Experiment 21
Intermolecular Forces and Boiling Points

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Name:

Date:

Key Objectives

1. What is responsible for each of the four Intermolecular Forces (IMF's)
2. Assign Intermolecular Forces to organic molecules.
3. Understand the relationship between Intermolecular Forces and physical properties (specifically boiling points in this lab)
4. Review of Accuracy and Precision (Lab 1) and Graphing Lab from last semester.

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 drawing molecules we show IMF's as dashed lines. There are four important IMF we will consider in this experiment.

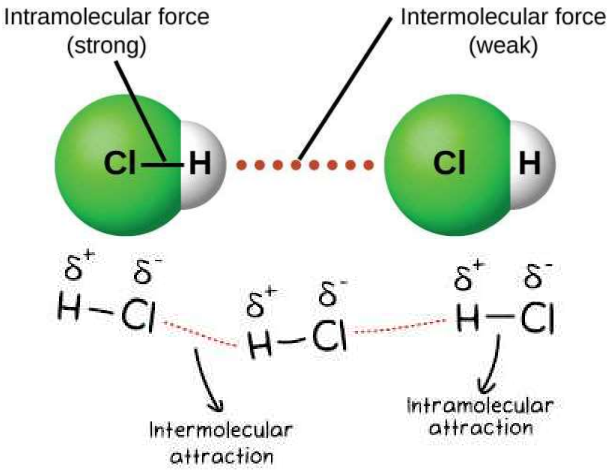

Figure 21.1: Intramolecular (Chemical Bonds) vs Intermolecular (Between molecules) credit: top - OpenStaxCollege. License: CC BY 4.0 bottom - `https://cdn.kastatic.org/ka-perseus-images/2596bfb5d1f501a78d6693d0a9bd79c17631ffb0.png`

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. 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.

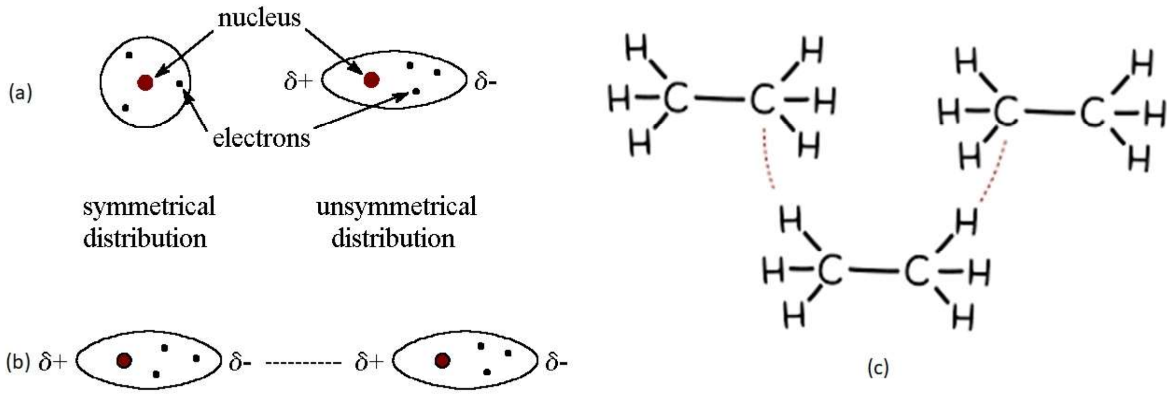

Figure 21.2: (a) Top - Symmetric and asymmetric distribution of electrons (b) Bottom - Temporary dipoles can be attracted to each other and are called London Dispersion Forces (LDF). (c) Right - For larger molecules that are symmetrical the attraction is generally shown as dotted lines between the molecules. 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.

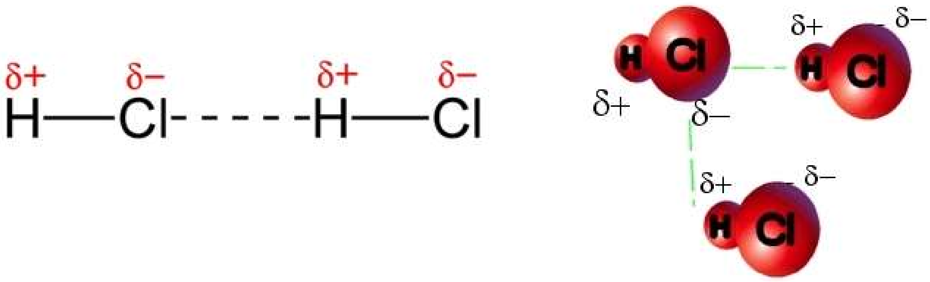

Figure 21.3: Dipole-Dipole (DD) interaction between two hydrochloric acid molecules. credit: left - `https://courses.lumenlearning.com/introchem/chapter/dipole-dipole-force/` right - `https://wyomingapchemistry.weebly.com/101-intermolecular-forces.html`

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.

