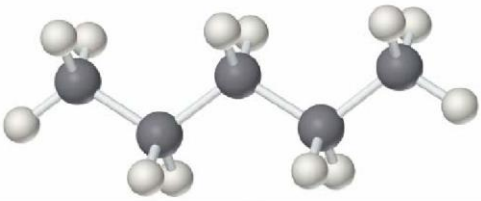
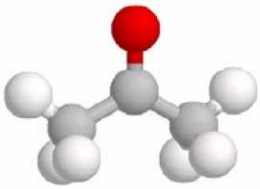
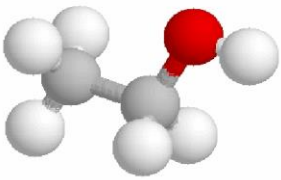
Intermolecular and Ionic Forces

Introduction: Molecules are attracted to each other in the liquid and solid states by **intermolecular**, or attractive, forces. These are the attractions that must be overcome when a liquid becomes a gas (vaporization) or a solid becomes a gas (sublimation). In the gas phase molecules are much farther apart and, therefore, do not interact as strongly as they do in the liquid (condensed) phase. Thus, the energy required to boil or vaporize a pure substance is a measure of the strength of the intermolecular forces in that substance. In general, intermolecular forces are much weaker than the ionic and covalent bonds that hold together the atoms and ions in a compound. For example, about **40 kJ** of energy are required to vaporize 18 grams of water molecules—i.e., completely convert 18 grams of water to water vapor or steam. In contrast, **930 kJ** of energy is required to break all of the O-H covalent bonds in the same amount of water!

Alkanes (molecules with C and H atoms), alcohols (molecules with C, H, and O atoms), and ketones (C double-bonded to O and single-bonded to two other C's) will be the three types of organic compounds compared in this lab. The alkane used in this activity is pentane, $CH_3CH_2CH_2CH_2CH_3$. Alcohols are similar to alkanes except one carbon in the chain is bonded to a hydroxyl (-OH) group (see the figures shown below). The alcohol used in this activity is ethanol, $CH_3\ CH_2OH$. The ketone used in this lab is acetone (the primary ingredient in fingernail polish remover), CH_3COCH_3.

	pentane $CH_3CH_2CH_2CH_2CH_3$
	acetone $CH_3C{=}OCH_3$
	ethanol CH_3CH_2OH

Types of Intermolecular Forces:

Refer to section 12.1 in your textbook for a thorough explanation of each type of Intermolecular force. Intermolecular interactions are generally classified as being London (dispersion) forces, dipole-dipole forces, hydrogen bridges, and ion-dipole forces.

London forces (also called dispersion or induced-dipole forces) exist in all molecular samples – both polar and nonpolar compounds. Consider a Br_2 molecule (see Figure 1), which has no permanent dipole moment. Remember that dipole means 2 poles, like a magnet they have a distinct positive end and a distinct negative end.

Figure 1: London forces for two Br_2 molecules. (a) Represents the two molecules before the induced dipoles appear; (b) shows the alignment of the induced dipoles.

Nonpolar molecules and single atoms have only London forces. The temporary shift of electrons (Figure 1b above) causes a temporary dipole in one molecule that will then cause (or induce) a temporary dipole in another molecule. The temporary dipoles cause quick and very weak attractions between atoms or molecules.

London forces exist in all compounds and will be stronger in larger molecules or atoms that have larger numbers of electrons to shift. Larger atoms or molecules are thus more polarizable (can experience a stronger temporary dipole). Dispersion forces are the only type of intermolecular forces experienced by nonpolar molecules.

Dipole-dipole forces exist between **polar molecules**. Polar molecules have a permanent dipole, so the charges (notated as + (partial positive) and □(partial negative) are separated from one another. See Figure 2 below for an example. In polar compounds, the positive end of one molecule is attracted to the negative end of another molecule. Because polar molecules have a permanent dipole, rather than a temporary dipole, the dipole-dipole forces they experience are generally stronger than London forces. However, the H -Cl covalent bond within the molecule is still far stronger than the dipole-dipole intermolecular forces between molecules. Figure 2a shows how multiple polar molecules will arrange themselves in solids and liquids.

