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Introduction:

In this unit we'll spend a lot of time looking at molecules with covalent bonds. This unit could also be called organic chemistry – which includes studying the kinds of compounds that are essential for living organisms.

In the first investigation we learn how to represent covalent bonds with Lewis Dot Diagrams. In investigation two we’ll examine the structures and shapes of molecules, and in the third investigation we’ll understand how the structure of the molecule relates to properties like solubility or attraction to other molecules.
Lesson Guide: Smells

Investigation I – Lesson 1

This lesson introduces the study of smell, which runs throughout the unit as a theme. You will be introduced to different smell categories used for identifying and grouping odors. Using these categories you will then classify the odors of various substances. Through the course of two lessons you will discover how smell is related to the name and chemical formula of a molecule.

ChemCatalyst

- What does “smell” mean? Write a definition.
- What do you think your nose is detecting when it smells something?
- What evidence do you have to support your answer?

Notes

Smell classifications are general terms that describe a group of smells. Smell chemists have to agree on a common language in order to talk about smells accurately. There are many smell classifications. The ones listed on the board are five of the categories, which show up consistently in smell research.

Putrid is a word that describes things that smell quite repulsive. Dead animal or really old leftovers from the refrigerator smell putrid.

Camphor is a word used to describe things that smell pungent and medicinal. Camphor is a substance and a word used back in your grandparents’ day. The smell of camphor is quite distinctive. Bengay® or Vicks Vapor Rub® smell like camphor.

Sweet smells include things that are flowery and fruity smelling.

Minty is a green herbal smell that most people are very familiar with and doesn’t usually require explanation.

Fishy smells are very distinctive and don’t usually require an explanation. Most seafoods smell fishy.

Activity

Part I: Smell the samples of common household items. Describe the smell using your new smell classification vocabulary (minty, putrid, fishy, camphor, sweet) in a table on the next page.

At the end of the part of this activity the instructor will make certain that the class comes to some sort of consensus regarding the classifications each household item fits into. While everyone may not agree on the assignment of these categories, it is important that we work from a common starting point.
Lesson Guide: Investigation I – Lesson 1

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Part I Table:

<table>
<thead>
<tr>
<th>Sample</th>
<th>Substance (Name or description)</th>
<th>Smell Classification</th>
<th>Class Consensus</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
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</tr>
</tbody>
</table>

Part II: Smell the five mystery smells and classify them using the smell classifications.

<table>
<thead>
<tr>
<th>Vial</th>
<th>Smell classification</th>
<th>Chemical Name</th>
<th>Molecular Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>L-carvone</td>
<td></td>
<td>C_{10}H_{14}O</td>
</tr>
<tr>
<td>B</td>
<td>phenylethylamine</td>
<td></td>
<td>C_{8}H_{11}N</td>
</tr>
<tr>
<td>C</td>
<td>amyl propionate</td>
<td></td>
<td>C_{8}H_{16}O_{2}</td>
</tr>
<tr>
<td>D</td>
<td>isoamyl acetate</td>
<td></td>
<td>C_{7}H_{14}O_{2}</td>
</tr>
<tr>
<td>E</td>
<td>menthone</td>
<td></td>
<td>C_{10}H_{18}O</td>
</tr>
</tbody>
</table>

Look for patterns in the mystery vial data. In your workbook, write down at least eight patterns you discover between the data and the various smells. Look for how smell relates to chemical structure- what elements are involved, how many of each, ratio of atoms, etc...

Making Sense Question:
Is there any evidence that smell, molecular formula, and chemical name are related? Explain.
Part III of this lesson involves making predictions and testing them. Below you will find some new data, for three new molecules. Use the patterns you discovered to predict how these new molecules would smell. After you make your predictions your instructor will allow you to smell the three new vials.

<table>
<thead>
<tr>
<th>Vial</th>
<th>Chemical Name</th>
<th>Molecular Formula</th>
<th>Predicted Smell</th>
<th>Actual Smell</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>ethyl valerate</td>
<td>C7H14O2</td>
<td></td>
<td></td>
</tr>
<tr>
<td>G</td>
<td>butyric acid</td>
<td>C4H8O2</td>
<td></td>
<td></td>
</tr>
<tr>
<td>H</td>
<td>ethyl acetate</td>
<td>C4H8O2</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Making Sense Question:
What was your reasoning for each of the three predictions you made?

Check-in

- How would you expect a molecule with the molecular formula C8H16O2 to smell? Explain.

Homework

Complete the following for homework:

1. Name two things that you have learned about smell so far.

2. List at least five questions you have about the sense of smell. Keep in mind the variations we observed when we smelled and named the substances in the vials. Make sure that at least two of your questions have something to do with molecules.

3. Predict the smell of the following molecules. Explain your reasoning.
   - methyl octanoate
   - monoethylamine
   - C9H18O2
   - C2H3N

3. What do you think 1,5 pentanedithiol (C5H12S2) would smell like? How confident of your answer are you? Explain why.

4. Let’s say you smelled 30 vials, all with substances that contain two oxygen atoms in their molecular formulas. Twenty-nine of them smell sweet and one smells really putrid. Can you still claim a connection between molecular formula and smell? Explain your reasoning.
In this lesson you are introduced to the **structural formulas** of the molecules you have smelled. These are two-dimensional drawings of molecules. These drawings show you how the atoms in a molecule are arranged and connected. During this class you will be introduced to the concept of **functional groups**.

**ChemCatalyst**

- Here are drawings of two molecules that you’ve already smelled. List at least three differences and three similarities between the two molecules.

```
H   O   H   H     H   H   O
H—C—C—O—C—C—H     H—C—C—C—O—H
H   H   H           H   H   H
Molecule #1          Molecule #2
```

**Hint:** *If you count the carbon, hydrogen, and oxygen atoms in each molecule you may find some similarities. If you notice where these different atoms are located within each molecule you may find some differences. These molecules are from the previous lesson, and they had very different smells.*

**A structural formula** is a drawing or diagram that a chemist uses to show how the atoms in a molecule are connected. It is a two-dimensional drawing of a molecule.

Chemists refer to the connections between atoms in a molecule as **bonds**. In structural formulas, the covalent bonds are represented as lines. A **double bond** is represented by two lines together.

**Activity**

**Instructions:** Use the functional group cards provided by your instructor and work with your group. Sort the cards any way you wish, as long as it is consistent and makes some sense. Then, complete the following:

1. Describe how your group sorted the structural formula cards:

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2. Look for patterns in your groupings. Write down as many patterns as you can find within the groups.

3. Look for differences between groupings. Write down as many differences between the groupings as you can find.

**Making Sense Question:**

What structural features seem to be the best predictors of the smell of a molecule?

**Functional groups** are structural features that show up repeatedly in molecules and seem to account for some of their chemical properties.

**Making Sense**

You may have noticed that certain structural patterns are repeated in the different molecules. Your instructor will introduce you to the names of some of the different functional groups in today’s molecules. You should take notes on your sheet of binder paper. An example of a functional group would be a carbon atom double-bonded to an oxygen atom.

**Check-in**

- If a molecule is sweet, what other things do you know about it?

- List at least four things that are probably true about that molecule.
  - 
  - 
  - 
  - 
  - 
1. Two kids each start out with 20 wooden boards of the same dimensions, 50 roofing tiles, and 100 nails. Working on their own and using a hammer and a saw they each build two very different doghouses. Explain how this story relates to what you learned in class today.

