Gas Pressure Encyclopedia Article

Gas Pressure

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Gas Pressure

According to the kinetic theory of gases, gas **pressure** is the **force** exerted by a gas on a unit area. In a gaseous systems the gas molecules are in continual **motion** and the hotter the gas the faster its molecules are moving. When these moving molecules (or atoms) strike an object they exert a force upon the object. Accordingly, gas pressure is an indirect measure of the kinetic **energy** of gas molecules and is a direct measure of the sum of the forces of collision over a defined unit area.

The speed with which a gas molecule moves is directly related to its **temperature**. The greater the rate of movement (molecular **velocity**) the greater number of molecular collisions with a given surface area. In addition, greater molecular velocity imparts a greater **momentum** at impact and means that the molecules carry a increased **kinetic energy** into collisions. All of these factors contribute to an increased gas pressure.

Just as pressure and temperature are connected, so are pressure and volume. Reducing the volume of a trapped gas means a shorter distance for the gas molecules to travel before colliding with the walls of the container. Trapping gas inside a container and then reducing the volume of the container will mean an increase in gas pressure. In fact, reducing the volume by one half will double the number of collisions and thus double the pressure. This is the basis of **Boyle's law** formulated by English physicist **Robert Boyle** (1627-1691) in 1662. Boyle's law states that, for an ideal gas at constant temperature, the pressure (p) of a gas varies inversely with volume (v). In equation form pv = k, a constant. Boyle's law is broadly applicable to real gases at lower pressure but fails to provide accurate results at high pressures or temperatures.