# CHEMISTRY 20 

## 2017 LAB <br> MANUAL

# Skill 1): AN INTRODUCTION TO PROCESS SKILLS 

There are six steps to the scientific inquiry model:

1) Questioning - Defining Problems (and variables)
2) Proposing Ideas - Background, Hypothesis and/or Prediction
3) Designing Experiments - Design, Variables, Pre-lab calculations \& questions
4) Observing and Measuring - Recording data
5) Processing Evidence - Analysis with charts, tables and graphs
6) Interpretation of Evidence - Evaluation, errors, applications, and new ideas

## QUESTIONING

VARIABLES: The first step in defining an appropriate, manageable problem is to identify the specific variables intended to be observed. "A variable is any part of an experiment that changes (varies) or could be changed. The three types of variables are:

1. manipulated variable - Change one variable in a systematic way. Ie) sunlight
2. responding variable - Corresponding changes in another variable. Ie) oxygen production in peas
3. controlled variables - Held constant (kept the same). Ie) size of plant, temperature, water

PROBLEM: Once the variables have been identified, the next step is to state the possible relationship between the variables or define the problem statement. A problem should be stated so:
a) It clearly describes what is to be solved
b) It can be solved in a reasonable amount of time.

Ie) $\boldsymbol{\rightarrow}$ Problem as a Question: "What effect does sunlight (manipulated) have on the production of oxygen in plants (responding)?"
Ie) $\boldsymbol{>}$ Problem as a Statement: "To determine the effect of sunlight (manipulated) on the production of oxygen in plants.(responding)."

## PROPOSING IDEAS

BACKGROUND INFORMATION: Background information is a necessary research tool that must precede any experimental work. The knowledge and information gained from this research will help the scientist keep on task and prevent them from duplicating previously established findings. Another important reason for researching a project is the information may be used to establish a hypothesis. Background information will be presented in question format or paragraph format.

HYPOTHESIS: A hypothesis is defined as a statement or idea which explains the relationship between variables based on existing knowledge or experience (background information). The hypotheses of others may be used because science is often a community effort. When applied to an investigation it must do three things:

1. Give a possible answer to the problem.
2. Give a reason or explanation for the answer.
3. It must be testable.

Ie) $\rightarrow$ I hypothesis that sunlight will cause oxygen to be produced in plants because plants photosynthesize in the presence of light. This will be tested by placing water plants in light and collecting the gas produced.

PREDICTIONS: While a hypothesis explains the relationship between variables, a prediction can be defined as a specific forecast as to possible observations or results of an experiment. No explanation of the prediction is required. Often predictions include numerical values.
Ie) $\rightarrow$ I predict that 12 hours of sunlight will produce 6 mol of oxygen in 1 test tube of water plants.

## DESIGNING EXPERIMENTS

INVESTIGATIVE DESIGN: Once a hypothesis and prediction have been made the next step is to devise a means to test the hypothesis and try to falsify the predictions. Often a brief overview is given -- a short explanation of what is to be done in the lab exercise. The experiment must have:

1. two parts for comparison:
a. the experimental or variable part testing a condition: (experimental group)
b. the check or control part which is kept constant: (control group) NOTE: The control group is used as a comparison for the experimental group.
2. only one variable - only one difference -between the experimental group and the control group.
3. have a large number of cases in both the experimental group and the control group.
4. be repeatable by any other person who wants to do so, and who has the necessary skills.
5. a conclusion that does not include a wider area than the experimental materials allow.

PRE - IAB CALCULATIONS: Any calculations required to do the lab procedure must be shown in this section. Examples include: mass of reactants, solutions preparation.

MATERIALS AND EQUIPMENT: All materials and equipment required to do the lab are listed in columns. Specifies are to be given when necessary. Ie) $\rightarrow \mathbf{5 0} \boldsymbol{m L}$ beaker \& $\mathbf{1 0 0} \mathbf{m L}$ graduated cylinder

PROCEDURE: State the procedure clearly and concisely following the principles listed below. This is a step by step account of how the experiment is to be completed.
1.Use complete sentences.
2.Choose appropriate intervals for the manipulated variables.
i.e. "How often will you measure the growth of your experimental plant:?"
3.Use an appropriate number of trials to give reasonable reliability.
4.State any precautions that should be taken by the experimenter.

## OBSERVING AND MEASURING DATA

OBSERVING ACCURATELY: Both qualitative (quality) and quantitative (numerical) observations must be made accurately. To insure this, the following criteria must be fulfilled:
a) Sampling and measuring instruments must be used with skill and precision.
b) The experimenter must be familiar with the procedures.
c) Proper laboratory techniques must be used to eliminate improper results from using dirty glassware, etc.
d) The experimenter must be able to estimate quantities of objects observed and distinguish between reasonable and unreasonable values.

RECORDING DATA: The recording of data must be clear and accurate. Methods of recording may vary including tables, diagrams, etc. Titles, labels, and units must be used.

Ie) $\boldsymbol{\rightarrow} \boldsymbol{D}$ Distance Traveled on September 11, 1993 from 8 a.m. to 4 p.m.

| TIME | TOTAL DISTANCE (km) |
| :---: | :---: |
| 8 am | 0 |
| 10 am | 23 |
| 12 p.m. | 42 |
| 2 p.m. | 57 |
| 4 p.m. | 68 |

## PROCESSING EVIDENCE

ANALYSIS: After the data is collected it must be processed before any interpretations or conclusions can be made. Data can be processed in a variety of ways including:
a) the mathematical calculations of data
b) the designing of charts and tables for processed data
c) the production of suitable graphs

NOTE: The Data and Analysis sections may be combined if suitable tables are designed.
Two types of graphs used in science are the line and the bar graphs. The four line graphs are:
exponential

sigmoid

curved


The following is a list of 10 rules to keep in mind when preparing a line and/or bar graph:

1. The title of the graph is written across the top of the graph paper. (NOTE: The title should indicate the relationship between the manipulated variable and the responding variable.)
2. On the horizontal axis place the name of the manipulated variable and its unit of measure.
3. On the vertical axis place the name of the responding variable and its unit of measure.
4. Calculate the values for the scales so the plotted data extends over most of the graph. (not over)
5. The scales need not be the same along both axis. In most cases the scales should start at zero.
6. Plot the coordinate points for the manipulated and responding variables on the graph.
7. The number pairs are plotted as a point. If the degree of error is known the point is circled. The circle represents your possible error in calculating the location of that dot. (line graph)
8. The points are joined by a smooth curve or line. (line graph) *Use a "line of beet fit" rather than joining the dots.
9. If you are plotting two or more variables, a legend should be placed in the lower right hand corner of the graph. The legend should identify each of the lines on the graph. NOTE: All work must be printed and in pencil.

## INTERPRETATION OF EVIDENCE

EVALUATION: In this section the experimenter evaluates if the analyzed data proves or disproves the hypothesis and verifies or falsified the predictions. The evaluation of the data and experiment should be written in paragraph form. In the paragraph the following questions should be answered:

1. Did the analysis of the data support the hypothesis and predictions? Prove with data.
2. What conclusions can be drawn from the data?
3. Make an evaluation of the experimental design. If the design is inadequate, what changes should be made to correct it?
4. Are there any sources of error in the evidence collected arising from limitation of the measuring instruments, limitations of the experimental design or human error?
5. What new questions arose from this experiment?

Cite all references alphabetically in scientific format. Refer to the following examples.
Article: Mohr, H. 1962. Primary effect of light on growth. Ann Rev. Plant Physiol. 13:465-88.
[Note that capitalization of the title of the article should be that used in a sentence.]
Book: Kramer, P.J., and Kozlowski, T.T. 1960. Physiology of Trees, p. 11. New York: McGraw-Hill. Internet:

## Process Skills Assignment

Name:

## Due Date:

## Score:

Go through the six scientific process steps to investigate how the amount of glucose burned affected the amount of gases produced.

PROBLEM: (2 marks)
VARIABLES:
MANIPULATED (1 mark):
RESPONDING: (1 mark):
2 CONTROLLED: (2 marks):
BACKGROUND INFORMATION: Give five piece of background information relevant to this study. (5 marks) HINT: An understanding of science 10 chemistry is beneficial.
-
-
-
-
-
HYPOTHESIS: (3 marks):

PREDICTION: Predict the types of gases produced and the number of moles produced if 1 mol of glucose was burned.(2 marks):

OVERVIEW: (2 marks):

MATERIALS: (2 marks):

PROCEDURE: (2 marks):

DATA \& OBSERVATION: Make a data table that would indicate how many moles of gas would be produced if $1 \mathrm{~mol}, 2 \mathrm{~mol}$ and 3 mol of glucose was burned. You can use data that either verifies or contradicts you hypothesis and predictions. (1 mark)

ANALYSIS: Graph the data. (3 marks)

CONCLUSION: Based on the data above, write a complete conclusion. (4 marks)

## Skill 2）：Significant Digits \＆Unit Conversions <br> \section*{Due Date：} <br> Score：

Name：

## Definition of Significant digits

Significant digits indicate how accurate a measurement is．Significant digits are the digits that are certain plus one uncertain digit （the last digit）．Significant digits are NOT defined as important digits．

## Counting Significant Digits

When counting significant digits，count all the digits from 1 to 9 plus zeroes in between and zeroes following these digits．DO NOT count zeroes in front of a 1 to 9 because they only serve to set the decimal place．Constants and exact numbers have infinite number of significant digits．
ie） 0.02050 kg The two zeros in front are NOT significant．This number has 4 significant digits．
仓仑けれれஈ $\uparrow=$ significant digit $\hat{\cup}=$ not significant digit

## Rounding off when using significant digits

When the next digit（after those that are kept as significant）is less than 5 ，all the digits remain the same．When the next digit is 5 or greater，the last digit that is kept is increased by one．
Ie．） 19.95 m with 3 significant digits would be rounded off to 20.0 m .129 .49 g with 3 significant digits would be not be rounded off and remain 129 g

## Scientific notation

Scientific notation is the method of expressing values as a number between 1 and 10 multiplied by a power of ten．（\＃．\＃\＃x $10^{\#}$ ）
Scientific notation is used for very large numbers or very small numbers with a few significant digits．
Ie） 1490 m with 2 significant digits would be expressed as $1.5 \times 10^{3} \mathrm{~m}$ The decimal moved 3 places to the left
0.0015678 g with 1 significant digit would be expressed as $2 \times 10^{-3} \mathrm{~g}$ The decimal moved 3 places to the right．

NOTE：There is always only one digit（other than 0 ）and then the decimal when using scientific notation．The digits in $10^{\#}$ are not significant．

