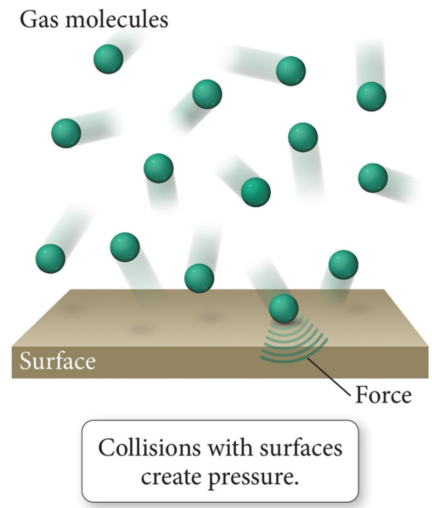
- Gases are composed of particles that are moving around very fast in their container(s).
- These particles move in straight lines until they collide with either the container wall or another particle, and then they bounce off.
- A lot of empty space in the container
  - ✓ The higher the concentration, the greater the pressure.



- $$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{F}{A}$$
- The pressure of a gas depends on several factors:
  - Number of gas particles in a given volume
  - Volume of the container
  - Average speed of the gas particles
- ► Pressure units:

1 atm = 760 mm Hg = 14.7 psi = 101,325 Pa

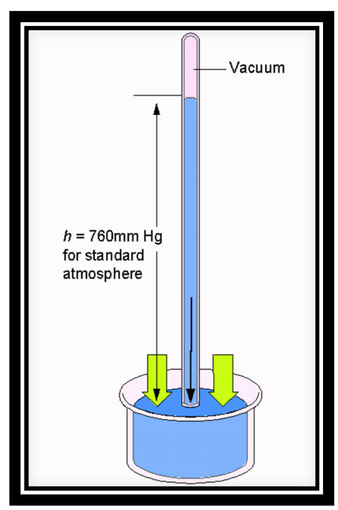
1 mm Hg = 1 torr = 133.32 Pa

Measuring Pressure:

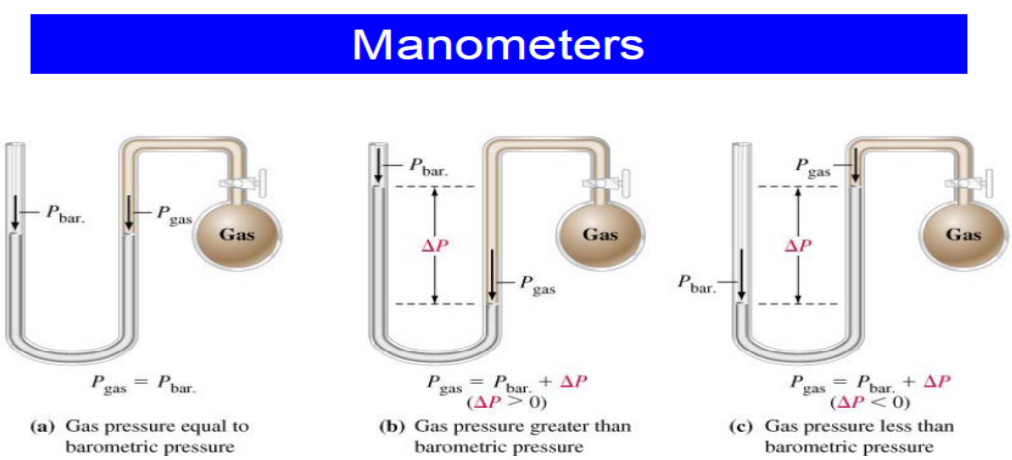
- ➢ **Sphygmomanometer:** An instrument for measuring blood pressure in the arteries, consists rubber cuff that wraps around the upper arm this is connected to a pressure gauge (a column of mercury next to a graduated scale, enabling the determination of systolic and diastolic blood pressure by increasing and gradually releasing the pressure in the cuff

- ► A **barometer** measures pressure as the height of a mercury column. **Atmospheric pressure** presses down on mercury in a dish and pushes it up a tube.

