Pearson

15.7: The Reaction Quotient: Predicting the Direction of Change

When the reactants of a chemical reaction mix, they generally react to form products—we say that the reaction proceeds to the right (toward the products). The amount of products formed when equilibrium is reached depends on the magnitude of the equilibrium constant, as we have seen. However, what if a reaction mixture not at equilibrium contains both reactants *and products*? Can we predict the direction of change for such a mixture?

To gauge the progress of a reaction relative to equilibrium, we use a quantity called the *reaction quotient*. The definition of the reaction quotient takes the same form as the definition of the equilibrium constant, except that the reaction need not be at equilibrium. So, for the general reaction:

$$a\mathrm{A} + b\mathrm{B} \rightleftharpoons c\mathrm{C} + d\mathrm{D}$$

we define the **reaction quotient** ($\mathbf{Q_c}$) as the ratio—at any point in the reaction—of the concentrations of the products raised to their stoichiometric coefficients divided by the concentrations of the reactants raised to their stoichiometric coefficients. For gases with amounts measured in atmospheres, the reaction quotient uses the partial pressures in place of concentrations and is called ($\mathbf{Q_P}$):

$$\mathrm{Q_c} = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b} \quad \mathrm{Q}_p = \frac{P_\mathrm{C}^c P_\mathrm{D}^d}{P_\mathrm{A}^a P_\mathrm{B}^b}$$

The difference between the reaction quotient and the equilibrium constant is that, at a given temperature, the equilibrium constant has only one value and it specifies the relative amounts of reactants and products *at equilibrium*. The reaction quotient, by contrast, depends on the current state of the reaction and has many different values as the reaction proceeds. For example, in a reaction mixture containing only reactants, the reaction quotient is zero ($\mathscr{Q}_\mathrm{c} = 0$):

$$Q_\mathrm{c} = \frac{[0]^c[0]^d}{[\mathrm{A}]^a[\mathrm{B}]^b} = 0$$

In a reaction mixture containing only products, the reaction quotient is infinite ($Q_\mathrm{c} = \infty$):

$$Q_\mathrm{c} = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[0]^a[0]^b} = \infty$$

In a reaction mixture containing both reactants and products, each at a concentration of 1 M, the reaction quotient is one ($Q_\mathrm{c} = 1$):

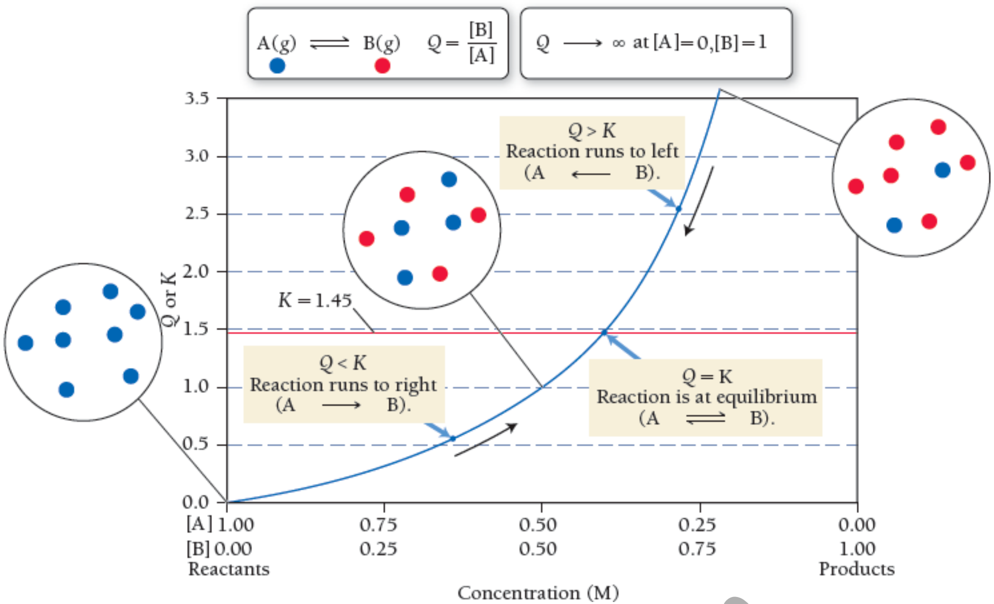
$$Q_\mathrm{c} = \frac{[1]^c[1]^d}{[1]^a[1]^b} = 1$$

