Chapter 6: Properties of Gases: The Air We Breathe

# Definitions: Phases of Matter

<table>
<thead>
<tr>
<th></th>
<th>Gas</th>
<th>Liquid</th>
<th>Solid</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Water as an example:</strong></td>
<td>[Image of gaseous water (steam)]</td>
<td>[Image of liquid water]</td>
<td>[Image of solid water (ice)]</td>
</tr>
<tr>
<td><strong>Shape</strong></td>
<td>Variable—same as a closed container</td>
<td>Variable—same as the bottom of the container</td>
<td>Constant—rigid, fixed</td>
</tr>
<tr>
<td><strong>Volume</strong></td>
<td>Variable—same as a closed container</td>
<td>Constant</td>
<td>Constant</td>
</tr>
<tr>
<td><strong>Particle Movement</strong></td>
<td>Completely independent (random); each particle may go anyplace in a closed container</td>
<td>Independent beneath the surface, limited to the volume of the liquid and the shape of the bottom of the container</td>
<td>Vibration in fixed position</td>
</tr>
</tbody>
</table>
Important Characteristics of Gases

1) Gases are highly compressible
   An external force compresses the gas sample and decreases its volume; removing the external force allows the gas volume to increase.

2) Gases are thermally expandable
   When a gas sample is heated, its volume increases; when it is cooled its volume decreases.

3) Gases have low viscosity
   Gases flow much easier than liquids or solids.

4) Most gases have low densities
   Gas densities are on the order of grams per liter, whereas liquids and solids are grams per cm$^3$ (mL), 1000 times greater.

5) Gases are infinitely miscible
   Gases mix in any proportion (air is a mixture of many gases).
Steam clean the inside of a train tank car, and then seal the top and allow the contents to thermally equilibrate with the surroundings.

What will happen??

Heat released to the surroundings:

\[ q = -108,000 \text{ kJ} \]

Initial Volume of the tank car:

\[ V_{\text{initial}} = 23,000 \text{ gal} \]

Final Volume of the tank car:

\[ V_{\text{final}} = 13,800 \text{ gal} \]

What is \( w \)?

What is \( DE \)?

http://www.youtube.com/watch?v=Zz95_VvTxZM

Tank Car Collapse!!

http://www.youtube.com/watch?v=Zz95_VvTxZM
Earth’s Atmosphere

The Earth’s atmosphere is a layer of gases about 50 km (~31 miles) thick, consisting primarily of nitrogen (78%), oxygen (21%), and trace gases (Ar, CO$_2$, water vapor, etc.).
Atmospheric Pressure

Gravity pulls the gases in the Earth’s atmosphere towards the planet’s surface.

→ The weight of these gases exerts *pressure* on surfaces in the environment.

**Atmospheric pressure** is exerted by gas molecules in the Earth’s atmosphere.

→ Anything entering the Earth's atmosphere from space comes into contact with all of these gas molecules → friction and heat.

→ Objects burn up in the atmosphere before reaching the Earth’s surface!
Atmospheric Pressure

Atmospheric Pressure
• depends on location, temperature, and weather conditions
• depends on the overlying mass of the atmosphere
  – decreases as altitude increases
  • Thus, air becomes thinner at higher altitudes.
• Atmospheric pressure is ~760 mmHg at sea level but 15-20% lower in Denver (~1 mile above sea level) and 65% lower at the top of Mt. Everest (~5.5 mi. above sea level)

Vacuum: empty space with no gas molecules present
  – gas pressure equals zero: \( P_{\text{gas}} = 0 \)
Pressure

• Pressure is force per unit area:
  \[ P = \frac{F}{A} \]

• SI units: pressure is expressed in newtons per square meter, \( \text{N/m}^2 \), the pascal (Pa).

  \[ 1 \text{ Pa} = 1 \text{ N/m}^2 \]

• How do we measure the force of a **gas**?
Barometer

The gases in the atmosphere at sea level on a fair day (no storms) exert “one atmosphere” of pressure.

The mercury column exerts a force over the cross-sectional area of the tube.

The pressure exerted by the mercury column is exactly balanced by the pressure of the atmosphere.

\[ P = h \cdot g \cdot d \]

- \( P \) = pressure
- \( h \) = height
- \( g \) = accel. due to gravity
- \( d \) = density
Measuring the height of the Hg column will tell you what the atmospheric pressure is.

1 atm of pressure = 760 mm Hg
...at sea level on a nice day (no storms)

In Breckenridge, CO (elev. 9600 ft) atmospheric pressure is only 520 mm Hg.

