

18.6: Spontaneity and Equilibrium

Learning Objectives

- To know the relationship between free energy and the equilibrium constant.

We have identified three criteria for whether a given reaction will occur spontaneously: $\Delta S_{\text{univ}} > 0$, $\Delta G_{\text{sys}} < 0$, and the relative magnitude of the reaction quotient Q versus the equilibrium constant K . (For more information on the reaction quotient and the equilibrium constant, see Chapter 15.) Recall that if $Q < K$, then the reaction proceeds spontaneously to the right as written, resulting in the net conversion of reactants to products. Conversely, if $Q > K$, then the reaction proceeds spontaneously to the left as written, resulting in the net conversion of products to reactants. If $Q = K$, then the system is at equilibrium, and no net reaction occurs. Table 18.6.1 summarizes these criteria and their relative values for spontaneous, nonspontaneous, and equilibrium processes. Because all three criteria are assessing the same thing—the spontaneity of the process—it would be most surprising indeed if they were not related. The relationship between ΔS_{univ} and ΔG_{sys} was described in Section 18.5. In this section, we explore the relationship between the standard free energy of reaction (ΔG°) and the equilibrium constant (K).

Table 18.6.1 Criteria for the Spontaneity of a Process as Written

| Spontaneous | Equilibrium | Nonspontaneous* |
|------------------------------|------------------------------|------------------------------|
| $\Delta S_{\text{univ}} > 0$ | $\Delta S_{\text{univ}} = 0$ | $\Delta S_{\text{univ}} < 0$ |
| $\Delta G_{\text{sys}} < 0$ | $\Delta G_{\text{sys}} = 0$ | $\Delta G_{\text{sys}} > 0$ |
| $Q < K$ | $Q = K$ | $Q > K$ |

*Spontaneous in the *reverse* direction.

Free Energy and the Equilibrium Constant

Because ΔH° and ΔS° determine the magnitude of ΔG° (Equation 18.5.5), and because K is a measure of the ratio of the concentrations of products to the concentrations of reactants, we should be able to express K in terms of ΔG° and vice versa. As you learned in Section 18.5, ΔG is equal to the maximum amount of work a system can perform on its surroundings while undergoing a spontaneous change. For a reversible process that does not involve external work, we can express the change in free energy in terms of volume, pressure, entropy, and temperature, thereby eliminating ΔH from the equation for ΔG . Using higher math, the general relationship can be shown as follows:

$$\Delta G = V\Delta P - S\Delta T \quad (18.6.1)$$

If a reaction is carried out at constant temperature ($\Delta T = 0$), then Equation 18.29 simplifies to

$$\Delta G = V\Delta P \quad (18.6.2)$$

Under normal conditions, the pressure dependence of free energy is not important for solids and liquids because of their small molar volumes. For reactions that involve gases, however, the effect of pressure on free energy is very important. What Equation 18.6.2 tells us is how the free energy changes as the pressure changes at constant temperature.

Assuming ideal gas behavior, we can replace the V in Equation 18.6.2 by nRT/P (where n is the number of moles of gas and R is the ideal gas constant) and express ΔG in terms of the initial and final pressures (P_i and P_f , respectively) as in Equation 18.4.1.

$$\Delta G = \left(\frac{nRT}{P}\right) \Delta P = nRT \frac{\Delta P}{P} = nRT \ln \left(\frac{P_f}{P_i}\right) \quad (18.6.4)$$

To actually show the final step one has to consider infinitesimal changes in both the free energy and the pressure and then integrate both sides of the equation using calculus, which is beyond the scope of this course. If the initial state is the standard state with $P_i = 1$ atm, then the change in free energy of a substance when going from the standard state to any other state with a pressure P can be written as follows:

$$G - G^\circ = nRT \ln P$$

This can be rearranged as follows:

$$G = G^\circ + nRT \ln P \quad (18.6.5)$$

As you will soon discover, Equation 18.6.5 allows us to relate ΔG° and K_p . Any relationship that is true for K_p must also be true for K because K_p and K are simply different ways of expressing the equilibrium constant using different units.

