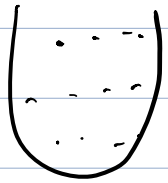
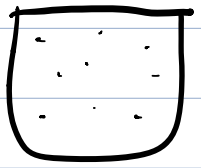
Gases



LIQUID

70% molecules themselves

30% free volume



GAS

0.01% molecules themselves

99.99% free volume

10x distance

for liquids & solids

intermolecular forces give the substance their identity/properties

for gas

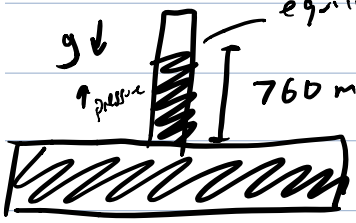
no intermolecular forces, distance too large. Same behaviour

Pressure: molecules hit the container

but what about atmospheric pressure? Due to gravity.

Highest atm. pressure at sea level

- Barometer



mercury is ideal b.c. of high density

IDEAL GAS LAW

Boyle's Law

Const. T

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

$V = \frac{k}{P}$ ~ proportionality constant

Avogadro's Law

Equal volumes of different gases at the same P, T, and vol. will contain the same number of particles.

$V \propto n$

$V = a \cdot n$

Charles's Law

const. P

$V \propto T$

$V = bT$

So $V = \frac{k \cdot b \cdot T \cdot a \cdot n}{P}$

$= \frac{R \cdot T \cdot n}{P}$

$PV = nRT$

Constant $0.8206 \frac{L \cdot atm}{k \cdot mol}$

or

$8.314 \quad L \cdot kPa$

$\frac{}{K \cdot mol}$

ex. $2\,NaN_{3(s)} \rightarrow 2\,Na_{(s)} + 3\,N_{2(g)}$

small volume to big volume (airbags)

What mass of sodium azide would be needed to inflate a 25.0 L airbag to a pressure of 1.3 atm at 25°C?

$PV = nRT$

$$n = \frac{PV}{RT} = \frac{1.3\ atm \cdot 25.0\ L}{0.0820\ \frac{L \cdot atm}{K \cdot mol} \cdot 298\ K} = 1.33\ mol\ N_2$$

$$1.33\ \cancel{mol\ N_2} \times \frac{2\ mol\ NaN_3}{3\ \cancel{mol\ N_2}} \times \frac{65.02\ g}{1\ mol\ NaN_3} = 58\ g\ NaN_3$$

ex. A sample of Ammonia (NH_3) gas with a volume of 3.5 L at a pressure of 1.68 atm is compressed to a volume of 1.35 L at a constant temp. Calculate the final pressure.

State 1 **State 2**

$P_1V_1 = nRT$ ← not changing, constants $\qquad P_2V_2 = nRT$

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2} = \frac{1.68\ atm \cdot 3.5\ \cancel{L}}{1.35\ \cancel{L}} = 4.4\ atm$$

V_2

ex. A sample of methane gas has a volume of 3.8 L at 5°C and is heated to 86°C at constant pressure. Calculate the new volume.

State 1

$PV_1 = nRT_1$

$\frac{P}{nR} = \frac{T_1}{V_1}$

State 2

$\frac{T_2}{V_2} = \frac{P}{nR}$

$\frac{T_1}{V_1} = \frac{T_2}{V_2}$

$V_2 = \frac{V_1 T_2}{T_1} = \frac{(3.8\ L)(359\ K)}{278\ K} = 4.9\ L$

ex. $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

$V = 2.80\ L$	$V = 35.0\ L$	$P = 2.50$ atm
$T = 25°C$	$T = 31°C$	$T = 125°C$
$P = 1.65$ atm	$P = 1.25$ atm	$V = ??$
n_{CH_4}	n_{O_2}	

CH_4: $n_{CH_4} = \frac{PV}{RT} = \frac{1.65\ \text{atm} \times 2.80\ L}{0.08206 \frac{L \cdot \text{atm}}{K \cdot \text{mol}} \times 298\ K}$

↑ Limiting reactant

$= 0.189\ \text{mol}\ CH_4 \times \frac{1\ \text{mol}\ CO_2}{1\ \text{mol}\ CH_4} = \boxed{0.189\ \text{mol}\ CO_2}$

Theoretical yield

$$n_{O_2} = \frac{PV}{RT} = \frac{1.25\ atm \times 35.0\ L}{0.08206\ \frac{L \cdot atm}{K \cdot mol} \cdot 304\ K}$$

$$= 1.75\ mol\ O_2 \times \frac{1\ mol\ CO_2}{2\ mol\ O_2} = 0.875\ mol\ CO_2$$

$$V_{CO_2} = \frac{n \cdot R \cdot T}{P} = \frac{0.189\ mol\ CO_2 \cdot 0.08206\ \frac{L \cdot atm}{K \cdot mol} \cdot 328\ K}{2.50\ atm}$$

$$= 2.47\ L$$

STP Standard temp & pressure

273K 1 atm

Molar Volume at STP

$$V = \frac{n \cdot R \cdot T}{P} = \frac{1\ mol \times 0.0826\ \frac{L \cdot atm}{K \cdot mol} \cdot 273\ K}{1\ atm}$$

$$= 22.4\ L$$

Density and Molar Mass

$$PV = n \cdot R \cdot T$$

$n = \frac{m}{M}$

$$P = \frac{n \cdot R \cdot T}{V} = \frac{m \cdot RT}{M \cdot V}$$

density

$$P = \frac{dRT}{M}$$

November 4, 2016

ex. The density of a gas was measured at 1.50 atm and 27°C and found to be 1.95 g/L. Calculate the molar mass of the gas.

