Bonding Models and Lewis Structures: Crash Course Chemistry #24

Crash Course: Chemistry
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So models are cool. [on screen: animation of swimsuit models on a beach] Not... those models. I actually find that particular cultural institution pretty peculiar. I'm talking about the kind of model that lets you experience things you otherwise couldn't experience. Maybe it's something that's too big to see all at once, like our solar system, or too small, like a cell, or something that unfortunately doesn't exist, like the Millennium Falcon. [Michael Aranda walks behind Hank with a model of the Millennium Falcon] Or something that would just be way too dangerous to have in your office, like a trebuchet. Sometimes it is difficult, maybe even impossible, to really understand things fully without making a model.

But models don't have to be three-dimensional objects. In scientific terms, a model is anything that represents something else, whether it's physical or conceptual. In the same way that musical notes on a page are a model for music, the same goes for chemistry. Chemists use many models, or simplified versions of reality, to help them understand atoms and their interactions. Because the universe is weird.

This ball-and-stick model is one idea of a molecule — perfectly spherical atoms connected by clearly defined bonds. This version of a molecule is a great way to begin understanding chemical bonds. But I gotta take you beyond beginners' stuff to understand some models that explain bonds with more delightful complexity. Taking in the details requires models that are more complicated, but also more fascinatingly awesome. So while this episode of Crash Course Chemistry won't involve any Brazilians in swimsuits, I promise you will not be disappointed.

(Intro)

=====Models (1:32)=====

It turns out that chemical bonds are not like little sticks at all. Bonded atoms — molecules — are more like groups of atoms hanging out just, like, close to each other because that's where their energy is minimized. So if I tossed all these models of atoms into the air and let them fly apart from each other, it's cool and fun for me, but those little balls would also have a lot more energy, and that's not the ideal situation for an atom, now is it? In reality, the only thing connecting two atoms together in a chemical bond is a bunch of electrons, and they don't sit still between the atoms gluing everything together. Instead, they're in constant motion zooming all around the nuclei in a somewhat predictable pattern. In a covalent bond, the bonding electrons spend most of their time between the nuclei, and the nuclei stay close together because they're attracted to the electrons.

And honestly, this concept of electrons holding everything together is itself just another model, an idea that represents molecules in a way that can be visualized, and it's a more accurate representation of reality than the ball-and-stick model. But the ball-and-stick model isn't useless. It helps us visualize and understand many important things about molecules. It also looks pretty cool. Imagine if, instead of making generalizations about how chemicals behave, like "water dissolves salt", you had to memorize every single behavior of every single substance. No one would ever get anything else done in their lives. We generalize so that we can free our minds to do bigger and better things.

So make no mistake, some models are great even though they're oversimplified to the point of being outright lies. It's important, in fact, to realize that all models are imperfect to some extent. I mean, think of the human models that people compare themselves to. Do you think that women in underwear catalogs and the guys in those black-and-white cologne ads look like that on Saturday morning after a long night of doing whatever models do on Friday nights? In fact, if a scientific model were a perfect representation of reality, it would cease being a model and become reality. So, in addition to understanding how a model represents reality, you also have to recognize the ways that it doesn't represent reality so that you don't base a bunch of incorrect assumptions on it.

=====History of the Chemical Bonding Model (3:28)=====

Unfortunately, sometimes models aren't merely oversimplified, sometimes they are just downright wrong, and the chemical bonding model is no exception. Over centuries, it's been updated as experimental results have provided information about how the universe actually works. Early scientists, including Isaac Newton, thought that atoms combined because they were literally sticky, or because they had tiny little hooks on them like a kind of Velcro that held them together. That was their bonding model.

In the 19th century, chemists like Berzelius discovered positive and negative charges associated with chemicals in certain situations. He and his contemporaries theorized that this was the force holding molecules together. That's a much better model than the first one, but still not entirely accurate, because they thought that atoms more or less attracted each other like magnets. Only after the discovery of electrons in the 1890s could chemists begin to understand the true nature of chemical bonds.

Then, in 1916, American chemist Gilbert Newton Lewis described a covalent bond as two atoms sharing electrons. Modern chemists still use this model as a simple way to represent chemical bonds on paper.

=====Lewis Structures (4:27)=====

A Lewis structure is a two-dimensional model that represents covalent bonds as straight lines and un-bonded, valence electrons, those in the outermost energy level of an atom, as dots. Inner electrons aren't shown at all, and although it was developed to explain covalent bonds, it also works for ionic bonds. In Lewis structures, bonds are formed by pairs of valence electrons, called bonding pairs, in the space between the two atoms. Pairs of electrons that are attached to only one atom are known as lone pairs.

As you probably remember, atoms are most stable when their outermost electron shells are filled. For many atoms, that takes eight electrons, so it's called the octet rule. And if you're bracing yourselves for some exceptions to this rule, well, you are correct. Tiny little hydrogen can only hold two electrons total, not eight — that'll make more sense when you watch our upcoming episode about atomic orbitals — but also elements in the third row of the periodic table and below often have more than eight valence electrons. Beryllium and boron are notorious for having weird numbers like six or 12 electrons. So yeah, the whole octet thing is really more of a gentle nudge than a rule. Let's stick to rows 1 and 2 of the table to keep it simple.

=====Ionic Bonds (5:30)=====

Let's say I want to draw the Lewis structure for sodium chloride. Sodium has one valence electron and chlorine has seven. Since sodium is a metal, this must mean that it's going to be an ionic bond, meaning that electrons are transferred. Sodium transfers its one valence electron to chlorine, creating a +1 charge on the sodium and a –1 charge on the chlorine. The two ions are attracted to each other due to their opposite charges. They're written slightly separated because the electrons aren't shared, but they're not too far apart because the ions stay close enough together to neutralize their charges. Lewis structures use a line to note covalent bonds, but we're not doing that here because there are no covalent bonds. It's an ionic bond, and there's no physical contact been the ions. So that is what salt looks like from a Lewis dot structure point of view.

