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We've seen how the idea of boiling, the observation of boiling, depends on the vapor pressure of the liquid pushing against the atmosphere, and that we can describe vapor pressure as an equilibrium between the liquid and vapor phases of a substance. And again, when that pressure becomes enough that it can push away the atmosphere, then boiling occurs. So we describe that on a graph, again, the following way. We look at the vapor pressure of a substance and how it depends on temperature. And again, we see this general dependence that, as we increase temperature, the vapor pressure increases. And finally, where that vapor pressure curve intersects 1 atmosphere, that's the point where we would have boiling, if we were at 1 atmosphere.

Well, let's play with this idea for a moment. What if we were not at 1 atmosphere? What if we were, for instance, in Denver, where Gordon lives, and the atmospheric pressure was .9 instead of 1? If we were in Denver, it would be .9, unless we were Gordon, in which case we'd be living in a vacuum. Anyway, if the atmospheric pressure is .9, we can see that the vapor pressure needed to boil is actually lower now than the temperature needed at 1 atmosphere. So, once again, here is the temperature that we need at 1 atmosphere pushing down on the liquid, here is the temperature that we need at .9 atmospheres. So what we find is, at higher altitudes, where there's less pressure pushing down on the liquid, we don't need as much temperature to reach the point where the vapor pressure equals that external atmosphere. We have a lower boiling point, in other words. You look on the sides of cake boxes and so forth and it tells you how to adjust for that phenomenon. If you cook spaghetti in Denver, you need to cook it a little longer, because the boiling temperature is actually going to be somewhat lower that the boiling temperature would be if we're at sea level.

Now, let's ask one other question. Suppose we're back here at 1 atmosphere again. Is it ever possible for the vapor pressure of a liquid to actually get above the pressure of the atmosphere? And the answer is yes. In fact, we call this super heating a liquid. Boiling is a kinetic phenomenon. What I mean by that is it depends on how fast, how rapidly the molecules can form these bubbles. And sometimes, when you don't have a good place to nucleate the bubble, and what I mean by that is the bubble needs to have its birth point some place. The bubble needs to start, and then it can rapidly expand. But sometimes that starting point can be very difficult to achieve. And if you have a very, very smooth surface on your container, very oftentimes you can raise the temperature to a point where the vapor pressure is actually above the atmosphere and you still don't have boiling. That's a very dangerous situation then, because your system wants to boil, but it can't. It can't seem to get off the ground. Where this shows up, for instance, is in a microwave oven. If you have a very, very clean coffee mug, let's say, you have to be careful, because you can raise the temperature well beyond boiling and not see any boiling until you drop you tea bag or your spoon in, and then the whole thing just takes off. So maybe some of you have had that experience. Anyway, you have to be very careful about super heating the liquid.

Okay, now, let's turn our attention now to looking at boiling points for different molecular materials, different chemical substances. And let's see if we can make some predictions on what's going to happen. We've actually looked at this a little bit before. In fact, let's return to the data that we've seen and see if we can make sense out of it. We said, in general, that, as we increase the forces between molecules, get more stickiness between those molecules, that we have to go to higher temperatures to make those substances boil. Now we can really make sense of why that would be. We consider the vapor pressure curves for several materials, so this graph is now the exact same graph that I showed you before. We're still looking at pressure as a function of temperature, and this is vapor pressure of a substance, but now we're comparing vapor pressure of several different materials.

So let's introduce our cast of characters here. We all know what water looks like. Mercury, again, is straightforward. This is just elemental mercury. But remember, we described mercury has having very strong attractive forces, even stronger than the forces between the water molecules. Diethyl ether, again, is a material we've talked about a little bit before, but it has this basic structure. The important thing to point out about it – I've got this kind of chain, and then I've got hydrogens off of my carbons here. So, if you look at the graphic here, it'll show you a better picture than my little scratching here. But the important thing for us to remember about ether is that it has much, much lower polarity than does water. So the attractive forces between ether molecules are much lower. Finally, there is methane on this chart, and we know that methane is a completely nonpolar material.

So if we ask, “Okay, what's the difference between something like methane and water, for instance?” Something very polar, strong attractive forces versus methane. We know that the forces are much higher, and so our vapor pressure at a given temperature is expected to be much lower, because what's holding the molecules in the vapor phase, that force is a lot higher. So, at a given temperature, there's going to be many fewer molecules that can get into the gas
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The stronger those forces are, the lower the vapor pressure will be. The higher the temperature will need to be then, in order to raise that vapor pressure up to the point that it equals the external atmosphere pushing against it. So there's the connection then between understanding the molecular level and, in this case, connecting that to the macroscopic quality of boiling point.