Laboratory Manual for General Chemistry I

(CHEM 110L)

New York Institute of Technology

Dr. John C. Ringen

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Table of Contents

Note: Experiments #8a (Chemical Nomenclature) and #8b (Specific Heat) are both done together in week #8.

Credits

Figures 1.1, 4.1, and 12.1 were taken from William E. Bull, William T. Smith, Jr., and Jesse H. Wood, ***Laboratory Manual for College Chemistry, Sixth Edition***, San Francisco, Harper & Row, 1980. Copyright © 1980 by William E. Bull, William T. Smith, Jr., and Jesse H. Wood. Used by permission.

Figures 4.4, 7.1, 10.1, and 11.1 were taken from John H. Nelson and Kenneth C. Kemp, ***Laboratory Experiments for Brown and LeMay / Chemistry: The Central Science, 2nd Edition***, Englewood Cliffs, N.J., Prentice-Hall, Inc., 1981. Copyright © 1981 by Prentice-Hall, Inc. Used by permission.

In addition, I would like to thank the following:

1. Dr. Delmar S. Larsen of the University of California at Davis for permission to use one illustration on his web page for Figure 3.1.

2. Dr. Robert Schneider of Stony Brook University for permission to use one illustration on his web page for Figure 4.2.

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Experiment 1. Introduction to Laboratory Work.

I. Overview

In Week #1 the following items will be covered:

1. CHECK-IN

You will be working with another student who will be your lab partner for the entire semester. The check-in procedure consists of the following steps:

a. Print both of your names on the sheet provided next to the number of an empty lab drawer. This will be your drawer for the entire semester.

b. Pick up a check-in sheet (one sheet per group). Write your *names* and *ID numbers* in the spaces provided on the check-in sheet.

c. Locate the key to your lab drawer (on the key ring) and open the drawer.

d. Your lab instructor will go through the check-in sheet to identify each item which is listed. Place a check mark on the sheet if the item is in your kit. ***Note which items (if any) are missing or broken; obtain these items from the lab technician as soon as possible.***

e. Hand in the check-in sheet to your lab instructor.

2. LABORATORY SAFETY

a. A DVD dealing with chemistry laboratory safety will be shown.

b. The location and use of items such as the hoods, fire extinguishers and blankets, eye wash stations, and safety showers will be explained.

c. Two additional sheets (a lab safety sheet and an OSHA sheet) will also be distributed. Please sign each sheet, and then hand in: (a) the bottom portion of the lab safety sheet and (b) the complete OSHA sheet to your lab instructor.

3. BALANCE INSTRUCTIONS

Balances are used in virtually all of the lab experiments. Chemists often use the term "weight" to refer to the *mass* of the object (in grams). Physicists are more careful to point out the distinction between these two terms. If we say "*weigh* the crucible," we really should say "obtain the *mass* of the crucible."

Your lab instructor will demonstrate how to use the electronic balances that are located in the chemistry laboratory. These balances make it possible for you to record all masses of objects to the nearest one-thousandth of a gram (0.001 g).

Whenever these balances are used, ***you will need to copy all of the numbers in the balance reading*** (and *never* round any of the balance readings) when recording the data in your lab manual. Instructions for using these balances appear at the top of the following page.

Taring the Balance. With the air shield in place, press the "TARE" button; the balance should read either 0.000 g or -0.000 g within a few seconds. If not, repeat this procedure as required.

Weighing an Object. Place the object on the pan, using the air shield if the object is not too tall or too wide, and wait several seconds for the reading to stabilize. Then record the mass of the object. ***Use the air shield whenever possible to make sure that your readings are accurate.***

Note: to weigh a chemical, put waxed paper (or a small beaker) on the balance, and then zero the balance with the paper or beaker on. The balance reading will then automatically give you the mass of the chemical that you have added.

