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Introduction

Scientific laws are critical to our ability to describe the behavior of matter in all phases - solids, liquids, and gases. However, characterizing gases can be challenging compared to solids and liquids, since they are highly influenced by changes in physical conditions. The basic gas laws describe the relationship between pressure, volume, temperature, and amount of gas, providing a way to determine unknown properties from known ones.

Boyle's Law

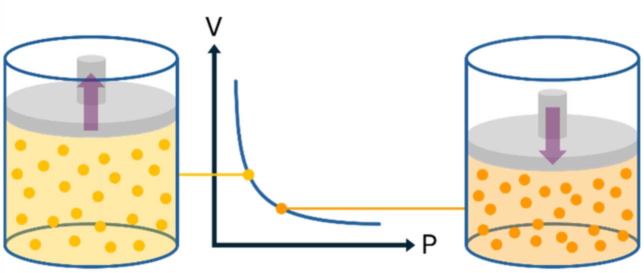
The relationship between pressure and volume is given by Boyle's Law, named after Robert Boyle, who was the first to publish the information in 1662. Boyle's Law states that the volume that a gas occupies is inversely proportional to its absolute pressure, given a closed system at constant temperature.

$$V \propto \frac{1}{P}$$

$$PV = Constant$$

Where:

- V = Volume (ft^3)
- P = Pressure (psia)



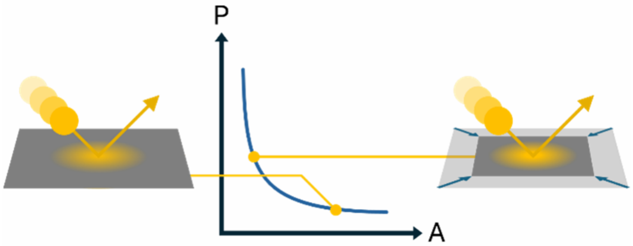
This correlation can be better understood using the definitions of temperature and pressure. Temperature is a measure of the kinetic energy in a system, which, for a gas, is the movement of the individual gas particles. Within a closed container, the force of the collisions between the particles and interior wall exert pressure on the surface.

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

Where:

- P = Pressure (psia)
- F = Force (lbf)
- A = Area (ft^2)

Since the temperature is constant and the kinetic energy/force of the collisions remains unchanged, reducing the system volume decreases the surface area, increasing the pressure (and vice versa).



A common use of Boyle's law is to determine gas volume at standard conditions. For example, given pipeline gas at 600 psig, in 100 ft of 6" pipe, the process to convert the volume of gas from actual cubic feet (ACF) to standard cubic feet (SCF) is as follows:

$$P_1V_1 = P_2V_2$$

Where:

- (1) = Actual Conditions
- (2) = Standard Conditions, at 60 °F, 14.73 psia

$$V_1 = \frac{\pi D^2}{4} \times L = \pi \frac{6^2}{4 \times 144} \times 100 = 19.63\ ACF$$

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(614.73)(19.63)}{(14.73)} = 819.22\ SCF$$

Charles' Law

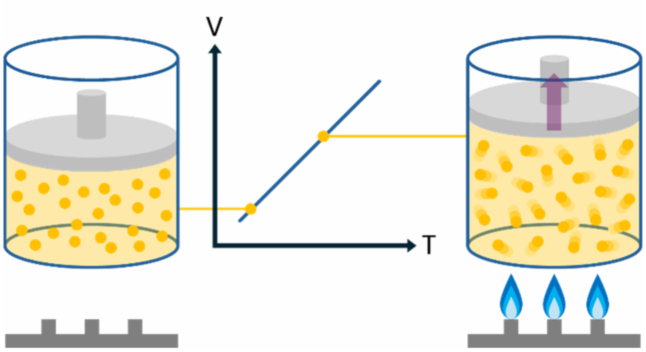
The relationship between temperature and volume is given by Charles' Law, named after Jacques Charles for his work from the 1780s. Charles' Law states that in a closed system at constant pressure, the volume that a gas occupies is directly proportional to its temperature.

$$V \propto T$$

$$\frac{V}{T} = Constant$$

Where:

- V = Volume (ft^3)
- T = Temperature (°R)



As noted previously, the temperature is a measure of the kinetic energy of the gas. As the temperature increases, the amount of surface collisions increases, but since pressure is held constant, the associated surface area/volume of the gas must increase.

Gay-Lussac's Law

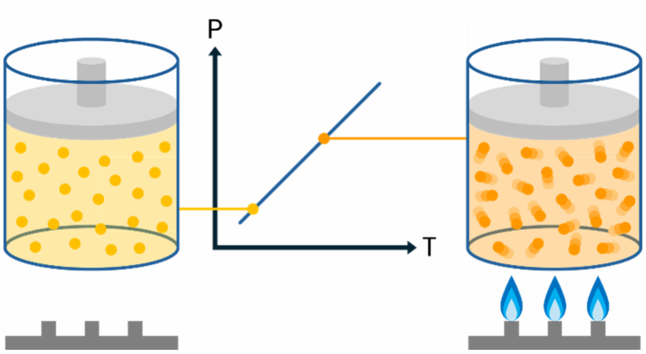
The relationship between pressure and temperature is given by Gay-Lussac's Law, named for Joseph-Lewis Gay-Lussac. Gay-Lussac's Law states that absolute pressure is directly proportional to the temperature in a closed system with a fixed volume.

$$P \propto T$$

$$\frac{P}{T} = Constant$$

Where:

- P = Pressure (psia)
- T = Temperature (°R)



When volume is held constant, changes in temperature will vary the number of molecular collisions with the interior surface of the container, resulting in a corresponding change in pressure.

