Chapter 13: Solids, Liquids, and Gases

Musician Bootsy Collins contains all three states of matter, plus large quantities of the yet uncharacterized state of matter known as "funk."

http://commons.wikimedia.org/wiki/File:Bootsy_Collins.jpg

Section 13.1: The States of Matter

Everything is made of one of the three states of matter. Either something is a solid like the Washington Monument, a liquid like that stuff that collects at the top of yogurt if you leave it in the fridge too long, or a gas like the stuff your Uncle Wally releases in large quantities at Thanksgiving dinner. Some things, such as a flaming toddler, contain more than one state of matter.[1]

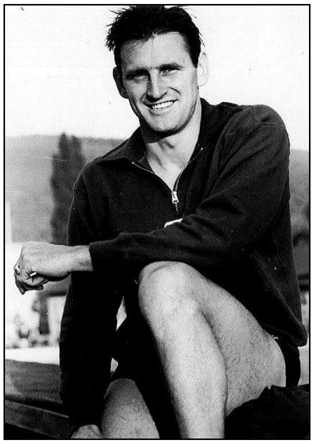

Figure 13.1: *Uncle Wally. Though he seems like a handsome and athletic guy, he's a little gassy after Thanksgiving dinner. He also eats all the pie, drinks too much, and makes the occasional racist comment about some ethnic group you've never heard of. You've really got to stop inviting that guy to family events.*

http://commons.wikimedia.org/wiki/File:Uncle_Wally.JPG

The properties of each phase of matter are directly related to their structure on an atomic/molecular level. The properties of solids (i.e. that they're hard, have a fixed shape) are due to the fact that the particles in a solid are locked into a fixed arrangement. The properties of liquids (i.e. they are squishy and don't have a fixed shape) are due to the fact that the particles in a liquid are weakly attracted to one another by something called intermolecular forces. The properties of gases (i.e. they expand and have no fixed volume or shape) is due to the fact that the particles in a gas really don't interact with each other and all and fly all over the place.

Section 13.2: Gases

As I mentioned in the last paragraph, the properties of gases are determined by the way that the particles all kind of fly all over the place. This idea, that the motion of gas particles are responsible for the way that they behave is known as the **kinetic molecular theory of gases**.[2]

The kinetic molecular theory of gases is a mathematical model that's intended to quantitatively describe the behavior of gases. Unfortunately, the math that goes into such a model is really hard to do, and requires annoyingly huge quantities of computing power. Because we usually don't feel like buying such gigantic computers (and because the model was invented in the mid-19th century, before computers were

[1] Plasmas, also known as ionized gases, are often referred to as a separate state of matter because they have properties that differ from that of other gases. Fire, for example, is a plasma. However, since the definition of a plasma is "ionized gas", that implies to me that they're just really unusual gases. You can decide what you think for yourself, but the bottom line is that they're in kind of a gray area between a gas and something else.

[2] I'll periodically abbreviate this as "KMT", because it's irritating to have to type out "kinetic molecular theory" all the time.

even born), the kinetic molecular theory uses some assumptions about the particles in gases to make the whole thing a little easier to deal with. The postulates on which the KMT is based are:

- **Gas molecules are infinitely small**. Because the particles in a gas are all really far away from each other, the volume of the particles in a gas is really small compared to its overall volume. Since the gas molecules take up a fantastically small percentage of its volume, we can just make the assumption that they're infinitely small to make the math easier.

Figure 13.2: *A scientist at work on gas simulations using the most modern supercomputers available to her.*

http://commons.wikimedia.org/wiki/File:CER12rt.jpg

- **Gas molecules are constantly moving in a random way.** Though it may seem obvious, this just means that 1) Gas molecules are always moving around; and 2) Gas molecules aren't preferentially moving in one direction or the other. The first statement is pretty clear, and the meaning of the second statement is simply to say that in a gas the particles aren't all travelling in the same direction. Instead, the particles in a gas are travelling independently of one another, in no particular direction.

- **Gas molecules undergo elastic collisions.** This assumption means that when two gas molecules hit each other, none of the energy of motion is lost. When the molecules fly apart, they have the same amount of energy as when they hit in the first place.[3]

- **The kinetic energy of a gas molecule is related to its temperature**. In short, when a gas is at high temperature, it travels very quickly and has lots of kinetic energy. When you cool it down, the gas slows down and it loses kinetic energy.[4]

- **Gases don't experience intermolecular forces**. We haven't talked about intermolecular forces yet (we'll talk about these when we get to liquids), but the easiest definition of intermolecular forces is that they're attractive forces between molecules. Taken this way, this postulate just says that there are no attractive forces between gas molecules.

