Experiment 1
Intermolecular Forces (IMF’s)

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Name:

Date:

Key Objectives

1. What is responsible for Intermolecular Forces (IMF’s)
2. Relationship between Intermolecular Forces and physical properties (specifically Vapor Pressure)
3. Assign Intermolecular forces to molecules
4. Make predictions of physical properties (boiling point, melting point and vapor pressure)

Discussion

Review: Intramolecular Forces (Chemical Bonds) vs Intermolecular Forces

Chemical bonds are the attractive forces holding atoms together to form molecules. This include forming ionic compounds (ionic bonds - gain and lose electrons to complete octets) or molecular compounds (covalent bonds - share electrons to complete octets). Chemical bonds are responsible for chemical properties, mainly the chemical reactions we study. Chemical bonds are much stronger than intermolecular forces.When drawing molecules we show bonds as solid lines.

Intermolecular forces (IMF’s) are the attractive forces between molecules. They are responsible for many of the physical properties of substances. Physical properties include boiling points, melting points, and solubility. When drawng molecules we show IMF’s as dashed lines. There are four important IMF we will consider in this experiment.

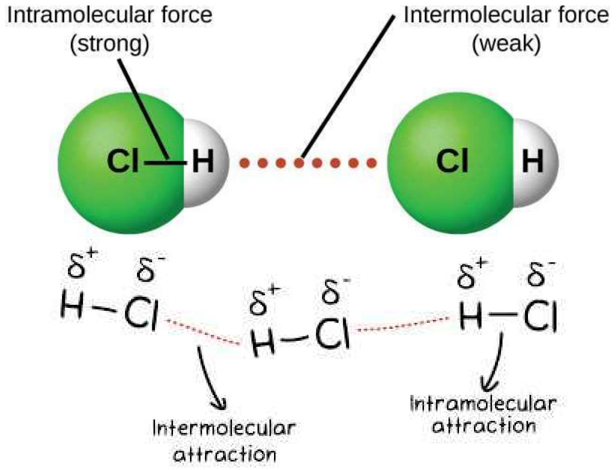


Figure 1.1: Intramolecular (Chemical Bonds) vs Intermolecular (Between molecules) credit: top - OpenStaxCollege. License: CC BY 4.0 bottom - `https://cdn.kastatic.org/ka-perseus-images/2596bfb5d1f501a78d6693d0a9bd79c17631ffb0.png`

Types of Intermolecular Forces (IMF)

For a detailed discussion of Intermolecular forces review your notes from last semester, look at the linked video on (www.chemhaven.org/che102) or borrow a textbook (Tro Section 11.2).

London Dispersion Forces (LDF)

London Dispersion Forces (LDF) occur between non-polar molecules due to the formation of instantaneous dipoles due to the temporary asymmetric distribution of electrons around an atom. The dipoles are temporary (very) lasting a fraction of second, therefore the attraction between molecules is very weak. They are the weakest IMF but are proportional to the size (or molecular weight) of the molecule so in large molecules can outweigh the effects of stronger IMF’s in smaller molecules. Non-polar molecules are often symmetrical.

