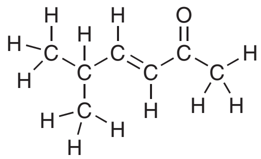
CHAPTER 1

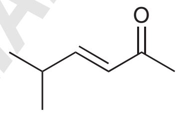
BOND-LINE DRAWINGS

To do well in organic chemistry, you must first learn to interpret the drawings that organic chemists use. When you see a drawing of a molecule, it is absolutely critical that you can read all of the information contained in that drawing. Without this skill, it will be impossible to master even the most basic reactions and concepts.

Molecules can be drawn in many ways. For example, below are three different ways of drawing the same molecule:



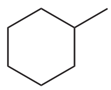
$(CH_3)_2CHCHCHCOCH_3$



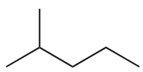
Without a doubt, the last structure (bond-line drawing) is the quickest to draw, the quickest to read, and the best way to communicate. Open to any page in the second half of your textbook and you will find that every page is plastered with bond-line drawings. It is essential that you learn to read these drawings fluently. This chapter will help you develop the skills that you will need to read these drawings quickly and fluently.

1.1 HOW TO READ BOND-LINE DRAWINGS

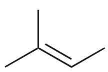
Bond-line drawings show the carbon skeleton (the connections of all the carbon atoms that build up the backbone, or skeleton, of the molecule) with any functional groups that are attached, such as —OH or —Br. Lines are drawn in a zigzag format, where each corner or endpoint represents a carbon atom. For example, the following compound has seven carbon atoms:

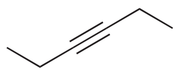


Don't forget that the ends of lines represent carbon atoms as well. For example, the following molecule has six carbon atoms (make sure you can count them):

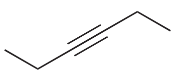


Double bonds are shown with two lines, and triple bonds are shown with three lines:

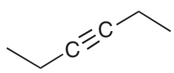




When drawing triple bonds, be sure to draw them in a straight line rather than zigzag, because triple bonds are linear (there will be more about this in the chapter on geometry). This can be quite confusing at first, because it can get hard to see just how many carbon atoms are in a triple bond, so let's make it clear:

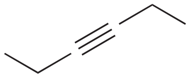


is the same as

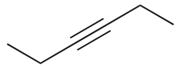


so this compound has 6 carbon atoms

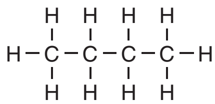
It is common to see a small gap on either side of a triple bond, like this:



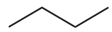
is the same as



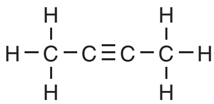
Both drawings above are commonly used, and you should train your eyes to see triple bonds either way. Don't let triple bonds confuse you. The two carbon atoms of the triple bond and the two carbons connected to them are drawn in a straight line. All other bonds are drawn as a zigzag:



is drawn like this:

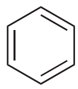


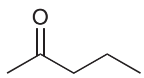
BUT



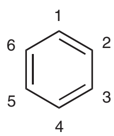
is drawn like this:

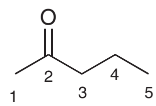
EXERCISE 1.1 Count the number of carbon atoms in each of the following drawings:



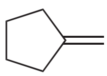


Answer The first compound has six carbon atoms, and the second compound has five carbon atoms:

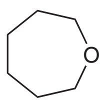




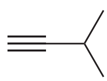
PROBLEMS Count the number of carbon atoms in each of the following drawings.



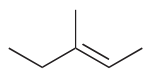
1.2 Answer: ______



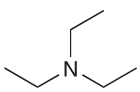
1.3 Answer: ______



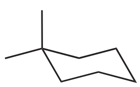
1.4 Answer: ______



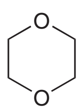
1.5 Answer: ______



1.6 Answer: ______

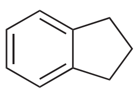


1.7 Answer: ______

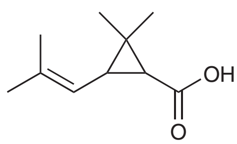


1.8 Answer: ______

1.9 Answer: ______

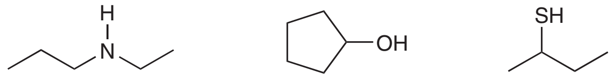


1.10 Answer: ______

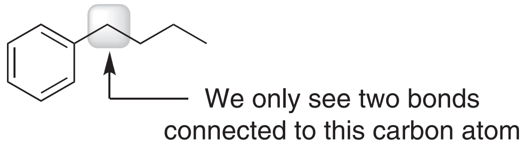


1.11 Answer: ______

Now that we know how to count carbon atoms, we must learn how to count the hydrogen atoms in a bond-line drawing. Most hydrogen atoms are not shown, so bond-line drawings can be drawn very quickly. Hydrogen atoms connected to atoms other than carbon (such as nitrogen or oxygen) must be drawn:



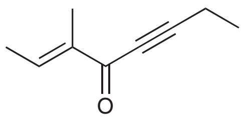
But hydrogen atoms connected to carbon are not drawn. Here is the rule for determining how many hydrogen atoms are connected to each carbon atom: *uncharged carbon atoms have a total of four bonds*. In the following drawing, the highlighted carbon atom is showing only two bonds:



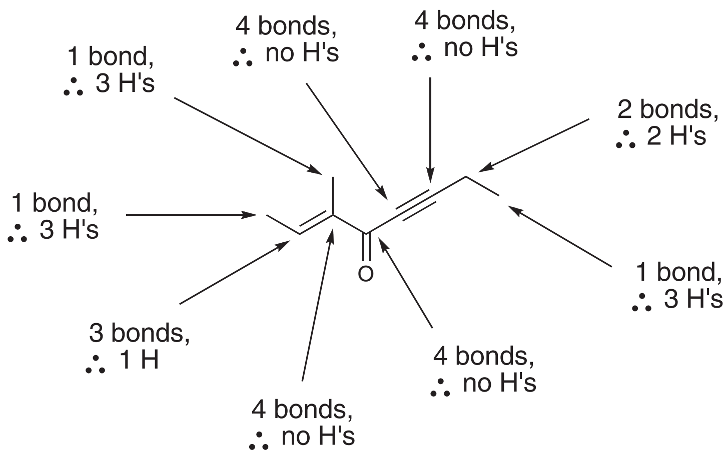
Therefore, it is assumed that there are two more bonds to hydrogen atoms (to give a total of four bonds). This is what allows us to avoid drawing the hydrogen atoms and to save so much time when drawing molecules. It is assumed that the average person knows how to count to four, and therefore is capable of determining the number of hydrogen atoms even though they are not shown.

So you only need to count the number of bonds that you can see on a carbon atom, and then you know that there should be enough hydrogen atoms to give a total of four bonds to the carbon atom. After doing this many times, you will get to a point where you do not need to count anymore. You will simply get accustomed to seeing these types of drawings, and you will be able to instantly "see" all of the hydrogen atoms without counting them. Now we will do some exercises that will help you get to that point.

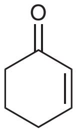
EXERCISE 1.12 The following molecule has nine carbon atoms. Count the number of hydrogen atoms connected to each carbon atom.

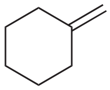


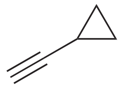
Answer

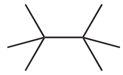


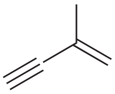
PROBLEMS For each of the following molecules, count the number of hydrogen atoms connected to each carbon atom.

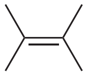
1.13 

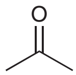
1.14 

1.15 

1.16 

1.17 

1.18 

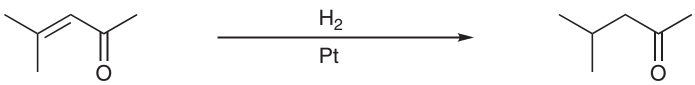
1.19 

1.20 

Now we can understand why we save so much time by using bond-line drawings. Of course, we save time by not drawing every C and H. But, there is an even larger benefit to using these drawings. Not only are they easier to draw, but they are easier to read as well. Take the following reaction for example:

$$(CH_3)_2C{=}CHCOCH_3 \xrightarrow[\text{Pt}]{H_2} (CH_3)_2CHCH_2COCH_3$$

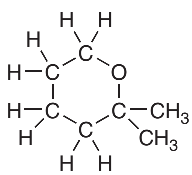
It is somewhat difficult to see what is happening in the reaction. You need to stare at it for a while to see the change that took place. However, when we redraw the reaction using bond-line drawings, the reaction becomes very easy to read immediately:



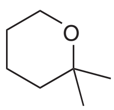
As soon as you see the reaction, you immediately know what is happening. In this reaction we are converting a double bond into a single bond by adding two hydrogen atoms across the double bond. Once you get comfortable reading these drawings, you will be better equipped to see the changes taking place in reactions.

1.2 HOW TO DRAW BOND-LINE DRAWINGS

Now that we know how to read these drawings, we need to learn how to draw them. Take the following molecule as an example:

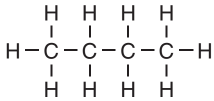


To draw this as a bond-line drawing, we focus on the carbon skeleton, making sure to draw any atoms other than C and H. All atoms other than carbon and hydrogen *must* be drawn. So the example above would look like this:

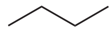


A few pointers may be helpful before you do some problems.

