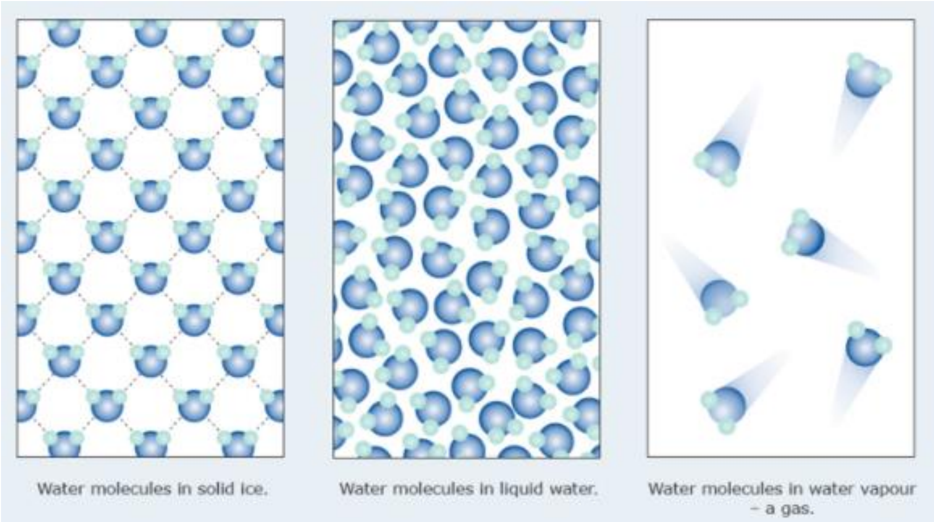
UNIT 2

PARTICLES, BONDING AND STRUCTURE

PART 1 – THE KINETIC MODEL OF MATTER



Contents

Key words: kinetic model, solid, liquid, gas, melting, freezing, evaporating, condensing, vapour pressure, melting point, boiling point, solution, solvent, solute, state symbol, aqueous

Units which must be completed before this unit can be attempted:
Unit 1 – Atomic Structure and the Periodic Table

1) Introduction

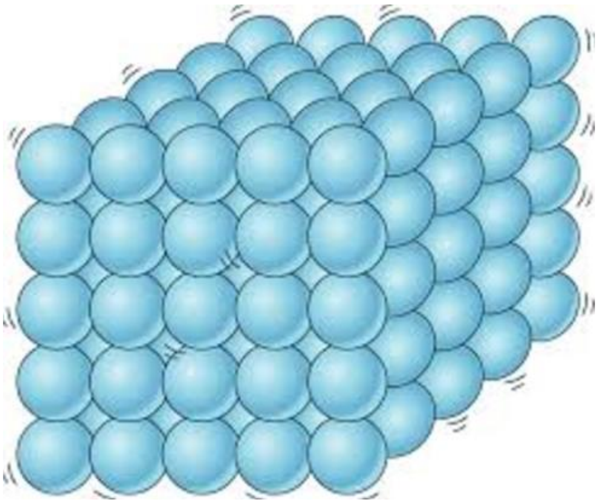
All matter is made up of particles. These particles as atoms, molecules or ions depending on the chemical structure of the material. In between these particles, there is empty space.

Except at absolute zero (0 K or -273 ^{o}C), all particles have kinetic energy. This means that all particles are always moving. This is known as the **kinetic model of matter**.

Matter can exist in one of three states, depending on how much energy the particles have, how they are moving and how the different particles interact.

2) Solids

In solids, all particles are **close together** and arranged in a **regular pattern (or lattice)** in **three dimensions**. The particles are fixed in place by the adjacent particles and so move by vibrating about their positions.



Solids have a fixed shape and cannot flow, because the particles are held in fixed positions by the chemical bonds between them. Solids cannot be compressed or squashed because the particles are already close together and have no space to move in to.

A solid is a state with high order and low randomness.

