All Things Being Equal!

MI HSCE

C5.3a Describe equilibrium shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).
C5.3b Predict shifts in a chemical systems by changing conditions (Le Chatelier's Principle).
C5.3c Predict the extent that reactants are converted into products using the value of the equilibrium constant.

NSES

As a result of activities in grades 9-12, all students should develop:

- Content Standard A: abilities necessary to do scientific inquiry.
- Content Standard B: an understanding of chemical reactions

TYPE OF INQUIRY

Guided Inquiry - students will proceed through each activity with tasks that are given, but each team of students will likely come up with unique responses. After each activity, a discussion takes place allowing the instructor ask additional questions to help ensure students come to the acceptable scientific explanation.

TIME

Prep – Varies: See below (1 time set-up)
Activity time: Prelab: 5-10 minutes
Activity: Part I (45min); Part II (35min); Part III (45 min);
Postlab: 15 minutes

EDUCATIONAL OBJECTIVES

<u>Students will be able to:</u>

- Define a chemical equilibrium system one where the rates of the forward and reverse reactions are equal.
- Write an equilibrium constant expression.
- Explain what an equilibrium constant tells you about the relative concentrations of reactants and products in a reaction mixture. .
- Explain how a reversible reaction can be different colors and still at equilibrium (LeChatelier's Principle).
- Demonstrate a particulate understanding of LeChatelier's Principle.

CONCEPTS ADDRESSED

Understanding an equilibrium system.
Understanding LeChatelier's Principle and how it affects an equilibrium system.

MISCONCEPTIONS

This activity addresses two major student misconceptions about equilibrium (Orvis & Orvis, 2005).

1) Many students erroneously assume that the condition of equilibrium means equal concentrations of reactants and products. This is addressed in Part I where students examine a physical equilibrium and recognize that eventually the system reaches a point where the amount of water in beakers A and B is not changing but the amount of water in beaker A is not equal to the amount of water in beaker B.
2) Students do not easily grasp the notion that one can approach an equilibrium state from either direction. This is addressed in Parts II and III where students examine what happens when they increase or decrease the concentration of a reactant.

PREREQUISITE KNOWLEDGE

Reversible reactions, particulate representations of reactions, collision theory.

TEACHER BACKGROUND

This activity is designed to help students understand the difficult concept of equilibrium. Many students can memorize that an equilibrium reaction is "when the rate of the forward and backward reaction are the same". Though they can state this, students generally do not understand what this means. Part I of this activity is designed to help students understand an equilibrium system by using a physical equilibrium model. As students watch this demonstration, graph the data, and answer the analysis questions they can see that they reach a point when the water levels in beakers A and B do not change which helps them construct the concept of equilibrium. The analysis questions also help students focus on the important idea that equilibrium does not mean equal concentrations. LeChatelier's Principle is another difficult concept for students. Again students can often indicate which direction a reaction "shifts" but do not really understand what is happening in the reaction. In Part II, students conduct experiments where they apply stresses to an equilibrium system and see color changes in the reaction. They then use these color changes to determine what is happening to the concentration of the reactant and product species. In Part III, the same reactions are studied using bingo chips to represent the reactant and product species from Part II. The purpose of Part III is to help students visualize what happens to the reactant and product species in an equilibrium system when a "stress" is applied. Students use the equilibrium constant (K) to determine whether a reaction is at equilibrium. Then by adding or removing "particles", students determine whether more reactants or products are made as a result of the changes.

A common misconception regarding equilibrium is when the amount of reactants and products is the same. This activity is designed so that whenever an equilibrium system is reached the amounts of reactants and products are not the same. In Part I, the amounts of water in the reactant and product beakers are not equal though the rate of formation is the same. This is also occurs in Part III of the activity.

SAFETY

No safety issues for Part I and III. Potassium thiocyanate causes irritation to eyes, skin and respiratory tracts. Iron (III) nitrate nonahydrate is a strong oxidizer. Disodium hydrogen phosphate causes irritation to eyes, skin and respiratory tracts.

MATERIALS AND PREPARATION

Materials are for 4 sections with 16 groups (pairs) per class

Part I: Prep time: 5 minutes

Materials: For class demonstration

a) 4 beakers: 100 mL and 50 mL and **two** 1000 mL beakers (A and B)
b) 700 mL of water with 3 drops of food coloring in beaker A
c) Graph paper

This part of the lab can be done as a class demonstration or an activity. If it is done as an activity each group will need four beakers.

