Diamond and graphite

Big Manny: Alright, boom.

Did you know that diamond and graphite are more similar than you might think?

That's right.

The sparkly stuff in jewelry and the stuff in your pencils are both made of carbon.

They are allotropes, which means different forms of the same element.

We're going to be looking at the different structures of diamond and graphite.

First, let's remind ourselves of the electronic configuration of a carbon atom.

It has four electrons in its outer shell and is in group four of the periodic table.

Now in a diamond, each carbon atom uses all four of its outer electrons to form covalent bonds with four neighbouring carbon atoms.

This creates a giant covalent structure with strong bonds throughout, creating one continuous lattice with no free electrons.

Now the covalent bonds are very strong, which makes diamond the hardest known natural substance, and it gives it an extremely high melting point.

But I can't lie, it can't conduct any electricity though.

That's because it doesn't have any free electrons.

These properties make it very useful for industrial purposes like cutting tools, drilling, grinding, and even polishing.

When it comes to graphite, things are a little bit different.

Unlike in diamond, each carbon atom only bonds to three others, leaving one delocalised electron free to move.

Graphite is made of layers that are held together with weak forces of attraction, which are easy to overcome.

Layers slide over each other, which makes graphite slippery and useful as a lubricant.

It's also used to make lead in pencils.

The delocalised electrons are free to move through the structure and carry a charge.

This makes it an excellent conductor of electricity.

It can be useful for electrolysis, or the electrodes in batteries and that.

Let's look at them side by side.

Diamond with four bonds per atom and no free-moving electrons makes it hard but a very poor conductor of electricity.

Graphite with three bonds per atom and delocalized electrons is an excellent conductor of electricity.

Both have very high melting points because of the covalent bonds in the giant structures.

Now it's over to you.

I've got a quick question for you.

Which of the following is a reason why graphite is a lubricant?

Is it A, because delocalised electrons can move between layers?

B, because of weak forces of attraction between layers?

Or C, because carbon atoms have strong covalent bonds in a giant structure?

You can pause while you have a little think.

The correct answer is B.

Weak forces of attraction between layers mean they can slide easily, making graphite useful as a lubricant.

So the next time you see a diamond or use a pencil, remember they're both made of carbon.

It's their structure that makes all the difference.

You done know, diamond and graphite ting.