

CHE 111 Lab Calculations Guide

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This guide will serve as a reference to assist you with the calculations needed to complete the laboratory assignments in CHE 111. This guide should be used as a supplement to the laboratory handouts.

Contents:

Using a Line of Linear Regression to Determine Concentration.....	2
(Density, Chemistry of Color)	
Determining the Empirical Formula of a Hydrate.....	4
(Empirical Formula, Limiting Reagent)	
Calorimetry Calculations.....	5
(Calorimetry)	
Preparing Solutions and Making Dilutions.....	8
(Chemistry of Color, Titration of Fruit Juice)	
Titration Calculations.....	9
(Titration of Fruit Juice)	

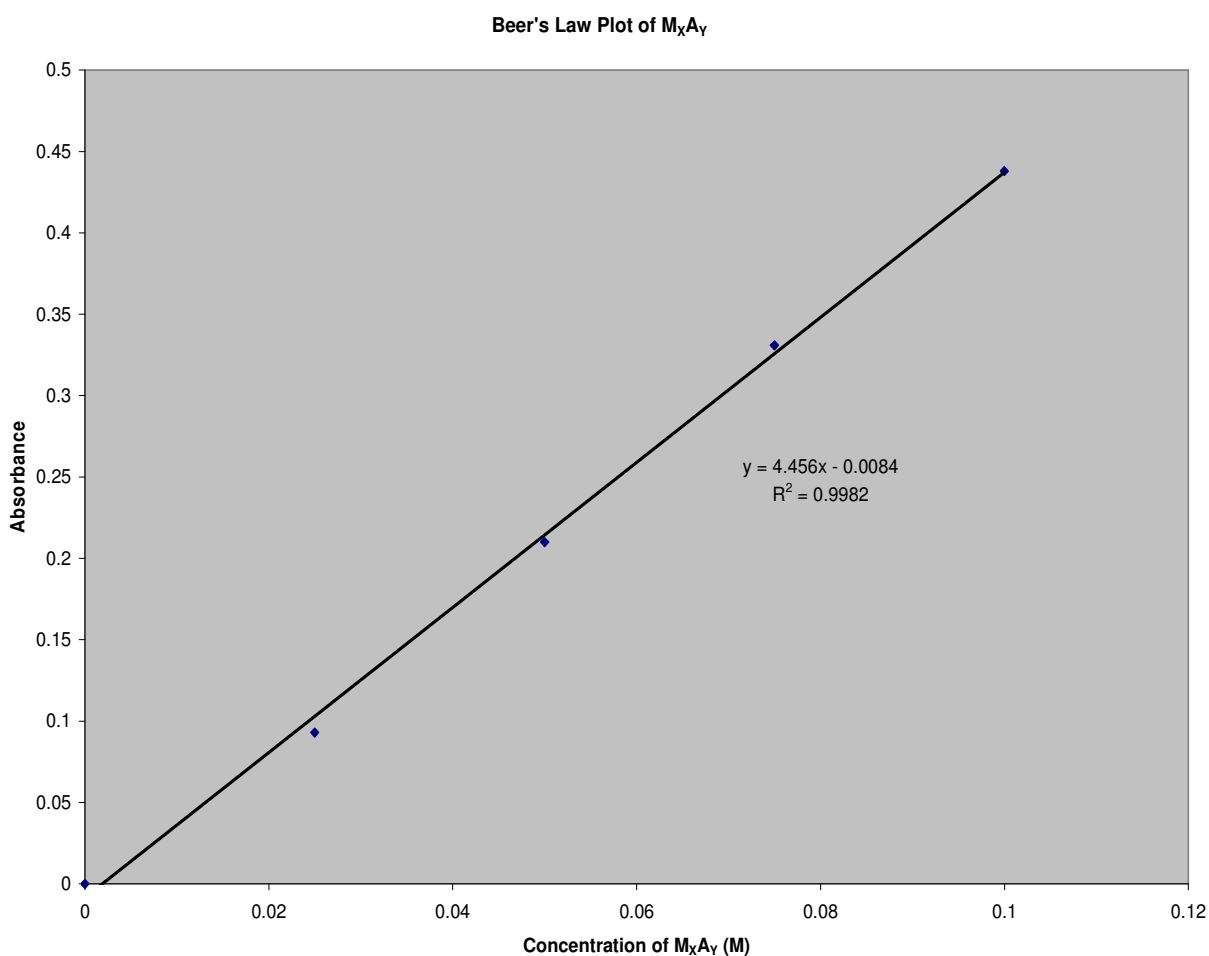
The labs listed under each subject utilize all or part of the calculations illustrated.

