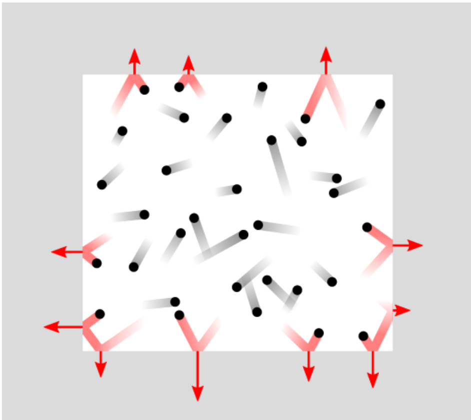
Kinetic-Molecular Theory

> **Skills to Develop**
>
> - Define the kinetic-molecular theory and its relationship to the ideal gas equation

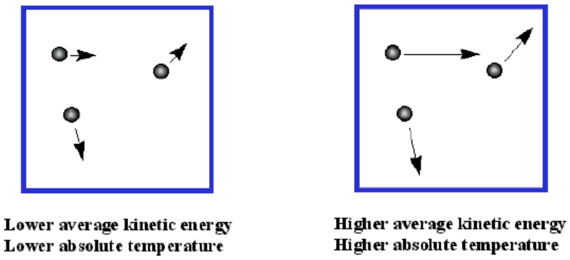
Boyle's Law was published around 1660. In 1718, a mathematician named Bernoulli proposed an explanation for Boyle's Law. Although this was almost a hundred years before Dalton's Atomic Theory, atomistic theories (also sometimes called "corpuscular theories") had been around a long time. Boyle himself had made some arguments similar to Dalton's that he borrowed from Sennert. Bernoulli assumed that the gas was made of many small particles moving quickly. They move straight until they bump into another particle or a wall, then they bounce off according to conservation of momentum. Pressure comes from the impact when these particles bump into the walls.



An illustration of Bernoulli's explanation of gas pressure.

How does the pressure depend on the volume? Bernoulli gives this explanation. If we compress a gas, the particles will bump the walls more often. This happens for 2 reasons: first, there are more particles in the layer next to the wall, where they can bump it. Second, particles moving away from the wall are more likely to bump into another particle, change direction, and bump the wall again. Imagine we have a cubic container, with each side length s. If volume is decreased from 1 to s^3, then the number of particles in the layer next to the wall increases by $s^2/1$. Also, the number of collisions between the wall-layer particles and the wall increases by s/1. Combining these, the number of collisions increases by s^3 when the volume decreases by $1/s^3$. The pressure is the number of impacts multiplied by the momentum of the particles, mv, where m is the mass of a particle and v is the average velocity. When you calculate the average momentum change from each collision and the average number of collisions per area of wall, the result is $P = nmv^2/3V$, where n is the number of particles and V is the volume. You can see that this matches Boyle's Law: $PV = nmv^2/3$ = constant. Later, mv^2, the kinetic energy, was shown to be proportional to the temperature: $kT = mv^2$. This is the ideal gas equation.

How fast do gas particles move? Because the average kinetic energy is proportional to the temperature, heavier gases move slower than light gases at the same temperature. There will be a big range of speeds for different molecules, because they change speed as they bump off each other. For N_2 at 0 °C, the range might be 0-1300 m/s, with the average speed about 500 m/s. As the temperature gets bigger, the range gets bigger and so does the average speed. So the particles are usually moving very fast! How far do they go between collisions? This depends on the conditions, but the average distance between collisions might be 10^{-7} m, so not far!



Thus, we imagine a gas as many small particles that bump off each other perfectly elastically (with conservation of momentum). They behave like hard little balls, and don't attract each other at all. Their kinetic energy depends on the temperature. In the derivation of the ideal gas law, we assume that there are no attractive forces between the particles and that the particles don't take up any space. These two assumptions are obviously incorrect: if there are no attractions between particles, there would be no liquids or solids. Likewise, the particles do take up a little space. Since we know that attractive forces become important at low temperatures, and that the volume of the particles will be important we the volume is relatively low (meaning pressure is high) we can predict that the ideal gas equation works best a high temperatures and low pressures.

Outside Link

- CrashCourse Chemistry: Passing Gases - Effusion, Diffusion, and Velocity of Gases (12 min)

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