

Exploring Chemical Equilibrium for Alcohol-Based Cobalt Complexation through Visualization of Color Change and UV–vis Spectroscopy

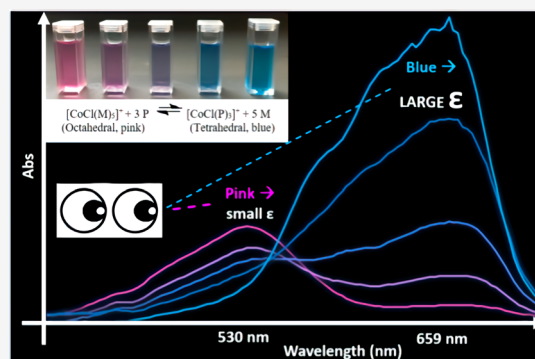
Jolie Ren, Tiffany Lin, Leonard W. Sprague,[✉] Iris Peng, and Li-Qiong Wang*[✉]

Department of Chemistry, Brown University, Providence, Rhode Island 02912, United States

Supporting Information

ABSTRACT: Introducing chemical equilibrium concepts in undergraduate general chemistry promotes improved understanding of chemical reactions. We have developed an engaging laboratory experiment exploring the equilibrium of cobalt complexation in alcohols using UV–vis spectroscopy and successfully implemented in a large general chemistry class of 378 students at Brown University. The octahedral to tetrahedral (pink to blue) cobalt complex transition generates vivid visualizations, increasing students' interest in learning. The equilibrium constants can be measured using UV–vis absorption spectroscopy and the Beer–Lambert law. Vast differences in molar absorptivity coefficients between octahedral and tetrahedral geometries of cobalt complexes prompt discussions on absorptivity, orbital splitting, and color change under the purview of learning Le Châtelier's principle. Additionally, the experimental results regarding the equilibrium constant allowed students to examine possible mechanistic pathways. Student responses to conducting the experiment were positive, most notably because this experiment encouraged them to analyze their experimental results critically and propose possible reaction mechanisms and equilibrium expressions while appreciating the sharp color transition that the complexation equilibrium undergoes.

KEYWORDS: First-Year Undergraduate/General, Laboratory Instruction, Physical Chemistry, Hands-On Learning, Problem Solving/Decision Making, Equilibrium, Coordination Compounds, UV–vis Spectroscopy



INTRODUCTION

Understanding Le Châtelier's principle is important for student learning in a general chemistry course, as it illustrates the stabilizing behavior of chemical reactions and how they are impacted by internal or external factors. To improve comprehension of Le Châtelier's principle, laboratory exercises often visualize concepts such as equilibrium using observable phenomena. For instance, color change in cobalt (Co) complexation^{1,2} is an ideal identifier of reaction progress because it is both qualitatively distinguishable (when significant color change occurs) and quantifiable via instrumentation. UV–visible spectrophotometers are commonly used instruments for probing color-change equilibria in laboratory environments, and such equipment holds continued relevance in academia and industry to this day.³ To reinforce learned concepts about how light interacts with matter, it is important for students to take advantage of such concepts using spectroscopic methods, such as the UV–visible spectrophotometer which is accessible in many undergraduate institutions.

Methods of shifting equilibrium are abundant, but varying concentration allows us to utilize the Beer–Lambert law (assuming 1 cm for path length):

$$A = \epsilon c \quad (1)$$

where A is absorbance, ϵ is the molar absorptivity, and c is the concentration of the absorbing complex. Calibrated UV–visible spectra are then used to identify concentrations of an ultraviolet or visible absorbing complex in solution. Although the absorbance and concentration are directly proportional, the molar absorptivity coefficient, ϵ , drastically impacts the scale of this relationship. A reactant with a large ϵ needs only a low concentration to be visible to the eye. However, if the product's ϵ is much smaller than the reactant's, pushing the reaction forward will not allow the product to be readily observed. Reactant versus product ϵ values can be different by multiple orders of magnitude, making observation of the color change difficult when the reactant $\epsilon \gg$ product ϵ . Conversely, it becomes much easier to visibly observe the color change when the reactant $\epsilon \ll$ product ϵ . This presents a crucial opportunity for instructors to engage and instruct students on the effects of molar absorptivity differences on research and observational techniques.

Received: March 31, 2019

Revised: December 8, 2019

In this study, we report a new and engaging general chemistry laboratory experiment that explores Le Châtelier's principle of chemical equilibrium for alcohol-based cobalt complexation through visualization of color change via UV–vis spectroscopy. This experiment takes advantage of a much higher molar absorptivity for a blue Co complex by shifting the chemical equilibrium from a pink octahedral to a blue tetrahedral cobalt complex through small, varying additions of 2-propanal, which allows students to clearly observe the vivid color changes as new equilibrium states are established without using corrosive HCl. Vast differences in molar absorptivity coefficients between octahedral and tetrahedral geometries of cobalt complexes prompt discussions on absorptivity, orbital splitting, and color change under the purview of learning Le Châtelier's principle. Additionally, the experimental results regarding equilibrium constants allow students to propose alternative mechanistic reaction pathways.

