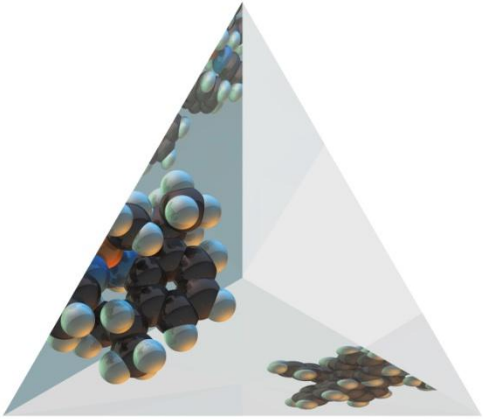


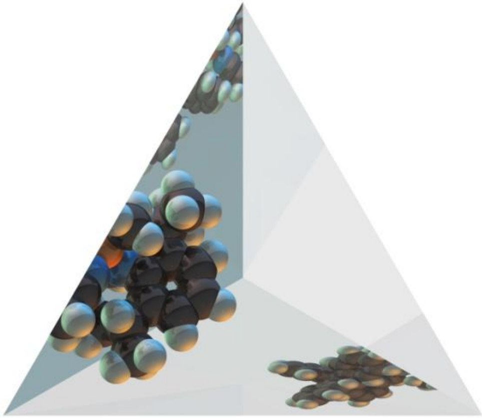
CHEMISTRY

The Central Science

8th Edition

Chapter 10
Gases

Kozet YAPSAKLI



• Elements that exist as gases at 25^{0}C and 1 atmosphere

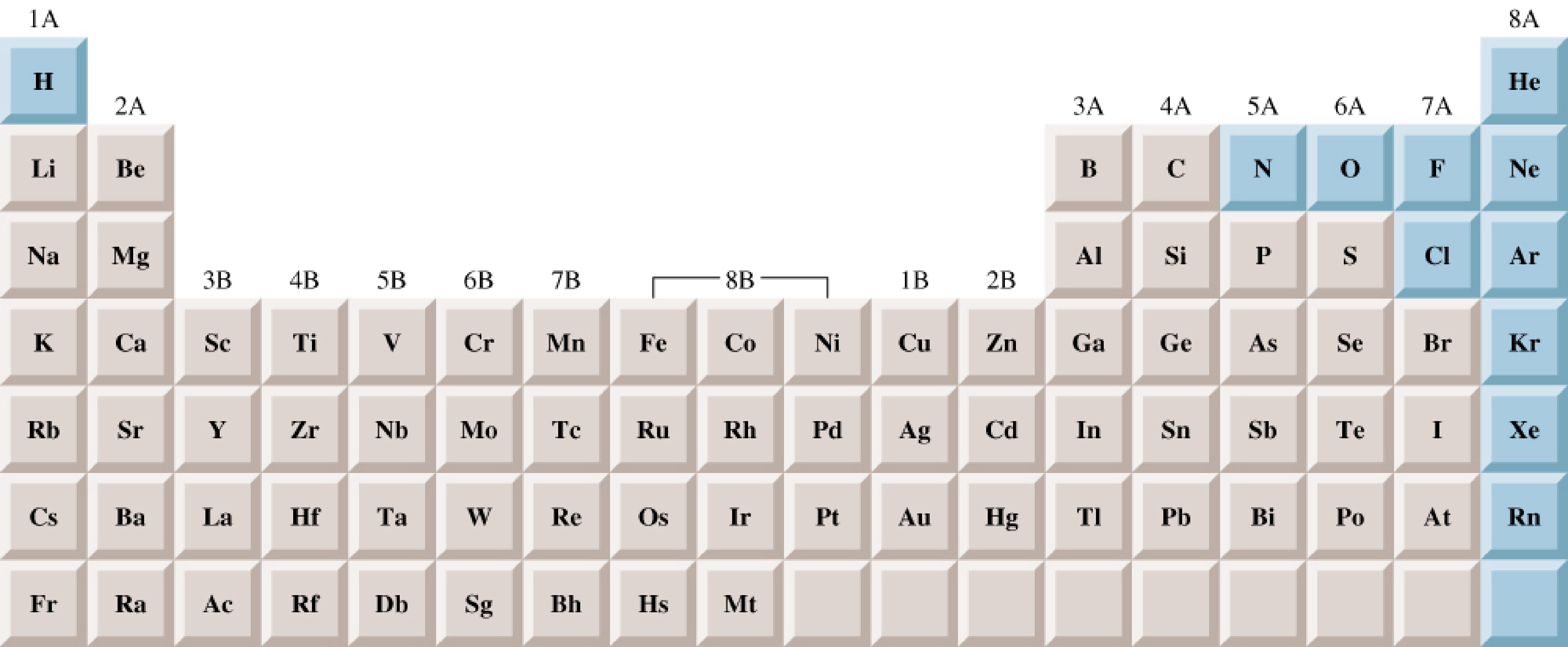


TABLE 10.1 Some Common Compounds That Are Gases At Room Temperature

Formula	Name	Characteristics
HCN	Hydrogen cyanide	Very toxic, slight odor of bitter almonds
H_2S	Hydrogen sulfide	Very toxic, odor of rotten eggs
CO	Carbon monoxide	Toxic, colorless, odorless
CO_2	Carbon dioxide	Colorless, odorless
CH_4	Methane	Colorless, odorless, flammable
C_2H_4	Ethylene	Colorless; ripens fruit
C_3H_8	Propane	Colorless; bottled gas
N_2O	Nitrous oxide	Colorless, sweet odor, laughing gas
NO_2	Nitrogen dioxide	Toxic, red-brown, irritating odor
NH_3	Ammonia	Colorless, pungent odor
SO_2	Sulfur dioxide	Colorless, irritating odor

Characteristics of Gases

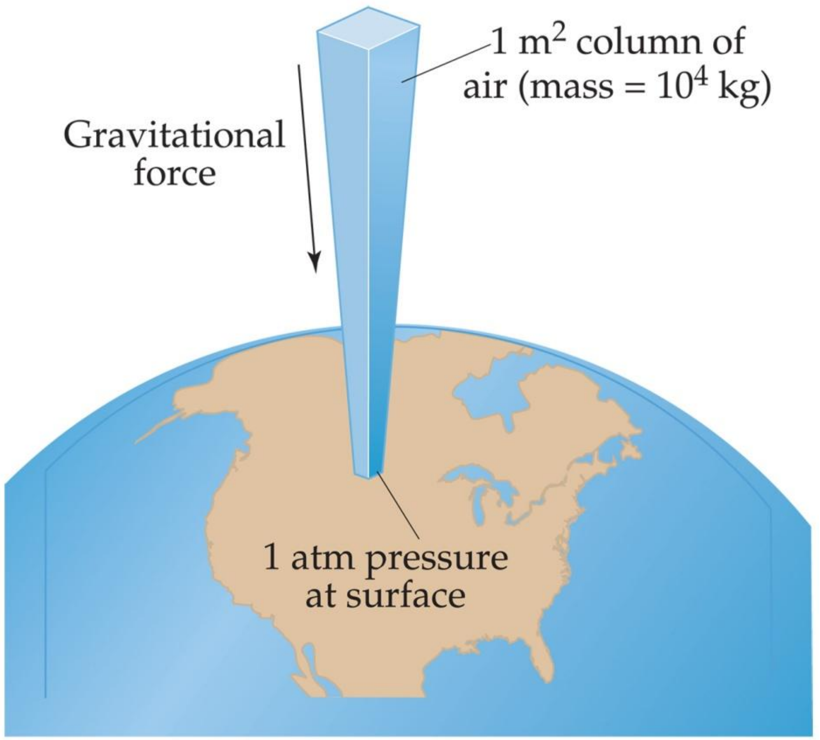
- Unlike liquids and solids, they
  - Expand to fill their containers.
  - Are highly compressible.
  - Have extremely low densities.
  - Form homogoneouns mixtures

