

CHAPTER 5: States of Matter

- 5.1 The Gaseous State
- 5.2 The Liquid State
- 5.3 The Solid State
- 5.4 Ceramics
- 5.5 Conserving Materials

Learning outcomes:

- (a) state the basic assumptions of the kinetic theory as applied to an ideal gas.
- (b) explain qualitatively in terms of intermolecular forces and molecular size:
 - (i) the conditions necessary for a gas to approach ideal behaviour.
 - (ii) the limitations of ideality at very high pressures and very low temperatures.
- (c) state and use the general gas equation $pV = nRT$ in calculations, including the determination of M_r .
- (d) describe, using a kinetic-molecular model: the liquid state, melting, vaporisation, vapour pressure.
- (e) describe, in simple terms, the lattice structure of a crystalline solid which is:
 - (i) ionic, as in sodium chloride, magnesium oxide.
 - (ii) simple molecular, as in iodine.
 - (iii) giant molecular, as in silicon(IV) oxide and the graphite and diamond allotropes of carbon.
 - (iv) hydrogen-bonded, as in ice.
 - (v) metallic, as in copper.

[the concept of the 'unit cell' is not required]
- (f) explain the strength, high melting point and electrical insulating properties of ceramics in terms of their giant molecular structure.
- (g) relate the uses of ceramics, based on magnesium oxide, aluminium oxide and silicon(IV) oxide, to their properties (suitable examples include furnace linings, electrical insulators, glass, crockery).
- (h) discuss the finite nature of materials as a resource and the importance of recycling processes.
- (i) outline the importance of hydrogen bonding to the physical properties of substances, including ice and water (for example, boiling and melting points, viscosity and surface tension).
- (j) suggest from quoted physical data the type of structure and bonding present in a substance.

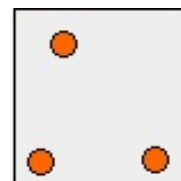
5.1 The Gaseous State

Kinetic theory of gases

- 1) Assumptions made in the kinetic theory of gases as applied to **ideal gases**:
 - The gas particles have **zero intermolecular forces** between them.
 - The gas particles behave as point particles which have **negligible volume**.
 - The gas particles are in constant random motion, colliding with each other and the wall of the container frequently.
 - All collisions between the gas particles are perfectly elastic.
 - Pressure is due to the collision of gas particles with the wall of container.

- 2) In the gaseous state, the particles can **move freely** and are **far apart**.

- 3) A gas has **no fixed shape and volume**, it takes the shape of container and always fills it. The volume of a gas depends on its pressure, temperature and number of moles.



gas

Ideal gas and real gas

- 1) A real gas is most like an ideal gas at:
 - i. **low pressures**
 - At low pressures, the distance between gas particles is large and the volume of the gas particles is negligible (very small compared to the volume of the container).
 - Intermolecular forces are also negligible at low pressures.
 - ii. **high temperatures (well above its boiling point)**
 - At high temperatures, the gas particles have negligible intermolecular forces between them because they have sufficient energy to overcome it.
- 2) However, an ideal gas **does not exist**, because:
 - i. there are intermolecular forces between the gas particles.
 - ii. the volume occupied by the gas particles is not zero.
- 3) A real gas shows biggest deviation from an ideal gas at:
 - i. **high pressures**
 - At high pressures, the gas particles are packed close together, thus the volume occupied by the gas particles is not negligible.
 - The intermolecular forces between them is also not negligible.

ii. low temperature

- At low temperatures, the gas particles are packed close together, thus the volume occupied by the gas particles is not negligible.
- The intermolecular forces between them is also not negligible because they do not have sufficient energy to overcome it.

The general gas equation

- The general gas equation gives the relationship between the *volume* of a given mass of gas and the prevailing conditions (*temperature, pressure*).

$$pV = nRT$$

where p = pressure in Pa (or N m^{-2})
 V = volume in m^3
 n = no. of moles of gas
 R = gas constant ($8.31 \text{ J K}^{-1} \text{ mol}^{-1}$)
 T = temperature in K

- The general gas equation, $pV = nRT$, can be expressed as

$$pV = \frac{m}{M}RT \quad \left[\text{since } n = \frac{m}{M} \quad \text{where } m = \text{mass in g} \right.$$

$$\Rightarrow M = \frac{m}{pV}RT \quad \left. M = \text{molar mass in g mol}^{-1} \right]$$

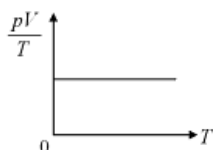
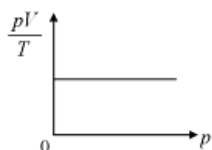
Hence, the general gas equation can be used to determine M_r .

