Cambridge (CIE) IGCSE Chemistry

Your notes

Giant Structures

Contents

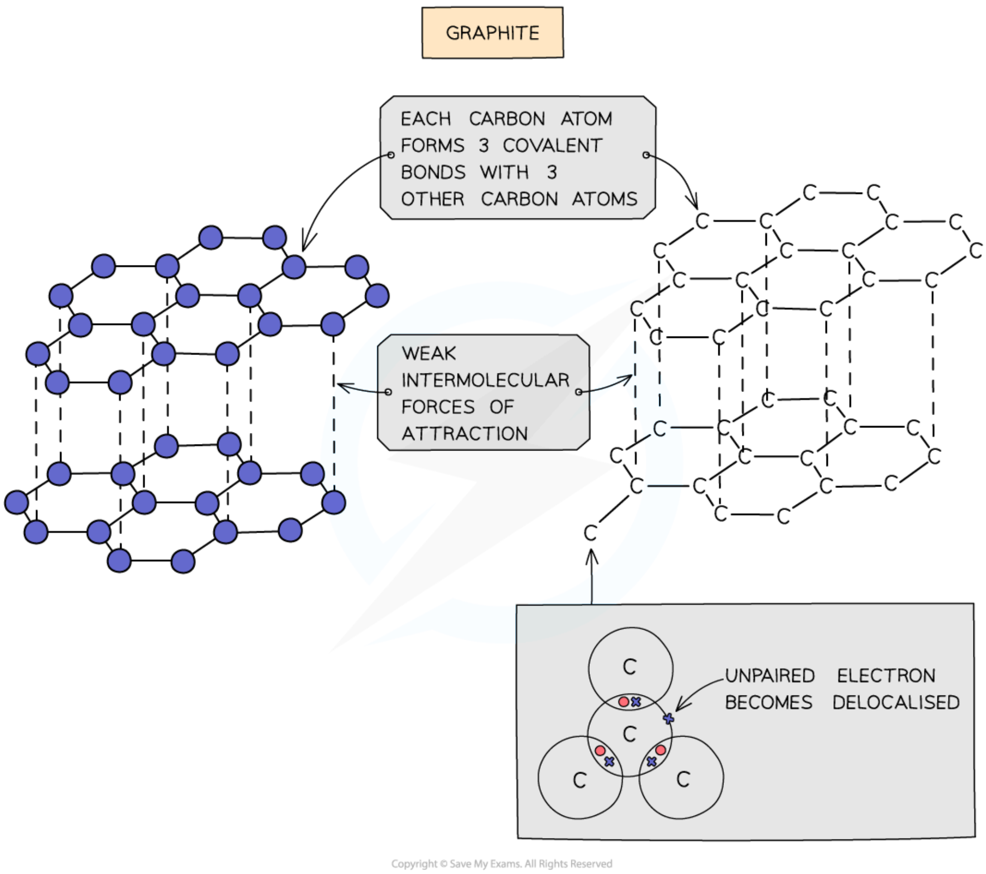
Your notes

Structure of graphite & diamond

What is the structure of graphite?

- Each carbon atom in graphite is bonded to **three** others forming **layers** of **hexagons**, leaving one free electron per carbon atom which becomes delocalised
- The covalent bonds within the layers are very strong, but the layers are attracted to each other by weak **intermolecular forces**

Diagram to show the bonding and structure in graphite



The structure and bonding in graphite

- Diamond and graphite are **allotropes** of carbon which have **giant covalent structures**
- Both substances contain only carbon atoms but due to the differences in bonding arrangements they are physically completely different
- Giant covalent structures contain billions of non-metal atoms, each joined to adjacent atoms by covalent bonds forming a giant lattice structure

What is the structure of diamond?

