**Lab 1**  **Lab Techniques and Observations**

**Purpose**:

* To acquaint the student with the laboratory and the equipment
* To practice various laboratory skills and develop observation skills
* To begin to see the complexity and orderliness of the Creation

**Apparatus**:

* Bunsen Burner
* Matches
* Graduated Cylinder
* Sugar
* Electronic Balance
* Spatula
* Various Types of Equipment
* Various toys and items for investigation
* Sandy Water

**Procedure**:

1. Set up the Bunsen burner and demonstrate ability of lighting it and adjusting the flame from a "cool" flame to a "hot" flame. Explain what a "cool" flame looks like and what a "hot" flame looks like.

See:

<http://ch185.semo.edu/labsafe/bunsen.html>

<http://en.wikipedia.org/wiki/Bunsen_burner>

2. Go to the various graduated cylinders around the room and record the volume of

the liquid inside it. Be sure to right down the number of the cylinder next to your volume reading.

3. Demonstrate to the teacher proper procedure for massing out 3 grams of sugar.

4. Go to the table with the toys and other items and make as many observations about them as possible. Record your observations and be ready to discuss them in class.

5. Practice decanting a mixture.

6. Practice pouring liquids from one beaker to another with the use of the stirring rod.

**Laboratory Techniques**:

**1. Measuring Volume in a Graduated Cylinder or Beaker.**

* Get your eye level with the liquid.
* Read the bottom of the meniscus. This indicates the correct volume. If you read

as high as the liquid "creeps" up the sides of the container, you will have an exaggerated volume.

Incorrect Method Correct Method

**2. Bunsen Burner**

* Connect the Bunsen burner hose to the gas outlet.
* Close the air vents on the Bunsen burner.
* Open the gas valve at the base of the Bunsen burner if it has a valve.
* Open the main gas valve outlet.
* Light a match and bring it from the bottom of the Bunsen burner up to the top of the burner, along the side of it. The gas should ignite. If it does not, open the gas valves some more. If it still does not open the air vents a little.
* When the gas ignites it should form a yellow flame.
* Open the air vents and adjust the gas flow to get a blue flame.

See:

<http://ch185.semo.edu/labsafe/bunsen.html>

**3. Handling Solids.**

* Read the label carefully to make sure you have the correct substance. Read for any precautions. Open the lid. Use the spatula to obtain some of the solid. Pour the solid on a piece of paper. Either roll the paper or funnel it so that the solid can be poured into the test tube or beaker.
* Never put chemicals back into the original bottle. Always take out less than you need and return for more if you need more. If you ever have excess solid or liquid chemicals it is better to discard them than to return them to the original bottle, otherwise you risk contaminating the entire stock bottle.

**4. Handling Liquids.**

* Always read the labels carefully to ensure you have the correct materials and so you are aware of safety hazards.
* Pour liquids into another bottle by using a stirring rod. Place the stirring rod over the lip of the bottle from which you are pouring. Tip the bottle. The liquid will run down the stirring rod into the other container. This will avoid dripping and spillages.
* It is always a good idea not to place the stopper or lid on the table top. These stoppers have chemicals on them and will either leave the chemical on the table top, or may be contaminated by something else on the table top. It is best to hold the stopper yourself as pictured below or give it to a partner to hold until you are completely done pouring the liquid. Be sure that the stopper is immediately returned to the stock bottle, and make sure it is returned to the correct stock bottle.



**5. How to Heat a Material in a Test Tube.**

* Always have a good grip on the test tube. Always point the top of the test tube away from yourself and do not point it at others. Continuously move the test tube.



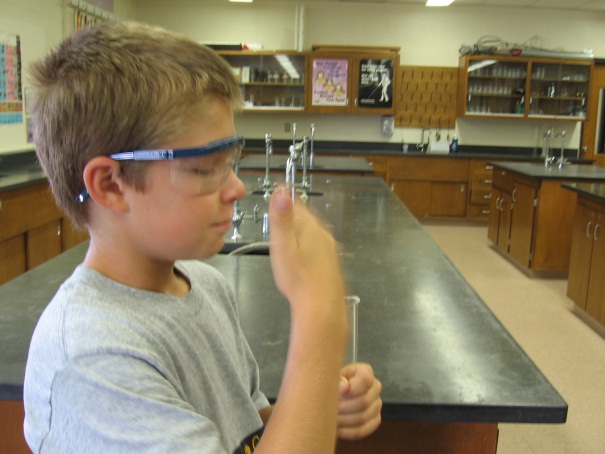
**6. How to Use a Mortar and Pestle**.

* This apparatus is used to grind things to powder or to smaller pieces. It should only be used for one substance at a time. Pound to reduce the object in size, and grind to pulverize it.



**7. How to Smell for Odors.**

* Never smell the substance directly, but waft the odor towards your nose.

**Lab 2 Candle Lab**

**A. Content/Context:**

* Observation Skills
* Connecting Observations with Conclusions

**B. Procedure:**

* Each group of students should obtain a small Petri dish, birthday candle, paper-clip stand, test tube (a variety of widths may be helpful) and a book of matches.
* Stand the candle up using the one-hole rubber stopper as a base.
* Place candle in the center of the Petri dish and add water to a depth of a few centimeters.
* Light the candle and let if burn for a moment. Carefully cover the candle with the test-tube, lowering the test-tube over the candle and into the water.
* Make observations.
* Repeat the experiment many times. Keep track of your procedure and what is kept constant and what you vary. (Change test-tubes perhaps, change speed at which you cover the candle, change number of candles which are covered).
* NOTE: To redo the test, smoke from previous trials must be expelled from the test tube.

**C. Equipment:**

* Petri Dish, test tube, matches, candle, one-hole rubber stopper

**D. Student Questions:**

* List all your observations
* Provide an explanation that accounts for all your observations. Many things happened – give a scientific explanation for what you observed.
* After you have written a thorough explanation that accounts for all the observations you have made, give your explanation to a classmate. The classmate must write a peer evaluation. Things to consider for the peer reviewer:
  + Does your classmate’s scientific explanation account for all the observations?
  + In what ways can the explanation improve?
  + Other suggestions for improving the explanation?

**Lab 3**  **Density**

**Purpose**:

* To see the relationship between mass and volume
* To develop abilities to graph and develop lab techniques
* To determine the density of the substance from a graph
* To compare the calculated density to the true density
* To see the orderliness and beauty of the Creation

**Apparatus**:

* electronic balance
* unknown samples
* graduated cylinders

**Procedure**:

1. Take a piece or a few pieces of sample A. Determine its mass tothe nearest 0.01 gram. Record the mass.

2. Take sample A and find its volume by displacement of water. Fill a graduated cylinder about a quarter full of water. Measure and record the initial volume. Tip the graduated cylinder sideways slightly and gently add sample A being sure not to break the graduated cylinder. Measure and record the final volume. What is the volume of sample A? Record it.

3. Dry off your sample A and return it to its place. Clean up your lab station.

4. Place your data for the average mass and average volume of sample A on the board. Copy down the class data

5. Using the class data plot a graph of mass versus volume. Determine a "best fit" line through the points.

6. Determine the slope of the line using points “picked” off from your “best-fit” line.

7. Once you have written your lab report, including the “Questions” section below, give your lab report to a classmate who will perform a peer review of it. The Peer Review will examine the following:

* Style, format, spelling, grammar, readability, etc. (10 pts)
* Conclusion: logical, consistent, thorough, good analysis, explains the results well, shows purpose was accomplished, etc. (20 pts)
* Questions: answered thoroughly and clearly, good, well-thought out answers, etc (10 pts)

8. Make necessary corrections according to the advice of the Peer Review. Submit original lab report, final lab report, and signed copy of Peer Review.

**Data/Results:**

1. Be sure to have your data recorded in lab notebook

2. Be sure to have a neat table of class data

3. Be sure to have the appropriate graphs, peer review, and corrections from peer review.

4. Answer the following questions:

a. Is density an extensive or an intensive property of matter? Why?

Prove your answer using the data collected in the lab.

b. On your graph, what does the slope represent? Remember what slope is as defined in mathematics. m = . . . . .

c. What does the experiment demonstrate about the density of a substance?

d. Use some serious thought to answer this question.