4. USE OF BUNSEN BURNERS

A diagram of your bunsen burner appears in Fig. 1.1 at the right. The amount of gas entering the burner may be adjusted either by using the gas valve or the gas flow adjustment near the bottom of the burner. Always use the gas lighters provided, *not* matches, to light the burners.

The barrel can be moved up or down to control the air/gas ratio. If the barrel is too low, the flame is yellow and it is a cool flame; not enough air is present and unburned carbon soot is produced. When the barrel is moved up, the flame becomes hotter and the yellow flame goes away; when the barrel is near the top, you will clearly see the inner and outer cones. ***The hottest portion of the flame is located just above the inner cone.***

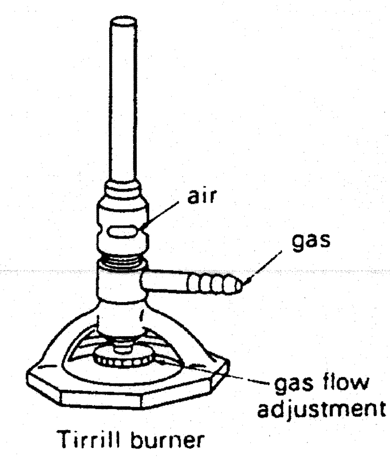


Fig. 1.1

Your laboratory instructor may ask you to practice using this bunsen burner during the first laboratory period.

5. SIGNIFICANT FIGURES, MATH, METRIC SYSTEM, AND CONVERSIONS

Please read the following topics in the textbook (Whitten et. al., *Chemistry: The Core*, 10th edition, Brooks/Cole Cengage Learning, 2014) as soon as possible:

a. ***Scientific Notation:*** Appendix A-1 (pages A-1 to A-3)

b. ***Measurements in Chemistry and Units of Measurement:*** Sections 1-7 and 1-8 (pages 19 to 21)

c. ***Significant Figures:*** Appendix A-4 (pages A-5 to A-8). ***You must study pages A-7 and A-8 of the textbook as well as the next two pages of this lab manual to learn the distinction between the two significant figure rules involving: (a) addition and subtraction and (b) multiplication and division.***

It is very important that you record all data and perform all calculations in these lab experiments using the proper number of significant figures. After reading the above pages, be sure to complete the Significant Figure Practice Exercise on page 5. Your lab instructor may ask you to hand in this exercise to be graded, go over the answers next week, and/or give you a lab quiz dealing with this material.

Please ask your laboratory instructor for assistance if required.

BASIC SIGNIFICANT FIGURE RULES

Exact numbers have an *infinite* number of significant figures (SF); they *never* limit the accuracy of a calculated result. There are 12 inches in one foot, 60 minutes in one hour, 1,000 milliliters in one liter, and 2.54 centimeters in one inch. There may be exactly 19 students performing an experiment in the CHEM 110 laboratory at one point in time. All of these numbers are *exact* numbers, and they have an *infinite* number of significant figures.

All of the rules that follow apply only to *measurements* or to quantities that you have *calculated* from such measured quantities. Such numbers always have a *finite* number of significant figures.

1. ***Non-zero digits*** are *always* significant. For example, **91** has *two* significant figures (9 and 1), while **123.45** has *five* significant figures (1, 2, 3, 4, and 5).

2. ***Zeroes appearing anywhere between two non-zero digits*** are *always* significant. For example, **102.034** has *six* significant figures (1, 0, 2, 0, 3, and 4).

3. ***Trailing zeroes*** (at the right of a number) are *always* significant because they indicate how accurately the measurement was taken. Examples:

   a. **34.2** g has *three* significant figures. The number indicates that the measurement was taken to the nearest *tenth* of a gram, so that the possible error is ±0.1 g.

   b. **34.20** g has *four* significant figures. The number indicates that the measurement was taken to the nearest *hundredth* of a gram, so that the possible error is ±0.01 g.

   c. **34.200** g has *five* significant figures. The number indicates that the measurement was taken to the nearest *thousandth* of a gram, so that the possible error is ±0.001 g.