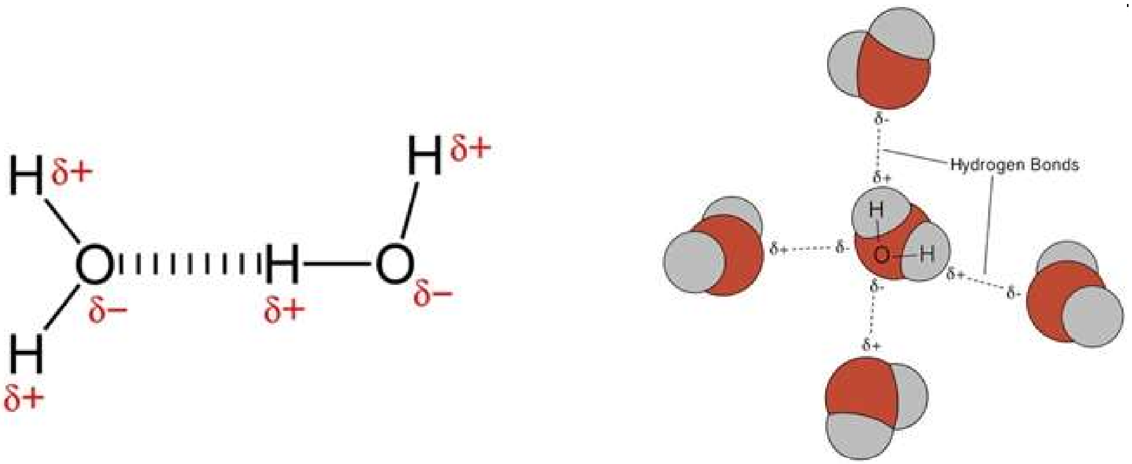

Figure 21.4: Hydrogen Bond (HB) between two water molecules. credit: `https://courses.lumenlearning.com/cheminter/chapter/hydrogen-bonding/`

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.

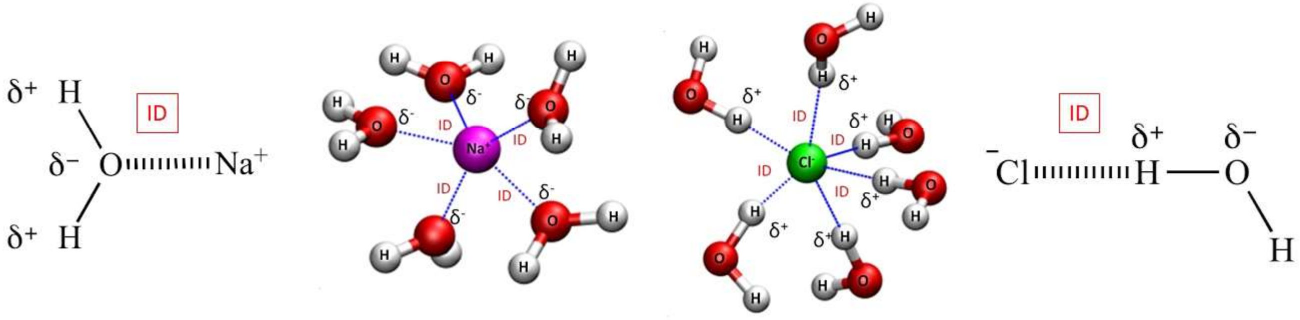

Figure 21.5: Ion Dipole interaction (ID) a cation (Na^+) and an anion (Cl^-) and water molecules. credit: author

IMF's and Boiling Points

IMF's are responsible for the attractive forces between molecules. They are important because they determine many of the physical properties we will explore for each functional group in organic chemistry. In this lab we will focus on Boiling Points and in the next lab look at Solubility.

