

Chapter 15. Chemical Equilibrium

Common Student Misconceptions

- Students confuse the arrows used for resonance and equilibrium.
- Students often incorrectly include concentrations of pure liquids and solids in equilibrium constant expressions.
- Students often have problems distinguishing between K and Q .
- Students struggle with the idea that K and Q have no units.
- Students often do not know (or check) whether an approximate equilibrium calculation is valid.

Teaching Tips

- Many students need to see how the numerical problems in this chapter are solved.
- Students who have difficulty with some of the mathematical manipulations in this chapter should be directed to Appendix A of the text.
- It is useful to stress that concentrations in equilibrium constant expressions are merely approximations of activities. Since activities of pure substances are 1, they do not appear in equilibrium constant expressions, and neither should their approximations, i.e. concentrations.
- Students might find it helpful to consider changing the concentration or pressure of a substance in a reaction at equilibrium as creating stress. The removal of something creates a "deficit stress," and adding something creates an "excess stress." Reactions will always respond to stress by reducing the deficit (making more of what was removed) or by reducing the excess (consuming some of what was added). This description works regardless of which side of the reaction is changed.
- Students should be reminded that principles of general equilibria taught in this chapter are very universal and need to be mastered in order to understand concepts in future chapters.

Lecture Outline

15.1 The Concept of Equilibrium^{1,2,3,4,5,6,7,8}

- Consider colorless frozen N_2O_4 .
 - At room temperature, it decomposes to brown NO_2 .

$$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$$
 - At some time, the color stops changing and we have a mixture of N_2O_4 and NO_2 .
 - **Chemical equilibrium** is the point at which the concentrations of all species are constant.
 - Opposing reactions proceed at equal rates.
 - Assume that both the forward and reverse reactions are elementary processes.
 - We can write rate expressions for each reaction.
 - Forward reaction: $\text{N}_2\text{O}_4(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$
 - $\text{Rate}_f = k_f[\text{N}_2\text{O}_4]$ $k_f = \text{rate constant (forward reaction)}$

¹ "Chemical Equilibrium" Activity from Instructor's Resource CD/DVD

² Figure 15.2 from Transparency Pack

³ "NO₂-N₂O₄ Equilibrium" Activity from Instructor's Resource CD/DVD

⁴ "NO₂-N₂O₄ Equilibrium" Animation from Instructor's Resource CD/DVD

⁵ "Nitrogen Dioxide and Dinitrogen Tetroxide" Movie from Instructor's Resource CD/DVD

⁶ "Dinitrogen Tetroxide" 3-D Model from Instructor's Resource CD/DVD

⁷ "Nitrogen Dioxide" 3-D Model from Instructor's Resource CD/DVD

⁸ "Oxygen" 3-D Model from Instructor's Resource CD/DVD

- Reverse reaction: $2\text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$
 - $\text{Rate}_r = k_r[\text{NO}_2]^2$ $k_r = \text{rate constant (reverse reaction)}$
- Place some pure N_2O_4 into a closed container.
 - As N_2O_4 reacts to form NO_2 , the concentration of N_2O_4 will decrease and the concentration of NO_2 will increase.
 - Thus, we expect the forward reaction rate to slow and the reverse reaction rate to increase.
 - Eventually we get to equilibrium where the forward and reverse rates are equal.
 - At equilibrium:

$$k_f[\text{N}_2\text{O}_4] = k_r[\text{NO}_2]^2$$
 - Rearranging, we get:

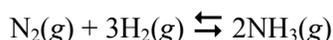
$$k_f/k_r = \text{a constant}$$
- At equilibrium the concentrations of N_2O_4 and NO_2 do not change.
 - This mixture is called an *equilibrium mixture*.
 - The equilibrium mixture results because the reaction is *reversible*.
 - This is an example of a dynamic equilibrium.
 - A dynamic equilibrium exists when the rates of the forward and reverse reactions are equal.
 - No further net change in reactant or product concentration occurs.
 - The double arrow \rightleftharpoons implies that the process is dynamic.

FORWARD REFERENCES

- Equilibria involving acids and bases will be the subject of Chapters 16 and 17.
- Solubility equilibria will be discussed in Chapter 17 (section 17.4).
- The concept of equilibrium will be important throughout Chapter 19.

