Protocol | Intermolecular Forces

Intermolecular Forces, Evaporation, and Vapor Pressure

Introduction

Watch the [conceptual introduction video](#) for this experiment.
Within a molecule, atoms are held in close proximity by chemical bonds. However, many of the macroscopic properties of compounds are dictated not by interactions within molecules, but *between* molecules. These [intermolecular forces (IMFs)](#) hold molecules relatively close to one another and account for the existence of condensed phases despite their lower entropy than the gas phase.

Intramolecular forces are electrostatic in nature, arising from the attraction of full or partial charges of opposite sign in neighboring ions or molecules. IMFs can be classified according to the extent of the charge (full or partial) and its lifetime (temporary or permanent) on the interacting particles. The six most common types of IMFs are:

- Ion-ion forces (ionic bonds)
- Ion-dipole forces
- Hydrogen-bonding interactions
- Dipole-dipole forces
- Dipole-induced dipole forces
- London dispersion forces

We represent the strengths of chemical bonds using the bond energy, the amount of energy input required to break the bond. The strengths of intermolecular forces can similarly be represented using amounts of energy. Because it takes some energy to disrupt intermolecular forces, non-zero temperature is required to cause melting or evaporation (phase transitions) in a substance. Furthermore, the *stronger* the intermolecular forces, the *larger* the temperatures required to effect phase transitions. Thus, observations of phase transitions give us insight into the relative strengths of IMFs in different substances. Working in the opposite direction, we can interpret Lewis structures to identify the dominant intermolecular forces in a series of substances and predict or deduce the properties of their phase transitions.

In this experiment, we will focus on evaporation, the liquid-gas phase transition. Because of the very high entropy of the gas phase, evaporation occurs spontaneously from liquids open to the air. The rate of evaporation, measured as the percent mass lost per unit time, depends on the strength of intermolecular forces in the liquid. After working through the exact form of this dependence and predicting relative evaporation rates for six liquids,

we will make simple measurements of the evaporation rates of the liquids and compare them to the predictions.

We will also examine how the temperature of a piece of filter paper changes as liquid on the paper evaporates from it. Evaporation is an *endothermic* process that requires an *investment* of enthalpy to drive the disruption of forces between the liquid molecules. This means that the temperature of the filter paper will *decrease* as liquid evaporates from it: energy is transferred from the paper to the liquid and that energy "goes toward" breaking intermolecular forces and converting the liquid to gas. How much the temperature decreases depends on how much energy must be invested to break the IMFs in the liquid phase. We will again work through the exact form of this dependence and make predictions of the relative temperature changes (ΔT) for the same six liquids. After making measurements of ΔT for the liquids, we will compare the experimental results to the predictions.

Finally, we will study the dependence of the vapor pressure of these six liquids on temperature. Vapor pressure is defined as the pressure exerted by gas that has evaporated from the surface of a liquid in a closed container. As the figure below illustrates, some of the molecules of a liquid evaporate spontaneously to form a vapor phase above the liquid surface that exerts a pressure. This is what we call vapor pressure.
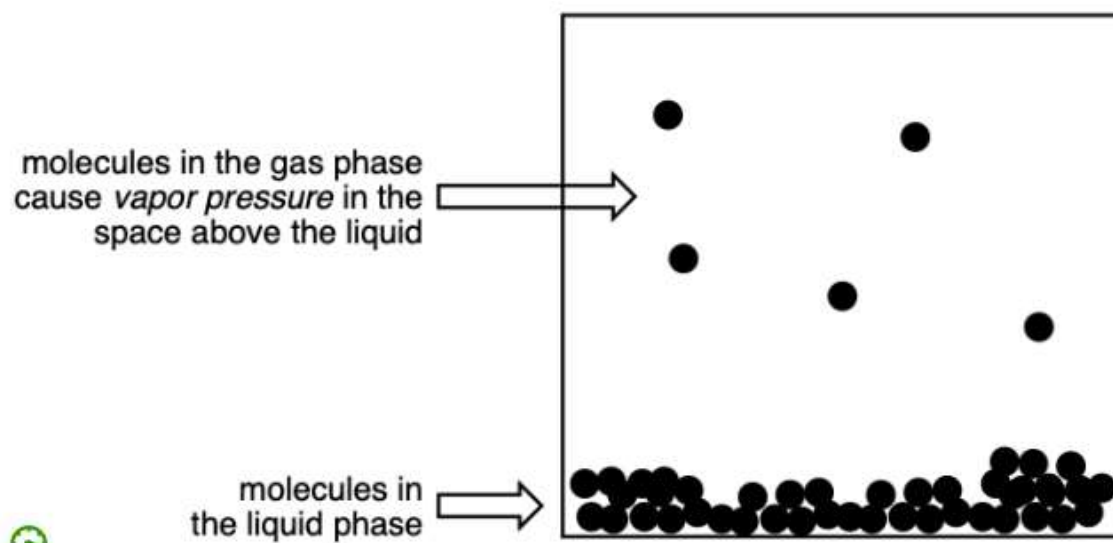
Because the gas phase is favored at higher temperatures, heating increases the vapor pressure of liquids. However, the rate at which vapor pressure increases with temperature depends on the strength of intermolecular forces within the liquid. If you consider how strong intermolecular forces stabilize the liquid phase, it isn't difficult to work out the qualitative relationship between IMF strength and vapor pressure.

