**Rowan College at Burlington County**

**CHE 116**

**Lab Assignments and Supplemental Information Packet**

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**Student Drawers Laboratory Equipment List**

**Mt. Laurel campus, Science building, room 100**

**Student Drawer 1**

Beaker, 50 mL (1)

Beaker, 100 mL (1)

Beaker, 150 mL or 100 mL (3)

Beaker, 250 mL (3)

Beaker, 400 mL (1)

Beaker, 600 mL (1)

Erlenmeyer flask, 125 mL, (3)

Erlenmeyer flask, 250 mL, (3)

Funnel (1)

Graduated cylinder, 10 mL (1)

Graduated cylinder, 100 mL (1)

Spatula (1)

Stir rod (1)

Test tube holder (1)

Test tube rack (1)

Watch glass (1)

**Student Drawer 2**

Bunsen burner with hose (1)

Clay triangle (1)

Crucible tongs (1)

Ring Clamp (1)

Striker (1)

Wire screen (1)

**Student Drawer 3 (left drawer)**

Hot plate / stirrer

Timer

**Student Center Cabinet**

Burette stand with clamp

**Student Drawers Laboratory Equipment List**

**Pemberton campus, Parker building, room 144**

**Student Drawer 1**

Beaker, 50 mL (1)

Beaker, 100 mL (1)

Beaker, 150 mL (4)

Beaker, 250 mL (1)

Beaker, 400 mL (1)

Beaker, 600 mL (1)

Erlenmeyer flask, 125 mL, (3)

Erlenmeyer flask, 250 mL, (3)

Graduated cylinder, 10 mL (1)

Graduated cylinder, 100 mL (1)

**Student Drawer 2**

Clay triangle (1)

Bunsen burner with hose (1)

Striker (1)

Wire screen (1)

**Student Drawer 3**

Crucible tongs (1)

Eye dropper (1)

Funnel (1)

Spatula (1)

Stir rod (1)

Test tube holder (1)

Test tube rack (1)

Watch glass (1)

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**CHE116 Rowan College at Burlington County**

**Week 1**

Assignment: MythBusters Episode Described Using the Scientific Method

MythBusters: Season 4 Episode title: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

The steps to the Scientific Method are listed below. Describe the parts of the MythBusters episode that corresponds to each step of the Scientific Method. Please use complete sentences for this assignment.

Question asked:

Background work (information)/Literature search:

Hypothesis (rewrite the ‘Question asked’ as a statement):

Test hypothesis with experiment (adjust experiment as needed):

Analyze results to make conclusion:

Adjust hypothesis and try again (if necessary):

Report final results (conclusion):

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**CHE116 Rowan College at Burlington County**

**Week 2**

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Lab Partner: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Report Sheet for Experiment 1 Unknown**

*This page should accompany the report sheets from your lab manual.*

Experiment 1additional work: Density and Identity of an Unknown Liquid

Temperature of the unknown = \_\_\_\_\_\_\_\_\_\_\_\_\_

Unknown sample #: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

|  |  |  |  |
| --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 |
| Mass flask + sample |  |  |  |
| Mass empty flask |  |  |  |
| Mass of sample |  |  |  |
| Density of sample |  |  |  |

Show the calculation of density for trial 1, in the space below:

Average density of the sample: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show the calculation of the average density (also called the mean density) below:

Average deviation from the mean: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show the calculation of the average deviation from the mean below:

Identity of the unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_; explain your answer below:

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**CHE116 Rowan College at Burlington County**

**Week 3**

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Lab Partner: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Experiment 2 Solubility, Density, and Melting Points of Solids**

Room temperature = \_\_\_\_\_\_\_\_\_\_\_\_\_

Unknown # = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Table 1: Solubility of Naphthalene and of the Unknown

|  |  |  |  |
| --- | --- | --- | --- |
|  | Solvent: Water | Solvent: Cyclohexane | Solvent: Ethyl Alcohol |
| Naphthalene |  |  |  |
| Unknown |  |  |  |

*(Use the terms soluble, sparingly soluble, insoluble to describe the solubility.)*

Table 2: Density of Naphthalene and of the Unknown

|  |  |  |
| --- | --- | --- |
|  | Naphthalene | Unknown |
| Type of liquid used for volume by difference |  |  |
| Mass of solid |  |  |
| Initial volume (only liquid) |  |  |
| Final volume (liquid + solid) |  |  |
| Volume of solid |  |  |
| Density of solid |  |  |

Show the calculation for the density of naphthalene below:

Table 3: Melting Point of the Unknown

|  |  |
| --- | --- |
|  | Unknown |
| Approximate melting point |  |

**Identity of the unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (Briefly explain below.)**

**Questions:** As assigned by your lab instructor.

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**CHE116 Rowan College at Burlington County**

**Week 4**

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Identification of Several Unknown Substances by Physical Properties**

At least one solid and at least one liquid unknown should be used; determine the identity of three unknown substances.