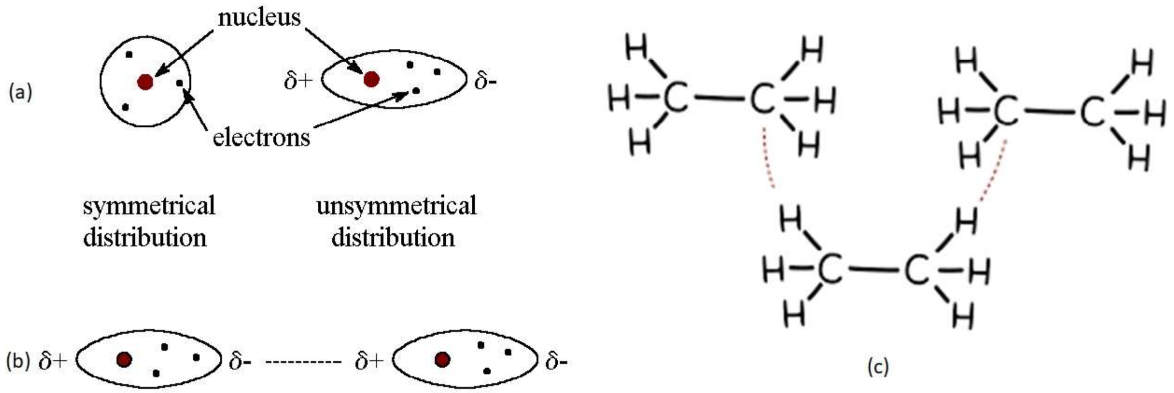


Figure 1.2: (a) Top - Symmetric and asymmetric distribution of electrons (b) Bottom - Temporary dipoles can be attracted to each other and are called London Dispersion Forces (LDF). (c) Right - For larger molecules that are symmetrical the attraction is generally shown as dotted lines between the molecules. credit: (a) and (b) - `https://www.chem.purdue.edu/gchelp/liquids/disperse.html` (c) `https://www.khanacademy.org/science/class-11-chemistry-india/xfbb6cb8fc2bd00c8:in-in-states-of-matter/xfbb6cb8fc2bd00c8:in-in-intermolecular-forces/a/intramolecular-and-intermolecular-forces`

Dipole-Dipole (DD)

Dipole-Dipole forces occur between polar molecules due to the asymmetric distribution of charge producing a partial charge (δ^{+}/δ^{-}), and electrostatic attraction between the molecules. Molecules that are not symmetrical often form DD interactions. The attractive force is stronger than LDF, but weaker than Hydrogen Bonds.

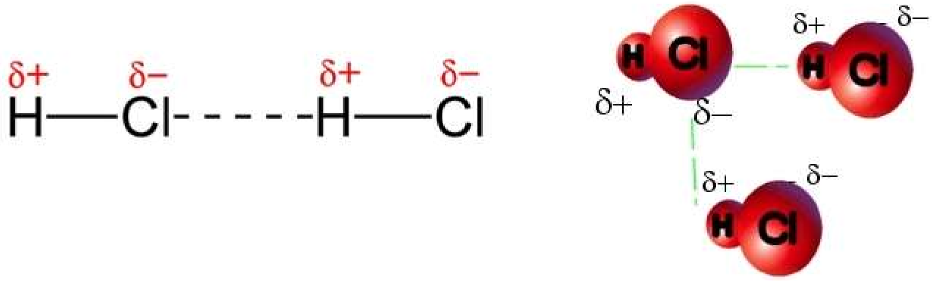


Figure 1.3: Dipole-Dipole (DD) interaction between two hydrochloric acid molecules. credit: left - `https://courses.lumenlearning.com/introchem/chapter/dipole-dipole-force/` right - `https://wyomingapchemistry.weebly.com/101-intermolecular-forces.html`

Hydrogen Bonds (HB)

A special case of Dipole-Dipole attraction in which a Hydrogen bonded to Oxygen, Nitrogen or Fluorine is attracted to the lone pair electrons on or a partial negative charge on a molecule. The asymmetric distribution of charge when H is bonded to O,N, or F is substantially larger than a normal dipole due to Hydrogen only having one electron, resulting in an attractive force much larger than the typical DD interaction.