Using a Line of Linear Regression to Determine Concentration.

Sample Data

Concentration	Absorbance
0 M	0
0.025 M	.093
0.050 M	.210
0.075 M	.331
0.10 M	.438

Graphing the data results in the following plot:



The absorbance of a solution with an unknown concentration was determined to be 0.0615. The concentration can be determined by substituting the absorbance into Y, and solving for X (concentration).

$$Y = 4.456x - 0.0084$$
$$0.0615 = 4.456x - 0.0084$$

Add 0.0084 to each side, to separate the terms with x from the terms without x.

$$0.0615 + 0.0084 = 4.456x$$

$$0.0699 = 4.456x$$

Divide both sides by 4.456 to determine the value of x (concentration).

$$\frac{0.0699}{4.456} = \frac{4.456x}{4.456}$$

$$x = 0.0157 \text{ M}$$

Determining the Empirical Formula of a Hydrate

Hydrates are ionic compounds that have a certain number of water molecules permanently associated with them. They have a general formula of $M_xA_y \cdot nH_2O$, where n is a whole number.

In order to determine the empirical formula of a hydrate you need to determine the moles of both the dehydrated ionic compound and the water.

Sample Data

Chemical:	$MgCO_3 \cdot nH_2O$
Mass of hydrate:	1.365 g
Mass of dehydrated compound:	0.660 g

The mass of water can be determined by subtracting the mass of the dehydrated compound from the mass of the hydrate.

$$\text{Mass of water} = 1.365 \text{ g} - 0.660 \text{ g} = 0.705 \text{ g}$$

To determine moles, divide mass by molar mass. Molar mass of a compound is the sum of the molar masses of the elements that comprise the compound. The molar masses of the elements can be found on the periodic table.

Molar mass of $MgCO_3$

$$\begin{aligned} \text{Mg: } & 1(24.31) = 24.31 \text{ g} \\ \text{C: } & 1(12.01) = 12.01 \text{ g} \\ \text{+ O: } & 3(16.00) = 48.00 \text{ g} \\ \hline \text{MgCO}_3 & = 84.32 \text{ g/mol} \end{aligned}$$

Molar mass of H_2O

$$\begin{aligned} \text{H: } & 2(1.01) = 2.02 \text{ g} \\ \text{+ O: } & 1(16.00) = 16.00 \text{ g} \\ \hline \text{H}_2\text{O} & = 18.02 \text{ g/mol} \end{aligned}$$

Moles of $MgCO_3$

$$0.660 \text{ g MgCO}_3 \left(\frac{1 \text{ mol}}{84.32 \text{ g}} \right) = 0.00783 \text{ mol MgCO}_3$$

Moles of water

$$0.705 \text{ g H}_2\text{O} \left(\frac{1 \text{ mol}}{18.02 \text{ g}} \right) = 0.0391 \text{ mol H}_2\text{O}$$

The value of n in the empirical formula is found by determining the ratio of moles of water to moles of dehydrate.