**Room Temperature: \_\_\_\_\_\_\_\_\_\_**

**Unknown Number: \_\_\_\_\_\_\_\_\_\_** *(solid or liquid)*

**Solubility of this Unknown**

|  |  |
| --- | --- |
| **Solvent Used** | **Result (soluble, insoluble, sparingly soluble)** |
| Solvent used = |  |
| Solvent used = |  |
| Solvent used = |  |

**Density of this Unknown**

|  |  |  |
| --- | --- | --- |
|  | Trial 1 | Trial 2 |
| Mass of vessel and unknown |  |  |
| Mass of empty vessel |  |  |
| Mass of unknown |  |  |
| *Volume of only liquid in grad. Cylinder (for solid sample)* |  |  |
| *Volume of liquid and solid in cylinder (for solid sample)*  *Use a liquid that does not dissolve your solid sample.* |  |  |
| Volume of unknown |  |  |
| Calculated density of unknown |  |  |

*Use the same balance for both trials.*

Show the calculations for the densities here:

Mean density of this unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show calculation here:

**Melting Point of this Unknown**  \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Identity of this Unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

*(Use list provided on page 9.)*

**Explain your choice for the identity of your unknown:**

**Unknown Number: \_\_\_\_\_\_\_\_\_\_\_**

**Physical state of this unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Solubility of this Unknown**

|  |  |
| --- | --- |
| **Solvent Used** | **Result (soluble, insoluble, sparingly soluble)** |
| Solvent used: |  |
| Solvent used: |  |
| Solvent used: |  |

**Density of this Unknown**

|  |  |  |
| --- | --- | --- |
|  | Trial 1 | Trial 2 |
| Mass of vessel and unknown |  |  |
| Mass of empty vessel |  |  |
| Mass of unknown |  |  |
| *Volume of only liquid in grad. Cylinder (if applicable)* |  |  |
| *Volume of liquid and solid in cylinder (for solid sample)*  *Use a liquid that does not dissolve your solid sample.* |  |  |
| Volume of unknown |  |  |
| Calculated density of unknown |  |  |

*Use the same balance for both trials.*

Show the calculations for the densities here:

Mean density of this unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show calculation here:

**Melting Point of this Unknown**  \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Identity of this Unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

*(Use list provided on page 9.)*

**Explain your choice for the identity of your unknown:**

**Unknown Number: \_\_\_\_\_\_\_\_\_\_\_**

**Physical state of this unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Solubility of this Unknown**

|  |  |
| --- | --- |
| **Solvent Used** | **Result (soluble, insoluble, sparingly soluble)** |
| Solvent used: |  |
| Solvent used: |  |
| Solvent used: |  |

**Density of this Unknown**

|  |  |  |
| --- | --- | --- |
|  | Trial 1 | Trial 2 |
| Mass of vessel and unknown |  |  |
| Mass of empty vessel |  |  |
| Mass of unknown |  |  |
| *Volume of only liquid in grad. Cylinder (if applicable)* |  |  |
| *Volume of liquid and solid in cylinder (for solid sample)*  *Use a liquid that does not dissolve your solid sample.* |  |  |
| Volume of unknown |  |  |
| Calculated density of unknown |  |  |

*Use the same balance for both trials.*

Show the calculations for the densities here:

Mean density of this unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show calculation here:

**Melting Point of this Unknown**  \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Identity of this Unknown: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

*(Use list provided on page 9.)*

**Explain your choice for the identity of your unknown:**

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**CHE116 Rowan College at Burlington County**

**Week 8**

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Lab Partners: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Experiment 8 (CHE116 supplement packet)** Gravimetric Analysis Experiment

# In the field of Analytic Chemistry, there are two major distinctions, Qualitative and Quantitative analysis. As the names imply, qualitative analysis involves determining what a substance is, and quantitative analysis involves determining how much of it there is or how pure it is.

# There are also specialties within these categories. Three specialties of quantitative analysis are gravimetric analysis, volumetric analysis, and instrumental analysis.

# This experiment involves the use of gravimetric analysis to determine the quantitative amount of an ion present in a sample. In addition, this determination may be used to find the molecular mass of the substance in question, which will identify the substance from among several possibilities.

# The purpose of this experiment is to determine the identity of a group I metal carbonate by gravimetric analysis. The unknown is weighed and dissolved in water. A solution of calcium carbonate is added to the metal carbonate solution to precipitate the carbonate ions as calcium carbonate. The precipitate is filtered, dried and weighed. From the data, the formula weight and identity of the unknown metal carbonate are determined.

# We know that the alkali metals in group I (Li, Na, K, etc.) form a carbonate of the molecular structure *M*2CO3. These compounds are soluble in water. If we add another water-soluble substance, CaCl2, to a solution of an alkali metal carbonate, a precipitateof CaCO3 is formed. By determining the amount of precipitate formed, we can determine the original amount of alkali carbonate with which we started. The overall reaction for this process is as follows, where “*M*” represents the alkali metal ion:

***M***2CO3(aq) + CaCl2(aq) 🡪 CaCO3(s) + 2***M***Cl(aq)

The stoichiometry of this reaction can tell us the number of moles of carbonate originally present, and using the original sample weight, the molecular mass of the alkali metal is calculated. Since we can determine the molecular mass of the alkali metal possibilities, we can identify which of the alkali metal carbonates we have.