What would you have to do differently in terms of procedure to obtain the density of these substances? Be sure to explain carefully how you would obtain these items’ volumes.

i.) Granulated sugar

ii.) Irregularly-shaped wood chips

**Discussion:**

Discuss your results, considering these questions:

1. What is the density of **your** Sample A?

2. What is the density of the class’s Sample A?

3. How did you determine the density of the class’s sample? What other ways are

there to determine the density of the class’s sample?

4. How do these values (from # 1 and 2) compare with the true values as found in

the CRC Handbook of Chemistry and Physics? Calculate the percent errors.

**Lab 4: Separation Techniques**

**Purpose:**

* To learn different methods to separate mixtures
* To develop laboratory skills and techniques
* To see God’s handiwork in the creation

**Procedure:**

**Part A: Simple Distillation**

a.) Mass a distillation flask.

b.) Add 30.0 mL of salt water to flask.

c.) Place rubber tubing on the side arm of the flask and wrap with wet paper

towels. Light Bunsen burner and heat flask slowly.

d.) Collect distillate in small beaker.

e.) Measure volume of distillate recovered.

f.) When distillation flask cools, find the mass of the flask and its contents.

**Part B: Paper Chromatography**

a.) Take four pieces of filter paper and draw a pencil mark one inch from the bottom of each piece of paper.

b.) At the center of each pencil line, place a small spot of ink (different color on each paper).

c) Place less than one inch of water into a wide test tube.

d.) Suspend the filter paper in the test tube so that the water does not rise above the pencil line. Place a stopper on the test tube to close the test tube and to help in suspending the filter paper.

e.) Watch the water rise up the filter paper. Observe what happens to the

ink.

f.) When the water level rises about ¾ the height of the paper, take the paper out of the test tube.

g.) Place your name on the top of the filter paper in pencil and allow it to dry.

h.) When it is dry, calculate the Rf value (Rf=Ds/Df; where Ds= distance solute rose and Df= distance solvent rose)

i.) Note how the different ink heights and colors compare to each other.

j.) Answer the following questions.

1. What did you find interesting about the three different dyes? Did they all act the same (have same Rf)?
2. Suggest reasons for differences in the behavior of the different dyes.
3. Why was a pencil used on the paper and not an ink pen?
4. Why do some colors “travel” further up the paper than others?

**Data/Results:**

1. Be sure to record the mass and volumes obtained from the distillation.

2. Be sure to attach the chromatography papers into your lab notebook.

3. Answer the questions regarding chromatography.

**Discussion:**

1. Summarize what you learned about separation techniques. Advantages of one over the other? When would you use one or the other?

**Lab 5 Spectrophotometry or Colorimetry: Absorption Spectra**

**Purpose:**

* To learn the basics of spectrophotometry
* To make an Absorption Spectrum
* To learn how to make a standard curve and determine the concentration of an unknown from the standard curve
* To learn to use technology
* To see the beauty and order of God’s creation

**Introduction:**

In this exercise, you will learn the basic principals of spectrophotometry and serial dilution and their practical applications. You will need these skills to complete other exercises throughout the semester. A spectrophotometer is a very powerful tool used in both the biological and chemical sciences yet operates by simply shining a beam of light, filtered to a specific wavelength (or very narrow range of wavelengths), through a sample and onto a light meter. Some basic properties of the sample can be determined by the wavelengths and amount of light absorbed by the sample (Using the Spectrophotometer, no date; accessed online on August 10, 2011 at: ).

Visible light consists of wavelengths ranging from 380 nm (blue violet) to 720 nm (red). When all wavelengths of visible light are present, the light appears "white" to our eyes. If any

wavelength is removed (absorbed), the remaining combination of wavelengths of light that fall

on our eyes is perceived by our brains as the "complimentary" color.

Many transition metal complexes absorb wavelengths of light in the visible range and, as a

result, their crystals and solutions are brightly colored. The energy of the absorbed photon is

used to excite d-orbital electrons from lower to higher energy levels.

If we allow white light to pass through a test tube containing a solution of copper (II) ion, the solution will appear blue colored because the Cu2+ ion absorbs a photon of light in the 570-600 nm range. In order to describe the amount of light passing through the solution, we use two

terms called "transmittance" and "absorbance".

Transmittance (T), or percent transmittance (%T) is defined as the ratio of the amount of light leaving the sample to the amount of light entering the sample (or transmitted light, I, to incident light, Io) – see equation 1.

(1) T = I/Io or %T = I/Io x 100

Absorbance (A) is a term that refers to the amount of light reduction as it passes through a

solution. Absorbance and transmittance are related according to equation 2.

(2) A = log (1/T) = −log T

An instrument that can measure the absolute or relative intensities of light passing through a sample is called a spectrophotometer or spectrometer. The spectrometer mechanism consists of

a light source, focusing lenses, a diffraction grating or prism to split light into different wavelengths, a sample holder or "cell", a photosensitive detector which measures the light

passing through the sample, an amplifier, and an output device such as a meter or recorder. (The Spec 20, 2002 – accessed online on August 10, 2011 at:)

The Beer-Lambert law (A=εCl; A = Absorbance, ε = molar absorbtion coeffient, C = molar concentration; l = sample length or path length) teaches that there is a linear relationship between the Absorbance and the Concentration of a solution. Therefore, a standard curve can be generated by making standard concentrations and measuring their corresponding absorbance values. From this standard curve it is possible to identify the concentration of an unknown solution by measuring its Absorbance value and using the Beer-Lambert law to calculate the corresponding concentration.

**Procedure:**

**Part 1 – Obtaining the Absorption Spectrum**

1. Obtain approximately a 5.0 mL sample of 0.150 M CoCl2 solution and place it in a cuvette.

2. Prepare a “blank” by placing 5.0 mL sample of distilled water in a second cuvette.

3. Set the spectrophotometer to 400 nm.

4. Zero the spectrophotometer by making sure the sample chamber is empty and closing the lid and adjusting the power knob until the spectrophotometer reads 0% T.

5. Using the “blank”, set the spectrophotometer to read 100% T (Place “blank” in the sample chamber and adjust the transmittance adjustment knob accordingly).

6. Place the CoCl2 sample into the sample chamber. Read and record the absorbance value.

7. Repeat steps 4-6 after increasing the wavelength by 25 nm. Continue to increase by 25 nm increments until 600 nm.

8. Plot a graph of Absorbance vs. wavelength. Draw a smooth curve to fit the experimental points. Identify and record the maximum in the absorption curve to the nearest multiple of 25 nm.

**Part 2 – Creating a Standard Curve**

1. Prepare a series of CoCl2 solutions according to the following table. Use a pipet (or buret or graduated cylinder if buret or pipet are not available) to add the specified amount of 0.150 M CoCl2 and distilled water to each cuvette.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Cuvette Number** | **1** | **2** | **3** | **4** | **5** |
| **Volume (mL) CoCl2** | **5.0** | **4.0** | **3.0** | **2.0** | **1.0** |
| **Volume (mL) water** | **0.0** | **1.0** | **2.0** | **3.0** | **4.0** |

2. Set the spectrophotometer (or colorimeter) to the wavelength of your maximum absorbance as found in Part 1 of the lab.

3. Adjust your spectrophotometer for zero %T (empty compartment) and for 100 %T (with “blank”) – see Part 1 steps 4 and 5.

4. Measure and record the Absorbance value for each solution.

5. Calculate the concentration of each solution using the dilution formula (McVc=MdVd).

6. Plot the Absorbance vs. Concentration. Find the best-fit line and its slope.

7. Measure the path length (width of cuvette). Calculate and record the value of ε.

8. Empty all solutions and clean up.

**Part 3 – Finding the Concentration of an Unknown**

1. Obtain 5.0 mL of an unknown concentration of CoCl2.

2. Using the same wavelength settings as Part 2, and ensuring the spectrophotometer has been set properly (Part 1, steps 4 and 5), put the unknown into the sample compartment.

3. Read and record the unknown’s Absorbance value.

4. Using the equation of best-fit, determine the Concentration of the unknown.

5. Using the Beer-Lambert equation (and the value for ε that was calculated in Part 2, step7), determine the concentration of the unknown.

