

Exercise 2

Lewis Structures, Isomers, Bond Polarity, Molecular Models and Geometry

Fig. X2-1



Gilbert N. Lewis (1875 - 1946)
One of America's greatest chemists.

<http://www.woodrow.org/teachers/chemistry/institutes/1992/Lewis.html>

Although it has recently become possible to image molecules and even atoms using a scanning tunneling microscope, most of our information about molecular structure comes from interpretation of physical, chemical and spectroscopic properties of substances. This information often enables us to piece together a 3-dimensional picture or model of the molecule. On paper, one of the best methods we have of representing this model is by drawing a Lewis structure of the molecule or ion. The ability to draw Lewis structures for covalently bonded compounds and polyatomic ions is essential for the understanding of polarity, resonance structures, chemical reactivity and isomerism. Lewis structures of organic molecules coupled with a consideration of bond polarities enable us to deduce the geometry of molecules and the molecular polarity. The use of molecular models considerably facilitates the visualization of the three dimensional shape of the molecule and further enhances our ability to understand chemical reactivity and mechanisms. This *Exercise* has been designed to give you more experience with the drawing of Lewis structures and the determination of molecular polarity and geometry from the structures. As with *Exercise 1*, many of the problems should lead to interesting learning and insight. Answers to the problems with an * are in *Appendix A*.

Lewis structures. Remember the **absolute** rule that the Lewis structure **must** show the correct number of electrons. For a molecule, the sum of the valence electrons is the correct number. For a positive ion, subtract one electron for each positive charge from the number of valence electrons. For a negative ion, add one electron for each negative charge to the number of valence electrons. For example, formaldehyde, CH_2O , must show $4 + 2 + 6 = 12$ electrons and nitrite, NO_2^- , must show $5 + 6 + 6 + 1 = 18$ electrons.

One general sequence to follow when constructing Lewis structures is:

1. Determine the correct number of electrons that must be showing (sum of the valence electrons and the negative of the charge on the ion).
2. Arrange the atoms in the correct sequence or sequences if there is more than one logical sequence. For polyatomic ions, the atom listed first (e.g. S in SO_4^{2-}) is usually the central atom (an atom attached to two or more atoms). For more information on choosing the correct sequence, read the discussion about formal charge below.
3. Connect each of the atoms with a line representing one bond and therefore two electrons.
4. Distribute the remaining electrons in pairs (subtract two each from the total for each bond already inserted) around the external atoms (the ones attached to the central atom) attempting to complete each atom's octet. Remember that the octet (duet for hydrogen) should **never** be exceeded for atoms through the second period (through neon) and should be exceeded for atoms beyond the second period when the atom is a central atom. If electrons still remain, attempt to fill the octets of central atoms or even exceed the octets for atoms beyond the second period.

5. If central atoms are short of achieving an octet, move electron pairs from external atoms to form multiple bonds between the external atom and the central atom in an attempt to provide all atoms with an octet (or, for atoms beyond the second period the octet might be exceeded).
6. If the total number of electrons is odd, it will be impossible to satisfy the octet rule for all atoms and in fact the species will be a free radical (it will have an unpaired electron) and probably will be a relatively reactive species.
7. Visually inspect the resulting structure and decide if it is consistent with other structures you have seen and the guidelines below (generalities are difficult but the following are usually true):
 - a. hydrogen has one bond
 - b. carbon has four bonds (except for carbon monoxide and cyanide ion)
 - c. halogens have one bond unless the halogen is a central atom
 - d. oxygen usually has two bonds (except for polyatomic ions)
 - e. nitrogen usually has three bonds

If the structure is not intuitively satisfying or if you cannot decide between different bonding sequences, determine the formal charge on each atom. You are probably familiar with the concept of oxidation number. Basically, the oxidation number method assumes that bonds are 100% ionic and assigns all the electrons in bonds to the more electronegative partner. For sodium chloride, this results in oxidation numbers of +1 for sodium and -1 for chloride. While this method yields a good description of the compound when the bonding is ionic, it should be considered nothing more than a bookkeeping method for covalent bonds. In other words, it is a useful method for determining if a reaction is a redox reaction and if so, what is oxidized and what is reduced. For covalent bonds, oxidation numbers give a very distorted view of the charges in the molecule. For instance, for HCl, the oxidation number method results in a +1 for hydrogen and -1 for chloride. In actuality, the hydrogen chlorine bond is best described as polar covalent with a partial positive charge on the hydrogen and a partial negative charge on the chlorine.

