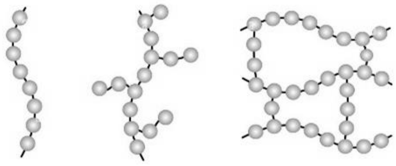
Polymers

Structure

A polymer is a **large (long-chain) molecule** made from many **repeating units (monomers)** joined together by **strong covalent bonds**.

- NM + NM
- Giant molecular (polymeric) structure consisting of long chains joined by strong covalent bonds.
- Strong covalent bonds within each polymer chain.
- Weak intermolecular forces between neighbouring polymer chains.
- Branched polymer chains cannot pack closely together, resulting in a **lower density** than polymers with more linear chains.
- Usually insoluble in water.
- Poor conductor of heat (low thermal conductivity).



Explaining the properties

Polymers consist of long chains of covalently bonded atoms with weak intermolecular forces between neighbouring chains.

Density

- Branched chains cannot pack together as efficiently as straight chains. This reduces the density of the polymer.
- If the density is less than that of water, the polymer will float.

Insolubility in water

- Water molecules cannot form attractions strong enough to separate the polymer chains.
- The weak attractions between water and the polymer are insufficient to overcome the many intermolecular forces between neighbouring polymer chains.
- Therefore, most polymers are insoluble in water.

Flexibility

- Weak intermolecular forces between neighbouring chains require relatively little energy to overcome.
- When a force is applied, the chains can slide past one another, allowing the polymer to bend or be moulded into shape.

Elasticity (cross-linked polymers)

- Cross-links join neighbouring polymer chains together. When stretched, the chains can move only a limited amount before the cross-links pull them back.
- As a result, the polymer returns to its original shape when the force is removed.
- Strong covalent cross-links prevent the chains from sliding, making the polymer stronger, more rigid, and more resistant to heat. Cross-linked polymers generally do not melt but instead decompose if heated sufficiently.

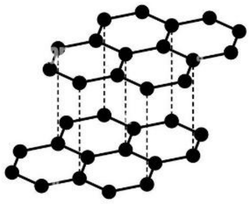
Thermal conductivity

- Polymers are poor conductors of heat because they **do not contain mobile (delocalised) electrons** to transfer thermal energy efficiently. The weak intermolecular forces between neighbouring chains reduce the transfer of vibrations from one chain to another, making polymers good thermal insulators.

Covalent Network Solids

Graphite

Structure



- a giant covalent (network) substance made entirely of carbon atoms.
- carbon atoms are arranged in layers of hexagonal rings (a 2D network).
- each carbon atom is covalently bonded to three other carbon atoms by strong covalent bonds.
- each carbon atom has one delocalised (mobile) electron, the delocalised electrons can move throughout each layer

Explaining the properties

Strength within the layers

- The carbon atoms within each layer are joined by strong covalent bonds.
- A large amount of energy is required to break these bonds. Therefore, each layer is very strong.

Softness / Lubricant

- Only weak intermolecular forces hold the layers together.
- These weak forces require little energy to overcome.
- The layers can slide over one another easily, making graphite soft and slippery.
- This is why graphite is used as a lubricant and in pencil "lead". As pencil "lead" the layers are transferred to the paper when a small force / little energy is applied.

Electrical conductivity

- Each carbon atom contributes one delocalised electron. These electrons are free to move throughout the layers.
- The movement of these electrons allows graphite to conduct electricity.

Thermal conductivity

- The carbon atoms are joined by strong covalent bonds within each layer.
- Vibrations are transferred rapidly from atom to atom through these bonds. The delocalised electrons also make a small contribution to heat transfer.
- Therefore, graphite is a good conductor of heat, particularly along the layers.

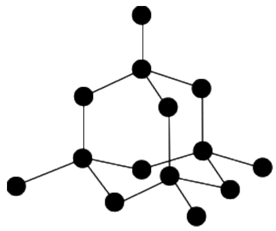
Insolubility

- Water molecules cannot form attractions strong enough to separate the carbon atoms.
- The strong covalent bonds are not broken by water.
- Therefore, graphite is insoluble in water.

