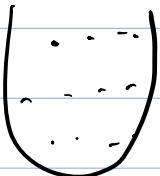
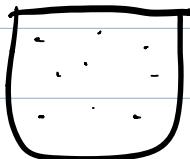


Gases



L I Q U I D



G A S

70% molecules themselves

0.01% molecules themselves

30% free volume

99.99% free volume

10x distance

for liquids & solids

intermolecular forces give the substances their identity/properties

for gns

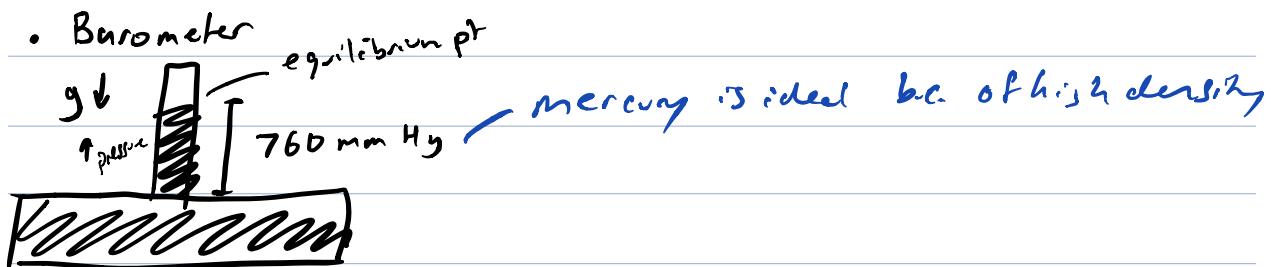
no intermolecular forces, distance too large. Same behaviour

Pressure: molecules hit the container

but what about atmospheric pressure? Due to gravity.

Highest atm. pressure at sea level

• Barometer



IDEAL GAS LAW

Boyle's Law

Const. T

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

$$V = \frac{k}{P} \quad \begin{matrix} \text{proportional} \\ \text{constant} \end{matrix}$$

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

$$V \propto T$$

$$V = b T$$

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

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

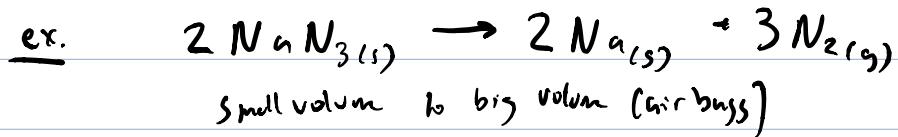
$$PV = nRT$$

Constant $0.8206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}$

or

$$8.314 \text{ L} \cdot \text{KPa}$$

k · mol



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?

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{1.3 \text{ atm} \cdot 25.0 \text{ L}}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \cdot 298 \text{ K}} = 1.33 \text{ mol N}_2$$

$$1.33 \text{ mol N}_2 \times \frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \times \frac{65.02 \text{ g}}{1 \text{ mol NaN}_3} = 58 \text{ g NaN}_3$$

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.

<u>State 1</u>	<u>State 2</u>
$P_1 V_1 = n(RT)$ <small>not changing, constants</small>	$P_2 V_2 = nRT$

$$P_1 V_1 = P_2 V_2$$

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{1.68 \text{ atm} \cdot 3.5 \text{ L}}{1.35 \text{ L}} = 4.4 \text{ atm}$$

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

$$\textcircled{P} V_1 = \textcircled{n} \textcircled{R} T_1$$

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

State 2

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

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

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(3.8 \text{ L})(353 \text{ K})}{278 \text{ K}} = 4.9 \text{ L}$$



$$V = 2.80 \text{ L} \quad V = 35.0 \text{ L} \quad P = 2.50 \text{ atm}$$

$$T = 25^\circ\text{C} \quad T = 31^\circ\text{C} \quad T = 125^\circ\text{C}$$

$$P = 1.65 \text{ atm}$$

$$P = 1.25 \text{ atm}$$

$$V = ??$$

$$n_{\text{CH}_4}$$

$$n_{\text{O}_2}$$

$$\text{CH}_4 : n_{\text{CH}_4} = \frac{PV}{RT} = \frac{1.65 \text{ atm} \times 2.80 \text{ L}}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} 298 \text{ K}}$$

