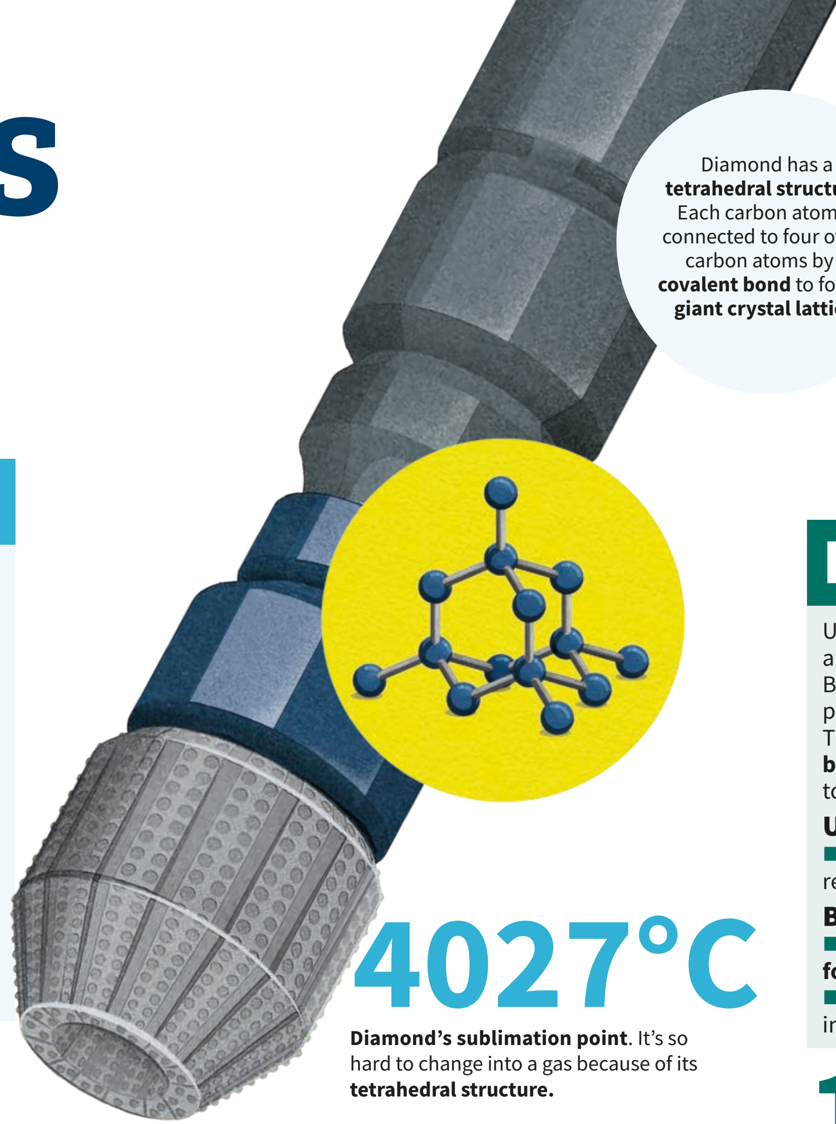
Allotropes of carbon

Some elements are able to exist in different structural forms, known as **allotropes**. Carbon does this very well because of its ability to form bonds with other neighbouring carbon atoms – something called **catenation**. The way in which carbon atoms are connected to each other makes a big difference to the **physical, chemical and electronic properties** of the material.

DIAMOND

Used for ...
- **drill bits** in oil exploration and for slicing through concrete
- **jewellery**: naturally-made diamonds are of higher purity and very expensive!

Because ...
- of its tetrahedral structure, diamond is **one of the hardest known materials**
- it has a high refractive index, light is reflected internally, so it sparkles
- all its electrons are used to create the bonding lattice, leaving none spare, it's **a poor conductor of electricity**



Diamond has a **tetrahedral structure**. Each carbon atom is connected to four other carbon atoms by a **covalent bond** to form a **giant crystal lattice**.

