The standard temperature and pressure for gases is:

\[ T = 273 \text{ K} \quad (0^\circ \text{C}) \]

\[ P = 1 \text{ atm} \]

\[ = 101.325 \text{ kPa} \]

\[ = 1.01325 \text{ bar} \]

At STP, 1 mol of any ideal gas occupies 22.4 L
A 5.00 L sample of Ar gas at **STP** is heated until its final pressure and temperature are 2.50 atm and 75.50°C. Calculate the new volume of the gas.

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}; \quad V_2 = ?
\]

\[
P_1 = 1 \text{ atm}, \quad V_1 = 5.00 \text{ L}, \quad T_1 = 273 \text{ K};
\]

\[
P_2 = 2.50 \text{ atm}, \quad T_2 = 273 + 75.50 = 348.5 \text{ K}
\]

\[
V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{(1.00 \text{ atm})(5.00 \text{ L})(348.5\text{K})}{(273 \text{ K})(2.50 \text{ atm})} = 2.55 \text{ L}
\]
Using PV = nRT

How many moles of NO₂ gas occupy a volume of 5.00 L at 50.00°C and 725 mmHg?

\[ PV = nRT; \quad n = \frac{PV}{RT} \]

\[ P = 725 \text{ mmHg} = 0.954 \text{ atm}; \quad V = 5.00 \text{ L}, \quad T = 273 + 50 = 323 \text{ K}; \]

\[ n = \frac{(0.954 \text{ atm})(5.00 \text{ L})}{(0.08206 \text{ Latm/molK})(323 \text{ K})} \]

\[ n = 0.18 \text{ moles} \]
Using \( PV = nRT \)

What mass of CO gas occupies a volume of 75.0 L at 35.00°C and 2.50 atm?

\[
P = 2.50 \text{ atm}; \quad V = 75.0 \text{ L}; \quad T = 273 + 35 = 308 \text{ K};
\]

\[
n = \frac{PV}{RT} = \frac{(2.5 \text{ atm})(75.0 \text{ L})}{(0.08206 \text{ Latm/molK})(308 \text{ K})} = 7.42 \text{ moles}
\]

\[
mw \text{ of CO} = 28.01 \text{ g/mol} \]

\[
\text{mass} = 7.42 \text{ moles} \times 28.01 \text{ g/mol} = 207.8 \text{ g}
\]
Gas Density

- We can use $PV = nRT$ to determine the density of gases.
- What are the units of density?

  mass/volume

- What does this suggest about gas density?

  *It will depend strongly on $T$, $P$, and the mass of the gas molecules.*

- Contrast with liquids and solids, whose densities depend somewhat on $T$, but *far* less on $P.$
Gas Density Example

What is the density of O₂ gas in g/L at 25°C and 0.850 atm?

Calculate # moles in 1 L, use MW of O₂ to get g, divide by V.

- P = 0.850 atm
- V = 1 L
- n = ?
- R = 0.08206 L·atm/mol·K
- T = 25°C + 273 = 298 K

\[
PV = nRT \quad \rightarrow \quad n = \frac{PV}{RT}
\]

\[
n = \frac{(0.850 \text{ atm})(1 \text{ L})}{(0.08206 \text{ L·atm/mol·K})(298 \text{ K})} = 0.0347 \text{ mol O}_2
\]

? g/L O₂ = 0.0347 mol O₂ \( \left( \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) \left( \frac{1}{1 \text{ L}} \right) = 1.11 \text{ g/L O}_2 \)

@ 25°C, 0.850 atm

Note: A value for gas density is meaningful only if accompanied by the T and P at which it was measured.
Gas Density and Molar Mass

We can develop an expression relating density and molecular weight using $PV = nRT$.

Substituting:

Therefore, the density of an ideal gas can be related to the $T$, $P$, $V$ and molecular weight AND we can use the density of a gas to determine its molar mass.
A 0.76 g sample of an unknown solid was vaporized in a 345-mL vessel. If the vapor has a pressure of 985 mmHg at 148°C, what is the molecular weight of the solid?

\[ P = 985 \text{ mmHg} = 1.296 \text{ atm} \]
\[ d = \frac{0.76 \text{ g}}{0.345 \text{ L}} = 2.20 \text{ g/L} \]
\[ R = 0.08206 \text{ L·atm/mol·K} \]
\[ T = 148°C = 421 \text{ K} \]

\[ MW = \frac{dRT}{P} = \frac{(2.20 \frac{\text{g}}{\text{L}})(0.08206 \frac{\text{L·atm}}{\text{mol·K}})(421 \text{ K})}{1.296 \text{ atm}} = 58.65 \text{ g/mol} \]
Chemical Equations and Calculations