The formal charge method assumes that the bonds are 100% covalent with bonded electrons equally shared by the two partners. As a result, the formal charge method gives an indication of the locations of charges in covalently bonded compounds. Although formal charges ignore differences in electronegativities, the method still provides useful information. In general, everything else being equal (e.g., octet rule is satisfied) **the structure with the minimum number of formal charges is favored**. The formal charge is calculated using the following formula:

$$\text{formal charge} = \text{valence electrons} - \text{bonds} - \text{nonbonded electrons}$$

After you have determined the formal charges, choose the structure with the minimum number of formal charges. For some structures with central atoms that are beyond the second period (commonly phosphorous, sulfur, chlorine, bromine, iodine), it is sometimes preferable to move electrons from external atoms to form multiple bonds with the central atom. Empirically it appears that minimizing formal charge is more important than maintaining an octet for these atoms.

8. If it is possible to draw more than one reasonable structure by moving electrons only, then all of these structures should be drawn and connected by double headed arrows \leftrightarrow . These structures are resonance structures and the actual structure is a hybrid of all of the resonance structures. It is very important to realize that the actual structure is not going back and forth between the resonance structures but is a hybrid of the structures. For instance, if formal charges are different from one structure to the other, the formal charges on each atom are probably best represented by an average of the values in each structure.
9. Determine the geometry of the structure by applying VSEPR theory or a hybridization model to the resulting Lewis structure. Both models, when appropriately used will predict with very few exceptions the same shape. The table below summarizes the theories.

<u>groups¹ of electrons around central atom</u>	<u>electronic shape²</u>	<u>bond angles</u>	<u>hybridization</u>
2	linear	180°	sp
3	planar	120°	sp ²
4	tetrahedral	109.47°	sp ³
5	trigonal bipyramid	90°, 120°, 180°	dsp ³
6	octahedral	90°, 180°	d ² sp ³

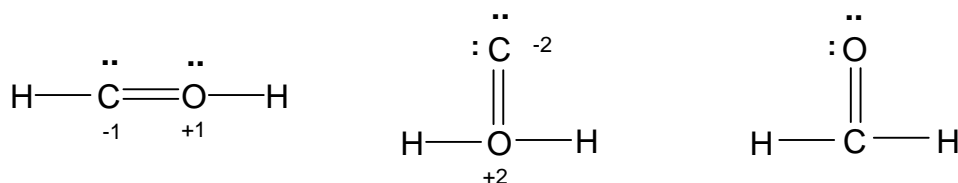
¹The number of groups of electrons is equal to the sum of the number of neighbor atoms and non-bonded electron pairs.

²Be sure to distinguish electronic shape from molecular shape. If one or more of the groups of electrons are non-bonded pairs, the molecule needs to be described by the relative positions of the atoms; not by the shapes of the electronic orbitals.

To illustrate the use of the preceding guidelines, the procedure will be applied to a couple of examples. Consider the molecule formaldehyde, CH₂O. First, determine the correct number of electrons that should show in the final structure which is 4 + 2 + 6 = 12. Now we must consider the sequence of bonding. Three possible choices are (other possibilities can be eliminated because they require two bonds to hydrogen which except for some unusual boron compounds should be absolutely avoided):



As symmetrical choices are often favored, the first structure could be eliminated because it has lower symmetry than the other two. We will leave it in for this discussion and continue by inserting the three single bonds (6 e⁻) and adding the remaining 6 e⁻ to the external atoms (C and O respectively for the second and third structures) or 6 e⁻ to the C and O of the first structure. To satisfy the octet rule in each structure, one non-bonded electron pair is moved from the external atom in the second and third structures to form a double bond with the central atom. For the first structure, the second non-bonded electron pair on either the C or O is moved between the two atoms to form a double bond.