- In diamond, each carbon atom bonds with four other carbons, forming a **tetrahedron**
- All the covalent bonds are identical, very strong and there are no **intermolecular forces**

Diagram to show the formation of a covalent bond

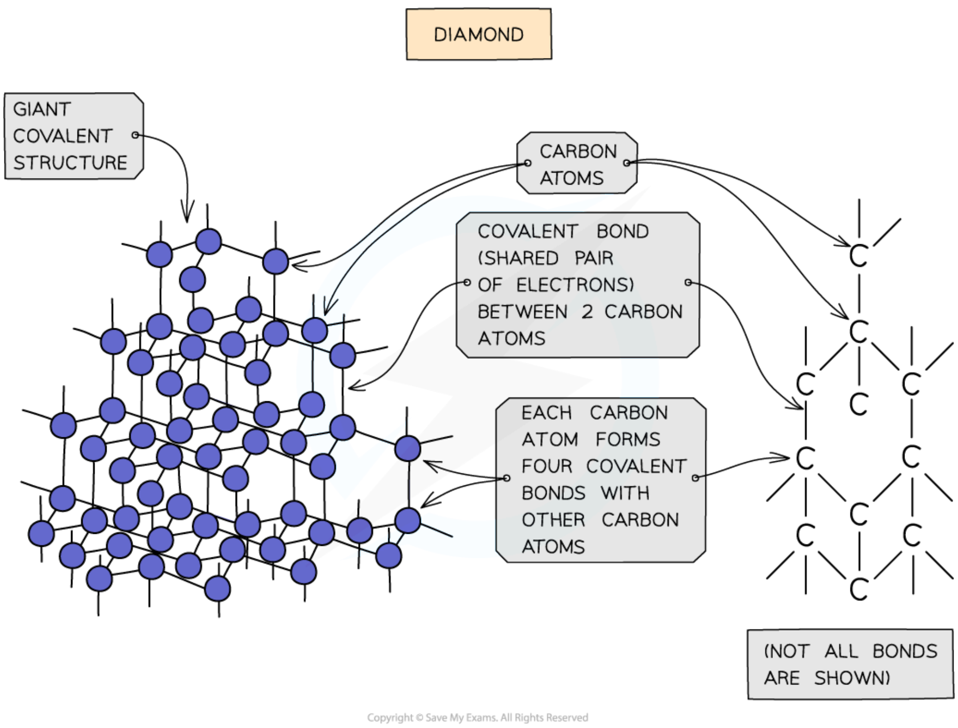


Diagram showing the structure and bonding arrangement in diamond

Uses of graphite & diamond

What are the properties and uses of graphite?

- Graphite **conducts electricity**
  - Each carbon atom is bonded to three others leaving one free electron per carbon atom
  - These free (delocalised) electrons exist in between the layers
  - They are free to move through the structure and carry charge
- Graphite has a **high melting point**
  - Graphite has a giant covalent structure
  - There are strong covalent bonds between the carbon atoms
  - These need lots of energy to break

Your notes

- Graphite is **slippery**
  - Graphite is arranged in layers
  - Although the atoms within the layers are joined by strong covalent bonds, the layers have only weak intermolecular forces between them
  - As a result the layers can slide over each other
  - This property allows graphite to be used in pencils and as an industrial lubricant

Examiner Tips and Tricks

Don't confuse pencil lead with the metal lead – they have nothing in common. Pencil lead is actually graphite, and historical research suggests that in the past, lead miners sometimes confused the mineral galena (lead sulfide) with graphite; since the two looked similar they termed both minerals 'lead'. The word graphite derives from the Greek word graphein, meaning "to write", so it is a well named mineral!

What are the properties and uses of diamond?

- Diamond does **not conduct electricity**
  - All the outer shell electrons in carbon are held in the four covalent bonds around each carbon atom
  - As a result, there are no freely moving particles to carry a charge
- Diamond has a very **high melting point**
  - Diamond has a giant covalent structure
  - There are strong covalent bonds between the carbon atoms
  - These need lots of energy to break
- It is extremely **hard** and **dense**
  - It has strong covalent bonds and each carbon atom is bonded to four other carbon atoms
  - Diamond's hardness makes it very useful in cutting tools like drills
- Diamond has the following physical properties:
  - It does not conduct electricity
  - It has a very high melting point
  - It is extremely hard and dense

Examiner Tips and Tricks

Your notes

Diamond is the hardest naturally occurring mineral, but it is by no means the strongest. Students often confuse **hard** with **strong**, thinking it is the opposites of weak. Diamonds are hard, but brittle – that is, they can be smashed fairly easily with a hammer. The opposite of saying a material is hard is to describe it as **soft**.

