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Chapter 6

PARTICLES IN MOTION

From the 17th century on, physicists were aided in their interpretation of nature by a "kinetic theory" of matter as consisting of very small particles, or molecules, in very rapid random motion. With this theory, external properties of matter could be explained by the internal behavior of particles in motion.

The easiest illustration is provided by gases, which have two major properties: 1) they tend to expand until they completely fill any container into which they are introduced, and 2) they are easily compressed under normal conditions, and have a density about 1000 times less than liquids.

These properties are readily explained in terms of the kinetic theory. A gas tends to fill its container because it is made of a huge number of molecules free to move very rapidly at random in all directions. A gas can be easily compressed because, normally, its molecules are very widely separated by empty space.

At the other extreme, the molecules of a solid are strongly bonded together by cohesive forces, which, as we will see, are electrical in nature. These cohesive forces are not as strong in the case of liquids, which, accordingly, have properties that are intermediate between gases and solids. The molecules of a liquid are almost in contact with one another, but are free to slide over one another. Consequently, a liquid is hard to compress, but free to assume the shape of its container.

Molecules have limited mobility in liquids, and essentially none in solids. At all times, however, there is motion *within* each molecule, with its constituent atoms vibrating back and forth. Typical vibrations are in the order of 10 million millions per second!

The Empirical Study of Gases

Important early clues about the nature of matter were provided by the empirical study of gases, which are much easier to investigate than liquids or solids.

In the 17th century, the Anglo-Irish chemist Robert Boyle experimentally studied the behavior of a gas contained in a cylinder with a movable piston, see **Figure 6.1**.

Later, physicists established empirically a general law, called the "ideal gas law". Provided the pressure is not too high nor the temperature too low, this law states that

> the product of (the pressure exerted by the gas) times
> (the volume of the gas)
> is proportional to (the absolute temperature of the gas).

Statistical Mechanics

The rather simple external behavior described by the ideal gas law is in sharp contrast with the chaotic situation that, according to the kinetic theory, exists inside a gas, where countless molecules colliding with one another move in all directions, in zigzag fashion.

If the velocities of all the molecules could be specified at a given instant, and if the interactions among the molecules were known, then, *in principle*, the entire future course of the molecules could be calculated using Newton's laws. In practice, however, this is totally impossible because of the huge number of molecules, even in a small volume of gas.

Taking advantage of the very fact that so many molecules are involved, a solution was found by adopting a statistical approach applying the mathematical laws of probability. This approach is the basis of "statistical mechanics". It cannot tell us what happens to individual molecules, but it can derive very useful conclusions about their collective behavior.

When the statistical behavior of the gas molecules inside a cylinder is analyzed mathematically, and the result is compared with the ideal gas law, a very important fact emerges. The external property of temperature is related internally to the random motion of the molecules: the absolute temperature of a gas is proportional to the average kinetic energy of its molecules. We talk of average kinetic energy because the molecules in their random motion have speeds that vary over some range of values.

Heat then is the kinetic energy of particles in random motion, and temperature is a measure of their average kinetic energy. The greater the molecular agitation, the higher is the temperature we measure.

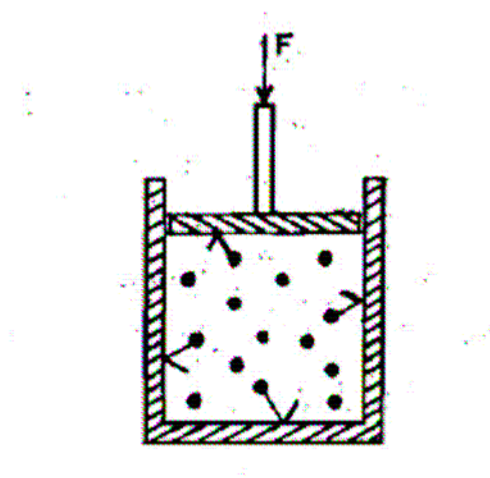

Figure 6.1 - Gas under pressure in a cylinder with a movable piston

If a force bears down on the piston, the piston applies to the gas a "pressure" measured, for instance, in pounds per square inch. The piston will settle at some position where the pressure exerted by the piston on the gas is balanced by the pressure exerted by the gas against the piston. Pressure is also exerted against all parts of the cylinder with which the gas is in contact.

Empirically, Boyle discovered that, provided the mass and the temperature of the gas are kept constant, the product of the pressure times the volume of the gas in the cylinder remains constant, as different forces are applied to the piston. This means that, if we double the pressure, for instance, the volume of the gas is reduced to half its previous value; if we halve the pressure, the volume doubles.

The kinetic theory provides an easy explanation of "Boyle's law". The pressure by the gas against the walls of the cylinder results from countless molecules in random motion continually colliding with, and bouncing away from, either the walls or other molecules. If the volume of the gas is reduced to, say, half, the molecules are packed more closely so that there are twice as many of them per cubic inch as before, anywhere in the cylinder. They will strike the walls of the cylinder, on the average, with the same impact as before (provided the temperature is kept constant), but twice as many collisions will occur, thus doubling the pressure.

Brownian Motion

As late as the first decade of the 20th century, the atomic nature of matter was not unanimously accepted. Among the dissidents were distinguished scientists who argued that it was not necessary to assume the existence of atoms.

Direct conclusive evidence was finally provided around 1908 when the French physicist Jean Perrin, using an ultramicroscope, confirmed experimentally Albert Einstein's theory for the so-called "Brownian motion", named after the Scottish botanist Robert Brown. In 1827, Brown observed under a microscope a rapid jiggling motion of pollen grains suspended in water. Later, he helped to demonstrate that the jiggling motion was not due to any living organisms.

Using statistical mechanics, Einstein was able to develop a quantitative theory, which attributed the jiggling of microscopic particles to their being continuously bombarded by the much smaller molecules of the medium in which the particles were suspended.

To convey some idea of how small and how numerous atoms are, Lord Kelvin used the following example. "Suppose that you could mark all the molecules in a glass of water; then pour the contents of the glass into the ocean and stir the latter thoroughly so as to distribute the marked molecules uniformly throughout the seven seas; if then you took a glass of water anywhere out of the ocean, you would find in it about a hundred of your marked molecules." [1]