PROPERTIES OF GASES

- **Gases** are the least dense and most mobile of the three phases of matter.

- **Particles of matter** in the gas phase are spaced far apart from one another and move rapidly and collide with each other often.

- Gases occupy much greater space than the same amount of liquid or solid. This is because the gas particles are spaced apart from one another and are therefore compressible. Solid or liquid particles are spaced much closer and cannot be compressed further.

- Gases are characterized by four properties. These are:
  1. Pressure (P).
  2. Volume (V)
  3. Temperature (T)
  4. Amount (n)

**Kinetic-Molecular Theory of Gases**

- Scientists use the kinetic-molecular theory (KMT) to describe the behavior of gases. The KMT consists of several postulates:

  1. Gases consist of small particles (atoms or molecules) that move randomly with rapid velocities.

  2. Gas particles have little attraction for one another. Therefore, attractive forces between gas molecules can be ignored.

  3. The distance between the particles is large compared to their size. Therefore the volume occupied by gas molecules is small compared to the volume of the gas.

  4. Gas particles move in straight lines and collide with each other and the container frequently. The force of collisions of the gas particles with the walls of the container causes pressure.

  5. The average kinetic energy of gas molecules is directly proportional to the absolute temperature (Kelvin).
PRESSURE AND ITS MEASUREMENT

- **Pressure** is the result of collision of gas particles with the sides of the container. Pressure is defined as the force per unit area.

- **Pressure** is measured in units of atmosphere (atm) or mmHg or torr. The SI unit of pressure is pascal (Pa) or kilopascal (kPa).

  \[
  \begin{align*}
  1 \text{ atm} & = 760 \text{ mmHg} \\
  1 \text{ mmHg} & = 1 \text{ torr} \\
  1 \text{ atm} & = 101.325 \text{ kPa}
  \end{align*}
  \]

- Atmospheric pressure can be measured with the use of a barometer. Mercury is used in a barometer due to its high density. At sea level, the mercury stands at 760 mm above its base.

- The pressure of a gas is directly proportional to the number of particles (moles) present.

**Examples:**
1. The atmospheric pressure at Walnut, CA is 740. mmHg. Calculate this pressure in torr and atm.

   \[
   1 \text{ mmHg} = 1 \text{ torr} \quad \text{therefore,} \quad 740. \text{ mmHg} = 740. \text{ mmHg} \times \frac{1}{\text{torr}} = \text{atm}
   \]

2. The barometer at a location reads 1.12 atm. Calculate the pressure in mmHg and torr.
RELATIONSHIP BETWEEN PRESSURE & VOLUME

BOYLE’S LAW

- At constant temperature, the volume of a fixed amount of gas is inversely proportional to its pressure.

\[ P_1V_1 = P_2V_2 \]

Examples:

1. A sample of H₂ gas has a volume of 5.0 L and a pressure of 1.0 atm. What is the new pressure if the volume is decreased to 2.0 L at constant temperature?

\[
P_1 = 1.0 \text{ atm} \quad P_2 = ??? \quad P_2 = \frac{P_1V_1}{V_2} = \frac{(1.0 \text{ atm})(5.0 \text{ L})}{2.0 \text{ L}} = 2.5 \text{ atm}
\]

\[
V_1 = 5.0 \text{ L} \quad V_2 = 2.0 \text{ L}
\]

2. A sample of gas has a volume of 12-L and a pressure of 4500 mmHg. What is the volume of the gas when the pressure is reduced to 750 mmHg?

\[
P_1 = \quad P_2 = \]

\[
V_1 = \quad V_2 = ???
\]

3. A sample of hydrogen gas occupies 4.0 L at 650 mmHg. What volume would it occupy at 2.0 atm?
RELATIONSHIP BETWEEN TEMP. & VOLUME

CHARLES’S LAW

- At constant pressure, the volume of a fixed amount of gas is directly proportional to its absolute temperature.