Pressure of a Gas

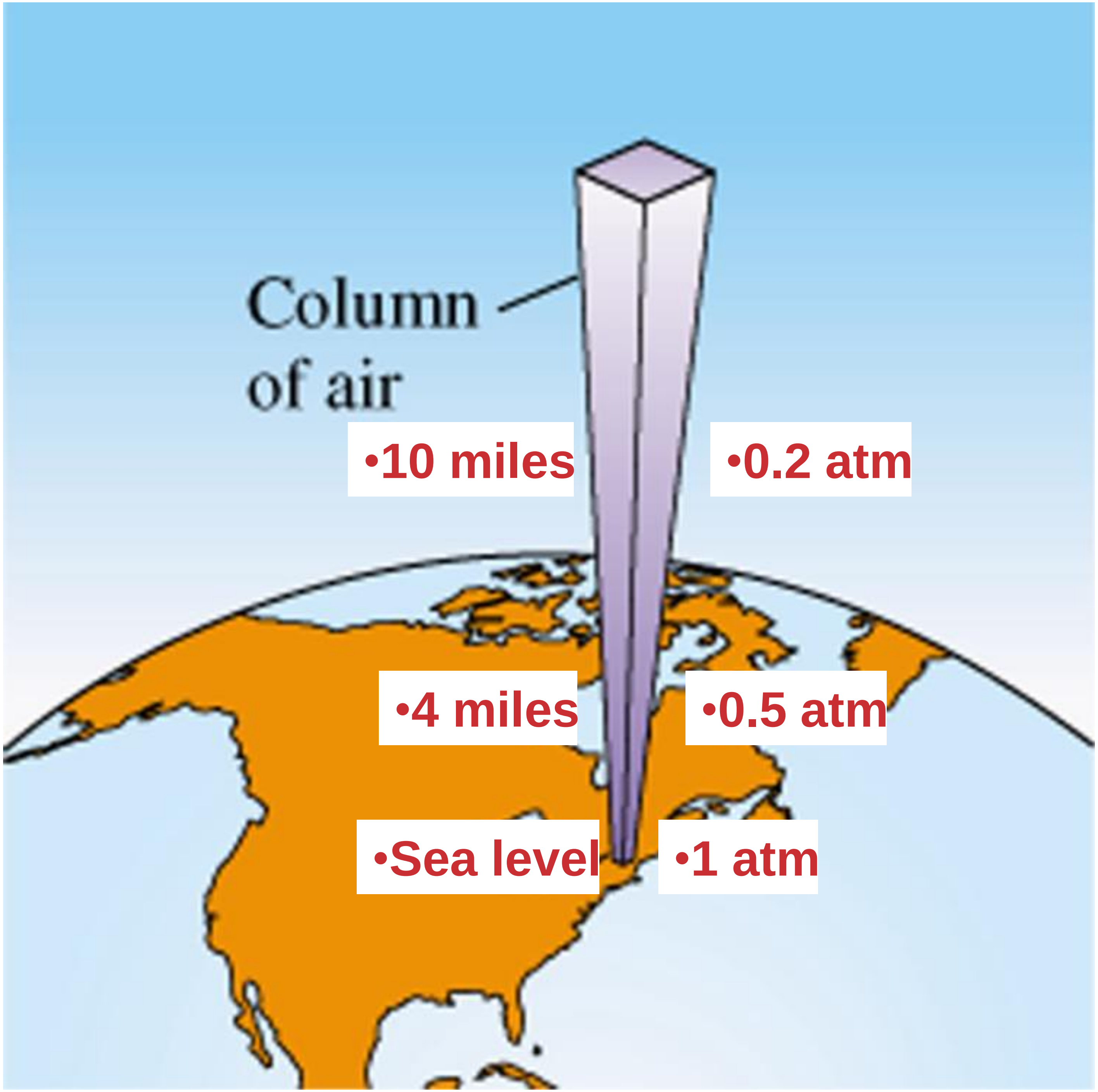
- Pressure is the amount of force applied to an area.

$$P = \frac{F}{A}$$

- Atmospheric pressure is the weight of air per unit of area.



Gases


Column of air
•10 miles
•0.2 atm
•4 miles
•0.5 atm
•Sea level
•1 atm

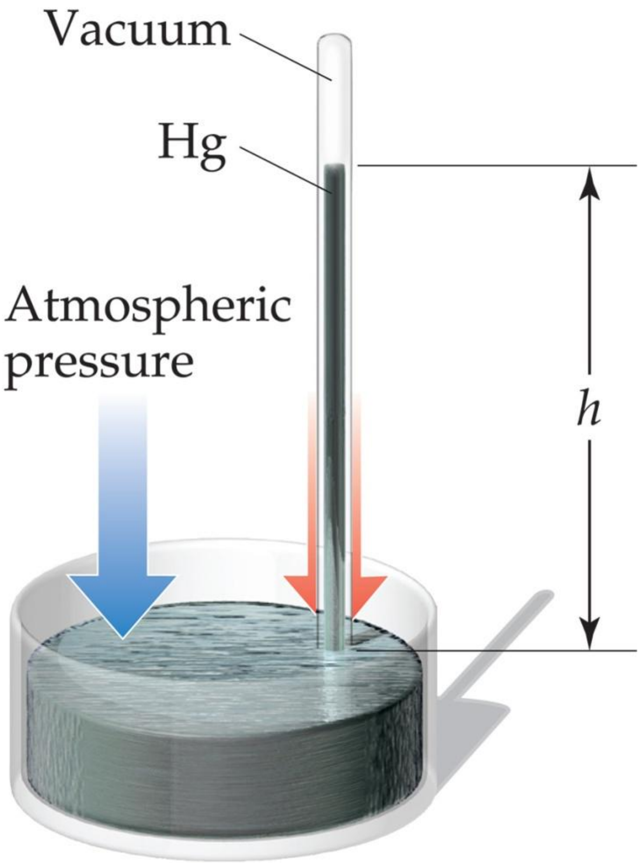
Units of Pressure

- Pascals
  - $1\ Pa = 1\ N/m^2$
- Bar
  - $1\ bar = 10^5\ Pa = 100\ kPa$
- Atmosphere

  1.00 atm = 760 torr

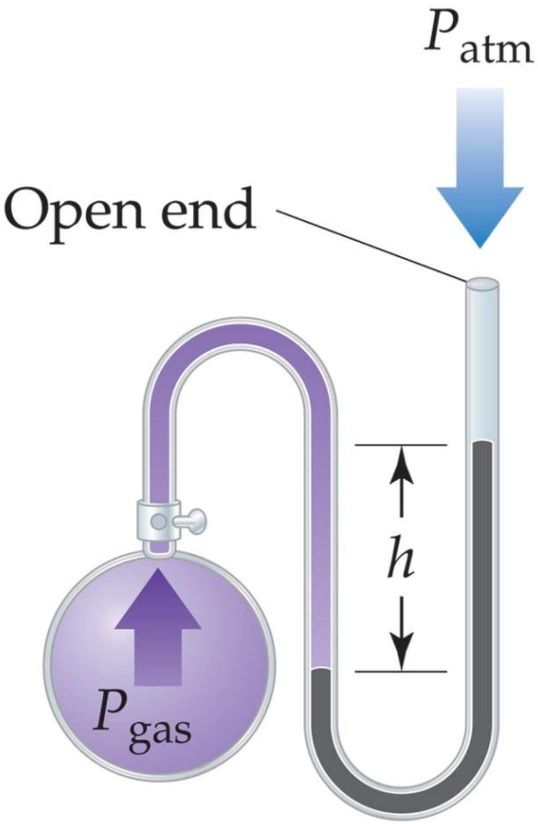
  = 14.7 psi

  =? m water



Gases

Manometer



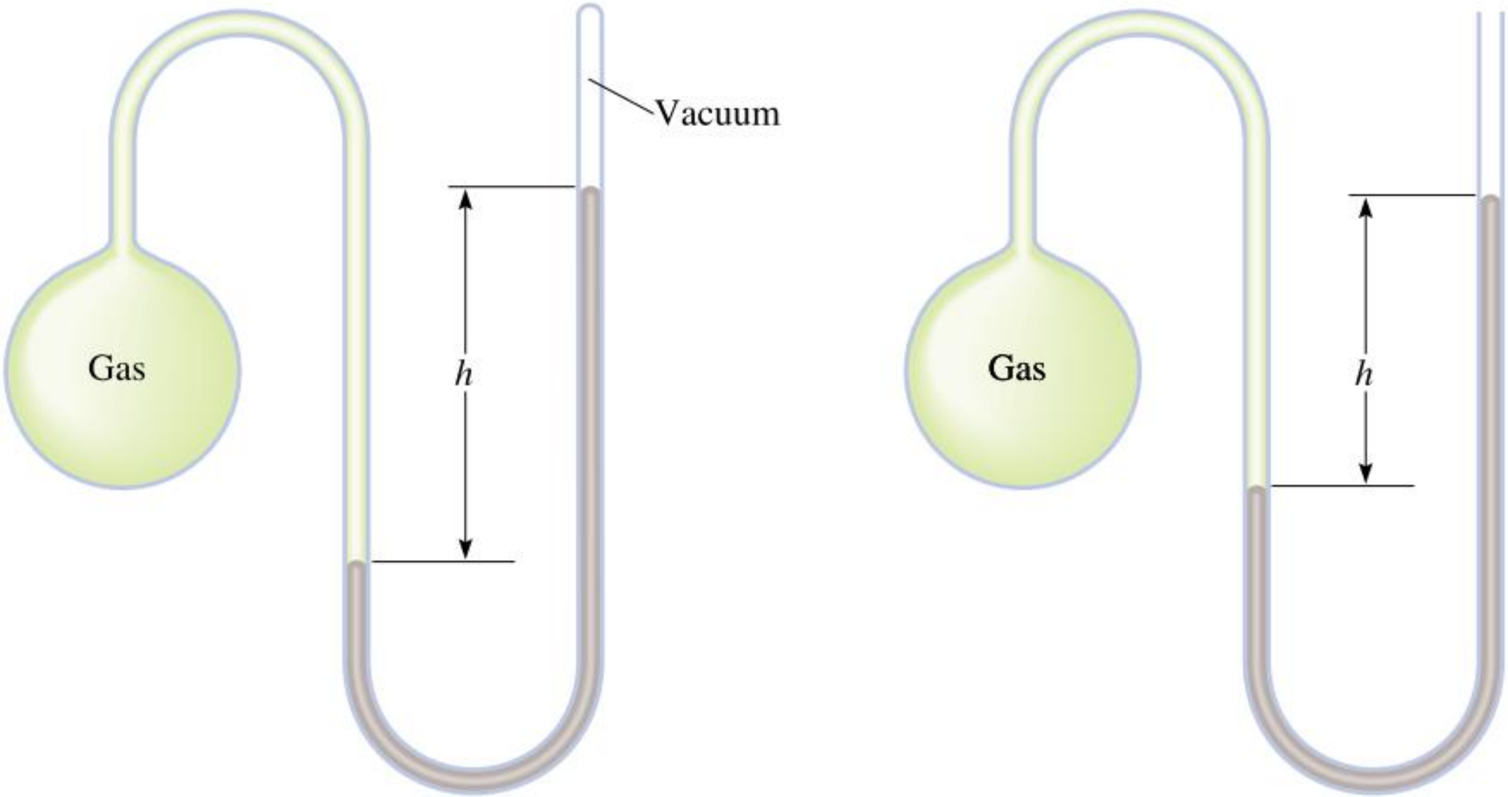
$$P_{gas} = P_{atm} + P_h$$

Figure 10.4

Used to measure the difference in pressure between atmospheric pressure and that of a gas in a vessel.