The reaction quotient is useful because *the value of Q relative to K is a measure of the progress of the reaction toward equilibrium. At equilibrium, the reaction quotient is equal to the equilibrium constant*. Figure 15.6 shows a plot of Q as a function of the concentrations of A and B for the simple reaction $\mathrm{A}(g) \rightleftharpoons \mathrm{B}(g)$, which has an equilibrium constant of $K = 1.45$. The following points are representative of three possible conditions:

Figure 15.6 ***Q*, *K*, and the Direction of a Reaction**

The graph shows a plot of Q as a function of the concentrations of the reactants and products in a simple reaction $\mathrm{A} \rightleftharpoons \mathrm{B}$ in which $K = 1.45$ and the sum of the reactant and product concentrations is 1 M. The far left of the graph represents pure reactant, and the far right represents pure product. The midpoint of the graph represents an equal mixture of A and B. When Q is less than K, the reaction moves in the forward direction $(A \rightarrow B)$. When Q is greater than K, the reaction moves in the reverse direction $(A \leftarrow B)$. When Q is equal to K, the reaction is at equilibrium.

Q, K, and the Direction of a Reaction



Q	K	Predicted Direction of Reaction
0.55	1.45	To the right (toward products)
2.55	1.45	To the left (toward reactants)
1.45	1.45	No change (at equilibrium)

For the first set of values in the table, Q is less than K and must therefore get larger as the reaction proceeds toward equilibrium. Q becomes larger as the reactant concentration decreases and the product concentration increases—the reaction proceeds to the right. For the second set of values, Q is greater than K and must therefore get smaller as the reaction proceeds toward equilibrium. Q gets smaller as the reactant concentration increases and the product concentration decreases—the reaction proceeds to the left. In the third set of values, $Q = K$, implying that the reaction is at equilibrium—the reaction will not proceed in either direction.

Summarizing Direction of Change Predictions:

The reaction quotient (Q) relative to the equilibrium constant (K) is a measure of the progress of a reaction toward equilibrium.

- $Q < K$ Reaction goes to the right (toward products).
- $Q > K$ Reaction goes to the left (toward reactants).
- $Q = K$ Reaction is at equilibrium.

Example 15.7 Predicting the Direction of a Reaction by Comparing Q and K

Consider the reaction and its equilibrium constant.

$$I_2(g) + Cl_2(g) \rightleftharpoons 2\,ICl(g) \quad K_p = 81.9$$

A reaction mixture contains $P_{I_2} = 0.114$ atm, $P_{Cl_2} = 0.102$ atm, and $P_{ICl} = 0.355$ atm. Is the reaction mixture at equilibrium? If not, in which direction will the reaction proceed?

SOLUTION To determine the progress of the reaction relative to the equilibrium state, first calculate Q.

SOLUTION To determine the progress of the reaction relative to the equilibrium state, first calculate Q.

$$\begin{aligned} Q_P &= \left(\frac{P_{ICl}^2}{P_{I_2}P_{Cl_2}}\right) \\ &= \frac{(0.355)^2}{(0.114)(0.102)} \\ &= 10.8 \end{aligned}$$

Compare Q to K.

$$Q_P = 10.8;\ K_P = 81.9$$

Since $Q_P < K_P$, the reaction is not at equilibrium and will proceed to the right.

FOR PRACTICE 15.7 Consider the reaction and its equilibrium constant.

$$N_2O_4(g) \rightleftharpoons 2\ NO_2(g) \quad K_c = 5.85 \times 10^{-3} \text{ (at some temperature)}$$

A reaction mixture contains $[NO_2] = 0.0255$ M and $[N_2O_4] = 0.0331$ M. Calculate Q_c and determine the direction in which the reaction will proceed.

Conceptual Connection 15.5 *Q* and *K*

Interactive