$$M = \frac{dRT}{P} = \frac{1.95\ g/L \cdot 0.08206 \frac{L \cdot atm}{K \cdot mol} \times 300K}{1.5\ atm} = 32.0\ g/mol$$

Gas problem worksheet

#3 Bottle + gas : 136.201 — Bottle + H_2O : 385.42

Bottle : 135.821 — Bottle : 135.821

gas : 0.380 g — 249.599 ≈ 249.60 g

$$V = 249.60\ g \times \frac{1\ ml}{1\ g} = 249.60\ mL$$

$$M = \frac{dRT}{P} = \frac{0.380\ g / 0.24960\ L \times 0.08206 \frac{L \cdot atm}{K \cdot mol} \times 273K}{1\ atm} = 34.0\ g/mol$$

MIXTURES OF GASES

DALTON'S LAW: $P_{Total} = P_1 + P_2 + P_3 + \dots$

$P_1 = n_1 \cdot \frac{RT}{V}$

$P_2 = n_2 \cdot \frac{RT}{V}$

$\vdots$

$P_{Total} = (n_1 + n_2 + n_3)\frac{RT}{V} = n_{Total}\frac{RT}{V}$

ex.	O_2	He	DIVING TANK
	$V = 46.0\ L$	$V = 12.0\ L$	$V = 5.0\ L$
	$T = 25.0^\circ C$	$T = 25.0^\circ C$ →	$T = 25^\circ C$
	$P = 1.0\ atm$	$P = 1\ atm$	P_{O_2}, P_{He} & P_{Total}

$n_{O_2} = \frac{PV}{RT} = \frac{(1)(46)}{(0.08206)(298)} = 1.9\ mol\ O_2$

$n_{He} = \frac{PV}{RT} = \frac{(1)(12)}{(0.08206)(298)} = 0.49\ mol\ He$

$P_{O_2} = n_{O_2}\frac{RT}{V} = \frac{(1.9)(0.08206)(298)}{5} = 9.3\ atm$

$P_{He} = n_{He}\cdot\frac{RT}{V} = \frac{(0.49)(0.08206)(298)}{5} = 2.4\ atm$

$P_{Total} = P_{O_2} + P_{He} = 11.7\ atm$

* Liquids evaporate before reaching boiling point

5) $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

$P_{total} = P_{H_2} + P_{H_2O} = 0.980$ atm ; $P_{H_2} = P_{Tot} - P_{H_2O}$

$= 0.980 \text{ atm} - 29.8 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}$

$= 0.949$ atm

$n_{H_2} = \frac{PV}{RT}$

$n_{H_2} = \frac{(0.949)(7.80)}{(0.08206)(298)} = 0.303 \text{ mol } H_2$

$0.303 \text{ mol } H_2 \times \frac{1 \text{ mol Zn}}{1 \text{ mol } H_2} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 19.8 \text{ g Zn}$

LAB NOTES

Boiling pt: Temp at which pressure of vapored water reaches 760 mmHg

- extensive bubbling at boiling pt (bubble formed of H_2O gas from liquid)

vapour pressure reaches atmospheric pressure

Water sticking to container: air dissolved in water, solubility is inversely proportional to T

↓ solubility of air, air has to leave water, gets to see the bubble

(nothing to do with boiling pt)

MOLE FRACTION X

Mixture of components 1, 2 & 3

$$\chi_1 = \frac{n_1}{n_1 + n_2 + n_3} = \frac{n_1}{n_{Total}}$$

↑ "Quai"

$$\chi_2 = \frac{n_2}{n_{Total}}$$

$$\vdots$$

$$\chi_1 + \chi_2 + \chi_3 \cdots = 1$$

$$n_1 = P_1 \frac{V}{RT}$$

$$n_2 = P_2 \frac{V}{RT}$$

$$\vdots$$

For gases...

$$\chi_1 = \frac{P_1 \frac{V}{RT}}{P_1 \frac{V}{RT} + P_2 \frac{V}{RT}} = \frac{P_1}{P_{Total}}$$

ex. The partial pressure of oxygen in air was observed to be 156 torr when the atmospheric pressure was 743 torr. Calculate the mole fraction of O_2 present.

$$\chi_{O_2} = \frac{P_{O_2}}{P_{atm}} = \frac{156 \text{ torr}}{743 \text{ torr}} = 0.21$$

November 16, 2016

ex. Calculate the pressure of 10.0 mol of ethane, C_2H_6, in a 4.86-L flask at 300K using ideal gas law and van der waals eq.
$\left(a = 3.489 \frac{L^2 atm}{mol^2}\ ;\ b = 0.0638 \frac{L}{mol}\right)$

Ideal Gas Law: $PV = nRT$

$$P = \frac{n \cdot R \cdot T}{V} = \frac{(10)(0.08206)(300)}{4.86} = 50.7 \text{ atm}$$

V.D.W: $\left[P_{obs} + a\left(\frac{n}{V}\right)^2\right](V - nb) = nRT$

$$P_{obs} = \frac{nRT}{V - nb} - a\left(\frac{n}{V}\right)^2$$

$$= \frac{(10)(0.08206)(300)}{4.86 - 10(0.0638)} - 5.489\left(\frac{10}{4.86}\right)^2$$

$$= 35.1 \text{ atm}$$