Let's consider the following hypothetical reaction, in which all the reactants and the products are ideal gases and the lowercase letters correspond to the stoichiometric coefficients for the various species:



Because the free-energy change for a reaction is the difference between the sum of the free energies of the products and the reactants, we can write the following expression for ΔG :

$$\Delta G = \sum mG_{\text{products}} - \sum mG_{\text{reactants}} = (cG_C + dG_D) - (aG_A + bG_B) \quad (18.6.7)$$

Substituting Equation 18.6.5 for each term into Equation 18.6.7,

$$\Delta G = [(cG_C + cRT \ln P_C) + (dG_D + dRT \ln P_D)] - [(aG_A + aRT \ln P_A) + (bG_B + bRT \ln P_B)]$$

Combining terms gives the following relationship between ΔG and the reaction quotient Q :

$$\Delta G = \Delta G^\circ + RT \ln \left(\frac{P_C^c P_D^d}{P_A^a P_B^b} \right) = \Delta G^\circ + RT \ln Q \quad (18.6.8)$$

where ΔG° indicates that all reactants and products are in their standard states. In Chapter 15, you learned that for gases $Q = K_p$ at equilibrium, and as you've learned in this chapter, $\Delta G = 0$ for a system at equilibrium. Therefore, we can describe the relationship between ΔG° and K_p for gases as follows:

$$0 = \Delta G^\circ + RT \ln K_p \quad (18.6.9)$$

$$\Delta G^\circ = -RT \ln K_p$$

If the products and reactants are in their standard states and $\Delta G^\circ < 0$, then $K_p > 1$, and products are favored over reactants. Conversely, if $\Delta G^\circ > 0$, then $K_p < 1$, and reactants are favored over products. If $\Delta G^\circ = 0$, then $K_p = 1$, and neither reactants nor products are favored: the system is at equilibrium.

Note the Pattern

For a spontaneous process under standard conditions, K_{eq} and K_p are greater than 1.

Example 18.6.1

In Example 18.5.3, we calculated that $\Delta G^\circ = -32.7$ kJ/mol of N_2 for the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$. This calculation was for the reaction under standard conditions—that is, with all gases present at a partial pressure of 1 atm and a temperature of 25°C. Calculate ΔG for the same reaction under the following nonstandard conditions: $P(\text{N}_2) = 2.00$ atm, $P(\text{H}_2) = 7.00$ atm, $P(\text{NH}_3) = 0.021$ atm, and $T = 100^\circ\text{C}$. Does the reaction favor products or reactants?

Given: balanced chemical equation, partial pressure of each species, temperature, and ΔG°

Asked for: whether products or reactants are favored

Strategy:

A Using the values given and Equation 18.6.8, calculate Q .

B Substitute the values of ΔG° and Q into Equation 18.6.8 to obtain ΔG for the reaction under nonstandard conditions.

Solution:

A The relationship between ΔG° and ΔG under nonstandard conditions is given in Equation 18.6.8. Substituting the partial pressures given, we can calculate Q :

$$Q = \frac{P_{\text{NH}_3}^2}{P_{\text{N}_2} P_{\text{H}_2}^3} = \frac{(0.021)^2}{(2.00)(7.00)^3} = 6.4 \times 10^{-7}$$

B Substituting the values of ΔG° and Q into Equation 18.6.8,

$$\Delta G = \Delta G^\circ + RT \ln Q = -32.7 \text{ kJ} + \left[(8.314 \text{ J/K}) (373 \text{ K}) (1 \text{ kJ}/1000 \text{ J}) \ln (6.4 \times 10^{-7}) \right]$$

$$= -32.7 \text{ kJ} + (-44 \text{ kJ}) = -77 \text{ kJ/mol } N_2$$

Because $\Delta G < 0$ and $Q < 1.0$, the reaction is spontaneous to the right as written, so products are favored over reactants.

Exercise

Calculate ΔG for the reaction of nitric oxide with oxygen to give nitrogen dioxide under these conditions: $T = 50^\circ\text{C}$, $P(\text{NO}) = 0.0100 \text{ atm}$, $P(\text{O}_2) = 0.200 \text{ atm}$, and $P(\text{NO}_2) = 1.00 \times 10^{-4} \text{ atm}$. The value of ΔG° for this reaction is -72.5 kJ/mol of O_2 . Are products or reactants favored?

Answer: -92.9 kJ/mol of O_2 ; the reaction is spontaneous to the right as written, so products are favored.

Example 18.6.2

Calculate K_p for the reaction of H_2 with N_2 to give NH_3 at 25°C . As calculated in Example 18.5.3, ΔG° for this reaction is -32.7 kJ/mol of N_2 .