## Pre-Lab Questions

1. Why would we choose calcium chloride for the reactant to precipitate the soluble carbonate?
2. How do we make use of stoichiometry to determine molecular mass of the original substance?
3. What is the mole ratio of the calcium carbonate formed to the original alkali metal carbonate?
4. What are the molecular formulas and molar masses of the possible alkali metal carbonates? (Li2CO3, Na2CO3, K2CO3)
5. An unknown metal carbonate was analyzed gravimetrically and yielded the following data. Calculate the values for lines d – i.
   1. Mass of dried ***M***2CO3 1.972g
   2. Mass of filter paper 0.598g
   3. Mass of filter paper + CaCO3 2.436g
   4. Mass of CaCO3 \_\_\_\_\_\_\_\_\_\_\_\_\_ (c – b)
   5. Moles of CaCO3  \_\_\_\_\_\_\_\_\_\_\_\_\_
   6. Moles of ***M***2CO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_ (= e)
   7. Molar mass of ***M***2CO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_ (a / f)
   8. Identity of ***M***2CO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_
   9. Percent error \_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show your calculations here:

#### Materials

# Calcium chloride solution, 1.0 M

# DI water

# Unknown alkali metal carbonate (which may be lithium carbonate, sodium carbonate, or potassium carbonate)

#### Equipment

# Balance, 0.001 g or 0.01 g

Beakers, various sizes

Crucible tongs (to carry hot watch glass)

Drying oven

Funnel

Filter paper

Glass stirring rod

Graduated cylinder

Ring stand with ring

Spatula

Wash bottle

Watch glass

#### Experimental procedure

1. *Set up a ring stand with a clean, dry crucible held by a clay triangle on a ring clamp*
2. *Put about 2 grams of the unknown group I metal carbonate into the clean dry crucible*
3. *Lightly heat the crucible with a Bunsen burner for about 2-3 minutes*
4. *Allow the crucible to cool*

*------------------------------------------------------------------------------------*

1. Weigh approximately 0.5 grams of the unknown into a 150 ml beaker. **Record the mass of the unknown to the nearest .001 gram or 0.01 gram.**
2. Add approximately 50 ml of water (deionized) to this beaker.
3. Stir with a clean glass stirring rod until the solid is dissolved.
4. Add 30 ml of 1.0 M calcium chloride solution to the beaker; stir.
5. Let the solution settle for at least 5 minutes
6. While the solutions settle, assemble a gravity filtration setup. Fold a piece of filter paper; tear a corner, and label the paper. Then record the mass of the filter paper.
7. Carefully filter the solution; pour the clear, top part of the liquid into the filter paper first, then pour in the cloudy, bottom part of the liquid. Make sure all of the solid material is transferred into the filter paper. Be careful not to over-fill the filter paper.
8. Rinse any remaining precipitate into the filter with water.
9. After the filtration is done, open the filter paper onto a clean watch glass, being careful to keep all of the solid on the filter paper.
10. Place the watch glass with filter paper into the drying oven.
11. After approximately 10 minutes of drying, remove the sample from the oven and break up the precipitate, using a clean spatula or stirring rod.
12. Return the sample to the drying oven for an additional 5 minutes.
13. Remove the sample from the drying oven, and allow the sample to cool. When cooled, find the mass of the filter paper and solid.
14. If time permits, repeat steps 16 and 17 to make sure the sample is completely dry.

### Data Table

**Unknown Number =\_\_\_\_\_\_\_**

|  |  |  |
| --- | --- | --- |
|  | Trial 1 | Trial 2  (from your lab partner) |
| a. Mass of dried ***M***2CO3 |  |  |
| b. Mass of filter paper |  |  |
| c. Mass of filter paper + CaCO3 after drying |  |  |
| d. Mass of CaCO3 |  |  |
| e. Moles of CaCO3 |  |  |
| f. Moles of ***M***2CO3 |  |  |
| g. Molar mass of ***M***2CO3 |  |  |
| h. Identity of ***M***2CO3 |  |  |
| j. Percent error |  |  |

Note; identity could be Li2CO3, Na2CO3, or K2CO3

## Calculations

1. Fill in all of the spaces in the data table above

Line d. = line c. – line b.

Line e. = line d. / molecular weight of CaCO3

Line f. = line e.

Line g. = line a. / line f.

Line h, compare to molar mass of three possible choices

Line j = ((true – exp.) / true ) x 100

1. Review all of the procedures that you used in this experiment. Give any possible sources of error that would make your calculated molar mass answer either too high or too low. Remember that moles of CaCO3 are in the denominator.

(mass of dried ***M***2CO3 / moles of CaCO3 = molar mass; units of g/mole)

* 1. Molar mass too high *(denominator too low)*:
  2. Molar mass too low *(denominator too high)*:

### Post Lab Questions

1. Although steps 1 – 4 of the procedure were not done, why would we heat the unknown carbonate at the start of the experiment?
2. For procedure steps 15 – 17, why do you think we remove the precipitate from the drying oven, break it up and heat it longer?
3. What possible effect would the use of tap water, rather than deionized water, have on your result?
4. If some of the CaCO3 solid was lost during the filtering step, how would this affect your result?
5. If some of the unknown carbonate solution was spilled before mixing with the calcium chloride solution, how would this affect your result?
6. What is the percentage, by mass, of sodium in pure table salt (NaCl)?

Show calculations.

1. How many **milligrams** of sodium are contained in 4.28 g of NaCl?

Show calculations.

1. Barium can be analyzed by precipitating it as BaSO4 and weighing the precipitate. When a 0.269 g sample of a barium compound, dissolved in an aqueous solution, was treated with an excess of sulfate ion, 0.0891 g of BaSO4 was formed. What is the percentage of barium in the original unknown compound?

Show calculations.

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**CHE116 Rowan College at Burlington County**

**Week 9**

**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

Assignment: Chemical Information Search

This is a written assignment only; you will not be preparing this solution.