6. Compare the value from #4 and #5.

**Data/Results:**

1. Be sure to prepare the appropriate graphs from Part 1 and Part 2.

2. Be sure to show the appropriate calculations from Part 1 – 3.

3. What is the concentration of the unknown?

**Discussion:**

1. Be sure to comment on the accuracy of the Beer-Lambert law.

2. Be sure to explain how you were able to calculate the unknown’s concentration.

3. Discuss how spectrophotometry is a useful tool.**Lab 6** **Introduction to Spectrophotometry**

**Purpose**:

* To learn how to use the spectrophotometer
* To use spectrophotometry as a means to identify unknown concentrations
* To develop graphing skills and graphing interpretation skills.
* To see the beauty and intricacies of the Creation

**Apparatus**:

* Spec 20
* 0.5 M FeCl3 solution
* Unknown concentrations of FeCl3
* 6 test tubes
* Squirt bottle
* Distilled water
* Pipettes
* Kleenex

**Procedure**:

**Part A: Calibration of the Spectrophotometer**

1. The Spectrophotometer should have been running for 10 - 20 minutes before class to warm up.

2. Close the sample compartment and be sure nothing was in it.

3. Adjust the Spectrophotometer to read 450 nm (This is the wavelength of light that the Spectrophotometer will emit and that will pass through all your samples.) 4. Adjust the Spectrophotometer to 0 % Transmittance. This means that all the light

being emitted at one end of the Spectrophotometer is being received on the other end. ***This should be the case since nothing is in the sample compartment.***

5. Place a dry clean cuvette in the sample compartment with a sample of the solvent (distilled water). Adjust the Spectrophotometer to read 100 % T.

Note: Always dry the outside of the cuvette to remove dirt, wetness, or fingerprints. It must be dried with a Kleenex and not with a paper towel or anything else. If this is not done you will get invalid answers. Hold the cuvette at the top since it will not be deep enough into the sample compartment to harm your readings.

**Part B: Making the Sample**

1. Fill the six labeled test tubes as follows.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Cuvette Number | 1 | 2 | 3 | 4 | 5 | 6 |
| Volume (mL) Water | 10.0 | 8.0 | 6.0 | 4.0 | 2.0 | 0.0 |
| Volume (mL) FeCl3 | 0.0 | 2.0 | 4.0 | 6.0 | 8.0 | 10.0 |
| Concentration of Solution |  |  |  |  |  |  |

2. Determine the concentration of each new solution considering that the original stock was 0.5 M. Use the formula McVc = MdVd. Note that in each case the volume of the new dilute solution is 10 ml.

**Part C: Creating Standard Curve**

1. Add solution 1 to the Spectrophotometer’s cuvette. Be sure the cuvette is clean and dry on the outside.

2. Place it in the sample compartment, close it, and record the Absorbance.

3. Remove the cuvette, throw away solution 1.

4. Repeat steps 1 – 3 for all the other solutions.

5. Plot the Absorbance values for solutions 1 – 6 vs. their concentrations.

6. Determine the best-fit line.

**Part D: Testing an Unknown**

1. Take a sample of the unknown and place it in the Spectrophotometer.

2. Record the Absorbance value.

3. Remove the cuvette from the Spectrophotometer, and throw away the solution.

**Data/Results:**

1. Make sure all the calculations and graphs are done carefully and shown.

2. Answer the following questions:

a.) Determine the concentration of the unknown.

b.) What are some advantages of spectrophotometry compared to other

analytical methods of identification?

c.) How would Absorbance change if a person placed yellow food coloring into your test tube with FeCl3?

**Discussion:**

Be sure to report what was the identity of your unknown (Unknown A, B, or C) and what its concentration was. Explain how you determined the unknown concentration. Comment on the value of spectrophotometry.

**Lab 7 Making Solutions Lab**

**Purpose:**

* To demonstrate knowledge and ability to make % Solution and Molarity solutions
* To practice laboratory skills
* To observe the order and beauty of God’s creation

**Apparatus:**

* Beakers
* Electronic Balance
* Spectrophotometer or Calorimeter
* Graduated Cylinder

**Procedure:**

**Part A: Solutions by Molarity**

1. Describe (calculate and show work in laboratory notebook) how to make 5.0 mL of a 0.34 M solution of copper (II) chloride. (Use distilled water!)
2. Show your teacher your calculations and receive approval to proceed.
3. Make the solution per your calculations and instructions.
4. Place a sample of water into your colorimeter (set to 635 nm) and calibrate it. Then place your solution into the colorimeter and read the absorbance value. Record the Absorbance value in your lab notebook.

5. Discard your solution into the #1 Waste Container.

**Part B: Solutions by % W-V**

1. Describe (calculate and show work in laboratory notebook) how to make 5.0 mL of a 3.5 % solution of copper (II) chloride. (Use distilled water!)

2. Show your teacher your calculations and receive approval to proceed.

3. Make the solution per your calculations and instructions.

4. Place a sample of water into your colorimeter (set to 635 nm) and calibrate it. Then place your solution into the colorimeter and read the absorbance value. Record the Absorbance value in your lab notebook.

5. Discard your solution into the #2 Waste Container.

**Discussion:** Explain your calculations and your process and accuracy in making solutions

**Lab 8**  **Empirical Formula Determination**

**Purpose**:

* To determine the empirical formula for a compound
* To develop lab techniques
* To see the beauty of the Creation; to see the complexity, yet order, of God's Creation

**Apparatus**:

* crucible and lid
* crucible tongs
* pipettes
* electronic balance
* Bunsen burner and apparatus
* Sandpaper
* distilled water
* magnesium

**Procedure**:

1. Obtain a piece of magnesium metal. Sand it down with sandpaper so that it is

shiny. Find the mass of the piece of magnesium by using the following conversion factor: ***1 meter of magnesium has a mass of 1.038 grams.***

2. Mass the crucible and lid and record these values.

3. Set up the Bunsen burner and ring stand apparatus, with the clay triangle on the top of the ring stand. Place the crucible on the clay triangle so that it is stable.

4. Roll the magnesium into a loose coil. Place it in the crucible. Put the lid completely on the crucible.

5. Begin to heat the crucible gradually. Heat slowly by moving the flame around underneath the crucible. Remove the heat if large amounts of smoke come out of the crucible. DO NOT LOOK INTO THE CRUCIBLE AT THIS STEP. BRILLIANT LIGHT IS EMITTED FROM THE REACTION AND WILL CAUSE BLINDNESS TO OCCUR.

6. After about 5 minutes of direct heating with no smoke, remove the lid slightly. Heat the crucible to redness for four minutes.

7. Heat vigorously for 5 minutes with the lid completely off.

8. Remove the heat. Replace the lid. Allow the crucible to cool to room temperature. Find the mass of the crucible, the cover, and the contents. Record the mass.

***The Following Steps Should Be Followed Only If You Have Adequate Time!***

9. Add 10 drops of distilled water to the contents of the crucible. Smell cautiously and record any odors. Put the crucible on the ring-stand and heat again with the lid on. Allow the crucible to cool.

10. Mass the crucible, lid, and contents again. Record the mass.

11. If time remains, heat the crucible again, allow it to cool, and mass it. Continue to do this until the mass does not differ significantly between massings.

12. Clean up. Dispose of contents into the garbage.

13. Wash your hands thoroughly with soap.

**Data/Results:**

In this section be sure to demonstrate that you can make the necessary calculations in order to determine the empirical formula for the material that now remains in your crucible at the end of the experiment. A few hints, or things to remember.

If you want the empirical formula you must calculate the number of moles of the various elements that compose your final product. We will assume that the product is some form of magnesium oxide (contains only magnesium and oxygen). Therefore any difference in mass that occurred from the time you began the lab to the time you ended the lab must have been a result of oxygen combining to the magnesium.

**Discussion**:

Be sure to indicate what the calculated empirical formula is, as well as what the theoretical empirical formula should have been. How do your results compare?

What kinds of things may have contributed to your percent error?

