CHEM105: BIOCHEMISTRY AND SOCIETY
Experiment 3
Intermolecular Forces and Evaporation

(ADAPTED FROM BRAD DELEEW)

BACKGROUND


Part I   Organic Compounds: Alcohols and Alkanes

You will work with two types of organic compounds in this experiment: alcohols and alkanes. The term *organic compound* was once used to mean a compound that came from a living organism. Today, chemists use this term to describe any compound that is composed of carbon (C) and hydrogen (H), and possibly a few other elements such as nitrogen (N), oxygen (O), sulfur (S), phosphorus (P), and so on.
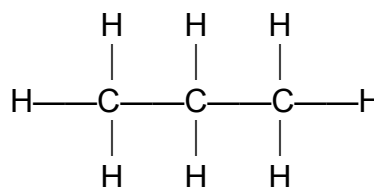
Alkanes are organic compounds composed of just C and H atoms with only single bonds present. Lewis dot structures and ball-and-stick models of the three simplest alkanes are shown below.
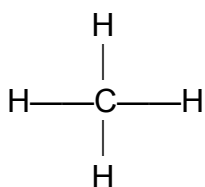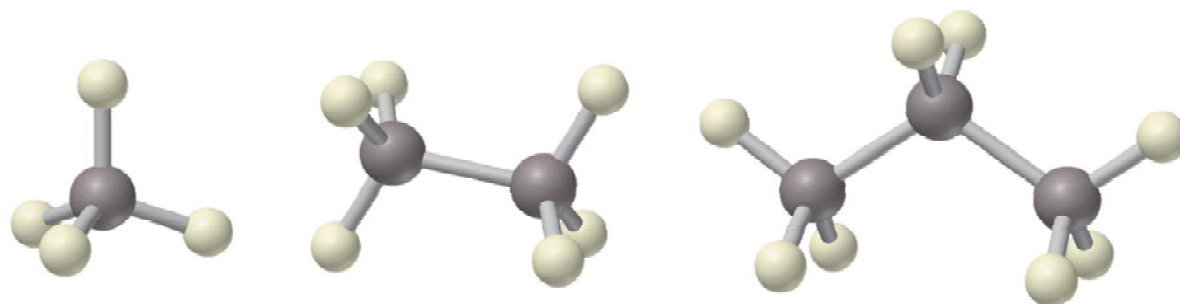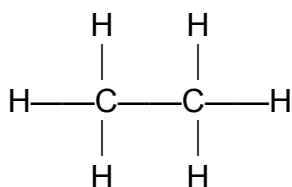
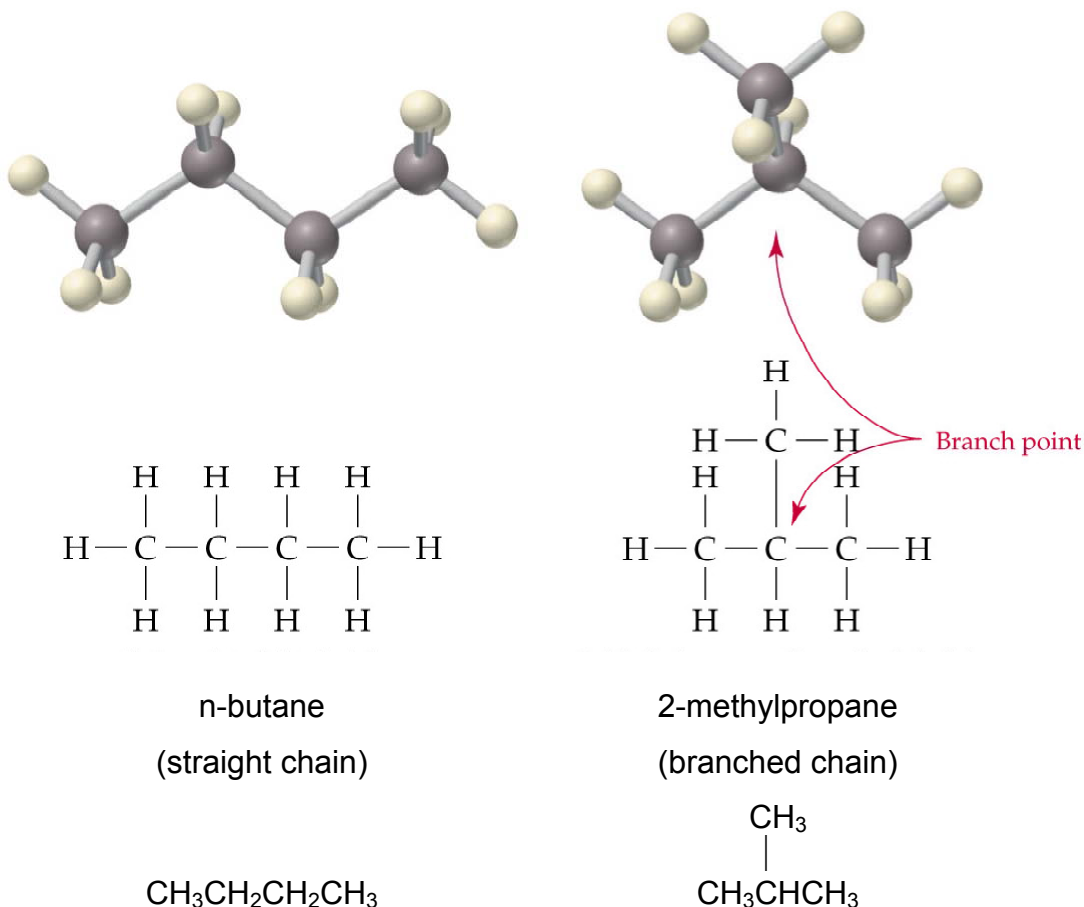|		
Methane	Ethane	Propane
CH_4	CH_3CH_3	$CH_3CH_2CH_3$

