Chapter 5: Gases

• Properties of Gases
• Gas Laws
• Mixture of Gases
• Kinetic-Molecular Theory
• Real Gases

Properties of Gases

• Gases are fluids. However unlike liquids the atoms or molecules are far apart and thus there is little interaction between molecules.
• Expand to fill the volume of any container.
• Have much lower densities than solids or liquids.
• Have highly variable densities, depending on conditions.
• Mix with one another readily and thoroughly (i.e. they are miscible).
• Change volume dramatically with changing temperature.
• Collision of gas particles on the container’s walls exerts a pressure.
• Gases are compressible (their volume decreases as the pressure increases).
• The pressure of a gas in a closed container increases as the number of gas molecules increases.

Properties of Gases: Gas Pressure

• Gas Pressure
\[ P \text{ (Pa)} = \frac{\text{Force (N)}}{\text{Area (m}^2\text{)}} \]

• Liquid Pressure
\[ P = g \cdot h \cdot d \]

Pressure

• Pressure results from molecular collisions between gas molecules and container walls.
  – Each collision imparts a small amount of force.
  – Summation of the forces of all molecular collisions produces the macroscopic property of pressure.

Measuring Pressure: Barometer

• A barometer is used to measure atmospheric pressure.
• Mercury flows out of the tube until the pressure of the column of mercury standing on the surface of the mercury in the dish is equal to the pressure of the air on the rest of the surface of the mercury in the dish.

• Units of Pressure
  – 1 torr = 1 mm Hg
  – 1 atm = 760 torr (exactly)
  – 1 atm = 101,325 Pa (exactly)

Standard Atmospheric Pressure
1.00 atm
760 mm Hg, 760 torr
101.325 kPa
1.01325 bar
14.7 lb in \(^2\)

Manometers

Device used for measuring the pressure of a gas in a container.
**Pressure Example**

- How high can water be “raised” by 1 atm of pressure. Mercury has a density of 13.60 g/mL at 25°C. Water has a density of 0.997 g/mL at 25°C.

**Charles’s Law**

- Jacques Charles studied relationship between volume and temperature at constant pressure and amount of gas.
  - Plots of $V$ versus $T$ for different gas samples converged to the same temperature at zero volume.
  - Basis of the Kelvin temperature scale.

\[ V \propto T \text{ (K)} \]
\[ V = bT \]

**The Effects of Increasing the Temperature of a Sample of Gas at Constant Pressure**

![Image showing temperature increase](image1.png)

**Volume-Temperature Examples**

A certain mass of gas at constant pressure has a volume of 22.0 ml at 25.0°C. What will the volume be at 75.0°C?

To what temperature (in °C) must a gas at 20.0°C be raised to double its volume?

**Boyle’s Law**

- Boyle 1662
- Constant $T$ and $n$

\[ PV = \text{constant} \]

Volume is inversely proportional to pressure

\[ P_1 V_1 = P_2 V_2 \]

**Increased Pressure due to Decreased Volume**

![Image showing increased pressure](image2.png)
Pressure-Volume Examples #1

A certain mass of gas at constant temperature has a volume of 45.0 ml at 755 mm Hg. What will the volume be at 375 mm Hg?

A certain mass of gas at constant temperature has a volume of 326 ml at 0.950 atm. At what pressure will the volume be 125 ml?

Pressure-Volume Example #2

Relating Gas Volume and Pressure – Boyle’s Law.

Filling a tank from another tank

STP Conditions

• Gas properties depend on conditions.
• Standard temperature and pressure, STP, for a gas is 0°C (273.15 K) and 1 atm.
  – For one mole of gas at STP, the standard molar volume is 22.41 L (calculated using ideal gas law)
  – This number provides a conversion factor for stoichiometric problems that include gases, provided the STP conditions are maintained.

Avogadro’s Law

• Gay-Lussac 1808
  – Small volumes of gases react in the ratio of small whole numbers.
• Avogadro 1811
  – Equal volumes of gases have equal numbers of molecules and
  – Gas molecules may break up when they react.