**Lab 9**  **Quantitative Analysis with Baking Soda and HCl**

**Purpose**:

* To practice laboratory techniques
* To collect a final product and to determine the percent yield.
* To make appropriate calculations.
* To see the beauty of the Creation

**Apparatus**:

* Bunsen burner and ring-stand apparatus
* evaporating dish
* watch glass
* matches
* baking soda
* 6 M HCl

**Procedure**: 1. Take an evaporating dish and sterilize it by heating it in the Bunsen burner

flame for about 5 minutes. Allow the dish to cool to room temperature.

2. Find the mass of the evaporating dish and a watch glass. Record the mass.

3. Mass out approximately 2.50 grams of baking soda (NaHCO3).

Record the mass.

4. Place the evaporating dish on the ring-stand apparatus.

5. Place the watch glass on top of the evaporating dish.

6. In a test tube, obtain about 5 ml of the 6 M HCl and a pipette.

7. Place the tip of the pipette into the lip of the evaporating dish, and slowly

add a few drops of the 6 M HCl.

8. Continue to do this with occasionally rockings of the evaporating dish so that the HCl reacts with all the baking soda. Continue to add HCl until there is no observable reactions occurring in the evaporating dish.

9. Remove the watch glass and with a squirt bottle, rinse any residue off the watch glass into the evaporating dish. Be sure not to lose any product by dripping beside the evaporating dish.

10. Place the watch glass on the lab table, wet side up.

11. Gently heat the evaporating dish. Use a low flame and move the flame

around below the evaporating dish. Rapid heating will result in splattering of your product - thus, a lost product.

12. When almost all the liquid is gone, replace the watch glass, leaving a small opening to allow vapor to escape. Continue to heat gently until the product looks very dry.

13. Allow the apparatus to cool.

14. Find the mass of the watch glass, evaporating dish, and contents.

Record the mass.

**Data/Results:**

Be sure that this section shows all the calculations necessary to demonstrate what you would expect to be the theoretical yield. Remember that in order to determine the theoretical yield you need to have a balanced chemical reaction and do the proper stoichiometrical calculations to determine the correct amount of grams that would be produced of the final product.

Also determine the percent yield by comparing the actual yield and theoretical yield as we have done in class.

**Conclusions:**

Be sure to discuss what happened in the experiment on the molecular level. Also discuss whether or not the experiment was a success and explain why.

Be sure to discuss what could have been some sources of error.

**Questions:**

1. What caused the bubbles when the HCl reacted with the baking soda?

Suggest a method for testing this.

2. What effect does the *loss of product* have on percent yield?

(That is, does percent yield increase or decrease ***and why***)

3. What effect do i*mpurities* have on percent yield?

4. What effect does *wet product* have on percent yield?

5. What effect does *unreacted baking soda* have on percent yield*?*

**Lab 10**  **Quantitative Analysis with Copper Solution and Iron Nail**

**Purpose**:

* To observe a single replacement reaction.
* To make stoichiometric calculations
* To determine and predict the amount of Copper solid produced in the reaction.
* To observe and use God ordained laws of Creation to determine the amount of a substance that is produced in a reaction

**Apparatus**:

* iron metal
* 0.5 M CuCl2 solution
* 250 ml beaker
* stirring rod
* electronic balance
* spatula
* tongs

**Procedure**:

1. Find the mass of a clean dry 250 ml beaker. Record it.

2. Determine the amount of CuCl2 necessary to dissolve in 50 ml of water to make a 0.5 M CuCl2 solution.

3. Add the amount from # 2 into the beaker with 50 ml of water and stir to dissolve the copper chloride.

4. Find the mass of an iron metal. Try to get a piece about 1 gram in size. Record its mass.

5. Place the metal in the solution for 30 minutes. Do not disturb it.

6. After 30 minutes check with the teacher if you can proceed. You may have to wait until next period to finish the lab.

7. Remove the metal with a set of tongs.

8. Scrape off any residue on the metal. Rinse the metal down well with distilled water. Get as much product off the metal as possible. Set the metal aside on a piece of paper towel to dry.

9. When the metal dries, find and record its mass.

10. Decant the liquid in the beaker into another beaker. Try to keep as much solid in the original beaker as possible.

11. Rinse the solid with distilled water and decant again.

12. Repeat # 11.

13. Wash the solid with 25 ml of 1 M HCl. Decant again.

14. Rinse the solid with 25 ml of distilled water and decant again.

15. Allow the solid to dry completely over night.

16. Find the mass of the solid and the beaker together. Record it.

17. Clean up.

**Data/Results:**

The goal in the lab is to determine how much copper solid should have been produced in theory and to compare this value with what actually was produced. Show all the appropriate balanced equations and stoichiometrical calculations that are necessary to give a theoretical yield of copper metal. You should assume in the balanced equation that the iron metal will turn into the iron + 3 ion.

**Conclusions**:

Be sure to explain what happened in the experiment from a molecular level. Be sure to indicate how much copper solid was actually produced compared to how much should have been produced in theory. Determine the percent yield. Give reasons why the percent yield is not 100%.

**Questions:**

1. How would your results have changed had the iron turned into Fe+2 rather than the Fe+3 ion? Be sure to give specific values.

2. Assume you have an unlimited supply of CuCl2. How many grams of iron would have been used up if 45.0 grams of copper were to be produced?

3. In this lab, why was the hydrochloric acid used to “wash” the product?

**Lab 11 Cumulative Lab Exam – Reactions and Stoichiometry Lab**

**Purpose:**

* To conduct a number of reactions and correctly identify each by type
* To calculate the theoretical yield for each reaction, using stoichiometrical calculations
* To determine the percent yield in each reaction
* To develop in laboratory skills and techniques
* To see the vast richness and variety in God’s creation – evidenced by the variety of reactions.

**Procedure:**Conduct the following chemical reactions. Each time: label the type of reaction it is, write a balanced chemical reaction, determine the amount of the specified product using stoichiometry, and calculate the percent yield.React a piece of aluminum (find its mass) with 20 mL of a 0.1 M copper (II) sulfate solution. Find the mass of the solid particles that form in this reaction. Make sure the solid particles dry well before massing them.React 10.0 ml of a 0.1 M KI solution with 10.0 mL of a 0.1 M lead (II) nitrate solution. Allow the solid precipitate to settle well and then make sure the product dries well before massing it.Place about 2.00 grams of baking soda into a previously massed crucible. Heat it for about 10 minutes with a Bunsen burner. Break up any clumps with a glass stirring rod. Once the crucible cools, weigh up the remaining product and crucible. Subtract the mass of the crucible to find the mass of the product.**Data/Results:**For each chemical reaction, be sure to:List all observationsShow all masses**Conclusion/Discussion:**For each chemical reaction:Write a balanced chemical reactionShow the stoichiometrical calculations to calculate the mass of the desired productCompare the theoretically calculated mass of the product to the actual product mass.Calculate percent yield in each case.**Lab 12 Colligative Properties**

**Purpose:**

* To measure the melting point of pure vs. impure samples of material.
* To practice laboratory skills
* To see the beauty of God’s creation

**Apparatus:**

* Stearic Acid (reagent and laboratory grade)
* Disposable plastic test-tubes or “weighing cups”
* Thermometers or LabQuests with thermometer probes

**Procedure:**

1. Place 150 ml of water in a 250 ml beaker and begin to warm on a hotplate.

2. Obtain two disposable plastic test tubes or two “weighing cups” and label as A and B.

3. Scoop some Stearic Acid (Reagent Grade) into test tube A and some Stearic Acid (Laboratory Grade) into test tube B. Use about ¾ of an inch of powder.

4. Set up LabQuest with two thermometers in Channels 1 and 2.

5. Set the settings to “Time Based” and Length **to “1800 seconds”**

6. Place Channel 1 thermometer into Test Tube A and Channel 2 thermometer into Test Tube B.

7. Using a system of clamps or other method to keep plastic tubes in the warm water, place test tubes into the warm water bath.

8. Start recording data.

9. Observe and record which sample melts first. (Also record the temperature when all the material is melted).

10. When both substances have fully melted, remove the test tubes from the warm water bath and place them in a test tube rack or into an empty beaker (keep thermometers in the liquids).