   The balances that you will use always give readings to the nearest thousandth of a gram. There will be three numbers after the decimal point in all of your balance readings. Always copy all of these numbers into your lab manual, and never round these balance readings!

   Similarly, a balance reading such as **34.000 g** has *five* significant figures (3, 4, 0, 0, and 0). Because the balance is capable of reading masses to the nearest *thousandth* of a gram, you need to include *all* of those zeroes to inform the reader that the possible error in the number was ±0.001 g.

4. ***Leading zeroes*** (at the left of a number) are *never* significant, because ***these zeroes always drop out when you write that number in scientific (exponential) notation.*** For example, **0.004050** has only *four* significant figures (4, 0, 5, and 0). ***When you write that number in scientific (exponential) notation, drop the leading zeroes and write only the numbers that are significant.*** The result is $\mathbf{4.050 \times 10^{-3}}$.

5. ***In a number greater than 1***, the presence of ***trailing zeroes*** (at the right) without a decimal point creates confusion, because we do not know how accurately the measurement has been taken. For example, **400 g** could mean (400 ± 100 g) [1 SF], which is equivalent to $\mathbf{4 \times 10^2}$ **g**. It could also mean (400 ± 10 g) [2 SF], which is equivalent to $\mathbf{4.0 \times 10^2}$ **g**. Finally, it could also mean (400 ± 1 g) [3 SF], which is equivalent to $\mathbf{4.00 \times 10^2}$ **g**. The use of scientific notation completely eliminates the confusion that had existed as to how accurately the original measurement had been taken.

   Important Note: if a decimal point is used at the end, then the presence of the decimal point means that *all* of the zeroes *are* significant. For example, **400. g** means (400 ± 1 g), or $\mathbf{4.00 \times 10^2}$ **g**.

SIGNIFICANT FIGURE RULES IN SIMPLE MATHEMATICAL OPERATIONS

1. **Multiplication and Division.** For multiplication and division, the result should have as many significant figures as the *measured* number with the *smallest* number of significant figures.

   Remember that numbers that are *exact* (such as those appearing in the conversion 1 L = 1000 mL) have an *infinite* number of significant figures, and they do *not* limit the accuracy of an answer.

   a. If you multiply **(0.00530) × (12.468)** on a calculator, the display reads **0.0660804**. The first number (0.00530) has *three* significant figures, and the second number (12.468) has *five* significant figures. The display reading (0.0**660**804) must therefore be rounded to *three* significant figures. ***You must round all answers up or down appropriately.*** In this case, round the answer *up* to **0.0661**, not *down* to 0.0660. The answer, written in scientific notation, becomes **6.61×10^{-2}**. Of course, if the unrounded number would have been 0.0**660**304, then you would have simply dropped the "304" at the end, and written down **0.0660** as the final answer.

   b. If you divide **45.678 milliliters** (5 SF) by 1000 to obtain **0.045678 liters** (on a calculator), remember that the number of milliliters in a liter (1000) is an *exact* number that has an *infinite* number of significant figures. The number of significant figures in the answer is therefore equal to the number of significant figures in the original number, or *five*. The answer may be left either as **0.045678 liters** or written as **4.5678×10^{-2} L**.

2. **Addition and Subtraction.** For addition and subtraction, the result should have as many decimal places as the *measured* number with the *smallest* number of decimal places. Examples:

   a. **12.345 g + 0.12 g** = 12.465 g (on the calculator). The first number has *three* decimal places, and the second number has only *two* decimal places. The answer must therefore be rounded to *two* decimal places. Because the number we want to drop is "5," we must round the preceding number *up*, to obtain **12.47 g**. ***Never just drop a "5"*** (which would have given you the incorrect answer, 12.46 g).

