Evaporation and Intermolecular Forces

Learning Outcomes

1. Draw alcohols and alkane chains, including skeletal structures
2. Determine whether a molecule is polar or nonpolar using the molecular shape and bond polarity.
3. Measure the temperature decrease from evaporation of several liquids.
4. Determine a relationship between temperature change and strength of intermolecular forces.
5. Use the relationship discovered between temperature change and the strength of intermolecular forces to predict temperature change of other liquids.

Introduction

Have you ever stepped out of the pool or shower and wrapped up in a towel quickly because you were freezing? It may seem peculiar to feel cold when getting out of the pool, since the outdoor temperature is most often warmer than the water. This feeling is a result of evaporative cooling, a phenomenon that will be explored in this lab.

Evaporation, or vaporization, is a type of phase change that results when liquid molecules have sufficient energy to escape the attractive forces of neighboring liquid molecules to become gas molecules. The energy needed for a molecule to be converted from a liquid to a gas is dependent on the strength of the attractive forces between molecules. Attractive forces between molecules, called intermolecular forces, are based on electrostatic forces. In a previous lab, *Oil and Water Don't Mix – Why?*, you explored hydrogen bonding and ion-dipole forces, which are the strongest intermolecular forces. In that same lab you also discovered how the presence of polar bonds and molecular structure influence the overall properties observed. The same is true for evaporation. The molecular structure influences the rate of evaporation.

Intermolecular forces in molecular compounds

A hydrogen-bonding interaction may be up to 10% of the strength of an ordinary covalent bond! This interaction occurs between polar molecules when one molecule has a protic hydrogen – hydrogen that is bonded to a fluorine, oxygen, or nitrogen atom – and the other molecule has fluorine, oxygen, or nitrogen with a lone electron pair. This is shown in Figure 1 between adenine and thymine, two base pairs of DNA. The partially positively-charged protic hydrogen is strongly attracted to the partially negatively-charged fluorine, oxygen, or nitrogen atom on the neighboring molecule and forms a strong electrostatic interaction between the molecules.

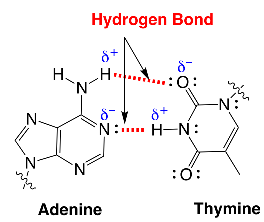


Figure 1. Hydrogen bonding between base pairs of DNA

Polar molecules or molecules with polar moieties, but without protic hydrogen atoms, exhibit dipole-dipole interactions. Thus, the partially positively-charged part of one molecule is attracted to the partially negatively-charged part of the other molecule and vice versa. These dipole-dipole interactions, while not generally as strong as hydrogen bonding interactions, provide stronger attractions between molecules than is seen with similarly sized non-polar molecules.

Permanent dipoles arise from the presence of polar bonds in the molecule. Remember that symmetry plays a role in whether a molecule is polar or non-polar. A molecule can have polar bonds, but if opposing bond dipoles cancel, the overall dipole of the molecule is zero, and therefore the molecule is non-polar. This is why carbon dioxide is nonpolar while water is polar. See Figure 2 for examples.

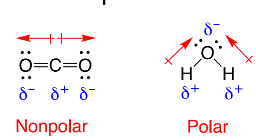


Figure 2. Comparison of the bond dipoles in carbond dioxide and water.

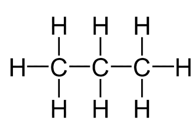
Instantaneous induced-dipole intermolecular forces, also called London dispersion forces, are proposed to account for the attraction of non-polar molecules for one another. These dispersion forces are caused by the intermittent unequal distribution of electrons in an atom or molecule resulting from the random motion of electrons. This random motion results in a <u>temporary</u> dipole in the atom or molecule. All molecules exhibit London dispersion forces. As molecules become larger, their electron clouds become more polarizable, meaning their electrons are more likely to distribute unequally; therefore, they exhibit larger London dispersion forces. While generally considered the weakest of intermolecular forces individually, very large molecules exhibit very large London dispersion forces because of the large number of possible, nearly simultaneous interactions throughout the molecule. Some large non-polar molecules may have larger intermolecular force interactions than smaller polar molecules or even those exhibiting hydrogen bonding interactions. More than one force can be at work in the interactions between molecules. For instance, a polar molecule capable of hydrogen-bonding would have hydrogen-bonding, dipole-dipole forces, and London dispersion forces. Non-polar molecules, those without a permanent dipole, can exhibit only London dispersion forces.

For this lab, you will investigate the Lewis structure of several molecules and the intermolecular forces present between the neighboring molecules in each liquid by noticing the structural features of the molecules and observing the temperature change of the liquid during evaporation. Stronger intermolecular forces will result in less evaporation because more energy will be required to overcome the intermolecular forces. Molecules interacting with neighboring molecules with weak intermolecular forces will more easily vaporize.

The following liquids will be evaluated.

Methanol	CH_3OH
Ethanol	CH_3CH_2OH
1-Propanol	$CH_3CH_2CH_2OH$
1-Pentanol	$CH_3CH_2CH_2CH_2CH_2OH$
n-Hexane	$CH_3CH_2CH_2CH_2CH_2CH_3$
n-Heptane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
n-Octane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Acetone	$CH_3(CO)CH_3$
2-Propanol	$CH_3CH(OH)CH_3$

Different formats can be used for the chemical formulas given. A structural formula shows the sequence of the bonding, for example $CH_3CH_2CH_3$ would have the structure given below. Other notations can show the sequence of functional groups where propane would be represented as CH_3-CH_2-CH_3. Last week you learned about skeletal structure where carbons are represented by the intersection or termini of lines.

