Properties of Gases

- Each state of matter has its own properties.
- Gases have unique properties because the distance between the particles of a gas is much greater than the distance between the particles of a liquid or a solid.
  - Although liquids and solids seem very different from each other, both have small intermolecular distances.
- In some ways, gases behave like liquids; in other ways, they have unique properties.
Properties of Gases, *continued*

**Gases are Fluids**

- Gases are considered *fluids*.

- The word *fluid* means “any substance that can flow.”

- Gas particles can flow because they are relatively far apart and therefore are able to move past each other easily.
Properties of Gases, *continued*

Gases Have Low Density

- Gases have much lower densities than liquids and solids do.
- Because of the relatively large distances between gas particles, most of the volume occupied by a gas is empty space.
  - The distance between particles explains why a substance in the liquid or solid state always has a much greater density than the same substance in the gaseous state does.
- The low density of gases also means that gas particles travel relatively long distances before colliding with each other.
Properties of Gases, *continued*

Gases Are Highly Compressible

- Suppose you completely fill a syringe with liquid and try to push the plunger in when the opening is plugged.
  - You cannot make the space the liquid takes up become smaller.
- The space occupied by the gas particles is very small compared with the total volume of the gas.
- Applying a small pressure will move the gas particles closer together and will decrease the volume.
Properties of Gases, continued

Gases Completely Fill a Container

• A solid has a certain shape and volume.

• A liquid has a certain volume but takes the shape of the lower part of its container.

• In contrast, a gas completely fills its container.

• Gas particles are constantly moving at high speeds and are far apart enough that they do not attract each other as much as particles of solids and liquids do.

  • Therefore, a gas expands to fill the entire volume available.
Gas Pressure

• Earth’s atmosphere, commonly known as air, is a mixture of gases: mainly nitrogen and oxygen.

• Because you cannot always feel air, you may have thought of gases as being weightless, but all gases have mass; therefore, they have weight in a gravitational field.

• As gas molecules are pulled toward the surface of Earth, they collide with each other and with the surface of Earth more often. Collisions of gas molecules are what cause air pressure.
Gas Pressure, *continued*

- The density of the air changes when you change altitudes.

- The atmosphere is denser as you move closer to Earth’s surface because the weight of atmospheric gases at any elevation compresses the gases below.
Gas Pressure, *continued*

**Measuring Pressure**

- The scientific definition of *pressure* is “force divided by area.” Pressure may also be defined as the amount of force exerted per unit area of surface.
  - To find pressure, you need to know the force and the area over which that force is exerted.
- The unit of force in SI units is the *newton*, N.
  - One newton is the force that gives an acceleration of 1 m/s^2 to an object whose mass is 1 kg.

\[
1 \text{ newton} = 1 \text{ kg} \times 1 \text{ m/s}^2 = 1 \text{ N}
\]
Gas Pressure, continued

Measuring Pressure, continued

- The SI unit of pressure is the pascal, Pa, which is the force of one newton applied over an area of one square meter.

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

- One pascal is a small unit of pressure. It is the pressure exerted by a layer of water that is 0.102 mm deep over an area of one square meter.
Gas Pressure, \textit{continued}
Measuring Pressure, \textit{continued}

- Atmospheric pressure can be measured by a barometer.

- The atmosphere exerts pressure on the surface of mercury in the dish. This pressure goes through the fluid and up the column of mercury.

- The mercury settles at a point where the pressure exerted downward by its weight equals the pressure exerted by the atmosphere.
Mercury Barometer

- Vacuum
- Pressure exerted by the column of mercury
- Atmospheric pressure
- Surface of mercury
- 760 mm
- $O_2$
- $N_2$
Gas Pressure, continued
Measuring Pressure, continued

• At sea level, the atmosphere keeps the mercury in a barometer at an average height of 760 mm, which is 1 atmosphere, atm.