15.2 The Equilibrium Constant^{9,10,11,12,13,14,15,16,17}

- Consider the **Haber process**:



- It is used for the preparation of ammonia from nitrogen and hydrogen.
- The process is carried out at high temperature and pressure.
 - Ammonia is a good source of fixed nitrogen for plants.
 - Much of the NH_3 produced industrially is used as a fertilizer.
- If we start with a mixture of nitrogen and hydrogen (in any proportions), the reaction will reach equilibrium with constant concentrations of nitrogen, hydrogen, and ammonia.
- However, if we start with just ammonia and no nitrogen or hydrogen, the reaction will proceed and N_2 and H_2 will be produced until equilibrium is achieved.
- No matter what the starting composition of reactants and products is, the equilibrium mixture contains the same relative concentrations of reactants and products.
 - Equilibrium can be reached from either direction.

⁹ “Fritz Haber” from Further Readings

¹⁰ “Equilibrium Constant” Activity from Instructor’s Resource CD/DVD

¹¹ “An Elementary Discussion of Chemical Equilibrium” from Further Readings

¹² “Equilibrium and Le Châtelier’s Principle” from Live Demonstrations

¹³ “Chemical Equilibrium in the General Chemistry Course” from Further Readings

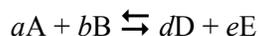
¹⁴ “The Complexity of Teaching and Learning Chemical Equilibrium” from Further Readings

¹⁵ “Equilibrium: A Teaching/Learning Activity” from Further Readings

¹⁶ “Introducing Dynamic Equilibrium as an Explanatory Model” from Further Readings

¹⁷ “Simulations for Teaching Chemical Equilibrium” from Further Readings

- We can write an expression for the relationship between the concentration of the reactants and products at equilibrium.
 - This expression is based on the **law of mass action**.
 - For a general reaction,



- The **equilibrium-constant expression** is given by:

$$K_c = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$

- Where K_c is the **equilibrium constant**.
- The subscript “c” indicates that molar concentrations were used to evaluate the constant.
- Note that the equilibrium constant expression has products in the numerator and reactants in the denominator.

Evaluating K_c

- The value of K_c does not depend on initial concentrations of products or reactants.
 - Consider the reaction:



- The equilibrium constant is given by:

$$K_c = \frac{[NO_2]^2}{[N_2O_4]}$$

- The value of this constant (at 100 °C) is 6.50 (regardless of the initial concentrations of $N_2O_4(g)$ or $NO_2(g)$).
- The equilibrium expression depends on stoichiometry.
 - It does not depend on the reaction mechanism.
 - The value of K_c varies with temperature.
- We generally omit the units of the equilibrium constant.

Equilibrium Constants in Terms of Pressure, K_p

- When the reactants and products are gases, we can write an equilibrium expression using partial pressures rather than molar concentrations.
- The equilibrium constant is K_p where “p” stands for pressure.
- For the reaction:



$$K_p = \frac{(P_D)^d (P_E)^e}{(P_A)^a (P_B)^b}$$

- They can be interconverted using the ideal gas equation and our definition of molarity:

$$PV = nRT \quad \text{thus} \quad P = (n/V)RT$$

- If we express volume in liters the quantity (n/V) is equivalent to molarity.
- Thus the partial pressure of a substance, A, is given as:

$$P_A = (n_A/V)RT = [A]RT$$

- We can use this to obtain a general expression relating K_c and K_p .

$$K_p = K_c(RT)^{\Delta n}$$

- Where $\Delta n = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants})$.
- The numerical values of K_c and K_p will differ if $\Delta n \neq 0$.

FORWARD REFERENCES

- Various equilibrium constants: K_a , K_b , K_w , K_{sp} will be used throughout Chapters 16 and 17, and later in select end-of-chapter problems in Chapters 19 and 20.
- K_p and K_c will be revisited in Chapter 19 (section 19.7).
- Haber process will be further discussed in Chapter 22 (sections 22.2 and 22.7).

Equilibrium Constants and Units

- Equilibrium constants are reported without units.
 - The equilibrium constant may be derived from thermodynamic measurements.
 - The constants are defined in terms of *activities* rather than concentrations or partial pressures.
 - The activity of a substance in an *ideal* mixture is the ratio of the concentration or pressure of the substance either to a reference concentration (1 M) or a reference pressure (1 atm).
 - The units in these ratios cancel, thus activities have no units.
 - The numerical value of the activity equals the concentration.
 - For pure solids and pure liquids, the activities equal 1.
 - What about “real” systems?
 - Here activities also have no units.
 - The activities are not exactly equal to concentration but we will ignore the differences.
 - *Thermodynamic equilibrium constants* derived from these activities also have no units.