11. Observe and record which sample freezes first. (Also record the temperature that crystals first appear).

12. Continue to record data until both substances have cooled to about 35oC.

13. Stop recording data and save data.

14. When opportunity arises – transfer data to library or media lab computers and print the graph and save a copy in your network folder.

a.) Connect USB to computer and LabQuest and open Logger Pro

b.) Click “File”; “Labquest Browswer”; “Import” and choose saved file

c.) Double click on graph and click on Axes Option tab

d.) Under Y axis column click on Run 1 and check Temperature 1 and Temp 2

e.) Under X axis column click on Time

f.) Give the graph a title and click Done

g.) Graphs should appear with a title and labeled axis.

h.) Print graph (no data table).

**Data/Results:**

1. Print a graph for each substance and trim and paste it into your lab notebook.

2. Determine the melting/freezing point for each substance.

**Discussion:**

1. Explain what is the difference between the two graphs and why.

2. Explain the difference between reagent and laboratory grade reagents.

3. How do your graphs show that one substance was pure and the other impure?

**Lab 13 Heats of Solution Lab**

**Purpose:**

* To learn how much heat is released or absorbed by dissolving substances in water
* To practice good laboratory skills and use technology
* To see the order and beauty of God’s creation

**Procedure:**

***Part A: Heat of solution for sodium hydroxide***

1. Measure out 100.0 ml of water. Record the exact amount you obtain.

2. Pour water into a Styrofoam cup.

3. Mass out approximately 10 g of sodium hydroxide. Record the exact amount you obtain.

4. Place Labquest thermometer into the water and let the temperature settle.

5. Make sure the Labquest has the following settings:

a.) Mode: Time based

b.) Rate: 2 samples/sec

c.) Length: 180 sec

6. Click the “Start” button.

7. After about 15-20 sec, pour the sodium hydroxide into the water.

8. Stir the solution with the thermometer constantly until the 180 sec is over.

9. When the Labquest stops collecting data, click on Analyze and Statistics.

10. Using the maximum and minimum temperatures, determine the ΔT for the water.

11. Calculate the heat absorbed by the water.

12. Determine the moles of sodium hydroxide dissolved in water.

13. Divide the heat absorbed by the water by the moles of sodium hydroxide that dissolved in water.

14. Calculate the percent error for the lab, using the true values for the Heat of Solution from the Text.

***Part B: Heat of solution for potassium nitrate***

1. Measure out 100.0 ml of water. Record the exact amount you obtain.

2. Pour water into a Styrofoam cup.

3. Mass out approximately 5 g of potassium nitrate. Record the exact amount you obtain.

4. Place Labquest thermometer into the water and let the temperature settle.

5. Make sure the Labquest has the following settings:

a.) Mode: Time based

b.) Rate: 2 samples/sec

c.) Length: 180 sec

6. Click the “Start” button.

7. After about 15-20 sec, pour the potassium nitrate into the water.

8. Stir the solution with the thermometer constantly until the 180 sec is over.

9. When the Labquest stops collecting data, click on Analyze and Statistics.

10. Using the maximum and minimum temperatures, determine the ΔT for the water.

11. Calculate the heat absorbed by the water.

12. Determine the moles of potassium nitrate dissolved in water.

13. Divide the heat absorbed by the water by the moles of potassium nitrate that dissolved in water.

14. Calculate the percent error for the lab, using the true values for the Heat of Solution from the Text.

**Data/Results:**

1. Print a copy of the graphs produced by the LabQuests.

2. Show calculations for the heat of solution for each substance.

3. Show the percent error calculations for each heat of solution.

**Discussion:**

1. Be sure to explain what happened for each trial and how the results were different.

2. Do all substances require heat to dissolve? Explain.

3. How well did your results match with the true values?**Lab 14**  **Qualitative Analysis Laboratory: “What is in my test tube?”**

**Purpose:**

* To develop a data base of observations on various reactions.
* To develop and practice separation schemes.
* To determine the cation identity from a random unknown sample
* To develop better laboratory techniques.
* To develop a familiarity for the centrifuge.
* To develop logic and problem solving skills and to work independently.
* To develop a better understanding and appreciation for God's Creation.

**Apparatus:**

* test-tubes
* pipettes
* centrifuge
* distilled water
* cation solutions (Pb(NO3)2, AgNO3, Ni(NO3)2, Fe(NO3)3, Cu(NO3)2, Ba(NO3)2)
* reactant solutions (NaOH, NaCl, Na2SO4, Na2CO3, NH4OH)
* confirmatory test solutions (KSCN, dimethylglyoxime, KI)

**General Procedure:**

In this lab the general idea is to react each of the six solutions with each of the five reactant solutions. Therefore, there will be 30 chemical reactions to observe. Each reaction should occur in a separate microcentrifuge tube. Initially, each student will **not** need 30 tubes, but rather should obtain six tubes. Ideally, each student should react the six solutions with ***two*** *of the five reactant solutions* in a class period. Therefore, after 3 class periods all of the reactions should be complete.

When reacting the solutions it is not necessary to use large quantities of solutions. Approximately 10 drops of solution will be adequate. Once you have a good idea of how much 10 drops of solution is, it will not be necessary to count the drops. **This lab is qualitative, not quantitative.**

Although this seems quite simply, it actually involves great observations and other work. After any two solutions are mixed, you should make and record observations (color changes, precipitate (ppt) formation, color and texture of ppt., ppt reactions, gas evolution etc.). The better your observations are, the better your data base of information will be. Every time there is a precipitate, you must centrifuge the solution and follow the specific procedural guidelines below to test the precipitate to see what kinds of reactions the precipitate will have.

Once a data base of information on the original 30 reactions as well as the precipitate reactions is collected, you need to do the confirmatory tests. Then you will be given a test tube with a number of cations (positive ions). Your job will be to test that solution using the knowledge of reactions obtained earlier in the lab and identify the cations that are present.

**Specific Procedure:**

**NaCl reactions:**

* React all 6 solutions with NaCl. Make and record observations. Centrifuge all solutions with a ppt. Remove the supernatant.

* Test ppts for solubility in hot water. Put a few drops of hot water into the test tube and place the test tube in a hot water bath. Observe whether or not the ppt dissolves in water.
* Repeat that specific reaction again to obtain the same ppt. Now test this ppt for solubility with a few drops of 15M NH4OH.

**Na2CO3 tests:**

* Test all 6 solutions with 7 drops of Na2CO3. Centrifuge all solutions that form a ppt. Remove the supernatant.
* Test the ppt with a few drops of 6 M HNO3. Record your observations. Watch for solubility, and the evolution of gases.

**NaOH tests:**

* Test all 6 solutions with 7 drops of NaOH. Centrifuge all solutions with a ppt.

Remove the supernatant.

* Test the ppt with a few drops of 6 M HNO3, mix well, and record observations.
* Repeat the specific reaction again to obtain the same ppt. Now test the ppt. with a few drops of 6 M NaOH, mix well, and record observations.

**Na2SO4 tests:**

* Test each solution with 7 drops of Na2SO4. Centrifuge all solutions with a ppt.

Remove the supernatant.

* Test the ppt with a few drops of 6 M NaOH, and place it in a hot water bath. If the ppt dissolves, add an excess of 3 M H2SO4, and observe any changes.

**NH4OH tests:**

* Test each solution with 7 drops of NH4OH. Centrifuge all solutions with a ppt.

Remove the supernatant.

* Test the ppt with a few drops of 15 M NH4OH, mix well, and record your observations.

**Confirmatory Tests Procedure:**

To make sure certain cations are present in a solution it is necessary to test our solutions with confirmatory tests.

Do the following tests:

1. Test for the presence of Fe+3, by taking a few drops of Fe(NO3)3 and adding a few drops of KSCN solution. Make observations.

2. Test for the presence of Ni+2, by taking a few drops of Ni(NO3)2 and adding a few drops of ***dimethylglyoxime solution***. Make observations.

Repeat this same procedure on all the cation solutions.

3. Test for the presence of Pb+2, by taking a few drops of Pb(NO3)2 and adding a few drops of KI. Make observations.

4. Test for the presence of Ag+, by taking a few drops of AgNO3, and adding a few drops of KI. Then test the ppt with 15 M NH4OH. Make observations.

5. Test for the presence of Cu+2, by taking a few drops of Cu(NO3)2, and adding an excess of KI. Make observations.

**Design a Scheme:**

See questions on the board. We will practice this together.

