Preparing and Diluting Solutions Lab

Introduction

The amount of solute that is dissolved in a given quantity of solvent is called the concentration of the solution. A dilute solution contains only a small amount of solute in a given amount of solution, while a concentrated solution contains a large amount of solute in a given amount of solution. The molarity, M, of a solution is defined as the number of moles of solute in one liter of solution and is the primary units of concentration used in the chemistry laboratory.

Combined with molar mass of a solute, molarity is used to calculate the number of grams of solute needed to prepare a given volume of a solution with a specific concentration. For example, consider the preparation of 500.0 ml of a 0.80 M solution of sodium chloride.

Step 1: Determine the number of moles necessary to prepare the solution.

$$\text{moles of solute} = \text{Molarity} \times \text{volume of solution}$$

$$\text{mol}_{\text{NaCl}} = \frac{0.80 \text{ mol NaCl}}{1 \text{ L}} \times 500.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.40 \text{ mol NaCl}$$

Step 2: Convert the number of moles to grams using the molar mass of solute.

$$\text{grams of solute} = \text{moles of solute} \times \text{molar mass of solute}$$

$$\text{grams}_{\text{NaCl}} = 0.40 \text{ mol NaCl} \times \frac{58.44 \text{ g}}{1 \text{ mol}} = 23 \text{ g NaCl}$$

The calculations show that 23 g of sodium chloride are required to prepare 500.0 ml of 0.80 M NaCl solution.

Once the calculations have been done to determine how much solute is needed to prepare a solution, precise analytical techniques must be followed to ensure accuracy in making the solution. One very important aspect of analytical technique involves choosing the right type of glassware. Volumetric glassware is glassware that has been calibrated and marked to hold a specific volume. The most common form of volumetric glassware used for preparing solutions is the volumetric flask, which has a long, narrow neck with a single, hairline marking on it. For a 100-mL volumetric flask, the mark on the neck indicates that when filled to the mark, the flask will contain precisely 100.0 mL at room temperature.

Experiments often require a solution that is more dilute than what is on hand in the stockroom. In this case, a more concentrated stock solution must be diluted to obtain the desired concentration. To carry out a dilution, the following equation is used:

$$M_1V_1 = M_2V_2$$

where M_1 equals the molarity of the more concentrated stock solution, V_1 equals the volume of that stock solution that will be taken to prepare the dilution, M_2 equals the desired molarity of the final diluted solution, and V_2 equals the final volume of the diluted solution. For example, assume that the 0.80 M sodium chloride solution prepared in the example above is in the stockroom, but for another experiment, 100.0 mL of a 0.20 M sodium chloride solution is needed. In performing a dilution calculation, M_1, M_2, and V_2 are generally known and the equation is rearranged to solve for the unknown V_1. Substituting the known values for this example into the equation determines the volume of the concentrated solution required to prepare the dilute solution.

$$V_1 = \frac{M_2V_2}{M_1} = \frac{0.20 \text{ M} \times 100.0 \text{ mL}}{0.80 \text{ M}} = 25 \text{ mL}$$

In order to prepare the 0.20 M solution, one would measure out 25 mL of the 0.80 M stock solution and bring the final volume to 100.0 mL with deionized water. Proper analytical technique for preparing the diluted solution requires that the initial and final volumes (V_1 and V_2) must be accurately measured using a graduated cylinder or, preferably, a pipet and volumetric flask.

Molarity and dilution calculations are used to know how to prepare solutions of known concentration. Another important problem chemists encounter in the lab is how to determine the concentration of an unknown solution. If the solution is colored, the concentration of an unknown solution can be determined by measuring the intensity of the color. A spectrophotometer is used to measure the absorbance of visible light that gives the solution its color. Generally, the more intense the color of the solution, the greater the absorbance of light will be. In using colorimetry, it is important to remember that the color of light transmitted by the solution (the color observed) is complementary to the color of light absorbed by teh solution (the color measured). Since the color of light depends on its wavelength, the wavelength of light absorbed by a colored substance in solution is complementary to the

wavelength of light transmitted by the substance. Copper (II) sulfate solutions, for example are blue. The absorbance of copper (II) sulfate solutions is measured at 635 nm, corresponding to red light.

