Gases - Pressure = $\frac{\text{Force}}{\text{Area}}$

units of pressure Pascals (Pa) $1 Pa = \frac{1 N}{m^2}$ force area

1 atm = 101325 Pa

1 bar = 100000 Pa

1 atm = 760 mmHg

- Volume
- Temperature
- how much (moles)

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Relationships btw these quantities

If T, moles are constant how are P and V related

$$P \propto \frac{1}{V}$$

$$P \times V = \text{constant}$$

$$P_i V_i = P_f V_f$$

} Boyle's Law

If P, moles are constant how are volume + Temp related

$$V \propto T$$

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

$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

} Charles' law

If P, T are constant, how is Volume + # moles related?

$$V \propto n \quad \text{\# of moles}$$

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

$$\frac{V_i}{n_i} = \frac{V_f}{n_f}$$

} Avogadro's Law

$V \propto \frac{1}{P}$
$V \propto T$
$V \propto n$

$V \propto \frac{Tn}{P}$ → need prop. constant $V = \ldots$

R = gas constant

$V = \left(\frac{Tn}{P}\right) R$

$\boxed{PV = nRT}$ - Ideal gas law

$R = 0.08205 \frac{L\cdot atm}{mol\cdot K}$

Ex: If 25.0g $N_2$ in a container w/ V = 1.45L @ 37°C what is the pressure?

* Use Kelvin

n? $25.0g\ N_2 \times \frac{1 mole\ N_2}{28g} = 0.893$ moles

T? $273.15 + 37 = 310.15\ K$

$PV = nRT \rightarrow P = \frac{nRT}{V} = \frac{(0.893\ \cancel{moles})(0.08205 \frac{\cancel{L}\cdot atm}{\cancel{mol}\cdot \cancel{K}})(310.15\ \cancel{K})}{1.45\ \cancel{L}}$

P = 15.7 atm

Ex: 100g $C_8H_{18}$ (liquid) in a container with V = 2.5L
also in container is $O_2$ and P = 4.0 atm
How many gram $CO_2$ from reaction? T = 25°C

$C_8H_{18} + \frac{25}{2}O_2 \rightarrow 8CO_2 + 9H_2O$

$100g\ C_8H_{18} \times \frac{1 mol}{114g} = \underline{0.88\ moles\ C_8H_{18}}$

$\underline{PV = nRT}$

$n_{O_2} = \frac{PV}{RT} = \frac{(4.0\ atm)(2.5L)}{(0.08206 \frac{L\cdot atm}{mol\cdot K})(298\ K)}$

$O_2 = LR$

$n_{O_2} = 0.41$ moles

moles $CO_2$ $\quad 0.41 \text{ moles } O_2 \times \frac{8 \text{ mol } CO_2}{12.5 \text{ moles } O_2} \times \frac{44 \text{ g } CO_2}{\text{mol}} = 11.5 \text{ g } CO_2$

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When you have a mixture of gases

$$P_{tot} = \underbrace{P_{N_2} + P_{O_2} + P_{CO_2} + P_{H_2O}}_{\text{"partial pressure"}}$$

$P_{CO_2}$ = "partial pressure" of $CO_2$

= pressure of that # moles of $CO_2$ if nothing else was around

$$\text{Mole fraction } (X) = \frac{\text{\# of moles of one thing}}{\text{total \# moles of everything}}$$

$P_{CO_2} = P_{tot} X_{CO_2}$

($P_{tot}$: total P of everything together)

25' x 40' x 10'

750 cm x 1200 cm x 300 cm

$V_{room} \approx 270{,}000$ L

$P_{tot} = 1$ atm

$T = 295$ K

$$n = \frac{PV}{RT} = \frac{(1 \text{ atm})(270{,}000 \text{ L})}{.08206 \frac{\text{L atm}}{\text{mol K}} \cdot 295 \text{ K}} = 11153 \text{ moles of gas}$$

2230 moles $O_2$

14 mol $CO_2$

8909 moles $N_2$

$$X_{CO_2} = \frac{14}{11153} = 0.00125$$

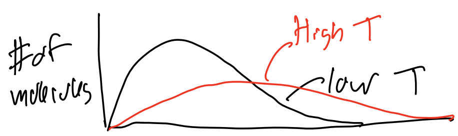
$$P_{CO_2} = (1 \text{ atm})(0.00125) = 0.00125 \text{ atm}$$

$$P_{O_2} = (1 \text{ atm})\left(\frac{2230}{11153}\right) = 0.199 \text{ atm}$$

Kinetic molecular theory of gases

- gas molecules are in constant motion
- Velocity of molecules $\propto$ Kinetic energy $\quad KE = \frac{1}{2}mv^2$
- Kinetic energy of gas molecules $\propto$ Temp

* At a given temp there is going to be a distribution of speed (KE)




For 1 molecule $\quad KE = \frac{1}{2}mv^2$

mass of one molecule

velocity of particular molecule under consideration

For 1 mole of molecules - can talk about average KE and average velocity

$$KE_{avg} = \frac{1}{2}mv_{avg}^2$$

It can be shown that $KE_{avg} = \frac{3}{2}kT$

Boltzmann constant $1.381 \times 10^{-23} \frac{J}{K}$

so $\frac{3}{2}kT = \frac{1}{2}mv_{avg}^2$

$$\frac{3kT}{m} = v_{avg}^2 \quad \text{or} \quad \sqrt{\frac{3kT}{m}} = \sqrt{v_{avg}^2}$$

Can relate KE and velocity to pressure
- Pressure due to molecules colliding with container
- higher KE $\rightarrow$ more pressure

Helps explain diffusion and effusion
- How rapidly gas molecules spread out

- Related to how fast they are moving

Larger molecules with larger M are moving slower (ie V is smaller) at given temp

Higher temp → more KE → higher velocity

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Non-ideality of gases

Ideal gas ~~ea~~ law assumes - gas molecules are infinitely small

- gas molecules don't interact

At high pressure - both of these assumption become invalid and ideal gas law becomes un-useful

VanderWaals eq

$$\left(P + \frac{an^2}{V^2}\right)(V - bn) = nRT$$

$\approx P$ $\quad$ $\approx V$

correction for interactions

correction to account for non-zero size

a, b are constants for each different gas