Liquids: Crash Course Chemistry #26

Crash Course: Chemistry
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Hank Green:
At the risk of sounding like I'm on drugs, aren't liquids like, super weird? Like, the idea that there is stuff out there that doesn't have a shape? How did they get away with not having a shape?!

And there aren't that many of them. Like, name a liquid. My first thoughts were juice, milk and blood, and it's weird that I went from milk to blood, but not that weird! They're both produced by the body... and both were consumed in great quantities by the Mongols...

[Mongoltage]

I've gotten off-topic; my point is that all those things are just stuff dissolved in water. The only pure non-water liquid that the average person in pre-modern times would run into would be liquid fats, produced only by plants and animals.

Liquids are super weird and super rare. Only two of the hundred-something known elements are liquid at room temperature -- bromine and mercury -- though there are a few that are nearly liquid at room temperature. Check this out; this is a little nugget of gallium that I got on eBay. Number 31 on the periodic table, gallium is a non-toxic brittle metal, not very useful. When it is 10 degrees above the temperature of this room, it will melt.

Don't believe me? Well, you probably do, but if you don't, if I wait long enough, this will actually melt in the palm of my hand, but I don't want to wait for that, so what's warmer than the palm of my hand?

Everyone, I'm going to put this gallium in my mouth.

[begins inserting the plastic-wrapped gallium into his mouth]

It's cold right now. I really hope I don't accidentally break the Saran wrap.

It's gonna take a little while for all the heat from my mouth to get into that little nugget, but every solid has a temperature and a pressure at which it will melt and become a liquid -- diamonds are liquid at 3,600 degrees Celsius -- and every gas, likewise, has a temperature and a pressure at which it will condense into a liquid. Titan is so cold that methane is liquid on its surface. It falls as rain and runs through riverbeds into methane seas.
Also, I should've mentioned that I'm a trained professional idiot. Do not do this at home.

This is an extremely weird feeling. It's starting to liquefy in my mouth, and it's like water, but way heavier. It's like, rea--it's metal, so it's heavy. Okay, I think, I feel like it's totally done now.

[removes the plastic pouch of molten gallium]

So there we have it. Liquid gallium. Suddenly, the power of my body and it becomes how liquids are really weird, right?! Like, gallium, what the heck are you even thinking?

[Intro sequence]

=====Intermolecular Forces=====

So it's time now to actually figure out for ourselves what's happening at a particle level that makes liquids liquid. First, we have to consider that no molecule is an island. Molecules and atoms interact with each other within substances. It's happening because of what we call "intermolecular forces." They're weaker than the forces that cause ionic or covalent bonds between atoms. That's why molecular substances like ice and water can be physically broken or separated into portions, while only a chemical reaction can break apart the molecules themselves, like how water can be split into hydrogen and oxygen using electricity. But, do not get me wrong, here. I'm not saying that intermolecular forces aren't important. On the contrary, liquids and solids could not exist without them.

There are basically two main types of intermolecular forces. There may be three, depending on how you count them.

=====London Dispersion Forces=====

The first is called "London dispersion forces," which are most common in London. I'm lying to you; they're named for the physicist Fritz London. These are the weakest of the intermolecular forces because they are based on the temporary clustering of electrons that takes place inside of molecules. They are most notable in noble gases like helium and neon, and non-polar molecules like carbon dioxide or any kind of oil. And it's not that London dispersion forces are especially strong in these substances, it's just that there's nothing else holding them together, so even these tiny forces are noticeable in them.

So what makes them work? Occasionally, in the course of their movements around their individual nuclei, electrons in a molecule become clustered together. The region where they cluster briefly acquires a slight negative charge, while other regions in the molecule experience a slight positive charge. Those positive and negative charges, although small, are strong enough to affect other nearby molecules. The positive side of the molecule attracts electrons on adjacent molecules while the negative side repels any electrons that are close to it.

These individual attractions are extremely weak and don't last very long because the clustering of electrons itself is short-lived, but the overall effect of this force is super important. London dispersion forces are the only thing that makes non-polar substances like methane and even otherwise non-reactive stuff like helium stick together well enough to condense from gases into liquids. And we should all remember the lesson of London dispersion forces: small, weak things can have important jobs, too! Which makes it sound like I think London is small and weak. I don't, London is awesome.

=====Dipole-Dipole Forces=====
The other main type of intermolecular forces are "dipole-dipole forces." As you should recall from our lesson on polar and non-polar molecules, a dipole is a separation of charges, like areas of partial positive and partial negative in a polar molecule such as water. Dipole-dipole forces occur when the partial charges in these molecules attract or repel each other. The molecules orient themselves so that the attraction is maximized and the repulsion is minimized. It's like two colleagues whose very different personalities improve each other's work, like Holmes and Watson. They're a great team, even though they get on each other's nerves.

For instance, partial charges occur in hydrogen bromide because the electrons are more attracted to the bromide than they are to the hydrogen. But the negative charges around the bromines are attracted to the positive charges around the hydrogens on other hydrogen bromide molecules, and that helps to hold the molecules close to each other.