Your notes

Structure of silicon(IV) oxide

What is the structure of silicon(IV) oxide?

Extended tier only

- Silicon(IV) oxide (also known as silicon dioxide or silica), SiO_2, is a macromolecular compound which occurs naturally as **sand** and **quartz**
- Each oxygen atom forms covalent bonds with **2** silicon atoms and each silicon atom in turn forms covalent bonds with **4** oxygen atoms
- A tetrahedron is formed with one silicon atom and four oxygen atoms, similar to diamond

The structure of silicon(IV) oxide

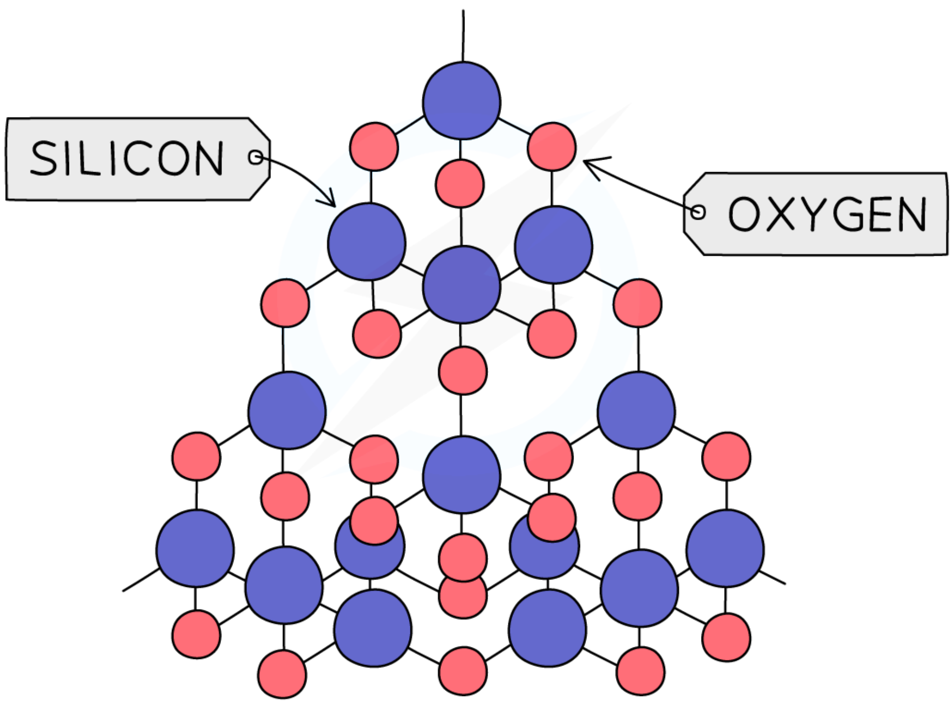


Diagram showing the structure of SiO_2 with the silicon atoms in blue and the oxygen atoms in red

Comparing diamond & silicon(IV) oxide

How does diamond compare to silicon(IV) oxide?

Your notes

Extended tier only

- SiO_2 has lots of very strong covalent bonds and no intermolecular forces so it has similar properties to diamond
- It is very **hard**, has a very **high** boiling point, is insoluble in water and does not conduct electricity
- SiO_2 is cheap since it is available naturally and is used to make sandpaper and to line the inside of furnaces

Metallic bonding

The structure of a metal

Extended tier only

- Metals consist of **giant** structures
- Within the metal lattice, the atoms lose their outer electrons and become positively charged metal **ions**
  - The outer electrons no longer belong to any specific metal atom and are said to be **delocalised**
  - This means they can move freely between the positive metal ions and act like a "*sea of electrons*"
- The metallic bond is the strong force of **attraction** between the positive metal ions and the delocalised electrons
- This type of bonding occurs in metals and metal alloys, which are mixtures of metal