The simplest alkane is methane, CH_4. It contains just one carbon atom and is the main component of natural gas. Ethane, C_2H_6, contains two carbon atoms. Propane, C_3H_8, contains three carbon atoms. Notice that alkanes have the general chemical formula $C_nH_{(2n+2)}$.

The formulas given with the Lewis dot structures shown above are examples of **condensed structural formulas** or **condensed structures**.

For alkanes with 4 or more carbon atoms, there is more than one way to connect the carbon atoms together using just single bonds. With 4 carbon atoms (C_4H_{10}), there are two different ways to connect the carbon atoms together:

n-butane

(straight chain)

$CH_3CH_2CH_2CH_3$

2-methylpropane

(branched chain)

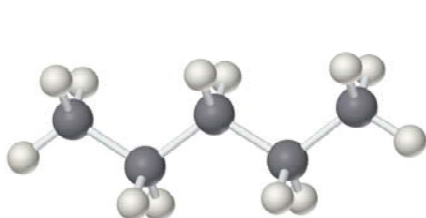$$CH_3CHCH_3 \quad \overset{\displaystyle CH_3}{\big|}$$

Because these two molecules have different structures, they have different physical and chemical properties. In other words, these are two completely different compounds. For example, n-butane boils at −0.5°C, while 2-methylpropane boils at −11.6°C.
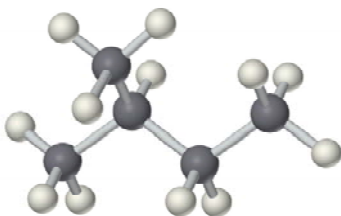
These two compounds are referred to as **structural isomers**. Structural isomers have the same chemical formulas, but their atoms are connected in different ways.

