Matter and its Properties

Key Words

Volume	condense	molecule	property
mass	evaporate	particle	extensive(general)
density	melt area	kinetic model	intensive(characteristic)
boiling point	freeze	state of matter	
melting point	sublime(verb)	change of state	

Properties of Matter

Matter is the 'stuff' that things are made of - it **occupies space** and we can **measure its properties**.

General or **extensive** properties of matter depend on the amount of matter that is being measured, and do not allow us to identify or distinguish one substance from another; for example mass or volume.

Characteristic or **intensive** properties of matter do not depend on the amount of matter and help us identify or distinguish one substance from another (- for example colour, boiling/freezing point or density).

We use intensive properties to help distinguish matter.

Mass is the quantity of matter in an object. The SI unit of mass is the kilogram (kg), although the gram is often used for smaller quantities and the tonne for larger quantities. **Volume** is the space an object occupies. The SI unit of volume is the cubic metre (m^3).
Density is the relation between the mass and the volume of an object. The density of an object is the quotient of its mass per unit of volume. The Si unit of density is kg/m^3. (It measures how concentrated the mass is).

Melting point is the temperature at which a solid, at standard pressure, completely changes into a liquid, and **boiling point** is the temperature at which a liquid boils and turns into a gas (or vapour), under standard pressure. The SI unit of temperature is the Kelvin (K), although the Celsius (ºC) temperature scale is often used.

(Remember that a liquid can change into gas at temperatures well below its boiling point by the process known as evaporation).

Kinetic Theory - Particles on the Move
States of Matter – Changing State

The kinetic particle theory explains the properties of the different states of matter. Everything is made of tiny particles (we will study them in the next unit). These particles are not free to move around in a solid, but they move freely in liquids and gases. As they randomly move they collide with each other and bounce off in all directions. Therefore, the particles in solids, liquids and gases have different amounts of energy. They are arranged differently and move in different ways.

These particles can establish forces of attraction between them. Depending on how strong these forces are the state of matter is determined.

There are four states of matter: solid, liquid, gas and plasma. We will study the first three.
Solids have a fixed shape and a fixed volume. They cannot be compressed and do not flow.
Liquids have a definite volume, but not a definite shape, as they take the shape of the container they are in. They flow easily and are not compressible.
Gases have no definite volume or shape. They assume the volume and shape of the container they are in. They flow easily and can be compressed.

But, how can we explain the different characteristics of solids, liquids and gases? It is the arrangement of the particles and how strong the forces of attraction between them are that makes a difference. Let's take a closer look!

Particles in Solids, Liquids, and Gases

- **Solid:** The particles are packed tightly in a fixed position. The forces of attraction which holds them together are very strong, allowing them only to vibrate. Therefore, solids have a fixed shape and volume.