      Note that the answer has *four* significant figures even though the least number of significant figures in the numbers being added together was only *two* (in the number 0.12 g). ***Never apply the significant figure rules that deal with multiplication and division when you are doing addition and subtraction!***

   b. **1.456 g – 1.326 g** = 0.13 g (on the calculator). Each mass reading has *three* decimal places and the possible error is ±0.001 g, so the answer must have *three* decimal places also. You need to add the zero that is missing on the calculator's display, and write the answer as **0.130 g**, not as 0.13 g!

      Note that the answer has only *three* significant figures, even though each of the original numbers had *four* significant figures. As was mentioned above, never apply the significant figure rules that deal with multiplication and division when you are doing addition and subtraction!

   c. **76°C – 74°C = 2 °C** Each temperature reading was known to the nearest *whole degree* (±1°C), and there are *no* digits to the right of the decimal point (had a decimal point been used). As a result, the answer can be known only to the nearest whole degree. ***Here, the answer will have only one significant figure, even though both of the original numbers had two significant figures.*** Do *not* write the answer as 2.0°C, because this would be incorrect.

CHEM 110 - GENERAL CHEMISTRY I - SIGNIFICANT FIGURE PRACTICE EXERCISE

Experiment 1. Introduction to Laboratory Work

Name______________________________ Lab Partner's Name______________________________

Section__________ Date______________________ Instructor______________________________

Important: Even though you will be working with a lab partner when doing experiments throughout the semester, each person must submit his or her own report each week.

Before completing this page, be sure to read the following sections in the textbook (Whitten et. al., *Chemistry*, 10th edition, Thomson Brooks/Cole, 2014): ***Scientific Notation*** (Appendix A-1, pages A-1 to A-3) and ***Significant Figures*** (Appendix A-4, pages A-5 to A-8).

Then answer the questions shown below.

1. For each number shown below, first write down how many significant figures the number has. Then write that number in scientific (exponential) notation in "standard form" (with the number to the left of the exponent being between 1 and 10).

   Example: 4.56×10^{-5} *is* in standard form, but neither 0.456×10^{-4} nor 45.6×10^{-7} are in standard form (since the number to the left of the exponent is either less than 1 or greater than 10).

Number	How many significant figures are in the number?	Write the number in scientific (exponential) notation in standard form
0.07		
0.070		
0.0700		

2. Perform each calculation. First express the answer using a regular number (such as 34.50), and then write the answer using scientific (exponential) notation in standard form (such as 3.450×10^{1}). ***In all cases, be sure to express all answers using the correct number of significant figures.***

Calculation to be Performed	Answer using a regular number*	Answer written in exponential notation in standard form*
Addition: $7.65 + 6.75$		
Subtraction: $7.65 - 6.75$		
Multiplication: 3.45×8.89		
Division: $7.65 \div 6.78$		
Multiplication: $(0.00450) \times (2.0000)$		

* - rounded properly, using the correct number of significant figures.

Pre-Laboratory Exercise - Preparation of Graphs

Please read the instructions which follow, and show the practice graph to your lab instructor at the very beginning of the laboratory period the day you are scheduled to perform this experiment.

A graph shows how one variable (called the "dependent variable") varies when you change the second variable (called the "independent variable"). If you find out how the solubility of an unknown varies with the temperature, the solubility (which you calculate) is called the dependent variable, and it is plotted on the ***y-axis*** (sometimes called the "*ordinate*") of a graph. The temperature is then called the independent variable and it is plotted on the ***x-axis*** (sometimes called the "*abscissa*") of the same graph.

Graphs are often prepared automatically using computer spreadsheet programs such as Microsoft Excel. We will *not* be using Microsoft Excel (or any similar computer spreadsheet program) in this course. This means that you need to learn how to prepare graphs *manually* on graph paper, and how to interpret those graphs. The practice graph described herein will help you acquire the needed graph-making skills.