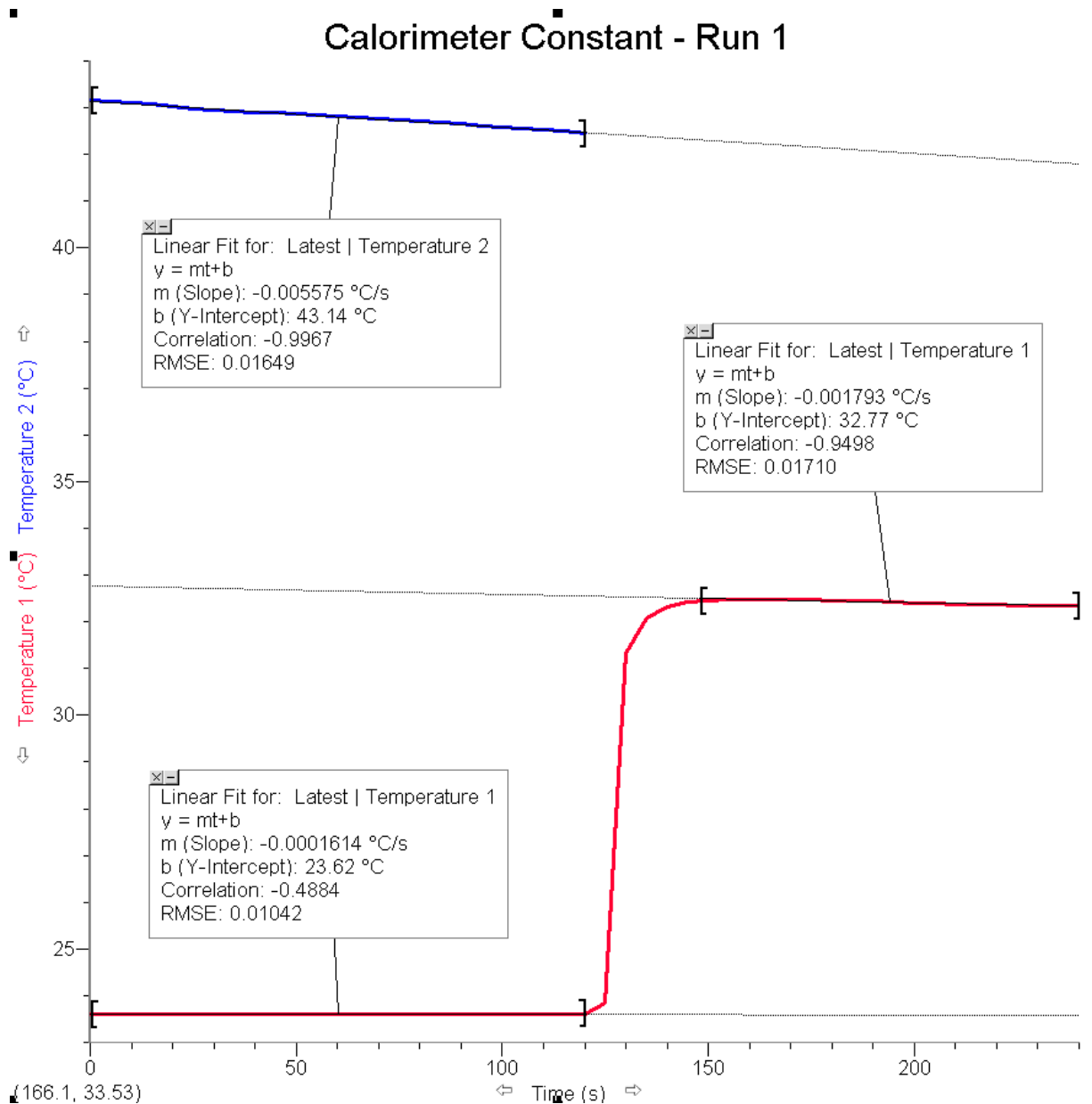
$$n = \frac{\text{moles water}}{\text{moles dehydrate}} = \frac{0.0391 \text{ mol}}{0.00783 \text{ mol}} = 5.00$$

The empirical formula can now be written as $MgCO_3 \cdot 5H_2O$

Calorimetry Calculations

The following data and graph were collected in order to determine the calorimeter constant.