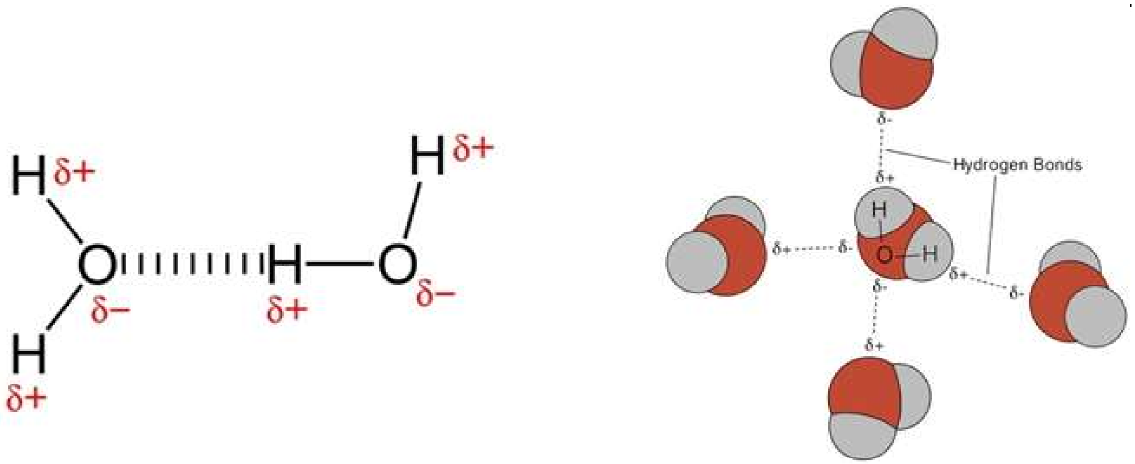


Figure 1.4: Hydrogen Bond (HB) between two water molecules. credit: `https://courses.lumenlearning.com/cheminter/chapter/hydrogen-bonding/`

Ion-Dipole(ID)

The electrostatic attraction between a molecule with a partial charge and an ion with a full charge. Not relevant in this experiment as we are dealing with pure solutions.

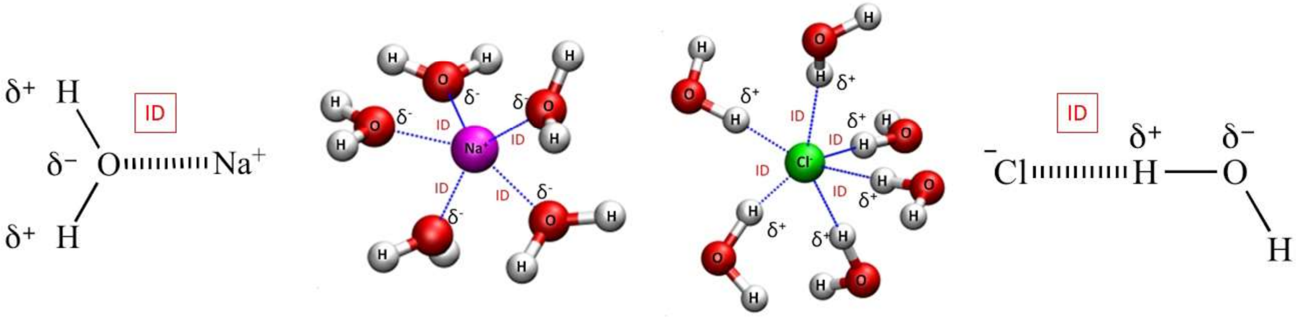


Figure 1.5: Ion Dipole interaction (ID) a cation (Na^+) and an anion (Cl^-) and water molecules. credit: author

Ionic Bonds (IB)

This is **NOT** an IMF, but the attraction between full charges responsible for ionic (chemical) bonds. This represents the strongest attraction one can have between charges and is considerably stronger than the Intermolecular Forces studied in this lab. Recall that ionic bonds are generally shown with either lewis structures or as models.

