9.1 Gas Pressure

Pressure is the force of gas particles colliding with a surface, thus is illustrated by:

$$P = \frac{F}{A}$$

- Using the units of Newtons per meter squared, the SI unit of pressure is pascal (Pa). It can also be expressed as atm, which is atmospheric pressure at sea level.

Unit Name and Abbreviation	Definition or Relation to Other Unit
pascal (Pa)	1 Pa = 1 N/m^2 recommended IUPAC unit
kilopascal (kPa)	1 kPa = 1000 Pa
pounds per square inch (psi)	air pressure at sea level is ~14.7 psi
atmosphere (atm)	1 atm = 101,325 Pa air pressure at sea level is ~1 atm
bar (bar, or b)	1 bar = 100,000 Pa (exactly) commonly used in meteorology
millibar (mbar, or mb)	1000 mbar = 1 bar
inches of mercury (in. Hg)	1 in. Hg = 3386 Pa used by aviation industry, also some weather reports
torr	$1 \text{ torr} = \frac{1}{760} \text{ atm}$ named after Evangelista Torricelli, inventor of the barometer
millimeters of mercury (mm Hg)	1 mm Hg ~1 torr

- Torr is a conversion from atm to mm Hg (millimeters of mercury) which corresponds to height in the barometer equation.

Barometers function as a measure of hydrostatic pressure which is the pressure exerted by a fluid due to gravity.

$$p = h\rho g$$

where h is the height of the fluid, ρ is the density of the fluid, and g is acceleration due to gravity.

Manometers function similarly to barometers. They are either closed-ended or open-ended; in open-ended systems, the pressure measured is the difference between the gas in the container and atmospheric pressure.

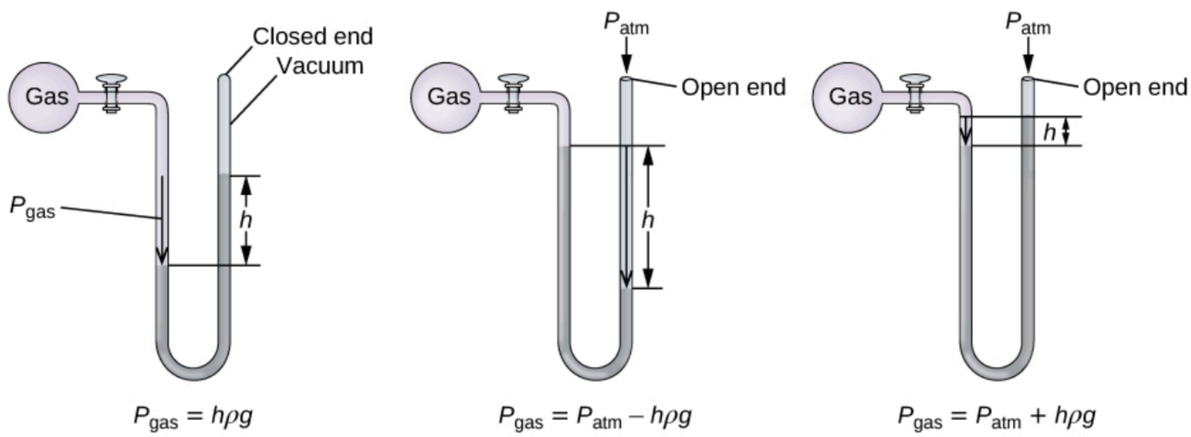

Figure 9.5. A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.

- Use second equation when atmospheric pressure is greater than the pressure of the gas
- Use third equation when atmospheric pressure is less than the pressure of the gas

<u>9.2 Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law</u>
<u>Pressure and Temperature: Amontons's Law</u>
- Pressure and temperature are directly proportional

<u>Volume and Temperature: Charles's Law</u>
- Volume and temperature are directly proportional

<u>Volume and Pressure: Boyle's Law</u>
- Pressure and volume are inversely proportional

<u>Moles of Gas and Volume: Avogadro's Law</u>
- Volume and number of moles are directly proportional

<u>9.3 The Ideal Gas Law</u>
Combining these laws creates the ideal gas law.

$$PV = nRT$$

where P is the pressure of a gas, V is its volume, n is the number of moles of the gas, T is its temperature on the kelvin scale, and R is a constant called the ideal gas constant or the universal gas constant. The units used to express pressure, volume, and temperature will determine the proper form of the gas constant as required by dimensional analysis, the most commonly encountered values being 0.08206 L atm mol^{-1} K^{-1} and 8.314 kPa L mol^{-1} K^{-1}.

- **PAY CLOSE ATTENTION TO THE UNITS**

The Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

using units of atm, L, and K. Both sets of conditions are equal to the product of n R (where n = the number of moles of the gas and R is the ideal gas law constant).

Standard Conditions of Temperature and Pressure

STP - standard temperature and pressure 273.15 K (0℃) and 1 atm. At standard temperature, conditions:

- Volume: 22.4L

9.4 Stoichiometry of Gaseous Substances, Mixtures, and Reactions

Density of a Gas

$$4.\ m/V = \rho = \frac{P\mathscr{M}}{RT}$$

Molar Mass of a Gas

$$\mathscr{M} = \frac{\text{grams of substance}}{\text{moles of substance}} = \frac{m}{n}$$

$$\mathscr{M} = \frac{mRT}{PV}$$

The Pressure of a Mixture of Gases: Dalton's Law

Gasses, unless reacting with each other, exert the same pressure whether alone or with other gasses. **Dalton's law of partial pressures**: The total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases.

$$P_{Total} = P_A + P_B + P_C + \ldots = \Sigma_i P_i$$

In the equation PTotal is the total pressure of a mixture of gases, PA is the partial pressure of gas A; PB is the partial pressure of gas B; PC is the partial pressure of gas C; and so on.

Mole fraction is the component of a certain substance's moles in relation to the mixture that it is in.

$$P_A = X_A \times P_{Total} \quad \text{where} \quad X_A = \frac{n_A}{n_{Total}}$$

where PA, XA, and nA are the partial pressure, mole fraction, and number of moles of gas A, respectively, and nTotal is the number of moles of all components in the mixture.

9.5 The Kinetic Molecular Theory

- Gasses are in continuous motion, only changing direction when colliding
- Gasses are neglibibly small compared to distance between them
- Pressure of container is due to collisions between container and particles within it
- Collisions are elastic (no energy loss) because gas particles exert neither attractive nor repulsive force
- Average kinetic energy of gas is proportional to temperature

Molecular Velocities and Kinetic Energy

Kinetic energy is expressed as:

$$KE = \frac{1}{2}mu^2$$

The room mean squared velocity is the square root of the average squared velocity per particles:

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u^2{}_1 + u^2{}_2 + u^2{}_3 + u^2{}_4 + \ldots}{n}}$$

- Represents average velocity

The average kinetic energy can be expressed as:

$$KE_{avg} = \frac{1}{2} M u^2{}_{rms}$$

where M is the molar mass in units of kg/mol, to get the units of KEavg to be in joules (J).

Average KE is also described as:

$$KE_{avg} = \frac{3}{2} RT$$

- Temperature is directly proportional to average kinetic energy of particles

Which allows us to deduce that average velocity can also be expressed in terms of:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

At a given temperature, all gas particles have the same kinetic energy. However lighter particles have greater u_{rms} values, while heavier have lower values.

<u>Ideal Gases</u>

Ideal gases behave under two assumptions:

- Gasses are point particles (zero volume)
- Gasses lack IM attraction and repulsion