GASES

Pressure

Pressure Units (SI):

$$P = \frac{\text{Force}}{\text{Area}} = \frac{\text{N}}{\text{m}^2} = \text{Pa}$$

Pa = pascal $1 \text{ Pa} = 1 \text{ N/m}^2$

N = newton $1 \text{ N} = 1 \text{ kg.m/s}^2$

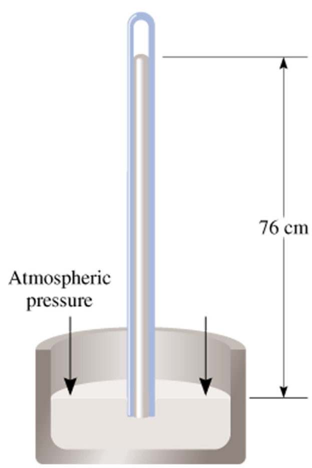
atmospheric pressure $\simeq 1 \times 10^5 \text{ N/m}^2$
$= 1 \times 10^5 \text{ Pa}$
$= 1 \times 10^2 \text{ kPa}$

Standard Atmospheric Pressure:

The pressure exerted by a column of Hg exactly 760 mm high at 0°C at sea level.

1 atm = 760 mmHg
= 760 torr
= 101.325 kPa
= 14.70 lb/in^2 (psi)

1 mmHg = 1 torr

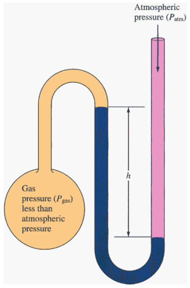


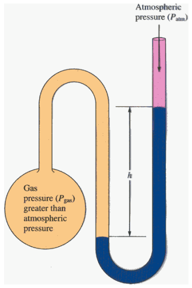
If the barometer reading is 688 mmHg, what is this pressure in units of atm, kPa, and torr?

$$688 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.905 \text{ atm}$$

How to measure the pressure of a gas?

Open-end manometer





$P_{gas} + h = P_{atm}$

$P_{gas} = P_{atm} - h$

$P_{gas} = P_{atm} + h$

Gas Laws

Four **variables define the state of a gas**

- T = temperature (K)
- P = pressure (atm)
- V = volume (L)
- n = amount of gas (moles)

Related to each other by the gas laws

Boyle's Law

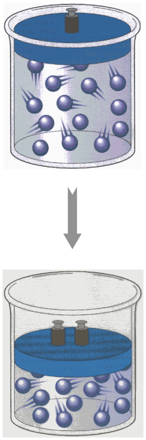
P-V **relationship:**

$P \alpha \frac{1}{V}$ @ constant n & T

$P = \text{constant} \times \frac{1}{V}$

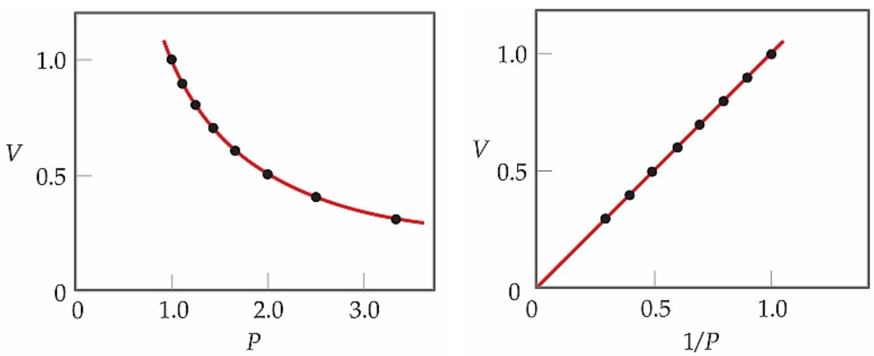
$PV = \text{constant}$

$P_1V_1 = P_2V_2$



Boyle's Law

$P = \text{constant} \times \frac{1}{V}$



Boyle's Law

A gas sample occupies a volume 946 mL at a pressure of 726 mmHg. What is the pressure if the volume is reduced at constant temperature to 154 mL ?