15.3 Understanding and Working with Equilibrium Constants**The Magnitude of Equilibrium Constants**

- The equilibrium constant, K , is the ratio of products to reactants.
 - Therefore, the larger K the more products are present at equilibrium.
 - Conversely, the smaller K the more reactants are present at equilibrium.
 - If $K \gg 1$, then products dominate at equilibrium and equilibrium *lies to the right*.
 - If $K \ll 1$, then reactants dominate at equilibrium and the equilibrium *lies to the left*.

The Direction of the Chemical Equation and K

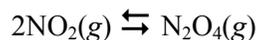
- An equilibrium can be approached from either direction.
- Consider the reaction:



- The equilibrium constant for this reaction (at 100 °C) is:

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.212$$

- However, when we write the equilibrium expression for the reverse reaction,



- The equilibrium constant for this reaction (at 100 °C) is:

$$K_c = \frac{[\text{N}_2\text{O}_4]}{[\text{NO}_2]^2} = 4.72$$

- The equilibrium constant for a reaction in one direction is the reciprocal of the equilibrium constant of the reaction in the reverse direction.

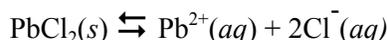
Relating Chemical Equation Stoichiometry and Equilibrium Constants

- It is possible to calculate the equilibrium constant for a reaction if we know the equilibrium constants for other reactions that add up to give us the one we want.
 - This is similar to using Hess’s law.

- The equilibrium constant of a reaction in the reverse direction is the inverse of the equilibrium constant of the reaction in the forward direction.
- The equilibrium constant of a reaction that has been multiplied by a number is the equilibrium constant raised to a power equal to that number.
- The equilibrium constant for a net reaction made up of two or more steps is the product of the equilibrium constants for the individual steps.

15.4 Heterogeneous Equilibria¹⁸

- Equilibria in which all reactants and products are present in the same phase are called **homogeneous equilibria**.
- Equilibria in which one or more reactants or products are present in a different phase are called **heterogeneous equilibria**.
- Consider the equilibrium established when solid lead(II) chloride dissolves in water to form a saturated solution:



- Experimentally, the amount of Pb^{2+} and Cl^{-} does not depend on the amount of PbCl_2 .
 - Why?
 - The concentration of a pure solid or pure liquid equals its density divided by its molar mass.
 - Neither density nor molar mass is a variable.
 - Thus the concentrations of solids and pure liquids are constant.
 - For the dissolution of PbCl_2 :

$$K_c = [\text{Pb}^{2+}][\text{Cl}^{-}]^2$$
- If a pure solid or pure liquid is involved in a heterogeneous equilibrium, its concentration is not included in the equilibrium constant expression.
- Note: Although the *concentrations* of these species are not included in the equilibrium expression, they *do* participate in the reaction and *must* be present for an equilibrium to be established!
- Other common examples of heterogeneous equilibria include:
 - systems involving solids and gases.
 - Example: $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$

$$K_c = [\text{CO}_2] \text{ and } K_p = P_{\text{CO}_2}$$
 - systems where the solvent is involved as a reactant or product and the solutes are present at low concentrations.
 - Example: $\text{H}_2\text{O}(l) + \text{CO}_3^{2-}(aq) \rightleftharpoons \text{OH}^{-}(aq) + \text{HCO}_3^{-}(aq)$

$$K_c = [\text{OH}^{-}][\text{HCO}_3^{-}] / [\text{CO}_3^{2-}]$$
 - Here the concentration of water is essentially constant and we can think of it as a pure liquid.

FORWARD REFERENCES

- Solubility equilibria will be discussed in detail in Chapter 17 (section 17.4).
- Exclusion of pure liquids and solids from equilibrium expressions will be revisited in Chapter 19.