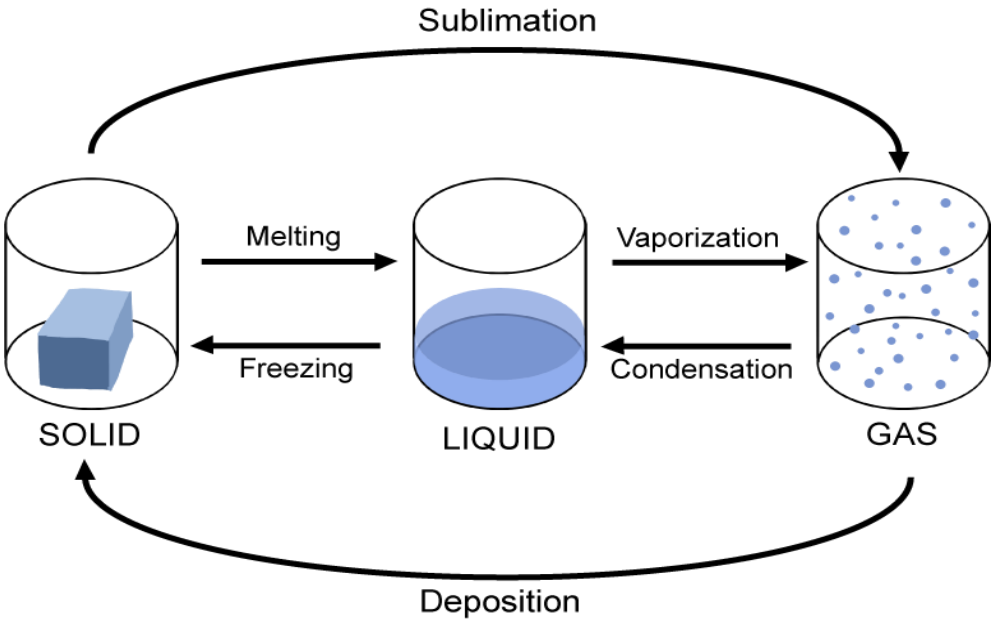
- **Liquid:** The particles in a liquid are close together, but are not in a fixed pattern. The forces holding them together are weaker than in a solid, giving them some freedom of movement. These particles can move about and slide past each other but not escape. Therefore, liquids have a constant volume (particles don't separate from each other), but not a fixed shape. Liquids take the shape of the receptacle in which they are contained, and they can be poured (they flow).

- **Gas:** There are almost no forces holding the particles of a gas together. The particles in a gas are far apart and they move about very quickly, occupying all the space available. Gases do not have a fixed shape or volume. Gases are easily compressed into a smaller volume. We say gases are compressible.

Changing states – Phase changes

How do substances go from one state to another? It is all about energy!

We use **heat (= thermal energy)** to melt a solid. Heat makes the particles move faster. Once the particles have enough energy, they can start to overcome those forces which hold them together. As we keep providing energy in the form of heat the particles will move faster and eventually the liquid will boil and change into a gas (vapour).



("2016")

In a solid body (e.g. ice), there is in general very little movement of the particles. As we heat the solid its particles get more energy and vibrate more. As the particles now have more energy, they can start to overcome those attractive forces between them breaking away from their position. The particles start to change places and the solid melts into a liquid (e.g. liquid water). If we continue heating the particles, they get more energy, moving even faster. Eventually the particles get enough energy to overcome the forces between them. Spaces will start to open up between the particles until they break way to form a gas (e.g. water vapour).

Now, if we do the opposite, that is, if we **cool down**, or remove energy from the system the reverse changes will take place. The gas will condense into a liquid and the liquid will freeze forming a solid.

The changes from a gas to a liquid, a liquid to a solid etc. are called **phase changes**.

During a phase change the temperature remains constant (lab demonstration)
So even though you keep on adding heat, the temperature stays steady while the substance melts, and again while it boils.

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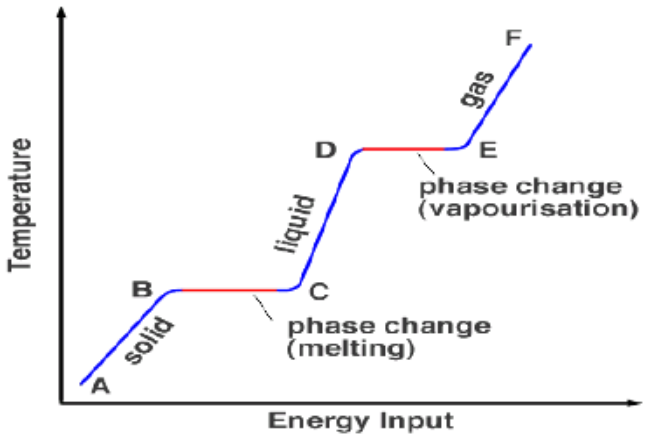


Phase change diagram.

There are two types of phase changes: progressive and regressive.

Progressive phase changes are those that take place when **energy is provided** to the system. Fusion, vaporization and positive (or progressive) sublimation.

Regressive phase changes are those that take place when **energy is removed** from the system. Solidification, condensation, and negative (or regressive) sublimation (deposition).