Part II: Prep time: 20 minutes

Materials: For 4 sections with 16 groups (activity done in pairs)

a) 0.0010 M KSCN (need 640 mL: 0.062 g of KSCN in 640 mL of solution)
b) 0.10 M KSCN (need 28 mL: 0.272 g of KSCN in 28 mL of solution)
c) 0.20 M $Fe(NO_3)_3$ (need 48 mL: 3.88 g of $Fe(NO_3)_3{\cdot}9H_2O$ in 48 mL of solution)
d) 0.10 M Na_2HPO_4 (need 28 mL: 0.396 g Na_2HPO_4 in 28 mL of solution)
e) 4 Small test tubes and 1 large test tube

If possible make a stock bottle of 0.0010 M KSCN and 0.20 M $Fe(NO_3)_3$ for the class. Then make individual dropper bottles of 0.10 M KSCN and 0.10 M Na_2HPO_4. This should allow the students to finish the reaction in 25 minutes and reduce confusion with the two KSCN solutions.

Part III: Prep time: Varies

Materials: For 16 groups (activity done in pairs)

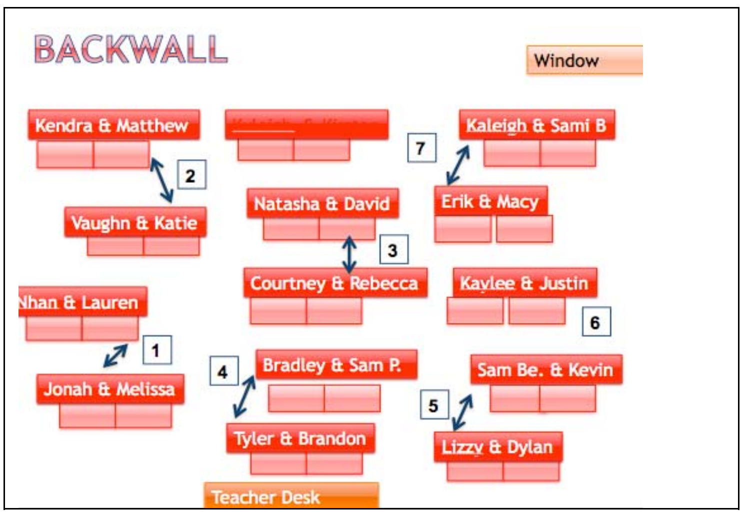
a) Red/Yellow chips: 160
b) Clear, colorless chips: 160

Assembly of Models The red and yellow chips are red on one side and yellow on the other side. These chips can be purchased from www.educationworkshawaii.com (two-color counting chips red and yellow, #LER7566). The 1 inch counting chips are $8.99 +tax/shipping for 200 chips. Clear, colorless chips can be cut from Plexiglass in 1inch rounds or squares. Another option for the colorless chips is to use key rings. They can be purchased from Gallagher Promotional Products (http://www.gppinc.com/). The site sells Chrome Plated Steel Key Rings 7/8" (Qty 250) for $5.50 + tax/shipping. The chips can also be made by gluing red and orange cardstock together and then laminating the sheets. The circles can be punched out of the laminated paper using a 1inch scrapbooking punch (any craft store for $9.00) or cut into squares. The clear and colorless chips can be made from laminating an acetate (overhead) and punched using the 1 inch punch or cut into squares.

Set up: For each group place 10 red/yellow chips and 10 colorless chips in a Ziploc bag. Also include the test tube reaction sheet.

RECOMMENATION FOR CLASSROOM SET-UP

If possible, place desks into pairs and set-up a seating chart with the assigned pairs. On the chart also assign the partner groups (see chart).



The arrows on the chart indicate the pair partners and the numbers indicate their assigned lab tables (Part II). Have students complete Part I in their desks. For section II, have students work with their partner and pair group (4 total) and complete the reactions at the lab tables. Though they are a group of 4, both groups should do the reactions separately. Then for Part III, have students return to their paired desks.

PRELAB ENGAGEMENT / QUESTIONS

Prior to the lab ask the students about reversible reactions they have seen before (physical or chemical). If necessary remind them about freezing/melting or any demonstrations you have done in class. Show the students an example of any reversible reaction. A possible demonstration is the Blue bottle demonstration (http://chemistry.about.com/od/chemistrydemonstrations/ss/bluebottle.htm).

Write a general reaction on the board:

$2MB + O_2 \rightleftharpoons 2MBO$

Colorless Blue

Ask students “what are reactants in the forward direction? What are the products in the forward direction?” Write the reactions on the board.

Tell them today they will be studying reversible reactions and learning about equilibrium systems.