Formation of Water
Avogadro’s Law

At an a fixed temperature and pressure:

\[ V \propto n \quad \text{or} \quad V = c \, n \]

At STP

1 mol gas = 22.4 L gas

Molar Volumes for Various Gases at 0°C and 1 atm

<table>
<thead>
<tr>
<th>Gas</th>
<th>Molar Volume (L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ideal</td>
<td>22.414</td>
</tr>
<tr>
<td>O₂</td>
<td>22.397</td>
</tr>
<tr>
<td>N₂</td>
<td>22.402</td>
</tr>
<tr>
<td>H₂</td>
<td>22.433</td>
</tr>
<tr>
<td>He</td>
<td>22.434</td>
</tr>
<tr>
<td>CO₂</td>
<td>22.260</td>
</tr>
<tr>
<td>NH₃</td>
<td>22.079</td>
</tr>
</tbody>
</table>

Note: A gas’s molar volume can be higher or lower than an ideal gas

Increased Volume due to Increased Moles of Gas at Constant Temperature and Pressure

Combining the Gas Laws: The Ideal Gas Equation and the General Gas Equation

• Boyle’s law \( V \propto 1/P \)
• Charles’s law \( V \propto T \)
• Avogadro’s law \( V \propto n \)

\[
V = R \left( \frac{nT}{P} \right) \quad \text{\( R \) is the ideal gas constant}
\]

\[
PV = nRT
\]

Units and the Ideal Gas Law

• Temperature must be expressed in Kelvin for all gas calculations!

• The unit for moles is always mol.

• The units for measuring pressure and volume can vary. In gas calculations, these units must agree with those of the gas constant
  - \( R = 8.31446 \, \text{J mol}^{-1} \, \text{K}^{-1} = 8.31446 \, \text{Pa m}^3 \, \text{mol}^{-1} \, \text{K}^{-1} \)
  - \( R = 0.0820575 \, \text{L atm} \, \text{mol}^{-1} \, \text{K}^{-1} \)

The General Gas Equation

\[
R = \frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}
\]

If we hold the amount and volume constant:

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]
**Increased Pressure due to Increased Temperature**

Temperature is increased

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**Ideal Gas Examples**

1. Calculate the volume occupied by 2.68 g of N₂O at 19.0°C and 755 mm Hg.

2. How many molecules are left in a 1.00 L vessel which has been evacuated to a pressure of 0.0100 torr at 25°C.

3. A sample of phosgene gas weighing 0.152 g has a volume of 48.3 mL at 100.0°C and 740. Torr. Calculate the molar mass of phosgene.

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**Mixtures of Gases**

- Gas laws apply to mixtures of gases.
  - Gas laws do not depend on identity of gases.
  - Pressure due to total moles gas present.

- Simplest approach is to use \( n_{\text{total}} \).
  - For an ideal gas all gases behave the same.

- Partial pressure
  - Each component of a gas mixture exerts a pressure that it would exert if it were in the container alone.
  
  \[
  P_{\text{total}} = P_1 + P_2 + P_3 + \ldots = \sum P_i
  \]

---

**Dalton’s Law of Partial Pressure**

A gas in a mixture behaves as if it was isolated in the same volume

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**Partial Pressure**

- Dalton’s Law can be expressed in terms of mole fraction.
  - Mole fraction \( (X) \) for a gas in a gas mixture is the moles of the gas \( (n) \) divided by the total moles gas present.
  - The partial pressure of each gas is related to its mole fraction.

\[
\frac{P_i}{P_{\text{total}}} = \frac{n_i}{n_{\text{total}}} = X_i \quad \text{or} \quad P_i = X_i P_{\text{total}}
\]
Partial Pressure Example

- A breathing mixture of 2.0\% oxygen and 98.0\% helium is sometimes used by divers. Calculate the partial pressure of oxygen in the mixture at a depth of 100. m. Assume the water has a density of 0.997 g/mL and an atmospheric pressure of 1.00 atm.