The three structures are technically correct Lewis structures for CH₂O but only one correctly represents formaldehyde. In addition to considering symmetry, it is possible to choose between structures by evaluating the formal charge on each atom in a structure. The minimization of formal charges for CH₂O enables us to strongly favor the third and correct structure.

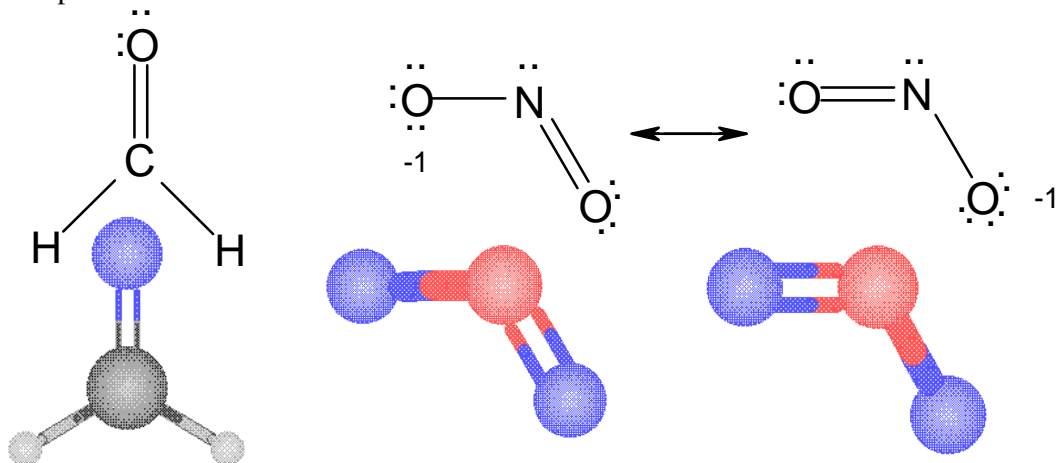
To draw Lewis structures of ions, the procedure is similar but a modification is needed. In determining the correct number of electrons, the number of valence electrons should be added to the negative of the charge. If the ion has a negative charge, it has extra electrons that it has acquired from its partners. The following example will illustrate the technique for nitrite, NO₂⁻. The correct number of electrons is 5 + 12 + 1 = 18. Use of the symmetry guideline leads to the sequence ONO rather than OON and, in addition, a formal charge analysis on the completed Lewis structures (try it!) also favors ONO.

In this case, you should draw two structures that differ by the position of electrons only and are resonance structures. The double headed arrow below is the convention used to indicate that the two Lewis structures are resonance structures.



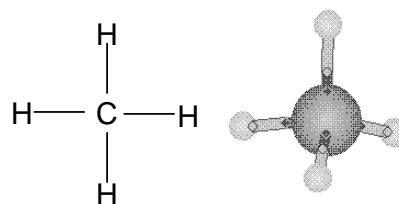
Remember that when resonance structures can be drawn, none of the Lewis structures correctly depicts the structure but one must try to imagine a hybrid (or enhanced average) as a better model for the species.

Now let's look at the geometry of formaldehyde and nitrite. For the formaldehyde molecule drawn earlier, there are 3 groups of electrons around the central carbon. Notice the double bond counts as 1 group of electrons and not 2! Focus on the shape around each atom and do not count electrons on the oxygen when determining the shape around carbon. The 3 groups result in a prediction that the atoms around carbon are 120° apart in a plane.