If the density, ρ , of the gas is given, then $M = \rho \frac{RT}{p}$ since $\rho = \frac{m}{V}$.

- An **ideal gas** obeys the general gas equation, $pV = nRT$, *exactly*.

For a fixed mass of ideal gas,

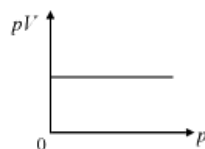
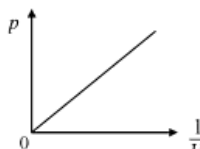
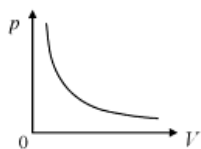
➤ $\frac{pV}{T} = nR$ OR $\frac{pV}{T} = \text{constant}$.



$$\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$$

➤ At constant temperature, $pV = \text{constant}$ OR $V \propto \frac{1}{p}$.

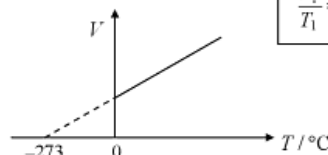
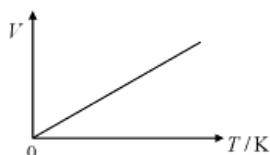
$$p_1V_1 = p_2V_2$$



← Boyle's Law

➤ At constant pressure, $\frac{V}{T} = \text{constant}$ OR $V \propto T$.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$



← Charles' Law

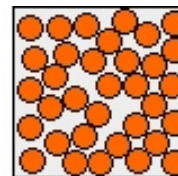
Side note

- 1 atm = 1.01×10^5 Pa
- $(x)^\circ\text{C} = (x + 273)$ K
- 1 $\text{m}^3 = 1000 \text{ dm}^3$
= 1000000 cm^3
- At s.t.p:
 - p = 1.01×10^5 Pa
 - T = 273 K

5.2 The Liquid State

Kinetic theory of liquids

1) In liquids, the particles are packed **quite closely** together but in a **fairly random arrangement** (gaps are present between them). So, the particles have limited movements.



liquid

2) A liquid has **fixed volume** but do not have fixed shape.

3) Intermolecular forces are present between the particles, its strength is stronger than in the gaseous state but weaker than in the solid state.

Melting and freezing

1) When a solid is heated, the energy is absorbed by the particles and they vibrate about their fixed positions **more vigorously**.

2) Then, a point is reached where the particles have energy high enough to overcome the attractive forces that hold them in fixed positions. They break away from their fixed positions and move freely.

3) The solid then becomes a liquid, this process is called **melting**. The temperature at which this process happens is called the **melting point**.

4) In freezing, the reverse happens. The liquid particles lose energy until they do not have enough energy to move freely. They are held together in fixed positions again. The liquid **solidifies**.

Boiling(vapourisation) and condensation

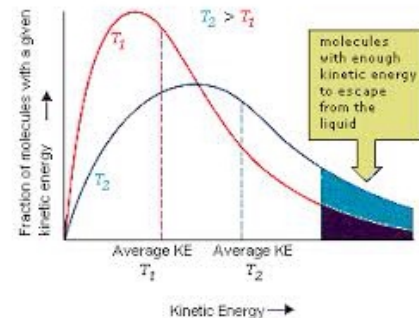
1) When a liquid is heated, the vapour pressure of the liquid increases until eventually it is **equal** to the atmospheric pressure. Bubbles of vapour will form in the body of liquid.

2) The bubbles then rise to the surface of the liquid, burst open and escape into the atmosphere as a gas. The liquid **boils**.