**Reactants**

- **Atoms (Molecules)**
  - Avogadro's Number
  - $6.022 \times 10^{23} \text{ mol}^{-1}$

- **Moles**

**Products**

- **Mass**
  - Molar Mass
  - g/mol

**Solutions**

- **Molarity**
  - moles / L

**Gases**

- **PV = nRT**
Gas Stoichiometry Example

What volume of \( \text{N}_2(\text{g}) \) is produced when 70.0 g NaN\(_3\) is decomposed at a pressure of 735 mmHg and at 26\(^\circ\)C?

\[
2 \text{ NaN}_3(s) \rightarrow 2 \text{ Na}(l) + 3 \text{ N}_2(g)
\]

70.0 g  
64.99 g/mol  

? L  
735 mmHg  
26\(^\circ\)C  
28.02 g/mol

1. **The Stoichiometry Part**: mass NaN\(_3\) \(\rightarrow\) mol NaN\(_3\) \(\rightarrow\) mol N\(_2\)

2. **The Gas Law Part**: mol N\(_2\), P, T \(\rightarrow\) V of N\(_2\)(g)
Example: The Stoichiometry Part

What volume of $N_2(g)$ is produced when 70.0 g NaN$_3$ is decomposed at a pressure of 735 mmHg and at 26°C?

$$2 \text{ NaN}_3(s) \rightarrow 2 \text{ Na}(l) + 3 \text{ N}_2(g)$$

\[ ? \text{ mol } N_2 = 70.0 \text{ g } \text{NaN}_3 \left( \frac{1 \text{ mol } \text{NaN}_3}{64.99 \text{ g } \text{NaN}_3} \right) \left( \frac{3 \text{ mol } \text{N}_2}{2 \text{ mol } \text{NaN}_3} \right) \]

\[ = 1.616 \text{ mol } N_2 \]
Example: The Gas Law Part

What volume of \( \text{N}_2(\text{g}) \) is produced when 70.0 g \( \text{NaN}_3 \) is decomposed? (\( P = 735 \text{ mmHg}, \ T = 26^\circ\text{C} \))

\[
P = 735 \text{ mmHg} \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.9671 \text{ atm}
\]

\( V = ? \text{ L} \)

\( n = 1.616 \text{ mol N}_2 \)

\( R = 0.08206 \text{ L.atm/mol.K} \)

\( T = 26^\circ\text{C} + 273 = 299 \text{ K} \)

\[
\frac{PV}{nRT} = V = \frac{nRT}{P}
\]

\[
V = \frac{(1.616 \text{ mol N}_2)(0.08206 \frac{\text{L atm}}{\text{mol K}})(299 \text{ K})}{0.9671 \text{ atm}}
\]

\[
V = 40.999 \text{ L} = 41.0 \text{ L}
\]
Example

The active agent in many hair bleaches is hydrogen peroxide, $\text{H}_2\text{O}_2$. The amount of $\text{H}_2\text{O}_2$ present can be determined by titration with a standard permanganate solution:

$$2 \text{MnO}_4^- (\text{aq}) + 5 \text{H}_2\text{O}_2 (\text{aq}) + 6 \text{H}^+ (\text{aq}) \rightarrow 5 \text{O}_2 (\text{g}) + 2 \text{Mn}^{+2} (\text{aq}) + 8 \text{H}_2\text{O} (\text{l})$$

Calculate the molarity of hydrogen peroxide if 28.75 mL of hydrogen peroxide produced 695 mL of oxygen gas at 0.950 atm and 315 K?

$$P = 0.950 \text{ atm}; \ V = 0.695 \text{ L}; \ T = 315 \text{ K}$$
$$n = \frac{PV}{RT} = \frac{(0.950 \text{ atm})(0.695 \text{ L})}{(0.08206 \text{ Latm/molK})(315 \text{ K})}$$
$$n = 0.0255 \text{ moles O}_2 = \text{ moles H}_2\text{O}_2 \ (1:1 \text{ ratio})$$
$$\text{Molarity} = \frac{0.0255 \text{ moles H}_2\text{O}_2}{0.02875 \text{ L}} = 0.888 \text{ M}$$
Example