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

Note: T must be in unit of K

Examples:
1. A 2.0-L sample of a gas is cooled from 298K to 278 K, at constant pressure. What is the new volume of the gas?

   \[ V_1 = 2.0 \text{ L} \quad V_2 = ??? \]

   \[ \frac{T_2}{T_1} = \frac{278 \text{ K}}{298 \text{ K}} = 0.931 \]

   \[ V_2 = V_1 \times \frac{T_2}{T_1} = 2.0 \text{ L} \times 0.931 = 1.9 \text{ L} \]

2. A sample of gas has a volume of 5.0-L and a temperature of 20°C. What is the volume of the gas when the temperature is increased to 50°C, at constant pressure?

   \[ V_1 = \quad V_2 = \quad \]

   \[ T_1 = \quad T_2 = \]

3. If 20.0 L of oxygen gas is cooled from 100 °C to 0 °C, what is the new volume?
At constant volume, the pressure of a fixed amount of gas is directly proportional to its absolute temperature.

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

**Examples:**

1. An aerosol spray can has a pressure of 4.0 atm at 25°C. What pressure will the can have if it is placed in a fire and reaches temperature of 400°C?

   \[P_1 = 4.0 \text{ atm} \quad P_2 = ??? \quad T_1 = 298 \text{ K} \quad T_2 = 673 \text{ K}\]

   \[P_2 = P_1 \times \frac{T_2}{T_1} = 4.0 \text{ atm} \times \frac{673 \text{ K}}{298 \text{ K}} = 9.0 \text{ atm}\]

2. A cylinder of gas with a volume of 15.0-L and a pressure of 965 mmHg is stored at a temperature of 55°C. To what temperature must the cylinder be cooled to reach a pressure of 850 mmHg?

   \[P_1 = \quad P_2 = \quad T_1 = 298 \text{ K} \quad T_2 = ???\]

3. The pressure of a container of helium is 650 mmHg at 25°C. If the container is cooled to 0°C, what will the pressure be?
VAPOR PRESSURE & BOILING POINT

- In an open container, liquid molecules at the surface that possess sufficient energy, can break away from surface and become gas particles or vapor. In a closed container, these gas particles can accumulate and create pressure called vapor pressure.

- Vapor pressure is defined as the pressure above a liquid at a given temperature. Vapor pressure varies with each liquid and increases with temperature. Listed below is the vapor pressure of water at various temperatures.

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Vapor Pressure (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>5</td>
</tr>
<tr>
<td>10</td>
<td>9</td>
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<td>20</td>
<td>18</td>
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<td>30</td>
<td>32</td>
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<td>37</td>
<td>47&lt;sup&gt;a&lt;/sup&gt;</td>
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<td>40</td>
<td>55</td>
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<td>50</td>
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<tr>
<td>60</td>
<td>149</td>
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<td>70</td>
<td>234</td>
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<tr>
<td>80</td>
<td>355</td>
</tr>
<tr>
<td>90</td>
<td>528</td>
</tr>
<tr>
<td>100</td>
<td>760</td>
</tr>
</tbody>
</table>

<sup>a</sup>At body temperature.

- A liquid reaches its boiling point when its vapor pressure becomes equal to the external pressure (atmospheric pressure). For example, at sea level, water reaches its boiling point at 100°C since its vapor pressure is 760 mmHg at this temperature.

- At higher altitudes, where atmospheric pressure is lower, water reaches boiling point at temperatures lower than 100°C. For example in Denver, where atmospheric pressure is 630 mmHg, water boils at 95°C, since its vapor pressure is 630 mmHg at this temperature.
RELATIONSHIP BETWEEN PRESSURE, VOL. & TEMP.

COMBINED GAS LAW

- All pressure-volume-temperature relationships can be combined into a single relationship called the combined gas law. This law is useful for studying the effect of changes in two variables.

\[
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}
\]

- The individual gas laws studied previously are embodied in the combined gas law.

\[
\begin{align*}
\frac{P_1V_1}{T_1} &= \frac{P_2V_2}{T_2} & \text{Boyle's Law} \\
\frac{P_1V_1}{T_1} &= \frac{P_2V_2}{T_2} & \text{Charles's Law} \\
\frac{P_1V_1}{T_1} &= \frac{P_2V_2}{T_2} & \text{Gay-Lussac's Law}
\end{align*}
\]

Examples:
1. A 25.0-mL sample of gas has a pressure of 4.00 atm at a temperature of 10°C. What is the volume of the gas at a pressure of 1.00 atm and a temperature of 18°C?