Several undergraduate laboratories and demonstrations^{1,2,4–12} have used Co complexation reactions for teaching equilibrium. Some experiments demonstrated the vivid color change by shifting the chemical equilibrium of Co complexes from pink octahedral to blue tetrahedral complexes by using HCl as a ligand.^{1,5,9} Other work investigated the blue-to-pink transition by changing temperature in an alcohol-based Co complexation where the temperature effects on thermodynamic parameters were the focus.² However, no prior work has used alcohol ligands to shift the chemical equilibrium from pink to blue Co complexes via alcohol concentration. This experiment allows a range of colors, from pink to blue, to be clearly visible as the reaction shifts from an alcohol-based pink complex to a blue complex. Notably, this clear color change benefits student engagement and is achievable without HCl.

This laboratory experiment was developed and successfully implemented to a large general chemistry class of 378 students through a collaborative effort between the faculty and the undergraduates supported by Brown University's Undergraduate Teaching and Research Award (UTRA).¹³ Involving outstanding undergraduates in this development process allowed us to create an experiment that would be at the appropriate skill and knowledge level of our general chemistry students, promoting academic engagement.

The general chemistry course (CHEM 0330) at Brown University is required for many science majors and consists of lectures, prelab lectures, and laboratory sessions. It is a fast-paced course that covers quantum mechanics, basic molecular orbital theory, thermodynamics, chemical equilibrium, and kinetics within one semester. The lab portion of the course is designed to correlate with the lecture content. Instructors often do not cover some chemistry topics such as traditional acid–base chemical equilibria as students are expected to have a relatively strong chemistry background, holding either an AP Chemistry score above 4 or prior enrollment in CHEM 0100, a lower-level introductory chemistry course.

As a result of the CHEM 0330 pedagogy, students are exposed to advanced concepts early in lectures and laboratories. For example, students were introduced to absorption spectroscopy, d–d orbital splitting, conjugated π orbitals, and plasma resonance in an "Origins of Color" lab given earlier in the semester to correlate with the lecture material on quantum mechanics, molecular orbitals, and spectroscopy. Due to the prior lab/lecture preparation and the prerequisites, this Co complexation equilibrium experiment given in the second half of the semester to complement the

chemical equilibrium lecture is ideal for our general chemistry students and is at an appropriate learning level. Based on the student feedback, this experiment did not overwhelm them, and on the contrary, the introduction and application of new concepts along with the visualization of vivid color increased students' motivation and stimulated their interests in learning.

This experiment can be adapted to a high school or to other undergraduate curriculum, emphasizing chemical equilibrium through visualizing color change with or without introducing advanced topics, depending on students' learning levels. For our general chemistry students, our key learning objectives for this experiment are to

- (1) Provide a tangible and visual example of chemical equilibrium concepts, including Le Châtelier's principle, taught in class; for example, in which direction does the cobalt complex solution's color and therefore the equilibrium shift if students add more (colorless) reactant?
- (2) Apply UV–vis absorption spectroscopy and the Beer–Lambert law to determine equilibrium constants.
- (3) Encourage students to carry out critical analyses of the results on the measured equilibrium constants through prelab discussion, in-class queries, and the lab report.
- (4) Introduce students to new concepts (coordination complexes, their geometries, and the Beer–Lambert law) and reinforce learned concepts/techniques (d–d orbital splitting, chemical origin of color, absorption spectroscopy, absorptivity) taught in prior lectures and laboratories.
- (5) Allow students to make connections among different concepts involved in this experiment including solution color, absorptivity, coordination complex geometries, d–d orbital splitting, and chemical equilibrium.

Background: Alcohol-Based Cobalt Complex Laboratory Exercise

For discussion purposes, the *forward* chemical equilibrium is defined by shifting the blue tetrahedral complex to the pink octahedral complex, and is described by



where P is isopropyl alcohol, M is methanol, $a + b = 6$, and $a + n = 4$, conserving observed coordination geometry. The number of ligands bonded to the metal Co^{2+} ion (coordination number) depends on several factors such as the type, charge, and size of the ligands. One Co^{2+} ion can accommodate six methanol (CH_3OH) molecules to form an octahedral coordinated complex. As a result of minimizing steric hindrance, a tetrahedral complex is formed when the Co^{2+} ion reacts with larger and bulkier isopropyl alcohol $\text{CH}_3\text{CHOHCH}_3$ molecules, as opposed to CH_3OH (see [lab manual](#) for more details).

The above *forward* chemical equilibrium has been developed and carried out in previous semesters of our general chemistry laboratory courses, prior to this new experiment. Calculations of equilibrium K_{eq} constants were accurate for this reaction, described by



and the blue color was prominent, making for simple and straightforward laboratory reports and discussion. However, no color change from blue to pink was observed when increasing amounts of methanol were used to shift the reaction to the

right. Due to the severe differences in molar absorptivity (explained by the aforementioned reactant $\epsilon \gg$ product ϵ), the pink cobalt complex was unobservable by visual inspection. Because the students were unable to perceive a concrete visual of the shift in equilibrium without using the UV–vis spectrophotometer, many were disappointed about the lack of aesthetic results and even baffled at the apparent failure of the experiment. As a consequence of the students' responses, we were motivated to design this lab to engage first-year undergraduates in general chemistry more effectively and thus developed a clearer visualization of the vivid colors in the cobalt equilibrium. This experiment prompts deeper analysis by encompassing visualization difficulties incurred in practical settings, such as variations in molar absorptivity or in reactants and products.