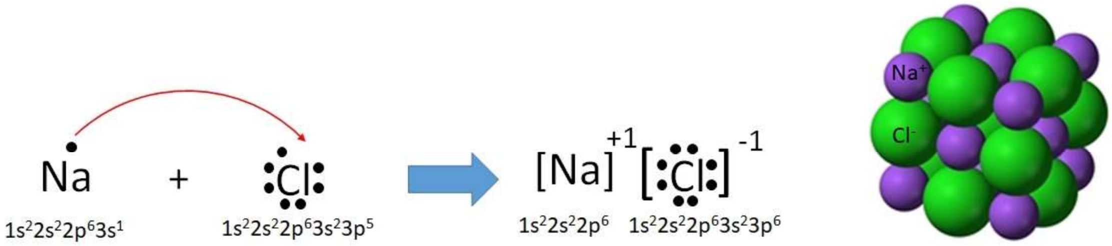


Figure 1.6: An Ionic Bond (IB) is the attractive force between full charges and is considered a chemical bond. credit: author

Physical Properties

Intermolecular Forces are responsible for many of the physical properties of substances as they are the primary interaction between two molecules. A detailed discussion can be found in your textbook (OER 10.2-10.3)

Evaporation is the process (or phase transition) by which a molecule in the liquid state becomes a molecule in the gas state. Molecules in the liquid phase are held together by the IMF between the molecules, as the temperature increases eventually the IMF are broken and we transition from the liquid phase to gas phase.

The process of evaporation is endothermic, (ΔH is negative), and the substance absorbs energy from the surroundings (in this experiment a temperature probe) to have enough energy to go from liquid to gas. One can imagine the flow of energy is from the surroundings (temperature probe) to the system (filter paper + solution) to the molecules in the solution causing them to evaporate.

In this experiment you will explore the relationship between the IMF present in the molecules and the change in temperature experienced as a substance evaporates.

Accuracy and Precision - a Brief Refresher

A more detailed discussion of Accuracy and Precision can be found in the first experiment you performed (last semester!) Refer to it if needed.

To determine if you have good/reproducible data one needs to look at the accuracy and precision of the measurement. To determine the accuracy one compares the measured value to the theoretical or known value. Due to the nature of the experiment (and the large number of controllable variables in taking the measurements, we are concerned only with the differences in magnitude of the measured temperature changes (ΔT). Therefore, we will not know how the measured values correlate to the accepted values.

To determine precision one takes multiple measurements and compares how close to each other they are. In this experiment we will be able to determine the precision of our measurements. Precision is often given by the Average and Standard Deviation of the measurements as detailed below.

Average and Sample Standard Deviation

After collecting your data the first step to determining the precision is to calculate the average ($\bar{x}$). The average is given below in two forms, on the left, is the formal mathematical expression, on the right is the more standard form.

$$\bar{x} = \frac{1}{n} \times \sum_{i=1}^{n} x_i$$

$\bar{x} = \text{Average}$

$i = \text{individual measurement}$

$n = \text{number of measurements}$

$$\text{Average} = \frac{x_1 + x_2 + x_3 ....}{n}$$

Figure 1.7: (a) Formal Mathematical Definition of an Average. (b) Standard Definition of Average.

The sample standard deviation (s) is used to measure how close each value is to the average value, and quantifies the amount of variation in set of data values. A low sample standard deviation indicates that the data points tend to be very close to the average value, while a high sample standard deviation indicates that the data points are far from the average value. Thus, a small sample standard deviation means that the experimental result is reproducible, while a large sample standard deviation means the experimental results are not very reproducible. The equation for the sample standard deviation is given below:

$$s = \left(\frac{\sum_{i=1}^{n} (x_i - \bar{x})^2}{n-1} \right)^{\frac{1}{2}}$$

Figure 1.8: Standard Deviation

The following example shows the calculation of the average and standard deviation for an example set of data. A student measures the change in temperature of a sample 5 times with the following results:

Trial	**ΔT (°C)**
1	25.0
2	24.0
3	23.0
4	23.5
5	26.0

$$\bar{x} = \frac{25.0 + 24.0 + 23.0 + 23.5 + 26.0}{5} = \frac{121}{5} = 24.3$$

$$\sigma = \left(\frac{(24.3-25.0)^2 + (24.3-24.0)^2 + (24.3-23.0)^2 + (24.3-23.5)^2 + (24.3-26.0)^2}{4} \right)^{\frac{1}{2}}$$

$$= 1.204$$

Figure 1.9: Example Calculation for Average and Standard Deviation

Based on the example above, one would report the measured value as 24.3 ± 1.2 °C.