**Analysis of unknown:**

1. Obtain a list from your teacher of 4 possible cations in your unknown.

2. Using your data base, and reasoning developed above, develop a scheme for the separation and identification of your cations.

3. Have your teacher check your qual scheme. Test it by placing all 4 possible cations into a solution and separate out each cation, one at a time.

4. When you are satisfied that your qual scheme is working, obtain your unknown from your teacher. This unknown will contain one to four of the cations in your scheme. Subject this unknown to the tests set forth in your qual scheme and report which cations are present in your unknown.

**Lab Report**:

Write a formal lab report in your lab notebook:

a.) Include in the ***data/results*** your data base of observations, including your confirmatory results. Use one page per cation test. Therefore, on one page have a chart that includes all your observations for the reactions with NaCl; another page for the observations for the reactions with Na2CO3; and so on.

b.) Include in your ***data/results*** your qual-scheme design.

c.) Include in your ***discussion*** a detailed explanation of the procedure and tests performed in order to determine the identity of your unknown as well as the results you observed and the inferences you drew from each result and the logic behind these inferences. This section should clearly explain why you concluded that certain ions were present and others were absent.

Be sure to identify what unknown number you were and what you believed were the ions present in the solution and why.

**Lab 15 Galvanic Corrosion**

**Purpose:**

* To observe galvanic corrosion.
* To begin to understand oxidation and reduction.
* To write chemical equations.
* To see the complexity of God’s Creation.
* To see how to prevent corrosion.

**Apparatus:**

* Agar-Agar solution
* Iron metal
* Petri dishes
* Phenolphthalein solution
* Strips of zinc and copper metal
* Sandpaper

**Procedure:** 1. Prepare 50 ml of agar-agar solution. Measure out a mass of 0.5 grams of

powdered agar-agar. Heat 50 ml of water to boiling. Remove the water from the heat and add the agar-agar powder slowly while constantly stirring. Once the agar has dissolved, add 5 drops of phenolphthalein solution.

2. Sand the copper and zinc strips. Sand the iron metal as well so that the pure metals are exposed.

3. Take two pieces of iron metal and wrap them in the strips of copper and zinc metals. (One iron piece is wrapped in zinc and the other in copper.) Place the two wrapped pieces of iron into a petri dish. ***Be sure the metals do not touch.***

4. Slowly pour the agar-agar solution into the petri dish to a depth of about 1 cm above the metals.

5. Allow the petri dish to remain untouched for a day. Make observations.

**Data/Results:** Be sure to record all observations here.

**Discussion:** Be sure to comment on the reason for the color changes. Be sure to explain why the one iron metal has different results than the other iron metal. Be sure to relate these answers to the concepts of corrosion, oxidation and reduction.

**Questions:** 1. What is a cathode and what is an anode?

2. What does the “pink” color indicate?

3. What is oxidation?

4. Explain “corrosion” or “rust” in an electrochemical point of view, that is,

in terms of loss or gain of electrons.

5. Explain why the nail corroded in the one situation and not the other.

6. How is “corrosion” prevented in modern industries, such as with

automobiles?

7. What does the Bible teach about corrosion? What does the Reformed believer understand about corrosion?

**Lab 16 Boyle's Law**

**Purpose**:

* To use a "sticky-tape" apparatus and see how effective it actually is
* To use data to graph and see if results approximate Boyle's Law
* To predict with the data, graph, and equation for the line, new volumes with given pressures
* To observe and appreciate the God ordained laws of Creation

**Apparatus**:

* Boyle's law apparatus
* Tire pressure gauge
* Bicycle pump
* Calculator

**Procedure**: 1. Pump air into the Boyle's Law apparatus until the syringe inside changes

its volume significantly. This will occur when it becomes difficult to pump air into the bottle and when the bottle is very firm. Do not over-inflate.

2. Record the volume in the syringe as V1.

3. Check the pressure in the bottle with the tire pressure gauge. Record the

pressure as P1.

4. Release some air out of the bottle. Do not release a lot. Just release enough air so that the syringe changes volume by a single marking.

5. Record the new volume in the syringe as V2.

6. Check the new pressure in the bottle with the tire pressure gauge.

Record the pressure as P2.

7. Continue with steps 5 and 6 until you have several sets of data points, at least until V7 or V8. The more valid points the better.

8. Repeat the entire experiment if time permits.

9. Graph the data with Logger Pro.

Plot pressure versus volume.

-Note: adjust your data before entering it into the calculator.

What does it mean that your tire pressure gauge reads 0 pounds per square inch?

Does that *really* mean that there is *no* pressure in the bottle?

10. Determine the "best-fit" curve for the line. Determine the equation for the line.

11. Plot Pressure vs. 1/V

12. Find the "best-fit" curve and the equation for the line.

**Discussion**: 1. What is the relationship between pressure and volume?

2. What is the relationship between P and 1/V?

3. What are the equations for the lines in # 1 and # 2?

**Lab 17 Determining “R” – The Ideal Gas Law Constant**

**Purpose**: - to use a "sticky-tape" apparatus and see how effective it actually is

- to use data to determine “R”

- to observe and appreciate the God ordained laws of Creation

**Apparatus**: - Boyle's law apparatus - tire pressure gauge

- bicycle pump - calculator

**Procedure: Part A:**

* Obtain a sealed syringe. Measure the volume of air trapped in the syringe
* Determine the barometric pressure and temperature in the laboratory.
* Using the Combined Gas Law, determine the theoretical volume (V2) of the syringe at STP.
  + P1 = Atmospheric pressure in the laboratory
  + V1 = Volume of trapped gas in the syringe **before** you perform any experiments
  + T1 = Temperature in the laboratory.
  + P2 = STP pressure
  + V2 = STP volume (this is your unknown)
  + T2 = STP temperature

Using this information, determine V2.

Knowing the Molar Volume of a gas at STP (22.4 L per one mole of gas), set up a ratio to determine the number of moles of gas trapped in the syringe (using V2) as the STP volume of gas in the syringe.

**Procedure: Part B:**

* Insert into a “Boyle’s Law Apparatus” a sealed syringe and a thermometer.
* Fill the bottle with air.
* Measure the new volume on the syringe and the pressure in the bottle (being sure to add to the pressure of the bottle the atmospheric pressure)
* Repeat to obtain at least 5 good sets of data.
* For each set of data points, calculate the Ideal Gas Law Constant “R”.
* From the trials performed, determine the average “R” value. Report it.

**Data/Results:**

* Provide a neat table with all data points
* Provide a set of calculations for Part A – showing how the number of moles of gas in the syringe were determined.
* Provide neat calculations for each trial – showing how “R” was calculated.
* Show your final “R” value.

**Discussion:**

* Explain the “theory” behind your calculations.

**Lab 18 Specific Heat of a Common Metal**

**Purpose**:

* To become acquainted with heat calculations
* To practice and develop calorimetry procedures
* To work with a calorimeter
* To compare the experimental specific heat of a metal to its "true" value
* To learn more about God's Creation

**Apparatus**:

* coffee cups
* thermometers
* stirring rods
* water
* hot plate
* beakers
* lead samples
* electronic balance

**Procedure**:

**Part A: Calibrating a Coffee Cup Calorimeter**

1. Trim the lip off one of the coffee cups. Put a hole for the thermometer and one for

the stirring rod in the center of the coffee cup that has the lip trimmed off.

2. Measure approximately 50 ml of water. Record the exact volume that you have. Pour it into the untrimmed coffee cup.

3. Measure and record the temperature of the water in the cup.

4. Measure approximately 50 ml of water. Record the exact volume that you have.

Pour it into a beaker and place it on the hot plate.

5. Heat the water in the beaker until it is about 60 o C.

6. Measure the exact water temperature in the beaker and record it.

7. Immediately pour the hot water into the cup with the cool water. Make sure you

do not tip over the cup.

8. Quickly cover the cup with the water, with the trimmed cup which is equipped with

a thermometer and a stir stick.