An aqueous solution needs to be prepared for use in a chemistry experiment. The solution will be used at room temperature. Listed below are four compounds to consider for use as the solute in the solution. Fill in the requested information for each compound. After the requested information is complete, choose the best compound to use as the solute. Explain your decision in detail. Please complete the two definitions on the last page.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Topic** | **Toluene** | **Lead (II) Nitrate** | **Sucrose** | **Nitrogen Dioxide** |
| Chemical Formula and Formula Weight |  |  |  |  |
| CAS Number |  |  |  |  |
| Physical State of Matter at Room Temperature (20°C) |  |  |  |  |
| *Physical Properties*  m.p. or b.p. =  color = |  |  |  |  |
| Solubility in H2O, at 20°C or 25°C  (*answer yes, no, or slightly*) |  |  |  |  |
| NFPA Ratings for  **Health =**  **Fire =**  **Reactivity =**  (special if applicable) = |  |  |  |  |
| Safety Equipment Needed |  |  |  |  |
| Cost to Purchase (Reagent or technical grade)  $ / unit of measure |  |  |  |  |
| Level of Difficulty Working with this Substance  (*your opinion: use*  *very difficult, difficult, or relatively easy*) |  |  |  |  |

Which is the best compound to use as the solute? Explain your answer.

Please give definitions for the following:

NFPA rating:

CAS Number:

**CHE116 Rowan College at Burlington County**

**Week 9**

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Lab Partner: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Experiment: Solution Concentration Units**

**Purpose** The purpose of this experiment is for students to gain experience with

concentration units. At some point in time, in a scientific job, a person will have to deal with concentration units. This person may need to know how much solute is in the portion of solution they are using, how much solution to use, or how to prepare the next batch of solution.

**Procedure** Students should work in groups of two; each group of students should complete some calculations during lab class. The remaining calculations should be completed before the report due date.

**Concentration Units** The concentration units listed below are the units that will be used during this lab period.

|  |  |  |
| --- | --- | --- |
| **Unit** | **Calculation Equation** | **Typically Used For….** |
| % weight | % wt. = (mass of solute / mass of solution) x 100 | General applications; mixing a solid or liquid into a liquid solvent |
| % volume | %(v/v) = (volume of solute / volume of solution) x 100 | General applications; mixing of two liquids |
| % g/mL | % (m/v) = (mass of solute, g / volume of solution, mL) x 100 | Not very common, specific units required |
| ppm | ppm = (mass of solute / mass of solution) x (1x106) | Environmental samples |
| ppb | ppb = (mass of solute / mass of solution) x (1x109) | Environmental samples |
| molarity | M = (moles of solute / Liters of solution) | Chemistry applications; all types |
| normality | N = (number of equivalences / Liter of solution) | Chemistry applications; acid,base rxns and redox rxns |
| molality | m = (moles of solute / kg of solvent) | Chemistry applications; f.p depression and b.p. elevation calcs. |
| Dilutions | Mc • Vc = Md • Vd | Any type of concentration unit; used when a dilution of a concentrated solution is done. |

**Part A: Percentage Units**

All percentage units involve a ratio of solute divided by solution, multiplied by 100. For a weight percentage, the abbreviation is typically % wt. or % (m/m). For a volume percentage, the abbreviation is typically % (v/v). For a mass to volume percentage, the abbreviation is % (m/v).

For % wt., the units for mass must be the same in the numerator and denominator.

For % (v/v), the units for volume must be the same in the numerator and denominator.

For % (m/v), the units for the numerator must be in g, the units for the denominator must be in mL.

Please answer the following questions involving percentage units; show all of your calculations.

1. If 10.0 mL of red dye are mixed with enough deionized water to prepare 250.0 mL of solution, what is the % (v/v) concentration of the red dye solution?

2. Rubbing alcohol is 70% (v/v) isopropyl alcohol mixed with water. If you used 64.0 mL of this rubbing alcohol solution, how many milliliters of isopropyl alcohol did you use?

3. What is the % (m/m) of Mg(NO3)2 in solution when 0.84 g of Mg(NO3)2 are dissolved in enough deionized water to prepare 500.0 mL of solution (assume the solution is dilute enough that the density of the solution is 1.0 g/mL)?

4. How many grams of CuSO4 would you need to use to prepare 250.0 g of a 0.10% (m/m) solution of CuSO4?

5. If a person drinks 250.0 g of lowfat 1% (m/m) milk, how many grams of fat did the person consume?

6. A saline solution (NaCl) is 0.90% (m/v). If a hospital patient received 150.0 mL of this solution, how many grams of NaCl did the patient receive, and how many moles of NaCl did the patient receive?

Grams of NaCl:

Moles of NaCl:

**Part B: Parts Per \_\_\_\_\_ Units**

Parts per million (ppm) and parts per billion (ppb) are common units for environmental samples. The calculations are very similar to the percentage units, (which are essentially parts per hundred units). The mass units must be the same for the numerator and denominator, so these units cancel during the calculations.

Please answer the following questions involving the parts per \_\_\_\_\_\_ units; show all of your calculations.

7. A sample of river water was analyzed, and it was determined that there was 0.000012 g of lead in 1000.0 g of river water. What is the concentration of lead, reported in units of ppm and ppb?

Pb concentration in ppm:

Pb concentration in ppb:

8. If a person’s tap water has a lead concentration of 8.0 ppb, and this person drinks 1.0 L of the tap water, how much lead did they consume (in units of grams)? Assume the density of the tap water is 1.0 g/mL.

9. The state of Washington set an action limit of 20 ppm for arsenic in soils. For a sample of soil taken from King County, Washington, the arsenic level in a soil sample was determined to be 86 ppm. How many grams of arsenic are in 1000 g of this soil?