Procedure

In this experiment you will be measuring the temperature change caused by the evaporation of different liquids from a Vernier Temperature Probe over a fixed time interval of 4 minutes. You will be collecting several different sets of data. In between each set of data you will answer several questions.

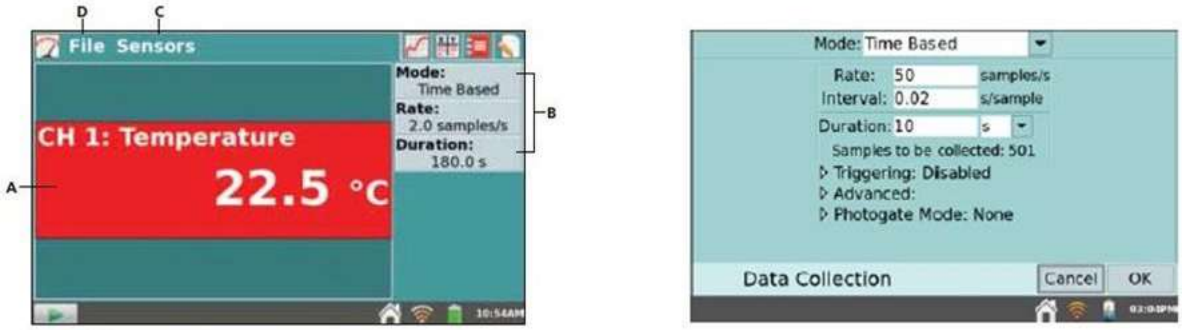
1. You will measure the evaporation rate of one compound 3 times in order to make an estimate of the precision of your measurements. (Data Table 1)
2. You will then measure 1-propanol and make several predictions (Data Table 2).
3. Last you will make several predictions, and then make the measurement (Data Table 3).

At the end of the experiment you will be able to graph the data to show some general trends.

Data Collection - Setting up the Vernier Temperature Probe

1. Plug in the Vernier Temperature Probe into the LabQuest II.
2. From the main menu click on the "Speedometer" symbol > Sensors > Data Collection. See Figure 1.10.
3. Make sure the LabQuest II is taking data every 5 seconds for a duration of 10 minutes total (you

will not need to collect data for this long, but it is best to be safe). Note, **DO NOT** used the settings pictured in the figure below, change them as instructed).



CH 1: Temperature
22.5 °C
Mode: Time Based
Rate: 2.0 samples/s
Duration: 180.0 s

Mode: Time Based
Rate: 50 samples/s
Interval: 0.02 s/sample
Duration: 10 s
Samples to be collected: 501
Triggering: Disabled
Advanced:
Photogate Mode: None

Figure 1.10: LabQuest II Setup. credit: `http://www2.vernier.com/manuals/labquest2_user_manual.pdf`

General Experimental Procedure

1. Many of the compounds used in the laboratory are flammable and poisonous. Avoid physical contact with all solution. Be sure there are no open flames in the laboratory during the experiment. Dispose of all mixtures in the appropriately labeled container in the hood.

2. Obtain a piece of filter paper, scissors, pencil, and ruler and cut several (1/trial) pieces of filter paper 3.0 cm x 3.0 cm.

3. Wrap the temperature probe with the square piece of filter paper and secure it in place using the small rubber bands or dental floss. The paper should be even with the probe end.

4. Obtain a large test tube and fill it with the solution to be tested to a depth of 4 cm. Be sure to fill all test tubes are filled to the same depth.

5. Place the temperature probe in the solution to be tested for at least 1 minute to reach equilibrium as shown in Figure **??**.