9. Move the stirrer up and down to mix the water thoroughly.

10. Record the highest temperature attained. It could take 30 seconds.

11. Empty the calorimeter. Clean up.

12. Determine the calorimeter constant C' for your calorimeter.

*Heat lost by the warm water = Heat gained by cool water + Heat gained by the calorimeter*

q warm water = (mass warm water) . (4.184 J/ g . o C) . ( Δt )

q calorimeter = q warm water - q cool water

q calorimeter = C' . (Δt cool water)

C’ = q calorimeter / Δt cool water

**Part B: Finding the Specific Heat of a Common Metal**

1. Fill a 600 ml beaker with roughly 200 ml of water. Begin to heat the water on the hot plate.

2. Get approximately 70 grams of lead “sinkers”. Record the exact mass.

3. Place the sinkers in a clean, dry 250 ml beaker.

4. Place the 250 ml beaker into the water in the 600 ml beaker. Do not let water

overflow into the 250 ml beaker.

5. Let the water in the 600 ml beaker boil for 10 minutes. We assume that at that point the sinkers will be the same temperature as the water.

6. Meanwhile, set up your coffee cup calorimeter. Measure out exactly 100 ml of

water. Record the exact volume of water and then pour it into the coffee cup.

7. Measure and record the water temperature in the coffee cup calorimeter.

8. Measure and record the exact water temperature of the boiling water.

9. As carefully, but as rapidly as possible, have your partner hold the coffee cup near the hot beakers. Lift the small beaker with the sinkers out of the beaker

of water and quickly pour the sinkers into the coffee cup. Have your partner quickly place the trimmed coffee cup which is equipped with a thermometer and stirring rod on top of the coffee cup with the water and sinkers.

10. Stir the solution.

11. Find the highest temperature reached in the coffee cup calorimeter.

12. Clean up.

**Data/Results:**

1.Complete tables similar to the ones below.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Procedure A:** | |  | | |
| Mass (volume) of Cold water | |  | | |
| Initial Temp. of Cold water | |  | | |
| Final Temp. of Cold water | |  | | |
| Change in Temp. of Cold water | |  | | |
| Mass (volume) of Hot water | |  | | |
| Initial Temp. of Hot water | |  | | |
| Final Temp. of Hot water | |  | | |
| Change in Temp. of Hot water | |  | | |
| **Procedure B:** | |  |
| Mass of lead “sinkers” | |  |
| Initial Temp. of lead “sinkers” | |  |
| Final Temp. of lead “sinkers” | |  |
| Change in Temp. of lead “sinkers” | |  |
| Mass (volume) of Cold water | |  |
| Initial Temp. of Cold water | |  |
| Final Temp. of Cold water | |  |
| Change in Temp. of Cold water | |  |

2. Show calculations necessary to determine the “specific heat constant” for the lead “sinkers”.

3. Show the calculations necessary to compare the experimentally determined “specific heat constant” for the lead sinkers to the “true” value of lead as found in the CRC Handbook of Chemistry and Physics.

**Discussion:**

For this experiment be sure to state what your experimental results were and what they teach you about chemistry concepts.

**Questions:**

1. Why is it necessary to calibrate the coffee cup?

2. The specific heat of a material is often a determining property in its practical use. What metals would be the best metals to use in electrical wiring? Why do you think that copper is used in most electrical wiring? What other properties may have been considered in choosing copper for electrical wiring over other metals, such as, gold, aluminum, lead, or sodium?

3. Explain how the large specific heat of water is responsible for the moderating effects that large bodies of water have on climate.

**Lab 19** **Chemical Equilibrium Analysis**

**Purpose**:

* To observe Le Chatelier's principle in action
* To understand better equilibrium systems, and thus, God's Creation
* To accurately predict the outcome of adding a particular stress to a system

**Apparatus**:

* Thymol blue
* NaOH solutions (0.1 M and 6 M)
* HCl solutions (0.01 M, .1 M, 6 M)
* FeCl3 solution (0.20 M)
* KSCN solution (0.20 M)
* Fe(NO3)3 solution (0.20 M)
* KCl solution (0.20 M)

**Procedure: Part A:**

Important Information:

Thymol blue (blue color) + H+ <===> Thymol blue+ (yellow color)

**Pre-lab:**

1. Predict what will happen if you add a few drops of 0.01 M HCl to test-tube 2.

Explain your reasoning on the basis of equilibrium studies thus far.

2. Predict what will happen if you add a few drops of 0.1 M HCl to test tube 3.

Explain your reasoning on the basis of equilibrium studies thus far.

3. Predict what will happen when the same volume of 0.1 M NaOH is added to test-tube 3 as was added of 0.1 M HCl. Explain your reasoning on the basis of your equilibrium studies.

**Lab:**

1. Take 3 test-tubes and add Thymol blue from the stock bottle to each test-tube.

Label each test-tube, 1, 2, 3.

***What color is the liquid in your test-tubes?***

2. Leave test-tube 1 alone. Only experiment with test-tubes 2 and 3.

***Why is it important to leave test-tube 1 alone?***

3. Add a few drops of 0.01 M HCl to test tube 2. Observe what happens.

4. Add a few drops of 0.1 m HCl to test tube 3. Observe what happens.

5. Add several drops of 0.1 M NaOH to test tube 2. Observe what happens.

6. ***Write chemical reactions and statements that support your observations in # 3,4,5***.

**Part B:** Important Information:

Fe+3 + SCN- <====> Fe(SCN)+2

**Caution**: 6 M NaOH is highly corrosive. Hand and face protection must be worn. Clean up spills immediately. Notify instructor of spills. As always, wash hands thoroughly after lab.

**Pre-Lab:**

1. Predict the outcome of adding 10 drops of 0.20 M Fe(NO3)3 to test-tube B.

Explain your reasoning.

2. Predict the outcome of adding 10 drops of 0.20 M KCl to test-tube C.

Explain your reasoning.

3. Predict the outcome of adding 10 drops of 0.20 M KSCN to test-tube D.

Explain your reasoning.

4. Predict the outcome of adding 10 drops of **6.0 M NaOH** (***caution***) to test-tube E.

Explain your reasoning. Use solubility and formation constant tables if necessary.

**Lab:**

1. Place 1 ml of 0.20 M FeCl3 into a 250 ml beaker.

Add 1 ml of 0.20 M KSCN to the same 250 ml beaker.

Record the color of each solution as it existed initially and record the color of the solution resulting from mixing the two solutions.

Dilute the new solution with distilled water (about 100 ml) to dilute the intense color.

*Why add* ***distilled*** *water, rather than regular water?*

Place 5 ml of the diluted solution into 5 different test-tubes.

Label the test-tubes A, B, C, D, E. Leave A alone. Test only on B through E.

*Why leave test-tube A alone?*

2. Add 10 drops of 0.20 M Fe(NO3)3 to test tube B. Record Observations.

3. Add 10 drops of 0.20 M KCl to test tube C. Record Observations.

4. Add 10 drops of 0.20 M KSCN to test tube D. Record Observations.

5. Add 10 drops of 6.0 M NaOH to test tube E. Record Observations.

6. For the above trials, write chemical reactions and statements which demonstrate your understanding of these concepts and support your observations.

**Data/Results:**

1. Be sure to list your observations for each reaction that occurred.

2. Write a chemical equation for each reaction that occurred

**Discussion:**

Explain what happened in each specific test tube, in the light of equilibrium studies.

**Questions:**

1. What is Le Chatelier’s Principle?

2. How is Le Chatelier’s Principle being applied in this lab?

**Lab 20 Acids/Bases Quantitative Laboratory: Titration**

**Purpose**:

* To learn the proper procedure and technique for a titration with burets
* To determine the molarity of an acid using titration theory
* To determine the endpoint
* To better appreciate and understand God and His Creation

**Procedure:**

1. Arrange the titration apparatus as demonstrated.

2. Measure with a graduated cylinder **exactly** 10ml of HCl and put it in a flask. Record the exact volume of HCl used. Place the flask aside to use later.

3. Clean out your buret with water. Dispose of the water down the drain. Clean out your

buret with about 5 ml of NaOH as demonstrated in class.

