Outline

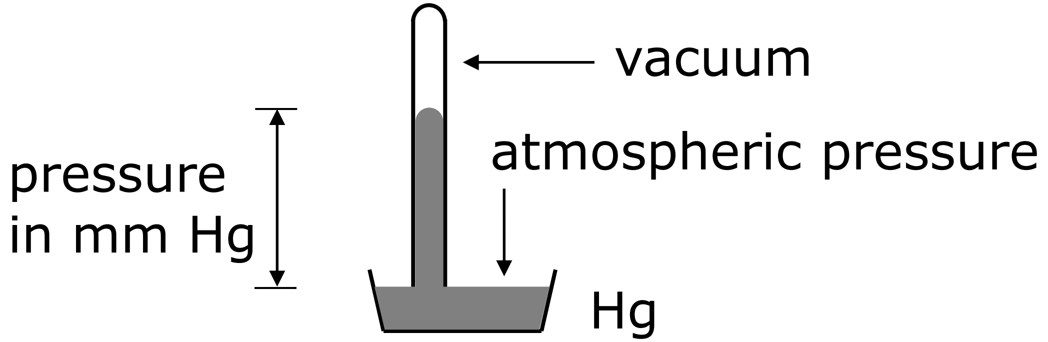
- Properties of Gases
- Gas Laws
- Gas Mixtures
- Kinetic Molecular Theory
- Real Gases

Properties of Gases

Gases have indefinite shape and volume, can be highly compressed

Gas pressure results from collisions of particles with walls of a container

Atmospheric pressure is measured with a <u>barometer</u>



Pressure of atmosphere can support a 760 mm column of Hg

Standard atmospheric pressure is 760 mm Hg…

1 atm = 760 mm Hg = 760 torr = 101.325 kPa

Other units for gases…

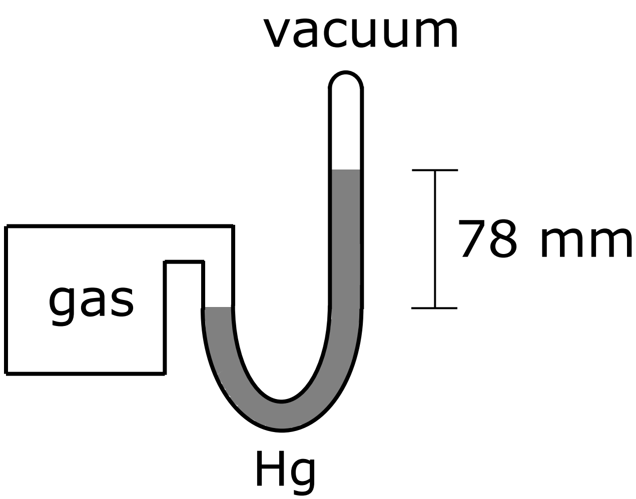
n moles (mol)

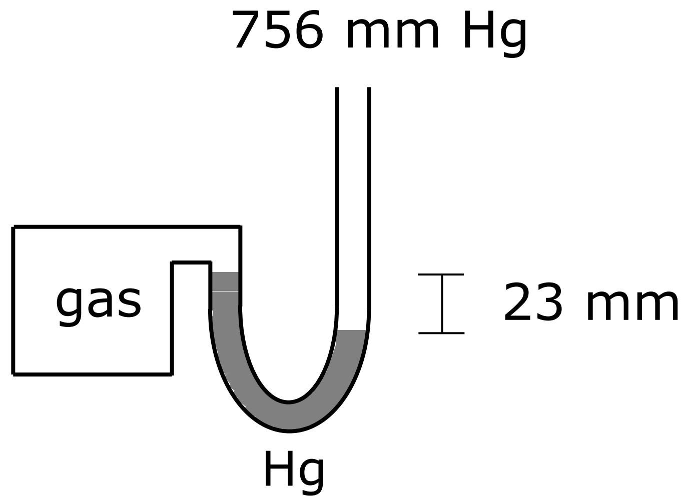
V volume (L)

T temperature (K)

Standard temperature is 273.15 K (0 ^{o}C)

Gas pressure is measured with a manometer





P_{gas} = 78 mm Hg

P_{gas} = 0.10 atm

P_{gas} = 733 mm Hg

P_{gas} = 0.964 atm

Gas Laws

Avogadro's Law

For constant T and P, V is proportional to n

$$V \propto n \quad \Rightarrow \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Boyle's Law