## SI（System International）Prefixes \＆Unit Conversions

SI prefixes are often used to replace the power of ten in scientific notation．Here are the most common prefixes．These and other prefixes are also located in your databook on page 1 ．

$$
\begin{aligned}
& \text { Giga }(G)=10^{9} \\
& \operatorname{Mega}(M)=10^{6} \\
& \operatorname{Kilo}(k)=10^{3}
\end{aligned}
$$

$$
\text { centi }(c)=10^{-2}
$$

$$
\operatorname{milli}(\mathrm{m})=10^{-3}
$$

$$
\operatorname{micro}(u)=10^{-6}
$$

Scientists need to be able to convert from one prefix to another．
Ie） $1.5 \times 103 \mathrm{~m} \rightarrow 1.5 \mathrm{~km} \quad 2 \times 10^{-3} \mathrm{~g} \rightarrow 2 \mathrm{mg}$

## Addition \＆Subtraction significant digit rules

Add／subtract and then round off the answer to the least number of decimal places contained in the question．
Ie） $26.5 \mathrm{~m}+7.01 \mathrm{~m}=33.51 \mathrm{~m}$ Rounded $=33.5$
（ 1 dec. ）（ 2 dec ．）
（1 dec．）

## Multiplication \＆Division significant digit rules

Multiply／divide and then round off the answer to the least number of total significant digits contained in the question．Decimal places are NOT considered for significant digits when you multiply or divide．
Ie） $100 \mathrm{~s} \mathrm{x} 5.0 \mathrm{~m} / \mathrm{s}=500 \mathrm{~m} \quad$ Rounded $=5.0 \times 10^{2} \mathrm{~m}$

1. Identify how many significant digits are in each of the following measurments:

Ie) $\mathbf{0 . 0 0 5 0 6 0} 4$ significant digits (zeros in front of the 5 are not significant)
a. 15.8 g $\qquad$
c. $1.50 \mathrm{~km} / \mathrm{h}$ $\qquad$
e. $0.0061 \mathrm{~mol} / \mathrm{L}$ $\qquad$
g. 1200 cm $\qquad$
i. 0.08 hectares $\qquad$
2. Perform the following calculations.

Ie. $35.7 \mathrm{~mol} \times 168.92 \mathrm{~g} / \mathrm{mol}=$
a. $\quad \mathbf{1 6 . 7 5} \mathrm{s} \mathrm{x} 85 \mathrm{~m} / \mathrm{s}$
b. $\quad 0.00085 \mathrm{~L} \times 1.3111 \mathrm{~g} / \mathrm{L}$
c. $\mathbf{0 . 0 0 0 1 1 8} \mathbf{~ m o l ~ x ~} \mathbf{1 8 . 0 2} \mathbf{g} / \mathbf{m o l}$
d. $0.12 \times 10^{6} \mathbf{~ m o l ~} \times 22.4 \mathrm{~L} / \mathbf{m o l}$
e. $\quad \mathbf{0 . 1 7 8} \mathbf{g} / \mathbf{1 2 . 0 1} \mathbf{g} / \mathbf{m o l}$
f. $\quad 0.1456 \mathbf{m o l} / 2.3 \mathrm{~L}$
g. $\quad \mathbf{4 5 2 . 6 5} \mathrm{g} / \mathbf{5 8 . 0 6} \mathrm{g} / \mathrm{mol}$
h. $1.12 \times 10^{-5} \mathbf{m o l} / 2.5 \mathrm{~mol} / \mathrm{L}$
i. $\quad 1.28 \times 10^{6} \mathrm{~g} \times 3.33 \times 10^{3} \mathrm{~J} / \mathrm{g}$
j. $\quad 0.0088 \mathbf{m o l} / 179 \mathrm{~L}$
k. $\quad \mathbf{7 6 0} \mathbf{m}+\mathbf{4 2 . 6} \mathbf{m}$

1. $\mathbf{9 . 9 9} \mathbf{~ m o l}+\mathbf{1 5 1 0 . 9} \mathbf{~ m o l}$
m. $\mathbf{1 4 . 7 6} \mathrm{mL}-4 \mathrm{~mL}$
n. $\mathbf{1 2 9} \mathbf{g}-\mathbf{2 9 . 5} \mathrm{g}$
b. $0.167 \mathrm{~m} / \mathrm{s}$ $\qquad$
d. $23.005 \mathrm{~g} / \mathrm{L}$ $\qquad$
f. $1.54 \times 10^{6} \mathrm{~km}$ $\qquad$
h. $5.00 \times 10^{-3} \mathrm{t}$ $\qquad$
j. 14.03 C $\qquad$

|  | 35.7 mol $\times 168.92 \mathrm{~g} / \mathrm{mol}=$ |  | Unrounded $6030.44 \mathrm{~g}$ |
| :---: | :---: | :---: | :---: |
| a. | 16.75 s x $85 \mathrm{~m} / \mathrm{s}$ | = |  |
| b. | $0.00085 \mathrm{~L} \times 1.3111 \mathrm{~g} / \mathrm{L}$ | = |  |
| c. | $0.000118 \mathrm{~mol} \mathrm{x} 18.02 \mathrm{~g} / \mathrm{mol}$ | = |  |
| d. | $0.12 \times 10^{6} \mathrm{~mol} \times 22.4 \mathrm{~L} / \mathrm{mol}$ | = |  |
| e. | $0.178 \mathrm{~g} / 12.01 \mathrm{~g} / \mathrm{mol}$ | $=$ |  |
| f. | $0.1456 \mathrm{~mol} / 2.3 \mathrm{~L}$ | $=$ |  |
| g. | $452.65 \mathrm{~g} / 58.06 \mathrm{~g} / \mathrm{mol}$ | = |  |
| h. | $1.12 \times 10^{-5} \mathrm{~mol} / 2.5 \mathrm{~mol} / \mathrm{L}$ | $=$ |  |
| i. | $1.28 \times 10^{6} \mathrm{~g} \mathrm{x} 3.33 \times 10^{3} \mathrm{~J} / \mathrm{g}$ | = |  |
| j. | $0.0088 \mathrm{~mol} / 179 \mathrm{~L}$ | $=$ |  |
| k. | $760 \mathrm{~m}+42.6 \mathrm{~m}$ | $=$ |  |
| 1. | $9.99 \mathrm{~mol}+1510.9 \mathrm{~mol}$ | $=$ |  |
|  | $14.76 \mathrm{~mL}-4 \mathrm{~mL}$ | $=$ |  |
|  | 129 g-29.5g | $=$ |  |


| Rounded | SI Prefix |
| :--- | :--- |
| $\mathbf{6 . 0 3 \times 1 0 ^ { 3 }} \mathbf{g}$ | $\mathbf{6 . 0 3} \mathbf{~ k g}$ |

6.03 kg
$\qquad$
$\qquad$
$\qquad$
$\qquad$
$\qquad$
$\qquad$

XXXXXXXXXXXX
$\qquad$
$\qquad$
$\qquad$
$\qquad$
$\qquad$
$\qquad$


## Skill 3) WHMIS Symbols and Safety

1. Please complete the following table. The first one is done for you.

| Class A |  |  |  |  |  |
| :---: | :---: | :---: | :---: | :---: | :---: |
| NAME: <br> Compressed <br> Gas |  |  |  |  |  |
| DANGGER: <br> May explode <br> when near <br> heat or <br> dropped |  |  |  |  |  |
| EXAMPLE: <br> acetylene |  |  |  |  |  |

2. What are the lab rules and guidelines?
3. Complete the following table on lab equipment

| Picture | Names | Purposes |
| :---: | :---: | :---: |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

## Chemistry 20 Lab 1.1: Conservation of Mass

Name: $\qquad$ Due Date: Score: $\qquad$
Pre-lab Discussion: Matter cannot be created of destroyed by a chemical change. This very important principle is known as the law of conservation of mass. This law applies to ordinary chemical reactions (as opposed to nuclear reactions, in which matter can be changed to energy). During a chemical change (reaction), the atoms of one or more substances (reactants) simply undergo some "rearrangements". The result off these rearrangements is the formation of new, different substances (products). All of the original atoms are still present. It is because of the law of conservation of mass that we are able to write balanced chemical equations. Such equations make it possible to predict the masses of reactants and products that will be involved in a chemical reaction.

In this experiment, aqueous solutions of three different compounds will be used to produce two separate and distinct chemical reactions. The fact that change occurs during each reaction will be readily observable. Balanced chemical equations for the two reactions are:

$$
\begin{aligned}
& \mathrm{Na}_{2} \mathrm{CO}_{3}(\mathrm{aq})+\mathrm{CaCl}_{2}(\mathrm{aq}) \rightarrow 2 \mathrm{NaCl}(\mathrm{aq})+\mathrm{CaCO}_{3}(\mathrm{aq})(\mathrm{Eq} .1) \\
& \mathrm{CaCO}_{3}(\mathrm{aq})+\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq}) \rightarrow \mathrm{CaSO}_{4}(\mathrm{~s})+\mathrm{H}_{2} \mathrm{O}+\mathrm{CO}_{2}(\mathrm{~g})(\mathrm{Eq} .2)
\end{aligned}
$$

The combined masses of the three solutions (and their containers) will be measured before and after each reaction has occurred. This experiment should give you a better understanding of the law of conservation of mass and its importance in chemistry.

Purpose: To recognize the law of conservation of mass and to learn the careful use of a laboratory scale in determining the masses of reactants and products.

## Equipment:

- Laboratory scale (balance)
- Labels
- 10 mL graduated cylinder
- 13 - 100 mm test tubes with corks
- 125 mL Erlenmeyer flask with rubber stopper
- lab coat and safety goggles

Materials: $1.00 \mathrm{~mol} / \mathrm{L}$ solutions of: $\mathrm{NaCO}_{3}(\mathrm{aq}), \mathrm{CaCl}_{2}(\mathrm{aq})$ and $\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq})$
Safety: Handle sulphuric acid with extra caution. Always wear safety goggles when handling acids. Report all acid spills to your teacher. Flush acid spills on your hands with cold water. Dilute acid spills with a solution of sodium hydrogen carbonate $\left(\mathrm{NaHCO}_{3}(\mathrm{aq})\right)$. Wear a lab coat and goggles at all times when working in the lab.