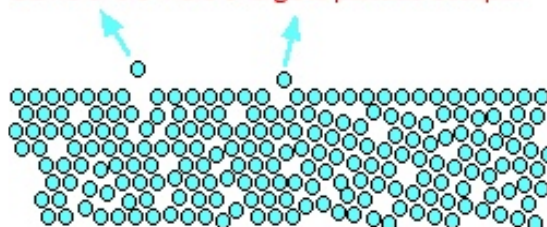
- 3) This happens because when a liquid is heated, the particles absorb energy until it is sufficient to overcome the forces of attraction between them. The particles break away from the fairly close arrangement of the liquid and boils.
- 4) This process is called **boiling**. The temperature at which this process happens is called the **boiling point**.
 - i. Boiling point depends on external pressure. If a liquid boils under a pressure lower than the atmospheric pressure (1 atm), it will boil faster.
 - ii. Conversely, if it boils at a pressure higher than the atmospheric pressure, it will boil slower.
- 5) In condensation, the reverse happens, the particles lose energy and experience increasing attractive force. They move slower and become closer together when temperature is sufficiently low. The gas **liquefies**.

Evaporation and vapour pressure

- 1) The energy distribution of particles in the liquid state follows a shape similar to the normal distribution.
- 2) The average energy is governed by the temperature. The higher the temperature, the higher the average energy.
- 3) Some particles have energy higher than the average while some have lower. The more energetic particles at the surface of the liquid can be moving fast enough and eventually overcome the attractive forces and escape into the atmosphere. They **evaporate** to form vapour (Vapour is the gas form of a particle below its boiling point).
- 4) Unlike boiling, evaporation only takes place on the **surface** of the liquid.
- 5) In an open container, the liquid will evaporate until none is left.

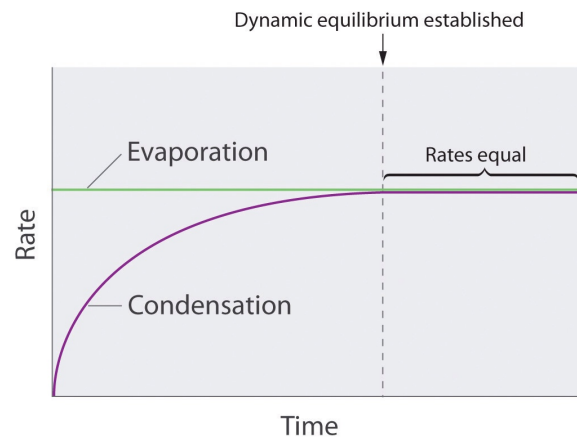


Some of the more energetic particles escape.