**Dispose the NaOH in the waste NaOH container. DO NOT PUT INTO THE STOCK BOTTLE.**

4. Fill the buret, using the funnel with the NaOH. Make sure the stop-cock is closed before filling. Fill it just past the 0 ml mark. ***Note the concentration of the NaOH and record it.***

5. Drain a little NaOH out of the buret into an empty beaker, making sure there are no air bubbles in the buret and that the level is exactly at the 0 ml mark (bottom of meniscus).

6. Add 1-2 drops of phenolphthalein to the HCl flask.

7. Place the flask below the buret, so the tip is just inside the mouth of the flask.

8. Open the stop-cock slowly and add NaOH to the HCl a few drops at a time.

9. Swirl the flask after each addition of 4-5 drops of NaOH.

10. Add more NaOH in groups of 4-5 drops a time, until a slight pink is observed in the HCl solution. At this point, add NaOH, one drop at a time.

11. Swirl-drop, Swirl-drop - until a very faint but permanent pink color is found in the HCl.

12. Record the volume on the buret. Calculate the amount of NaOH used.

**13. Determine the concentration of the HCl.**

**14. Repeat the procedure for 3 good trials.**

15. Repeat the experiment again, now using bromthymol blue as the indicator.

**Lab 21 Primary Standards Lab**

**Purpose:**

* To understand how to make primary standards and test them
* To see the order and beauty of God’s creation

**Apparatus:**

* Tartaric or Ascorbic Acid
* Buret
* Erlenmeyer flask
* Phenolphthanlein

**Procedure:**

1. Prepare 100 mL of a standard acid (either Tartaric (150.09 g/mol) or Ascorbic Acid (176.13 g/mol)) which has a concentration of 0.10 M. Do this by calculating the number of moles necessary to make that volume of solution with that concentration. Then, knowing the acid’s molar mass, calculate the number of grams necessary of that acid to make the solution. **Show your teacher the calculation at this point and receive approval to continue.**  Dissolve the powder in 100.0 mL of distilled water.
2. Prepare 100 mL of a standard base (NaOH (40.0 g/mol)) which has a concentration of 0.15 M. Follow a similar procedure as in #1. **Show your teacher the calculation at this point and receive approval to continue.** Dissolve the powder in 100.0 mL of distilled water.
3. Fill the buret with the base. Save your extra base for future trials.
4. Take a 20.0 mL sample of your acid and place it in an Erlenmeyer flask.
5. Add phenolphthalein indicator to the Erlenmeyer flask with your acid.
6. Titrate to the phenolphthalein endpoint.
7. Repeat for three good titrations.

**Data/Results:**

1. Show calculations of how made standard acid and standard base.

2. Using the volume of base used and its theoretical concentration and the volume of acid used, determine the concentration of the acid. Compare the lab results’ concentration to the actual concentration you made.

**Discussion:**

1. Explain the concept of primary standards and comment on the accuracy of yours.

**Lab 22 Percent Acetic Acid in Vinegar**

**Purpose:**

* To learn to be stewardly in choices that one has to make
* To determine the percent acetic acid in vinegar
* To determine which vinegar is the better buy
* To practice and develop in titration techniques
* To see how Chemistry is a part of "everyday-life" experiences

**Apparatus**:

* various store-bought vinegars
* standard NaOH solution
* phenolphthalein
* buret and ring-stand
* Erlenmeyer flask

**Information:**

**Heinz Vinegar - $ 1.99 for 1.892 L of vinegar Meijer Brand - $ 1.79 for 3.78 L of vinegar**

**Procedure**:

1. Measure out approximately 10 ml of one of the vinegars. Make sure to record the exact volume of the vinegar. ***Assume the density of vinegar is 1 g/ml.*** Find the mass of the vinegar sample.

2. Fill the buret with the standardized NaOH solution. Prepare the buret for the titration.

3. Add 1 or 2 drops of phenolphthalein to the Erlenmeyer flask with the vinegar.

4. Begin proper titration techniques.

5. Proceed with the titration until you arrive at the endpoint. Record the amount of NaOH added to the vinegar to reach the phenolphthalein endpoint.

6. Repeat the procedure until you have three valid trials.

7. Do the calculations below for each trial and then find the average percent acetic acid in your vinegar sample.

8. Report your results on the board. Copy down the class data.

9. Repeat the procedure with the other vinegar.

**Data/Results:**

1. Determine which vinegar is the best buy. Provide and show all calculations necessary to do so.

To do this: Remember that you will want to compare the amount of acetic acid per dollar present in the Heinz container to the amount in the Save Rite container.

Calculate the grams of acetic acid that was in the sample. This can be calculated using regular stoichiometry.

Calculate the percent acetic acid in the vinegar sample:

grams of acetic acid

grams of vinegar X 100

NB: the grams of vinegar are calculated using the density of vinegar (see #1 in Procedure) and the volume of vinegar used (see #1 Procedure)

Use the percent acetic acid to determine how much actual acetic acid is present in the bottle of Heinz or Save Rite, and then, how much acetic acid one can purchase for $ 1.00.

**Discussion**:

Which vinegar is the best buy? Why?

**Lab 23 Antacid Analysis**

**Purpose:**

* To learn to be stewardly in choices that one has to make
* To determine which antacid is the best buy
* To practice and develop in titration techniques
* To see how Chemistry is a part of "everyday-life" experiences

**Apparatus**:

* various antacid tablets
* standard HCl solution
* standard NaOH solution
* buret and stand
* bromophenol blue indicator
* hot plate

**Information:**

**Tums – 1.37 grams per tablet; $ 1.89 for 36 tablets**

**Spartan – 1.32 grams per tablet; $ 3.09 for 150 tablets**

This lab requires us to perform a back titration because we are beginning with a basic tablet which is buffered. Therefore we will swamp the antacid with an excess amount of acid. This will neutralize the tablet but will make the ensuing solution to be acidic. Then we will need to titrate that solution with a basic solution to find the endpoint.

The tablet does not necessarily contain NaOH, but we will consider it to contain "NaOH equivalents".

The tablet is not 100 % "NaOH equivalents". It contains materials to bind the tablet together and other materials for taste. The "NaOH equivalents" will react with the acid. Some of the "binders" may dissolve, but do not be distressed if everything does not completely dissolve.

Bromphenol blue indicator turns blue in basic solutions and yellow in acidic solutions.

**Procedure**: **Part A:**

1. Crush an antacid tablet with a mortar and pestle.

2. Mass approximately 0.7 grams of the tablet. Record the exact mass.

Place it in an Erlenmeyer flask.

3. Add 25 ml of the HCl solution to the tablet in your Erlenmeyer flask.

Record the exact amount of HCl added and its concentration.

4. Swirl the contents and make sure it is mixed well.

5. It is possible that the solution has carbon dioxide dissolved in it. Heat the solution on the hot plate to its boiling point and boil it for at least 2 minutes to expel dissolved carbon dioxide. Allow the flask to cool.

6. Add 4 - 5 drops of bromphenol blue indicator solution to the solution in the Erlenmeyer flask. If the solution is still blue, add 15 ml more of HCl acid and boil it again. Continue to add HCl until the solution turns yellow. Remember to keep track of how much HCl you have added.

**Part B: Titration**

1. Clean the buret. Fill it with the Standardized NaOH solution. Record the NaOH's concentration.

2. Place the Erlenmeyer flask with the yellow solution below the buret.

3. Begin to titrate.

4. When the solution turns blue, you have arrived at the endpoint. Record the total volume of NaOH that was added to the solution.

5. Repeat Parts A and B again if time permits.

6. Clean up your lab station appropriately.

7. Put your information on the board and write down the class data.

**Data/Results:**

In this section show all the calculations necessary to determine which antacid tablet is the better buy. To help in these calculations consider these hints:

a.) Calculate the total number of moles of HCl that were added to the tablet.

b.) Remember that we “swamped” the tablet with excess HCl. Therefore if you subtract the number of moles of NaOH that were added to the “HCl / tablet” mixture from the total number of moles of HCl added, you should end up with the number of moles of HCl that actually reacted with the tablet.

c.) Determine the number of moles of “NaOH equivalents” in the tablet. This should be equal to the number of moles of HCl that actually reacted with the tablet.

d.) Determine which antacid gives more “NaOH equivalents” per dollar. This is the better buy.

**Discussion:**

Be sure to explain which antacid tablet is the better buy and why.