PROCEDURE

Part I (as a demonstration): The goal of Part I is to help students understand that an equilibrium system is a dynamic process. It is also have students graph and interpret data and understand that the amounts of reactants and products are not necessarily equal when a system is at equilibrium. Before class determine which groups will share answers during the Analysis section of Part I. At the beginning of the demonstration, have the students read Part I and construct a table to record their data on a separate sheet of paper. If necessary, help students develop the data table. In order to complete the demonstration quickly, warn the students that you will call out the volumes and they must record them. In other words you (the teacher) are not writing them down nor are you stopping the demonstration until the system has reached equilibrium. Follow the steps of the demonstration until the volume of Beaker A and Beaker B stay the same for 4 continuous cycles. Ask the students if you should keep going. You will have students that say no. Ask them why and they should say the volumes of water are not changing. In groups, have students complete the Analysis sections (graphing and analysis). While students work on this section it is important for the teacher to continuously check for conceptual understanding. Teachers should walk around and ask student groups’ questions such as “why was it important for the water levels to remain the same for 4 continuous cycles?” While answering question #2 students may think that the correct points on the graph are where the two lines intersect. This is because they believe that the reaction is at equilibrium when the reactants and products are equal. Make sure to walk around and correct students of this misconception. Students may need assistance developing a definition for equilibrium. Encourage them to focus on the term “equilibrium” and focus on the graph. Eventually they should recognize that equilibrium is when the amounts of water in Beaker A and Beaker B don’t change. After students have met with their partner groups, talk with individual groups to make sure they understand the definition of equilibrium.

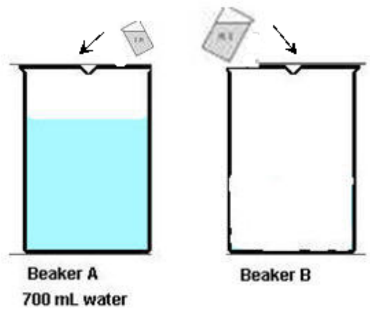
Part II: The purpose of this section is for students to investigate a macroscopic equilibrium system. Before starting, remind students to stir the reactions before making observations. In addition there are two different KSCN solutions (0.0010M and 0.10 M). For the introduction students may struggle with what is meant by the forward and reverse reactions. Remind them that a reversible reaction occurs in both directions and you want them write each reaction separately. If students struggle with questions 4c, ask them "what are the colors of the reactants and products (from introduction)". Then ask them "based on the color of test tube #2 what must be happening to the concentrations of Fe^{3+} and $FeSCN^{2+}$"? Since the test tube #2 is now red, there must be more products ($FeSCN^{2+}$) and presumably less Fe^{3+} as some of it would have to be used up to produce the $FeSCN^{2+}$.

Part III: Before starting Part III, students need to be taught how to write the equilibrium expression (Keq) for an equilibrium system. It is not necessary to teach them how to do any calculations before this part of the activity. The purpose of this section is to give students a particulate understanding of the equilibrium system from Part II. It is also to help students understand equilibrium constants and how they can be used to determine whether a reaction is at equilibrium or not. When introducing Part III, review what it means to have an equilibrium system (Part I). Also review the results of Part II, reminding students that there were three equilibrium systems in Part II. Remind students the colors of each system (orange red, dark red, light orange). Inform them that today they will learn how it is possible to have an equilibrium system even though the reaction is different colors. Some students may question the stoichiometry of this section (2 orange + 6 colorless $\rightleftharpoons$ 3 red). If this is the case, explain that this is an analogy and that the particles were chosen to simplify the calculations. At the beginning of the make sure that students make the correct amount of reactants (2 orange chips and 6 colorless), products (3 products – colorless over red chip) and correctly fill in the chart for question #5. While the students are working, walk around and make sure students have the correct equilibrium expression (Keq) for the reaction. A common mistake is for students to write the ion charges as the exponents. Make sure they use the reaction coefficients from the balanced chemical reaction, not the ions charges in the equilibrium expression. When they are calculating the Keq value (question #7) make sure students are multiplying the concentrations (not adding) and reducing the fractions to lowest terms. Also when manipulating the reactants and products students should not add or remove any addition chips from the reaction. For Question #12, students may not recognize that they need to make reactants into products. Once they realize this, it may be necessary to remind them that the product is one colorless chip on top of one red chip. Once they start making products ask them, "how will you know when to stop making products?". Students should recognize that the system has reestablished equilibrium when the ratio of products to reactants is 1:4. Tell them to make one product at a time and recalculate the ratios (space for work on Question #12), until they have reestablished equilibrium.

PURPOSE: To study equilibrium and understand what happens to the concentration of reactants and products in an equilibrium system.