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[3] Here's something to think about: If the first postulate says that gas molecules are infinitely small, then how would they ever be able to hit each other? As far as I know, this isn't really addressed.

[4] Kinetic energy refers to the energy of motion that something has. For example, a fired bullet has more kinetic energy than one that's thrown because it's moving a lot faster. Likewise, a bowling ball has more kinetic energy than a bullet travelling at the same speed because it weighs more.

What does all of this mean? It means that, using this model (which works pretty well), gas molecules fly all over the place and don't have much to do with each other. This explains, among other things, why gases have low densities (the molecules aren't packed together tightly) and why gases can be squished (because there's a lot of space between the molecules to squish them into).

Graham's Law[5]

How can we keep track of how a gas moves around? Well, it depends on what the gas is doing. **Diffusion** is the process by which one gas mixes with another gas. **Effusion**, on the other hand, occurs when a gas travels into a vacuum containing no other gas molecules.

To illustrate the difference between diffusion and effusion, imagine Uncle Wally letting loose some of his famous monster flatulence. If he does this in the living room, which is otherwise full of air, the movement of his flatus through the room is a diffusion process. On the other hand, if Uncle Wally lets loose in outer space, the movement of his gas in space is an example of effusion.

Figure 13.3: *The example above is the reason that Uncle Wally was asked to leave NASA's astronaut program in the mid-1990's.*

http://commons.wikimedia.org/wiki/File:Atlantis_taking_off_on_STS-27.jpg

The KMT tells us that all gas molecules, regardless of environment and makeup, have the same kinetic energies at the same temperature. What this means to us is that heavy molecules will tend to move slower than light ones when they're at the same temperature – this is because light molecules have to move faster to achieve the same kinetic energy as a heavy one.[6]

Because we love equations, we can relate the speeds of two gases to each other using Graham's law:

$$\frac{\text{Rate of gas A}}{\text{Rate of gas B}} = \sqrt{\frac{\text{molar mass of gas B}}{\text{molar mass of gas A}}}$$

The left side of the equation is making reference to how many times faster gas A is moving than gas B, and the right is just referring to the ratios of the molar masses of the two gases. Again, we'd expect the particles of the light gas to move quicker than that of the heavy gas to achieve the same amount of kinetic energy.

[5] Often referred to as "Graham's Law of Diffusion" or "Graham's Law of Effusion."

[6] To illustrate this, imagine two vehicles hitting the side of your house. In order to do the same amount of damage as a big rig (i.e. put the same amount of kinetic energy into the house), a motorcycle will have to drive a whole lot faster than the big rig does.

Let's do a sample problem: How much faster do nitrogen molecules diffuse through air than oxygen molecules? To solve that, plug in the molar masses of each compound into the equation to find that:

$$\frac{\text{rate of } N_2}{\text{rate of } O_2} = \frac{\text{molar mass of } O_2}{\text{molar mass of } N_2} = \frac{32.00 \text{ g/mol}}{28.02 \text{ g/mol}} = 1.142 \text{ times faster}$$

Focus on Technology

Figure 13.4: *Gas diffusion processes were used to purify the uranium that was used in the atom bomb that was dropped on Hiroshima in August, 1945. I'm sure Thomas Graham would have been proud to see his work used in this way.*

http://commons.wikimedia.org/wiki/File:AtomicEffects-Hiroshima.jpg

Section 13.2: Intermolecular Forces

Covalent molecules are weakly attracted to one another via something known as **intermolecular forces**. These intermolecular forces are different than the electrostatic interactions in ionic compounds, covalent bonding, or metallic bonding, because they happen *between* molecules in a compound and not *within* the molecules of a compound.[7]

There are three intermolecular forces. Let's have a look!

- **Dipole-dipole forces** are attractive forces between polar molecules. If you think of polar molecules as being little magnets, it makes sense that polar molecules will be attracted to each other.

- **Hydrogen bonds**[8] are extra super strong dipole-dipole forces that occur in molecules with H-O, H-N, and H-F bonds. Because these molecules are extremely polar due to the high difference in electronegativity between these two molecules, and because the lone pairs on the O, F, or N atoms can get really close to the hydrogen atom on another molecule, these forces are way stronger than regular dipole-dipole forces. You can see how hydrogen bonds work by checking out Figure 13.5:

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[7] Ionic compounds and metals don't actually have molecules, so this section doesn't apply to them.

