1.10.3: The Equilibrium Constant - A Measure of How Far a Reaction Goes

Learning Objectives

- Write equilibrium constant expressions.
- Use equilibrium constant expressions to solve for unknown concentrations.
- Use known concentrations to solve for the equilibrium constants.
- Explain what the value of K means in terms of relative concentrations of reactants and products.

In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

The Equilibrium Constant

Equilibrium reactions are those that do not go to completion, but are in a state where the reactants are reacting to yield products and the products are reacting to produce reactants. In a reaction at equilibrium, the equilibrium concentrations of all reactants and products can be measured. The **equilibrium constant (K)** is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants. Sometimes, subscripts are added to the equilibrium constant symbol K, such as K_{eq}, K_{c}, K_{p}, K_{a}, K_{b}, and K_{sp}. These are all equilibrium constants and are subscripted to indicate special types of equilibrium reactions.

There are some rules about writing equilibrium constant expressions that need to be learned:

1. Concentrations of products are multiplied on the top of the expression. Concentrations of reactants are multiplied together on the bottom.
2. Coefficients in the equation become exponents in the equilibrium constant expression.
3. Solids, liquids, and solvents are assigned a value of 1, so their concentrations do not affect the value of K.

Example 1.10.3.1

Write the equilibrium constant expression for:

$$\text{CO}\,(g) + 3\text{H}_2\,(g) \rightleftharpoons \text{CH}_4\,(g) + \text{H}_2\text{O}\,(g)$$

Solution

$$K = \frac{\left[\text{CH}_4\right]\left[\text{H}_2\text{O}\right]}{\left[\text{CO}\right]\left[\text{H}_2\right]^3}$$

*Note that the coefficients become exponents. Also, note that the concentrations of products in the numerator are *multiplied*. The same is true of the reactants in the denominator.

Example 1.10.3.2

Write the equilibrium constant expression for:

$$2\text{TiCl}_3\,(s) + 2\text{HCl}\,(g) \rightleftharpoons 2\text{TiCl}_4\,(s) + \text{H}_2\,(g)$$

Solution

$$K = \frac{\left[\text{H}_2\right]}{\left[\text{HCl}\right]^2}$$

*Note that the solids have a value of 1, and multiplying or dividing by 1 does not change the value of K.

> ✓ Example 1.10.3.2
>
> Write the equilibrium constant expression for:
>
> $$P_4\ (s) + 6Cl_2\ (g) \rightleftharpoons 4PCl_3\ (s)$$
>
> **Solution**
>
> $$K = \frac{1}{[Cl_2]^6}$$
>
> *Note that the only product is a solid, which is defined to have a value of 1. That leaves just 1 on top in the numerator.

> ✓ Example 1.10.3.3
>
> Write the equilibrium constant expression for:
>
> $$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$$
>
> **Solution**
>
> $$K = [H^+]\ [OH^-]$$
>
> *Note that the water is the solvent, and thus has a value of 1. Dividing by 1 does not change the value of K.

Equilibrium Constant Expressions

The equilibrium constant value is the ratio of the concentrations of the products over the reactants. This means that we can use the value of K to predict whether there are more products or reactants at equilibrium for a given reaction. What can the value of K_{eq} tell us about a reaction?

- **If K_{eq} is very large**, the concentration of the products is much greater than the concentration of the reactants. The reaction essentially "goes to completion"; all, or most of, the reactants are used up to form the products.
- **If K_{eq} is very small**, the concentration of the reactants is much greater than the concentration of the products. The reaction does not occur to any great extent—most of the reactants remain unchanged, and there are few products produced.
- **When K_{eq} is not very large or very small** (close to a value of 1) then roughly equal amounts of reactants and products are present at equilibrium.

Here are some examples to consider:

Reaction	Chemical Equations	Equilibrium Constant
the decomposition of ozone, O_3	$2\,O_3(g) \rightleftharpoons 3\,O_2(g)$	$K_{eq} = 2.0 \times 10^{57}$
	K_{eq} is very large, indicating that mostly O_2 is present in an equilibrium system, with very little O_3.	
production of nitrogen monoxide	$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$	$K_{eq} = 1.0 \times 10^{-25}$
	Very little NO is produced by this reaction; N_2 and O_2 do not react readily to produce NO (this is lucky for us—otherwise we would have little oxygen to breath in our atmosphere!).	
reaction of carbon monoxide and water	$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$	$K_{eq} = 5.09$ (at 700 K)
	The concentrations of the reactants are very close to the concentrations of the products at equilibrium.	

If the equilibrium constant is 1 or nearly 1, it indicates that the molarities of the reactants and products are about the same. If the equilibrium constant value is a large number, like 100, or a very large number, like 1×10^{15}, it indicates that the products (numerator) are a great deal larger than the reactants. This means that at equilibrium, the great majority of the material is in the form of products and it is said that the "products are strongly favored". If the equilibrium constant is small, like 0.10, or very small, like 1×10^{-12}, it indicates that the reactants are much larger than the products and the reactants are strongly favored. With large K

values, most of the material at equilibrium is in the form of products and with small K values, most of the material at equilibrium is in the form of the reactants.

The equilibrium constant expression is an equation that we can use to solve for K or for the concentration of a reactant or product.

Example 1.10.3.4

Determine the value of K for the reaction

$$\mathrm{SO_2}(g) + \mathrm{NO_2}(g) \rightleftharpoons \mathrm{SO_3}(g) + \mathrm{NO}(g)$$

when the equilibrium concentrations are: $[\mathrm{SO_2}] = 1.20\ \mathrm{M}$, $[\mathrm{NO_2}] = 0.60\ \mathrm{M}$, $[\mathrm{NO}] = 1.6\ \mathrm{M}$, and $[\mathrm{SO_3}] = 2.2\ \mathrm{M}$.

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{[\mathrm{SO_3}][\mathrm{NO}]}{[\mathrm{SO_2}][\mathrm{NO_2}]}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)(1.6)}{(1.20)(0.60)} = 4.9$$

Example 1.10.3.5

Consider the following reaction:

$$\mathrm{CO}(g) + \mathrm{H_2O}(g) \rightleftharpoons \mathrm{H_2}(g) + \mathrm{CO_2}(g)$$

with $K = 1.34$. If the $[\mathrm{H_2O}] = 0.100\ \mathrm{M}$, $[\mathrm{H_2}] = 0.100\ \mathrm{M}$, and $[\mathrm{CO_2}] = 0.100\ \mathrm{M}$ at equilibrium, what is the equilibrium concentration of CO?

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{[\mathrm{H_2}][\mathrm{CO_2}]}{[\mathrm{CO}][\mathrm{H_2O}]}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)(0.100)}{[\mathrm{CO}](0.100)}$$

Solving for $[\mathrm{CO}]$, we get: $[\mathrm{CO}] = 0.0746\ \mathrm{M}$

Summary

- The equilibrium constant expression is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants.
- If the value of K is greater than 1, the products in the reaction are favored. If the value of K is less than 1, the reactants in the reaction are favored. If K is equal to 1, neither reactants nor products are favored.

Vocabulary

- **Equilibrium constant** (K) - A mathematical ratio that shows the concentrations of the products divided by the concentrations of the reactants.

Contributions & Attributions

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- Modified by Tom Neils (Grand Rapids Community College)