Carbon fibre

- Carbon fibre consists mainly of layers of graphite.
- The strong covalent bonds within each layer make the material very strong.
- A large force is required to break these bonds, allowing carbon fibre to withstand large forces without deforming.

Diamond



Structure

- Diamond is a giant covalent (network) substance made entirely of carbon atoms.
- Each carbon atom is covalently bonded to four other carbon atoms.
- The atoms form a three-dimensional (3D) tetrahedral network.
- All four valence electrons are involved in covalent bonds. There are no delocalised electrons.

Explaining the properties

Hardness

- Every carbon atom is joined to four neighbouring carbon atoms by strong covalent bonds.
- These bonds extend throughout the entire 3D network.
- A very large amount of energy is needed to break these bonds.
- Therefore, diamond is extremely hard.

High melting point

- Strong covalent bonds exist throughout the giant network.
- Many strong covalent bonds must be broken for melting to occur.
- Therefore, diamond has a very high melting point.

Electrical conductivity

- All four valence electrons are involved in covalent bonds. There are no mobile (delocalised) electrons or ions.
- Therefore, diamond does not conduct electricity.

Thermal conductivity

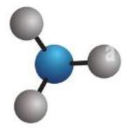
- The strong covalent bonds allow vibrations to be transferred rapidly from atom to atom throughout the 3D network.
- Therefore, diamond is an excellent conductor of heat, despite being an electrical insulator.

Insolubility

- Water molecules cannot form attractions strong enough to separate the carbon atoms.
- The strong covalent bonds throughout the network are not broken by water (and the attractions between the water molecules are also not broken).
- Therefore, diamond is insoluble in water.

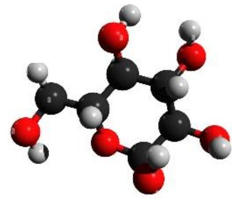
Covalent Molecular Substances

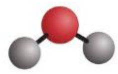
Structure



Covalent molecular substances consist of **small, discrete molecules**.

- Atoms within each molecule are joined by **strong covalent bonds**.
- There are **weak intermolecular forces** between neighbouring molecules.
- They are **simple molecular (NM + NM)** substances.
- Usually have **low melting points and low boiling points.**
- They are **insulators** (of heat and electricity).
- Solubility in water depends on the molecule.





Explaining the properties

Low melting point and boiling point

- The atoms within each molecule are held together by **strong covalent bonds**.
- However, only **weak intermolecular forces** exist between neighbouring molecules.
- Only a small amount of energy is needed to overcome these weak intermolecular forces.
- Therefore, covalent molecular substances have **low melting points and low boiling points**.
  (The covalent bonds within the molecules are not broken during melting or boiling.)

Electrical conductivity

- The molecules contain **no mobile (delocalised) electrons** and **no ions** that are free to move.
- Covalent molecular substances **do not conduct electricity** in the solid, liquid or gaseous state. They are insulators.

Thermal conductivity

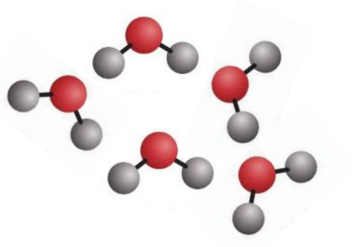
- Covalent molecular substances are **poor conductors of heat**. The molecules are held together by **weak intermolecular forces**, so thermal energy is transferred inefficiently between molecules.
- They also lack **mobile electrons**, which could help transfer thermal energy.
- Therefore, they have **low thermal conductivity**.

Gaseous state

- Molecules are **far apart** with little attraction between them.
- The molecules move freely and rapidly in all directions.
- Gases fill all the available space in a container.
- The molecules are far apart, gases have a **low density** / low mass per unit volume, so have a very low density. The large spaces between molecules mean gases can be **compressed easily**.