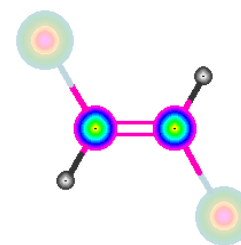
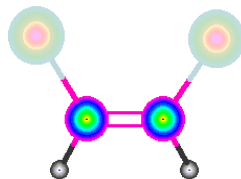
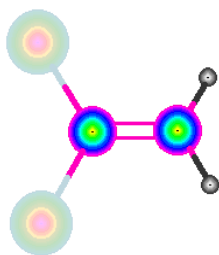
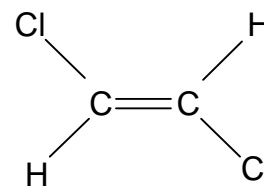
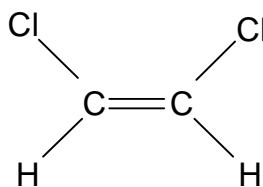
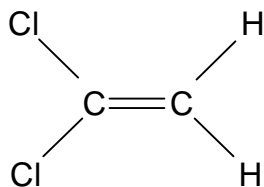


For nitrite, there are also 3 groups of electrons around the central nitrogen as the non-bonded electron pair counts as a group. Again VSEPR theory predicts 120° bond angles and a bent ion.

When 4 groups of electrons surround the central atom, a tetrahedral structure results which can be represented on paper as a projection. While projections are commonly used, CH_4 is often written in planar form with bond angles that appear to be 90° . You should recognize when you see the planar drawing that the bond angles are actually 109.5° and that the molecule is tetrahedral. The 3-dimensional nature of most molecules plays a significant role in determining physical and chemical properties of a substance. For example, a substrate must exactly fit into a portion of an enzyme if the enzyme is to perform its catalytic function, just as a key must have just the right shape to fit a lock.



Isomerism. The importance of the sequence and spatial arrangement of the atoms is established with the concept of isomerism. Consider the formula $\text{C}_2\text{H}_2\text{Cl}_2$. It is possible to draw three acceptable Lewis structures for this formula that correspond to the compounds 1,1-dichloroethene, *cis*-1,2-dichloroethene and *trans*-1,2-dichloroethene.



1,1-dichloroethene

cis-1,2-dichloroethene*trans*-1,2-dichloroethene.

b.p. 37°C
m.p. -122°C
density 1.218 g/mL

60°C
-80°C
1.284 g/mL

48°C
-50°C
1.257 g/mL

The three possible **isomers** exist and as can be observed from the data above, have different properties. This demonstrates the fact that the spatial arrangement of the atoms has significant effects on the physical and chemical properties of the molecule. The first isomer (1,1-dichloroethene) has a different sequence of bonding than the other two and is a structural isomer of the other two. The second and third isomers have the same sequence of bonding but are spatially different. Because they differ in geometry, they are called geometric isomers. Notice that because of the presence of the double bond, rotation from the *cis* to the *trans* is not possible unless substantial energy input is provided. The absorption of light by rhodopsin in the eye causes a *cis* - *trans* isomerization that leads to the nerve impulse that is sent to the brain and results in vision.

Bond and Molecular Polarity. Once the Lewis structure has been drawn and the geometry resolved, it is possible to determine if the molecule is polar. For a molecule to be polar, it first must have polar covalent bonds and it must have a geometry that does not result in the cancellation of bond dipoles. **An understanding of the location of the sites of partial charges that result from bond polarities in a molecule often enables the organic chemist to predict solubility, relative boiling points and what kinds of reagents will attack the substance, the mechanism of the reaction that will occur and the structure of the product that should result.** You have probably learned that bond polarities can be determined by comparing the electronegativities of the two partners in a bond. The classical Pauling scale and a more modern set of electronegativities for elements encountered in most organic compounds are included in the table below.

element	H	Li	B	C	N	O	F	Mg	Si	P	S	Cl	Cu	Br	I
en ¹	2.1	1.0	2.0	2.5	3.0	3.5	4.0	1.2	1.8	2.1	2.5	3.0	1.9	2.8	2.5
en ²	2.1	0.91	2.05	2.54	3.07	3.61	4.19	1.29	1.92	2.25	2.59	2.87	1.8	2.69	2.36

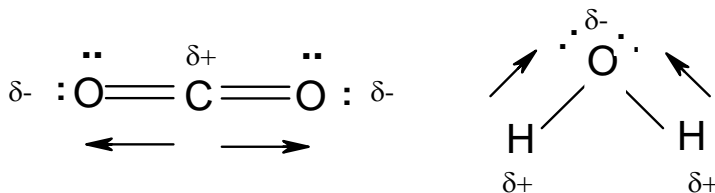
¹ Pauling electronegativity scale ² L. C. Allen, E. T. Knight, *Journal of Mol.St.*, **1992**, 261, 313.

