Gases

Pressure is key quantity $= \frac{\text{Force}}{\text{Area}}$

units of pressure

Pascals (Pa) $\quad 1\ Pa = \frac{1\ N}{m^2}$

1 atmosphere (atm) = 101325 Pa

1 bar = 100,000 Pa

1 atm = 760 mmHg

Volume
Temperature (K)
Moles

Relationships btw P, V, T, moles (n)

(1) Constant T, moles – how are P and V related

P T V ↓ $\quad P \propto \frac{1}{V}$ (P is proportional to 1/V)

$P \times V = \text{constant} \quad P_i V_i = P_f V_f$

"Boyle's Law"

(2) constant P, moles – how V and T are related

V T T ↑ $\quad V \propto T$

$\frac{V}{T} = \text{const} \quad \frac{V_i}{T_i} = \frac{V_f}{T_f}$ NO – T_f

"Charles' Law"

(3) const P, T – how are V and n related

$n \uparrow \; V \uparrow \qquad V \propto n$

$\frac{V}{n} = \text{constant} \qquad \frac{V_i}{n_i} = \frac{V_f}{n_f}$

"Avogadro's Law"

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

$V \propto \frac{Tn}{P}$ → need proportionality constant

R = gas constant

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

$\boxed{PV = nRT}$ – Ideal gas Law

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

*Need to use Kelvin

Ex: If we have 25.0 g of N_2 in a container w/ V = 1.45 L at 37 °C, what is the pressure of this gas

$PV = nRT \rightarrow P = \frac{nRT}{V}$

n: $25.0\ g\ N_2 \times \frac{1\ mol}{28\ g} = 0.893\ moles\ N_2$

V: 1.45 L

T: 37 °C + 273.15 → 310.15 K

$P = \frac{(0.893\ moles)(0.08205 \frac{L \cdot atm}{mol \cdot K})(310.15\ K)}{1.45\ L} = 15.7\ atm$

Ex: Container w/ V = 2.5 L

100 g of C_8H_{18} in container

O_2 inside container also — P = 4.0 atm

Temp 25°C

Do combustion

How many grams of CO_2 gas is formed?

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

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

$$n_{O_2} = \frac{PV}{RT} = \frac{(4\text{ atm})(2.5\text{ L})}{\left(0.0820\frac{\text{L atm}}{\text{mol K}}\right)(273.15+25\text{ K})} = 0.41 \text{ moles } O_2 \quad \text{(LR)}$$

$$0.41\text{ mol } O_2 \times \frac{8\text{ mol } CO_2}{12.5\text{ mol } O_2} \times \frac{44\text{ g}}{1\text{ mol}} = 11.5\text{ g } CO_2 \text{ produced}$$

If we let container cool back to 25°C what is pressure inside container after reaction

only CO_2 gas now, no more O_2

$$11.5\text{ g } CO_2 \times \frac{1\text{ mol}}{44\text{ g}} = 0.26\text{ mol } CO_2$$

$$PV = nRT \rightarrow P = \frac{nRT}{V} = \frac{(0.26\text{ mol})\left(0.0820\frac{\text{L atm}}{\text{mol K}}\right)(298.15\text{ K})}{2.5\text{ L}}$$

$$P = 2.56 \text{ atm}$$

Mixtures of gases in this room

$P_{tot} = P_{N_2} + P_{O_2} + P_{CO_2} + P_{H_2O}$

"partial pressures"

P_{O_2} is partial pressure of O_2 — what pressure would be if all other gases were absent

X_n = mole fraction of "n" = $\frac{\#\text{ moles of "n"}}{\text{total moles of everything}}$

$P_{tot} X_n = P_n$

If 21 moles O_2 and 79 moles N_2 in room and $P_{tot} = 1.0$ atm

$X_{N_2} = \frac{79}{100} = 0.79$ $P_{N_2} = (1.0\text{ atm})(0.79) = 0.79$ atm

$X_{O_2} = \frac{21}{100} = 0.21$ $P_{O_2} = (1.0\text{ atm})(0.21) = 0.21$ atm

Room 40 × 40 × 17 Feet

$X_{O_2} \approx 0.21$

1220 × 1220 × 366 cm

$V \approx 5.5 \times 10^5$ L

$P_{tot} = 1$ atm

$T = 298.15$ K

$PV = nRT$

$P_{O_2} = 0.21$ atm

$\text{moles } O_2 = \frac{(0.21\text{ atm})(5.5 \times 10^5\text{ L})}{(0.08205 \frac{\text{L atm}}{\text{mol K}})(298.15\text{ K})}$

≈ 4700 moles O_2

$\text{moles total} = \frac{PV}{RT} = \frac{(1\text{ atm})(5.5 \times 10^5\text{ L})}{.08205 \times 298.15}$

22500 moles of gas

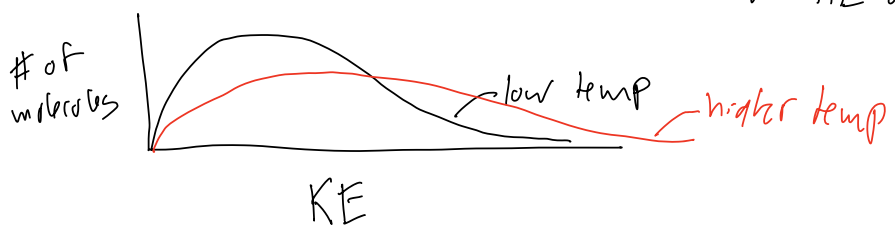
$\downarrow \times 0.21$

≈ 4700 mol O_2

Kinetic molecular theory of gases

- gas molecules are in constant motion
- kinetic energy of gas molecules is related to their ~~mass~~ velocity $KE = \frac{1}{2}mv^2$
- kinetic energy of gas molecules is also related to temperature

- At a given temp there will be a particular distribution of gas molecules in terms of KE or velocity



For 1 molecule we can talk about indiv KE, velocity

For 1 mole, we talk about KE_{avg} and v_{avg}

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

It can be shown that $KE_{avg} = \frac{3}{2}kT$ — Boltzmann constant 1.381×10^{-23} J/K

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

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

high temp → high velocity

high mass → low velocity

Pressure is really gas molecules crashing into walls of container

More KE in collisions → higher pressure

Helps to explain effusion/diffusion — mass and velocity

KE is related to temp $KE_{avg} = \frac{3}{2}kT$

At a given temp 2 diff gases will have same KE

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

gas 1 $\frac{1}{2}m_1 v_{1\,avg}^2 = \frac{1}{2}m_2 v_{2\,avg}^2$ gas

If $m_1 \neq m_2$ then the 2 gases will have diff v_{avg}

Ideal gas law is an OK approximation but it has 2 major flaws

- Assumes atoms/molecules are infinitely small but they are not
- Assumes no interactions

Especially at higher pressures these assumptions are not good

Non-ideal gas laws - specific to particular gases

Vander Waals equation

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

a, b are constants for individual gases

$\frac{an^2}{V^2}$: $\approx P$, account for interactions

bn: $\approx V$, account for non-zero size