9.1: Gasses and Atmospheric Pressure

In Chapter 2, we learned about the three principle states of matter; solids, liquids and gasses. We explained the properties of the states of matter using the *kinetic molecular theory* (KMT). Substances in the gaseous state, according to the KMT, have enough kinetic energy to break *all* of the attractive forces between the individual gas particles and are therefore free to separate and rapidly move throughout the entire volume of their container. Because there is so much space between the particles in a gas, a gas is *highly compressible*. High compressibility and the ability of gases to take on the shape and volume of its container are two of the important physical properties of gasses.

The gas that we are all most familiar with is the mixture of elements and compounds that we call the “atmosphere”. The air that we breath is mostly nitrogen and oxygen, with much smaller amounts of water vapor, carbon dioxide, noble gasses and the organic compound, methane (Table 9.1).

Table 9.1. Approximate Composition of the Atmosphere

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Gas	Concentration, Parts per Billion	Percentage
N_2	7.8×10^8	78%
O_2	2.0×10^8	20%
H_2O	About $10^6 – 10^7$	< 1%
Ar	9.3×10^6	< 1%
CO_2	3.5×10^5	< 0.05%
Ne	1.8×10^4	*trace*
He	5.2×10^3	*trace*
CH_4	1.6×10^3	*trace*

A gas that is enclosed in a container exerts a *pressure* on the inner walls of that container. This pressure is the result of the countless collisions of the gas particles with the container wall. As each collision occurs, a small amount of energy is transferred, generating a net pressure. Although we are generally unaware of it, the gasses in the atmosphere generate a tremendous pressure on all of us. At sea level, atmospheric pressure is equal to 14.7 pounds per square inch. Putting this in perspective, for a person of average height and build, the total pressure from the atmosphere pressing on their body is about 45,000 pounds! Why aren’t we squashed? Remember, we also have air *inside* our bodies and the pressure from the inside balances the pressure outside, keeping us nice and firm, not squishy!

The proper SI unit for pressure is the Pascal (Pa), where 1 Pa = 1 kg m^{-1} s^{-2}. In chemistry, however, it is more common to measure pressure in terms of atmospheres (atm) where 1 atm is atmospheric pressure at sea level, or 1 atm = 14.7 pounds per square inch (1 atm = 101,325 Pa). Atmospheric pressure is typically measured using a device called a *barometer*. A simple mercury barometer (also called a *Torricelli barometer*, after its inventor) consists of a glass column, about 30 inches high, closed at one end and filled with mercury. The column is inverted and placed in an open, mercury-filled reservoir. The weight of the mercury in the tube causes the column to drop to the point that the mass of the mercury column matches the atmospheric pressure exerted on the mercury in the reservoir. The atmospheric pressure is then read as the *height* of the mercury column. Again, working at sea level, 1 atmosphere is *exactly* equal to a column height of 760 mm of mercury. The units for the conversion are 1 atm = 760 mm Hg, and this is an *exact* relationship with regard to significant figures. The unit *torr* (after Torricelli) is sometimes used in place of mm Hg.