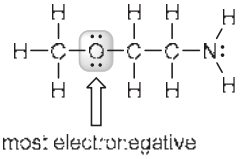
1.52. (b) The highlighted carbon atom has three attached atoms (three σ bonds) and no lone pairs (steric number = 3). This carbon atom is sp^2 hybridized, with trigonal planar geometry. Each of the other three carbon atoms has four σ bonds (steric number = 4). Those three carbon atoms are all sp^3 hybridized,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



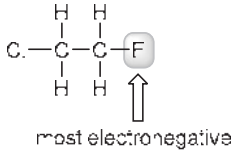
1.53. Each of the highlighted carbon atoms has four σ bonds (steric number = 4), and is sp^3 hybridized, with tetrahedral geometry. Each of the other fourteen carbon atoms in this structure has three attached atoms (three σ bonds) and no lone pairs (steric number = 3). Each of these fourteen carbon atoms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



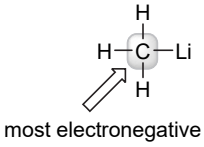
1.54. (a) Oxygen is the most electronegative atom in this compound. See Table 1.1 for electronegativity values.



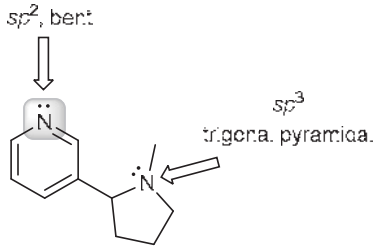
1.54. (b) Fluorine is the most electronegative atom. See Table 1.1 for electronegativity values.



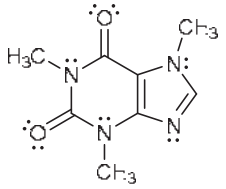
1.54. (c) Carbon is the most electronegative atom in this compound. See Table 1.1 for electronegativity values.



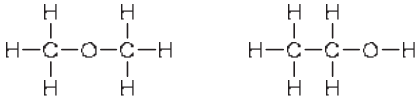
1.55. The highlighted nitrogen atom has two attached atoms (two σ bonds) and one lone pair (steric number = 3). This nitrogen atom is sp^2 hybridized. It is electronically trigonal planar, but one of the sp^2 hybridized orbitals is occupied by a lone pair, so the geometry (arrangement of atoms) is bent.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

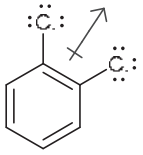


1.57. As seen in Section 1.2, the following two compounds have the molecular formula C_2H_6O.

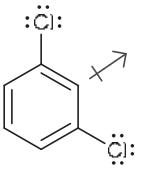


The second compound will have a higher boiling point because it possesses an OH group which can exhibit hydrogen bonding interactions.

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:

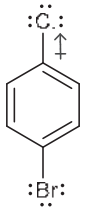


1.58. (b) Each C─Cl bond has a dipole moment, and the two dipole moments do not fully cancel each other because the polar bonds are not pointing in opposite directions. As such, there will be a net molecular dipole moment, as shown here:



1.58. (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.

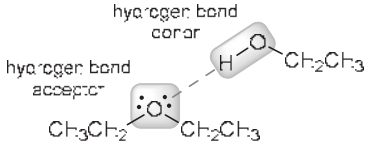
1.58. (d) The C─Cl bond has a dipole moment, and the C─Br bond also has a dipole moment. These two dipole moments are in opposite directions, but they do not have the same magnitude. The C─Cl bond has a larger dipole moment than the C─Br bond, because chlorine is more electronegative than bromine. Therefore, 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.

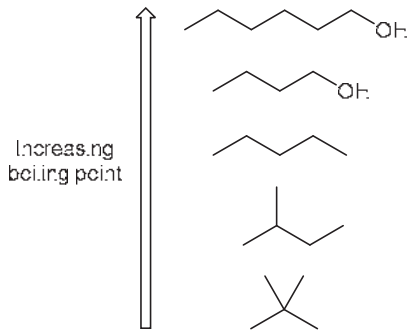
1.60. $CHCl_3$ is expected to have a larger molecular dipole moment than $CBrCl_3$, because the bromine atom in $CBrCl_3$ serves to partially cancel out the dipole moments of the three C─Cl bonds (as is the case for CCl_4).

1.61. Ethanol has a hydrogen atom attached to an oxygen atom, so it serves as the hydrogen bond donor. The lone pair on the oxygen atom in diethyl ether serves as the hydrogen bond acceptor. A dashed line is used to represent the hydrogen bond that forms between the two molecules:



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:

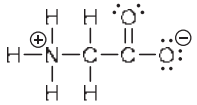
H—C≡N:

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.

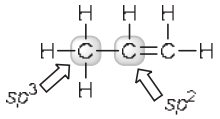
1.64. The molecular formula of cyclobutane is C_4H_8. Of the four structures shown, only structure (c) has the same molecular formula (C_4H_8).