15.5 Calculating Equilibrium Constants^{19,20}

- Proceed as follows:

¹⁸ “Carbon Dioxide” 3-D Model from Instructor’s Resource CD/DVD

¹⁹ “Using an Equilibrium Table” Activity from Instructor’s Resource CD/DVD

²⁰ “Applying the Reaction Table Method for Chemical Reaction Problems (Stoichiometry and Equilibrium)” from Further Readings

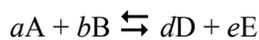
- Tabulate initial and equilibrium concentrations (or partial pressures) for all species in the equilibrium.
- If an initial *and* an equilibrium concentration are given for a species, calculate the change in concentration.
- Use the coefficients in the balanced chemical equation to calculate the changes in concentration of all species.
- Deduce the equilibrium concentrations of all species.
- Use these to calculate the value of the equilibrium constant.

FORWARD REFERENCES

- Methods of finding equilibrium constants from thermodynamic or electrochemical will be discussed in Chapters 19 and 20, respectively.

15.6 Applications of Equilibrium Constants**Predicting the Direction of Reaction**

- For a general reaction:



- We define Q , the **reaction quotient**, as:

$$Q = \frac{[D]^d [E]^e}{[A]^a [B]^b}$$

- Where $[A]$, $[B]$, $[D]$, and $[E]$ are molarities (for substances in solution) or partial pressures (for gases) at any given time.
- We can compare Q_c to K_c or Q_p to K_p :
 - If $Q = K$, then the system is at equilibrium.
 - If $Q < K$, then the forward reaction must occur to reach equilibrium.
 - If $Q > K$, then the reverse reaction must occur to reach equilibrium.
 - Products are consumed, reactants are formed.
 - Q decreases until it equals K .

Calculating Equilibrium Concentrations

- The same steps used to calculate equilibrium constants are used to calculate equilibrium concentrations.
- Generally, we do not have a number for the change in concentration.
 - Therefore, we need to assume that x mol/L of a species is produced (or used).
- The equilibrium concentrations are given as algebraic expressions.

FORWARD REFERENCES

- Q will be used in Chapter 19 (section 19.7) to determine Gibb's free energy change at nonstandard conditions.
- Q will be used in Chapter 20 (section 20.6) to determine cell potentials at nonstandard conditions (Nernst equation).

15.7 Le Châtelier's Principle^{21,22,23,24,25}

- Consider the Haber process:

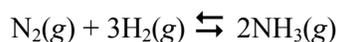
²¹ Figure 15.9 from Transparency Pack

²² "Formation of Water" Movie from Instructor's Resource CD/DVD

²³ "Nitrogen" 3-D Model from Instructor's Resource CD/DVD

²⁴ "Hydrogen" 3-D Model from Instructor's Resource CD/DVD

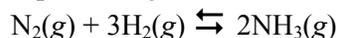
²⁵ "Ammonia" 3-D Model from Instructor's Resource CD/DVD



- As the pressure increases, the amount of ammonia present at equilibrium increases.
- As the temperature increases, the amount of ammonia at equilibrium decreases.
- Can this be predicted?
 - Yes! We can use Le Châtelier's principle to make qualitative predictions about how a system at equilibrium responds to various changes in external conditions.
 - Le Châtelier's principle: if a system at equilibrium is disturbed by a change in temperature, a change in pressure, or a change in the concentration of one or more components, the system will shift its equilibrium position in such a way as to counteract the effects of the disturbance.

Change in Reactant or Product Concentration^{26,27,28,29,30,31,32}

- If a chemical system is at equilibrium and we add or remove a product or reactant, the reaction will shift so as to reestablish equilibrium.
 - For example, consider the Haber process again:



- If H_2 is added while the system is at equilibrium, $Q < K$.
 - The system must respond to counteract the added H_2 (Le Châtelier's principle).
 - That is, the system must consume the H_2 and produce products until a new equilibrium is established.
 - Therefore, $[\text{H}_2]$ and $[\text{N}_2]$ will decrease and $[\text{NH}_3]$ will increase until $Q = K$.
- We can exploit this industrially.
 - Suppose that we wanted to optimize the amount of ammonia we formed from the Haber process.
 - We might flood the reaction vessel with reactants and continuously remove product.
 - The amount of ammonia produced is optimized because the product (NH_3) is continuously removed and the reactants (N_2 and H_2) are continuously being added.

Effects of Volume and Pressure Changes³³

- Consider a system at equilibrium.
- If the equilibrium involves gaseous products or reactants, the concentration of these species will be changed if we change the volume of the container.
 - For example, if we decrease the volume of the container, the partial pressures of each gaseous species will increase.
 - Le Châtelier's principle predicts that if pressure is increased, the system will shift to counteract the increase.
 - That is, the system shifts to remove gases and decrease pressure.
 - An increase in pressure favors the direction that has fewer moles of gas.
- Consider the following system:



- An increase in pressure (by decreasing the volume) favors the formation of colorless N_2O_4 .