Stoichiometry of Reactions Involving Gases

- For reactions involving gases, the ideal gas law is used to determine moles of gas involved in the reaction.
  - Use mole ratios (stoichiometry)
  - Connect number of moles of a gas to its temperature, pressure, or volume with ideal gas law

Gases Reactions Examples

Using the Ideal gas Equation in Reaction Stoichiometry Calculations.

1. The decomposition of sodium azide, NaN\(_3\), at high temperatures produces N\(_2\)(g). What volume of N\(_2\)(g), measured at 735 mm Hg and 26°C, is produced when 70.0 g NaN\(_3\) is decomposed.
   \[ 2 \text{NaN}_3(s) \rightarrow 2 \text{Na}(l) + 3 \text{N}_2(g) \]

2. NH\(_4\)NO\(_3\)(s) decomposes on heating to form N\(_2\)O and water as shown below:
   \[ \text{NH}_4\text{NO}_3(s) \rightarrow \text{N}_2\text{O}(g) + 2\text{H}_2\text{O}(g) \]

   A 2.00 g sample of NH\(_4\)NO\(_3\)(s) was placed in an evacuated 800.0 ml container and completely decomposed by heating to 155°C.
   Calculate the pressure when the contents of the container were at (a) 155°C and (b) cooled to 25°C.

Kinetic-Molecular Theory and Ideal versus Real Gases

- In many important practical settings, gases do not always behave ideally, especially at very high pressure and/or very low temperature.
  - Nonideal gas behavior can be explained using Kinetic Molecular Theory.
    - Provides connections between observed macroscopic properties of gases, the gas law equation, and the behavior of gas molecules on a microscopic scale.

Postulates of the Kinetic Molecular Theory

- Gases are made up of large collections of particles, which are in constant, random motion.
- Gas particles are infinitely small and occupy negligible volume.
- Gas particles move in straight lines except when they collide with other particles or with the container walls. These collisions are elastic, so kinetic energy of particles is conserved.
- Particles interact with each other only when collisions occur.
- The average kinetic energy of a gas is proportional to the absolute temperature of the gas but does not depend upon the identity of the gas.
Molecular View of Pressure

Pressure depends on:

• Random collisions of the gas particles with the walls create the pressure exerted by the gas
• Collision frequency of the gas particles with the walls
• The momentum of the gas particles

Collision with a Wall

• Consider the collisions of particles with a wall in the xy plane of a Cartesian space (z=0).
• Only those particles above the wall with v_z<0 will strike the wall.
• Only those particles with v_z<0 closer that |v_z|Δt will strike the wall in the time interval Δt.
• Let A be the cross-sectional area and n* be the molecules per unit volume then the number of collisions with the wall with |v_z| is :
  \# collisions have |v_z| = n*A |v_z| Δt f(v_z)
  where f(v_z) is the fraction of particles that have velocity |v_z|
• The total collisions are calculated by summing over all negative velocities
  \[
  \text{total collisions} = n* A \Delta t \sum_{v_z} |v_z| f(v_z)
  \]

Pressure of a Gas

• Pressure is force (F) per unit area (A); and F is mass times acceleration
  \[
  F = \frac{dp}{dt} = \frac{dp}{dt}
  \]
  where p = mv is the momentum
• A particle heading toward the wall will have momentum before collision:
  \[
  p_x = mv_x, \quad p_y = mv_y, \quad p_z = mv_z
  \]
• A particle heading toward the wall will have momentum after collision:
  \[
  p_x = mv_x, \quad p_y = mv_y, \quad p_z = -mv_z
  \]
• So the change in momentum for each collision is 2m|v_z|