An unknown gas having a mass of 6.150 g occupies a volume of 5.00 L at 874 torr and 23.50°C. Calculate the molar mass of the unknown gas.

\[ P = 874 \text{ torr} = 1.15 \text{ atm}; \quad V = 5.00 \text{ L}; \quad T = 23.50 + 273.15 = 296.65 \text{ K} \]
\[ n = \frac{PV}{RT} = \frac{(1.15 \text{ atm})(5.00 \text{ L})}{(0.08206 \text{ Latm/molK})(296.65 \text{ K})} \]
\[ n = 0.236 \text{ moles gas} \]
\[ MW = \frac{6.150 \text{ g gas}}{0.236 \text{ moles gas}} = 26.04 \text{ g/mol} \]

\[ \text{C}_2\text{H}_2 \text{ (acetylene)}? \]
Dalton’s Law of Partial Pressures

Recall that according to the ideal gas law, gas molecules are non-interacting point particles.

Increasing the number of point particles increases the pressure by an amount that is proportional to the number of particles.

For a mixture of ideal gases in a container:

\[ \text{total pressure} = \text{the sum of the individual gas pressures}. \]
Dalton’s Law of Partial Pressures

Say we have a container with some amount of three different gases inside, at a certain $T$ and $P$.

\[ n_{\text{total}} = n_1 + n_2 + n_3 \]

Dalton’s Law says that the total pressure exerted by the three gases is the sum of the individual pressures.

\[ P_{\text{total}} = P_1 + P_2 + P_3 \]

\[ P_{\text{total}} = \frac{n_1RT}{V} + \frac{n_2RT}{V} + \frac{n_3RT}{V} \]

\[ P_{\text{total}}V = n_1RT + n_2RT + n_3RT \]

\[ P_{\text{total}}V = (n_1 + n_2 + n_3)RT \]

\[ P_{\text{total}}V = n_{\text{total}}RT \]
Partial Pressures Example

Mixtures of He and O\textsubscript{2} are used in scuba tanks to help prevent “the bends.” For a particular dive, 12 L of O\textsubscript{2} at 25\degree C and 1.0 atm was pumped along with 46 L of He at 25\degree C and 1.0 atm into a 5.0-L tank. What is the partial pressure of each gas? What is the total pressure?

1. Find the number of moles of each gas that were delivered to the tank.

2. Find the partial pressure of each gas in the tank.

3. Add them up!
Mixtures of He and O$_2$ are used in scuba tanks to help prevent “the bends.” For a particular dive, 12 L of O$_2$ at 25°C and 1.0 atm was pumped along with 46 L of He at 25°C and 1.0 atm into a 5.0-L tank. What is the partial pressure of each gas? What is the total pressure?

<table>
<thead>
<tr>
<th>O$_2$ Data:</th>
<th>He Data:</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>P =</td>
<td>P =</td>
<td></td>
</tr>
<tr>
<td>V =</td>
<td>V =</td>
<td></td>
</tr>
<tr>
<td>n =</td>
<td>n =</td>
<td></td>
</tr>
<tr>
<td>R =</td>
<td>R =</td>
<td></td>
</tr>
<tr>
<td>T =</td>
<td>T =</td>
<td></td>
</tr>
</tbody>
</table>

\[
n_{O_2} = \frac{(1.0 \text{ atm})(12 \text{ L})}{(0.08206 \text{ L atm/mol K})(298 \text{ K})} = 0.49 \text{ mol O}_2
\]

\[
n_{\text{He}} = \frac{(1.0 \text{ atm})(46 \text{ L})}{(0.08206 \text{ L atm/mol K})(298 \text{ K})} = 1.9 \text{ mol He}
\]
Using the moles of each gas, the temperature, and volume of the tank we can now calculate the partial pressure of each gas, then add them to get the total pressure.

\[ P = \frac{nRT}{V} \]

\[
\begin{align*}
P_{O_2} &= \frac{(0.49 \text{ mol } O_2) \left(0.08206 \text{ L atm/mol K}\right)(298 \text{ K})}{5.0 \text{ L}} = 2.4 \text{ atm} \\
P_{\text{He}} &= \frac{(1.9 \text{ mol He}) \left(0.08206 \text{ L atm/mol K}\right)(298 \text{ K})}{5.0 \text{ L}} = 9.3 \text{ atm}
\end{align*}
\]

\[ P_{\text{total}} = 11.7 \text{ atm} \]

Partial pressures of \(O_2\) and He in the 5.0 L scuba tank
Mole Fraction and Partial Pressure

• In the last example, we determined the total pressure by adding the partial pressures.