4027°C

Diamond's sublimation point. It's so hard to change into a gas because of its **tetrahedral structure.**

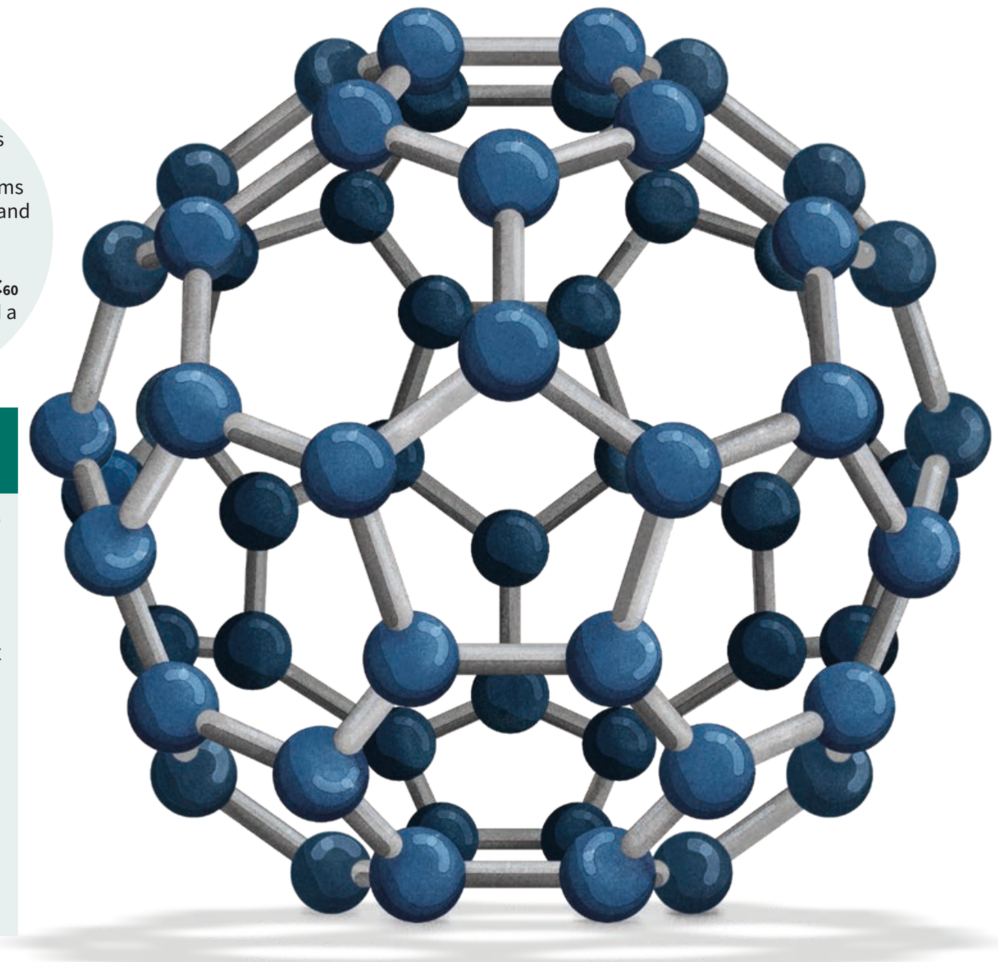
FULLERENES

Until 1985 it was thought there were only two allotropes of carbon: **diamond** and **graphite**. But scientists thought they had detected the presence of another form of carbon in space. That mysterious new allotrope is **C_{60}**, or **buckminsterfullerene**. Other fullerenes exist too, like C_{70}, as well as ellipsoids and tubes.

Used for ...
- **drug delivery** potentially – many researchers are currently working on this

Because ...
- buckminsterfullerene's **intermolecular forces** are weak, its melting point is low
- fullerenes have a sea of free electrons inside, they can **conduct electricity**

Buckminsterfullerene's **spherical structure** comprises 60 carbon atoms arranged as 10 hexagons and 12 pentagons. The same shape as a football – which is why **C_{60}** is also sometimes called a **buckyball**.



1985

A team headed by Professor Sir Harry Kroto **discovered and named C_{60}**, or buckminsterfullerene.

