Chapter 11

The Gaseous State

by Christopher Hamaker
Properties of Gases

• There are five important properties of gases:
  1. Gases have an indefinite shape.
  2. Gases have low densities.
  3. Gases can compress.
  4. Gases can expand.
  5. Gases mix completely with other gases in the same container.

• Let’s take a closer look at these properties.
1. Gases have an indefinite shape.
   - A gas takes the shape of its container, filling it completely. If the container changes shape, the gas also changes shape.

2. Gases have low densities:
   - The density of air is about 0.001 g/mL compared to a density of 1.0 g/mL for water. Air is about 1000 times less dense than water.
3. Gases can compress.
   • The volume of a gas decreases when the volume of its container decreases. If the volume is reduced enough, the gas will liquefy.

4. Gases can expand.
   • A gas constantly expands to fill a sealed container. The volume of a gas increases if there is an increase in the volume of the container.
5. Gases mix completely with other gases in the same container.

- Air is an example of a mixture of gases. When automobiles emit nitrogen oxide gases into the atmosphere, they mix with the other atmospheric gases.
- A mixture of gases in a sealed container will mix to form a homogeneous mixture.
The Greenhouse Effect

• Several gases contribute to the greenhouse effect.
• High energy radiation strikes Earth’s surface, and is converted to heat.
• This heat is radiated from the surface as infrared radiation.
• This infrared radiation is absorbed by the gases, and released as heat in all directions, heating the atmosphere.
• Gas pressure is the result of constantly moving gas molecules striking the inside surface of their container.
  
  – The more often the molecules collide with the sides of the container, the higher the pressure.

  – The higher the temperature, the faster the gas molecules move.
Atmospheric pressure is a result of the air molecules in the environment.

Evangelista Torricelli invented the barometer in 1643 to measure atmospheric pressure.

Atmospheric pressure is 29.9 inches of mercury or 760 torr at sea level.
Units of Pressure

• Standard pressure is the atmospheric pressure at sea level, 29.9 inches of mercury.
  – Here is standard pressure expressed in other units.

<table>
<thead>
<tr>
<th>TABLE 11.1 UNITS OF GAS PRESSURE</th>
</tr>
</thead>
<tbody>
<tr>
<td>UNIT</td>
</tr>
<tr>
<td>atmosphere</td>
</tr>
<tr>
<td>inches of mercury</td>
</tr>
<tr>
<td>centimeters of mercury</td>
</tr>
<tr>
<td>millimeters of mercury</td>
</tr>
<tr>
<td>torr¹</td>
</tr>
<tr>
<td>pounds per square inch</td>
</tr>
<tr>
<td>kilopascal²</td>
</tr>
</tbody>
</table>

¹A torr is defined as exactly equal to 1 mm Hg pressure; thus, standard pressure is 760 torr.
²The kilopascal (kPa) is the official SI unit of pressure.
Gas Pressure Conversions

• The barometric pressure is 27.5 in. Hg. What is the barometric pressure in atmospheres?

• We want atm; we have in. Hg.

• Use 1 atm = 29.9 in. Hg:

\[
27.5 \text{ in. Hg} \times \frac{1 \text{ atm}}{29.9 \text{ in. Hg}} = 0.920 \text{ atm}
\]
Variables Affecting Gas Pressure

There are three variables that affect gas pressure:

1. The *volume* of the container.
2. The *temperature* of the gas.
3. The *number of molecules* of gas in the container.
Volume Versus Pressure

- When volume decreases, the gas molecules collide with the container more often, so pressure increases.

- When volume increases, the gas molecules collide with the container less often, so pressure decreases.
Temperature Versus Pressure

• When temperature decreases, the gas molecules move slower and collide with the container less often, so pressure decreases.

• When temperature increases, the gas molecules move faster and collide with the container more often, so pressure increases.
Molecules Versus Pressure

- When the number of molecules decreases, there are fewer gas molecules colliding with the side of the container, so pressure decreases.

- When the number of molecules increases, there are more gas molecules colliding with the side of the container, so pressure increases.
Boyle’s Gas Experiment

- Robert Boyle trapped air in a J-tube using liquid mercury.
- He found that the volume of air decreased as he added more mercury.
- When he halved the volume, the pressure doubled.
Boyle’s Law

- *Boyle’s law* states that the volume of a gas is inversely proportional to the pressure at constant temperature.

- *Inversely proportional* means two variables have a reciprocal relationship.

- Mathematically, we write: 

\[ V \propto \frac{1}{P} \]
Boyle’s Law, Continued

• If we introduce a proportionality constant, \( k \), we can write Boyle’s law as follows:

\[ V = k \times \frac{1}{P} \]

• We can also rearrange it to \( PV = k \).