Avogadro's Law

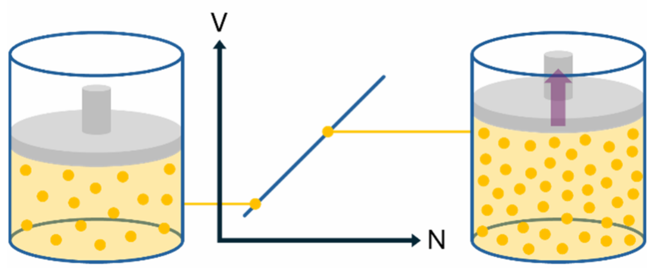
Boyle's, Charles', and Gay-Lussac's Laws all assume a fixed amount of gas in a closed system. Avogadro's Law (named for Amedeo Avogadro, 1811) provides the relationship between volume and the amount of gas, stating that the two are directly proportional at the same temperature and pressure.

$$V \propto N$$

$$\frac{V}{N} = Constant$$

Where:

- V = Volume (ft^3)
- N = Number of Molecules



Another way to describe Avogadro's Law is that a fixed amount of gas at the same temperature and pressure will always occupy the same volume.

Ideal Gas Law

The various gas laws provide a useful way to compare the change in one gas property to another, but assume that the other gas properties are held constant. For more dynamic systems, Boyle's, Charles', Gay-Lussac's, and Avogadro's Laws can be combined into a single expression known as the Ideal Gas Law.

$$\frac{PV}{NT} = Constant$$

$$PV = NT \times Constant$$

$$PV = Nk_bT$$

Where:

- P = Pressure (psia)
- V = Volume (ft^3)
- N = Number of Molecules
- T = Temperature (°R)
- k_b = Boltzmann's Constant (ft·lb/°R)

The constant in this form is known as the Boltzmann's Constant and typically noted as k_b. However, when quantifying the amount of gas in a system, it is more common to use the number of moles instead of the number of molecules. Therefore, the Ideal Gas Law is most often given in its molar form.

$$PV = nRT$$

Where:

- P = Pressure (psia)
- V = Volume (ft^3)
- n = Number of Moles
- R = Universal Gas Constant (psi·ft^3/lbmol·°R)
- T = Temperature (°R)

Supercompressibility/Real Gas Behavior

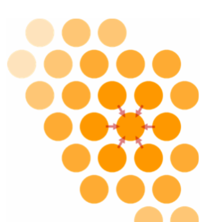
The Ideal Gas Law is known as such because it accurately describes the behavior of a hypothetical, "ideal" gas. An ideal gas differs from real gases in that molecular size and interparticle interactions are ignored. Despite this generalization, the Ideal Gas Law is still useful since many real gases exhibit "ideal" behavior under certain conditions. Generally, the behavior of a real gas will align with an ideal gas at high temperature and low pressure.

However, since gas systems cannot always operate under these ideal conditions, the term supercompressibility is used to describe how real gas behavior deviates from ideal. At low temperatures, the decrease in molecular motion allows the attractive forces between individual gas molecules to play a factor, causing the gas to occupy a smaller volume than an ideal gas. At high pressures, the reduced proximity of the individual gas molecules allows the repulsive forces between molecules as well as the molecular volume become significant, resulting in a larger volume of real gas compared to an ideal gas.

Low Temperatures

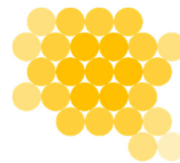
Ideal Gas

Real Gas

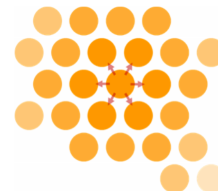


High Pressures

Ideal Gas



Real Gas



The compressibility factor (Z) is added to the Ideal Gas Law to account for these forces. When attractive forces dominate, the gas is more compressible than an ideal gas, and Z < 1. When repulsive forces are significant, the gas is less compressible than an ideal gas, and Z > 1.

$$PV = ZnRT$$

Where:

- P = Pressure (psia)
- V = Volume (ft^3)
- Z = Compressibility Factor
- n = Number of Moles
- R = Universal Gas Constant (psi·ft^3/lbmol·°R)
- T = Temperature (°R)

While the compressibility factor can be estimated mathematically, the most accurate methods are fairly complex, requiring gas composition data and iterative calculations. Examples of industry standard methods for estimating the compressibility factor include AGA Report Number 8 and the older AGA NX19.

Conclusion

These basic gas laws provide simple mathematical relationships between the fundamental properties of gases, describing their behavior as operating conditions change. The laws form the basis for the design of gas measurement systems, since accurate measurement of pressurized gas flows must compensate for variations in both pressure and temperature.