Liquid state

- The molecules are close together but are still able to move past one another.
- Weak intermolecular forces begin to form and are sufficient to hold the molecules together.
- Therefore, liquids can **flow** and take the shape of their container.



Solid state

- Molecules are packed closely together in a regular arrangement, with weak intermolecular forces hold neighbouring molecules in fixed positions.
- As only weak intermolecular forces need to be overcome, these substances melt relatively easily.

Volatility

- Weak intermolecular forces are easily overcome.   Molecules can escape from the surface of the liquid into the gas phase. Therefore, many covalent molecular substances are **volatile** (they evaporate easily).

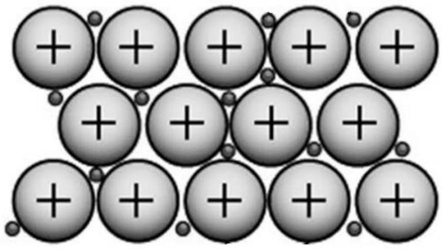
Solubility (general rule)

- **Most covalent molecular substances are insoluble in water** because water molecules cannot form attractions strong enough to separate the molecules.
- **Some polar covalent molecules** (such as ethanol or sugar) are soluble because they can form strong attractions (including hydrogen bonds, where applicable) with water molecules.

Metallic Solids

Structure

Metallic substances consist of a **giant metallic structure**.



- Metals are **metal + metal (M + M)** substances.
- They consist of a **3D lattice of positive metal ions (cations)** surrounded by a "sea" of **delocalised (mobile) valence electrons**.
- The metal atoms lose their outer electrons, forming positive metal ions.
- The metal cations are held together by **strong electrostatic attractions** between the positive ions and the negative delocalised electrons.
- These attractions are called **metallic bonds**.
- Metallic bonds are **non-directional**, meaning the attraction acts in all directions around the metal ions.

Explaining the properties

Electrical conductivity

- Metals contain **delocalised electrons** that are free to move throughout the structure.
- When a potential difference is applied, these electrons can flow through the metal.
- Therefore, metals are **good conductors of electricity**.

Thermal conductivity

- Metals are **good conductors of heat**.
- The delocalised electrons can move rapidly through the structure.
- These electrons transfer kinetic energy by colliding with metal ions, spreading thermal energy through the metal. Therefore, metals transfer heat efficiently.

Malleability and ductility

- Metallic bonds are **non-directional**, meaning the attraction between metal ions and delocalised electrons is maintained even when the layers of ions move.
- When a force is applied, layers of metal ions can slide past each other without breaking the metallic bonds.
- Therefore, metals can be shaped (**malleable**) and drawn into wires (**ductile**).

Pure metals vs alloys

- Pure metals are generally more malleable and ductile than alloys.
- In a pure metal, all atoms/ions are the same size, allowing layers to slide easily.
- In alloys, different-sized atoms disrupt the regular arrangement.
- This makes it more difficult for layers to slide past each other, making alloys harder and less malleable.

Melting point

- Most metals have high melting points.
- A large amount of energy is needed to overcome the strong electrostatic attraction between metal cations and delocalised electrons.
- Therefore, most metals are solids at room temperature. *(Some metals, such as mercury, have low melting points and are liquid at room temperature.)*

Density

- The density of a metal depends on density depends on the mass of the atoms per unit volume
  - the mass of the metal atoms/ions
  - how closely packed the particles are in the metallic lattice.
- Metals with heavier atoms and more closely packed structures have higher densities.

Solubility

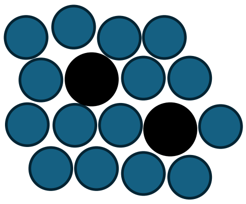
- Metals are generally **insoluble in water**.
- The forces of attraction between the metal atoms / cations / positive ions and water molecules are weaker than the forces of attraction between the metal atoms (and their valence electrons)
- Therefore, water cannot separate the metal ions from the metallic lattice.