This new experiment reverses the reaction in eq 2 and can be described by the following general equation (variables are constrained, just as with eq 2):



or, by assuming equivalent stoichiometry to the forward equilibrium (eq 3), would be



Reversal of the complex equilibrium, and thus reversal of the ϵ problem, resolves the difficulty present in previous laboratory exercises regarding visualization of the blue complex: both pink and blue complexes are now observable via straightforward visual inspection.

Along with better qualitative determination via improved color clarity, which allows for the aforementioned molar absorptivity analysis and discussion, students are exposed to a new quantitative perspective. In our accompanying lab report (see Supporting Information), we presented data from our previous experiment investigating blue to pink complexation in order to allow students to analyze the discrepancy between the forward (eqs 2 and 3) and reverse reactions (eq 4), most notably in differing K_{eq} values. This new laboratory exercise will focus on the reverse reaction (eq 4), but students are asked in their lab reports (see Supporting Information) to discuss the mechanisms for both reactions: (1) the forward blue complex \rightarrow pink complex and (2) the reverse pink complex \rightarrow blue complex. Through introducing students to the relationship between K_{eq} values (highlighting the K_{eq} values' consistency at constant temperature) and stressing the impact a proposed mechanism has on K_{eq} , this quantitative analysis acts as a primer for posing and investigating authentic research problems in the future.

LABORATORY DESCRIPTION

Prelaboratory Preparation

Before their laboratory sections, students are required to finish the prelaboratory exercise questions consisting of preparatory material on Le Châtelier's principle, read the lab manual to learn about light and matter interactions (see Supporting Information), prepare their lab notebooks, and attend a 50 min prelaboratory lecture led by a faculty member. During this lecture, the professor explains the experimental procedures in detail, reviews safety precautions, and explains any new concepts that were not covered in the lecture part of the course (e.g. coordination complex geometry and the Beer–Lambert law).

Laboratory Details

This experiment was designed to fit comfortably in the students' 4 h laboratory period, with all students completing the lab within 2–3 h. The spectrometers used were SpectroVis models, purchased from Vernier Software and Technology (Beaverton, OR), which were paired with Vernier LabQuest for data collection. Prior to the experiment, a stock solution was created with 0.1792 g of $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ and 15 mL of methanol to make a magenta 0.05 mol L^{-1} solution. Students were given the molar absorptivity coefficient, ϵ , for the blue tetrahedral complex, $338.51 \text{ L mol}^{-1} \text{ cm}^{-1}$, in their laboratory reports in order to consolidate laboratory time and prompt comparison with the pink ϵ for discussion. Students were also provided data from the tetrahedral–octahedral equilibrium (eqs 2 and 3) for comparison. The detailed experimental procedures can be found in the lab manual (see Supporting Information).

In our experiment, we made use of two different modes of absorbance measurement: full-spectrum measurement for λ_{max} identification and single-point measurement for determination of ϵ (wherein students are reminded of ϵ 's correlation to electron transition probability). To determine ϵ , students were instructed to plot concentration, c , against absorbance, A , for each solution.

Identifying λ_{max} and ϵ for the Octahedral Complex

Students shifted the equilibrium from the pink octahedral complex to the tetrahedral coordination by adding 2-propanol to the $\text{Co}^{2+}/\text{MeOH}$ stock solution. To identify the λ_{max} and ϵ for the octahedral complex, students were given concentrations of three 3.0 mL calibration solutions to be created, labeled A, B, and C: 0.05, 0.033, and 0.017 mol L^{-1} , respectively (Table 1). To visualize this octahedral–tetrahedral (pink to blue)

Table 1. Sample Absorbance Data and Color Observation for Solutions A, B, and C

solution	volume of stock CoCl_2 solution (mL)	volume of methanol (mL)	solution color	$[\text{Co}(\text{oct})]$, mol L^{-1}	absorbance at λ_{max}
A	3.000	0.000	pink	0.050	0.433
B	2.000	1.000	light pink	0.033	0.292
C	1.000	2.000	very light pink	0.017	0.164

equilibrium, students were asked to make new 3.0 mL solutions, as outlined in Table 2. Samples of each solution were extracted and transferred to a well plate for visual comparison.

In order to observe the reverse equilibrium (eq 4), students prepared Table 2 solutions and analyzed the full spectrum on the LabQuest. The absorbance values and the given

Table 2. Compositions for Solutions 1–5

solution	$0.05 \text{ mol L}^{-1} V_{\text{stock}}$ mL	$V_{2\text{-propanol}}$ mL	V_{total} mL
1	3.000	0.000	3.000
2	2.000	1.000	3.000
3	1.500	1.500	3.000
4	1.000	2.000	3.000
5	0.500	2.500	3.000

equilibrium constant expression,² derived from eq 4, were used to determine the equilibrium constant for each solution in order to quantify and compare both directions of the cobalt complex equilibrium. For solution 3, representing the coexistence of two different complex species (and thus two distinct ϵ) in solution, students were asked to plot absorbance A against wavelength λ .