Part I: Investigating a simple physical equilibrium system. (Jordan)
Materials: 4 beakers: 100 mL and 50 mL and **two** 1000 mL beakers (A and B)

PROCEDURE:



www. genchem1.chem.okstate.edu

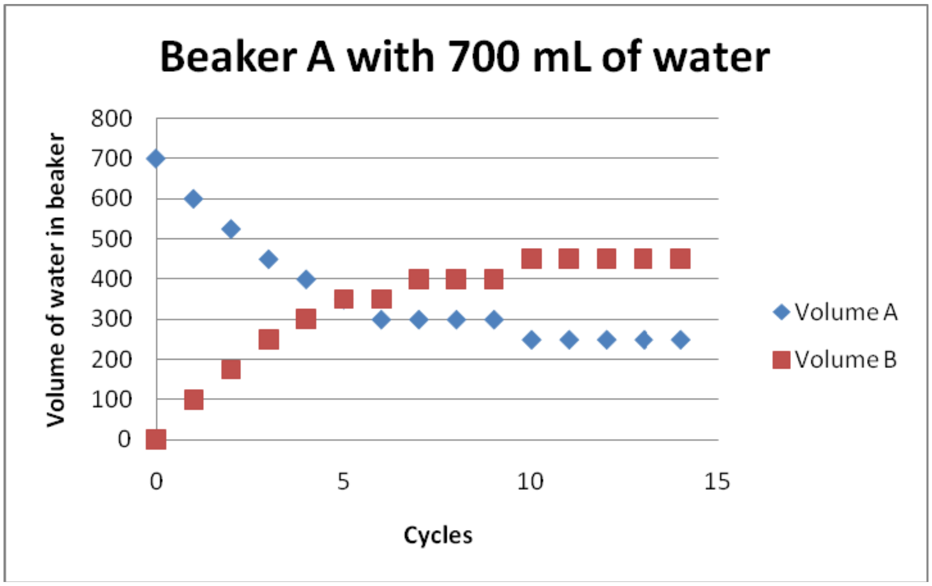
a) Put 700 ml of colored water in the large beaker “A”. Leave the other beaker “B” empty.
b) Record the volume of water in the beakers in your data table (Cycle 0).
c) Transfer water between the large beakers using the following “rules”
    - Use the 100 mL beaker to transfer water from A to B;
    - Use the 50 mL beaker transfer water from B to A.
    - Fill the small beakers as full as possible **without tipping the large beakers** in any way.
    - One cycle consists of one A → B transfer and one B → A transfer.
    - **For each cycle**, record the volume of water in beakers A and B.
d) Continue cycles and recording the volumes, until the level of water in beakers A and B are unchanging for at least four cycles.

DATA TABLE *Sample Data*

Beaker A filled with 700 ml of water		
Cycle	Volume Beaker A	Volume Beaker B
0	700	0
1	600	100
2	525	175
3	450	250
4	400	300
5	350	350
6	300	350
7	300	400
8	300	400
9	300	400
10	250	450
11	250	450
12	250	450
13	250	450
14	250	450
15	250	450

Note: Depending on how carefully water is transferred, it may take more (or less) than 15 cycles to reach equilibrium.

A. Graphing
Graph the volume of water (in beakers A and B) versus cycle. Make sure you label the axis and title the graph. Be sure the reader can distinguish between the A and B points on the graph (use different colors or symbols for points). Trace the points to make smooth curves for each plot.



B. Questions

1. What do think is meant by equilibrium?

Answers will vary somewhat but should focus on the idea of amounts of reactants and products not changing or forward and reverse rates being equal. Inclusion of the dynamic aspect (forward and reverse reactions not stopping) is also an important idea. If students do not include this in their definition you should ask them "does this mean the reaction stops when the amount of reactants and products remain constant?"Encourage students to refer to their graphs.

2. On the graph, circle the points where you believe the reaction has reached equilibrium.

Points on the graph should be when Volume A = 250 mL; Volume B = 450 mL If students choose where the lines cross, that is ok as long as they modify this by question 4.

3. What cycle(s) of the experiment did the reaction reach equilibrium? B) How do you know?

Cycles 10-15

B. *The volumes of water in each beaker do not change.*

4. Compare your ideas about equilibrium with another group (see your teacher for your partner group). Look up the textbook definition of equilibrium and write it here. Does this definition match your definition? Why or why not? With your partner group come to an agreement about where the system reaches equilibrium on your graph(s). If this position is not the same position you circled for question 2, circle the new point on your graph in a different color and indicate the color here.