## Procedure:

1. In a graduated cylinder, measure exactly 10.0 mL of sodium carbonate solution $\left(\mathrm{NaCO}_{3}(\mathrm{aq})\right)$. Pour this into a clean, dry 125 mL Erlenmeyer flask and stopper the flask. Rinse and dry the graduated cylinder.
2. Measure exactly 3.0 mL of $1.00 \mathrm{~mol} / \mathrm{L}$ calcium chloride solution $\left(\mathrm{CaCl}_{2}(\mathrm{aq})\right)$ and pour into a clean, dry test tube. Cork and label the test tube. Rinse and dry the graduated cylinder.
3. Repeat step 2 with 3.0 mL of $1.00 \mathrm{~mol} / \mathrm{L}$ sulphuric acid solution $\left(\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq})\right)$. CAUTION: Handle this acid with care.
4. Place the stoppered flask and the corked test tubes together on the pan of the laboratory scale. Tilt the test tubes to prevent liquids from touching the corks. Measure the combined mass of these containers, stoppers and solutions. This will be Mass A in your data table.
5. Remove the flask and the test tube containing the $\mathrm{CaCl}_{2}(\mathrm{aq})$ solution from the laboratory scale. Pour the $\mathrm{CaCl}_{2}(\mathrm{aq})$ solution into the $\mathrm{Na}_{2} \mathrm{CO}_{3}(\mathrm{aq})$ solution in the flask. Swirl the flask to thoroughly mix the two solutions. Record your observations.
6. Replace the stopper and cork in their proper containers. Once again, measure the combined mass of the three containers, stoppers and contents. (Mass B in the data table.)
7. Remove the flask and the test tube containing $\mathrm{H}_{2} \mathrm{SO}_{4}$ from the laboratory scale. Carefully pour the $\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq})$ solution into the flask. WITH THE STOPPER OFF, swirl the flask until all bubbling stops. Record your observations. Allow the flask to cool to room temperature.
8. Replace the stopper and cork. Measure again the mass of the three containers, stoppers and contents. (Mass C in data table)

## Observations \& Data Table: (5 marks)

Record your observations below:
Step 5: $\qquad$
Step 7: $\qquad$

## Data Table: Masses of Chemicals

Mass A (before mixing): g

Mass B (after mixing $\mathrm{Na}_{2} \mathrm{CO}_{3}(\mathrm{aq}) \& \mathrm{CaCl}_{2}(\mathrm{aq})$ ) g

Mass C (after adding $\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq})$ to the mixture) g

## Analysis Questions: (8 marks)

1. What evidences of a chemical reaction can be observed in step 5 ? (1 mark)
2. What evidences of a chemical reaction can be observed in step 7 ? (1 mark)
3. Why were you instructed to leave the flask open after the $\mathrm{H}_{2} \mathrm{SO}_{4}(\mathrm{aq})$ solution was added? (What might have happened if this was not done?) (1 mark)
4. Compare mass A, mass B, and mass C. Account for the difference. (Why are the masses different?) (2 marks)
5. In your opinion does this experiment verify the law of conservation of mass? (1 mark)
6. How might the experiment be improved to bring its results in line with the law of conservation of mass? (1 mark)
7. When you burn a $\log$ in a fireplace the resulting ashes have a mass less than that of the original log. Account for the difference. (Why are the masses different?) (1 mark)

## Chemistry 20 Lab 1.2: Studying Chemical Change

## Date

 Name $\qquad$ Score: $\qquad$PURPOSE:
To identify and examine examples of five types of chemical reactions.
To review balancing techniques.

## BACKGROUND INFORMATION:

a. There are six types of chemical reactions listed below. Identify the first 5 reactions:

1. $\qquad$ - element + element ---> compound

$$
(\mathrm{A}+\mathrm{B}--->\mathrm{AB})
$$

2. $\qquad$ - compound ----> element + element

$$
(\mathrm{AB}-->\mathrm{A}+\mathrm{B})
$$

3. $\qquad$ - element + compound -----> element + compound

$$
\left.\overline{(\mathrm{A}}+\mathrm{BC}_{(\mathrm{aq})}--->\mathrm{B}+\mathrm{AC}_{(\mathrm{aq})}\right)
$$

4. $\qquad$ - compound + compound $--->$ compound + compound

$$
\left(\mathrm{AB}_{(\mathrm{aq})}+\mathrm{CD}_{(\mathrm{aq})} \cdots \mathrm{CB}_{(\mathrm{aq})}+\mathrm{AD}_{(\mathrm{aq})}\right)
$$

5. $\qquad$ - hydrocarbon + oxygen ---> carbon dioxide + water $\left(\mathrm{C}_{\mathrm{x}} \mathrm{H}_{\mathrm{y}}+\mathrm{O}_{2(\mathrm{~g})}-->\mathrm{CO}_{2(\mathrm{~g})}+\mathrm{H}_{2} \mathrm{O}_{(\mathrm{l})}\right.$
6. Other (O)
b. There are five evidences of a chemical reaction, which are:
7. $\qquad$
8. $\qquad$
9. $\qquad$
10. $\qquad$
11. $\qquad$
c. Is making tea a chemical reaction? Why or why not?
d. Is boiling water a chemical reaction? Why or why not?

## PROCEDURE:

1. Try each of the reactions below as instructed. Balance the chemical reaction by placing a coefficient in front of the elements or compounds. Add the states of matter.
2. Identify the reaction type. Indicate if the reaction is endothermic or exothermic
3. Record all the evidences of a chemical change.
4. Answer the questions that follow.

Reaction \#1: Add water to the Hofmans apperatus. Put electricity through the water.
a) ___ $\left.\mathrm{H}_{2} \mathrm{O}_{( }\right)+$electricity $\rightarrow$ (Have you balanced? states?)
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences:
$\qquad$
d) How do you know which gas is hydrogen? $\qquad$

Reaction \#2: CAUTION: Acids are corrosive. Place a small piece of magnesium into a flask of hydrochloric acid. Place a balloon on top of the flask for another reaction.
a) $\qquad$ $\mathrm{Mg}_{( }{ }^{+}$ $\qquad$ $\mathrm{HCl}_{(\mathrm{aq})}$---> $\qquad$ (Have you balanced? states?)
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$
d) What is the WHMIS symbol that will appear on the acid? $\qquad$
Reaction \#3: CAUTION: Methane is explosive; WEAR safety goggles and rubber gloves for this reaction. Attach a bunsen burner to the outlet on your lab bench. Turn on the methane gas outlet and light the bunsen burner with a match. Adjust the flame for the next reaction.
a) __ $\left.\mathrm{CH}_{4()}+\ldots \mathrm{O}_{2( }\right) \rightarrow$ (HINT: Balance the $\mathrm{O}_{2}$ last.)
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$
What is the WHMIS symbol that would appear on a bottle of methane gas?

Reaction \#4: CAUTION: Hydrogen is explosive; WEAR safety goggles and rubber gloves for this reaction. Tie off the balloon from reaction \#2. Put the balloons into some tongs. CAREFULLY place the balloon into the bunsen burner from reaction \#3.
a) $\qquad$ $\left.\mathrm{H}_{2( }\right)+$ $\qquad$ $\mathrm{O}_{2( }$ ) $\rightarrow$
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$
d) What is the WHMIS symbol that would appear on a bottle of hydrogen gas?

Reaction \#5: CAUTION: Long term exposure to lead is poisonous; dispose of chemicals in the waste beaker at the front of the class. Place 5 drops of lead (II) nitrate onto a spot plate. Add 5 drops of potassium iodide.
a) __ $\left.\left.\mathrm{Pb}\left(\mathrm{NO}_{3}\right)_{2( }\right)+\ldots \mathrm{KI}_{( }\right) \quad{ }^{--->}$
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences:
d) What is the WHMIS symbol that would appear on a bottle of lead (II) nitrate?
e) What is the name and color of the precipitate? $\qquad$
Reaction \#6: Burn the candle. Place a cold beaker of water over the flame and observe what forms on the outside of the beaker.
a) $\qquad$ $\left.\mathrm{C}_{25} \mathrm{H}_{52( }\right)+$ $\qquad$ $\mathrm{O}_{2( }$ ) --->
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$

OPTIONAL Reaction \#7: Place 10 mL of silver nitrate into a beaker and add some copper. Wait a few minutes for the reaction. Indicate the colors of each substance. Dispose of waste in the waste container.
a) $\qquad$ $\left.\mathrm{Cu}_{( }\right)+$ $\mathrm{AgNO}_{3}($ ) $--->$ $\qquad$ $+$ $\qquad$ colors: bronze
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$
d) Why is the formula for copper (II) nitrate not $\mathrm{CuNO}_{3}$ ? $\qquad$
CHALLENGE Reaction \#8: CAUTION: Acids and bases are corrosive. Sometimes it is difficult to see evidences of a reaction- so indicators (chemicals that change color) are used. One such reaction is a neutralization reaction. Take a small beaker with about 10 mL of hydrochloric acid and another small beaker with about 10 mL of sodium hydroxide. Add a few drops of bromothymol blue to each beaker and record the color under each compound below. Using a dropper, slowly add the sodium hydroxide to the hydrochloric acid. Swirl to mix each time you add. Stop adding sodium hydroxide when the hydrochloric acid solution changes color. (If you do it perfectly, you should get a green color)
a) __ $\left.\left.\mathrm{HCl}_{( }\right)+{ }_{-} \mathrm{NaOH}_{( }\right)$---> $\qquad$ (with HBb indicator) colors:
b) Reaction type: $\qquad$ Exothermic or endothermic.
c) Evidences: $\qquad$
d) Why is this called a neutralization reaction?

## Chemistry 20 Lab 2.1: Mole/Particle Lab

Name: $\qquad$ Due Date:

Score: $\qquad$
Purpose: To find the mass, moles and number of particles of a common compounds based on calculated masses.

Materials: laboratory scale, sucrose, $\mathrm{NaCl}, \mathrm{CaCO}_{3}$, potassium permanganate, H 2 O , water, methanol, zinc powder, suphur, NaOH , iron fillings.

## Variables:

Manipulated: $\qquad$
Responding: $\qquad$
3 Controlled: $\qquad$ , $\qquad$ ,

## Procedure:

1. Place a paper or glass container onto the laboratory scale and zero the scale.
2. Measure out 1 level spoonful of the required substance.
3. Place onto a laboratory scale and record the mass of the sample. Never put chemicals directly on the pan of the laboratory scale.
4. Complete the data table.

Observations and Interpretations ( 35 marks):

| FORMULA | NAME | Mass (g) | Molar mass <br> (g/mol) | Total Moles <br> (mol) | Moles of <br> each <br> element | atoms of <br> each <br> element |
| :--- | :--- | :--- | :--- | :--- | :--- | :--- |
|  | sucrose |  |  |  |  |  |
| $\mathrm{NaCl}(\mathrm{s})$ |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
| $\mathrm{CaCO}_{3}(\mathrm{~s})$ |  |  |  |  |  |  |
|  |  |  |  |  |  |  |


| FORMULA | NAME | Mass (g) | Molar mass <br> (g/mol) | Total Moles <br> (mol) | Moles of <br> each <br> element | atoms of <br> each <br> element |
| :--- | :--- | :--- | :--- | :--- | :--- | :--- |
|  | potassium <br> permanganate |  |  |  |  |  |
| $\mathrm{H}_{2} \mathrm{O}(1)$ |  |  |  |  |  |  |
| $\mathrm{CH}_{3} \mathrm{OH}(1)$ |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |
|  |  |  |  |  |  |  |

## Chemistry 20 Lab 2.2: Stoichiometry: Mole Ratios in a Chemical Reaction

Name: Due Date: Score:

Background: Chemists can use a balanced equation to determine the quantities of any chemical in the equation. The use of these numerical relationships is called stoichiometry. The large numbers in front of a balanced equation are called coefficients. These numbers represent the number of moles that each chemical of the reaction has. A mole ratio can be established between any two of these numbers.