$$P_1 V_1 = P_2 V_2$$

V_1 = 946 mL P_1 = 726 mmHg

V_2 = 154 mL P_2 = ? mmHg

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{726 \text{ mmHg} \times 946 \text{ mL}}{154 \text{ mL}}$$

= 4460 mmHg

Charle's Law

***V-T* Relationship:**

V α ***T*** **@ constant *P* & *n***

***V* = constant x T**

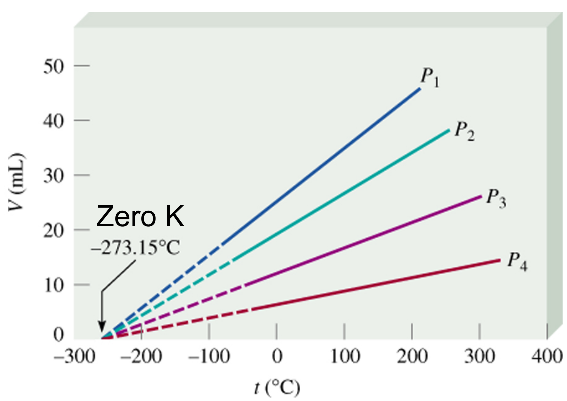
$$\frac{V}{T} = \text{constant}$$

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

***T* must be in K**

Charle's Law

V = constant x T



Charle's Law

A gas sample occupies 3.20 L at 125°C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

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

***T* must be in K**

$$T_1 = 125 + 273 = 398 \text{ K}$$

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{1.54 \text{ L} \times 398 \text{ K}}{3.20 \text{ L}} = 192 \text{ K}$$

Avogadro'e Law

V-n **Relationship:**

$V \alpha n$ @ constant T & P

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

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Avogadro'e Law

At constant T & P:

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

Molar volume

For any ideal gas:

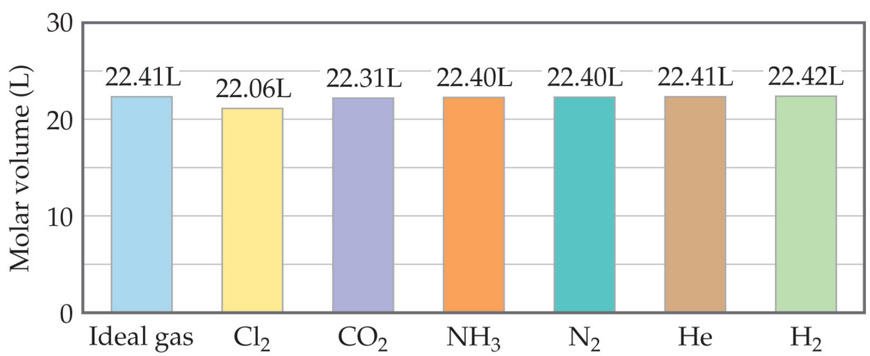
Molar volume = 22.41 L at STP

STP = Standard T & P Memorize

Standard T = 0°C = 273 K

Standard P = 1 atm

Molar Volumes



A mole of any gas occupies a volume of approximately 22.4 L at STP

Molar Volumes

$3\ H_2\ (g)$ +	$N_2\ (g)$	⟶	$2\ NH_3\ (g)$
3 mol	1 mol		2 mol
3 x 22.4 L	1 x 22.4 L		2 x 22.4 L
3 volumes	1 volume		2 volumes
3 L	1 L		2 L

Remember constant T & P

How many liters of O_2 are required to completely burn 2.0 L of ammonia at the same *T* & *P*?

$$4NH_3\ (g) + 5O_2\ (g) \rightarrow 4NO\ (g) + 6H_2O\ (g)$$

$$2\ \text{L NH}_3 \times \frac{5\ \text{L O}_2}{4\ \text{L NH}_3} = 2.5\ \text{L O}_2$$