- Pressure= Density of Hg X Height of column X acceleration of Gravity



The Manometer: This device is used to measure the difference in pressure between atmospheric pressure and that of a gas in a vessel.



Ideal Gas Law

Ideal gases are imaginary gases that perfectly fit all of the assumptions of the kinetic molecular theory.

- Gases consist of tiny particles that are far apart relative to their size.
- Collisions between gas particles and between particles and the walls of the container are elastic collisions
- No kinetic energy is lost in elastic collisions
- Gas particles are in constant, rapid motion. They therefore possess kinetic energy, the energy of motion
- There are no forces of attraction between gas particles
- The average kinetic energy of gas particles depends on temperature, not on the identity of the particle.

Ideal Gas Law: PV = nRT

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m^3-Pa/mol-K*	8.314
L-torr/mol-K	62.36

*SI unit

The ideal gas; law works reasonably well for nonreactive gases:
IF the temperature is not too low and
IF the pressure is not too high

Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25 °C, (R=0.0821 atm.L/mol.K), (1 atm =14.7 psi).

0.22

Dinitrogen monoxide (N_2O), laughing gas, is used by dentists as an anesthetic. If a 20.0 L tank of laughing gas contains 2.86 moles N_2O at 23°C, what is the pressure (mm Hg) in the tank?

$2.64X10^3$

- ✓ Under standard temperature and pressure (**STP for gases, T = 0^{o}C (273K); P = 1 atm**)
- I mol of gas occupies a volume of 22.4L (we call it **Molar volume** at **STP**)
- PV = nRT (substitute and get the R value), OR substitute all values to get V =22.4L

- **Gas Density:** Density of a gas is generally given in **g/L**.

Density of a Gas at STP = $\dfrac{\textbf{Molar Mass}}{\textbf{Molar Volume}}$

- The mass of 1 mole = molar mass.
- The volume of 1 mole at STP = 22.4 L.
-

What volume (L) of O_2 gas is needed to completely react with 15.0 g of aluminum at **STP**?

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

The density of an unknown gas is 1.25g/L at STP. Calculate its molar mass

9.33

28.0

A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55 °C and a pressure of 886 mmHg. Find its molar mass.

9.74×10^{-3} mol

31.99

Gas Stoichiometry: Reactions Involving Gases:

Sodium azide decomposes to NaN_3 (s) → $2Na(s) + 3N_2$ (g)
If the mass of NaN_3 is 45.0 g what is the volume of the produced N_2 gas at 303.15 K and 1.15 atm

1.04 mol

22.5 L

Practice

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*) → $6CO_2$ (*g*) + $6H_2O$ (*l*)

0.187 mol CO_2

4.76 L

To decompose 0.11 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 ^oC? what is the pressure Of H_2O? $H_2O_2(l) \longrightarrow 2\ H_2O(g) + O_2(g)$

0.0016 mol O_2

0.016 atm O_2

For water

Calcium hydride reacts with water to form calcium hydroxide (aqueous) and hydrogen gas. How many grams of calcium hydride are needed to form 205 L of hydrogen gas at 23°C and 850 mmHg?

9.44 mol H_2

198.5

Practice

How many liters of oxygen gas form when 294 g of $KClO_3$ completely reacts in the following reaction? Assume the oxygen gas is collected at *P* = 755 mmHg and *T* = 305 K.

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

3.60 ml

90.7 L

Combined Gas Law

When we have more complex situation where there are more than two variables
We can rearrange PV= nRT to get R = PV/nT

In general, if we have a gas under two sets of conditions, then:

- We often have a situation in which *P*, *V,* and *T* all change for a fixed number of moles of gas.

A sample of helium gas has a volume of 0.180 L, a pressure of 0.800 atm and a temperature of 29°C. At what temperature (°C) will the helium have a volume of 90.0 mL and a pressure of 3.20 atm (n constant)?