2. Molecules G and H each have the same molecular formula yet one smells sweet and the other smells putrid. Explain how you think this might be possible.

3. In each of the six molecules below... (1) circle the functional group present, (2) write the name of the functional group, (3) write the molecular formula

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a)</td>
<td>d)</td>
</tr>
<tr>
<td>b)</td>
<td>e)</td>
</tr>
<tr>
<td>c)</td>
<td>f)</td>
</tr>
</tbody>
</table>
3.1.2 NOTES

**Functional groups:** Structural features that show up repeatedly in molecules and seem to account for some of their chemical properties (like smell!).

![Structural formulas of various functional groups](image)

Examples of molecules with these functional groups:

- **Ketone**
- **Ester**
- **Carboxylic acid**
- **Amine**

**Structural Formula:** A drawing or diagram that a chemist uses to show how the atoms in a molecule are connected. It is a 2-dimensional picture of the molecule. Structural formulas show how the atoms in a molecule are put together.

**Bond:** A connection between atoms in a molecule. In structural formulas, the covalent bonds are represented as lines. Double bond = 2 lines together.

Do you see any patterns between the smells and structures in the pictures below?

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>1-carvone</td>
<td>minty</td>
</tr>
<tr>
<td>J</td>
<td>pulegone</td>
<td>minty</td>
</tr>
<tr>
<td>E</td>
<td>menthone</td>
<td>minty</td>
</tr>
<tr>
<td></td>
<td>Molecule Name</td>
<td>Flavor</td>
</tr>
<tr>
<td>---</td>
<td>---------------------</td>
<td>---------</td>
</tr>
<tr>
<td>K</td>
<td>diisobutylamine</td>
<td>fishy</td>
</tr>
<tr>
<td>B</td>
<td>phenylethylamine</td>
<td>fishy</td>
</tr>
<tr>
<td>I</td>
<td>isopentanoic acid</td>
<td>putrid</td>
</tr>
<tr>
<td>G</td>
<td>butyric acid</td>
<td>putrid</td>
</tr>
<tr>
<td>C</td>
<td>amyl propionate</td>
<td>sweet</td>
</tr>
<tr>
<td>D</td>
<td>isoamyl acetate</td>
<td>sweet</td>
</tr>
<tr>
<td>F</td>
<td>ethyl valerate</td>
<td>sweet</td>
</tr>
<tr>
<td>L</td>
<td>hexyl acetate</td>
<td>sweet</td>
</tr>
<tr>
<td>H</td>
<td>ethyl acetate</td>
<td>sweet</td>
</tr>
</tbody>
</table>
Lesson Guide: HONC If You Like Molecules

Investigation I – Lesson 3

Some structural formulas have long chains of atoms and others have rings or are branched. Now that you have looked at two-dimensional pictures of molecules, you will begin to learn why atoms arrange themselves within a molecule in certain ways.

ChemCatalyst

Answer the following question:

- Examine the following molecules. What patterns do you see in the bonding of hydrogen, oxygen, carbon and nitrogen?

![Molecules with bonding patterns](image)

Vial K - fishy
diisobutylamine

Vial E - minty
menthone

Your teacher will show you additional examples, what bonding patterns do H, O, N, and C follow?

Activity

The purpose of this activity is to give you practice in creating structural formulas from molecular formulas and to help you begin to understand why atoms end up in the specific arrangements we find them in.

Instructions: Your task is to create structural formulas for the following three molecular formulas.

\[ \text{C}_3\text{H}_8 \quad \text{C}_3\text{H}_8\text{O} \quad \text{C}_3\text{H}_9\text{N} \]

1. Start with the carbon atoms. Connect them together.
2. Now insert the nitrogen, oxygen, or other atoms. Remember, they may connect on the ends of the carbon chain, or somewhere in the middle.
3. Fill in with the hydrogen atoms.
4. Problem-solve until you have the correct number of bonds for each element. Keep working until your structures follow the HONC 1234 rule.
Making Sense:
The third molecular formula has at least three possible structures. Are these all the same molecule? Explain.

**Isomers** are two or more molecules that are composed of the same elements in the same proportions but differ in properties because of differences in structural formulas.

### Check-in

Answer the following question:

- Are the following molecules correct according to the HONC 1234 rule? If not, what specifically is wrong with them?

![Molecules A and B](image)

### Homework

1. Use the HONC 1234 rule to create possible molecules with the following molecular formulas. Remember it is easiest to start with the carbon atoms.
   a. $C_3H_8O_2$
   b. $C_4H_{11}N$
   c. $C_4H_{10}$
   d. $C_5H_{12}O$

2. Draw one alternate structural formula (isomer) for one of the molecules a-d in problem #1. This structure must be different from the one you drew in the table above!

3. Answer the following questions for the “molecules” pictured below:

![Molecules a to e](image)

   a. Is each structural formula correct according to the HONC 1234 rule?
   b. For any molecules that don’t follow the HONC 1234 rule, repair the incorrect molecules. *Note: there may be more than one correct way to repair a molecule!*
   c. Write the molecular formula for each molecule.
Lesson Guide: Connect the Dots

Investigation I – Lesson 4

In this lesson you will begin to understand why atoms connect to each other the way they do. You will be introduced to a tool that will assist you in building molecules and predicting how many bonds an element will have.

ChemCatalyst

Answer the following question:

- This is a drawing of the structural formula of a methane molecule. The lines represent bonds. Explain what you think a bond is.

![Structural formula of methane molecule]

Notes

This particular unit focuses on one specific type of bond, called a covalent bond. A methane molecule has four covalent bonds. We can replace each of the lines in the structural formula of methane with a pair of electrons to show the covalent bonds. A covalent bond is defined as a link between two atoms that are sharing a pair of electrons. Finally, we can explode this drawing to illustrate from where each electron originated (see below).

A covalent bond is a connection that forms between two atoms when those atoms are sharing a pair of electrons between them.
We see that four electrons came from the carbon atom and one electron came from each one of the hydrogen atoms. This corresponds to the number of valence electrons for carbon and hydrogen. If we look at the periodic table we see that hydrogen is in Group 1 and carbon is in Group 4. As you may recall from the first unit, the valence electrons are the ones involved in bonding and reactivity. When we draw an atom using dots to represent the valence electrons it is called a Lewis dot symbol. When we draw a molecule using dots to represent the valence electrons it is called a Lewis dot structure.

The drawing in the middle is called the Lewis dot structure for the molecule, methane. It shows the valence electrons for each atom bonded together. The drawing on the right shows the Lewis dot symbols of the individual atoms of hydrogen and carbon. This means that nitrogen has 3 electrons that can be paired up with electrons from other atoms:

\[
\text{\textbullet} \quad \text{\textbullet} \quad \text{\textbullet} \quad \text{\textbullet} \quad \text{\textbullet} \quad \text{\textbullet} \\
\bullet \quad \text{N} \quad \bullet
\]

A Lewis dot structure is a way to represent a molecule using dots in place of the valence electrons of the atoms. Lewis dot structures keep track of the electrons involved in bonding and show where they came from.

A Lewis dot symbol shows the valence electrons for a single atom.

