Graphite

4.2.3.2 · Bonding, structure and the properties of matter · revision notes

Graphite is a form of carbon made of layers of hexagonal rings. You need to explain why it is soft and slippery, has a very high melting point and conducts electricity, and why it is similar to metals. A 4 to 6 mark question comparing graphite with diamond is common.

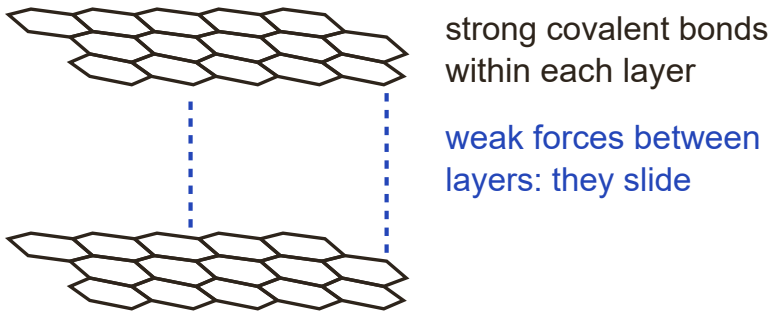
Key facts

- Graphite: each carbon forms 3 covalent bonds → layers of hexagonal rings
- One electron per carbon atom is delocalised → conducts electricity and thermal energy
- Like metals, graphite has delocalised electrons
- No covalent bonds between layers (only weak forces) → layers slide → soft and slippery
- Melting breaks strong covalent bonds within the layers → very high melting point
- Uses: pencils, lubricants, electrodes

Notes

Structure and bonding

- Graphite is made of **carbon atoms only**, in a giant covalent structure.
- Each carbon atom forms **three covalent bonds** with three other carbon atoms, making flat **layers of hexagonal rings**.
- There are **no covalent bonds between the layers**, only weak forces.



Graphite: layers of hexagons. Each carbon atom bonds to 3 others; the fourth electron is delocalised.

- Each carbon atom has four outer electrons but uses only three in bonds. The fourth electron is **delocalised**: one electron from each carbon atom is free to move through the structure.

Properties explained

- **Soft and slippery**: the forces between the layers are weak, so the **layers can slide over each other**. This is why graphite is used in pencils and as a lubricant.
- **Very high melting point**: to melt graphite, the **strong covalent bonds within the layers** must be broken. There are many of them, so a lot of energy is needed.
- **Conducts electricity** (and thermal energy): the delocalised electrons can move through the structure, carrying charge.
- Graphite is **similar to metals** because it has delocalised electrons. This is one reason it is used for electrodes.

Graphite compared with diamond

- Both are giant covalent structures of carbon atoms, and both have very high melting points.
- Diamond: 4 bonds per atom in a rigid 3D network → very hard; no delocalised electrons → does not conduct.
- Graphite: 3 bonds per atom, in layers with weak forces between them → soft; one delocalised electron per atom → conducts.

How to answer each type of question

Explain why graphite is soft and slippery 3 marks **Grade 5**

1. Say graphite is made of layers (of hexagonal rings of carbon atoms).
2. Say there are only weak forces (no covalent bonds) between the layers.
3. Say the layers can slide over each other.

Example. When you write with a pencil, layers of graphite rub off onto the paper.
Explain why graphite is soft.

Model answer. Graphite is made of layers of carbon atoms (1). There are only weak forces between the layers / no covalent bonds between the layers (1). So the layers can easily slide over each other (1).

Explain why graphite conducts electricity 3 marks **Grade 5**

1. Say each carbon atom forms only three covalent bonds.
2. So one electron from each carbon atom is delocalised.
3. The delocalised electrons can move through the structure and carry charge.

Example. Graphite is used to make electrodes.
Explain why graphite conducts electricity.

Model answer. Each carbon atom forms only three covalent bonds (1), so one outer electron from each carbon atom is delocalised (1). The delocalised electrons can move through the structure and carry charge (1).