=====Covalent Bonds (6:10)=====

Covalent bonds: a little more complicated, and it works best if you follow some specific steps. Let's try it out with water. First, figure out how many total valence electrons are available. It doesn't matter which atoms they come from or how they're arranged before you start. Each hydrogen atom has one valence electron, and oxygen has six, for a total of eight electrons. Neither hydrogen nor oxygen has enough valence electrons to be stable, so they share electrons to make up the lack. And, just like how sharing a snack can make a new friend, the sharing greets a bond between the atoms.

Lay out the molecules and create the bonds. Both hydrogens are bound to oxygen, so the oxygen goes in the middle. Remember, in Lewis's model each bond requires a pair of electrons. So we used four of the eight available electrons to form the bonds. Now fill in the outermost energy levels. Hydrogen atoms only need two electrons total, so they're already full. Oxygen, on the other hand, needs an octet — eight electrons. So place the remaining electrons around it in pairs to complete the structure. Finally, again, Lewis structures use lines to represent covalent bonds, so put those in the place of the bonding electrons, and you have it. Water contains two covalent bonds and two lone pairs. Simple enough.

Well, it can get a little trickier. Let's do carbon dioxide, another tremendously important molecule on our planet. Carbon has four valence electrons, and each oxygen has six, for a total of 16 electrons. Both oxygens bond to the carbon, so put the carbon in the middle and make the bonds using four of the 16 electrons. All three of these atoms, they need a full octet, so fill that in next. So yay! Now all the atoms have a full octet, but take a closer look; 20 electrons are needed to fill the octet, but only 16 are available. So, what to do?

=====Double Bonds (7:45)=====

When there aren't enough electrons to fill all the octets with normal sharing, atoms have to share more of them. In this case, they form a double bond by putting two pairs of electrons, for a total of four, between each pair of atoms. All four bonding electrons count toward both atoms' octets, so we need fewer lone pairs to fill all the octets. The double bonds allow us to use only 16 electrons. Replace the bonds with double lines, and the Lewis structure is complete, with two double bonds and two lone pairs on each oxygen.

=====Triple Bonds (8:14)=====

That was a little weird, so surely we can do something simple like molecular nitrogen. Just two atoms of nitrogen bonded together, right? Well, nitrogen has five valence electrons, and there are two atoms, so that's a total of 10 electrons. Let's put in the bond and complete the octets, and we've used 14 electrons. That — we don't have that many. We can try a double bond, but that still uses 12 electrons, two more than we got, so we still don't have enough electrons to go around. Let's kick it up one more notch and make it a triple bond. That's when atoms share three pairs of electrons. And that does it. Just 10 electrons, and all the atoms have a full octet. Switch the lines to bonds, and it's done. A triple bond, and a lone pair on each atom. And this is why molecular nitrogen is really hard to break up to form, like, fertilizer and stuff. In case you're wondering, three is the maximum number of bonds. There is no such thing as a quadruple covalent bond.

=====Linus Pauling (9:03)=====

So that's the Lewis model — separate, discrete bonds formed by sharing specific electrons. It's a good model, and pretty close to the modern definition of a covalent bond, but it is still grossly oversimplified and unfortunately not 100% accurate. But here's the thing about models: even when they're partially wrong, you can build on the good parts to create even better models, and that is what Linus Pauling did. While he was in college, Pauling read Gilbert Lewis's chemical bond research, published just three years earlier. It was his model that inspired Pauling to spend the rest of his life studying the relationships between the properties of substances and their molecular structures.

After getting his PhD in physical chemistry, he traveled to Europe to study the new field of quantum mechanics with great physicists like Arnold Sommerfeld, Niels Bohr, and Erwin Schrödinger. Quantum mechanics basically involves the idea that some things, like light and electrons, are both particles and waves of energy. So Pauling applied the quantum mechanics model to chemical bonds, and this was the birth of the bonding model that we know today, which conceives of chemical bonds as a sort of overlap of atoms' individual electron clouds, rather than the simple sharing of specific electrons. We'll explore that in more detail when we talk about orbitals, but this is the electrons-holding-everything-together model that I mentioned at the beginning of the episode, and that we pretty much take for granted today.

Pauling's contributions to the model of chemical bonding made such an impact on how we understand the universe that he won a Nobel Prize for it in 1954. It may not sound totally mind-blowing today, but imagine figuring this out when your only concept of atoms was little bits of stuff like Newton, or even just a vague idea of charges in an atom like Berzelius. It's mind-boggling. Our ability to understand chemistry at all is a direct result of the models that scientists like Lewis and Pauling have provided, so to them I say thank you.

=====Conclusion/Credits (10:45)=====
And thanks to you for watching this episode of Crash Course Chemistry. I think you're all... model students. If you paid attention today you learned that a scientific model is anything that represents something else in a different way, that we often learn the most when we try to understand why things don't work the way that we expect, and that you can build new models on the foundation of old ones. You also learned that the chemical bonding model developed by the great Linus Pauling is crucial to our understanding of chemistry, and also the universe, and you learned how to draw Lewis structures. This episode of Crash Course Chemistry was written by Edi González and edited by Blake de Pastino. Our chemistry consultant is Dr. Heiko Langner. It was filmed, edited, and directed by Nicholas Jenkins, and our script supervisor and sound designer is Michael Aranda, and of course our graphics team is Thought Café.