Chapter 19
Thermal Properties of Matter
A PowerPoint Presentation by
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Objectives: After finishing this unit, you should be able to:

- Write and apply relationships among pressure, volume, temperature, and quantity of matter for ideal gases undergoing changes of state.
- Define and apply concepts involving molecular mass, moles, and Avogadro’s number.
- Write and apply the general gas law for a particular state of an ideal gas.
The thermodynamic state of a gas is defined by four coordinates:

- Absolute pressure, $P$
- Absolute temperature, $T$
- Volume, $V$
- Mass $m$ or quantity of matter $n$
Boyle’s Law, Charles’ Law, and Gay-Lusac’s Law can be combined into a single formula for an ideal gas that changes from State 1 to another State 2.

\[ \frac{P_1 V_1}{m_1 T_1} = \frac{P_2 V_2}{m_2 T_2} \]

Any Factor that remains constant divides out.
**Example 1:** An auto tire has an gauge pressure of 28 psi in the morning at 20°C. After driving for hours the temperature of air inside the tire is 30°C. What will the gauge read? (Assume 1 atm = 14.7 psi.)

\[
T_1 = 20 + 273 = 293 \text{ K}
\]

\[
T_2 = 30 + 273 = 303 \text{ K}
\]

\[
P_{\text{abs}} = P_{\text{gauge}} + 1 \text{ atm}; \quad P_1 = 28 + 14.7 = 42.7 \text{ psi}
\]

\[
\frac{P_1}{V_1T_1} = \frac{P_2}{V_2T_2}
\]

Same air in tires: \( m_1 = m_2 \)

Same volume of air: \( V_1 = V_2 \)
**Example 1:** What will the gauge read?

**Given:** \( T_1 = 293 \, \text{K}; \quad T_2 = 303 \, \text{K}; \quad P_1 = 42.7 \, \text{psi} \)

\[
\frac{P_1 \sqrt{T_1}}{\sqrt{m_1}} = \frac{P_2 \sqrt{T_2}}{\sqrt{m_2}},
\]

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

\[
P_2 = \frac{P_1 T_2}{T_1} = \frac{(42.7 \, \text{psi})(303 \, \text{K})}{293 \, \text{K}} = 44.2 \, \text{psi}
\]

Gauge pressure is 14.7 psi less than this value:

\[
P_2 = 44.2 \, \text{psi} - 14.7 \, \text{psi} = 29.5 \, \text{psi}
\]
The Composition of Matter

When dealing with gases, it is much more convenient to work with relative masses of atoms.

Building blocks of atoms.

Atoms contain protons and neutrons, which are close to the same mass, surrounded by electrons which are almost negligible by comparison.
To understand relative scales, let’s ignore electrons and compare atoms by total number of nuclear particles.

<table>
<thead>
<tr>
<th>Element</th>
<th>Particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen, H</td>
<td>1 particle</td>
</tr>
<tr>
<td>Helium, He</td>
<td>4 particles</td>
</tr>
<tr>
<td>Lithium, Li</td>
<td>7 particles</td>
</tr>
<tr>
<td>Carbon, C</td>
<td>12 particles</td>
</tr>
<tr>
<td>Oxygen, O</td>
<td>16 particles</td>
</tr>
</tbody>
</table>
The atomic mass of an element is the mass of an atom of the element compared with the mass of an atom of carbon taken as 12 atomic mass units (u).

<table>
<thead>
<tr>
<th>Atomic masses of a few elements:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen, H = 1.0 u</td>
</tr>
<tr>
<td>Helium, He = 4.0 u</td>
</tr>
<tr>
<td>Lithium, Li = 7.0 u</td>
</tr>
<tr>
<td>Beryllium, Be = 9.0 u</td>
</tr>
<tr>
<td>Carbon, C = 12.0 u</td>
</tr>
<tr>
<td>Nitrogen, N = 14.0 u</td>
</tr>
<tr>
<td>Neon, Ne = 20.0 u</td>
</tr>
<tr>
<td>Copper, Cu = 64.0 u</td>
</tr>
</tbody>
</table>
The molecular mass $M$ is the sum of the atomic masses of all the atoms making up the molecule.

Consider Carbon Dioxide (CO$_2$)

<table>
<thead>
<tr>
<th>The molecule has one carbon atom and two oxygen atoms</th>
<th>1 C = 1 x 12 u = 12 u</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>2 O = 2 x 16 u = 32 u</td>
</tr>
<tr>
<td></td>
<td>CO$_2$ = 44 u</td>
</tr>
</tbody>
</table>
Definition of a Mole

One **mole** is that quantity of a substance that contains the same number of particles as there are in **12 g of carbon-12**. \(6.023 \times 10^{23} \text{ particles}\)