In this experiment sodium hydrogen carbonate will react with hydrochloric acid and produce sodium chloride, carbon dioxide and water. From the balanced equation, the mole ratio of sodium hydrogen carbonate and sodium chloride can be determined and then compared to the actual mole ratio. The two ratios will be slightly different because of errors, such as an incomplete reaction or contaminants. The percent error formula is:

$$
\% \text { error }=\frac{[\text { actual }- \text { theoretical }] \times 100}{\text { theoretical }}
$$

1. Write out a balanced equation for the reaction occurring in this lab.
2. What is the theoretical mole ratio between sodium hydrogen carbonate and sodium chloride?
3. Define Stoichiometry.

Purpose: To compare the theoretical mole ratio of a chemical reaction with the actual (experimental) mole ratio.

## Materials:

sodium hydrogen carbonate
10 mL volumetric pipet with bulb
filter and filter paper
scale
$3.0 \mathrm{~mol} / \mathrm{L}$ hydrochloric acid
watch glass \& evaporating dish
150 mL beaker
Optional: burner with stand \& tongs

CAUTION: Wear lab coats. Notify your instructor of any spills. Wash acid off under running water for three minutes. The flame of the Bunsen burner should be regulated to prevent splattering.

## Procedure:

1. Using a scale, measure the mass of the watch glass \& evaporating dish. Record the mass in the table below.
2. With the watch glass/evaporating dish still on the scale, measure about two grams of sodium hydrogen carbonate. Record the exact mass in the table below.
3. Practice using the pipet and bulb with some tape water. When you feel comfortable, add 10 mL of 3.0 $\mathrm{mol} / \mathrm{L}$ hydrochloric acid to the watch glass, using the volumetric pipet.
4. After the reaction is complete, heat the contents of the watch glass under a Bunsen burner or let the liquid evaporate in a drying oven for a few nights.
5. Measure the mass of the watch glass/evaporating dish and the dry solid that remains. Record the mass in the table below. Flush the residue down the drain.

## Data \& Observations:

Day 1:

Day 2:

TITLE:

| Mass of watch glass or evaporating dish |  |
| :--- | :--- |
| Mass of sodium hydrogen carbonate |  |
| Mass of watch glass or evaporating dish and solid <br> after heating or drying |  |

CALCULATIONS TABLE (show any appropriate calculations)

| Moles of sodium hydrogen carbonate |  |
| :--- | :--- |
| Mass of sodium chloride |  |
| Moles of sodium chloride |  |
| Mole ratio (experimental) |  |
| Mole ratio (theoretical) |  |

## Analysis:

1. What was the gas that was produced during the reaction?
2. What was the liquid that was evaporated off?
3. What was the solid that remained?
4. Is the theoretical mole ratio close to the experimental mole ratio? Why or why not?

## Conclusion:

# Chemistry 20 Lab 2.3: Stoichiometry - Mass-Mass Relationships 

Name:
Due Date:
Score:
Background: If chemists know the amount of one chemical in a reaction, then they can use this to determine, theoretically, the amount of another chemical in that reaction.

In this experiment a known mass of zinc nitrate will be mixed with a solution of sodium hydrogen phosphate. Then a precipitate will be collected and its mass will be measured and compared to what theoretically it should have been.

1. Describe what a precipitate is.
2. Identify the type of reaction that will be used in this lab
3. Using stoichiometry, determine the theoretical mass of the precipitate that should be collected. Use 1.50 g of zinc nitrate as your known quantity. Show all your work.

## Problem:

## Variables:

Manipulated:
Responding:
Controlled (2):

## Materials:

1.50 g of zinc nitrate $\quad 2.00 \mathrm{~g}$ of sodium hydrogen phosphate

2-150 mL beaker
100 mL graduated cylinder
filter paper, iron ring, stand \& funnel
stirring rod with a rubber policeman electronic scale
distilled water
Procedure:

1. Using a scale, measure 1.50 g of zinc nitrate and place it into a 150 mL beaker . Record the exact mass in the table below.
2. With the graduated cylinder, add 50 mL of water to the zinc nitrate and stir with a stirring rod until all the zinc nitrate dissolves.
3. Using a scale, measure 2.0 g of sodium hydrogen phosphate and place it into a 150 mL beaker. Record the exact mass in the table below.
4. With the graduated cylinder, add 50 mL of water to the sodium hydrogen phosphate and stir with a stirring rod until all the sodium hydrogen phosphate is dissolved.
5. Pour the zinc nitrate solution into the beaker with the sodium phosphate solution. Record your observations below.
6. Determine the mass of the filter paper. Fold the filter paper and place it into a funnel, wetting it with a few drops of distilled water.
7. Filter the solution into a 150 mL beaker and collect the precipitate in the filter paper. Use a rubber policeman and distilled water to remove all the precipitate from the beaker. Be patient and do not overfill the filter paper.
8. When the filtration is complete, remove the filter paper, opening it to promote drying, and place it on a labeled paper towel. Allow the precipitate to dry overnight in a drying oven or in the fumehood.
9. Determine the mass of the filter paper and precipitate. Determine the mass of the precipitate by subtracting the mass of the filter paper.
10. Dispose of the filter paper into the garbage and clean up your station.

## Data \& Observations: <br> DATA TABLE

## Analysis:

1. What was the precipitate that was produced during the reaction?
2. Determine the number of moles of precipitate collected. Show your work.

Conclusion: (4 parts)

## Chemistry 20 Lab 2.4: Stoichiometry: Limiting Reagents \& Percent Yield Due Date: Score:

Name:
Background: The amount of product collected in a chemical reaction depends on the moles of the limiting reactant used. For industrial processes, using the correct amounts of reactants is crucial to maintaining a profitable business. For example, concentrated sulphuric acid is one of the key reactants in the production of laundry detergent. Using the wrong amount of acid could result in detergent that could damage your laundry.

In this experiment, you will combine masses of reactants to determine which reactant controls the mass of product collected.

1. Write a balanced chemical equation for the reaction of a solution of copper (II) sulphate with iron (steel wool). Identify the type of reaction this is. (2 marks)
2. 2.00 g of solid copper (II) sulphate pentahydrate is mixed with water to make a copper (II) sulphate solution. The copper (II) sulphate is then mixed with 0.30 g of iron (steel wool). Calculate the theoretical mass of copper that should be produced. (5 marks)
3. Solutions with copper ions are blue in color. Iron is grey while copper is bronze colored. Predict what will happen as the reaction occur. (2 marks)

## Problem:

## Variables:

Manipulated: $\qquad$
Responding:
Controlled (2):
Materials:

CAUTION: Copper (II) sulphate is an irritant as well as a poison. Avoid skin contact. Keep your hands away from your face. Wash your hands at the end of the experiment. Wear a lab coat. Remove contacts.

## Procedure:

1. Using a spatula and a weight scale, place 2.00 g of copper (II) sulphate pentahydrate into 15 mL of distilled water in a 250 mL beaker. Record the mass of copper (II) sulfate pentahydrate used in the table below.
2. Measure 0.30 g of steel wool. Break up the steel wool into small pieces and add it to the copper (II) sulphate solution. Record the mass of steel wool used in the table below.
3. With a stirring rod, stir and gently crush the steel wool against the sides of the beaker until the reaction is complete
4. Determine the mass of the filter paper and record in the table below.
5. Fold and place the filter paper into a funnel. Moisten the filter paper with distilled water so it fits snugly into the funnel. Use a stand and clamp to hold the funnel.
6. Decant (pour off) any excess solution from the beaker. Guide the flow of liquid with a glass rod and make sure that none of the solid is lost.
7. Filter the remaining mixture. Use a rubber policeman and wash bottle to remove and solid residue that remains.
8. After filtering, remove the filter paper from the funnel, place the filter paper onto a labeled paper towel and dry in the fumehood overnight.
9. The next day, determine the mass of the residue. Dispose of the filter paper and contents into the garbage. Wash solutions down the sink.
10. Calculate the moles of residue, percent error and percent yield. Show your work here.

## Data \& Calculations:

## DATA TABLE

## Analysis:

1. What reactant do the observations indicate is limiting and excess? Explain why? HINT: look at the color changes.
2. What is the actual mole ratio between the limiting reagent and copper produced? Compare this to the mole ratio found in the chemical equation. (Write both ratios down)

## Extension:

The burning of sulphur-containing coal releases sulphur dioxide into the air, causing acid rain. Sulphur dioxide can be removed from smoke stack emissions by passing it over "slaked lime" or wet calcium hydroxide.
Calcium sulphite and water are produced. What mass of calcium hydroxide is required to remove $6.96 \times 10^{11} \mathrm{~g}$ of sulphur dioxide?

Conclusion: (4 parts)

## Chemistry 20 Lab 3.1: Preparing and Diluting Solutions

Name: Due Date: $\qquad$ Score: $\qquad$
Purpose:
PART A: To watch and practice correct pipetting techniques.
PART B: To accurately prepare a solution of known concentration.
PART C: To subsequently dilute the solution prepared in Part A.
Background: : There are two main ways that chemists use to prepare solution of a particular concentration. One method is to obtain a given mass of the solid chemical and then add a predetermined amount of water to the solid. This method involves the following 4 steps:

1) Calculate the number of moles of solution needed with the formula, $\mathrm{n}=\mathrm{CV}$
2) Calculate the mass of solid needed with the formula, $m=n x M$. Use the Molar mass of the chemical formula in its solid form, including hydrates. Measure the mass with a scale.
3) Mix the mass in a beaker with about $1 / 2$ the water.
4) Place the mass into a volumetric flask of a desired volume (e.g. 100 mL ). Add water to the meniscus line, using an eye dropper to accurately fill the flask. Cap the flask, invert and mix until all the solid is dissolved.

Another method is to dilute a solution that is already prepared. This method involves removing a portion of the solution and adding water to it. This method involves 4 steps:

1) Calculate the volume of the original solution needed with the formula, $\mathrm{V}_{1}=\mathrm{C}_{2} \mathrm{~V}_{2} / \mathrm{C}_{1}$
2) Obtain the volume using a volumetric pipet(measure one specific volume) or graduated pipets(measure incremental volumes) Pipets measure a small volume ( 25 mL or less) to high precision ( 0.1 mL to 0.01 mL )
3) Place the solution in the pipet into a volumetric flask of a desired volume. Add water to the meniscus line, using an eye dropper to accurately fill the flask.
4) Cap the flask, invert and mix.

Prelab:

1) Calculate the mass of solid copper (II) sulphate pentahydrate required to prepare 100 mL of a $0.200 \mathrm{~mol} / \mathrm{L}$ solution
2) Calculate the volume of the $0.200 \mathrm{~mol} / \mathrm{L}$ solution needed to make a 100 mL of a $0.0200 \mathrm{~mol} / \mathrm{L}$ solution.