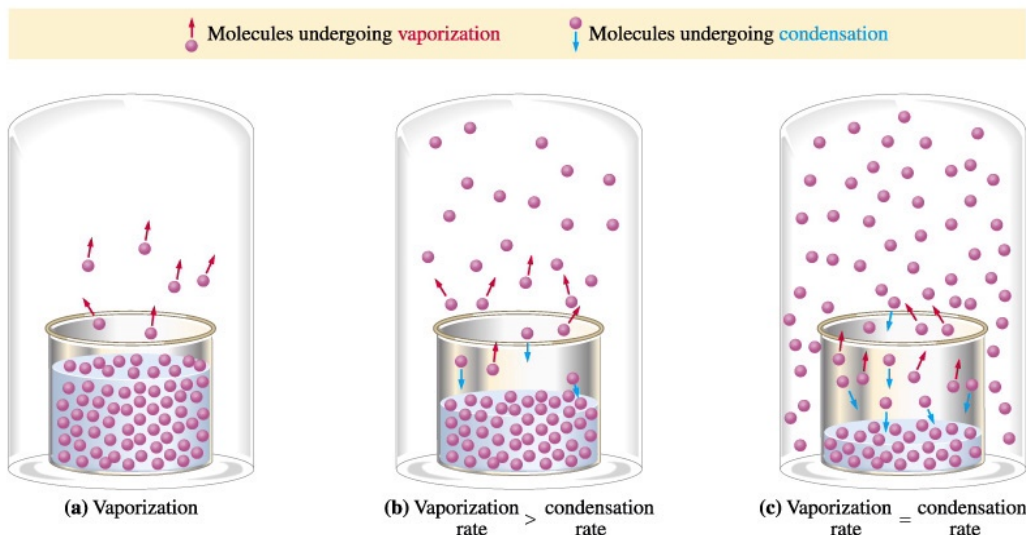


- 6) However, a different thing happens when the liquid is evaporated in a closed container.
- 7) At first, liquid particles with higher energy escape from the surface of the liquid to become vapour. The vapour particles will collide with the wall of container. The collisions exert a pressure called **vapour pressure**.
- 8) As more and more particles escape, the vapour particles become close together. Eventually the particles with lower energy will not be able to overcome the attractive forces between them. The vapour begins to condense and return to liquid.

- 9) Eventually the vapour particles return to liquid at the **same rate** as the liquid particles evaporate to form vapour. An **equilibrium** is reached. At this equilibrium, the concentration of liquid particles and vapour particles remains constant.



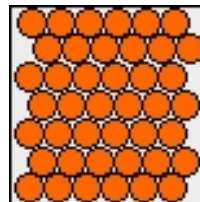
- 10) In this situation, the vapour pressure is **maximum** and is called the **saturated vapour pressure**.
- 11) Vapour pressure will increase when temperature is increased. At higher temperature, more liquid can undergo evaporation. The vapour particles are more energetic and collide with the wall of container harder and more frequently. This causes the pressure exerted to be higher.



5.3 The Solid State

Kinetic theory of solids

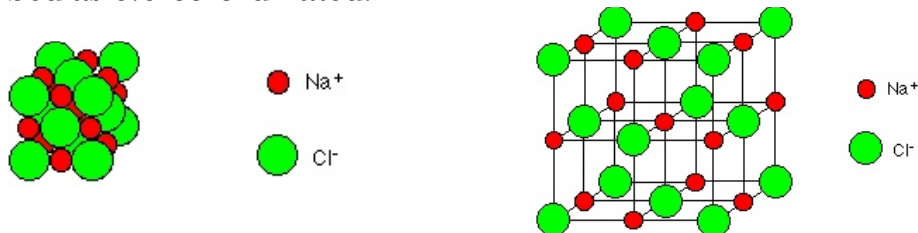
- 1) In solids, very strong forces of attraction hold the particles in **fixed positions** and **close to each other**. Hence, particles in the solid state can only rotate and cannot translate.
- 2) A solid has **fixed shape and volume**.
- 3) The structure of solids can be crystalline or amorphous.
 - i. In **crystalline solids**, the particles are arranged in a regular and orderly pattern called a lattice structure.
 - ii. In **amorphous solids**, the particles do not have a regular and orderly arrangement. Examples are rubber and plastic.



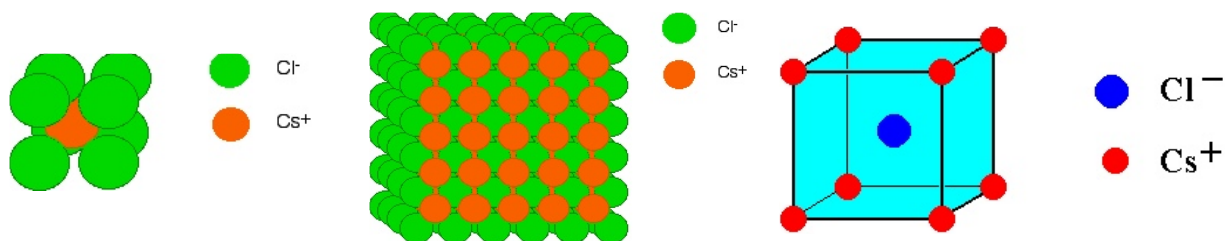
solid

Ionic lattices

- 1) An example of ionic lattice is sodium chloride, NaCl, where strong ionic bonds hold the Na⁺ and Cl⁻ ions rigidly in place in the solid lattice.
- 2) Sodium chloride has a **face-centred cubic structure**. In this structure, each ion is surrounded by **six** other oppositely-charged ions. So, sodium chloride is described as 6:6-co-ordinated.



