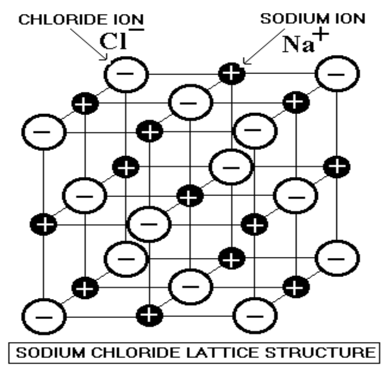
Ionic compounds

- Physical properties of ionic compounds – Pg 60 table 3.3
- Ionic compounds are solids at room temperature. The ions arrange themselves into a regular lattice structure.
- In this regular arrangement each ion is surrounded by ions of the opposite charge (why?).
- The entire structure is held together by the electrostatic forces of attraction that occur between particles of opposite charge.



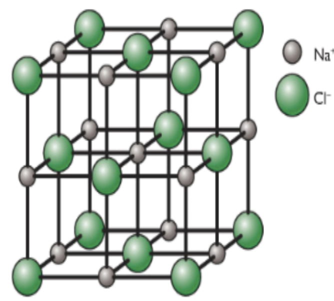
Giant structures

The four different types of solid physical structure are:

1. **Giant ionic lattice structures (ionic bonding)**: a lattice held together by the electrostatic forces of attraction between +ve and −ve ions.
2. **Giant covalent structures (covalent bonding)**: a substance where large numbers of atoms are held together by covalent bonds forming a strong lattice structure.
3. **Giant metallic lattice (metallic bonding)**: a regular arrangement of positive metal ions held together by the mobile 'sea' of electrons moving between the ions.
4. The reason why they are called 'giant' structures is because the structure repeats itself in all directions. The forces involved are the same in all directions holding the whole structure together.

Giant ionic lattice - Ionic crystals

- In these structures the atoms are arranged in an **ordered** and **repeating** fashion.
- The lattices formed by ionic compounds consist of a **regular arrangement** of **alternating** positive and negative ions.
- Ionic crystals are hard but brittle because of the structure of the layers. In an ionic crystal, pushing one layer against another brings ions of the same charges next to each other. The repulsions force the layers apart.
- Water can also disrupt an ionic lattice. Many ionic compounds dissolve in water. Water molecules are able to interact with both +ve and −ve ions.
- Ions in solution are able to move, so **the solution can carry an electric current**. They can conduct electricity when dissolved in water and when melted because the ions are free to move. They **cannot conduct electricity** when in solid form because ions are not free to move.



Giant covalent structures

- Giant molecular lattice (crystals) are held together by strong **covalent bonds.** This type of structure is shown by some elements (such as carbon in the form diamond and graphite, and also by some compounds (for example, SiO_2).
- Diamond and graphite are **allotropes** of carbon which have **giant covalent structures**.
- Allotropes are different atomic or molecular arrangements of the same element in the same physical state.
- These classes of substance contain a lot of non-metal atoms, each joined to adjacent atoms by covalent bonds forming a giant lattice structure.

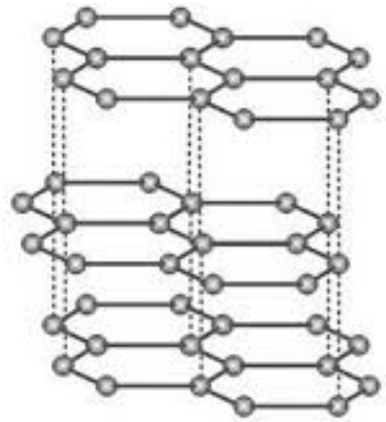
Properties of diamond

- In diamond, the strong covalent bonds extend in all directions through the whole crystal.
- Each carbon atom bonds with four other carbons, forming a **tetrahedron**. All the outer electrons of each carbon atom in used to form covalent bonds. There are no free electrons hence **diamond does not conduct electricity**.
- All the covalent bonds are identical and strong with no weak **intermolecular forces**.
- Diamond thus:
  - Does not conduct electricity.
  - Has a very high melting point.
  - Is extremely hard and dense.
- Diamond is used in **jewelry** and as **cutting tools**.