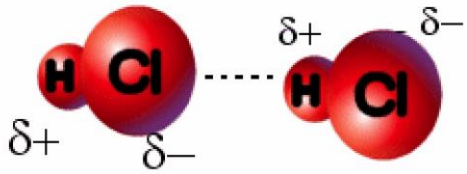


Figure 2: Dipole-dipole intermolecular interaction between two HCl molecules.

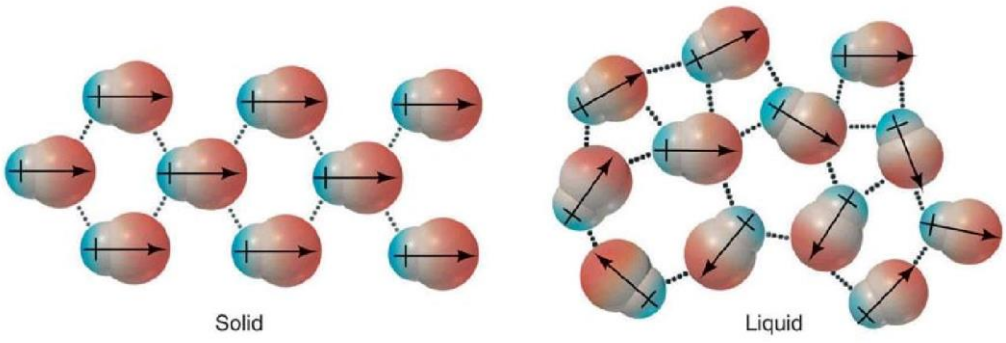


Figure 2a: How polar molecules align themselves in bulk.

Hydrogen bridges (or bonds) are basically very strong dipole-dipole forces between the hydrogen atom bonded to a N, F, or O atom within one molecule to an N, F, or O in another molecule. Only molecules with a H-N, H-O, or H-F bond in them can form hydrogen bridges. These bonds are very

polar with the H having a partial positive charge which is attracted to the lone pairs on an N, F, or O of a second molecule. Thus the two molecules orient so that the H from the first is near the N, O or F of the second. Two ammonia (NH_3) molecules are depicted in Figure 3 below with the Hydrogen bridge shown as a dotted line between the H atom in one molecule and the N atom in the second molecule. Remember the covalent bond between N and H within an ammonia molecule is considerably stronger than the hydrogen bridges holding a sample of ammonia molecules together.

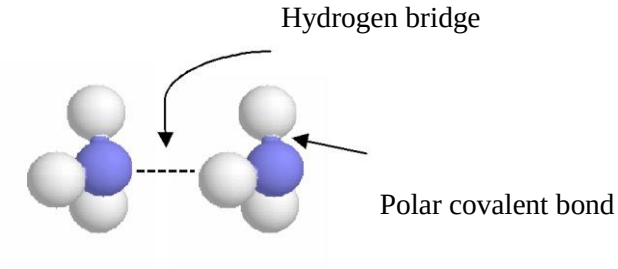


Figure 3: Hydrogen bridge (dotted line) between two ammonia molecules.

Ion-dipole forces occur between an ionic (ion) substance and a polar liquid, or solvent (dipole). When table salt, NaCl*(s)*, dissolves in water, the positive end of the water molecule is attracted to the negatively charged chloride ion (Cl^-) and the negative end of the water molecules is attracted to the positively charged sodium ion (Na^+). These attractions are strong enough to separate some ionic compounds into individual ions into solution.

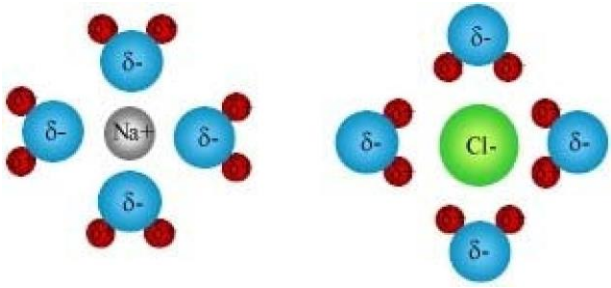


Figure 4: Ion-dipole forces showing polar water molecules orienting themselves around (left) a positively charged Na^+ ion and (right) a negatively charged Cl^- ion.

This experiment focuses on the Intermolecular Forces that are broken when substances undergo phase (or physical) changes. During these processes, the **chemical structure** of a substance doesn't change, only its phase changes. For example, boiling water is the same as evaporation. In a chemical equation, we can write it as: H_2O (l) □ H_2O (g). Notice that we wrote the formula the same on both sides of the equation, so the chemical structure is the same: no **BONDS** were broken during this process.