The purpose of this experiment is to prepare a series of blue copper (II) sulfate solutions of known concentration using the molarity and dilutions equations. The relationship between the concentration of the solutions and their absorbance will be investigated. The accuracy of the solution preparation and dilution procedures will then be determined by analyzing an unknown solution of copper (II) sulfate.

Pre-Lab Calculations

1. Calculate the number of grams of copper (II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$, required to prepare 100.0 mL of a 0.350 M $CuSO_4$ solution.
2. Calculate the number of milliliters of 0.350 M $CuSO_4$ solution that must be diluted to prepare 10.0 mL of a 0.0550 M $CuSO_4$ solution.
3. If you take 3.50 ml of the 0.0550 M $CuSO_4$ solution and dilute it to a final volume of 10.0 mL, what will be your final concentration?

In order to save valuable laboratory time, you should complete the necessary laboratory calculations. These calculations should be written in the Calculations section of the Results part of the lab.

Materials

balance, spatula, 100-mL volumetric flask, 250-mL beaker, 250-mL plastic bottle, funnel, test tube rack, five 16 x 125 mm test tubes, 10-mL graduated cylinder, spectrophotometer with cuvet, Kim wipes, weigh boats, $CuSO_4 \cdot 5H_2O$, wash bottle filled with deionized water, disposable Beral-type pipets, labeling tape and markers

Procedure

The spectrophotometer must warm up for at least 15 minutes. At the start of the lab, you should turn the spectrophotometer on and adjust the wavelength to 635 nm.

Part 1. Preparing the Stock Solution

1. Prepare 100.0 mL of a 0.150 M $CuSO_4$ solution. To do this: first review your pre-lab calculations and then calculate the amount of $CuSO_4 \cdot 5H_2O$ you will have to measure out. Clearly show your work in the Calculations section of your Results for this lab and indicate that that is the amount of $CuSO_4 \cdot 5H_2O$ you will be using to make 100.0 mL of a 0.150 M $CuSO_4$ solution. Check the value with your instructor before proceeding.
2. Once your calculation has been approved, weigh out the required amount of $CuSO_4 \cdot 5H_2O$ on a balance in a clean, dry weighing dish. Transfer the solid to a clean beaker. Use a wash bottle filled with deionized water to rinse any remaining solid from the weighing dish into the beaker.
3. Dissolve the solid in the beaker in a minimum amount of deionized water, then transfer the solution to a 100-mL volumetric flask using a funnel. Rinse the beaker with deionized water using a wash bottle. Pour the rinse water through the funnel and into the volumetric flask so that every bit of solution is transferred to the volumetric flask.
4. Slowly add more deionized water to the volumetric flask until the liquid level is almost to the 100-mL mark. Fill to the mark with a pipe or wash bottle drop-by-drop so that the bottom of the meniscus is exactly at the 100-mL mark. Cap the volumetric flask and invert it 10-15 times to give a homogeneous solution.
5. Once the solution is thoroughly mixed, transfer it to a clean, labeled bottle. Your label should indicate the concentration of your solution with the proper number of significant figures, the date the solution was made, and the initials of the solution's maker. Clean and put away the volumetric flask.

Part 2. Preparing Diluted Solutions

6. Place five clean and dry test tubes in a test tube rack and label them #1-#5. Label one pipet "stock" and use it to transfer the stock solution only.
7. Using a 10-mL graduated cylinder, measure and pour 10-mL of the 0.150 M stock solution into test tube #1. Record the necessary data for this solution in the Data Table.
8. Using a clean pipet, fill the 10-mL graduated cylinder exactly to the 3.80-mL mark with the stock solution. Try not to get any drops of solution on the sides of the cylinder. Make sure that the bottom of the meniscus sits exactly at the 3.80-mL mark. Carefully fill the graduated cylinder to the 10.0-mL mark with deionized water. Do not overfill!
9. Mix the solution in the graduated cylinder by repeatedly filling and emptying a clean pipet with the solution three times. The agitation caused by filling and emptying the pipet will mix the solution. Transfer the mixed solution to test tube #2. Determine the concentration of the solution in test tube #2 by performing the necessary calculations. Show this work in the calculations section of your Results and record the necessary data for the solution in the Data Table.