Limiting reagent

$$= 0.189 \text{ mol CH}_4 \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CH}_4} = 0.189 \text{ mol CO}_2$$

Theoretical yield

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

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

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

$$= 2.47 \text{ L}$$

STP Standard temp & pressure
 ↓
 273K 1 atm

Molar Volume at STP

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

$$= 22.4 \text{ L}$$

Density and Molar Mass

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

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

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

density

$$P = \frac{d RT}{M}$$

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.

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

Gas problem worksheet

#3 Bottle + gas : 136.201

$$\begin{array}{rcl} \text{Bottle} & : & 135.821 \\ \text{gas} : & 0.380 \text{ g} \end{array}$$

$$\text{Bottle} + \text{H}_2\text{O} : 385.42$$

$$\begin{array}{rcl} \text{Bottle} & : & 135.821 \\ 249.59 \text{ g} & \approx & 249.60 \text{ g} \end{array}$$

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

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

MIXTURES OF GASES

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

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

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

:

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

ex. O₂ H_e DIVING TANK

$$V = 46.0 \text{ L}$$

$$V = 12.0 \text{ L}$$

$$V = 5.0 \text{ L}$$

$$T = 25.0^\circ\text{C}$$

$$T = 25.0^\circ\text{C}$$

$$T = 25^\circ\text{C}$$

$$P = 1.0 \text{ atm}$$

$$P = 1 \text{ atm}$$

$$P_{O_2}, P_{H_e} \text{ & } P_{\text{Total}}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{(1)(46)}{(0.08206)(298)} = 1.9 \text{ mol O}_2$$

$$n_{H_e} = \frac{PV}{RT} = \frac{(1)(12)}{(0.08206)(298)} = 0.49 \text{ mol H}_e$$

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

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

$$P_{\text{Total}} = P_{O_2} + P_{H_e} = 11.7 \text{ atm}$$

* Liquids evaporate before reaching boiling point



$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} = 0.980 \text{ atm}; P_{\text{H}_2} = P_{\text{Total}} - P_{\text{H}_2\text{O}}$$

$$= 0.980 \text{ atm} - 23.8 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}$$

$$= 0.949 \text{ atm}$$

$$n_{\text{H}_2} = \frac{PV}{RT}$$

$$n_{\text{H}_2} = \frac{(0.949)(7.80)}{(0.08206)(298)} = 0.303 \text{ mol H}_2$$

$$0.303 \text{ mol H}_2 \times \frac{1 \text{ mole Zn}}{1 \text{ mole H}_2} \times \frac{65.38 \text{ g}^{2+}}{1 \text{ mole Zn}} = 19.8 \text{ g Zn}$$

LAB NOTES

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

• extensive bubbling at boiling pt (bubble formed of H₂O gas from liquid)

vapour pressure matches atmospheric pressure

solvability is

water sticking to container; air dissolved in water, inversely proportional to T

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

(nothing to do with boiling pt)

MOLE FRACTION X

Mixture of components 1, 2 & 3

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

"Quai"

$$\chi_2 = \frac{n_2}{n_{\text{total}}}$$

:

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

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

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

:

For gases...

$$\chi_1 = \frac{P_1 / k_T}{P_1 / k_T + P_2 / k_T} = \frac{P_1}{P_{\text{total}}}$$

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 O₂ present.

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

November 16, 2016

ex Calculate the pressure of 10.0 mol of ethene, C_2H_4 , in a 4.86-L flask at 300K using ideal gas law and van der waals eq.

$$(a = 5.489 \frac{L^2 \text{ atm}}{\text{mol}^2}; b = 0.0638 \frac{L}{\text{mol}})$$

Ideal Gas Law: $PV = nRT$

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

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

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

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

$$\approx 35.1 \text{ atm}$$