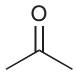
1. Don't forget that carbon atoms in a straight chain are drawn in a zigzag format:



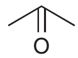
is drawn like this:



2. When drawing double bonds, try to draw the other bonds as far away from the double bond as possible:

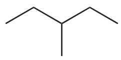


is much better than

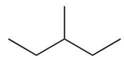


BAD

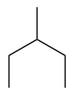
3. When drawing zigzags, it does not matter in which direction you start drawing:



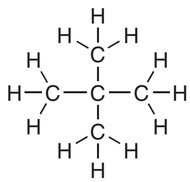
is the same as



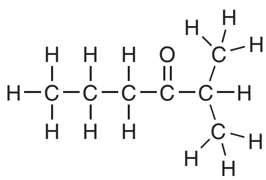
is the same as



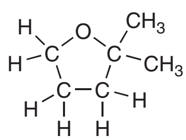
PROBLEMS For each structure below, draw a bond-line drawing in the box provided.



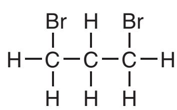
1.21



1.22



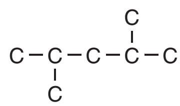
1.23



1.24

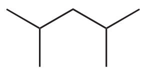
1.3 MISTAKES TO AVOID

1. *Never* draw a carbon atom with more than four bonds. This is a big no-no. Carbon atoms only have four orbitals; therefore, carbon atoms can form only four bonds (bonds are formed when orbitals of one atom overlap with orbitals of another atom). This is true of all second-row elements, and we will discuss this in more detail in the upcoming chapter.
2. When drawing a molecule, you should either show all of the H's and all of the C's, or draw a bond-line drawing where the C's and H's are not drawn. You *cannot* draw the C's without also drawing the H's:

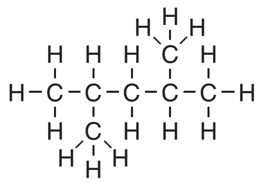


NEVER DO THIS

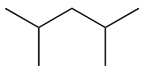
This drawing is insufficient. Either leave out the C's (which is preferable) or put in the H's:



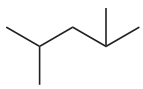
or



3. When drawing each carbon atom in a zigzag, try to draw all of the bonds as far apart as possible:



is better than



1.4 MORE EXERCISES

First, open your textbook and flip through the pages in the second half. Choose any bond-line drawing and make sure that you can say with confidence how many carbon atoms you see and how many hydrogen atoms are attached to each of those carbon atoms.

Now examine the following transformation, and think about the changes that are occurring:

Don't worry about *how* these changes occur. That will be covered much later (in Chapter 11), when we explore this type of transformation in more detail. For now, just focus on describing the changes that you see. In this case, two hydrogen atoms have been installed, and a double bond has been converted into a single bond. It is certainly clear to see that the double bond has been converted into a single bond, but you should also clearly see that two hydrogen atoms have been installed during this process.

Consider another example:

In this example, H and Br have been removed, and a single bond has been converted into a double bond (we will see in Chapter 10 that it is actually H^+ and Br^- that are removed). If you cannot see that an H was removed, then you will need to count the number of hydrogen atoms in the starting material and compare it with the product:

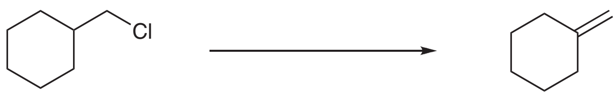
Now consider one more example:

In this example, a bromine atom has been replaced with a chlorine atom (as we will see in Chapter 9). Inspection of the bond-line drawings clearly indicates that no other changes occurred in this case.

PROBLEMS For each of the following transformations, describe the changes that are occurring.

1.25
Answer:

1.26
Answer:



1.27
Answer:

1.28
Answer:

1.29
Answer:

1.30 

Answer: ____

____

1.31 

Answer: ____

____

1.32 

Answer: ____

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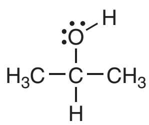
1.5 IDENTIFYING FORMAL CHARGES

Formal charges are charges (either positive or negative) that we must often include in our drawings. They are extremely important. If you don't draw a formal charge when it is supposed to be drawn, then your drawing will be incomplete (and wrong). So you must learn how to identify when you need formal charges and how to draw them. If you cannot do this, then you will not be able to draw resonance structures (which we will see in the next chapter), and if you can't do that, then you will have a very hard time passing this course.

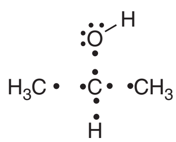
A formal charge is a charge associated with an atom that does not exhibit the expected number of valence electrons. When calculating the formal charge on an atom, we first need to know the number of valence electrons the atom is *supposed* to have. We can get this number by inspecting the periodic table, since each column of the periodic table indicates the number of expected valence electrons (valence electrons are the electrons in the valence shell, or the outermost shell of electrons—you probably remember this from high school chemistry). For example, carbon is in Column 4A, and therefore has four valence electrons. This is the number of valence electrons that a carbon atom is supposed to have.