Notice that n-butane has all four carbon atoms connected together in a continuous chain. This is referred to as a **straight-chain** alkane. In 2-methylpropane, however, the longest continuous chain is only 3 carbon atoms long. The fourth carbon atom is connected to the middle (#2) C atom of this 3-carbon chain. This isomer of C_4H_{10} is referred to as a **branched-chain** alkane.
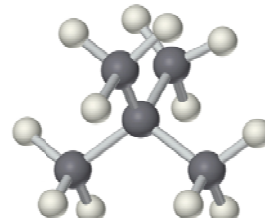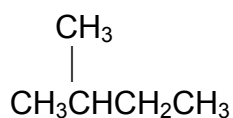
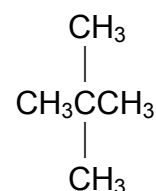With 5 carbon atoms (C_5H_{12}), there are *three* different ways to connect the carbon atoms together:

n-pentane	2-methylbutane	2,2-dimethylpropane
(straight chain)	(branched chain)	(branched chain)

$$CH_3CH_2CH_2CH_2CH_3 \qquad CH_3\overset{\displaystyle CH_3}{\underset{|}{C}}HCH_2CH_3 \qquad CH_3\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}CH_3$$

As the number of C atoms increases, the number of possible structural isomers for alkanes increases dramatically.

In this lab, you will work with 4 different straight-chain alkanes: n-pentane, n-hexane, n-heptane, and n-octane.

Alcohols are alkanes in which one of the hydrogen atoms has been replaced by a hydroxyl group (—O—H). Ball-and-stick models and condensed structures of the two simplest alcohols are shown below.
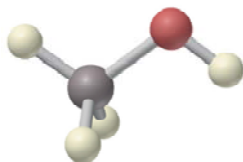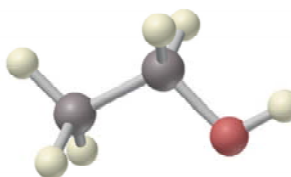
Methanol	Ethanol
CH_3OH	CH_3CH_2OH

For alcohols with 3 or more carbon atoms, there is more than one way to attach the –OH group to the carbon chain. In 1-propanol, the –OH group is attached to the first (#1) C atom of the carbon chain. In 2-propanol, the –OH group is attached to the middle (#2) C atom of the carbon chain.
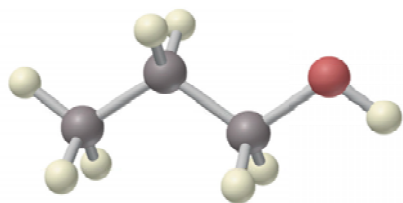
1-propanol
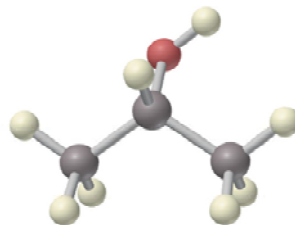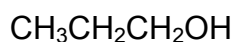
2-propanol
(isopropyl alcohol)

$$OH$$
$$|$$
$CH_3CH_2CH_2OH$ $CH_3CH_2CH_3$
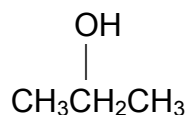
In this lab, you will work with 4 different alcohols: methanol, ethanol, 1-propanol, and 1-butanol.

A ball-and-stick model and the condensed structure of 1-butanol are shown below.

1-butanol

$CH_3CH_2CH_2CH_2OH$

Part II   Evaporation as a Cooling Process

Water (H_2O) molecules are attracted to each other by several different types of intermolecular forces. London (dispersion) forces are present between all molecules. In addition, because water molecules are polar, they are attracted to each other by dipole-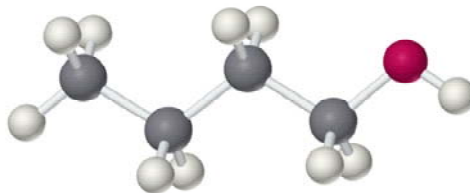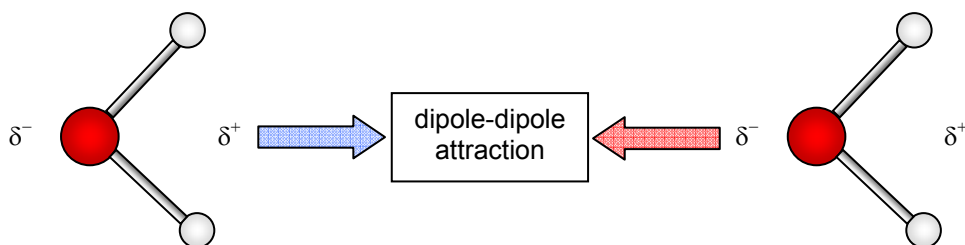dipole attractions when the δ^+ end of one water molecule is pointed towards the δ^- end of a neighboring water molecule.