10. Rinse the graduated cylinder several times with deionized water before proceeding to the next step.
11. Repeat steps 8-10 using 2.40 ml of stock solution and transfer the mixed solution to test tube #3. Perform the necessary calculations and record the data for this solution in the Data Table.
12. Determine the amount of 0.150 M stock solution that will be needed to prepare 10.0-mL of 0.075 M and 0.023 M $CuSO_4$ solutions, respectively. Write these calculations in the Calculations section and check the values with your instructor before proceeding.
13. Using your calculations and analytical technique described in steps 8-10, prepare 10.0-mL of a 0.075 M solution by diluting the stock solution. Transfer this solution to test tube #4 and record the necessary data in the Data Table.
14. Using your calculations and analytical technique described in steps 8-10, prepare 10.0-mL of a 0.023 M solution by diluting the stock solution. Transfer this solution to test tube #5 and record the necessary data in the Data Table.
15. Compare the color of the stock solution and each of the dilutions in tests tubes #1-5. Rank them in terms of color from deepest blue to lightest blue. Record these observations in the Data Table.

Part 3. Colorimetry Measurements

16. Rinse a cuvet several times with deionized water, handling the cuvet by its frosted side to avoid getting fingerprints on the surface of the glass. Use lint-free wipes to clean off any residue from the glass.
17. Fill the cuvet ¾ full with deionized water and place the cuvet in the spectrophotometer compartment.
18. The spectrophotometer must have warmed up for at least 15 minutes. Make sure the spectrophotometer is set to measure absorbance at 635 nm and set the instrument so it is reading zero for the blank. Record the value for the water blank in your results section. Then dump the water from blank into a waste container.
19. You will now analyze your samples in <u>increasing</u> concentration or darkness. Take the lightest sample and pour ½ to 1-mL of the sample into the cuvet. Then dump the contents into the waste beaker. This removes any residual water from the cuvet so it does not dilute your sample. Now pour the rest of the contents of your test tube into the cuvet. Make sure there are not fingerprint smudges on the cuvet and place it in the spectrophotometer. Record the absorbance value in your Data Table. Return the solution back to its original test tube in case you need to have to reanalyze the sample.
20. Repeat step 19 with other samples in order of increasing darkness in order to prevent contamination.
21. Rinse the cuvet several times with deionized water. Obtain an unknown sample of $CuSO_4$ and similarly analyze the sample in the spectrophotometer. Record the absorbance of the unknown sample in your results section.
22. Rinse the cuvet several times with deionized water and place it upside down atop a clean wipe in a test tube rack to dry. If you are the last group using the spectrophotometer, you should shut it off and put it away.
22. Return to your station and make a graph of your data with absorbance on the x-axis and concentration on the y-axis. The graph should be big enough that you can justly plot the data and determine the best-fit line for the points. In addition, enter the data into a graphing calculator and determine the equation of the line and the R^2 value for the linear regression. Record both the equation of the line and its R^2 value next to the graph. If your R^2 value is greater than 0.95, you may dump your solutions and clean the equipment. If the R^2 value is less than 0.95, you must retest and/or remake solutions until you have a good linear regression for the standards.
23. Using the equation for the line, determine the concentration of your unknown. Obtain the true value from your instructor and determine percent error for the lab.

<u>Discussion/Conclusion</u>

Write a meaningful discussion and conclusion for the lab. Discuss the making of the solutions, dilutions, and use of the spectrophotometer. Reflect on your recorded observations and/or difficulties during the lab. What is the relationship between concentration and color intensity? Between concentration and absorbance? Does it make sense that the relationship between concentration and absorbance should include the origin (0,0) as a point? Explain your reasoning. Discuss how you plotted the data and determined the concentration of the unknown solution. What techniques did you learn and what difficulties did you encounter? What would you do differently?

<u>Post-Lab Questions</u> (Read the Appendix pages A16-A19 of your textbook to help you.)

1. How is it that a spectrophotometer can send a particular wavelength of light through a sample?
2. Write the Beer-Lambert Law and define the variables.
3. Most cuvets are 1.0 cm in width. Why do you think is?
4. Beer's Law states that absorbance is proportional to the concentration of the absorbing species and works very well for dilute solutions. Beer's law fails when solutions become to concentrated. As solutions become too concentrated, how is this reflected in a graph of absorbance vs. concentration?