• We could have also added the moles of each gas, and determined a total pressure.

• These two approaches suggest that a relationship exists between the moles of each gas and the total pressure.
Mole Fraction and Partial Pressure

Mole Fraction ($\chi$): ratio of the number of moles of a component in a mixture to the total number of moles in the mixture.

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

$$\chi_i = \frac{n_i}{n_{\text{total}}} = \frac{P_i \left( \frac{V}{RT} \right)}{P_{\text{total}} \left( \frac{V}{RT} \right)} = \frac{P_i}{P_{\text{total}}} \implies P_i = \chi_i P_{\text{total}}$$

Dalton’s Law

The fraction of moles of a certain gas in a mixture is equal to the ratio of its partial pressure to the total pressure of the mixture.
Mixtures of He and O$_2$ are used in scuba tanks to help prevent “the bends.” For a particular dive, 12 L of O$_2$ at 25°C and 1.0 atm was pumped along with 46 L of He at 25°C and 1.0 atm into a 5.0-L tank. What is the partial pressure of each gas? What is the total pressure?

\[ n_{O_2} = 0.49 \text{ mol} \quad n_{\text{total}} = 2.39 \text{ mol} \]
\[ n_{\text{He}} = 1.9 \text{ mol} \quad P_{\text{total}} = 11.7 \text{ atm} \]

\[ \chi_{O_2} = \frac{0.49 \text{ mol O}_2}{2.39 \text{ mol gas}} = 0.205 \]
\[ \chi_{\text{He}} = \frac{1.9 \text{ mol He}}{2.39 \text{ mol gas}} = 0.795 \]

\[ P_{O_2} = 0.205(11.7 \text{ atm}) = 2.39 \text{ atm} \]
\[ P_{\text{He}} = 0.795(11.7 \text{ atm}) = 9.28 \text{ atm} \]

Partial pressures of O$_2$ and He in the 5.0 L scuba tank
Example

A mixture of gases contains 4.465 mol of neon, 0.741 mol of argon, and 2.154 mol of xenon. Calculate the partial pressures of all the gases if the total pressure is 2.00 atm at a given temperature.

\[
\begin{align*}
n_{\text{Ne}} &= 4.465 \text{ mol} \\
n_{\text{Ar}} &= 0.741 \text{ mol} \\
n_{\text{Xe}} &= 2.154 \text{ mol} \\
n_{\text{Total}} &= 7.36 \text{ mol}
\end{align*}
\]

\[
\begin{align*}
\chi_{\text{Ne}} &= \frac{4.465 \text{ mol}}{7.36 \text{ mol}} = 0.607 \\
\chi_{\text{Ar}} &= \frac{0.741 \text{ mol}}{7.36 \text{ mol}} = 0.101 \\
\chi_{\text{Xe}} &= \frac{2.154 \text{ mol}}{7.36 \text{ mol}} = 0.293
\end{align*}
\]

\[
\begin{align*}
P_{\text{Ne}} &= 0.607 \times (2.00 \text{ atm}) = 1.213 \text{ atm} \\
P_{\text{Ar}} &= 0.101 \times (2.00 \text{ atm}) = 0.201 \text{ atm} \\
P_{\text{Xe}} &= 0.293 \times (2.00 \text{ atm}) = 0.586 \text{ atm}
\end{align*}
\]
Kinetic Molecular Theory (KMT)

- The gas laws of Boyle, Charles, and Avogadro are *empirical*, meaning they are based on observation of a macroscopic property.

- These laws offer a general description of behavior based on many experiments.

- The empirical gas laws can tell you what happens to an ideal gas under certain conditions, but not why it happens.

- KMT is a theoretical, *molecular-level model* of ideal gases, which can be used to predict the macroscopic behavior of a gaseous system.

Postulates of KMT

Gas particles are so small that their volume is negligible.

Gas particles are in constant, random motion. Gas molecules constantly collide with each other and with the container walls. The collisions of the particles with the container walls are the cause of the pressure exerted by the gas. Collisions are elastic.