Finally, when a water molecule approaches a neighboring water molecule oriented in just the right way, a hydrogen bond can form. Covalent O—H bonds are very polar because of the large difference in electronegativities. A hydrogen bond forms between two water molecules when a hydrogen atom in one water molecule is pointing at the lone pair of electrons of an oxygen atom in a neighboring water molecule.
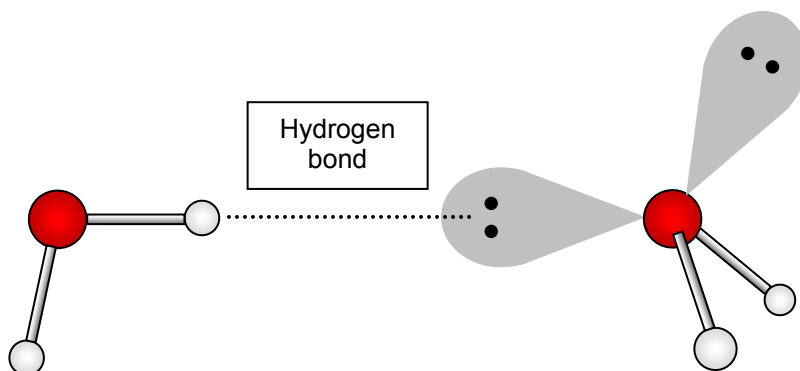
If an open glass of water is left standing on a desk, the liquid level in the glass slowly decreases as water (H_2O) molecules gradually escape to the gas phase. When a liquid changes into a gas at a temperature below its boiling point, that phase change is called ***evaporation***.

Water molecules can evaporate when they are at the boundary between the liquid and gas phases. In other words, evaporation involves the escape of H_2O molecules from the *surface* of liquid water.
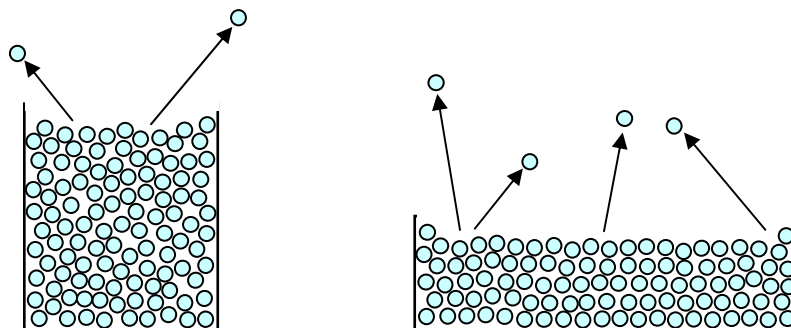
The larger the surface area of the liquid water, the faster it tends to evaporate. Most people have noticed that a puddle of water tends to dry up (evaporate) more quickly if it is spread out over a larger area of the floor.

An H_2O molecule at the surface of the liquid is pulled back towards the liquid by the intermolecular forces acting between water molecules. In order to escape from the liquid phase, an H_2O molecule must be moving fast enough to overcome these intermolecular attractions.

Fast-moving molecules *can* escape.