- **Fusion (melting):** the change in state from a solid to a liquid.
- **Solidification (freezing):** the change in state from a liquid to a solid.
- **Sublimation:** the direct change of state from a solid to a gas.
- **Negative sublimation:** the direct change of state from a gas to a solid.
- **Vaporization:** the change in state from a liquid to a gas.

Evaporation: is the process of a liquid changing into a vapour at temperatures below its boiling point. It only occurs on the surface of a liquid when a particle has by chance sufficient energy to escape.
Boiling: is the rapid change in state from a liquid to a gas usually caused by heating. When heating, the particles in a liquid gain energy. Eventually they have sufficient energy to overcome the binding forces in the liquid, and spread away as particles of gas.

- **Condensation:** the change in state from a gas to a liquid at room temperature.

Pressure affects the temperatures at which the phases change.
For instance, the boiling point of water lowers down when pressure decreases. Similarly, **the boiling point rises when pressure increases**. In a pressure cooker, we increase the pressure to obtain a situation where the elevated boiling point (≈ 120 °C) allows us to prepare our food at a higher temperature and thereby reduce the cooking time.

Ideal Gases: Boyle-Mariotte, Charles and Gay-Lussac perfect gases laws

An **ideal or perfect gas** is that which consists of molecules that occupy negligible space, and there is no attraction, nor interaction between its molecules (except at collisions).

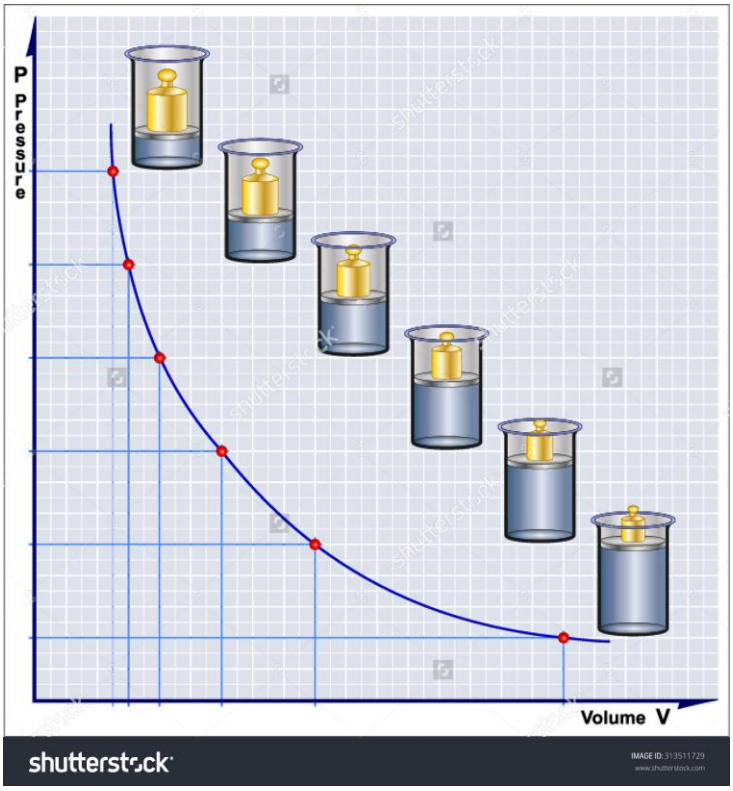
When dealing with a fixed mass of gas, there are always three factors to consider: pressure, volume and temperature. A change in one of these factors affects at least one of the others.

There are three laws that link the different factors in ideal gases:

- **Boyle's Law** (linking pressure and volume, at constant temperature)

For a fixed mass of gas at constant temperature, the pressure is inversely proportional to the volume.

So, as we increase the pressure the volume decreases and as we decrease the pressure the volume increases.



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Another way of expressing **Boyle's law** is as follows:

For a fixed mass of gas at constant temperature, the product of the pressure by the volume remains constant.