For constant n and T, V is inversely proportional to P

$$V \propto \frac{1}{P} \quad \Rightarrow \quad P_1V_1 = P_2V_2$$

A sample of gas has V = 268 mL at 2.1 atm. What's V at 1.7 atm?

$$P_1V_1 = P_2V_2 \quad \Rightarrow \quad V_2 = \frac{P_1V_1}{P_2} = \frac{(2.1\ \text{atm})(268\ \text{mL})}{(1.7\ \text{atm})} = \underline{\underline{330\ \text{mL}}}$$

Charles' Law

For constant n and P, volume is proportional to T

$$V \propto T \quad \Rightarrow \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Gay-Lussac's Law

For constant n and V, pressure is proportional to T

$$P \propto T \quad \Rightarrow \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A sample of gas has V = 3.68 L at 201 K. What's T if V = 9.67 L?

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \Rightarrow \quad T_2 = \frac{V_2 T_1}{V_1} = \frac{(9.67\,\mathrm{L})(201\,\mathrm{K})}{(3.68\,\mathrm{L})} = \underline{528\,\mathrm{K}}$$

"Char-Boyled" Law

Combination of Charles' Law and Boyle's Law

$$V \propto \frac{T}{P} \quad \Rightarrow \quad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

<u>Ideal Gas Law</u> is the combination of all gas laws…

$$V \propto \frac{n\,T}{P} \quad \Rightarrow \quad PV = nRT$$

R is the universal gas constant…

$R = 0.08206\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}} = 8.314\ \mathrm{J\ mol^{-1}\ K^{-1}}$

85.0 mL of a gas is collected at 17 °C and 700. torr. What will its volume be at STP?

$$85.0\ \text{mL} \Rightarrow 0.0850\ \text{L}$$

$$17\ ^{o}\text{C} \Rightarrow 29\underline{0}.15\ \text{K}$$

$$700.\ \text{torr} \Rightarrow 0.92\underline{1}05\ \text{atm}$$

$$\text{STP} \Rightarrow T = 273.15\ \text{K} \quad \text{and} \quad P = 1\ \text{atm}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \Rightarrow \frac{(0.92105\ \text{atm})(0.0850\ \text{L})}{29\underline{0}.15\ \text{K}} = \frac{(1.00\ \text{atm})(V_2)}{273.15\ \text{K}}$$

$$V_2 = 0.0737\ \text{L} = \underline{73.7\ \text{mL}}$$

How many moles of a gas are in a flask with a capacity of 500. mL, when the pressure is 570. torr at 22 ^oC?

$$500.\ \text{mL} \Rightarrow 0.500\ \text{L}$$

$$22\ ^oC \Rightarrow 29\underline{5}.15\ \text{K}$$

$$570.\ \text{torr} \Rightarrow 0.750\ \text{atm}$$

$$PV = nRT$$

$$(0.750\ \text{atm})(0.500\ \text{L}) = n(0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1})(29\underline{5}.15\ \text{K})$$

$$n = \underline{0.0155\ \text{mol}}$$

$$PV = nRT \quad \Rightarrow \quad PV = \left(\frac{m}{MM}\right)RT \quad \Rightarrow \quad \frac{m}{V} = \frac{P\,(MM)}{RT}$$

What is the pressure of oxygen gas if 0.200 g is confined in a 1.00 L flask at 25 ^{o}C?

$$P(1.00\ \mathrm{L}) = \left(\frac{0.200\ \mathrm{g}}{32.00\ \mathrm{g/mol}}\right)(0.08206\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}})(29\underline{8}.15\ \mathrm{K})$$

$$\Rightarrow \quad P = \underline{0.153\ \mathrm{atm}}$$

What is the density of nitrogen gas at 20. ^{o}C and 1.00 atm pressure?

$$\frac{m}{V} = \frac{P\,(MM)}{RT} = \frac{(1.00\ \mathrm{atm})(28.02\ \mathrm{g\ mol^{-1}})}{(0.08206\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}})(29\underline{3}.15\ \mathrm{K})} = \underline{1.16\ \mathrm{g\ L^{-1}}}$$

The volume of 1 mole of a gas at STP is 22.41 L

…22.41 L mol^{-1} is the <u>Standard Molar Volume</u>

…is used to convert between liters and moles (at STP)!