Pressure of a Gas

• The change in momentum due to all collisions with molecules with velocity |v_z| is:
  \[
  \Delta p = 2mn* A \Delta t |v_z|^2 f(v_z)
  \]
• The total change in momentum for all collisions is obtained by summing over all negative velocities
  \[
  \text{total } \Delta p = 2mn* A \Delta t \sum_{v_z} |v_z|^2 f(v_z)
  \]
• Because the motion is isotropic the sum over the negative velocities is equal to the sum over the positive velocities so the sum over the negative velocities is just half the sum over all velocities:
  \[
  \sum_{v_z} |v_z|^2 f(v_z) = \sum_{v_z} v_z^2 = \text{average squared velocity}
  \]
  \[
  \text{total } \Delta p = \frac{mn* A \Delta t}{2} v_z^2
  \]

Kinetic Energy of a Gas

• The translational kinetic energy of N gas particles of mass m is
  \[
  KE = \frac{1}{2} \text{molar KE}
  \]
  so
  \[
  PV = \frac{1}{3} \text{molar KE} = \frac{2}{3} \left( \frac{1}{2} \text{molar KE} \right) = \frac{2}{3} KE
  \]
• Combining with the ideal gas law
  \[
  PV = \frac{nRT}{2} \frac{3}{2} KE
  \]
• so
  \[
  KE = \frac{3}{2} nRT \quad \text{or} \quad KE = \frac{3}{2} RT \text{ per mole}
  \]
• Since R = N_A k, \quad KE = \frac{3}{2} RT \text{ per molecule}
Translational Kinetic Energy Example

- Calculate the translational kinetic energy of a mole of any ideal gas at 300 K.
- For comparison, calculate at what speed a 1.00 kg mass would have the same amount of kinetic energy.

Average Velocity of a Gas

\[ PV = nRT = \frac{nMRT}{3} \]

so \( v = \sqrt[3]{\frac{3RT}{M}} \) \( \text{(} k \text{ Boltzmann constant } = \frac{R}{N_A} \text{)} \)

\[ v_{rms} = \sqrt[3]{\frac{3RT}{M}} = \sqrt[3]{\frac{3kT}{m}} \]

Note:
- Lighter gases move faster at a given temperature
- All gases move faster at higher temperatures

\[ v_{rms} \propto \sqrt[3]{T} \]

Velocity Distribution

- Let \( v_x \) be the velocity in the x direction, \( v_y \) be the velocity in the y direction, and \( v_z \) be the velocity in the z direction.

\[ v^2 = v_x^2 + v_y^2 + v_z^2 \]

since \( v_x, v_y, \) and \( v_z \) are independent variables

\[ 2v = 2v_x, dv = v_x \, dv_x = v_y \, dv_y = v_z \, dv_z \]

- Let \( F(v_x, v_y, v_z) \) be the probability distribution function. In an isotropic gas the velocity components \( v_x, v_y, v_z \) are independent of each other. So

\[ F(v_x, v_y, v_z) = f(v_x) \cdot f(v_y) \cdot f(v_z) \]

then

\[ \ln[F(v_x, v_y, v_z)] = \ln[f(v_x)] + \ln[f(v_y)] + \ln[f(v_z)] \]

This gives

\[ f(v_x) = \frac{m}{2\pi kT} e^{-\frac{m}{kT} v_x^2} \]

Similarly

\[ f(v_y) = \frac{m}{2\pi kT} e^{-\frac{m}{kT} v_y^2} \]

and

\[ f(v_z) = \frac{m}{2\pi kT} e^{-\frac{m}{kT} v_z^2} \]

then

\[ F(v_x, v_y, v_z) = \frac{m^3}{(2\pi kT)^{3\frac{3}{2}}} e^{-\frac{m}{2kT} \left( \frac{v_x^2}{kT} + \frac{v_y^2}{kT} + \frac{v_z^2}{kT} \right)} \]

\[ F(v_x, v_y, v_z) = \frac{m^3}{(2\pi kT)^{\frac{3}{2}}} e^{-\frac{m}{2kT} \left( \frac{v_x^2}{kT} + \frac{v_y^2}{kT} + \frac{v_z^2}{kT} \right)} \]
Speed Distribution