HAZARDS

Cobalt(II) chloride (CoCl_2) is toxic and an irritant. 2-Propanol is an irritant, and methanol is toxic. Methanol and 2-propanol are volatile and flammable. To minimize vaporization, which will alter concentrations, cover all beakers with watch glasses. Avoid skin contact with solutions, and promptly wash any contaminated skin. Wear goggles and aprons at all times and perform all steps in a fume hood.

RESULTS AND DISCUSSION

Visualization of Equilibrium through Color Change

Figures 1 and 2 display varying colors corresponding to the solutions prepared based on the compositions given in Tables

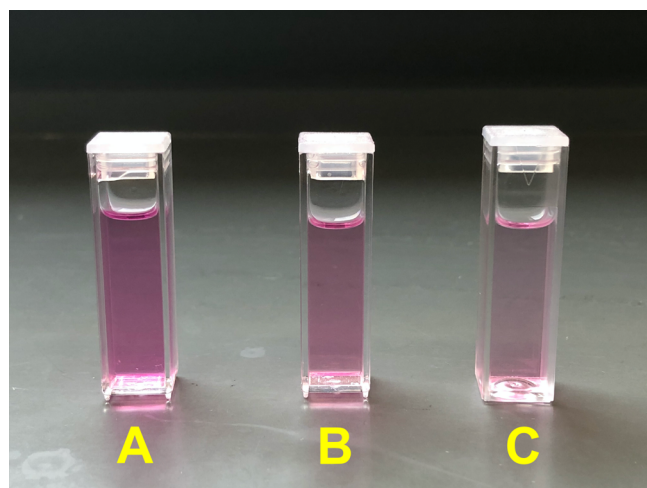


Figure 1. Colorful solutions prepared using the composition given in Table 1.

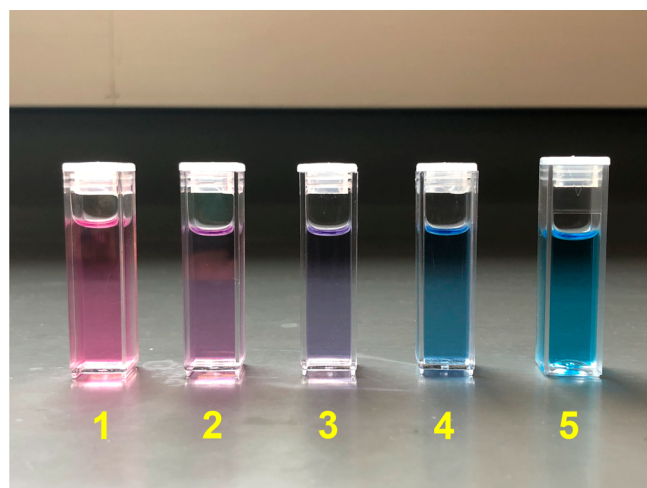


Figure 2. Colorful solutions prepared using the compositions given in Table 2.

1 and 2, respectively. For the pink octahedral complex solution, most students determined the λ_{max} to be 529.52 nm (blue tetrahedral λ_{max} was given as ~ 659.04 nm; see [Supporting Information: Instructor's Notes](#)).

Students were then able to use their values of λ_{max} for the octahedral coordination to find the concentration of the octahedral complex, $[\text{Co}_{(\text{oct})}]$, in mol L^{-1} , and plot $[\text{Co}_{(\text{oct})}]$ against absorbance at λ_{max} on a computer graphing program. Figure 3 (top) displays a fitted sample absorbance vs concentration curve of solutions A to C for the pink complex calibration. The values of ϵ obtained from the students' data ranged from $8.37\text{--}9.81 \text{ L mol}^{-1} \text{ cm}^{-1}$, which are reasonable values for the pink octahedral coordination.

Regarding the intensity of color, most students speculated that the spectrometer "was more sensitive" to blue than pink. Though the tetrahedral solution was determined to be 10 times less concentrated than the octahedral solution, the absorbance of this blue solution was about three times greater than the pink (Figure 3). The key difference between the two solutions is in their molar absorptivity coefficients. As expected, the absorbance at 650.16 nm for the blue complex stock solution as a function of the concentrations gives a much larger ϵ of $353.6 \text{ L mol}^{-1} \text{ cm}^{-1}$. Despite having a lower concentration than the octahedral pink solution, the blue tetrahedral solution has an absorbance and molar absorptivity coefficient that is much higher than that of the octahedral complex (~ 1.2 vs ~ 0.5 Abs)—thus reinforcing the notion that d-electrons are transitioning much more frequently in the blue complex than the pink upon absorbing photons.

Students were asked to compare this purple mixture to the blue and pink stock solution spectra given in the lab report (see Figure 4 and [Supporting Information](#)). At pink $\lambda_{\text{max}} \sim 529.52$ nm for solution 1, the absorbance was observed as ~ 0.433 nm, whereas the maximum absorbance at blue $\lambda_{\text{max}} \sim 659.04$ nm for Co tetrahedral stock solution was 1.035. Unlike the distinct peaks of the blue and pink solutions, the purple solution 3 contained two overlapping peaks rather than a singular peak. Through this observation, students accurately concluded that these separate peaks indicated that both pink and blue complex species with two distinct molar absorptivity coefficients were present in solution 3.