Explain why graphite is soft but has a very high melting point

3 marks | Grade 7

1. Softness: layers slide because the forces between the layers are weak.
2. Melting: the strong covalent bonds within the layers must be broken.
3. There are many of these bonds, so a lot of energy is needed.

Example. Graphite is soft, but it has a very high melting point.
Explain why.

Model answer. Graphite is soft because there are only weak forces between the layers, so the layers slide over each other (1). To melt graphite, the strong covalent bonds within the layers must be broken (1). There are many of these bonds, so a lot of energy is needed (1).

6-mark: compare diamond and graphite

6 marks | Grade 8

1. Structure: both are giant covalent structures of carbon; diamond has 4 bonds per atom in a 3D network, graphite has 3 bonds per atom in layers.
2. Hardness: diamond is rigid, so hard; graphite's layers slide, so it is soft.
3. Conductivity: diamond has no delocalised electrons; graphite has one per carbon atom.
4. Write point by point, linking each property to the structure, and mention both substances each time.

Example. Diamond and graphite are both forms of carbon.
Compare the structure and bonding of diamond and graphite, and explain how they affect the hardness and electrical conductivity of each substance.

Model answer.

Marked in levels. A top-level answer (5–6 marks) links structure to both properties for both substances, using points such as:

• both are giant covalent structures made of carbon atoms
• in diamond each carbon atom forms four covalent bonds; in graphite each forms three
• graphite has layers of hexagonal rings, with weak forces (no covalent bonds) between the layers
• diamond is very hard because it is a rigid 3D network of strong covalent bonds
• graphite is soft because its layers can slide over each other
• graphite has one delocalised electron per carbon atom, which can move and carry charge, so it conducts electricity
• diamond has no delocalised electrons, so it does not conduct

Level 2 (3–4 marks): some correct links between structure and properties. Level 1 (1–2 marks): simple correct statements with few links.

Don’t lose marks

✗ Saying graphite has a low melting point because it is soft. Melting breaks the strong covalent bonds within the layers.

- ✗ Saying there are covalent bonds between the layers, or that the layers slide because the covalent bonds are weak.
- ✗ Saying graphite conducts because of free ions, or because the layers move.
- ✗ Writing that each carbon atom in graphite forms four bonds.
- ✗ Calling graphite a metal. It is a non-metal that has delocalised electrons like a metal.

More tips

Memory tricks

- Graphite: 3 bonds + 1 delocalised electron = conducts. Diamond: 4 bonds + 0 delocalised electrons = doesn't.
- Pencil 'lead' is graphite: its layers slide off onto the paper, which reminds you why graphite is soft.
- Soft but hard to melt: sliding only overcomes the weak forces between layers; melting has to break covalent bonds.

Exam technique

- Say 'one electron from each carbon atom is delocalised', not just 'it has free electrons'.
- For softness you need three ideas: layers, weak forces between them, layers slide.
- In 6-mark comparisons, organise your answer by property (structure, hardness, conductivity) and mention both substances in each part.

What each grade needs

What you need to be able to do, from the first marks up to the top grade.

Grade 3 — **Describe graphite as layers of carbon atoms**
Graphite is made of layers of carbon atoms arranged in hexagonal rings.

Grade 4 — **State the number of bonds each carbon forms**
Each carbon atom forms three covalent bonds with three other carbon atoms.

Grade 5 — **Explain why graphite is soft and slippery**
There are no covalent bonds between the layers, so the layers can slide over each other.

Grade 5 — **Explain why graphite conducts electricity**
One electron from each carbon atom is delocalised and can move through the structure, carrying charge.

Grade 6 — **Explain graphite's very high melting point**
The many strong covalent bonds within the layers must be broken.

Grade 8 — **Compare the structures and properties of diamond and graphite**
Four bonds in a rigid 3D network compared with three bonds, layers and delocalised electrons.

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