As solids are heated, the particles gain more energy and vibrate more. Eventually, the particles have sufficient energy to break some of the bonds holding the particles together, so some of these bonds break and the particles are able to move past each other. When the particles are able to move past each other, a change of state from solid to liquid has taken place. This change of state is known as **melting** (the reverse is known as **freezing**).

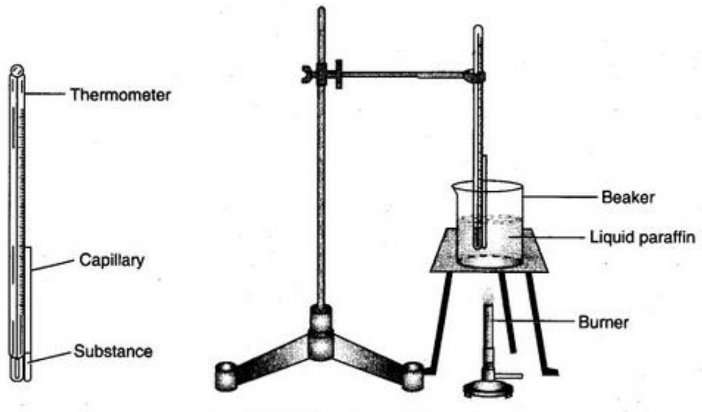
Demonstration: Observe ice and sulphur and melting

In some cases, all of the bonds holding the particles together break at the same time, and the solid changes directly into a gas. This change of state is known as **subliming**.

Demonstration: Observe iodine subliming

The temperature at which a solid turns into a liquid is known as the **melting point** of a solid.

The melting point of a solid can be determined by packing a small quantity of the solid into a capillary tube until it is approximately 1 cm full. This capillary tube should then be strapped to the stem of a thermometer, which should be placed into a heating bath. The bath should be heated slowly and the temperature at which the solid starts to melt and finishes melting should be recorded.



Practical: Determine the melting point of naththalene

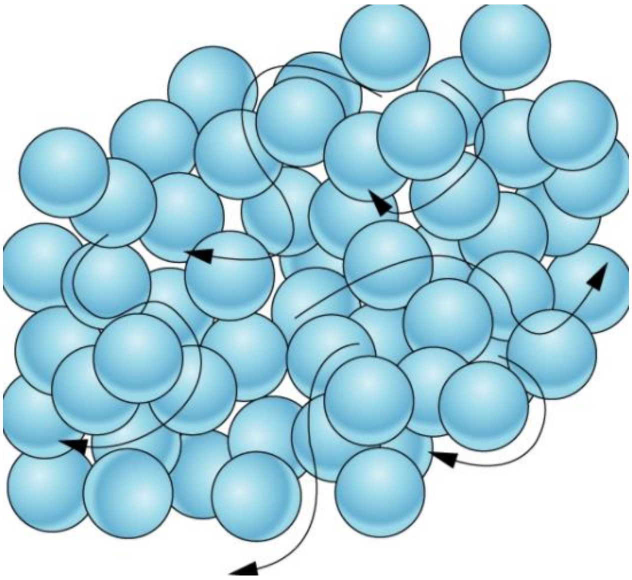
Many solids contain impurities. Impurities interfere with the bonding between the particles and makes them easier to break. As a result, impurities lower the melting point of substances.

You can test how pure a substance is by finding its melting point and comparing it to the known melting point of the pure substance. The closer the melting point to that of the pure substance, the greater the purity of the sample. Impure solids also tend to melt over a range of temperatures. Hence melting points are an indicator of the purity of solids.

The physical properties of a solid depend on the type of particles present and the type and strength of the forces between them.

3) Liquids

In solids, all particles are **close together** but able to move past each other because some of the bonds between the particles have been broken. The particles are therefore vibrating as well as moving around. Most of the bonds holding the particles together, however, are still intact.



Liquids do not have a fixed shape and can flow, because the particles can move past each other. Liquids cannot be compressed or squashed because the particles are already close together and have no space to move in to.