Generally, when the electronegativity difference between two bonding partners is very small (0.4<), the bond behaves as though it is non-polar. When the electronegativity difference between two bonding partners is large (>1.7), the bond behaves as though it is predominantly ionic. Despite the common usage of these criteria in general and organic chemistry courses, use of electronegativities often results in incorrect predictions of bond character and polarity. For example, bonds between transition metals and the halides

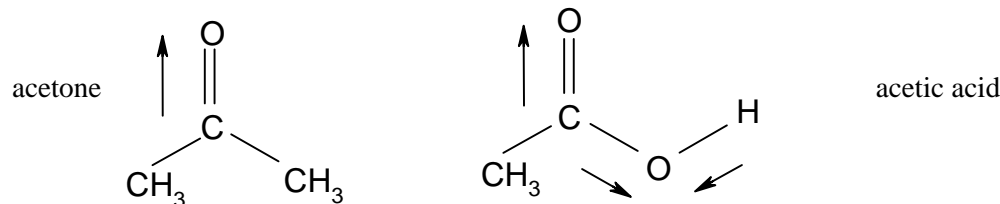
(except for fluorine) with electronegativity differences about 1 are predicted to be covalent. While compounds such as CuBr_2 probably have some covalent character, their behavior in aqueous solution is much better modeled by considering the bond to be ionic. Perhaps even more important to the organic chemist, carbon bonds to sulfur, bromine and iodine are predicted by electronegativity differences to be non-polar. In fact, the electronegativity difference using the modern scale for the carbon - iodine bond predicts a dipole with the positive charge on iodine and the negative charge on carbon. Dipole moment calculations contradict this prediction with CH_3F , CH_3Cl , CH_3Br and CH_3I all having similar dipole moments. Additionally, the assumption that the carbon to halide bond is polar with the partial positive charge on the carbon and the partial negative charge on the halide is consistent with the reaction mechanisms observed for alkyl halides. Finally, use of Pauling electronegativities does not explain why ammonia hydrogen bonds but hydrogen chloride does not demonstrate effective hydrogen bonding.

In this text, instead of using electronegativities to determine bond polarity, it will be assumed that the most common bonds in organic chemistry C - C and C - H bonds are non-polar. Further, it will be assumed that bonds between C and N, O, F, S, P, Br or I are polar covalent with the partial positive charge on the C and bonds between C and metals such as Mg or Li are polar covalent with the partial negative charge on the carbon.

If none of the bonds is polar covalent (all bonds are C-C and C-H), the molecule will not be polar. If there are polar covalent bonds, then it is necessary to consider the geometry of the molecule and the orientation of the dipoles to determine if the molecule is polar. For instance, consider carbon dioxide and



water. Although carbon dioxide has polar bonds, as shown above, the bond vectors or dipoles are equal in magnitude but opposite in direction and cancel out. The result is that CO_2 is non-polar. On the other hand, the bond dipoles for water add in the vertical direction giving the molecule a net dipole. Thus water is very polar. When you open a bottle of soda, carbon dioxide immediately begins to escape from the liquid. As carbon dioxide is non-polar and water is polar, carbon dioxide has a low solubility in water and escapes as soon as the pressure is lowered. For 3-dimensional molecules, the analysis is a little more difficult but is done the same way. For most complex molecules that do have polar covalent bonds, the molecule will be at least somewhat polar as complex molecules usually lack complete symmetry and the bond dipoles seldom cancel.



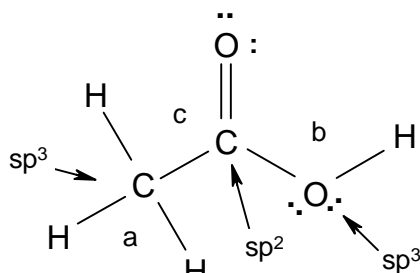
As examples, both acetone and acetic acid are polar molecules.