**Activity**

In this lesson you will begin to understand why atoms connect to each other the way they do. You will be introduced to a tool, called Lewis dot symbols, which will assist you in building molecules and predicting how many bonds an element will have.

**Part I**

1. In the space provided below, draw Lewis dot symbols on the elements below. When you are done, answer the following questions.

   C       N       O       F       Ne

   Si      P       S       Cl       Ar
2. Do you notice any patterns in the Lewis dot symbols above? Explain.

3. Determine how many bonds each group of elements will make.

<table>
<thead>
<tr>
<th>Group number</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
<th>VIII</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of valence electrons</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Number of bonds</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>First row elements</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
</tr>
<tr>
<td>Second row elements</td>
<td>Si</td>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td>Ar</td>
</tr>
</tbody>
</table>

4. What relationship do you notice between the group number and the number of bonds that the elements in that particular group make?

5. Explain why nitrogen, N, in Group 5 does not make five bonds even though it has five valence electrons.

**Part II**
The following drawing represents four hydrogen atoms combining with a carbon atom to make a methane molecule. Use it as an example to guide you in answering the questions below.

6. Create a drawing that would represent three hydrogen atoms combining with one nitrogen atom to create an ammonia molecule, NH$_3$.

7. Create a drawing that would represent two hydrogen atoms combining with one oxygen atom to create a water molecule, H$_2$O.
If you finish early... 
Draw a Lewis dot structure for HF.

Making Sense

You may have noticed that some of the electrons in a Lewis dot structure are involved in bonding while others are not. The electrons that are involved in a covalent bond are called **bonded pairs**. The electrons that are not involved in bonding are referred to as **lone pairs**. Below is an illustration showing the Lewis dot structure of ammonia with a lone pair and a bonded pair labeled.

\[ \text{lone pair} \]  
\[
\begin{array}{c}
\text{H} \\
\vdots \\
\text{N} \\
\vdots \\
\text{H} \\
\end{array}
\]

\[ \text{bonded} \]

Based on what you’ve learned in this lesson, explain why the HONC 1234 rule works.

Check-in

Answer the following questions:

- Draw the Lewis dot symbol for the element I, iodine. Explain how you arrived at your particular drawing.
- How many covalent bonds does iodine make?

Homework

1. Draw the Lewis dot symbols for the following elements:

   Se  Br  K  As  Ga  Ge

   a) Arrange them in order of their group numbers.

   b) Determine how many covalent bonds each would make.

2. In your own words explain what you know about the HONC 1234 rule. How does it help you in creating molecular structures
Lesson Guide: Eight Is Enough

Investigation I – Lesson 5

In this lesson you will use Lewis dot structures to create structural formulas of molecules containing elements in addition to \( \text{H} \), \( \text{O} \), \( \text{N} \), and \( \text{C} \). You will look for patterns in the number of electrons surrounding each atom in a Lewis dot structure in order to develop further understanding of bonding.

Answer the following questions:

- **Draw the Lewis dot structure** for the following covalently bonded molecule \( \text{Cl}_2 \) in the chart below. **Explain** how you arrived at your answer.
- Try to come up with other molecules that might have similar Lewis dot structures. To answer this question, think about where chlorine is located on the periodic table.

<table>
<thead>
<tr>
<th>Lewis Dot Structure for ( \text{Cl}_2 )</th>
<th>Explanation</th>
<th>Other molecules which might be similar</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Activity**

**Part I:** Draw Lewis dot structures for the following molecules. First, draw the Lewis dot symbols for all the atoms involved. Put your final drawings in the boxes below.

<table>
<thead>
<tr>
<th>( \text{Br}_2 )</th>
<th>( \text{H}_2\text{S} )</th>
<th>( \text{PH}_3 )</th>
<th>( \text{SiH}_4 )</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
1. Look at your drawings of Br₂, H₂S, PH₃, and SiH₄ on the previous page:
   a) How many electrons surround the bromine atom on the left, including the
      ones it is sharing?
   b) How many electrons surround the bromine atom on the right, including
      the ones it is sharing?
   c) How many electrons surround sulfur, phosphorus, and silicon?
   d) How many electrons does each hydrogen have access to in these
      molecules?

Part II: Answer the following questions. Begin your drawings with the central
atom(s).

2. How many chlorine atoms would bond covalently with a single carbon atom?
   Explain. (Hint: Use Lewis dot symbols to help you figure this out.)

3. Create a drawing representing the bonding of chlorine atoms with a single
   carbon atom. Start with the Lewis dot symbol for carbon.

4. How many fluorine atoms would bond covalently with a single phosphorous
   atom? Explain.

5. Create a drawing representing the bonding of fluorine atoms with a single
   phosphorous atom. Start with the Lewis dot structure of phosphorous.

6. Draw the structural formula of the molecule you created using Lewis dots
   symbols in question 5.

7. What is the molecular formula of the molecule you created in question 5?

Making Sense:
The noble gases do not form bonds with other atoms (except under very extreme
conditions). Explain why you think this might be true (use your Lewis dot
structures).

If you finish early......
Draw the Lewis dot structure for a molecule of ethane, C₂H₆. Start by drawing the
Lewis dot structures of the central atoms, then add hydrogen atoms as needed.
In general, elements bond in such a way as to create a more stable form. The more stable the form (or compound), the more likely we are to find that it exists naturally. The most stable form for bonded atoms appears to be one in which each atom is surrounded by eight valence electrons. This is called the octet rule. (See the illustration of phosphorus trifluoride at right.)

The octet rule states that atoms tend to form bonds by sharing valence electrons until eight valence electrons surround each atom.

Check-in

Answer the following questions:

1. Which of the following compounds satisfy the HONC 1234 rule?

2. Which of the following satisfy the octet rule?

3. Which of the following would result in stable compounds we might find in the world around us?

   (A) CH₃  (B) NH₃  (C) CH₄  (D) NH₄
Homework

1. The following compounds can be found in nature or prepared in a laboratory. They are all compounds made from a main group element and the maximum number of hydrogen atoms that will bond with that element:

   HBr    SiCl\textsubscript{2}H\textsubscript{2}    SbH\textsubscript{3}    H\textsubscript{2}Se    GeH\textsubscript{4}

   a) Draw Lewis dot structures for each molecule.

   b) Which molecules have the same total number of electrons?

   c) Which molecules have the same number of valence electrons?

   d) What patterns do you notice in the number of H atoms?

2. Draw Lewis dot structures for the following molecules.

   F\textsubscript{2}    H\textsubscript{2}Te    HOCl    CH\textsubscript{2}Cl\textsubscript{2}

3. Consider the following formulas and answer the questions:

   NH\textsubscript{3}    NH\textsubscript{4}

   a) Which of the formulas satisfy the HONC 1234 rule?

   b) Which of the formulas satisfy the octet rule?

   c) Which of the formulas would result in a stable compound we might find in the world around us?
Lesson Guide: Dots, Dots, and More Dots

Investigation I – Lesson 6

In this lesson you will work to create structural formulas for various molecules. You will start with the Lewis dot symbols of individual atoms. These atoms can then be arranged in more than one way to create molecules. Finally, structural formulas will be translated from the Lewis dot structures.

