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.

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

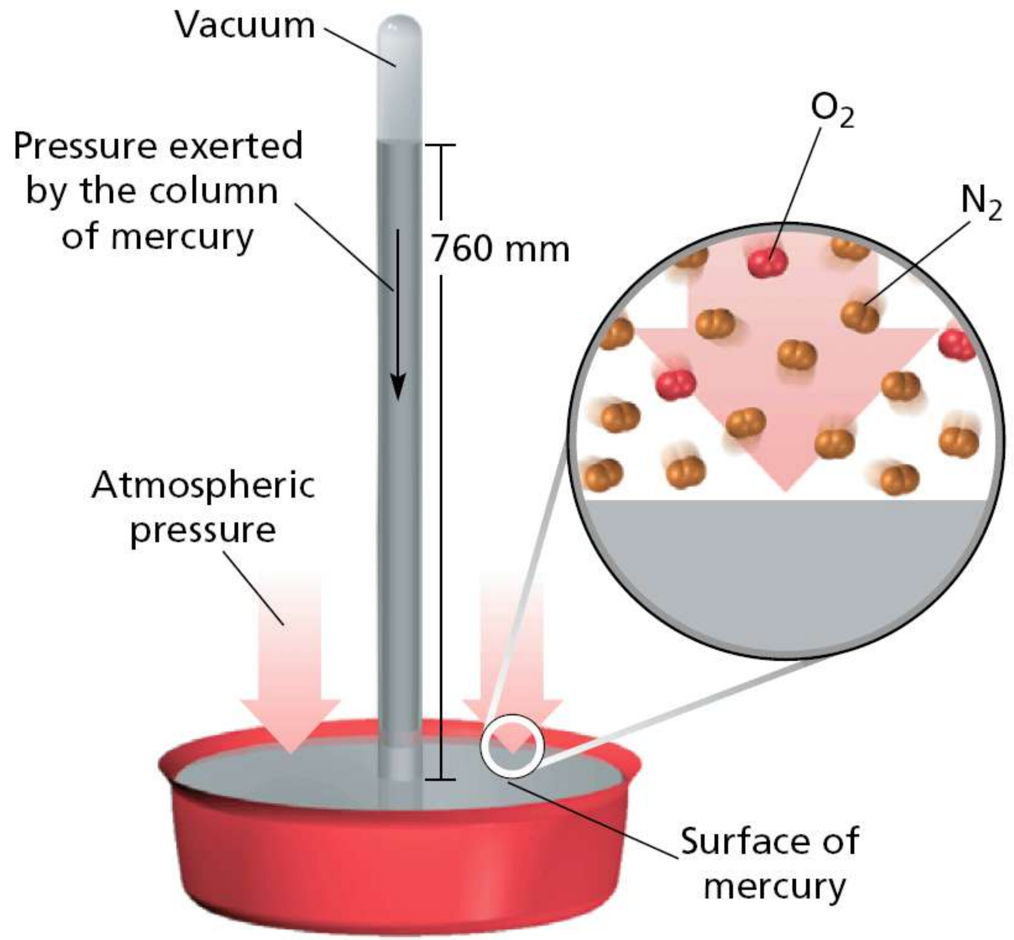
- One pascal is a small unit of pressure. It is the pressure exerted by a layer of water that is 0.102 mm deep over an area of one square meter.

Gas Pressure, *continued*

Measuring Pressure, *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



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.

Gas Pressure, *continued*

Measuring Pressure, *continued*

- 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

Unit	Abbreviation	Equivalent number of pascals
Atmosphere	atm	1 atm = 101 325 Pa
Bar	bar	1 bar = 100 025 Pa
Millimeter of mercury	mm Hg	1 mm Hg = 133.322 Pa
Pascal	Pa	1
Pounds per square inch	psi	1 psi = $6.892\ 86 \times 10^3$ Pa
Torr	torr	1 torr = 133.322 Pa