Molecular models. Most molecular model kits will help you visualize the structures of molecules in three dimensions. However, the most common ball and stick models give slightly incorrect bond angles when multiple bonds are present. For instance, for formaldehyde, the carbon used will have its holes in a tetrahedral arrangement and the H-C-H and H-C-O bond angles appear to be 109.5° and 125.3° respectively. The correct bond angles are very close to 120° . Also the model will show the double bond to consist of two identical bent bonds while bonding models consider the double bond to be composed of two different kinds of bonds: a σ bond and a π bond.

A. Review of General Chemistry

- For each molecule below, draw a Lewis structure, determine the bond angle(s) as predicted by VSEPR or hybridization models (Note: the angle for water predicted by models is 109.5° but is experimentally found to be 104.5° . Give the predicted value rather than the experimental value.), molecular polarity and the hybridization of the central atoms and construct a model.

Example:
 $\text{HC}_2\text{H}_3\text{O}_2$
 acetic acid



$a = 109.5^\circ$
 $b = 109.5^\circ$
 $c = 120^\circ$
 polar molecule

Molecule

Lewis Structure

Molecule

Lewis Structure

- F_2
- N_2
- ICl

- NH_3
- COCl_2 (phosgene)
- HCN

- For the molecules below, draw the two reasonable possible Lewis structures. As above, determine the bond angle(s), molecular polarity and the hybridization of the central atoms. Calculate and indicate values of non-zero formal charges. Based on formal charges, circle the preferred structure and construct a model of it

Molecule

Lewis Structures

- CH_4O
- N_2O
- HClO

- For each of the molecules below, draw the two reasonable resonance structures and indicate the non-zero formal charges that are present (if any) in each of the structures. For each, construct a model of one of the resonance structures. (Hint: sulfur and nitrogen are the central atoms respectively and the hydrogen in nitric acid is bonded to an oxygen.)

Molecule

Lewis Structures

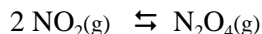
- SO_2
- HNO_3

4. For each polyatomic ion in the chart below, draw all the reasonable resonance structures, indicate nonzero formal charges and construct a model of one of the resonance structures of each ion. Determine the indicated bond angle(s).

	Polyatomic Ion	Lewis Structures
a.	CH_3O^-	
b.	ClO_2^-	
c.	ClO_3^-	
d.	CO_3^{2-}	
e.	N_3^-	
f.	BF_4^-	
g.	HCO_2^-	
h.	NH_2^-	
i.	H_3O^+	
j.	NO_2^+	

Problems 5-7 attempt to demonstrate the importance of Lewis structures. In addition to providing a view of the shape and polarity of the molecule, Lewis structures sometimes provide insight into chemical reactivity. When the Lewis structure indicates some unusual or undesirable characteristic such as high formal charges or unusual oxidation numbers, strained bond angles, unpaired electrons or lack of an octet, there is a strong possibility that the molecule will exhibit extraordinary behavior. Basically this exercise is designed to show you that you can apply your knowledge of chemistry to new situations and think like a chemist.

5. Ozone (O_3) is needed in the stratosphere to absorb (and filter out) potentially damaging ultraviolet light. However, in the lower atmosphere it is a dangerous pollutant as it is a very reactive form of oxygen and as a result very toxic and destructive. Draw the two reasonable resonance structures (Hint: it is not a ring) and indicate the bond angle and the nonzero formal charges. Suggest a reason for its high reactivity.
6. At the high temperatures generated in your car cylinder by the combustion of gasoline, an undesired combination of nitrogen and oxygen takes place to produce some nitrogen monoxide. After emission from the exhaust, the nitrogen monoxide is oxidized by oxygen to nitrogen dioxide. NO_2 is one of the brown colored gases present in smog and is a very dangerous pollutant because of its very high reactivity. NO_2 establishes an equilibrium with its dimer N_2O_4 . Draw Lewis structures of nitrogen dioxide and its dimer, dinitrogen tetroxide.