**ChemCatalyst**

Answer the following questions:

1. Describe the “octet rule”:

2. Here are the structural formulas for $\text{N}_2$ (nitrogen gas), $\text{O}_2$ (oxygen gas), and $\text{F}_2$ (fluorine gas). Draw the Lewis dot structures for these three molecules; represent all the bonded pairs and lone pairs of electrons as dots!

| Structural formulas: | $\text{N}=\text{N}$ | $\text{O}=\text{O}$ | $\text{F}−\text{F}$ |

**Notes**

Sometimes atoms share more than one pair of electrons in order to satisfy the octet rule. Atoms can share four, or even six electrons, in a single covalent bond. These bonds are called **double** and **triple bonds**, respectively.

For example, carbon dioxide ($\text{CO}_2$) requires double bonds in order to achieve a stable octet for all of the atoms involved. Your teacher will walk you through this, take notes in the space provided below:
In this lesson you will work to create Lewis structures for various molecules. Use the strategies outlined below to assist you. Remember to satisfy the octet rule (H is an exception!), and use HONC 1234 and other bonding patterns you noticed in the last 3 lessons.

**Strategies for drawing Lewis Dot Structures:**

1. Identify the central atom of the molecule and work from there.
   - The central atom is usually the one that can form the most bonds
   - Carbon can make 4 bonds, so it is often central
   - Hydrogen is never the central atom (it can only form 1 bond!)
   - There may be more than one central atom- you may end up with a carbon "chain" in the center, for example.

2. Determine the number of valence electrons for each atom and draw the Lewis dot symbols for the individual atoms.

3. Begin to combine atoms with the central atom, keeping HONC 1234 in mind. When drawing bonds, you may replace two shared electrons (dots), with one line, if you wish.

4. Check to see that all the atoms satisfy the octet rule (they are sharing eight electrons; two electrons for hydrogen atoms).

5. Count the total number of valence electrons in your molecule. See if it matches the total number of valence electrons in the individual atoms.

6. If your structure does not satisfy HONC or octet, you may need to consider double or triple bonds.

<table>
<thead>
<tr>
<th>Molecular Formula</th>
<th>Draw lewis structure for starting atoms</th>
<th>Identify the central atom(s)</th>
<th>Draw the lewis structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>$C_2H_6$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$C_2H_4$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Chemical</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>----------</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C$_2$H$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_4$O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_2$O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_5$N</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_3$N</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>HCN</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_4$O$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CH$_2$O$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Making Sense Question:
Explain how HONC 1234 helps you to check the structural formulas you create.

If you finish early...
Predict how the molecules in your data table will smell from what you know so far about molecular formula, structural formula, and smell. Explain your choices.

Check-in
Answer the following question:
- We know two things about a certain molecule. We know that its molecular formula is C₂H₄O₂ and we know that it has one C=O in it. Using Lewis dot symbols and the octet rule to guide you, draw at least one possible structure for this molecule. (There are a total of three possible.)

Homework
Complete the following for homework:
1. What can you do to check that your Lewis dot structures are correct? Name at least three ways.
2. Hydrogen is an example of an element that forms one bond. Which other elements form one bond? Explain why.
3. Using Lewis dot structures and the octet rule, draw all the stable molecules with the molecular formula C₃H₈O. (Hint: There are 3 molecules total.)
4. Draw the Lewis dot structures for the following molecules.
   a. N₂O₂F₂H₃
   b. N₂O₂F₂H₂
5. Draw the Lewis dot structures for the following molecules:
   HCN     CH₃OH     H₂O₂     H₂N₂
   CH₃Cl    AsH₃     I₂       H₂S
Lesson Guide: New Smells, New Ideas

Investigation I – Lesson 7

In this lesson you will be introduced to five new molecules. These molecules will lead you in the direction of new discoveries about the relationship between smell and chemistry.

ChemCatalyst

Answer the following question:

- Do you think any of these molecules will smell similar? What evidence do you have to support your prediction?

- citronellol $\text{C}_{10}\text{H}_{20}\text{O}$

- geraniol $\text{C}_{10}\text{H}_{18}\text{O}$

- menthol $\text{C}_{10}\text{H}_{20}\text{O}$
## Activity

**Instructions:** Predict the smell of the following molecules, then answer the questions.

<table>
<thead>
<tr>
<th>Vial</th>
<th>Molecular formula and name</th>
<th>Functional group</th>
<th>Structural formula</th>
<th>Predicted Smell</th>
<th>Actual Smell</th>
</tr>
</thead>
<tbody>
<tr>
<td>O</td>
<td>C₁₀H₂₀O citronellol</td>
<td>alcohol</td>
<td><img src="image1" alt="Structural formula" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>P</td>
<td>C₁₀H₁₈O fenchol</td>
<td>alcohol</td>
<td><img src="image2" alt="Structural formula" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Q</td>
<td>C₁₀H₁₈O geraniol</td>
<td>alcohol</td>
<td><img src="image3" alt="Structural formula" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>R</td>
<td>C₁₀H₂₀O menthol</td>
<td>alcohol</td>
<td><img src="image4" alt="Structural formula" /></td>
<td></td>
<td></td>
</tr>
<tr>
<td>S</td>
<td>C₁₀H₁₈O borneol</td>
<td>alcohol</td>
<td><img src="image5" alt="Structural formula" /></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
1. For each structural formula, find the location of the functional groups.

2. Can functional group alone be used to classify these five molecules according to their smell? Why or why not?

3. Can molecular formula alone be used to classify these five molecules according to their smell? Why or why not?

4. Are there any structural similarities besides functional group that could be used to classify these molecules? If so, what are they?

**Making Sense**

Review the results of the smell investigation to date by adding a Smells Summary Chart in your workbook. Your teacher will make available the functional group cards and other resources you’ve used so far this unit.

Draw a central circle and write: “How does structure affect smell” in the center. Draw five spokes out of the central circle for each of the five functional groups or smells that we have studied. Answer the following questions by adding information to each spoke:

(1) how molecular formulas can be used to predict smell,

(2) how name can be used to predict smell,

(3) how functional group can be used to predict smell,

(4) what other information might be important.
Homework

1. Predict the smell of the following molecules. Give evidence to support your predictions.
   a. $C_{10}H_{16}O$
   b. $C_{8}H_{19}N$
   c. ethyl butyrate
   
   \[
   \begin{array}{c}
   \text{H} \\
   \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{O} - \text{H} \\
   \text{H} \quad \text{H} \quad \text{H}
   \end{array}
   \]
   d. 
   e. $C_{6}H_{12}O_{2}$

2. What kind of generalization could we make about the smell of molecules with alcohol functional groups?

3. If a molecule smells minty, what can you say about its chemistry?

4. If you have a molecule with 10 carbon atoms and 1 oxygen atom (plus some hydrogen atoms), what could it smell like? Draw a molecule like this with a correct structural formula. Predict the molecule's smell based on the structural formula you have drawn.

5. What could a molecule with 8 carbon atoms and 2 oxygen atoms (plus hydrogen atoms) smell like? Draw 2 different structural formulas for this particular molecular formula and label the smells.

6. What are some possible things that might be different about two molecules with the same functional group and the same molecular formula?
In this class you will gain practice creating three-dimensional models of some small molecules. The concept of electron domains helps to explain why molecules exist in crooked and bent shapes, rather than straight lines.

