10.4: Using Molarity in Equilibrium Calculations

As we have pointed out several times in the preceding sections, the Ideal Gas Laws (Chapter 10) tell us that the **partial pressure** of a gas and the **molar concentration** of that gas are directly proportional. We can show this simply by beginning with the combined gas law:

$$P_{gas}V = nRT$$

If we divide both sides by the volume, *V*, and state that *V* must be expressed in liters, the right side of the equation now contains the term (n/V_{liters}). Realizing that the number of moles of gas (*n*) divided by the volume in liters is equal to *molarity*, *M*, this expression can be re-written as:

\[P_{gas}=MRT \nonumber \]

Using this expression, molar concentrations can easily be substituted for partial pressures, and *visa versa*.

Exercise 10.4.1

1. For the reaction shown below, if the molar concentrations of SO_3, NO and SO_2 are all 0.100 *M*, what is the equilibrium concentration of NO_2?
2. For the reaction between carbon monoxide and chlorine to form phosgene, the equilibrium constant calculated from partial pressures is $K = 0.20$. How does this value relate to the equilibrium constant, K_C, under the same conditions, calculated from molar concentrations?

$$\mathbf{CO\ (g) + Cl_2\ (g) \rightleftarrows COCl_2\ (g)}$$