9.S: The Gaseous State (Summary)

- Gasses are *compressible* because there is so much space between individual gas particles. Energy transferred from the collision of gas particles with their container exerts a *gas pressure*. In chemistry, we typically measure gas pressure using units of *atmospheres* (atm) or in *mm Hg* (also referred to as torr). One atmosphere of pressure equals exactly 760 mm Hg.
- The volume of a gas varies *inversely* with the applied pressure; the greater the pressure, the smaller the volume. This is relationship is referred to as *Boyles's Law*. For a two-state system where the number of moles of gas and the temperature remain constant, Boyle's Law can be expressed as

$$P_1V_1 = P_2V_2$$

- The volume of a gas varies *directly* with the absolute temperature; the higher the temperature, the larger the volume. This is relationship is referred to as *Charles's Law*. For a two-state system where the number of moles of gas and the pressure remain constant, Charles's Law can be expressed as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- In this equation, the *absolute temperature* in Kelvin (K) must be used. Kelvin is defined as (degrees centigrade + 273.15). Zero degrees Kelvin is referred to as "absolute zero" and it is the temperature at which (theoretically) all molecular motion would cease.
- The volume of a gas varies *directly* with the number of moles of the gas that are present; the greater the number of moles, the larger the volume. This is relationship is referred to as *Avogadro's Law*. For a two-state system where the temperature and the pressure of a gas remain constant, Avogadro's Law can be expressed as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

- Because the volume of a gas varies directly with the number of moles of the gas that are present and with the absolute temperature (in Kelvin), and inversely with the pressure, the gas laws can be combined into a single proportionality;

$$V \propto \left(\frac{nT}{P}\right)$$

- This proportionality can be converted to an equality by inserting the proportionality constant *R* (the universal gas constant), where R = 0.082057 L atm mol^{-1} K^{-1}, and can be re-written as:

$$V = R\left(\frac{nT}{P}\right) \quad or \quad PV = nRT$$

- This is referred to as the Ideal *Gas Law* and is valid for most gasses at low concentrations. For a two-state system where the identity of the gas does not change, the Ideal Gas Law can be expressed as:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

- The gas constant *R*, is calculated based on the experimental observation that *exactly* one mole of any gas at *exactly* 1 atm and at 0 ˚C (273 K) has a volume of **22.414 L**. The conditions, 1 atm and 0 ˚C, are called standard temperature and pressure, or *STP*.
- The ideal gas laws allow a quantitative analysis of whole spectrum of chemical reactions involving gasses. When you are approaching these problems, remember to *first* decide on the *class* of the problem:
  - If it is a "single state" problem (a gas is produced at a single, given, set of conditions), then you want to use $PV = nRT$.
  - If it is a "two state" problem (a gas is changed from one set of conditions to another) you want to use

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

  - If the volume of gas is quoted at STP, you can quickly convert this volume into moles with by dividing by 22.414 L mol^{-1}.