Answers will vary: In a reversible reaction when the rate of the forward and backward reactions are equal. If they do not choose the correct points in question 2, they should choose those points now.

5. At equilibrium, is the amount of water in Beaker A equal to the amount of water in Beaker B?

No the amount of water in Beaker A and Beaker B are not the same.

6. Based on the graph what must be the same in order for a reaction to be at equilibrium?

The amount of water in Beaker A and Beaker B stays the same (doesn't change from one cycle to another.

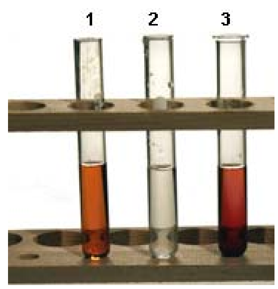
PART II: A macroscopic view of a chemical equilibrium system.
INTRODUCTION: An equilibrium system can be formed in the solution with the following ions:

$$Fe^{+3} \quad + \quad SCN^{-} \quad \rightleftharpoons \quad FeSCN^{+2}$$

Color: of ions	light orange	Colorless	Red

Sample Results: Solutions

Test Tube #1: $Fe(NO_3)_3$
Test Tube #2: KSCN
Test Tube #3: $Fe(SCN)^{+2}$



The above reaction is a reversible reaction.
1. A) Write the balanced equation for the reaction in the forward direction. B. What are the colors of the reactants and products for this reaction?

$$Fe^{+3} \quad + \quad SCN^{-} \quad \rightarrow \quad FeSCN^{+2}$$

B) Reactants: Fe^{3+} *light orange/yellow;* *SCN^{-} colorless; Product: $FeSCN^{2+}$ red*

2. A) Write the balanced equation for the reaction in the reverse direction. B) What are the colors of the reactants and products for this reaction (in the reverse direction)?

$$FeSCN^{+2} \quad \rightarrow \quad Fe^{+3} \quad + \quad SCN^{-}$$

B) *Reactant: $FeSCN^{2+}$ red; Products:* Fe^{3+} *light orange/yellow;* *SCN^{-} colorless*

PROCEDURE

1. Fill a **small** test tube with SCN^- (KSCN) solution. In another **small** test tube put in 10 drops of Fe^{+3} ($Fe(NO_3)_3$) solution. Record the color of the solutions and confirm that the colors are the same as listed above.

SCN^- = colorless; Fe^{+3} = light orange

2. Mix the two solutions together in **LARGE** test tube. Record the color of the solution when the Fe^{+3}, SCN^- and $FeSCN^{+2}$ are all present at the same time.

Red or red-orange (depending on the amount of Fe^{+3} ions added)

3. Take this solution and pour equal amounts into three **small** test tubes (label 1,2,3,). Test tube 1 is the reference tube and tests will be run on test tubes 2,3.

4. To test tube #2 add 5 drops of 0.10 M KSCN(make sure you add the correct solution). Stir and record the results.

Dark Red

b. How does the color of test tube #2 compare to test tube #1?

Much darker than test tube #1

c. Based on the color of the solution in the test tube, what do you think is happening to the concentration of Fe^{3+}? $FeSCN^{2+}$? Justify your answer. *Concentration of $FeSCN^{2+}$ is increasing because solution is getting more red. Concentration of Fe^{3+} is decreasing because you need that to make the $FeSCN^{2+}$.*

5. To test tube #3, add 4 drops of 0.10 M Na_2HPO_4, one at time. This compound will react with Fe^{+3} to form $Fe(HPO_4)^{2+}$ (this is the same effect as removing the Fe^{+3} ion from the reaction). Stir and record the results.

Turned light orange or colorless (if add too much Na_2HPO_4*)*

b. How does the color of test tube #3 compare to test tube #1?

Lost the red color – turned orange

c. Based on the color of the solution in the test tube, do you see a change in the concentration of reactants or products? How do you know?

More reactants are made – reaction becomes orange
Concentration of products decreases because red color disappears.

______ Teacher check

Part III: A Particulate View of a Chemical Equilibrium System (Convery, 2010)

The purpose of Part III of the activity is to understand what is happening at the particulate level in the reactions in Part II of the activity. If known, an equilibrium constant (K) may be used to determine if a system is at equilibrium or not. The following is the generic formula for writing the equilibrium constant.

$aA + bB \rightleftharpoons cC + dD$ a = coefficients of a balanced equation
A = compounds/elements

$$K = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$$

Materials per group: Bingo chips: 2 different colors (10 of two sided chips – red/orange, 10 of the colorless chips), Reaction paper – see last page of Teacher's handout

In this reaction Fe^{+3} is represented by Orange chips and the SCN^- is represented by the colorless chip and the product, $FeSCN^{+2}$ is presented by the red chips.

