10.3: Pressure: The Result of Particle Collisions

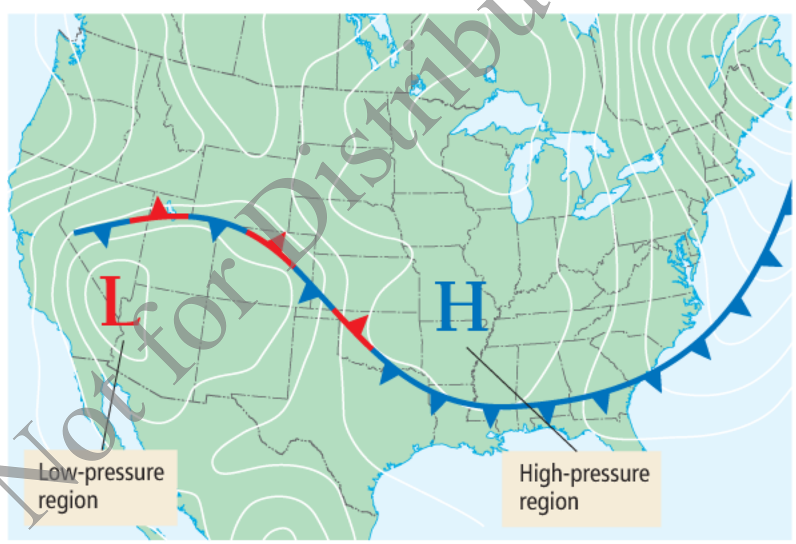
As we just discussed, particles in a gas collide with each other and with the surfaces around them. Each collision exerts only a small force, but when the forces of the many particles are summed, they quickly add up. The result of the constant collisions between the atoms or molecules in a gas and the surfaces around them is pressure. Because of pressure, we can drink from straws, inflate basketballs, and breathe. Variation in pressure in Earth's atmosphere creates wind, and changes in pressure help us to predict weather (Figure 10.4). Pressure is all around us and inside of us. The pressure that a gas sample exerts is the *force* that results from the collisions of gas particles divided by the *area* of the surface with which they collide:

[10.1]

$$\text{pressure} = \frac{\text{force}}{\text{area}} = \frac{F}{A}$$

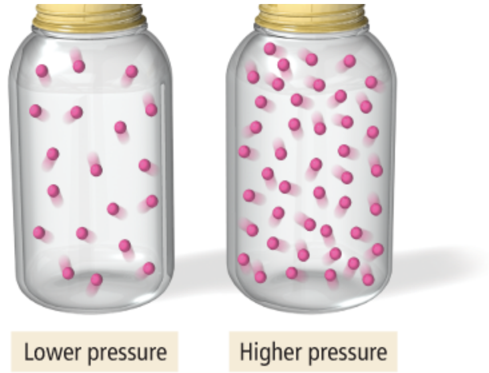
Figure 10.4 **Pressure and Weather**

Pressure variations in Earth's atmosphere create wind and weather. The H in this map indicates a region of high pressure, usually associated with clear weather. The L indicates a region of low pressure, often associated with unstable weather.



The pressure exerted by a gas sample, therefore, depends on the number of gas particles in a given volume—the fewer the gas particles, the lower the force per unit area and the lower the pressure (Figure 10.5). Because the number of gas particles in a given volume generally decreases with increasing altitude, *pressure decreases with increasing altitude.* Above 30,000 ft (about 5.6 mi or 9 km), for example, where most commercial airplanes fly, the pressure is so low that a person could pass out due to a lack of oxygen. For this reason, airplane cabins are artificially pressurized. At 24 mi, the altitude from which Baumgartner jumped, the pressure is less than 1% of the pressure at sea level.