Converting Presssure 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

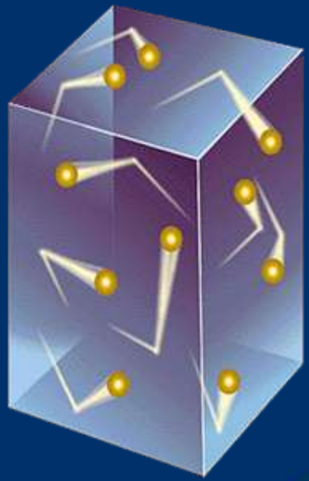
The conversion factors are $\dfrac{101\,325\ \text{Pa}}{1\ \text{atm}}$ and $\dfrac{1\ \text{mm Hg}}{133.322\ \text{Pa}}$.

$$1.000\ \text{atm} \times \frac{101\,325\ \text{Pa}}{1\ \text{atm}} \times \frac{1\ \text{mm Hg}}{133.322\ \text{Pa}} =$$

760.0 mm Hg

The Kinetic-Molecular Theory

- 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, *continued*

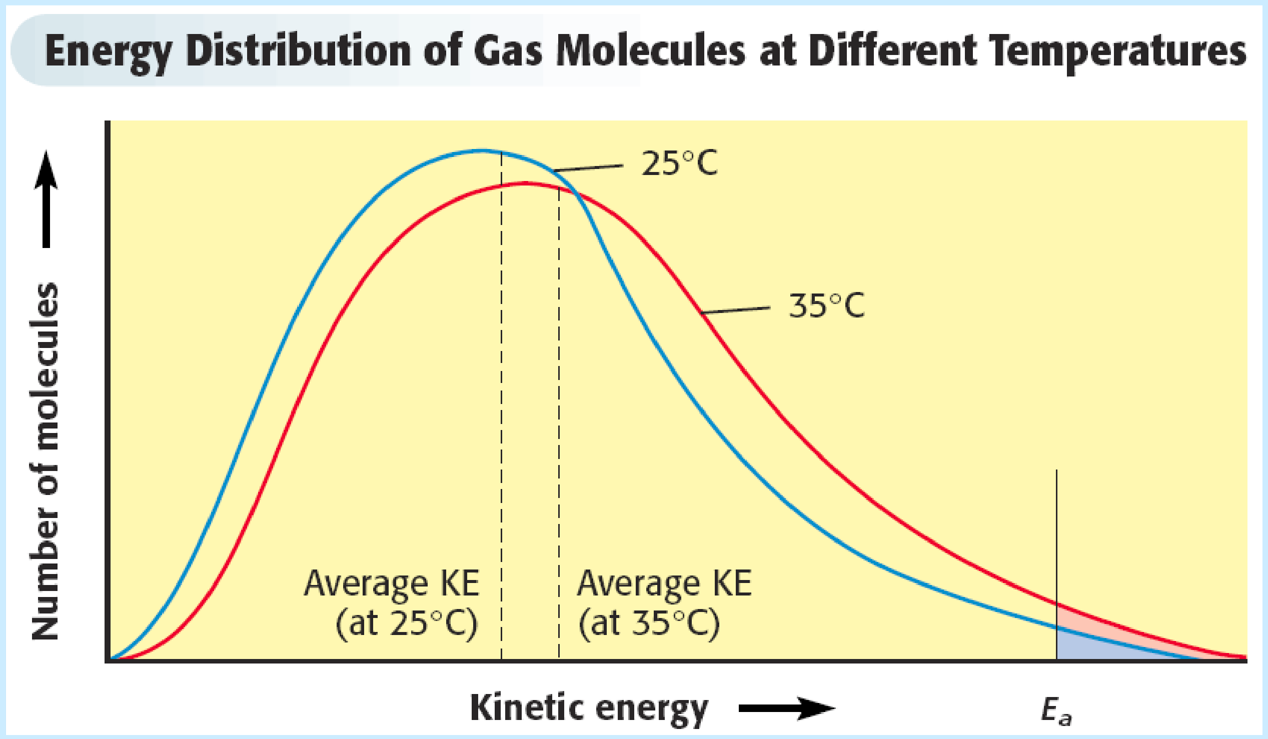
- Compared to the space through which they travel, the particles that make up the gas are so small that their volume can be ignored.
- The individual particles are neither attracted to one another nor do they repel one another.
- When particles collide with one another (or the walls of the container) 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.

The Kinetic-Molecular Theory, *continued*



Increasing the temperature of a gas shifts the energy distribution in the direction of greater average kinetic energy.