Evaporation: When you step out of a swimming pool, you feel cool as the water evaporates off your skin. Similarly, when a metal object (e.g., a metal temperature probe) is immersed in a volatile liquid (a liquid that evaporates easily like an alcohol) and is then exposed to the air, the surface of the metal object becomes colder as the liquid evaporates. This effect is called evaporative cooling. Both the alcohol and the metal object are at room temperature when the object is immersed in the alcohol. In general, liquid molecules with stronger intermolecular attractions require more energy to escape to the gas phase, so they undergo a **slower** rate of evaporation. **In other words, a substance with stronger intermolecular forces will evaporate more slowly and will have more mass left after a set amount of time.**

Intermolecular and Ionic Forces

Comparing IMF Using Percent Evaporation:
This lab will explore changing masses during evaporation of organic liquids. The decrease in mass is related to the strength of the intermolecular forces of attraction. Consider the strength of the forces that hold molecules together in each liquid. Will a greater decrease in mass as each substance evaporates indicate stronger or weaker intermolecular forces?

Part I: Comparison of evaporation rates
Intermolecular forces influence the ability of the molecules to evaporate. You will measure the mass of a sample that evaporates after 3 minutes, then take the difference between that mass and the initial mass of the sample to determine the percent by mass of the sample that evaporated. You will explain the results of your data based on the intermolecular forces of the different substances.

One group will work at a time in each hood to complete this portion of the lab. While your group is waiting for a hood, you can work on Parts II and III. Part I should take approximately 15-20 minutes to complete.

Caution: Organic liquids are highly flammable and volatile. The three organic liquids MUST remain in the hood at all times.

The data for every experiment needs to be in your lab notebook.

Procedure:

1. In this experiment, you will use three organic liquids: ethanol, acetone, and pentane. In the hoods, there are three 10-mL beakers labeled "E" (for ethanol), "A" (for acetone), and "P" (for pentane) for use with each of the liquids.
2. Use a Kimwipe to completely dry the inside of the beaker labeled "E". Weigh the empty beaker, and record its mass in your lab notebook.
3. ***In the hood***, carefully add **only 10 drops** of the ethanol liquid into the beaker. Reweigh the beaker, and record its mass in your lab notebook.
4. Take the beaker back to the hood, and set a timer for 3 minutes.
5. When the timer goes off for each liquid, weigh the sample again, and record its mass in your lab notebook.
6. Repeat the procedure (steps 2-5) for the acetone sample.
7. Finally, repeat the procedure (steps 2-5) for the pentane sample.
8. Calculate the mass of the sample at the start for each liquid, then calculate the mass of the sample that evaporated for each liquid by subtracting the "mass of the beaker + sample after 3 minutes" from the "mass of the beaker + sample at the start".
9. Calculate the percent by mass of the sample that evaporated for each liquid using the formula below:

$$\%\text{ by mass of sample that evaporated} = \frac{\text{mass of sample evaporated}}{\text{initial mass of sample}} \text{x} 100$$

****Lab Notebook****

A table with all the data collected (including values derived from calculations) should contain the following for each liquid tested (think about the best way of including all these data in *one* table):

- Identity of liquid
- Mass of empty beaker
- Mass of beaker and sample before experiment
- Mass of beaker and sample after 3 minutes
- Mass of sample (only) at beginning of experiment
- Mass of sample evaporated
- % of sample evaporated

******Also be sure to show your calculations within your **lab notebook**.

Part II: Model structure of ice

The back bench of the lab has been set up with multiple model kits to build the structure of solid water (ice). This kit is specifically designed to show the difference between covalent bonds O-H bonds within a water molecule and the hydrogen bridges holding two different H_2O molecules together. Groups will rotate through this station. **The water molecule models must stay intact**. After you are finished with the lab, check to make sure that all water molecule models are still intact.

1. Another type of model showing the structure of ice has been assembled. In this model the red ball representing oxygen and the smaller white balls representing hydrogen. This is an approximate model showing the overall structure water makes when frozen. However, you will assemble a more accurate model using the plastic molecular models.

2. Use the ruler provided to measure (in cm) the approximate distance from the **center of one hydrogen atom** to the **center of the oxygen atom**. Record these data in your **lab notebook**.