Next we ask how many electrons are attributed to each atom in the Lewis structure. For each atom, if the electron count does not match the expected valence, then the atom bears a formal charge.

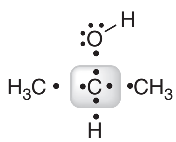
Let's see an example. Consider the central carbon atom in the compound below:



Remember that every bond represents two electrons being shared between two atoms. Begin by splitting each bond apart equally, like this:



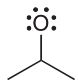
Now count the number of electrons immediately surrounding the central carbon atom:



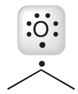
There are four electrons. This is the number of electrons that are attributed to this carbon atom.

In this case, the electron count matches the expected valence, so this carbon atom has no formal charge. This will be the case for most of the atoms in the structures you will draw in this course. But in some cases, there will be a difference between the electron count and the expected valence. In those cases, there will be a formal charge. So let's see an example of an atom that has a formal charge.

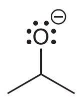
Consider the oxygen atom in the structure below:



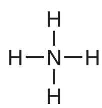
Let's begin by determining the number of valence electrons that are expected for an oxygen atom. Oxygen is in Column 6A of the periodic table, so oxygen should have six valence electrons. Next, we count the number of valence electrons that are attributed to this oxygen atom in the Lewis structure. To do this, we redraw the structure by splitting up the C—O bond equally, like this:



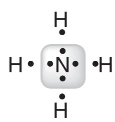
One electron of the C—O bond is attributed to the oxygen atom, and in addition, this oxygen atom also has three lone pairs. Recall that lone pairs are unshared electrons (electrons that are not being used to form bonds), and therefore, we attribute both electrons of each lone pair to the oxygen atom. As such, we attribute seven electrons to this oxygen atom (one from the C—O bond, and six more from the lone pairs). In this case, the electron count (seven) does NOT match the expected the valence (six). This oxygen atom has one extra electron electron and will therefore bear a negative charge:



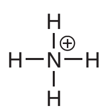
EXERCISE 1.33 Consider the nitrogen atom in the structure below and determine if it has a formal charge:



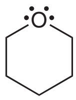
Answer Nitrogen is in Column 5A of the periodic table so it should have five valence electrons. Now we count the number of electrons that are attributed to the nitrogen atom in this structure:

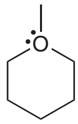


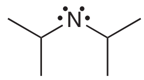
There are only four electrons attributed to this nitrogen atom (one electron for each bond). This nitrogen atom appears to be missing an electron, and therefore, this nitrogen atom has a positive charge:

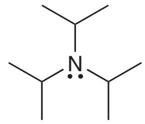


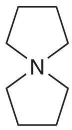
PROBLEMS For each of the structures below determine if the oxygen or nitrogen atom has a formal charge. If there is a charge, draw the charge.

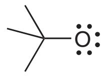
1.34 

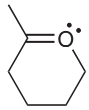
1.35 

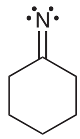
1.36 

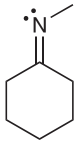
1.37 

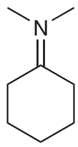
1.38 

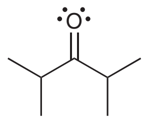
1.39 

1.40 

1.41 

1.42 

1.43 

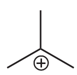
1.44 

1.45 $H_3C{-}C{\equiv}N{:}$

This brings us to the most important atom of all: carbon. We saw before that carbon forms four bonds. This allows us to ignore the hydrogen atoms when drawing bond-line structures, because it is assumed that we know how to count to four and can figure out how many hydrogen atoms are there. When we said that, we were only talking about carbon atoms without formal charges (most carbon atoms in most structures will not have formal charges). But now that we have learned what a formal charge is, let's consider what happens when carbon has a formal charge.

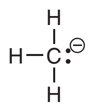
If carbon bears a formal charge, then we cannot just assume it will still have four bonds. In fact, it will have only three. Let's see why. Let's first consider C^+, and then we will move on to C^-.

If carbon has a *positive* formal charge, then it has only three valence electrons (it is supposed to have four valence electrons, because carbon is in Column 4A of the periodic table). Since it has only three valence electrons, it can form only three bonds. That's it. So, a carbon atom with a positive formal charge will have only three bonds, and you should keep this in mind when counting hydrogen atoms:



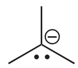
No hydrogen atoms
on this C^+

Now let's consider what happens when we have a carbon atom with a *negative* formal charge. The reason it has a negative formal charge is because it has one more electron than it is supposed to have. That is, it has five electrons attributed to it, rather than four. Two of these electrons form a lone pair, and the other three electrons are used to form bonds:



We have the lone pair, because we can't use each of the five electrons to form a bond. Carbon can *never* have five bonds. Why not? Electrons exist in regions of space called orbitals. These orbitals can overlap with orbitals from other atoms to form bonds, or the orbitals can contain two electrons (which is called a lone pair). Carbon has only four orbitals in its valence shell, so there is no way it could possibly form five bonds—it does not have five orbitals to use to form those bonds. This is why a carbon atom with a negative charge will have a lone pair (if you look at the drawing above, you will count four orbitals—one for the lone pair and then three more for the bonds).