• One millimeter of mercury is also called a torr, after Evangelista Torricelli, the Italian physicist who invented the barometer.
In studying the effects of changing temperature and pressure on a gas, one will find a standard for comparison useful.

Scientists have specified a set of standard conditions called **standard temperature and pressure**, or STP, which is equal to 0°C and 1 atm.
**Gas Pressure, continued**

**Pressure Units**

<table>
<thead>
<tr>
<th>Unit</th>
<th>Abbreviation</th>
<th>Equivalent number of pascals</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atmosphere</td>
<td>atm</td>
<td>$1 \text{ atm} = 101,325 \ \text{Pa}$</td>
</tr>
<tr>
<td>Bar</td>
<td>bar</td>
<td>$1 \text{ bar} = 100,025 \ \text{Pa}$</td>
</tr>
<tr>
<td>Millimeter of mercury</td>
<td>mm Hg</td>
<td>$1 \text{ mm Hg} = 133.322 \ \text{Pa}$</td>
</tr>
<tr>
<td>Pascal</td>
<td>Pa</td>
<td>1</td>
</tr>
<tr>
<td>Pounds per square inch</td>
<td>psi</td>
<td>$1 \text{ psi} = 6.892 \ 86 \times 10^3 \ \text{Pa}$</td>
</tr>
<tr>
<td>Torr</td>
<td>torr</td>
<td>$1 \text{ torr} = 133.322 \ \text{Pa}$</td>
</tr>
</tbody>
</table>
Converting Pressure Units

Sample Problem A
Convert the pressure of 1.000 atm to millimeters of mercury.
Gas Pressure, *continued*

**Sample Problem A Solution**

1 atmosphere = 101 325 Pa, 1 mm Hg = 133.322 Pa

The conversion factors are \( \frac{101 325 \text{ Pa}}{1 \text{ atm}} \) and \( \frac{1 \text{ mm Hg}}{133.322 \text{ Pa}} \).

\[
1.000 \text{ atm} \times \frac{101 325 \text{ Pa}}{1 \text{ atm}} \times \frac{1 \text{ mm Hg}}{133.322 \text{ Pa}} = 760.0 \text{ mm Hg}
\]
The properties of gases stated earlier are explained on the molecular level in terms of the kinetic-molecular theory. (The kinetic-molecular theory is a model that is used to predict gas behavior.)

- The kinetic-molecular theory states that gas particles are in constant rapid, random motion.
- The theory also states that the particles of a gas are very far apart relative to their size.
- This idea explains the fluidity and compressibility of gases.
The Kinetic-Molecular Theory, continued

- Gas particles can easily move past one another or move closer together because they are farther apart than liquid or solid particles.
- A gas is composed of particles that are in constant motion and that collide with each other and with the walls of their container.
- The pressure exerted by a gas is the result of collisions of the molecules against the walls of the container.
- The average kinetic energy depends on temperature, the higher the temperature, the higher the kinetic energy and the faster the particles are moving.
The Kinetic-Molecular Theory, \textit{continued}

- Compared to the space through which they travel, the particles that make up the gas are so small that their volume can be ignored.
- The individual particles are neither attracted to one another nor do they repel one another.
- When particles collide with one another (or the walls of the container) they bounce rather than stick. These collisions are elastic; if one particle gains kinetic energy another loses kinetic energy so that the average remains constant.
The Kinetic-Molecular Theory, continued

Gas Temperature Is Proportional to Average Kinetic Energy

- The average kinetic energy of random motion is proportional to the absolute temperature, or temperature in kelvins.

- Heat increases the energy of random motion of a gas.

- Not all molecules are traveling at the same speed.
  - As a result of multiple collisions, the molecules have a range of speeds.
  - For a 10°C rise in temperature from STP, the average energy increases about 3%, while the number of very high-energy molecules approximately doubles or triples.
Increasing the temperature of a gas shifts the energy distribution in the direction of greater average kinetic energy.