• Let’s take a sample of gas at \( P_1 \) and \( V_1 \), and change the conditions to \( P_2 \) and \( V_2 \). Because the product of pressure and volume is constant, we can write:

\[ P_1 V_1 = k = P_2 V_2 \]
Applying Boyle’s Law

• To find the new pressure after a change in volume:

\[ V_1 \times \frac{P_1}{P_2} = V_2 \]

• To find the new volume after a change in pressure:

\[ P_1 \times \frac{V_1}{V_2} = P_2 \]
Boyle’s Law Problem

• A 1.50 L sample of methane gas exerts a pressure of 1650 mm Hg. What is the final pressure if the volume changes to 7.00 L?

\[ P_1 \times \frac{V_1}{V_2} = P_2 \]

1650 mm Hg × \( \frac{1.50 \text{ L}}{7.00 \text{ L}} \) = 354 mm Hg

• The volume increased and the pressure decreased as we expected.
Charles’s Law

• In 1783, Jacques Charles discovered that the volume of a gas is directly proportional to the temperature in Kelvin. This is **Charles’s law**.

\[ V \propto T \] at constant pressure.

• Notice that Charles’s law gives a straight line graph.
Charles’s Law, Continued

• We can write Charles’s law as an equation using a proportionality constant, $k$.

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

• Again, let’s consider a sample of gas at $V_1$ and $T_1$, and change the volume and temperature to $V_2$ and $T_2$. Because the ratio of volume to temperature is constant, we can write:

$$\frac{V_1}{T_1} = k = \frac{V_2}{T_2}$$
Illustration of Charles’s Law

- Below is an illustration of Charles’s law.
- As a balloon is cooled from room temperature with liquid nitrogen (–196 °C), its volume decreases.
Applying Charles’s Law

• To find the new volume after a change in temperature:

\[ V_1 \times \frac{T_2}{T_1} = V_2 \]

• To find the new temperature after a change in volume:

\[ T_1 \times \frac{V_2}{V_1} = T_2 \]
Charles’s Law Problem

• A 275 L helium balloon is heated from 20 °C to 40 °C. What is the final volume at constant $P$?

• We first have to convert the temp from °C to K:

  $20 \, ^\circ\text{C} + 273 = 293 \, \text{K}$

  $40 \, ^\circ\text{C} + 273 = 313 \, \text{K}$

\[
V_1 \times \frac{T_2}{T_1} = V_2
\]

\[
275 \, \text{L} \times \frac{313 \, \text{K}}{293 \, \text{K}} = 294 \, \text{L}
\]
Gay-Lussac’s Law

• In 1802, Joseph Gay-Lussac discovered that the pressure of a gas is directly proportional to the temperature in Kelvin. This is Gay-Lussac’s Law.

• $P \propto T$ at constant temperature.

• Notice that Gay-Lussac’s law gives a straight line graph.
Gay-Lussac’s Law, Continued

• We can write Gay-Lussac’s law as an equation using a proportionality constant, \( k \).

\[ P = k T \quad \text{or} \quad \frac{P}{T} = k \]

• Let’s consider a sample of gas at \( P_1 \) and \( T_1 \), and change the volume and temperature to \( P_2 \) and \( T_2 \). Because the ratio of pressure to temperature is constant, we can write:

\[ \frac{P_1}{T_1} = k = \frac{P_2}{T_2} \]
Illustration of Gay-Lussac’s Law

- Here is an illustration of Gay-Lussac’s law.

- As the temperature of a gas in a steel cylinder increases, the pressure increases.
Applying Gay-Lussac’s Law

• To find the new volume after a change in temperature:

\[ \frac{T_2}{T_1} = \frac{P_2}{P_1} \]

• To find the new temperature after a change in volume:

\[ \frac{P_2}{P_1} = \frac{T_2}{T_1} \]
Gay-Lussac’s Law Problem

- A steel container of nitrous oxide at 15.0 atm is cooled from 25 °C to −40 °C. What is the final volume at constant $V$?

- We first have to convert the temp from °C to K:

  \[25 \, ^\circ\text{C} + 273 = 298 \, \text{K}\]
  \[−40 \, ^\circ\text{C} + 273 = 233 \, \text{K}\]

  \[P_1 \times \frac{T_2}{T_1} = P_2\]

  \[15.0 \, \text{atm} \times \frac{298 \, \text{K}}{233 \, \text{K}} = 11.7 \, \text{atm}\]
Combined Gas Law

• When we introduced Boyle’s, Charles’s, and Gay-Lussac’s laws, we assumed that one of the variables remained constant.

• Experimentally, all three (temperature, pressure, and volume) usually change.

• By combining all three laws, we obtain the *combined gas law*:

\[
\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}
\]
Applying the Combined Gas Law

• To find a new volume when \( P \) and \( T \) change:

\[ V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} \]

• To find a new pressure when \( V \) and \( T \) change:

\[ P_2 = P_1 \times \frac{V_1}{V_2} \times \frac{T_2}{T_1} \]

• To find a new temperature when \( P \) and \( V \) change:

\[ T_2 = T_1 \times \frac{P_2}{P_1} \times \frac{V_2}{V_1} \]
Combined Gas Law Problem

- In a combined gas law problem, there are three variables: $P$, $V$, and $T$.