## Materials:

- electronic scale
- copper (II) sulphate pentahydrate
- stirring rod
- 100 mL volumetric flask
- medicine dropper

Procedure:
PART A: Watch and practice pipetting

1. Clean and rinse the pipet with distilled water.
2. Take a small sample of the practice solution(tap water) and run it through the pipet. This is known as rinsing with the sample solution.
3. Hold the pipet near the top between the thumb and the last three fingers of one hand, leaving the index finger free. (This allows for quick finger action and does not cover the calibrated lines on the pipet.) Squeeze the pipet bulb or hold the dispenser with the other hand. (right hand if you're right handed.)
4. Apply the pipet bulb/dispenser to the large end of the pipet. Slowly release the pressure on the bulb or roll the dispenser to draw liquid up into the pipet. (The bulb/dispenser should be held firmly to create an air seal. Do not force the pipet into the bulb/dispenser. Never use your mouth. To slow the rise of the liquid, press the end of the pipet into the bottom of the beaker.)
5. When the liquid level rises above the calibrated mark on the pipet, remove the bulb and place your index finger over the end of the pipet. Your index finger is used rather than the thumb because the index finger has better control. Gradually roll the index finger breaking the air seal and allowing the fluid to flow out of the pipet. Allow the level to drop until the bottom of the meniscus is on the calibrated mark. Hold the level there by replacing the air seal with your index finger. (NOTE: If you are using the dispenser, locate the release valve and push it until the fluid has reached the calibrated mark. You can remove the dispenser and use your index finger for practice.)
6. Place the tip of the pipet against the inside wall of the receiving container and allow the contents to flow out of the pipet. Stop the flow at the desired calibrated mark if you are using graduated pipets. (NOTE: Pipets are calibrated to include the fluid that remains after the pipet has been touched to the inside of the receiving container.)

## Part B: Preparing solutions

1. Measure and record the mass of copper (II) sulphate pentahydrate required to prepare 100 mL of 0.200 $\mathrm{mol} / \mathrm{L}$ solution.
2. Add 40- 60 mL of distilled water to the copper (II) sulphate pentahydrate in the beaker. Stir the solution with a clean stirring rod until the copper (II) sulphate pentahydrate is all dissolved.
3. Transfer the solution from the beaker into a clean 100 mL volumetric flask. Use the wash bottle ( distilled water) to rinse any solution from the stirring rod, and the beaker into the volumetric flask.
4. Use a medicine dropper to carefully bring the bottom of the solution meniscus to the 100 mL mark on the volumetric flask. Stopper the volumetric flask and invert several times.

## Part C: Proceed with the following steps to dilute the solution from Part A.

1. Pour the $0.200 \mathrm{~mol} / \mathrm{L}$ copper (II) sulphate solution from Part A into a clean, dry 250 mL beaker.
2. Use a 25 mL graduated pipette to transfer the desired volume (determined in the prelab) of the $0.200 \mathrm{~mol} / \mathrm{L}$ copper sulphate into a clean 100 mL volumetric flask.
3. Add distilled water into the volumetric flask until the bottom of the meniscus reaches the 100 mL mark. Use an eye dropper when the water nears the meniscus line.
4. Stopper the volumetric flask and invert several times.
5. Take the final solution to the teacher who will check it for color intensity or conductivity against a set of standard colors.

## OBSERVATIONS:

PART A: The two types of pipets are $\qquad$ and $\qquad$
PART B: Mass of copper (II) sulphate pentahydrate required $\qquad$
PART B: Observations

PART C: Observations

## Analysis:

PART A

1. When must a pipet be used instead of a graduated cylinder?
2. Graduated pipets are used for $\qquad$ volumes while volumetric pipets are used for
$\qquad$ volumes.
3. Why must the pipet be rinsed with the sample solution?
4. List two reasons why the index finger is used rather than the thumb.
5. List two ways that the flow of the liquid can be slowed down.
6. How might the accuracy of the pipet be checked?

## PART B \& C

7. What is the solvent and what is the solute in the solution in PART B.
8. Calculate the concentration of the solution after dilution.
9. What property of the copper (II) sulphate solution changes noticeably upon dilution?
10. The original and diluted solution will react with zinc metal. Predict which solution would react with zinc at a faster rate? Why? Write a balanced equation and the type of reaction.
11. Which solution would contain a greater number of moles of solute, 8.00 mL of the concentrated copper (II) sulphate solution or 40.0 mL of the diluted solution? Show calculations.

## Conclusion:

## Chemistry 20 Lab 3.2: Solution Stoichiometry

## Name:

 Due Date: Score:Background: Pure carbonate compounds are very difficult and expensive to obtain. Because of this, it is difficult to prepare a sodium carbonate solution of a particular concentration to any degree of accuracy. In this experiment, you will use solution chemistry to make a sodium carbonate solution, dilute a calcium nitrate solution and determine the mass of the precipitate formed when these solutions are mixed.

1. Calculate the mass needed to make 100 mL of $0.20 \mathrm{~mol} / \mathrm{L}$ solution of sodium carbonate.
2. Calculate the volume of a $0.50 \mathrm{~mol} / \mathrm{L}$ solution of calcium nitrate needed to make 100 mL of $0.20 \mathrm{~mol} / \mathrm{L}$ of calcium nitrate dilute solution.
3. Write a balanced equation for the reaction of sodium carbonate and calcium nitrate. Identify the type of reaction that occurs.
4. Identify the limiting and excess reagents. Read the procedure to determine the volumes used.
5. Predict the theoretical mass of the precipitate that forms in this reaction and record it in the data table.

## Problem:

## Variables:

Manipulated: $\qquad$
Responding:
Controlled (2): $\qquad$

## Materials:

| $\bullet$ | - |
| :--- | :--- |
| $\bullet$ | - |
| $\bullet$ | $\bullet$ |
| $\bullet$ | $\bullet$ |

## Procedure:

1. Measure enough grams of sodium carbonate to prepare a 100 mL of $0.20 \mathrm{~mol} / \mathrm{L}$ solution. Place the solid sodium carbonate into a 100 mL volumetric flask and add water.
2. Using a graduated cylinder put 50.0 mL of the sodium carbonate solution you prepared into a 250 mL beaker.
3. Dilute a $1.0 \mathrm{~mol} / \mathrm{L}$ solution of calcium nitrate to make a $100 \mathrm{~mL} 0.20 \mathrm{~mol} / \mathrm{L}$ solution of calcium nitrate.
4. Slowly add the $100 \mathrm{~mL} 0.20 \mathrm{~mol} / \mathrm{L}$ solution of calcium nitrate to the beaker with sodium carbonate.
5. Determine the mass of a filter paper. Fold the filter paper, place it into a funnel and wet with a few drops of distilled water.
6. Filter the solution and collect the precipitate in the filter paper. When the filtering is done, carefully remove the filter paper, opening it to promote drying. Place the filter paper onto a labeled paper towel and let it dry overnight.
7. Determine the mass of the precipitate.
8. Dispose of the solution down the sink and throw the solids into the garbage.

Data \& Observations:
DATA TABLE

| Mass of filter paper (before) |  |
| :--- | :--- |
| Mass of filter paper and precipitate |  |
| Mass \& moles of precipitate (actual - exp.) |  |
| Predicted mass of precipitate (theoretical) |  |
| Percent yield and percent error |  |

Analysis:

1. Determine the number of moles of precipitate collected and place the value in the Data table above.
2. Determine the percent yield and percent error and place the value in the table above.

Conclusion: (4 parts)

## Chemistry 20 Lab 3.3: Chemistry 20: Five Day Copper Recovery Lab

Name: $\qquad$ Due Date: $\qquad$ Score: $\qquad$
Objective: Recover 1.50 g of Cu after making 5 chemical conversions

## Materials:

- Copper
- Assorted beakers
- $3.00 \mathrm{~mol} / \mathrm{L}$ nitric acid
- Watch glass
- pH paper
- $3.00 \mathrm{~mol} / \mathrm{L}$ sodium hydroxide
- Ice/Snow bath
- Distilled water
- $3.00 \mathrm{~mol} / \mathrm{L}$ sulphuric acid
- Zinc metal


## Procedure:

## Conversion 1: Changing $\mathrm{Cu}_{(\mathrm{s})}$ to $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{aq})}$

1. Take approximately 1.50 g of copper wire or copper turnings, clean off any corrosion, weight accurately and record the mass on the data sheet.
2. Place the sample in a clean 400 mL beaker and carefully add 20.0 mL of dilute nitric acid. $\left(\mathrm{HNO}_{3(\mathrm{aq})}\right)$.
3. Place a watch glass on top. Briefly observe the reaction that is taking place and record observations on the data sheet.
4. Set the beaker in the fume hood until the copper has dissolved. The brownish-orange gas produced by the reaction is nitrogen dioxide, $\mathrm{NO}_{2(\mathrm{~g})}$. The blue colour of the solution is characteristic of many copper compounds dissolved in water.
5. Expressed in words, the reaction is:

$$
\text { copper + nitric acid ---> copper(II) nitrate }+ \text { nitrogen dioxide }+ \text { water }
$$

6. Complete the attached data sheet for Conversion 1.

Conversion 2: Changing $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{aq})} \mathrm{to} \mathrm{Cu}(\mathrm{OH})_{2(\mathrm{~s})}$

1. Test your solution from conversion 1 with pH paper. Obtain 20.0 mL of $\mathrm{NaOH}_{(\mathrm{aq})}$ and test this solution with pH paper also. Record both pH values on the data sheet.
2. Fill a 600 mL beaker one third full of ice water or cold tap water. Carefully place the 400 mL beaker containing the conversion 1 solution inside the 600 mL beaker so that the beaker floats in water. (looks like a double boiler)
3. Cautiously and carefully add the 20.0 mL of sodium hydroxide solution to your copper(II) nitrate solution in the 400 mL beaker.
4. Mix the solutions with a gently \swirling motion in the 400 mL beaker. The sodium hydroxide neutralizes or destroys the acid properties of the excess nitric acid still present. This is an exothermic reaction and the beakers may become very HOT. BE CAREFUL...
5. Test the resulting solution with pH paper and record the pH on the data sheet.
6. If this pH test does not match the original colour of the basic sodium hydroxide, add additional sodium hydroxide to the 400 mL beaker while mixing until it does. A pale blue solid precipitate should have formed.
7. Here is how the reaction is expressed. copper (II) nitrate + sodium hydroxide --> copper (II) hydroxide + sodium nitrate
8. Complete the attached data sheet.