Units: “mm Hg” (milimeters of mercury)
Also known as “torr” after Evangelista Torricelli, inventor of the barometer.
# Common Units of Pressure

<table>
<thead>
<tr>
<th>Unit</th>
<th>Atmospheric Pressure</th>
<th>Scientific Field Used</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pascal (Pa) = N/m²; kilopascal (kPa)</td>
<td>1.01325 x 10⁵ Pa 101.325 kPa</td>
<td>SI unit; physics, Chemistry</td>
</tr>
<tr>
<td>bar</td>
<td>1.01325 bar</td>
<td>Meteorology, Chemistry</td>
</tr>
<tr>
<td>atmosphere (atm)</td>
<td>1 atm</td>
<td>Chemistry</td>
</tr>
<tr>
<td>millimeters of mercury (mmHg), also called “torr”</td>
<td>760 mmHg 760 torr</td>
<td>Chemistry, medicine, biology</td>
</tr>
<tr>
<td>pounds per square inch (psi or lb/in²)</td>
<td>14.7 lb/in²</td>
<td>Engineering</td>
</tr>
</tbody>
</table>
Example

The atmospheric pressure in the lab was measured to be 29.22 inHg. Express this pressure in units of mmHg, torr, atm, and kPa.

\[
29.22 \text{ inHg} \times \frac{25.4 \text{ mm}}{1 \text{ inch}} = 742.2 \text{ mmHg} = 742.2 \text{ torr}
\]

\[
742.2 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.977 \text{ atm}
\]

\[
742.2 \text{ torr} \times \frac{101.325 \text{ kPa}}{760 \text{ torr}} = 98.95 \text{ kPa}
\]
Typical Gases

Know a few substances are gases at "normal atmospheric conditions" (25°C and 1 atm)

• Elements that are gases: H₂, N₂, O₂, F₂, Cl₂, ozone (O₃), all Noble Gases

• Some molecules are gases (CO, CO₂, HCl, NH₃, CH₄); most are solids or liquids.

• No ionic compounds exist as gases.
Manometers operate on the same principle as barometers, but they measure the pressure of an isolated gas sample rather than the whole atmosphere.

\[ P_{\text{gas}} + h = P_{\text{atm}} \]

\[ P_{\text{gas}} = P_{\text{atm}} + h \]

**Note:** \( P_{\text{atm}} \) is obtained from a barometer.
Example

Gas pressure is higher than atmospheric pressure for which one?

left  right

Gas pressure is lower than atmospheric pressure for which one?

left  right
Example

If the height difference for the example on the left is 95 mmHg and atmospheric pressure is 725 mmHg, calculate the gas pressure.

\[ P_{\text{gas}} = P_{\text{atm}} - \Delta h \]
\[ = 725 \text{ mmHg} - 95 \text{ mmHg} \]
\[ = 630 \text{ mmHg} \]
Example

If the height difference for the example on the right is 85 torr and atmospheric pressure is 0.975 atm, calculate the gas pressure.

\[ P_{\text{gas}} = P_{\text{atm}} + \Delta h; \quad P_{\text{atm}} = 0.975 \text{ atm} = 741 \text{ torr} \]
\[ = 741 \text{ torr} + 85 \text{ torr} \]
\[ = 826 \text{ torr} \]
The First Gas Laws

• Gases are relatively easy to measure and observe in a laboratory.

• This made the physical properties of gases a popular object of study in the 17th, 18th, and 19th centuries.

• Boyle, Charles, Avogadro (yes, *that* Avogadro), and Amontons determined fundamental connections between P, V, T and n (# of moles) for gases.
Boyle’s Law (c. 1650)

Boyle studied the connection between $P$ and $V$ of gases.

$T$ and $n$ held constant.

$PV = \text{constant}$
Boyle’s Law

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

For an ideal gas, PV will be a constant as the pressure increases. For real gases, this is approximately true at low pressures.
Example of Boyle’s Law

A gas sample at a pressure of 1.23 atm has a volume of 15.8 cm³, what will be the volume (in L) if the pressure is increased to 3.16 atm?

Do you expect volume to increase or decrease?

\[ P_{\text{initial}} = 1.23 \text{ atm} \quad P_{\text{final}} = 3.16 \text{ atm} \]
\[ V_{\text{initial}} = 15.8 \text{ cm}^3 \quad V_{\text{final}} = \text{unknown} \]
\[ T \text{ and } n \text{ remain constant} \]

\[ P_i V_i = P_f V_f \]

\[ V_i = 15.8 \text{ cm}^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0158 \text{ L} \]

\[ V_f = V_i \times \frac{P_i}{P_f} = 0.0158 \text{ L} \times \frac{1.23 \text{ atm}}{3.16 \text{ atm}} = 0.00615 \text{ L} \]
Example of Boyle’s Law

A 250.0-mL sample of helium at 722 mmHg is compressed until the new pressure is 3.60 atm. Calculate the new volume.