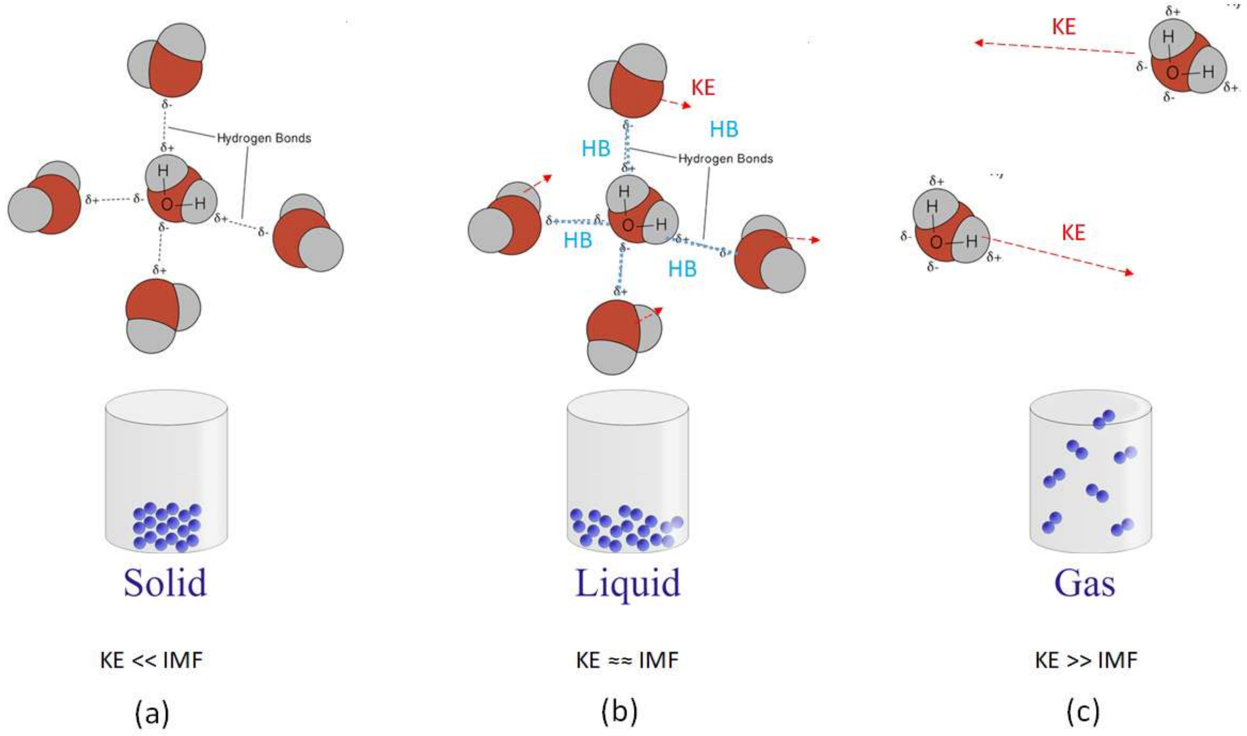

Figure 21.6: Going from a solid to liquid to gas is a study of the IMF holding the molecules together. credit: author

The boiling point of a compound is the temperature at which it changes from a liquid to a gas. This is a physical property related to the IMF's between molecules. As show in Figure 21.6 (a) the Kinetic Energy of the molecules is much less that the strength of the Intermolecular Forces holding the molecules together resulting in a solid state with the molecules held rigidly in place. Figure 21.6 (b) the Kinetic Energy of the molecules is about the same as strength of the Intermolecular Forces holding the molecules together resulting in a liquid in which the molecules are held near each other but are free to move. In Figure 21.6 (c) the Kinetic Energy of the molecules is much larger than the Intermolecular Forces holding them together resulting in a gas, the molecules are far apart, move rapidly and do not interact. The transition from the liquid to solid state occurs at the boiling point and is thus related to the Intermolecular Forces holding the molecules together, the stronger the Intermolecular forces between molecules the more energy (or higher temperature) is required to overcome the attractive forces and separate the molecules far enough apart to be in the gas state.

The molecules of compounds that exist in the liquid state are relatively close together, compared with molecules of gaseous compounds. The close proximity of molecules in the liquid state allow these molecules to interact via IMF's. In general, these interactions are favorable and help to hold the molecules together in a defined volume, but still allow free motion or "flow". Conversely, molecules of a gaseous compound are much farther away from each other and are not confined to a specific volume and thus there is no interaction between the molecules. If enough energy (often in the form of heat) is provided to the liquid, the molecules begin to move away from each other by "breaking" the non-covalent forces that hold the compound in the liquid state. Boiling Points are thus a very good measure of the IMF's between molecules.