## Conversion 3: Changing $\mathrm{Cu}(\mathbf{O H})_{2(\mathrm{~s})}$ into $\mathrm{CuO}_{(\mathrm{s})}$

1. Add 100 mL of distilled water to the beaker containing the copper (II) hydroxide precipitate.
2. Heat to a gently boil and stir until all the material is converted to a brown - black substance. This substance is copper (II) oxide.
3. Remove the stirring rod from the beaker and let the beaker and the solution cool for about 5 minutes.
4. Decant the extra liquid. (Be sure not to lose any solid!!).
5. Wash the precipitate remaining in the beaker by adding about 100 mL of distilled water and stirring gently.
6. Let the precipitate settle for another 5 minutes and pour off the wash water again leaving all the solid particles in the beaker.
7. The conversion of copper (II) hydroxide to black copper (II) oxide is presented this way.
copper (II) hydroxide ---> copper (II) oxide + water
8. Complete the attached data sheet.

## Conversion 4: Changing $\mathrm{CuO}_{(\mathrm{s})}$ to $\mathrm{CusO}_{4(\mathrm{aq})}$

1. Add 50.0 mL of sulphuric acid $\left(\mathrm{H}_{2} \mathrm{SO}_{4(\mathrm{aqq})}\right)$ to the black copper (II) oxide in the beaker. Stir gently.
2. The oxide will dissolve within a minute. The blue solution results from the formation of copper (II) sulphate. The conversion is presented below.
copper (II) oxide + sulphuric acid --> copper (II) sulphate + water
3. Complete the attached data sheet.

## Conversion 5: Changing $\mathrm{CuSO}_{4(\mathrm{aq})}$ to $\mathrm{Cu}_{(\mathrm{s})}$

1. Add about 7.00 g of granular zinc to the copper (II) sulphate solution.
2. Immediately cover the beaker with a watch glass and allow it to stand until the blue colour disappears. (Swirl gently).
3. Your may add 20 mL more of the acid after about 5 minutes if the reaction seems to be proceeding too slowly.
4. Let the beaker stand until the zinc metal completely disappears. Since the reaction produces hydrogen gas, the absence of bubbles will indicate when the zinc metal is all consumed.
5. This is the reaction that brings us back the solid copper with which we began.
```
copper (II) sulphate + zinc ---> copper + zinc sulphate
```

6. Finally the second step of this chemical change removes any solid zinc which is still remaining in the beaker.
zinc + sulphuric acid ---> hydrogen gas + zinc sulphate
7. Allow the solid copper to settle. Decant and discard the clear liquid.
8. Wash the copper in the beaker three times with 50 mL portions of hot distilled water, stir vigorously and let settle. Decant after each wash. Don't lose any solid.
9. Accurately weigh an evaporating dish. Record this value on the data sheet.
10. Then transfer all the copper metal to the dish, using a small amount of wash water.
11. Evaporate the water from the copper in an oven or over a steam bath made with a beaker. Heat the sample only long enough to dry the copper thoroughly. Hot dry copper will react with oxygen from the air to form copper(II)oxide. Remember we have already been there!!
12. Weigh the cool dry dish plus the copper.
13. Complete the attached data sheet.

## Five Day Copper Recovery Lab Data Sheet

Name: $\qquad$
Start Date: Hypothesis:

Partner:
End Date: $\qquad$

Observations and calculations for Copper Recovery Lab:
Initial mass of $\mathrm{Cu}_{(\mathrm{s})}$ g (Beginning of Conversion 1)
Final mass of Cu plus dish
$\qquad$
mass of evaporating dish alone
$\ldots \mathrm{g}$ (End of Conversion 5)
Final mass of recovered $\mathrm{Cu}_{(\mathrm{s})} \quad \mathrm{g} \overline{(\text { End of Conversion 5) }}$
$\%$ recovery $=$ Final Mass of $\mathrm{Cu}_{(\mathrm{s})} \times 100=$ $\qquad$ \%
initial mass of $\mathrm{Cu}_{(\mathrm{s})}$
Conclusions: Write a complete conclusion. Account for any difference(s) between the initial amount of copper used and the amount of copper recovered in your experiment.

Analysis Questions: Please answer these questions as you finish each conversion.
Conversion 1: Changing $\mathrm{Cu}(\mathrm{s})$ to $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\text { aq })}$

1. Write non-ionic, total ionic, and net ionic equations for the reaction taking place in conversion: copper + nitric acid ---> copper(II) nitrate + nitrogen dioxide + water
2. Calculate the volume of $\mathrm{NO}_{2(\mathrm{~g})}$ produced at STP using the net ionic equation.
3. Calculate the mass of $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\text { aq })}$ that would be dissolved in the resulting solution.
4. What caused the solution to be blue?

## Conversion 2: Changing $\mathrm{Cu}\left(\mathrm{NO}_{3}\right)_{2(\text { aq })}$ to $\mathrm{Cu}(\mathrm{OH})_{2(\mathrm{~s})}$

1. Write the non-ionic, total ionic \& net ionic equations for the reaction taking place in conversion 2 : copper (II) nitrate + sodium hydroxide --> copper (II) hydroxide + sodium nitrate
2. Calculate the mass of $\mathrm{Cu}(\mathrm{OH})_{2(\mathrm{~s})}$ produced in this step using the net ionic equation.
3. What is a precipitate and identify it?
4. What did the pH paper tell you about your solution?

## Conversion 3: Changing $\mathrm{Cu}(\mathbf{O H})_{2(\mathrm{~s})}$ into $\mathrm{CuO}_{(\mathrm{s})}$

1. Write the balanced equation for the reaction taking place in conversion 3. copper (II) hydroxide ---> copper (II) oxide + water
2. What is the black solid produced in this step?
3. Calculate the mass of copper (II) oxide produced in this step using the non-ionic equation.

## Conversion 4: Changing $\mathrm{CuO}_{(\mathrm{s})}$ to $\mathrm{CuSO}_{4(\mathrm{aq})}$

1. Write the balanced equation for the reaction taking place in conversion 4. copper (II) oxide + sulphuric acid --> copper (II) sulphate + water
2. Determine the mass of $\mathrm{CuSO}_{4(\mathrm{aq})}$ dissolved in the solution using the non-ionic equation.

## Conversion 5: Changing $\mathrm{CuSO}_{4(\text { (aq) }}$ to $\mathrm{Cu}_{(\mathrm{s})}$

1. Write the non-ionic, total ionic, and net ionic equation for the two reactions used in Conversion 5. copper (II) sulphate + zinc $-->$ copper + zinc sulphate

## zinc + sulphuric acid ---> hydrogen gas + zinc sulphate

2. Calculate the mass of zinc solid that should have been consumed in this step to just recover the copper using the correct net ionic equation above.
3. Calculate the mass of $\mathrm{Cu}_{(\mathrm{s})}$ that should have been recovered in this step using the net ionic equation above.
4. Why was it necessary to add an excess of sulphuric acid to remove any leftover zinc.

## Chemistry 20 Lab 3.4: Properties of Acids \& Bases

Name:
Due Date:
Score:
Background: Slogans on soap and shampoo advertising like "low pH " and " pH neutralized" imply that pH is an additive which should be avoided due to its harmful properties. If a shampoo had a very low pH , hair loss would result. Therefore understanding pH is important.

The pH is a measure of the acidity or basicity of a solution and can be measured with a pH meter or acid-base indicators. Acid-base indicators are substances, which change color over a specific pH range. The pH of a solution can be calculated if the hydronium concentration or the hydroxide concentration are known.

In this experiment, you will extract a naturally occurring acid-base indicator and then compare its pH color range with that of a group of synthetic acid-base indicators. Finally, you will use these indicators to estimate the acidity or basicity of a set of common household products and foods. Please answer the following questions.

1. What is the formal definition of pH ?
2. What are the abbreviations, pH ranges and color changes for the indicators used in this lab (don't include universal indicator or red cabage)?
3. What is the pH if the hydronium concentration is $6.78 \times 10^{-7} \mathrm{M}$ ?
4. What pH values suggest a substance is acidic? Basic? Neutral?
5. List one common household substance that is acid and one that is basic.
6. Hydrochloric acid is a common industrial acid.
a) Write the ionization equation for hydrochloric acid.
b) Determine the concentration of hydrochloric acid used if the pH was $1.5 ; 3.5$
c) Using a solution that has a pH of 1.5 , describe the dilution steps you would use to prepare 100 mL solution with a pH of 3.5. Include calculations and equipment you would need.
7. Sodium hydroxide is a common industrial base.
a) Write the dissociation equation for sodium hydroxide
b) Calculate the pH of a $0.300 \mathrm{~mol} / \mathrm{L}$ solution of sodium hydroxide.

## Write a lab report that includes the following:

## 1) Prelab:

- Complete the background information questions.
- From the background information and the procedure, formulate purpose statements for Part I \& Part III and formulate a problem statement for Part II.
- For Part II only, determine
a) the manipulated variable
b) the responding variable
c) two controlled variables
- List the equipment and materials necessary for this experiment.
- Fill in the indicators and substances into the data tables attached at the back of this lab. Keep the indicators and substances in the same order as you find them in the lab.
- Bring your household item to class.

2) Postlab:

- Complete the experiment and fill in the data tables.
- Answer the analysis questions below

1) Which of the indicators used in this experiment, other than the universal indicator or cabbage indicator, could best identify a neutral solution? Explain.
2) What characteristics should an indicator have if it is to be used in a universal indicator mixture?
3) Why is the use of acid-base indicators not ideal for determining the pH of substances such as blood, soft drinks, and coffee?
4) Some of the solutions used in Part II are buffer solutions. Buffers have the ability to resist pH changes when acids or bases are added to them. However, the amount of acid or base that can be absorbed has a limit called the buffering capacity. Briefly describe an experiment you could use to determine the buffering capacity of a buffer solution.

- Write a one or two paragraph conclusion on this lab.


## CAUTIONS:

- Remove contacts, wear closed shoes and put on a lab coat.
- Read labels and note WHMIS symbols.
- Avoid skin contact with low and high pH solutions because they are corrosive. If the chemical comes in contact with your skin, wash with running water for 5 minutes. If the chemical comes in contact with your eye, flush your eyes with water for 20 minutes.
- Report spills. Add baking soda (sodium bicarbonate) to acid spills and vinegar (acetic acid) to base spills.
- Do not touch the hot beaker of red cabbage leaves in Part I.


## Procedure:

PART I: Red Cabbage Acid-Base Indicator

1. Shred two red cabbage leaves and place them in a 250 mL beaker.
2. Add enough tap water to submerge the cabbage. The red dye of the cabbage will gradually dissolve in the hot water.
3. Let the extract cool after most of the red dye has been removed from the cabbage leaves.
4. Record your observations

## PART II: The pH scale and Acid-Base Indicators

1. Place one drop of each of the following solutions in the wells of a spot plate. $\mathrm{pH}=1 / 2 \mathrm{pH}=3$
$\mathrm{pH}=5$
$\mathrm{pH}=7$
$\mathrm{pH}=11$
$\mathrm{pH}=12 / 13$
2. Add one drop of methyl orange to each well. Record your observations in a table.
3. Clean \& dry the spot plate.
4. Repeat steps 1-3 for bromothymol blue, phenolphthalein, alizarin yellow, universal indicator and your red cabbage extract from PART I.
5. For the first four indicators determine the pH ranges and the color changes for those pH ranges.
6. OPTION: Have each group complete the procedure above for one indicator. Have other groups look at their results.