²⁶ "Amounts Tables as a Diagnostic Tool for Flawed Stoichiometric Reasoning" from Further Readings

²⁷ "Calculating Equilibrium Concentrations by Iteration: Recycle Your Approximations" from Further Readings

²⁸ "Man of Principle" from Further Readings

²⁹ Figure 15.10 from Transparency Pack

³⁰ "Chemical Equilibrium and Polynomial Equations: Beware of Roots" from Further Readings

³¹ "Le Châtelier's Principle" from Further Readings

³² "Effect of Concentration on Equilibrium: Cobalt Complex" from Live Demonstrations

³³ "Le Châtelier's Principle" Activity from Instructor's Resource CD/DVD

- The instant the pressure increases, the concentration of both gases increases and the system is not at equilibrium.
- The system changes to reduce the number moles of gas.
- A new equilibrium is established.
 - The mixture is lighter in color.
 - Some of the brown NO_2 has been converted into colorless $\text{N}_2\text{O}_4(\text{g})$
- In a reaction with the same number of moles of gas in the products and reactants, changing the pressure has no effect on the equilibrium.
- In addition, no change will occur if we increase the total gas pressure by the addition of a gas that is not involved in the reaction.

Effect of Temperature Changes^{34,35,36,37,38,39}

- The equilibrium constant is temperature dependent.
- How will a change in temperature alter a system at equilibrium?
 - It depends on the particular reaction.
 - For example, consider the endothermic reaction:

$$\text{Co}(\text{H}_2\text{O})_6^{2+}(\text{aq}) + 4\text{Cl}^-(\text{aq}) \rightleftharpoons \text{CoCl}_4^{2-}(\text{aq}) + 6\text{H}_2\text{O}(\text{l}) \quad \Delta H > 0$$
 - $\text{Co}(\text{H}_2\text{O})_6^{2+}$ is pale pink and CoCl_4^{2-} is a deep blue.
 - At room temperature, an equilibrium mixture (light purple) is placed in a beaker of warm water.
 - The mixture turns deep blue.
 - This indicates a shift toward products (blue CoCl_4^{2-}).
 - This reaction is endothermic.
 - For an endothermic reaction ($\Delta H > 0$), heat can be considered as a reactant.
 - Thus, adding heat causes a shift in the forward direction.
 - The room-temperature equilibrium mixture is placed in a beaker of ice water.
 - The mixture turns bright pink.
 - This indicates a shift toward reactants (pink $\text{Co}(\text{H}_2\text{O})_6^{2+}$).
 - In this case, by cooling the system we are removing a reactant (heat).
 - Thus, the reaction is shifted in the reverse reaction.
- A change in temperature causes a change in the value of K .
 - If we increase the temperature of an endothermic reaction, K increases.
 - If we increase the temperature of an exothermic reaction, K decreases.

The Effect of Catalysts

- A catalyst lowers the activation energy barrier for the reaction.
 - Therefore, a catalyst will decrease the amount of time needed to reach equilibrium.
 - A catalyst *does not* affect the composition of the equilibrium mixture.

FORWARD REFERENCES

- Le Châtelier's principle versus percent ionization will be mentioned in Chapter 16 (section 16.6).
- Le Châtelier's principle will be brought up in the common ion effect in Chapter 17 (sections 17.1 and 17.3).

³⁴ "Effect of Temperature Change on Equilibrium: Cobalt Complexes" from Live Demonstrations

³⁵ "Equilibrium in the Gas Phase" from Live Demonstrations

³⁶ "From Chicken Breath to the Killer Lakes of Cameroon: Uniting Seven Interesting Phenomena with a Single Chemical Underpinning" from Live Demonstrations

³⁷ "The Fizz Keeper, a Case Study in Chemical Education, Equilibrium, and Kinetics" from Further Readings

³⁸ "Temperature Dependence of Equilibrium" Animation from Instructor's Resource CD/DVD

³⁹ "Tetrachlorocobaltate(III)" 3-D Model from Instructor's Resource CD/DVD

- Le Châtelier's principle versus solubility will be discussed in Chapter 17 (section 17.5).
- Predictions of the equilibrium shifts will be made in Chapter 19 (section 19.7).
- Haber process will be further discussed in Chapter 22 (sections 22.2 and 22.7).