Answer the following question:

- Here is the structural formula of ethanol. Which is the correct ball-and-stick model for ethanol? Explain your reasoning.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{C} - \text{C} - \text{O} - \text{H} \\
\text{H} & \quad \text{H}
\end{align*}
\]

(1) \hspace{1cm} (2) \hspace{1cm} (3) \hspace{1cm} (4)
Sets of electrons that remain together in bonds or in lone pairs are referred to as **electron domains**. Electron domains “prefer” to be as far apart as possible from each other within a molecule.

**Activity**

**Purpose:** In this class you will gain practice creating three dimensional models of some small molecules. The concept of electron domains helps to explain why molecules actually exist in crooked and bent shapes, rather than straight lines.

**Draw Lewis dot structures for the following molecules:**

- CH$_4$
- NH$_3$
- H$_2$O

1. How many electron domains are located around the central atom of each molecule?

2. If each central atom is sharing eight valence electrons, what is different about these three molecules?

**Using gumdrops and toothpicks, create ball-and-stick models of NH$_3$ and H$_2$O.**

3. Did you remember lone pairs? How could you represent lone pairs in your model?

**Fix your models if you need to so that lone pairs are represented.**

4. Do the lone pairs have an effect on the shape of your molecule?

**Fix your models if the lone pairs are not yet affecting the shape of the model.**

5. Compare your three models – are there any similarities?

**Making Sense Question:**

Explain how the lone pairs affect the shape of your molecules.
Check-in

Complete the following task:

- Build a model for C₂H₆. Be sure to show all of the lone pairs.

Homework

1. If you were going to predict the 3-dimensional structure of a small molecule what would you want to know?
2. Using the periodic table as a reference, predict the 3-D structure of H₂S.
3. Draw the Lewis dot structure for C₂H₂. How many electron domains does the molecule have? What shape would you predict for C₂H₂?
4. What shape do you predict for carbon dioxide, CO₂? Draw the Lewis dot structure and a picture of the molecular model.
Lesson Guide: Let’s Build It!

Investigation II – Lesson 2

In this lesson you gain practice creating ball-and-stick models from molecular formulas, using Lewis dot structures to assist you.

**ChemCatalyst**

Answer the following question:
(Your instructor will have molecular models on hand for you to refer to.)

- Remove the lone-pair paddles from all five models. Now describe the remaining geometric shape.

```
CH₄      NH₃      H₂O      HF      Ne

H     |     H     |     N     |     O     |     F
H     |     C     |     H     |     H     |     Ne
H     |     H     |     H     |     H     |      |
```

Your instructor will pass around ball-and-stick models of these molecules for you to examine. The paddle shaped extensions on the models help to show the space occupied by the lone pairs of electrons in these molecules.

Following the ChemCatalyst, your instructor will introduce the appropriate names for shapes of small molecules.

**Activity**

**Purpose:** In this lesson you gain practice creating actual ball-and-stick models from molecular formulas, using Lewis dot structures to assist you. For each molecule, first draw the Lewis dot structure. Then create a 3-D molecular model, using the Lewis Dot structures as a guide. Some molecules have more than one shape in them – if this is the case name the shapes around the central atoms.
<table>
<thead>
<tr>
<th>Molecular Formula</th>
<th>Lewis dot structure</th>
<th>Describe/ Draw Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>methane: CH₄</td>
<td>![Lewis dot structure for methane]</td>
<td>Tetrahedral</td>
</tr>
<tr>
<td>water: H₂O</td>
<td>![Lewis dot structure for water] or ![Lewis dot structure for water]</td>
<td>Bent</td>
</tr>
<tr>
<td>carbon dioxide: CO₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>chloromethane: CH₃Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ethane: C₂H₆</td>
<td></td>
<td></td>
</tr>
<tr>
<td>methanol: CH₃OH</td>
<td></td>
<td></td>
</tr>
<tr>
<td>methyl amine: CH₃NH₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>formaldehyde: CH₂O</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Lesson Guide: Investigation II – Lesson 2

Let’s Build It!

<table>
<thead>
<tr>
<th>Molecular Formula</th>
<th>Lewis dot structure</th>
<th>Describe/ Draw Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>ethene (ethylene):</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{C}_2\text{H}_4$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen cyanide:</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{HCN}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ethyne (acetylene):</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{C}_2\text{H}_2$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Check-in

Answer the following questions:

- What is the shape of the following molecule?

  $\text{H}_2\text{S}$

  Draw Structure

  Name of Shape:

Homework

1. Describe the shape of each of the following molecules. (Use Lewis dot structures, the periodic table, and HONC 1234 to assist you.)

   $\text{Cl}_2$  $\text{CO}_2$  $\text{H}_2\text{O}$  $\text{H}_2\text{S}$  $\text{CH}_2\text{Cl}_2$

2. Which of the above molecules have the same shape?
Lesson Guide: Investigation II – Lesson 3

Lesson Guide: Shape Matters

ChemCatalyst

Answer the following questions:

- Write chemical formulas for the following two molecules.
- Are these two representations of the same molecule? Why or why not?
- Do you expect these two molecules to have similar properties? Why or why not?

![fumaric acid](image)

![maleic acid](image)

Activity

Purpose: To compare the properties of maleic acid and fumaric acid, two compounds with identical molecular formulas.

Materials

- maleic acid
- fumaric acid
- sodium carbonate
- thymol blue indicator
- 2 test tubes (10 mL)
- 2-cm strip of magnesium
- spatula
- distilled water
- large well plates (with at least 6 wells)

Safety Note: Everyone must wear safety googles at all times.

Procedure

1. Label a micro test tube for maleic acid. Use the tip of a spatula to measure a small amount of maleic acid. Add it to the tube.
2. Label a micro test tube for fumaric acid. Use the tip of a spatula to measure a small amount of fumaric acid. Add it to the tube.
3. Pour 5 mL of distilled water into each test tube. Swirl the tubes to dissolve as much of the solids as possible. Allow any solid that does not dissolve to settle to the bottom of the tube.
4. Record your observations in the table on the worksheet.
5. Obtain a well plate with at least six wells. Keep track of the six wells you will be using as follows:
6. Fill three wells halfway with the maleic acid solution that you prepared in Step 3.
7. Fill three wells halfway with the fumaric acid solution that you prepared in Step 3.
8. Add a drop of thymol blue indicator to a well with maleic acid and one with fumaric acid. Observe what happens.
9. Add a small strip of magnesium to a well with maleic acid and one with fumaric acid. Observe carefully what happens to the magnesium.
10. Add a tiny crystal of sodium carbonate, Na$_2$CO$_3$, to a well with maleic acid and one with fumaric acid. Observe what happens.
11. Record your observations in the table provided.

**Data Table**

<table>
<thead>
<tr>
<th>Property</th>
<th>Maleic Acid (C$_4$H$_4$O$_4$)</th>
<th>Fumaric Acid (C$_4$H$_4$O$_4$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solubility</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction with thymol blue</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction with magnesium</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction with Na$_2$CO$_3$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

1. Maleic acid and fumaric acid both have the same molecular formula, but their chemical properties are different. Explain why.

2. Draw Lewis dot structures for maleic acid and fumaric acid based on the structural formulas in the chem catalyst. Include the lone pairs.

3. Provide evidence to support the following statement: It is not possible to rotate one end of a molecule around a carbon double bond, C=C. (Hint: If the molecule could rotate freely around the double bond, what properties would you expect from the two samples you studied in the lab?)
4. What evidence do you have that molecular shape is related to chemical properties?