- Let’s apply the combined gas law to 10.0 L of carbon dioxide gas at 300 K and 1.00 atm. If the volume and Kelvin temperature double, what is the new pressure?

<table>
<thead>
<tr>
<th>Conditions</th>
<th>$P$</th>
<th>$V$</th>
<th>$T$</th>
</tr>
</thead>
<tbody>
<tr>
<td>initial</td>
<td>1.00 atm</td>
<td>10.0 L</td>
<td>300 K</td>
</tr>
<tr>
<td>final</td>
<td>$P_2$</td>
<td>20.0 L</td>
<td>600 K</td>
</tr>
</tbody>
</table>
Combined Gas Law Problem, Continued

\[ P_2 = P_1 \times \frac{V_1}{V_2} \times \frac{T_2}{T_1} \]

\[ P_2 = 1.00 \text{ atm} \times \frac{10.0 \text{ L}}{20.0 \text{ L}} \times \frac{600 \text{ K}}{300 \text{ K}} \]

\[ P_2 = 1.00 \text{ atm} \]
Vapor Pressure

- **Vapor pressure** is the pressure exerted by the gaseous vapor above a liquid when the rates of evaporation and condensation are equal.

- Vapor pressure increases as temperature increases.
Dalton’s Law

- **Dalton’s law of partial pressures** states that the total pressure of a gaseous mixture is equal to the sum of the individual pressures of each gas.

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

- The pressure exerted by each gas in a mixture is its **partial pressure**, \( P_n \).
Dalton’s Law Calculation

• An atmospheric sample contains nitrogen, oxygen, and argon. If the partial pressure of nitrogen is 587 mm Hg, oxygen is 158 mm Hg, and argon is 7 mm Hg, what is the barometric pressure?

\[ P_{\text{total}} = P_{\text{nitrogen}} + P_{\text{oxygen}} + P_{\text{argon}} \]

\[ P_{\text{total}} = 587 \text{ mm Hg} + 158 \text{ mm Hg} + 7 \text{ mm Hg} \]

\[ P_{\text{total}} = 752 \text{ mm Hg} \]
Collecting a Gas Over Water

• We can measure the volume of a gas by displacement.
• By collecting the gas in a graduated cylinder, we can measure the amount of gas produced.
• The gas collected is referred to as “wet” gas since it also contains water vapor.
Ideal Gas Behavior

• An *ideal gas* is a gas that behaves in a predictable and consistent manner.

• Ideal gases have the following properties:
  – Gases are made up of very tiny molecules.
  – Gas molecules demonstrate rapid motion in straight lines and in random directions.
  – Gas molecules have no attraction for one another.
  – Gas molecules undergo elastic collisions.
  – The average kinetic energy of gas molecules is proportional to the Kelvin temperature, $KE \propto T$. 
Absolute Zero

- The temperature where the pressure and volume of a gas theoretically reaches zero is *absolute zero*.

- If we extrapolate $T$ versus $P$ or $T$ versus $V$ graphs to zero pressure or volume, the temperature is 0 Kelvin, or $-273 \, ^\circ C$. 
Ideal Gas Law

- Recall that the pressure of a gas is inversely proportional to volume and directly proportional to temperature and the number of molecules (or moles):

  \[ P \propto \frac{nT}{V} \]

- If we introduce a proportionality constant, \( R \), we can write the equation:

  \[ P = \frac{RnT}{V} \]
Ideal Gas Law, Continued

• We can rearrange the equation to read:

\[ PV = nRT \]

• This is the ideal gas law.

• The constant \( R \) is the ideal gas constant, and has a value of 0.0821 atm\( \cdot \)L/mol\( \cdot \)K.
Ideal Gas Law Problem

• How many mole of hydrogen gas occupy 0.500 L at STP?

• At STP, \( T = 273 \text{K} \) and \( P = 1 \text{ atm} \). Rearrange the ideal gas equation to solve for moles:

\[
n = \frac{PV}{RT}
\]

\[
n = \frac{(1 \text{ atm})(0.500 \text{ L})}{(0.0821 \text{ atm} \cdot \text{L/mol} \cdot \text{K})(273 \text{K})}
\]

\[
n = 0.0223 \text{ moles}
\]
Chapter Summary

• Gases have variable shape and volume.

• The pressure of a gas is *directly proportional* to the temperature and the number of moles present.

• The pressure of a gas is *inversely proportional* to the volume it occupies.

• Standard temperature and pressure are exactly 1 atmosphere and 0 °C (273 K).
Chapter Summary, Continued

• **Boyle’s law** is: \( P_1 V_1 = P_2 V_2 \).

• **Charles’s law** is: \( \frac{V_1}{T_1} = \frac{V_2}{T_2} \).

• **Gay-Lussac’s law** is: \( \frac{P_1}{T_1} = \frac{P_2}{T_2} \).

• **The combined gas law** is: \( \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \).
• *Dalton’s law of partial pressures* is:

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

• The *ideal gas law* is: \( PV = nRT \).

• \( R \) is the *ideal gas constant*: 0.0821 atm \cdot L/mol \cdot K.