Alloys

Structure

An alloy is a mixture of a metal with one or more other elements. Alloys can be:

- Metal + metal (M + M) e.g. brass (copper + zinc)
- Metal + non-metal (M + NM) e.g. steel (iron + carbon)
- Alloys contain a 3D metallic lattice of positive metal ions surrounded by delocalised electrons.
- The different atoms in the alloy are usually different sizes, which disrupts the regular arrangement of the metal lattice.
- Metallic bonds still form between the metal ions and the delocalised electrons.



Explaining the properties

Malleability and ductility

- Alloys contain different types of atoms/ions, which are usually different sizes.
- The different-sized atoms disrupt the regular arrangement of the metallic lattice.
- This makes it harder for layers of atoms/ions to slide past each other, so more force is required to move the layers without disrupting the metallic bonding.
- Therefore, alloys are **less malleable and less ductile** than pure metals.

Strength / hardness

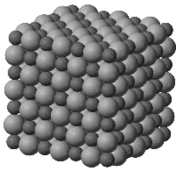
- The irregular arrangement caused by different-sized atoms prevents layers from sliding easily.
- More energy is required to deform the structure, so alloys are *usually* **harder and stronger** than pure metals.

Density

- The density of an alloy depends on:
    - the mass of the atoms present
    - how closely packed the atoms are in the metallic lattice.
- Adding heavier atoms can increase density, while changes to packing may increase or decrease density.

Ionic Solids

Structure



Ionic substances consist of a **giant ionic lattice**.

- They are made of a **3D lattice of oppositely charged ions**:
  - positive ions (**cations**)
  - negative ions (**anions**)
- The ions are arranged in a regular repeating pattern.
- The ions are held together by **strong electrostatic attractions** between oppositely charged ions.   These attractions are called **ionic bonds**.
- Each ion is surrounded by ions of the opposite charge.

Explaining the properties

High melting point

- Ionic solids have **strong electrostatic attractions** between oppositely charged ions.
- A large amount of energy is required to overcome these attractions and separate the ions.
- Therefore, ionic compounds generally have **high melting points**.

Brittleness

- Ionic solids are **brittle** because the ions are arranged in fixed positions in a regular lattice.
- When a force is applied, layers of ions may shift slightly. This can bring ions with the **same charge** next to each other.
- Like charges repel strongly. The repulsion causes the ionic lattice to crack and shatter.

Electrical conductivity

Solid ionic compounds:

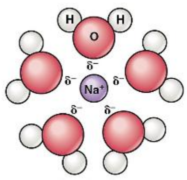
- The ions are held tightly in fixed positions by strong ionic bonds. The ions cannot move freely.
- Therefore, solid ionic compounds **do not conduct electricity**.

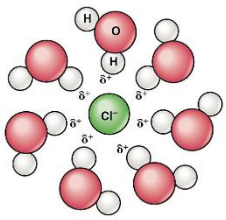
Molten ionic compounds or ionic solutions (aqueous):

- The ionic lattice has been broken apart. The ions are free to move. Moving ions can carry electrical charge.   Therefore, molten ionic compounds and ionic solutions **conduct electricity** and are called **electrolytes**.

Solubility in water

(Many ionic compounds are soluble, although not all.)





- Water is a polar molecule with:
  - a slightly negative oxygen end
  - slightly positive hydrogen ends
- The negative oxygen end of water molecules attracts and surrounds positive ions (**cations**).
- The positive hydrogen ends of water molecules attract and surround negative ions (**anions**).
- If the attraction between the water molecules and the + and – ions is stronger than the forces of attraction between the oppositely charged ions within the 3D lattice and the weaker attraction between neighbouring water molecules, the water molecules pull ions out from the lattice structure, to overcome the attractions holding the lattice together. The ions separate, become surrounded by water molecules, and spread throughout the solution. The solid is no longer visible because the ions are dispersed.