- If $P_{gas} < P_{atm}$ then $P_{gas} + P_{h2} = P_{atm}$
- If $P_{gas} > P_{atm}$ then $P_{gas} = P_{atm} + P_{h2}$

Gases



$P_{gas} = P_h$

(a)

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

(b)

Standard Pressure

- Normal atmospheric pressure at sea level.

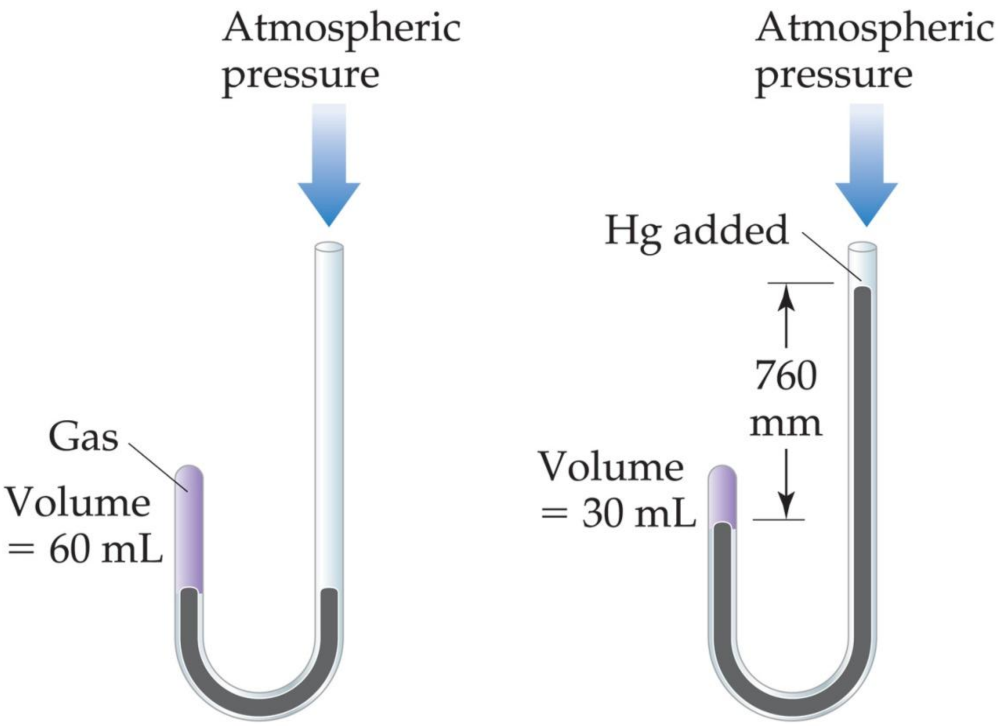
1.00 atm

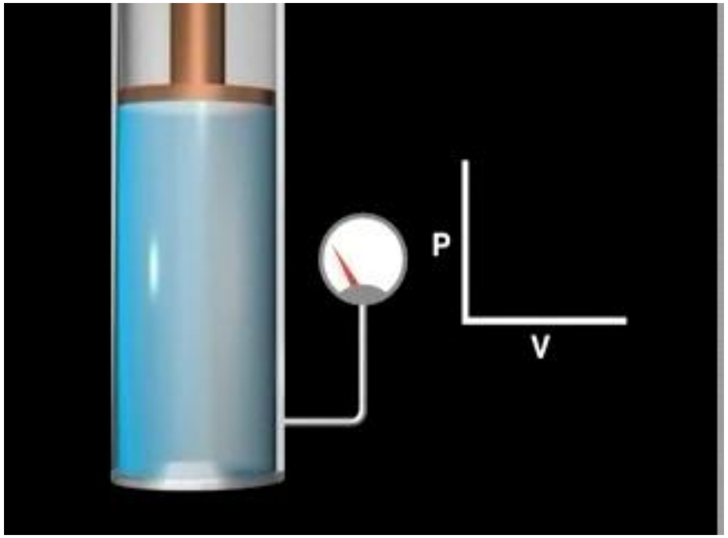
760 torr (760 mm Hg)

101.325 kPa

Boyle's Law:Pressure-Volume

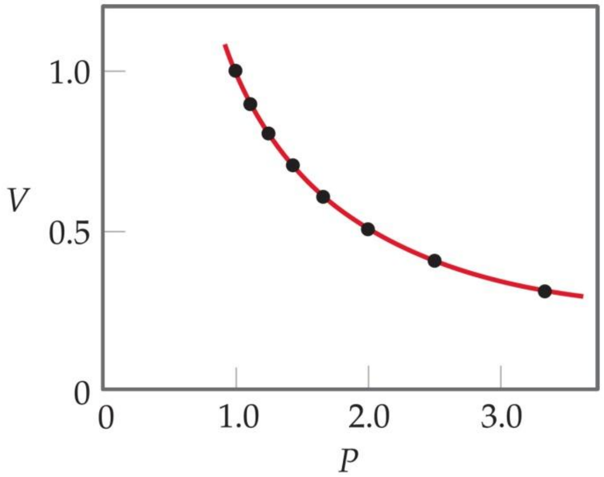
The volume of a fixed quantity of gas at constant temperature is inversely proportional to the pressure.





As *P* and *V* are inversely proportional

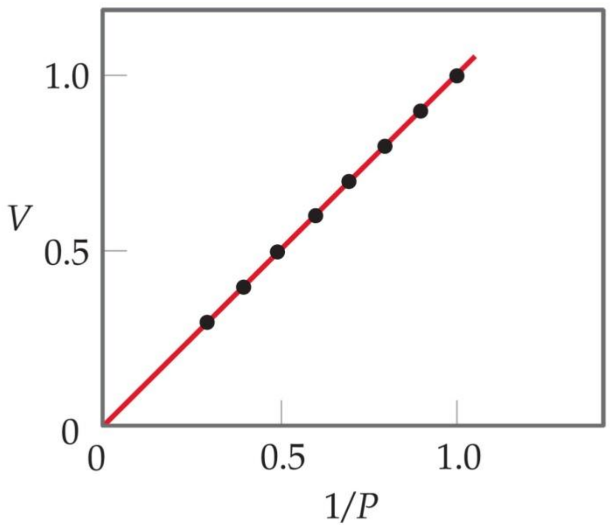
A plot of *V* versus *P* results in a curve.



Since $PV = k$

$V = k\,(1/P)$

This means a plot of *V* versus 1/*P* will be a straight line.

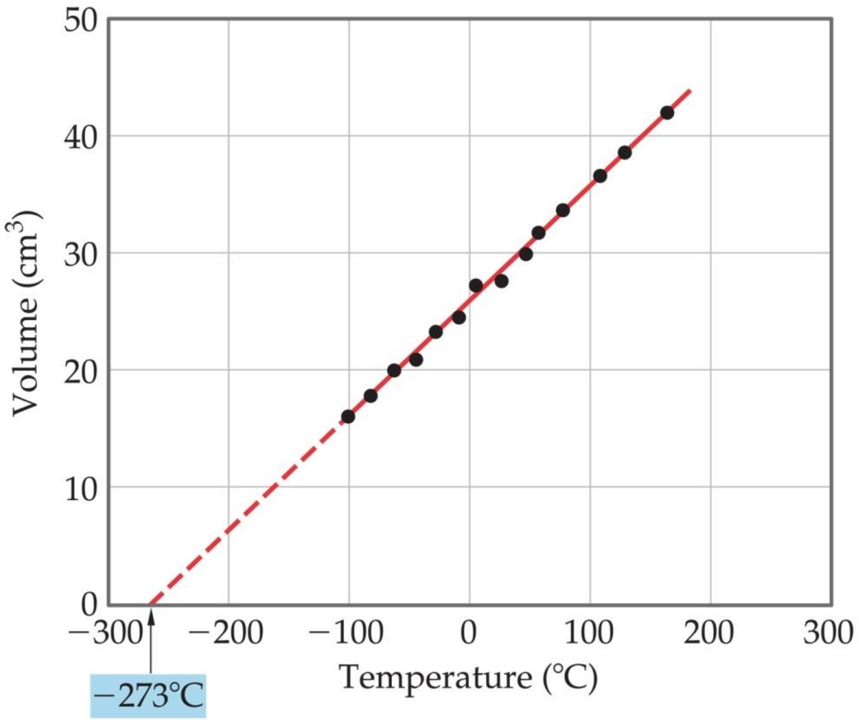


Gases

Charles's Law: Temperature-Volume

- The volume of a fixed amount of gas at constant pressure is directly proportional to its **absolute temperature.**

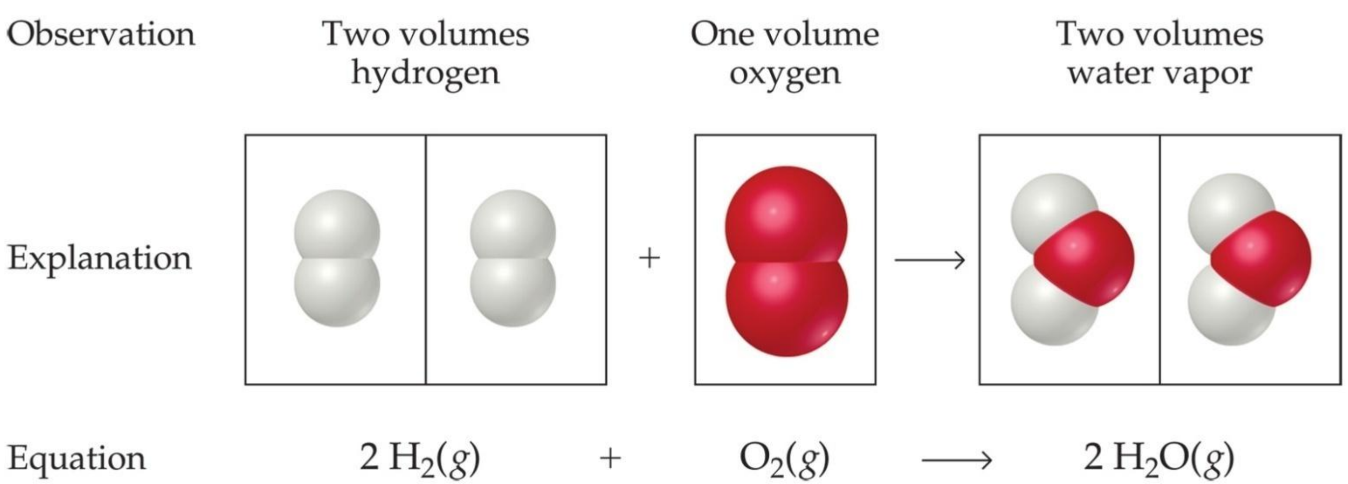
- i.e., $$\frac{V}{T} = k$$



A plot of *V* versus *T* will be a straight line.