[8] Hydrogen bonds aren't actually bonds. Back in the old days, people thought they were, so they named them hydrogen bonds. However, keep in mind that these are really intermolecular forces, and are way weaker than actual bonds.

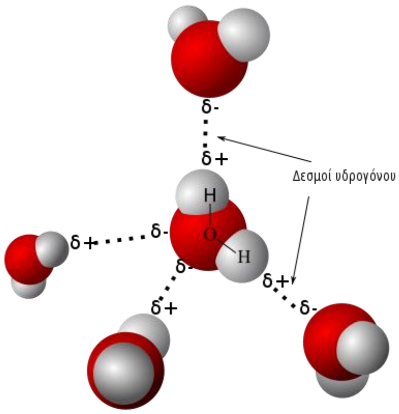

Figure 13.5: *Hydrogen bonding in water. The phrase written in Greek says "hydrogen bond." Or at least I'm assuming so. I got this picture off of the Internet because that's where you can get public domain images for free. The downside of this is that you end up with weird stuff like Greek writing in the diagrams. Still, it's better than paying for a diagram.*

http://commons.wikimedia.org/wiki/File:%CE%94%CE%B5%CF%83%CE%BC%CE%BF%CE%AF_%CE%97_%CF%83%CF%84%CE%BF_%CE%BD%CE%B5%CF%81%CF%8C.svg

- **London dispersion forces**[9] are attractions caused by temporarily induced dipoles in nonpolar molecules. How can a nonpolar molecule be polar? Let's take a look:

  1. Random movements of the electrons in a nonpolar molecule cause it to be temporarily and weakly polar. This is because there are more electrons on one side of the molecule than the other.
  2. This polar molecule causes the electrons to shift in an adjacent molecule, making it polar, too.
  3. These two molecules are now polar, so they stick to each other in the same manner of dipole-dipole forces.
  4. It doesn't take long for the electrons to re-randomize, making the molecules nonpolar and removing this weak attraction.

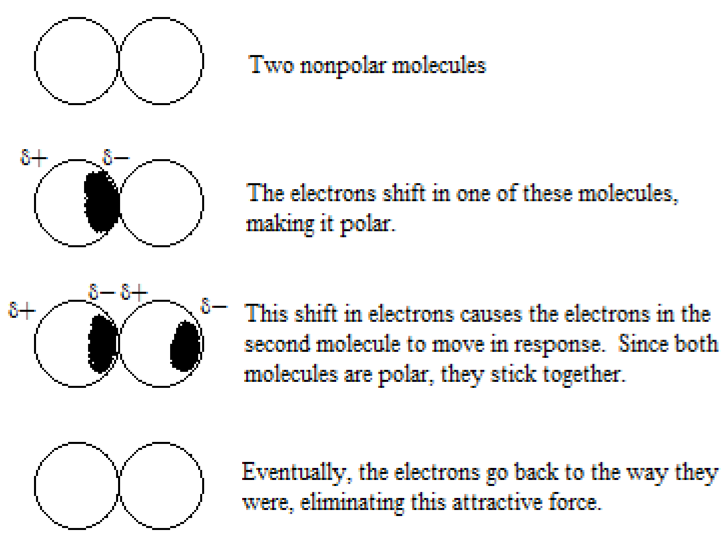

Figure 13.6: *How London dispersion forces work.*

[9] The term Van der Waals forces is usually used synonymously with London dispersion forces, though there are some differences between the two that we're not going to worry about.

Great Men of Science

Figure 13.7: *Fritz London was the guy who came up with the idea of dispersion forces. Though he only lived to be 54, he was generally considered a kick-ass dude during his lifetime. Fun fact: Dr. London didn't work with graduate students, probably because they're a bunch of beer-drinking lazy mooches (reference needed).*

http://commons.wikimedia.org/wiki/File:London,Fritz_1928_M%C3%BCnchen.jpg

Section 13.3: Solids and Liquids

Liquids and solids are states of matter in which the attractions between the molecules, ions, or atoms keep them from moving apart from each other.

What's the Deal with Solids?

