Experiment 13: EFFECT OF STRUCTURE ON PHYSICAL PROPERTIES

**Purpose:** The relationships between the structures of molecules and their physical properties, such as volatility, viscosity, and solubility are examined.

**Introduction:**

In the first semester of General Chemistry you had learned to tell from the chemical formula whether a substance is ionic or molecular. For molecular substances you had learned to predict molecular geometry and molecular polarity of a substance based on its molecular formula. You have probably wondered why you needed to know how to make such predictions. The reason is much of the chemical and physical nature of a substance is based on whether it is ionic or molecular, and if molecular, its nature is based on its molecular geometry and polarity. These attributes determine the type of *intermolecular forces* of the substance.

The term, *intermolecular forces*, refer to the interaction between a molecule and its neighboring molecules. They explain many of the common phenomena we see around us, such as why oil and water do not mix, why table salt dissolves easily in water and not so well in oil, why rubbing alcohol evaporates much faster than water, and why honey is more viscous than water. Based on the type of intermolecular forces, we can make predictions on the physical properties of the substance and the chemical outcome of a reaction. Or, having observed the outcome, we could explain it based on chemical theories. After all, that is what much of chemistry is about, to be able to make predictions and to offer explanations to phenomena observed.

In this experiment, we will focus on the physical properties based on the type of intermolecular forces present in a substance. Do not confuse *intermolecular* forces with *intra*molecular forces. *Intramolecular* forces refer to the forces that hold atoms together within a molecule or polyatomic ion, namely, the covalent bonds. *Intermolecular* forces are between different molecules, also referred to as *van der Waals forces*. Technically speaking, we should be talking about *interparticle* forces because sometimes we will be looking at the interaction of species other than molecules, such as the interaction between ions, or interaction between ions and molecules. The term, *particle*, is generic, referring to molecules as well as ions. Often, as is the case in this experiment, the term, *intermolecular forces*, is used with the understanding that it extends to ionic species as well.

The relative strengths of the intermolecular forces are of particular interest. When given two substances, they help us predict which has the higher boiling point, viscosity and surface tension, and to a certain extent, the solubility of a substance in a given solvent.

Review the chapter in your textbook for details relating to intermolecular forces. The following is a few major points that you should remember:

- The three types of intermolecular forces are **London forces, dipole forces and hydrogen bonding (H-bonding)**.
London forces (also known as London dispersion forces) exist for all substances, regardless of whether it is an ion, atom, molecule, polar molecule or nonpolar molecule. For nonpolar species, London dispersion forces are their only intermolecular forces.

Dipole forces (also known as dipole-dipole forces) exist only for polar molecules. It is the attraction between the dipoles of one molecule to the dipoles of a neighboring molecule. The more polar a molecule, the stronger is the dipole force.

H-bonding (also known as hydrogen bonds) refers specifically to the intermolecular force between molecules containing H-F, H-N or H-O bonds. It is the attraction of these special hydrogen atoms to the dipole of a neighboring molecule. Remember these are not the covalent bonds within the molecule.

If the sizes of the species are comparable, intermolecular forces and the corresponding species, in increasing order of strength are as follows:

- London forces < dipole forces < H-bonding << ionic bonds/covalent bonds/metalllic bonding
- nonpolar molecules < polar molecules < molecules with H-bonding << ionic compounds, network solids, metals
- The larger the size of the species, the stronger the intermolecular force is. This is because electrons of larger species are more polarizable, thus forming stronger London forces. A very large nonpolar molecule could have a stronger intermolecular force than a small polar molecule even though London forces are listed as being “weaker” than dipole forces. Remember that the order of strength of intermolecular forces listed above is only if we are comparing particles of similar sizes.

In this experiment you will make observations on physical properties of various compounds and try to relate them to the type and relative strengths of intermolecular forces.

Safety Precautions

- Be sure that your safety goggles are in place during all parts of this experiment.
- The compounds n-pentane, n-hexane, n-octane, n-propanol, n-octanol, acetone, are all flammable and should not be used in the presence of an open flame.
- All mixtures in Part IV should be disposed of in the designated waste container in the hood.

Procedure: You may work individually or with one partner.

First work through the procedure and record careful and complete observations in your lab notebook. When the entire class is finished, we will go over the discussions together. While you are waiting for others to finish their work, you can begin completing the Discussion pages at the back of this experiment.

At the end of the lab you will be turning in your carbon copies of your lab notebook which contains your observations and answers to some of the questions (as specified in the procedure).