- 3) Some other examples with face-centred cubic structure are magnesium chloride and magnesium oxide.
- 4) Another example is caesium chloride, CsCl, which has a **body-centred cubic structure**. In this structure, each ion is surrounded by **eight** other oppositely-charged ions. So, caesium chloride is described as 8:8-co-ordinated.
- 5) This is just like layers of oppositely-charged ions stacking on each other.



6) Properties of ionic compounds:(For more detailed discussions, see Chapter 4)

i. **They are hard**

- This is because the ionic bonds within are very strong. A lot of energy is required to split them apart.

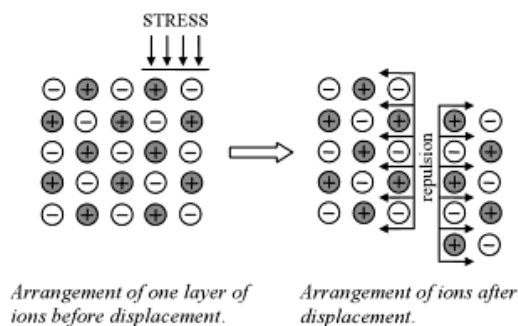
ii. **They have high melting and boiling points**

- This is because the ionic bonds within are very strong. A lot of energy is required to break the bonds.

iii. **They are brittle**

- This is because when a stress is applied, the layers of ions may be **displaced** by the force.

- This brings ions with the same charge together and the repulsion between them causes the crystal to split.



iv. **They conduct electricity in the molten or aqueous state**

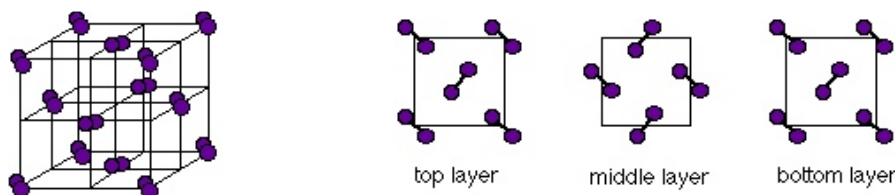
- This is because in the molten or aqueous state, free moving ions are present to carry the current.

v. **Many of them are soluble in water(polar solvents) but insoluble in non-polar solvents**

- This is because when ion-polar solvent attraction is set up, the energy released is enough to compensate the energy needed to break the ionic bonds.

Simple molecular lattices

1) An example of simple molecular lattice is iodine, I₂. Solid iodine has a face-centred cubic structure.

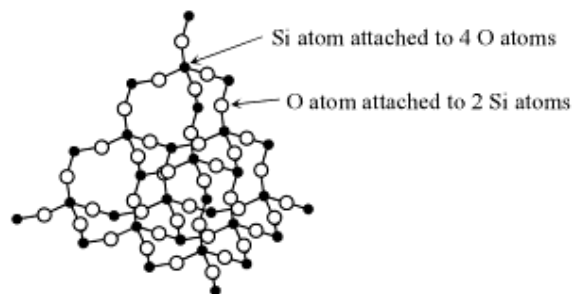


- 2) Weak van der Waals' forces of attraction hold the individual iodine molecules together.
- 3) Properties of simple molecular solids:(For more detailed discussions, see Chapter 4)
 - i. **They do not conduct electricity in any state**
 - This is because there are no free electrons to carry the current.
 - ii. **They have low melting and boiling points**
 - This is because the van der Waals' forces of attraction between them is weak, little energy is required to overcome it.
 - iii. **They are soluble in non-polar solvents but insoluble in water(polar solvents)**
 - This is because the molecule-non-polar solvent attraction set up is enough to compensate the energy needed to break the weak van der Waals' forces of attraction between the simple molecules.

Giant covalent lattices

- 1) In giant covalent structures, **all the atoms are covalently-bonded to each other**, linking the whole structure. An example of giant covalent lattice is silicon(IV) oxide or silicon dioxide, SiO_2 .

- 2) In silicon dioxide, each silicon atom is covalently-bonded to **four** oxygen atoms in a tetrahedral arrangement while each oxygen atom is covalently-bonded to **two** silicon atoms in a V-shaped(bent) arrangement.