6. Ready a piece of tape (or a book) to hold the temperature probe in place as shown in Figure **??**. Optionally you may use a ring stand and clamp to hold the temperature probe.

7. Make an initial temperature measurement lasting between 15-30 seconds long with the probe still in the liquid. At the end of the measurement remove the probe from the liquid and suspend it as shown in Figure **??**, making sure the end extends beyond the table by 2-3 inches. Continue collecting data for at least 4 minutes **OR** until the temperature has stopped dropping and has started to rise. If the temperature has started to increase after 4 minutes, make a note of it on your data table.

8. Record the initial and final temperatures and calculate the change in temperature for the sample.

9. Remove the filter paper and dispose of it in the provided solids container. Empty the test tube in the "Volatile Waste" liquids container.

10. Repeat the experiment for the next listed substance. Use a new piece of filter paper for each new liquid tested.

Dispose of all mixtures in the appropriately labeled container in the hood.

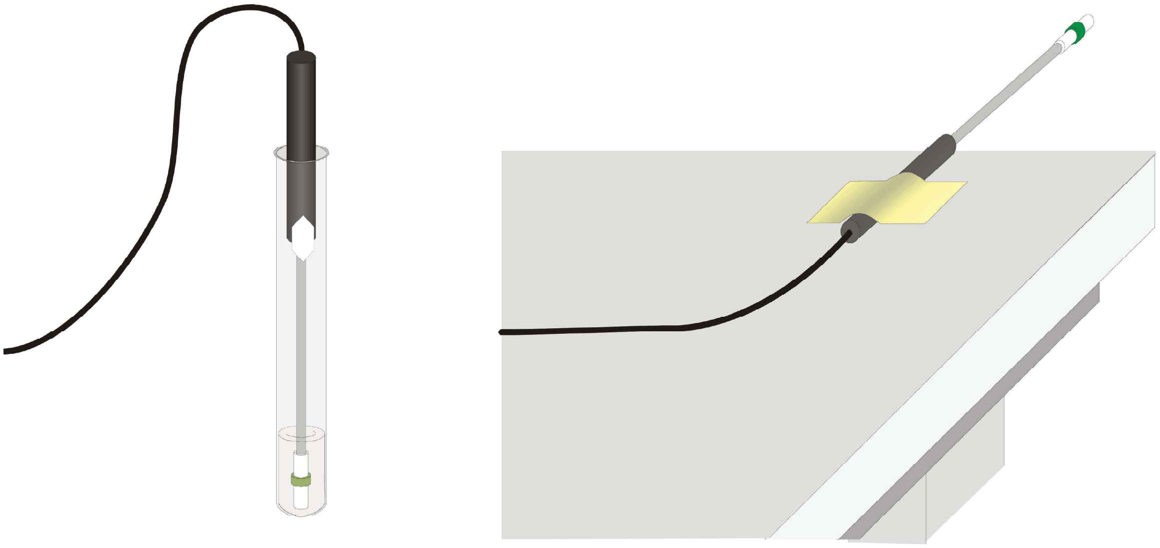


Figure 1.11: Left - Temperature probe wrapped in filter paper soaking in the solution to be tested. Right - Suspending the temperature probe to allow evaporation to occur. credit: `https://www.vernier.com/experiment/cwv-9_evaporation-and-intermolecular-attractions/`

Name: ____________________ Class: ________________ Date: ________________

Results

Data Table 1

Trial	**Initial Temp (℃)**	**Final Temp (℃)**	**Measured ∆T (℃)**
Ethanol - 1			
Ethanol - 2			
Ethanol - 3			
		Average:	
		Standard Deviation:	

1. For the trials on Ethanol, do all of your measurement fall within the calculated Standard Deviation? Do you feel confident in your data or do you feel you should take more data? Explain.

2. Based on your experience explain what types of error’s or bias’s that might occur in this experiment? What effect are would they have on the experiment? What you did to minimize those errors. (Discuss at least two errors.)