**If you finish early. . .**

Draw another molecule with the same molecular formula as these two molecules, but with a shape that differs from both maleic acid and fumaric acid.

**Making Sense**

Your instructor may show you several ball-and-stick models to demonstrate the restrictions a carbon=carbon, C=C, double bond places on certain molecules.

This activity has highlighted a special form of molecular structure. Maleic acid and fumaric acid are called isomers of one another. Isomers have the same molecular formula, but different shapes and hence different properties. The shape with both H atoms pointing in the same direction is referred to as the cis isomer. When the H atoms point in opposite directions, the isomer is referred to as a trans isomer. Thus, maleic acid is the trans form of C₄H₄O₄ while fumaric acid is the cis form of C₄H₄O₄.

**Homework**

Complete the following for homework:

Do a lab write-up for the experiments you did with maleic and fumaric acid. Be sure to include the following parts: Title, Purpose, Data, Results, Discussion, and Conclusions.

1. There are three different forms of C₂H₂F₂.
   
   a. Draw the structural formulas of the three forms. Include the lone pairs.
   
   b. Do you expect the properties of these molecules to be the same? Why or why not?
Lesson Guide: Attractive Molecules

Investigation III – Lesson 1

The big question for today: How do molecules interact with each other?

ChemCatalyst

This demonstration shows that different molecules interact with water molecules in different ways, and how molecular interactions might be connected to the process of dissolving.

Demonstration: Water and Charge. Describe what happens when a charged object is brought close to a stream of water?

![Diagram of attractive and repulsive forces]

<table>
<thead>
<tr>
<th>Attractive and Repulsive Forces</th>
</tr>
</thead>
<tbody>
<tr>
<td>Like charges <strong>REPEL</strong></td>
</tr>
<tr>
<td>Opposite charges <strong>ATTRACTION</strong></td>
</tr>
<tr>
<td>Neutral = No attraction OR repulsion</td>
</tr>
</tbody>
</table>

40
Purpose: In this lesson, you will observe the response of certain liquids to a charged wand, the behavior of the same liquids as droplets on waxed paper, and their ability to dissolve in water. These activities give you information about possible interactions between molecules.

Part I:
With your group, test the four liquids at the stations in your classroom with a charged wand. Record your observations in the Data Table below. Answer the following questions before moving on to Part II.

1. What evidence do you have that some of the molecules you just tested may have a charge on them?
2. How would you explain any liquids that were not attracted to the charged wand?

Part II:
Place a single droplet of each of the four liquids onto a piece of waxed paper. Take care to leave space between each droplet. Record your observations in the Data Table. Answer the following questions before moving on to Part III.

3. Were there any differences between the four droplets? Describe the differences you noticed. (Hints: How flat or round is the drop? Did it evaporate quickly?)
4. Why didn’t all of the droplets behave the same way when placed on the waxed paper?

Part III:
Place a single droplet of each of the four liquids on a piece of paper towel. Take care to leave space between the droplets. Record your observations of the droplets in the Data Table. Answer the following question.

5. Why do the droplets behave differently on the waxed paper and on the paper towel?

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Results of charged wand test</th>
<th>Results of waxed paper test</th>
<th>Results of paper towel test</th>
</tr>
</thead>
<tbody>
<tr>
<td>water</td>
<td>attracted</td>
<td></td>
<td></td>
</tr>
<tr>
<td>acetic acid</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>isopropanol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>hexane</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Making Sense:

1. Here is an artist’s interpretation of what is happening between the water molecules and the charged wand. Describe in your own words what you think is happening in the picture.

2. If water molecules are carrying a partial charge, as shown in the following picture, how do you think a group of water molecules would behave towards each other? Draw a picture of several water molecules interacting below to illustrate your thinking. (Draw at least two more).

Check-in

The charged wand experiment provides evidence of molecules interacting with a charge. Not only do some molecules interact with a charge, they also interact with each other. The interactive forces between molecules are called intermolecular forces or intermolecular attractions. As part of the Making Sense Discussion your instructor will introduce you to polar and nonpolar molecules.

1. What is a polar molecule?
2. What is a non-polar molecule?
3. Ammonia is a polar molecule. Based on what you learned from this lesson
   a. How will it react to a charged wand? (Refer to Data Table)
   b. How will it behave when placed onto waxed paper?

Homework

1. Explain why hexane does not bend in the direction of a charged wand. (Is hexane polar or non-polar?)
2. Predict why vegetable oil and water do not mix (like in Italian salad dressings that have two liquid layers).
Lesson Guide: Polar Bears and Penguins

Investigation III – Lesson 3

In this lesson you will be exploring polarity and bonding between atoms in greater detail. A comic book will provide new information about these topics and will introduce you to the concept of electronegativity, which helps us to understand partial charges.

**Purpose:** How can we explain partial charges on molecules?

Answer the following questions:

1. Draw the Lewis dot structure for HCl:

2. If the penguin represents a hydrogen atom and the polar bear represents a chlorine atom, what does the ice cream represent in the drawing?

3. What do you think the picture is trying to illustrate?

4. Would HCl, liquid hydrochloric acid, be attracted to the charged wand? Explain your thinking.
Activity

Read the comic book called “The Bare Essentials of Polarity”. Answer the following questions.

1. How does the comic book define a “polar molecule?”

2. Define electronegativity as you understand it, after reading the first two pages of the comic book.

3. (Optional) Interpret the picture at the bottom of the first page. Explain how the iceberg, penguins, and polar bears represent trends in electronegativity.

4. What is the artist trying to represent when there are two polar bears arm wrestling together, or two penguins arm wrestling together?

5. What three types of bonds are represented on the third page of the comic book? What happens to the bonding electrons in each type of bond?

6. Explain why there are four scoops of ice cream in the illustration of O₂ on page 3 of the comic strip.

7. What do the six scoops of ice cream represent in the illustration of N₂ on page 4 of the comic strip?

8. Describe what you think is happening to the penguin in the CO₂ molecule in the picture on page 4.

9. Name three things that the picture of CO₂ on page 4 illustrates about the molecule.

10. Describe what you think is happening to the penguins in the illustration of H₂O on page 4?

11. Explain what you think the crossed arrow represents in the comic book.

12. What are the two definitions of “dipole” given in the comic book?

If you finish early. . . (Optional)

- Using polar bears and penguins, create an illustration showing a hydrogen sulfide molecule, H₂S. (Hint: You may wish to start with a Lewis dot structure.)

- Go back to the previous lesson and look at the structures for 16-mercaptophexadecanoic acid and 1-hexadecanethiol. Which one is polar and which one is non-polar? Does this make sense based on your observations of their interactions with water?
**Making Sense:** When two atoms with different electronegativities are bonded together, they tend to attract the bonded electrons to different degrees. This causes the electrons to spend more time around one of the atoms, resulting in a partial negative charge on this atom. This tendency of an atom to attract electrons shared between two atoms is called **electronegativity**. An atom that strongly attracts the shared electrons is considered highly electronegative. The atom with lower electronegativity will end up with a partial positive charge on it. The result is a polar bond. Chemists have a specific name for a molecule that has two poles - it is called a **dipole**.

Polar molecules are also called **dipoles**. The prefix di- means two. A dipole is a molecule with two partially charged ends, or poles.