Given: balanced chemical equation from Example 10, ΔG° , and temperature

Asked for: K_p

Strategy:

Substitute values for ΔG° and T (in kelvins) into Equation 18.6.9 to calculate K_p , the equilibrium constant for the formation of ammonia.

Solution:

In Example 18.5.3, we used tabulated values of ΔG° to calculate ΔG° for this reaction (-32.7 kJ/mol of N_2). For equilibrium conditions, rearranging Equation 18.6.9,

$$\begin{aligned}\Delta G^\circ &= -RT \ln K_p \\ \frac{-\Delta G^\circ}{RT} &= \ln K_p\end{aligned}$$

Inserting the value of ΔG° and the temperature ($25^\circ\text{C} = 298 \text{ K}$) into this equation,

$$\begin{aligned}\ln K_p &= -\frac{(-32.7 \text{ kJ}) (1000 \text{ J/kJ})}{8.314 \text{ J/K} (298 \text{ K})} = 13.2 \\ K_p &= 5.4 \times 10^5\end{aligned}$$

Thus the equilibrium constant for the formation of ammonia at room temperature is favorable. As we saw in [Chapter 15](#), however, the *rate* at which the reaction occurs at room temperature is too slow to be useful.

Exercise

Calculate K_p for the reaction of NO with O_2 to give NO_2 at 25°C . As calculated in the exercise in Example 10, ΔG° for this reaction is -70.5 kJ/mol of O_2 .

Answer: 2.2×10^{12}

Although K_p is defined in terms of the partial pressures of the reactants and the products, the equilibrium constant K is defined in terms of the concentrations of the reactants and the products. We described the relationship between the numerical magnitude of K_p and K in [Chapter 15](#) and showed that they are related:

$$K_p = K(RT)^{\Delta n} \quad (18.6.10)$$

where Δn is the number of moles of gaseous product minus the number of moles of gaseous reactant. For reactions that involve only solutions, liquids, and solids, $\Delta n = 0$, so $K_p = K$. For all reactions that do not involve a change in the number of moles of gas present, the relationship in [Equation 18.36](#) can be written in a more general form:

$$\Delta G^\circ = -RT \ln K \quad (18.6.11)$$

Only when a reaction results in a net production or consumption of gases is it necessary to correct [Equation 18.38](#) for the difference between K_p and K . Although we typically use concentrations or pressures in our equilibrium calculations, recall that equilibrium

constants are generally expressed as unitless numbers because of the use of activities or fugacities in precise thermodynamic work. Systems that contain gases at high pressures or concentrated solutions that deviate substantially from ideal behavior require the use of fugacities or activities, respectively.

Combining Equation 18.26 and Equation 18.6.11 provides insight into how the components of ΔG° influence the magnitude of the equilibrium constant:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -RT \ln K \quad (18.6.12)$$

Notice that K becomes larger as ΔS° becomes more positive, indicating that the magnitude of the equilibrium constant is directly influenced by the tendency of a system to move toward maximum disorder. Moreover, K increases as ΔH° decreases. Thus the magnitude of the equilibrium constant is also directly influenced by the tendency of a system to seek the lowest energy state possible.

Note the Pattern

The magnitude of the equilibrium constant is directly influenced by the tendency of a system to move toward maximum disorder and seek the lowest energy state possible.

Temperature Dependence of the Equilibrium Constant

The fact that ΔG° and K are related provides us with another explanation of why equilibrium constants are temperature dependent. This relationship is shown explicitly in Equation 18.6.12, which can be rearranged as follows:

$$\ln K = \frac{\Delta H^\circ}{RT} + \frac{\Delta S^\circ}{R} \quad (18.6.13)$$

Assuming ΔH° and ΔS° are temperature independent, for an exothermic reaction ($\Delta H^\circ < 0$), the magnitude of K decreases with increasing temperature, whereas for an endothermic reaction ($\Delta H^\circ > 0$), the magnitude of K increases with increasing temperature. The quantitative relationship expressed in Equation 18.6.13 agrees with the qualitative predictions made by applying Le Chatelier's principle, which we discussed in Chapter 15. Because heat is produced in an exothermic reaction, adding heat (by increasing the temperature) will shift the equilibrium to the left, favoring the reactants and decreasing the magnitude of K . Conversely, because heat is consumed in an endothermic reaction, adding heat will shift the equilibrium to the right, favoring the products and increasing the magnitude of K . Equation 18.6.13 also shows that the *magnitude* of ΔH° dictates how rapidly K changes as a function of temperature. In contrast, the magnitude and sign of ΔS° affect the magnitude of K but not its temperature dependence.

