

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:

$$P_{gas} V = nRT$$

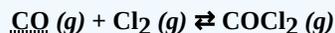
If we divide both sides by the volume, V , and state that V must be expressed in liters, the right side of the equation now contains the term ($\frac{n}{V_{liters}}$). Realizing that the number of moles of gas (n) divided by the volume in liters is equal to *molarity*, M , this expression can be re-written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

$$P_{gas} = MRT$$

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?



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