Ideal Gas Law

Boyle's Law	Charle's Law	Avogadro's Law
$V \alpha \frac{1}{P}$	$V \alpha T$	$V \alpha n$

$$V = \text{constant} \times \frac{1}{P} \times T \times n$$

$$V = \frac{RTn}{P}$$ R = Gas constant

$PV = nRT$ Ideal Gas Law

Ideal Gas Law

$PV = nRT$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} = R$$

gas laws are special cases of this formula

Boyle's Law	Charles' Law	Avogadro's Law
$\frac{P_1V_1}{\cancel{nT}} = \frac{P_2V_2}{\cancel{nT}}$	$\frac{\cancel{P}V_1}{\cancel{n}T_1} = \frac{\cancel{P}V_2}{\cancel{n}T_2}$	$\frac{\cancel{P}V_1}{n_1\cancel{T}} = \frac{\cancel{P}V_2}{n_2\cancel{T}}$

The Gas Constant *R*:

$$R = \frac{PV}{nT}$$ take 1 mole of a gas at STP

$$R = \frac{(1\ \text{atm})(22.41\ \text{L})}{(1\ \text{mol})(273\ \text{K})} = 0.08206\ \frac{\text{L . atm}}{\text{mol . K}}$$

$$R = 8.314\ \frac{\text{Pa.m}^3}{\text{mol.K}} = 8.314\ \frac{\text{J}}{\text{mol . K}}$$

What is the volume (in liters) occupied by 49.8 g HCl at STP?

$PV = nRT$

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

$P = 1$ atm

$T = 0°C = 273$ K

$$n_{HCl} = 49.8 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.45 \text{ g HCl}} = 1.37 \text{ mol}$$

$$V = \frac{1.37 \text{ mol} \times 0.0821 \frac{\text{L.atm}}{\text{mol.K}} \times 273 \text{ K}}{1 \text{ atm}} = 30.7 \text{ L}$$

At 46 °C and 669 mm Hg pressure, a gas occupies a volume of 0.600 L. How many liters it will occupy at 0.0 °C and 0.205 atm?

$$\frac{P_1V_1}{nT_1} = \frac{P_2V_2}{nT_2} \rightarrow \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$V_2 = \frac{669 \text{ mmHg} \times 0.600 \text{ L} \times 273 \text{ K}}{0.205 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} \times 319 \text{ K}} = 2.20 \text{ L}$$

Gas Densities & Molar Mass

$PV = nRT$

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

m = mass of the gas in g

$\mathcal{M}$ = molar mass of the gas

$PV = \frac{mRT}{\mathcal{M}}$

$\mathcal{M} = \frac{mRT}{VP} = \frac{dRT}{P}$

d = density of the gas in g/L

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

Gas Densities & Molar Mass

A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0 °C. What is the molar mass of the gas?

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

$$\mathcal{M} = \frac{4.65 \text{ g} \times 0.0821 \frac{\text{L.atm}}{\text{mol.K}} \times 300 \text{ K}}{2.10 \text{ L} \times 1 \text{ atm}}$$

$\mathcal{M} = 54.6$ g/mol

What is the volume of CO_2 produced at 37.0°C and 1.00 atm when 5.60 g of glucose are used up in the reaction:

$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

$g\ C_6H_{12}O_6 \rightarrow mol\ C_6H_{12}O_6 \rightarrow mol\ CO_2 \rightarrow V\ CO_2$

$$5.60\ g\ C_6H_{12}O_6 \times \frac{1\ mol\ C_6H_{12}O_6}{180\ g\ C_6H_{12}O_6} \times \frac{6\ mol\ CO_2}{1\ mol\ C_6H_{12}O_6}$$

$= 0.187\ mol\ CO_2$

$g\ C_6H_{12}O_6 \rightarrow mol\ C_6H_{12}O_6 \rightarrow mol\ CO_2 \rightarrow V\ CO_2$

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

$$V = \frac{0.187\ mol \times 0.0821\ \frac{L \cdot atm}{mol \cdot K} \times 310.\ K}{1\ atm}$$