After analyzing each solution, students were encouraged to array their extracted solutions in a well plate to better appreciate the five solutions taken from this equilibrium, noting the gradation in color among these solutions (Figure 5) and giving them an aesthetic view of the chemical equilibrium. A clear representation of the equilibrium shift is also apparent when overlaying all full spectrum scans for each solution (Figure 6).

Comparing the Forward and Reverse Equilibrium Constants and Analyzing Mechanisms

To calculate K_{eq} for each solution, students were first instructed to calculate the concentration of the octahedral complex, $[\text{Co}_{(\text{oct})}]$ and the concentration of the tetrahedral complex, $[\text{Co}_{(\text{tet})}]$, separately, using their previously determined absorbances for the two corresponding λ_{max} values and the Beer–Lambert law. Students were also asked to calculate the concentrations of 2-propanol and methanol and then were given the following equilibrium constant expression, based on the eq 5, for their laboratory reports:

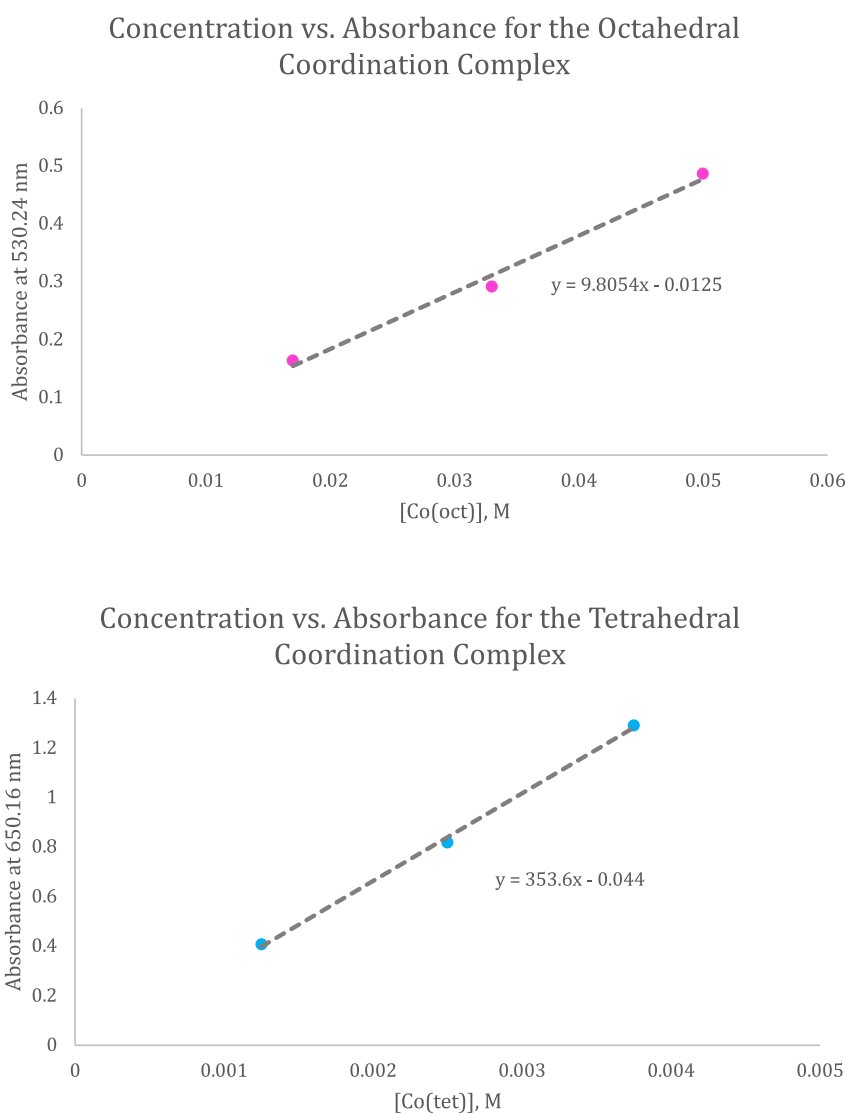


Figure 3. A fitted absorbance vs concentration curve of solutions A–C for the pink Co complex calibration (top) along with the fitted curve (bottom) for blue Co complex (prepared by dissolving CoCl_2 in 2-propanol) calibration.