604 K
331C

❖ Mixtures of Gases:

Use ideal gas law to characterize gas mixtures.

- ✓ Gases behave independently

The pressure of a single gas in a mixture of gases is called its **partial pressure**

✓ **Dalton's Law of Partial Pressures:**

$$P_{\text{total}} = P_a + P_b + P_c + \ldots$$

The concentration of any individual component in the gas mixture us usually expressed as **Mole Fraction ($\mathcal{X}$)**

In a gas mixture made from gases **a, b, c**

X_a =

Sum of Xs (mol fractions)= 1

The pressure of each gas in a mixture is a fraction of the total pressure

Find the Partial Pressure of Neon in a Mixture of Ne and Xe with Total Pressure 3.9 atm, Volume 8.7 L, Temperature 598 K, and 0.17 moles Xe

0.69 mol total

2.9 atm

- Also For gases, the mole fraction of a component in a gas mixture is equivalent to its **fractional composition** in a mixture (**percent divided by 100%**).

Nitrogen has a 78% composition of air; find its partial pressure if the atmospheric pressure is 1 atm.

0.78 atm

If the atmospheric pressure today is 745 mm Hg, what is the partial pressure (mm Hg) of O_2 in the air (O_2 has a 21% composition in air)?

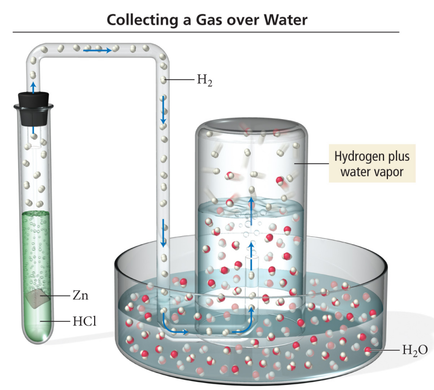
156.45

Collecting Gases over water (Water displacement)

In the laboratory, hydrogen gas is usually made by the following reaction:

$$Zn(s) + 2\,HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

How many liters of H_2 gas, collected over water at an atmospheric pressure of 752 mm Hg and a temperature of 21.0°C, can be made from 1.566 g of Zn and excess HCl? The partial pressure of water vapor is 18.65 mm Hg at 21.0°C.



733.35 mmHg

0.0239 mol H_2

0.599 L

Acetylene may be produced from the reaction between calcium carbide and water. How many moles of acetylene are present in a 500.0 mL sample of gas collected over water at 20.0 ºC and a total pressure of 758 torr? The vapor pressure of water at 20.0 ºC is 17.54 mm Hg.

740.46 torr

0.0203 mol

Practice

What mass of water vapor is produced from the gas phase reaction of O_2 with H_2 in a 2.5 L container at 28 ºC if the partial pressures of O_2 and H_2 are P_{O2} = 0.98 atm and P_{H2} = 1.24 atm?