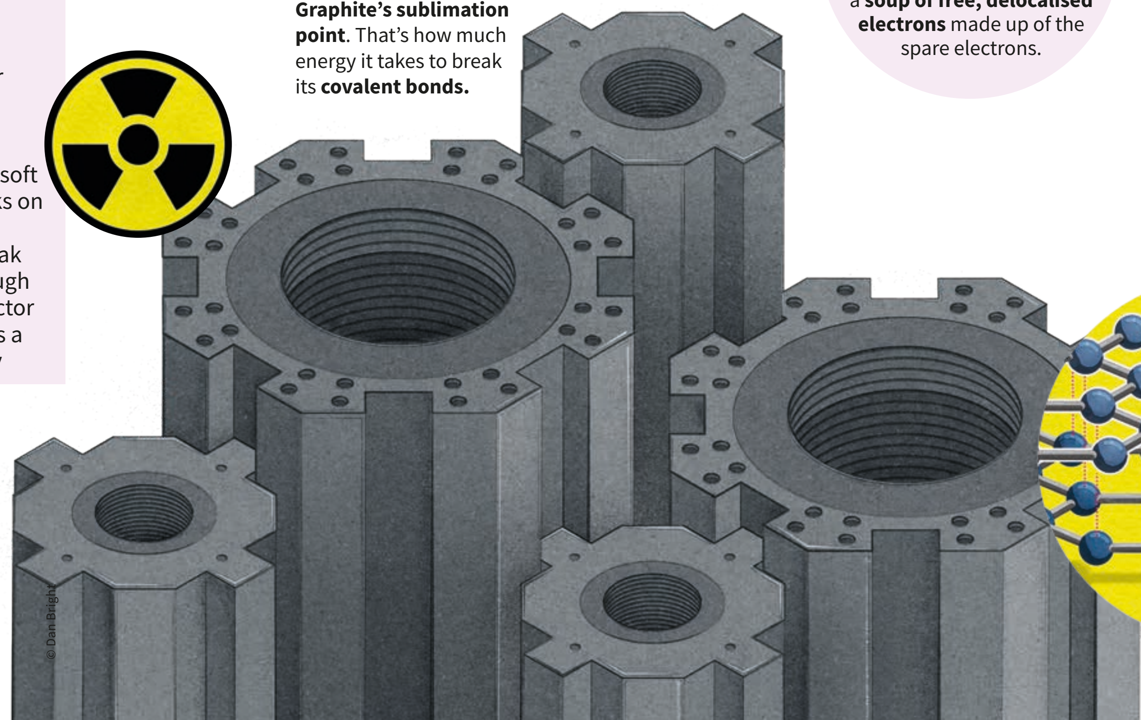
GRAPHITE

Used for ...
- **pencil leads**
- **nuclear reactor cores**, to stop or slow the nuclear reaction

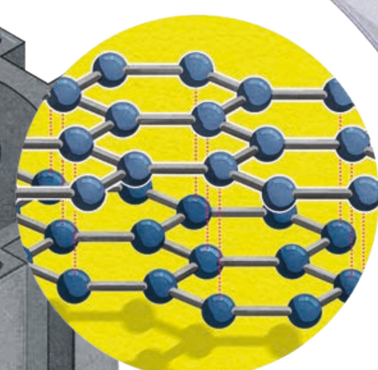
Because ...
- its **layer-like structure** makes it soft and flaky, as a pencil it leaves marks on your paper
- so much energy is needed to break the **covalent bonds**, graphite is tough enough to be used in a nuclear reactor
- of its soup of spare electrons, it is a very good **conductor of electricity**

3600°C

Graphite's sublimation point. That's how much energy it takes to break its **covalent bonds.**



Each carbon atom is **covalently bonded** to three others, leaving one electron spare. This results in **atoms arranged in flat layers of hexagons**, between which is a **soup of free, delocalised electrons** made up of the spare electrons.



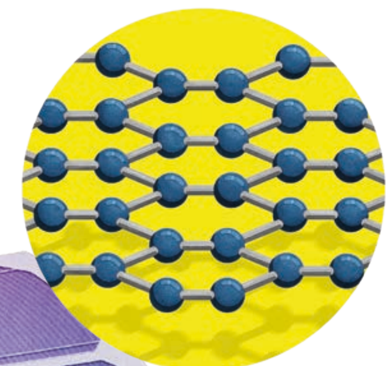
GRAPHENE

Graphene was a theoretical concept before it was isolated and studied in 2004 by Andre Geim and Konstantin Novoselov at the University of Manchester. They were awarded the Nobel prize in physics in 2010 for their discovery. It's the **thinnest, lightest, strongest, most stretchy material** we've ever created.

Used for ...
- **solar cells** that are both transparent and flexible
- **smart windows** that can control heat and light transmittance
- **electronic displays**

Because ...
- of its spare electrons, graphene is an excellent **conductor of electricity and heat**

Think of graphene as a single layer extracted from graphite. In its **hexagonal lattice**, each carbon atom is bonded with three others, leaving a spare electron.



1 g

A single sheet of graphene the **size of a football pitch** would weigh less than 1 g!

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