Mass of calorimeter	3.164 g
Mass of cold water + calorimeter	51.934 g
Mass of mixed water + calorimeter	97.651 g



To determine the temperature at mixing ($t = 125$ s), solve each line for Temperature (Y) at this time (X). These temperatures are then used to determine the temperature change of the hot water and the cold water.

Finding the temperature of the cold water at 125 s.

$$Y = -0.0001614(^{\circ}\text{C}/\text{s})x + 23.62 ^{\circ}\text{C}$$

$$Y = -0.0001614(^{\circ}\text{C}/\text{s})(125 \text{ s}) + 23.62 ^{\circ}\text{C} = 23.60 ^{\circ}\text{C}$$

Doing the same for the hot water and the mixed water gives the following temperatures:

$$\text{Hot: } 42.44 ^{\circ}\text{C}$$

$$\text{Mixed: } 32.55 ^{\circ}\text{C}$$

Temperature change is found by subtracting the initial temperature (hot or cold water) from the final temperature (mixed water), at the time of mixing.

$$\Delta T_{\text{cold}} = T_{\text{mixed}} - T_{\text{cold}} = 32.55 ^{\circ}\text{C} - 23.60 ^{\circ}\text{C} = 8.95 ^{\circ}\text{C}$$

$$\Delta T_{\text{hot}} = T_{\text{mixed}} - T_{\text{hot}} = 32.55 ^{\circ}\text{C} - 42.44 ^{\circ}\text{C} = -9.89 ^{\circ}\text{C}$$

Since the cold water is in the calorimeter, the temperature change of the calorimeter is the same as the temperature change of the cold water.

$$\Delta T_{\text{cal}} = \Delta T_{\text{cold}} = 8.95 ^{\circ}\text{C}$$

The mass of the cold water is found by subtracting the mass of the calorimeter from the mass of the calorimeter and cold water.

$$m_{\text{cold}} = 51.934 \text{ g} - 3.164 \text{ g} = 48.770 \text{ g}$$

The mass of the hot water is determined by subtracting the mass of the cold water and calorimeter from the mass of the mixed water and calorimeter.

$$m_{\text{hot}} = 97.561 \text{ g} - 51.934 \text{ g} = 45.627 \text{ g}$$

Knowing that the heat lost by the hot water is equal to the sum of the heat gained by the cold water and the calorimeter, we can write the following equation.

$$q_{\text{hot}} = -(q_{\text{cold}} + q_{\text{cal}})$$

This equation can be rewritten as:

$$m_{\text{hot}}\Delta T_{\text{hot}}C_{\text{hot}} = -m_{\text{cold}}\Delta T_{\text{cold}}C_{\text{cold}} - C_{\text{cal}}\Delta T_{\text{cal}}$$

Rearranging this equation to determine the calorimeter constant (C_{cal}) gives

$$C_{\text{cal}} = \frac{m_{\text{hot}}\Delta T_{\text{hot}}C_{\text{hot}} + m_{\text{cold}}\Delta T_{\text{cold}}C_{\text{cold}}}{-\Delta T_{\text{cal}}}$$

Substituting in data gives:

$$C_{\text{cal}} = \frac{(45.627 \text{ g})(-9.89 \text{ }^\circ\text{C})(4.184 \text{ J/g }^\circ\text{C}) + (48.770 \text{ g})(8.95 \text{ }^\circ\text{C})(4.184 \text{ J/g }^\circ\text{C})}{-8.95 \text{ }^\circ\text{C}}$$

$$C_{\text{cal}} = 6.93 \text{ J/ }^\circ\text{C}$$

The same general procedure is used to determine the specific heat of the unknown metal (C_{metal}).

