If you take final

≤ avg → all scores stand

score > avg of midterms → lower of 2 exam scores is replaced by final exam score

Exam 1 85%
Exam 2 95%

If final > 90% (95%)

Gases – need to know about pressure $\text{Pressure} = \frac{\text{Force}}{\text{Area}}$

Units for pressure

Pascals (Pa)	$1\ \text{Pa} = 1\ \frac{N}{m^2}$
atmospheres	1 atm = 101325 Pa
bar	1 bar = 100000 Pa
mm Hg	1 atm = 760 mm Hg

We will describe gases with pressure, volume, temperature, moles
need to know relationships between them all

(1) T constant, # moles constant → what is relationship btw P and V

$P \propto \frac{1}{V}$ ("proportional") $P \times V = \text{constant}$ $P_i V_i = P_f V_f$

"Boyle's law"

(2) P constant, # moles constant → what is relationship between temp and volume?

$V \propto T$ $\frac{V}{T} = \text{constant}$ $\frac{V_i}{T_i} = \frac{V_f}{T_f}$

"Charles' Law"

(3) P constant, T constant → How is volume related to # of moles

$V \propto n$ (# of moles)

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

"Avogadro's Law"

$V \propto \frac{1}{P}$, $V \propto T$, $V \propto n$ $\quad V \propto \frac{Tn}{P}$

R = proportionality constant "gas constant"

$V = \frac{RTn}{P}$

↓ rearrange

$\boxed{PV = nRT}$ Ideal Gas Law

$R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$ * Always use Kelvin

Ex: If we have 75.0 g of N_2 gas in a container with V = 1.45L and the temp = 37°C, what is the pressure

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

n: $75.0\ g\ N_2 \times \frac{1\ mol\ N_2}{28.0\ g} = 0.893\ moles\ N_2$

T: $37°C + 273.15 \rightarrow 310.15\ K$

V = 1.45 L

$P = \frac{(0.893\ mol)(0.08206 \frac{L \cdot atm}{mol \cdot K})(310.15\ K)}{1.45\ L}$

P = 15.7 atm

Ex: 2.5 L container with 100 g $C_8H_{18}(\ell)$ - ignore volume

also in container is O_2 gas @ pressure of 4.0 atm

Temp initially 75°C

spark → combustion

(1) Immediately after combustion what is temp and pressure of gas in container

(2) ~~once~~ Once everything cools to 25°C what is pressure in container

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

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

$O_2 \rightarrow PV = nRT \rightarrow n = \frac{PV}{RT}$ $\quad n = \frac{(4.0\,atm)(2.5\,L)}{(0.08200\,\frac{L\,atm}{mol\,K})(25+273.15\,K)} = 0.41$ moles O_2 (LR)

$0.41\,mol\,O_2 \times \frac{1\,mol\,C_8H_{18}}{12.5\,moles\,O_2} = 0.0328$ mole C_8H_{18} combusts

$0.0328\,mole\,C_8H_{18} \times \frac{8\,mole\,CO_2}{1\,mol\,C_8H_{18}} = 0.262$ moles CO_2 ← only gas b/c all O_2 is gone

$\Delta_R H^\circ = \left[\left(8 \times -393\,\frac{kJ}{mol}\right) + \left(9 \times -285\,\frac{kJ}{mol}\right)\right] - \left[1 \times -250\,kJ/mol\right] =$

$-5709 - -250 = -5459\,\frac{kJ}{mol\,rxn} \times 0.0328\,moles = 179$ kJ given off for us

$0.262\,moles\,CO_2 \times \frac{44\,g}{mol} = 11.5\,g\,CO_2$

$q = +179000\,J = (11.5\,g)(1.1\,\frac{J}{g^\circ C})(T_f - 75^\circ C)$

$\frac{179000\,J}{(11.5\,g)(1.1\,\frac{J}{g^\circ C})} = T_f - 25^\circ C \qquad T_f = 14{,}175^\circ C$