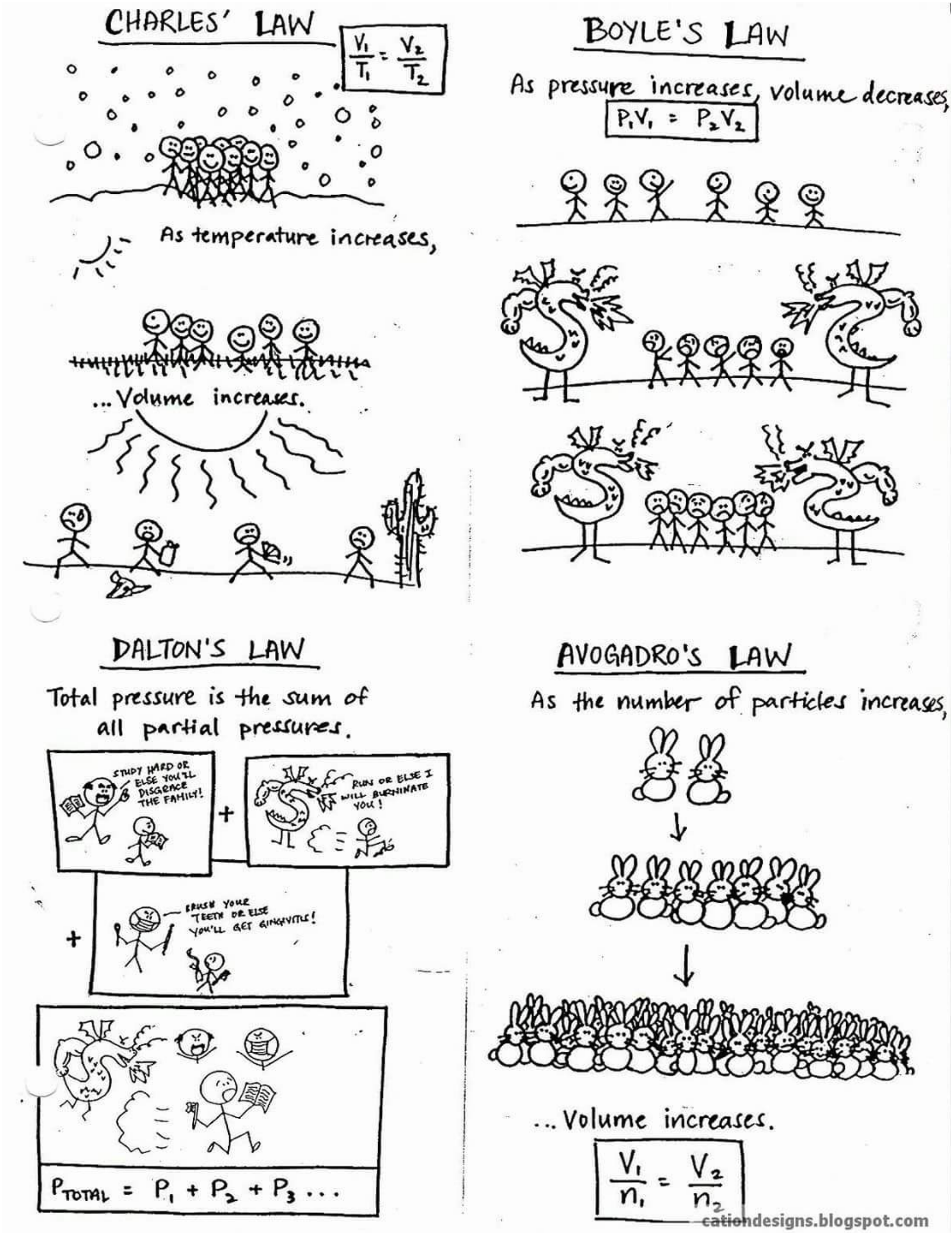
$$2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)$$

0.0991 mol O_2

0.125 mol H_2

2.25 g

Gas Laws


CHARLES' LAW
$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
As temperature increases,
...Volume increases.
BOYLE'S LAW
As pressure increases, volume decreases,
$P_1V_1 = P_2V_2$
DALTON'S LAW
Total pressure is the sum of all partial pressures.
STUDY HARD OR ELSE YOU'LL DISGRACE THE FAMILY!
RUN OR ELSE I WILL BURNINATE YOU!
BRUSH YOUR TEETH OR ELSE YOU'LL GET GINGIVITIS!
$P_{TOTAL} = P_1 + P_2 + P_3 \ldots$
AVOGADRO'S LAW
As the number of particles increases,
...Volume increases.
$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
cationdesigns.blogspot.com