Do you expect volume to increase or decrease?

\[
P_{\text{initial}} = 722 \text{ mmHg} \quad P_{\text{final}} = 3.60 \text{ atm}
\]

\[
V_{\text{initial}} = 250.0 \text{ mL} \quad V_{\text{final}} = \text{unknown}
\]

\[
T \text{ and } n \text{ remain constant}
\]

\[
P_i V_i = P_f V_f
\]

\[
V_i = 250 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.250 \text{ L}
\]

\[
P_i = 722 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.950 \text{ atm}
\]

\[
V_f = V_i \times \frac{P_i}{P_f} = 0.250 \text{ L} \times \frac{0.95 \text{ atm}}{3.60 \text{ atm}} = 0.066 \text{ L}
\]
Decreasing the pressure in the marshmallow vessel will cause the marshmallows to...

1. Get bigger
2. Get smaller
3. Stay the same

Marshmallow Video
Charles’ Law (c. 1800)

Charles studied the connection between T and V of gases.

P and n held constant

Liq. N$_2$, T = 77 K

Blue balloon is placed in the liquid N$_2$...

Cooling the contents of the blue balloon decreases its volume dramatically.

The volume of a sample of an ideal gas is directly proportional to its temperature measured in *Kelvin*. 
Temperature on Two Scales

\[
K = ^\circ C + 273.15
\]
Example of Charles’ Law

A sample of carbon monoxide occupies 3.20 L at 125 °C. If the sample is heated at constant pressure, calculate the temperature (°C) at which the gas will occupy 1.54 L.

\[ V_{\text{initial}} = 3.20 \text{ L} \]
\[ V_{\text{final}} = 1.54 \text{ L} \]

\[ T_{\text{initial}} = 125^\circ \text{C} + 273.15 = 398.15 \text{ K} \]

\[ T_{\text{final}} = ? \]

**P and n are held constant**

*Volume decreases, so what will happen to T?*

\[ \frac{V_i}{T_i} = \frac{V_f}{T_f} \]

\[ T_f = T_i \times \left( \frac{V_f}{V_i} \right) \]

\[ T_f = 398.15 \text{ K} \times \frac{1.54 \text{ L}}{3.20 \text{ L}} = 191.6 \text{ K} \]

\[ ^\circ \text{C} = K - 273.15 = 191.6 - 273.15 \]
\[ ^\circ \text{C} = -81.55^\circ \text{C} = -82^\circ \text{C} \]
Example of Charles’ Law

When 20.0 L of hydrogen gas are heated from 25.00°C to 450.50°C, the volume changes. Calculate the new volume.

\[ V_{\text{initial}} = 20.0 \text{ L} \quad T_{\text{initial}} = 25.00^\circ\text{C} + 273.15 = 398.15 \text{ K} \]
\[ V_{\text{final}} = ? \quad T_{\text{final}} = 450.5^\circ\text{C} + 273.15 = 723.65 \text{ K} \]

\( P \) and \( n \) are held constant

\textit{Temperature increases, so what will happen to} \( V \)?

\[
\frac{V_i}{T_i} = \frac{V_f}{T_f}
\]

\[ V_f = T_f \times \left( \frac{V_i}{T_i} \right) = 723.65 \text{ K} \times \frac{20.0 \text{ L}}{398.15 \text{ K}} \]
\[ = 36.35 \text{ L} \]
Avogadro’s Law (c. 1800)

Avogadro studied the connection between V and n of gases

P and T held constant

Volumes of gases that react do so in small whole number ratios:

\[
\text{2 vol } H_2 + \text{1 vol } O_2 = 2 \text{ vol } H_2O
\]

\[
V \propto n
\]

The same volume of two different gases at the same T and P will have the same number of particles.
Amontons’ Law (c. 1700)

Amontons studied the connection between $T$ and $P$ of gases.