The probability of finding a velocity in the range \(v_x\) to \(v_x + dv_x\), \(v_y\) to \(v_y + dv_y\), and \(v_z\) to \(v_z + dv_z\) is

\[
F(v_x, v_y, v_z) dv_x, dv_y, dv_z = \left(\frac{m}{2\pi kT}\right)^{2\over 3} e^{-mv^2/2kT} \, dv_x \, dv_y \, dv_z
\]

But what we really want to know is the probability of a speed between \(v\) and \(v + dv\). To do this we switch to polar coordinates and note that the distribution is isotropic (does not depend on \(\theta\) and \(\phi\)). So we can integrate over \(\theta\) and \(\phi\), then

\[
dv_x, dv_y, dv_z = 4\pi v^2 \, dv
\]

\[
F(v) \, dv = 4\pi \left(\frac{m}{2\pi kT}\right)^{2\over 3} \, v^2 \, e^{-mv^2/2kT} \, dv
\]

This is called the Maxwell-Boltzmann distribution.

**Speed Distribution**

- The most probable speed, \(v_{\text{mp}}\), corresponds to the maximum in the speed distribution function. \(v_{\text{mp}}\) can be found by setting the 1st derivative of the speed distribution function to 0 and solving for \(v\).
- The average speed can be calculated from

\[
\bar{v} = \int v F(v) \, dv = \frac{8kT}{\pi m} = \frac{8RT}{\pi M}
\]

- Recall root-mean-squared speed is

\[
v_{\text{rms}} = \sqrt{\frac{3RT}{m}} \, \sqrt{\frac{3RT}{M}}
\]

**Molecular Speed**

- For hydrogen gas at 0°C

\[
u_m = \text{the modal speed,} \quad u_{av} = \text{the simple average,} \quad u_{\text{rms}} = \sqrt{\pi}
\]

**Molecular Speed Dependence on Temperature**

- As temperature increases, average speed increases.
- As temperature increases, the fraction of molecules moving at higher speeds increases.

CO₂ Gas

\[
V_{\text{max}} = \frac{3RT}{\sqrt{M}}
\]

**Molecular Speed Dependence on Molecular Weight**

- For a fixed temperature, as the molecular weight increases, the average speed for the gas molecules decreases.

\[
T = 300 \text{ K} \quad V_{\text{rms}} = \frac{3RT}{M}
\]
Determining Molecular Speed

Molecular Speed Example

• Calculate the most probable, average and root-mean-squared speed of nitrogen molecules at 300 K and 1000 K.

Real Gases and Limitations of the Kinetic Theory

• Kinetic molecular theory implies that the volume of a gas molecule is insignificant compared to the “empty space” volume for a gas sample.
  – Mean free path used to test validity of assumption.
    • Average distance a particle travels between collisions with other particles.
    • The mean free path for air at room temperature and atmospheric pressure is 70 nm.
      – This value is 200 times larger than the typical radius of a small molecule like N₂ or O₂.
    • Volume of empty space in a gas is 1 million times that of gas particle volume.

Real Gases and Limitations of the Kinetic Theory

• The volume of a gas particle is significant compared to the “empty space” volume under high pressure conditions.
  – Mean free path decreases as pressure increases.
  – Gas molecules are very close together.
  – Therefore, volume of the gas particles becomes significant.

Real Gases and Limitations of the Kinetic Theory

• Kinetic molecular theory asserts that gas molecules move in straight lines and interact only through perfectly elastic collisions.
  – Gas molecules neither attract nor repel.
  – Strength of attractive forces small compared to kinetic energy of gas molecules.

• Attractive and repulsive forces are significant under conditions of low temperature.
  – Kinetic energy decreases with temperature.
    – Gas molecules experience “sticky” collisions.
    – Collision rate decreases, decreasing the pressure.