Slow-moving molecules *cannot* escape.
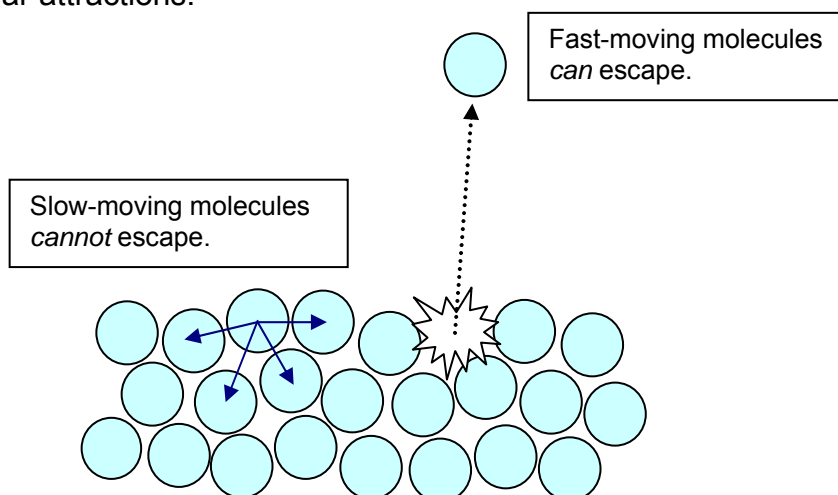
The H_2O molecules in a sample of water are all moving at different speeds. The *temperature* of the water is actually a measurement of the *average speed* of its H_2O molecules. The average speed of H_2O molecules is greater in a glass of hot water than the average speed of H_2O molecules in a glass of cold water. Water evaporates more quickly when it is hot than it does when it is cold. This is because more H_2O molecules in a sample of hot water are moving at speeds fast enough to escape into the gas phase when they are at the surface of the liquid.

Evaporation tends to *lower* the temperature of a sample of water. At the molecular (submicroscopic) level, this cooling effect can be explained by the fact that only the fastest H_2O molecules tend to escape into the gas phase. If a large number

of fast-moving H_2O molecules are removed from a sample of liquid water, the average speed of the remaining molecules in the liquid will decrease.

One way to picture this effect is to imagine a classroom with 31 people in it. Suppose you measure the height of every person and use that information to calculate the average height of people in the room. If the 5 tallest people leave the room, the average height of the remaining 26 people will be less than the average height of the original group of 31 people.

Part III   Evaporation and Intermolecular Attractions

Different compounds have intermolecular forces of different strength. As we have learned in class, the stronger the intermolecular forces in a compound, the higher its boiling point will be. Differences in the strength of intermolecular forces also cause different liquids to evaporate at different rates. The *weaker* the intermolecular forces in a liquid, the *more rapidly* it will tend to evaporate.

In this experiment, you will place a thermometer in a test tube containing a compound that is a liquid at room temperature. A piece of filter paper will be wrapped around the end of thermometer to soak up some of that liquid. When the thermometer is removed from the test tube, the liquid will begin to evaporate from the filter paper, causing a temperature decrease. The more rapidly the liquid evaporates, the larger the observed temperature decrease will be.

PROCEDURE

Work in groups of two in the hoods. Wear goggles and gloves at all times.
Record your data and observations neatly in the worksheets provided.
The 8 different compounds are available in large test tubes in each hood.

Alcohols

Test Tube	*Name*	*Condensed Structure*
1	methanol	CH_3OH
2	ethanol	CH_3CH_2OH
3	1-propanol	$CH_3CH_2CH_2OH$
4	1-butanol	$CH_3CH_2CH_2CH_2OH$

Alkanes

Test Tube	*Name*	*Condensed Structure*
5	n-pentane	$CH_3CH_2CH_2CH_2CH_3$
6	n-hexane	$CH_3CH_2CH_2CH_2CH_2CH_3$
7	n-heptane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
8	n-octane	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

1. Trim a round piece of filter paper with scissors to make a piece that is roughly rectangular.

    Make 2 cuts (along the dashed lines)…
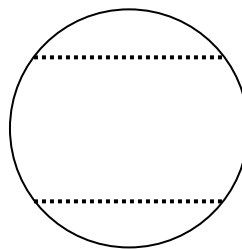
    Do this with a total of 8 round pieces of filter paper. Make sure that you trim all 8 of them down to a fairly consistent final size.

2. Attaching a piece of filter paper to the end of a thermometer correctly is tricky.

    Watch the demonstration by your instructor carefully and make sure that you see how to do this properly.