=====Hydrogen Bonds=====

Then, there's hydrogen bonding, a special type of dipole-dipole force, one that occurs on polar molecules that contain hydrogen and a strongly electronegative element such as nitrogen, oxygen, or fluorine. Because of the very small size of the hydrogen atom and the extreme attraction of its electron to those other atoms, the dipoles in these molecules are extra strong, so we actually it a bond. Really though, hydrogen bonds aren't chemical bonds in the same sense that ionic and covalent bonds are, but they are the strongest interactions that occur between molecules. Water, for example, is very well-known for its tendency to form hydrogen bonds.

These intermolecular forces, while relatively weak, are why phases of matter exist, and, in fact, it's the relative weakness of these forces that allows a substance to change from one phase to another fairly easily.

So, the molecules of a solid are just stuck in one spot, spinning and vibrating. Intermolecular forces are what's holding them there. But there are always ways to break free. What phase a material is in has a lot to do with its kinetic energy, and kinetic energy can come from heat. So, if enough thermal energy is added to a solid, the kinetic energy of the particles increases enough to allow them to overcome some of those intermolecular forces and flow more freely. This is the liquid state. Like, I can smash some ice with a hammer all day if I want to to break it into smaller and smaller pieces, but only a big enough increase in heat will change its kinetic energy enough to make it change phases and become water.

The particles in a liquid are still pretty close together, though. Both liquids and solids are known as "condensed states" because there's a lot of interaction among the molecules. But, if the particles in a liquid acquire enough kinetic energy to escape the surface of the liquid, usually by absorbing thermal energy, those freed particles comprise a gas. And the particles in a gas experience intermolecular forces, too, it's just that the forces are far weaker than in the condensed states, so they allow them to spread much further apart and interact very little.

Because the gas particles are so far apart, gases have low density and are highly compressible. Conversely, the particles in a solid are very close together, so they have high density and can be compressed only a teeny tiny bit, even under huge amounts of pressure. Obviously, liquids fall between these two extremes, but they're much more similar to solids than to gases. Most liquids are less dense and more compressible than their corresponding solids, but only slightly so. Makes sense when you think about it, because the amount of energy it takes to move particles just far enough apart that they can flow past each other is much less than the energy needed to separate them completely.

=====Cohesion=====

Now, the attraction or "cohesion" between the molecules in liquids give them some important properties. For one thing, because all of the particles are pulled toward the others, they tend to merge into the most compact shape possible, which, just floating in the air, is a sphere. But when a liquid can't form a perfect sphere because it's not in space, like when it's resting on a surface or filling a container, its free edge will curve as much as possible to approach that spherical state. This is easy to observe by overfilling a glass of water. Liquid will pile up on top as long as it can, until the weight of the water is greater than the intermolecular forces holding it in place.

The overall effect of this behavior--called "surface tension"--is so strong in water that small objects like paper clips can actually rest on top of it. Cohesion manifests itself in other ways, too. For example, in molecules with very large intermolecular forces--like the molecules in honey, which have tons of oxygens and hydrogens just itching to form hydrogen bonds with each other--the cohesion is so great that it makes the liquid flow very slowly. That resistance to flow is called "viscosity," and you can see it in liquids like honey, syrup, oils, glycerol.

A third effect of cohesion can be observed by placing one end of a very narrow tube, called a "capillary tube" in a liquid. If the liquid isn't too viscous, it will spontaneously rise into the tube. This is called "capillary action," and happens partly because the molecules of liquid that are outside the tube are attracted to the ones that are inside of it, so they sort of follow them in.

=====Adhesion=====

The capillary action also depends on another phenomenon known as "adhesion," or the attraction of molecules in the liquid to the container. The liquid can't rise in the tube if it's too attracted to itself. It must also be attracted to the glass, causing it to cling to the sides. If the molecules in the liquid are more attracted to the container than each other, they will form a concave meniscus or a crescent-shaped curve at the surface as we see here with water in glass. And if the molecules in the liquid are more attracted to each other than to the container, as is the case with mercury and glass, they'll form a convex meniscus.

=====Summary=====

So, once again, mystery is answered. The bizarre reality of liquids understood at the most basic level. Final question for you all, an opportunity to use your critical thinking skills: Why, do you think, are liquids so much more rare than solids and gases? Discuss in the comments. And thank you for watching this episode of Crash Course. If you paid attention, you learned that intermolecular forces--including London dispersion forces, dipole-dipole forces, and hydrogen bonds attract molecules to each other, especially in solids and liquids. You learned about some of the effects of these forces, like cohesion, where one substance is attracted to itself, and adhesion, where it's attracted to something else, and that those effects cause some unique behaviors in liquids,
like viscosity, capillary action, and surface tension.

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=====Credits=====

This episode of Crash Course Chemistry was written by Edi Gonzalez, the script was edited by Blake de Pastino, and our chemistry consultant was Dr. Heiko Langner. It was filmed, edited and directed by Nicholas Jenkins. Our script supervisor is Katherine Green, and our sound designer is Michael Aranda, and, of course, our graphics team is Though Café.