Diagram to show metallic bonding

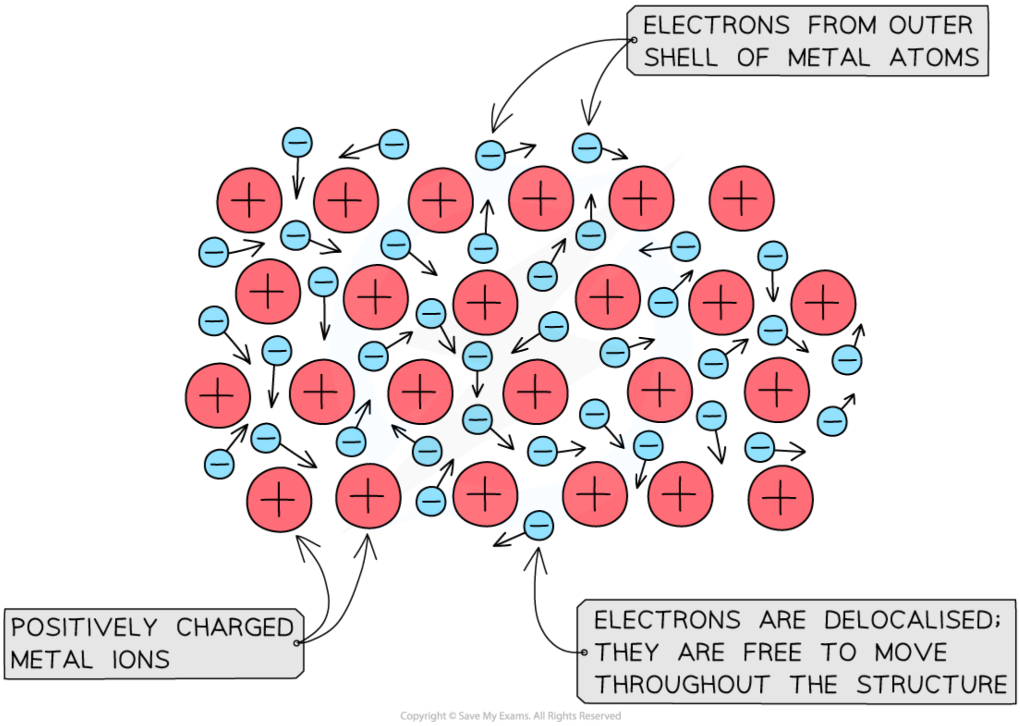


Diagram showing metallic lattice structure with delocalised electrons

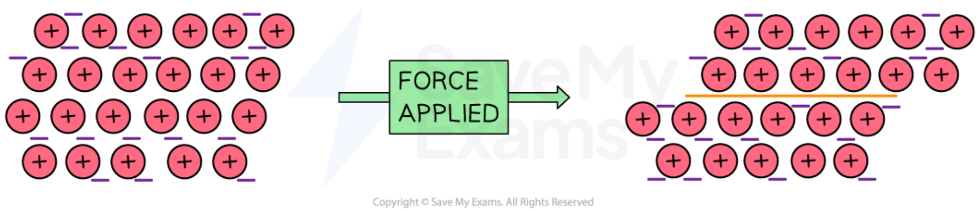
Properties of metals

What are the properties of metals?

Extended tier only

- Most metals have **high melting and boiling points**
  - There are strong electrostatic forces of attraction between the positive metal ions and the negative delocalised electrons within the metal lattice structure
  - These needs lots of energy to be broken
- Metals are **good conductors of heat and electricity**
  - The delocalised electrons are free to move and carry a charge through the whole structure
- Most metals are **malleable**
  - This means they can be hammered into shape
  - This is because the atoms are arranged in layers which can slide over each when force is applied

Malleability of metals



When a force is applied, the layers of positive ions slide over each other

- Most metals are **ductile**
  - This means they can be pulled into wires
  - This is also because the atoms are arranged in layers which can slide over each when force is applied

Examiner Tips and Tricks

When explaining why metals can conduct electricity, be careful of the terminology you use. Don't get confused with ionic compounds.

Your notes

Metals can conduct electricity as they have free **electrons** that can carry charge whereas molten or aqueous **ionic** compounds can conduct electricity because they have free **ions** that can carry charge.

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