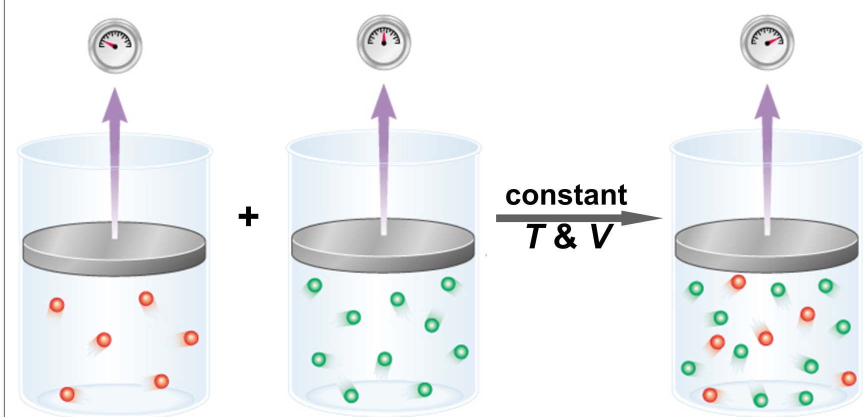
$= 4.76\ L$

Gas Mixtures & Partial Pressures

Dalton's Law of Partial Pressures:

The total pressure of a mixture of gases in a container is equal to the sum of the partial pressures of the individual gases in the mixture. (constant *V* & *T*)

Gas Mixtures & Partial Pressures



P_1 P_2 $P_t = P_1 + P_2$

Dalton's Law of Partial Pressures

$P_t = P_1 + P_2 + P_3 + ..$ constant V, T

P_t = total pressure in container

P_i = partial pressure of gas i.

Partial pressure: the pressure exerted by the gas if it was alone in the container

$$P = n\frac{RT}{V}$$

At constant V, T:

$$P_1 = n_1\frac{RT}{V} \qquad P_2 = n_2\frac{RT}{V}$$

$$P_1 \propto n_1 \qquad P_2 \propto n_2$$

In gas phase reactions at constant V & T, we can use partial pressures as we use number of moles.

At constant V & T:

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

$$\frac{P_1}{P_t} = \frac{n_1}{n_t} = X_1$$ mole fraction of gas 1

$$P_1 = X_1 P_t$$

$$X_1 + X_2 + ... = 1$$

A sample of natural gas contains 8.24 mol of CH_4, 0.421 mol of C_2H_6, and 0.116 mol of C_3H_8. If the total pressure of the gases is 1.37 atm, what is the partial pressure of propane C_3H_8?

$$P_{prop} = X_{prop} P_t$$

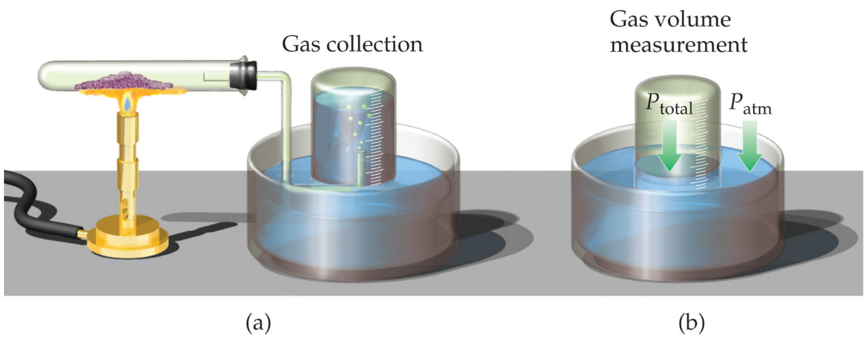
$$X_{prop} = \frac{n_{C3H8}}{n_{C3H8} + n_{CH4} + n_{C2H6}}$$