How many liters of oxygen are produced at STP when 2.5 g of potassium chlorate decompose?

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

$$2.5\ g\ KClO_3 \times \frac{1\ mol}{122.55\ g} \times \frac{3\ mol\ O_2}{2\ mol\ KClO_3} \times \frac{22.41\ L}{1\ mol} = \underline{0.69\ L}$$

Gas Mixtures

Partial Pressure…

The pressure exerted by a single gas in a mixture of gases

The total pressure of a gas mixture equals the sum of the partial pressures

$$P_{total} = P_a + P_b + P_c + \ldots \qquad \text{(Dalton's Law)}$$

Partial pressures are proportional to number of moles of gas in a mixture

$$p = \chi P_{total}$$

Mole Fraction (χ) is the moles of one gas in a mixture divided by the total moles of gas

At 27 ^{o}C a 2.00 L flask has 4.40 g of CO_2 and 1.00 g N_2. Find the partial pressures and total pressure.

$$4.40\ \text{g}\ CO_2 \quad \Rightarrow \quad 0.100\ \text{mol}\ CO_2 \quad \text{and} \quad 1.00\ \text{g}\ N_2 \quad \Rightarrow \quad 0.0357\ \text{mol}\ N_2$$

$$\chi_{CO_2} = \frac{0.100\ \text{mol}}{0.136\ \text{mol}} = 0.735 \quad \text{and} \quad \chi_{N_2} = \frac{0.0357\ \text{mol}}{0.136\ \text{mol}} = 0.263$$

$$P_{total} = \frac{(0.136\ \text{mol})(0.08206\ \text{L atm mol}^{-1}\ \text{K}^{-1})(30\underline{0}.15\ \text{K})}{2.00\ \text{L}} = \underline{1.67\ \text{atm}}$$

$$p_{CO_2} = (0.735)(1.67\ \text{atm}) = \underline{1.23\ \text{atm}}$$

$$p_{N_2} = (0.263)(1.67\ \text{atm}) = \underline{0.439\ \text{atm}}$$

Kinetic Molecular Theory

Simple model to explain properties of an ideal gas

Postulates...

1. Volume of particles is negligible

2. Particles are in constant motion

3. Average kinetic energy of particles is proportional to temperature

4. Particles do not exert forces on one another

Pressure results from collisions between molecules and walls of container…

Pressure increases with increasing collision frequency!

1. Boyle's Law ($V \propto P^{-1}$)

   dec. volume $\rightarrow$ inc. collisions $\rightarrow$ inc. pressure

2. Gay-Lussac's Law ($P \propto T$)

   inc. temp $\rightarrow$ inc. velocity/collisions $\rightarrow$ inc. pressure

3. Charles' Law ($V \propto T$)

   inc. temp $\rightarrow$ inc. velocity/collisions $\rightarrow$ inc. volume

Kinetic Energy (E_K) is the energy of motion...

$$E_K = \frac{1}{2}mu^2 \qquad (m = \text{mass and } u = \text{velocity})$$

Mathematically, for 1 mol of gas...

$$E_K = N_A\left(\frac{1}{2}m\overline{u^2}\right) = \frac{3}{2}RT$$

Temperature is a measure of the average kinetic energy of molecules

Molecules at the same T have the same average E_K!