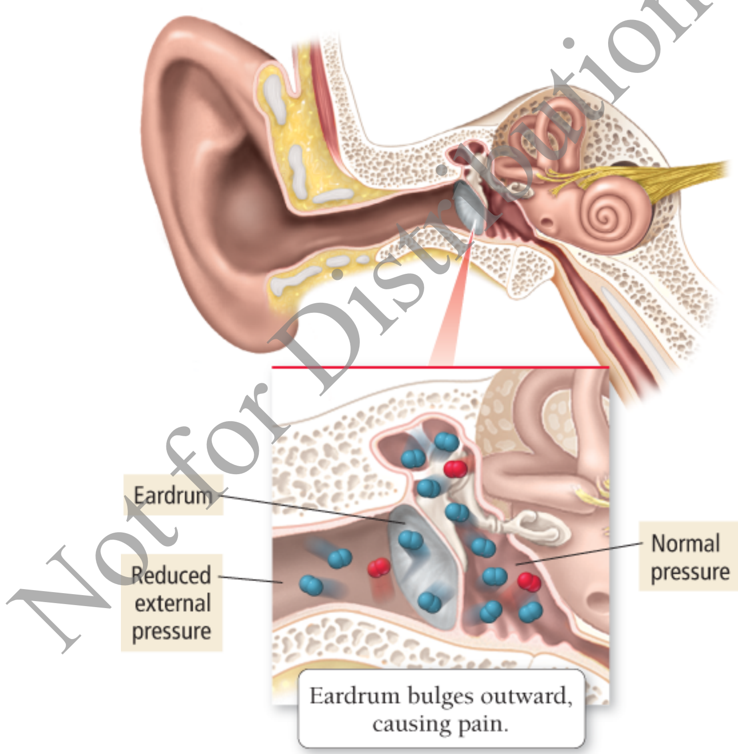
Figure 10.5 **Pressure and Particle Density**



You may sometimes feel the effect of a drop in pressure as a brief pain in your ears. This pain arises within the air-containing cavities in your ear (Figure 10.6). When you ascend in a plane or hike up a mountain, the external pressure (the pressure that surrounds you) drops, while the pressure within your ear cavities (the internal pressure) remains the same. This creates an imbalance—the greater internal pressure forces your eardrum to bulge outward, causing pain. With time, and with the help of a yawn or two, the excess air within your ear's cavities escapes, equalizing the internal and external pressure and relieving the pain.

Figure 10.6 Pressure Imbalance

The discomfort you may feel in your ears upon ascending a mountain is caused by a pressure imbalance between the cavities in your ears and the outside air.

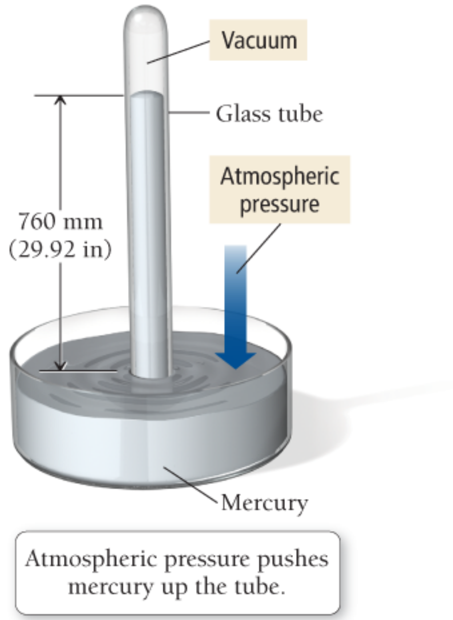


Pressure Units

We measure pressure in several different units. A common unit of pressure, the **millimeter of mercury (mmHg)**, originates from how pressure is measured with a **barometer** (Figure 10.7). A barometer is an evacuated glass tube, the tip of which is submerged in a pool of mercury. Atmospheric pressure on the liquid mercury's surface forces the mercury upward into the evacuated tube. Because mercury is so dense (13.5 times more dense than water), atmospheric pressure can support a column of Hg that is only about 0.760 m or 760 mm (about 30 in) tall. By contrast, atmospheric pressure can support a column of water that is about 10.3 m tall (about 405 in). This makes a column of mercury a convenient way to measure pressure.

Figure 10.7 The Mercury Barometer

Average atmospheric pressure at sea level can support a column of mercury 760 mm in height.



In a barometer, when the atmospheric pressure rises, the height of the mercury column rises as well. Similarly, when atmospheric pressure falls, the height of the mercury column falls. The unit *millimeter of mercury* is often called a **torr**, after the Italian physicist Evangelista Torricelli (1608–1647) who invented the barometer.

$$1 \text{ mmHg} = 1 \text{ torr}$$

A second unit of pressure is the **atmosphere (atm)**, the average pressure at sea level. One atmosphere of pressure pushes a column of mercury to a height of 760 mm; 1 atm and 760 mmHg are equal:

$$1 \text{ atm} = 760 \text{ mmHg}$$

A fully inflated mountain bike tire has a pressure of about 6 atm, and the pressure at the top of Mount Everest is about 0.31 atm.