Part I: Relative Volatilities

Volatility is a measure of how easily a liquid vaporizes. You should think about how the strength of intermolecular forces would affect the ease of a liquid to convert to a gaseous phase. Would a volatile substance have a strong or weak intermolecular force?
1. Turn an extra large (800-mL or 1-L) beaker upside down, and make sure the bottom of the beaker is clean and dry. Place 2 drops each of n-octane, n-hexane and n-pentane on the bottom of the beaker far apart enough that they do not run into each other. Gently blow across these drops. The substance that evaporates fastest is the most volatile.

*Remember to answer these questions on the Discussion page provided.*

a) Record the names of the compounds in order of increasing volatility.

b) Examine the structures on the Structure Page. What type of intermolecular forces is present for each of the molecules?

c) What is the relationship between the structure and the volatility of the molecules? Explain why fully.

2. Repeat the above procedure using 2 drops each of n-propanol, water, and n-pentane. Which of these seems to be the most and which the least volatile? List the compounds in order of increasing volatility.

Does the size of the molecule explain the order of volatility? (Relative size can be estimated from the molecular weights of these substances.) Include their molar masses in your answer. Explain fully.

What are all the intermolecular forces for each of the three compounds? Can they be used to explain the order of volatility? Explain fully.

**Part II: Relative Viscosities of Materials**

*Viscosity* is a measure of a liquid’s resistance to flow. A liquid with high viscosity is thick and moves slowly. You should think about how the strength of intermolecular forces would affect the ease of flow of a liquid. Would a viscous liquid have a strong or weak intermolecular force?

1. Examine the three vials of substances labeled as 124-IIA, 124-IIB and 124-IIC. DO NOT OPEN THE VIALS! By swirling or gently shaking the tubes, observe which of these liquids has the highest viscosity and which, the lowest. Record your observations in your lab notebook.

The compounds in the vials are n-hexane, glycerol and ethylene glycol, listed here in random order. Examine their structures and identify the content of each vial based on their viscosities.

a) Write the full structure of each of these compounds in the Discussion page and explain how their intermolecular forces helped you identify the compounds.

**Part III: Solubility**

Much of solubility can be predicted by the general rule “Like Dissolves Like.” It means that a substance dissolves best in a solvent that is most similar to it in character (in the type of intermolecular forces involved).
Examine the six vials labeled 124-III A, 124-III B, 124-III C, 124-III D, 124-III E and 124-III F. DO NOT OPEN THE VIALS! Each vial contains a pair of liquid substances, some of which are miscible (mutually soluble) and others, immiscible. Examine each vial and record in your notebook as to whether the two liquids are miscible or immiscible. If the two liquids form two layers, they are considered immiscible. Look closely from the side. It is difficult to see the phase boundary if both liquids are colorless.

<table>
<thead>
<tr>
<th>Vial</th>
<th>Contents</th>
</tr>
</thead>
<tbody>
<tr>
<td>124-III A</td>
<td>n-hexane and n-octane</td>
</tr>
<tr>
<td>124-III B</td>
<td>n-hexane and water.</td>
</tr>
<tr>
<td>124-III C</td>
<td>ethylene glycol and water</td>
</tr>
<tr>
<td>124-III D</td>
<td>sucrose and water</td>
</tr>
<tr>
<td>124-III E</td>
<td>acetone and water.</td>
</tr>
<tr>
<td>124-III F</td>
<td>n-octanol and water</td>
</tr>
</tbody>
</table>

a) Identify the predominant type of intermolecular force for each compound based on its structure and complete the table on the Discussion page.

b) For each pair of compounds that was miscible, state the predominant intermolecular force between the two compounds that allowed them to be miscible.

c) Use your observations to explain the rule “Like dissolves like.” Be specific.

d) IN YOUR LAB NOTEBOOK, draw the structure of acetone and water. Orient them in such a way that it shows the hydrogen bonding between them. Use dotted lines to show the hydrogen bonding.

e) For n-octanol and water, explain why you might have expected them to be miscible and offer an explanation why they are not.

f) How does the structure of sucrose explain why table sugar dissolves easily in water?

Part IV: Solubility of Solids in Liquids

1. Obtain six clean small test tubes from your drawer and make sure three of them are dry. Place 20 drops of n-hexane in the three dry test tubes and approximately the same volume of deionized water in the other three.