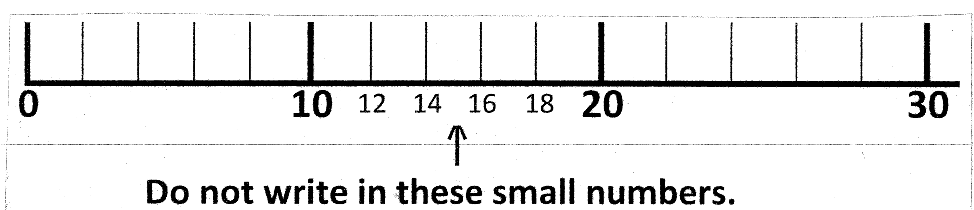
When graphs are plotted on graph paper, the quantity being plotted, as well as the units, must be clearly shown on each axis. ***Major lines on each axis must represent only whole numbers (such as 0, 10, 20, 30, etc.).*** The numerical values that correspond to the very small lines on each axis need to be determined as follows:

1. Subtract two numbers on adjacent major lines on a given axis to find out what the numerical difference is between those two major lines.

2. Count how many small lines are between each of these two major lines. Add the number "1" to this value to obtain the number of *divisions* between each of these two major lines.

3. Divide the number obtained in step 1. above by the number of divisions obtained in step 2. above to determine the numerical value that each small line represents.

Please see the illustration below. One major line represents 10˚C, and the next major line represents 20˚C. There are *four smaller lines* between those two major lines, causing there to be *five divisions* between these two major lines. Each small line therefore represents (10/5)˚C or 2˚C. This means that the six lines between 10°C and 20°C represent temperatures of 10˚C, 12˚C, 14˚C, 16˚C, 18˚C, and 20˚C:



Experimental values are then placed on the graph using dots. It is recommended that you draw a small circle around each dot so that the dots will still be clearly visible after the graph is drawn in. Sometimes the final graph should be a ***straight line***, as will be the case with the practice graph you will prepare in the lab. In such a case, use only a *ruler* to draw in the best straight line that goes through those data points. ***Never use a zig-zag line to connect each pair of dots separately!!*** It is normal for some data points to fall slightly higher and/or lower than the straight line that you have drawn in.

Alternatively, if the final graph should be a ***curve***, as will be the case with the actual solubility curve that you will plot for your unknown, use a *pencil* to draw in the curve by hand. Correct what you have done using the pencil eraser (if necessary) so that the final result will be a smooth curve. Again, individual data points may fall slightly higher and/or lower than this curve, due to experimental error.

Graphing Exercise

Using the graph paper on the next page, make a plot of solubility vs. temperature using the sample data shown below, carefully following these instructions. Then show the graph to your laboratory instructor to obtain his or her approval before you begin performing the actual experiment.

1. Whenever we say "vs." (versus), it always means that the quantity shown *before* the "vs." is plotted on the *y-axis*, and the quantity shown *after* the "vs." is plotted on the *x-axis*.

2. Label each axis to show the quantity being plotted, and always include the *units* being used. The graph is incorrect unless *both* items (quantity plotted *and* the units) are clearly shown on each axis.

3. Start numbering the axes at the *bottom left corner* of the graph. Some instructors want you to leave blank squares on both axes in the left and bottom portions of the graph, so that the axes are indented in both directions. *Do not do this.* Just begin numbering the major lines of each axis at the bottom left hand corner of the graph paper where the squares actually begin.

4. Use ***only whole numbers*** such as (0, 5, 10, 15, etc.), (0, 10, 20, 30, etc.), or (0, 20, 40, 60, etc.) to represent the major lines on each axis. If the numbers are all less than 1, you might select numbers such as 0.000, 0.010, 0.020, 0.030, 0.040, and 0.050 to represent the major lines on that axis. ***Never write a number such as 34.3 or 0.0147 on any major line of the graph!*** Select the numbers that represent these major lines so that the graph will occupy most of the space on the graph paper.