The Boiling point of a liquid is defined as the temperature when its vapor pressure is equal to the atmospheric pressure. Normally, the boiling point is measured at one atmosphere (101 kPa or 760 mmHg or 760 torr), however, we live in Rangely CO and will have to determine them at around 0.85 atmosphere. This will result in our values being slightly below the accepted literature values.

Effect of Pressure on Boiling Points

As you should recall from Chapter 13.5 the Boiling Point depends on the atmospheric pressure and thus on the elevation at which it is measured. You will need to correct your experimental data using the following two equations.

$$\Delta T_b = \frac{P_{Rangely} - P_{Sea-Level}}{} \times \frac{0.5\,^{\circ}\mathrm{C}}{10\,\mathrm{mmHg}} \tag{1}$$

$$T_b(\mathrm{Sea-Level}) = T_b(\mathrm{Measured}) - \Delta T_b \tag{2}$$

The first equation determines how much lower your measured pressure is than the pressure at sea-level. The second equation corrects your measured value at elevation to the measured value at sea level.

For example, if you measured the Boiling Point of water in Rangely as 94.0 °C at a pressure of 650.0 mmHg then the equations would look like:

$$\Delta T_b = \frac{650\ \mathrm{mmHg} - 760\ \mathrm{mmHg}}{} \times \frac{0.5\,^\circ\mathrm{C}}{10\,\mathrm{mmHg}} = -5.5\,^\circ\mathrm{C}$$

$$T_b(\mathrm{Sea-Level}) = 94.0\,^\circ\mathrm{C}\ -\ -5.5\,^\circ\mathrm{C}\ = 99.5\,^\circ\mathrm{C}$$

Not quite the expected value of 100.0 °C but well within the margin of error for the measurement.

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 precision the 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.

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

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

$\bar{x}$ = Average

i = individual measurement

n = number of measurements

Figure 21.7: (a) Standard Definition of Average. (b) Formal Mathematical Definition of an 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 21.8: 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 (℃)**
1	25.0
2	24.0
3	23.0
4	23.5
5	26.0

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

$$\sigma = \left(\frac{(24.3-25.0)^2 + (24.3-24.0)^2 + (24.3-23.0)^2 + (24.3-23.5)^2 + (24.3-26.0)^2}{5-1} \right)^{\frac{1}{2}} = 1.204$$

Figure 21.9: Example Calculation for Average and Standard Deviation

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

Percent Error

The percent error is a measure of how close to the correct value your measurements are. It is given by the following equation:

$$\text{PercentError} = \left| \frac{\text{Experimental} - \text{Theoretical}}{\text{Theoretical}} \right| \times 100$$

For the example above if the correct (theoretical) value was 25.0 then the percent error would be calculated as:

$$\text{PercentError} = \left| \frac{24.3 - 25.0}{25.0} \right| \times 100 = 2.8\%$$

Graphing Review

Review how to make good graphs by examining your Labs from last semester. A quick list of things to include in a graph are:

1. Graph Title - meaningful title
2. Label x-axis + Units
3. Label y-axis + Units
4. Major and Minor Tick marks
5. NO gridlines
6. Legend (inside graph)
7. Curve-Fit/Line-Fit (y=mx+b)
8. Goodness of fit (R^2)
9. Graph takes up all available space
10. No wasted space
11. Name and Date
12. Check spelling and CaPitaLIZatioN

Procedure

In this experiment you will be measuring the Boiling Point of several organic substances to see the effects of different Functional Groups (IMF's) and the effect of Size on the Boiling Point.

1. You will measure the Boiling Point of one compound 3 times in order to make an estimate of the precision of your measurements. This will include calculating your average and standard deviation.

2. You will then collect data from the class to look at the effect of molecular weight (molecular size) on the Boiling Point of compounds.

3. Last you will take the data provided by your instructor and look at the effect of different IMF's on the Boiling Point. (We will skip this part for now and maybe look at it after we have explored all the functional groups!)

Group Number	Assigned Liquid
Group 1	Pentane
Group 2	Hexane
Group 3	Heptane
Group 4	Methanol
Group 5	Ethanol
Group 6	1-Propanol

Figure 21.10: Assigned Liquids for Measuring the Boiling Point

Part 1: Measuring the Boiling Point of a Liquid (Done in Class)

One student can assemble the Ring Stand, Bunsen Burner and Clamps as shown in Figure 21.11(a), while the other student sets up the test tube, sample and thermometer as shown on Figure 21.11(b).

