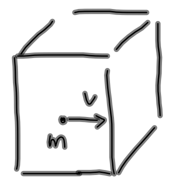
Macroscopic behaviour of an ideal gas

How do the microscopic properties of an ideal gas explain the macroscopic properties?

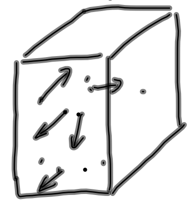
Consider a gas confined in a container with particles moving back & forth and colliding with the walls.

Consider the effect of one particle and then the overall effect for all the particles.

one particle

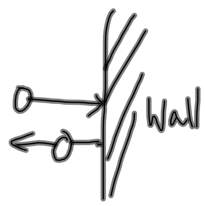


a large number of particles.



<u>Pressure</u>

- caused by the particles colliding with the wall and rebounding elastically.
- move at a constant speed since no force between them as they move across the container (Newton's First Law)
- when the particle collides with the wall:



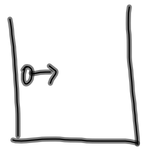
the force of the wall on the particle causes it to rebound

(this force is equal to the rate of change in the momentum of the particle) (Newton's 2nd Law)

- By Newton's 3rd Law the particle exerts an equal but opposite force on the wall
- The average force on the wall is the sum of all the forces averaged over time.
- the pressure is the average force per surface area. (due to the collisions with the wall)

The pressure Law ($P \propto T$, M and V are constant, T is in kelvin)

- increase the temperature of a fixed mass of gas at constant volume
- the mean kinetic energy increases
- the speed increases.
- increases the momentum of the particles. ($p \propto v$)
  (in proportion to the speed, v)
- as the particles move back & forth it takes less time to reach the other side (i.e. less time between collisions or the frequency of the collisions increases)



time to cross the container → $\Delta t \propto \frac{1}{v}$

- when a particle collides with the wall, the rate of change in momentum is increased $\left(\frac{\Delta P}{\Delta t}\right)$ ← time in contact with the wall
  The change in momentum is proportional to v^2

$\Delta t \propto \frac{1}{v}$ $\quad v^2$

$\frac{mv-(-mv)}{1/v} = 2mv(v)$

- by Newton's 2nd Law $F = \frac{\Delta P}{\Delta t}$,
  there for the force $\propto v^2$ (of the wall)
- the force of the particles on the wall $\propto v^2$
- the pressure on the wall by the particle $\propto v^2$
- the absolute (Kelvin) temperature of the gas depends on the mean translational kinetic energy, $\frac{1}{2}mv^2$ (i.e. $T \propto v^2$)

$\therefore$ $P \propto T$ ← kelvin

Pressure Law

If the temperature of a gas increases, the pressure of the gas increases because the particles move faster and so they hit the wall ① harder and ② more often.

Boyles law ($P \propto \frac{1}{V}$ where m and T are constant)

- if decrease the volume of the container, decreasing the time interval between collisions with the walls (less distance to travel)
- the momentum change (Δp) of each particle with the wall remains the same (since T is constant, v is constant)
- since the time interval between collisions has decreased, the particles collide more frequently with the wall.
- Since the collisions are more frequent, the average force on the particle increases.
- By Newton's 3rd Law, the force on the wall also increases
- Therefore the pressure increases as the volume decreases

$$\left(P \propto \frac{1}{V}\right)$$

Charles's Law ($V \propto T$, m and P constant, T is in kelvin)

* container that is able to expand

- increase the temperature → increase in KE (mean) → increase in v (mean)
- increase in momentum → increase in $\frac{\Delta p}{\Delta t}$ → increase in the force on the particle → increase in the force on the wall → increase the pressure if the container remained the same size.

to "relieve" this pressure, the container is allowed to expand.

① (less collisions per unit surface area) (force)

also increases the time interval between collisions with the wall → collisions occur less frequently ②

Adiabatic Process

An adiabatic process is one in which there is no flow of heat into or out of the system.

If the temperature does not stay constant, then Boyle's Law ($p \propto \frac{1}{V}$) does not apply. => bicycle pump.

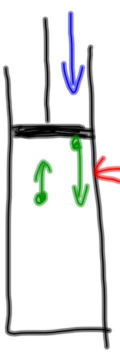
If we want to compress the air in a bicycle pump adiabatically -> none of the heat generated passes out of the pump.

This would happen if the air was compressed quickly or if the pump were thermally insulated.

The temperature of the air increases (from the added KE energy from motion of piston)

The pressure increases, but Boyle's Law does not apply. (T not constant)

=> the air has been adiabatically compressed



particle rebounds from the piston at a higher speed $\therefore$ more KE $\therefore \uparrow T$