The six liquids we will study in this experiment are water, ethanol, 1-propanol, acetonitrile, pentane, and toluene.

Safety and Materials

The last part of this experiment involves gently heating a closed system. Although the heat used is gentle, there is some risk of the stopper popping off, particularly for the liquids with intrinsically high vapor pressure. (Do you know which liquids these are?) Resting a finger on the stopper ensures that it will not pop off as the apparatus is heated.

The following reagents will be available:

- [Ethanol](#)
- [1-Propanol](#)
- [Acetonitrile](#)
- [Pentane](#)
- [Toluene](#)

Relevant Experimental Techniques

[Measuring Temperature or Pressure with a LabQuest 2](#)

Research Question

Do the relative strengths of dominant intermolecular forces between solvent molecules determine the macroscopic properties associated with evaporation?

Procedures

A. Lewis Structures and Intermolecular Forces

Both Partners (Work Independently)

1. Molecular formulas are provided below for the six compounds we will investigate in this experiment. In your lab notebook, draw a Lewis structure for each compound.
2. Identify the dominant intermolecular force (IMF) in each compound.
3. Before beginning measurements, predict the relative strengths of the intermolecular forces in these compounds. Order them from strongest to weakest IMFs. For compounds with the same dominant IMF, carefully consider other effects that could influence the strength of intermolecular interactions.

4. Make a prediction regarding the effect of IMFs on evaporation rates. Order the compounds from largest to smallest evaporation rate.
5. Make a prediction regarding the effect of IMFs on the temperature change during evaporation. Order the compounds from largest to smallest expected temperature change.

Compound	Molecular Formula
Ethanol	CH_3CH_2OH
1-Propanol	$CH_3CH_2CH_2OH$
Acetonitrile	CH_3CN
Pentane	$CH_3CH_2CH_2CH_2CH_3$
Toluene	$C_6H_5CH_3$ (see model)
Water	H_2O

B. Measurement of Vapor Pressure as a Function of Temperature

1. Your bench will be assigned one of the six **liquids** for this part of the experiment. ***Make sure to obtain temperature and vapor pressure data for the other five liquids from the other lab benches.***
2. On a hot plate, heat about 250 mL of water in a separate 400 or 600 mL beaker to approximately 60 °C. This water will be added to the bath of the apparatus above in small increments to increase its temperature.
3. Set up an apparatus to measure the pressure of gas inside a 125 mL Erlenmeyer flask as a function of temperature (see below). Ensure that the flask is closed and that a water bath surrounds the flask. Do not place the apparatus on a hot plate—we will use additions of hot water to the bath to heat the apparatus. Add enough water to the bath initially to cover the volume of the flask accessible to the gas. Do not add your assigned liquid to the Erlenmeyer flask yet. Throughout this part of the experiment, just use the LabQuest display for measurements and record manually in your notebook.
4. Record the atmospheric pressure before adding your assigned liquid to the Erlenmeyer flask. Transfer approximately 3 mL of the assigned liquid into the flask, then replace the stopper to re-seal the system. Once the pressure

returns to equilibrium, record the pressure and the temperature in Kelvin or degrees Celsius. This pressure minus atmospheric pressure is equal to the vapor pressure of the liquid.
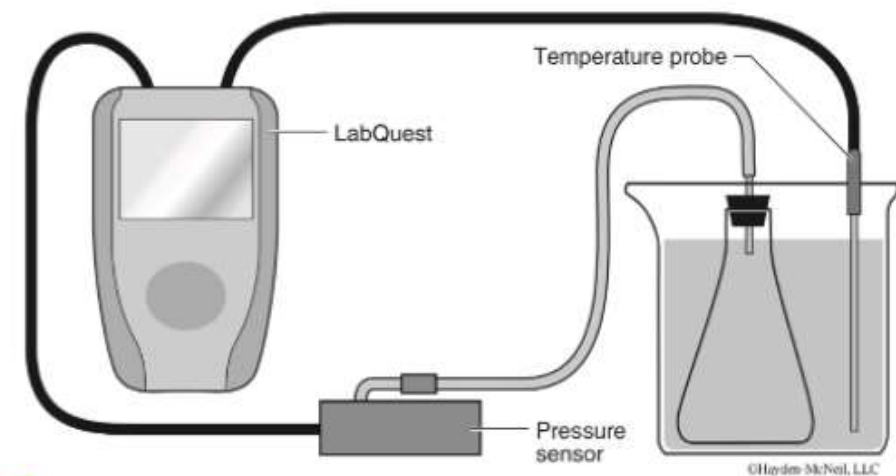

