Introduction

In CHM 110, we used **kinetic** theory to explain the behavior of gases. Now, we will discuss solids and liquids. While we don't have easy equations like the ideal gas law to describe all solids and liquids, we **can** qualitatively discuss liquids and solid properties and how they follow from the **structure and size** of their molecules. We will first quickly review what we know about the three phases of matter - solid, liquid, and gas. We will then discuss what happens during a **phase transition** (the change of a substance from one phase to another).

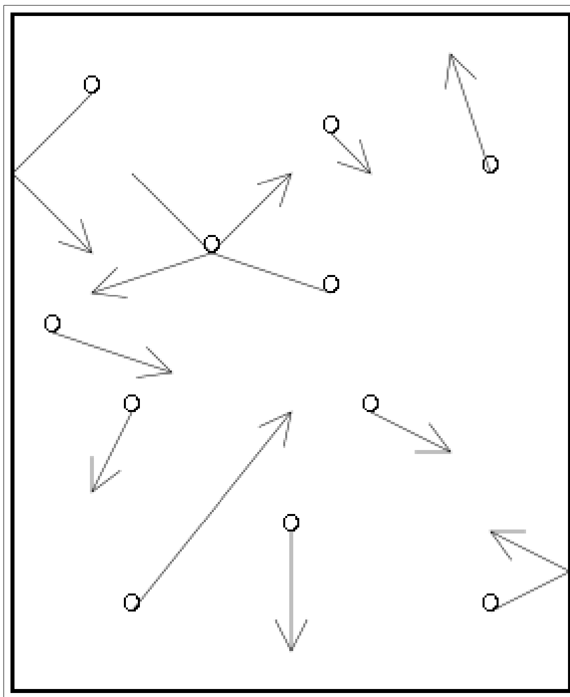
The states/phases of matter: Gases

First, since we have discussed gases in detail, we will begin with a brief review of what we know about gases.

All gases have much in common. Gases are:

1. **compressible** - You can change the volume of a gas easily by increasing the pressure you apply to it.
2. **fluid** - A gas will flow easily from one area to another. Gases flow **so** easily that they cannot maintain a definite shape or volume without the influence of a container.
3. ***not* dense** - Gases are mostly empty space, with molecules spread far apart in comparison to the size of the molecules.

We described gases using the **kinetic theory**, as illustrated in the picture below:



Postulates of kinetic theory, in brief:

1) Gas molecules are small compared to the empty space between them.
2) Gas molecules move in straight lines until they hit something (another gas molecule or container walls).
3) There are no forces between the molecules - either attracting them together or pushing them apart - unless they collide.
4) When gas molecules collide, energy may be transferred, but none is lost as heat.
5) The temperature is related to the average kinetic energy (speed) of the molecules.

Illustration 1 - The kinetic theory of gases

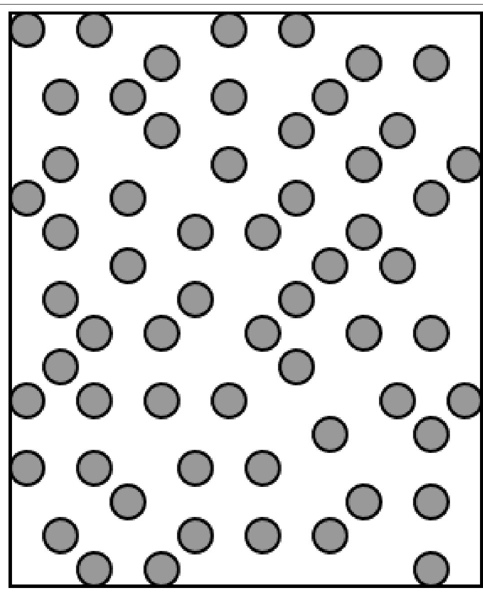
Kinetic theory leads us to the conclusion that most gases have very similar physical properties. At the conditions we're used to seeing gases in, they do. Let's compare with liquids and solids.

The states/phases of matter: Liquids

What about liquids? When compared to gases, liquids are:

1. **incompressible** - You can't change the volume of a liquid easily by increasing the pressure you apply to it.
2. **fluid** - A liquid will flow easily from one area to another and take the shape of the container they are put in. Liquids do have a definite volume, however.
3. **dense** - The molecules in a liquid are very close together.