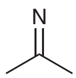
Therefore, a carbon atom with a negative charge can also form only three bonds (just like a carbon with a positive charge). When you count hydrogen atoms, you should keep this in mind:



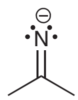
No hydrogen atoms
on this C⁻

1.6 FINDING LONE PAIRS THAT ARE NOT DRAWN

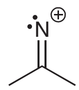
From all of the cases above (oxygen, nitrogen, and carbon), you can see why you have to know how many lone pairs there are on an atom in order to figure out the formal charge on that atom. Similarly, you have to know the formal charge to figure out how many lone pairs there are on an atom. Take the case below with the nitrogen atom shown:



could either be



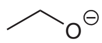
or



If the lone pairs were drawn, then we would be able to figure out the charge (two lone pairs would mean a negative charge and one lone pair would mean a positive charge). Similarly, if the formal charge was drawn, we would be able to figure out how many lone pairs there are (a negative charge would mean two lone pairs and a positive charge would mean one lone pair). So you can see that drawings must include either lone pairs or formal charges. *The convention is to always show formal charges and to leave out the lone pairs*. This is much easier to draw, because you usually won't have more than one charge on a drawing (if even that), so you get to save time by not drawing every lone pair on every atom.

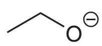
Now that we have established that formal charges must *always* be drawn and that lone pairs are usually *not* drawn, we need to get practice in how to see the lone pairs when they are not drawn. This is not much different from training yourself to see all the hydrogen atoms in a bond-line drawing even though they are not drawn. If you know how to count, then you should be able to figure out how many lone pairs are on an atom where the lone pairs are not drawn.

Let's see an example to demonstrate how you do this:

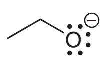


In this case, we are looking at an oxygen atom. Oxygen is in Column 6A of the periodic table, so it is supposed to have six valence electrons. Then, we need to take the formal charge into account. This oxygen atom has a negative charge, which means one extra electron. Therefore, we attribute $6 + 1 = 7$ electrons to this oxygen atom. Now we can figure out how many lone pairs there are.

The oxygen atom has one bond, which means that it is using one of its seven electrons to form a bond. The other six must be in lone pairs. Since each lone pair is two electrons, this must mean that there are three lone pairs:



is the same as

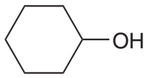
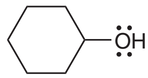


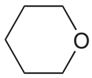
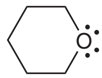
Let's review the process:

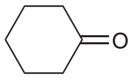
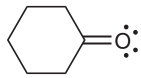
1. Identify the expected valence (using the periodic table).
2. Determine the number of electrons attributed to the atom in the Lewis structure. To do this, take the formal charge into account. A negative charge means one more electron, and a positive charge means one less electron.
3. Use this number to figure out how many lone pairs there are.

Now we need to get used to the common examples. Although it is important that you know how to count and determine numbers of lone pairs, it is actually much more important to get to a point where you don't have to waste time counting. You need to get familiar with the common situations you will encounter. Let's go through them methodically.

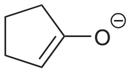
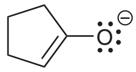
When oxygen has no formal charge, it will have two bonds and two lone pairs:

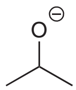
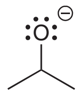
 is the same as 

 is the same as 

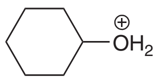
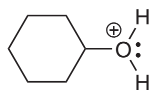
 is the same as 

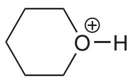
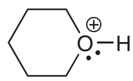
If oxygen has a negative formal charge, then it must have one bond and three lone pairs:

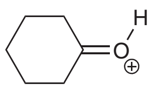
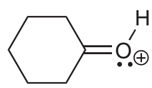
 is the same as 

 is the same as 

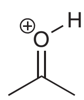
If oxygen has a positive charge, then it must have three bonds and one lone pair:

 is the same as 

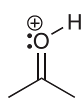
 is the same as 

 is the same as 

EXERCISE 1.46 Draw all lone pairs in the following structure:



Answer The oxygen atom has a positive formal charge and three bonds. You should try to get to a point where you recognize that this must mean that the oxygen atom has one lone pair:

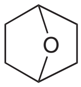


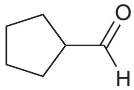
Until you get to the point where you can recognize this, you should be able to figure out the answer by counting.