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 (it is the only compound with both a hydrogen bond donor and a hydrogen bond acceptor).

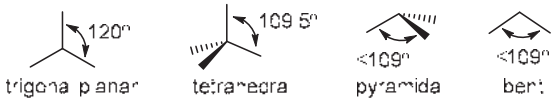
1.66. The first statement (a) is the correct answer, because an oxygen atom has a negative charge, and the nitrogen atom has a positive charge, as shown here:



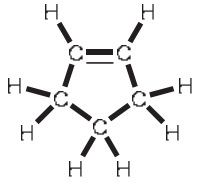
1.67. The correct answer is (d). Carbon-carbon σ bonds are formed by overlapping hybrid orbitals, so we must determine the hybridization of both carbon atoms (highlighted below). The carbon on the left has four σ bonds, so it is sp^3-hybridized, and the carbon on the right has three σ bonds, so it is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



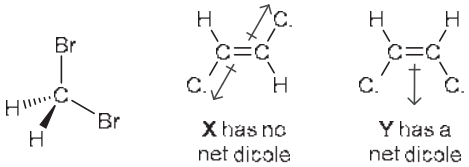
1.68. The correct answer is (c). An atom with a trigonal planar arrangement of bonds has a bond angle of 120°, the largest bond angle among the choices listed. A tetrahedral arrangement of bonds corresponds to an angle of 109.5°, while trigonal pyramidal and bent geometries typically have bond angles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

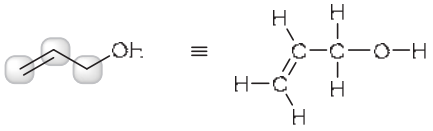


1.70. The correct answer is (b). Each C─Cl bond in **Y** has a dipole moment, and compound **Y** has a molecular dipole moment. Each C─Cl bond in **X** also has a dipole moment, but in this case, the two polar bonds are pointing in opposite directions. As such, the dipole moments fully cancel each other, 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**).

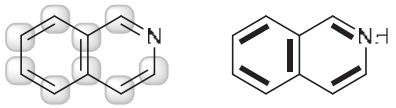


1.71. The correct answer is (b). In the bond-line drawing, each corner and each endpoint represents a carbon atom, so this compound has three carbon atoms (highlighted below). Each carbon atom will have enough hydrogen atoms to have exactly four bonds. Together with the

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.

1.73. (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).

1.73. (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.

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

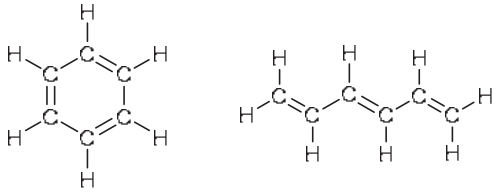
1.73. (e) A double bond represents one σ bond and one π bond, while a triple bond represents one σ bond and two π bonds. As such, compound C exhibits two π bonds.

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

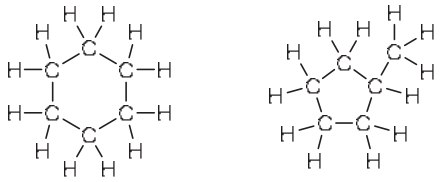
1.73. (g) Each of the carbon atoms in compound B is sp^3 hybridized with four σ bonds (steric number = 4). Similarly, the nitrogen atom in compound B has three σ bonds and one lone pair (steric number = 4). This nitrogen atom is also sp^3 hybridized.

1.73. (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.

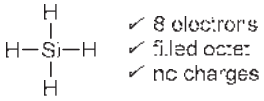
1.74. (a) In each of the following two compounds, all of the carbon atoms are sp^2 hybridized (each carbon atom has three σ bonds and one π bond). There are certainly many other possible compounds for which all of the carbon atoms are sp^2 hybridized.



1.74. (b) 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.

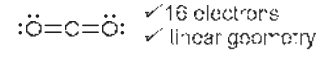


1.75. (a) Silicon is just below carbon in the periodic table. Like carbon, silicon has four valence electrons, and it will not have a formal charge when it forms four σ bonds. The Lewis structure of silane (SiH_4) exhibits eight valence electrons (four electrons from silicon and one electron from each of the four hydrogen atoms).



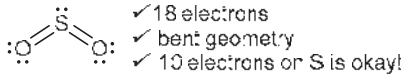
1.75. (b) Silane has four Si─H bonds, each of which exhibits a dipole moment, toward the more electronegative hydrogen atom (see Table 1.1). However, the central silicon atom has four attached atoms (four σ bonds) so it is expected to have tetrahedral geometry. As such, the four dipole moments completely cancel each other out, and there is no net molecular dipole moment (much like CCl_4, as seen in Figure 1.44). Therefore, silane is a nonpolar molecule.