10. The amount of lead in a child’s blood is expressed in units of ug/dL. If a child’s blood is determined to have 8 ug/dL of lead, what is this amount of lead expressed in units of ppb?

The density of blood is 1.05 g/mL.

**Part C: Molarity and Normality**

Molarity is a very common unit used in chemistry; normality is a multiple of molarity. Molarity, M, is calculated as the number of moles of solute per liter of solution. When using the unit of molarity, a person should be able to calculate any one of the values used in the molarity definition when they are given the other two values. (M = moles / L)

Normality is a multiple of molarity. Normality is typically used for acid base work and oxidation reduction work. The normality of a particular substance in the solution is a multiple of the molarity of a more general substance. The normality value should specifically state the substance being reported to avoid confusion.

Please answer the following questions involving the molarity and normality units. Please show all of your calculations.

11. A solution of copper(II)chloride is prepared with 2.6 g of copper(II)chloride and enough deionized water to make 250.0 mL of solution. What is the molarity of this copper(II)chloride solution?

12. If you need to prepare 500.0 mL of a 0.10 M solution of hydrochloric acid, how much hydrochloric acid do you need to use? Express your answer in units of moles of HCl needed. Question 13 is a continuation of this question.

Moles HCl:

13. Calculate the grams of HCl concentrated solution that would be needed to prepare the solution in question 12. The commercially available, concentrated hydrochloric acid solution that would be used for this preparation is 37% (m/m) HCl in water. (Hint: convert moles HCl in question 12 to grams of HCl, then calculate the mass of commercial solution needed to get these grams of HCl.)

14. What is the maximum volume of 0.20 M zinc(II)nitrate solution that can be prepared with 12.6 g of zinc(II)nitrate?

15. A 0.80 M solution of H2SO4 is to be completely neutralized by a solution of NaOH. The balanced acid base neutralization equation is:

\_\_ H2SO4 (aq) + \_\_ NaOH (aq) → \_\_\_\_\_\_\_ + \_\_\_\_\_\_\_ (Molecular Equation format)

\_\_\_\_\_\_\_ + \_\_\_\_\_\_ → \_\_\_\_\_\_\_ (Net Ionic Equation format)

Taking into account this neutralization reaction, what is the normality of this 0.80 M solution of H2SO4, with respect to the H+ ion?

16. A piece of zinc metal is placed into a 0.60 M solution of CuCl2. An oxidation reduction reaction takes place. Complete and balance the net ionic equation for this reaction:

Zn(s) + Cu2+(aq) → \_\_\_\_\_ (aq) + \_\_\_\_\_ (s) (Net Ionic Equation format)

Taking into account this oxidation reduction reaction, what is the normality of this 0.60 M solution of CuCl2, with respect to the electrons transferred for the reaction.

**Part D: Molality Unit**

Molality, m, is a ratio of moles of solute divided by kg of solvent. This is kg of **solvent**, not solution. One advantage that the molality unit has is that it is not temperature dependent; volume is not in the denominator.

Please answer the following questions involving the molality unit; show all of your calculations.

17. A 0.20 molality K2O solution needs to be prepared for a Quality Control lab located in a manufacturing plant. How many grams of K2O will be needed when 268 g of solvent will be used to prepare the solution?

18. How much solvent is needed, in units of g, to prepare a 0.030 molality solution of K2O that will be prepared with 0.048 g of K2O?

**Part E: Dilutions**

The dilution equation is often shown with the molarity concentration unit. However, any concentration unit can be used in this dilution equation. This equation mathematically shows that the amount of solute taken from the concentrated, stock solution is equal to the amount of solute transferred into the new, diluted solution.

19. What is the CuCl2 concentration of the new, diluted solution that is prepared by diluting 10.0 mL of a 2.0 M solution of CuCl2 to a new volume of 1000.0 mL? (This is referred to as a 1/100 dilution of the 2.0 M solution; 10.0 mL / 1000.0 mL = 1/100.)

20. What is the concentration of the new, diluted HNO3 solution that is prepared by a 1/50 dilution of 3.0 M HNO3 solution?

**Atomic Emission Lines**

Observe the emission lines for hydrogen and for helium; fill out the requested information in the tables below.

Table 1 Emission Lines for Hydrogen

|  |  |
| --- | --- |
| **Color of Line** | **Wavelength, nm (as read from spectroscope scale)** |
| Red |  |
| Blue (blue/green) |  |
| Violet (blue/violet) |  |

Table 2 Emission Lines for Helium

|  |  |
| --- | --- |
| **Color of Line** | **Wavelength, nm (as read from spectroscope scale)** |
| Red |  |
|  |  |
|  |  |
| Violet |  |

**CHE116 Rowan College at Burlington County**

**Week 10**

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Lab Partner: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Experiment** Solution Preparation

**Purpose** The purpose of this experiment is for students to gain experience with solution preparation by preparing a concentrated solution, which will be used to make a series of less concentrated solutions. The concentrated, “stock”, solution will be used to prepare standards used for the construction of a calibration curve.

**Reagents** The substance to be used for this experiment is a liquid red dye. This dye will stain skin and clothing; wear goggles, gloves, and an apron. The diluting solvent is deionized water, available at all of the lab benches.

**Procedure** Students should work in groups of two.

1. Each group of students should make one concentrated stock solution as directed

2. Use this solution to prepare the diluted solutions to use for constructing the calibration curve. These diluted solutions should have a concentration of 0.010 % (v/v), 0.020 % (v/v), 0.030 % (v/v), and 0.040 % (v/v). Deionized water is the solvent, and this solvent should also be used as a blank for the spectrophotometer.