$P_{of\ hot\ gas} = \frac{nRT}{V} = \frac{(0.262\,mol\,CO_2)(0.08206\,\frac{L\,atm}{mol\,K})(14448\,K)}{2.5\,L} = 124$ atm

When CO_2 cools to 25°C what is pressure

$$P_{\text{cool } CO_2} = \frac{(0.262 \text{ mol})\left(0.08206 \frac{\text{L atm}}{\text{mol K}}\right)(298.15 \text{ K})}{2.5 \text{ L}} = 2.56 \text{ atm}$$

$PV = nRT$ (const: V, n, R, T)

$P = \frac{nRT}{V} \rightarrow \frac{P}{T} = \text{const}$

Mixtures of gases → use "partial pressure"

↳ pressure you would have if only that gas was present

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

↳ "partial pressure of N_2"

X_{O_2} = "mole fraction of O_2"

$= \frac{\#\text{ moles } O_2}{\text{total moles of all things}}$

$P_g = P_{tot} X_g$

$P_{N_2} = P_{tot} X_{N_2}$

$P_{O_2} = P_{tot} X_{O_2}$

Room → $P_{tot} = 1.0$ atm

$X_{N_2} = 0.79$

$X_{O_2} = 0.21$

→ $P_{N_2} = 0.79 \times 1 \text{ atm} = 0.79 \text{ atm}$

$P_{O_2} = 0.21 \times 1 \text{ atm} = 0.21 \text{ atm}$

Room is $1700 \text{ cm} \times 1200 \text{ cm} \times 350 \text{ cm} = 5 \times 10^8 \text{ cm}^3$

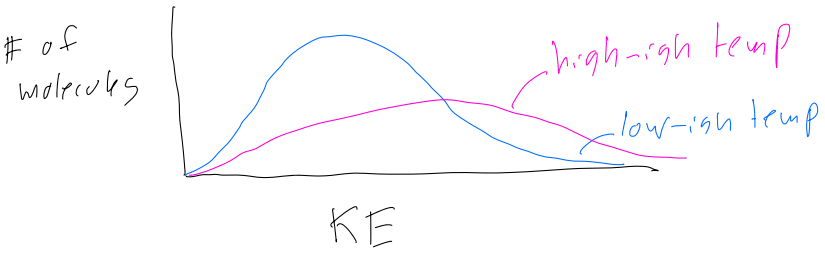
$= 5 \times 10^5 \text{ L}$

$$n_{O_2 \text{ in classroom}} = \frac{(0.21 \text{ atm})(5 \times 10^5 \text{ L})}{\left(0.08206 \frac{\text{L atm}}{\text{mol K}}\right)(294.8 \text{ K})} = 4300 \text{ moles } O_2$$

Kinetic molecular theory of gases

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

- At any given temperature there will be a distribution of molecules with varying amounts of KE



For a collection of molecules we want to describe KE_{avg} and v_{avg}

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

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

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

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

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

higher temp → higher velocity
higher mass → lower velocity

* consider pressure as the $\frac{\text{force}}{\text{area}}$ caused by gas molecules crashing into sides of container

higher KE for collisions → higher pressure

Diffusion - moving from higher conc → lower conc
Effusion - moving from higher to lower conc through small hole

KE is related to mass ($\frac{1}{2}mv^2$) but $KE_{avg} = \frac{3}{2}kT$ and has no mass dependence

consider 2 different gases at same temp
both have same KE $\frac{1}{2}m_1v_{1,avg}^2 = \frac{1}{2}m_2v_{2,avg}^2$

since $m_1 \neq m_2$ $\quad v_{1,avg} \neq v_{2,avg}$

Ideal gas law is a fine approximation but it has 2 major flaws

- Assumes atoms/molecules are infinitely small
- Assumes no interactions between atoms/molecules

Becomes especially problematic at higher pressures

use a non-ideal gas law to be more accurate

Various versions of these non-ideal gas laws all with specific parameters for each type of gas

Van der Waals equation

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

$\frac{an^2}{V^2}$: accounts for interactions

bn: accounts for size of molecule not being infinitely small

a, b are constants for individual gases