Real Gases and Limitations of the Kinetic Theory

• The ideal gas model breaks down at high pressures and low temperatures.
  – High pressure: volume of particles no longer negligible.
  – Low temperature: particles move slowly enough to interact.
**Real Gases**

- We must correct for ideal gas behavior when
  - the pressure of the gas is high.
  - the temperature is low.
- Smaller volume
- Attractive forces become important (gas molecules are attracted to each other and the walls)

- Compressibility factor, \( Z \) (\( Z = 1 \) for an ideal gas)
  \[ Z = \frac{PV}{nRT} \]

- Deviations occur for real gases.
  - \( Z > 1 \) – molecular volume is significant.
  - \( Z < 1 \) – intermolecular forces of attraction.

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**Plots of \( \frac{PV}{nRT} \) Versus \( P \) for Several Gases (200 K)**

- For nitrogen gas at three temperatures:
  - 203 K
  - 293 K
  - 673 K

**Illustration of Pairwise Interactions Among Gas Particles**

For \( n \) particles there are \( n\binom{n-1}{2} \) pairwise interactions

For large \( n \), \( N_a \), the number of interactions go as \( n^2 \).
van der Waals Equation

\[ P_{\text{obs}} + \alpha \frac{n}{V^2} \times (V - nb) = nRT \]

- \( \alpha \) corrects for attractive forces.
  - Molecules with stronger attractive forces have larger \( \alpha \) values.
- \( b \) corrects for the volume occupied by gas molecules.
  - Large molecules have larger \( b \) values.

van der Waals Constants for Some Common Gases

<table>
<thead>
<tr>
<th>Gas</th>
<th>( a ) (atm L^2 mol^{-2})</th>
<th>( b ) (L mol^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia, NH₃</td>
<td>4.170</td>
<td>0.0376</td>
</tr>
<tr>
<td>Argon, Ar</td>
<td>1.345</td>
<td>0.0321</td>
</tr>
<tr>
<td>Carbon dioxide, CO₂</td>
<td>3.192</td>
<td>0.0423</td>
</tr>
<tr>
<td>Helium, He</td>
<td>0.014</td>
<td>0.0277</td>
</tr>
<tr>
<td>Hydrogen, H₂</td>
<td>0.2484</td>
<td>0.0266</td>
</tr>
<tr>
<td>Hydrogen bromide, HF</td>
<td>9.431</td>
<td>0.0721</td>
</tr>
<tr>
<td>Methane, CH₄</td>
<td>2.253</td>
<td>0.0478</td>
</tr>
<tr>
<td>Nitrogen, N₂</td>
<td>1.390</td>
<td>0.0391</td>
</tr>
<tr>
<td>Oxygen, O₂</td>
<td>1.360</td>
<td>0.0310</td>
</tr>
<tr>
<td>Sulfur dioxide, SO₂</td>
<td>6.714</td>
<td>0.0561</td>
</tr>
<tr>
<td>Water, H₂O</td>
<td>5.464</td>
<td>0.0364</td>
</tr>
</tbody>
</table>

- The van der Waals constants \( \alpha \) and \( b \) are compound specific.
  - Both are zero in gases behaving ideally.

Real Gas Example

- Calculate the pressure exerted by 2.00 mol of oxygen gas at 27.0°C in a 24.0 L container
  - Using the ideal gas equation
  - Using the van der Waals equation
    \( \alpha = 1.36 \text{ atm L}^2 \text{ mol}^{-2} \) \( b = 0.0318 \text{ L mol}^{-1} \)
- Repeat the calculations with the volume changed to 0.240L
- Discuss the values obtained

Gas Sensors

- The concentration of air pollutants is monitored by the EPA.
  - The concentration of a gas is proportional to the partial pressure of the gas.
- Gas pressure sensors are used to monitor changes in partial pressure or concentration of gases.

