Network Solids and Carbon: Crash Course Chemistry #34

Crash Course: Chemistry
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=====Introduction (0:00) + Sheet & 3D Networks (0:48)=====

The diamond in this ring will cut into any substance on Earth. It's the hardest, natural material on the planet. But it's just carbon, the same element that makes up the graphite in this pencil lead which is so soft, that it's intended purpose is to rub off on stuff. The atoms are all of the same element and they're all bonded together with covalent bonds, so what's the difference?

Well the atoms are arranged differently, they form different atomic networks and therefore, different network solids. In chemistry we place networks in various categories because each type of networks have characteristics that make them useful for all sorts of handy applications.

You can have a chain of identical molecules, each molecule is connected on two sides. So glycogen which is just a chain of glucose molecules is bigger than glucose but it's not super durable as substances go.

A sheet of particles like the proteins that make up silk is a more sophisticated kind of network because the particles are connected on all sides but its connections are still two-dimensional so it's only stronger in those directions.

Now, you can stack sheets on top of each other in order to build a network into the third dimension but then the sheets just slide around on each other and they're not much more useful or durable than they were before.

A truly three-dimensional network branches out in multiple directions, forming covalent bonds in a structure that resist forces much better than a chain or a sheet of the same material.

So in the most basic sense, the way atoms and molecules bond to form network is what make many materials what they are. It explains why diamond is so hard while graphite is so soft and how the main ingredient in sand can also be made into the silicon wafers that make our electronics. It's also the key to how we can turn this, into this.

Though before you send the kiss off email to your student loan officer I should warn you that you won't be able to do that at home.

[intro music]

=====Solid Networks (01:54)=====

Network solids are one of the three types of atomic solid we've talked about, materials that are made of individual atoms rather than molecules or ions.

The cool and special thing about network solid is that the atoms exist in, you guessed it, a network structure that is each atoms is linked to several others in various directions. Like all networks, this makes the arrangement both strong and stable. It also provide some other properties that may surprise you.

If you've seen our episode on atomic orbitals you know the way the electron in a chemical bond orient themselves, that is the way the orbitals hybridize, has a big effect on the properties of a substance.

If water's orbitals hybridize differently for example, it might not be polar and life as we know it wouldn't be possible. It's the same with network solids. Orbitals and their hybridizations make a huge difference. Thankfully though, orbitals can't be different. They are the way they are because of the laws of physics.

So first, let's look at carbon. Pure carbon can bond with itself in two different ways. In the first scenario, the outer electrons of each atom are arranged into SP2 hybridized orbitals with two other atoms for a total of three lobes in a flat trigonal arrangement.

This creates a sheet with a hexagonal structure and with one unhybridized P orbital left on each atom. These orbitals form pi bonds that merge into an extensive network on their own and this gives the structure its real strength.

=====Diamond & Graphite Network Structures (3:03)=====
You know this carbon network by its street name, graphite. The pi bonds within each sheet is really strong which allows graphite to withstand a lot of pressure. But as with any sheet-like formation, only in two dimensions. The sheets are only held to other sheets with really weak van der Waals forces. They allow the sheets to slide on top of each other so they can easily be removed layer by layer.

That's why graphite is so fantastic for writing with. When you use a pencil, layer after layer of graphite is being transferred into the paper. That slipping sheet structure also makes graphite an excellent lubricant. It can break down into an extremely fine powder of slippery platelets that almost any substance can slide easily over. Which is why it's great for things like sticky locks.

And because the electrons that form the pi bonds in the graphite are able to move from one atom to another, graphite also conducts electricity. Most of the properties of graphite, you'll notice, don't apply to the other type of network that carbon can form, diamond. Even though the graphite in the pencil is made of the exact same thing as diamond, pure carbon atoms, they're different from each other in almost every possible way. And all of these differences are simply because of the way their atoms are bonded.

In diamond, the bonding electrons in each atom are arranged in SP3 orbitals. Four lobes that are as far away from each other as possible. Again, information you should have in you from the orbital episode. Each lobe in that tetrahedral structure overlaps and bonds with one on an adjacent carbon atom creating a totally uniform, three dimensional network of carbon in every direction. So any stress on this network will be resisted by multiple atoms and multiple bonds. This stability is what makes diamonds so famously, incredibly hard.

But the downside is that this rigid structure also makes them quite brittle. In other words, they break before they bend. And like any crystal and solid they tend to break along the seam between atoms called a cleavage plane. And these are the places that the diamond cutters attack first.

And here's another way that the diamond's network structure makes it totally different from graphite. Diamonds are electrical insulators, not conductors. This is because the carbon atoms in diamonds share sigma bonds not pi bonds which doesn't allow the electrons much freedom of movement. And if an electron can't move around, it can't be used to transfer energy in the form of electricity or anything else.

But even though diamond is terrible at conducting electricity, both graphite and diamond are good conductors of heat. That's because of the strength of their covalent bonds. These type bonds require that when one atom vibrates, all the atoms around it vibrate in exactly the same way.

So when thermal energy increases in one region, it spreads quickly throughout the diamond or graphite. Even though there is some space between the sheets of graphite, the structure of the sheets is so rigid that the vibrations transfer through the whole mass. And as you may have guessed, diamond conducts heat even better than graphite because each atom spreads that energy in three directions rather than two.

Graphite has another interesting response to heat. When it's heated to very high temperatures, around 3000 degrees Celsius and placed under extreme pressure, about 15 million kilopascals, the atoms rearrange themselves from a two dimensional sheet network into a 3D network. Now that's roughly the amount of pressure of 15 metric tons on one square centimeter. Basically five elephants sitting on the heel of a high-heeled shoe.

That tremendous heat provides energy for the breaking and reforming of bonds which allows the atoms to reorganize themselves to relieve all of that pressure. And that my friends is how you can turn common graphite into diamonds. So no, you cannot do that at home. But just like with you and me sometimes what seems like an excessive amount of pressure can end up forcing changes that stabilize things in the long run.

Now it seems kinda odd but the opposite reaction, diamonds turning into graphite, is even more difficult. So much so that it's functionally impossible for it to happen naturally. At least on this planet.

Even though it is possible under those incredibly hot elephants on heel circumstances to break the covalent bonds in graphite, the activation energy required to break those bonds in diamond is so high that the rate of its breakdown is essentially zero for all earthly purposes.

The similarities and especially the differences between diamonds and graphite are an amazing display of the power of chemistry. Even though they're both made of the exact same material, the mere difference in the nature and arrangement of their chemical bonds completely changes their physical and chemical and electrical characteristics.

Also, one is pretty and one is pencil lead. They explain why diamond can be used to cut through glass and drill through rock while graphite is soft enough to write and lubricate with. And yes, diamonds are more valuable, but you could argue that graphite with its easily erasable mass producible pencils might have done more good for the world. The pencil is maybe mightier than the diamond.

Thank you for watching this episode of Crash Course Chemistry, by the way. If you paid attention, you learned that networks lend strength and widely distributed stability to network solids. You learned that both diamonds and graphite are network solids made up of pure carbon atoms but that the arrangement of those atoms in two and three dimensions respectively give them completely different properties.

This episode of Crash Course was written by Edi Gonzales, edited by Blake de Pastino. Our chemistry consultant is Dr. Heiko Langner. It is filmed, edited, and directed by Nicholas Jenkins. The script supervisor was Michael Aranda, who is also our sound designer and our graphics team, of course, is Thought Café.