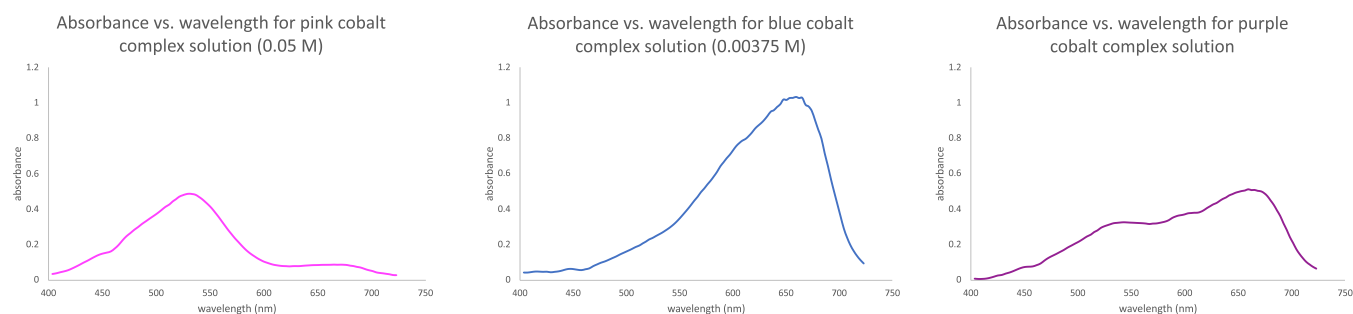


Figure 4. Absorbance vs wavelength for solution 1 (left, 0.05 M stock pink octahedral complex), the blue stock solution (center, 0.00375 M tetrahedral complex), and solution 3 (right, near-equal pink and blue complex concentrations).

$$K_{\text{eq}} = \frac{[\text{CoCl}(\text{P})_3][\text{M}]^5}{[\text{CoCl}(\text{M})_5][\text{P}]^3} \quad (6)$$

It was found that the K_{eq} values for this equation were not consistent at a constant temperature (Table 3), and therefore eq 5 was not an accurate representation for the pink-to-blue complex reaction mechanism. In addition, K_{eq} values were not consistent for all other reactions with a different combination

of a , b , and n numbers given in the general equation (eq 4). Therefore, we proposed that this reverse equilibrium (eq 4) may involve multiple reactions or entail a different coordination than the aquo-cobalt reactant being mostly solvated in methanol.

In our lab report, we provided students with previously collected data from the forward equilibrium experiment (eqs 2 and 3), which looks at the equilibrium from a blue to a pink

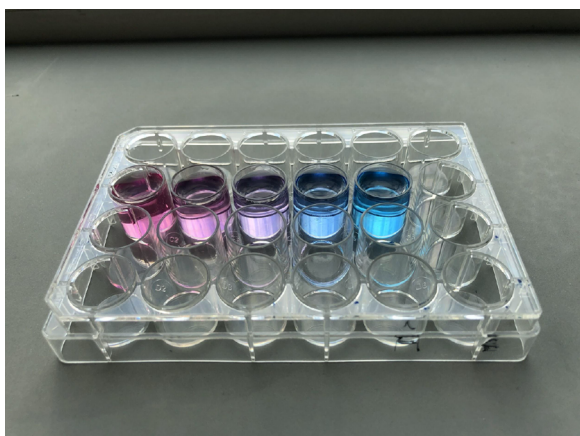


Figure 5. Solutions 1–5, arranged in a well plate to highlight the visual aspect of our cobalt complex equilibrium experiment.

solution. By allowing students to examine two experiments (one given and one self-enacted), students gain information about both the forward and the reverse equilibrium reactions while having carried out only one. The K_{eq} values for the reverse equilibrium were far less consistent than the values for the forward equilibrium (see [Supporting Information: Lab Report](#)), as large discrepancies between solutions taken at the same temperature were apparent. The inconsistency of the K_{eq} values at a given temperature for the reaction from pink to blue complexes is the “mystery” that invited students to think more critically, as they were asked to consider the possible phenomenon of this equilibrium.

Students were tasked with brainstorming reasons for why this inconsistency might exist, recalling previously covered topics such as experimental error, coordination chemistry, and equilibrium topics in general. At first, they attributed this K_{eq} inconsistency to poor spectrophotometer detection of pink and blue colors. Students then regularly determined or learned that the large discrepancy in K_{eq} for the reverse reaction (eq 4) was not due to experimental errors since fairly consistent results were observed in the forward reaction (eq 2). Eventually, through guided discussions with teaching assistants or

professors, many realized that the K_{eq} values’ inconsistency in the reverse mechanism may be due to multiple reactions occurring simultaneously. As K_{eq} calculations are highly accurate if only *one* mechanism is occurring in solution, the presence of multiple simultaneous mechanisms and multiple other possible complex species would interfere with our K_{eq} calculations for this reverse equilibrium reaction.

Additionally, we realized that students may benefit from exploring what combination of stoichiometric values enforce the K_{eq} consistency for the blue to pink transition, using expected cobalt geometries including all possible ligands, by varying the values in a more general expression:

$$K_{\text{eq}} = \frac{[\text{CoCl}_a(\text{H}_2\text{O})_b(\text{P})_n]^d [\text{M}]^e}{[\text{CoCl}_a(\text{H}_2\text{O})_b(\text{M})_e]^d [\text{P}]^n} \quad (7)$$

With $a + b + n = 4$ and $a + b + e = 6$ to match the tetrahedral and octahedral geometries, respectively. Students are then invited to work backward from their findings to try and figure out why that ratio works by rebalancing the equilibrium equations (see [Supporting Information: Instructor’s Notes](#) for a more detailed discussion of stoichiometry).