If we were to make a picture of the liquid phase to compare with the kinetic theory of gases, we might come up with a picture like this:

<table>
<tr><td></td><td>In the liquid state:
• Molecules are much closer together than in a gas.
• Molecules are free to move around each other (liquids can flow).
• Molecules are held together by what we will call intermolecular forces</td></tr>
<tr><td>Illustration 2 - How a liquid is put together</td><td></td></tr>
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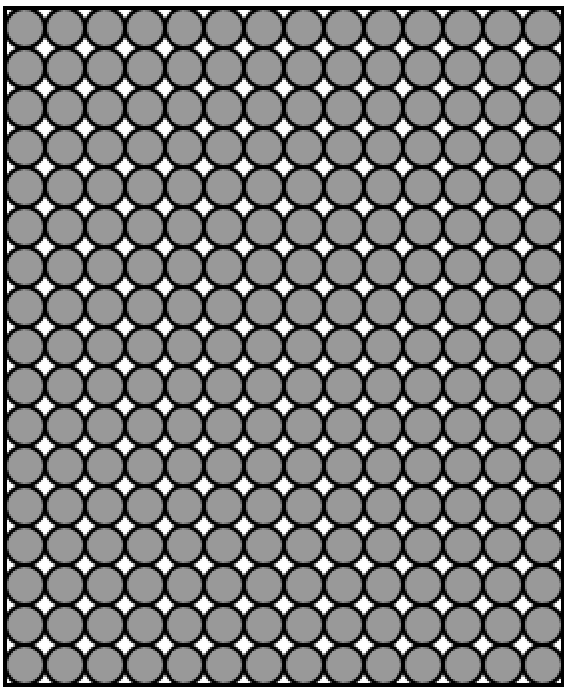
Liquids are not as similar in their properties as gases are. It's easy to see the difference in properties between water and motor oil. This is due to the fact that the molecules in liquids are close enough to interact with each another. The interactions are different between different kinds of molecules.

The states/phases of matter: Solids

Solids? When comparing to gases and liquids, solids are:

1. **incompressible** - You can't change the volume of a solid easily by increasing the pressure you apply to it. It's even more difficult in most cases to do this to a solid than a liquid.
2. **rigid** - A solid will not flow easily from one area to another. Solids have definite shape and volume.
3. **dense** - The molecules in a solid are very close together - usually closer together than in a liquid.

We can visualize a solid like this:

	In the solid state: • Molecules tend to be packed more closely together than in the liquid or gas state. (Exception: In ice, water molecules are actually farther apart than in liquid water.) • Molecules can't flow freely around one another. Molecular motion is limited to vibration. • Molecules are tightly held together by **intermolecular forces**. • Most solids have a definite structure - a regular ordering of molecules called a **crystal lattice**. This is not true for *all* solids.
Illustration 3 - How a solid is put together	

Like liquids, solids also differ greatly among themselves. Compare candle wax to a diamond, for example. These differences are caused by differences in the nature of the intermolecular forces holding the solids together.

<u>Phase transitions: Going from one state of matter to another ...</u>

Before we discuss the nature of the **intermolecular forces** that make liquids and solids behave differently from gases, we will first discuss changes between the various phases of matter. A change from one phase of matter to another is called a **phase transition.** There are **six** possible phase transitions listed in the table below

Transition name	***Type***	***Example***
Melting	Solid to liquid	Ice melts to form liquid water.
Sublimation	Solid to gas	Dry ice (frozen CO_2) sublimes to form CO_2 gas.
Freezing	Liquid to solid	Liquid water freezes to form ice.
Vaporization	Liquid to gas	When you boil a pot of water, liquid H_2O changes to water vapor (water in the gaseous state).
Deposition	Gas to solid	Frost is formed by the deposition of water vapor onto the ground as ice.
Condensation	Gas to liquid	Water will condense onto a cold glass left in a humid room. (Humid means there is a significant amount of water vapor in the air.)

What happens during a phase transition?

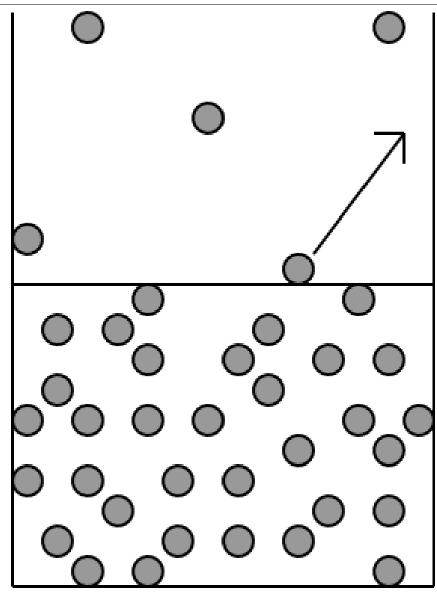
It will help us understand the effect of intermolecular forces and molecular structure on phase changes if we first have an understanding of what goes on during a phase change (at the microscopic level). We will discuss a pair of phase changes in some detail - vaporization/condensation. We'll extend that discussion to the others.