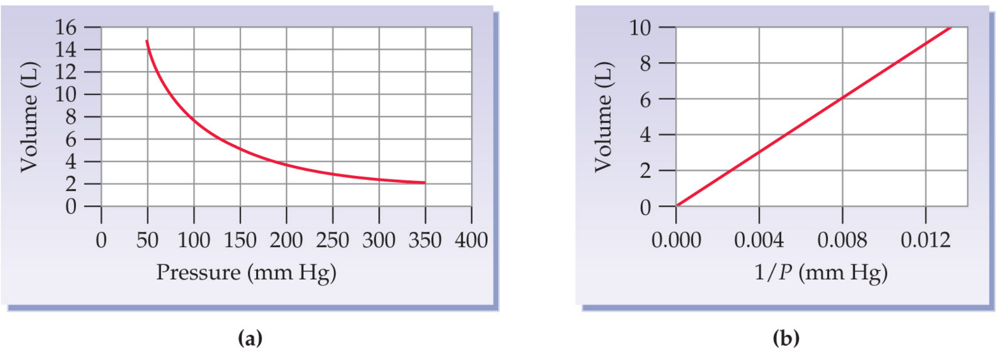
Gas Laws:

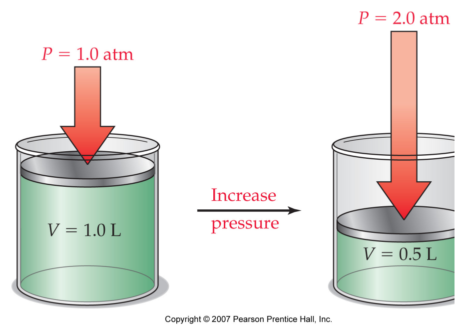
- There are four basic properties of a gas: pressure (*P*), volume (*V*) , temperature (*T*), and amount in moles (*n*).

Boyle's Law: The Relation Between Volume and Pressure

Constant temperature (T) and moles (n) of gas

✓ The pressure of a gas is inversely proportional to its volume





A Balloon Is Put in a Bell Jar and the Pressure Is Reduced from 782 torr to 0.500 atm. If the Volume of the Balloon Is Now 2780 mL, What Was It Originally?
(1 atm = 760 torr)

1350 ml

Practice

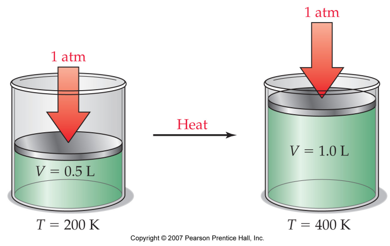
If a sample of helium gas has a volume of 120 mL and a pressure of 850 mm Hg, what is the new volume if the pressure is changed to 425 mm Hg ?

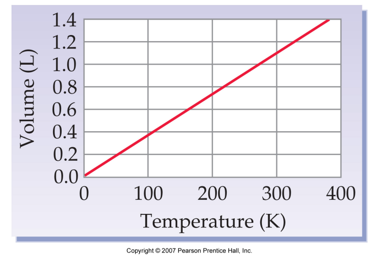
240 ml

Charles' Law: The Relation Between Volume and Temperature

Constant pressure (p) and moles (n) of gas

✓ The volume of a gas increases with increasing temperature.





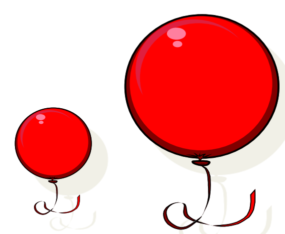

A balloon has volume of 785 mL at 21°C. If the temperature drops to 0°C, what is the new volume of the balloon (P constant)?

729 ml

Avogadro's Law: The Relation Between Volume and Amount of gas

Constant pressure (p) and temperature (T)

- ✓ Volume directly proportional to the number of gas molecules



A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L? (Assume constant pressure and temperature.)

0.29 mol

0.07

Reading Assignment P224-226

Kinetic-Molecular Theory

Theory used to explain gas laws.

