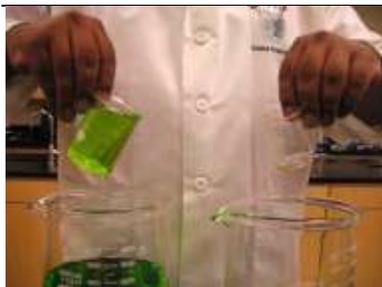


Week 4 • Equilibrium

LABS and DEMONSTRATIONS

Students have difficulty seeing the big picture with equilibrium. Hands-on activities that allow students to process these ideas help them understand. Demonstrations to which you can refer help students visualize the concepts. Here are a few labs and activities that I use with my classes.

The Two Beaker Demonstration:



I mentioned in the lecture notes that I use this demonstration to introduce equilibrium. I make four important points:



- **equilibrium never stops**

Throughout the demonstration, I keep moving water back and forth. Hopefully, my constant motion drives home the idea that equilibrium involves constant change.

- **you can start anywhere and come to the same equilibrium**

You can begin with a full reactant beaker or a full product beaker or half and half and you will come to the same equilibrium. I demonstrate this so that when we see it again in an equilibrium problem, my students will buy into the idea.

- **observable properties are constant because the forward and reverse changes are equal**

If equal amounts of change take place, it is obvious that there is no net change. Measuring the amount of water carried in the two different-sized beakers brings this home to the student.



- **the initial and equilibrium amounts do not have to match the stoichiometry of the equation**

It is a common misconception with these situations that if the reaction is a 1:3:2 ratio, the final products will also be that ratio. Only the change occurs in that ratio.

NOTE: I am using water colored with “Neon” food colors found at the grocery store. They are more expensive, but the colors are great. If I am going to make the students stare at colored water for 10 minutes, it might as well be pretty.

M.O.M. to the Rescue:



Begin by putting some **Milk of Magnesia** (unflavored) into a beaker on a stirrer. Add some water because the stuff is thick (and kind of expensive). Add some universal indicator until you have an observable color. The M.O.M. is purple because of the hydroxide ions in solution. This is a good time to note that even though we say magnesium hydroxide is insoluble, it must be **dissolving a little** or else the solution would be neutral.



Add some 6 M HCl. You will see a flash of pink and then orange, green, blue, and then the solution will turn back to purple. The concentrated acid neutralizes the few OH⁻ ions in solution and lowers the pH. Quickly, however, we have an example of Le Chatelier's Principle as the $\text{Mg(OH)}_2(\text{s}) \rightleftharpoons \text{Mg}^{2+}(\text{aq}) + 2 \text{OH}^{-}(\text{aq})$.



Continue to add HCl and students notice that the color does not return completely to purple and the suspension is noticeably less opaque. Eventually, the solution is clear when all the magnesium hydroxide has dissolved. If you are careful (I never am) you can get the solution to end up green (neutral). Patience is the key (that's why mine never works). You could have a squeeze bottle of 1 M HCl handy for the last few tries. Add some acid and then talk about something else. Come back and observe the beaker throughout the period.

Rainbow Cylinder:



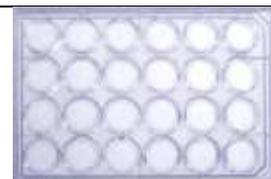
When it comes time to introduce universal indicator, I take a **0.10 M solution of HCl** (usually I dilute whatever molarity acid I happen to have right in the cylinder... it does not need to be perfect) and add enough universal indicator to give it some color. I *carefully* pour some **saturated Na₂CO₃** down the side of the cylinder (see technique at left). The dense, basic solution flows to the bottom. Additionally, small bubbles of CO₂(g) are formed and mix the solution. A pH gradient is formed with purple (basic) solution at the bottom and pink (acidic) solution at the top. This works just fine in a 100-mL cylinder, too.



NOTE: I make a solution of saturated Na₂CO₃ and keep it in my chemical cabinet and this year is the first time I have had to make more solution even though I have been doing this demonstration for about five years.

Serial Dilution & Universal Indicator:

I have students make a serial dilution of 1 M HCl and another of 1 M NaOH in a 24-well plate beginning with 20 drop of 1 M solution. By moving 2 drops to the next well and adding 18 drops of distilled water, each cell is 1/10th as concentrated as the one before (and one pH point different). Repeat this 5 times with each solution. To each well add a drop of universal indicator and you can get an array of colors showing each pH value.



Alka Seltzer Labette:



This is one of my favorite labettes. It allows students to **discover** what a buffer is before you **define** it. Having experience with a buffer first, helps students make sense of all the discussion. Begin with four plastic cups (or beakers) half-filled with distilled water. Add the same amount of universal indicator to each until they all look the same with



a neutral green color (if your water is a little acidic, that's OK as long as all four are the same color). Ask students to add 1 M HCl dropwise to one beaker to see how many drops it takes to get a color change (it takes one drop). Do the same in another beaker with 1 M NaOH. Now, break an Alka Seltzer in half (I use the cheaper generic brand) and add it to the other two beakers. They usually turn pink and then come back to a champagne color. Repeat the dropwise addition of acid to the Alka Seltzer solution looking for the same color as the other acidic cup. It will take between 60 and 120 drops to get to the same color. The same occurs with the base. We now have an **operational definition** (a test) of a buffer. It takes a lot of acid or a lot of base to change its pH. Later in the labette, we make a buffer using a **half-titration** of a diluted vinegar solution. We can use this same test to see whether it acts as a buffer. It is not as good a buffer because of the diluted vinegar (but I found that using pure vinegar uses up a LOT of NaOH). This gives us a chance to observe the property of **buffering capacity**, too.