Vaporization and condensation

Before we explain what's going on in each of these processes, let's first make an observation. Vaporization and condensation occur **continually** over a liquid surface, no matter the temperature. You don't have to boil water for it to vaporize, though boiling it speeds the process. At room temperature, water will evaporate. - liquid water molecules are vaporized. On the other hand, water vapor condenses on the liquid surface. A simple piece of evidence for this is the observation that if you **cover** the glass of water, the water can't leave the glass. So the evaporating water must condense back into the liquid surface.

Why? At given conditions, only so much of a substance can be in the vapor phase. On the weather report, they call this the humidity. On a day with 100% humidity, the air is holding as much water vapor as it can at that temperature. We will use **vapor pressure** in CHM 111. The **vapor pressure** is the partial pressure of vapor over a liquid surface at equilibrium (that is, when the rates of evaporation and condensation are equal) at a given temperature.

Now - how can we explain vaporization? Let's look at the kinetic theory of gases and our picture of the liquid state. For water to get from the liquid to the gas state, the molecules have to be accelerated to a speed where they can break away from the other water molecules (and their attractive forces). The faster the molecule is going, the easier time it has at breaking away. This would explain more water evaporating at high temperatures, but how does the water evaporate at **low** temperature?

	• The more dense liquid state is shown on the bottom. • The horizontal line in the middle of the diagram represents the liquid surface. • The molecule with the arrow has acquired enough kinetic energy (speed) to break out of the liquid state and into the gas.
Illustration 4 - Vaporization	

We can account for vaporization at room temperatures this way: The definition of temperature, according to kinetic theory, is the **average kinetic energy** (speed) of the molecules. Remember that all molecules are **not** moving at the same speed. The distribution of molecules looks something like this:

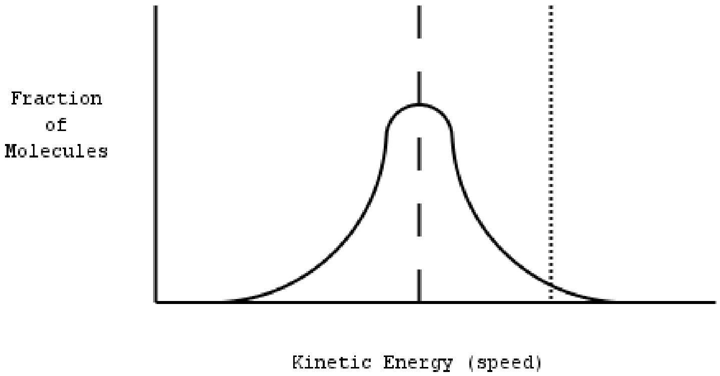


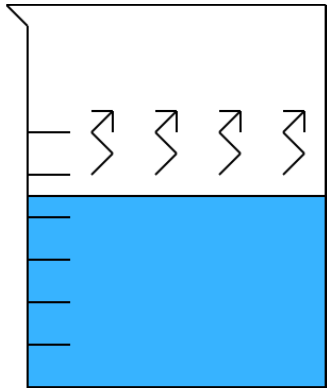
Illustration 5 - Distribution of molecular speeds at a given temperature

- This distribution is called a **Gaussian** in statistics.
- The dashed line represents the average speed (the temperature).
- The dotted line on the right represents the speed necessary for a molecule to escape into the gas phase.
- Note that even at low temperature **a fraction of the molecules are moving fast enough to escape**!
- A temperature change will shift the curve - lower temperatures will shift the curve to the left, and higher temperatures will shift it to the right.

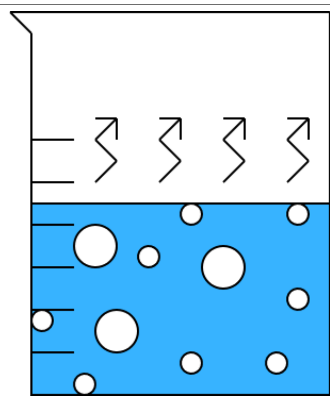
So, even at a low temperature, some molecules still have enough kinetic energy to vaporize. This is how evaporation at room temperature occurs.