1. Use the syringe in the gas accessories kit at your lab bench to remove approximately 40 mL of water from the room temperature bath.
2. Using the same syringe, add approximately 40 mL of hot water back to the bath. Gently stir using the temperature probe to mix the bath.
3. Continue stirring until the temperature and the pressure stabilize. The temperature should rise by a few degrees. Note: the pressure may rise slowly! Watch the LabQuest display carefully. The temperature should rise by 3 – 5 °C.
4. Record the temperature and pressure displayed on the LabQuest.
5. Repeat steps 5 – 8 five more times to obtain six total data points of vapor pressure and temperature. With each addition of hot water to the bath, the temperature should rise by 3 – 5 °C. It may be necessary to replace more than 40 mL of water in the bath to observe a significant temperature increase. Avoid increasing the temperature beyond 40 °C (313 K) to avoid popping off the stopper.
6. Calculate the vapor pressure at each temperature by subtracting your measured atmospheric pressure. Add your temperature and vapor pressure data to a table on the board and record data from the other five benches.

C. Measuring Temperature Changes During Evaporation

1. Set up a LabQuest at your bench with two temperature probes, carefully noting which probe corresponds to each temperature displayed on the LabQuest. Program the LabQuest to collect temperature measurements every 2 seconds for 400 seconds total.

2. Cut two pieces of filter paper to a width of about an inch and wrap them around the ends of the temperature probes. Secure the filter paper in place with a rubber band.
3. Add a few milliliters of **your assigned liquid** to a medium test tube and submerge the probe in the liquid so that the filter paper is completely covered.
4. With the probe submerged, start data collection on the LabQuest. Allow the probe to sit in the test tube for about 15 seconds to obtain a reliable initial temperature.
5. Remove the probe from the test tube and tape or hold it to the end of your lab bench so that the end extend beyond the end of the bench.
6. As liquid evaporates from the filter paper, you will observe a temperature decrease. Once both probes have stopped showing a decrease in temperature, stop data collection and record the initial and final temperatures for the liquid. Add this value to the class data.
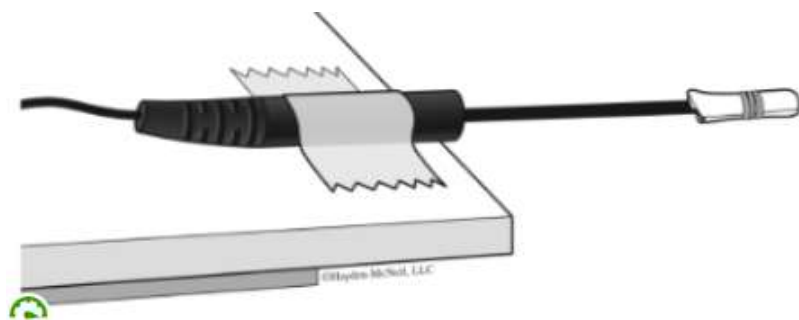
D. Measuring Evaporation Rates

1. Record the mass of a clean and dry 30 mL beaker. Add 20 drops of ethanol to the beaker and re-weigh it. Subtract the mass of the empty beaker to calculate the initial mass of ethanol alone.
2. Return the beaker to your lab bench and set a timer for 2 minutes. If all of the substance evaporates, repeat step 1 waiting for only 1 minute instead of 2.
3. Once 2 minutes have elapsed, return the beaker to the balance and re-weigh it. Calculate the mass lost due to evaporation and the percentage of the initial mass lost per minute (evaporation rate).
4. Empty the beaker into the waste bottle, clean and dry it with a Kimwipe, then repeat steps 1 – 3 using each of the other 5 solvents.

Group Analysis and Argumentation

With the other students at your bench, compile class and/or group data and analyze any trends. Return to the research question and arrive at a conclusion addressing the

question together. Generate an argument from evidence that can be used in your group post-lab poster. The posters will be presented in class during the checkout week and will be peer evaluated. Record the trends and argument in your lab notebooks before leaving. Decide who will complete each portion of the poster template.

Post-lab Calculations and Data Workup

Compare your predictions in Part A to your results. Does vapor pressure follow the trend you expected? What about temperature change upon evaporation and evaporation rate? Are any of your results surprising?

Tabulate your data using the data workup template.