Equilibrium and Le Chatelier's Principle

Strand	Nomenclature, Chemical Formulas, and Reactions	
Topic	Investigating chemical reactions and equations	
Primary SOL	CH.3	The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include f) reaction rates, kinetics, and equilibrium.
Related SOL	CH.1	The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include a) designated laboratory techniques; b) safe use of chemicals and equipment; c) proper response to emergency situations.

Background Information

In many chemical reactions, the reactants are not completely converted into products. When all apparent chemical change has ceased, often significant amounts of the reactant and product species remain in the reaction mixture. Just as the reactants can interact to form products at a particular speed, called the "forward rate," the products can interact to re-form the reactants at a speed called the "reverse rate."

In a typical chemical reaction, initially the reactant concentrations are high, and the product concentrations are low. Thus the initial forward rate is much faster than the initial reverse rate. However, the forward reaction slows down as the reactants are depleted, and at the same time, the reverse reaction speeds up due to the increasing amounts of products. The two rates become exactly equal at a particular combination of reactant and product concentrations. The reactants are then replenished by the reverse reaction just as quickly as they are consumed in the forward reaction. As there is no net change in the amounts of reactants and products present in the reaction mixture, the reaction is said to be in a state of dynamic equilibrium.

Henri Le Chatelier's principle states that if a stress is applied to a system in equilibrium, the system will react in such a way as to relieve that stress and restore equilibrium under a new set of conditions. To relieve stress, a system can do only one of two things: (1) form more products, using up reactants, or (2) reverse the reaction and form more reactants, using up products. Le Chatelier proposed that the effect of changes in pressure, temperature, and reactant and product concentrations upon a chemical equilibrium would always be such as to oppose the applied change.

Materials

- Distilled water
- Test tubes
- Stirrer
- Pipette
- Ice water bath

- Hot water bath
- Ethanol
- Solid NaCl
- Concentrated HCl
- Bromothymol blue indicator
- *M* HCl
- *M* NaOH
- Solid $CoCl_2 \bullet 6\ H_2O$
- *M* silver nitrate solution
- *Safety goggles*

Vocabulary

concentrated, equilibrium, indicators, ion, products, reactants, saturated solution

Student/Teacher Actions (what students and teachers should be doing to facilitate learning)

Introduction

1. Put students into groups of three or four. Tell the groups that as you do the following demonstrations, they are to watch the demonstrations closely, ask necessary questions, and write down their observations. Have students then discuss within their groups the results of the demonstrations and come to conclusions about what is happening to the equilibrium systems. Each group will then present their results to the class.

Procedure

Demonstration 1: Equilibrium in a Saturated Sodium Chloride Solution

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$

1. Pour some solid NaCl into a test tube, and fill the tube ¾ full of distilled water.
2. Cork and shake the tube to form a saturated solution.
3. If all the NaCl dissolves, add some more until excess solid is obtained.
4. Decant the solution into a second test tube.
5. To this saturated solution of NaCl, add some Cl^- ions in the form of concentrated HCl.
6. Have students make observations and explain the results. (Results: The addition of the Cl^- ion from the HCl shifts the equilibrium to the left, resulting in solid NaCl forming in the solution and settling to the bottom of the test tube.)

Demonstration 2: Equilibrium with an Acid-Base Indicator

Acid-base indicators like bromothymol blue have large organic molecules that can gain and lose hydrogen ions to form substances that have different colors. The reaction of the indicator bromothymol blue can be illustrated as follows:

$HIn(aq) \rightarrow H^+(aq) + In^-(aq)$

yellow *blue*

In this reaction, HIn is the neutral indicator molecule, and In^- is the indicator ion after the molecule has lost a hydrogen ion. This demonstration also shows that equilibrium reactions can easily be caused to go in either direction.

1. Fill a test tube about half full of distilled water.
2. Add several drops of bromothymol blue indicator solution. Then, add drops of 0.1 *M* HCl while stirring. This will cause the H^+ ion concentration to *increase*.
3. Have students note the color of the indicator.
4. Next, stir the solution constantly while adding 0.1 *M* NaOH drop by drop until no further color change occurs. Adding OH^- ions causes the H^+ ion concentration to *decrease*.
5. Now, try to add just the right amount of acid back to this solution to cause it to be neutral or green (half of the indicator molecules are blue and the other half are yellow) in color after it is stirred.
6. Have students make observations and explain the results. (Results: The addition of the H^+ ion from the HCl shifts the equilibrium to the left, resulting in the HIn form of the indicator molecule. The addition of the OH^- ion from the NaOH shifts the equilibrium to the right as the H^+ ion is removed during its reaction with the added OH^- ion to form water. The shift to the right results in the In^- form of the indicator molecule.)