If we know the value of K at a given temperature and the value of ΔH° for a reaction, we can estimate the value of K at any other temperature, even in the absence of information on ΔS° . Suppose, for example, that K_1 and K_2 are the equilibrium constants for a reaction at temperatures T_1 and T_2 , respectively. Applying Equation 18.6.13 gives the following relationship at each temperature:

$$\begin{aligned} \ln K_1 &= \frac{\Delta H^\circ}{RT_1} + \frac{\Delta S^\circ}{R} \\ \ln K_2 &= \frac{\Delta H^\circ}{RT_2} + \frac{\Delta S^\circ}{R} \end{aligned}$$

Subtracting $\ln K_1$ from $\ln K_2$,

$$\ln K_2 - \ln K_1 = \ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) \quad (18.6.13)$$

Thus calculating ΔH° from tabulated enthalpies of formation and measuring the equilibrium constant at one temperature (K_1) allow us to calculate the value of the equilibrium constant at any other temperature (K_2), assuming that ΔH° and ΔS° are independent of temperature.

Example 18.6.3

The equilibrium constant for the formation of NH_3 from H_2 and N_2 at 25°C was calculated to be $K_p = 5.4 \times 10^5$ in Example 13. What is K_p at 500°C ? (Use the data from Example 10.)

Given: balanced chemical equation, ΔH° , initial and final T , and K_p at 25°C

Asked for: K_p at 500°C

Strategy:

Convert the initial and final temperatures to kelvins. Then substitute appropriate values into [Equation 18.41](#) to obtain K_2 , the equilibrium constant at the final temperature.

Solution:

The value of ΔH° for the reaction obtained using Hess's law is -91.8 kJ/mol of N_2 . If we set $T_1 = 25^\circ\text{C} = 298 \text{ K}$ and $T_2 = 500^\circ\text{C} = 773 \text{ K}$, then from [Equation 18.6.13](#) we obtain the following:

$$= \frac{(-91.8 \text{ kJ}) (1000 \text{ J/kJ})}{8.314 \text{ J/K}} \left(\frac{1}{298 \text{ K}} - \frac{1}{773 \text{ K}} \right) = -22.8$$

$$\frac{K_2}{K_1} = 1.3 \pm \times 10^{-10}$$

$$K_2 = (5.4 \times 10^5) (1.3 \pm \times 10^{-10}) = 7.0 \times 10^{-5}$$

Thus at 500°C , the equilibrium strongly favors the reactants over the products.

Exercise

In the exercise in Example 13, you calculated $K_p = 2.2 \times 10^{12}$ for the reaction of NO with O_2 to give NO_2 at 25°C . Use the ΔH_f° values in the exercise in Example 10 to calculate K_p for this reaction at 1000°C .

Answer: 5.6×10^{-4}

Summary

For a reversible process that does not involve external work, we can express the change in free energy in terms of volume, pressure, entropy, and temperature. If we assume ideal gas behavior, the ideal gas law allows us to express ΔG in terms of the partial pressures of the reactants and products, which gives us a relationship between ΔG and K_p , the equilibrium constant of a reaction involving gases, or K , the equilibrium constant expressed in terms of concentrations. If $\Delta G^\circ < 0$, then K or $K_p > 1$, and products are favored over reactants. If $\Delta G^\circ > 0$, then K or $K_p < 1$, and reactants are favored over products. If $\Delta G^\circ = 0$, then K or $K_p = 1$, and the system is at equilibrium. We can use the measured equilibrium constant K at one temperature and ΔH° to estimate the equilibrium constant for a reaction at any other temperature.

Key Takeaway

- The change in free energy of a reaction can be expressed in terms of the standard free-energy change and the equilibrium constant K or K_p and indicates whether a reaction will occur spontaneously under a given set of conditions.