Capacitance Manometer

- Changes in pressure cause deflections in the diaphragm, changing the capacitance.
  - Used to measure pressures from 0.001 - 1000 torr

Thermocouple Gauge

- Measures pressure by the cooling effect of colliding gas molecules.
  - Higher pressure, more collisions with heated filament, lowers filament temperature.
- Used to measure pressures from 0.01 to 1.0 torr
**Ionization Gauge**

- Pressure measured by producing gaseous cations with the electrons emitted from a hot filament.
  - Higher pressure, more gas cations, more current collected at the grid.
  - Used to measure pressures as low as $10^{-11}$ torr.

**Gas Sensors**

- A thermocouple gauge, a capacitance manometer, and an ionization gauge.

**Mass Spectrometers**

- Mass spectrometers can be used to measure partial pressures for gas mixtures.
  - Mass spectrometers ionize gas like an ionization gauge, but can select the mass of the gas being analyzed with the use of a magnetic field.
  - Several masses can be scanned simultaneously allowing for multiple gas analyses.
  - Current generated can be used to determine the partial pressure of gas.

**Gas Example**

Gasoline is a mixture of a large number of hydrocarbons (compounds that contain only H and C atoms). The principal ones are various octanes, which have the formula C$_8$H$_{18}$ (hence the term octane rating for fuels). When we burn gasoline in an automobile engine we try to get complete conversion of these hydrocarbons to CO$_2$ and H$_2$O. This requires that we use the correct mass ratio of air to gasoline (the air/fuel ratio). If the mixture is fuel-rich there is insufficient oxygen for complete combustion of the gasoline and unburned hydrocarbons and carbon monoxide are produced. These are unwanted air pollutants. If the mixture is fuel lean, there is an excess of air. This completes the combustion process, reducing the emissions of hydrocarbons and carbon monoxide. However, under these conditions the formation of nitrogen oxides is favored. These are unwanted air pollutants. The ideal air/fuel ratio is one in which oxygen (from air) and gasoline hydrocarbons are in stoichiometric proportions.

Major components of dry air by volume are 78.08% N$_2$, 20.95% O$_2$, 0.93% Ar and 0.03% CO$_2$.

**Gas Example**

1. Write a chemical equation for the complete combustions of octane (C$_8$H$_{18}$)
2. What is the mass of O$_2$ consumed in the complete combustion of 100. g C$_8$H$_{18}$?
3. What is the volume of CO$_2$(g), at 22°C and 745 mm Hg pressure, that is produced in the complete combustion of 100. g C$_8$H$_{18}$?
4. What is the percent mass of oxygen in air?
5. Calculate the mass ratio of air to octane for a mixture in which the octane burns stoichiometrically?

**Gas Example**

The molar mass of O$_2$ is to be determine similar to as was done in lab.

For O$_2$ the van der Waals constants are

- $\alpha = 1.360$ atm L$^2$ mol$^{-2}$
- $\beta = 0.03183$ L mol$^{-1}$

The following data was obtained

- $R = 0.0820575$ L atm mol$^{-1}$ K$^{-1}$
- $M_{O2} = 31.9988$ g
- $T_{0} = 20.00$°C
- $V_{0} = 24.055$ L

- Calculate the pressure using the van der Waals equation.
- Using the pressure from the van der Waals equation calculate the O$_2$ molar mass using the ideal gas equation.
Gas Example

$N_2O_4$ gas can break about into $NO_2$ and $NO_2$ can recombine to form $N_2O_4$. So a gas will contain a mixture of both $NO_2$ and $N_2O_4$ molecules.

A 1.00 L gas sample containing only $NO_2$ and $N_2O_4$ at 1 atm pressure and 25°C weighs 3.22 g.

1. Calculate the partial pressures of $NO_2$ and $N_2O_4$ and the ratio $P_{NO_2}/P_{N_2O_4}$.
   Assume ideal gas behavior.

2. If this system was an ideal gas, what would be the total pressure when the volume is expanded to 2.00 L?

3. When the volume is expanded to 2.00 L the final pressure is 0.527 atm. Calculate the partial pressures of $NO_2$ and $N_2O_4$ and the ratio $P_{NO_2}/P_{N_2O_4}$.

4. How do you explain this result?