**Chemical Equilibrium Lab: The Equilibrium of Copper(II) Chloride**

**Introduction**

Chemical equilibrium is one of the most important aspects of chemistry. Understanding how to manipulate the rates of reactions at equilibrium to force a reaction to yield certain products can be a very useful tool for chemists in the field. We can apply Le Chatelier’s principle in order to force a reaction one way or another. We have not learned about Le Chatelier’s principle yet. However, this lab will serve as a wonderful introduction into exploring and understanding this amazing phenomenon.

**Materials**

|  |  |  |
| --- | --- | --- |
| .1M CuCl4-2 Solution  | Distilled Water | 50 ml Graduated Cylinder |
| 12M HCl | 125ml Flask | 10 ml Graduated Cylinder |

**Useful Information**

The reaction we will be investigating is given below

$$CuCl\_{4}^{-2}\left(green\right)(aq)+4H\_{2}O(l) \leftrightarrow Cu\left(H\_{2}O\right)\_{4}^{2+}(blue)(aq)+4 Cl^{-}(aq)$$

We will be using this information to keep track of the concentrations of our reactants and products as the experiment proceeds.

**Procedure**

1. First acquire a 10 ml graduated cylinder, a 50 ml graduated cylinder, and a 125 ml flask.
2. Add 20ml of the .1M CuCl4-2 Solution to the 125 ml flask.
3. Acquire some distilled water, add distilled water to the flask using your graduated cylinders until the solution turns light blue. Keep track of the volume of water added in your data table.
4. Once the solution is light blue, add the 12 M HCl to your flask using your graduated cylinder until the solution turns dark green.
5. Continue steps 3 and 4 repeatedly until you reach the 125 ml mark on your flask.
6. Once you have reached the 125 mL mark on your flask record the final color.
7. Rinse your flask out with water and pour it down the drain.
8. Follow the solution with a plethora of water.
9. Clean all the glassware and put it back where you found it.
10. Answer the following questions in your lab notebook using your final observation and your data table.

|  |  |  |
| --- | --- | --- |
| **Step Number** | **mL of H2O** | **mL of 12 M HCl** |
| **1** |  |  |
| **2** |  |  |
| **3** |  |  |
| **4** |  |  |
| **5** |  |  |
| **6** |  |  |
| **7** |  |  |
| **8** |  |  |
| **9** |  |  |

Initial volume: 20 mL

**Final Color:**

1. How many moles of CuCl4 did you initially start with?
2. How many moles of HCl was added to your total solution?
3. How many moles of H2O was added to your total solution?
4. Based on the final color of your solution was this reaction product favored or reactant favored?
5. Since we know the total mols of CuCl4, H20, and HCl, we theoretically know the total number of mols of copper atoms (Cu), Chloride atoms (Cl), and H2O atoms. Depending on the color of your final solution, which will vary from group to group, we can predict the general value of our K. Using the total number of moles for all the species of atoms and the final volume, predict the actual value of K for your reaction.

*Hint:* *you will want to split the moles of your atoms in such a way that follows your prediction on whether the equilibrium concentrations are product favored or reaction favored.*

1. Why do you think adding water to the reaction turned in blue? Why did adding HCl turn it back to green?
2. Does adding concentrations during the reaction change its equilibrium position for this reaction? If so why is that the case?