Avogadro's Law

- The volume of a gas at constant temperature and pressure is directly proportional to the number of moles of the gas.
- Mathematically, this means $V = kn$



The Gas Laws

The Quantity-Volume Relationship: Avogadro's Law

	He	N_2	CH_4
Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Ideal-Gas Equation

- So far we've seen that

$V \propto 1/P$ (Boyle's law)

$V \propto T$ (Charles's law)

$V \propto n$ (Avogadro's law)

- Combining these, we get

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

Ideal-Gas Equation

The relationship

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

then becomes $V = R\dfrac{nT}{P}$ or $PV = nRT$

$$PV = nRT$$

$$V = \frac{nRT}{P} = \frac{(1\,\text{mol})(0.08206\,\text{L}\cdot\text{atm/mol}\cdot\text{K})(273.15\,\text{K})}{1.000\,\text{atm}} = 22.41\,\text{L}$$

Ideal-Gas Equation

The constant of proportionality is known as *R*, the gas constant.

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.

R=1atm*22.414L/(1mol*273.15K)

Densities of Gases

If we divide both sides of the ideal-gas equation by *V* and by *RT*, we get

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

Densities of Gases

- We know that

moles $\times$ molecular mass = mass

$$n \times M = m$$

- So multiplying both sides by the molecular mass (M) gives

$$\frac{m}{V} = \frac{PM}{RT}$$

Densities of Gases

- Mass ÷ volume = density
- So, $$d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$
- Note: One only needs to know the molecular mass, the pressure, and the temperature to calculate the density of a gas.

Molecular Mass

We can manipulate the density equation to enable us to find the molecular mass of a gas:

$$d = \frac{PM}{RT}$$

Becomes

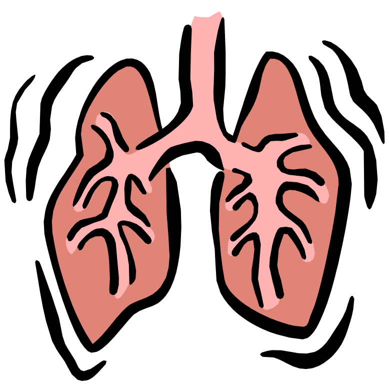
$$M = \frac{dRT}{P}$$

Gases

•Chemistry in Action:

•Scuba Diving and the Gas Laws

Depth (ft)	Pressure (atm)
0	1
33	2
66	3

•P ↓  •V ↑

Dalton's Law of Partial Pressures

- The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone.