A liquid is a state with lower order (more disorder) and more randomness than a solid.

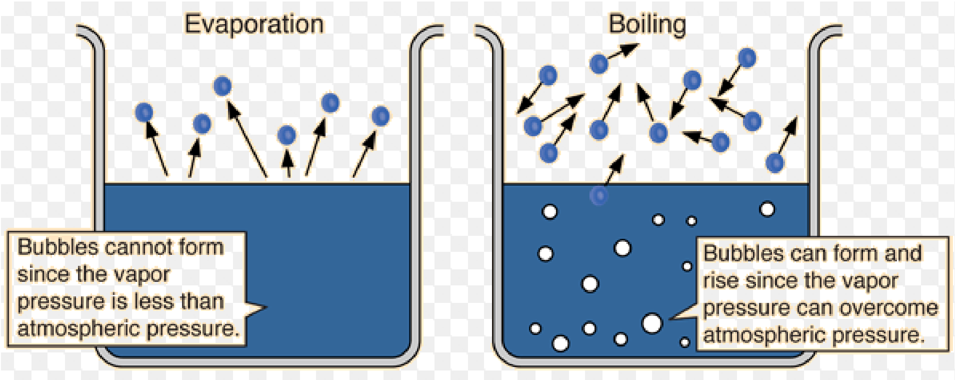
As liquids are heated, the particles gain more energy and move around faster. Eventually, the particles have sufficient energy to break the remaining bonds holding the particles together, so the particles are no longer attracted to each other and move apart. When the particles are no longer close together, a change of state from liquid to gas has taken place. This change of state is known as **boiling** (the reverse is known as **condensing**).

Demonstration: Observe water boiling

Not all particles in a liquid are moving at the same speed at the same time – the speed changes every time two particles collide. Sometimes, particles at the surface of a liquid have enough energy to escape from the liquid and temporarily become gas particles. This process is known as **evaporation**. As a result, all liquids contain a small amount of vapour immediately above it. The pressure exerted by this vapour is known as the **vapour pressure** of the liquid.

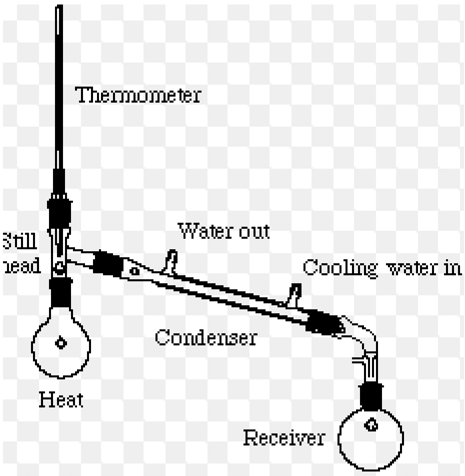
As the temperature increases, the vapour pressure of the liquid increases. When the vapour pressure of a liquid equals the atmospheric pressure, the liquid turns into a gas.

The physical properties of a solid depend on the type of particles present and the type and strength of the forces between them.



The temperature at which a liquid turns into a gas is known as the **boiling point** of the liquid. The **standard boiling point** of a liquid is the temperature at which the vapour pressure of the liquid is equal to 1 atm pressure (100,000 Pa – this is normal atmospheric pressure).

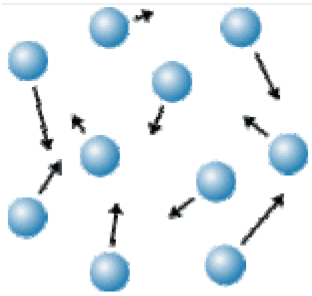
The boiling point of a liquid can be determined by adding a small amount of liquid to a flask and connecting it to a three-way head with a thermometer inserted into it. The head should also be attached to a condenser set up for distillation. The bulb of the thermometer should be just above the level of the liquid. Record the maximum temperature reached – this is the boiling point of the liquid.