Thermometer:

Wrap the filter paper into a tight cylinder around the bulb end of a thermometer Hold the middle of a piece of copper wire against the top of the paper cylinder.
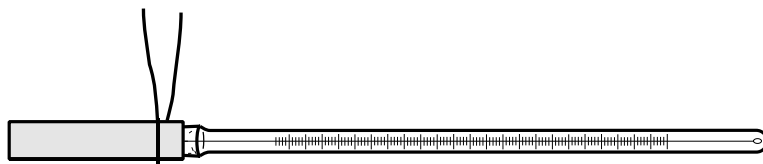
Twist the ends of the wire around each other until you have a tight wire loop holding the paper firmly in place on the thermometer.

If the wire loop isn't tight enough to hold the filter paper in place when you hold the thermometer in a vertical position, you need to twist the wire more so the loop is drawn tighter. The dangling ends of the twisted copper wire can be wrapped around the thermometer and paper.
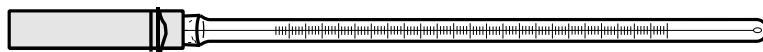
3.   Uncork the first test tube (methanol). Leave the test tube standing upright in the plastic rack. Make sure that the plastic test tube rack is near the front of the hood, just inside the safe work zone marked by the striped yellow tape lines.

Put the thermometer into the test tube and wait about 1 minute until the temperature reading stabilizes.
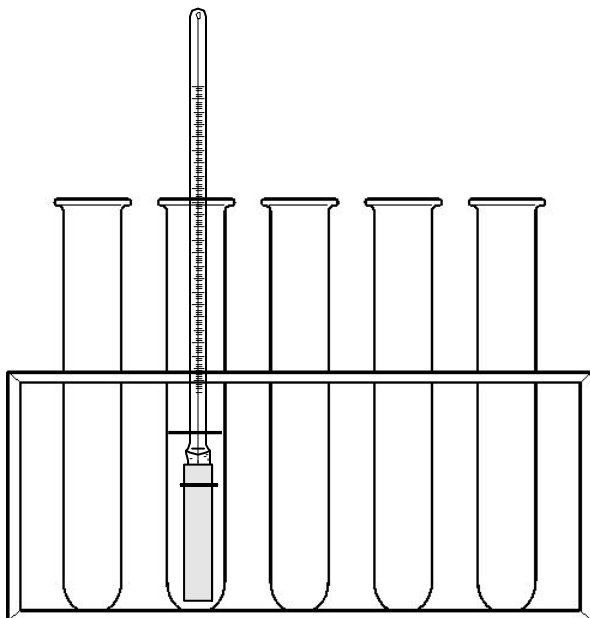
When the thermometer reading is stable, record this initial temperature (T_i).

4. Remove the thermometer from the test tube.

    Be careful. Some of the methanol soaked up by the filter paper may drip.

    Set the thermometer down horizontally across top of the test tube rack and allow the methanol to evaporate from the filter paper.

    Put the cork back into the test tube.

5. As the methanol evaporates, it cools the thermometer and the temperature reading will drop.

    Watch the thermometer reading closely until the temperature reading stops decreasing and *stabilizes* at a minimum temperature.

    Be patient! For liquids that evaporate rapidly, it may take 2−3 minutes for the temperature to stop decreasing. For the liquids that evaporate very slowly, however, it may take 6−7 minutes for the temperature to stop decreasing.

    Record the final temperature (T_f) when the thermometer reading stops decreasing.

6. The temperature drop (cooling) produced by evaporation of the methanol will be calculated as:

$$|\Delta T| \ = \ |\ T_f \ - \ T_i\ |$$

Calculate and record $|\Delta T|$ for methanol.

7. Untwist the copper wire and remove the used filter paper from the thermometer. Wipe the thermometer off with a clean piece of lab tissue.

Used filter paper should be collected in a beaker in your hood.

Do NOT throw used filter paper in the trash.

8. Repeat **Steps 2−7** for each of the other liquids.

Note: Make sure that you work with the liquids in the order described in the lab report worksheets below

Time-Saving Tip:

While one person carefully watches the thermometer being cooled by evaporation in **Step 5**, the other person can work on attaching a fresh piece of filter paper to a second thermometer as described in **Step 2** to get it ready for use with the next liquid.