   Important note: there is absolutely ***no*** requirement that you must start numbering the two axes at zero! As is noted above, your goal is to have the graph occupy most of the space on the graph paper. To do this, feel free to begin numbering the x-axis at 10°C or some other number to make the graph spread out as large as possible. Similarly, on the y-axis, you may want to begin the numbering at 20 g unknown/100 g H_2O or something else, depending on how your calculations have turned out.

5. Place dots on the graph to represent each data value (x–y pair). It is recommended that you draw a small circle around each dot so that the dots will still be visible after the line or curve is drawn in.

6. If the graph should be a straight line, draw in that line using a ruler. If the graph should be a curve, draw the curve (freehand) to show the relationship between the two variables that you have plotted.

SAMPLE DATA FOR PRACTICE GRAPH

Temperature (°C)	Solubility (g unknown/100 g H_2O)
23	45
34	69
49	87
65	125
87	163

Practice Graph

Plot the sample data on page 8 here, following the instructions on pages 7 and 8. Then show the completed graph to your laboratory instructor before you perform the actual experiment.

Solubility Experiment

I. Purpose:

1. To learn how to prepare graphs (on graph paper) showing your experimental data.

2. To study how the solubility of a substance varies with temperature.

3. To use solubility data obtained at some temperatures to estimate the solubility of the substance at other temperatures.

4. To become familiar with the use of pipets to transfer liquids accurately.

II. Theory:

A *solution* results when a *solute* (often a solid) is dissolved in a *solvent* (usually a liquid) to produce a uniform (homogeneous) mixture called a *solution*. For example, if you dissolved salt in water, salt is the *solute*, water is the *solvent*, and the final mixture (salt water) is the *solution*.

One of the physical properties of a substance is the *solubility* of the substance—i.e. the maximum number of grams of the substance which can be dissolved in a fixed amount (usually 100 grams) of the solvent. For example, if 3.00 g of your unknown will dissolve in 25.0 g of water at a given temperature, then the solubility of the unknown at that temperature would be equal to ***12.0 g of the unknown per 100 g of water***. The solubility of a substance will always depend on the temperature of the solution. The unknown solid studied in this experiment becomes more soluble in water as the temperature increases.

In this experiment, you will determine the solubility of your unknown solid at a number of temperatures. A graph will then be plotted to show how the solubility of the unknown varies with the temperature.

III. Chemicals Required:

Solubility unknown.

IV. Special Equipment Needed:

Rubber suction bulb, thermometer, large test tube (8"), and either a transfer pipet (5 mL) or a graduated pipet (5 mL or 10 mL).

VI. Procedure

1. Weigh accurately between 7.0 g and 7.5 g of your unknown using the electronic balance and its tare feature. To do so, place an empty ***weighing boat*** (*not* a piece of flat waxed paper) on the pan of the balance, cover it with the air shield, and tare (zero) the balance with the weighing boat on. Then remove the air shield and add between 7.0 g and 7.5 g of the unknown to the weighing boat using a clean spatula. Put the air shield back on and ***record the exact balance reading to the nearest 0.001 g, such as 7.231 g or 7.462 g.*** (line 1) ***Whenever the balance is tared in this manner, the balance reading will always equal the mass of the solid that was added.*** Carefully transfer the weighed unknown from the weighing boat to a large (8") test tube (*not* a test tube from your lab drawer).

2. Pipet 5.00 mL of deionized water into the test tube (instructor will demonstrate). Clamp the test tube to a ring stand. Set up a hot water bath using a 400-mL beaker, and heat the test tube in this bath until all of the solid has dissolved. (Carefully stir the mixture using your thermometer.)

3. "***Dry Run.***" To become familiar with the change in the appearance when crystallization begins, first

perform a "dry run" in which no attempt is made to record the exact crystallization temperature. After all of the unknown has dissolved, shut off the gas, take the test tube out of the water bath, and allow the tube to cool, carefully stirring the solution using the thermometer as the solution cools. When crystallization begins, the solution has become saturated with the solute, and white particles of solid will begin to appear throughout the solution. As the temperature decreases, the number of such particles will significantly increase, the clear solution will turn into a milky-white suspension, and particles of excess solute will begin to sink to the bottom of the test tube. Do not confuse small dark specks which may be present at higher temperatures (which might be dirt or other impurities) with the actual white crystals which are present when crystallization begins.