3. Determine the absorbance of each solution using the spectrophotometer (λ = 560 nm). Also measure the absorbance of the unknown solution provided; do not dilute the unknown.

4. Use the diluted solutions of known concentrations and known absorbances to construct a calibration curve.

The laboratory report should consist of the completed tables, completed questions, and an original calibration curve constructed by hand. The best-fit line should be hand-drawn. Graphing data is discussed in Appendix C in your lab manual. Each student should use their own calibration curve to determine the concentration of the unknown solution; this determination should be shown on the calibration curve.

**1. Preparation of the stock solution (concentrated solution):**  Prepare a 1.0% (v/v) stock solution by diluting 0.50 mL of the liquid red dye with deionized water; use a 50 mL volumetric flask to prepare exactly 50.0 mL of this stock solution. The bottom of the meniscus should be at the calibration mark on the neck of the volumetric flask. Cover and seal the volumetric flask with Parafilm® and invert 50 times to mix the solution.

Concentration of stock solution: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

*Show your calculation here.*

**2. Preparation of the diluted solutions (these will be the standard solutions of known concentration used to construct the calibration curve):** Calculate the volume of stock (concentrated) solution needed to prepare each of the diluted solutions listed in Table 1.Pipet the appropriate volume of stock solution into clean volumetric flasks. Dilute with deionized water; each volumetric flask has its own, unique calibration mark.

**Table 1: Volume of Stock Solution Needed** (Volume of Volumetric Flask = 100.0 mL )

|  |  |
| --- | --- |
| **Concentration of Diluted Solution, % (v/v)** | **Volume of Stock Solution Needed, mL** |
| 0.010 |  |
| 0.020 |  |
| 0.030 |  |
| 0.040 |  |

Show all calculations here:

**3. Experimental Results:** The absorbance of each standard solution and the unknown solution should be listed in the table below:

**Table 2: Absorbance Values**

|  |  |
| --- | --- |
| **Concentration, % (v/v)** | **Absorbance** |
| 0.010 |  |
| 0.020 |  |
| 0.030 |  |
| 0.040 |  |
| Unknown sample |  |

**4. Calibration Curve:** Construct and attach your original calibration curve to this report. Make sure you show, on the calibration curve, how you determined the concentration of the unknown (show the horizontal and vertical lines).

**Checklist for the construction of a calibration curve.**

1. Label each axis of the graph, include units. (Absorbance does not use units.) Put absorbance on the y axis and concentration on the x axis.

2. Determine the scale to use for each axis. The y and x scales should be quick and easy to figure out by any person looking at the graph. Each square may be worth 0.5, 1, 2, or 5. **It is not good practice to have 3 squares equal 1 or 10, because then each square is worth 0.333 or 3.33 which is difficult to visualize.** The scales should allow the data points to be spread out across most of the graph, not squished in one corner.

3. Plot each data point on the graph. Each point is made with a concentration value and the corresponding absorbance reading. Use a darkened circle for each data point.

4. Use a ruler to draw the best-fit, straight line through the data points. Not all of the data points may be on the line, some data points may be above the line and some may be below the line. DO NOT connect the dots to produce a zigzag line. DO NOT draw the line “free hand”; use a ruler.

5. Give the graph a title.

**Determining the concentration of an unknown solution:**

1. Determine the absorbance reading of the unknown solution using the spectrophotometer.

2. Find the position of the unknown’s absorbance reading on the y axis of the calibration curve.

3. Use a ruler to draw a horizontal, straight line from the absorbance reading on the y axis to the calibration curve (the best fit line).

4. At the point where this horizontal line intersects the calibration curve, draw a straight line vertically down to the x axis.

5. Read the concentration from the x axis where the vertical line intersects the x axis. This is the concentration of unknown in the solution.

The concentration of solute in the unknown solution is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Calibration Curve**

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(Questions are on the next page.)

**Questions**

1) Why is it necessary to use a blank when using the spectrophotometer?

2) Why do you have to use a volumetric flask instead of an Erlenmeyer flask or beaker when preparing the solutions in this experiment?

3) If you need to prepare 100.0 mL of a 0.080 M solution, how much of a 1.6 M stock solution do you need to use? (Show your calculations)

4) If you need to prepare 100.0 mL of a 0.12 M solution, how much of a 0.80 M stock solution do you need to use? (Show your calculations)

5) For the dilution equation, Mc • Vc = Md • Vd , what can be said about the relative values of the concentrations and the relative values of the volumes? Use “less than” or “greater than” in your answers.

The value of Vc is always \_\_\_\_\_\_\_\_\_\_\_ the value of Vd because \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

The value of Mc is always \_\_\_\_\_\_\_\_\_\_\_\_ the value of Md because

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

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**CHE116 Rowan College at Burlington County  
Week 11**

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Lab Partners: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Experiment** Death of A Businessman (This is adapted from Professor Peter Jeschofnig at Colorado Mountain College.)

**Purpose** The purpose of this experiment is for students to determine the killer of the businessman. The laboratory technique of solution preparation, the spectrophotometer, and a calibration curve will be used.

**Summary** A businessman was found dead in his backyard. There was a large amount of water in the dead man’s lungs; therefore the cause of death was concluded to be drowning. The police suspected one of his three grown children. Each of his children had access to a large body of water in which to drown someone. The daughter had access to her swimming pool, the rancher son had a large pond on his property, and the engineer son had access to a mine-waste settling pond. The coroner obtained a sample of water from the dead man’s lungs. The police obtained water from each of the water sources: the swimming pool, the ranch pond, and the settling pond.