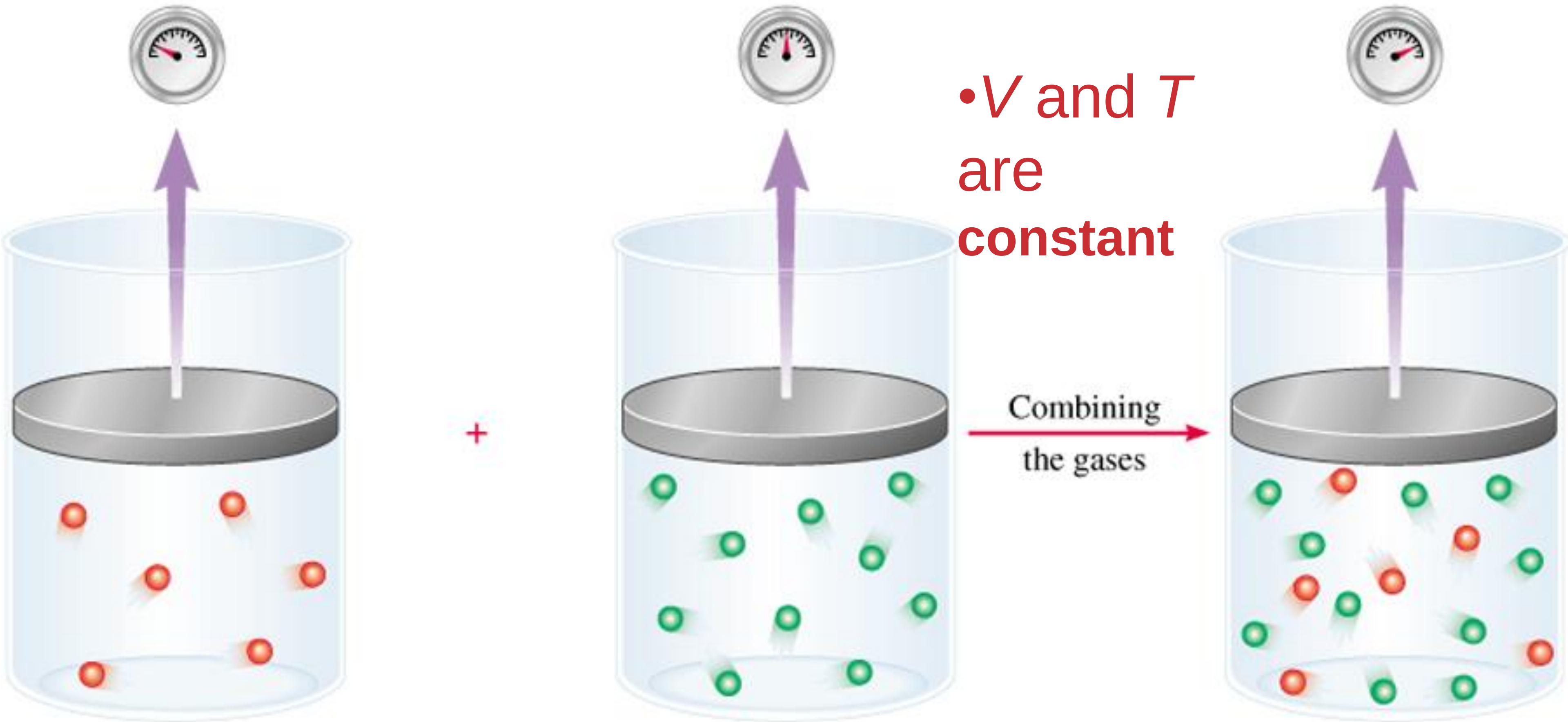
- In other words,

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

$$P_i = n_i \left(\frac{RT}{V} \right)$$

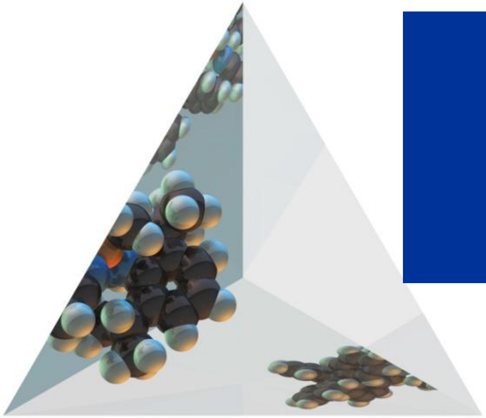
•Dalton's Law of Partial Pressures

•V and T are **constant**



•P_1 •P_2 •$P_{total} = P_1 + P_2$

Gases



Gas Mixtures and Partial Pressures

- Combining the equations

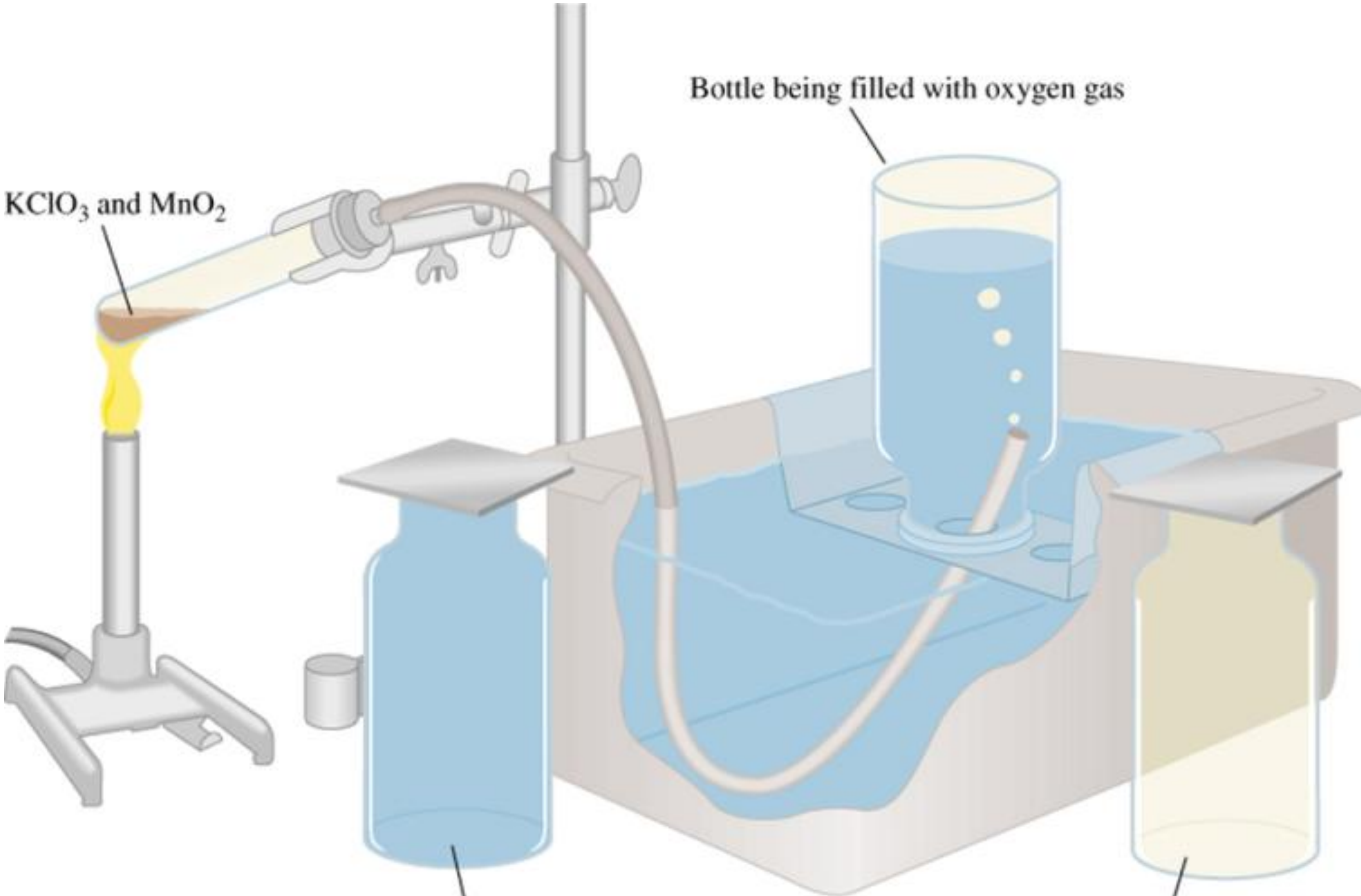
$$P_{\text{total}} = (n_1 + n_2 + n_3 + \cdots)\left(\frac{RT}{V}\right)$$

Partial Pressures and Mole Fractions

- Let n_i be the number of moles of gas i exerting a partial pressure P_i, then

$$P_i = X_i P_{\text{total}}$$

where X_i is the **mole fraction** (n_i/n_t).

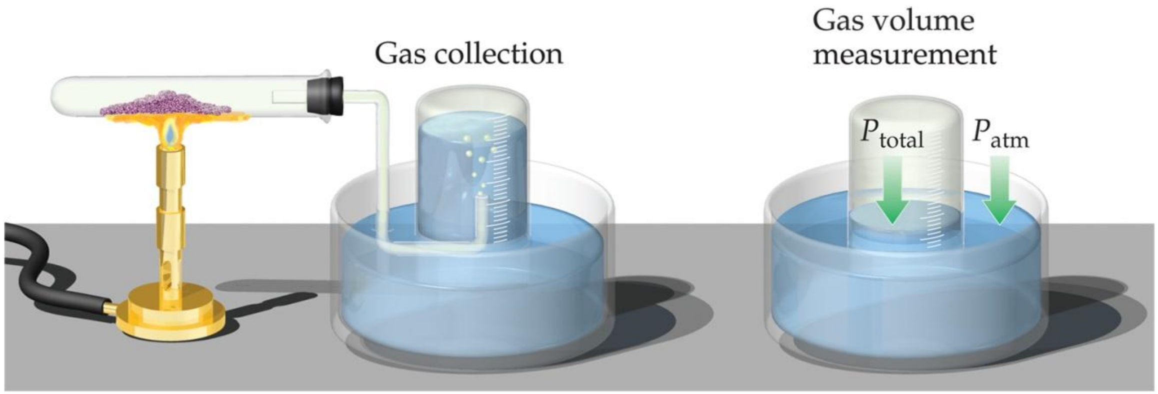


•Bottle full of oxygen gas and water vapor

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

•$P_T = P_{O_2} + P_{H_2O}$

Partial Pressures



- When one collects a gas over water, there is water vapor mixed in with the gas.
- To find only the pressure of the desired gas, one must subtract the vapor pressure of water from the total pressure. (Example 10.11)

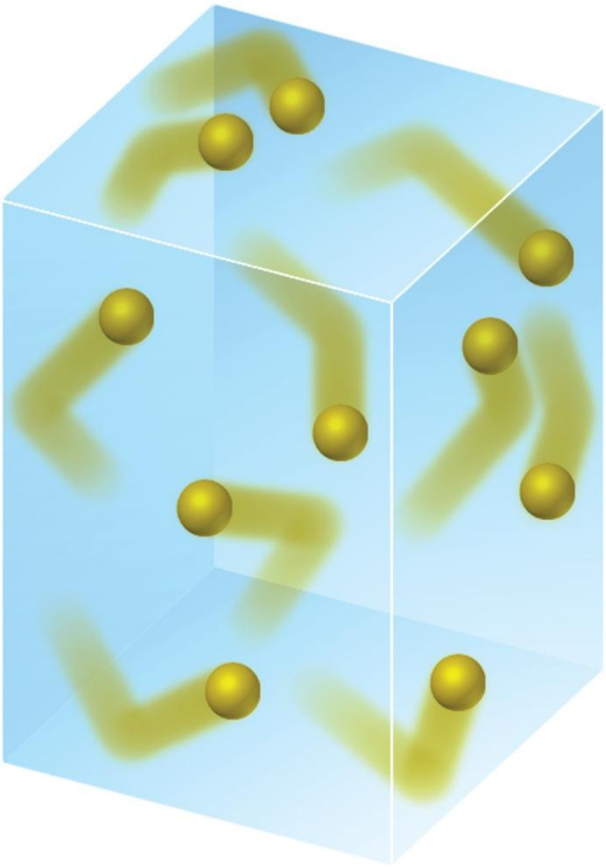
$$P_{total} = P_{gas} + P_{water}$$

Table of Vapor Pressures for Water

Temperature, °C	Pressure, mmHg	Temperature, °C	Pressure, mmHg
0	4.6	27	26.7
5	6.5	28	28.3
10	9.2	29	30.0
11	9.8	30	31.8
12	10.5	35	42.2
13	11.2	40	55.3
14	12.0	45	71.9
15	12.8	50	92.5
16	13.6	55	118.0
17	14.5	60	149.4
18	15.5	65	187.5
19	16.5	70	233.7
20	17.5	75	289.1
21	18.7	80	355.1
22	19.8	85	433.6
23	21.1	90	525.8
24	22.4	95	633.9
25	23.8	100	760.0
26	25.2	105	906.1

ses

Kinetic-Molecular Theory



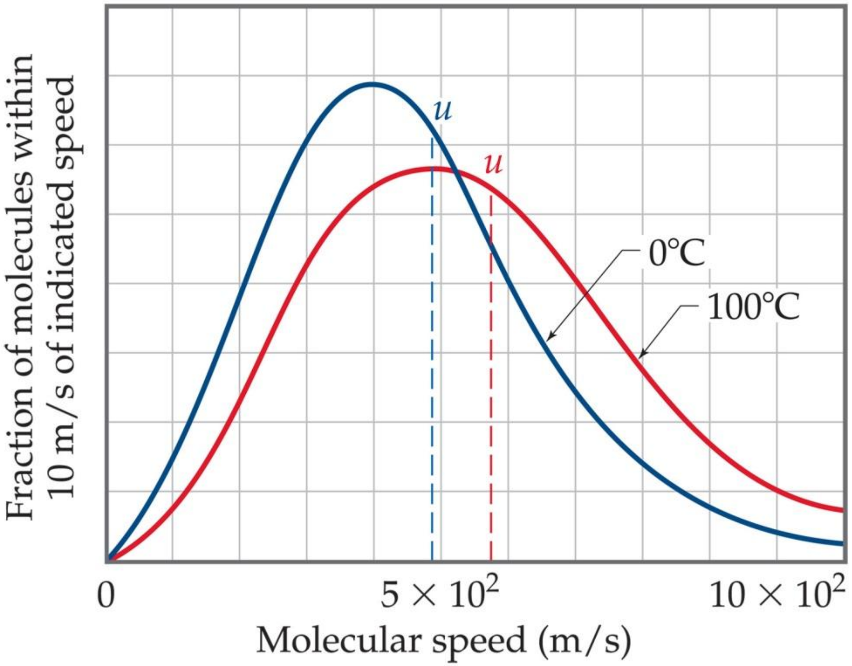
- This is a model that aids in our understanding of what happens to gas particles as environmental conditions change.
- It gives us an understanding of pressure and temperature on the molecular level.

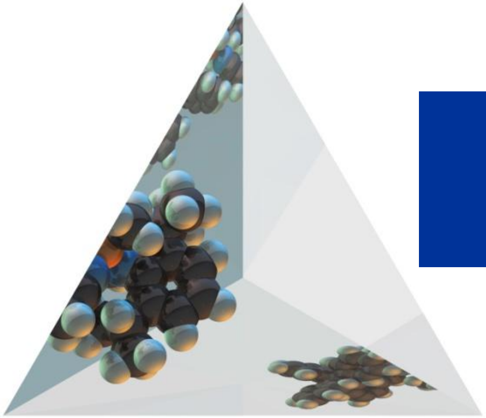
Kinetic-Molecular Theory

- Assumptions:
  - Gases consist of a large number of molecules that are in constant random motion.
  - Volume of individual molecules negligible compared to volume of container.
  - Attractive and repulsive forces between gas molecules are negligible.
  - Energy can be transferred between molecules, but total kinetic energy is constant at constant temperature.
  - Average kinetic energy of molecules is proportional to temperature.