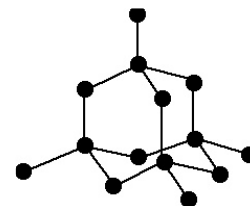


- 3) Properties of silicon dioxide:
 - i. **Very high melting and boiling points**
 - The covalent bonds holding the atoms are very strong, a lot of energy is required to break these strong bonds.
 - ii. **Very hard**
 - This is because of the need to break the strong covalent bonds
 - iii. **Does not conduct electricity**
 - This is because there are no delocalised electrons to move around and carry the current. All the electrons are held tightly to the atoms.

4) Another two examples are diamond and graphite. Both of them are **allotropes of carbon**.

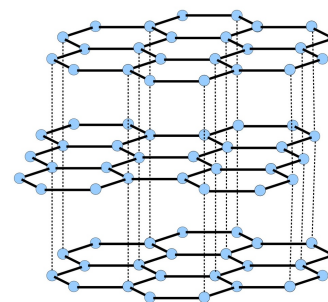
5) *Allotropes* are different crystalline or molecular forms of the same element in the same physical state, having different atomic arrangement.

6) In diamond, all the carbon atoms undergo sp^3 hybridisation and covalently-bonded to **four** other carbon atoms in a tetrahedral arrangement. This network of carbon extends throughout the whole structure.



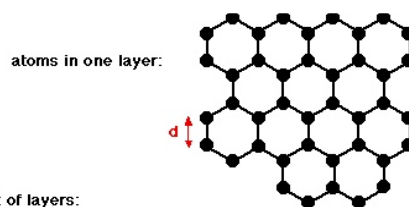
7) The properties of diamond are same as the ones in silicon dioxide.

8) In graphite, the carbon atoms undergo sp^2 hybridisation. Each carbon atom is covalently-bonded to **three** other carbon atoms in a trigonal planar arrangement **within each layer**, forming hexagons.



9) The unhybridised p orbital in each planar layers overlap sideways **to form a giant molecular orbital above and below each layer**. The electrons are then **delocalised**, free to move between the layers of carbon. However, the electrons cannot move from one layer to another.

10) The distance between carbon atoms between the layers is **greater** than the distance between carbon atoms in each layer.



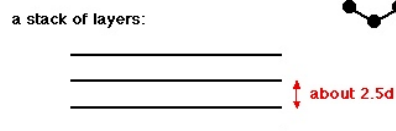
11) Properties of graphite:

i. **Very high melting and boiling points**

- This is because the covalent bonds holding the carbon atoms with each layer is very strong, a lot of energy is required to overcome the strong forces of attraction.
- In fact, it has a higher melting point than diamond because of the additional attraction between the delocalised electrons and the nucleus. Dispersion forces are also present, this is because the delocalised electrons can set up a temporary dipole easily.

ii. **Soft and slippery**

- Layers of carbons can slide over easily, this is because the intermolecular forces between the layers are very weak.



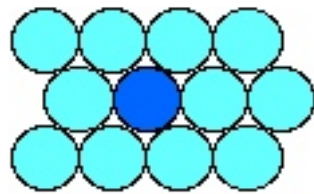
- Because of this, graphite is often used as pencil 'leads' and lubricants.

iii. Conducts electricity

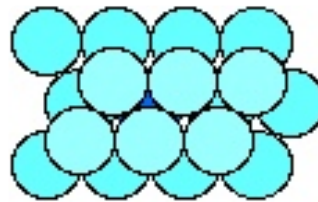
- It can conduct electricity in the direction parallel to each layer but not perpendicular to it.
- This is because the delocalised electrons are free to move between the layers but not from one layer to another.

Metallic lattices

- 1) An example of metallic lattice is copper, Cu. In copper, each copper atom is surrounded by **twelve** other copper atoms. Therefore, copper is a 12-co-ordinated metal.



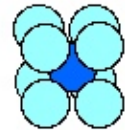
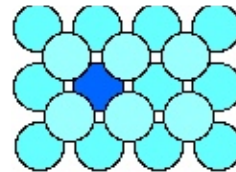
Six copper atoms surround it in a layer



Three other copper atoms surround it up and down.