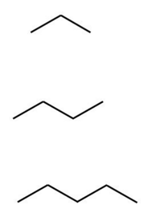


Structural formula

Propane CH_3-CH_2-CH_3

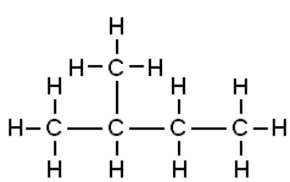
Butane CH_3-CH_2-CH_2-CH_3

Pentane CH_3-CH_2-CH_2-CH_2-CH_3

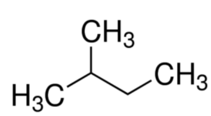


Skeletal structure

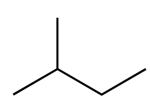
The chemical formulas that contain atoms in parenthesis represent branching groups from the main chain. For example, $CH_3CH(CH_3)CH_2CH_3$ would be represented by any of the formats shown below.



Structural formula



Line-bond structure



Skeletal structure

CH_3-$CH(CH_3)$-CH_2-CH_3

Procedure

For each of the liquids, determine the shape around the carbon atoms (bent, tetrahedral, trigonal planar, etc.), determine the molecular polarity, strongest intermolecular forces present, and the molar mass and fill in the first table on the lab report. This can be done while you are performing the experiments.

The liquids used in this lab are highly flammable and poisonous. Wear appropriate eye protection and gloves. Avoid breathing vapors from the liquids. Keep all liquid containers closed when not in use.

Hexane, heptane, and octane should remain in the fume hoods, so perform the experiment for these liquids there. All other liquids can be assessed at your bench.

Tell your instructor immediately if there are spills or accidents.

Fold a Kimwipe in half 3 times to make a small square. Wrap the folded Kimwipe around the probe end of a yellow 'Traceable ®' thermometer so that the end is not visible. Secure the Kimwipe with tape. Make sure the tape is attaching the Kimwipe to the thermometer but is not completely covering the Kimwipe. Hold the thermometer by the yellow head. Press the "ON/OFF" button and place the wrapped probe in the first liquid solution and soak the Kimwipe for 10 – 30 seconds until the temperature stops fluctuating. When the temperature stops fluctuating, record the temperature as the initial temperature in Table 2. (This initial temperature should not change by more than one degree Celsius throughout the experiment.) Remove the probe from the solution and hold the thermometer by the yellow head so that the covered end extends from the edge of the lab bench. The Kimwipe and probe end of the thermometer should not be touching anyone or any surface. Observe the temperature from the probe until the temperature stabilizes (stops changing), which may take about 2 minutes. After the temperature stabilizes, record the temperature value as the final temperature in Table 2 of your lab notebook. Repeat the procedure with each liquid. Be sure to use a new Kimwipe and dispose of the used Kimwipes as instructed. You will need to wait until the thermometer returns to the room temperature to begin the next liquid!

Data

Complete Table 1 with the list of compounds provided in the introduction.

Table 1. Physical properties of alkanes and alcohols

Liquid	Molar Mass (g/mol)	Polarity (P, NP)	Strongest type of Intermolecular Force	Shape around each carbon	Skeletal structure of molecule

Complete Table 2 by finding the change in temperature during evaporation over the 2-minute time period. To do this, subtract the temperature after evaporation from the initial temperature for each liquid.

Table 2. Initial and final temperatures recorded for the evaporation of a series of alkanes and alcohols.

Liquid Name	**Initial Temperature (°C)**	**Final Temperature (°C)**	$T_{initial} - T_{final}$

Create a graph of the molar mass of methanol, ethanol, 1-propanol, and 1-pentanol vs. the ΔT calculated. The Molar mass should be the x-axis and the ΔT should be the y-axis. Add a second data series to the graph that includes n-hexane, n-heptane, and n-octane vs the ΔT calculated for each. Add a trendline for each data series. There should be two lines on your graph. Display the equation for the line of best-fit for each of the data series on the graph. Display the Legend for your graph that identifies the two data series.

Results

1. Using the best-fit line equations from your graph, calculate the estimated ΔT value for n-pentane and 1-butanol. Show your work in the boxes below. Be sure to use the appropriate equation for each molecule.

n-pentane	1-butanol

2. Based on your calculations, did n-pentane or 1-butanol have a larger ΔT value? Use your understanding of intermolecular forces to explain these results.

3. Why is the ΔT value of 2-propanol different than the ΔT value of 1-propanol?

4. Acetone is neither an alkane or an alcohol. Does the ΔT value of acetone fit better with the alkane data or the alcohol data? Explain your reasoning.

5. Which of the 9 molecules had the strongest intermolecular forces? Use your data to determine your answer. drop

6. As molar mass increases, the ΔT value of a molecule with similar types of intermolecular forces ______________.

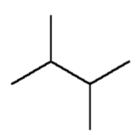
7. The presence of an alcohol group (-OH), __________ the ΔT value of a molecule compared to the presence of a methyl group ($-CH_3$).

8. 2-propanol had a _________ ΔT value compared to 1-propanol because _________________.

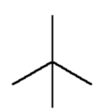
9. Kinetic energy is _____________ related to the temperature of a substance.

10. As liquid molecules are converted to gas molecules, the kinetic energy of the remaining liquid molecules is _________ than before the evaporation.

11. How many carbons are in the following molecule?



12. How many carbons are in the following molecule?



13. How many carbons are in the following molecule?