Demonstration: Determine the boiling point of water

4) Gases

In gases, the particles are moving around rapidly and all bonds between them have been broken. There are significant spaces between the particles.



Gases have no fixed shape and will fill the entire volume available to them, because they are able to move rapidly. Gases can be compressed or squashed because there are spaces between the particles.

A gas is a state with the lowest order (highest disorder) and most randomness.

Two basic assumptions can be made about gases:
- the strength of the forces between gas particles is negligible
- the volume occupied by gas particles is negligible compared with the volume of the container

Any gas for which these two assumptions are true is called an **ideal gas**. Most gases behave like ideal gases except at very high pressures and very low temperatures.

As a result, unlike solids and liquids, the physical properties of a gas do not depend on the type of particle involved or the forces between them. All gases therefore have very similar physical properties. These properties can be summarised in three gas laws:

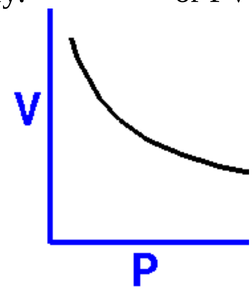
(a) Boyle's Law

If the volume occupied by a fixed amount of gas increases at a fixed temperature, the particles will collide with the sides of the container less often and therefore the pressure of the gas will increase. Similarly, if a fixed amount of gas is compressed into a smaller volume, its pressure will increase.

Boyle's Law states that "**the pressure exerted by a given mass of an ideal gas is inversely proportional to the volume it occupies, if the temperature remains unchanged in a closed system**".

(a closed system is a system from which particles cannot escape, and into which particles cannot enter)

Mathematically: $P \propto \frac{1}{V}$ or PV = k or $P_1V_1 = P_2V_2$

Graphically: 

Example: If a sample of gas exerts a pressure of 100,000 Pa in a container of fixed volume 0.05 m^3, what pressure will it exert if the volume is compressed to 0.02 m^3?

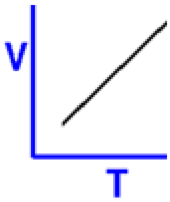
Answer: P_1V_1 = 100,000 x 0.05 = 5,000 = P_2V_2, so P_2 = 5000/V = 5000/0.02 = **250,000 Pa**

(b) Charles' Law

If a fixed amount of gas is heated at constant pressure, the particles have more energy and move faster. This will cause the gas to expand. Similarly, if a fixed amount of gas is cooled at constant pressure, the particles will have less energy and move more slowly, causing the gas to contract.

Charles' Law states that "**the volume occupied by a given mass of an ideal gas is directly proportional to its absolute temperature, if the pressure remains unchanged in a closed system**".

Mathematically: $V \propto T$ or V/T = k or $V_1/T_1 = V_2/T_2$

Graphically: 

Example: If a sample of gas occupies a volume of 100 dm^3 at a temperature of 300 K, what volume will the gas occupy if it is heated to a temperature of 360 K? at constant pressure?

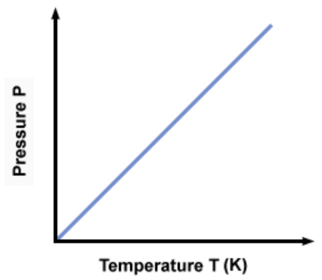
Answer: V_1/T_1 = 100/300 = 1/3 = V_2/T_2, so V_2 = 1/3 x 360 = **120 dm^3**

(c) Gay-Lussac's Law

If a fixed amount of gas is heated at constant volume, the particles have more energy and move faster. This will cause them to collide with the sides of the container with more force, resulting in an increase in pressure. Similarly, a fixed amount of gas is cooled at constant volume, the particles will collide with the sides of the container with less force, resulting in a decrease in pressure.

Gay-Lussac's Law states that "**the pressure exerted by a given mass of an ideal gas is directly proportional to its absolute temperature, if the volume remains unchanged in a closed system**".