Key Equations

Relationship between standard free-energy change and equilibrium constant

[Equation 18.6.9:](#) $\Delta G^\circ = -RT \ln K$

Temperature dependence of equilibrium constant

[Equation 18.6.11:](#) $\ln K = \frac{-\Delta H^\circ}{RT} + \frac{-\Delta S^\circ}{R}$

Calculation of K at second temperature

[Equation 18.6.11:](#) $\ln \frac{K_2}{K_1} = \frac{-\Delta H^\circ}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$

Conceptual Problems

- Do you expect products or reactants to dominate at equilibrium in a reaction for which ΔG° is equal to
 - 1.4 kJ/mol ?
 - 105 kJ/mol ?
 - -34 kJ/mol ?
- The change in free energy enables us to determine whether a reaction will proceed spontaneously. How is this related to the extent to which a reaction proceeds?

3. What happens to the change in free energy of the reaction $\text{N}_2(\text{g}) + 3\text{F}_2(\text{g}) \rightarrow 2\text{NF}_3(\text{g})$ if the pressure is increased while the temperature remains constant? if the temperature is increased at constant pressure? Why are these effects not so important for reactions that involve liquids and solids?
4. Compare the expressions for the relationship between the change in free energy of a reaction and its equilibrium constant where the reactants are gases versus liquids. What are the differences between these expressions?

Numerical Problems

1. Carbon monoxide, a toxic product from the incomplete combustion of fossil fuels, reacts with water to form CO_2 and H_2 , as shown in the equation $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$, for which $\Delta H^\circ = -41.0 \text{ kJ/mol}$ and $\Delta S^\circ = -42.3 \text{ J cal}/(\text{mol}\cdot\text{K})$ at 25°C and 1 atm.
 1. What is ΔG° for this reaction?
 2. What is ΔG if the gases have the following partial pressures: $P(\text{CO}) = 1.3 \text{ atm}$, $P(\text{H}_2\text{O}) = 0.8 \text{ atm}$, $P(\text{CO}_2) = 2.0 \text{ atm}$, and $P(\text{H}_2) = 1.3 \text{ atm}$?
 3. What is ΔG if the temperature is increased to 150°C assuming no change in pressure?
2. Methane and water react to form carbon monoxide and hydrogen according to the equation $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$
 1. What is the standard free energy change for this reaction?
 2. What is K_p for this reaction?
 3. What is the carbon monoxide pressure if 1.3 atm of methane reacts with 0.8 atm of water, producing 1.8 atm of hydrogen gas?
 4. What is the hydrogen gas pressure if 2.0 atm of methane is allowed to react with 1.1 atm of water?
 5. At what temperature does the reaction become spontaneous?
3. Calculate the equilibrium constant at 25°C for each equilibrium reaction and comment on the extent of the reaction.
 1. $\text{CCl}_4(\text{g}) + 6 \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CO}_2(\text{g}) + 4\text{HCl}(\text{aq})$; $\Delta G^\circ = -377 \text{ kJ/mol}$
 2. $\text{Xe}(\text{g}) + 2\text{F}_2(\text{g}) \rightleftharpoons \text{XeF}_4(\text{s})$; $\Delta H^\circ = -66.3 \text{ kJ/mol}$, $\Delta S^\circ = -102.3 \text{ J}/(\text{mol}\cdot\text{K})$
 3. $\text{PCl}_3(\text{g}) + \text{S} \rightleftharpoons \text{PSCl}_3(\text{l})$; $\Delta G_f^\circ(\text{PCl}_3) = -272.4 \text{ kJ/mol}$, $\Delta G_f^\circ(\text{PSCl}_3) = -363.2 \text{ kJ/mol}$
4. Calculate the equilibrium constant at 25°C for each equilibrium reaction and comment on the extent of the reaction.
 1. $2\text{KClO}_3(\text{s}) \rightleftharpoons 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$; $\Delta G^\circ = -225.8 \text{ kJ/mol}$
 2. $\text{CoCl}_2(\text{s}) + 6 \text{H}_2\text{O}(\text{l}) \rightleftharpoons 6\text{CoCl}_2 \cdot 6 \text{H}_2\text{O}(\text{s})$; $\Delta H_{\text{rxn}}^\circ = -352 \text{ kJ/mol}$, $\Delta S_{\text{rxn}}^\circ = -899 \text{ J}/(\text{mol}\cdot\text{K})$
 3. $\text{PCl}_3(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{POCl}_3(\text{g})$; $\Delta G_f^\circ(\text{PCl}_3) = -272.4 \text{ kJ/mol}$, $\Delta G_f^\circ(\text{POCl}_3) = -558.5 \text{ kJ/mol}$
5. The gas-phase decomposition of N_2O_4 to NO_2 is an equilibrium reaction with $K_p = 4.66 \times 10^{-3}$. Calculate the standard free-energy change for the equilibrium reaction between N_2O_4 and NO_2 .
6. The standard free-energy change for the dissolution $\text{K}_4\text{Fe}(\text{CN})_6 \cdot \text{H}_2\text{O}(\text{s}) \rightleftharpoons 4\text{K}^+(\text{aq}) + \text{Fe}(\text{CN})_6^{4-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ is 26.1 kJ/mol . What is the equilibrium constant for this process at 25°C ?
7. Ammonia reacts with water in liquid ammonia solution (am) according to the equation $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{am}) \rightleftharpoons \text{NH}_4^+(\text{am}) + \text{OH}^-(\text{am})$. The change in enthalpy for this reaction is 21 kJ/mol , and $\Delta S^\circ = -303 \text{ J}/(\text{mol}\cdot\text{K})$. What is the equilibrium constant for the reaction at the boiling point of liquid ammonia (-31°C)?
8. At 25°C , a saturated solution of barium carbonate is found to have a concentration of $[\text{Ba}^{2+}] = [\text{CO}_3^{2-}] = 5.08 \times 10^{-5} \text{ M}$. Determine ΔG° for the dissolution of BaCO_3 .
9. Lead phosphates are believed to play a major role in controlling the overall solubility of lead in acidic soils. One of the dissolution reactions is $\text{Pb}_3(\text{PO}_4)_2(\text{s}) + 4\text{H}^+(\text{aq}) \rightleftharpoons 3\text{Pb}^{2+}(\text{aq}) + 2\text{H}_2\text{PO}_4^-(\text{aq})$ for which $\log K = -1.80$. What is ΔG° for this reaction?
10. The conversion of butane to 2-methylpropane is an equilibrium process with $\Delta H^\circ = -2.05 \text{ kcal/mol}$ and $\Delta G^\circ = -0.89 \text{ kcal/mol}$.
 1. What is the change in entropy for this conversion?
 2. Based on structural arguments, are the sign and magnitude of the entropy change what you would expect? Why?
 3. What is the equilibrium constant for this reaction?