3. Now attach the water molecules to each other using the purple hydrogen bonds provided. The purple connectors will be used to connect the white H atoms to the red oxygen atoms.
Your instructor may help you to build the structure in stages: a) build six member rings in the right configuration (chair form), b) fuse rings to make a sheet and c) stack sheets to create the ice model. You should notice 6-sided, open channels within the structure of ice. (Compare your model to the pre-built read and white structure at this station.). Use as many water molecules as possible.

****Lab Notebook****

Get your lab instructor's signature confirming that your model of ice is correct.

4. Use the ruler to measure (in cm) the approximate distance between the center of a hydrogen atom in one molecule and the center of the oxygen atom in a neighboring molecule. Record these data in your **lab notebook**.

5. When you are done at this station, disconnect the hydrogen bridges (but leave the water molecules intact).

Part III: Tutorial of online animations for Intermolecular Forces

- View the interactive tutorial by opening the following link:
- http://www.wwnorton.com/college/chemistry/chemistry3/ch/10/chemtours.aspx Or, follow the link from the chemistry website.
- Scroll down until you see the icon for the green **"Intermolecular Forces Tutorial**" link.
- The 13 sections each contain multiple pages. You can click on the "Previous Section" button at the bottom left to return to an earlier section if you would like to review it.
- Be sure to click on the "Real-World Connections", "Science Connections", and "Concept Question" links.

Part IV: Phase Diagrams

- Follow this link:
- http://www.wwnorton.com/college/chemistry/chemistry3/ch/10/chemtours.aspx Or follow the link from the chemistry website.
- Scroll to the bottom of the page and select the green **"Phase Diagrams Tutorial"** link. There are 7 sections in this tutorial. Answer the questions on your lab report corresponding to the correct sections.

Be sure to click on the "Real-World Connections", "Science Connections", and "Concept Question" links

Post-Lab Questions

1. Draw the electron dot structures for each molecule.

   a. acetone

   b. pentane

   c. ethanol

2. a. Identify all type(s) of intermolecular force(s) between molecules of each substance using **LF**=London forces, **DDF**=Dipole-dipole forces, and **HB**=Hydrogen bonds.

ethanol: _____ **acetone:_____** **pentane: _____**

b. Which is the strongest intermolecular force in each molecule?

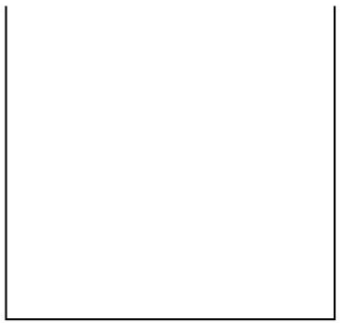
ethanol: _____ **acetone:_____** **pentane: _____**

3. Which molecule evaporated most rapidly?

_____

4. Explain how the strength of intermolecular forces in different molecules relate to their evaporation rates.

5. Draw 6 water molecules in each beaker below. On the left should be 6 liquid water molecules. On the right draw the gas particles that are present after evaporation, but before any substances escape the beaker.

Before evaporation	**After evaporation**
	

Part II: Model structure of ice

6. Sketch two water molecules (labeling the oxygen and hydrogen atoms) and identify the covalent bonds (as solid lines) holding atoms together and the hydrogen bridge (as a dotted line) attracting one molecule to another. **Pay attention to the orientations of the water molecules!**

7. Explain why your model (the one you constructed with the plastic atoms and bonds) more closely resembles the structure of solid water than does the other pre-built red and white model. What makes your model more accurate?

Part III: Tutorial of online animations for Intermolecular Forces

8. Which ionic compound would have stronger ion-ion forces (higher lattice energy) NaF or KBr? Explain your reasoning.

9. Which ionic compound would have the stronger ion-ion forces (higher lattice energy) CaS or LiF? Explain your reasoning.