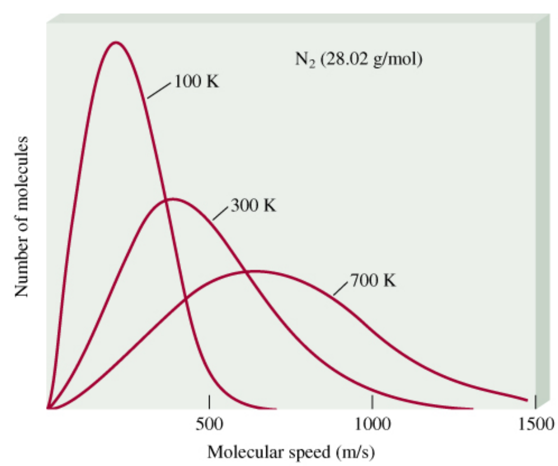
KMT assumptions are

- Gases consist of tiny particles (size negligible) in constant, random motion.
- Most of the volume of a gas is empty space (Volume of gas molecules is negligible).
- Gas particles move independently (ignore each other). No interaction between them
- The collision of one particle with another (or with the walls of its container) is completely elastic (there is no overall *loss of energy*)

❖ The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature.

✓ As you raise the temperature of the gas, the average speed of the particles increases.

➢ But not all the gas particles are moving at the same speed!



- $KE_{avg} = ½ N_A mu^2$
- $KE_{avg} = (3/2)RT$

✓ N_A is Avogadro's number.

✓ ***m** is mass of particle in Kg*

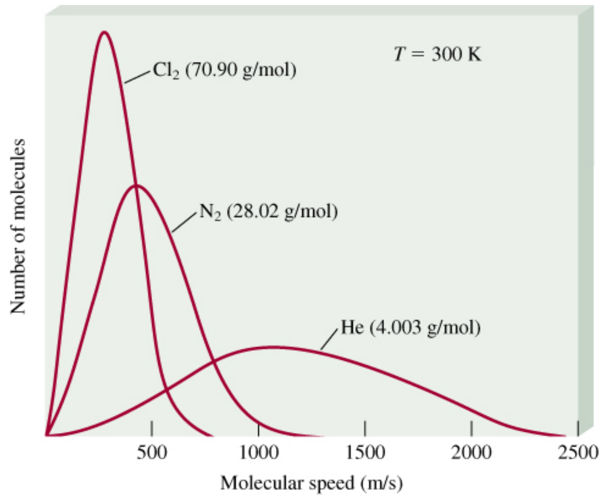
✓ **N_A m** = *molar mass ($\mathcal{M}$) in Kg/mol*

✓ ***R* is the gas constant in energy units, 8.314 J/mol.K**

✓ **1 J = 1 kg · m²/s²**

❖ Average kinetic energy of the gas molecules depends on the average mass (molar mass M) and velocity.

$$KE_{avg} = \frac{1}{2}N_A m\overline{u^2}$$



$$u_{rms} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

($\mathcal{M}$) is molar mass of gas in Kg/mol

rms=root mean square speed

� How the individual gas laws follow the kinetic molecular theory:

Boyle's Law:
Constant n and T , P inversely proportional to V.
If we decrease the volume then the number of collisions will increase, resulting in greater pressure.

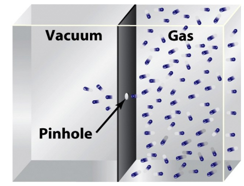
Charles' Law:
Constant n and P, V directly proportional to T.
KMT says when T increases the average speed of the gas particles increase thus the average kinetic energy icreases resulting more frequent collisions, thus increasing the pressure. But Charles' law states that P is constant, so to keep P constant th volume has to increase.

Avogadro's law
Constant T and P, V is directly proportional to n.
KMT : as n increase then the number of collisons icrease resulting in increasing the pressure . But the pressure is constant according to Avogadro's law. So to keep p constant the volume has to increase.