1. From the previous page rewrite the balanced reaction and colors for each ion.

	Fe^{+3}	+	SCN^-	$\rightleftharpoons$	$FeSCN^{+2}$
Color:	<u>*Orange*</u>		<u>*Colorless*</u>		<u>*Red*</u>

2. Write the equilibrium expression (K) for the reaction below.

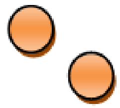
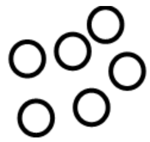
$$K = \frac{[FeSCN^{+2}]}{[Fe^{+3}] \times [SCN^-]}$$

3. A) Based on your equilibrium constant, what particles (reactants or products) are always on top of the equilibrium expression K? B) Which particles (reactants or products) are always on the bottom of the equilibrium expression?
   A) *Products - top*

   B) *Reactants - bottom*

4. This reaction is at equilibrium when there are 2 Fe^{+3} (orange chips), 6 SCN^- (colorless chips) and 3 $FeSCN^{+2}$ (3 color chips on top of 3 red chips). On the reaction paper use the chips to represent this reaction at equilibrium.

5. Fill in the table below with the number of each type of ion in your "test tube."

$[Fe^{+3}]$	$[SCN^{-}]$	$[FeSCN^{+2}]$
Orange Particles	Colorless Particles	Red Particles
2	*6*	*3*
		

6. Look at your reaction does this reaction have equal numbers of reactant particles and product particles? *No. There are more reactants than products.*

7. Using the equilibrium expression (K) and the number of each particles (Question #5), calculate the equilibrium constant for this reaction (leave in fraction form – reduced to lowest terms).

$$K = \frac{3}{(2x6)} = \frac{1}{4}$$

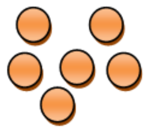
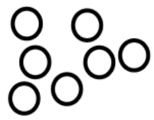
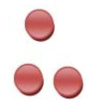
8. The color of a reaction is determined by which colored particles are most prevalent in the container. (In other words if there are 10 green particles and 8 blue particles in a container, the reaction will look green-blue). If two colors have the same number of particles then the color of the reaction is a mixture of both of them (10 red particles and 10 blue particles will look purple). Based on the ions on the paper, what color would we observe if we had moles of these ions in the same ratio as given by the chips?

   The reaction should look reddish orange.

9. One stress that can be placed on an equilibrium system is to change the concentration of the reactants or products. Add 4 orange chips and 1colorless chips to your system. At this time do not touch or adjust the products. Now how many of each chip is in the reaction? Fill in the number on the table for question #10.

Target Inquiry GVSU-2010, Alice Putti, Jenison High School

10. Based on the chips in your "test tube" now, B) If this new reaction is at equilibrium the ratio of reactants and products should equal the value you found for K earlier (that's why it's called a constant). Calculate the value. C) Is this reaction at equilibrium?

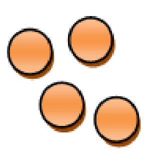
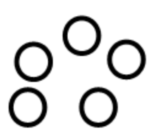
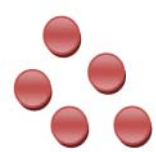
$[Fe^{+3}]$	$[SCN^-]$	$[FeSCN^{+2}]$
Orange	Colorless	Red
6	*7*	*3*
		
B) $\frac{3}{(6x7)}$	$= \frac{3}{42}$ $= \frac{1}{14}$	

C) *Not at equilibrium because it is not at a 1:4 ratio*

11. Is the ratio you calculated from question #10 bigger or smaller than the equilibrium constant K (question #7). Use what you know about fractions and decide if you need more reactants or products to reestablish equilibrium.
The reaction will need more products in order to reestablish equilibrium.

12. Without removing any ions (chips) from the test tube (paper), reorganize (make reactants into products or products into reactants) them such that the system will be at equilibrium. Use the space below to complete any necessary calculations.

13. Once you have successfully reestablished equilibrium, A) Fill in the chart below B) Calculate your equilibrium constant to be sure. C) What did you need to do to change your model so that it showed a system at equilibrium?

$[Fe^{+3}]$	$[SCN^-]$	$[FeSCN^{+2}]$
Orange	Colorless	Red
4	*5*	*5*
		
B) $K = \frac{5}{(4x5)}$	$= \frac{5}{20}$ $= \frac{1}{4}$	

C) *Needed to make more reactants into products.*

14. Based on the information from Question #8, what is the color of the reaction now?

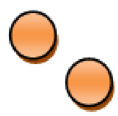
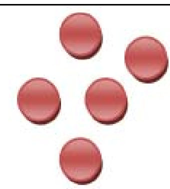
    Reaction should look more red.