$$\overline{u^2} = \frac{3RT}{M}$$

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$$

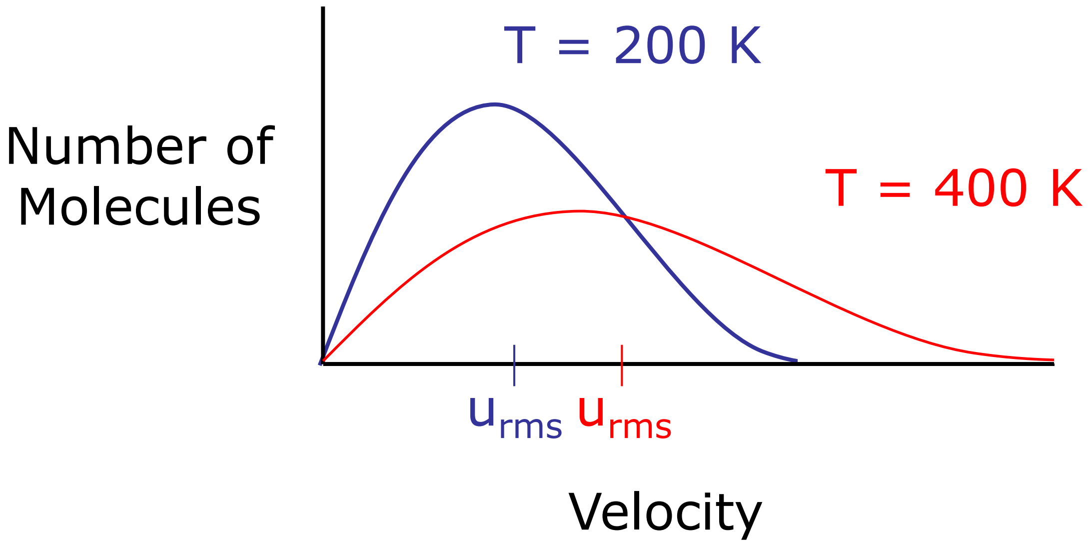
Root-Mean-Square-Velocity (u_{rms}) is the average velocity of the molecules

u_{rms} and M are inversely related

For same T, the bigger the molecules, the smaller their velocities

R is 8.314 J / mol K and M is in kg / mol

The u_{rms} is just the “average velocity” of a group of molecules at a specific temperature…



These curves represent the velocities of molecules at a given T, and are called Maxwell-Boltzmann Distributions

Find the u_{rms} of oxygen molecules at 25 ^oC.

$$u_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3\left(8.314 \text{ J mol}^{-1}\text{ K}^{-1}\right)(29\underline{8}.15 \text{ K})}{0.03200 \text{ kg mol}^{-1}}} = \underline{482 \text{ m s}^{-1}}$$

Determine the molar mass of molecules with a root-mean-square-velocity of 312 m/s at 25 ^oC.

$$u_{rms} = \sqrt{\frac{3RT}{M}} \quad \Rightarrow \quad M = \frac{3\,RT}{u_{rms}^{\;2}}$$

$$M = \frac{3\left(8.314 \text{ J mol}^{-1}\text{ K}^{-1}\right)(29\underline{8}.15 \text{ K})}{\left(312 \text{ m s}^{-1}\right)^2} = 0.0764 \text{ kg/mol} = \underline{76.4 \text{ g/mol}}$$

Effusion is the movement of gaseous molecules into a chamber through a pinhole…

Graham’s Law of Effusion…

rates of effusion are inversely proportional to square root of molar mass

$$\frac{\text{rate}_1}{\text{rate}_2} = \sqrt{\frac{M_2}{M_1}}$$

Diffusion is the migration of molecules as a result of random molecular motion…

Real Gases

Gases that conform to Kinetic Molecular Theory are ideal gases

Real gases deviate from ideal behavior because...

1. molecules attract each other (reducing pressure)
2. molecules take up space (reducing volume)

Deviation from ideal behavior is reduced for…

1. low pressures…
   attractive forces decrease with increasing distance
2. high temperatures…
   attractive forces more easily overcome with higher speeds

The van der Waals equation corrects for these deviations!

van der Waals Equation:

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

“a” is a correction factor for molecular attraction

“b” is a correction for molecular size

Find the pressure of 1.000 moles of ammonia at 0.00 ^{o}C and 22.41 L. For ammonia, a = 4.170 L^2 atm / mol^2 and b = 0.03707 L / mol. What is the percent deviation from ideality?

$$\left[P_{obs} + a\left(\frac{n}{V}\right)^2\right] \times (V - nb) = nRT \qquad \Rightarrow \qquad P_{obs} = \frac{nRT}{(V - nb)} - a\left(\frac{n}{V}\right)^2$$

$$P_{obs} = \left[\frac{(1.000)(0.08206)(273.15\,K)}{(22.41 - 1.000 \times 0.03707)} - 4.170\left(\frac{1.000}{22.41}\right)^2\right] atm$$

$$P_{obs} = \underline{0.994\ atm}$$

$$\%\ deviation = \frac{|0.994\ atm - 1.000\ atm|}{1.000\ atm} \times 100 = \underline{0.600\%}$$