11. The reaction of $\text{CaCO}_3(\text{s})$ to produce $\text{CaO}(\text{s})$ and $\text{CO}_2(\text{g})$ has an equilibrium constant at 25°C of 2×10^{-23} . Values of ΔH_f° are as follows: CaCO_3 , -1207.6 kJ/mol ; CaO , -634.9 kJ/mol ; and CO_2 , -393.5 kJ/mol .
1. What is ΔG° for this reaction?
 2. What is the equilibrium constant at 900°C ?
 3. What is the partial pressure of $\text{CO}_2(\text{g})$ in equilibrium with CaO and CaCO_3 at this temperature?
 4. Are reactants or products favored at the lower temperature? at the higher temperature?
12. In acidic soils, dissolved Al^{3+} undergoes a complex formation reaction with SO_4^{2-} to form $[\text{AlSO}_4]^+$. The equilibrium constant at 25°C for the reaction $\text{Al}^{3+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightleftharpoons \text{AlSO}_4^+(\text{aq})$ is 1585.
1. What is ΔG° for this reaction?
 2. How does this value compare with ΔG° for the reaction $\text{Al}^{3+}(\text{aq}) + \text{F}^-(\text{aq}) \rightleftharpoons \text{AlF}^{2+}(\text{aq})$ for which $K = 10^7$ at 25°C ?
 3. Which is the better ligand to use to trap Al^{3+} from the soil?

Answers

1. -28.4 kJ/mol
2. -26.1 kJ/mol
3. -19.9 kJ/mol
- 2.
3. 1. 1.21×10^{66} ; equilibrium lies far to the right.
2. 1.89×10^6 ; equilibrium lies to the right.
3. 5.28×10^{16} ; equilibrium lies far to the right.
- 4.
5. 13.3 kJ/mol
- 6.
7. 5.1×10^{-21}
- 8.
9. 10.3 kJ/mol
- 10.
11. 1. 129.5 kJ/mol
2. 6
3. 6.0 atm
4. Products are favored at high T ; reactants are favored at low T .
- 12.

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