Chapter 5
Gases

Gases are really different than the other states of matter...

State Variables of Gases
To describe a gas, we use certain state variables. \((P, V, T, n)\)

\[
\text{pressure} = \frac{\text{force}}{\text{area}}
\]

SI unit: Pascal. \(1 \text{ Pa} = 1 \text{ N} \cdot \text{m}^{-2}\); tiny! \(1 \text{ bar} \equiv 100 \text{ kPa} = 100,000 \text{ Pa}\)

1 atm = pressure exerted by a column of air at sea level and 0 °C

1 atm ≡ 760 torr = 760 mm Hg = 29.92 in Hg = 101.325 kPa = 14.7 psi

Gas Laws

Boyle’s Law (P-V relationship)
1660: Robert Boyle examined how \(P\) and \(V\) are related when holding \(n\) and \(T\) constant. Pressure and volume inversely proportional, i.e., \(PV = k\).

Considering two different sets of conditions, \(P_1V_1 = P_2V_2\)

Charles’s Law (V-T relationship)
Early 1800s: Jacques-Alexandre-César Charles examined how \(V\) and \(T\) are related when holding \(n\) and \(P\) constant. Temperature and volume directly proportional, i.e., \(V = kT\) or \(V/T = k\).

Considering two different sets of conditions, \(\frac{V_1}{T_1} = \frac{V_2}{T_2}\)

PROTIP
In all of these equations that involve \(T\), you must use Kelvin. If not: EPIC FAIL, FACEPALM, YOU’RE DOING IT WRONG.
Combining Boyle’s and Charles’s Laws
Since V is inversely proportional to P and directly proportional to T,
\[ \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \]

Amonton’s/Redenbacher’s Law (P-T relationship)
From the combined gas law above, we see that at constant volume,
\[ \frac{P_1}{P_2} = \frac{T_1}{T_2} \]

Avogadro’s Law (V-n relationship)
Volume proportional to moles of gas
\[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \]

**Ideal Gas Law**
Combining the combined gas law and Avogadro’s Law, we see that volume is inversely proportional to pressure and directly proportional to moles and temperature.
\[ \therefore PV = nRT \quad \text{where} \quad R = 0.08206 \text{ L·atm·mol}^{-1}·\text{K}^{-1} \]

**Ideal Gas Law and Stoichiometry**

\[ \text{GRAMS of A} \]
\[ \frac{\text{PV}}{\text{RT}} = n \]
\[ \text{MOLES of A} \]
\[ \text{Molarity} \times \text{Volume} \]
\[ \text{Solution A} \]

\[ \div \text{mole ratio of A} \]
\[ \text{GRAMS of B} \]
\[ \times \text{mole ratio of B} \]
\[ \text{MOLES of B} \]
\[ \text{n} = \frac{\text{PV}}{\text{RT}} \]
\[ \text{Gas B} \]
\[ = \text{Molarity} \times \text{Volume} \]
\[ \text{Solution B} \]
Gas Density

Kinetic-Molecular Theory (KMT)

Postulates of KMT

1. Particle Volume
   - Volume of each gas particle *extremely small* compared with volume of container.
   - Limiting case: Each gas particle has mass, but no volume.

2. Particle Motion
   - Gas particles move in constant, random, straight-line motion until they collide with other gas particles, or container walls.

3. Particle Collisions
   - All collisions are elastic – energy exchanged when gas particles collide with other gas particles, or with the container walls, but no energy lost through frictional forces.
   - Between collisions, gas particles do not experience forces of attraction with each other, or with the container walls.

\[
\therefore \text{Total kinetic energy of collection of gas particles is constant at a certain temperature.}
\]

\[
\therefore \text{The average kinetic energy of collection of gas particles is proportional to the absolute temperature, i.e., at a given temperature, all gases have the same average kinetic energy.}
\]

Gaseous Effusion & Diffusion
**Real Gases**

<table>
<thead>
<tr>
<th>KMT Assumption</th>
<th>Reality</th>
<th>Result</th>
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<tbody>
<tr>
<td>• gas particles have no volume.</td>
<td></td>
<td></td>
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<tr>
<td>• gas particles do not attract each other.</td>
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van der Waals equation: \( P + \frac{n^2a}{V^2} (V - nb) = nRT \)

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**Dalton’s Law of Partial Pressures**

Over a mixture of non-reacting gases: \( P_{\text{TOT}} = P_1 + P_2 + P_3 + \ldots = \Sigma P_i \)

The fraction of the pressure contributed by each gas equals the mole fraction of that gas.

\[
\frac{P_i}{P_{\text{TOT}}} = \frac{n_i}{n_{\text{TOT}}}
\]