Chemists refer to polar molecules as dipoles and they also say that molecules with polar bonds have dipoles. These multiple definitions can be a bit confusing.

**Check-in**

Answer the following questions:

1. Is the bond between these atoms polar? Explain your reasoning.
2. How would the atoms be portrayed in the comic book – as polar bears, penguins, or both? Explain.
3. What does electronegativity have to do with polarity?
The Bare Essentials of Polarity

You don’t have to go to the ends of the earth to find polar molecules. They’re all over the place. A polar molecule is just a molecule with a difference in electrical charge between two ends.

The electrical imbalance of polarity is caused by differences in electronegativity between atoms. Electronegativity is the ability of an atom/nucleus to attract bonding electrons toward itself.

In HCl, the bonded pair of electrons spends more time near the chlorine’s nucleus because chlorine is more electronegative than hydrogen.

The periodic table shows a general trend in the electronegativity of the elements. Electronegativity tends to rise as you move “northeast” on the periodic table, and fall as you move “southwest.”

Note: The noble gasses, in the periodic table’s far right column, are often assigned an electronegativity value of zero because they are relatively nonreactive.
When two atoms with unequal electronegativity values bond, they do not share the bonding electrons evenly. The bonding electrons spend more time around the more electronegative atom, creating a PARTIAL NEGATIVE CHARGE on that atom. The other atom then has a PARTIAL POSITIVE CHARGE, and the bond is polar.

Polarity is unfair!

So the polarity of a bond is a function of the difference between the electronegativity values of two bonding atoms. Bonded atoms with equal electron-attracting strength will have nonpolar bonds.

However, if the electronegativity of two bonded atoms is unequal, then their bond will be polarized—maybe a little...

...maybe a lot.
Because the elements have such varying electronegativities and can come together in so many different combinations, there is really a **continuum of polarity in bonding**. For convenience, we can break the continuum down into three categories: (1) nonpolar covalent, (2) polar covalent, and (3) ionic.

**Nonpolar Covalent**

The clearest examples of nonpolar covalent bonds are those between identical atoms, such as H₂, N₂, O₂, or Cl₂. Bonds between atoms with nearly the same electronegativity value, such as carbon and hydrogen atoms, are usually also considered nonpolar. Remember, this is really a continuum, and conventional distinctions are somewhat artificial.

**Polar Covalent**

In a polar covalent bond, two atoms still share bonded pairs of electrons, but those electrons are decidedly more attracted to one atom than the other. Examples include bonds between carbon and oxygen atoms, or between hydrogen and fluorine atoms.

**Ionic**

At the extreme of difference in electronegativity, polar covalence shades into the winner-take-all situation of ionic bonding. The more electronegative atom seizes all the bonding electrons and becomes a negative ion, while the other atom becomes a positive ion. The opposite charges on the ions attract each other.
Polar bonds between atoms constitute DIPOLES. Actually, the word "dipole" can refer to several different things that are relevant here: (1) the polarity of an individual polar bond between atoms, (2) the net polarity of a polar molecule that may have several polar covalent bonds within it, and (3) the polar molecule itself.

Confusing? Let's look at some examples:

An \( \text{N}_2 \) molecule isn't a dipole (it's not a polar molecule), and it doesn't have any dipoles (polar bonds) within it.

\( \text{HCl} \) has a dipole (a polar bond) and it is a dipole (a polar molecule).

On the other hand, \( \text{CO}_2 \) has two dipoles (two polar bonds), but the \( \text{CO}_2 \) molecule itself is not a dipole because its polar bonds cancel each other out and make the molecule nonpolar overall.

Like \( \text{CO}_2 \), \( \text{H}_2\text{O} \) has two dipoles (two polar bonds). But because of \( \text{H}_2\text{O} \)'s bent shape (caused by lone pairs of electrons on the oxygen atom), \( \text{H}_2\text{O} \) also has a dipole in the sense of an overall polarity. So \( \text{H}_2\text{O} \) is a dipole in the sense of being a polar molecule.

The polarity of molecules can affect many of their other properties, such as their solubility, their boiling and melting points, and their odor.
Lesson Guide: Thinking (Electro)Negatively

Investigation III – Lesson 4

This lesson explores electronegativity in a quantitative fashion – that is, it applies numbers to our investigation of polarity. Using the electronegativity scale it is possible to compare atoms and find out which ones will attract electrons more strongly in a bond.

Describe how the illustration and the table might relate to each other.
Electronegativity determines the polarity of bonds

In a molecule such as H₂, the atoms are identical, so there is no difference in the degree to which each atom in the bonded pair attract the shared electrons. These molecules are covalent and nonpolar.

Nonpolar covalent bonds are really the only true covalent bonds. If the electronegativities between two atoms are even slightly different they form what is called a polar covalent bond. In polar covalent bonds the bonding electrons are located closer to the more electronegative atom.

When the electronegativities between two atoms are greatly different the bond is called an ionic bond. In the case of an ionic bond the electron of one atom is completely given up to the other atom. The result of this kind of bond is two charged ions. The more electronegative atom becomes an ion with a full negative charge and the less electronegative atom (sometimes called the electropositive atom) becomes an ion with a full positive charge.

Very generally speaking, when the difference between the electronegativities of two bonded atoms is greater than 2.1, the bond is considered ionic. NaCl, sodium chloride (table salt), is an example of a compound with an ionic bond. The difference in the electronegativities of Na and Cl is greater than 2.1. It is important to note that there is no hard and fast dividing line between these three different types of bonds – they are on a continuum.
Chemists wanted a way to measure the relative attraction for the pair of electrons shared between two atoms. In 1932 Linus Pauling created the scale for electronegativity given on the previous page. The scale indicates how strongly an atom attracts shared electrons. **An atom with a large electronegativity attracts shared electrons very strongly.**

Discuss as a class:
- What happens to the electronegativity values across each period from left to right?
- What happens to the electronegativity values of each group from the bottom to the top?

1. If you have a bond between a metal atom and a nonmetal atom, which of the two is more electronegative? Explain your thinking.
2. If the difference in electronegativity is greater than 2.1, then the bond is considered ionic. List three examples of pairs of atoms with ionic bonds.
3. If the difference in electronegativity between two bonded atoms is smaller than 2.1, but greater than 0.5, then the bond is polar covalent. List three examples of pairs of atoms with polar covalent bonds.
4. Metal atoms tend to form cations (ions with positive charges). Is this consistent with the electronegativity of metal atoms? Why or why not?
5. Do nonmetal atoms tend to form cations with positive charges or anions with negative charges? Explain your thinking.
6. List the following molecules in order of increasing partial positive charge on the H atom from the smallest to largest: CH₄, NH₃, H₂O, and HF.
7. An acid is a compound in which the H atom has a large partial positive charge. Which of the molecules in question 16 is the most acidic? Explain your reasoning.

**Making Sense**
8. Explain how you would determine both the direction and degree of polarity of a bond between two different atoms using the electronegativity scale.
9. Arrange them in order from least polar bond to most polar bond.

   C - H   C - S   H - F   C - O   H - Br   C - N
1. List two similarities and two differences between covalent and ionic bonds.

2. Describe or draw what happens to the electrons in a polar covalent bond, a nonpolar bond, and an ionic bond.

3. Which molecule would be more polar – HI (hydrogen iodide) or HCl (hydrogen chloride)? Show how you arrived at your answer.