Demonstration 3: Equilibrium with Cobalt Complex Ions

$[Co(H_2O)_6]^{2+}(aq) + 4\ Cl^-(aq) \rightarrow [CoCl_4]^{2-} + 6\ H_2O\ \Delta H = +\ 50$ kJ/mol

pink *blue*

1. Dissolve a small amount of solid $CoCl_2 \bullet 6\ H_2O$ in a beaker containing ethanol. The solution should be purple; if it is pink, add a little concentrated HCl, drop by drop while stirring, until it is purple.
2. Put some of the solution into each of two test tubes.
3. To one of the test tubes, add drops of concentrated HCl, one drop at a time while stirring. Have students note the result.
4. To the second tube, add drops of 0.1 M silver nitrate solution, one drop at a time while stirring.
5. Have students make observations and explain the results. (Results: The addition of the concentrated HCl shifts the equilibrium to the right, resulting in a deep blue color. This is due to the addition of the Cl^- ion. The addition of the $AgNO_3$ shifts the equilibrium to the left, resulting in a pink color. The added Ag^+ ion reacts with the Cl^- ion to remove it by forming the insoluble compound AgCl. The equilibrium shifts left, producing a cloudy pink color. The cloudy appearance is due to the precipitate of AgCl.)

The Effect of Changing the Temperature

1. Take some of the cobalt chloride and HCl solution from step 3 above. The solution should be violet—i.e., between the original pink-red and the bright blue.
2. If it is not violet, adjust the color by carefully adding distilled water, one drop at a time while stirring, until a violet color is obtained. Alternatively, add concentrated HCl, depending on the color of the solution and the direction you need to go. You now have a solution that contains both $[Co(H_2O)_6]^{2+}$ and $[CoCl_4]^{2-}$ ions.

3. Divide this violet solution into three medium-size test tubes. Place one tube in an ice bath. Place the second tube in a hot water bath maintained at between 80°C and 90°C, and maintain the third tube at room temperature.
4. After the tubes have reached the desired temperatures, have students note and record the colors of each of the three solutions. *(Results: The solution in the hot water bath is deep blue. The reaction is endothermic as noted by the +ΔH value. Heat is a reactant in this reaction. Addition of heat shifts the reaction equilibrium to the right, forming the blue complex. The solution in the ice water bath is pink. Removal of heat shifts the equilibrium to the left, forming the pink complex. The solution at room temperature has an equilibrium position between the pink and the blue; it remains purple.)*
5. Determine whether the color changes are reversible by switching the test tubes in the hot and cold water baths and allowing them to come to room temperature.

Assessment

- **Questions**
  The following chemical system is at equilibrium: $2\ W(aq) + X(g) \leftrightarrow 3\ Y(g) + 2\ Z(g)$ + heat
  - How would the equilibrium change if the concentration of W were increased?
  - How would the equilibrium change if the temperature were increased?
  - In what direction would the equilibrium shift if a catalyst were added to the system?
  - What is your prediction for what would happen to the concentration of W if the temperature were decreased?
  - If the temperature of the system were decreased, what would happen to Z?
  - If the concentration of W were decreased, how would the equilibrium change?
  - In what direction would the equilibrium shift if the concentration of Y were increased?
  - What would happen to the concentration of Y if the concentration of W were increased?
  - What are four specific things that could be applied to this system that would favor the reverse reaction?
- **Other**
  - Novelty devices for prediction of rain contain cobalt(II) chloride and are based on the following equilibrium: $CoCl_2(s) + 6\ H_2O(g) \leftrightarrow CoCl_2 \bullet 6\ H_2O(s)$. What color will such an indicator be if it is about to rain? *(purple or pink)*
  - If you wanted to use the reaction $CO(g) + 2\ H_2(g) \leftrightarrow CH_3OH(g)$ + heat for producing methanol (CH_3OH) commercially, would high temperature or low temperature favor a maximum yield?
  - Arsenic can be extracted from its ores by first reacting the ore with oxygen (called "roasting") to form solid As_4O_6, which is then reduced by using carbon: $As_4O_6(s) + 6\ C(s) \leftrightarrow As_4(g) + 6\ CO(g)$. Predict the direction of the equilibrium shift in response to each of the following changes in conditions:
    - Addition of carbon monoxide
    - Addition of carbon
    - Removal of As_4
  - For each of the following reactions, predict how the equilibrium concentration of the products will be affected if the temperature is increased:

- $N_2(g) + O_2(g) \leftarrow\rightarrow 2\ NO(g)$ $\Delta H = +181$ kJ
- $2\ SO_2(g) + O_2(g) \leftarrow\rightarrow 2\ SO_3(g)$ $\Delta H = -198$ kJ

Extensions and Connections (for all students)

- Have the class determine a local industry and arrange for the class or a team of students to visit, or have a guest speaker from a local industry that uses the Le Chatelier's principle in product manufacturing.

Strategies for Differentiation

- Le Chatelier's principle can be demonstrated through a student-movement demonstration: If a student pushes another student, a discussion can evolve around the reaction. How does one individual affect the movement of another?
- Have students create a color-coded flow chart to represent the reactions observed in the demonstrations.