$V$ and $n$ held constant

The pressure of a gas is directly proportional to its temperature measured in Kelvin.

$$P \propto T$$
Empirical Gas Law Summary

The variables $P$, $T$, $V$, and $n$ can be used to describe the state of a gas. If two of these variables are held constant, we can observe how the remaining two variables are related:

Boyle’s Law ($T$, $n$ const):

$$V \propto \frac{1}{P}$$

Charles’s Law ($P$, $n$ const):

$$V \propto T$$

Avogadro’s Law ($T$, $P$ const):

$$V \propto n$$

Amontons’s Law ($V$, $n$ const):

$$P \propto T$$

Since $V$ is directly proportional to $1/P$, $T$, and $n$, $V$ must also be directly proportional to the product:

$$V \propto \frac{nT}{P}$$
Ideal Gas Law

\[ V = R \left( \frac{nT}{P} \right) \]

\[ PV = nRT \]

P = pressure
V = volume
n = number of moles
R = “gas constant”
T = temperature \textit{in Kelvin}

\textbf{Common R values:}

\[ 0.082057 \text{ L \cdot atm/mol\cdot K} \]
\[ 8.3145 \text{ J/mol\cdot K} \]

\textit{It’s the Swiss Army knife of gas laws!!}
What is Ideality?

Recall that the molecules in a gaseous substance are very, very far apart. So we can make a couple of assumptions...

- Molecules of an ideal gas do not attract or repel one another
- The volume of an ideal gas molecule is negligible with respect to the container

An ideal gas is essentially a collection of non-interacting particles.

Under what conditions would you expect ideality to fail?

- high P → molecules get too close, start interacting
- low T → same thing
PV = nRT Example

What is the pressure exerted by 1.00 \times 10^{20} \text{ molecules of N}_2 \text{ gas in a 305 mL flask at 175}^\circ\text{C}?

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

\[ \begin{align*}
\text{P} & = \\
\text{V} & = \\
n & = \\
R & = \\
T & = 
\end{align*} \]
What is the pressure exerted by $1.00 \times 10^{20}$ molecules of $\text{N}_2$ gas in a 305 mL flask at 175°C?

\[ P = ?, \quad V = 0.305 \text{ L} \]
\[ n = 0.000166 \text{ mol N}_2 \]
\[ R = 0.08206 \text{ L.atm/mol.K} \]
\[ T = 448 \text{ K} \]

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

\[ P = \frac{(0.000166 \text{ mol N}_2)(0.08206 \text{ L.atm/mol.K})(448 \text{ K})}{0.305 \text{ L}} \]

\[ P = 0.02000 \text{ atm} \]
What if you forget Boyle’s Law?

A gas sample at a pressure of 1.23 atm has a volume of 15.8 cm$^3$, what will be the volume (in L) if the pressure is increased to 3.16 atm?

\[ P_1 V_1 = nRT = P_2 V_2 \]
\[ 1.23 \text{ atm} \times 15.8 \text{ mL} = 3.16 \text{ atm} \times V_2 \]
\[ V_2 = 6.15 \text{ mL} = 0.00615 \text{ L} \]
Example of Avogadro’s Law

If the volume of 2.67 g of SF\textsubscript{6} gas at 1.143 atm and 28.5°C is 2.93 m\textsuperscript{3}, what mass of SF\textsubscript{6} will you find in a container with a volume of 543.9 m\textsuperscript{3} at the same pressure and temperature?

\[ PV = nRT; \]
\[ \text{mw of SF}_6 = 146.07 \text{ g/mol} \]
\[ \text{moles of SF}_6 = 0.0183 \text{ mol} \]
\[ \frac{V_1}{n_1} = \frac{V_2}{n_2}; n_2 = n_1 \frac{V_2}{V_1} \]
\[ n_2 = \frac{(0.183 \text{ mol})(543.9 \text{ m}^3)}{(2.93 \text{ m}^3)} \]
\[ n_2 = 3.4 \text{ mol SF}_6 = 496 \text{ g SF}_6 \]
Example of Amontons’ Law

The air pressure in the tires of an automobile is adjusted to 28 psi at a gas station in San Diego, CA, where the air temperature is 68°F (20.°C). The automobile is then driven east along a hot desert highway. Along the way, the temperature of the tire reaches 140°F (60.°C). What is the pressure in the tires?

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

\[
P_2 = \frac{(28 \text{ psi})(333 \text{ K})}{(293 \text{ K})}
\]

\[
P_2 = 32 \text{ psi}
\]
Another Boyle’s Law Example

An inflated balloon has a volume of 0.55 L at sea level. It is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming the temperature remains constant, what is the final volume of the balloon?

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

\[
V_2 = \frac{(1 \text{ atm})(0.55 \text{ L})}{(0.40 \text{ atm})} = 1.4 \text{ L}
\]
Another Charles’ Law Example

A 452-mL sample of fluorine gas is heated from 22°C to 187°C at constant pressure. What is the final volume of the gas?

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

\[
V_2 = \frac{(452 \text{ mL})(460 \text{ K})}{(295 \text{ K})}
\]

\[V_2 = 705 \text{ mL}\]