The SI unit of pressure is the **pascal (Pa)**, defined as 1 newton (N) per square meter:

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

The pascal is a much smaller unit of pressure than the atmosphere:

$$1 \text{ atm} = 101,325 \text{ Pa}$$

Other common units of pressure include the bar, inches of mercury (in Hg), and pounds per square inch (psi):

$$1 \text{ atm} = 1.013 \text{ bar} \quad 1 \text{ atm} = 29.92 \text{ in Hg} \quad 1 \text{ atm} = 14.7 \text{ psi}$$

Table 10.1 summarizes these units.

Table 10.1 Common Units of Pressure

Unit	Abbreviation	Average Air Pressure at Sea Level
Pascal (1 N/m^2)	Pa	101,325 Pa

Pounds per square inch	psi	14.7 psi
Torr (1 mmHg)	torr	760 torr (exact)
Inches of mercury	in Hg	29.92 in Hg
Bar	bar	1.013 bar
Atmosphere	atm	1.00 atm

Example 10.1 Converting between Pressure Units

A cyclist inflates her high-performance road bicycle tire to a total pressure of 132 psi. What is this pressure in mmHg?

SORT The problem gives a pressure in psi and asks you to convert the units to mmHg.

GIVEN: 132 psi

FIND: mmHg

STRATEGIZE Table 10.1 does not have a direct conversion factor between psi and mmHg, but it does provide relationships between both of these units and atmospheres, so you can convert to atm as an intermediate step.

CONCEPTUAL PLAN

$$\boxed{\text{psi}} \xrightarrow[\frac{1\ \text{atm}}{14.7\ \text{psi}}]{} \boxed{\text{atm}} \xrightarrow[\frac{760\ \text{mmHg}}{1\ \text{atm}}]{} \boxed{\text{mmHg}}$$

RELATIONSHIPS USED

$1\ \text{atm} = 14.7\ \text{psi}$
$760\ \text{mmHg} = 1\ \text{atm}$ (both from Table 10.1)

SOLVE Follow the conceptual plan to solve the problem. Begin with 132 psi and use the conversion factors to arrive at the pressure in mmHg.

SOLUTION

$$132\ \cancel{\text{psi}} \times \frac{1\ \cancel{\text{atm}}}{14.7\ \cancel{\text{psi}}} \times \frac{760\ \text{mmHg}}{1\ \cancel{\text{atm}}} = 6.82 \times 10^3\ \text{mmHg}$$

CHECK The units of the answer are correct. The magnitude of the answer, 6.82×10^3 mmHg, is greater than the given pressure in psi. This is reasonable because mmHg is a much smaller unit than psi.

FOR PRACTICE 10.1 Your local weather report announces that the barometric pressure is 30.44 in Hg. Convert this pressure to psi.

FOR MORE PRACTICE 10.1 Convert a pressure of 23.8 in Hg to kPa.

The Manometer: A Way to Measure Pressure in the Laboratory

We can measure the pressure of a gas sample in the laboratory with a **manometer**. A manometer is a U-shaped tube containing a dense liquid, usually mercury. In a manometer such as the one shown in Figure 10.8, one end of the tube is open to atmospheric pressure and the other is attached to a flask containing a gas sample. If the pressure of the gas sample is exactly equal to atmospheric pressure, then the mercury levels on both sides of the tube are the same. If the pressure of the sample is *greater than* atmospheric pressure, the mercury level on the left side of the tube is *higher than* the level on the right (which is the case illustrated in Figure 10.8). If the pressure of the sample is *less than* atmospheric pressure, the mercury level on the left side is *lower than* the level on the right. This type of manometer always measures the pressure of the gas sample relative to atmospheric pressure. The difference in height between the two levels (h in Figure 10.8) is equal to the difference between the sample's pressure and atmospheric pressure. To accurately calculate the absolute pressure of the sample, we also need a barometer to measure atmospheric pressure (which can vary from day to day).

Figure 10.8 The Manometer

A manometer measures the pressure exerted by a sample of gas.