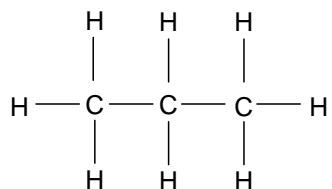


Suggest a reason for the high reactivity of NO_2 and its dimerization reaction. Use formal charges to explain why the equilibrium does not lie far to the right.

7. Hydrogen peroxide (H_2O_2) is a reactive molecule, often used as an antiseptic (perhaps inappropriately) and sometimes used for bleaching. Draw a Lewis structure for hydrogen peroxide. Calculate the oxidation number of oxygen in hydrogen peroxide and suggest a reason for its reactivity.

B. Organic Chemistry

1. Draw a Lewis structure and construct a model of a methane molecule (CH_4). Record the bond angles.
2. Draw a Lewis structure and construct a model of ethane (C_2H_6). Observe the free rotation about the carbon - carbon single bond. The different rotational positions (eclipsed, staggered) are called conformers. Record the bond angles.
3. Draw Lewis structures and construct two models of propane (C_3H_8). Again notice the rotational possibilities and record the bond angles. Compare the structure to the Lewis structure below. Remember that the Lewis structures of 3-dimensional molecules will not correctly show bond angles and the 90° bond angles typically included in Lewis structures are seldom correct.



Although the Lewis structure of propane is technically the figure on the left, it is sometimes more convenient to use partially condensed formulas, $\text{CH}_3\text{-CH}_2\text{-CH}_3$ or even fully condensed formulas, $\text{CH}_3\text{CH}_2\text{CH}_3$.

4. Remove one of the hydrogens from the first carbon on one of the propane models and replace it with a methyl group ($-\text{CH}_3$). This is a model of butane. With the other propane model, replace a hydrogen on the second carbon with a methyl group. This compound is 2-methylpropane (its common name is isobutane). Notice both models have the same formula, C_4H_{10} , but different structures and therefore different chemical and physical properties. Draw Lewis structures of both models and look up the boiling and melting points of the two isomers in the *Handbook of Chemistry and Physics* or on the Internet. The fully condensed formulas for the two isomers are $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$ (*n*-butane) and $\text{CH}_3\text{CH}(\text{CH}_3)\text{CH}_3$ (2-methylpropane).
5. Make all the possible isomers of pentane (C_5H_{12}). Write all the partially condensed formulas for the isomers of pentane.
6. Make all the possible isomers of chlorobutane ($\text{C}_4\text{H}_9\text{Cl}$). Write the partially condensed formulas for the isomers of chlorobutane.
7. Construct two models of methane.
8. In each of the models from 7, replace one hydrogen with one chlorine. Verify that both models are identical or superimposable.
9. Replace a second hydrogen on both of the models from 8 with a bromine to make bromochloromethane. Again verify that the models are identical.
10. Replace a third hydrogen on both of the models from 9 with a fluorine to make bromochlorofluoromethane. Are the models still superimposable? Be careful before you answer this. The students who obtain superimposable models should raise their hands (author's prediction: about one-half of the class will raise their hands). Those who raise their hands should take one model and switch the hydrogen and chlorine atoms. Now no one in the class should have superimposable models. You should observe that the models are nonsuperimposable mirror images (enantiomers) of each other. This phenomenon most commonly occurs when at least one carbon in the molecule has four different groups attached to it (in this case H, Br, Cl, F). This creates a stereogenic center and in this case, a chiral molecule. The phenomenon of chirality is quite common in biological systems. For example, amino

acids are chiral and only one of the two enantiomers of each amino acid is biologically active in animals. Although enantiomers have identical physical properties (melting point, boiling point, density, color), they rotate the plane of polarized light in equal but **opposite** directions. Enantiomers also interact differently with other chiral molecules. Since many molecules in your body are chiral, enantiomers of compounds behave differently in your body. Only one of the two enantiomers of each amino acid is of use to your metabolic system. The compounds in spearmint and caraway (*d* and *l*-carvone) are enantiomers and yet smell differently to most people because the receptors in our noses are chiral.