Properties of graphite

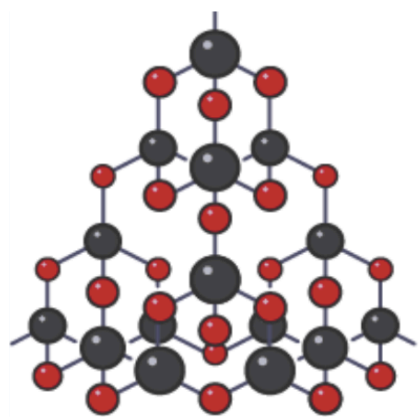
- Each carbon atom is bonded to **three** others forming **layers** of hexagonal shaped forms, leaving one free electron per carbon atom.
- These free electrons exist **in between** the layers and are free to move and carry charge, hence graphite can conduct electricity.
- The covalent bonds within the layers are very strong but the layers are connected to each other by weak **intermolecular forces** only, hence the layers can **slide** over each other making graphite **slippery** and **smooth**.
- Graphite thus:
  - Conducts electricity.
  - Has a very high melting point.
  - Is soft and slippery, less dense than diamond.
- Graphite is used in **pencils** and as an industrial **lubricant**, in engines and in locks.
- It is also used to make non-reactive **electrodes** for **electrolysis**.

Diamond and graphite



The Structure of Silicon(IV) Oxide (Silicon Dioxide)

- 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 as in diamond.

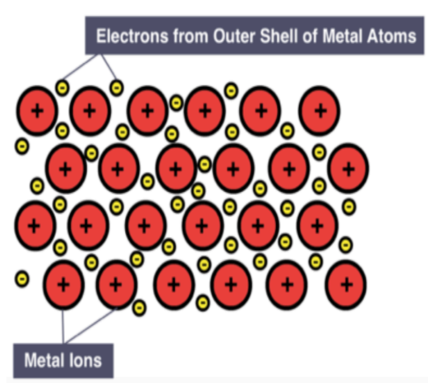


Similarity in properties between diamond and silicon(IV) oxide, related to their structures

- SiO_2 has lots of very strong covalent bonds so it has similar properties as diamond. Both have a rigid tetrahedral structure.
- It is very **hard**, has a very **high** melting point, is insoluble in water and does not conduct electricity (no free electrons).
- SiO_2 is cheap since it is available naturally.

Giant metallic lattice – metal crystals

- Metal atoms are held together strongly by metallic bonding.
- Within the metal lattice, the atoms lose their valence electrons and become positively charged.
- The valence electrons no longer belong to any metal atom and are said to be **delocalised (not restricted to orbiting one positive ion)**.
- They move freely between the positive metal ions like a sea of electrons. The form a kind of electrostatic 'glue' holding the structure together in what is called **metallic bonding.**
- Metallic bonds are strong and are a result of the attraction between the positive metal ions and the negatively charged delocalised electrons.



The links between metallic bonding and the properties of metals:

1. Metals have **high** melting and boiling points because:
   - There are many **strong metallic bonds** in giant metallic structures.
   - A lot of heat energy is needed to overcome forces between the positive metal ions and ‘sea’ of electrons and break these bonds.

The links between metallic bonding and the properties of metals:

2. Metals **conduct** electricity because:
   - There are **free electrons** available to move and carry charge.
   - Electrons entering one end of the metal cause a delocalised electron to displace itself from the other end.
   - Hence electrons can flow so electricity is conducted.

The links between metallic bonding and the properties of metals:

3. Metals are **malleable** and **ductile** because:
   - Layers of positive ions can **slide** over one another and take up different positions.
   - Metallic bonding is not disrupted as the valence electrons do not belong to any particular metal atom so the delocalised electrons will move with them.
   - Metallic bonds are thus not broken and as a result metals are strong but **flexible**.
   - They can be hammered and bent into different shapes without breaking.
   - Definition of malleable and ductile (pg 65)