Further Readings:

1. Martin R. Feldman and Monica L. Tarver, "Fritz Haber," *J. Chem. Educ.*, Vol. 60, **1983**, 463–464.
2. Carl W. David, "An Elementary Discussion of Chemical Equilibrium," *J. Chem. Educ.*, Vol. 65, **1988**, 407–409.
3. Vladimir E. Fainzilberg and Stewart Karp, "Chemical Equilibrium in the General Chemistry Course," *J. Chem. Educ.*, Vol. 71, **1994**, 769–770.
4. Louise Tyson, David F. Treagust, and Robert B. Bucat, "The Complexity of Teaching and Learning Chemical Equilibrium," *J. Chem. Educ.*, Vol. 76, **1999**, 554–558.
5. Audrey H. Wilson, "Equilibrium: A Teaching/Learning Activity," *J. Chem. Educ.*, Vol. 75, **1998**, 1176–1177.
6. Jan H. Van Driel, Wobbe de Vos, and Nico Verloop, "Introducing Dynamic Equilibrium as an Explanatory Model," *J. Chem. Educ.*, Vol. 76, **1999**, 559–561.
7. Penelope A. Huddle, Margie W. White, and Fiona Rogers, "Simulations for Teaching Chemical Equilibrium," *J. Chem. Educ.*, Vol. 77, **2000**, 920–926.
8. Steven F. Watkins, "Applying the Reaction Table Method for Chemical Reaction Problems (Stoichiometry and Equilibrium)," *J. Chem. Educ.*, Vol. 80, **2003**, 658–661.
9. John Olmsted III, "Amounts Tables as a Diagnostic Tool for Flawed Stoichiometric Reasoning," *J. Chem. Educ.*, Vol. 76, **1999**, 52–54.
10. Michael Sutton, "Man of Principle," *Chemistry in Britain*, **June 2000**, 42–44.
11. William R. Smith and Ronald W. Missen, "Chemical Equilibrium and Polynomial Equations: Beware of Roots," *J. Chem. Educ.*, Vol. 66, **1989**, 489–490.
12. E. Weltin, "Calculating Equilibrium Concentrations by Iteration: Recycle Your Approximations," *J. Chem. Educ.*, Vol. 72, **1995**, 36–38.
13. Richard S. Treptow, "Le Châtelier's Principle," *J. Chem. Educ.*, Vol. 57, **1980**, 417–420.
14. Reed A. Howald, "The Fizz Keeper, a Case Study in Chemical Education, Equilibrium, and Kinetics," *J. Chem. Educ.*, Vol. 76, **1999**, 208–209.

Live Demonstrations:

1. Lee R. Summerlin and James L. Ealy, Jr., "Equilibrium in the Gas Phase," *Chemical Demonstrations, A Sourcebook for Teachers* (Washington: American Chemical Society, **1988**), pp. 85–86. Color changes in a mixture of NO_2 and N_2O_4 as a sealed tube of gas is heated or cooled are used to demonstrate Le Châtelier's principle.
2. Lee R. Summerlin, and James L. Ealy, Jr., "Effect of Temperature Change on Equilibrium: Cobalt Complex," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1* (Washington: American

Chemical Society, **1988**), pp. 79–80. An equilibrium system containing the dehydrated-hydrated cobalt complex is shifted in response to changes in temperature.

3. Lee. R. Summerlin, and James. L. Ealy, Jr., “Equilibrium and Le Châtelier's Principle,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1* (Washington: American Chemical Society, **1988**), pp. 77–78. An overhead projector demonstration of the effect of reactant concentration on equilibrium.

4. Lee. R. Summerlin, and James. L. Ealy, Jr., “Effect of Concentration on Equilibrium: Cobalt Complex,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1* (Washington: American Chemical Society, **1988**), pp. 81–82.

5. Ron DeLorenzo, “From Chicken Breath to the Killer Lakes of Cameroon: Uniting Seven Interesting Phenomena with a Single Chemical Underpinning,” *J. Chem. Educ.*, Vol. 78, **2001**, 191–194. A collection of demonstrations dealing with equilibria associated with a sodium bicarbonate solution.