So, if the pressure of a gas changes from P_1 to P_2 when the volume is changed from V_1 to V_2:

$$P_1 \times V_1 = P_2 \times V_2 \quad \textbf{(at constant temperature)}$$ (Pople, 2007)

Heating gases

A gas does not necessarily expand when heated as its volume depends on the container it is in. When we change the temperature of a gas, its pressure or volume can change depending on the circumstances.

You are familiar with the Celsius (°C) temperature scale, however, when dealing with gases we must talk about **absolute zero** and the **kelvin** (K) temperature scale.

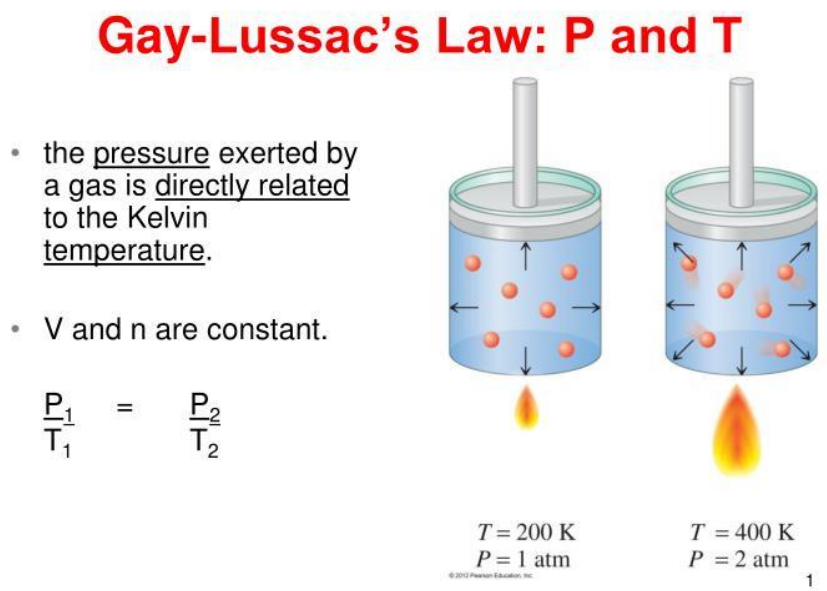
Absolute zero and the kinetic theory: according to the kinetic theory if the temperature of a gas is reduced, its molecules slow down striking the walls of the container they are in with less force; therefore the pressure drops. If you keep on cooling the gas down, its molecules would eventually stop moving and cause no pressure. This would happen at **-273 °C**, which is the lowest possible temperature and it is called **absolute zero**. (Below that temperature an ideal gas will turn liquid, and real gases turn liquid before absolute zero is reached). The kelvin temperature scale uses absolute zero as its zero (0 K), so to convert from Celsius (°C) to kelvin (K), you just add 273. (eg. 20 °C => 20 + 273 = 293 K).

- **Gay – Lussac's law or the pressure law** (linking pressure and temperature , at constant volume)

If you place a gas in a closed flask of fixed volume and increase the temperature of the gas, so does the pressure, and this relation is directly proportional, so:

For a fixed mass of gas at constant volume, the pressure is directly proportional to the Kelvin (absolute) temperature.