The particles are assumed to exert no forces on each other; they neither attract or repel their neighbors.

This motion is associated with an average kinetic energy that is directly proportional to the Kelvin temperature of the gas.
KMT: Central Points

• The main ideas you should take from KMT are that we can describe temperature and pressure from a molecular perspective.

• **Pressure:** arises from molecules banging into the container walls.

• **Temperature:** is directly related to the kinetic energy of the gas molecules. The more KE they have, the greater their temperature.
Let’s consider the average KE per molecule and see how it determines molecular speed.

Note: $T$ is measured in *Kelvin*; $R = 8.3145 \text{ J/mol} \cdot \text{K}$; ($J = \text{kg} \cdot \text{m}^2/\text{s}^2$)

Average KE per molecule: \[ \frac{1}{2} mu^2 \]

Where $u$ is an average of molecular velocity, and $m$ is the mass of one molecule.

The root-mean-square speed: \[ u_{\text{rms}} = \sqrt{\frac{3RT}{MM}} \]

We are apportioning the total KE in the mole of gas among all the molecules in an average fashion.

$u_{\text{rms}}$ is the speed of a molecule that has the average KE.

$u_{\text{rms}}$ gives us a formal connection between average gas speed, $T$, and $M$. 

Distribution of Molecular Speeds

“Maxwell-Boltzmann” curve (a statistical distribution)

This plot represents the fraction of gas molecules in a sample that are traveling at a given velocity.

- $u_m$ – most probable speed
- $u_{avg}$ – average speed
- $u_{rms}$ – the speed of a molecule with the average molecular kinetic energy

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

Increased T $\rightarrow$ increased average KE $\rightarrow$ increased $u_{rms}$

Increased M $\rightarrow$ decreased $u_{rms}$

The higher the molar mass of a particle, the slower the particle moves.
Increased $T \rightarrow \text{increased average KE} \rightarrow \text{increased } u_{\text{rms}}$

- Maximum of curve shifts to higher $u$, and distribution spreads out.
- Distribution of speeds will be “shorter” and “fatter” at higher temperatures.

Increased $M \rightarrow \text{decreased } u_{\text{rms}}$

- Heavier molecules have lower average speed than lighter molecules at a given temperature.
- Distribution of speeds for heavier gases will be “taller” and “skinnier” than for lighter molecules.

\[ u_{\text{rms}} = \sqrt{\frac{3RT}{M}} \]

**NOTE:**
There are always some molecules with low velocity in a Boltzmann distribution!!
Molecular Speed (cont.)

• Let’s determine $u_{rms}$ for N\textsubscript{2} at 298 K.

\[
\begin{align*}
   u_{rms} &= \sqrt{\frac{3RT}{M}} = \left(\frac{3 \times (8.314 \text{ J mol}^{-1} \text{ K}^{-1}) \times (298 \text{ K})}{(0.0280 \text{ kg mol}^{-1})}\right)^{\frac{1}{2}} = 515 \text{ m s}^{-1}
\end{align*}
\]

• For a sense of scale, this is on the order of the speed of sound (~320 m s\textsuperscript{-1})….which isn’t a coincidence.

Smaller mass = greater speed
Comparison of $u_{\text{rms}}$ for He and $N_2$

At 25°C, which gas will travel faster, He or $N_2$?

- $M(\text{He}) = 4.0$ g/mol
- $M(N_2) = 28$ g/mol

A car travelling at 60 mph,

- $u_{\text{car}} = 26.8$ m/s

*If gases travel so fast, why does it take so long for you to smell perfume from across the room?*
Example

Match each of the following gases at the same temperature with its distribution curve in the figure at the left: $\text{N}_2\text{O}$, $\text{Kr}$, and $\text{H}_2$.

Kr – highest MW, lowest rms speed

$\text{N}_2\text{O}$ – mid-range MW, mid-range rms speed

$\text{H}_2$ – lowest MW, highest rms speed
Diffusion

- Gas molecules travel in a straight line only until they collide with a container wall or another gas molecule.

- Gas molecules do not have an uninterrupted path in front of them.

- They are constantly colliding with other gas molecules.