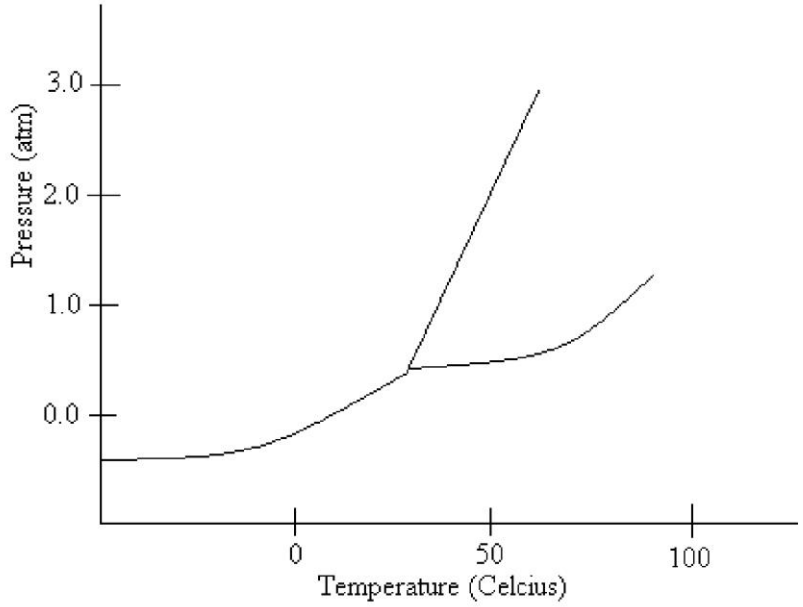
10. Which compound has the weakest intermolecular force? Cl_2, Br_2 or I_2? Explain.

11. Is KBr to be solid or liquid at room temperature? Explain why. (Hint: Think about the type and strength of intermolecular force you would find in a sample of KBr.)

Part IV: Phase Diagrams

12. Section 4: Explain specifically why the temperature change stops at each of the black lines on the phase diagram as the water sample is being heated?

13. Section 5: Use the phase diagram below (this is **NOT** water's phase diagram) to label the 3 open areas as solid (S), liquid (L), or gas (G); label the three segments as solid-liquid (SL), liquid-gas (LG), or solid-gas (SG) equilibria; label the triple point (TP).



What is the normal boiling point (at 1 atm) of this substance? _____

What is the normal freezing point (at 1 atm) of this substance? _____

14. Circle all of the following molecules that can participate in hydrogen bonding:

H_2O H_2S CH_4 CO_2 HF

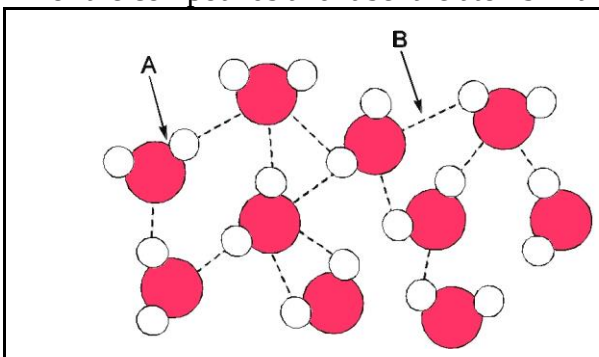
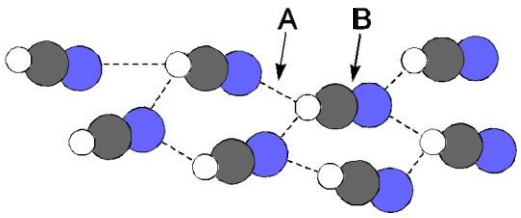
15. Rank the following in order of **increasing** boiling point (Hint: what must be broken to allow ionic substances boil? What must be broken to allow covalent substances to boil?)

a. KCl, Al_2S_3, H_2O, F_2, HCl, H_2.

_____ < _____ < _____ < _____ < _____ < _____

lowest b.p. highest b.p.

16. The box on the left is H_2O, the box on the right is HCN. Identify what the letters A and B refer to in the structures below – a chemical bond (ionic, polar covalent, or nonpolar covalent) or an intermolecular force. For IMF's, identify the **strongest intermolecular force**. (Hint: Write formulas of the compounds and label the atoms in the pictures!)

	
A:	A:
B:	B:

For number 5 below, consider the following six choices:

A. ionic bond
B. polar covalent bond
C. nonpolar covalent bond
D. London (dispersion) forces
E. dipole-dipole forces
F. hydrogen bond

17. Identify the type of bond or intermolecular force between the following:

_____ a. I and I in an I_2 molecule.
_____ b. Na^+ and Cl^- in solid NaCl.
_____ c. HBr and another HBr molecule.
_____ d. NH_3 and another NH_3 molecule.
_____ e. H and F in a HF molecule.
_____ f. Br_2 and another Br_2 molecule.