What about boiling? Boiling is a form of vaporization. But why is there such a thing as a boiling point? Before you got to chemistry class, you knew that the boiling point was the temperature at which a substance starts to boil. You know that water boils at 212^{o}F (or 100^{o}C). If you've done much cooking , you might know that water boils at a lower temperature at high altitudes ("high altitude" cooking instructions account for this). What is boiling to a scientist?

We defined the **vapor pressure** as the pressure of the gas phase of a substance over the liquid surface. We will now define the **boiling point** of the substance as the point at which the **vapor pressure** of the substance equals the external pressure. When you boil a liquid in an open container, the external pressure is **atmospheric pressure**. Let's look at what happens when we heat a liquid to boiling. We will use a beaker of water as an example.

	As we heat the beaker of water, we can observe the following: • The temperature of the water rises. • As the temperature rises, the rate of vaporization (evaporation) increases. • We may see steam - water vapor - rise from the water surface.
Illustration 6 - A beaker of water.	

As we heat the water, its temperature increases - just as we'd expect. The rate of vaporization also increases. We start to see steam form when the water gets hot. Eventually, the water begins to boil:

	As the water boils: • We must continue to heat the beaker, or else boiling will stop. • Bubbles of water vapor are visible in the water. • The **temperature of the water remains constant**. For pure water at sea level, this is 100ºC.
Illustration 7 - A beaker of boiling water	

When we reach the boiling point, the vapor pressure of the water is equal to the external pressure. That means that vapor bubbles can (and do) form in the liquid. These bubbles that form in the liquid are pure water vapor. Since they're less dense than the liquid phase, they float to the top and burst, giving us boiling action.

During the boiling process, the temperature remains constant. But **why** does the temperature remain constant when we are constantly supplying heat? There is energy involved in the phase change itself. It **requires energy to break up the intermolecular forces** holding the liquid water together. (This energy is called the **heat of vaporization**.) So during boiling, the heat added to the system is used up by the boiling process. The temperature can't rise again until **after boiling is completed** and only water vapor remains.

Boiling is an **endothermic** process - it cools the environment. The environment in this case is the beaker and the remaining water. Since the beaker is being constantly cooled by the boiling of water, the temperature **does not** rise above the boiling temperature.

This seems counterintuitive at first, but consider this: Your body cools itself by sweating. The water secretions don't cool you off directly - they're at the same temperature your skin is! Sweat doesn't cool you off until it evaporates. Since the process of evaporation (vaporization) is **endothermic**, the surroundings (your arms and legs, etc.) are cooled. This is also why you feel hotter on a humid day. On humid days, there is **already** a lot of water vapor in the air, so evaporation from your skin is **slowed**. You feel warmer because your body can't cool itself off as fast as it normally would.

How about condensation? Take our discussion of vaporization and flip it around. Vaporization requires energy. Condensation releases energy. When water vapor condenses on the side of a cold glass, the glass is warmed. The liquid state is a lower energy state than the gas state. Energy is released when attractive forces begin to hold the molecules together in the liquid state.

General observations about phase changes

What happens during the transition:

- In condensation, freezing, and deposition, intermolecular forces are formed or strengthened.
- In vaporization, melting, and sublimation, intermolecular forces are weakened or broken.

Heat:

- Processes that go from a lower energy state to a higher energy one are **endothermic**. Examples of these are vaporization, melting, and sublimation.
- Processes that go from a higher energy state to a lower energy one are **exothermic**. Examples of these are condensation, freezing, and deposition.

Temperature:

- During a bulk phase change (a liquid boiling, a liquid freezing, etc.), the temperature remains constant. This is so because the phase change itself involves an energy change. In the case of freezing, the phase change releases heat, keeping the freezing liquid at a constant temperature. In the case of boiling, the phase change absorbs heat, keeping the boiling liquid at a constant temperature.

Summary

In this note pack, we've discussed the three phases of matter and changes from one phase to another. You should now have a picture in your mind about how each phase is put together. You should also be able to visualize what happens during a phase change and why heat is involved. Remember that we are dealing with attractive forces between the molecules.

We dealt with vaporization in depth. Apply the same logic to the **other** phase changes on your own. The details are very similar.