Setting up the Ring Stand

1. Obtain a ring stand, a test tube clamp, a thermometer clamp, a Bunsen burner, several capillary and test tube.

2. Clamp the test tube high enough that the Bunsen burner can be waved below the apparatus so you can control the rate of heating.

Setting up the Test Tube and Sample

1. Add the liquid to be tested to a small test tube to a depth of approximately 2-3 cm. It should be enough liquid to cover the bulb (full of mercury) at the bottom of the thermometer it will be attached too.

2. In the test tube place a small capillary tube (sealed on one end) with the open end in the liquid (sealed end up).

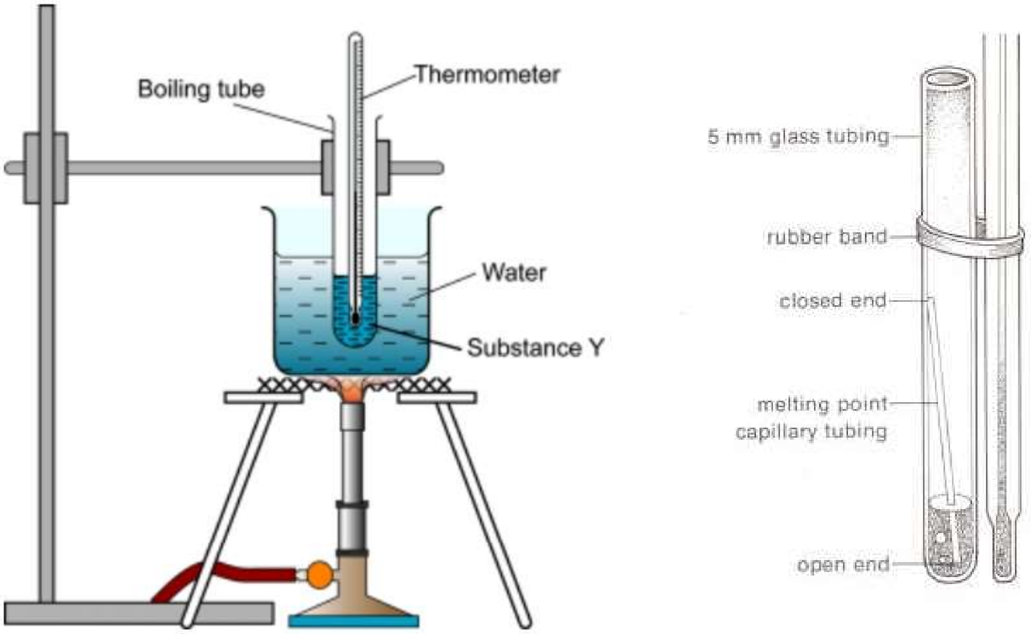

Figure 21.11: (a) Ring Stand Setup (b) Sample and Thermometer Setup. credit: (a) - `https://www.pngegg.com/en/png-ekvnv` (b) - credit: `https://www.chem.ucalgary.ca/courses/351/laboratory/boilingpoint.pdf`

General Experimental Procedure

1. Using the Bunsen burner slowly heat up the sample by waving the flame below the test tube. Aim for a rate around 10 ℃ per minute.
2. When a rapid and continuous stream of bubbles emerges from the inverted capillary tube stop heating the sample (remove the bunsen burner and turn it off).
3. Raise the sample test-tube out of the water bath and allow it to cool faster..
4. Observe the sample carefully, when the stream of bubbles escaping slows to a stop **AND** the liquid begins to enter the capillary tube the Boiling Point has been reached. Record the temperature (to the appropriate number of Significant Figures) in column 1 of the data table.
5. Allow the sample to cool down (run it under the water in the sink and/or place it in another beaker of room temperature water for a few minutes).
6. Allow the water bath too cool slightly (add some cool water).
7. Place a new capillary tube into the test tube.
8. Repeat the experiment at least 3 times so you can calculate the average and standard deviation.
9. Dispose of all mixtures in the appropriately labeled container in the hood. ☠

Calculations

1. Record the Atmospheric Pressure in the laboratory (Ask your instructor).
2. Using the equations in the discussion section to correct your measured values taken at elevation to the values measured at sea-level.
    (a) First calculate the Atmospheric pressure in mmHg.
    (b) Second calculate the correction factor ΔT_b using equation 1. Place the result in column 2.
    (c) Third calculate the Boiling Point (B_p) at sea level for each trial using equation 2. Place the result in column 3.
3. Show a sample calculation for Trial 1 in the space below the Table.
4. Calculate the Average and Standard Deviation of your Data using the corrected values. Show the calculation in the space below the Table.
5. Look up the accepted value for your liquid and calculate the Percent Error. Show your calculation below the Table.
6. Double check all Significant Figures and Units

Part II - Examining the Effect of Size on Boiling Points (Outside of Class

1. Copy the data for all groups into Data Table 2. For the values not measured in lab you will have to look up the measured values.
2. Make a nice graph using all you learned in Lab 2.0 and 2.5 in first semester.
3. Be sure to include: Axis Labels and Units, Major and Minor tick marks, A Legend, a Curve-Fit and R^2 value. Also label each of the liquids measured.