Mathematically: $P \propto T$ or P/T = k or $P_1/T_1 = P_2/T_2$
Graphically:



Example: If a sample of gas occupies exerts a pressure of 100,000 Pa at a temperature of 300 K, what pressure will the gas exert if it is heated to a temperature of 450 K whilst maintaining its volume constant?

Answer: P_1/T_1 = 100000/300 = 1000/3 = P_2/T_2, so P_2 = 1000/3 x 450 = **150,000 Pa**

(d) Combined Gas Law

Boyle's Law, Charles' Law and Gay-Lussac's Law can be combined as follows:

PV/T = k or $P_1V_1/T_1 = P_2V_2/T_2$

This law is called the **combined gas law**.

Example: If a sample of gas occupies exerts a pressure of 100,000 Pa in a volume of 5 dm^3 at a temperature of 300 K, what pressure will the gas exert if it is heated to a temperature of 450 K whilst also being compressed into a volume of 4 dm^3?

Answer: P_1V_1/T_1 = 100000 x 5/300 = 5000/3 = P_2V_2/T_2, so P_2 = 5000/3 x 450/4 = **187,500 Pa**

Test Your Progress: Topic 2 Part 1 Exercise 1

1. If a sample of gas occupies exerts a pressure of 200,000 Pa at a temperature of 350 K, what pressure will the gas exert if it is heated to a temperature of 400 K whilst maintaining its volume constant?
2. If a sample of gas exerts a pressure of 150,000 Pa in a container of fixed volume 0.04 m^3, what pressure will it exert if the volume is expanded to 0.06 m^3?
3. If a sample of gas occupies a volume of 50 dm^3 at a temperature of 300 K, what volume will the gas occupy if it is cooled to a temperature of 275 K?
4. If a sample of gas occupies exerts a pressure of 100,000 Pa in a volume of 2 dm^3 at a temperature of 320 K, what pressure will the gas exert if it is cooled to a temperature of 300 K whilst also being compressed into a volume of 1.5 dm^3?
5. In each of questions 1 – 4, state which Law you used to calculate your answer.

(e) Dalton's Law of Partial Pressures

The gas laws apply even if the gas is a mixture of different gases.

The partial pressure of a gas in a mixture of gases is the pressure which would be exerted if that gas alone occupied the entire gaseous volume.

For example, dry air consists of a mixture of different gases, mainly nitrogen, oxygen and argon. If the gas is at normal atmospheric pressure, the partial pressure of each gas is as follows:

Nitrogen:	79,000 Pa
Oxygen:	20,000 Pa
Argon:	1000 Pa

According to Dalton's Law of partial pressures, the total pressure in a system is the sum of the partial pressures of the different gases in the system.

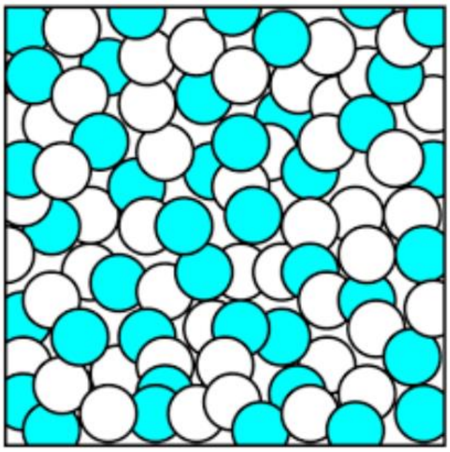
So the overall pressure of air at normal atmospheric pressure is 79000 + 20000 + 1000 = 100,000 Pa.

Like the other gas laws, this law works because the gas particles do not interact with each other.

5) Solutions

Pure substances can exist either as solids, liquids or gases, depending on the temperature and the nature of the forces between the particles.

It is also common for two or more different substances to mix uniformly together. This usually happens when a solid, liquid or gas dissolves in another liquid. The particles in the dissolving substance separate and distribute themselves evenly throughout the liquid. The resulting mixture is known as a **solution**.