Kinetic-Molecular Theory

- The average kinetic energy of the molecules is proportional to the absolute temperature.
- Each gas molecule has a different energy.





Kinetic Molecular Theory

- As kinetic energy increases,
  - velocity of the gas molecules increases.
- Root mean square speed (rms), *u*, is the average molecular speed
- Average kinetic energy, ε :

$$\varepsilon = \frac{1}{2}mu^2$$

Application to Gas Laws

As -***volume*** increases- at -constant ***temperature***-,

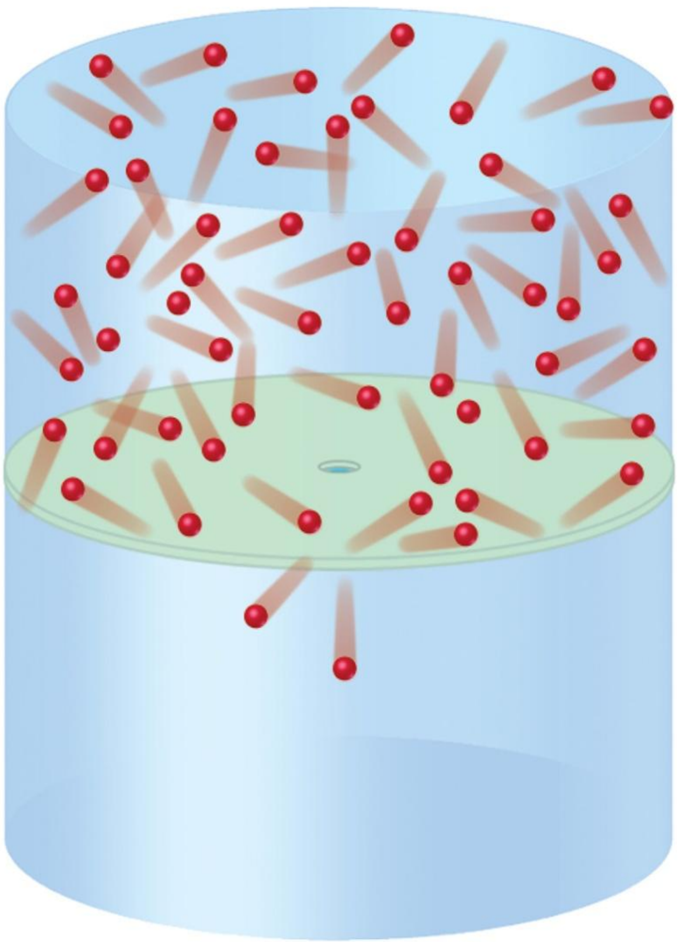
the ***average kinetic*** of the gas remains constant and *u* is constant.

Molecules must move a longer distance between collusions

Therefore, ***pressure*** decreases.

- If ***temperature*** *increases* at *constant* ***volume***,
  - the average ***kinetic energy*** of the gas molecules *increases*.
  - Therefore, there are more collisions with the container walls and the ***pressure*** *increases*.

Effusion



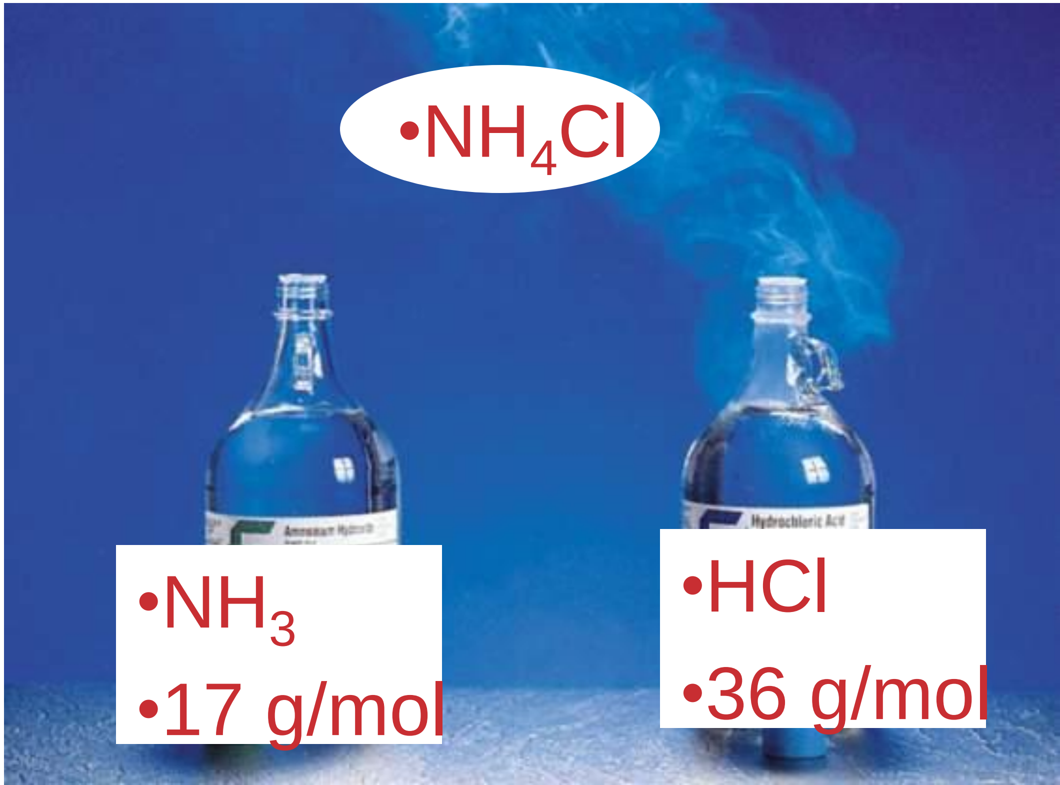
The escape of gas molecules through a tiny hole into an evacuated space.

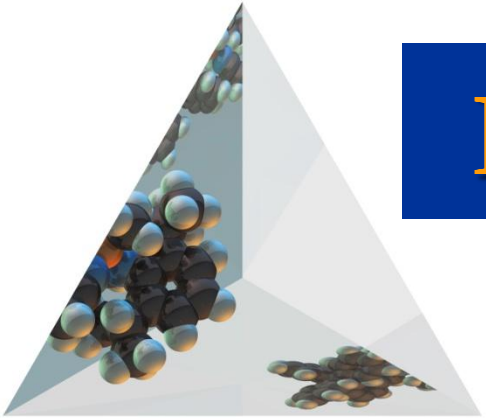
Diffusion

The spread of one substance throughout a space or throughout a second substance.

- **Gas diffusion** is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties.

- Gases meet to form NH_4Cl
- HCl heavier than NH_3
- Therefore, NH_4Cl forms closer to HCl end of tube.