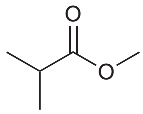
Oxygen is supposed to have six valence electrons. This oxygen atom has a positive charge, which means it is missing an electron. Therefore, we attribute $6 - 1 = 5$ electrons to this oxygen atom. Now, we can figure out how many lone pairs there are.

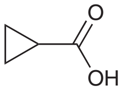
The oxygen atom has three bonds, which means that it is using three of its five electrons to form bonds. The other two must be in a lone pair. So there is only one lone pair.

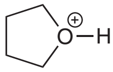
PROBLEMS Review the common situations above, and then come back to these problems. For each of the following structures, draw all lone pairs. Try to recognize how many lone pairs there are *without* having to count. Then count to see if you were right.

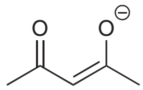
1.47 

1.48 

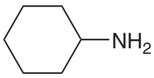
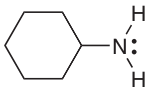
1.49 

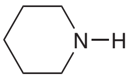
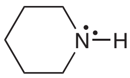
1.50 

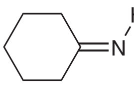
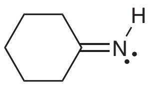
1.51 

1.52 

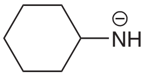
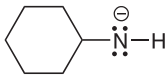
Now let's look at the common situations for nitrogen atoms. When nitrogen has no formal charge, it will have three bonds and one lone pair:

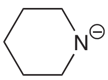
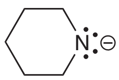
 is the same as 

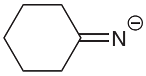
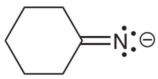
 is the same as 

 is the same as 

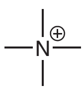
If nitrogen has a negative formal charge, then it must have two bonds and two lone pairs:

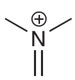
 is the same as 

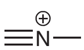
 is the same as 

 is the same as 

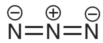
If nitrogen has a positive charge, then it must have four bonds and no lone pairs:

 has no lone pairs

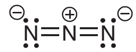
 has no lone pairs

 has no lone pairs

EXERCISE 1.53 Draw all lone pairs in the following structure:



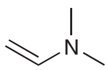
Answer The central nitrogen atom has a positive formal charge and four bonds. You should try to get to a point where you recognize that this nitrogen atom does not have any lone pairs. Each of the other nitrogen atoms has a negative formal charge and two bonds. You should try to get to a point where you recognize that each of these nitrogen atoms has two lone pairs:

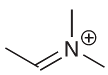


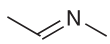
Until you get to the point where you can recognize this, you should be able to figure out the answer by counting. Nitrogen is supposed to have five valence electrons. The central nitrogen atom has a positive charge, which means it is missing an electron. In other words, we attribute $5 - 1 = 4$ electrons to this nitrogen atom. Now, we can figure out how many lone pairs there are. Since it has four bonds, it is using all of its electrons to form bonds. So there is no lone pair on this nitrogen atom.

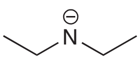
For each of the remaining nitrogen atoms, there is a negative formal charge. That means that each of those nitrogen atoms has one extra electron, $5 + 1 = 6$ electrons. Each nitrogen atom has two bonds, which means that each nitrogen atom has four electrons left over, giving two lone pairs.

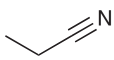
PROBLEMS Review the common situations for nitrogen, and then come back to these problems. For each of the following structures, draw all lone pairs. Try to recognize how many lone pairs there are *without* having to count. Then count to see if you were right.

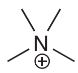
1.54 

1.55 

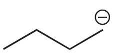
1.56 

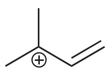
1.57 

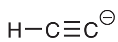
1.58 

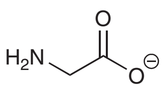
1.59 

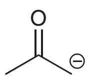
For each of the following structures, draw all lone pairs.

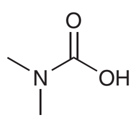
1.60 

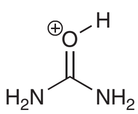
1.61 

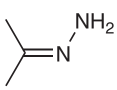
1.62 

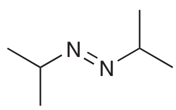
1.63 

1.64 

1.65 

1.66 

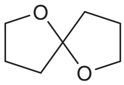
1.67 

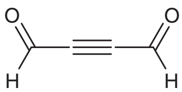
1.68 

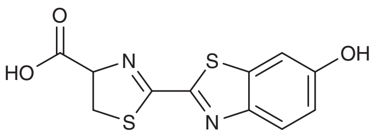
END-OF-CHAPTER PROBLEMS

PRACTICE PROBLEMS *(Problems that involve only one skill)*

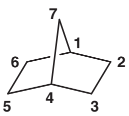
Problems 1.69–1.71. Count the number of carbon atoms in each of the following structures:

1.69  Number of C atoms: ____

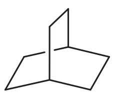
1.70  Number of C atoms: ____

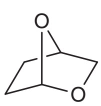
1.71 
Number of C atoms: ____

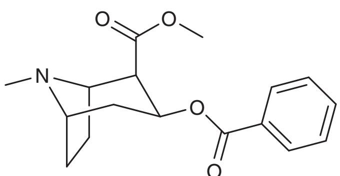
Problems 1.72–1.74. The structure below has seven carbon atoms and is said to be bicyclic (because converting this compound into an acyclic compound, with no ring at all, would require that we break two C—C bonds).



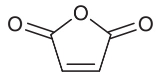
Notice that the bond between C1 and C6 is drawn with a gap to indicate that this bond is behind the bond between C4 and C7. Bicyclic compounds are drawn in this way (with one of the bonds drawn with a gap) to illustrate their three-dimensional nature. With this in mind, count the number of carbon atoms in each of the following structures:

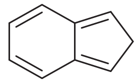
1.72  Number of C atoms: ____

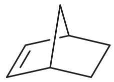
1.73  Number of C atoms: ____

1.74 
Cocaine
Number of C atoms: ____

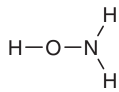
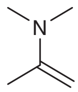
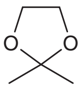
Problems 1.75–1.77. Count the number of hydrogen atoms in each of the following structures:

1.75  Number of H atoms: ____

1.76  Number of H atoms: ____

1.77  Number of H atoms: ____

Problems 1.78–1.80. Each of the following structures has one or more atoms with lone pairs. We will encounter each of these structures as we study organic chemistry, and in each case, one of the lone pairs is responsible for the reactivity that we will see. For each of the following structures, identify all lone pairs:

1.78  **1.79**  **1.80** 

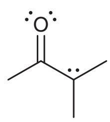
Problems 1.81–1.90. Draw any missing formal charges in each of the following structures. Also, determine the net (total) charge on each structure by adding together all the individual charges. For example, a structure will have a net charge of zero if the structure has one positive charge and one negative charge.

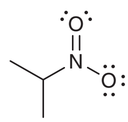
1.81 $:C\equiv N:$

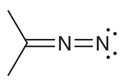
1.82 $:C\equiv O:$

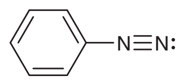
1.83 $:N\equiv O:$

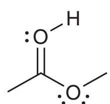
1.84 $:N\equiv N:$

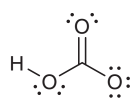
1.85 

1.86 

1.87 

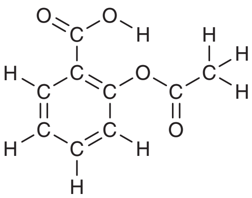
1.88 

1.89 

1.90 

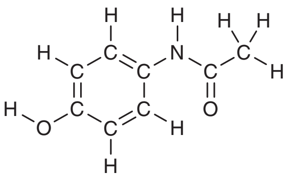
Problems 1.91–1.93. Below are the structures of three common analgesics (pain relievers). In each of the following cases, use the box provided to draw a bond-line structure for the compound shown.

1.91



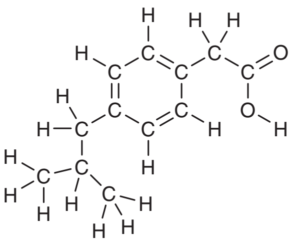
Aspirin

1.92



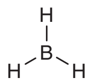
Tylenol (acetaminophen)

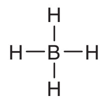
1.93

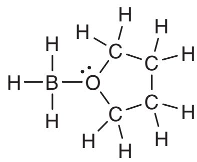


Advil (ibuprofen)

Problems 1.94–1.96. Boron is in column 3A of the periodic table and therefore has three valence electrons (as compared with carbon, which has four valence electrons). Based on this information, draw any missing formal charges in each of the following structures. Note: All lone pairs in the given structures are shown. Also, determine the net (total) charge on each structure, by adding together all the individual charges. For example, a structure will have a net charge of zero if the structure has one positive charge and one negative charge:

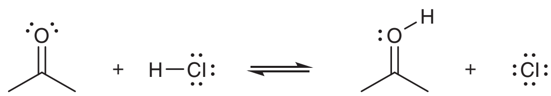
1.94 

1.95 



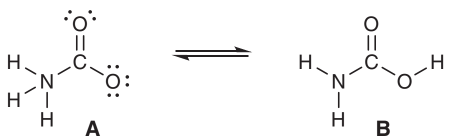
1.96

1.97 Below is the reaction between a ketone and HCl. The products are ionic, but the formal charges have not been shown. Draw the formal charges on the products.