An analysis must be made of the phosphate content of each of these water samples to determine which one of the children killed their father; the assumption is that finding where the businessman was drowned will identify the killer.

**Chemicals and Equipment Needed**

Spectrophotometer

Glassware: Erlenmeyer flasks, beakers, pipets, volumetric flasks, test tubes

Ammonium molybdate, (stock solution already prepared)

Tin(II)chloride (stannous chloride), (stock solution already prepared)

Phosphate solution (stock solution already prepared)

Deionized water

Disposable droppers

Parafilm®

**General Procedure**

1. Rinse all glassware with deionized water to avoid phosphate contamination.

2. Prepare the standard solutions that will be used to make the calibration curve. The phosphate solution is the solute and deionized water is the solvent. There will be a total of 4 solutions prepared in this step.

3. Prepare the color in the standard solutions, the water samples, and the blank (directions are on page 3). There will be a total of 9 solutions prepared in this step.

4. Use the spectrophotometer to find the absorbance of all of the solutions; the four standard solutions and the four water solutions. Use the blank to zero the spectrophotometer. (λ = 650 nm.)

5. Construct a calibration curve using the concentration values and absorbance values of the standard solutions.

6. Use the calibration curve and absorbance values of the four water samples to determine the phosphate concentration in each of the water samples.

7. Determine who killed the businessman; use the concentration of phosphate in the lung water and the concentration of phosphate in the other three water samples to make your conclusion.

Students should work in groups of four; each student can prepare one standard solution, then all four students can prepare the color in the solutions. The color preparation for all solutions must be done at the sample time, not one after another. All four students should use the spectrophotometer together. Each student should construct their own calibration curve and make their own conclusion.

**Table 1 Standards for the Calibration Curve**

Concentrated phosphate solution concentration: 20.0 ppm

|  |  |  |
| --- | --- | --- |
| Standard Solution Concentration | Volume Prepared | Volume of stock solution needed (pipetted) |
| 1.0 ppm | 100.0 mL |  |
| 2.0 ppm | 100.0 mL |  |
| 3.0 ppm | 100.0 mL |  |
| 5.0 ppm | 100.0 mL |  |

**Color Development for the Standard Solutions and Water Samples**

In order to use the spectrophotometer, all of the solutions must absorb visible light. The solutions (four standard solutions, four water samples, and one blank) must be put through an oxidation/reduction reaction to form a colored product.

**Colored Solution Preparation**

1. Place 25.0 mL of each solution into a separate Erlenmeyer flask (use a graduated cylinder to measure the volume of each solution). There are 9 solutions that will each be put into individual Erlenmeyer flasks (4 standards + 1 blank + 4 samples = **9 Erlenmeyer flasks**).

2. Add 1.00 mL of ammonium molybdate solution into each flask and swirl (pipet).

3. Add 2 drops of tin(II) chloride solution (stannous chloride) to each flask and swirl (disposable dropper).

4. Let the color develop for at least 5 minutes.

5. Find the absorbance of each solution **before** the color has developed for 15 minutes.

Remember to use the blank solution you prepared to zero the spectrophotometer.

**Experimental Results**

The absorbance of each solution should be listed in the table below:

**Table 2 Absorbance Data**

|  |  |
| --- | --- |
| **Solution ID** | **Absorbance** |
| Standard solution 1.0 ppm |  |
| Standard solution 2.0 ppm |  |
| Standard solution 3.0 ppm |  |
| Standard solution 5.0 ppm |  |
| Sample: Pool Water |  |
| Sample: Ranch Pond Water |  |
| Sample: Settling Pond Water |  |
| Sample: Lung Water |  |

Prepare your calibration curve with the concentration on the x axis and the absorbance on the y axis. Show (on your graph) how you determined the phosphate concentration in each water sample.

**Table 3 Phosphate Concentrations**

|  |  |
| --- | --- |
| **Sample ID** | **Phosphate Concentration, ppm** |
| Sample: Pool Water |  |
| Sample: Ranch Pond Water |  |
| Sample: Settling Pond Water |  |
| Sample: Lung Water |  |

**Who Killed the Businessman: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Explain Your Answer:**

**Calibration Curve**

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
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(Questions are on the next page.)

**Questions**

1) What would happen to your absorbance readings if you did not rinse your glassware with deionized water before beginning this experiment? Explain in detail.

2) Why is it acceptable to use a clean, wet volumetric flask, instead of a clean, dry volumetric flask, when you are preparing your standard solutions by dilution? Explain in detail.

3) Why did you have to perform the color development step for all of the solutions at the same time? Explain in detail.

4) When calculating the concentration of a diluted solution, a person can use the dilution equation: Mc Vc = Md Vd

Why does Mc Vc equal Md Vd ? What is this mathematical statement saying? Explain in detail.

5) If 10.0 mL of a 10.0 ppm solution is used to prepare 50.0 mL of a new solution, what is the concentration of this new solution?

CHE116 Rowan College at Burlington County

Week 12

Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Partner: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Experiment Introduction to Titrations**

**Purpose** This experiment introduces the usefulness of a titration and how a titration is done.

A titration is a wet chemical technique performed to determine information about an unknown substance. The determination of the molarity of an acid solution is a very common type of titration, and that is the type of titration for this experiment. A solution of sodium hydroxide with a known concentration is used to determine the molarity of a hydrochloric acid solution.