Dalton's Law:
Overall pressure id the sum of all partial pressures.
KMT says size is negligible and no interactions between the gas particles. And the only property that distinguishes between the different particles is the mass. However different masses have the same average KE at a given T, so they exert the same pressure. (even adding different kinds of gas has the same effect as simply adding more particles)

Graham's Law: Diffusion and Effusion of Gases:

<u>Effusion:</u> The escape of a gas through a pinhole into a vacuum without molecular collisions.



Effusion is the escape of a gas through a pinhole into a vacuum without molecular collisions.

Figure 9-13 Chemistry, 5/e
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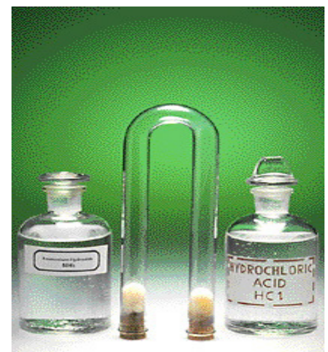
Gas Diffusion:
relation of mass to rate of diffusion

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{\mathcal{M}_B}{\mathcal{M}_A}}$$



Gaseous diffusion of $NH_3(g)$ and $HCl(g)$

- In a certain experiment the rate of diffusion of a certain unknown gas is found to be 1.47 times as fast as that of hydrogen chloride gas. What is the molar mass of the unknown gas?

16.9 g/mol

Reading Assignment Section 5.10 (P232-236)

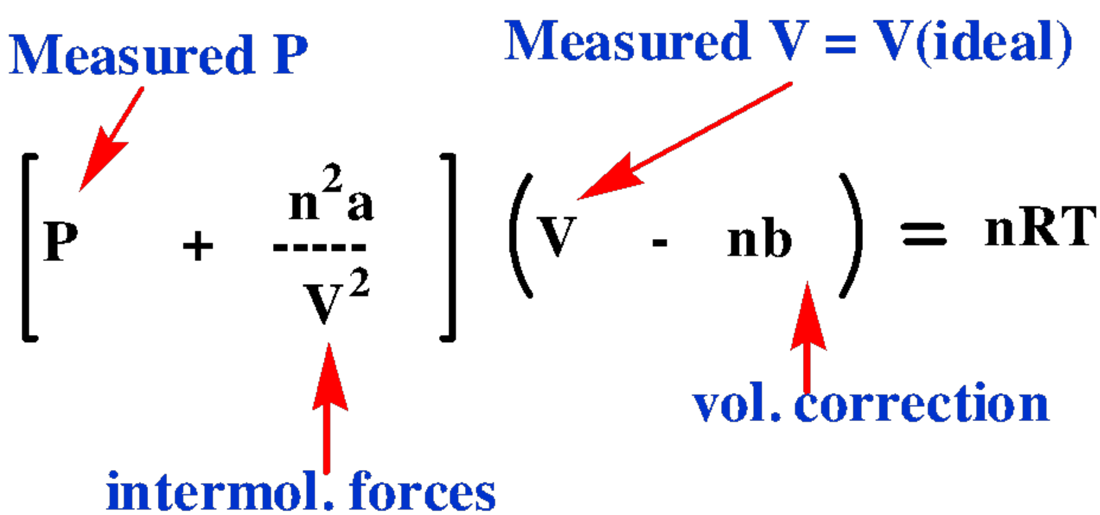
✓ Deviations from Ideal Behavior Real Gases Do Not Behave Ideally

Likely to behave nearly ideally
Gases at high temperature and low pressure
Small non-polar gas molecules

Likely not to behave ideally
Gases at low temperature and high pressure
Large, polar gas molecules

- Real gases DO experience inter-molecular attractions (otherwise a gas could not become a liquid).
- Real gases DO have volume
- Real gases DO NOT have elastic collisions

Real Gases and the Van der Waals's Equation:



Cl_2 gas **has a** = 6.49, b = 0.0562

For 4.00 mol Cl_2 in a 4.00 L tank at 100.0 ^oC.

P (ideal) = nRT/V = 30.6 atm

P (van der Waals) = 26.0 atm