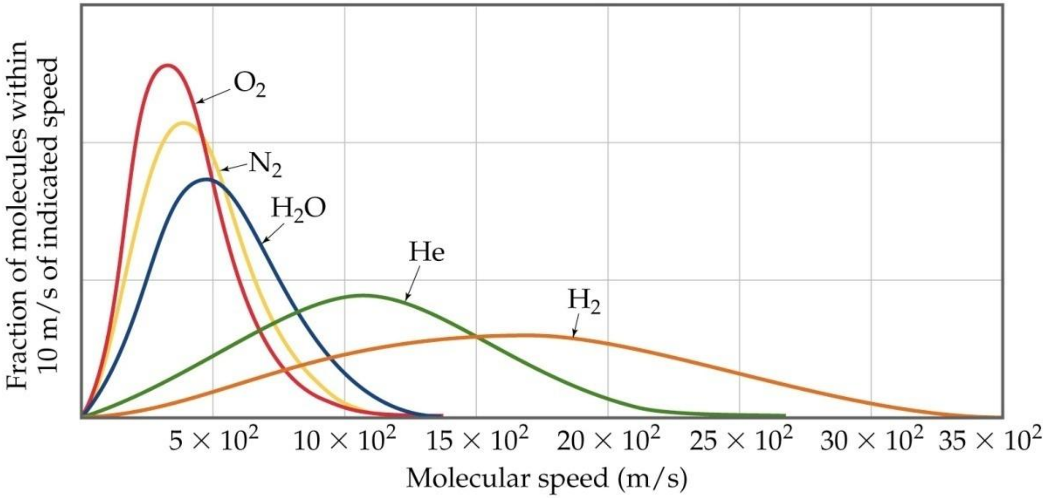


Kinetic Molecular Theory

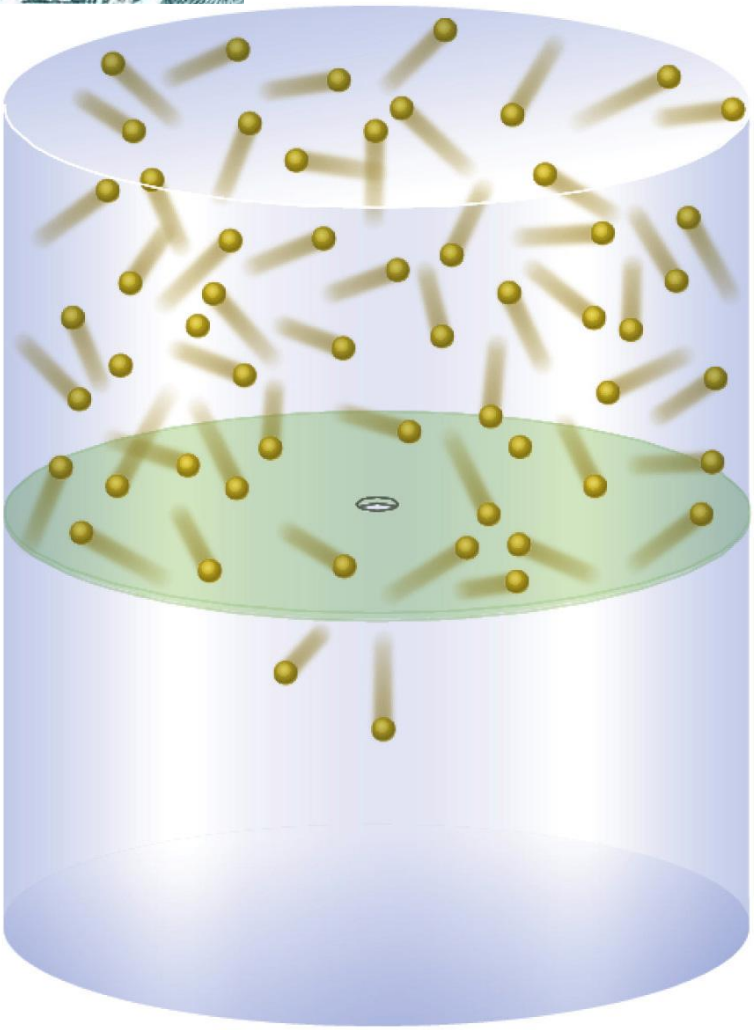
Molecular Effusion and Diffusion

- Consider two gases at the same temperature: the lighter gas has a higher rms than the heavier gas.
- Mathematically: $u = \sqrt{\frac{3RT}{\mathcal{M}}}$

The lower the molar mass, *M*, the higher the rms.

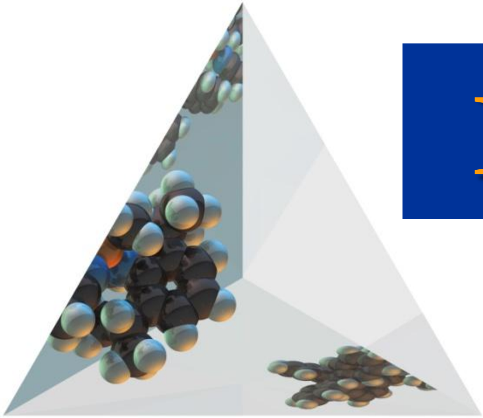


Kinetic Molecular Theory



Graham's Law of Effusion

- As kinetic energy increases, the velocity of the gas molecules increases.
- The rate of effusion can be quantified.



Kinetic Molecular Theory

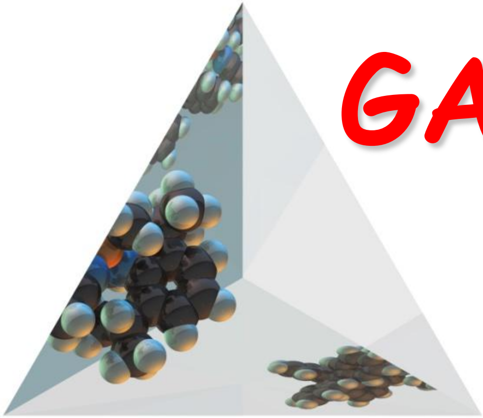
Graham's Law of Effusion

- Consider two gases with molar masses $\mathcal{M}_1$ and $\mathcal{M}_2$, the relative rate of effusion is given by:

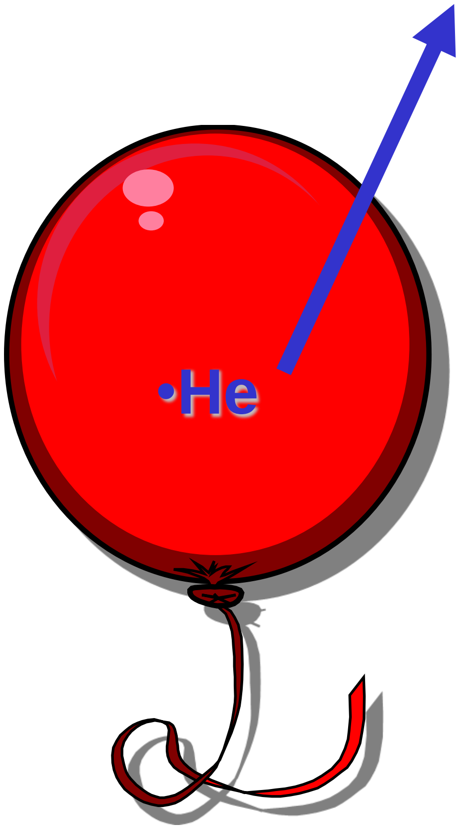
$$\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{3RT/\mathcal{M}_1}{3RT/\mathcal{M}_2}} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

- Only those molecules that hit the small hole will escape through it.
- Therefore, the higher the rms the more possibility of a gas molecule hitting the hole.



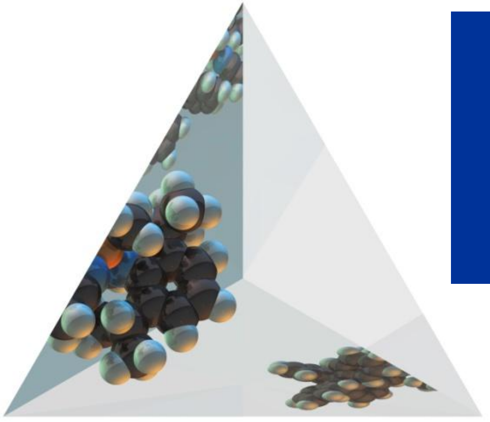
GAS DIFFUSION AND EFFUSION



Molecules effuse thru holes in a rubber balloon, for example, at a rate (= moles/time) that is

- proportional to T
- inversely proportional to M.

Therefore, He effuses more rapidly than O_2 at same T.



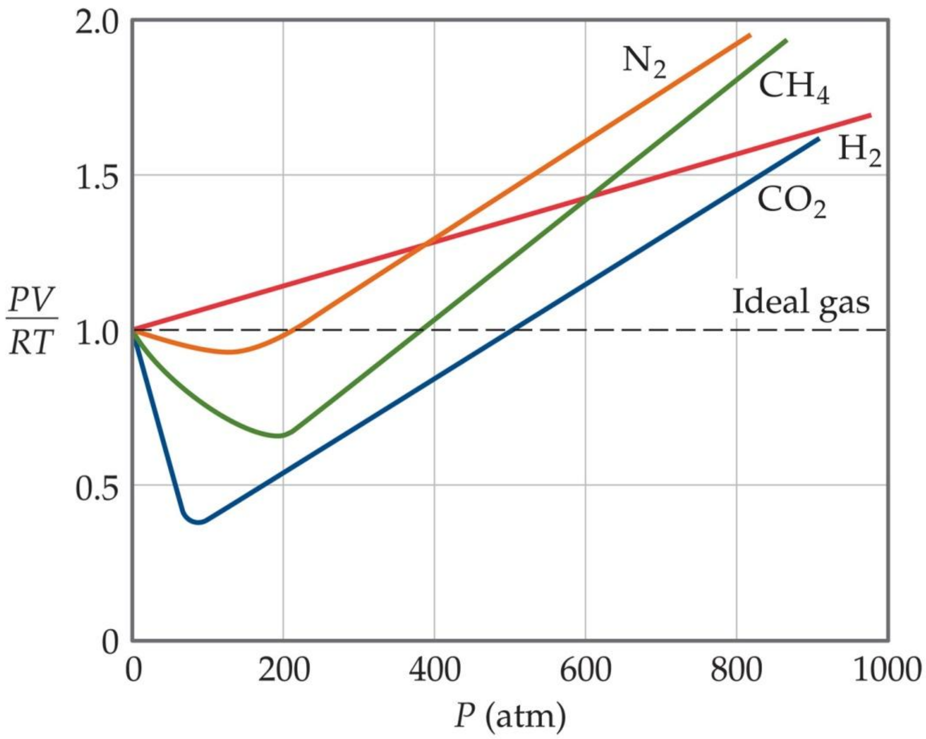
Real Gases: Deviations from Ideal Behavior

- From the ideal gas equation, we have

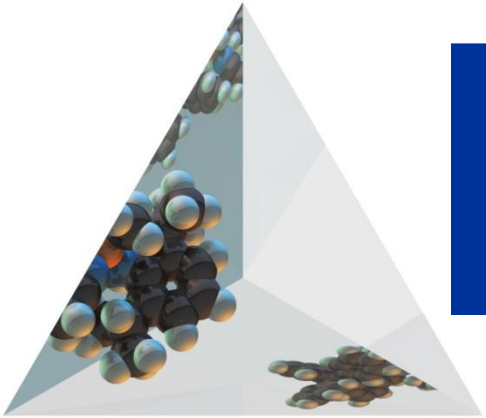
$$\frac{PV}{RT} = n$$

- For 1 mol of gas, $PV/RT = 1$ for all pressures.
- In a real gas, PV/RT varies from 1 significantly.
- The higher the pressure the more the deviation from ideal behavior.