When mixed together, the sodium hydroxide and hydrochloric acid undergo an acid base neutralization reaction. The clear and colorless solution of sodium hydroxide reacts with the clear and colorless solution of hydrochloric acid to produce a clear and colorless solution containing the products of the reaction. The titration is stopped when the moles of OH1- added is equal to the moles of H1+ in the titration flask; this is the equivalence point. In order to “see” the equivalence point, an indicator is used to produce an observable change in the appearance of the hydroxide and acid mixture; this is the end point. The end point should occur at the equivalence point.

Chemicals and Equipment

0.10 M NaOH solution

Approximately 0.1 M HCl solution

Phenolphthalein Indicator solution (dropper bottles)

One 50 mL buret

One buret stand with clamp

One Erlenmeyer flask, 125 mL size

One 10 mL pipet and pipetter

One narrow stem funnel

Beakers, various sizes

Recommended YouTube Titration Video

<https://www.youtube.com/watch?v=9DkB82xLvNE>

**Procedure**

Do one trial at a time, from start to finish. After the trial is complete, dispose of the titration solution in the appropriate waste container. Fill out Table 1 as you do each trial of the experiment.

1. Obtain the buret stand, buret clamp, and buret. Assemble as shown by the instructor.

2. Turn the stopcock to the open position, and rinse the buret with approximately 10 ml of DI water. Let the rinse water drain into a large waste beaker. Repeat this rinsing step two more times.

3. Rinse the buret with approximately 5 mL of 1.0 M sodium hydroxide solution. Let the rinse drain into the waste beaker. Repeat this rinse step two more times.

4. Fill the buret with the 1.0 M sodium hydroxide solution; open the stopcock to drain out some of the sodium hydroxide solution so the top of the solution is within the calibration marks and the air bubble in the buret tip is removed. Have the waste beaker under the buret tip while this is done.

5. Record the volume of the sodium hydroxide solution in the buret; this is the initial volume of sodium hydroxide solution. Remember to use two digits after the decimal point when you record a buret volume.

6. Pipet 10.0 mL of the hydrochloric acid solution into the Erlenmeyer flask.

7. Add three drops of phenolphthalein solution and approximately 10 mL of deionized water to the hydrochloric acid solution in the Erlenmeyer flask. Swirl to mix well. The acidic solution should be colorless.

8. Place the Erlenmeyer flask under the tip of the buret, as shown in the titration video.

9. Perform the titration by opening the stopcock to allow the sodium hydroxide solution to go into the Erlenmeyer flask and react with the hydrochloric acid solution; swirl the flask while this is being done. There should be some pink color at the point where the sodium hydroxide enters the acid solution. When this pink color starts to be present for a longer period of time, slow down the rate at which the sodium hydroxide is added. Remember to swirl. When the pink color lingers for a longer period of time, add the sodium hydroxide dropwise; first fast drops and then slow drops. When the very faint pink color does not disappear, stop the titration. The color should exist in the solution for at least one minute; do not titrate during this minute.

10. Record the final volume of sodium hydroxide in the buret. Use two digits after the decimal point for the volume.

11. This first titration was the practice trial. Repeat the experiment three more times to obtain data for trials one, two, and three. The buret does not need to be rinsed in between trials. Add more sodium hydroxide solution to the buret when there is not enough solution in the buret for the next trial. Always make sure there is enough solution in the buret before you start a trial.

Molarity of the NaOH solution: \_\_\_\_\_\_\_\_\_\_\_\_\_\_

Volume of HCl solution pipeted into flask for each trial: \_\_\_\_\_\_\_\_\_\_\_\_

**Table 1: Titration Data for All Trials**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Practice Trial | Trial 1 | Trial 2 | Trial 3 |
| Final NaOH volume reading |  |  |  |  |
| Initial NaOH volume reading |  |  |  |  |
| Volume of NaOH used |  |  |  |  |

\*\*Do not use the practice trial data for calculations.

**Table 2: Titration Calculations**

|  |  |  |  |
| --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 |
| Moles of NaOH used |  |  |  |
| Moles of HCl in the Erlenmeyer flask |  |  |  |
| Molarity of the original HCl solution |  |  |  |

Average Molarity of HCl : \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Write the acid base neutralization reaction below:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Show calculations for each trial below:

**Trial 1:**

moles NaOH used:

moles of HCl in flask:

Molarity of HCl:

**Trial 2:**

moles NaOH used:

moles of HCl in flask:

Molarity of HCl:

**Trial 3:**

moles NaOH used:

moles of HCl in flask:

Molarity of HCl:

**Average Molarity of HCl:**

**Questions**

1. Write the definitions for the end point and the equivalence point for a titration.

End point:

Equivalence point:

2. After you rinse the buret with DI water, you are instructed to rinse the buret with the NaOH solution. Why do you need to rinse the buret with the NaOH solution?

Use the following description to answer questions 3 – 6; **show all calculations:**

A titration is done using 1.0 M NaOH in the buret and 15.0 mL of 1.8 M HCl in the Erlenmeyer flask.

3. How many moles of HCl are in the flask before the titration is started?

4. How many moles of HCl are left in the flask after a total of 5.0 mL of the 1.0 M NaOH solution were added to the flask?

5. How many moles of HCl are left in the flask after a total of 15.0 mL of the 1.0 M NaOH solution were added to the flask?

6. How many milliliters of NaOH solution are needed to reach the equivalence point?