The reaction quotient \( (Q) \) measures the relative amounts of products and reactants present during a reaction at a particular point in time. The reaction quotient aids in figuring out which direction a reaction is likely to proceed, given either the pressures or the concentrations of the reactants and the products. The \( (Q) \) value can be compared to the Equilibrium Constant, \( (K) \), to determine the direction of the reaction that is taking place.

### K vs. Q

The main difference between \( (K) \) and \( (Q) \) is that \( (K) \) describes a reaction that is at equilibrium, whereas \( (Q) \) describes a reaction that is not at equilibrium. To determine \( (Q) \), the concentrations of the reactants and products must be known. For a given general chemical equation:

\[
[\text{aA + bB } \rightleftharpoons \text{ cC + dD}] \tag{1}
\]

the Q equation is written by multiplying the activities (which are approximated by concentrations) for the species of the products and dividing by the activities of the reactants. If any component in the reaction has a coefficient, indicated above with lower case letters, the concentration is raised to the power of the coefficient. \( (Q) \) for the above equation is therefore:

\[
(Q_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}) \tag{2}
\]

Note

This equation only shows components in the gaseous or aqueous states. Each pure liquid or solid has an activity of one and can be functionally omitted. Equilibrium constants really contain a ratio of concentrations (actual concentration divided by the reference concentration that defines the standard state). Because the standard state for concentrations is usually chosen to be 1 mol/L, it is not written out in practical applications. Hence, the ratio does not contain units.

A comparison of \( (Q) \) with \( (K) \) indicates which way the reaction shifts and which side of the reaction is favored:

- If \( (Q > K) \), then the reaction favors the reactants. This means that in the \( (Q) \) equation, the ratio of the numerator (the concentration or pressure of the products) to the denominator (the concentration or pressure of the reactants) is larger than that for \( (K) \), indicating that more products are present than there would be at equilibrium. Because reactions always tend toward equilibrium (Le Châtelier's Principle), the reaction produces more reactants from the excess products, therefore causing the system to shift to the LEFT. This allows the system to reach equilibrium.

- If \( (Q < K) \), then the reaction favors the products. The ratio of products to reactants is less than that for the system at equilibrium—the concentration or the pressure of the reactants is greater than the concentration or pressure of the products. Because the reaction tends toward reach equilibrium, the system shifts to the RIGHT to make more products.

- If \( (Q = K) \), then the reaction is already at equilibrium. There is no tendency to form more reactants or more products at this point. No side is favored and no shift occurs.

### Activity

Another important concept that is used in the calculation of the reaction quotient is called an activity. For example, consider the \( (Q) \) equation for this acid/base reaction:
\[\text{CH}_3\text{CH}_2\text{CO}_2\text{H}^{(aq)} + \text{H}_2\text{O}^{(l)} \rightleftharpoons \text{H}_3\text{O}^+^{(aq)} + \text{CH}_3\text{CH}_2\text{CO}_2^-^{(aq)} \tag{3}\]

The \(Q\) equation is written as the concentrations of the products divided by the concentrations of the reactants, but only including components in the gaseous or aqueous states and omitting pure liquid or solid states. The \(Q\) equation for this example is the following:

\[Q_c = \dfrac{[\text{H}_3\text{O}^+^{(aq)}][\text{CH}_3\text{CH}_2\text{CO}_2^-^{(aq)}]}{[\text{CH}_3\text{CH}_2\text{CO}_2\text{H}^{(aq)}]} \tag{4}\]

Example 1

What is the \(Q\) value for this equation? Which direction will the reaction shift?

Given: \(\text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2(g)\)

\(K_c = 1.0\)

- \([\text{CO}_2(g)] = 2.0\ \text{M}\)
- \([\text{H}_2(g)] = 2.0\ \text{M}\)
- \([\text{CO}(g)] = 1.0\ \text{M}\)
- \([\text{H}_2\text{O}(g)] = 1.0\ \text{M}\)

**SOLUTION**

Step 1: Write the \(Q\) formula:

\[Q_c = \dfrac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}\]

Step 2: Plug in given concentration values:

\[Q_c = \dfrac{(2.0)(2.0)}{(1.0)(1.0)}\]

\(Q = 4.0\)

Step 3: Compare \(Q\) to \(K\):

Because \(4.0 > 1.0\), then \(Q > K\) and the reaction shifts left toward the reactants.

Answer: \(Q = 4.0\) and the reaction shifts left.

Example 2

Find the value of \(Q\) and determine which side of the reaction is favored.

Given \(K = 0.5\)

\(\text{HCl}(g) + \text{NaOH}(aq) \rightleftharpoons \text{NaCl}(aq) + \text{H}_2\text{O}(l)\)

\([\text{HCl}] = 3.2\)
\[ [\text{NaOH}] = 4.3 \]  
\[ [\text{NaCl}] = 6 \]

**SOLUTION**

Step 1: Write the Q formula. Because the activity of a liquid is 1, we can omit the water component in the equation.

\[ Q_c = \dfrac{[\text{NaCl (aq)}]}{[\text{HCl (g)}][\text{NaOH (aq)}]} \]

Step 2: Plug in given concentrations into the \( Q \) formula:

\[ Q_c = \dfrac{[6]}{[3.2][4.3]} \]

Step 3: Calculate using the given concentrations:

\[ Q = 0.436 \]

Step 4: Compare \( Q \) to \( K \). The \( Q \) value, 0.436, is less than the given \( K \) value of 0.5, so \( Q < K \).

Because \( Q < K \), the reaction is not at equilibrium and proceeds to the products side to reach dynamic equilibrium once again.

Answer: \( Q = 0.436 \) and the reaction favors the products.

**Example 3**

Given the equation, \( \text{N}_2(g) + 3\text{H}_2(aq) \rightleftharpoons 2\text{NH}_3(g) \) find \( Q \) and determine which direction the reaction will shift in order to reach equilibrium.

Given:
\[ [\text{N}_2] = 0.04 \text{M} \]
\[ [\text{H}_2] = 0.09 \text{M} \]
\[ K = 0.040 \]

**SOLUTION**

Step 1: Write the \( Q \) formula:

\[ Q_c = \dfrac{[\text{NH}_3(g)]^2}{[\text{N}_2(g)][\text{H}_2(g)]^3} \]

Step 2: Plug in values. Because the concentrations for \( [\text{N}_2] \) and \( [\text{H}_2] \) were given, they can be inserted directly into the equation. However, no concentration value was given for \( \text{NH}_3 \) so the concentration is assumed to be 0.

\[ Q_c = \dfrac{(0)^2}{(0.04)(0.09)^3} \]

Step 3: Solve for \( Q \):

\[ Q = 0 \]

Step 4: Compare \( Q \) to \( K \). Because \( K = 0.040 \) and \( Q = 0 \), \( K > Q \) and the reaction will shift right to regain equilibrium.
Answer: \(Q=0\), the reaction shifts right.

Outside Links

- [http://web.mst.edu/~gbert/LeChatelier/RxQuotient.html](http://web.mst.edu/~gbert/LeChatelier/RxQuotient.html)

References


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