Chemical equilibrium is the state where the concentrations of all reactants and products remain constant with time.

Consider the following reaction:

\[ \text{H}_2\text{O} + \text{CO} \rightarrow \text{H}_2 + \text{CO}_2 \]

Suppose you were to start the reaction with some amount of each reactant (and no \( \text{H}_2 \) or \( \text{CO}_2 \)). As the reaction proceeds, the concentrations of CO and \( \text{H}_2\text{O} \) will decrease as they are converted into \( \text{CO}_2 \) and \( \text{H}_2 \). Eventually, the reaction reaches a point at which the concentrations of reactants and products stop changing. Reactant is still present (see the non-zero concentration of reactants in the illustration below?), but it looks like the reaction has stopped.

\[ \text{In fact, the reaction has NOT stopped.} \]

Reactions occur when molecules collide with one another. As the concentrations of CO and \( \text{H}_2\text{O} \) decrease, it becomes less likely for these two molecules to collide, even though they are both still present. As the concentration of the products increases, the probability of a collision between \( \text{H}_2 \) and \( \text{CO}_2 \) increases. In this case, the reaction can occur in both the forward (as written) and reverse (products converting to reactants) directions. At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction.
The equilibrium constant, $K$, is used to determine the relative concentrations of products and reactants at equilibrium. According to the law of mass action, if a chemical reaction has the form
\[ aA + bB \leftrightarrow cC + dD \]
then the equilibrium constant can be expressed as
\[
K = \frac{[C]^c[D]^d}{[A]^a[B]^b}
\]

1. In an experiment conducted at 74°C, the equilibrium concentrations of reactants and products for the equation shown below were $[CO] = 1.2 \times 10^{-2}$ M, $[Cl_2] = 0.054$ M and $[COCl_2] = 0.14$ M.

\[ CO(g) + Cl_2(g) \leftrightarrow COCl_2(g) \]

a) What is the equilibrium expression for this reaction?

b) Calculate the value of the equilibrium constant

2. One student performed an experiment in which nitrogen and hydrogen gas were mixed together to form ammonia. Equilibrium concentrations of the three species were $[NH_3] = 0.157$ M, $[N_2] = 0.921$ M and $[H_2] = 0.763$ M.

\[ N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) \]

a) What is the equilibrium expression for this reaction?

b) Calculate the value of the equilibrium constant
3. Another student performed the same reaction (to form ammonia) with the same equilibrium constants, but wrote the balanced equation as
\[
\frac{1}{2} N_2(g) + \frac{3}{2} H_2(g) \leftrightarrow NH_3(g)
\]

a) What is the equilibrium expression when the reaction is written this way?

b) Calculate the value of the equilibrium constant

c) How are the equilibrium constants from questions 2b and 3b related?

4. The equilibrium constant for the following reaction at 100°C is 7.10. The equilibrium concentrations of the reactants were found to be \([Br_2] = 2.3 \times 10^{-3} \text{ M}\) and \([Cl_2] = 1.2 \times 10^{-2} \text{ M}\).

\[
Br_2(g) + Cl_2(g) \leftrightarrow 2BrCl(g)
\]

a) What is the equilibrium concentration of BrCl?

b) If the reaction were performed in a 5.0 L container, how many moles of BrCl would be produced?

5. Which of the above reactions had the highest value of \(K\)? What does this indicate about the extent of the reaction?
When using gases, the equilibrium expression can also be written using pressures of the reactants and products rather than concentrations. Remember that concentration has units of moles (n) per liter (V). The ideal gas law can be rearranged to the form

\[ P = \frac{nRT}{V} = \left(\frac{n}{V}\right)RT = CRT \]

where \(C\) represents concentration. Using the last example again, we can derive an equilibrium expression involving pressures:

\[ Br_2(g) + Cl_2(g) \leftrightarrow 2BrCl(g) \]

\[ K = \frac{[BrCl]^2}{[Br_2][Cl_2]} = \frac{\left(\frac{P_{BrCl}}{RT}\right)^2}{\left(\frac{P_{Br_2}}{RT}\right)\left(\frac{P_{Cl_2}}{RT}\right)} = \frac{P_{BrCl}^2}{P_{Br_2}P_{Cl_2}} \left(\frac{RT}{P_{Br_2}}\right)^2 = \frac{P_{BrCl}^2}{P_{Br_2}P_{Cl_2}} = K_p. \]

This equilibrium expression is called \(K_p\) because it uses pressures instead of concentrations. \(K_p\) has the same form as \(K\), but with pressures.

6. The following equilibrium pressures at a certain temperature were observed for the reaction shown below. Calculate \(K_p\) for this reaction.

\[ 2NO_2(g) \leftrightarrow 2NO(g) + O_2(g) \]

\[ P_{NO_2} = 0.55 \text{ atm} \quad P_{NO} = 6.5 \times 10^{-5} \text{ atm} \quad P_{O_2} = 4.5 \times 10^{-5} \text{ atm} \]

In some cases, \(K = K_p\). This is only true if the number of moles of gas on the product side of the reaction are equal to the number of moles of gas on the reactant side.

In general, \(K_p = K(RT)^{\Delta n}\), where \(\Delta n = \text{moles of product gas} - \text{moles of reactant gas}\) and \(R = 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}\)

7. Calculate \(K\) for the reaction in problem 6 at 100°C
All of the examples considered at this point have involved gases. **If a chemical reaction involves solids or pure liquids, they are not included in the equilibrium expression.** The fundamental reason for this is that the concentration of a solid or pure liquid cannot change.

The equilibrium expression for the reaction \( CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g) \) would be written simply as \( \text{K} = [CO_2] \) because the CaCO\(_3\) and CaO are solids and must be excluded.

8. Write the equilibrium expression for the reaction

\[
CO_2(g) + C(s) \leftrightarrow 2CO(g)
\]

9. Write the expressions for \( K \) and \( K_p \) for the following process:

Solid phosphorus pentachloride decomposes to liquid phosphorus trichloride and chlorine gas