4. Circle the atom in each pair below that will attract shared electrons more strongly:
   a) C or Cl  
   b) Rb or Br  
   c) I or In  
   d) Ag or S  
   e) As or Na  
   f) H or Se

5. Which two atoms from question 4 form the most polar bond?

6. List at least three examples of pairs of atoms with nonpolar covalent bonds.

7. Sulfur forms both ZnS and SF₂. Is sulfur the most electronegative element in both compounds? Why or why not?
Lesson Guide: I Can Relate

Investigation III – Lesson 5

The goal of this lesson is to give you practice in determining the polarity of small molecules with more than two atoms.

ChemCatalyst

Answer the following question:

1. HCl, (hydrogen chloride) and NH₃ (ammonia) dissolve easily in water. O₂, N₂, and CH₄ (oxygen, nitrogen, and methane) do not dissolve easily in water. What is one major difference between the two molecules dissolve and the ones that do not?

2. Now, consider three molecules Cl₂, and Br₂, and I₂, which are all made of halogens with 7 valence electrons on each atom and are all non-polar. However, Cl₂ is a gas, Br₂ is a liquid, and I₂ is a solid at room temperature and atmospheric pressure. Why do you think this is so?
**Intermolecular Forces**

Intermolecular forces (IMF) are the forces that attract molecules to one another. These forces are much weaker than intramolecular bonds such as ionic and covalent bonds. The strength of the IMF between molecules affects a number of properties of the compound.

All intermolecular forces are thought to result from the attraction of opposite charges. In other words, the positive end of one molecule attracts the negative end of another molecule and vice versa. All molecules contain positive and negative charges within them, so all can attract each other but the strength of this attraction can vary greatly from compound to compound. Intermolecular forces are therefore broken into three categories:

1. **Dispersion Forces**
   
   Even nonpolar molecules attract each other. This is thought to be due to temporary dipoles lining up and attracting the temporary or induced dipoles of nearby molecules. You may recall that electrons are thought to move around randomly within atoms. As a result, the electron density in different parts of a molecule can vary over time as electrons jump around from one part of the molecule to another. This causes temporary dipoles: temporary separations of charge within the molecule. This temporary separation of charge can cause or “induce” polarity in a nearby molecule. The forces of attraction between neighboring temporary dipoles are called dispersion forces. They are sometimes called London forces. All molecules have temporary dipoles, so all molecules have dispersion forces acting.
   
   **FOR NON-POLAR MOLECULES, THE ONLY TYPE OF INTERMOLECULAR FORCES IS DISPERSION FORCES.** The greater the number of electrons in a molecule, the more ways there are for temporary dipoles to exist. For this reason, the dispersion forces are greater for larger, more massive molecules.

2. **Dipole-Dipole Forces**
   
   The attractive forces polar molecules have with other polar molecules. A permanent dipole means a separation of charge. **ONLY POLAR COMPOUNDS HAVE DIPOLE-DIPOLE FORCES.** Therefore, the composition of the molecule and the molecular shape are important to determining whether a molecule will have dipole-dipole forces; the molecule must have some asymmetry for a permanent dipole to exist; i.e. different surrounding atoms or an asymmetric geometry. You have seen examples of these kinds of polar molecules in Lesson 3.

3. **Hydrogen Bonding**
   
   A “Hydrogen bond” refers to the attraction of one molecule (which contains Hydrogen) for another molecule. Hydrogen bonds are not actual bonds, they are just particularly strong dipole-dipole forces that occur in certain compounds. When hydrogen is involved in a very polar bond (with a very electronegative partner), the hydrogen forms a region of particularly dense positive charge within a molecule. This happens when H is (polar covalently) bonded to F, O or N within the molecule, the large difference in electronegativity causes hydrogen’s only electron to be pulled relatively far away from the single proton that makes up hydrogen’s nucleus. And because hydrogen is so small, this charge is concentrated in a small volume. For example, if a Phosphorus atom had a slight positive charge due to a polar bond, the charge would be spread out over a whole big atom with 15 protons and 15 electrons. But if Hydrogen has a slight positive charge, that charge is very concentrated. This makes it very attractive to its slightly negative neighbors.

**Relative Strength of Intermolecular Forces**

In general, **H-bonds > dipole-dipole > dispersion.** If two substances have the same type of intermolecular force acting, the substance with the greater molar mass will typically have a greater overall attractive force.

Adapted from Chemistry Study Guide Unit 5, Harvard-Westlake school (HWSCIENCE.COM)
Activity

Using the cut-out molecules, create a poster that shows each of the three types of interactions: Dispersion Forces, Dipole-Dipole Forces, and Hydrogen Bonding.

Instructions
1. Classify the 7 molecules on the handout as polar or non-polar.
2. Cut the molecules apart and look for pairs that will interact. You may have pairs of molecules of the same substance (e.g. two water molecules) or you may have pairs of molecules of different substances (e.g. a water molecule interacting with a molecule of ammonia).
3. Pay special attention to the molecules that contain H, N, O, or F. These molecules may have hydrogen bonds.
4. When you have found pairs of molecules that illustrate each type of intermolecular force, glue the molecules onto your poster and draw in where the attractions between molecules would occur. Include at least 3 molecules of each type, and accurately show how they would interact with each other.
5. For each type of interaction, include a brief written explanation (2-3 sentences minimum) describing how your cut outs demonstrate that intermolecular force.

Making Sense:
1. Review at least 3 posters made by other groups.
2. As you look at the poster, answer the following questions:
   - What molecule was used for each type of force?
   - Is this an appropriate choice (think: does the molecule have a permanent dipole? is H bonded to O, N or F? is the molecule non-polar?)?
3. Be prepared to share the information you gathered with the class.

Check-in

Due to differences in electronegativity we expect HCN, hydrogen cyanide, to be polar. Since water is a polar molecule as well, which way do you think water and hydrogen cyanide molecules would orient with each other? Explain your reasoning.

If you finish early: What kind of intermolecular force would be present between a water molecule and an oxygen molecule? Draw a diagram of this interaction.
For each molecule (#1-4):

- Draw the Lewis dot structure
- Name the geometric shape of the molecule
- Identify the molecule as polar or non-polar
- Identify the intermolecular force present between molecules and explain how you determined which IMF was present

<p>| | |</p>
<table>
<thead>
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<tbody>
<tr>
<td>1. H₂O</td>
<td>2. NF₃</td>
</tr>
<tr>
<td>3. CCl₄</td>
<td>4. HBr</td>
</tr>
</tbody>
</table>

5. Which of the substance would have a higher boiling point, a substance which is polar or a substance which is non-polar? Explain your reasoning.

6. Rank the following substances in order of strength of intermolecular attractions, from weakest to strongest: CH₄, H₂O, or H₂S? Explain your reasoning.

7. Which molecule would you predict would be most likely to dissolve in water, ethane (C₂H₆) or hydrogen cyanide (HCN)? Explain your reasoning. Include structures of all 3 molecules.
Cut-out Molecules

<table>
<thead>
<tr>
<th>NH₃</th>
<th>H₂S</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₃F</td>
<td>CF₄</td>
</tr>
<tr>
<td>CH₄O</td>
<td></td>
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<tr>
<td>HCl</td>
<td>H₂CO</td>
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Note: The images of the molecules are shown in the document.