4. ***Determination of the Exact Crystallization Temperature.*** Having observed what happens when crystallization takes place, place the test tube back into the hot water bath and stir with the thermometer until all of the excess solid goes back into solution. Then remove the test tube out of the water bath and allow the test tube to cool, carefully stirring the solution with the thermometer as it cools. This time, record the exact temperature at which crystallization just begins to occur.

5. ***Second Data Point.*** Pipet an additional 5.00 mL portion of deionized water into the same test tube, so that the total volume of water is now 10.00 mL. Now place the test tube again in the hot water bath and repeat the above procedure.

6. ***Additional Data Points:*** Repeat with three more additions of 5.00 mL of deionized water, making the total volume of water 15.00 mL, 20.00 mL, and 25.00 mL. At this point, you may find it necessary to cool the test tube using a beaker of cold water or an ice-water mixture in order to get the excess solute to crystallize. Be sure to stir the mixture continuously as it cools so that the temperature of the solution is uniform throughout.

7. ***Waste Disposal:*** Put the test tube back into the hot water bath, heat until all solid has dissolved, and then pour the solution into the special waste container provided in the hood.

VII. Calculations and Graph:

1. Calculate the solubility of your compound at each temperature where crystallization occurred. A sample calculation to illustrate the method follows.

   Suppose that you started with 7.231 g of unknown and that when only 5.00 mL of water had been added, crystallization just began at 65°C. At that temperature, 7.231 g of unknown just dissolves in 5.00 g of water. (Since the density of water is 1.00 g/mL, its mass in grams equals its volume in milliliters.) To calculate how many grams of the unknown will dissolve in 100 mL (i.e. 100 g) of water, the easiest way is to set up a proportion:

   $$\frac{7.231 \text{ g unknown}}{x \text{ g unknown}} = \frac{5.00 \text{ g water}}{100 \text{ g water}}$$

   To solve for x, cancel out the units and cross-multiply to obtain: (5.00)(x) = (7.231)(100). Then: x = (7.231)(100)/(5.00) = 144.62, which rounds to 145. Thus the solubility at 65°C is ***145 g unknown per 100 g of water*** or ***145 g unknown/100 g H_2O***. The same method is used to determine the solubility of the unknown at the other temperatures when the total mass of water is 10.0 g, 15.0 g, 20.0 g, and 25.0 g.

2. Using the graph paper on page 15, make a plot of *solubility* (in grams unknown/100 g water) on the y-axis vs. *temperature* (in °C) on the x-axis. ***Follow the instructions on pages 8 and 9.*** *Important note:* your graph will be a *smooth curve*, and not a straight line. Just draw in the best smooth curve that you can through your data points.

3. From this curve, estimate the solubility of your unknown at 30°C and 50°C.

CHEM 110 - GENERAL CHEMISTRY I - LABORATORY REPORT

Experiment 2. Solubility

Name____________________ Lab Partner's Name____________________

Section________ Date______________ Instructor____________________

BALANCE READING

1. Mass of unknown . . . . ________________

TEMPERATURE AND SOLUBILITY DATA

Volume of water used	(x-axis) Crystallization temperature	(y-axis) Solubility (g unknown/100 g H_2O)
5.00 mL		
10.00 mL		
15.00 mL		
20.00 mL		
25.00 mL		

GRAPH

Use the sheet on the next page to prepare a graph (see instructions, pages 7 to 10), properly drawn with axes labeled and major divisions numbered. Attach the graph to this sheet.

Estimated solubilities (from graph):

1. At 30 °C ________________ 2. At 50 °C ________________