2. Add a small sample (size of a pea) of nickel(II) sulfate hexahydrate in a test tube containing n-hexane, and add a sample to a test tube containing water. Mix each by holding the test tube between your thumb and index finger and rock the test tube back and forth. DO NOT PUT YOUR THUMB OVER THE TOP AND SHAKE IT! (Your instructor can show you the technique if you have not learned it from your previous chemistry courses.) Did the solid dissolve in n-hexane? in water? Record your observations.

3. In two other test tubes, repeat with a small sample of iodine first in n-hexane, and then in water in two other test tubes. What is the formula for iodine? Remember it is a diatomic molecule! Did it dissolve in n-hexane? in water?

4. Repeat with a small sample of naphthalene. Did it dissolve in n-hexane? in water?

5. Dispose of all mixtures in the designated container in the hood.

Note: In your discussion of nickel(II) sulfate hexahydrate, you can ignore the six water molecules of hydration as it plays no part in the phenomenon you are observing.

ANSWER THESE QUESTIONS IN YOUR LAB NOTEBOOK:

a) Is the nickel(II) sulfate more soluble in n-hexane or in water?
b) Is the iodine more soluble in n-hexane or in water?
c) Is the naphthalene more soluble in n-hexane or in water?
d) Classify the type of compound (ionic? polar molecule? nonpolar molecule?) and the predominant type of intermolecular force for each compound.
e) Use your observations to explain the rule “Like Dissolves Like.”
f) Suppose you had a mixture of solid iodine and nickel(II) sulfate and to this solid mixture you added a mixture of water and n-hexane. Make a sketch of a test tube and its contents to show what you would expect to see.
   Test your prediction experimentally. Be sure to use enough nickel(II) sulfate so that one of the layers is distinctly colored. Mix well. Record your observations fully.
g) How did the results compare with your prediction? Explain your observations using intermolecular forces and the rule “Like dissolves like.”

Part V: Relationship of Boiling Point and Molar Mass
Calculate the molar mass of each of the 4 substances to 3 sig. fig. and write them in the table on the Discussion Page. On the graph paper provided, prepare a graph of boiling point versus molar mass of the compounds shown below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Boiling Point</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>100.0°C</td>
</tr>
<tr>
<td>H₂S</td>
<td>−60.3°C</td>
</tr>
<tr>
<td>H₂Se</td>
<td>−41.2°C</td>
</tr>
<tr>
<td>H₂Te</td>
<td>−2.2°C</td>
</tr>
</tbody>
</table>

a) Based on intermolecular forces, what do you expect to be the relationship between molar mass and the boiling point?
b) How does your graph compare with your prediction? Why is one point seemingly out of line? Explain fully.

Pre-Lab Assignment: (There will be no quiz. Instead prepare your notebook as described.)
1) First review the Introduction section and the section in your textbook on intermolecular forces. Complete the tables specified first in the lab manual, then copy the table and your entries in your lab notebook. These will be collected at the beginning of class.

A Structure Page showing the structures of the organic (i.e., carbon-containing) compounds you will be using is provided. Use this page to help you determine the types of intermolecular forces present, and develop explanations for the phenomena observed.
1. Tables 1b) and 2 b) on p. 130
2. Table in Part III a) on p. 131
3. Table in Part IV d) on p.133
4. Table in Part V. Plot the graph on p. 134 (not in your notebook), but write the answers to the questions in your lab notebook.

2) Preparation of Notebook:
On a new page, set up separate tables in your lab notebook for each part of the experiment to record your observations clearly. Each table should contain two columns. In the left column
you should list the reagent/reagents you are observing, and on the right column, you will be recording your observations.

Thought Questions (These will not be collected, but you would want to think about them in preparation for your lab final.)

1. Would you expect calcium chloride to be more soluble in methanol or in carbon tetrachloride? Explain your reasoning in terms of the structures and predominant intermolecular forces in these three compounds.

2. Suppose that you were given a mixture of sand, iodine, sodium chloride, and n-pentane to separate. Using the information gathered on these and similar compounds in this experiment, and using what you know about the properties of these materials from other contexts, outline a procedure that could be used to do a quantitative separation of these four materials.