Real Gases

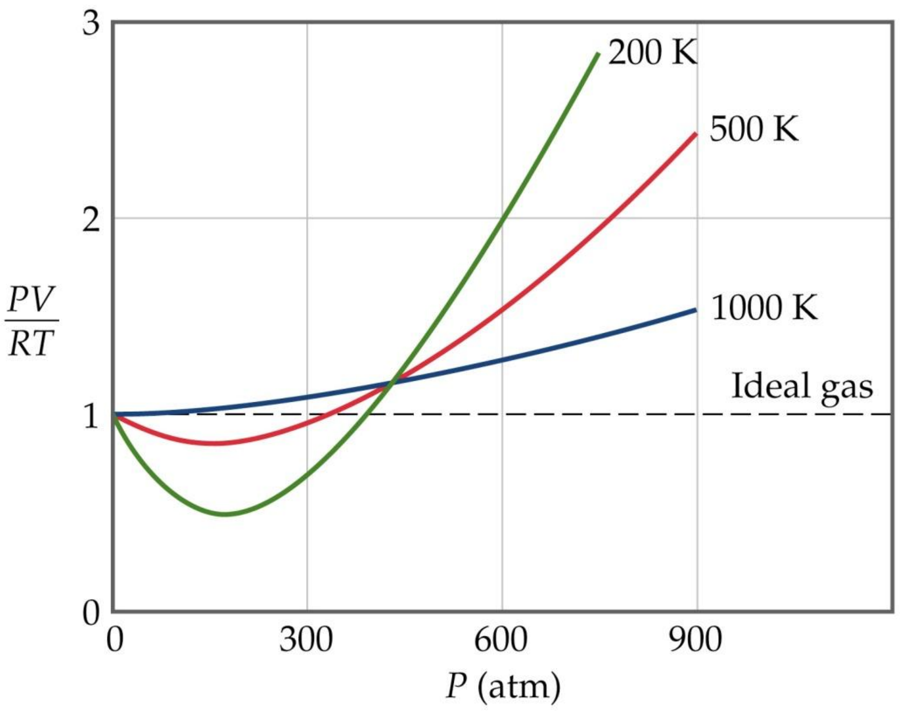


In the real world, the behavior of gases only conforms to the ideal-gas equation at relatively high temperature and low pressure.

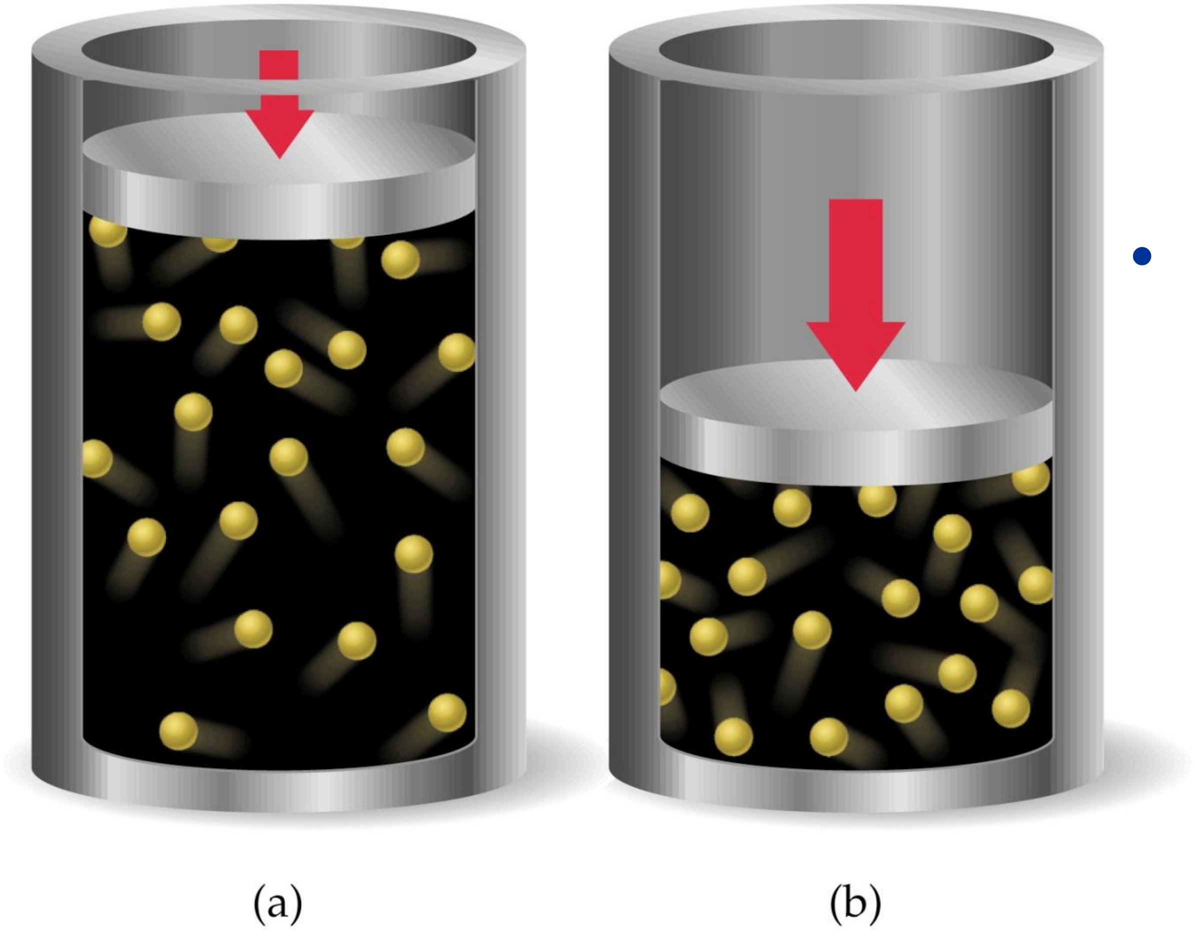


Real Gases: Deviations from Ideal Behavior

- The assumptions in kinetic molecular theory show where ideal gas behavior breaks down:
  - real molecules of a gas have finite volume;
  - molecules of a gas do attract each other.

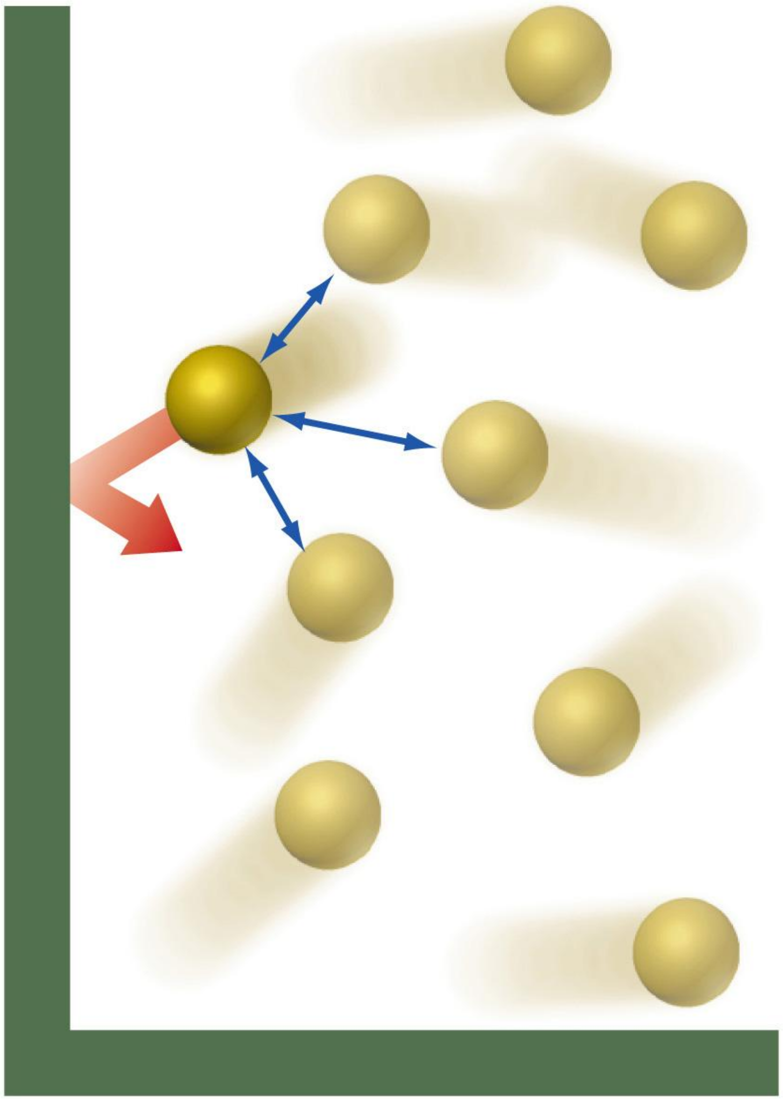


Real Gases: Deviations from Ideal Behavior



- As the gas molecules get closer together, the smaller the intermolecular distance.

Real Gases: Deviations from Ideal Behavior



As temperature increases, the gas molecules move faster and further apart.

Also, higher temperatures mean more energy available to break intermolecular forces.

- the higher the temperature, the more ideal the gas.

Corrections for Nonideal Behavior

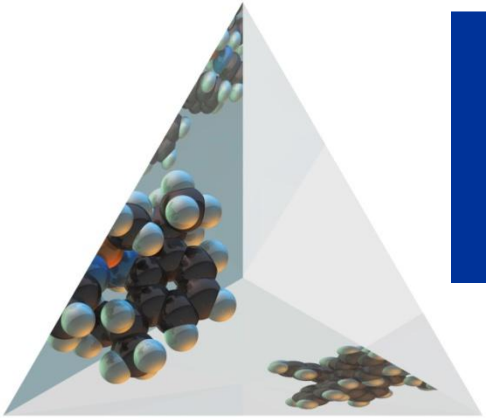
- The ideal-gas equation can be adjusted to take these deviations from ideal behavior into account.
- The corrected ideal-gas equation is known as the *van der Waals* equation.

The van der Waals Equation

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

a and *b* are empirical constants

Substance	a (L^2-atm/mol^2)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383



Real Gases: Deviations from Ideal Behavior

The van der Waals Equation

$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}$$

Corrects for molecular volume

Corrects for molecular attraction

- General form of the van der Waals equation:

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