NOTE: I buy plastic cups in bulk at Sam's Club and use them as "disposable beakers". This is a time saver and I don't have to worry about contamination from lab to lab. In a lab like this, a good rinse is usually sufficient to remove the chemicals, so I can re-use the cups from period to period. We use beakers, too.

Hydrogen Chloride and Ammonia Gas:



You need to keep your nose away from an open bottle of concentrated HCl because of the HCl(g) that leaves the bottle. In the same way, concentrated ammonia gives off NH₃(g). When they mix in the air, however, they make great examples of a **Brønsted-Lowry acid and base**.

Arrhenius only dealt with **aqueous solutions**. When NH₃ and HCl meet in the air, a proton (H⁺) is transferred from the HCl to the NH₃. The two resulting ions immediately are attracted to each other to form the solid, NH₄Cl(s) which we see as smoke. HCl(g) is a B-L **acid** because it **donates a proton**. NH₃(g) is a B-L **base** because it **accepts the proton**.

If you focus on the bond formed between the lone pair on the NH₃ and the empty orbital on the H⁺ ion, this is also an example of a **Lewis acid-base reaction**. NH₃ (“have pair will share”) is the **Lewis base**. The bond formed is a **coordinate covalent bond** with the shared pair of electrons being supplied solely by the nitrogen atom.

K_{sp} of Ca(OH)₂ Lab:

In the Sally Ann Vonderbrink lab book, there is an elegant little lab to find the K_{sp} of Ca(OH)₂. In a well-plate you make a serial dilution of calcium nitrate and then add sodium hydroxide to each well. The calcium ion concentration increases in each well. At some point, a precipitate of Ca(OH)₂ forms. You can calculate the concentration of the ions and use them to determine the K_{sp} of Ca(OH)₂. You repeat the procedure making a serial dilution of the other reactant.

Note: I am writing this from memory and am not certain of the procedure.

Le Châtelier’s labs and demonstrations are plentiful.

I actually have three (maybe four) in my repertoire. I use them all, but at different times for different reasons.

Chromate Dichromate Equilibrium:



Looking for a picture of this equilibrium I found **another** great web site that describes this equilibrium, is interactive, and has movies! Check it out:

<http://www.chem.uiuc.edu/webFunChem/chromate/ChromateIndex.htm>

I use this equilibrium to help students understand why these two strong oxidizers are pH dependent. That is, you can have an acidic, basic, or neutral solution of permanganate and there are different products depending on the situation.

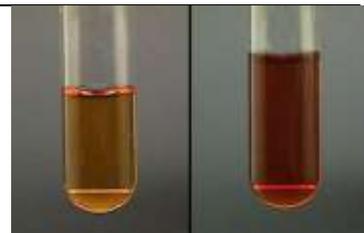
You cannot have an **acidic solution of yellow chromate**, however, because it turns into orange dichromate. In the same way, there is **no basic solution of dichromate** because it turns into chromate. This is an easy demonstration to set up, so I often use it with my general chemistry class for Le Châtelier’s principle, but I use it when I am discussing strong oxidizing agents in AP chemistry.

Iron Thiocyanate Equilibrium:

Pale yellow Fe³⁺(aq) reacts with colorless SCN⁻(aq) to form the red-brown complex ion, FeSCN²⁺(aq). This complex ion doesn’t follow the general rules of other complex ions, however, it does appear on the AP exam rather frequently. It is worth doing a quick labette to make certain students have seen this example. I do it in four Petri dishes. See my handout.

Here is the somewhat interactive website where I stole the pictures:

<http://www.capital.net/com/vcl/equil/fescn.htm>



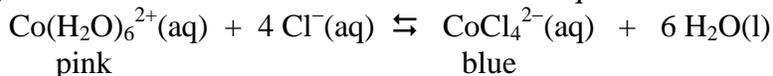
Cobalt(II) – Tetrachlorocobalt(II) Ion Equilibrium:



This is my favorite of the Le Chatelier's demonstrations and one that I use with my AP classes. Flinn published a great version of this demonstration in one of their Flinn Fax. and sells a laboratory version. Look for the write-up in the Week 4 Course Content.

I like it because it is easy to set up (always one of my criteria), visually pleasing, and gets into some great concepts including complex ions (always thinking about predicting reactions), exothermic/endothemic, and Le Châtelier's Principle.

I especially enjoy alerting students to a common error. Consider the equilibrium involved:



When water is added, students often think that the stress is too much $\text{H}_2\text{O}(\text{l})$ which will cause the equilibrium to shift to the left changing from blue to pink. This is what occurs. However, liquids cannot change their concentration and therefore **do not affect the equilibrium**. The actual stress on the system is that you have diluted the Cl^- ion causing the equilibrium to shift toward the reactants.

Copper(II) – Copper Chloride Complex Equilibrium:

This is a similar system to the cobalt ion above. Blue $\text{Cu}^{2+}(\text{aq})$ can be turned into the green complex ion by adding $\text{NaCl}(\text{s})$ to the solution. You adjust the tube to a moderate color and then test the solution in hot and cold water to determine whether the reaction is exothermic or endothermic. It is easy and relatively safe. The cobalt demonstration uses concentrated HCl . The equilibrium is $\text{Cu}(\text{H}_2\text{O})_4^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightleftharpoons \text{Cu}(\text{H}_2\text{O})\text{Cl}_2(\text{aq})$

blue

green