- 3) The copper atoms are bonded to each other through strong metallic bonding. (For more information about metallic bonding, refer Chapter 4)

- 4) Another example is sodium. In sodium, each sodium atom is surrounded by **eight** other sodium atoms only. Therefore, sodium is a 8-co-ordinated metal.



In the first diagram, no atoms are touching each other in a layer.

- 5) Properties of metals:

i. High melting and boiling points

- This is because the metallic bonds are very strong, a lot of energy is required to break these strong bonds.

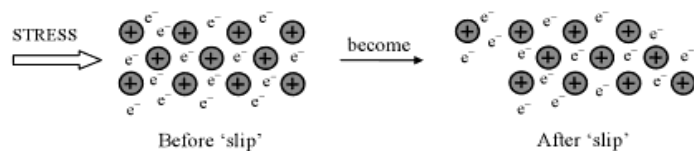
ii. Conducts heat and electricity

- This is because delocalised electrons are present and free to move.

iii. Ductile and malleable

- Metals are **ductile**(can be pulled into wires) and **malleable**(can be hammered into different shapes).

This is because the layers of metals can slide over each other; the broken metallic bonds can be immediately replaced by the new ones.



5.4 Ceramics

Ceramics

- 1) A *ceramic* is an inorganic non-metallic solid which is prepared by heating a substance or mixture of substances to a high temperature.
- 2) Ceramics often contain silicon dioxide, magnesium oxide and aluminium oxide. This gives ceramics their giant covalent or ionic structures.
- 3) Properties of ceramics:
 - i. **Very high melting and boiling points**
 - This is because most ceramics contain giant covalent or ionic structures. The covalent or ionic bonds holding them together is very strong, a lot of energy is required to overcome it.
 - ii. **Does not conduct electricity or heat**
 - This is because there are no delocalised electrons or free moving ions present. Therefore most of them are electrical insulators.
 - iii. **Chemically unreactive**
 - This is because all the electrons are held firmly in strong covalent bonds and not available for a reaction.
 - iv. **Very hard**
 - This is because the ionic or covalent bonds holding them is very strong.

Uses of ceramics

- 1) Ceramics containing **magnesium oxide** are used:
 - as electrical insulators in industrial electrical cables.
 - as a refractory in furnace linings because it has a high melting point.
 - in fire-resistant wall boards.
- 2) Ceramics containing **aluminium oxide** are used:
 - as a refractory in furnace linings because it has a high melting point.
 - as an abrasive for grinding hard materials because they do not conduct heat or melt when heat is given off during grinding.
 - in transparent aluminium oxide-scandium windows.
 - in high temperature and high voltage electrical insulators.
 - in the replacement of artificial hip joints.

- 3) Ceramics containing **silicon dioxide** are used:
- as a refractory in furnace linings because it has a high melting point.
 - as a abrasive, for example in sandpaper.
 - in the manufacture of glass.

5.5 Conserving Materials

Why conserve materials?

- 1) There is only a limited supply of metal ores in the Earth. Therefore metals are **finite resources**. They do not get replaced once they are used up.

Element	Mineral sources	Rel. abundance in Earth's crust
Aluminium	bauxite $Al_2O_3 \cdot 2H_2O$	8.13 %
Iron	haematite Fe_2O_3 magnetite Fe_3O_4 siderite $FeCO_3$ pyrites FeS_2	5.00 %
Calcium	limestone $CaCO_3$ gypsum $CaSO_4 \cdot 2H_2O$	3.63 %
Manganese	pyrolusite MnO_2	0.10 %
Chromium	chromite $FeO \cdot Cr_2O_3$	0.02 %

Ways to conserve materials

- 1) One way to conserve materials is through **recycling**.
- 2) Advantages of recycling:
- i. Recycling saves new resources.
 - ii. Recycling reduces the amount of waste materials to be disposed off.
 - iii. Recycling saves energy because less energy is needed to recycle metals than to extract it from their ores.
 - iv. Recycling protects the environment because it reduces pollution associated with product manufacture, disposal and littering. For example, the process of extracting metals produces gases which can lead to acid rain.
- 3) Two metals that can be recycled easily are copper and aluminium. Glass can also be recycled easily.
- 4) Another way to conserve materials is **using renewable resources**. Examples are like water, wood, sunlight, wind and etcetera.