3. The compounds dimethyl ether and ethanol have the same molecular formula, C2H6O. (That is, they are isomers of one another.) Their normal boiling points are significantly different, however, with one boiling at –23.6°C and the other boiling at 78.5°C. Show the structures of these two compounds and explain, in terms of their structures and the resulting intermolecular forces, why these two isomers boil at such different temperatures. Which isomer has which boiling point?
STRUCTURE PAGE

(Note: These give only the arrangement of the atoms; they are not meant to be complete Lewis structures. Nonbonding electrons may be present but are not shown here.)

\[ \text{\begin{align*}
\text{n-pentane} & \quad \text{n-hexane} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}} \]

\[ \text{\begin{align*}
\text{n-octane} & \quad \text{n-propanol} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}} \]

\[ \text{\begin{align*}
\text{glycerol} & \quad \text{ethylene glycol} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}} \]

\[ \text{\begin{align*}
\text{acetone} & \quad \text{naphthalene} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}} \]

\[ \text{\begin{align*}
\text{sucrose} & \quad \text{n-propanol} \\
\text{HOCH}_2 & \quad \text{HO} \\
\text{C} & \quad \text{H} \\
\text{OH} & \quad \text{H} \\
\text{CH}_2 & \quad \text{H} \\
\end{align*}} \]
Part I: Relative Volatilities

1a) List the compounds in order of increasing volatility:

\[ \text{______________} < \text{______________} < \text{______________} \]

1b) | Type of Intermolecular Force |
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>n-octane</td>
<td></td>
</tr>
<tr>
<td>n-hexane</td>
<td></td>
</tr>
<tr>
<td>n-pentane</td>
<td></td>
</tr>
</tbody>
</table>

1c) What is the relationship between the structure and the volatility of the molecules? Explain why.

2a) List the compounds in order of increasing volatility:

\[ \text{______________} < \text{______________} < \text{______________} \]

2b) | Molar Mass | All Types of Intermolecular Force for each |
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>n-propanol</td>
<td></td>
</tr>
<tr>
<td>water</td>
<td></td>
</tr>
<tr>
<td>n-pentane</td>
<td></td>
</tr>
</tbody>
</table>

2c) Does the size of the molecule explain the order of volatility for these three compounds? Can the intermolecular forces explain the order of volatility? Explain fully.
Part II: Relative Viscosities of Materials

<table>
<thead>
<tr>
<th>Name of Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>124-IIA</td>
</tr>
<tr>
<td>124-IIB</td>
</tr>
<tr>
<td>124-IIC</td>
</tr>
</tbody>
</table>

a) Write the full structure of each compound and explain how their intermolecular forces helped you identify the compounds.

Part III: Solubility

a) 

<table>
<thead>
<tr>
<th>Predominant Type of Intermolecular Force for each</th>
</tr>
</thead>
<tbody>
<tr>
<td>n-hexane</td>
</tr>
<tr>
<td>n-octane</td>
</tr>
<tr>
<td>water</td>
</tr>
<tr>
<td>ethylene glycol</td>
</tr>
<tr>
<td>sucrose</td>
</tr>
<tr>
<td>acetone</td>
</tr>
<tr>
<td>n-octanol</td>
</tr>
</tbody>
</table>

b) For each pair of compounds that was miscible, state the predominant intermolecular force between the two compounds that allowed them to be miscible.
c) Use your observations to explain the rule “Like Dissolves Like.” Be specific.

d) Draw the structure of acetone and water. Orient them in such a way that it shows the hydrogen bonding between them. Use dotted lines to show the hydrogen bonding.

e) For n-octanol and water, explain why you might have expected them to be miscible and offer an explanation why they are not.

f) How does the structure of sucrose explain why table sugar dissolves easily in water?

Part IV: Solubility of Solids in Liquids

a) Nickel(II) sulfate is more soluble in ______________.

b) Iodine is more soluble in ______________.

c) Naphthalene is more soluble in ______________.
d) Formula | Type of Compound | Predominant Type of Intermolecular Force
---|---|---
nickel(II) sulfate |  |  
iodine |  |  
naphthalene |  |  

e) Use your observations to explain the rule “Like Dissolves Like.”

f) Suppose you had a mixture of solid iodine and nickel(II) sulfate and to this solid mixture you added a mixture of water and n-hexane. Make a sketch of a test tube and its contents to show what you would expect to see.

g) How did the results compare with your prediction? Explain your observation using intermolecular forces and the rule “like dissolves like.”
Part V: Relationship of Boiling Point and Molar Mass

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<tr>
<td>H\textsubscript{2}O</td>
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<td>–60.3°C</td>
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<td>– 2.2°C</td>
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a) Based on intermolecular forces, what do you expect to be the relationship between molar mass and the boiling point?

b) How does your graph compare with your prediction? Why is one point seemingly out of line? Explain fully.