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

\[
\begin{align*}
P_1 &= \text{atm} \\
P_2 &= \text{atm} \\
V_1 &= \text{mL} \\
V_2 &= \text{mL} \\
T_1 &= \text{°C} \\
T_2 &= \text{°C}
\end{align*}
\]

\[
V_2 = 25.0 \text{ mL} \times \frac{4.00 \text{ atm}}{1.00 \text{ atm}} \times \frac{291 \text{ K}}{283 \text{ K}} = 103 \text{ mL}
\]

2. A sample of ammonia has a volume of 20.0 mL at 5°C and 700 mmHg. What is the volume of the gas at 50°C and 850 torr?

\[
\begin{align*}
P_1 &= \text{atm} \\
P_2 &= \text{atm} \\
V_1 &= \text{mL} \\
V_2 &= \text{mL} \\
T_1 &= \text{°C} \\
T_2 &= \text{°C}
\end{align*}
\]


### RELATIONSHIP BETWEEN VOLUME & MOLES

**AVOGADRO’S LAW**

- At constant temperature and pressure, the volume of a fixed amount of gas is directly proportional to the number of moles.

- As a result of Avogadro’s Law, equal volumes of different gases at the same temperature and pressure contain equal number of moles (molecules).

- This relationship also allows chemists to relate volumes and moles of a gas in a chemical reaction. For example:

\[
\begin{align*}
2 \text{H}_2 (g) & \quad + \quad \text{O}_2 (g) \quad \rightarrow \quad 2 \text{H}_2\text{O} (g) \\
2 \text{molecules} & \quad 1 \text{molecule} \quad 2 \text{molecules} \\
2 \text{moles} & \quad 1 \text{mole} \quad 2 \text{moles} \\
2 \text{Liters} & \quad 1 \text{Liter} \quad 2 \text{Liters}
\end{align*}
\]

**Examples:**

1. A sample of helium gas with a mass of 18.0 g occupies 1.6 Leters at a particular temperature and pressure. What mass of oxygen would occupy 1.6 L at the same temperature and pressure?

   \[
   \text{mol of He} = \frac{18.0 \text{ g}}{4.00 \text{ g/mol}} = 4.50 \text{ mol He}
   \]

   \[
   \text{mol of O}_2 = \text{mol He} = 4.50 \text{ mol} \quad \text{(since at same T & P)}
   \]

   \[
   \text{mass of O}_2 = 4.50 \text{ mol} \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 144 \text{ g}
   \]

2. How many Liters of NH\textsubscript{3} can be produced from reaction of 1.8 L of H\textsubscript{2} with excess N\textsubscript{2}, as shown below?

\[
\begin{align*}
\text{N}_2 (g) & \quad + \quad 3 \text{H}_2 (g) \quad \rightarrow \quad 2 \text{NH}_3 (g) \\
1.8 \text{ L H}_2 & \quad x \quad \frac{2 \text{ L NH}_3}{3 \text{ L H}_2} = 1.2 \text{ L NH}_3
\end{align*}
\]
STP & MOLAR VOLUME

- To better understand the factors that affect gas behavior, a set of *standard conditions* have been chosen for use, and are referred to as *Standard Temperature and Pressure (STP)*.

**STP: 0°C (273 K) and 1 atm (760 mmHg)***

- At STP conditions, *one mole of any gas* is observed to occupy a *volume of 22.4 L*. This quantity is referred to as *Molar Volume*.