PART III: The pH values of some household products.

1. In a spot plate, place a small amount of each of the following household substances: lemon juice vinegaraspirin distilled water tea baking soda laundry detergent tums other substance

The substance you will bring from home is:

## NOTE: If the substances are not liquids, suspend them in a few drops of distilled water.

2. Using the universal indicator and your red cabbage extract, estimate the pH of each substance. Confirm your pH with another indicator. Record your results in a table.
3. Indicate in the table if the product is an acid, base or neutral.
4. Research and record the name and the chemical formula of the ingredient responsible for each product's acidic or basic property. Look at the containers and/or look in the library for extra help.
5. Dispose of the contents as directed by your teacher.

Data \& Observations:
PART I

PART II
Title:
NOTE: Use the first letter of the color to represent that color and then make a key COLOR KEY:

|  | pH's |  |  |  |  |  |  |  |  |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| Indicator | 1 | 3 | 5 | 7 | 9 | 11 | 13 | pH range | Colour Changes |
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## PART III

Title:

| Substance | Estimated pH | $\mathbf{A} / \mathbf{B} / \mathbf{N}$ | Ingredient | Formula |
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## Chemistry 20 Lab 3.5: Reactions with Acids

Name:
Due Date:
Score:
Background: Acids exhibit characteristics reactions with some metals, carbonates, hydrogen carbonates and bases. These reactions can be destructive, when acid rain reacts with metals in bridges and vehicles and with the limestone in buildings and statues.

To prevent the acids in soils from reacting with the steel in pipelines, the pipes are coated with tar and wrapped. Acids, such as those in pineapples, would quickly corrode the steel cans in which they are packed if it were not for a thin coating of tin on the can. The corrosive properties of acids can be useful. For example, automobile bumpers are chemically cleaned by dipping them in acid before they are plated. Concrete floors are often "etched" with acid to improve paint-concrete adhesion. Lastly, scale buildup in kettles and water pipes can be removed with acids.

When acids and bases react together a neutralization reaction occurs. According to Arrhenius an acid produces $\mathrm{H}+$ ions and a base produces OH - ions. When acids and bases are mixed, the $\mathrm{H}+$ ions and the $\mathrm{OH}-$ ions combine to form water, destroying the acidic and basic properties of the solution. The other product is a salt that can be recovered by evaporating the water. Neutralization reactions are used to make fertilizers. Lime (calcium hydroxide) removes acid in your garden that has been produced by decaying material. Magnesium hydroxide in Milk of Magnesia neutralizes acid build-up in your stomach. Lastly in the lab, acid spills are neutralized with baking soda (sodium bicarbonate) and base spills are neutralized with vinegar (acetic acid)

In this experiment you will observe some of these reactions of acids.
CAUTIONS:

- Remove contacts, wear closed shoes and put on a lab coat.
- Read labels and note WHMIS symbols.
- Avoid skin contact with low and high pH solutions because they are corrosive. If the chemical comes in contact with your skin, wash with running water for 5 minutes. If the chemical comes in contact with your eye, flush your eyes with water for 20 minutes.
- Report spills. Add baking soda (sodium bicarbonate) to acid spills and vinegar (acetic acid) to base spills.


## PART I: Reactions with metals, carbonates and hydrogen carbonates

## Problem:

## Variables:

Manipulated:
Responding:
Controlled:(2)
Equipment \& Materials:

## Procedure:

1. Label six clean test tubes from A to F and place them in a test-tube rack.
2. Using a 150 mL beaker, pour a $6.00 \mathrm{~mol} / \mathrm{L}$ solution of hydrochloric acid into the six test tubes to a depth of 2 cm .
3. Using a scoopula, add one of the following materials to each test tube.

- A 2 cm iron strip
- B 2 cm zinc strip
- C 2 cm copper strip
- D 2 cm magnesium strip
- E pea-size scoop of sodium hydrogen carbonate
- F pea-size scoop of calcium carbonate

4. Observe the reactions for several minutes. For the test tubes where a gas evolves, hold a lighted wooden splint to the mouth of the test tube. Record your observation in a table.
5. For the test tubes with little or no reaction, use a test tube holder and place the test tubes into a beaker of hot water. Proceed as in Step 4.
6. Dispose of the materials as instructed by your teacher.

## Data \& Observations

Record your observations for the seven test tubes

## Analysis:

1. List the metals from the least reactive to the most reactive with hydrochloric acid.
2. What gas was produced in the reaction between a metal and the hydrochloric acid? What evidence supports this? Write an equation for the test with the splint and the gas.
3. Write balanced equations for the reaction of each metal with hydrochloric acid.
4. Carbon dioxide, water and a salt are the products for the reactions between sodium hydrogen carbonate or calcium carbonate and hydrochloric acid. Write out balanced equations for these two reactions.

## PART II: Neutralization reactions

Problem:

## Variables:

Manipulated:
Responding:
Controlled:(2)

## Equipment \& Materials:

## Prelab Calculations

1. Write a non ionic, total ionic and net ionic equation for the reaction between hydrochloric acid and potassium hydroxide.
2. If 5.0 ml of $1.00 \mathrm{~mol} / \mathrm{L}$ solution of potassium hydroxide is neutralized, what volume of a $1.00 \mathrm{~mol} / \mathrm{L}$ solution of hydrochloric acid is needed?
3. Calculate the mass of the salt that theoretically should be produced.

## Procedure:

1. Measure the mass of a clean, dry evaporating dish. Record the mass in a data table.
2. Place the evaporating dish in the mouth of a 250 mL beaker to steady the dish.
3. Use a 10 mL graduated pipet to measure 5.0 mL of a $1.00 \mathrm{~mol} / \mathrm{L}$ solution of potassium hydroxide. Carefully pour the solution into the evaporating dish.
4. Add one drop of bromothymol blue indicator to the potassium hydroxide solution. Record the color produced in an observation table.
5. Clean and use a 10 mL pipet to measure the volume of a $1.00 \mathrm{~mol} / \mathrm{L}$ solution of hydrochloric acid that you calculated you needed. (Look in Prelab calculations.) Record any color changes in an observation table.
6. Label the evaporating dish and place in the fumehood overnight to evaporate.
7. Measure and record the mass of the evaporating dish and salt. Record this in a data table.
8. Determine the actual mass of the salt. Determine the percent error. Record these values in a data table.
9. Examine the salt and record and observations. DO NOT taste the salt. Record an observations in an observation table.
10. Dispose of the salt down the sink.

## Data \& Observations

Make a data table to record your masses and percent error

Write down your observations.

## Analysis:

1. Is neutralization an endothermic or exothermic reaction? Support your answer.
2. What is the name of the salt produced?
3. Why should the salt prepared not be tasted?
4. Compare your salt to a commercially prepared salt. Did any bromothymol blue remain in your salt. Support your answer.

Conclusion: Write a four part conclusion for part I and part II of this lab.
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## Chemistry 20 Lab 3.6: Acid Base Titration Lab

Name: Due Date: Score:
$\qquad$
Purpose: $\quad$ To determine the concentration and the pH of a sample of nitric acid solution.
Materials: $\quad 10.0 \mathrm{~mL}$ of nitric acid solution in a 100 mL beaker
50 mL of sodium hydroxide solution in a 100 mL beaker
A 250 mL waste beaker
Burette and burette stand
10 mL pipette and bulb
Bromothymol blue indicator solution

## Procedure:

1) Prepare a 100 mL solution of $0.108 \mathrm{~mol} / \mathrm{L}$ solution of sodium hydroxide if your instructor has not done so already.
2) Using a pipette, take 10.0 mL of nitric acid solution and place it into a 100 mL beaker. Add 5 drops of bromothymol blue indicator solution.
3) Fill a clean burette with the titrant (sodium hydroxide). Let a little of the titrant out to remove any air bubbles.
4) Complete two or three Titration trials according to how you learned in class and record your data below.

Data, Observations and Calculations:

1. Calculate the pH of the original sodium hydroxide solution. (2 marks)
2. Fill in Data Table 1. (4 marks)

| Trial | Final Burette Reading <br> $(\mathrm{mL})$ | Initial Burette Reading <br> $(\mathrm{mL})$ | Volume of NaOH(aq) <br> added at Endpoint(mL) |
| :---: | :--- | :--- | :--- |
| 1 |  |  |  |
| 2 |  |  |  |
| 3 | Average Volume of NaOH used |  |  |

3. During trial 2, determine what affect the volume of titrant added, in 1.0 mL segments, has on the pH of the solution. Continue adding titrant 5.0 mL past the end point. Construct Data Table 2 below to compare the volume of titrant to the pH . (2 marks)

|  |  |
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4. Construct a titration curve using the data in table 2. Label the endpoint and equivalence point on the graph. (4 marks)

5. How does the average volume compare to the endpoint volume on your graph? (1 mark)
6. Use Acid/Base Stoichiometry to calculate the concentration and the pH of the nitric acid solution. (5 marks)
7. You will be given an accuracy mark of $0-10$ depending on how accurate your titration is. (10 marks)

## Chemistry 20 Lab 4.1: The Ideal Gas Constant

Name: $\qquad$ Due Date: $\qquad$ Score: $\qquad$
Background: Most scientific constants, such as molar mass and molar volume, are determined empirically. This investigation is an experimental design to determine the value of the universal gas constant.
Prelab:

1. Formulate a problem statement.
2. Write out the variables

Manipulated:
Responding:
2 Controlled:
3. Determine the at least 12 materials needed for this lab.

4. Make a prediction of what the ideal R value should be for all three pressure units.
5. Determine the formula for calculating " R ".
6. Calculate the concentration of the diluted hydrochloric acid assuming that the graduated cylinder can hold 120 mL ?
7. Write out a balanced reaction. What type of reaction is this?

## Procedure

CAUTION: Handle the acid with care. If you spill some acid please report it immediately to your teacher. Acid on your skin should be washed for 5 minutes under cold water. Acid in your eyes should be washed with an eye wash for 10 minutes. Acid on the desk or floor should be neutralized with baking soda.

1. Obtain a strip of magnesium ribbon about 40 mm long.
2. Measure and record the mass of the magnesium.
3. Fold the magnesium ribbon to make a small compact bundle (size of a pencil eraser).
4. Wrap a fine copper wire around the magnesium, making a cage to hold it but leaving 30 mm to 50 mm of the copper wire free for a handle.
5. Carefully pipette 15 mL of $3.0 \mathrm{~mol} / \mathrm{L}$ hydrochloric acid into the graduated cylinder.
6. Slowly fill the graduated cylinder to the brim with water from a beaker. As you fill the cylinder, pour down the side of the cylinder to minimize mixing of water with acid.
7. Half-fill a 600 or 1000 mL beaker with water
8. Bend the copper-wire handle through the holes of a stopper so that the magnesium hangs about 10 mm below the bottom of the stopper.
9. Insert the stopper into the graduated cylinder. The liquid in the cylinder will overflow a little.