Preparing Solutions and Making Dilutions

Accurately preparing solutions and making dilutions are essential skills that all chemists need to master. Before you make a solution in lab, you first need to determine the mass of solute that you need to measure.

Example : Determine the mass of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$ needed to prepare a 250 mL solution with a concentration of 0.400 M

First determine the moles of solute needed.

$$0.250 \text{ L} \left(\frac{0.400 \text{ mol}}{1 \text{ L}} \right) = 0.100 \text{ mol}$$

Now the mass of copper (II) chloride dihydrate* can be calculated.

$$0.100 \text{ mol} \left(\frac{170.49 \text{ g}}{1 \text{ mol}} \right) = 17.0 \text{ g}$$

* 170.49 g/mol is the molar mass of $\text{CuCl}_2 \cdot 2\text{H}_2\text{O}$

To make the solution you need to measure out 17.0 g (recording the exact mass) of copper (II) chloride dihydrate. Add the solid to a 250 mL volumetric flask which is half filled with DI water. Swirl until all of the solid has dissolved, then fill the flask to the mark with DI water. ***You need to calculate the exact concentration of the solution you made, start with the measured mass and work backwards through the calculations.***

Example: Determine the volume of 0.750 M stock NaCl solution needed to make a 25.00 mL solution with a concentration of 0.200 M.

Use the Dilution formula:

$$C_1 V_1 = C_2 V_2$$

Where C_1 and V_1 represent the stock concentration and volume respectively. The diluted concentration and volume are represented by C_2 and V_2 respectively.

Rearranging the dilution formula to solve for volume of stock.

$$V_1 = \frac{C_2 V_2}{C_1} = \frac{(0.200 \text{ M})(25.00 \text{ mL})}{0.750 \text{ M}} = 6.67 \text{ mL}$$

Thus 6.67 mL of stock solution needs to be pipetted into a 25.00 mL volumetric flask, which is then filled to the mark with DI water.

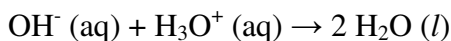
Titration Calculations

Standardizing NaOH by titrating it into 25.00 mL of 0.200 M HCl

Sample Data:

Volume of HCl titrated	25.00 mL
Concentration of HCl	0.200 M
Final burette level (NaOH)	23.21 mL
Initial burette level (NaOH)	1.27 mL

The reaction that occurs between sodium hydroxide (OH^-) and hydrochloric acid (H_3O^+) occurs in a one to one ratio.



In order to determine the concentration of the sodium hydroxide, you first need to determine the moles of acid that reacted.

$$0.02500 \text{ L} \left(\frac{0.200 \text{ mol}}{1 \text{ L}} \right) = 0.00500 \text{ mol H}_3\text{O}^+$$

Convert to moles of sodium hydroxide using the molar ratio from the balanced reaction.

$$0.00500 \text{ mol H}_3\text{O}^+ \left(\frac{1 \text{ mol OH}^-}{1 \text{ mol H}_3\text{O}^+} \right) = 0.00500 \text{ mol OH}^-$$

The concentration of the hydroxide ions can be determined by dividing moles of hydroxide by volume of hydroxide. The volume of hydroxide ions is equal to the volume of sodium hydroxide added, this is found by subtracting the initial burette level from the final burette level. The concentration of NaOH is equal to the concentration of OH^- since hydroxide forms in a one-to-one ratio from sodium hydroxide.

$$\text{Volume of NaOH} = 23.21 \text{ mL} - 1.27 \text{ mL} = 21.94 \text{ mL}$$

$$[\text{NaOH}] = [\text{OH}^-] = \frac{0.00500 \text{ mol OH}^-}{0.02194 \text{ L}} = 0.228 \text{ M OH}^-$$

Once the sodium hydroxide has been calculated it may be used to determine the concentration of an acid with an unknown concentration. First calculate the moles of hydroxide added. Next convert to moles of acid titrated. Finally determine the concentration of the acid titrated.