STUDENT FEEDBACK

According to the postlab survey that was distributed to 378 students, roughly 70% of respondents “strongly agreed” or “agreed” that they enjoyed this experiment and learned new concepts and methods ([Figure 7](#)). In their comments, students appreciated the nuances of cobalt complexation and the complex nature of equilibrium, and the emphasis on equilibrium visualization allowed them to engage more readily with the process of creating solutions for analysis. Many students appreciated the visual nature of this experiment, acknowledging that this experiment was a refreshing take compared to the formula-filled textbook. Because of the higher level of student engagement and the more prominent visuals, the relationship between the molar attenuation coefficient ϵ and color intensity through spectrophotometry was more easily understood. Some students suggested creating an additional online instructional video to help them to make the connection among different concepts learned and to understand the reason

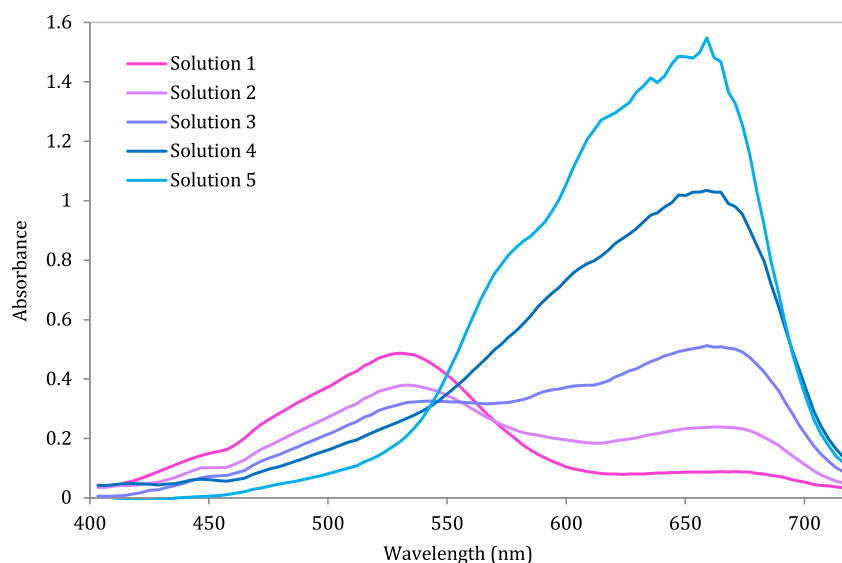


Figure 6. UV–vis spectra for solutions 1–5.

Table 3. Sample of Students' Lab Report Data for the Equilibrium Constant for Solutions 1–5

solution	stock, mL	2-propanol, mL	absorbance at $\lambda_{\max} = 529.52$ nm	absorbance at $\lambda_{\max} = 659.04$ nm	$[\text{Co}_{(\text{oct})}]$, M	$[\text{Co}_{(\text{tet})}]$, M	[2-propanol], M	[methanol], M	K_{eq}
1	3.000	0.000	0.435	0.069	0.049	0.000203	0.00	24.7	N/A
2	2.000	1.000	0.364	0.246	0.041	0.000727	4.36	16.5	259.27
3	1.500	1.500	0.310	0.530	0.035	0.001566	6.54	12.3	45.94
4	1.000	2.000	0.334	1.174	0.038	0.003470	8.72	8.23	5.27
5	0.500	2.500	0.194	1.659	0.022	0.004901	10.90	4.11	0.203

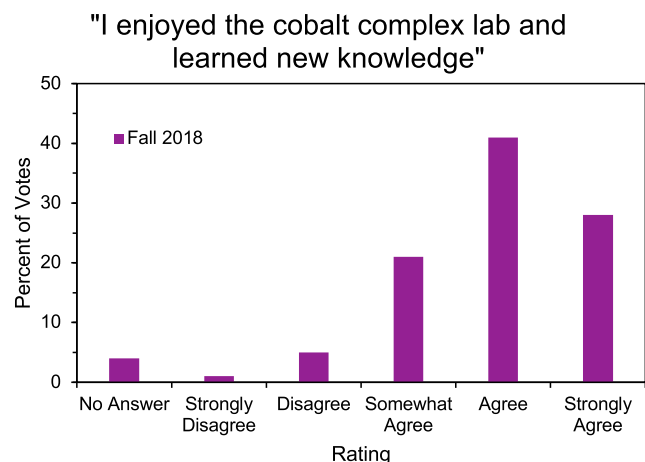


Figure 7. Quantitative student survey results regarding this laboratory experiment from 287 out of 378 general chemistry students.

behind the inconsistent values of K_{eq} . The visual highlights of this experiment successfully supplement and strengthen concepts taught in Brown University's general chemistry course, ranging from molecular geometries to Le Châtelier's principle, which are topics that can be difficult to conceptualize in a real-life context.

CONCLUSION

Through a collaborative effort between undergraduates, graduates, and faculty, we have successfully developed and implemented a new and engaging general chemistry laboratory experiment exploring the equilibrium of cobalt complexation using UV–vis spectroscopy. Our use of alcohols in place of HCl and selection of the pink-to-blue equilibrium direction improved both the safety and visual appeal of the experiment for a large number of general chemistry students. Additionally, these changes bring added challenges that simulate a real research environment for students.