$$P_1 / T_1 = P_2 / T_2 \qquad \text{(at constant volume)}$$



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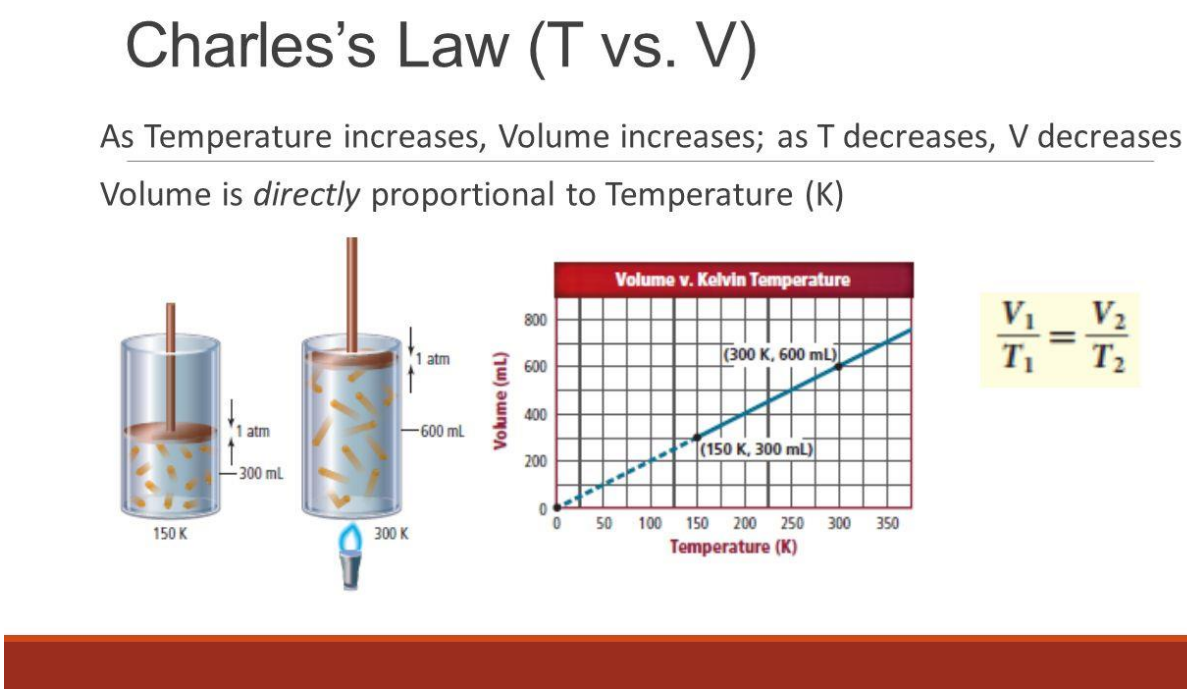
- **Charles's Law** (linking volume and temperature, at constant pressure)

If we take the same gas as in the example above, but now heat it at constant pressure, as the temperature of the gas rises, the volume of the gas also increases, so:

For a fixed mass of gas at constant pressure, the volume is directly proportional to the Kelvin (absolute) temperature.

$$V_1 / T_1 = V_2 / T_2 \quad \textbf{(at constant pressure)}$$ (Pople, 2007)



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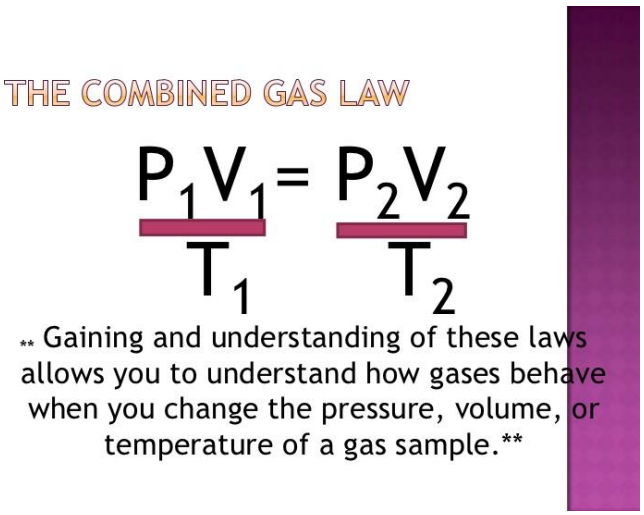
The three laws can be summarised as one mathematical equation where none of the variables remain constant.



