Gases $\rightarrow$ up to this point, we've been dealing with compounds in condensed phases; strong IM Forces keep these molecules interacting

As you give molecules energy (Kinetic Energy via heat), IMF are broken.

Gases Form and have KE $\quad KE = \frac{1}{2}mv^2$

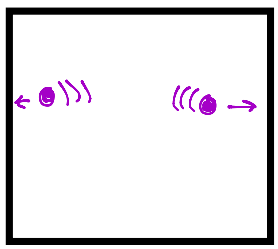
As molecules gain energy (KE), the velocity increase $\rightarrow$ Keep this in mind!

The most fundamental measurable property of gases is Pressure

conceptually:

pressure is how hard a molecule in motion "hits" a surface

- dependent on how big the surface is



Mathematically: $P = \frac{\text{Force}}{\text{Area}}$ $\quad F = ma$

surface area exposed to "hits" by gas molecules

mass kg $\quad$ acceleration $= \frac{\Delta v}{\Delta t} \rightarrow \frac{m/s}{s} = m/s^2$

$$\text{Force} = \frac{kg\,m}{s^2}$$

$$P = \frac{ma}{A} = \frac{kg\,m/s^2}{m^2} = \frac{kg}{m\,s^2} = \text{Pascal (Pa)}$$

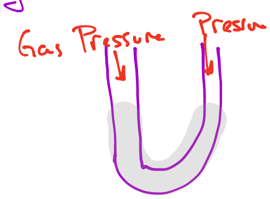
Recall that $\text{Joule} = \frac{kg\,m^2}{s^2}$ $\rightarrow$ $\frac{J}{m^3} = \frac{kg}{m^2 s^2} = \text{Pascal}$

$(KE = \frac{1}{2}mv^2)$

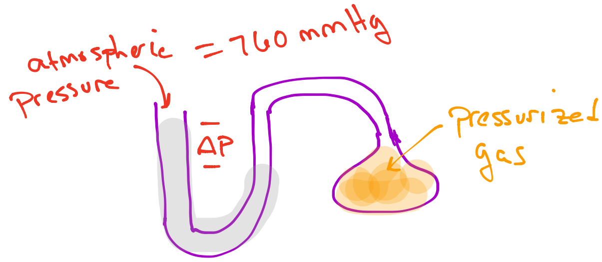
So Pressure is $\frac{\text{Energy}}{\text{Volume}}$

$P = \frac{KE}{V} \implies$ We'll come back to this... all day

Measuring Pressure: Manometer: U-shaped tube filled with liquid



- Gases strike the surface of the liquid on each side of tube.
  - If both pressures are the same, then the level of liquid is the same on both sides



If the pressure on one side of the tube is higher, the level of liquid will adjust

The difference in the level between the two side is directly related to the difference in pressure.

If the liquid is Hg, then the pressure difference will be in mmHg (common unit of pressure)

* atmospheric pressure = 760 mmHg *

So lets say you have the above experiment set up + the ΔP = 42 mmHg. What is the pressure in the flask?

ΔP = higher pressure − lower pressure

since the gas in the flask is "pushing" harder, it is @ higher pressure

42 mmHg = P − 760 mmHg

atmospheric pressure

P = 802 mmHg

Common Units of Pressure

mmHg → because it is historically the way pressure could be measured

Torr → 1 Torr = 1 mmHg

Atmospheres (atm) 1 atm = 760 mmHg = atmospheric pressure @ sea level

Bar (a scientist needs a place to hang out) 1 atm = 1.01325 bar

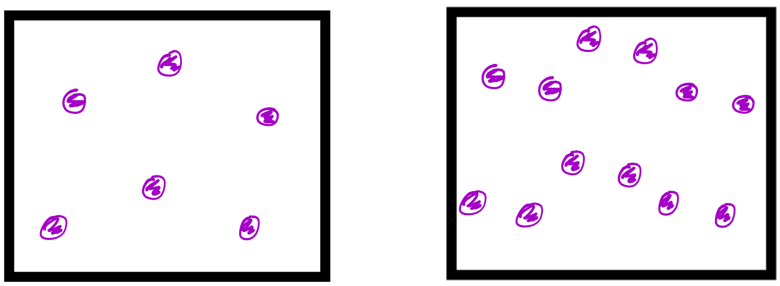
Pascal (Pa): SI unit 1×10^5 Pa = 1 bar

OK, let's think about ways to chang pressure.

$$P = \frac{F}{A}$$

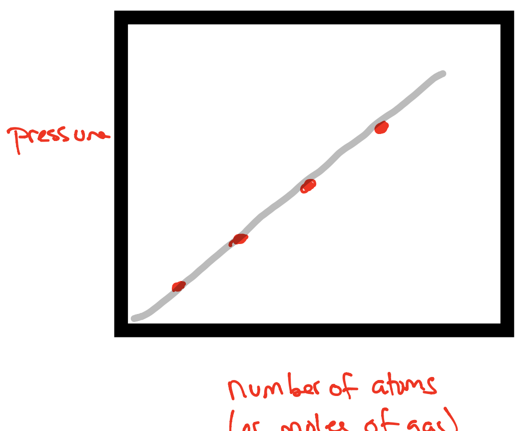
1. Change the number of collisions with the surface
2. Change the Area that gets hit
3. Change the Force that molecules strike the surface with

1 → this can happen if you increase the number of gas molecules contained in the box



Low Pressure

more atoms = more collisions = higher pressure



As the number of molecules in a chamber increases, the Pressure also increases

$$P \propto n$$

or

$$P = (\text{constant})\, n$$

this would be the slope of the line!

$$P = (\text{constant})\, n$$

$$\frac{P}{n} = \text{constant}$$

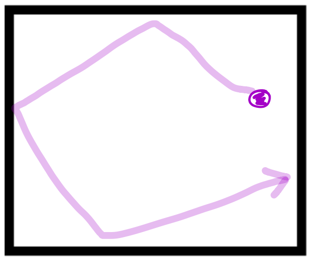
$$\frac{P_1}{n_1} = \frac{P_2}{n_2}$$

Sample problem: If a flask contains 100 moles of gas at 4 Pa, how many moles would be needed to increase the pressure to 400 Pa?