Data Table 2

Substance	**Initial Temp (℃)**	**Final Temp (℃)**	**Measured ΔT (℃)**
1-propanol			

3. Based on your measurement of Ethanol and 1-Propanol what is the relationship between the Intermolecular Forces (IMF's) and the measured change in temperature (ΔT? Are they directly proportional or inversely proportional?

Data Table 3

When explaining your predicted ΔT value, make sure you reference your previous observations in lab and what you know about IMF's. (I am not looking to see if your predicted value is correct, only that you have the right trend (ie higher or lower than some other measured value and a good reason why.) If you need more room to write attach a separate sheet of paper. **Show this to your instructor before making the next four measurements**

Substance	**Predicted ΔT (℃)**	**Explanation**	**Initial Temp (℃)**	**Final Temp (℃)**	**Measured ΔT (℃)**
1-butanol					
n-pentane					
methanol					
n-hexane					

Post Lab Questions

4. Attach a graph showing the effect of molar mass on the measured change in temperature for methanol, ethanol, 1-propanol, and 1-butanol. (Plot the molar mass on the x-axis and change in temperature on the y-axis). Based on your graph what trend did you observe (Directly Proportional, Inversely Proportional, No Trend). (Answer using a mathematical expression and using words).

5. Explain the trend observed in terms of the Intermolecular Forces (IMF) present between the molecules. (The change in temperature was (Directly Proportional, Inversely Proportional, No Trend) the strength of the IMF.

6. From your graph, make a prediction of the value (ie a number!) for the ΔT of 1-hexanol. Explain.

7. n-pentane (72 g/mol) and 1-butanol (74 g/mol) have nearly identical molecular weights but should have had very different ΔT values. Explain the difference in terms of the Intermolecular Forces (IMF) present between the molecules.

8. Which IMF's present between two methanol molecules? Draw a picture illustrating the interaction between two methanol molecules

9. Based on your observations in lab (and general knowledge of chemistry) IMF's are:
   (a) Directly proportional to the change in temperature or
   (b) Inversely proportional to the change in temperature?
   Explain.

10. Based on your observations in lab which molecule will have the highest Boiling Point? Explain.

11. For the alcohols we compared methanol, ethanol, 1-propanol, and 1-butanol which allowed us to examine trends based on molecular weight. Why didn't we do the same for the alkanes (methane, ethane, propane, butane, pentane and hexane)?

12. Based on your observations which predictions were correct and incorrect. For the incorrect ones explain why your prediction was wrong.

12(a) 1 - butanol

12(b) n-pentane

12(c) methanol

12(d) n-hexane

.

Name: ____________________________ Class: _____________________ Date: _____________________

Prelab Questions

1. For each of the IMF and bond types discussed draw an example of the interaction between two molecules. **DO NOT** use the examples given in the lab, lecture, or book, come up with your own examples!

   London Dispersion Forces (LDF)

   Dipole-Dipole (DD)

   Hydrogen Bond (HB)

   Ion-Dipole (ID)

   Molecular Bond

   Ionic Bond

2. For the data below calculate the average and standard deviation. Show work in the space provided.
   Trial 1: 15.5 ℃
   Trial 2: 16.0 ℃
   Trial 3: 14.4 ℃
   Trial 4: 15.2 ℃

   Average:
   Standard Deviation:

3. Complete the following table. For the Intermolecular Forces (IMF) present use the abbreviations given in lab to save space.

Compound	**Lewis Structure**	**Molecular Weight**	**IMF's Present**
ethanol CH_3CH_2OH			
1-propanol $CH_3CH_2CH_2OH$			
1-butanol $CH_3CH_2CH_2CH_2OH$			
n-pentane $CH_3CH_2CH_2CH_2CH_3$			
methanol CH_3OH			
n-hexane $CH_3CH_2CH_2CH_2CH_2CH_3$			