INTEGRATED PROBLEMS *(Problems that involve more than one skill)*

1.98 Shown below is the amino acid called glycine. Notice that glycine exists as an equilibrium between two structures, labelled **A** and **B** below:



(a) Structure **A** has formal charges which have not been shown. Draw those formal charges.

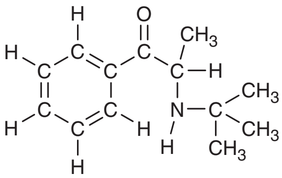
(b) What is the net charge on structure **A**?

(c) Draw all lone pairs in structure **B**.

(d) Does structure **B** have any formal charges that need to be drawn?

(e) Does structure **B** have the same net charge as structure **A**?

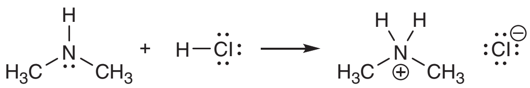
1.99 Marketed under the trade name Zyban, the drug below is used as a smoking cessation aid for people who want to quit smoking.



Zyban

(a) In the box provided below, draw a bond-line structure for Zyban:

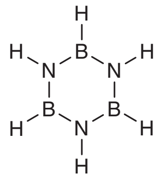
(b) Zyban is described as an amine because of the part of its structure that contains a nitrogen atom. As we will see in an upcoming chapter, when an amine is treated with HCl, an acid–base reaction occurs in which H^+ is transferred to the amine, thereby generating an ammonium ion and a chloride ion:



An amime An ammonium chloride

This product is an ionic species (much like NaCl), and is called an ammonium chloride. Zyban is actually sold as an ammonium chloride, rather than as an amine. In the box provided, draw the structure of this ammonium chloride:

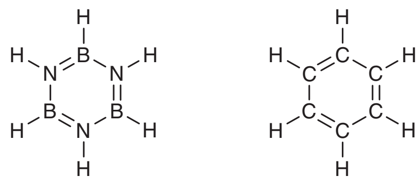
1.100 The following compound, called borazine, is inorganic (it contains no carbon atoms).



Borazine

(a) Boron is in column 3A of the periodic table and therefore has three valence electrons (as compared with carbon, which has four valence electrons). Based on this information, and assuming that none of the atoms bears a formal charge, draw all lone pairs in the structure above.

(b) Borazine is remarkably stable, as it shares features with a very stable organic compound, called benzene (later in the course, we will see the source of this stability). Indeed, borazine can alternatively be drawn as shown below, which illustrates its structural similarity with benzene:



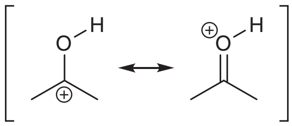
Borazine Benzene

In this drawing of borazine (which has three double bonds, but no lone pairs), there are missing formal charges that have not been drawn. Draw these missing formal charges, and then indicate the net charge on this structure.

CHALLENGE PROBLEM

1.101 In the next chapter, we will see that some structures cannot be adequately represented with only one bond-line drawing, because it is difficult to pinpoint the exact location of some of the electrons. In such cases, we will see that we must draw more than one drawing (called resonance structures), and then meld those images together in our minds. For example, in the previous problem, we saw two different resonance structures for borazine (one resonance structure that had only single bonds, and another resonance structure that had three double bonds but no lone pairs). Neither of these structures alone can adequately represent the nature of the compound. That is, the B—N bonds are not purely single bonds, nor are they purely double bonds. Rather, the B—N bonds are somewhere in between these two extremes, having some double-bond character and some single-bond character.

With this in mind, consider the cation shown below, which has two resonance structures.



(a) Draw all of the lone pairs in each of the resonance structures above.

(b) The two resonance structures above are used to represent a single chemical entity, although these two resonance structures do not contribute equal character to the overall nature of this chemical entity. That is, one of these resonance structures contributes more character than the other resonance structure. For reasons that we will see in the next chapter, the resonance structure with fewer lone pairs is the structure that contributes more character. Based on this information, determine whether the C—O bond has more single-bond character or more double-bond character.

(c) Let's explore the reason why one of these resonance structures contributes more character than the other. Recall that a lone pair is a pair of electrons that is NOT shared between atoms. Therefore, the resonance structure with fewer lone pairs (fewer unshared electrons) will be the resonance structure with more bonds (more shared electrons). With more shared electrons, it is more likely that all atoms will have filled octets (will be surrounded by eight electrons, which is one of the most important features for making a resonance structure stable). Indeed, one of the resonance structures shown above has an atom that is missing an octet of electrons, while the other resonance structure has filled octets for every atom. Identify the resonance structure with the unfilled octet, and identify which atom in that resonance structure lacks an octet of electrons.