In solids, the attractions between particles may be intermolecular forces, the electrostatic interactions in ionic compounds, covalent bonds, or metallic bonds. Let's have a look at the types of solids and how they differ from each other:

- **Ionic solids:** We talked about these crystalline solids a lot a few chapters back. Go look it up, lazybones.
- **Molecular solids:** These are crystalline solids in which molecules are stuck to each other via intermolecular forces. The melting and boiling points of these compounds will depend largely on how strong these forces are, as well as how big the molecules are.[10] Examples include sugar and ice.
- **Covalent network solids:** Crystalline solids held together by a bunch of covalent bonds. Though there are exceptions, these solids have very high melting and boiling points and tend to be very hard. Examples include quartz, diamond, and silicon.
- **Metallic solids:** Remember that chapter on ionic compounds? Well, it talked about those wacky crystalline solids, the metals, at the end. Go look it up, lazybones.

[10] Dispersion forces tend to be stronger in large molecules, due to the larger area over which these interactions can occur.

- **Amorphous solids:** Unlike crystalline solids, there's no long-range order in an amorphous solid. Amorphous solids contain compounds as varied as glasses, rubbers, and plastics.

Figure 13.8: *"Kiki" from the kids show "The Fresh Beat Band" contains many amorphous solids. Though you wouldn't know it from this picture, the actress who plays Kiki, Yvette Gonzalez-Nacer is really, really hot. Seriously, check out her website: http://www.yvettegonzaleznacer.com.*

http://commons.wikimedia.org/wiki/File:Fresh_Beat_Band4.JPG

The Magical World of Liquids

Liquids are the state of matter in which the particles are attracted to each other enough that they don't fly apart (like a gas) or stick in place (like a solid). Liquids generally have lower densities than solids,[11] but have much higher densities than gases. As a result, if you have a chunk of a solid, it will grow in volume when it melts, and the molecules will go over the place when you boil it.

Figure 13.9: *"Kiki" from the kids show "The Fresh Beat Band" contains many liquids. As mentioned before, the actress who plays Kiki, Yvette Gonzalez-Nacer, is really hot.*

http://commons.wikimedia.org/wiki/File:Fresh_Beat_Band4.JPG

Some of the cool properties of liquids include:

- They can't be compressed. The particles in a liquid are so close together that it's really hard to cram them together any further.

- They can flow. The particles have enough attraction to one another that they stick together, but not so much that they won't flow all over the place.

[11] An exception to this is ice, which is less dense than water. This is why ice floats in your soda instead of sinking.

- They have viscosity. Viscosity is a fancy word for "thickness," and the stronger the forces between particles, the higher the viscosity.

- They have surface tension. The stronger the intermolecular forces of a liquid, the higher the surface tension. Surface tension is the property that allows those water bugs to walk over the surface of water, totally freaking me out.

- They can experience capillary action. Capillary action is the process by which some liquids are pulled up a thin tube – the thinner the tube, the higher the liquids will be pulled. Though scientists will tell you that this is due to adhesion and cohesion, this process is actually caused by witchcraft.

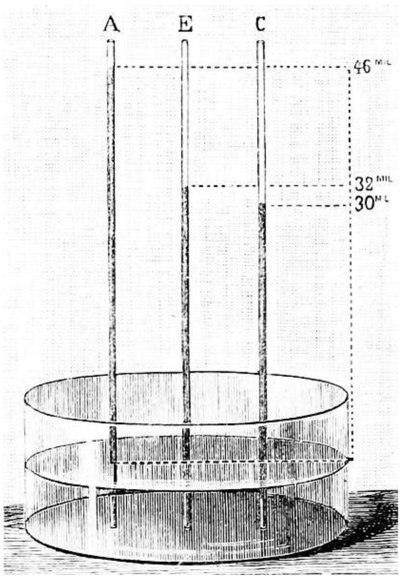

Figure 13.10: *This diagram explains how demons from the underworld are the main course of capillary action. At least, that's what I assume it explains. I would have read the caption, but I didn't feel like it.*

http://commons.wikimedia.org/wiki/File:PSM_V35_D624_Capillary_action_between_water_and_air_ether_or_camphor.jpg

Section 13.4: Phase Changes and Their Corresponding Diagrams

If you've ever sat in a room for several hours, just staring at a piece of ice as it melts, you know that compounds can undergo changes from one phase of matter to another.[12] It turns out that we can go between all three phases of matter, as discussed here:

- **Melting**: This is when you turn a solid into a liquid. This is done by adding enough energy to disrupt the forces between the particles in the solid, but not so much that the particles in the solid fly apart into a gas. An example of something melting is the example of ice above.

- **Freezing:** This is when a liquid turns into a solid. As you suck the energy out of a liquid, there's less energy available for the molecules to overcome the forces between the particles. As a result,

[12] You probably also have a mental disorder of some kind. Ask your doctor about Uncrazify, the latest anti-insanity medication from DeLoc Pharmaceuticals.

the particles lock into place as a solid. This is what happens when you turn water into ice by cooling it down.