15. Generalize your results: Adding reactants to the system resulted in an increase/decrease (circle one) in the concentration of products.

    Add to the reactant side of the reaction resulted in an increase in the concentration of the products.

Removing particles from the equilibrium:

16. To the system in Question #13, remove 2 orange particles and 4 colorless particles (don't touch any of the product particles). A) Now in the table fill in how many of each particle is left. B) Calculate the ratio and compare it to your equilibrium constant K C) Is the system at equilibrium?

$[Fe^{+3}]$	$[SCN^{-}]$	$[FeSCN^{+2}]$
Orange	Colorless	Red
2	*1*	*5*
		
B) $K = \frac{5}{(2x1)}$	$= \frac{5}{2}$	

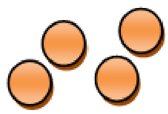
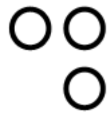
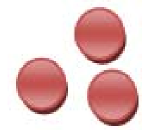
C) *Not at equilibrium because it is not at a ratio of 1:4.*

17. Is the ratio you calculated from question #16 bigger or smaller than the equilibrium constant K (question #7). Use what you know about fractions and decide if you need more reactants or products to reestablish equilibrium.

    Since the ratio is larger than ¼, the reaction will need more reactants in order to reestablish equilibrium.

18. Without removing any ions (chips) from the beaker (paper), reorganize them such that the system will be at equilibrium (make reactants into products or products into reactants). Use the space below for any necessary calculations.

19. Once you have successfully reestablished equilibrium, A) Fill in the table. B) Calculate your equilibrium constant to be sure. C) What did you need to do to change your model so that it showed a system at equilibrium?

$[Fe^{+3}]$	$[SCN^{-}]$	$[FeSCN^{+2}]$
Orange	Colorless	Red
4	*3*	*3*
		
B) $K = \frac{3}{(3x4)} = \frac{3}{12} = \frac{1}{4}$		

C) *Needed to make more reactants into products.*

20. What color is the reaction? Why?
The reaction should look more orange than red.

21. Generalize your results: Removing the reactants from the system resulted in an increase/decrease (circle one) in the concentration of products.
Removing the reactants resulted in a decrease in the concentration of products.

POST LAB DISCUSSION

After completing the activity review the results of Part I. Have students state in their own words what it means to have an equilibrium system. Ask them "what the beakers look like when the equilibrium is established?" "Do the amounts of reactants and products need to be the same when equilibrium is reached?" Then review the results for Part II of the activity. Finally tie the results from Part II to the reaction in Part III. Have students look at the first reactions of Part II. In this reaction KSCN is added to the reaction and as a result the reaction turns red. Ask students "why does the reaction turn red?" Students should recognize that the reaction is making more ($FeSCN^{+2}$) which is why it is red. Now have the students look back at the first reaction of Part III. Ask them "How do you know the reaction in Part III is at equilibrium?" Ask students "why was more ($FeSCN^{+2}$) made in the reaction?" After completing Part III students should know that the reaction makes more products in order to re-establish equilibrium. Students should also understand that the equilibrium constant can help tell us whether or not a reaction is at equilibrium and which direction a reaction will proceed (making more reactants or more products) to reach equilibrium. In addition to having students make connections between the macroscopic in Part II and the particulate in Part III, it is also important for students to be able to explain Le Chatelier's principle using collision theory (why does increasing the concentration of a reactant produce more products). The follow-up questions focus students on these aspects of the activity.

FOLLOW-UP QUESTIONS

1. A student states that an equilibrium system is when the amounts of reactants and products are equal. Why isn't this comment correct? What would you tell the student to correct their error? What is the factor that is the "same"?
The amounts of reactants and products do not need to be equal, the rates of the forward and reverse reactions have to be equal. The ratio of reactants to products must be the same, if a reaction is at equilibrium.

2. A) For Part III, when the reaction is at equilibrium, what is the color of the reaction (remember you had 3 different equilibrium conditions for this reaction, be sure to list the colors for each equilibrium). B) How is it possible for the same reaction to be at equilibrium yet be different colors?
   A) Equilibrium 1: Dark orange (reddish orange) Equilibrium 2: dark red Equilibrium 3: light orange

   B) The reaction can be different colors as long as the ratio of products to reactants is the same. (K = ¼)

3. In parts II and III of this activity we examined something called Le Chatelier's principle which states that when a stress is applied to an equilibrium system, that system shifts to offset the stress.