1.76. The carbon atom of CO_2 has two attached atoms (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 attached atoms), which means that it has bent geometry:



Notice that the sulfur atom has no formal charge (it has six electrons attributed to it, as expected for an element in group 6A), but it also has TEN total electrons around it (four bonds, plus a lone pair)! Normally, you might think this is a violation of the octet rule, but the octet cannot be exceeded *for elements in the second row only*. Carbon, nitrogen, and oxygen have only four orbitals in their valence shell and cannot accommodate more than eight electrons. Sulfur, however, is in the third row and can handle 10 or even 12 electrons around it. Because SO_2 is bent, the individual dipole moments of each S═O bond do NOT cancel each other completely, and the molecule *does* have a molecular dipole moment.

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

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

1.77. (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°.

1.77. (d) The nitrogen atom in **B** has two attached atoms (two σ bonds) and one lone pair (steric number = 3). It is sp^2 hybridized. The highlighted carbon atom has three attached atoms (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.

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

1.77. (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.

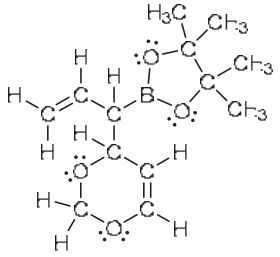
1.79. (a) In the first compound, the fluorine isotope (^{18}F) has no formal charge. Therefore, it must have three lone pairs (see Section 1.4 for a review of how formal charges are calculated). Since it has one σ bond and three lone pairs, it must have a steric number of 4, and is sp^3 hybridized. The 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 attached atoms (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.

1.79. (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°.

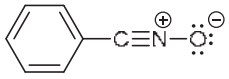
1.80. (a) Boron is in group 3A of the periodic table and is therefore expected to be trivalent. That is, it has three valence electrons, and it uses each one of those valence electrons to form a bond, giving rise to three bonds. It does not have any electrons left over for a lone pair (as in the case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

1.80. (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.

1.80. (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).

1.81. (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.

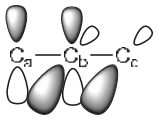
1.81. (c) In compound **1**, the carbon atom (attached to nitrogen) has three σ bonds and no lone pairs (steric number = 3). That carbon atom is sp^2 hybridized, with trigonal planar geometry. As such, the C─C─N bond angle in compound **1** is expected to be approximately 120°. However, in compound **2**, 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 attached atoms (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 attached atoms (two σ bonds) and no lone pairs, so it has a steric number of 2 and is therefore *sp* hybridized.

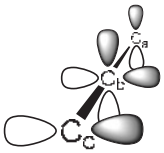
1.82. (b) Since C_a is sp^2 hybridized, we expect its geometry to be trigonal planar, so the bond angle should be approximately 120°.

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

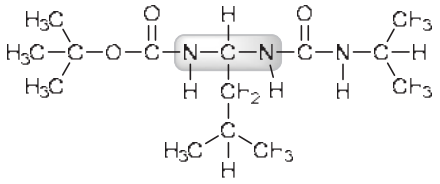
1.82. (d) The central carbon atom (C_b) is *sp* hybridized, so it is using two *sp* hybrid orbitals to form its two σ bonds, which will be arranged in a linear fashion. The remaining two *p* orbitals of C_b (used for π bonding) will be 90° apart from one another (just as we saw for the carbon atoms of a triple bond; see Figure 1.33).



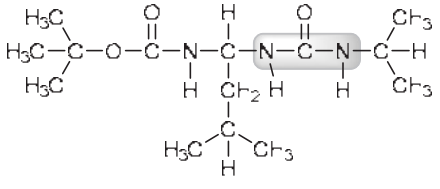
As a result, the two π systems are orthogonal (or 90°) to each other. Therefore, the *p* orbitals on C_a and C_c are orthogonal. The following is another drawing from a different perspective (looking down the axis of the linear C_a─C_b─C_c system.



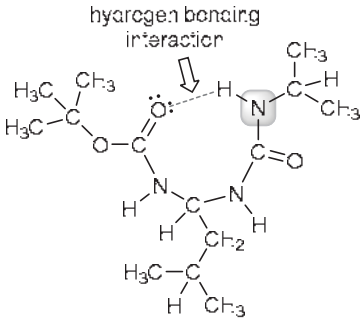
1.83. (a) The N─C─N unit highlighted below exhibits a central carbon atom that is sp^3 hybridized and is therefore expected to have tetrahedral geometry. Accordingly, the bond angles about that carbon atom are expected to be approximately 109.5°.



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°.



1.83. (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:



1.84. (a) The two esters readily combine to form a single solution because they have very similar polarities. Each molecule has a small polar region (with two oxygen atoms) and a very long nonpolar carbon chain. Because "like dissolves like" these two liquids are miscible.

1.84. (b) Glycerol has three O─H groups, making it an extremely polar molecule that is capable of forming hydrogen bonds. As such, it will have poor interactions with the largely nonpolar mixture of esters. Much like water creates a separate layer when mixed with oil (see Figure 1.53), the liquid glycerol separates from the liquid mixture of esters.

IMAGE CREDITS

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