- **Vaporization**: This is when you turn a liquid into a gas by adding energy to it - this added energy breaks apart the forces between the particles, allowing them to fly wild and free as gases.[13] One example of vaporization is evaporation, which occurs when only a few molecules of the liquid have enough energy to become a gas – the pressure of the gas caused by the evaporation of these particles is called its vapor pressure. As the temperature of the liquid increases, evaporation increases and the vapor pressure of the liquid increases, too. When the vapor pressure of the liquid increases to where it's the same as the atmospheric pressure, the liquid boils. An example of evaporation is when your goldfish dies because his bowl dried out. An example of boiling is when you turn water into steam while cooking something that contains water.[14]

Spotlight on Science Stuff

Figure 13.11: *You can see how the vapor pressure of a liquid increases as its temperature increases using the example of your shower. When you take a cold shower, there's very little water vapor in the air because there's not much energy around. However, if you take a hot shower, the water molecules have more energy and can vaporize more easily – this increased amount of water in the air is why your bathroom mirror gets all steamy.*

http://commons.wikimedia.org/wiki/File:GuentherZ_2010-01-24_0059_Brennpunkt_Museum_der_Heizkultur_Duschkabine.jpg

- **Condensation**: When you cool a gas enough, the particles no longer have enough energy to overcome the forces between them. At this point, the gas condenses into a liquid. An example of condensation is when water vapor makes water build up on the side of a cold drink on a hot day.

- **Sublimation**: Sublimation is when a solid turns directly into a gas. This occurs when the vapor pressure of the liquid phase of the material is so high that the material goes directly from a solid to a gas without first turning into a liquid. You know how ice cream gets all disgusting and rubbery after it's been in the freezer for a long time. That's because the water sublimed out of it and only the rubbery other stuff is left over.

[13] This happens to a lesser extent when something evaporates. At temperatures lower than the boiling point, some of the particles in a liquid have enough energy to vaporize, causing them to evaporate. As the temperature increases, more of the particles

[14] I don't do a lot of cooking.

- **Deposition**: When a gas turns into a solid. This happens when the gas loses energy and goes from the gas to the solid state. You know how ice cream grows ice crystals when it gets rubbery in the freezer? That's because the water was deposited as ice crystals.

Phase Diagrams

Let's say that, for some reason, you want to know what phase of matter a substance will be under some conditions of pressure and temperature. If you want this info, you need a phase diagram. Let's take a look at one of these fancy diagrams so we can learn how it works:

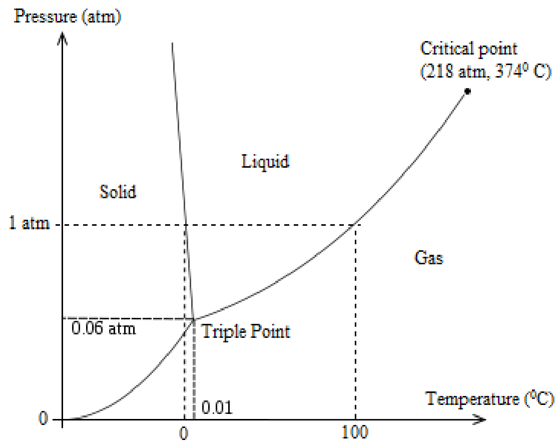

Figure 13.12: *The phase diagram of water.*

Image Copyright © 2007 Petr Sládek alias slady

Some of the main features of phase diagrams include the following:

- The lines: Each line that marks the border between two phase changes denotes the conditions under which both phases of matter can stably exist.[15]

- The **normal melting point**: The temperature at which the compound melts at a pressure of one atmosphere. In this diagram, the normal melting point is 0^0 C.

- The **normal boiling point**: The temperature at which the compound boils at a pressure of one atmosphere. In this diagram, the normal boiling point is 100^0 C.

- **Triple point**: The conditions of temperature and pressure at which all three phases of matter can stably exist. For water, the triple point is 0.06 atm and 0.01^0 C, which is why you've never seen all three phases of water in equilibrium.

- **Critical point**: The conditions of temperature and pressure past which it's impossible to distinguish between the liquid and gas phase of the material. This occurs because the material has too much energy to stick together (which is true of gases) but is crammed so tightly together that intermolecular forces between the particles are strong. Under these conditions, the material is said to be a supercritical fluid.

[15] The phases are said to be in equilibrium.