The minor component in a solution is called the **solute**. The major component in a solution is called the **solvent**. The most common solvent is water.

Solutions in which the solvent is water are known as **aqueous** solutions.

6) State Symbols

It is often important to indicate the physical state of a material. This is shown by showing an abbreviation for the physical state of the substance in brackets, after the chemical formula:

State	symbol
Solid	s
Liquid	l
Gas	g
Dissolved in water	aq

Eg Ice is represented by the formula H_2O (s)
Eg Water is represented by the formula H_2O (l)
Eg Steam or water vapour is represented by the formula H_2O (g)

7) Brownian Motion

Brownian motion is the random movement of particles suspended in a liquid or gas as a result of collisions with the liquid or gas particles they are suspended in.

When visible particles are suspended in a liquid or a gas, they are observed to move randomly, sometimes changing direction suddenly. These particles move because they are colliding with gas or liquid particles which are too small to see.

Brownian motion is evidence for the particle theory of matter in liquids and gases.

Demonstration: View Brownian motion of pollen grains or powdered sulphur in water, by observing through a microscope (www.youtube.com/watch?v=hy-clLi8gHg)

Demonstration: Collect some smoke in a glass container, and view Brownian motion of the smoke particles by shining a strong light on the glass container from the side (www.youtube.com/watch?v=hy-clLi8gHg)

Demonstration: Sweep a dusty room during the day, then look into the room from outside and observe Brownian motion of dust particles

8) Diffusion

(a) Particle Model of Diffusion

Particles added to a specific region of a liquid or a gas will generally spread out until they are uniformly distributed throughout the liquid or gas. This is because they move gradually from regions in which their concentration is high to regions in which their concentration is low, until their concentration everywhere is the same. This process is known as **diffusion**.

Diffusion results from the random movement of particles. The higher the concentration of particles in a space, the more likely particles are to move out of that space.

It can often be a slow process because the particles are constantly colliding with the other particles in the liquid or gas, which causes them to change direction. Diffusion occurs more quickly in gases than in liquids, because the particles are less densely packed and so collide with each other less frequently.

Demonstration: Fill a test tube with a coloured gas or vapour such as bromine, iodine or NO_2. Now connect this tube to an empty test tube. Observe the colour gradually move from one tube into the other until the colour is evenly distributed. (www.youtube.com/watch?v=H7QsDs8ZRMI)

Demonstration: Open a bottle of concentrated ammonia in a corner of the room. How long does it take before you can smell it? Why can you smell it?

(b) Graham's Law of diffusion

Not all particles diffuse at the same rate. At the same temperature, all particles have the same average kinetic energy. But because KE = ½ mv^2, then particles with a larger mass must be moving more slowly, and particles with a larger mass must be moving more quickly.

Shown mathematically, if KE = ½ mv^2 and all particles at the same temperature have the same energy, then:

$m_1v_1^2 = m_2v_2^2$, so $v_2^2/v_1^2 = m_1/m_2$ and $v_2/v_1 = \sqrt{(m_1/m_2)}$

In other words, the relative rate of diffusion of two gases is inversely proportional to the square root of their relative masses. This is known as **Graham's Law of Diffusion**.

Example: Argon atoms have a relative atomic mass of 40. Neon atoms have a relative atomic mass of 20. How fast will neon diffuse compared to Argon?

Answer: $(m_1/m_2) = 40/20 = 2$, so $\sqrt{(m_1/m_2)} = 1.4 = v_2/v_1$. **So Neon will diffuse 1.4 times faster than Argon.**

Test Your Progress: Topic 2 Part 1 Exercise 2

1. Helium, neon, argon and krypton have relative atomic masses of 4, 20, 40 and 84 respectively. Which gas will diffuse the most quickly and which gas will diffuse the most slowly?
2. How many times faster will helium diffuse compared to argon?
3. How many times faster will neon diffuse compared to neon?