- 1 mole of Carbon has a mass of 12 g
- 1 mole of Helium has a mass of 4 g
- 1 mole of Neon has a mass of 20 g
- 1 mole of Hydrogen \((H_2)\) = \(1 + 1 = 2\) g
- 1 mole of Oxygen \((O_2)\) is \(16 + 16 = 32\) g
The unit of molecular mass $M$ is grams per mole.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Molecular Mass (g/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen, H</td>
<td>1.0</td>
</tr>
<tr>
<td>Helium, He</td>
<td>4.0</td>
</tr>
<tr>
<td>Carbon, C</td>
<td>12.0</td>
</tr>
<tr>
<td>Oxygen, O</td>
<td>16.0</td>
</tr>
<tr>
<td>Hydrogen, H$_2$</td>
<td>2.0</td>
</tr>
<tr>
<td>Oxygen, O$_2$</td>
<td>16.0</td>
</tr>
<tr>
<td>Water, H$_2$O</td>
<td>18.0</td>
</tr>
<tr>
<td>Carbon Dioxide, CO$_2$</td>
<td>44.0</td>
</tr>
</tbody>
</table>

Each mole has $6.23 \times 10^{23}$ molecules.
Moles and Number of Molecules

Finding the number of moles \( n \) in a given number of \( N \) molecules:

\[
n = \frac{N}{N_A}
\]

Avogadro’s number: \( N_A = 6.023 \times 10^{23} \text{ particles/mol} \)

**Example 2:** How many moles of any gas will contain \( 20 \times 10^{23} \) molecules?

\[
n = \frac{N}{N_A} = \frac{20 \times 10^{23} \text{ molecules}}{6.023 \times 10^{23} \text{ molecules/mol}} = 3.32 \text{ mol}
\]
Moles and Molecular Mass M

Finding the number of moles \( n \) in a given mass \( m \) of a substance:

\[
 n = \frac{m}{M}
\]

Molecular mass \( M \) is expressed in grams per mole.

Example 3: How many moles are there in 200 g of oxygen gas \( \text{O}_2 \)? (\( M = 32 \text{ g/mol} \))

\[
 n = \frac{m}{M} = \frac{200 \text{ g}}{32 \text{ g/mol}}
\]

\( n = 6.25 \text{ mol} \)
Example 4: What is the mass of a single atom of boron (M = 11 g/mol)?

We are given both a number \( N = 1 \) and a molecular mass \( M = 11 \text{ g/mol} \). Recall that:

\[
n = \frac{N}{N_A} \quad \text{and} \quad n = \frac{m}{M} \quad \Rightarrow \quad \frac{m}{M} = \frac{N}{N_A}
\]

\[
m = \frac{NM}{N_A} = \frac{(1)(11 \text{ g/mol})}{6.023 \times 10^{23} \text{ atoms/mol}}
\]

\[
m = 1.83 \times 10^{-23} \text{ g}
\]
Ideal Gas Law

Substituting moles \( n \) for mass \( m \), we know that:

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

In other words, the ratio \( PV/nT \) is a constant, and if we can find its value, we can work with a single state.

Since a mole of any gas contains the same number of molecules, it will have the same volume for any gas.

Volume of one mole of a gas:

\[ V = 22.4 \text{ L or } 22.4 \times 10^{-3} \text{ m}^3 \]
The Universal Gas Constant $R$

The universal gas constant $R$ is defined as follows:

$$\frac{PV}{nT} = R$$

$$PV = nRT$$

Evaluate for one mole of gas at 1 atm, 273 K, 22.4 L.

$$R = \frac{PV}{nT} = \frac{(101,300 \text{ Pa})(22.4 \times 10^{-3} \text{ m}^3)}{(1 \text{ mol})(273 \text{ K})}$$

$$R = 8.314 \text{ J/mol} \cdot \text{K}$$
Example 5: Two hundred grams of oxygen \((M = 32 \text{ g/mol})\) fills a 2-L tank at a temperature of \(25^\circ\text{C}\). What is the absolute pressure \(P\) of the gas?

\[
T = 25^\circ\text{C} + 273^\circ\text{K} = 298 \text{ K}
\]

\[
V = 2 \text{ L} = 2 \times 10^{-3} \text{ m}^3
\]

\[
PV = nRT
\]

\[
n = \frac{m}{M}
\]

\[
PV = \frac{m}{M} RT
\]

\[
P = \frac{mRT}{MV} = \frac{(200 \text{ g})(8.314 \text{ J/mol} \cdot \text{K})(298 \text{ K})}{(32 \text{ g/mol})(2 \times 10^{-3} \text{ m}^3)}
\]

\[
P = 7.74 \text{ MPa}
\]
Example 6: How many grams of nitrogen gas (M = 28 g/mol) will occupy a volume of 2.4 m³ if the absolute pressure is 220 kPa and the temperature is 300 K?

\[ PV = \frac{m}{M} RT \]

\[ m = \frac{PVM}{RT} = \frac{(220,000 \text{ Pa})(2.4 \text{ m}^3)(28 \text{ g/mol})}{(8.314 \text{ J/mol} \cdot \text{K})(300 \text{ K})} \]

\[ m = 5930 \text{ g} \]

or

\[ m = 5.93 \text{ kg} \]
Summary of Formulas

\[ \frac{P_1 V_1}{m_1 T_1} = \frac{P_2 V_2}{m_2 T_2} \]

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

\[ n = \frac{N}{N_A} \]

\[ n = \frac{m}{M} \]

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

\[ P V = n R T \]
CONCLUSION: Chapter 19
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