- Rate of diffusion is proportional to $u_{rms}$. So lighter particles will have a higher rate of diffusion, and vice versa.
Diffusion is the process of mixing gases. In a closed container, diffusion will eventually lead to a homogeneous mixture. http://www.youtube.com/watch?v=H7QsDs8ZRMI
Diffusion Examples

• Circle the pair of gases in each set below that diffuse faster.

  a. Ne & F₂ or Ar & Cl₂

  b. Kr & Ar or O₂ & Cl₂
Nitrous oxide, $\text{N}_2\text{O}$, also known as laughing gas, is a colorless gas and has been used as a weak anesthetic. Hydrogen cyanide, HCN, is a poisonous, colorless gas that can cause a quick death. If both of these gases were accidentally released at the same time in the front of a theater full of moviegoers, would the people die laughing?

**MW of $\text{N}_2\text{O} = 44.02$ g/mol**
**MW of HCN = 27.029 g/mol**

HCN is lighter, so faster. **No.**
Effusion is a special case of diffusion, which exploits the difference in velocities of lighter gas molecules. This process was used during the Manhattan Project to separate $^{235}\text{U}$ and $^{238}\text{U}$ isotopes.
• Effusion is dependent on molecular speed. The molecular speed is in turn inversely dependent on the atomic or molar mass. Recall:

\[ u_{rms} = \sqrt{\frac{3RT}{M}} \]

• Graham’s Law of Effusion: Rates of effusion are inversely dependent on the square root of the mass of each gas:

\[ \frac{(\text{Rate of effusion})_1}{(\text{Rate of effusion})_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}} \]
Diffusion Reactions

- By similar arguments, the distance a molecule travels is inversely proportional to mass:

\[
\frac{(\text{distance})_1}{(\text{distance})_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}
\]

- Example: reaction of \( \text{NH}_3 \) with \( \text{HCl} \)

\[
\frac{(distance)_{\text{NH}_3}}{(distance)_{\text{HCl}}} = \frac{\sqrt{M_{\text{HCl}}}}{\sqrt{M_{\text{NH}_3}}} = \sqrt{\frac{36.5}{17}} = 1.5
\]

Exp. 1.3
Real Gases

• Generally speaking, there is no such thing as an “Ideal Gas.”

• There are conditions under which a gas will behave ideally...
  – low P
  – moderate to high T

• van der Waals developed some corrections to the Ideal Gas law, based on a molecular picture, to explain these observed deviations.
Real Gases (cont.)

- At high $P$, the volume of the individual gas molecules becomes non-negligible.

- Macroscopic gas is compressible, individual gas molecules are not.

- Under high $P$ conditions, the space available for a gas molecule to move is decreased by its neighbors, so the volume of the system is reduced relative to the ideal case.

$$V_{\text{eff}} = V_{\text{ideal}} - nb$$

- Number of moles of gas

- Empirical constant... different for each gas; increases with size of molecule.
Real Gases (cont.)

- In the Ideal Gas theory, we assume that gas molecules do not interact.
- But under **high P**, gas molecules get very close to each other and interact.
- Further, at **low T** the molecular speed drops also increasing the importance of intermolecular interactions.

*Under high P and/or low T conditions, the molecules don’t collide with the container as frequently, so the pressure of the system is reduced relative to the ideal case.*

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

Empirical constant… different for each gas; increases with increasing intermolecular attraction

Concentration of the gas
Real Gases (cont.)

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

- \(b\) generally increases with the size of the molecule
- \(a\) generally increases with the strength of intermolecular forces.

vdW equation corrects two major flaws in ideal gas theory:
- Gas molecules have finite volume which becomes important at high \(P\).
- Gas molecules have non-trivial attractions that become important at low \(T\) and high \(P\).
<table>
<thead>
<tr>
<th>Substance</th>
<th>$a , (L^2 \cdot atm/mol^2)$</th>
<th>$b , (L/mol)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>0.0341</td>
<td>0.02370</td>
</tr>
<tr>
<td>Ar</td>
<td>1.34</td>
<td>0.0322</td>
</tr>
<tr>
<td>$H_2$</td>
<td>0.244</td>
<td>0.0266</td>
</tr>
<tr>
<td>$N_2$</td>
<td>1.39</td>
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</tr>
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<td>$O_2$</td>
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<td>$CH_4$</td>
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<td>$CO_2$</td>
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<td>$H_2O$</td>
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<tr>
<td>NO</td>
<td>1.34</td>
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<td>$NO_2$</td>
<td>5.28</td>
<td>0.04424</td>
</tr>
<tr>
<td>HCl</td>
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<td>$SO_2$</td>
<td>6.71</td>
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