CHEMICAL INFORMATION SHEET
EXPERIMENT 7: EVAPORATION AND INTERMOLECULAR ATTRACTIONS

methanol ethanol 1-propanol 1-butanol	FLAMMABLE. Do not use near open flame or source of heat. POISONOUS. Do not ingest. Avoid skin contact. Avoid breathing fumes.
n-pentane n-hexane n-heptane n-octane	FLAMMABLE. Do not use near open flame or source of heat. POISONOUS. Do not ingest. Avoid skin contact. Avoid breathing fumes.

Disposal Procedures:	Used filter paper should be collected in a beaker and left in the hood.

EXPERIMENT 7
EVAPORATION AND INTERMOLECULAR ATTRACTIONS

Name: _____

Lab Partner: _____

LABORATORY REPORT

88 Points Total

(34 pts)   ***Sample Data and Results***

Test Tube	Liquid	T_i (°C)	T_f (°C)	$|\Delta T|$ (°C)
1	methanol			
2	ethanol			
3	1-propanol			
4	1-butanol			
5	n-pentane			
6	n-hexane			
7	n-heptane			
8	n-octane			

Predicted $|\Delta T|$ (°C)	Explanation

(12 pts) *Graph Information*

Liquid	Chemical Formula	$|\Delta T|$ (°C)	Total # of e^- per molecule
methanol	CH_4O		
ethanol	C_2H_6O		
1-propanol	C_3H_8O		
1-butanol	$C_4H_{10}O$		
n-pentane	C_5H_{12}		
n-hexane	C_6H_{14}		
n-heptane	C_7H_{16}		
n-octane	C_8H_{18}		

(20 pts) *Graph*

Make a single graph of the $|\Delta T|$ values for the four alcohols and the four alkanes plotted versus their respective total number of electrons per molecule. In other words, plot the total number of electrons per molecule along the horizontal axis and $|\Delta T|$ along the vertical axis. Your graph should have a descriptive title and the axes should be clearly labeled, including units (if any). You should use two different symbols for data points, one symbol for the alcohols and the other symbol for the alkanes. Label each data point with the name of the corresponding compound. Draw line segments connecting the alcohol data points to show the trend, and do the same for the alkane data points.

Questions

(3 pts)　**1.**　Describe the relationship between the size of $|\Delta T|$ observed for a liquid and the strength of the intermolecular attractions in that liquid. [In other words, does a larger $|\Delta T|$ value indicate stronger or weaker intermolecular attractions?] Explain briefly.
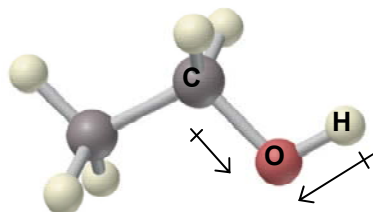
(2 pts)　**2.**　Are alkanes polar molecules? Explain briefly.

(2 pts)　**3.**　What type(s) of intermolecular attractions are present in the alkanes?

(2 pts)　**4.**　Which type of intermolecular attractions is changing in strength as the alkanes grow larger? Explain briefly.

(2 pts)   **5.**   Based upon your observations in lab, which of the alkanes studied had the strongest intermolecular attractions? Which one had the weakest intermolecular attractions?
Explain briefly.

(2 pts)   **6.**   Are the alcohols polar molecules? Explain briefly using the ball-and-stick model of ethanol shown below.

(2 pts)   **7.**   What type(s) of intermolecular attractions are present in the alcohols?

(2 pts) **8.** Which type of intermolecular attractions is changing in strength as the alcohols grow larger? Explain briefly.

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(2 pts) **9.** Based upon your observations in lab, which of the alcohols studied had the strongest intermolecular attractions? Which one had the weakest intermolecular attractions?
Explain briefly.

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(3 pts) **10.** Two of the liquids studied, n-pentane and 1-butanol, have identical numbers of electrons per molecule. However, these two liquids have significantly different $|\Delta T|$ values. Briefly explain the difference in the $|\Delta T|$ values observed for these two liquids by comparing the type(s) of intermolecular forces present in each of these liquids.

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