**Molar Volume of a gas at STP = 22.4 L**

**Examples:**

1. If 2.00 L of a gas at STP has a mass of 3.23 g, what is the molar mass of the gas?

\[
\text{mol of gas} = 2.00 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.0893 \text{ mol}
\]

\[
\text{molar mass} = \frac{\text{g}}{\text{mol}} = \frac{3.23 \text{ g}}{0.0893 \text{ mol}} = 36.2 \text{ g/mol}
\]

2. A sample of gas has a volume of 2.50 L at 730 mmHg and 20°C. What is the volume of this gas at STP?

\[
P_1 = \quad P_2 =
\]

\[
V_1 = \quad V_2 =
\]

\[
T_1 = \quad T_2 =
\]
IDEAL GAS LAW

- Combining all the laws that describe the behavior of gases, one can obtain a useful relationship that relates the volume of a gas to the temperature, pressure and number of moles.

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

\[ R = \text{Universal Gas Constant} = 0.0821 \ \text{L atm} / \text{mol K} \]

- This relationship is called the **Ideal Gas Law**, and commonly written as:

\[ \text{P} \ \text{V} = \text{n} \ \text{R} \ \text{T} \]

\[ \uparrow \ \uparrow \ \uparrow \ \uparrow \]

\[ \text{atm} \ \text{L} \ \text{mol} \ \text{K} \]

**Examples:**

1. A sample of H\(_2\) gas has a volume of 8.56 L at a temperature of 0°C and pressure of 1.5 atm. Calculate the moles of gas present.

   \[ P = 1.5 \ \text{atm} \quad \text{n} = \frac{PV}{RT} \]

   \[ V = 8.56 \ \text{L} \quad T = 273 \ \text{K} \]

   \[ n = \frac{(1.5 \ \text{atm})(8.56 \ \text{L})}{(0.0821 \ \text{L atm/mol K})(273 \ \text{K})} = 0.57 \ \text{mol} \]

2. What volume does 40.0 g of N\(_2\) gas occupy at 10°C and 750 mmHg?

   \[ P = \quad \text{n} = \]

   \[ V = ??? \quad T = \quad V = \frac{nRT}{P} = \]

3. A 23.8-L cylinder of contains oxygen at 20.0 °C and 732 mmHg. How many moles of oxygen does it contain?
PARTIAL PRESSURES
DALTON’S LAW

- Many gas samples are mixtures of gases. For example, the air we breathe is a mixture of mostly oxygen and nitrogen gases.

- Since gas particles have no attractions towards one another, each gas in a mixture behaves as if it is present by itself, and is not affected by the other gases present in the mixture.

- In a mixture, each gas exerts a pressure as if it was the only gas present in the container. This pressure is called partial pressure of the gas.

- In a mixture, the sum of all the partial pressures of gases in the mixture is equal to the total pressure of the gas mixture. This is called Dalton’s law of partial pressures.

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

- The partial pressure of each gas in a mixture is proportional to the amount (mol) of gas present in the mixture. For example, in a mixture of gases consisting of 1 mole of nitrogen and 1 mol of hydrogen gas, the partial pressure of each gas is one-half of the total pressure in the container.
PARTIAL PRESSURES

Examples:
1. Two 10-L tanks, one containing propane gas at 300 torr and the other containing methane at 500 torr, are combined in a 10-L tank at the same temperature. What is the total pressure of the gas mixture?

\[ P_{\text{total}} = P_1 + P_2 = 300 \text{ torr} + 500 \text{ torr} = 800 \text{ torr} \]

2. A scuba tank contains a mixture of oxygen and helium gases with total pressure of 7.00 atm. If the partial pressure of oxygen in the tank is 1140 mmHg, what is the partial pressure of helium in the tank?

\[ P_{\text{oxygen}} \text{(in atm)} = 1140 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = \]

\[ P_{\text{total}} = P_{\text{oxygen}} + P_{\text{helium}} \]

\[ P_{\text{helium}} = \]

3. A mixture of gases contains 2.0 mol of O\textsubscript{2} gas and 4.0 mol of N\textsubscript{2} gas with total pressure of 3.0 atm. What is the partial pressure of each gas in the mixture?