$$= \frac{0.116}{0.116 + 8.24 + 0.421} = 0.0132$$

$$P_{prop} = 0.0132 \times 1.37 \text{ atm} = 0.0181 \text{ atm}$$

Collection of Gases Over Water

$$2KClO_3\,(s) \rightarrow 2KCl\,(s) + 3O_2\,(g)$$



$$P_{atm} = P_t = P_{O_2} + P_{H_2O}$$

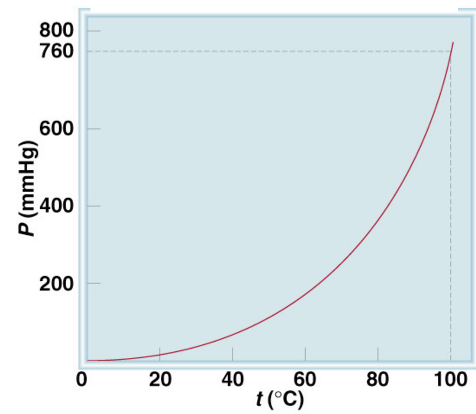


TABLE 5.3 Pressure of Water Vapor at Various Temperatures

Temperature (°C)	Water Vapor Pressure (mmHg)
0	4.58
5	6.54
10	9.21
15	12.79
20	17.54
25	23.76
30	31.82
35	42.18
40	55.32
45	71.88
50	92.51
55	118.04
60	149.38
65	187.54
70	233.7
75	289.1
80	355.1
85	433.6
90	525.76
95	633.90
100	760.00

$$2KClO_3\,(s) \rightarrow 2KCl\,(s) + 3O_2\,(g)$$

A 0.250 L of O_2 were collected over water at 26 °C and a total pressure of 765 torr. How many moles of O_2 will be produced?

P (H_2O) at 26 °C = 25 torr.

$$P_t = P_{O2} + P_{H2O}$$

$$P_{O2} = P_t - P_{H2O} = 765 - 25 = 740 \text{ torr}$$

$$n_{O_2} = \frac{P_{O_2}V_{O_2}}{RT} = \frac{740 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times 0.25 \text{ L}}{0.0821 \frac{\text{L . atm}}{\text{mol . K}} \times 299 \text{ K}}$$

$$= 9.9 \times 10^{-3} \text{ mol}$$

Molecular Effusion & Diffusion

Molecular Speeds

$$PV = nRT \qquad PV = \frac{n\mathcal{M}\overline{u^2}}{3}$$

$$\overline{u} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

u = Average speed of molecule

$\mathcal{M}$ = molar mass of the gas

$$\overline{u} \propto \sqrt{\frac{1}{\mathcal{M}}}$$

Under the same conditions, lighter gases travel faster

Calculate the average speed (m/s) of NH_3 molecules at 25 °C.

$$\bar{u} = \sqrt{\frac{3RT}{M}} \qquad 1\ J = 1\ kg\ .\ m^2\ .\ s^{-2}$$

$$\bar{u} = \sqrt{\frac{3\ (8.314\ J.K^{-1}.mol^{-1})\ (298\ K)}{17.01 \times 10^{-3}\ kg.mol^{-1}}}$$

$$= 661\ \sqrt{J\ .\ kg^{-1}} \quad = 661\ \sqrt{(kg.m^2.s^{-2})\ kg^{-1}}$$

661 m/s

Graham's Law of Diffusion & Effusion

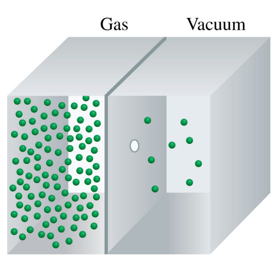
Diffusion

The mixing of different gases by random molecular motion with frequent collisions.

Graham's Law of Diffusion & Effusion

Effusion

Escape of a gas under pressure from one compartment of a container to another through a tiny hole.



Graham's Law

Under the same conditions of *T* & *P*, rates of diffusion of gases are inversely proportional to the square roots of their molar masses.

$$r \propto \sqrt{\frac{1}{M}}$$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \qquad \frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}}$$

It takes 192 s for 1.4 L of an unknown gas to effuse through a porous wall and 84 s for the same volume of N_2 to effuse at the same *T* and *P*. What is the molar mass of the unknown gas?

$$\frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}}$$

$$\frac{192}{84} = \sqrt{\frac{\mathcal{M}_{unk}}{28}}$$

$\mathcal{M}_{unk}$ **= 146 g/mol**

If methane (CH_4) effuses 3.3 times faster than $Ni(CO)_x$. What is the value of *x*

$$\frac{r_{CH_4}}{r_{Ni(CO)_x}} = 3.3 = \sqrt{\frac{M_{Ni(CO)_x}}{16}}$$

$$M_{Ni(CO)_x} = 16 \times (3.3)^2 = 174.2 \text{ g/mol)}$$

$$M_{Ni(CO)_x} = 58.7 + (28x) = 174.2$$

$$x = 4.1 \approx 4$$