   A. What was the stress applied in Part II step 4 and Part III step 9?
   Adding a reactant.

   B. What kind of shift did that cause in the equilibrium system?
   It caused the concentration of the products to increase.

   C. Use collision theory to explain why this stress would cause the equilibrium system to shift this way.
   Adding reactants would cause the rate of the forward reaction to increase and if the forward reaction is faster than the reverse reaction more products will be produced than reactants. However, as reactants are used up and more products are produced the reverse reaction will speed up and the forward reaction will slow down until they are at the same rate. At this point the reaction will reach equilibrium again.

   D. What was the stress applied in Part II step 5 and Part III step 16?
   Removing a reactant.

   E. Use collision theory to explain why this stress caused the equilibrium system to shift.
   Removing reactants would cause the rate of the forward reaction to decrease and if the forward reaction is slower than the reverse reaction more reactants will be produced than products. However, as products are used up and more reactants are produced the

forward reaction will speed up and the reverse reaction will slow down until they are at the same rate. At this point the reaction will reach equilibrium again.

F. Can you think of any other stress you could apply to this equilibrium system?
Answers will vary but many will say adding or removing a product.

G. What effect do you predict it would have on the equilibrium system?
Answer depends on answer to F.

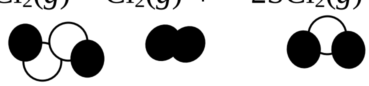
REFERENCES

1. Peg Convery (2010) Personal Contact: Farmington High School, Farmington, MI.
2. J. Orvis and J.A. Orvis, (2005). Throwing paper wads in the chemistry classroom, *Journal of College Science Teaching*, 35(3), pages 23-25.
3. Jordan, A. Introduction to Equilibrium (2009). Retrieved on March 28, 2010, from http://phet.colorado.edu/index.php

ASSESSMENT QUESTIONS

For questions 1 and 2 consider the following reaction

$$S_2Cl_2(g) + Cl_2(g) \rightleftharpoons 2SCl_2(g) \quad K_{eq} = 4$$



1. What is the correct equilibrium expression for this reaction?

   A. $K_{eq} = \dfrac{2[SCl_2]}{[S_2Cl_2][Cl_2]}$

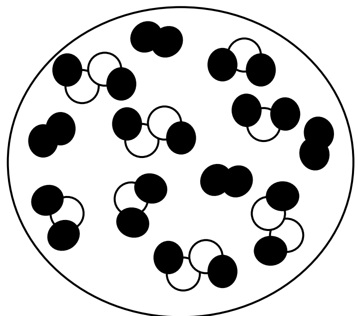
   B. $K_{eq} = \dfrac{[SCl_2]}{[S_2Cl_2]+[Cl_2]}$

   C. $K_{eq} = \dfrac{[S_2Cl_2][Cl_2]}{[SCl_2]^2}$

   D. $K_{eq} = \dfrac{[SCl_2]^2}{[S_2Cl_2][Cl_2]}$

   E. $[S_2Cl_2][Cl_2][SCl]^2$

2. Examine the figure, and determine if the system is at equilibrium. If it is not, in which direction will it proceed to reach equilibrium?

   A. It is not at equilibrium, it will proceed towards products.
   B. It is not at equilibrium, it will proceed towards reactants.
   C. The reaction is at equilibrium.
   D. It is not possible to tell if the reaction is at equilibrium.
   E. To reach equilibrium, the value of K_{eq} must change.

3. Which of the following is always true of a chemical system at equilibrium?

   A. The concentrations of the reactants and products decrease.
   B. The equilibrium shifts to favor the forward reaction.
   C. The equilibrium constant decreases.
   D. The concentration of the products equals the concentration of the reactants.
   E. The forward and reverse reactions proceed at equal rates.

4. Which of the following statements is correct for a reaction that has $K >> 1$?

   A. The forward reaction is faster than the reverse reaction.
   B. The reverse reaction is faster than the forward reaction.
   C. At equilibrium the reaction has a much higher concentrations of products as compared to reactants.
   D. At equilibrium the reaction has a much higher concentrations of reactants as compared to products.
   E. None of these statements is correct.

5. If the following reaction is carried out in a sealed container: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ a state of equilibrium can be reached if the container initially contains:

   A. NH_3 only.
   B. N_2 and H_2 only.
   C. NH_3 and N_2 only.
   D. NH_3, N_2, and H_2.

E. any of these combinations of reactants and product.

Reaction Paper – Test tube