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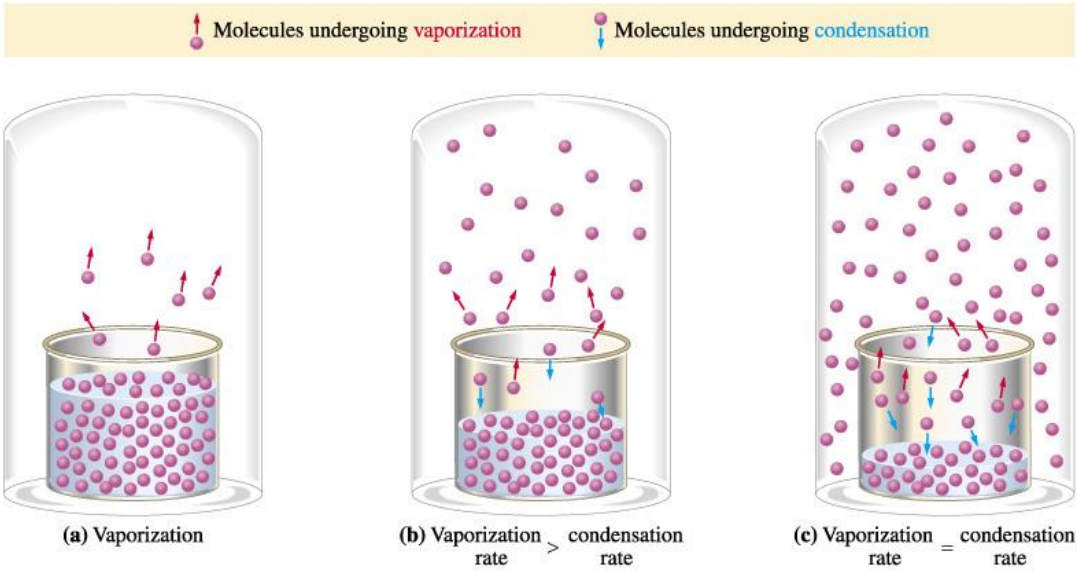
Vapour Pressure

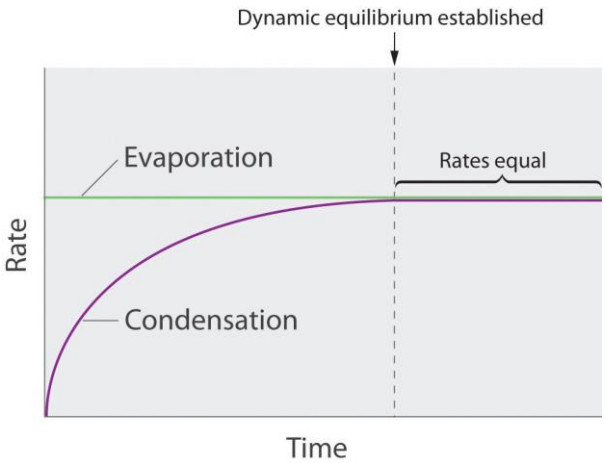
The vapour pressure of a liquid is the pressure exerted by the vapour when equilibrium is reached between its vapour phase and its liquid phase.



("CH105: Lesson 2 - Intermolecular Bonds", 2016)

So, at any given temperature, for a particular substance there is a pressure at which the vapour of that substance is in equilibrium with its liquid form (the same number of particles leaving the liquid by evaporation return to the liquid by condensation). This is termed the vapour pressure of that substance at that temperature.

Why does this happen? Well, as not all particles have the same kinetic energy, in all liquids at a given temperature, some particles will move faster than others. The faster ones and near the surface have enough energy to escape and turn into a gas – **evaporation** – As the number of these gas particles increases, so will the vapour pressure. Eventually a point will be reached where the inverse process will take place; gas particles will lose energy and will turn back into the liquid state – **condensation** - , reaching an equilibrium, when evaporation rate = condensation rate. At this point the pressure exerted by the vapour is called the vapour pressure.



("Vapor Pressure", 2016)

As temperature increases more molecules are able to escape the liquid, and as a consequence vapour pressure increases and vice versa. When the pressure reaches atmospheric pressure the entire liquid will start to boil. We can therefore say that the boiling point of a liquid is that at which its vapour pressure equals atmospheric pressure.

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