11. Draw a Lewis structure and construct a model of ethylene (or ethene, $\text{CH}_2=\text{CH}_2$). With most model kits, the model will incorrectly make it appear that the double bond in ethylene consists of two identical bonds. Better model kits show that the double bond consists of a σ and a π bond. Additionally, many kits provide only sp^3 hybridized carbons (4 symmetrically spaced holes 109.5° apart). Thus the model of ethylene will incorrectly show an H-C-H bond angle of 109.5° . Carbon is sp^2 hybridized with 120° bond angles when it has one π bond. Record the correct bond angles. Is rotation possible about the carbon - carbon double bond?
12. Replace a hydrogen on the ethylene with a methyl group to make propene (C_3H_6) and draw its Lewis structure.
13. There are four different types of hydrogens in propene thus there are four different ways to replace a hydrogen with a methyl group. Make all four butenes and draw the Lewis structure of each.
14. Make all the isomers of chloropropene ($\text{C}_3\text{H}_5\text{Cl}$). Write the partially condensed formulas for the isomers of chloropropene.
15. Draw a Lewis structure and construct a model of acetylene (C_2H_2). Record the bond angles.
16. Draw a Lewis structure and construct a model of cyclohexane (C_6H_{12}). What are the bond angles for a planar regular hexagon? With four groups of electrons about each carbon of cyclohexane, what bond angles do the carbons attempt to achieve? Notice that the model does not assume a planar form but is most stable in either a chair or boat conformation. Cyclohexane molecules spend most of their time in the chair conformation. Give the bond angles for the chair conformation.
17. Use of an atomic orbital approach to bonding without adding modifications for electron promotion and hybridization leads to the naive conclusion that carbon atoms should combine with hydrogen atoms to give CH_2 . Draw a Lewis structure for this result. CH_2 (usually called methylene or carbene) actually can be made as a transient species and plays a very important role in synthetic organic chemistry. It has been described as one of the most indiscriminate reagents in organic chemistry. For example, it reacts with ethylene (C_2H_4) to give cyclopropane (C_3H_6). Show this reaction using Lewis structures for reactants and products and suggest a reason for the very high reactivity of methylene.
18. Bromomethane (CH_3Br) reacts with hydroxide ion to give methanol (CH_3OH - see *Problem A-2-a* in this Exercise) and bromide ion. Show this reaction using Lewis structures and use an explanation involving polarity to give a reason for the site of attack of hydroxide on bromomethane.

19. Sodium borohydride (NaBH_4) and lithium aluminum hydride (LiAlH_4) are very useful reducing reagents in organic chemistry. They are commonly used to reduce carbonyl compounds (aldehydes and ketones) to alcohols [e.g., acetone (CH_3COCH_3) to isopropyl alcohol ($\text{CH}_3\text{CHOHCH}_3$)].
- Draw the Lewis structures of acetone and isopropyl alcohol.
 - Draw the Lewis structures of the borohydride and aluminum hydride ions. Calculate the oxidation numbers (Hint: consider electronegativities in chart on page X2-5) of boron, aluminum and hydrogen and suggest a reason for the reducing capability of the two ions.
20. Consider the following information about saturated acyclic hydrocarbons:

Formula	# of isomers
CH_4	1
C_2H_6	1
C_3H_8	1
C_4H_{10}	2
C_5H_{12}	3
C_6H_{14}	5
C_7H_{16}	9
C_8H_{18}	18
C_9H_{20}	35
$\text{C}_{10}\text{H}_{22}$	75
$\text{C}_{15}\text{H}_{32}$	4347
$\text{C}_{20}\text{H}_{42}$	366,319
$\text{C}_{30}\text{H}_{62}$	4,111,846,763
$\text{C}_{40}\text{H}_{82}$	62,491,178,805,831

- Do the numbers above suggest a reason for the existence of the field of organic chemistry?
- Does the rapid increase in the number of isomers after 10 carbons seem intuitively correct? Briefly explain your answer.