Chapter 5

States of Matter

You will be familiar with the three states of matter from your own experience and your work in previous years. We now take this model and explain some of the observed macroscopic properties of matter in more detail.

5.1 Solids, liquids and gases

We describe all of these in terms of ‘particles’, a deliberately vague term. Remember that we explain the large-scale properties (volume etc) in terms of the microscopic behaviour of the particles.

Solids

Solid substances have fixed shapes (assuming no resultant force is applied) and fixed volumes. We explain this by saying that the particles in solids are held together in a regular arrangement and that the movement of the particles is limited to vibration about a fixed point. The posh word for a fixed, regular arrangement is *lattice*. The particles must be very close together as the solid can not be compressed.

Liquids

Liquids have fixed volumes but will take the shape of the thing in which they are contained. The particles are close together, but unlike in solids they can move around one another - this is sometimes called *translation*. This motion of the particles means that substances can *diffuse* in liquids but not in solids.

Gases

Gases do not have fixed shapes or fixed volumes. They will expand to fill whatever they are put in. The particles in gases are not close together and are free to move around one another. Gases do exert *pressure* on their container - the constant motion of the particles against the walls of the container produces pressure. We will see later that the volume of a gas is mostly 'empty space', very little of the total volume is made up of the particles themselves. As in liquids the motion of the particles allows diffusion to take place.

5.1.1 Changes of state

The common changes of state are freezing / melting (for solids ↔ liquids) and boiling / condensing (for liquids ↔ gases). You should be able to describe these changes of state in terms of the changes in the arrangement, spacing and forces between the particles. Be careful when you do this! For instance, when a liquid boils the resultant gas will normally occupy a larger volume than the liquid because the spaces between the particles get larger - the particles themselves *do not expand.* You have also met sublimation, which is when a solid turns straight into a gas (we see this when iodine is heated and when aluminium chloride is heated - one of the clues that it is covalent rather than ionic).

The **temperature** of a substance is proportional to its (average) kinetic energy:

$$\mathrm{Temp}\alpha\frac{1}{2}mv^2 \tag{5.1}$$

If we heat up e.g. a solid we increase the average kinetic energy of the particles. The temperature increases. The particles eventually have enough energy to overcome the forces which were previously holding them together - the substance melts. While energy is being used to overcome these forces the kinetic energy of the particles does not increase - the temperature of the substance does not go up. Similar changes occur when a liquid boils.

5.1.2 The Maxwell-Boltzmann distribution

Consider a group of people ice-skating. They are inebriated. They travel in straight lines and bump into each other, and the walls of the rink, a lot. Some of them will travel very quickly until they crash, others will move inexorably slowly before colliding with something. After a collision they may set off in another straight line, quickly or slowly, until they crash into something else. Although they have different individual velocities (these are straight-line speeds) the group will have a roughly constant average velocity. This is similar to the motion of the particles in a liquid or gas. The average velocity gives us a measure of their average kinetic energy or their temperature, but even at this particular temperature the individual velocities of the particles can vary a lot. This distribution of velocities is shown by the Maxwell-Boltzmann curve[1] (figure 5.1).

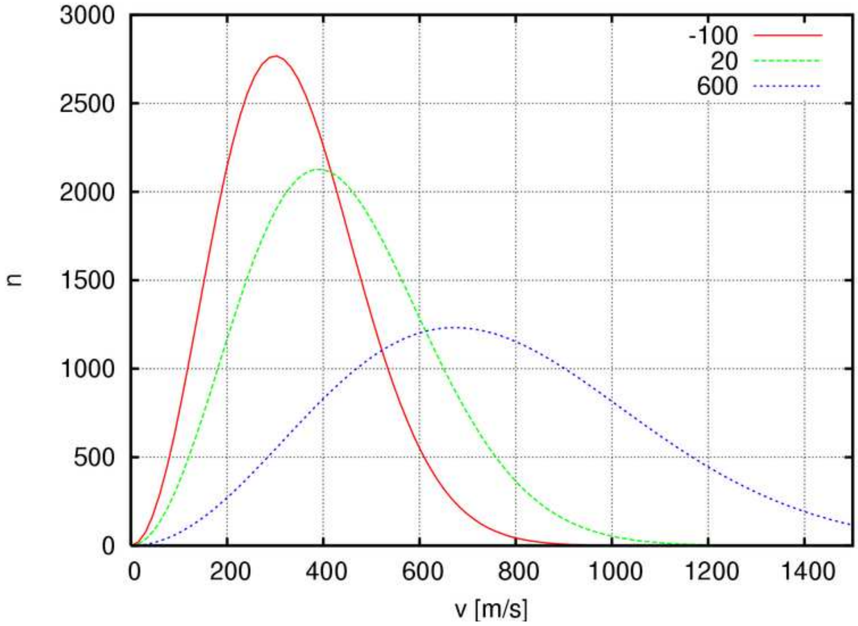

Figure 5.1: The Maxwell-Boltzmann distribution for different temperatures

The distribution shows how many particles have a certain velocity or kinetic energy. Heating the substance up has the effect of squashing and broadening the curve - the total number of particles does not change so the area under the curve must remain constant. There will be a minimum energy required for the particles in a liquid to escape from the liquid and form a vapour, and there will always be some particles which have this amount of energy. Heating increases this number, so hot water will vaporise more rapidly than cold water. When particles escape to form vapour the average energy of the particles in the remaining liquid is less than before, which is why evaporation has a cooling effect - you may have experienced this when using nail polish remover which is very volatile, it is also the reason that you sweat (although I'm sure that *you* don't).

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[1] $f(E) = A\mathrm{e}^{-kT}$

5.2 Gas laws and the Ideal Gas Equation

There are 3 main gas laws, **Boyle's law**

$$P_1V_1 = P_2V_2 \text{ at constant temperature} \tag{5.2}$$

Charles' law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ at constant pressure} \tag{5.3}$$

and the **Pressure law**

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ at constant volume.} \tag{5.4}$$

These can be combined in an equation which connects pressure, volume, temperature and amount of substance. This is called the **ideal gas equation**:

$$PV = nRT \tag{5.5}$$

In this equation it is usual to use the following units for the quantities:

pressure	kPa
temperature	K
volume	dm^3
n	mol

in which case $R = 8.314$ J K^{-1} mol^{-1}.

Ideal gases

An ideal gas is one in which the particles occupy no space, have no attractive forces between them and in which the kinetic energy of the particles is proportional to the absolute temperature. Using the ideal gas equation is always going to be an approximation because no real gases have these properties. It becomes an especially poor approximation for gases with strong forces between the molecules (such as hydrogen bonds), and at low temperature and high pressure, when the particles are squashed close together but have little energy compared to the energy of the forces between them.

The ideal gas equation can obviously be used to find one term if all of the others are known (or remain constant).

5.3 Phase diagrams

The phase of a substance may easily be changed by altering temperature or pressure. A phase diagram shows how the state changes as a function of these two variables. These diagrams are specific for particular substances.

Phase diagrams show that there are conditions under which all three phases are stable and in equilibrium - this is called the *triple point.* There is also a relatively high temperature beyond which the liquid and gas phases become ill-defined. The temperature and pressure at which this happen are called the critical temperature (T_c) and the critical pressure (P_c). Beyond this *critical point* it is sometimes said that the substance exists as a supercritical fluid.