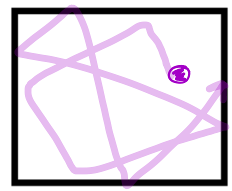
$$\frac{4\ \text{Pa}}{100\ \text{moles}} = 0.04\ \text{Pa/mol} = \frac{400\ \text{Pa}}{n}$$

n = 10,000 mols

② Change the area that gets hit : we can do this by decreasing volume



This molecule is travelling @ a certain velocity ($KE = \frac{1}{2}mv^2$) in a certain amount of time, it strikes a wall 4 times



If the volume decreases, the number of collisions increase. ↑ # of collisions = ↑P

So... as Volume ↑, Pressure ↓

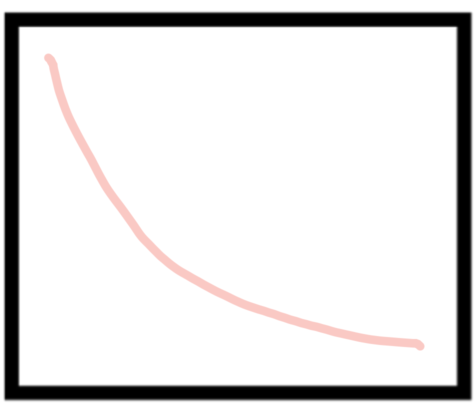
$P \propto \frac{1}{V}$ or $P = \frac{\text{constant}}{V}$

Remember above, we saw that $P = \frac{\text{energy}}{V}$

yeah!

So this constant has the unit of energy!

This is Boyle's Law



Pressure

Volume

So $P = \frac{\text{constant}}{V}$

∴ PV = constant

We can use this relationship to calculate the change in pressure as volume is changed

Example: P = 14 atm V = 1 L

if we change the volume to 4 L, what is the pressure?

14 atm (1L) = 14 L·atm ← this would be the constant

14 L·atm = P (4L) P = 3.5 atm

Boyle's Law $P_1V_1 = P_2V_2$

③ Change the force that molecules strike the surface with

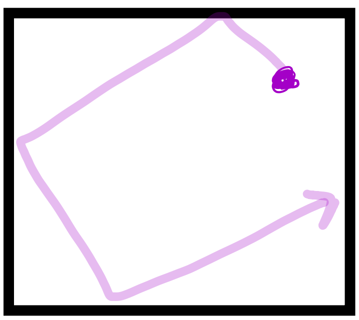
- Remember that $F = m\left(\frac{\Delta v}{\Delta t}\right)$ ← So $\Delta v > 0 \Rightarrow \uparrow F$

How can we increase the velocity of a gas?

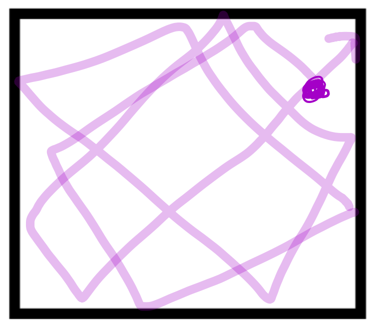
Give it more energy!

$KE = \frac{1}{2} m v^2$

We can give molecules more KE by heating them up



lower Temp, less energy, lower velocity, fewer collisions

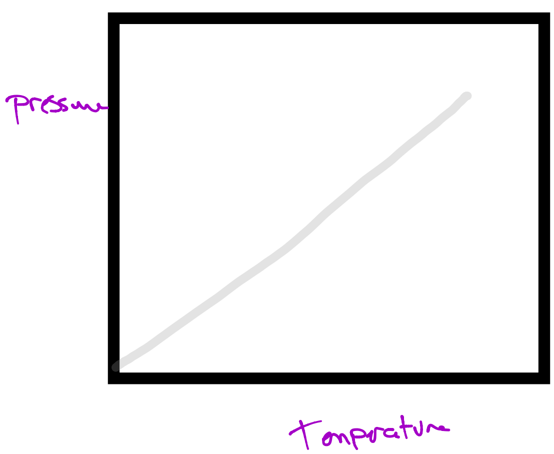


higer temperature, more KE = higher velocity = more collisions

$\uparrow T = \uparrow P$ $\qquad P = (\text{constant}) T$ $\qquad \frac{P}{T} = \text{constant}$

Charles' Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$



So collectively, our relationships are :

$PV = \text{constant}$

$\frac{P}{n} = \text{constant}$

$\frac{P}{T} = \text{constant}$

Grouped together : $\frac{PV}{nT} = \overbrace{\text{constant}}^{\text{let's call it } R}$

From this expression, we can see all of our predicted relationships and make a few more

$PV = RnT$

$\frac{P}{T} = \frac{Rn}{V}$

$\frac{P}{n} = \frac{RT}{V}$

* if the variables in red are held constant, our predicted relationships are true

Avogadro's Law

$\frac{V}{n} = \frac{RT}{V}$

$\frac{V}{T} = \frac{Rn}{P}$