This experiment takes advantage of a much higher molar absorptivity for a blue Co complex: by shifting the chemical equilibrium from a pink octahedral to a blue tetrahedral cobalt complex through small, varying additions of 2-propanol, students clearly observe the vivid color changes as new equilibrium states are established. The equilibrium constants were measured using UV–vis spectroscopy through application of the Beer–Lambert law, and the experimental results on equilibrium constants allowed students to examine alternative mechanistic pathways. Positive student responses can be attributed to the thought-provoking nature of the experiment and the vibrancy of the color transition for visualizing complexation equilibrium. Adaptations are readily incorporated to allow for application in high school or alternative upper-level undergraduate programs.

ASSOCIATED CONTENT

Supporting Information

The Supporting Information is available at <https://pubs.acs.org/doi/10.1021/acs.jchemed.9b00264>.

Lab manual consists of background information and instructions for the laboratory experiment and is intended as a student tool for concept reinforcement and technical guidance (PDF, DOCX)

Prelaboratory questions and answers, containing a series of relevant questions that help assess student understanding prior to the start of the experiment (PDF, DOCX)

Laboratory report for students, containing a postlaboratory assessment document for students to navigate and complete, prompting deeper inquiry of factors behind determining and driving K_{eq} (PDF, DOCX)

Instructor's notes, containing commentary on the complexity of this reaction direction, followed by possible student discussions over brainstorming different mechanisms for consistent K_{eq} (PDF, DOCX)

AUTHOR INFORMATION

Corresponding Author

*E-mail: li-qiong_wang@brown.edu.

ORCID

Leonard W. Sprague: 0000-0003-1926-8827

Li-Qiong Wang: 0000-0002-9501-6739

Notes

The authors declare no competing financial interest.

ACKNOWLEDGMENTS

The authors thank Brown University's Karen T. Romer Undergraduate Teaching and Research Award for allowing outstanding undergraduates to work on this project.

REFERENCES

- DeGrand, M. J.; Abrams, M. L.; Jenkins, J. L.; Welch, L. E. Gibbs Energy Changes during Cobalt Complexation: A Thermodynamics Experiment for the General Chemistry Laboratory. *J. Chem. Educ.* **2011**, *88* (5), 634–636.
- Nyasulu, F.; Nothing, D.; Barlag, R.; Wise, L.; Arthasery, P. Thermodynamics of the $[\text{CoCl}(\text{iPrOH})_3]^+$ to $[\text{CoCl}(\text{iPrOH})_2(\text{MeOH})_3]^+$ Reaction: A General Chemistry Laboratory Exercise. *J. Chem. Educ.* **2012**, *89* (4), 536–539.
- Ojeda, C. B.; Rojas, F. S. Process Analytical Chemistry: Applications of Ultraviolet/Visible Spectrometry in Environmental Analysis: An Overview. *Appl. Spectrosc. Rev.* **2009**, *44*, 245–265.
- Van Driel, J. H.; Verloop, N.; de Vos, W. Introducing Dynamic Equilibrium as an Explanatory Model. *J. Chem. Educ.* **1999**, *76* (4), 559.
- Grant, A. W. Cobalt complexes and Le Châtelier. *J. Chem. Educ.* **1984**, *61* (5), 466.

- (6) Shakhshiri, B. *Chemical Demonstrations*, 1st ed.; University of Wisconsin Press: Madison, WI, 1983; pp 280–285.
- (7) Cobalt Chloride Demonstration [Online]; Department of Chemistry and Biochemistry, California State University Bakersfield, Bakersfield, CA; https://www.csub.edu/Chemistry/_files/Cobalt%20DemoAO.pdf (Accessed November 18, 2019).
- (8) Tyson, L.; Treagust, D. F.; Bucat, R. B. The Complexity of Teaching and Learning Chemical Equilibrium. *J. Chem. Educ.* **1999**, *76* (4), 554.
- (9) Brown, J. L.; Battino, R. The CoCl_2 Thermosiphon. *J. Chem. Educ.* **1993**, *70* (2), 153.
- (10) Curtis, N. F.; Hay, R. W.; House, D. A.; Searle, G. B. Characterization of the unusual product from the reaction between cobalt (II) chloride, ethane-1,2-diamine, and hydrochloric acid: An undergraduate project involving an unknown metal complex. *J. Chem. Educ.* **1986**, *63* (10), 899.
- (11) Olson, G. L. Synthesis and reactions of cobalt complexes: A laboratory experiment. *J. Chem. Educ.* **1969**, *46* (8), 508.
- (12) Pratt, J. M.; Birk, J. P.; Tierney, D. L.; Yezierski, E. J. Combining Novel Visualizations and Synthesis To Explore Structure–Property Relationships Using Cobalt Complexes. *J. Chem. Educ.* **2017**, *94* (12), 1952–1955.
- (13) Klara, K.; Hou, N.; Lawman, A.; Wang, L.-Q. Developing and Implementing a Collaborative Teaching Innovation in Introductory Chemistry. *J. Chem. Educ.* **2013**, *90*, 401–404.