Wipe the liquid up. Make sure the stopper seals the edges of the cylinder.
10. Plug the hole of the stopper with your thumb, quickly invert the cylinder, \& immediately lower it so that the stopper is below the surface of the water. Remove your thumb.
11. Record your observations.
12. When the reaction has stopped, raise or lower the graduated cylinder so that the level of liquid inside the beaker is the same as the level of liquid in the graduated cylinder.
13. Measure and record the volume of gas in the graduated cylinder. (Measure at eye level.)
14. Record the laboratory temperature (convert to K ) and pressure (convert to kPa ).
15. Pour the liquids down the sink and clean up.

Observations:
Record your observations in a data table .

## Interpretation:

1. Using stoichiometry calculate the number of moles of Hydrogen produced
2. Determine the actual R value.
3. Using the calculated R value and the theoretical R value from your notes, determine the percent error for this investigation.

Conclusion: Write out a complete conclusion.

## Chemistry 20 Lab 4.1Production of Hydrogen Gas Due Date: <br> $\qquad$

Name: $\qquad$ Score: $\qquad$
Purpose: To produce hydrogen gas by reacting hydrochloric acid and zinc and observe the flammable nature of the hydrogen gas.
Safety: Wear PPE. Handle the acid with care and follow acid safety protocol.

## Procedure:

1) In a 250 mL flask, add 75 mL of $4.00 \mathrm{~mol} / \mathrm{L}$ hydrochloric acid to close to 7.00 g of zinc.
2) Immediately place a balloon over the top of the flask to trap the hydrogen gas as it is being generated
3) While waiting for the reaction to be completed, work on the questions below.
4) If time permits do the extension on the next page and answer the extension questions.

## Observations \& Questions:

1. Record the exact mass of zinc used: $\qquad$
2. Calculate the number of moles of zinc used.
3. Write out the balanced equation for the reaction of zinc with hydrochloric acid.
4. Using moles of zinc used, calculate the mass of hydrogen gas that should be produced.
5. Describe the smell of the hydrogen gas produced.
6. At SATP what volume of hydrogen gas should be produced?
7. Write out the balanced equation for the combustion of hydrogen gas.
8. Calculate the theoretical mass of water vapour that should be produced by the combustion of the hydrogen.
9. Calculate the theoretical volume of gas at SATP that reacted with the hydrogen gas.
10. Describe the combustion of hydrogen.

## 11. CHALLENGE:

a. Calculate the number of moles of HCl in the original solution.
b. If the concentration of $\mathrm{HCl}(\mathrm{aq})$ was $6.00 \mathrm{~mol} / \mathrm{L}$ would more or less hydrogen gas have been produced? Carefully explain your reasoning.
c. If the concentration of $\mathrm{HCl}(\mathrm{aq})$ was $2.00 \mathrm{~mol} / \mathrm{L}$ would more or less hydrogen gas have been produced? Carefully explain your reasoning.

EXTENSION: Measuring the volume of a gas and combustion of hydrogen.

## PROCEDURE:

1) Put a small pail inside a larger pail.
2) Fill the smaller pail right to the brim with water, being careful not to spill any of the water into the larger pail.
3) Take the balloon and hold it by the knot and submerge the balloon into the smaller pail until the whole balloon except the knot is submerged. The water displaced will equal the volume of hydrogen gas in the balloon.
4) Remove the balloon and dry it off
5) Remove the small pail and pour the water remaining in the larger pail into a 1.0 L graduated cylinder to measure the volume of water. There may be more than 1.0 L so you may have to measure more than once. Record the volume of hydrogen gas in question 1 below.
6) Place the balloon on a ring stand attached to a clamp
7) Place a candle at the end of a meter stick or large stick and light the candle
8) Make sure you are wearing PPE. Place the lit candle under the balloon with hydrogen and observe the combustion of hydrogen.
9) Answer the question below.

## EXTENSION QUESTIONS;

1. How does the actual volume of hydrogen gas compare to the expected volume of hydrogen gas collected at SATP in question 6 ?
2. Why might the two numbers not agree with each other?
3. Calculate the theoretical volume of gas expected using the ideal gas law
4. Which of the three volumes do you think is the most accurate? Explain.

## Chemistry 20 Lab 5.1: Ionic Models

Name:
Due Date:
Score: $\qquad$
Background Information:
Ionic compounds form specific crystal structures with related formula units. A crystal lattice is the rigid arrangement of ions in a regular pattern determined by the ions' sizes and charges. The formula unit is the smallest ratio of ions found in a crystal lattice with a net charge of zero. Below are four different crystal lattice structures. In this investigation, the lattice structure of sodium chloride will be built and analyzed. The crystal you will build in this investigation is very small in real life. There are roughly $3 \times 10^{18}$ ions of sodium and chloride in each salt crystal. If a person represented each ion, the ions would populate 1 trillion Alberta provinces. The structure of the crystal lattice will be used to determine the formula unit and some of the properties of sodium chloride.

## Prelab Questions

1. Write out a balanced reaction for the formation (simple composition) of sodium chloride?
2. List five physical characteristics that describe ionic compounds like sodium chloride.
3. What is the formula unit of a sodium chloride crystal?
4. What is the shape of a sodium chloride crystal?
5. What is the Lewis dot diagram of sodium chloride?

Materials

* 450 g bag of small, multi-coloured, soft jelly candies (jujubes or gumdrops)
* Box of toothpicks
* A few crystals of sodium chloride

Procedure: Work in groups of $\qquad$ .

1. Separate the jelly candies by colour. Let one colour represent the sodium cation (colour a) and another contrasting colour represent the chloride anion (colour b). You will need 14 of each colour.
2. Take one sodium and one chloride ion and make a sodium chloride atom by using toothpicks as your bond. Draw what you built.
3. Build a cube with your candies so that colour a and colour b are always in opposite corners. You should end up using 4 of each coloured candies. Draw what you built.
4. Build a 3 by 3 by 3 cube. Record how many anions surround the central cation ion. Record how many cations could potentially surround each anion.
5. OPTION: If there is more candies of your color, get 18 more of each colour and build a 4 by 4 cube.

## Observations \& Data:

## Analysis and interpretation Questions

1. How does the crystal lattice models you built correspond to its formula unit.
2. Explain two of the properties of the ionic crystal you built. (They should correspond to the properties of ionic compounds.)
3. Why are there so many different possible ionic crystal shapes?
4. Identify each of the following statements as true or false
a. The shape of crystal lattice structures depends on the size and charge of the ions.
b. All crystal lattice structures have the same basic shape as the sodium chloride crystal lattice.
c. In a crystal of an ionic solid, the ratio of cations to anions is the same as that of the formula unit.
d. In a crystal lattice, cations are arranged to be as far apart from each other as possible.
e. In a crystal lattice, anions are arranged to be as close to as many cations as possible, while being as far from as many anions as possible.
5. Examine the following crystal and or look at the crystal in class The larger (green) balls represent fluorine and the smaller (red) balls represent calcium. What is the unreduced ratio of calciums to fluorines? What is the reduced ratio? What is the formula unit of calcium fluoride?

6. Write the correct formula unit for
a. copper (I) sulphate
b. magnesium oxide

Conclusions:

Extension: Research ONE of the following situations.

1. Alberta and Saskatchewan have very high concentrations of salts in the soil. That is why the soil is so alkaline. In fact near Fort Saskatchewan, salt is mined from the earth. Describe how salt deposits are formed, how salt is mined and what salt is used for.
2. Most minerals get their properties from ionic substances. Describe what minerals are, how they are classified and why they maintain their same basic shape. Relate this to what you have learned about ionic compounds.

## Chemistry 20 Lab 5.2: VSEPR MODELS

Name: $\qquad$ Due Date: $\qquad$ Score: $\qquad$

## PURPOSE:

- To use the VESPR theory to predict shapes around central atoms.
- To construct models of molecules using a kit.


## PROCEDURE:

Construct models of each molecule using the correct colored balls for each atom and one spring for each shared electron pair. Draw a diagram of the model and name the shape. Also draw structural diagrams with bond dipoles (delta regions) and state if the molecule is polar or nonpolar.
OBSERVATIONS: Each row is worth 3 marks( $1 / 2$ mark off for wrong answers)

| molecule <br>  <br> name | Lewis Diagram | \# of lone pairs | \# of bond areas | VESPR shape diagram | name of shape | Structural diagram with bond dipoles | Polar or non polar |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| $\mathrm{NH}_{3(\mathrm{~s})}$ |  |  |  |  |  |  |  |
| $\mathrm{C}_{2} \mathrm{Cl}_{4(1)}$ tetrachloro ethene |  |  |  |  |  |  |  |
| $\mathrm{CF}_{4(\mathrm{~g})}$ tetrafluoro methane |  |  |  |  |  |  |  |
| $\mathrm{OCl}_{2(1)}$ |  |  |  |  |  |  |  |
| $\mathrm{C}_{2} \mathrm{~F}_{2(\mathrm{~g})}$ <br> difluoro ethyne |  |  |  |  |  |  |  |
| $\mathrm{HOF}_{(1)}$ |  |  |  |  |  |  |  |
| $\mathrm{NHF}_{2(\mathrm{~g})}$ |  |  |  |  |  |  |  |
| $\mathrm{C}_{2} \mathrm{H}_{6(1)}$ |  |  |  |  |  |  |  |
| $\mathrm{CHClBr}_{2(1)}$ dibromo chloro methane |  |  |  |  |  |  |  |


| molecule <br>  <br> name | Lewis Diagram | \# of <br> lone <br> pairs | \# of <br> bond <br> areas | VESPR shape <br> diagram | name of <br> shape | Structural <br> diagram with <br> bond dipoles | Polar or <br> non <br> polar |
| :--- | :--- | :--- | :--- | :--- | :--- | :--- | :--- |
| $\mathrm{C}_{2} \mathrm{HF}_{3(\mathrm{l})}$ <br> trifluoro ethene |  |  |  |  |  |  |  |
| $\mathrm{H}_{2} \mathrm{O}_{2(\mathrm{l})}$ |  |  |  |  |  |  |  |
| $\mathrm{CO}_{2(\mathrm{~g})}$ |  |  |  |  |  |  |  |
| $\mathrm{N}_{2} \mathrm{H}_{3} \mathrm{~F}_{(\mathrm{g})}$ |  |  |  |  |  |  |  |
| $\mathrm{C}_{2} \mathrm{H}_{5} \mathrm{OH}_{(\mathrm{l})}$ |  |  |  |  |  |  |  |
| $\mathrm{C}_{3} \mathrm{H}_{6(\mathrm{~g})}$ |  |  |  |  |  |  |  |

Bonus: Draw the structural shape diagram of a glucose molecule. State the shape of the molecule. (This could be cyclic or non-cyclic. State whether it is polar or nonpolar.)