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Name: ______________________________ Date: ______________________ Score: ______/75

Results

Part I - Boiling Point of Assigned Liquid

Group Number: _______ Assigned Liquid: __________________ Pressure in Rangely: _______

Group Members: ______________________ ______________________ ______________________

Trial	**Measured B$_p$ (Rangely)**	**ΔT$_b$ (Correction)**	**Correct B$_p$ (Sea-Level)**
1			
2			
3			
4			
5			
		Average:	
		Standard Deviation:	
		Literature Value:	
		Percent Error:	

Table 21.1: Measured Boiling Point of Assigned Liquid.

[15 pt] 1. Show a sample calculation for Trial 2 - Corrected B$_p$ (Sea-Level) **AND** calculation of Standard Deviation and Percent Error.

Part II - Relationship Between Molecular Weight and Boiling Point

Note: Fill the data out for **ALL** compounds, but you only need to make **ONE** graph for this part. If you were assigned an Alkane graph the Alkane data. If you were assigned an Alcohol graph the Alcohol Data.

Substance	**MW (g/mol)**	**B$_p$ (℃)**	**Substance**	**MW (g/mol)**	**B$_p$ (℃)**
Methane			Methanol (Lab)		
Ethane			Ethanol (Lab)		
Propane			1-Propanol (Lab)		
Butane			1-Butanol		
Pentane (lab)			1-Pentanol		
Hexane (Lab)			1-Hexanol		
Heptane (Lab)			1- Heptanol		
Octane			1-Octanol		

Table 21.2: Molecular Weight and Boiling Point Data for Graph

[25 pt] 2. Attach a graph showing the effect of molar mass on the measured change in temperature for the set of liquids assigned to you. (Plot the molar mass on the x-axis and boling point on the y-axis). Based on your graph what trend did you observe (Directly Proportional, Inversely Proportional, No Trend). (Answer using words **AND** a mathematical equation.)

[5 pt] 3. What is responsible for the trend observed. Explain.

Additional Post-Lab Questions

[5 pt] 1. From your graph, make a prediction of the value (ie a number!) for the Boiling Point of either Decane or 1-Decanol. Explain.

[5 pt] 2. Pentane (72 g/mol) and 1-butanol (74 g/mol) have nearly identical molecular weights but should have had very different Boiling Points. Explain the difference boiling points in terms of the Intermolecular Forces (IMF) present between the molecules.

[5 pt] 3. Which IMF's are present between two methane molecules? Draw a picture illustrating the interaction between two methane molecules.

[5 pt] 4. Which IMF's are present between two methanol molecules? Draw a picture illustrating the interaction between two methanol molecules

[5 pt] 5. If you were assigned Alkanes, why didn't we measure the boiling point values for Methane, Ethane, Propane, and Butane? If you were assigned the Alcohols why didn't we measure the boiling point for 1-Butanol, 1-Pentanol, and 1-Hexanol?

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Name: ______________________________ Date: __________________ Score: _____/25

Prelab Questions

[5 pt] 1. Complete the table below illustrating the differences between Chemical Bonds and Intermolecular Forces (IMF's).

Property	Chemical Bonds	Intermolecular Forces
Attraction between:		
Relative Strength:		
Represented by (in LS):		
Determine Properties like:		

[10 pt] 2. Define (using sentences) each of the following Intermolecular Forces (IMF's) including their relative strengths. In addition draw an example illustrating the attractive force between **TWO** molecules. Properly label all charges (+/-) and partial charges (δ^+/δ^-).

(a) London Dispersion Forces (LDF)

(b) Dipole-Dipole Forces (DD)

(c) Hydrogen Bonding (HB)

(d) Ion-Dipole (ID)

[5 pt] 3. 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:

[5 pt] 4. If the correct (Literature value) answer for the data measured above was 18.0℃. What is the Percent Error for the measured data? Show work in the space provided.